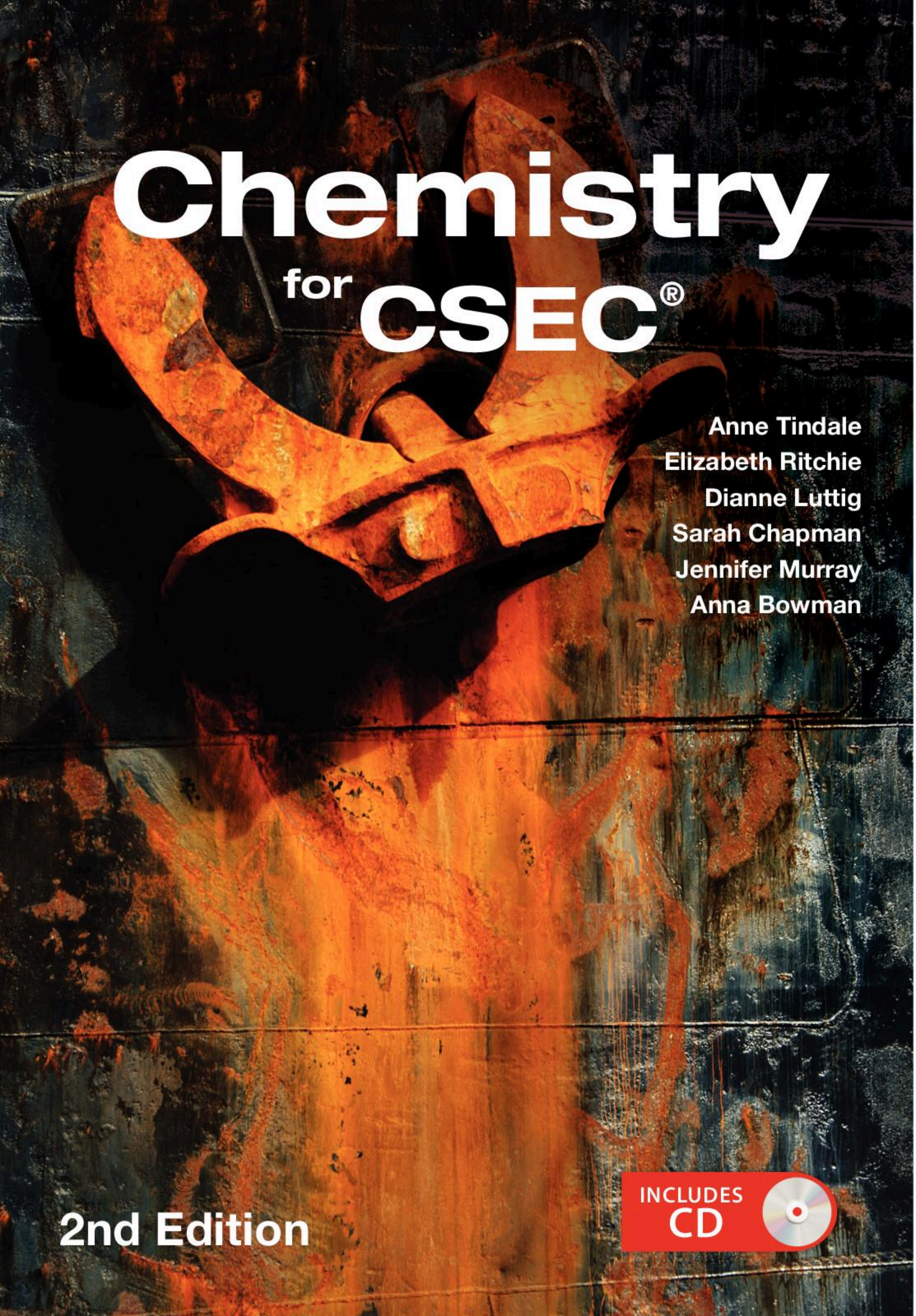

Chemistry for CSEC®

Anne Tindale
Elizabeth Ritchie
Dianne Luttig
Sarah Chapman
Jennifer Murray
Anna Bowman

2nd Edition

INCLUDES CD

Nelson Thornes
Part of Oxford University Press

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Contents

Introduction

To the student

Chemistry for CSEC® is a comprehensive course designed to help you achieve your best in the examination. It has been written by experienced teachers who have included features to make it easier for you to master the key concepts.

The Chemistry syllabus is divided into three sections; Section A, **Principles of Chemistry**, Section B, **Organic Chemistry**, and Section C, **Inorganic Chemistry**. Chapters 1–12 of this book cover topics in Section A, Chapters 13–16 cover topics in Section B, and Chapters 17–22 cover topics in Section C.

Key terms are highlighted throughout the text to help you to 'home in' on the key concepts. Important definitions and laws which you must be able to quote are given in **Key fact** boxes. Each chapter also includes **Did you know?** boxes which contain facts to stimulate your interest, and **Exam tip** boxes which contain valuable tips to help you pass your examination.

You will also find a large number of clear diagrams and colour photographs to enliven and enrich your learning, and a wide range of different **practical activities**. These activities have been designed to help you develop your practical skills and to enhance your learning using a 'hands-on' approach. Your teacher may also use some of these activities to assess your School-Based Assessment (SBA) skills.

Each chapter is broken down into several clearly defined units. The **learning objectives** for each unit are clearly given at the beginning of the unit so that you can see what you are expected to learn in the unit. These learning objectives relate fully to the specific objectives given in the syllabus. Each unit then ends with a selection of **summary questions** to test your comprehension of the material covered in the unit.

At the end of each chapter you will find a list of the **key concepts** to help you to revise the important content of the chapter. This list is followed by a variety of **multiple-choice** questions and usually two **exam-style** questions to help you to apply the knowledge you gained in the chapter to answer the different question types that you will encounter during your examination. The first of the exam-style questions is a structured question requiring short answers and in your CSEC® Examination you will be given spaces on the question paper for your answers. The second is an extended response question requiring a greater element of essay writing in your answers. The marks allocated for the different parts of each question are clearly given.

On the CD you will find a complete multiple-choice test composed of 60 questions to test your knowledge and understanding of material taken from all sections of the syllabus. If you work through all the end of chapter questions and the complete multiple-choice test you will be well prepared for your examination.

On the CD you will also find a section on **Data Analysis** which provides you with details about the Data Analysis question you will be required to answer as well as three questions with mark allocations for you to practise. There is also a section which gives you information about what is required from you in terms of **School-Based Assessment** (SBA). It is extremely important that you read this section before you embark on your two-year course of study.

To the teacher

The writing of *Chemistry for CSEC®* has been driven by experienced Caribbean science educators to ensure that both the requirements of the CSEC® syllabus are met, and that the content is appropriate and relevant for Caribbean students.

It provides a complete coverage of the CSEC® syllabus with the material presented in the order of the syllabus and divided into chapters based on the topics in the syllabus. To help the students grasp important concepts, and for you as a teacher to manage your teaching, each chapter is then broken down into several clearly defined units, each of which develops one of these important concepts.

All important definitions and laws which the students are expected to be able to quote are given in 'Key fact' boxes for students to clearly identify and then learn.

A range of practical activities which closely relate to those suggested in the syllabus can be found throughout the text. You can use these to help you develop a practical approach to your teaching and also to assess the School-Based Assessment (SBA) skills of your students.

Clear, explanatory diagrams and colour photographs which are designed to help students gain a better understanding of each topic are found in all chapters.

'Did you know?' boxes with facts to stimulate the student's interest, and 'Exam tip' boxes containing valuable tips to prepare students for their examination appear throughout the text.

Each unit begins with a list of learning objectives for students to clearly see what they are expected to master in the unit. These relate fully to the specific objectives in the syllabus.

Revision questions are provided at the end of each unit which are designed for students to consolidate their knowledge of the key material found in the unit.

A list of key concepts is given at the end of each chapter to help students revise the key content of the chapter.

Multiple-choice, structured and extended response questions similar to those on the CSEC® Examination are included at the end of each chapter for students to test their knowledge and understanding of the material covered in the chapter and hone their examination skills.

A complete 60-question multiple-choice test with answers is to be found on the CD. This will enable your students to test their knowledge and understanding of the material covered in the entire text.

A Data Analysis section on the CD provides information on answering the Data Analysis question together with three practice questions for students to develop the skills required in answering this question.

The CD also includes a section which provides detailed information about SBA for students to gain an understanding of what is required from them over their two-year programme of study.

Chemistry for CSEC® adopts a practical approach wherever possible. Each practical activity begins with a suggestion indicating how the activity may be used as a basis for SBA. The table which follows the SBA section provides the location in the book of the activities which may be used for the different types of assessment.

Section A

A1 States of matter

Objectives

By the end of this topic you will be able to:

- give a definition of matter
- give the four main ideas of the particulate theory of matter
- explain why scientists find the particulate theory of matter useful
- identify the three main states of matter
- explain the relationship between temperature and the motion of particles.

Chemistry is the study of the structure and behaviour of matter. Everything around us is made of matter. Matter has both mass and volume. Air, water, sand, human beings and animals are all matter. Matter exists in various states. The three main states of matter are solid, liquid and gas.

A1.1 The particulate nature of matter

Matter

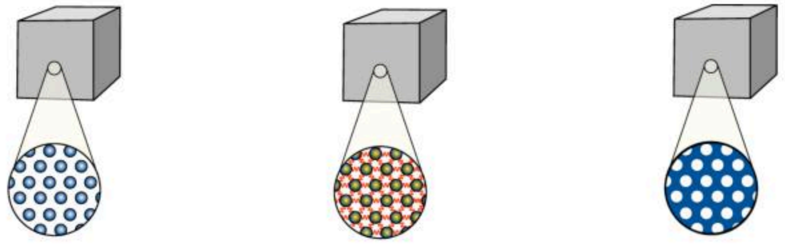
As far back as 460 BC a Greek philosopher called Democritus developed the idea that **matter** consisted of particles. He asked this question: 'If you cut a piece of matter, for example, a piece of gold, in half and then cut it in half again, how many cuts will you have to make before you can cut it no further?' Democritus thought that it ended at some point, the smallest bits of matter, and that these smallest bits of matter, or **particles**, would be the basic building blocks of matter. Today scientists have added to Democritus' idea and now describe matter and its properties using the particulate theory of matter.

The particulate theory of matter

The **particulate theory of matter** states that all matter is made of particles. This theory is very useful because it helps us to explain both the physical properties of matter and also the differences between the three states of matter. We will be looking at the three states of matter in detail in Unit A1.3.

The particulate theory of matter has *four* main ideas:

- all matter is made of particles
- the particles are in constant, random motion
- there are spaces between the particles
- there are forces of attraction between the particles.



matter is made of particles

particles are in constant motion

there are spaces between particles

▲ **Figure 1.1.1** The particulate theory of matter

The particulate theory of matter can be used to explain many physical phenomena that we observe in our everyday lives. Examples of some of these are given on the next page. It explains:

Exam tip

It is important that you know the **definitions** of key terms used in Chemistry. These definitions are provided for you in the **'Key fact'** boxes throughout the book.

Key fact

Matter is anything that has mass and occupies space.

Key fact

The **particulate theory of matter** states that all matter is made up of particles.

- The difference in density of solids, liquids and gases, e.g. why pebbles sink and bubbles rise in water.
- How cooling a liquid can cause it to change into a solid, e.g. when water is placed in a freezer it forms ice.
- Why a smell can move throughout a room, e.g. when chicken is frying, it can be smelt at the other side of the kitchen.
- Why the pressure of a gas increases with an increase in temperature, e.g. car tyres get harder as you drive.
- Why certain vegetables become crisper when soaked in water, e.g. raw potatoes.
- Surface tension in liquids, e.g. certain insects can 'walk' on water.


▲ **Figure 1.1.2** Pebbles sink in water

States of matter

Matter can exist in various forms or **states**. The three states of matter that are the most common are the **solid**, **liquid** and **gaseous** states. The difference between these states lies in the energy and arrangement of the particles.

Particles in the solid state have the least amount of energy, they simply vibrate in their fixed position and they are packed closely together. Particles in the liquid state have medium amounts of energy, they move about slowly and they have small spaces between them. Particles in the gaseous state have the greatest amount of energy, they move about rapidly and they have large spaces between them. You will study this in greater detail in Unit 1.3.

The energy of the particles is directly related to the temperature of the particles and matter can change from one physical state to another by changing its temperature. This change of state occurs because increasing the temperature of a substance increases the **kinetic energy** of the particles in the substance. The greater the kinetic energy the particles possess, the faster they move.

Changing state by changing temperature is a **physical change**. A physical change occurs when the form of the substance is changed without changing its chemical composition, for example, water as a solid, i.e. ice, has exactly the same chemical particles as water in the liquid state and as water in the gaseous state, i.e. water vapour.


▲ **Figure 1.1.3** The drinks illustrate the three states of matter

Summary questions

1 State the three main ideas of the particulate theory of matter.

2 If a crystal of potassium manganate(VII) is dropped into a beaker of water, the purple colour spreads throughout the water. What features of the particulate theory of matter does this observation provide evidence for?

3 Explain why scientists find the particulate theory of matter useful.

4 What are the three states of matter?

5 What is the relationship between temperature and the movement of particles?

A1.2 Evidence for the particulate theory of matter

Objectives

By the end of this topic you will be able to:

- explain evidence which supports the particulate theory of matter
- explain the processes of diffusion and osmosis
- describe experiments which demonstrate diffusion and osmosis
- explain the uses of salt and sugar to control garden pests and preserve food items.

In the previous unit we mentioned that scientists find the particulate theory of matter very useful because it allows them to explain the physical properties of matter. At the same time though, scientists have to provide evidence to support their ideas. There are simple practical activities involving **diffusion** and **osmosis** which we can perform to provide evidence for the existence and movement of particles.

Diffusion

We have all had experience of being aware of a smell, whether it is walking into a bakery, a cosmetic shop or climbing into a car that contains an air freshener. All of these smells are produced at a point in the shop or car, but the smell seems to travel through the air. This process of the smell travelling through the air is as a result of **diffusion**. Diffusion occurs because particles of matter are in constant motion and will move from a region of higher concentration to one of lower concentration.

Key fact

Diffusion is the movement of particles from an area of higher concentration to an area of lower concentration until they are evenly distributed.

Practical activity

Investigating the particulate theory of matter

Your teacher may use this activity to assess:

- observation, recording and reporting
- analysis and interpretation.

You will be supplied with a straw, a beaker containing distilled water and a potassium manganate(VII) crystal.

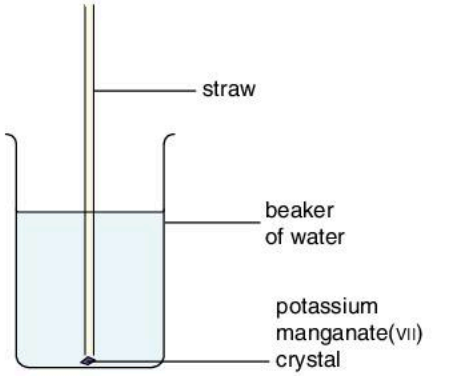
Method

1. Place the straw vertically in the beaker of water until it touches the bottom of the beaker.
2. Drop the crystal of potassium manganate(VII) into the straw without moving the straw.
3. Very carefully remove the straw trying to disturb the water as little as possible.
4. Observe how the purple colour immediately begins to spread throughout the water.
5. Leave the beaker and observe after a few days. Note that the purple colour has spread throughout all the water in the beaker. What conclusion can you draw about:
   - **a** the spaces between the water particles
   - **b** the movement of the potassium manganate(VII) particles?

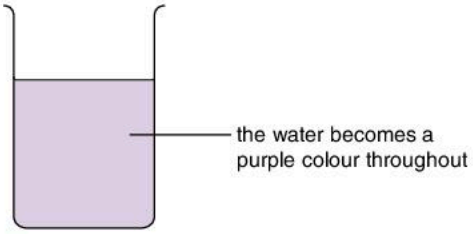
Exam tip

It is very important when answering questions in tests or examinations to distinguish between **observations** and **conclusions**. If you are asked to give your observations, then you must describe what you would see while the experiment is being performed. If you are asked to state what you would conclude from the experiment, then you must give what you can deduce from the observations. A deduction is made by using data from the experiment to arrive at a conclusion.

The potassium manganate(VII) crystal and the water used in the experiment illustrated in Figures 1.2.1 and 1.2.2 are both composed of *minute particles*. The particles in the crystal are packed closely together and those in the water have very small spaces between them. When the crystal is in the water, the minute crystal particles slowly separate from each other and diffuse into the spaces between the water particles. This continues until all the particles have separated from the crystal and have diffused between the water particles.

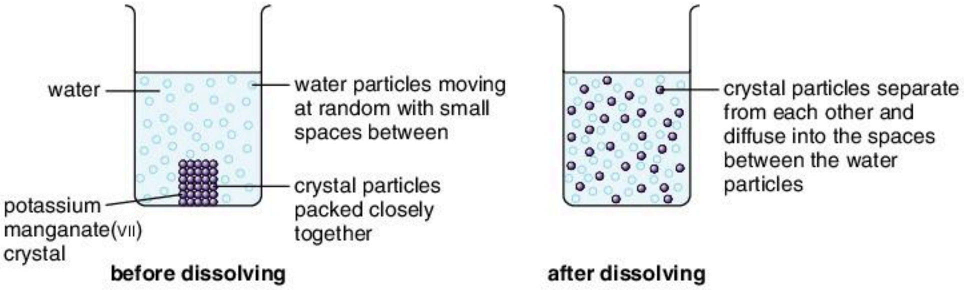


▲ **Figure 1.2.1** Potassium manganate(VII) crystal being placed in water



▲ **Figure 1.2.2** After a while the purple colour spreads throughout the water

The crystal gradually **dissolves** and an aqueous solution is formed, i.e. a solution where water is the solvent. Figure 1.2.3 illustrates the process of diffusion occurring in the experiment.



◀ **Figure 1.2.3** Explanation of diffusion

Practical activity

Investigating diffusion in gases

Your teacher may use this activity to assess:

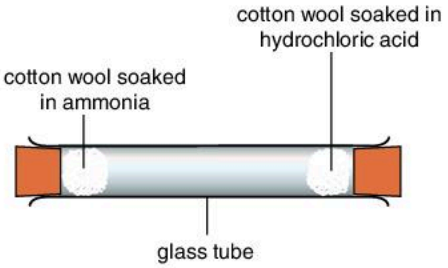
- observation, recording and reporting
- analysis and interpretation.

Your teacher will perform the following experiment:

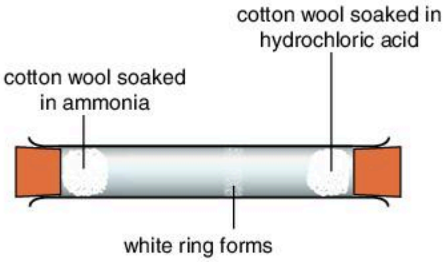
Method

1. Place a glass tube at least 1 m in length between two retort stands.
2. Soak separate pieces of cotton wool in concentrated ammonia solution and concentrated hydrochloric acid and place them simultaneously at each end of the glass tube.
3. Seal the ends of the glass tube with rubber stoppers.
4. Allow time for the ammonia and hydrogen chloride vapours to diffuse. Observe any changes.
5. Use your observations to explain what happened during the experiment.

During the experiment illustrated in Figures 1.2.4 and 1.2.5, the ammonia solution gives off a gas called *ammonia* gas and the hydrochloric acid gives off a gas called *hydrogen chloride* gas. The ammonia and hydrogen chloride particles diffuse through the air in the glass tube towards each other. When the particles meet, they collide and react to form a white solid known as *ammonium chloride*. The ammonium chloride forms a ring inside the glass tube.



▲ **Figure 1.2.4** Ammonia and hydrogen chloride gas diffuse towards each other



▲ **Figure 1.2.5** After a while a white ring forms inside the glass tube

We can represent the reaction between the ammonia and hydrogen chloride as a chemical equation where (g) and (s) indicate the state of the chemicals involved, (g) indicating a gas and (s) a solid:

$$\text{ammonia} + \text{hydrogen chloride} \longrightarrow \text{ammonium chloride}$$
$$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$$

The ammonium chloride forms closer to the cotton wool soaked in hydrochloric acid because the ammonia particles are lighter than the hydrogen chloride particles. Therefore, the ammonia particles move much *faster* through the air than the hydrogen chloride particles.

This experiment provides the following evidence for the particulate theory of matter:

- Particles are able to move – the ammonia and hydrogen chloride particles must have moved towards each other to form the white ring.
- There are spaces between particles – there must have been spaces between the air, ammonia and hydrogen chloride particles to allow them to move between each other.

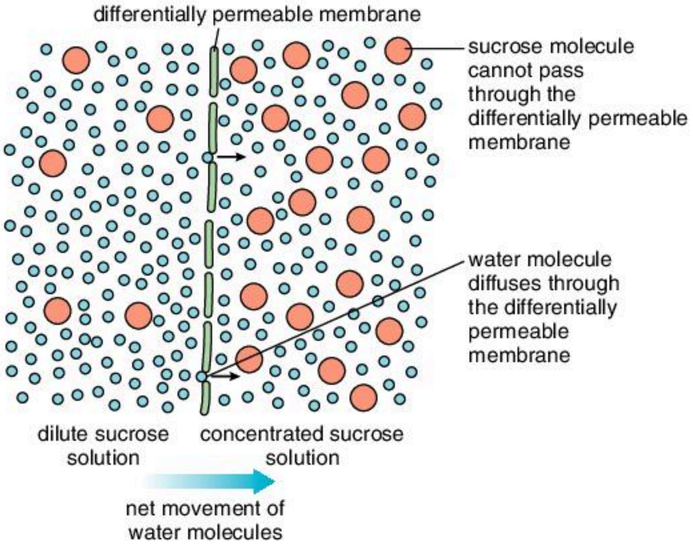
Osmosis

Osmosis is a special case of diffusion, which involves the movement of water molecules through a **differentially permeable membrane** from a region with a lot of water molecules to a region with fewer water molecules. A differentially permeable membrane is a membrane that allows some substances to pass through but not others. You may also find the membrane being called a semi-permeable or selectively permeable membrane. The cell membrane that surrounds biological cells is differentially permeable.

! Key fact

Osmosis is the movement of water molecules from a region with a lot of water molecules, e.g. a dilute solution or pure water, to a region with fewer water molecules, e.g. a concentrated solution, through a differentially permeable membrane.

A differentially permeable membrane contains minute pores. Water molecules are able to pass through these pores. However, the particles of many other substances, which may be dissolved in the water, are unable to pass through. When two solutions, e.g. sucrose solutions, which have different concentrations, are separated by a differentially permeable membrane, the water molecules will diffuse through the pores in the membrane from the more dilute solution to the more concentrated solution. The sucrose molecules, however, do not move because they are unable to pass through the pores in the membrane. The volume of the more dilute solution decreases and the volume of the more concentrated solution increases.



▶ **Figure 1.2.6** Theoretical explanation of osmosis

Practical activity

Investigating osmosis in green paw-paw

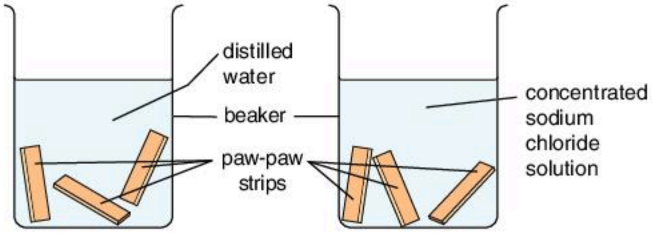
Your teacher may use this activity to assess:

- manipulation and measurement
- analysis and interpretation.

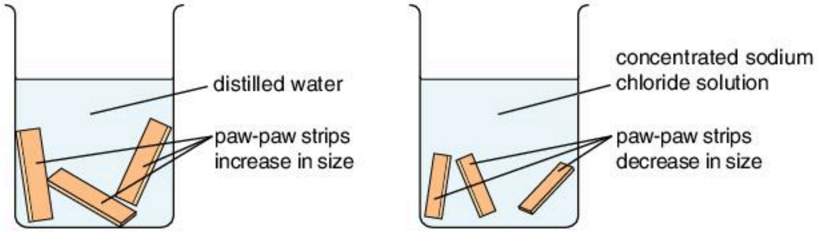
You will be supplied with a piece of green paw-paw (the experiment may be done with potato or yam), one beaker filled with distilled water and one beaker filled with concentrated sodium chloride solution.

Method

1. Cut the piece of green paw-paw into six strips of equal length.
2. Measure and record the length of each strip.
3. Place three of the strips into the beaker containing distilled water and place the other three strips into the beaker containing the concentrated sodium chloride solution.
4. Allow the strips to remain in the solutions for one hour.
5. Remove the strips from the beakers. Feel the strips and take note of the texture of each strip.
6. Measure and record the length of each strip.
7. Explain why the paw-paw strips placed in distilled water become more rigid and have increased in length (consider the direction in which the water molecules move, from the paw-paw into the distilled water or from the distilled water into the paw-paw).
8. Explain why the paw-paw strips placed in concentrated sodium chloride solution become floppy and softer and decrease in length (consider the direction in which the water molecules move, from the paw-paw into the sodium chloride solution or from the sodium chloride solution into the paw-paw).



◀ **Figure 1.2.7** Paw-paw strips in distilled water and concentrated sodium chloride solution



▲ **Figure 1.2.8** After a while the strips in water increase in length and those in the concentrated sodium chloride solution decrease in length

During the experiment illustrated in Figures 1.2.7 and 1.2.8, the cell membranes of the paw-paw cells act as differentially permeable membranes. Water can pass through the cell membranes, either into or out of the cells:

- Distilled water has a higher water content (or lower sodium chloride concentration) than the paw-paw cells, therefore water moves *into* the cells by osmosis, resulting in the paw-paw strip becoming longer and more rigid.
- The paw-paw cells have a higher water content than the concentrated sodium chloride solution, therefore water moves *out of* the cells by osmosis, resulting in the paw-paw strip becoming shorter and softer.

Did you know?

Osmosis works in the same way in your cells as it does in the paw-paw. If you **sweat** a lot you lose water. This lowers the amount of water in your blood and osmosis takes place and starts to pull water out of your cells. For this reason it is very important to drink lots of water on a hot day or when you exercise.

Practical uses of osmosis

We use the principles of osmosis in various ways. These include controlling garden pests and preserving food items.

Controlling garden pests

Slugs and snails, being herbivores, are serious garden pests which ravage many of our precious plants. The skin of these pests is a lot more permeable than the skin of most other animals. This means that they need to keep themselves moist to prevent water evaporating from their bodies causing them to dehydrate and die. We make use of these facts to control slugs and snails in our gardens by using sodium chloride (table salt).

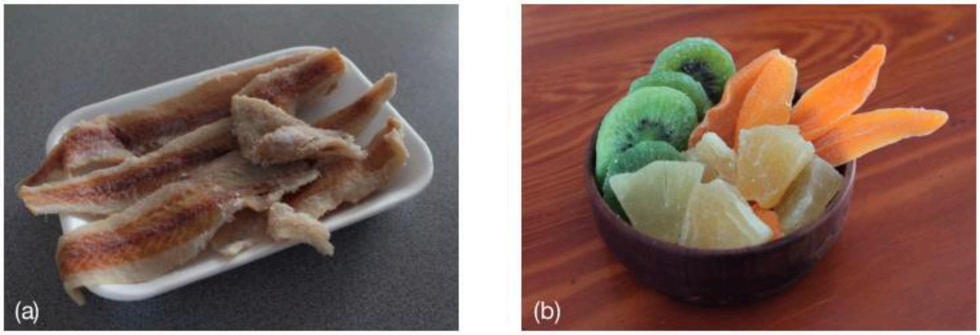
▲ **Figure 1.2.9** Snails are serious garden pests

Sodium chloride is **deliquescent**, which means that it absorbs water readily and dissolves. When sodium chloride is sprinkled on slugs and snails, it absorbs some of the moisture surrounding their bodies and dissolves forming a concentrated solution. This causes water inside their bodies to move out and into the solution by osmosis. If the slugs and snails lose more water than their bodies can tolerate, they die from dehydration.

Preserving food items

Both sodium chloride and sugar are used to preserve food items, e.g. meat, fish, fruits and vegetables. We are all familiar with salt fish, salt pork, crystallised fruits, guava jelly and glacé cherries. Both salt and sugar work in the same way to preserve foods:

- They withdraw water from the cells of the food items by osmosis. This makes the water unavailable for the chemical reactions in cells which cause decay. Without these reactions occurring, the food items do not decay.
- They also withdraw water from the microorganisms that bring about decay, e.g. bacteria and fungi. This prevents these organisms from growing and causing the food items to decay.



▲ **Figure 1.2.10 (a)** Salt fish and **(b)** crystallised fruits are examples of preserved food items

Summary questions

1. For each of the experiments described below, state the feature(s) of the particulate theory of matter for which the experiment provides evidence:
   - **a** If a crystal of potassium manganate(VII) is dropped into a beaker of water, the purple colour spreads throughout the water.
   - **b** If bottles of concentrated ammonia solution and concentrated hydrochloric acid are opened and placed next to each other, a white cloud forms between the two bottles.
2. State any similarities and differences between osmosis and diffusion.
3. Why do we say that osmosis is a special case of diffusion?
4. When a 5.0 cm long raw potato chip was placed in water and left for 30 minutes, its length was found to have increased to 5.5 cm. Use your knowledge of the particulate theory of matter to explain this.
5. Explain how sodium chloride is used:
   - **a** to control slugs in a garden
   - **b** to preserve fish.

A1.3 The three states of matter

You have learnt already that matter exists in three states: solid, liquid and gas. The three states of matter have noticeable differences in their physical properties. **Physical properties** are characteristics that can be observed or measured without changing the chemical composition of a substance. Shape, volume, density, compressibility, solubility, melting point and boiling point are all examples of physical properties. The different physical properties of the three states can be explained by the particulate theory of matter.

Table 1.3.1 summarises the physical properties of the three states of matter and the arrangement of particles in the three states.

We can use the arrangement of particles in the three states of matter to explain the physical properties of the different states:

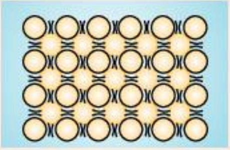
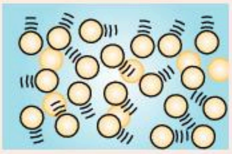
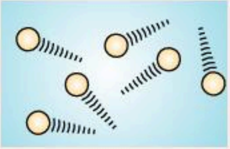
- Solids have a fixed shape because their particles are arranged in a regular way and they are unable to move out of their fixed positions.
- Solids have a fixed volume and are very difficult to compress because their particles are packed very closely together and cannot be pushed any closer together.
- Liquids flow because their particles are able to move past each other.
- Liquids can be compressed slightly because their particles have small spaces between them enabling them to be pushed closer together.
- Gases take the shape and volume of the container they are in because their particles move freely and rapidly. They have only weak forces of attraction between them, therefore they spread out to fill any available space.
- Gases are easy to compress because their particles have large spaces between them, therefore they can be pushed closer together.

Objectives

By the end of this topic you will be able to:

- explain the difference between the three states of matter in terms of energy and arrangement of particles, and forces of attraction between particles
- account for the physical properties of the three states of matter in terms of energy and arrangement of particles, and forces of attraction between particles
- explain the changes of state in terms of energy and arrangement of particles
- understand melting, evaporation, boiling, condensation, freezing and sublimation
- understand and interpret heating and cooling curves.

Table 1.3.1 The properties of the three states of matter

Property	Solid	Liquid	Gas
Shape and volume	Solids have a fixed shape and a fixed volume.	Liquids do not have a fixed shape, but they have a definite volume. Liquids take the shape of the part of the container that they occupy and the surface is always horizontal.	Gases do not have a fixed shape or volume. A gas will take up the space of the container it is placed in. The shape and volume of a gas is, therefore, the shape and volume of the entire container it is in.
Density	Most solids have a high density.	The density of liquids is usually lower than the density of solids.	Gases have a low density.
Compressibility	Solids are very difficult to compress.	Liquids can be compressed very slightly when pressure is applied.	Gases are easy to compress.
Arrangement of the particles	The particles are packed closely together, usually in a regular pattern.	The particles are randomly arranged and have small spaces between them.	The particles are randomly arranged and have large spaces between them.
Forces of attraction between the particles	The particles have very strong forces of attraction between them.	The forces of attraction between the particles are not as strong as those between the particles of a solid.	The particles have very weak forces of attraction between them.
Energy and movement of the particles	Particles in a solid have very small amounts of kinetic energy. The particles vibrate in their fixed position.	Particles in a liquid have more kinetic energy than particles in a solid. The particles move about slowly.	Particles in a gas have large amounts of kinetic energy. The particles move about freely and rapidly.
Arrangement of particles			

Changing state

Matter can be changed from one state to another by heating or cooling. A change of state is, therefore, caused by a change in temperature and consequently a change in the kinetic energy of the particles. For example, in order to change water into ice we need to put the water into the freezer, i.e. we need to remove heat energy. Changing the state of a substance without changing its chemical composition is a physical change. The different changes of state are summarised in Figure 1.3.1.

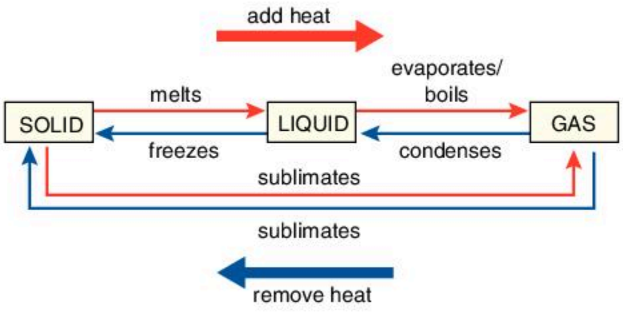


Figure 1.3.1 Summary of the changes of state

We will now look in more detail at each of the processes that cause the changes in state in Figure 1.3.1.

? Did you know?

SCUBA divers make use of the fact that gases are very easy to compress. An average sized SCUBA diving tank holds about 2250 litres of **compressed air**. To understand this, think of a milk carton. Most milk cartons hold one litre of milk, therefore, a SCUBA diving tank holds the same volume of air as 2250 empty milk cartons!

Melting

When a solid is heated, the particles gain kinetic energy and begin to vibrate more vigorously. Eventually the particles are able to overcome the strong forces of attraction between them and they move more freely and further apart forming a liquid, i.e. the solid **melts**. The temperature remains constant while the solid is melting because all the heat energy being supplied is used to overcome the forces of attraction between the solid particles. This constant temperature is known as the **melting point**.

Key fact

Melting point is the constant temperature at which a solid changes into a liquid.

Evaporation

When a liquid is heated, the particles gain kinetic energy and move faster. Some of the particles near the surface of the liquid have enough kinetic energy to overcome the forces of attraction between them and are able to leave the liquid and become a vapour. These particles are said to **evaporate**. The particles that leave the liquid take lots of energy with them, leading to a cooling of the liquid.

Did you know?

When we sweat and the water in the sweat **evaporates** from our skin, it takes energy with it causing our bodies to feel cooler. If we put alcohol on our skin, it evaporates even faster than water because it has a lower boiling point than water. This makes our skin feel even colder than when we sweat.

Boiling

When a liquid is heated its temperature eventually reaches a certain point where it starts to **boil**. At this point the liquid particles have gained enough kinetic energy and started to move fast enough to change into a gas both within the liquid and at its surface. The temperature remains constant while the liquid is boiling because the heat energy being supplied is used to overcome the forces of attraction between the liquid particles.This constant temperature is known as the **boiling point**.

Boiling differs from evaporation in two ways.

- Boiling occurs at a specific temperature, whereas evaporation can take place at any temperature.
- Boiling takes place throughout the liquid, whereas evaporation takes place only at the surface of the liquid.

Key fact

Boiling point is the constant temperature at which a liquid changes into a gas.

Exam tip

If you are asked to give a **difference** between two things, it is essential that you describe the specific property of each, using the word 'whereas' to link the two. Do not describe just one.

Condensation

When the temperature of a gas is lowered, the particles lose kinetic energy and begin to move more slowly. The forces of attraction between the particles become stronger causing the particles to move closer together forming a liquid, i.e. the liquid **condenses**.

Freezing

When the temperature of a liquid is lowered, the particles lose kinetic energy and begin to move more slowly. The forces of attraction between the particles become stronger causing the particles to move even closer together forming a solid, i.e. the liquid **freezes**. The temperature at which this occurs is called the **freezing point**.

The freezing point of a pure substance has the same value as the melting point, e.g. water has a melting point and a freezing point of 0 °C.

Key fact

Freezing point is the constant temperature at which a liquid changes into a solid.

Sublimation

When the forces of attraction between the particles in a solid are weak, the addition of a small amount of heat can cause the solid to change directly into a gas, without passing through the liquid state. If the gas is then cooled it will change directly back to the solid. When a substance changes directly from a solid to a gas or a gas to a solid it is said to **sublime**.

Examples of substances which undergo sublimation are iodine, carbon dioxide (known as ‘dry ice’), ammonium chloride and naphthalene. Moth balls or camphor balls are made of naphthalene. Solid air fresheners also sublime releasing their fragrances into the air.

▶ **Figure 1.3.2** Solid air fresheners sublime

Practical activity

Observing sublimation in iodine

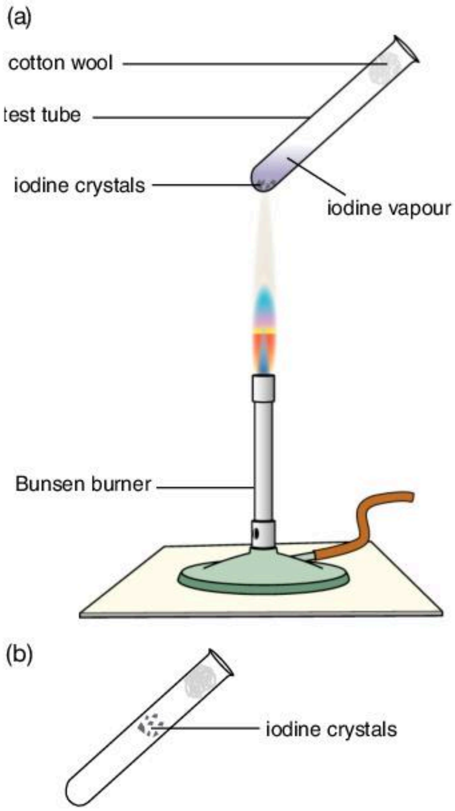
Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with a test tube, a small iodine crystal, a piece of cotton wool and a pair of tongs.

Method

1. Place the iodine crystal into the test tube and place the cotton wool into the mouth of the test tube.
2. Hold the test tube with tongs at a 45° angle and heat the bottom of the tube in the flame of a Bunsen burner until all the iodine crystal has sublimed.
3. Observe what happens as the iodine is heated.
4. Remove the tube from the flame and let it cool.
5. Observe what happens as the tube is cooling.



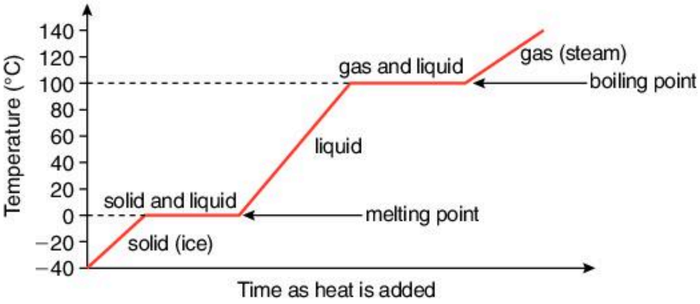
▲ **Figure 1.3.3** Iodine sublimes **(a)** when heated and **(b)** when cooled

During the experiment illustrated in Figure 1.3.3, as the iodine crystal is heated, it sublimes and forms purple iodine vapour which diffuses up the test tube. The top of the tube is much cooler and when the vapour reaches the top, it sublimes back to a solid, forming a ring of iodine crystals around the inside of the tube.

Heating and cooling curves

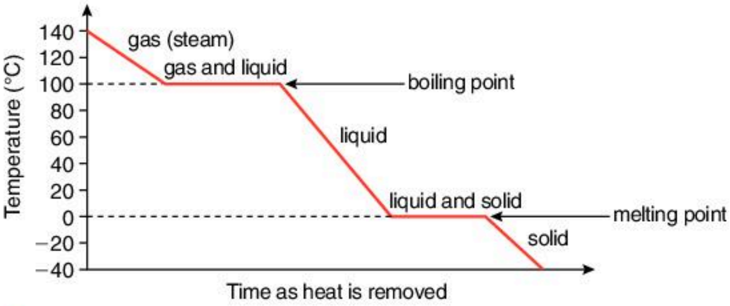
If the temperature of a pure solid is measured at intervals as it is heated and changes state to a liquid and then to a gas, and the temperature is plotted on a graph against time, a **heating curve** is obtained. The heating curve for water is shown in Figure 1.3.4.

The curve shows that as heating occurs, the temperature of the substance increases. However, the graph has two horizontal sections where the temperature remains constant for a period of time even though heating continues. These happen when there is a *change of state*. The first change of state is where *melting* occurs and the temperature remains constant at the melting point of the substance until all the substance has melted, e.g. for water this is 0 °C. The second change of state is where *boiling* occurs and the temperature remains constant at the boiling point of the substance until all the substance has boiled, e.g. for water this is 100 °C.



▲ **Figure 1.3.4** The heating curve for water

If the temperature of a gas is measured at intervals as it is cooled and changes state to a liquid and then to a solid, and the temperature is plotted on a graph against time, a **cooling curve** is obtained. The cooling curve for water is shown in Figure 1.3.5.



▲ **Figure 1.3.5** The cooling curve for water

Like heating curves, cooling curves have two horizontal sections. The first is where the state changes from gas to liquid and the second is where it changes from liquid to solid.

Summary questions

1 Complete the table below. The first row is completed as an example of what is required.

Change of state	Name given to change of state	Energy added or removed to change state
Solid to liquid	Melting	Added
Liquid to gas		
Gas to liquid		
Liquid to solid		
Solid to gas		

2 What are the main differences between evaporation and boiling?

3 **a** Explain what occurs during sublimation.
b Give three examples of solids which undergo sublimation.

4 Explain what a heating curve shows.

Key concepts

- Matter is defined as anything that has mass and occupies space.
- The properties of matter can be explained by the **particulate theory of matter**, which states that:
  - all matter is made of particles
  - the particles are in constant, random motion
  - there are spaces between the particles
  - there are forces of attraction between the particles.
- Matter exists in three states: **solid**, **liquid** and **gas**.
- The difference between the three states is due to the arrangement and energy of the particles.
- The energy of the particles is directly related to the temperature of the particles.
- **Diffusion** and **osmosis** provide evidence to support the particulate theory of matter.
- Diffusion is the movement of particles from an area of higher concentration to an area of lower concentration until they are evenly distributed.
- Osmosis is the movement of water molecules from a region with a lot of water molecules, e.g. a dilute solution or pure water, to a region with fewer water molecules, e.g. a concentrated solution, through a differentially permeable membrane.
- The principles of osmosis are used to control garden pests with sodium chloride and preserve food items with sodium chloride and sugar.
- The energy and arrangement of particles in solids, liquids and gases account for the physical properties of the three states.
- Matter can change from solid to liquid to gas with the addition of heat.
- Matter can change from gas to liquid to solid with the removal of heat.
- The names given to the processes which cause the change from one state to another are melting, evaporation, boiling, condensation, freezing and sublimation.
- Melting and boiling occur at specific temperatures known as the **melting point** and the **boiling point**, respectively.
- Evaporation of a liquid occurs at any temperature and leads to the cooling of the liquid.
- **Sublimation** is the change in state straight from a solid to a gas, or vice versa, without passing through the liquid state. It occurs in compounds with weak forces of attraction between their molecules.
- When the temperature of a pure substance is measured as it is heated and changes state from solid to liquid to gas, and the temperature is plotted against time, a **heating curve** is obtained.
- When the temperature of a pure substance is measured as it is cooled and changes state from gas to liquid to solid, and the temperature is plotted against time, a **cooling curve** is obtained.

Practice exam-style questions

Multiple-choice questions

1 Which of the following provide(s) evidence that matter is made of particles?

I Osmosis
II Decomposition
III Diffusion

A I, II and III
B II and III only
C I and III only
D I only

2 Particles in a solid:

A are packed closely together
B are capable of random movement
C have large spaces between them
D have weak forces of attraction between them.

3 Liquids differ from gases in that:

A the particles in a liquid can move more freely than those in a gas
B the particles in a gas are closer together than those in a liquid
C the particles in a liquid possess more kinetic energy than those in a gas
D the particles in a gas have weaker forces of attraction between them than those in a liquid.

4 The conversion of a gas to a liquid is described as:

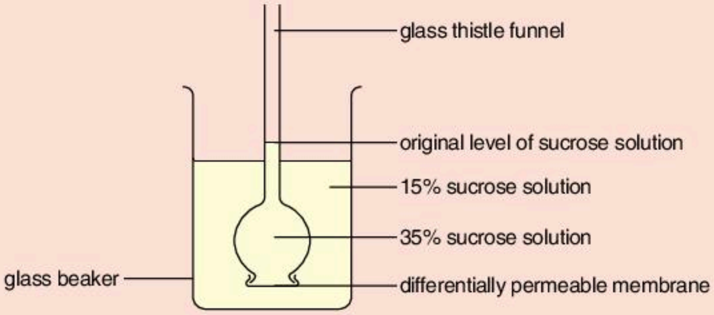
A melting
B evaporation
C condensation
D boiling.

5 Which of the following lists consist of substances which show sublimation?

A Aluminium chloride, iodine, carbon dioxide.
B Ammonium chloride, iodine, carbon monoxide.
C Ammonium chloride, naphthalene, iodine.
D Aluminium sulfate, carbon dioxide, naphthalene.

Structured question

6 a In order to investigate the particulate nature of matter, a student set up the apparatus shown in Figure 1 below and left it for 30 minutes.



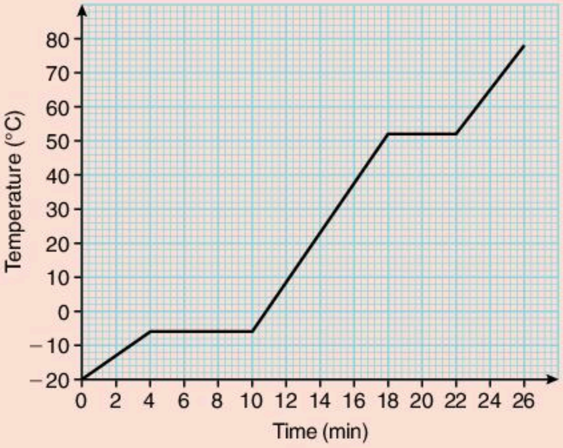
▲ **Figure 1** Apparatus used to investigate the particulate nature of matter

i) State what you would expect to have happened to the level of the sucrose solution in the thistle funnel after 30 minutes. *(1 mark)*

ii) Explain the reason for your observations. *(4 marks)*

iii) Name the process occurring in the apparatus. *(1 mark)*

b A second student measured the temperature of an unknown substance X as it was heated from the solid state through the liquid state to the gaseous state. Her results are plotted in Figure 2. Use the information in the Figure to answer the following questions.



▲ **Figure 2** Graph showing the temperature of substance X during heating against time

i) What state is X in at 68 °C? *(1 mark)*

ii) Over what temperature range does X exist in the liquid state? *(2 marks)*

iii) From the time X starts to melt, how long does it take until it has completely melted? *(1 mark)*

iv) Describe, in terms of energy and behaviour of particles, what is happening to substance X as it melts. *(3 marks)*

v) Using **o** to represent a particle of substance X, draw TWELVE particles as they would appear in X at −15 °C. *(2 marks)*

Total 15 marks

Extended response question

7 a Describe TWO pieces of evidence to support the particulate theory of matter. *(4 marks)*

b Water can exist as a solid, a liquid and a gas. Explain the differences between these three states in terms of arrangement of their particles, forces of attraction between them and their kinetic energy. *(6 marks)*

c By reference to particles, explain clearly the reason for EACH of the following:

i) steam can be converted to liquid water by reducing its temperature *(3 marks)*

ii) a crystal of potassium chloride has a very definite shape. *(2 marks)*

Total 15 marks

A2 Mixtures and their separation

Objectives

By the end of this topic you will be able to:

- distinguish between pure substances and mixtures
- explain the difference between an element, a compound and a mixture
- give examples of elements, compounds and mixtures
- explain the difference between a homogeneous and a heterogeneous mixture.

Elements, compounds and mixtures form a part of our everyday lives. When we wrap our food in aluminium foil we are using an element. When we place salt on our food we are eating a compound. When we drink a cold soda we are drinking a mixture. It may be useful to know how to separate some of these mixtures into their component parts. An example of this is the purification of drinking water.

A2.1 Elements, compounds and mixtures

Matter can be classified into two main groups: pure substances and mixtures.

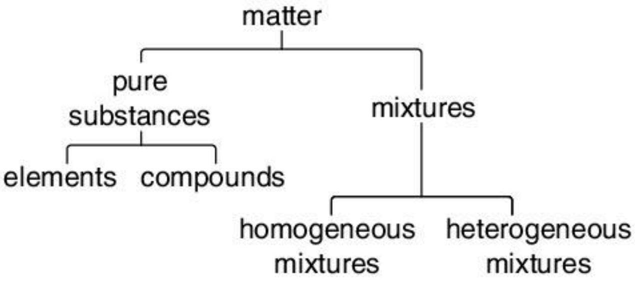
Pure substances have the following general characteristics.

- They have a fixed, constant composition.
- Their properties are fixed and constant.
- The component parts of a pure substance cannot be separated by physical means.

Mixtures have the following general characteristics.

- They have a variable composition.
- Their properties are variable since their components retain their own, individual properties.
- The component parts of mixtures can be separated by physical means.

Pure substances can be further classified into **elements** and **compounds**. Mixtures can be further classified into **homogeneous mixtures** and **heterogeneous mixtures**. The tree diagram in Figure 2.1.1 shows the breakdown of these groups.



▲ **Figure 2.1.1** Classification of matter

Pure substances

A **pure substance** is composed of only one type of material and has the following fixed properties:

- a sharply defined, constant melting point or freezing point
- a sharply defined, constant boiling point
- a constant density.

To determine if a substance is pure or not, its melting point or boiling point is determined. Any impurities in a pure substance will usually *lower* its melting point and cause it to melt over a wider temperature range. Impurities will usually *raise* the boiling point of a pure substance and cause it to boil over a wider temperature range.

Paper chromatography (Unit 2.4) may also be used. If a substance is pure it will produce only one single spot on a chromatogram. If it is not pure it will produce more than one spot.

Did you know?

In many countries where snow and ice pose a problem on the roads in winter, **rock salt** (sodium chloride) is spread on the roads to melt the ice. The salt dissolves in the film of water on the surface of the ice, this lowers its freezing point to below the temperature of the ice and the ice starts to melt. Salt can lower the freezing point of water to about −18 °C, the freezing point of saturated sodium chloride solution.

Practical activity

Comparing the boiling points of pure water and sodium chloride solution

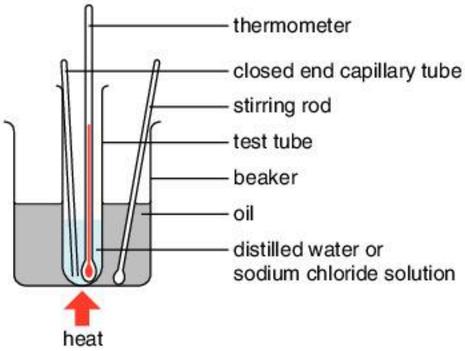
Your teacher may use this activity to assess:

- observation, recording and reporting.

Your teacher will perform the following demonstration.

Method

1. Place 2 cm^3 of distilled water in a test tube.
2. Place an inverted closed end capillary tube into the test tube with the open end facing downwards.
3. Place a thermometer in the tube.
4. Half fill a 250 cm^3 beaker with oil and place the above test tube assembly in the oil bath so that the surface of the water in the test tube is beneath the surface level of the oil.
5. Heat the beaker gently over a Bunsen burner, stirring constantly to ensure that heating is even. Continue heating until a rapid stream of bubbles emerges from the capillary tube. This stream of bubbles indicates that the water in the test tube is boiling.
6. Remove the heat source and observe the stream of bubbles. When the last bubble emerges from the capillary tube, record the temperature.
7. Reheat the oil bath and repeat the cooling process twice more. Record the temperature reading after each trial and average all three temperatures. This is the boiling point of distilled water.
8. Repeat the procedure using a sodium chloride solution in place of the water.
9. What conclusion can you draw about the boiling points of pure water and sodium chloride solution?



▲ **Figure 2.1.2** Determining the boiling point of a liquid

Elements

An **element** is the simplest form of matter. It cannot be broken down into anything simpler by ordinary chemical or physical means. We say 'ordinary chemical means' to exclude nuclear reactions. The smallest particle in an element that has the same properties as the element is an **atom**. Each element is composed of only *one kind* of atom.

Key fact

An **element** is a pure substance that cannot be broken down into any simpler substances by any ordinary chemical or physical means.

Examples of elements are iron (Fe), which is composed of only iron atoms, copper (Cu), which is composed of only copper atoms and oxygen (O_2), which is composed of only oxygen atoms.

Key fact

A **compound** is a pure substance that contains two or more different types of element which are bonded together chemically in fixed proportions and in such a way that their properties have changed.

Compounds

Compounds are composed of *more than one* kind of atom. These atoms are combined together chemically, they are always present in the same proportions by mass and they cannot be separated by physical means. A compound can be represented by a **chemical formula**, which indicates the elements that the compound is made up of and the ratio in which they have combined, e.g. the chemical formula of water is H_2O.

Examples of compounds are water, which is composed of hydrogen and oxygen in a ratio of 2 : 1, sodium chloride (NaCl), which is composed of sodium and chlorine in a ratio of 1 : 1 and methane (CH_4), which is composed of carbon and hydrogen in a ratio of 1 : 4.

The properties of a compound are fixed and are different from the properties of the individual elements that form the compound. For example, hydrogen and oxygen are both gases at room temperature, however, water is a liquid.

Key fact

A **mixture** consists of two or more substances (elements and/or compounds) physically combined together in varying proportions. Each component retains its own independent properties and has undergone no chemical reaction with any other substance in the mixture.

Mixtures

Mixtures are composed of two or more substances which are not chemically combined, therefore their components can be separated by physical means. Some of the physical methods for separating mixtures, which we will be investigating in Unit 2.4, are filtration, evaporation, crystallisation, distillation, fractional distillation and chromatography. In a mixture the component parts are not in a fixed ratio and they retain their own, individual physical properties.

Homogeneous mixtures

A **homogeneous mixture** is one in which the properties and composition are uniform throughout the mixture. The component parts cannot be distinguished from each other. A solution is a homogeneous mixture.

Examples of homogeneous mixtures are air, salt dissolved in water and metal alloys such as brass, a mixture of copper and zinc.

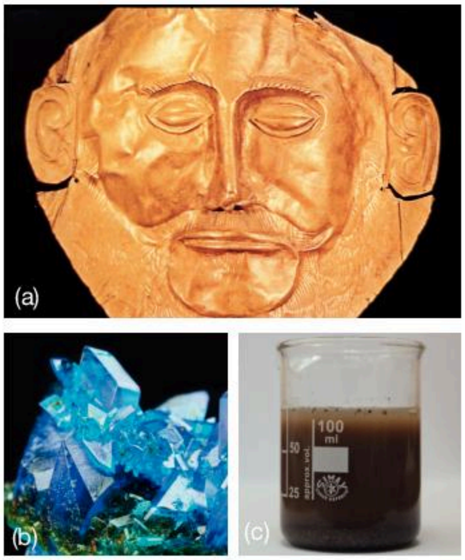
Heterogeneous mixtures

A **heterogeneous mixture** is a non-uniform mixture, for example, a mixture in which the component parts are in different states. The component parts can be distinguished from each other, although not always with the naked eye. Suspensions and colloids are heterogeneous mixtures.

Examples of heterogeneous mixtures are salt and sand, mayonnaise, and muddy water.

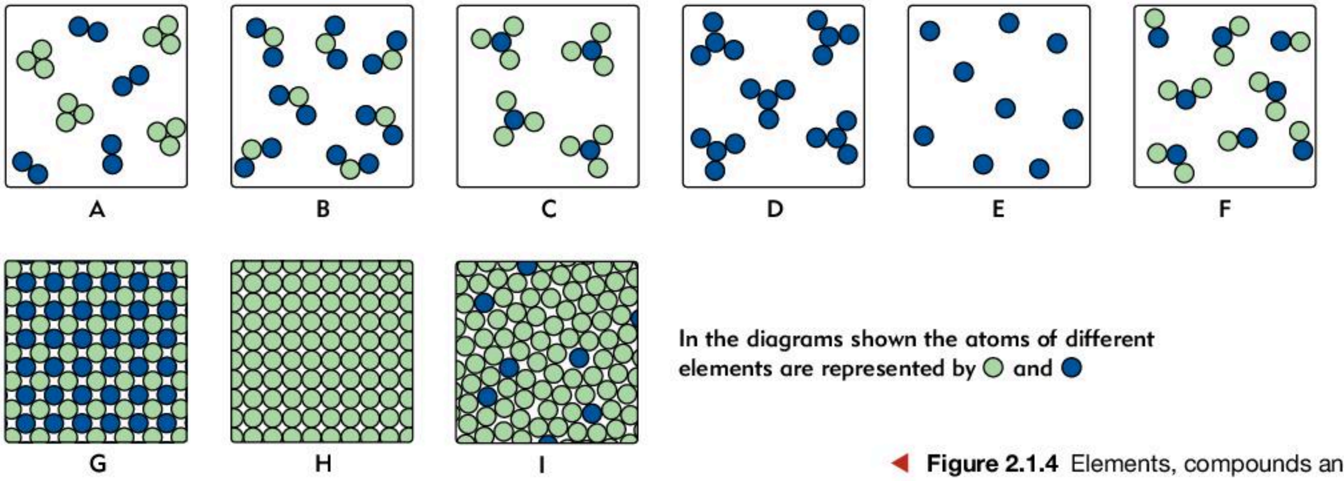
The diagrams in Figure 2.1.4 show how elements, compounds and mixtures can be distinguished by looking at the particles that make up the substance.

- If there is only one kind of atom, then it is an element.
- If there are two or more kinds of atoms joined together in the same ratio, then it is a compound.
- If there is a combination of two or more elements and/or compounds, then it is a mixture.



▲ **Figure 2.1.3** **(a)** Gold is an example of an element, **(b)** copper sulfate is an example of a compound and **(c)** muddy water is an example of a mixture.

In Figure 2.1.4, elements are represented in diagrams D, E and H, compounds are represented in diagrams B, C and G and mixtures are represented in diagrams A, F and I. I is a mixture and not a compound because the blue atoms are not arranged regularly among the green atoms.



In the diagrams shown the atoms of different elements are represented by 🟢 and 🔵

◀ **Figure 2.1.4** Elements, compounds and mixtures

Summary questions

1 Copper is described as an element. What does this tell us about copper?
2 What is a compound?
3 What are the differences between a pure substance and mixture?
4 What is the difference between a homogeneous mixture and a heterogeneous mixture?

A2.2 Solutions, suspensions and colloids

Solutions, suspensions and colloids form part of our everyday lives. For example, sea water is a solution, muddy water is a suspension and milk and fog are both colloids.

Solutions

A **solution** is a homogeneous mixture. The major component of a solution is known as the **solvent** and the minor component is known as the **solute**. Some solutions may contain more than one solute, e.g. sea water. The solute and solvent can be gases, liquids or solids. When a gas or a solid dissolves in a liquid, the liquid is always the solvent, e.g. in a mixture of salt in water, salt is the solute and water is the solvent.

Table 2.2.1 gives examples of various types of solutions.

▼ **Table 2.2.1** Examples of different types of solutions

Solute	Solvent	Example	Components of the solution
solid	liquid	sea water	sodium chloride in water
gas	liquid	soda water	carbon dioxide in water
solid	solid	brass	zinc in copper
liquid	liquid	white rum	ethanol in water
gas	gas	air	oxygen, water vapour, argon and carbon dioxide in nitrogen

Objectives

By the end of this topic you will be able to:

- explain the terms solvent, solute and solution
- identify different types of solutions
- distinguish between a solution, a suspension and a colloid
- describe some of the distinguishing properties of a solution, a suspension and a colloid
- give examples of solutions, suspensions and colloids.

! Key fact

A **solution** is a homogeneous mixture consisting of two or more components, one of which is usually a liquid.

▲ **Figure 2.2.1** **(a)** Soda water and **(b)** brass are examples of solutions

A saturated solution is one that contains as much solute as can be dissolved at a particular temperature in the presence of undissolved solute. You will study saturated solutions in more detail in Unit 2.3.

Suspensions

A suspension is a heterogeneous mixture containing minute particles which are visible to the naked eye. If left undisturbed, the particles in a suspension eventually settle. The components of a suspension can be separated by filtration.

▲ **Figure 2.2.2** Muddy water is an example of a suspension

Examples of suspensions

- Dust in air is a suspension of a solid in a gas.
- Powdered chalk in water is an example of a suspension of a solid in a liquid.
- Muddy water is another example of a solid suspended in a liquid.

Key fact

A **suspension** is a heterogeneous mixture in which minute, but visible, particles are dispersed in another substance, usually a liquid.

Colloids

A colloid is a heterogeneous mixture containing particles that are intermediate in size between those of a solution and those of a suspension. The particles in a colloid cannot be seen even with a microscope, and if left undisturbed they do not settle. The properties of a colloid are intermediate between those of a solution and those of a suspension.

Key fact

A **colloid** is a heterogeneous mixture in which the particles of one substance are dispersed in another substance, usually a liquid. The dispersed particles are smaller than those of a suspension, but larger than those of a solution.

Examples of colloids

- Smoke in air is a colloid of a solid in a gas, also known as a solid aerosol.
- Fog and aerosol sprays in air are colloids of a liquid dispersed in a gas, also known as liquid aerosols.
- Milk and mayonnaise are colloids of a liquid dispersed in a liquid, also known as emulsions.
- Gelatine and jelly are colloids of a solid dispersed in a liquid, also known as gels.

A comparison of the distinguishing properties of solutions, suspensions and colloids is given in Table 2.2.2.

Did you know?

1 nanometre = 10^{-6} millimetres
1 000 000 nanometres = 1 millimetre

▼ **Table 2.2.2** A comparison of the properties of solutions, colloids and suspensions

Property	Solution	Colloid	Suspension
Particle size	very small (less than one nanometre in diameter)	greater than that of a solution but they are not visible to the naked eye (between 1 and 1000 nanometres in diameter)	large enough so that they are visible to the naked eye (greater than 1000 nanometres in diameter)
Type of mixture	homogeneous	heterogeneous	heterogeneous
Appearance	generally transparent	usually opaque, some are translucent	opaque
Can the components be separated by filtration?	no	no	yes
Do the components separate out after the mixture has been standing for a while?	no	no	yes
Transmission of light	transmits light appearing transparent	will scatter light	does not transmit light; it is opaque

Practical activity

Comparing the properties of a solution, a suspension and a colloid

Your teacher may use this activity to assess:

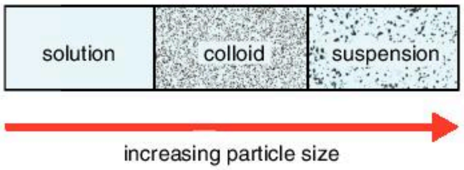
- observation, recording and reporting
- manipulation and measurement.

You will be supplied with two beakers, a filter funnel held in a retort stand, filter paper, distilled water, calcium hydroxide powder, copper(II) sulfate and gelatine.

Method

1. Half fill the beaker with water.
2. Place a large spatula full of calcium hydroxide powder into the water and mix vigorously.
3. Hold the mixture up to the light and look through it. Is it transparent or opaque?
4. Place the filter paper in the filter funnel and place the second beaker below it. Filter half the mixture. Could you separate the components by filtering?
5. Leave the other half of the mixture for two hours. Does the mixture settle out?
6. Using your answers from points **3**, **4** and **5**, decide whether the mixture is a solution, a suspension or a colloid.
7. Repeat the experiment with the copper(II) sulfate and gelatine, making sure that you stir each of the mixtures until there is no longer any solid remaining. You may need to heat the gelatine gently.
8. Record your findings in a table.

▲ **Figure 2.2.3** **(a)** Milk and **(b)** mayonnaise are examples of colloids



▲ **Figure 2.2.4** A comparison of the size of particles in solutions, colloids and suspensions

Summary questions

1. How does the particle size of:
   a. a solution compare with that of a suspension
   b. a suspension compare with that of a colloid?
2. Which of the following is true of a solution, a suspension or a colloid?
   a. does not transmit light
   b. can be separated by filtration
   c. will not settle out after standing?
3. Classify each of the following as a solution, a suspension or a colloid:
   a. mayonnaise
   b. chalk dust in water
   c. fog
   d. white vinegar.

Objectives

By the end of this topic you will be able to:

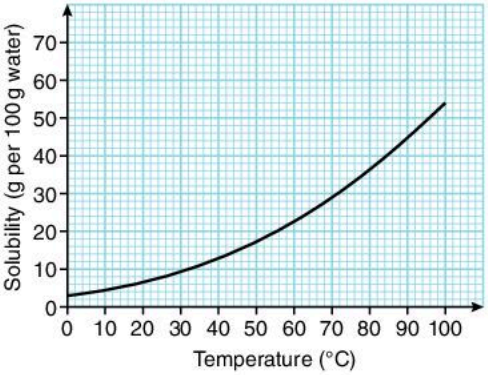
- define the term solubility
- explain what a saturated solution is
- explain the effect of temperature on solubility
- describe the solubility of a substance from a solubility curve
- perform calculations using a solubility curve.

Key fact

Solubility is the mass of solute that will saturate 100 g of solvent at a given temperature.

Key fact

A **saturated solution** is a solution which contains as much solute as can be dissolved at a given temperature, in the presence of undissolved solute.



▲ **Figure 2.3.1** Graph showing the solubility of potassium chlorate(v) against temperature

Exam tip

Before reading any graph you must ensure that you can interpret the **scale** of each axis. The scales of the two axes will not necessarily be the same.

A2.3 Solubility

The **solubility** of a solute is an indication of how much of the solute can dissolve in a fixed mass of solvent at a particular temperature. For example, we can find the solubility of sodium chloride in water by determining how much sodium chloride can dissolve in 10 cm^3 of water at a particular temperature. When no more solute can be dissolved in the solvent, the solution reaches saturation point and we say the solution is **saturated**. If any more solute is added to the solvent, the solute will remain in solid form and will be mixed with the saturated solution. The solubility of a solute in water is usually measured as the mass of solute that will saturate 100 g of water. Each solute and solvent combination has a specific solubility at a given temperature.

The solubility of a solute in a solvent is determined by the structure of the solute and solvent, and the temperature.

The effect of temperature on solubility

For most solid solutes in water, solubility *increases* as temperature increases. This means that as the temperature increases, a greater mass of solute will saturate a fixed mass of water. A solution which is saturated at one temperture will not be saturated at a higher temperature, and if a solution which is saturated at one temperature is cooled, crystals of the solute will form since less of the solute can dissolve at a lower temperature.

You learnt earlier in this unit that solubility in water is usually measured as the mass of solute which will saturate 100 g of water. The unit for solubility is **g per 100 g water**. However, it is not practical to weigh 100 g of water. In determining solubility of a solute in water in the laboratory, we make use of the fact that pure water has a density of 1 g cm^{-3}. In other words, 1 cm^3 of water has a mass of 1 g, or 1 g of water has a volume of 1 cm^3. This means that 100 g of water has a volume of 100 cm^3, and it is easy to measure 100 cm^3 of water in the laboratory.

When we plot solubility of a solute against temperature, we draw a graph known as a **solubility curve**. The solubility curve for potassium chlorate(v), $KClO_3$, is shown in Figure 2.3.1.

Looking at the solubility curve in Figure 2.3.1, you can see that the solubility of potassium chlorate(v) increases with an increase in temperature. Solubility curves such as this can be used to obtain various useful pieces of information, for example:

- The solubility of a solid at any temperature within the range of the graph.

Example

What is the solubility of potassium chlorate(v) at 78 °C?

Solubility of $KClO_3$ at 78 °C = **35 g per 100 g water**.

- The temperature at which crystals would just start to form if an unsaturated solution containing a fixed mass of solute is cooled to the point where it is just saturated.

Example

At what temperature would crystals just begin to form if an unsaturated solution of potassium chlorate(V) containing 20 g of potassium chlorate(V) dissolved in 100 g of water is cooled from 80 °C?

Temperature at which 20 g of $KClO_3$ saturates 100 g of water = 55 °C.

∴ temperature at which crystals just begin to form = **55 °C**.

- The mass of solute that would crystallise out of a saturated solution if its temperature is decreased.

Example

What mass of potassium chlorate(V) would crystallise out of a saturated solution containing 100 g of water when the temperature of the solution is decreased from 64 °C to 22 °C?

At 64 °C, **25 g** of $KClO_3$ saturates 100 g of water.

At 22 °C, **7 g** of $KClO_3$ saturates 100 g of water.

∴ mass of $KClO_3$ crystallising out of a saturated solution containing 100 g of water = 25 g − 7 g = **18 g**.

- The mass of solute to be added to resaturate a solution if its temperature is increased.

Example

What mass of potassium chlorate(V) must be added to resaturate a solution containing 250 g of water if the temperature of the saturated solution is increased from 32 °C to 82 °C?

At 32 °C, **10 g** of $KClO_3$ saturates 100 g of water.

At 82 °C, **38 g** of $KClO_3$ saturates 100 g of water.

∴ mass of $KClO_3$ to be added to resaturate a solution containing 100 g of water = 38 g − 10 g = **28 g**

and mass of $KClO_3$ to be added to resaturate a solution containing

250 g of water = $\frac{28}{100} \times 250\text{ g} = \mathbf{70\text{ g}}$.

- The minimum mass of water required to dissolve a fixed mass of solute at a given temperature.

Example

What is the minimum mass of water required to dissolve 40 g of potassium chlorate(V) at 74 °C?

At 74 °C, **32 g** of $KClO_3$ saturates 100 g of water.

∴ 40 g of $KClO_3$ saturate $\frac{100}{32} \times 40$ g of water = 125 g of water.

Minimum mass of water required = **125 g**.

Practical activity

Investigating the effect of temperature on solubility

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

You will be placed in groups by your teacher and each group will be supplied with a boiling tube, a thermometer, a beaker with water (for a water bath), a burette, a balance and potassium nitrate.

Method

Each group will determine the temperature at which a specific mass of potassium nitrate, varying from 6 g to 16 g, just saturates 10 g of water. Your teacher will tell you the mass that your group is to use. You will then share your results with the class.

1. Using the balance, weigh your group's mass of potassium nitrate.
2. Using the burette, place 10 cm^3 of distilled water in the boiling tube (this is equivalent to 10 g).
3. Add the potassium nitrate you weighed to the boiling tube and, using the thermometer, stir the solution carefully to dissolve as much potassium nitrate as possible.
4. Carefully heat the solution in the water bath, stirring it constantly until all the potassium nitrate has just dissolved.
5. Take the boiling tube out of the water bath and stir the solution gently.
6. Observe the solution as you stir it and as it cools. Look for signs of crystallisation, especially at the bottom of the tube as that is where it is easiest to see the crystals forming.
7. Note the temperature at which the crystals just start to form.
8. To ensure that your temperature is accurate, reheat the solution and cool it again and note the temperature at which the crystals just start to form.
9. Average your two temperature readings and record it in the table outlined on the board by your teacher (the table is given below).

Temperature (°C)						
Mass of KNO_3 saturating 10 g of water (g)	6.0	8.0	10.0	12.0	14.0	16.0
Solubility of KNO_3 (g per 100 g water)						

10. Once all the results from the class have been recorded, copy the table and calculate the mass of potassium nitrate that dissolved in 100 g of water at each temperature, i.e. the solubility of potassium nitrate at each temperature. Record these solubilities in your table.
11. On a sheet of graph paper, draw a solubility curve for potassium nitrate.

Summary questions

1. Define solubility.
2. What effect does temperature have on the solubility of solids in liquids?
3. What is the unit used for measuring solubility?

A2.4 Separating mixtures

Since mixtures form a part of our everyday lives, the separation of these mixtures into their component parts is important. Examples include the purification of drinking water and the making of filter coffee. If you were stuck on a desert island without any drinking water, your knowledge of separating mixtures could help you to make pure water from sea water.

Mixtures are a type of matter where the components can be separated by physical means. The method by which a mixture can be separated into its component parts is determined by the physical properties of these parts, e.g. particle size, boiling point, solubility. Table 2.4.1 summarises the methods of separation that you will be studying, together with the physical properties of the component parts that allows the mixtures to be separated by each method.

▼ **Table 2.4.1** A summary of the methods used to separate mixtures

Separation method	Physical properties of component parts
Filtration	A mixture of a solid and a liquid where the solid does not dissolve in the liquid. The components are separated based on their different particle sizes.
Evaporation and crystallisation	A mixture of a solid which is dissolved in a liquid where the boiling point of the liquid is lower than that of the solid. The methods only allow for collection of the solid. The components are separated based on their different boiling points.
Simple distillation	A mixture of a solid which is dissolved in a liquid where the boiling point of the liquid is lower than that of the solid. Both the liquid and the solid can be collected. The components are separated based on their different boiling points.
Fractional distillation	A mixture of two (or more) miscible liquids which have different boiling points, i.e. there is a difference in volatility. Miscible liquids are ones that mix completely. The components are separated based on their different boiling points.
Separating funnel	A mixture of two (or more) immiscible liquids which have different densities. Immiscible liquids are liquids which do not mix. The components are separated based on their different densities.
Chromatography	A mixture of dissolved substances which will travel through a material. The components are separated based on their different solubilities in a solvent and attraction to the material.

Objectives

By the end of this topic you will be able to:

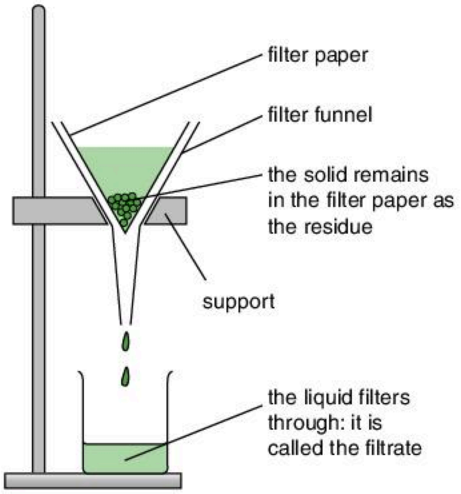
- identify suitable techniques to separate the components of mixtures
- describe the types of mixtures separated by filtration, evaporation, crystallisation, simple distillation, fractional distillation, a separating funnel and chromatography
- describe the processes involved in separating mixtures by filtration, evaporation, crystallisation, simple distillation, fractional distillation, a separating funnel and chromatography
- explain the terms filtrate, residue and distillate
- explain the difference between miscible and immiscible liquids.

Filtration

Filtration is used to separate a mixture of a solid and a liquid where the solid does not dissolve in the liquid, i.e. it is used to separate the components of a suspension. Filtration makes use of filter paper to separate the solid from the liquid. The filter paper works like a sieve, keeping back the solid particles. The solid particles are too big to pass through the filter paper, whereas the liquid particles are small enough to pass through.

The part of the mixture that stays behind in the filter paper is called the residue. The part of the mixture that passes through the filter paper is called the filtrate.

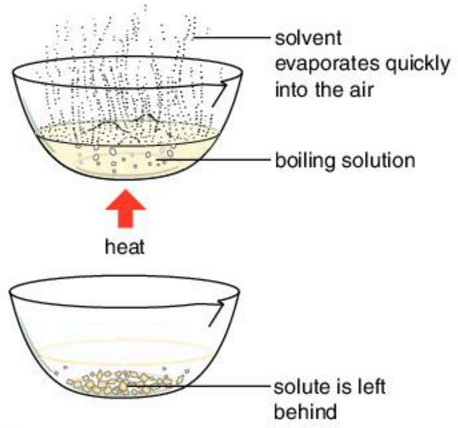
Filtration is one of the steps used in the purification of drinking water. Most simple water purification devices that you might use in your homes use filtration as their main method of purifying the water. They separate the solid impurities from the water. A coffee machine makes use of a filter in the form of filter paper to separate the coffee granules from the coffee.



▲ **Figure 2.4.1** Filtration apparatus

Evaporation

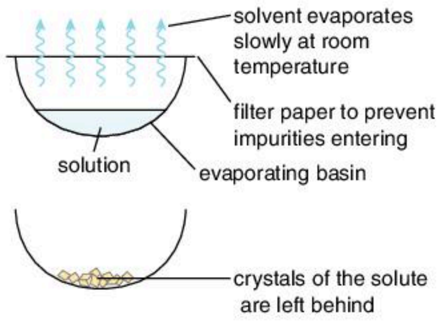
Evaporation is used to separate a solution of a solid dissolved in a liquid, but it only allows the solid to be collected. During evaporation, the solution is boiled allowing the liquid to vaporise into the air. The solute is left behind in the container. Evaporation is a fairly rapid process and if all the liquid evaporates, the solid remaining lacks any crystalline structure. The method is not suitable if the solid to be collected is decomposed by heat. Evaporation is useful to obtain the sodium chloride from a sodium chloride solution.



▲ **Figure 2.4.2** Evaporation

Crystallisation

Like evaporation, **crystallisation** is used to separate a solution of a solid dissolved in a liquid, and it only allows the solid to be collected. Unlike evaporation, the solution is not boiled, it is left in a container at room temperature for the liquid to vaporise into the air. Crystallisation is a slow process and the solid remaining has a distinct crystalline structure. It is used if a hydrated solid containing **water of crystallisation** is required. Water of crystallisation is water that is incorporated within the crystalline structure and you will learn more about this in Unit 8.4. Crystallisation is useful to obtain hydrated copper(II) sulfate crystals ($CuSO_4.5H_2O$) from a copper(II) sulfate solution.



▲ **Figure 2.4.3** Crystallisation

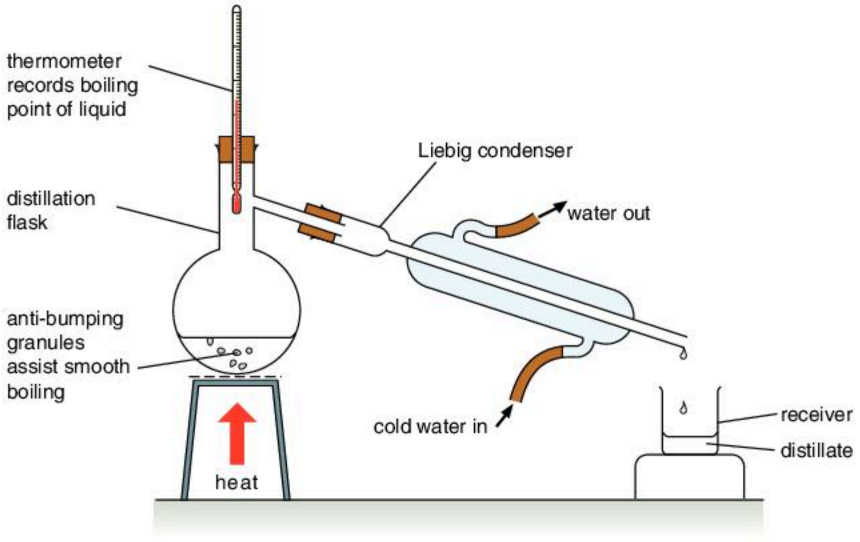
Simple distillation

Simple distillation is also used to separate a solution of a solid dissolved in a liquid. It allows both the solid and the liquid to be collected. Simple distillation is an appropriate method of separation only if the liquid has a lower boiling point than the solid, i.e. the liquid becomes a vapour before the solid. One of the key components of the apparatus used in distillation is a Liebig condenser.

The method by which the distillation apparatus works is as follows:

- The mixture is heated in the distillation flask until it boils and vaporisation occurs.
- The vapour rises up the distillation flask and as it passes into the Liebig condenser it cools and condenses back to a liquid. The liquid passes down the condenser and is collected as the **distillate**.
- The concentration of the solution in the distillation flask gradually increases and when most of the liquid has vaporised, the solution can be poured into an evaporating basin and left to crystallise to obtain crystals of the solute if required.

▶ **Figure 2.4.4** Apparatus for simple distillation



The thermometer is used to monitor the temperature of the vapour entering the condenser. If the temperature remains constant at the boiling point of the pure liquid, then the distillate is pure.

Distillation is a very useful method of separation if you need to collect both the solid and the liquid after separation. However, it is also used to make distilled water from tap water and can be used to obtain pure water from sea water. In both these cases, any solution remaining in the distillation flask is discarded since it is not required.

Exam tip

It is very important that you are able to draw line diagrams of the **apparatus** used in the different separation processes.

Fractional distillation

Fractional distillation is used to separate a mixture of two or more liquids which are miscible and have different boiling points that are close together. **Miscible** liquids are liquids that are able to mix. The apparatus used in fractional distillation is similar to that used in simple distillation. However, a **fractionating column** is attached between a round-bottomed flask and the Liebig condenser. Fractional distillation can be used to separate a solution of ethanol and water since the boiling point of ethanol is 78 °C and that of water is 100 °C.

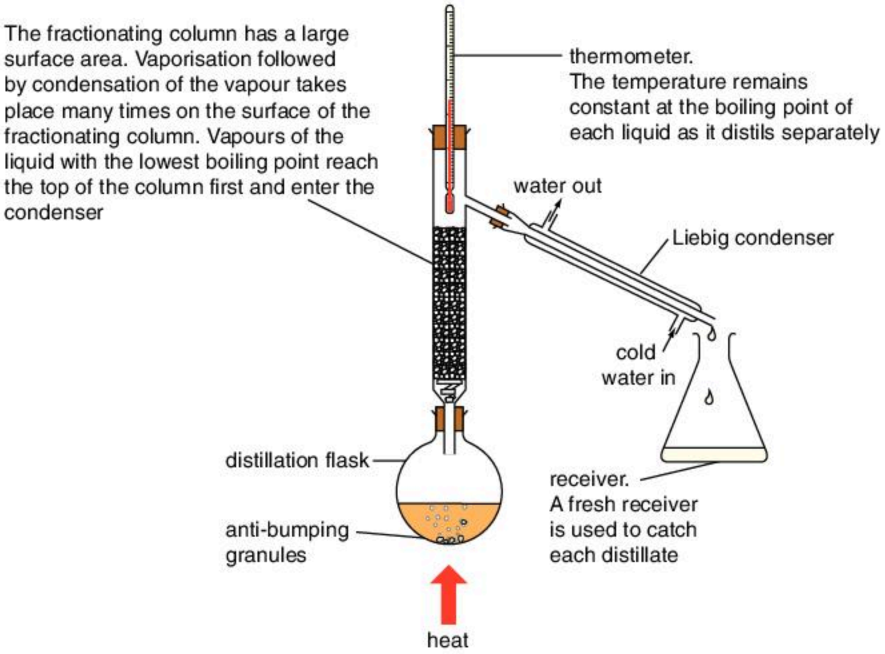


Figure 2.4.5 Fractional distillation using a fractionating column

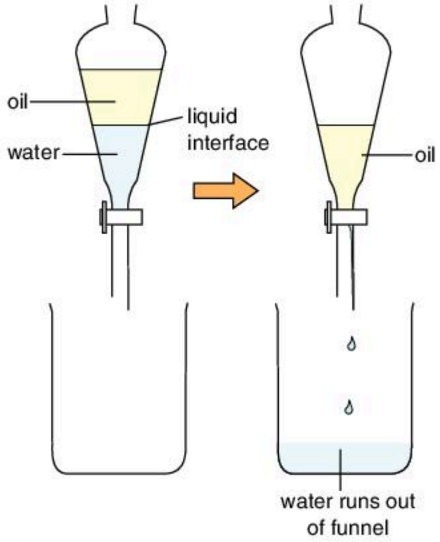
The method by which the apparatus for fractional distillation works is as follows.

- The mixture boils and vapours of both liquids enter and move up the fractionating column where they condense and vaporise many times.
- As the mixture of vapours moves up the fractionating column, it becomes increasingly richer in the more volatile component, i.e. the one with the lower boiling point (ethanol), until the vapour reaching the top of the column consists only of the more volatile component. This vapour passes into the condenser, condenses and is collected as the **distillate**.
- The vapour of the less volatile liquid, i.e. the one with the higher boiling point (water), condenses in the fractionating column and returns to the round-bottomed flask.

Did you know?

The process of **fractional distillation** is used in several industries that you might have come across. It is used in the rum industry to separate the ethanol from the fermentation mixture (Unit 15.1), and in the petroleum industry to separate crude oil into different fractions such as gasoline, kerosene and diesel (Unit 14.1). Also liquid air can be fractionally distilled into oxygen and nitrogen gases which have various uses (Unit 20).

- When almost all of the more volatile liquid has distilled over, the temperature begins to rise, showing that a mixture of both vapours is reaching the top of the column and distilling over. This is collected in a second container and discarded.
- Once the temperature reaches the boiling point of the less volatile liquid, that liquid is then distilled into a third container.



▲ **Figure 2.4.6** Separation using a separating funnel

Separating funnel

A separating funnel is used to separate a mixture of liquids that are immiscible and have different densities. Immiscible liquids are liquids which are unable to mix. A separating funnel is a container that has a tap at the bottom, allowing one liquid to be drained off before the other. It is usually conical in shape to reduce the amount of liquids lost at the interface of the two liquids. Oil and water are two liquids that are immiscible and they have different densities. If a mixture of oil and water is placed in a separating funnel, the oil with a lower density floats on the water which has a higher density.

The method by which the separating funnel works is as follows:

- The tap is opened to allow the liquid with the higher density (water) to run into the container below.
- The tap is closed as the liquid interface almost reaches it and the first container is replaced with a second.
- The tap is opened again to allow a very small amount of the liquid with the lower density (oil) to run into the container and then closed. The contents of the second container are discarded.

Paper chromatography

Paper chromatography is used to separate a mixture of dissolved substances which are coloured or can be coloured, and which will travel through a material, e.g. filter paper. The substances are separated based on:

- how soluble they are in the solvent used
- how strongly they are attracted to the paper.

Many inks and food colourings are mixtures of two or more dyes, which can be separated by paper chromatography.

The method by which paper chromatography works is as follows:

- A drop of the dye mixture is placed 1 cm from the bottom of a rectangular piece of absorbent paper, e.g. filter paper. The paper is then hung in a beaker containing solvent so that the lower edge is below the surface of the solvent and the dye is above.
- The solvent moves up the paper and dissolves the dyes in the mixture, carrying them with it. However, the different dyes travel up the paper at different rates.
- The dyes that are the most soluble in the solvent and least attracted to the paper travel the fastest and the furthest distance.
- The dyes that are the least soluble in the solvent and most attracted to the paper travel the slowest and the least distance.
- Once the solvent has completed its movement up the paper, the paper is allowed to dry. There will be a pattern of different coloured dyes on the paper, each one representing a part of the mixture. This pattern is known as a chromatogram.

? Did you know?

Chromatography is used extensively by forensic scientists in their work. For example, they use it to analyse body fluids for the presence of illegal drugs, to analyse blood from crime scenes, and, at airports, to detect residues from explosives. If the spots that form are colourless, a locating agent is sprayed on the chromatogram. This reacts with the colourless spots, making them coloured.

Practical activity

Separating mixtures using chromatography

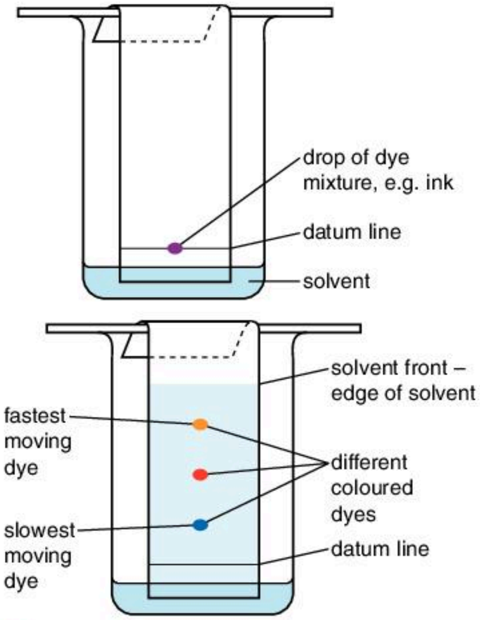
Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

You will be supplied with a piece of filter paper, a beaker, a capillary tube, two water-based markers of different colours and food colouring.

Method

1. Measure the depth of the beaker with a ruler. The depth will be the length of your filter paper.
2. Cut the filter paper into a rectangular strip, 4 cm wide and the length as measured in point **1** above.
3. Draw a pencil line across the strip 2 cm up from the bottom and mark the line at 1 cm intervals with a pencil dot.
4. Attach the top of the strip to a glass rod so that it can be suspended in the beaker with its lower edge just above the bottom of the beaker.
5. Make a dot with one of the markers on the first pencil dot. Make a dot with the second marker on the second pencil dot. Use the capillary tube to place a drop of food colouring on the third pencil dot.
6. Place enough water into the beaker so that when the strip of filter paper is hung in the beaker, the bottom of the strip will be immersed in the water, but the coloured dots will be above the water.
7. Hang the strip of filter paper in the beaker so that the water wets the bottom of the paper and watch the colours in the three dots separate and spread up the filter paper as the water rises up the paper.
8. Compare the results from your two markers. Did they contain any colours in common?
9. If members of the class used different coloured markers from you, compare your results with theirs to see if different coloured markers had any colours in common with yours.



▲ **Figure 2.4.7** Separation using paper chromatography

Did you know?

Using chromatography it can be shown that the green pigment in plants, known as **chlorophyll**, is actually made of three different coloured pigments, orange, yellow and green.

Summary questions

1. Which properties of the components of a mixture are considered when deciding on a separation technique?
2. What is the main difference between using distillation and evaporation as a separation method?
3. For each of the following mixtures, explain how you would separate them into their component parts:
   a. a mixture of salt, black pepper and water
   b. a mixture of oil, sugar and water.

A2.5 Extraction of sucrose from sugar cane

Objectives

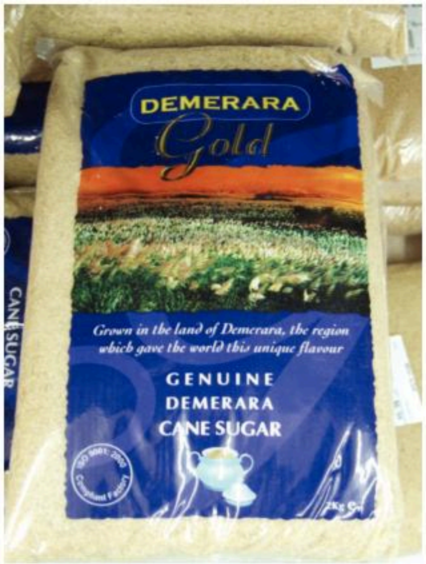
By the end of this topic you will be able to:

- describe the extraction of sucrose from sugar cane
- draw a flow diagram showing the steps which are involved in the separation of sucrose from sugar cane.

Sugar cane was first introduced to the Caribbean by the Dutch in about 1625 and has been a very important part of its economy ever since. The production of sucrose from the sugar cane plant is an industrial process that makes use of several separation techniques.

The processes involved in the separation of sucrose from sugar cane are as follows:

- The sugar cane stalks are harvested, transported to the factory, cleaned and cut into small pieces by revolving knives in the shredder.
- The pieces are then **crushed** in the crushers as water is sprayed on them to dissolve the sugar present. This produces cane juice and cane fibre, or bagasse. The bagasse is taken to the boiler furnace where it is burnt to supply heat for the boilers.
- The cane juice, which is acidic and contains impurities, enters the clarifier where **precipitation** occurs. The juice is heated and calcium hydroxide is added which neutralises any acids in the juice and causes the impurities to precipitate out, i.e. they are converted into larger, insoluble particles.
- The juice then moves into the rotary filter where continuous **filtration** takes place to remove the insoluble impurities. This produces factory mud and clarified juice. The factory mud is returned to the fields.
- The clarified juice, which is about 85% water, goes into a series of three or four boilers or evaporators where **vacuum distillation** occurs. These boilers are under successively lower pressures so that as the juice passes from one to the next it boils at successively lower temperatures. In this way the water evaporates and the juice is concentrated but not charred or caramelised by the boiling process. The juice from the last boiler is a thick syrup containing about 35% water.

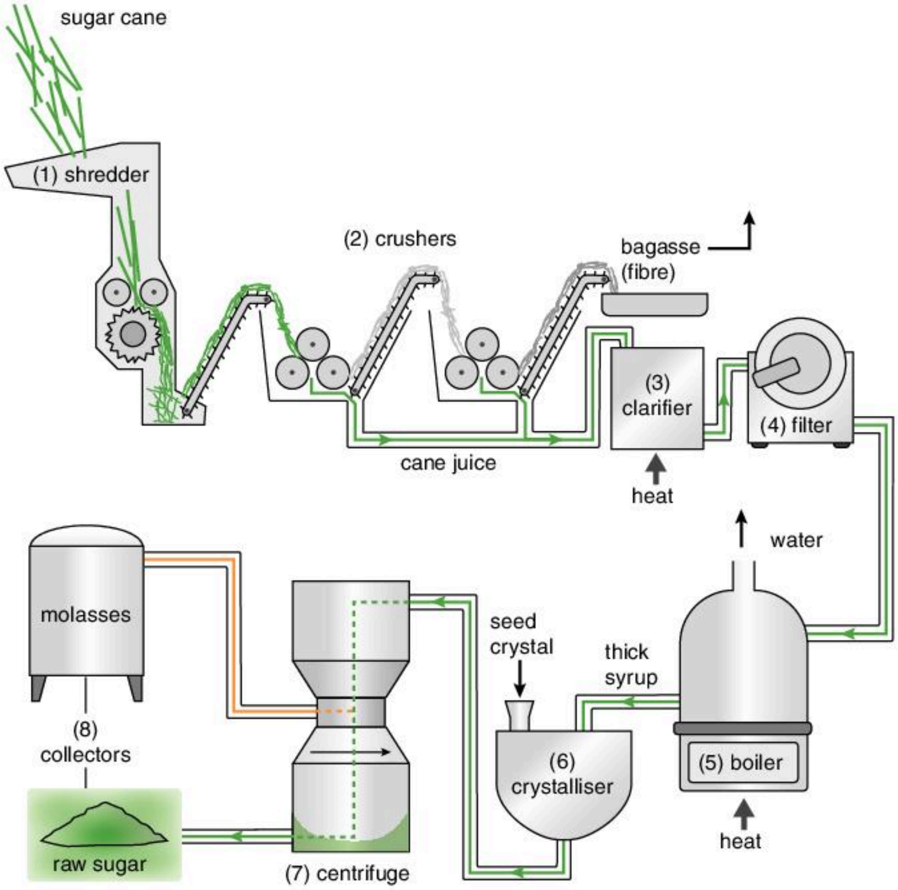


▲ **Figure 2.5.1** Sugar

▲ **Figure 2.5.2** Sugar cane in flower

▲ **Figure 2.5.3** Andrew's Sugar Factory in Barbados

- The thick syrup moves into the crystalliser where **crystallisation** takes place. Here the syrup is evaporated until it is saturated with sugar. As soon as the saturation point is exceeded, small grains of sugar, called 'seed', are added to serve as nuclei for the formation of sugar crystals. As the crystals form, the remaining syrup becomes thick and viscous and is called molasses. The mixture of crystals and molasses forms massecuite.
- The sugar crystals and molasses in the massecuite are then separated by **centrifugation** in the centrifuges. Each centrifuge contains a perforated basket. The massecuite is placed in the basket and this revolves at high speed. The molasses are forced out through the holes in the basket and are collected in the outer drum of the centrifuge. The sugar crystals remain behind in the basket.
- The damp, unrefined sugar crystals are collected and dried by being tumbled through heated air.



▲ **Figure 2.5.4** Flow diagram of the various stages in the extraction of sucrose

Summary questions

1 Draw a simple flow diagram of the steps which are involved in the separation of sucrose from sugar cane.

2 Why is it important to use a series of boilers in the extraction of sucrose from sugar cane?

Key concepts

- Matter can be classified into **pure substances** and **mixtures**.
- A pure substance is composed of only one type of material. It has a fixed, constant composition, its properties are fixed and constant and the component parts cannot be separated by physical means.
- Pure substances can be classified into **elements** and **compounds**.
- An element is a pure substance that cannot be broken down into any simpler substances by any ordinary chemical or physical means.
- A compound is a pure substance that contains two or more different types of element which are bonded together chemically in fixed proportions and in such a way that their properties have changed.
- A mixture consists of two or more substances combined together in varying proportions. Each component retains its own independent properties and has undergone no chemical reaction with any other substance in the mixture. The components of mixtures can be separated by physical means.
- Mixtures can be classified into **homogeneous mixtures** and **heterogeneous mixtures**.
- A homogeneous mixture is a mixture in which the properties and composition are uniform throughout the sample and the components cannot be distinguished from each other.
- A **solution** is a homogeneous mixture consisting of two or more components, one of which is usually a liquid. The particles in a solution are extremely small.
- A heterogeneous mixture is a non-uniform mixture.
- A **suspension** is a heterogeneous mixture in which minute, but visible, particles are dispersed in another substance, usually a liquid.
- A **colloid** is a heterogeneous mixture in which the particles of one substance are dispersed in another substance, usually a liquid. The dispersed particles are smaller than those of a suspension, but larger than those of a solution.
- **Solubility** is a measure of the mass of solute which will saturate a fixed mass of solvent at a given temperature. The unit of solubility is **g per 100 g water**.
- A **saturated solution** is a solution that contains as much solute as can be dissolved at a given temperature, in the presence of undissolved solute.
- Temperature also affects solubility. The solubility of most solids in water increases as temperature increases.
- A **solubility curve** is a graph showing how the solubility of a solute varies with temperature.
- The physical properties of the component parts of a mixture are used to determine the method by which a mixture can be separated.
- Possible separation methods are filtration, evaporation, crystallisation, simple distillation, fractional distillation, a separating funnel and chromatography.
- The extraction of sucrose from sugar cane involves the following processes: crushing, precipitation, filtration, vacuum distillation, crystallisation and centrifugation.

Practice exam-style questions

Multiple-choice questions

1 Which of the following is correct?
- **A** The components of a mixture can be separated by physical means.
- **B** A mixture contains two or more pure substances chemically combined together.
- **C** A mixture has a fixed melting point.
- **D** The component parts of a mixture are in a fixed ratio.

2 The solubility of a solid in water usually:
- **A** increases as temperature increases
- **B** increases as temperature decreases
- **C** remains constant as temperature increases
- **D** decreases as temperature increases

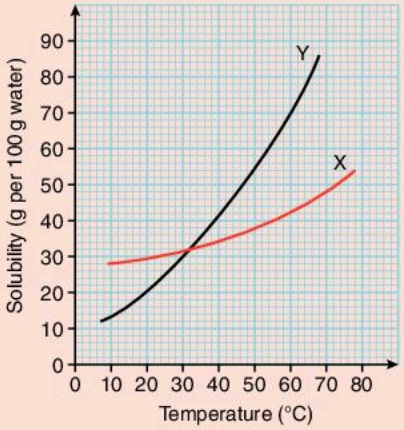
3 The following are all colloids except:
- **A** fog
- **B** air
- **C** milk
- **D** mayonnaise

4 The process of fractional distillation depends on a difference in the:
- **A** melting points of the components of the mixture
- **B** densities of the components of the mixture
- **C** solubilities of the components of the mixture in water
- **D** boiling points of the components of the mixture

5 The best way to separate sodium chloride solution and sand is by:
- **A** evaporation
- **B** carefully pouring off the sodium chloride solution
- **C** filtering the mixture
- **D** simple distillation

Structured question

6 a Figure 1 shows the solubility curves for two ionic solids, X and Y. Use this information to answer the following questions.



▲ **Figure 1** Solubility curves for ionic substances X and Y

i) Define the term 'solubility'. *(2 marks)*

ii) Temperature has the greatest effect on the solubility of which compound? *(1 mark)*

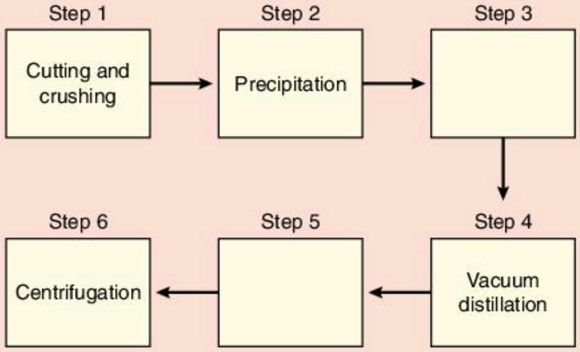
iii) Explain how you arrived at your answer to part **ii)** above. *(2 marks)*

iv) At what temperature are the solubilities of X and Y the same? *(1 mark)*

v) What is the solubility of X at 46 °C? *(1 mark)*

vi) What mass of Y would crystallise out of a saturated solution containing 100 g of water if this solution is cooled from 62 °C to 24 °C? *(3 marks)*

b Figure 2 shows the main steps in the extraction of sucrose from sugar cane.



▲ **Figure 2** Flow chart showing the extraction of sucrose from sugar cane

i) Complete the chart by labelling the boxes in steps 3 and 5. *(2 marks)*

ii) Describe the process taking place in step 2. *(2 marks)*

iii) Why is it necessary to have the boilers under successively lower pressures in step 4? *(1 mark)*

Total 15 marks

Extended response question

7 Chemical substances can combine in various ways to form mixtures, the components of which can be separated by various means.

a What do you understand by the term 'mixture'? *(2 marks)*

b Solutions, suspensions and colloids are examples of mixtures. Explain the differences between these three mixtures in terms of particle size, sedimentation when left to stand and transmission of light. *(6 marks)*

c You are given a mixture of three solid chlorides: sodium chloride, lead chloride and silver chloride. The solubilities of these three chlorides in water are given in the table below.

Compound	Solubility in cold water	Solubility in hot water
sodium chloride	soluble	soluble
lead chloride	insoluble	soluble
silver chloride	insoluble	insoluble

i) Use the information given in the table to devise a scheme to separate the mixture into its component parts. *(5 marks)*

ii) Draw a labelled diagram to show one of the pieces of apparatus you used in your separation scheme. *(2 marks)*

Total 15 marks

A3 Atomic structure

Objectives

By the end of this topic you will be able to:

- give the atomic symbols of the common elements
- describe the structure of an atom
- give the relative charge and mass of a proton, a neutron and an electron
- define the terms atomic number and mass number
- represent atoms by nuclear notation
- interpret nuclear notation to determine the number of subatomic particles in an atom.

Key fact

An **atom** is the smallest component of an element that can exist and still have the same chemical properties as the element.

Exam tip

It is essential that you know the **atomic symbols** of all the common elements and that you write them with care, making sure that you form your capital and common letters correctly.

For a long time, people believed that all matter consisted of the four basic elements: earth, fire, air and water. Scientists now know that the smallest identifying particle of matter is an atom. Atoms are the basic building blocks of matter. On a macroscopic scale we classified matter into mixtures and pure substances; we are now going to investigate matter on a microscopic scale by looking at the structure of the atom.

A3.1 The structure of atoms

Regardless of whether you have one kilogram, one gram or one atom of a particular element, all these masses will exhibit the same properties. For example, if one gram of gold is broken into smaller parts, it will still be gold. The smallest identifying particle of an element is an **atom**.

Elements can be classified into **metals** and **non-metals**. The properties of metals and non-metals are given in Units 4.4, 17.1 and 20.1. Each element can be represented by an **atomic symbol**, which represents one atom of the element. The atomic symbols of some of the commonest elements are given in Tables 3.1.1 and 3.1.2.

▼ **Table 3.1.1** The atomic symbols of the common metals

Metals			
Element	**Atomic symbol**	**Element**	**Atomic symbol**
Aluminium	Al	Lithium	Li
Barium	Ba	Magnesium	Mg
Beryllium	Be	Manganese	Mn
Calcium	Ca	Mercury	Hg
Chromium	Cr	Nickel	Ni
Cobalt	Co	Potassium	K
Copper	Cu	Silver	Ag
Gold	Au	Sodium	Na
Iron	Fe	Tin	Sn
Lead	Pb	Zinc	Zn

▼ **Table 3.1.2** The atomic symbols of the common non-metals

Non-metals			
Element	**Atomic symbol**	**Element**	**Atomic symbol**
Argon	Ar	Neon	Ne
Boron	B	Nitrogen	N
Bromine	Br	Oxygen	O
Carbon	C	Phosphorus	P
Chlorine	Cl	Silicon	Si
Fluorine	F	Sulfur	S
Helium	He		
Hydrogen	H		
Iodine	I		
Krypton	Kr		

We cannot see atoms because they are incredibly small. Compared to the thickness of a human hair, an atom is more than a million times smaller. The average diameter of an atom is about 2.5×10^{-7} mm, which means that, on average, an atom is 25 million times smaller than a millimetre. However, there are particles that are even smaller than an atom. These fundamental particles are called subatomic particles and atoms are made up of these particles.

Subatomic particles

Atoms are made up of three subatomic particles: protons, neutrons and electrons.

- Protons and neutrons are found in the centre of an atom in an area known as the nucleus. Protons and neutrons are collectively known as nucleons. Nearly all the mass of an atom is concentrated in its nucleus.
- Electrons are found at quite a distance from the nucleus, spinning around the nucleus in a series of levels known as energy shells (sometimes just called shells). Most of the atom is in fact empty space, but the movement of the electrons around the nucleus constitutes the volume of an atom. When chemical reactions occur, it is the electrons that are involved in the reaction.

Mass and charge of subatomic particles

Protons are *positively* charged particles and electrons are *negatively* charged. The size of the charge on a proton is equal to the size of the charge on an electron. Neutrons are electrically *neutral* so have no charge. In an atom, the number of protons is always the same as the number of electrons. Therefore, an atom is electrically *neutral* and has no charge.

The mass of a proton is equal to the mass of a neutron. The mass of a proton is extremely small, only 1.67×10^{-24} g. The mass of an electron is even smaller. Comparing the mass of a proton to that of an electron, a proton is about 1836 times heavier than an electron.

The properties of a proton, neutron and electron are summarised in Table 3.1.3.

▼ **Table 3.1.3** The properties of a proton, a neutron and an electron

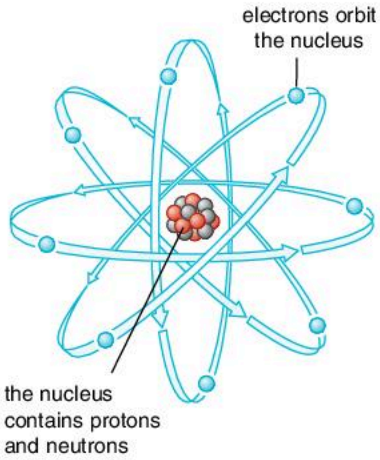
Subatomic particle	Relative charge	Relative mass
Proton	+1	1
Neutron	0	1
Electron	−1	$\frac{1}{1836}$

Atomic number

The number of protons in an atom is known as the atomic number and is unique to a particular element, i.e. there are no two elements in the world that have the same atomic number. For example, the atomic number of hydrogen is 1, which means that every hydrogen atom has 1 proton. The atomic number of calcium is 20, therefore, every calcium atom has 20 protons.

Did you know?

The **atomic symbols** of some of the elements are derived from the Latin name of the element. For example, the Latin name for iron is ferrum and the symbol for iron is Fe, copper is cuprum and its symbol is Cu, lead is plumbum and its symbol is Pb, and silver is argentum, symbol Ag.



▲ **Figure 3.1.1** The structure of an atom

Key fact

Atomic number is the number of protons in the nucleus of one atom of an element.

In an atom, since the number of electrons is always equal to the number of protons, the number of electrons is equal to the atomic number.

Key fact

Mass number is the total number of protons and neutrons in the nucleus of one atom of an element.

Mass number

The number of protons and neutrons in an atom is known as the **mass number**. The mass number is not unique to a particular element, for example polonium and astatine both have a mass number of 210.

The number of neutrons can be calculated by subtracting the atomic number from the mass number, i.e.

$$\text{number of neutrons} = \text{mass number} - \text{atomic number}$$

Nuclear notation

We can represent an atom (or ion) of an element using the **nuclear notation**:

$$^{A}_{Z}\mathrm{X}$$

where X = atomic symbol, A = mass number and Z = atomic number.

This notation is very useful because it allows us to calculate the number of protons, neutrons and electrons in an atom.

- The number of protons is equal to Z.
- The number of neutrons is equal to $A - Z$.
- The number of electrons is equal to the number of protons, which is equal to Z.

Examples are given in Table 3.1.4.

Table 3.1.4 The number of subatomic particles in the atoms of three elements

Element	Nuclear notation	Number of		
		Protons (Z)	Neutrons (A–Z)	Electrons
Nitrogen	$^{14}_{7}\mathrm{N}$	7	7	7
Aluminium	$^{27}_{13}\mathrm{Al}$	13	14	13
Chlorine	$^{37}_{17}\mathrm{Cl}$	17	20	17

The nuclear notations for all elements can be found in the periodic table of elements on page 360.

Summary questions

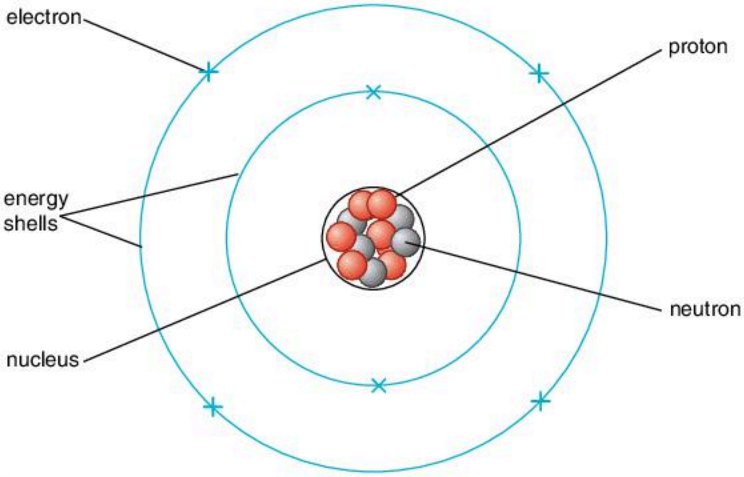
1 Which particle(s) in the atom is/are responsible for the:
  a mass of the atom
  b volume of the atom?

2 Define the following terms:
  atomic number, mass number.

3 For each of the following nuclear notations, give the number of protons, electrons and neutrons.
  a $^{11}_{5}\mathrm{B}$
  b $^{23}_{11}\mathrm{Na}$
  c $^{40}_{20}\mathrm{Ca}$

A3.2 The electronic configuration of an atom

Electrons spin around the nucleus of an atom in a series of levels known as energy shells. Each energy shell is a fixed distance from the nucleus of the atom and the electrons in the same shell have the same amount of energy. The electrons are arranged in the energy shells in a specific way. We can represent the arrangement of electrons in an atom, known as the electronic configuration (or structure), using another scientific model.



▲ **Figure 3.2.1** Model of the arrangement of electrons in an atom

In this scientific model, the energy shells are drawn as concentric circles around the nucleus, but as this is only a model, this is not how they actually look. The shells are numbered according to their distance from the nucleus. The shell closest to the nucleus is assigned number 1, the next furthest away is number 2, and so on.

Each shell can hold up to a certain maximum number of electrons:

- shell number 1 can hold a maximum of **2** electrons
- shell number 2 can hold a maximum of **8** electrons
- shell number 3 may be considered to hold a maximum of **8** electrons.

Further shells hold more electrons, but it is not necessary to know their maximum numbers since you will not have to work out the electron configuration of an atom with more than 20 electrons.

The negative electrons are constantly being attracted to the positive nucleus by an electrostatic force of attraction. The electrons maintain their spin around the nucleus and their distance from the nucleus because of the energy they possess. Electrons in the shells closest to the nucleus have less energy than those in the shells further from the nucleus. In fact, electrons can only occupy a specific shell if they have the required energy.

The chemical properties of an element depend on the arrangement of electrons in an atom of that element, in particular, the number of electrons in the outermost energy shell.

Objectives

By the end of this topic you will be able to:

- give the maximum number of electrons allowed in the first three energy shells of an atom
- work out the electronic configuration of the first 20 elements
- represent the electronic configuration of an atom in writing
- represent the electronic configuration of an atom by a shell diagram
- explain the term valence electron.

? Did you know?

The maximum number of electrons that an energy shell can hold is given by the formula $2n^2$, where n is the shell number**.** Shell number 3 can, therefore, actually hold a maximum of 18 electrons and shell number 4 can hold a maximum of 32 electrons. However, when filling up shell number 3, something unusual happens. When 8 electrons have been placed in shell number 3, the next 2 electrons are placed in shell number 4. Only after these 2 electrons have been placed in shell number 4 are any remaining electrons (up to 10) placed in shell number 3.

! Key fact

An **electrostatic force of attraction** is the force that exists between two oppositely charged particles.

Shell diagrams

The electronic configuration of an atom of an element can be represented by drawing a shell diagram or in writing using numbers. In order to draw a shell diagram of an atom, the following steps must be followed.

1) Determine the number of protons, neutrons and electrons in the atom using the atomic number and mass number (Unit 3.1).
2) Draw a small circle for the nucleus and write the number of protons and neutrons inside this circle.
3) Work out the electronic configuration by filling up the shells with electrons, in order, i.e. fill up shell 1 first and, when that is full, start to fill up shell 2, and so on.
4) Represent the electronic configuration by drawing concentric circles around the nucleus. The number of circles you should draw is determined by the number of shells that you are going to fill or partially fill. Use crosses or dots to represent the electrons. Draw the 2 electrons in shell number 1 as a pair. Draw the electrons in the other shells as pairs when the shell contains 5 or more electrons.

Key fact

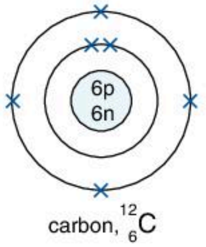
The **electronic configuration** of an atom is the arrangement of electrons in the energy shells of the atom.

Examples

1 The structure of a carbon atom, $^{12}_{6}C$.

A carbon atom has:

6 protons
6 neutrons (12 − 6)
6 electrons.



▲ **Figure 3.2.2** Electronic configuration of a carbon atom

▼ **Table 3.2.1** Determining the electronic configuration of a carbon atom

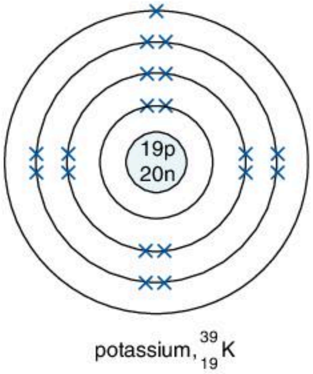
Shell number	Maximum number of electrons	Number of electrons in carbon
1	2	2
2	8	4

The electronic configuration of a carbon atom can be written by writing the symbol of the element followed by the number of electrons in each shell separated by commas, i.e. **C (2,4)**.

2 The structure of a potassium atom, $^{39}_{19}K$.

A potassium atom has:

19 protons
20 neutrons (39 − 19)
19 electrons.



▲ **Figure 3.2.3** Electronic configuration of a potassium atom

▼ **Table 3.2.2** Determining the electronic configuration of a potassium atom

Shell number	Maximum number of electrons	Number of electrons in potassium
1	2	2
2	8	8
3	8	8
4	not necessary	1

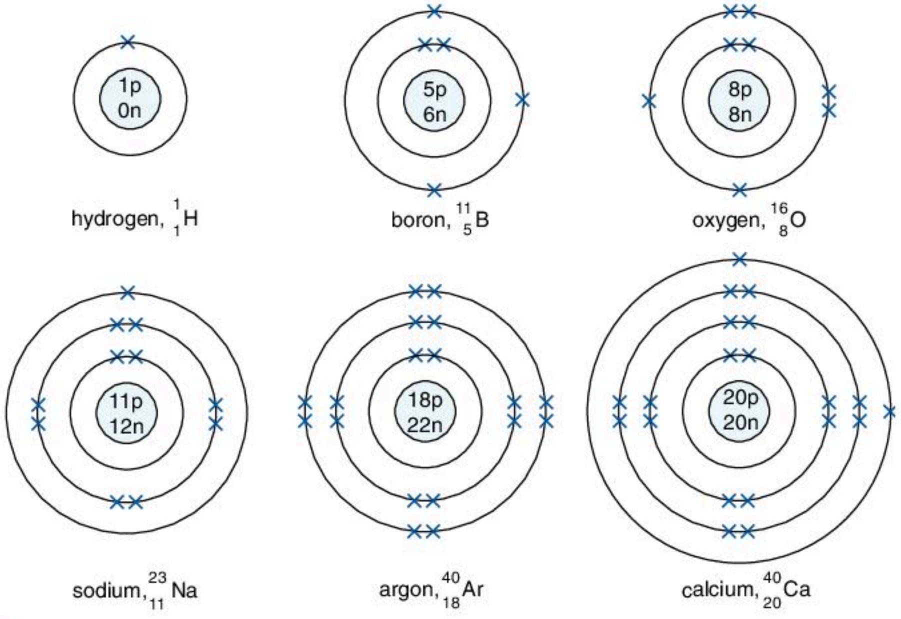
The electronic configuration of a potassium atom is **K (2,8,8,1)**.

3 The structure of other atoms.

Examples of the electronic configurations of atoms of other elements are given in the shell diagrams in Figure 3.2.4 and are written in Table 3.2.3.

Element	Electronic configuration
hydrogen	H (1)
boron	B (2,3)
oxygen	O (2,6)
sodium	Na (2,8,1)
argon	Ar (2,8,8)
calcium	Ca (2,8,8,2)

▲ **Table 3.2.3** Written electronic configurations of the atoms shown in Figure 3.2.4



hydrogen, $^{1}_{1}H$ boron, $^{11}_{5}B$ oxygen, $^{16}_{8}O$

sodium, $^{23}_{11}Na$ argon, $^{40}_{18}Ar$ calcium, $^{40}_{20}Ca$

▲ **Figure 3.2.4** Shell diagrams showing the electronic configurations of atoms of different elements

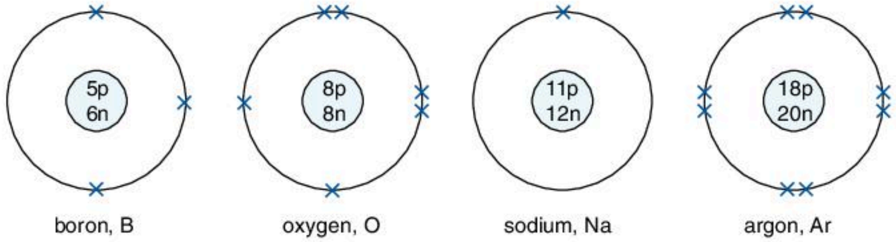
Exam tip

It is important that you can draw **shell diagrams** of atoms of atomic numbers 1 to 20, i.e. hydrogen to calcium in the periodic table.

Valence electrons

The electrons in the outermost energy shell are known as **valence electrons**. For example, potassium (2,8,8,1) has one valence electron.

The outermost electrons are the ones that are involved in chemical reactions, i.e. the valence electrons. A diagram of an atom can be drawn that shows only the valence electrons. Examples are given in Figure 3.2.5.



boron, B oxygen, O sodium, Na argon, Ar

▲ **Figure 3.2.5** Diagrams of boron, oxygen, sodium and argon atoms showing the valence electrons only

Summary questions

1 What is the maximum number of electrons allowed in shell number 2?

2 Represent the electronic configuration of each of the following atoms using both a shell diagram and writing:

a magnesium, $^{24}_{12}Mg$ **b** chlorine, $^{35}_{17}Cl$ **c** neon, $^{20}_{10}Ne$.

3 For each of the elements in question **2**, give the number of valence electrons.

A3.3 Isotopes and radioactivity

Objectives

By the end of this topic you will be able to:

- define the terms isotope and isotopy
- define relative atomic mass
- explain what is meant by a radioactive isotope
- describe the uses of radioactive isotopes.

Key fact

Isotopes are different atoms of the same element that have the same number of protons and electrons but different numbers of neutrons.

Key fact

Isotopy is the occurrence of atoms of the same element that have the same number of protons and electrons but different numbers of neutrons.

Isotopy

The atomic number is unique to a particular element. For example, all atoms of magnesium have an atomic number of 12, i.e. they all have 12 protons and, therefore, 12 electrons. The number of neutrons in atoms of the same element is not unique. For example, most magnesium atoms have 12 neutrons, however, some have 13 neutrons and others have 14 neutrons. Atoms with the same number of protons and electrons but different numbers of neutrons are called **isotopes**. The occurence of these different isotopes is known as **isotopy**.

In other words, isotopes have the same atomic number but different mass numbers. Isotopes have:

- the *same* chemical properties and electrical properties because they have the same number and arrangement of electrons, e.g. if the different isotopes of magnesium react with oxygen, the chemical reaction would be the same
- slightly *different* physical properties because of the different numbers of neutrons, e.g. the different isotopes of magnesium have slightly different masses.

Most elements have more than one isotope, but not all of these isotopes are stable. The unstable ones decay into other isotopes. The percentage of each isotope of an element in a sample of the element is referred to as the **abundance** of the isotope.

Examples

1 Carbon

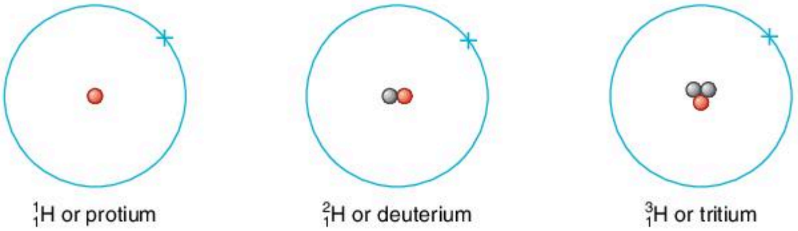
Carbon, atomic number 6, has *three* naturally occurring isotopes, one with a mass number of 12, one with a mass number of 13 and one with a mass number of 14. There are two main methods of representing these isotopes:

- Using nuclear notation:
  $^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$
- By name:
  carbon-12 or C-12
  carbon-13 or C-13
  carbon-14 or C-14.

Carbon-12 has 6 neutrons, carbon-13 has 7 neutrons and carbon-14, which is unstable, has 8 neutrons. Carbon-12 is the most abundant carbon isotope. In fact 98.89% of all naturally occurring carbon is the carbon-12 isotope.

2 Hydrogen

Hydrogen has *three* naturally occurring isotopes.



▲ **Figure 3.3.1** The three isotopes of hydrogen

Hydrogen is the only element that has different names for its isotopes. These are given in Figure 3.3.1. Protium, or normal hydrogen as it is usually

Did you know?

The deuterium isotope of hydrogen is heavier than normal hydrogen since it has an extra neutron. Water made from deuterium in place of normal hydrogen is known as **heavy water**. This heavy water is used in nuclear reactors to slow down the fast-moving neutrons, thus enabling a sustained chain reaction.

called, has a single proton in its nucleus. It is the most abundant hydrogen isotope, making up 99.985% of all naturally occurring hydrogen.

3 Chlorine

Chlorine has *two* naturally occurring isotopes, ${}^{35}_{17}\text{Cl}$, or chlorine-35, and ${}^{37}_{17}\text{Cl}$, or chlorine-37. Approximately 75% of naturally occurring chlorine is the chlorine-35 isotope and 25% is the chlorine-37 isotope.

Relative atomic mass

Chemists could use grams to measure the *mass* of atoms, but because the mass of an atom is so small, it becomes very clumsy. They needed a much smaller unit than a gram to measure the mass of atoms. They designed a system that gives us numbers that are easier to work with, but they had to ensure that these numbers are representative of the mass of each atom and in the same ratio as the actual masses of atoms.

To design this system the carbon-12 isotope was chosen as a basis to work from because of its great stability and abundance in nature.

To make the numbers representative of the mass, they chose the number for carbon-12 to be 12, because carbon-12 has a mass number of 12, i.e. it has 6 protons and 6 neutrons. They defined this new unit as one-twelfth the mass of a carbon-12 atom. A carbon-12 atom was assigned a mass of 12.00 atomic mass units or amu. Therefore, one-twelfth the mass of a carbon-12 atom has a mass of 1.00 amu.

The mass of an atom of an element compared to one-twelfth the mass of a carbon-12 atom was then assigned a value known as relative atomic mass. Relative atomic mass is given the symbol $\boldsymbol{A_r}$.

Relative atomic mass, therefore, compares the mass of an atom of an element to one-twelfth the mass of a carbon-12 atom. Because it is a comparative value, relative atomic mass has *no units*.

When calculating the relative atomic mass of an element, the relative abundance of each isotope is taken into account. For example, the relative atomic mass of chlorine, which is 35.5, is the average mass of the two isotopes of chlorine according to their relative abundance. In a sample of chlorine gas, approximately 75% of the sample consists of the chlorine-35 isotope and 25% consists of the chlorine-37 isotope.

$$\text{Relative atomic mass of chlorine} = \left(\frac{75}{100} \times 35\right) + \left(\frac{25}{100} \times 37\right) = \mathbf{35.5}$$

Relative atomic mass gives a rough idea of the number of nucleons (protons and neutrons) an atom has. However, because we are looking at a mixture of isotopes when calculating relative atomic mass, the relative atomic mass is not a whole number, e.g. the relative atomic mass of carbon is 12.011, but no real atom of carbon has 12.011 nucleons.

Radioactivity

Some isotopes have unstable nuclei. These are known as radioactive isotopes. These isotopes undergo radioactive decay by spontaneously ejecting particles and radiation from their nuclei. Radioactive isotopes eject these particles to become more stable and in the process they may produce an atom of a different element.

The time taken for half of the nuclei in a sample of a radioactive isotope to undergo radioactive decay is known as the half-life of the isotope.

! Key fact

Relative atomic mass, $\boldsymbol{A_r}$, is the average mass of one atom of an element compared to one-twelfth the mass of an atom of carbon-12.

! Key fact

A **radioactive isotope** has an unstable nucleus which decays spontaneously to a more stable form by emitting particles and radiation.

? Did you know?

The three main types of particles and radiation emitted by radioactive isotopes are:

- **alpha (α) particles**, which consist of 2 protons and 2 neutrons and have a charge of +2
- **beta (β) particles**, which consist of an electron and have a charge of −1
- **gamma (γ) radiation**, which is a form of high energy electromagnetic radiation and has neither mass nor a charge.

Uses of radioactive isotopes

Carbon-14 dating

Carbon-14 dating is used to determine the *age* of plant and animal remains, up to about 60 000 years old, which have been discovered in places such as archeological sites. It makes use of the fact that the half-life of radioactive carbon-14 is 5700 years.

▲ **Figure 3.3.2** Carbon-14 dating is used to determine the age of some fossils

All living organisms contain carbon-14 and carbon-12 in a specific ratio by mass. This is the same ratio as in carbon dioxide in the atmosphere. As the carbon-14 present in a living organism's body decays, it is replaced because the organism constantly takes in molecules containing carbon in the form of food (animals) or carbon dioxide (plants).

When a living organism dies, it stops taking in these carbon containing molecules and the carbon-14 is not replaced as it decays. The ratio of carbon-14 to carbon-12, therefore, decreases. The age of a specimen can be determined by comparing the ratio of carbon-14 to carbon-12 in the specimen with the expected ratio for that organism if it was alive. For example, if the amount of carbon-14 was half of the expected amount, then the specimen would be 5700 years old, if it was a quarter of the expected amount, then the specimen would be 11 400 years old.

Radiotherapy

▲ **Figure 3.3.3** Radiotherapy is used to treat cancerous growths

Radiotherapy uses radiation to cure or control **cancer**. Cancerous growths or tumours are composed of rapidly dividing cells and these cells are particularly sensitive to damage by radiation. They can be controlled or eliminated by irradiating the area containing the growth with external or internal radiation.

External irradiation can be carried out by directing a beam of gamma radiation from a radioactive cobalt-60 source at the cancerous growth. Internal irradiation involves either injecting a radioactive isotope, usually a gamma or beta emitter, into the target area, or placing a radiation source in a tiny protective capsule or wire (a radioactive seed) directly at the site of the cancerous tumour. For example, iodine-131 is used to treat thyroid cancer and radioactive seeds containing iridium-192, palladium-103 and iodine-125 are used to treat breast and prostate cancer and brain tumours.

Tracers

Radioactive isotopes are used as **tracers** in the medical field as a diagnostic tool. These tracers generally have a short half-life so they do not remain in the body for very long. A small amount of the tracer is usually injected into the bloodstream and as it moves around the body it emits gamma rays which can be observed using special equipment. This allows the physician to observe the functioning of specific organs. Examples of medical tracers are:

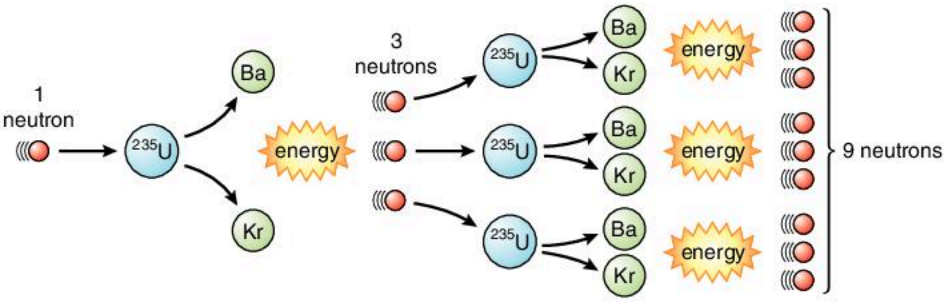
- Technetium-99 is used to produce images and carry out functional studies of most organs in the body, especially the brain, heart muscle, skeleton, lungs, liver, gall bladder and kidneys.
- Iodine-131 is used to image and carry out functional studies of the thyroid gland.

Another example of the use of radioactive tracers is in determining the uptake of atmospheric carbon dioxide by plants. Plants can be grown in an environment containing radioactive carbon dioxide, i.e. carbon dioxide that has been made by replacing the carbon-12 atoms in the molecules with carbon-14 atoms.

Using a Geiger counter (a machine that detects radiation) the parts of the plant that absorb the radioactive carbon dioxide can be identified.

Energy generation

Radioactive isotopes are used to **generate electricity** in nuclear power stations. The isotopes used are ones that release very large amounts of energy when they undergo **nuclear fission**, i.e. when the nucleus splits, usually into two smaller nuclei. Nuclear power stations use uranium-235 or plutonium-239. If an atom of either uranium-235 or plutonium-239 is struck by a free neutron, it splits into two smaller atoms and releases two or three neutrons and a large amount of energy. The released neutrons can then strike other atoms, which also split releasing more neutrons and energy. This sets up a chain reaction which releases enormous amounts of energy.



▲ **Figure 3.3.4** The chain reaction in splitting a uranium atom

If the chain reaction is controlled, the energy produced can be used to generate electricity. If the chain reaction is not controlled, it results in a nuclear explosion such as occurs in an atom bomb.

Did you know?

One type of **atom bomb** derives its destructive force from nuclear fission. Only two atom bombs have been used during warfare, both by the USA towards the end of World War II. The first, a uranium fission bomb, code named 'Little Boy', was exploded over the Japanese city of Hiroshima on 6 August, 1945. The second, a plutonium fission bomb, code named 'Fat Man', was detonated over Nagasaki in Japan three days later.

Heart pacemakers

A **pacemaker** is a medical device that is used to regulate the beating of the heart. It is implanted under the skin just below the collar bone and is connected to the heart muscle by electrodes.

In the early 1970s, some pacemakers were developed that were powered by batteries containing plutonium-238. As the plutonium-238 undergoes natural radioactive decay, it emits energy. In the pacemaker battery, this energy is used to generate an electric current which delivers about 0.2 milliwatt of power. Pacemakers powered by chemical batteries have a lifespan of about 10 years. To replace the batteries, the pacemaker has to be surgically removed and then replaced. Since the half-life of plutonium-238 is about 87 years, pacemakers powered by plutonium-238 batteries should be able to deliver current for a patient's lifetime without having to be replaced.

Did you know?

It is thought that only about 50 to 100 people living today have **pacemakers** that are powered by plutonium-238 batteries. Their use was never really developed for several reasons, one of the main ones being the reluctance of people to have a device containing radioactive plutonium in their body for life.

Summary questions

1 Define the following terms:
  a isotope
  b relative atomic mass.
2 Give the similarities and differences between the two isotopes of sodium, Na-23 and Na-22.
3 Explain why relative atomic mass has no units.
4 Explain what a radioactive isotope is.
5 How is carbon-14 used in the dating of fossil samples?

Key concepts

- Atoms are the basic building blocks of matter. An **atom** is the smallest identifying particle of an element that has the properties of the element.
- Atoms consist of three fundamental particles, known as subatomic particles: **protons**, **neutrons** and **electrons**.
- Protons and neutrons are found in the centre of the atom in an area known as the **nucleus**. Protons and neutrons are collectively known as **nucleons**.
- Electrons are found at quite a distance from the nucleus in a series of levels known as **energy shells**. Electrons in the same shell have the same amount of energy.
- Protons are positively charged, electrons are negatively charged and neutrons have no charge.
- The size of the charge on a proton is equal to the size of the charge on an electron. In an atom, the number of protons equals the number of electrons.
- The mass of a proton is equal to the mass of a neutron. An electron has a mass of $\frac{1}{1836}$ the mass of a proton or a neutron.
- Protons and neutrons are responsible for the mass of an atom. Electrons are responsible for the volume of an atom.
- The number of protons in an atom is known as the **atomic number**.
- The total number of protons and neutrons in an atom is known as the **mass number**.
- The electrons are arranged in the energy shells in a specific way known as the **electronic configuration**.
- Each electron shell can hold up to a certain maximum number of electrons.
- Shells closest to the nucleus fill up first.
- The electronic configuration of an atom can be represented in writing using numbers or by a shell diagram.
- Electrons in the outermost shell are known as **valence electrons**.
- The chemical properties of an element depend on the arrangement of electrons in the atoms of the element.
- **Isotopes** are different atoms of the same element that have the same number of protons and electrons but different numbers of neutrons. They have the same atomic number but different mass numbers.
- **Isotopy** is the occurrence of atoms of the same element that have the same number of protons and electrons but different numbers of neutrons.
- Isotopes of an element have the same chemical properties but different physical properties.
- The **relative atomic mass, A_r**, is a number that compares the average mass of an atom with one-twelfth the mass of a carbon-12 atom.
- Some isotopes have unstable nuclei. These are known as **radioactive isotopes**.
- Radioactive isotopes have many uses. Some of these are carbon-14 dating, radiotherapy, tracers, energy generation and pacemakers.

Practice exam-style questions

Multiple-choice questions

1 If the mass number of an atom is 59 and the number of neutrons in the nucleus of the atom is 32, then the atomic number is:

A 27
B 32
C 59
D 91

2 A potassium atom is written as $^{39}_{19}K$. Which of the following statements about a potassium atom are true?

I It has an atomic number of 19.
II It has 20 neutrons.
III It has 19 electrons.
IV The total number of protons and electrons is 39.

A I and IV only
B II and III only
C I, II and III only
D I, III and IV only

3 Which of the following is the electronic configuration of fluorine, $^{19}_{9}F$?

A 2,8,8,1
B 2,8,1
C 2,8
D 2,7

4 Isotopes of an element have:

I the same atomic number but different mass numbers
II the same number of electrons but different numbers of neutrons
III the same physical properties but different chemical properties

A I only
B I and II only
C II and III only
D I, II and III

5 Which of the following is not a use of radioactive isotopes?

A energy generation
B carbon dating
C chemotherapy
D radiotherapy

Structured question

6 a i) Complete the following table by filling in the gaps.

Particle	Mass number	Atomic number	Number of neutrons	Number of electrons
^{56}Fe		26		
^{58}Fe				

(7 marks)

ii) A sample of iron was found to contain a mixture of ^{56}Fe and ^{58}Fe in a ratio of 7 to 3. Calculate the relative atomic mass of the sample of iron. *(2 marks)*

b Students were asked questions based on the information given in the table below.

Element	Mass number	Atomic number	Electron configuration
A	25	–	2,8,2
B	17	8	–

The following statements were taken from one of the student's exercise books. These statements are INCORRECT. In each case, explain why the statement is incorrect and then give the correct statement.

i) INCORRECT STATEMENT:
'The atomic number of element A is 2 because it has 2 valence electrons.'
– Explain why the statement is incorrect.
– Give the correct statement. *(3 marks)*

ii) INCORRECT STATEMENT:
'The electron configuration of element B is 2,8,7 because it has a mass number of 17.'
– Explain why the statement is incorrect.
– Give the correct statement . *(3 marks)*

Total 15 marks

Extended response question

7 a The following nuclear notations refer to atoms of aluminium and fluorine:

$^{27}_{13}Al$ and $^{19}_{9}F$

Draw shell diagrams to show the structure of an aluminium atom and a fluorine atom. *(4 marks)*

b A naturally occurring element X consists of 85% $^{21}_{10}X$ and 15% $^{22}_{10}X$.

i) What can you deduce about naturally occurring element X? *(2 marks)*

ii) Determine the relative atomic mass of naturally occurring element X. *(2 marks)*

c i) What is a radioactive isotope? *(2 marks)*

ii) Explain how:
– uranium-235 is used to generate electricity
– carbon-14 is used to determine the age of plant and animal remains. *(5 marks)*

Total 15 marks

A4 The periodic table and periodicity

Objectives

By the end of this topic you will be able to:

- describe the historical development of the periodic table
- explain how elements are arranged in the modern periodic table
- identify metals, non-metals and metalloids in the periodic table
- explain the difference between groups and periods
- give the similarities in the electronic configuration of elements in the same group
- give the similarities in the electronic configuration of elements in the same period
- explain how the properties of elements change down a group and along a period.

Key fact

Periodicity is the recurrence of elements with similar physical and chemical properties at regular intervals in the periodic table.

Elements are arranged in the periodic table in increasing order of atomic number. The arrangement of the elements in specific rows and columns provides a wealth of information about the properties of elements in the same group and in the same period of the periodic table. One of the main values of the periodic table is the ability to predict the chemical properties of an element based on its location in the table.

A4.1 Arrangement of elements in the periodic table

Everyone is familiar with the **periodic table** of elements as most chemistry labs have one on the wall. The periodic table is a **classification** of all the elements and it is incredibly important to chemists. It shows the symbol, full name, mass number and atomic number of each element and it organises the elements into columns and rows based on the structure of their atoms and their properties.

Elements in the periodic table show **periodicity**. When all the elements are arranged in order of increasing atomic number, elements that are physically and chemically similar to each other occur at regular intervals, i.e. those with similar properties recur *periodically* throughout the series.

Historical development of the periodic table

Early in the 19th century, scientists started to try to classify elements based on the similarities between them. Four scientists made important contributions to the development of the modern periodic table: Johann Döbereiner, John Newlands, Dmitri Mendeleev and Henry Moseley.

Johann Döbereiner

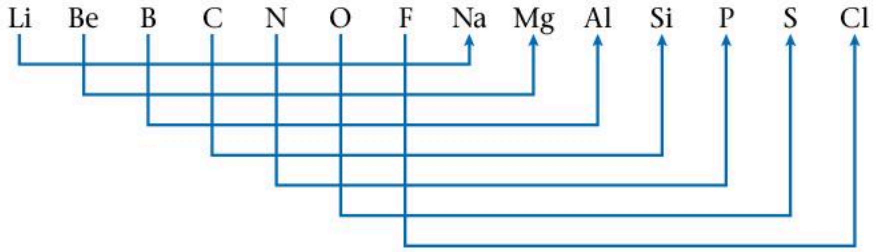
In 1829, Johann Döbereiner noted that certain groups of *three* elements possessed similar chemical and physical properties, e.g. he found that lithium, sodium and potassium were all soft, reactive metals. He also observed that if the three elements in a group were arranged in order of increasing relative atomic mass, then the relative atomic mass of the middle element was close to the average of the other two elements. He called these groups **triads**.

▼ **Table 4.1.1** Döbereiner's triads of elements

Elements and their relative atomic masses			Average relative atomic mass
lithium 7.0	sodium 23.0	potassium 39.0	23.0
calcium 40.0	strontium 88.0	barium 137.0	88.5
chlorine 35.5	bromine 80.0	iodine 127.0	81.25

John Newlands

In 1865, John Newlands put forward his Law of Octaves. He arranged the 56 elements that had been discovered at the time in increasing order of relative atomic mass and discovered that similar chemical and physical properties recurred every *eighth* element. For example, starting with lithium and arranging the elements in increasing order of relative atomic mass, he found that sodium was the eighth element and both lithium and sodium showed similar properties.



▲ **Figure 4.1.1** Newlands' octaves of elements

Dmitri Mendeleev

In 1869, Dmitri Mendeleev published his 'Periodic Classification of Elements' in which he arranged elements in increasing relative atomic mass and placed elements with similar chemical and physical properties together in vertical columns. His table was widely accepted at the time for two reasons:

- he left gaps when it seemed that the corresponding elements had not yet been discovered
- he occasionally ignored the order suggested by relative atomic mass and switched adjacent elements to better classify them into chemical families. For example, he placed tellurium, relative atomic mass 128, before iodine, relative atomic mass 127, so that iodine fell in the same group as fluorine, chlorine and bromine, all of which showed similar chemical properties.

Using his table, Mendeleev was able to predict the properties of the missing elements, e.g. he predicted the properties of gallium, which was not discovered until 1875. Mendeleev is credited with being the creator of the first version of the periodic table.

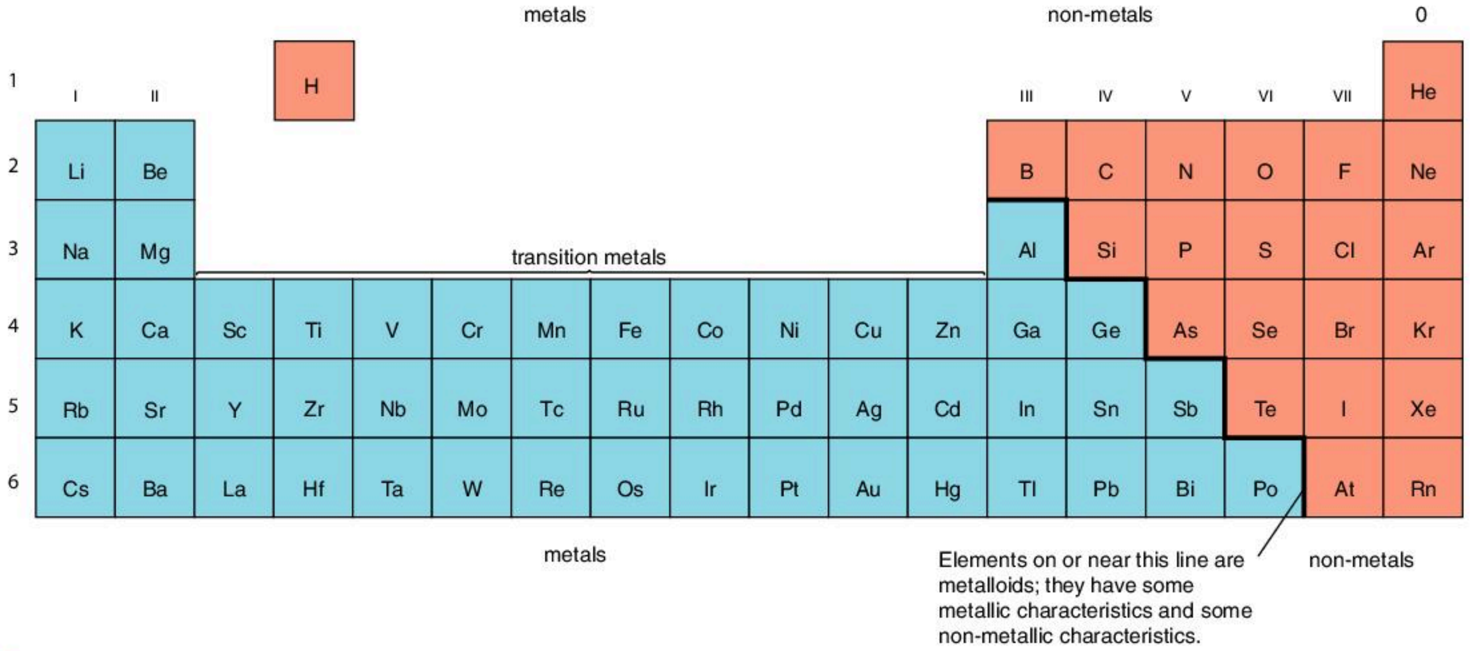
Henry Moseley

In 1914, Henry Moseley rearranged the elements in the table based on their atomic numbers instead of their relative atomic masses. In this new order, elements with similar chemical properties all fell in the same groups.

The modern periodic table

The modern periodic table consists of 103 elements arranged in order of increasing atomic number. A copy of the periodic table can be found on page 360.

Elements in the modern periodic table can be divided into metals and non-metals. The properties of metals and non-metals are given in Unit 4.4. Looking at the periodic table in Figure 4.1.2, the solid line which runs from above aluminium (Al) stepping down to astatine (At) divides the metals and the non-metals. All the elements to the left of the line, excluding hydrogen (H), are metals and the elements to the right of the line, including hydrogen, are non-metals. Some of the elements which border the line are known as **metalloids**. The properties of metalloids are intermediate between those of metals and non-metals. The metalloids include boron (B), silicon (Si), germanium (Ge), arsenic (As), antimony (Sb), tellurium (Te) and polonium (Po).



	I	II											III	IV	V	VI	VII	0
1				H														He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn

Elements on or near this line are metalloids; they have some metallic characteristics and some non-metallic characteristics.

▲ **Figure 4.1.2** Metals and non-metals in the periodic table

The periodic table is divided into vertical columns called **groups** and horizontal rows called **periods**.

Groups

Groups are vertical columns of elements. There are 18 groups in the periodic table. Eight of these are numbered using Roman numerals from I to VII, with the last group being Group 0. Elements in the same group show similar properties and, because of this, some groups have been assigned special names:

- Group I – alkali metals
- Group II – alkaline earth metals
- Group VII – halogens
- Group 0 – noble gases.

Between Groups II and III there are ten groups of elements called the **transition elements** or **transition metals**.

Periods

Periods are horizontal rows of elements numbered using Arabic numerals from 1 to 7. Moving along a period, the properties of the elements gradually change from metals on the left-hand side to non-metals on the right-hand side.

Electronic configuration and the periodic table

Looking at the electronic configuration of the elements in the periodic table, immediately we can see *similarities* between elements in the same group and between elements in the same period.

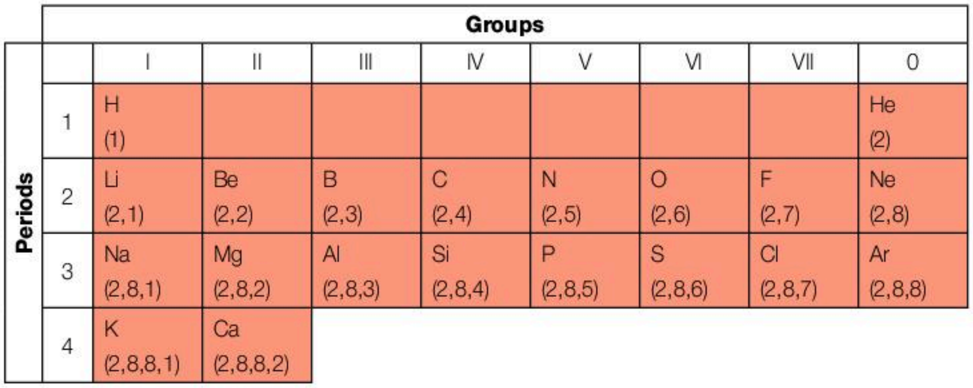
? Did you know?

The **periodic table** found in most textbooks consists of 103 elements. To date, however, 118 elements have actually been discovered or synthesised. Elements up to and including californium, atomic number 98, exist naturally, the rest have been synthesised in laboratories or nuclear accelerators. In some cases only very few atoms of the element have been synthesised.

? Did you know?

Naming newly discovered or synthesised elements is controlled by the **International Union of Pure and Applied Chemistry (IUPAC)**. Many are named after famous scientists or places, e.g. einsteinium and californium.

Figure 4.1.3 summarises the electronic configuration of the first 20 elements in the periodic table.



Periods	Groups I	II	III	IV	V	VI	VII	0
1	H (1)							He (2)
2	Li (2,1)	Be (2,2)	B (2,3)	C (2,4)	N (2,5)	O (2,6)	F (2,7)	Ne (2,8)
3	Na (2,8,1)	Mg (2,8,2)	Al (2,8,3)	Si (2,8,4)	P (2,8,5)	S (2,8,6)	Cl (2,8,7)	Ar (2,8,8)
4	K (2,8,8,1)	Ca (2,8,8,2)						

Figure 4.1.3 The electronic configuration of the first 20 elements in the periodic table

Groups

The similarities in the electronic configuration of elements in the same group are as follows:

- all elements in the same group have the same number of electrons in their outermost electron shell, i.e. they have the same number of valence electrons, for example, beryllium (Be), magnesium (Mg) and calcium (Ca) are all in the same group and they all have two valence electrons
- the number of valence electrons is the same as the group number, for example, sulfur (S) has six valence electrons and is in Group VI
- all elements in Group 0 have a full outermost electron shell, for example, helium (He) has two electrons in shell number 1 which is the maximum number of electrons this shell can hold, neon (Ne) has eight electrons in shell number 2, once again the maximum this shell can hold.

Moving down any group, each element has one more electron shell than the element directly above it. For example, lithium (Li) has two shells and sodium (Na) has three shells.

Periods

The similarities in the electronic configuration of elements in the same period are as follows:

- all elements in the same period have their valence electrons in the same electron shell, for example, sodium (Na), magnesium (Mg) and aluminium (Al) are all in the same period and they all have their valence electrons in shell number 3
- the number of occupied electron shells, i.e. the number of shells that contain electrons, is the same as the period number, for example, calcium (Ca) has four shells that contain electrons and it is in Period 4.

Moving along any period, each element has one more valence electron than the element directly before it. For example, silicon (Si) has four valence electrons and phosphorus (P) has five valence electrons.

Determining the electronic configuration of atoms

The electronic configuration of an atom can be determined from the group number and the period number, remembering that:

- the number of valence electrons is the same as the group number
- the number of occupied electron shells is the same as the period number.

Exam tip

Correctly representing the **electronic configuration** of atoms of elements is extremely important. It is used to determine the **reactivity** of an atom and the type of **bonding** an atom will be involved in when it forms a compound.

For example, if you know that phosphorus is in Group V and Period 3 of the periodic table, you can determine that a phosphorus atom has five valence electrons and three occupied electron shells. Therefore, the electronic structure of a phosphorus atom is (2,8,5).

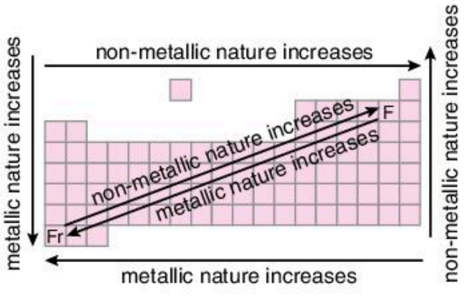
Alternatively, the group number and period number of an element can be determined if its electronic configuration is known. For example, if you know that the electronic configuration of a calcium atom is (2,8,8,2), you can determine that calcium is in Group II and Period 4 of the periodic table.

General trends in the periodic table

In Unit 3.2 you learnt that the **chemical properties** of an element depend on the arrangement of electrons in the atoms of the element, in particular, the number of valence electrons they have. The elements in the periodic table are arranged in such a way that there are trends in the physical and chemical properties of the elements in a particular group or in a particular period. These trends are directly related to electronic configurations of elements in the same group and elements in the same period. These trends can be summarised as follows.

- Moving down a group, the metallic nature of the elements *increases*.
- Moving down a group, the non-metallic nature of the elements *decreases*.
- Moving along a period, the metallic nature of the elements *decreases*.
- Moving along a period, the non-metallic nature of the elements *increases*.

These trends are shown in Figure 4.1.4.



Francium (Fr) is the most reactive metal

Fluorine (F) is the most reactive non-metal

▲ **Figure 4.1.4** Trends in the periodic table

It is important to note that elements in Group 0 are chemically *unreactive*. Their atoms all have full outer electron shells and, because of this, they have a stable electronic configuration. Atoms of all other elements do not have full outer electron shells and are, therefore, unstable. In order to become stable, they need to attain the electronic configuration of a noble gas, i.e. a full outer electron shell. They can do this by losing or gaining valence electrons.

The willingness of an atom of an element to lose or gain valence electrons is a strong indication of its metallic or non-metallic nature. Metals tend to *lose* valence electrons and non-metals tend to *gain* electrons into their valence shell when they react.

When atoms lose or gain valence electrons they form positively or negatively charged particles known as **ions**. They are said to have ionised. The ease with which atoms lose or gain valence electrons, known as **ease of ionisation**, determines how reactive they are.

Key fact

An **ion** is an electrically charged particle formed when an atom loses or gains electrons.

In general:

- The ease with which **metal** atoms lose electrons increases going *down* a group and going from *right to left* along a period. The reactivity of metals, therefore, increases going down a group and going from right to left along a period.
- The ease with which **non-metal** atoms gain electrons increases going *up* a group and going from *left to right* along a period. The reactivity of non-metals, therefore, increases going up a group and going from left to right along a period.

You will be studying the reasons for this in more detail in the next three units.

Summary questions

1 Name three scientists who made important contributions to the development of the periodic table.

2 What is the name given to the following elements in the periodic table?
a Group II **b** Group VII **c** Group 0

3 For elements in Group II of the periodic table, give the similarities and differences in their electronic configuration.

4 For elements in Period 3 of the periodic table, give the similarities and differences in their electronic configuration.

5 **a** Give the electronic arrangement of an atom of potassium, given that it is in Group I and Period 4.
b Give the group number and period number of sulfur, given that its electronic configuration is S (2,8,6).

6 How does the metallic nature of elements change when moving down a group and along a period?

A4.2 Trends in Group II of the periodic table

Group II of the periodic table is composed of the following elements: beryllium, magnesium, calcium, strontium, barium and radium. These elements are all metals which display very similar properties. They are chemically very reactive and are collectively known as **alkaline earth metals**. Because they are so reactive, they are never found naturally in their free state but are always found in nature combined with other elements in compounds. The elements in Group II are listed in Figure 4.2.1, together with their symbols, atomic numbers and electron arrangements.

beryllium ${}_{4}\text{Be}$ (2,2)
magnesium ${}_{12}\text{Mg}$ (2,8,2)
calcium ${}_{20}\text{Ca}$ (2,8,8,2)
strontium ${}_{38}\text{Sr}$ (2,8,18,8,2)
barium ${}_{56}\text{Ba}$ (2,8,18,18,8,2)
radium ${}_{88}\text{Ra}$ (2,8,18,32,18,8,2)

▲ **Figure 4.2.1** Group II elements and their electronic configurations

Objectives

By the end of this topic you will be able to:

- give the general properties of Group II elements
- give the chemical properties of Group II elements
- compare the reactivity of Group II elements with oxygen, water and dilute hydrochloric acid
- predict the properties of unknown elements in Group II based on group trends
- explain why elements in Group II show similar chemical properties
- explain why the reactivity of Group II elements increases going down the group.

Key fact

A **cation** is a positively charged ion frmed when a metal atom loses electrons.

General properties of Group II elements

Group II elements have the following general properties in common.

- They are fairly soft metals.
- They are shiny, silvery-white when freshly cut, but turn dull very quickly as they react with the oxygen in the air. This reaction forms a dull oxide layer on the surface of the metal.
- They have fairly high melting points and boiling points.
- Their atoms have *two* valence electrons.
- Their atoms lose their valence electrons readily to form metal **cations** with a charge of +2, e.g. Mg^{2+}, Ca^{2+} and Ba^{2+} (see Unit 5.2).
- Going *down* the group, the radii of the atoms (atomic radii) *increase* and it becomes easier for the atoms to lose their two valence electrons.

Chemical reactivity of Group II elements

Group II elements have the following chemical properties in common.

- They react with oxygen to form basic oxides (see Unit 8.2).
- They react with water to form metal hydroxides and hydrogen gas.
- They react with dilute hydrochloric acid to form salts known as metal chlorides (see Unit 8.4) and hydrogen gas.

The ease with which the metals react with oxygen, water and dilute hydrochloric acid is, however, different for the different metals and is known as the **reactivity** of the metal. If we rank the metals from the most reactive to the least reactive, based on how fast or how vigorously they react, the reactivity increases going *down* the group. Beryllium, at the top of the group, is the *least* reactive and radium, at the bottom of the group, is the *most* reactive.

Practical activity

Comparing the reactivity of magnesium and calcium

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with test tubes, strips of freshly cleaned magnesium ribbon, pieces of freshly cleaned calcium, dilute hydrochloric acid and distilled water.

Method

1 Pour 2 cm^3 of dilute hydrochloric acid into each of two test tubes. Add a piece of magnesium ribbon to the first tube and a piece of calcium to the second tube. Compare the strength of effervescence occurring in the two tubes.

2 Pour 2 cm^3 of distilled water into each of two test tubes. Add a piece of magnesium ribbon to the first tube and a piece of calcium to the second tube. Compare the strength of effervescence occurring in the two tubes. You might need to leave the tubes for a while and observe them periodically.

3 Based on the relative strength of effervescence that you observed in steps **1** and **2**, which metal would you consider is more reactive, magnesium or calcium?

4 If you added beryllium to dilute hydrochloric acid and distilled water, would you expect the effervescence to be more vigorous or less vigorous than with magnesium?

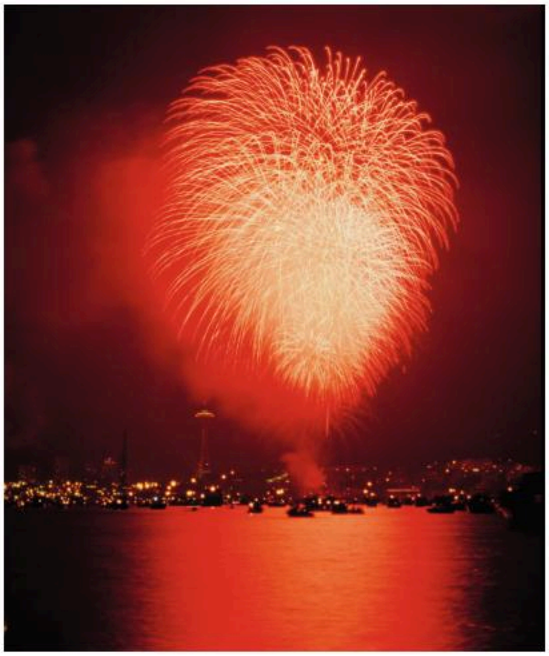
5 If you added barium to dilute hydrochloric acid and distilled water, would you expect the effervescence to be more vigorous or less vigorous than with calcium?

A comparison of the reactivity of magnesium, calcium and barium with oxygen, water and dilute hydrochloric acid is given in Table 4.2.1. Examples of the chemical equations for the reactions with magnesium are given. Similar equations can be written for the reactions with calcium and barium. You have already come across (g) and (s) to indicate a gas and a solid in a chemical equation. In these equations, (l) indicates a liquid and (aq) indicates an aqueous solution, i.e. a solution where water is the solvent.

▼ **Table 4.2.1** Reactions of magnesium, calcium and barium

Reaction	Magnesium (Mg)	Calcium (Ca)	Barium (Ba)
Oxygen (or air)	Reacts *slowly* to form a coating of magnesium oxide (MgO) on exposure to air. If ignited, magnesium burns with a blinding white flame, producing white, solid magnesium oxide. magnesium + oxygen → magnesium oxide $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$	Reacts *readily* to form a coating of calcium oxide (CaO) on exposure to air. if ignited, calcium burns with a brick red flame, producing white, solid calcium oxide.	Reacts *very readily* to form a coating of barium oxide (BaO) on exposure to air. If ignited, barium burns with an apple green flame, producing white, solid barium oxide.
Water	Very clean magnesium reacts *very slowly* with cold water, producing magnesium hydroxide ($Mg(OH)_2$) and hydrogen gas (H_2). magnesium + water → magnesium hydroxide + hydrogen $Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$	Reacts *vigorously* with cold water, producing calcium hydroxide ($Ca(OH)_2$) and hydrogen gas (H_2).	Reacts *very vigorously* with cold water, producing barium hydroxide ($Ba(OH)_2$) and hydrogen gas (H_2).
Dilute hydrochloric acid	Reacts *vigorously* to produce magnesium chloride ($MgCl_2$) and hydrogen gas (H_2). magnesium + hydrochloric acid → magnesium chloride + hydrogen $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$	Reacts *very vigorously* to produce calcium chloride ($CaCl_2$) and hydrogen gas (H_2).	Reacts *violently* to produce barium chloride ($BaCl_2$) and hydrogen gas (H_2).

From the trends in the reactivity of the Group II metals, we can predict the reactivity of beryllium, strontium and radium. We would expect beryllium to react *less* vigorously than magnesium, radium to react even *more* vigorously than barium, and the reactivity of strontium to be *more* vigorous than calcium, but *less* vigorous than barium.



▲ **Figure 4.2.2** Compounds of Group II elements are often used as fireworks

? Did you know?

Group II metals are often used in **fireworks** because they burn in air or oxygen producing coloured flames. Magnesium is used to produce a bright white light effect and calcium produces a brick red colour. Compounds of Group II metals may also be used, e.g. strontium carbonate produces a crimson red colour and barium chloride produces an apple green colour.

An explanation of the trends in Group II

Elements in Group II all have very similar chemical properties because their atoms all have two valence electrons. Being metals, when they react they ionise by losing their two valence electrons and form positively charged ions, i.e. **cations**. In doing this, their outer electron shell is now full.

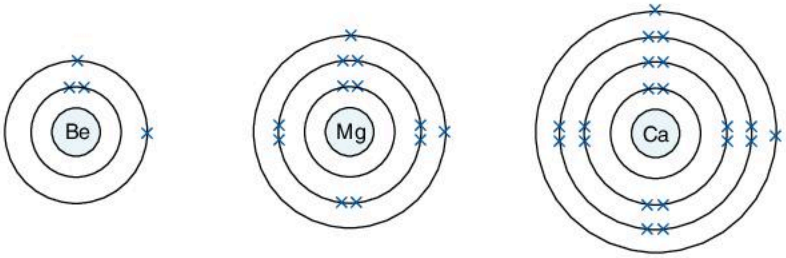
Since all the elements have the same number of electrons to lose, i.e. 2, they all have very similar chemical properties, however, moving *down* Group II the reactivity of the elements *increases*. Moving down the group, the number of occupied electron shells increases. As a result:

- the atomic radii get bigger due to the extra filled electron shells, therefore, the valence electrons become further away from the attractive pull of the positive nucleus
- the valence electrons become more shielded from the positive nucleus by the extra full electron shells.

The combination of these two factors means that the valence electrons are less attracted to the positive nucleus going down the group and it becomes easier for the atoms to lose their valence electrons. The ease of ionisation, therefore, increases going *down* the group.

Beryllium atoms, being the smallest, ionise the least easily. Beryllium is, therefore, the *least* reactive Group II element. Radium atoms, being the largest, ionise the most easily. Radium is, therefore, the *most* reactive Group II element.

The electronic structure of the first three elements in Group II is given in Figure 4.2.3.



▲ **Figure 4.2.3** Shell diagrams of the first three elements in Group II

Summary questions

1 What is the collective name for the elements in Group II?

2 a Does the reactivity of the elements in Group II increase or decrease going down the group?

b Give the reason for your answer to **a** above.

3 Compare the reactivity of beryllium and calcium.

A4.3 Trends in Group VII of the periodic table

Objectives

By the end of this topic you will be able to:

- give the general properties of Group VII elements
- give the chemical properties of Group VII elements
- explain why elements in Group VII show similar chemical properties
- explain why the reactivity of the elements in Group VII increases going up the group
- describe displacement reactions of elements in Group VII
- explain the term oxidising strength
- predict the properties of unknown elements in Group VII based on group trends.

Group VII in the periodic table is composed of the following elements: fluorine, chlorine, bromine, iodine and astatine. These elements are all non-metals. They display similar properties and are chemically very reactive. They are collectively known as the **halogens**. The elements in Group VII are listed in Figure 4.3.1, together with their symbols, atomic numbers and electron arrangements.

fluorine ${}_{9}\text{F}$ (2,7)
chlorine ${}_{17}\text{Cl}$ (2,8,7)
bromine ${}_{35}\text{Br}$ (2,8,18,7)
iodine ${}_{53}\text{I}$ (2,8,18,18,7)
astatine ${}_{85}\text{At}$ (2,8,18,32,18,7)

▲ **Figure 4.3.1** Group VII elements and their electronic configurations

General properties of Group VII elements

Group VII elements have the following general properties in common.

- They are poisonous elements.
- They exist as non-polar, diatomic molecules, e.g. F_2, Cl_2, Br_2 and I_2.
- They are generally soluble in non-polar solvents and slightly soluble in water.
- They have low melting points and boiling points.
- Their atoms have *seven* valence electrons.
- Their atoms readily accept an electron into their valence shells to form non-metal **anions** with a charge of −1, e.g. F^-, Cl^-, Br^- and I^- (see Unit 5.2).
- They share an electron readily with other non-metal atoms (see Unit 5.4).

Key fact

An **anion** is a negatively charged ion formed when a non-metal atom gains electrons.

Group VII elements have different colours and states at room temperature, as summarised in Table 4.3.1.

▼ **Table 4.3.1** Some physical properties of the first four elements in Group VII

Element	State at room temperature	Colour
fluorine	gas	pale yellow
chlorine	gas	yellow green
bromine	liquid	red-brown – the liquid evaporates readily at room temperature forming an orange vapour
iodine	solid	grey-black – the solid sublimes readily when heated forming a purple vapour

Chemical reactivity of Group VII elements

Group VII elements react with most metals to form ionic compounds (Unit 5.2) and they react with most other non-metals to form covalent compounds (Unit 5.4).

Examples

1 Chlorine reacts with sodium to form sodium chloride.

$$\text{Sodium} + \text{chlorine} \longrightarrow \text{sodium chloride}$$
$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

2 Chlorine reacts with hydrogen gas to form hydrogen chloride gas.

$$\text{Hydrogen} + \text{chlorine} \longrightarrow \text{hydrogen chloride}$$
$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

The reactivity of the Group VII elements increases going *up* the group. Fluorine, at the top of the group, is the *most* reactive and astatine, at the bottom of the group, is the *least* reactive.

An explanation of the trends in Group VII

Elements in Group VII all have very similar chemical properties because their atoms all have seven valence electrons. Being non-metals, when they react they ionise by gaining one valence electron and they form negatively charged ions, i.e. **anions**. In doing this, their outer electron shell is now full.

Since all the elements have the same number of electrons to gain, i.e. 1, they all have very similar chemical properties. However, moving *up* Group VII the reactivity of the elements *increases*. Moving up the group, the number of occupied electron shells decreases. As a result:

- the radii of the atoms get smaller due to the decrease in number of filled electron shells, therefore, the valence electrons become closer to the attractive pull of the positive nucleus
- the valence electrons become less shielded from the positive nucleus by the inner full electron shells.

The combination of these two factors means that the valence electrons are more attracted to the positive nucleus going up the group, as are any electrons being gained by the atoms. It becomes easier for the atoms to gain an electron into their valence shell going up the group. The ease of ionisation, therefore, increases going *up* the group.

Fluorine atoms, being the smallest, ionise the most easily. Fluorine is, therefore, the *most* reactive Group VII element. Astatine atoms, being the largest, ionise the least easily. Astatine is, therefore, the *least* reactive Group VII element.

The tendency for atoms to attract electrons is known as **electronegativity**. The electronegativity of the atoms, therefore, increases going up Group VII. Fluorine is the most electronegative element and astatine is the least electronegative.

The electronic structure of the first two elements in Group VII is given in Figure 4.3.2.

Key fact

Electronegativity is a measure of how strongly an atom attracts electrons.

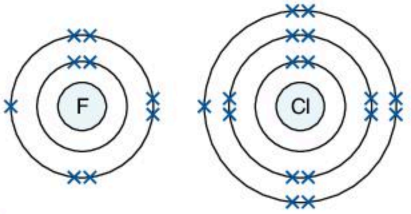
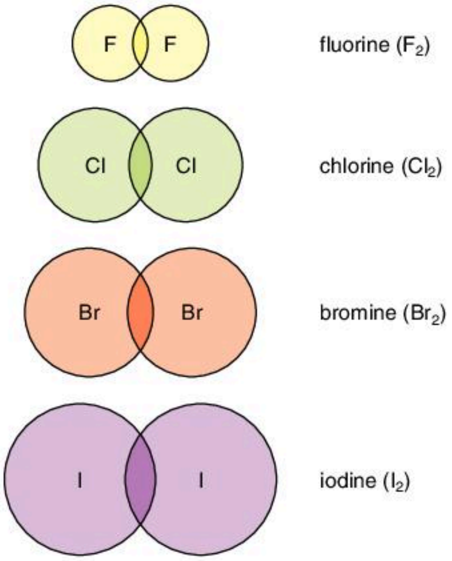


Figure 4.3.2 Shell diagrams of the first two elements in Group VII

The melting points and boiling points of the Group VII elements can also be explained by looking at the relative size of the molecules. Fluorine, having the smallest atoms, has the smallest molecules. The forces of attraction between these small molecules, known as **intermolecular forces**, are the weakest, which means that the melting point and boiling point of fluorine are the lowest. Astatine, having the largest atoms, has the largest molecules. The intermolecular forces between the large astatine molecules are the strongest, which means that the melting point and boiling point of astatine are the highest.



▲ **Figure 4.3.3** Comparing the sizes of the first four halogen molecules

Displacement reactions

A **displacement reaction** occurs when an element in its free state takes the place of another element in a compound. A more reactive element will displace a less reactive element from a compound containing the less reactive element. However, a less reactive element will not displace a more reactive element from a compound.

For example, if chlorine gas is bubbled into a solution of potassium bromide, the chlorine will displace the bromine from the potassium bromide forming potassium chloride and bromine:

$$\text{chlorine} + \text{potassium bromide} \longrightarrow \text{potassium chloride} + \text{bromine}$$
$$Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$$

However, if bromine is added to potassium chloride solution, the bromine will not displace the chlorine and no reaction occurs:

$$\text{bromine} + \text{potassium chloride} \longrightarrow \text{no reaction}$$

To determine whether a displacement reaction has occurred, an aqueous solution of a halogen is mixed with a sodium or potassium halide solution and the colour of the resulting solution is observed. The halide solutions are all colourless, e.g. potassium chloride (KCl), potassium bromide (KBr) and potassium iodide (KI) are all colourless. If a displacement reaction occurs and a different halogen forms, it can be identified by its distinctive colour. Chlorine is *yellow-green*, bromine is *red-brown* and iodine will range from a *brown* solution to a *grey-black* precipitate depending on the amount produced.

Practical activity

Investigating displacement reactions of the halogens

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with test tubes and aqueous solutions of chlorine, bromine, iodine, potassium chloride, potassium bromide and potassium iodide.

Method

1. Pour 2 cm^3 of potassium bromide solution into a test tube and 2 cm^3 of potassium iodide solution into another test tube. Carefully add chlorine solution to each tube and shake. Observe any colour changes.
2. Pour 2 cm^3 of potassium chloride solution into a test tube and 2 cm^3 of potassium iodide solution into another test tube. Carefully add bromine solution to each tube and shake. Observe any colour changes.
3. Pour 2 cm^3 of potassium chloride solution into a test tube and 2 cm^3 of potassium bromide solution into another test tube. Carefully add iodine solution to each tube and shake. Observe any colour changes.
4. Record your results in a table.
5. Using your results, place the elements in increasing order of reactivity.

From the practical activity above we can see that:

- Chlorine displaces bromine from potassium bromide producing potassium chloride and bromine:

  chlorine + potassium bromide ⟶ potassium chloride + bromine

  $$Cl_2(aq) + 2KBr(aq) \longrightarrow 2KCl(aq) + \underset{\text{red-brown}}{Br_2(aq)}$$

- Chlorine displaces iodine from potassium iodide producing potassium chloride and iodine:

  chlorine + potassium iodide ⟶ potassium chloride + iodine

  $$Cl_2(aq) + 2KI(aq) \longrightarrow 2KCl(aq) + \underset{\text{brown}}{I_2(aq)}$$

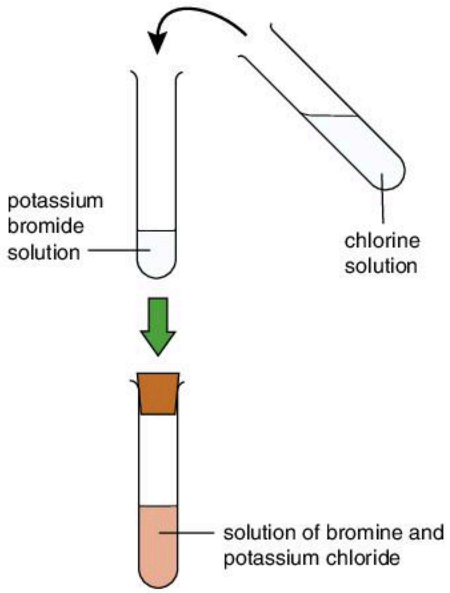
- Bromine displaces iodine from potassium iodide producing potassium bromide and iodine:

  bromine + potassium iodide ⟶ potassium bromide + iodine

  $$Br_2(aq) + 2KI(aq) \longrightarrow 2KBr(aq) + \underset{\text{brown}}{I_2(aq)}$$

- No displacement reaction occurs between bromine and potassium chloride.
- No displacement reactions occur with iodine.

From the trends in the displacement reactions involving chlorine, bromine and iodine, we can predict that fluorine, at the top of Group VII, will displace all the other halogens from compounds containing them. However, astatine, at the bottom of Group VII, will not displace any of the other halogens from their compounds.



▲ **Figure 4.3.4** Chlorine solution is added to potassium bromide solution

Exam tip

If a question asks you to describe a chemical reaction, you must provide all observations, including any colour changes.

Displacement reactions and oxidising strength

Displacement reactions can be explained by considering the relative oxidising strengths of the Group VII elements. **Oxidising strength** is a measure of how easily one substance takes electrons from another substance. The oxidising strength of the elements in Group VII increases moving *up* the group because the ability to ionise by gaining an electron increases moving up the group.

Key fact

An **oxidising agent** brings about the oxidation of another reactant by causing an atom or ion in that reactant to **lose** electrons. In the process, the oxidising agent gains electrons.

Comparing the relative oxidising strengths of chlorine, bromine and iodine, we see that chlorine has the strongest oxidising strength and iodine has the weakest. We say that chlorine is the strongest **oxidising agent** and iodine is the weakest.

A stronger oxidising agent will accept electrons from a weaker oxidising agent. This is why chlorine displaces the bromide ions in the potassium bromide. Each chlorine atom in a chlorine molecule accepts an electron from a bromide ion (Br^- ion). As a result, each chlorine atom forms a chloride ion (Cl^- ion) and each bromide ion forms a bromine atom. Two bromine atoms then pair to form a bromine molecule:

$$\begin{array}{ccccccc} \text{chlorine molecule} & + & \text{bromide ion} & \longrightarrow & \text{chloride ion} & + & \text{bromine molecule} \\ Cl_2(aq) & + & 2Br^-(aq) & \longrightarrow & 2Cl^-(aq) & + & Br_2(aq) \end{array}$$

Chlorine is said to have **oxidised** the potassium bromide. In the other displacement reactions, both chlorine and bromine oxidised the potassium iodide. You will learn a lot more about oxidation reactions in Unit 9.

Again, from the trends in the displacement reactions involving chlorine, bromine and iodine, we can predict that fluorine, at the top of Group VII, will be a *stronger* oxidising agent than chlorine and astatine, at the bottom of Group VII, will be a *weaker* oxidising agent than iodine.

Summary questions

1 What is similar in the electronic configuration of the elements in Group VII?

2 **a** Does the reactivity of the elements in Group VII increase or decrease going up the group?
b Give the reason for your answer to **a** above.

3 What is meant by the term 'oxidising strength'?

4 List the elements in Group VII in order of increasing oxidising strength.

5 **a** What would you observe if an aqueous solution of chlorine is added to potassium iodide solution?
b Give a reason for your answer to **a** above.

A4.4 Trends in Period 3 of the periodic table

Objectives

By the end of this topic you will be able to:

- give the properties of metals and non-metals
- describe how the properties of the elements change moving along Period 3
- explain why the properties of the elements change from metal to non-metal moving along Period 3.

Period 3 in the periodic table is composed of the following elements: sodium, magnesium, aluminium, silicon, phosphorus, sulfur, chlorine and argon. The physical and chemical properties of these elements change moving along the period. Moving from left to right, the **metallic nature** of the elements *decreases* and the **non-metallic nature** *increases*. The elements to the left are metals, those to the right are non-metals and silicon, in the middle, is a semi-metal or metalloid.

Physical properties of metals and non-metals

Metals have the following physical properties in common.

- They are solid at room temperature, except mercury, which is a liquid.
- They have high melting points and boiling points.
- They have high densities.
- They are good conductors of heat and electricity.
- They are shiny in appearance.
- They are malleable, i.e. they can be hammered into different shapes, and ductile, i.e. they can be drawn out into wires.
- They are sonorous, i.e. they make a ringing sound when hit.

Non-metals have the following physical properties in common.

- They are usually gases at room temperature, however, some are solid and bromine is a liquid.
- They have low melting points and boiling points.
- They have low densities.
- They are poor conductors of heat and electricity.
- In the solid state they are dull in appearance.
- In the solid state they are brittle.
- In the solid state they make a dull sound when hit.

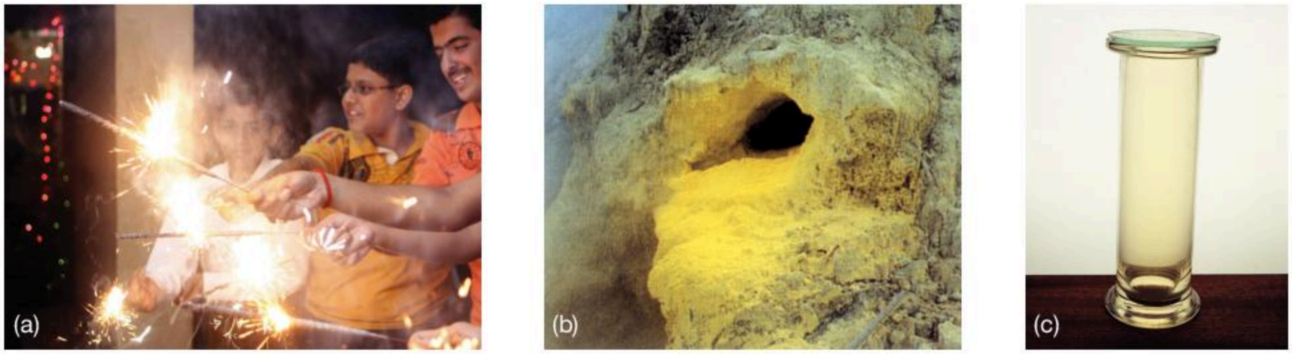
A summary of the trends of the elements in Period 3 is given in Table 4.4.1.

? Did you know?

Silicon is the second most abundant element on Earth after oxygen, making up 28% of the Earth's crust. It is a semiconductor, a property which has led to its use in the manufacture of silicon chips which are found in electronic devices from pocket calculators to computers. The first computer using silicon chips made its debut in 1961. The silicon chip forms the foundation of modern computer technology and without its invention such devices would not exist today.

▼ **Table 4.4.1** Trends in Period 3

	Na	Mg	Al	Si	P	S	Cl	Ar
Group number	I	II	III	IV	V	VI	VII	0
Atomic number	11	12	13	14	15	16	17	18
Electronic configuration	(2,8,1)	(2,8,2)	(2,8,3)	(2,8,4)	(2,8,5)	(2,8,6)	(2,8,7)	(2,8,8)
Loses/gains/shares electrons	loses $1e^-$	loses $2e^-$	loses $3e^-$	shares $4e^-$	gains $3e^-$	gains $2e^-$	gains $1e^-$	none
State at 25 °C	solid	solid	solid	solid	solid	solid	gas	gas
Electrical conductivity	conductor	conductor	conductor	semi-conductor	insulator	insulator	insulator	insulator
Metal/non-metal	metal	metal	metal	metalloid	non-metal	non-metal	non-metal	non-metal
Ease of ionisation	increasing ease of ionisation ←				increasing ease of ionisation →			



▲ **Figure 4.4.1** **(a)** Magnesium burning in sparklers, **(b)** sulfur around a volcano vent and **(c)** chlorine gas

An explanation of the trends in Period 3

As you learnt in Unit 4.1, the ease with which atoms lose or gain valence electrons is a strong indication of their metallic or non-metallic nature. Metals tend to lose electrons and non-metals tend to gain electrons. Moving along Period 3, the metallic nature of the elements decreases because their ability to *lose* electrons decreases and their non-metallic nature increases because their ability to *gain* electrons increases.

Each element in Period 3 has its valence electrons in the same electron shell, i.e. shell number 3. However, moving from left to right along the period, each element has one more electron and one more proton than the element before it. As a result:

- the number of positive protons in the nuclei of the atoms gets greater, causing the attraction between the positive nucleus and the valence electrons to get stronger
- the radii of the atoms get smaller due to the increasing number of positive protons in the nucleus, pulling the outer electrons closer to the nucleus.

The combination of these two factors means that the valence electrons are more attracted to the positive nucleus going along the period. It becomes harder for the atoms to lose their valence electrons and easier for them to gain electrons into their valence shell moving along the period. In other words, the electronegativity of the elements increases along Period 3.

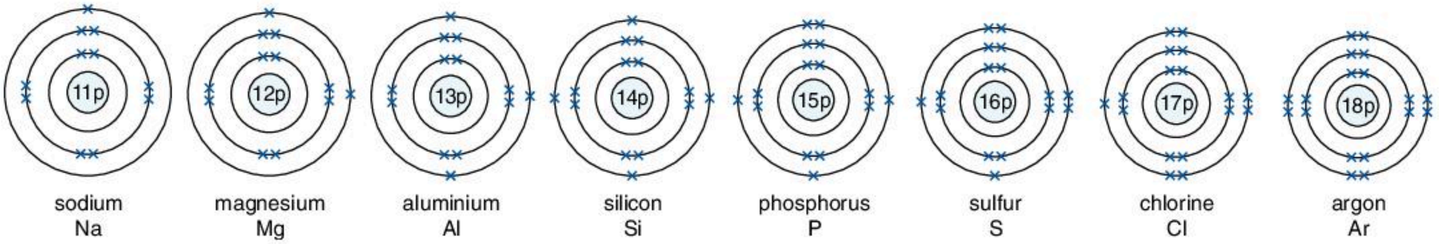
Since metal atoms ionise by losing electrons, the ease of ionisation of the metals sodium, magnesium and aluminium *decreases* along the period. Sodium is the largest of the three atoms and has 11 protons in its nucleus. Aluminium is the smallest and has two more protons in its nucleus, i.e. 13 (see Figures 4.4.2 and 4.4.3). Sodium ionises the most easily and is the *most* reactive. Aluminium ionises the least easily and is the *least* reactive.

Non-metal atoms ionise by gaining electrons, therefore, the ease of ionisation of the non-metals phosphorus, sulfur and chlorine *increases* along the period. Phosphorus is the largest of the three atoms and has 15 protons in its nucleus. Chlorine is the smallest and has two more protons, i.e. 17. Phosphorus ionises the least easily and is the *least* reactive. Chlorine ionises the most easily and is the *most* reactive.

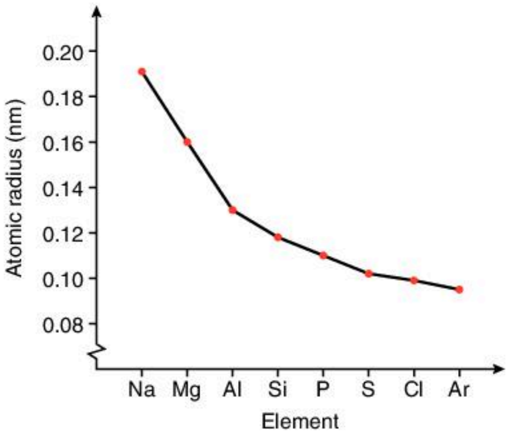
Argon is a non-metal found to the right of chlorine. However, its valence electron shell is full, i.e. it contains eight electrons. Because of this, argon has a stable electronic configuration, it does not ionise and it is chemically unreactive.

Silicon is found in Group IV and is a **metalloid** because it has a combination of metallic and non-metallic properties. It has four valence electrons and usually reacts by sharing electrons with other non-metal atoms. It does not usually ionise.

The electronic structure of the elements in Period 3 is given in Figure 4.4.2.



▲ **Figure 4.4.2** Electronic configuration of the elements in Period 3



▲ **Figure 4.4.3** Trends in the atomic radii of the elements in Period 3

Summary questions

1 Chlorine and magnesium are both elements which are found in Period 3 of the periodic table. One is a metal and one is a non-metal.
   - **a** Which one is the metal? Give THREE reasons for your answer.
   - **b** Which one is the non-metal? Give THREE reasons for your answer.

2 How does the metallic nature and the non-metallic nature of elements in Period 3 change from left to right?

3 **a** Which is more reactive, sodium or magnesium?
   - **b** Give the reason for your answer to **a** above.

4 **a** Which is more reactive, sulfur or chlorine?
   - **b** Give the reason for your answer to **a** above.

5 Which element in Period 3 has the greatest metallic nature?

Key concepts

- Scientists started classifying elements in the early 19th century.
- **Dmitri Mendeleev** is credited with being the creator of the first version of the **periodic table** in 1869.
- **Periodicity** is the recurrence of elements with similar physical and chemical properties at regular intervals in the periodic table.
- The elements in the modern periodic table are arranged in order of increasing atomic number.
- The periodic table consists of vertical columns called **groups** and horizontal rows called **periods**.
- All elements in the same group have the same number of valence electrons. This number is the same as the group number.
- All elements in the period have the same number of occupied electron shells. This number is the same as the period number.
- The chemical properties of an element are determined by the arrangement of electrons, in particular, the number of valence electrons.
- Elements in Group II have very similar properties because they all have two valence electrons.
- Group II elements all have similar reactions with oxygen, water and dilute hydrochloric acid.
- The reactivity of Group II elements, being metals, depends on the ability of their atoms to lose their valence electrons.
- The reactivity of Group II elements increases going *down* the group because the ability of their atoms to lose their valence electrons increases going down the group.
- Elements in Group VII have very similar properties because they all have seven valence electrons.
- The reactivity of Group VII elements, being non-metals, depends on the ability of their atoms to gain an electron into their valence shell.
- The reactivity of Group VII elements increases going *up* the group because the ability of their atoms to gain an electron increases going up the group.
- The elements in Group VII undergo **displacement reactions**. A more reactive element will displace a less reactive element from a compound containing the less reactive element.
- Displacement reactions of Group VII elements depend on the relative **oxidising strength** of the elements, where oxidising strength is a measure of how easily a substance takes electrons from another substance.
- The oxidising strength of Group VII elements increases going up the group.
- Going along Period 3 from left to right, the metallic nature of the elements decreases because their ability to lose electrons decreases.
- Going along Period 3 from left to right, the non-metallic nature of the elements increases because their ability to gain electrons increases.
- The **electronegativity** of elements in Period 3 increases going from left to right.

Practice exam-style questions

Multiple-choice questions

1 Elements in the modern periodic table are arranged in order of increasing:

A chemical reactivity
B atomic number
C mass number
D relative atomic mass

2 An element of atomic number 20 is expected to be:

I in Group II of the periodic table
II a transition element
III in Period 4 of the periodic table
IV a metal

A I and III only
B I, II and IV only
C I, III and IV only
D III and IV only

3 Which element shows very similar chemical properties to barium?

A sodium
B aluminium
C sulfur
D magnesium

4 X is an element below calcium in Group II of the periodic table. X would be expected to:

I react very vigorously with cold water
II ionise more readily than calcium
III react slowly with hydrochloric acid

A I only
B I and II only
C II and III only
D I, II and III

5 In which group of the periodic table are the halogens found?

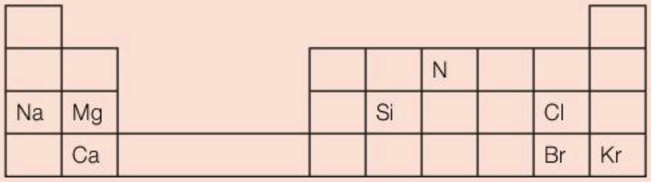
A Group I
B Group II
C Group VII
D Group 0

6 Which of the following is the most electronegative?

A chlorine
B silicon
C sulfur
D phosphorus

Structured question

7 The position of certain elements in the periodic table is given in the figure below. Use the information in the figure to answer the following questions.



The periodic table showing the position of some elements

a **i)** Name two elements which occur in the same period. *(1 mark)*
ii) State the reason for placing these two elements in the same period. *(1 mark)*

b **i)** Name two elements which occur in the same group. *(1 mark)*
ii) State the reason for placing these two elements in the same group. *(1 mark)*

c Element X has 3 valence electrons and 4 electron shells. Place element X in its correct position in the figure above. *(1 mark)*

d Give the electronic configuration of a calcium atom. *(1 mark)*

e Both magnesium and calcium react with hydrochloric acid.
i) Which element would you expect to react more vigorously? *(1 mark)*
ii) Explain your answer to **e i)** above. *(2 marks)*

f **i)** What would you expect to observe if an aqueous solution of chlorine is added to an aqueous solution of potassium bromide? *(1 mark)*
ii) Explain the reason for your observation in **f i)** above. *(2 marks)*

g Explain the reason for the fact that, when moving from left to right along Period 3, the metallic nature of the elements decreases and the non-metallic nature increases. *(3 marks)*

Total 15 marks

Extended response question

8 **a** Outline the contributions of Döbereiner and Mendeleev to the development of the periodic table. *(4 marks)*

b The reactivity of elements in Group II of the periodic table increases going down the group. Support this statement by reference to the reactions of magnesium and calcium with oxygen, water and hydrochloric acid. *(6 marks)*

c Element X was found to displace iodine from potassium iodide solution, but was found to have no reaction with potassium chloride solution.
i) Place the following elements in order as they would be found going down Group VII: element X, chlorine, iodine, fluorine. *(2 marks)*
ii) Explain, in terms of oxidising strength, why X was able to displace iodine from the potassium iodide solution. *(3 marks)*

Total 15 marks

A5 Structure and bonding

Most substances occur as compounds in nature. Compounds are formed when atoms combine chemically with each other. The compounds which are formed are more stable than the free elements were. The noble gases are an exception to this; they do occur in nature as atoms. The electronic configuration of the noble gases is the reason behind their stability.

Objectives

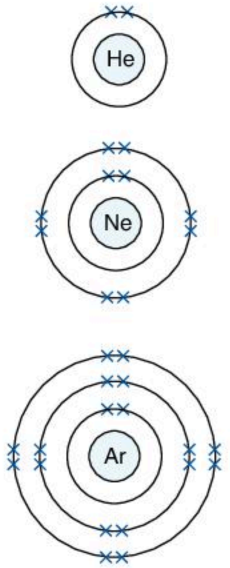
By the end of this topic you will be able to:

- explain why atoms form chemical bonds
- state how atoms form chemical bonds
- name the three types of chemical bonding
- describe three types of chemical formulae
- work out the empirical formulae of binary compounds from their names.

A5.1 Chemical bonding

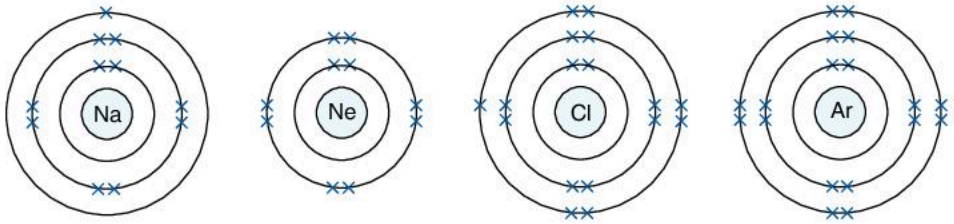
Introduction to chemical bonding

When atoms bond with each other, only the outermost electrons, known as **valence electrons**, are usually involved in the process. Atoms combine with each other to attain the most stable electronic configuration (Unit 3.2). The noble gases are stable atoms because their outermost electron shell is full and this gives them a stable electronic configuration. This is called the noble gas configuration. Atoms, therefore, bond with each other to fill their outermost electron shell or valence shell. Figure 5.1.1 shows the electronic configuration of the first three noble gases.



▲ **Figure 5.1.1** The electronic structures of helium, neon and argon

Atoms gain stability if they attain the noble gas configuration of the nearest noble gas to them in the periodic table, i.e. the element that is closest by atomic number. For example, the nearest noble gas to sodium is neon and the nearest noble gas to chlorine is argon. This is illustrated in Figure 5.1.2.



▲ **Figure 5.1.2** Comparing the electronic structure of sodium and chlorine to their nearest noble gases, neon and argon

Atoms can achieve a noble gas configuration by bonding with other atoms during a chemical change. This can occur in *three* ways. They can:

- donate one or more valence electrons to another atom
- gain one or more valence electrons from another atom
- share valence electrons with another atom.

Key fact

A **chemical bond** is a force of attraction between atoms that results from the redistribution of their electrons.

During chemical bonding, **chemical bonds** are formed between the atoms involved. These bonds are forces of attraction between the atoms which are caused by the redistribution of their electrons, and they join the atoms together. There are *three* main types of chemical bonding:

- **ionic bonding**, which occurs when a metal bonds with a non-metal
- **covalent bonding**, which occurs when two or more non-metals bond
- **metallic bonding**, which occurs within metals.

Ionic and covalent bonding both result in the formation of chemical compounds.

Chemical formulae of compounds

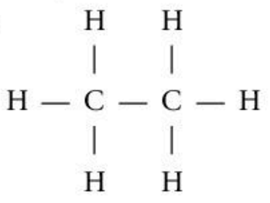
Just as elements can be represented by atomic symbols, compounds formed by ionic or covalent bonding can be represented by **chemical formulae**. Chemical formulae are important because they tell us which elements are found in a compound as well as the ratio between the elements in the compound. There are various types of chemical formulae. The three main types are given below.

- The **molecular formula**, which uses subscripts to give the actual number of atoms of each element present in one molecule of a compound. For example, the molecular formula of water is H_2O, which means one molecule of water contains 2 hydrogen atoms and 1 oxygen atom. Other molecular formulae include CO_2 for carbon dioxide and $C_6H_{12}O_6$ for glucose.
- The **structural formula**, which is a diagrammatic representation of one molecule of the compound using lines between the atoms to represent bonds. For example, the structural formula for carbon dioxide is $O{=}C{=}O$ which shows us that there are two bonds between each oxygen atom and the carbon atom.
- The **empirical formula**, which gives the simplest whole number ratio between the elements in the compound using subscripts. The compound consists of multiples of these smallest units. For example, the empirical formula for magnesium chloride is $MgCl_2$, which tells us that magnesium and chlorine are present in a ratio of 1 to 2.

Example

The three types of chemical formulae for ethane are:.

- Molecular formula: C_2H_6
- Structural formula:



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

- Empirical formula: CH_3

Key fact

Valence number (valency) is the maximum number of bonds an atom can form when it bonds with other atoms.

Writing empirical formulae of binary compounds

A **binary compound** is composed of two different elements only. We can write the empirical formulae of binary compounds using the concept of **valence number** or **valency**, which is the maximum number of bonds an atom can form.

The number of valence electrons an atom has determines its valence number. This is because it is the valence electrons which participate in the formation of chemical bonds. Since atoms lose, gain or share their valence electrons in forming bonds, the valence number can be thought of as the number of electrons an atom has to lose, gain or share during bonding. The valence number of an element is related to its position in the periodic table as shown in Table 5.1.1.

▼ **Table 5.1.1** Valence numbers

Group number	I	II	transition metals	III	IV	V	VI	VII	0
Valence number	1	2	variable, often 2	3	4	3	2	1	0

In other words:

- valence number of elements in Groups I to IV = the group number
- valence number of elements in Groups V to VII = 8 − the group number.

Since both types of atom forming a compound must lose, gain or share the same number of valence electrons, the sum of the valence numbers of each element in the compound must be equal.

Example

The chemical formula of aluminium oxide is $\mathbf{Al_2O_3}$

- Number of atoms of each element present: Al = 2,
  O = 3.
- Valence number of each element: Al = 3 (Group III),
  O = 2 (Group VI).
- Sum of the valence numbers: Al = $\mathbf{2} \times 3 = 6$,
  O = $\mathbf{3} \times 2 = 6$,
  i.e. the sum of the valence numbers of the two elements is equal.

To write the chemical formula of a binary compound, follow the steps given below.

1) Determine the valence numbers of the two elements present.
2) Write down the symbol of the first element. If a metal is present, write its symbol first.
3) Write the valence number of the second element as a subscript after the symbol of the first element.
4) Write the symbol of the second element directly after the subscript.
5) Write the valence number of the first element as a subscript after the symbol of the second element.

Points to note

- There must be no spaces between the symbols and subscripts.
- If the valence number is one, write only the symbol without the number 1 as a subscript (see example 2).
- Some of the transition metals have more than one valence number. The valence number of these is indicated in the name of the compound by a Roman numeral placed in brackets after the name of the metal, e.g. copper(II) indicates that copper has a valence number of 2 (see example 3).

- The valence number of hydrogen and the transition element, silver, is 1.
- To write the empirical formula, if the ratio of the subscripts is not in its simplest form, then cancel to its simplest form (see example 4).

Examples

1 Magnesium nitride

- Determine the valence numbers of the two elements present:
  Mg = 2 (Group II),
  N = 3 (Group V).
- Write down the symbol of the first element: Mg
- Write the valence number of the second element as a subscript after the symbol of the first element: Mg_3
- Write the symbol of the second element: Mg_3N
- Write the valence number of the first element as a subscript after the symbol of the second element: Mg_3N_2

The empirical formula of magnesium nitride is $\mathbf{Mg_3N_2}$.

2 Potassium sulfide

- Valence numbers of the elements: K = 1 (Group I),
  S = 2 (Group VI).
- Empirical formula of potassium sulfide is $\mathbf{K_2S}$.

3 Iron(III) bromide

- Valence numbers of the elements: Fe = 3 (given in the name),
  Br = 1 (Group VII).
- Empirical formula of iron(III) bromide is $\mathbf{FeBr_3}$.

4 Carbon dioxide

- Valence numbers of the elements: C = 4 (Group IV),
  O = 2 (Group VI).
- Empirical formula of carbon dioxide is $\mathbf{CO_2}$. (C_2O_4 is cancelled to its simplest ratio.)

Summary questions

1 Explain why atoms form compounds.

2 Explain how atoms form compounds.

3 Name the three types of chemical bonding.

4 Using your periodic table to help you, write the formula of each of the following binary compounds:
  a sodium oxide
  b copper(II) bromide
  c tetrachloromethane, a compound formed between carbon and chlorine
  d aluminium sulfide
  e ammonia, a compound formed between nitrogen and hydrogen.

A5.2 Formation of ionic bonds

Ionic bonding involves the transfer of valence electrons from metal atoms to non-metal atoms. Compounds formed by ionic bonding are known as **ionic compounds**.

The principles of ionic bonding are explained as follows.

- Metals are found to the left of the periodic table and most have 1, 2 or 3 valence electrons. If they lose their valence electrons they can attain a noble gas configuration.
- Non-metals are found to the right of the periodic table and most have 5, 6 or 7 valence electrons. If they gain electrons into their valence shell they can attain a noble gas configuration.
- When an atom loses or gains electrons it forms an **ion**. The ions formed are now charged particles because they no longer contain the same number of electrons as the number of protons.
- When a metal atom loses electrons, the ion formed has more protons than electrons. The ion has a positive charge and is called a **cation**.
- When a non-metal atom gains electrons, the ion formed has more electrons than protons. The ion has a negative charge and is called an **anion**.

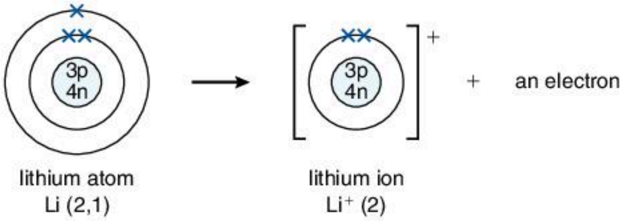
Objectives

By the end of this topic you will be able to:

- explain the principles of ionic bonding
- explain how metal atoms form cations and non-metal atoms form anions
- name ionic compounds
- describe the formation of ionic bonds
- represent ionic bonding using shell diagrams
- deduce the empirical formulae of ionic compounds
- describe the structure of a sodium chloride crystal.

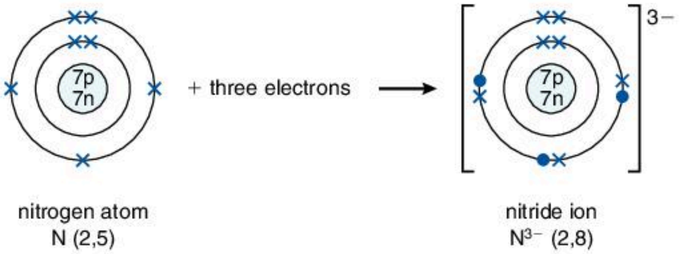
Examples

1 **Lithium** is in Group I of the periodic table and has *one* valence electron. When bonding with a non-metal, lithium loses its valence electron and now has one more proton than electron. Therefore it carries a single positive charge, +, and is represented as Li^+. It is known as the lithium ion or Li^+ ion.



▲ **Figure 5.2.1** Formation of a lithium ion, Li^+

2 **Nitrogen** is in Group V of the periodic table and has *five* valence electrons. When bonding with a metal, nitrogen gains three electrons and now has three more electrons than protons. Therefore it carries a triple negative charge, 3−, and is represented as N^{3-}. It is known as the nitride ion or N^{3-} ion.



▲ **Figure 5.2.2** Formation of a nitride ion, N^{3-}

Key fact

Ionic bonds are chemical bonds created by the electrostatic forces of attraction between the positively charged cations and negatively charged anions in ionic compounds.

When naming the negative anions formed when non-metal atoms gain electrons, the name changes from that of the atom. The name of the anion always ends in '**ide**'. For example, the chloride ion (Cl^-), the oxide ion (O^{2-}) and the bromide ion (Br^-).

In the compounds formed by ionic bonding, there are strong electrostatic forces of attraction between the positive cations and negative anions. These forces of attraction are known as ionic bonds and they hold the ions together. The strength of the ionic bonds accounts for the properties of ionic compounds (Unit 5.5).

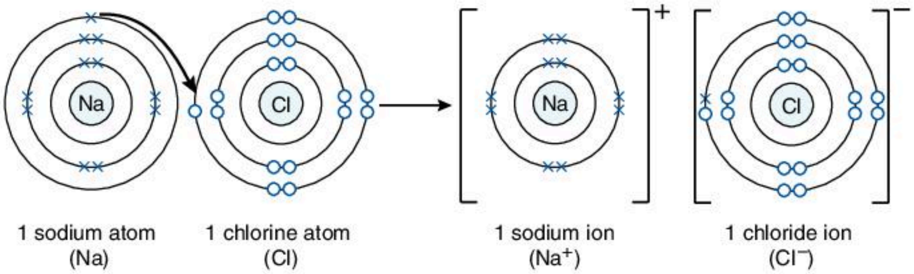
Examples to show the formation of ionic compounds

The formation of sodium chloride, magnesium fluoride, aluminium oxide and potassium nitride by ionic bonding is shown next by means of dot and cross diagrams.

Sodium chloride

Sodium is a metal with *one* valence electron. It loses this electron to form a positive sodium ion which has a single positive charge, i.e. Na^+. Chlorine is a non-metal with *seven* valence electrons. It gains one electron to form a negative chloride ion which has a single negative charge, i.e. Cl^-. For each sodium atom that loses one electron there will need to be one chlorine atom to accept this electron. The empirical formula of sodium chloride is, therefore, **NaCl**.

Figure 5.2.3 shows how to use shell diagrams to represent the ionic bonding in sodium chloride.



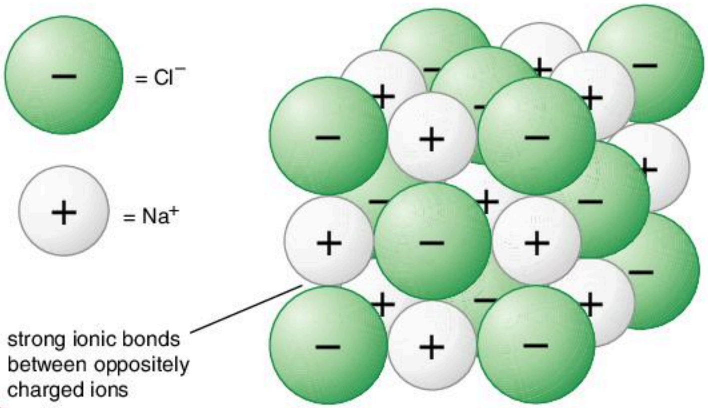
▲ **Figure 5.2.3** Ionic bonding in sodium chloride

Did you know?

The term '**lattice**' refers to a regular, three-dimensional arrangement of particles.

It is important to note that NaCl is not a separate, individual entity. NaCl simply represents the ratio of Na^+ ions to Cl^- ions in sodium chloride, i.e. it is the empirical formula. It can also be called the formula unit of sodium chloride.

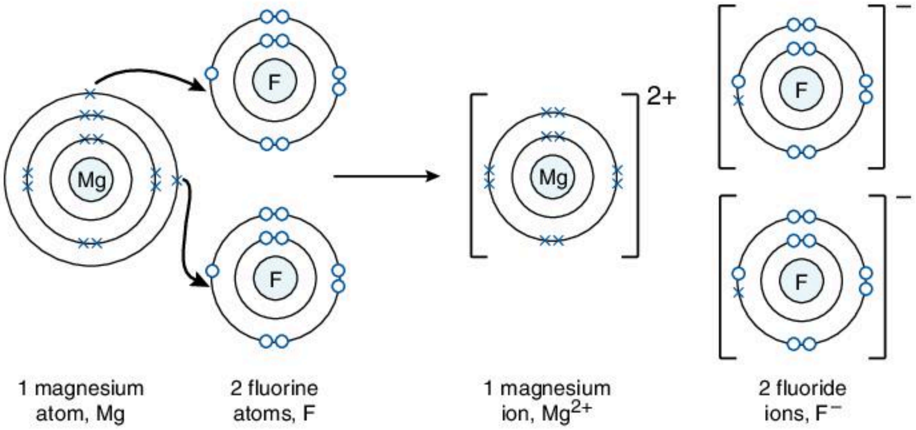
Sodium chloride that we sprinkle on our food is composed of tiny crystals that are cubic in shape. Each of these crystals is made up of millions of Na^+ ions and millions of Cl^- ions in a ratio of one Na^+ ion to one Cl^- ion. These ions are arranged in a very regular way; they occur alternately in rows and columns. Each Na^+ ion is bonded to *six* Cl^- ions and each Cl^- ion is bonded to *six* Na^+ ions by strong ionic bonds. This three-dimensional, ordered arrangement of ions forms a crystal lattice.



◀ **Figure 5.2.4** Crystal lattice of sodium chloride

Magnesium fluoride

Magnesium is a metal with *two* valence electrons. It loses these electrons to form a positive magnesium ion which has a double positive charge, i.e. Mg^{2+}. Fluorine is a non-metal with *seven* valence electrons. It gains one electron to form a negative fluoride ion which has a single negative charge, i.e. F^-. For each magnesium atom that loses two electrons there will need to be two fluorine atoms to accept these two electrons since each fluorine atom only accepts one electron. The empirical formula or formula unit of magnesium fluoride is, therefore, $\mathbf{MgF_2}$.

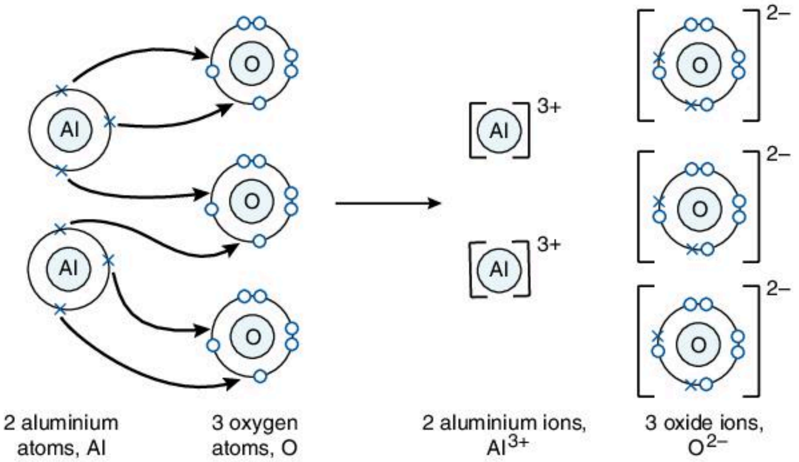


▲ **Figure 5.2.5** Ionic bonding in magnesium fluoride

Aluminium oxide

Aluminium is a metal with *three* valence electrons. It loses these electrons to form a positive aluminium ion which has a triple positive charge, i.e. Al^{3+}. Oxygen is a non-metal with *six* valence electrons. It gains two electrons to form a negative oxide ion which has a double negative charge, i.e. O^{2-}. Since each aluminium atom has to lose three electrons and each oxygen atom has to gain two electrons, two aluminium atoms will be needed to lose a total of six electrons and three oxygen atoms will be needed to gain these six electrons. The empirical formula or formula unit of aluminium oxide is, therefore, $\mathbf{Al_2O_3}$.

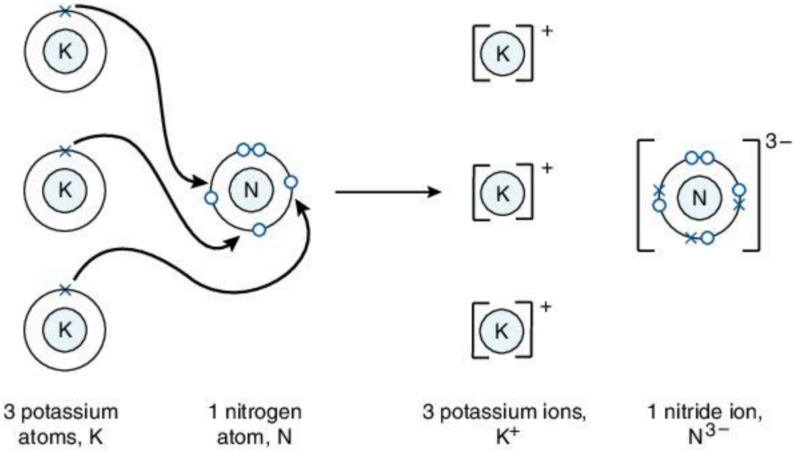
To simplify the diagrams, sometimes only the valence electrons are shown, as in Figure 5.2.6.



▲ **Figure 5.2.6** Ionic bonding in aluminium oxide

Potassium nitride

Potassium has *one* valence electron. It loses this electron to form a positive potassium ion with a single positive charge, i.e. K^+. Nitrogen has *five* valence electrons. It gains three electrons to form a negative nitride ion which has a triple negative charge, i.e. N^{3-}. Three potassium atoms are needed to lose three electrons to one nitrogen atom. The empirical formula or formula unit of potassium nitride is, therefore, $\mathbf{K_3N}$.



▲ **Figure 5.2.7** Ionic bonding in potassium nitride

Summary questions

1 Explain what happens during ionic bonding.

2 Use shell diagrams to represent the ionic bonding between:
  a sodium and oxygen to form sodium oxide, Na_2O
  b magnesium and nitrogen to form magnesium nitride, Mg_3N_2.

3 Calcium forms an ionic compound with fluorine.
  a Give the name of the compound.
  b Give the formula unit of the compound.

4 Describe the structure of the crystal lattice of sodium chloride.

Objectives

By the end of this topic you will be able to:

- give the names and formulae of the common cations and anions
- work out the formula of an ionic compound from the name
- write the name of an ionic compound from the formula.

A5.3 Writing chemical formulae of ionic compounds

Ionic compounds are represented using **empirical formulae**. The empirical formula of an ionic compound is also known as the **formula unit** of the compound. There are important points to note when writing formulae of ionic compounds.

- Metals generally lose electrons forming positive cations. Non-metals generally gain electrons forming negative anions.
- The magnitude of the charge on an ion is the same as the valence number. Therefore, when writing formulae of ionic compounds, the magnitude of the charge on the ion can be used in place of the valence number.
- Not all ions are formed from one atom. Many ions are made from several atoms bonded together. These ions are known as **polyatomic ions** (these were previously known as radicals).

- When writing the formula of an ionic compound, the sum of the positive charges must equal the sum of the negative charges, since both types of atom or group of atoms bonded together must lose or gain the same number of electrons.

The names and formulae of the ions that you are required to know are given in Tables 5.3.1 and 5.3.2.

▼ **Table 5.3.1** Common cations

Monovalent cations		Divalent cations		Trivalent cations	
hydrogen	H^+	magnesium	Mg^{2+}	iron(III)	Fe^{3+}
lithium	Li^+	calcium	Ca^{2+}	aluminium	Al^{3+}
sodium	Na^+	barium	Ba^{2+}		
potassium	K^+	iron(II)	Fe^{2+}		
copper(I)	Cu^+	copper(II)	Cu^{2+}		
silver	Ag^+	zinc	Zn^{2+}		
ammonium	NH_4^+	tin(II)	Sn^{2+}		
		lead(II)	Pb^{2+}		

▼ **Table 5.3.2** Common anions

Monovalent anions		Divalent anions		Trivalent anions	
fluoride	F^-	oxide	O^{2-}	nitride	N^{3-}
chloride	Cl^-	sulfide	S^{2-}	phosphate	PO_4^{3-}
bromide	Br^-	carbonate	CO_3^{2-}		
iodide	I^-	sulfite (sulfate(IV))	SO_3^{2-}		
hydride	H^-	sulfate (sulfate(VI))	SO_4^{2-}		
hydroxide	OH^-	dichromate(VI)	$Cr_2O_7^{2-}$		
nitrite (nitrate(III))	NO_2^-				
nitrate (nitrate(V))	NO_3^-				
manganate(VII)	MnO_4^-				
hydrogensulfate	HSO_4^-				
hydrogencarbonate	HCO_3^-				
ethanoate	CH_3COO^-				

Exam tip

Success in chemistry depends on being able to write the names and formulae of compounds. Knowing the **formulae of ions** is a very important step towards achieving this success.

In order to write the chemical formula of an ionic compound from the name, follow the steps given below.

1) Write down the formulae of the two ions that are involved from Tables 5.3.1 and 5.3.2.
2) Rewrite the formula of the cation, omitting its charge.
3) Write the magnitude of the charge on the anion as a subscript after the formula of the cation. If it is a polyatomic cation and more than one is required, place the polyatomic cation in brackets and write the subscript outside the bracket (see example 4).
4) Write the formula of the anion directly after the subscript, omitting its charge.
5) Write the magnitude of the charge on the cation as a subscript after the formula of the anion. If it is a polyatomic anion and more than one is required, place the polyatomic anion in brackets and write the subscript outside the bracket (see example 5).

Points to note

- If the magnitude of the charge is one, do not write the number 1 as a subscript (see examples 2 and 3).
- If the ratio of the subscripts is not in its simplest form, then cancel to its simplest form (see example 6).

Examples

1 Aluminium sulfide

- Write down the ions involved: Al^{3+}, S^{2-}
- Write down the formula of the cation without its charge: Al
- Write the magnitude of the charge on the anion as a subscript after the formula of the cation: Al_2
- Write the formula of the anion without its charge: Al_2S
- Write the magnitude of the charge on the cation as a subscript after the formula of the anion: Al_2S_3

The formula unit of aluminium sulfide is $\mathbf{Al_2S_3}$.

2 Calcium iodide

- Ions involved: Ca^{2+}, I^-
- Magnitude of the charges: $Ca^{2+} = 2$, $I^- = 1$
- Formula unit of calcium iodide is $\mathbf{CaI_2}$.

3 Sodium phosphate

- Ions involved: Na^+, PO_4^{3-}
- Magnitude of the charges: $Na^+ = 1$, $PO_4^{3-} = 3$
- Formula unit of sodium phosphate is $\mathbf{Na_3PO_4}$.

4 Ammonium sulfate

- Ions involved: NH_4^+, SO_4^{2-}
- Magnitude of the charges: $NH_4^+ = 1$, $SO_4^{2-} = 2$
- Formula unit of ammonium sulfate is $\mathbf{(NH_4)_2SO_4}$.

(The polyatomic ammonium ion, NH_4^+, is placed in brackets with the subscript outside.)

5 Copper(II) nitrate

- Ions involved: Cu^{2+}, NO_3^-
- Magnitude of the charges: $Cu^{2+} = 2$, $NO_3^- = 1$
- Formula unit of copper(II) nitrate is $\mathbf{Cu(NO_3)_2}$.

(The polyatomic nitrate ion, NO_3^-, is placed in brackets with the subscript outside.)

6 Magnesium carbonate

- Ions involved: Mg^{2+}, CO_3^{2-}
- Magnitude of the charges: $Mg^{2+} = 2$, $CO_3^{2-} = 2$
- Formula unit of magnesium carbonate is $\mathbf{MgCO_3}$.

($Mg_2(CO_3)_2$ is cancelled to its simplest ratio.)

Naming ionic compounds

The name of an ionic compound is built from the cation and anion names.

Examples

1 $Al(OH)_3$ is composed of the aluminium cation and hydroxide anion. Its name is aluminium hydroxide.

2 K_2O is composed of the potassium cation and the oxide anion. Its name is potassium oxide.

3 $Ca_3(PO_4)_2$ is composed of the calcium cation and phosphate anion. Its name is calcium phosphate.

If the cation is that of a transition metal, determine the magnitude of its charge by looking at the subscript after the anion and use this to determine its name.

Examples

1 $FeCl_2$

- The subscript after the chloride ion is 2.
- The magnitude of the charge on the iron ion must be 2.
- The cation must be the iron(II) ion, Fe^{2+}.
- The name of the compound is iron(II) chloride.

2 $Fe(NO_3)_3$

- The subscript after the nitrate ion is 3.
- The magnitude of the charge on the iron ion must be 3.
- The cation must be the iron(III) ion, Fe^{3+}.
- The name of the compound is iron(III) nitrate.

Summary questions

1 Write the chemical formula of each of the following ionic compounds:
  a iron(III) sulfide
  b calcium hydroxide
  c sodium sulfate
  d ammonium phosphate
  e silver oxide
  f zinc nitrate
  g potassium manganate(VII)
  h copper(II) carbonate

2 Write names for each of the following compounds:
  a $Mg_3(PO_4)_2$
  b AgCl
  c ZnI_2
  d $Al(HCO_3)_3$
  e $FeBr_3$
  f $FeSO_4$
  g Cu_2O
  h $Pb(NO_3)_2$

A5.4 Formation of covalent and metallic bonds

Objectives

By the end of this topic you will be able to:

- describe the formation of covalent bonds
- represent covalent bonding using shell diagrams
- write formulae of covalent compounds
- describe the formation of metallic bonds
- relate the properties of metals to the bonding in metals.

Key fact

A **covalent bond** is a chemical bond formed by the sharing of a pair of electrons between two atoms.

Key fact

A **molecule** is a group of two or more atoms which are covalently bonded together strongly enough to behave as a single unit in a chemical reaction.

Covalent bonding

Covalent bonding involves the sharing of valence electrons between non-metal atoms. Compounds formed by covalent bonding are known as **covalent compounds**.

Non-metals are found to the right of the periodic table and have 4, 5, 6 or 7 valence electrons. They need to gain electrons into their valence shell to attain a noble gas configuration.

Since all the atoms involved in covalent bonding need to gain electrons, they achieve this by approaching each other so that their outermost electron shells overlap. Any valence electrons that were not in pairs in the original atoms are then shared, in pairs, between the two overlapping atoms. Each shared pair of electrons forms a **covalent bond** and the electrons in a covalent bond are known as a **bonding pair**. Each bonding pair spends time with both atoms, orbiting around each nucleus. Each atom now has a completed outer electron shell and is stable. Two atoms may share one, two or three pairs of electrons between them.

Two or more non-metal atoms can bond together by sharing electron pairs. When they do this, they form separate, individual entities known as **molecules**. The atoms within a molecule are held together by the strong electrostatic forces of attraction between the nuclei of the atoms and the bonding pair or pairs of electrons. Atoms of the same element or of different elements may bond together by covalent bonding.

Examples to show the formation of covalent substances

The formation of a variety of substances by covalent bonding is shown next by means of dot and cross diagrams.

Chlorine gas

Chlorine is a non-metal with *seven* valence electrons. It needs one more electron to fill its valence electron shell and attain a noble gas configuration. During the formation of a chlorine molecule, two chlorine atoms come close enough for their outer electron shells to overlap. The two atoms share one pair of electrons between them. The shared pair orbit around both atoms forming a **single covalent bond** between the two atoms. The formula for chlorine is $\mathbf{Cl_2}$.

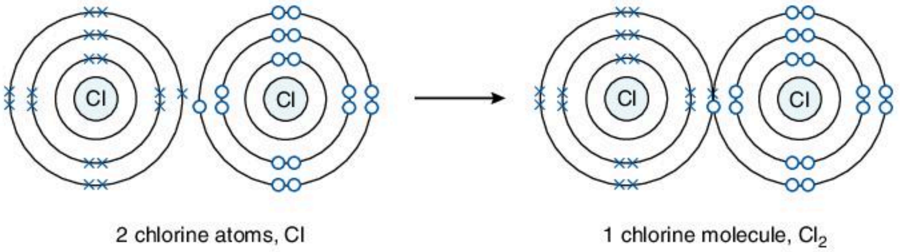
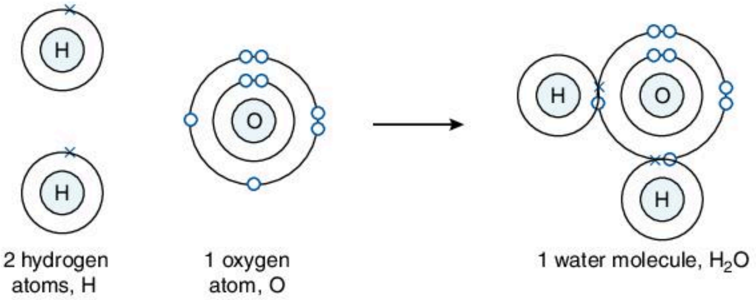


Figure 5.4.1 Covalent bonding in chlorine

Water

Water is a compound formed from hydrogen and oxygen atoms. Hydrogen is a non-metal with *one* valence electron. It needs one more electron to fill its valence electron shell. Oxygen is a non-metal with *six* valence electrons. It needs two more electrons to fill its valence electron shell. Two hydrogen atoms are needed to provide the two electrons required by oxygen. Each hydrogen atom shares one electron pair with the oxygen atom. The formula for water is $\mathbf{H_2O}$.



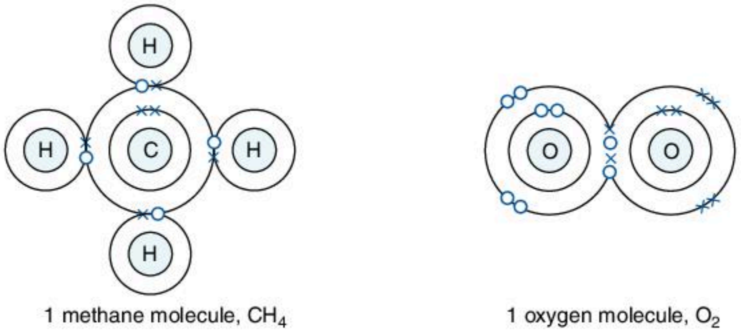
▲ **Figure 5.4.2** Covalent bonding in water

Methane

Methane is a compound formed from carbon and hydrogen. Carbon is a non-metal with *four* valence electrons. It needs four electrons to fill its valence electron shell. Hydrogen needs one electron to fill its valence electron shell. Four hydrogen atoms are needed to provide the four electrons required by carbon. Each hydrogen atom shares one electron pair with the carbon atom. The formula for methane is $\mathbf{CH_4}$.

Oxygen gas

Oxygen needs two electrons to fill its valence electron shell. During the formation of an oxygen molecule, two oxygen atoms come close enough for their outer electron shells to overlap. The two atoms share two pairs of electrons between them, forming a **double covalent bond**. The formula for oxygen is $\mathbf{O_2}$.



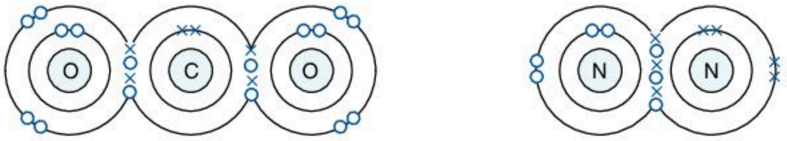
▲ **Figure 5.4.3** Covalent bonding in methane and oxygen

Carbon dioxide

Carbon needs four electrons and oxygen needs two electrons to fill their valence electron shells. Two oxygen atoms are needed to provide the four electrons required by carbon. Each oxygen atom shares two electron pairs with the carbon atom, forming a double covalent bond. The formula for carbon dioxide is $\mathbf{CO_2}$.

Nitrogen

Nitrogen is a non-metal with *five* valence electrons. It needs three more electrons to fill its valence electron shell. During the formation of a nitrogen molecule, two nitrogen atoms come close enough for their outer electron shells to overlap. The two atoms share three pairs of electrons between them, forming a **triple covalent bond**. The formula for nitrogen is $\mathbf{N_2}$.

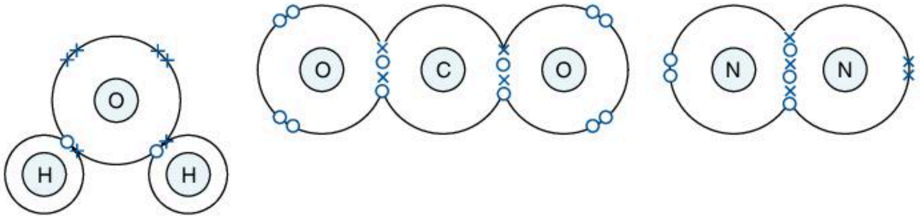


1 carbon dioxide molecule, CO_2

1 nitrogen molecule, N_2

▲ **Figure 5.4.4** Covalent bonding in carbon dioxide and nitrogen

The drawing of dot and cross diagrams to represent covalent bonding can be simplified by showing only the valence electrons. Examples of this type of shell diagram are shown in Figure 5.4.5.



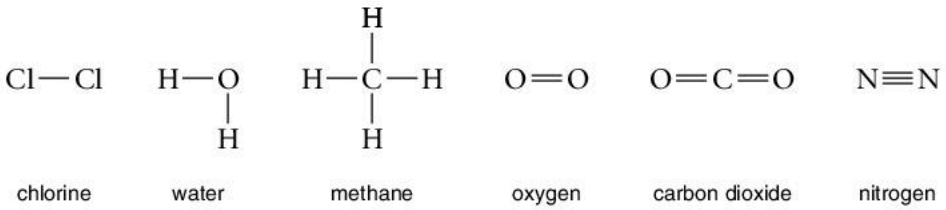
1 water molecule

1 carbon dioxide molecule

1 nitrogen molecule

▲ **Figure 5.4.5** Covalent bonding in water, carbon dioxide and nitrogen showing valence electrons only

Structural formulae may also be used to represent molecules in which each covalent bond is shown by a line between the two atoms. Examples of structural formulae of molecules are given in Figure 5.4.6.



$$\mathrm{Cl{-}Cl} \qquad \begin{array}{c}\mathrm{H{-}O}\\ \quad| \\ \quad\mathrm{H}\end{array} \qquad \begin{array}{c}\mathrm{H}\\ |\\ \mathrm{H{-}C{-}H}\\ |\\ \mathrm{H}\end{array} \qquad \mathrm{O{=}O} \qquad \mathrm{O{=}C{=}O} \qquad \mathrm{N{\equiv}N}$$

▲ **Figure 5.4.6** Structural formulae of chlorine, water, methane, oxygen, carbon dioxide and nitrogen

In Unit 5.1 you learnt that the number of bonds that an atom can form when it bonds with other atoms is known as its **valence number**. This is shown clearly by these dot and cross diagrams. Hydrogen has a valence number of one and the diagrams show it forming one covalent bond. Oxygen has a valence number of two and the diagrams show it forming two covalent bonds. The valence numbers and number of covalent bonds formed by the common non-metals when bonding to form covalent compounds are given in Table 5.4.1.

Polar and non-polar covalent molecules

Molecules may be polar or non-polar as a result of the **electronegativity** of the atoms in the molecule. Electronegativity is a measure of how strongly an atom attracts electrons. Oxygen, nitrogen, fluorine and chlorine are more strongly electronegative than atoms of other elements.

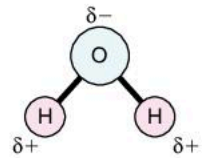
Exam tip

It is very important that you are able to draw **dot and cross diagrams** to show the bond formation in compounds. To do this, you must first decide if the compound is ionic or covalent. If the compound is formed from a metal and a non-metal it will be ionic, if it is formed from non-metals only it will be covalent. Remember that a metal usually has 1, 2 or 3 valence electrons and a non-metal usually has 4, 5, 6 or 7 valence electrons.

Element	Valence number	Number of covalent bonds formed
hydrogen	1	1
fluorine	1	1
chlorine	1	1
bromine	1	1
iodine	1	1
oxygen	2	2
sulfur	2	2
nitrogen	3	3
carbon	4	4

▲ **Table 5.4.1** Valence number and number of covalent bonds formed by non-metal atoms

- In a **polar molecule**, the atoms at each side of a covalent bond attract the electrons in the bond with different strengths. The result is that one side of the molecule has a slightly positive charge and one side has a slightly negative charge. Water (H_2O), ammonia (NH_3), hydrogen chloride (HCl), ethanol (C_2H_5OH) and glucose ($C_6H_{12}O_6$) molecules are all polar.
- In a **non-polar molecule**, the atoms at each side of a covalent bond attract electrons with equal strengths such that the molecule does not have any charged sides. Oxygen (O_2), nitrogen (N_2), chlorine (Cl_2), carbon dioxide (CO_2) and methane (CH_4) molecules are all non-polar.



▲ **Figure 5.4.7** A polar water molecule

Formulae of covalent compounds

The formula of a covalent compound gives the exact number of atoms of each element in one molecule of the compound, i.e. it is the **molecular formula**. In some covalent compounds the molecular formula and the ratio of elements are the same, e.g. water (H_2O), carbon dioxide (CO_2), ammonia (NH_3) and methane (CH_4). In these compounds the molecular and empirical formulae are the same. In other compounds they are not the same, e.g. the molecular formula of glucose is $C_6H_{12}O_6$ but its empirical formula is CH_2O.

The name of a covalent compound can sometimes indicate its molecular formula. This is done by using the prefixes di, tri, tetr, and so on. These are placed before the name of the element to which they are referring. These prefixes are given in Table 5.4.2.

Prefix	Number of atoms
mono	1
di	2
tri	3
tetr	4
pent	5
hex	6

▲ **Table 5.4.2** Prefixes sometimes used in naming covalent compounds

Examples

- CO is the formula for carbon monoxide.
- SO_2 is the formula for sulfur dioxide.
- SO_3 is the formula for sulfur trioxide.
- CCl_4 is the formula for carbon tetrachloride, more correctly known as tetrachloromethane.
- N_2O_4 is the formula for dinitrogen tetroxide.

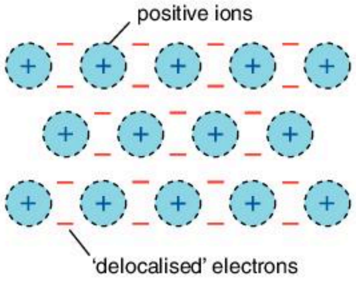
It is a good idea to learn the molecular formulae of covalent compounds as you come across them, especially since some cannot be derived from their names, e.g. ammonia (NH_3) and methane (CH_4).

Metallic bonding

Metals have very distinct chemical and physical properties as a result of the type of bonding that occurs between the metal atoms. The type of bonding between metal atoms in a metal is known as **metallic bonding**.

The atoms in a metal are packed very closely together to form a **metal lattice**. Because of this tight packing, the outer electron shells overlap and the valence electrons become **delocalised**, i.e. they are no longer associated with any particular atom. These delocalised electrons, also known as mobile electrons, form a 'sea' of electrons which move around within the metal lattice and belong to the lattice as a whole. Positive metal cations are formed from the metal atoms when the electrons become delocalised. The metal lattice is held together by the strong electrostatic force of attraction, known as the **metallic bond**, which exists between the positive cations and the moving, delocalised, negative electrons. Figure 5.4.8 illustrates the bonding in metals.

It is important to note that metals are still made up of atoms and not ions because the valence electrons have not been lost. They are still there in the structure even though they are no longer attached to a particular atom.



▲ **Figure 5.4.8** The structure of metals

Key fact

The **metallic bond** is the chemical bond created by the electrostatic force of attraction between the positive cations and the moving, delocalised negative electrons.

The way in which the metal atoms are bonded in metal lattices helps us to understand some of the properties of metals given in Unit 4.4. These are explained below.

- Metals have high melting points and boiling points because of the strong forces of attraction between the cations and delocalised electrons. Large amounts of heat energy are required to separate the atoms.
- Metals have high densities because the atoms are packed very closely together.
- Metals are good conductors of electricity because the delocalised electrons are free to move. The moving electrons act as charge carriers, which allows an electric current to be carried through the metal.
- Metals are good conductors of heat because of the delocalised electrons. These moving electrons act as heat carriers, which allows heat to be carried through the metal.
- Metals are malleable and ductile because the atoms in a metal are all of the same type, therefore, all of the same size. If force is applied to the metal, the atoms can roll over each other into new positions without breaking the metallic bond.

Summary questions

1 Use shell diagrams to represent covalent bonding in the following:
  a hydrogen, H_2
  b phosphorus trifluoride, PF_3
  c hydrogen bromide, HBr
  d carbon disulfide, CS_2.
2 Draw the shell diagram of the covalent bonding between carbon and chlorine and give the formula of the compound.
3 Explain the terms 'delocalised electron' and 'metallic bond'.
4 Explain why metals are electrical conductors.

A5.5 Structure and properties of solids

Objectives

By the end of this topic you will be able to:

- describe the structure and give examples of ionic crystals
- relate the properties of sodium chloride to its structure
- describe the structure and give examples of simple molecular crystals
- distinguish between ionic and simple molecular crystals
- describe the structure and give examples of giant molecular crystals
- describe the structure of diamond and graphite
- relate the properties and uses of diamond and graphite to their structures
- explain the term allotropy.

The properties of an element or compound, whether it is composed of atoms, molecules or ions, depend on the forces of attraction between the individual particles. This means that the properties depend on the type of chemical bond between the particles. In addition, the properties depend on the arrangement of the particles, for example, if the particles are packed very tightly together, the element or compound will have a high density.

When considering solids, we can identify *four* different structures based on the way in which the particles are bonded within the solid. They are known as **crystals** because their particles are arranged in a three-dimensional, ordered pattern throughout the structure. The four types of solid are:

- ionic crystals
- simple molecular crystals
- giant molecular crystals
- metallic crystals.

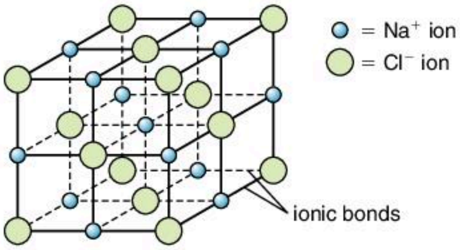
We have already studied the structure and properties of metals in Unit 5.4. We will be studying the structure and properties of the other three in this unit.

Ionic crystals

Ionic crystals are solids formed as a result of ionic bonding. An ionic crystal is composed of an **ionic lattice** in which the cations and anions are held together in a regular, repeating, three-dimensional arrangement by strong electrostatic forces of attraction called **ionic bonds**. Ionic crystals are represented by empirical formulae or formula units. All compounds formed by ionic bonding have an ionic crystal structure, for example, sodium chloride.

The structure and properties of sodium chloride

In Unit 5.2 you studied the ionic bonding between sodium and chlorine to form sodium chloride. You may remember that the sodium chloride crystal lattice is composed of millions of sodium ions or Na^+ ions and millions of chloride ions or Cl^- ions held together by strong ionic bonds. Also that the ions are in a ratio of 1 Na^+ ion to 1 Cl^- ion and that the empirical formula or formula unit for sodium chloride is NaCl.

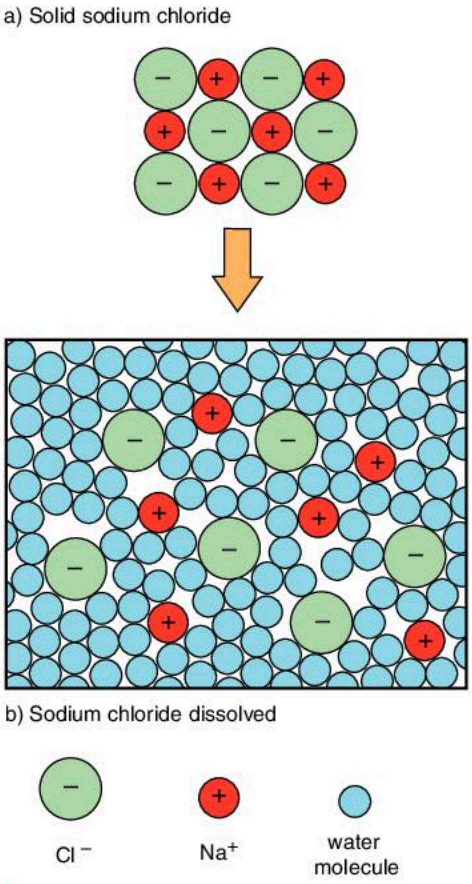


▲ **Figure 5.5.1** Crystal lattice of sodium chloride

The properties of sodium chloride can be explained by its structure, i.e. the arrangement and bonding of its particles, as summarised in Table 5.5.1.

▼ **Table 5.5.1** Properties of sodium chloride

Property	Explanation
High melting point – about 801 °C	The high melting point is due to the very strong electrostatic forces between the sodium ions and chloride ions. A large amount of heat energy is needed to weaken these forces and separate the ions from each other, allowing the solid to melt.
Hard and brittle	The sodium and chloride ions are closely packed and held together by strong electrostatic forces making the solid hard. When pressure is applied to the crystal lattice, the ion layers move slightly with respect to each other and ions with the same charges come to lie next to each other. Repulsion occurs between the like charges and the lattice breaks apart, i.e. it is brittle.
Soluble in water	When sodium chloride is added to water the ions can separate resulting in sodium chloride being soluble in water. The partial positive ends of the polar water molecules attract the negative chloride ions and the partial negative ends attract the positive sodium ions. This pulls the ions out of the lattice and the crystal dissolves. The ions become surrounded by the water molecules and new forces of attraction now exist between the ions and the water molecules.
Conducts electricity when molten, i.e. NaCl(l), or dissolved in water, i.e. NaCl(aq)	For a substance to conduct an electric current it must contain charged particles which are able to move. In molten (melted) sodium chloride, the ions are no longer held together by ionic bonds and they are able to move allowing sodium chloride to conduct electricity in the liquid state. When sodium chloride dissolves in water, the ions move freely throughout the water, allowing an aqueous solution of sodium chloride to conduct an electric current.



▲ **Figure 5.5.2** What happens when sodium chloride dissolves in water

Simple molecular crystals

Simple molecular crystals are solids composed of small molecules. Each molecule is composed of only a few atoms bonded together by strong covalent bonds. These molecules are then arranged in a regular, three-dimensional way to create a **simple molecular lattice**. The molecules within the lattice have weak forces of attraction between them known as **intermolecular forces**. These intermolecular forces hold the small molecules together. Molecular formulae are used to represent simple molecular crystals.

Examples of simple molecular crystals include:

- ice, composed of water molecules, molecular formula H_2O
- iodine, composed of iodine molecules, molecular formula I_2
- dry ice, composed of carbon dioxide molecules, molecular formula CO_2
- glucose, composed of glucose molecules, molecular formula $C_6H_{12}O_6$.

Distinguishing between ionic and simple molecular solids

Ionic and simple molecular solids have very different properties because of the different way in which their particles are bonded. These differences are summarised in Table 5.5.2.

▼ **Table 5.5.2** Comparison of the properties of ionic and simple molecular solids

Property	Ionic solids	Simple molecular solids
	Examples: NaCl, KBr	Examples: I_2, CO_2, H_2O
Type of chemical bond	Ionic bonds are found between ions. These are very strong.	Covalent bonds occur between the atoms in the molecules. These bonds are very strong. Intermolecular forces of attraction are found between the molecules. These are weak.
Melting point	Have a **high** melting point because of the strong ionic bonds between the ions. A large amount of heat energy is needed to weaken these bonds and separate the ions from each other, allowing the solid to melt.	Have a **low** melting point because of the weak intermolecular forces between the molecules. Not much heat energy is needed to weaken these forces and separate the molecules from each other, allowing the solid to melt. Some simple molecular solids sublime when heated, e.g. iodine and carbon dioxide.
Solubility	Due to being composed of positive and negative ions, most are soluble in polar solvents, e.g. water, but are insoluble in non-polar solvents, e.g. kerosene, tetrachloromethane, gasolene and other organic solvents.	Follow the rule 'like dissolves like'. Polar substances dissolve in polar solvents, e.g. glucose dissolves in water. Non-polar substances dissolve in non-polar solvents, e.g. iodine dissolves in tetrachloromethane.
Conductivity	Do not conduct electricity when solid because the ions are held together by strong ionic bonds and are not free to move. They do conduct electricity when molten (melted) or dissolved in water because the ions are no longer held together by ionic bonds and they are free to move, which allows an electric current to be carried.	Do not conduct electricity in any state because they do not have any charged particles which are free to move.

Practical activity

Investigating the solubility and electrical conductivity of ionic and simple molecular solids

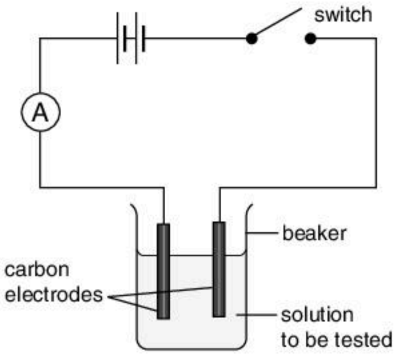
Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

You will be supplied with sodium chloride (ionic solid), glucose and iodine (simple molecular solids), distilled water (polar solvent), tetrachloromethane (non-polar solvent) and the apparatus used to test electrical conductivity shown in Figure 5.5.3.

Method

1. Investigate the solubility of sodium chloride, iodine and glucose by following the instructions below.
   - Place 5 cm^3 of distilled water into each of three test tubes.
   - Add a small spatula of sodium chloride to the first tube, a small spatula of glucose to the second tube and a crystal of iodine to the third tube.
   - Shake the tubes thoroughly and observe if each substance dissolves or not.
   - Repeat the experiment using tetrachloromethane instead of distilled water.
2. Investigate the electrical conductivity of solutions using the apparatus shown in Figure 5.5.3 and by following the instructions below.
   - Make three solutions by adding three small spatulas of sodium chloride to 25 cm^3 of distilled water, three small spatulas of glucose to 25 cm^3 of distilled water and one crystal of iodine to 25 cm^3 of tetrachloromethane in three separate beakers.
   - Place the electrodes into the sodium chloride solution and close the switch. If the ammeter registers a reading, the solution conducts an electric current, if it does not register a reading, then the solution does not conduct an electric current. Remove the electrodes and clean them, and test the other two solutions for electrical conductivity.
3. Draw up a table to record all your results.
4. What can you conclude about:
   - the type of solid (ionic or simple molecular) and its solubility
   - the type of solid and its electrical conductivity?



▲ **Figure 5.5.3** Apparatus to measure the electrical conductivity of a solution

Giant molecular crystals

Giant molecular crystals are composed of non-metal atoms bonded by strong **covalent bonds** in a regular, three-dimensional arrangement to form a **giant molecular lattice**. The covalent bonds exist between the atoms throughout the lattice such that a giant molecular crystal is also known as a **macromolecule**, i.e. it is a molecule composed of millions of atoms. The empirical formula is used to represent giant molecular crystals.

Examples of giant molecular crystals include:

- diamond, composed of carbon atoms, empirical formula C
- graphite, also composed of carbon atoms, empirical formula C
- silicon dioxide, composed of silicon and oxygen atoms, empirical formula SiO_2.

Did you know?

Silicon dioxide, also known as silica, is the most abundant mineral in the Earth's crust. It is a component of almost every rock type and is the main component of sand. Almost all of this silica occurs in a crystalline form known as quartz. Quartz that is almost pure forms large and beautiful crystals. These are very popular as ornamental stones and semi-precious gemstones. For example, amethyst is purple quartz, citrine is yellow orange and rose quartz is a pink to rose-red variety.

Structure and properties of diamond

Diamond is composed of carbon atoms. Each atom is bonded covalently to *four* other carbon atoms arranged in a *tetrahedron* around it. A crystal of diamond consists of millions of these carbon atoms bonded to one another throughout the crystal by strong covalent bonds. Because of the strength of the bonds, diamond is extremely hard and has a very high melting point. The properties of diamond can be explained by its structure, as summarised in Table 5.5.3.

▼ **Table 5.5.3** Properties of diamond

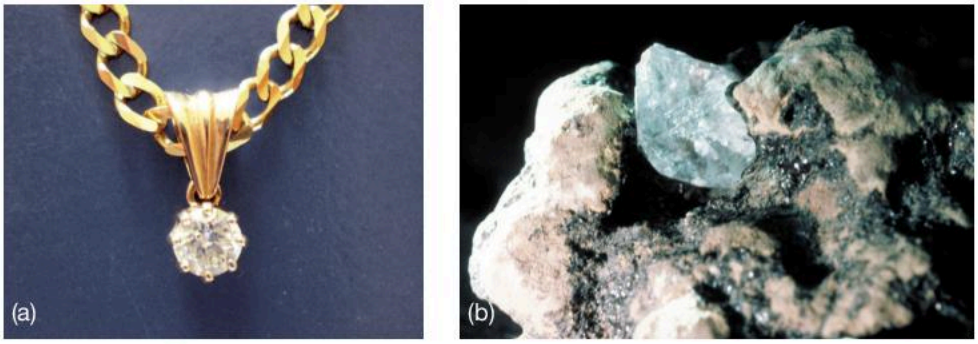
Property	Explanation
Very high melting point – about 3550 °C	The covalent bonds between the carbon atoms are very strong. A very large amount of heat energy is needed to weaken these forces and separate the atoms from each other, allowing the solid to melt.
Extremely hard	Diamond is the hardest naturally occurring solid. The hardness is due to the high strength of the covalent bonds between the carbon atoms. This property is very useful in industry as diamond is often used in cutting and drilling.
Does not conduct electricity	The valence electrons of the carbon atoms are all shared between the atoms. Carbon does not have any 'free' electrons to conduct an electric current.

? Did you know?

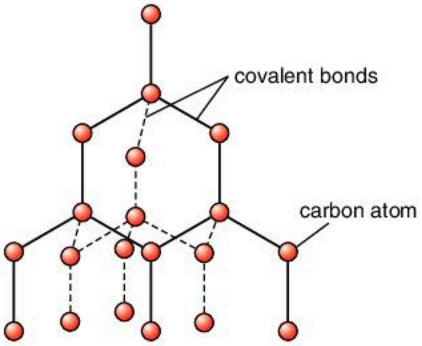
The word '**diamond**' comes from the ancient Greek word 'adámas', which means 'unbreakable'. Because diamond is so hard, a diamond is cut using the corner of another diamond, a technique known as bruiting. After cutting, the diamond is then polished using diamond powder to create beautiful, sparkling gems used to make jewellery.

? Did you know?

Most **diamonds** we see today were formed millions (if not billions) of years ago at depths of about 150 kilometres below the Earth's surface in the molten rock of the Earth's mantle where the pressure and temperature are both extremely high. Powerful magma eruptions then brought the diamonds to the Earth's surface.



▲ **Figure 5.5.4 (a)** A cut diamond and **(b)** an uncut diamond



▲ **Figure 5.5.5** Structure of diamond

Structure and properties of graphite

Graphite is also composed of carbon atoms. Each atom is bonded covalently to *three* other carbon atoms, forming *hexagonal rings* of atoms which link up to form flat sheets or layers of carbon atoms. The fourth electron of each atom is not bonded to another atom and becomes delocalised. These layers of carbon atoms then lie on top of each other. The covalent bonds between the carbon atoms in the layers are very strong, however, the layers are held together by weak forces of attraction. The properties of graphite can be explained by its structure, as summarised in Table 5.5.4.

Table 5.5.4 Properties of graphite

Property	Explanation
Very high melting point – about 3600 °C	The covalent bonds between the carbon atoms are very strong. A very large amount of heat energy is needed to weaken these forces and separate the atoms from each other, allowing the solid to melt. Because of this, graphite can be used in containers for molten metals, e.g. crucibles.
Soft and lubricating	The weak forces between the layers of carbon atoms allow the layers to slide over each other resulting in a 'soft' crystal which feels slippery. Because of this, graphite can act as a lubricant.
Good conductor of electricity	The fourth electron of each carbon atom is delocalised allowing it to move within the crystal. These moving electrons can carry electricity through the crystal. Because of this, graphite is used as electrodes during electrolysis.

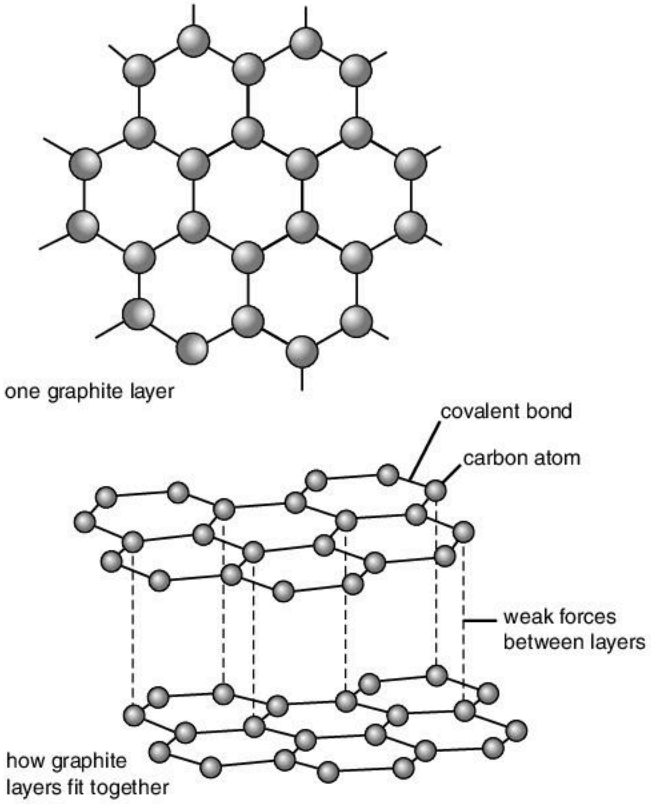


Figure 5.5.6 Structure of graphite

Did you know?

Graphite is a more stable form of carbon than diamond.

Practical activity

Constructing models of sodium chloride, diamond and graphite

Your teacher may use this activity to assess:

- manipulation and measurement.

You will be supplied with toothpicks and small, different coloured pieces of plasticine (or similar round, soft material to use as atoms or ions).

Method

Using the diagrams in Unit 5.5 to help you, make models of the following to show their structure:

- sodium chloride
- diamond
- graphite.

Use the plasticine to represent atoms or ions and the toothpicks to represent the bonds between them.

Exam tip

It is important that you can draw and describe the **structure** of sodium chloride, diamond and graphite and that you can relate their properties and uses to their structures.

Key fact

Allotropes are different structural forms of the same element in the same physical state.

Key fact

Allotropy is the existence of different structural forms of the same element in the same physical state.

Did you know?

Converting inexpensive graphite to diamond has been a long held dream of scientists throughout the world. However, the extremely high pressures and temperatures required to force carbon to change its bonding structure have meant that the process is extremely time-consuming and energy-intensive. Consequently it has not yet been achieved with any real success.

Allotropy

Diamond and graphite are known as **allotropes** of carbon. This is because they are both made out of the same element, carbon, but, in the solid state, their carbon atoms are bonded differently. The occurence of these different allotropes is known as **allotropy**.

Because diamond and graphite are both composed of carbon atoms, they have the *same* chemical properties. However, their atoms are bonded differently, therefore, they have different crystal structures which result in them having *different* physical properties.

Several other elements exhibit allotropy. These include sulfur and phosphorus, when in the solid state, and oxygen, when in the gaseous state.

Summary questions

1 What type of crystal structure does each of the following have?
   a Potassium bromide
   b Iodine
   c Ice
   d Graphite
   e Calcium carbonate

2 Why can sodium chloride dissolve in water?

3 Compare the melting points of an ionic solid and a simple molecular solid.

4 Why are diamonds extremely hard?

5 Why can graphite conduct an electric current?

Key concepts

- Atoms bond with each other to attain a more stable electronic configuration, i.e. that of the nearest noble gas in the periodic table.
- Atoms can lose, gain or share valence electrons to attain a stable electronic configuration.
- There are three main types of chemical bonding: ionic, covalent and metallic.
- A **chemical bond** is a force of attraction between atoms that results from the redistribution of their electrons.
- Chemical compounds formed as a result of ionic or covalent bonding can be represented by chemical formulae.
- There are three main types of chemical formulae: molecular formulae, structural formulae and empirical formulae.
- **Valence number** or **valency** is the number of bonds an atom can form with other atoms. Valence number can be used to write empirical formulae of compounds composed of two different elements.
- **Ionic bonding** involves the transfer of electrons from metal atoms to non-metal atoms, forming ions.
- Metal atoms form positive ions called **cations** and non-metal atoms form negative ions called **anions**.

- In ionic compounds, the strong electrostatic forces of attraction between the positive cations and negative anions, known as **ionic bonds**, hold these ions together in a crystal lattice.
- A crystal lattice is a three-dimensional regular arrangement of particles.
- Ionic compounds are represented using an empirical formula or formula unit, both of which give the ratio of the ions present.
- Many ions in ionic compounds are made from several atoms bonded together. These ions are known as **polyatomic ions**.
- **Covalent bonding** involves the sharing of electrons between non-metal atoms to form molecules. A shared pair of electrons forms a **covalent bond**.
- A **molecule** is a group of atoms which are bonded together strongly enough to behave as a single unit in a chemical reaction.
- Atoms within a molecule are held together by the strong electrostatic forces of attraction between the nuclei of the atoms and the bonding pair or pairs of electrons.
- In metals, the valence electrons become delocalised and are able to move between the metal cations which remain. These delocalised electrons are also known as mobile electrons.
- The strong electrostatic force of attraction between the delocalised electrons and the positive cations forms the **metallic bond** which holds the metal lattice together.
- The properties of metals are due to the bonding within the metal lattice.
- Four different solid structures can be identified based on the way the particles are bonded within the solids.
- The properties of the different solid structures depend on the bonding between their particles and the arrangement of the particles.
- **Ionic crystals** result from ionic bonding. They are composed of an ionic lattice in which the cations and anions are held together in a regular, repeating, three-dimensional arrangement by strong ionic bonds.
- **Simple molecular crystals** are composed of small molecules arranged in a regular, three-dimensional way to create a simple molecular lattice. Weak **intermolecular forces** between the molecules hold them together.
- Ionic crystals have high melting points, most dissolve in polar solvents and they conduct electricity when molten or dissolved in water.
- Simple molecular crystals have low melting points, most dissolve in non-polar solvents and they do not conduct electricity in any state.
- **Giant molecular crystals** are composed of non-metal atoms bonded by covalent bonds in a regular, three-dimensional arrangement to form a giant molecular lattice, also known as a macromolecule.
- Diamond and graphite are examples of giant molecular crystals.
- Diamond and graphite are **allotropes**. They are composed of the same element, carbon, but their atoms are bonded differently.
- Diamond is very hard, has a high melting point and does not conduct electricity.
- Graphite is ‘soft’ and lubricating, has a high melting point and does conduct electricity.

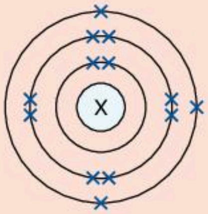
Practice exam-style questions

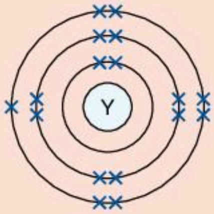
Multiple-choice questions

1 Which of the following is the correct formula for calcium phosphate?

A $Ca_2(PO_4)_3$
B $Ca(PO_4)_2$
C Ca_3PO_4
D $Ca_3(PO_4)_2$

2 The formula of the compound formed between atoms X and Y would be:





A XY
B X_3Y
C XY_3
D X_7Y_3

3 Ionic bonding occurs between:

A two metals
B a metal and a non-metal
C a metalloid and a non-metal
D two non-metals

4 An atom with an atomic number of 15 and an atom with an atomic number of 9 would:

A form a covalent compound
B form an ionic compound
C form a metallic lattice
D not bond with each other

5 Ionic compounds can only conduct electricity when in aqueous solution because:

A water molecules conduct electricity
B electrons become free when they are in aqueous solution
C the ions become free when they are in aqueous solution
D protons become free when they are in aqueous solution

6 Which of the following is/are the properties of a simple molecular compound?

I They are usually soluble in water.
II They conduct electricity in any state.
III They have low melting points.

A I and II only
B I and III only
C II and III only
D III only

7 Diamond and graphite are:

A isotopes of carbon
B isomers of carbon
C allotropes of carbon
D allomers of carbon

Structured question

8 The table gives the mass number and atomic number of four different elements.

Element	Mass number	Atomic number
fluorine	19	9
neon	20	10
aluminium	27	13
phosphorus	31	15

a **i)** Which element would not be capable of forming compounds? *(1 mark)*
ii) Give a reason for your answer. *(2 marks)*

b Two elements in the table are capable of bonding to form an ionic compound.
i) Name these elements. *(1 mark)*
ii) Give the formula of this compound. *(1 mark)*
iii) By means of a dot and cross diagram, show how this compound is formed. *(3 marks)*

c Two elements in the table are capable of bonding to form a covalent compound.
i) Name these elements. *(1 mark)*
ii) Give the formula of this compound. *(1 mark)*
iii) Draw a dot and cross diagram to show the bonding in this compound. *(2 marks)*

d **i)** Describe the bonding in aluminium. *(2 marks)*
ii) Why is aluminium able to conduct an electric current? *(1 mark)*

Total 15 marks

Extended response question

9 **a** With the aid of a diagram in EACH case, describe the lattice structure of:
i) sodium chloride *(3 marks)*
ii) graphite. *(3 marks)*

b With reference to the structure and type of bonding, explain each of the following:
i) sodium chloride is able to conduct electricity when molten *(2 marks)*
ii) graphite is able to conduct electricity in the solid state *(2 marks)*
iii) both sodium chloride and graphite have high melting points. *(3 marks)*

c Some types of sand consist of silica, SiO_2. Silica is relatively hard, does not conduct electricity in any state and has a high melting point. Describe the bonding which you would expect to find in silica. *(2 mark)*

Total 15 marks

A6 Chemical equations and reactions

Chemical equations use symbols and formulae to represent chemical change. There are universal standards for writing chemical equations which make it possible for chemists all over the world to interpret the equations written by others. The theory behind chemical reactions is very important in the chemical industry. It reduces the risk of accidents, can improve productivity and can increase the yield of chemicals.

Objectives

By the end of this topic you will be able to:

- understand the different conventions used when writing chemical equations
- write balanced chemical equations
- predict whether a compound is soluble or insoluble in water
- write balanced ionic equations.

A6.1 Writing and balancing chemical equations

A chemical equation is a shorthand representation of a chemical reaction using symbols and formulae. The chemicals that are reacting, i.e. the **reactants,** are shown on the left side of the equation and the chemicals formed during the reactions, i.e. the **products**, are shown on the right side. The reactants and products are separated by an arrow:

$$\text{reactants} \longrightarrow \text{products}$$

Understanding chemical equations

Chemists all use the same conventions when writing chemical equations so that any person reading a chemical equation will understand exactly what it means. The conventions used when writing chemical equations are as follows:

- The symbol or chemical formula of each reactant is placed on the left side of the equation and the symbol or formula of each product is placed on the right side.
- An arrow ($\longrightarrow$) is used to separate the reactants from the products. The arrow means 'produces'.
- A plus sign (+) is used to separate each reactant and each product. The plus sign on the left means 'reacts with' and on the right it means 'and'.
- A state symbol is written in brackets after the symbol or formula of each reactant and product. This indicates the physical state of the reactants and products. The state symbols used are as follows:
  (s) = solid
  (l) = liquid
  (g) = gas
  (aq) = aqueous solution, i.e. a solution where water is the solvent.
- Any conditions which are required for the reaction to occur can be written above the arrow, e.g. a specific temperature, a specific pressure or a specific catalyst. A catalyst is a chemical substance which speeds up a reaction without being used up in the reaction. You will learn more about catalysts in Unit 11.2. For example,

  $$CaCO_3(s) \xrightarrow{898\,°C} CaO(s) + CO_2(g)$$

  This shows that calcium carbonate requires heating to 898 °C for it to decompose into calcium oxide and carbon dioxide.

- If the reaction is reversible, a double arrow ($\rightleftharpoons$) is used. For example,

  $$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$

  You will learn more about reversible reactions in Unit 6.2.
- The equation must balance so that the number of atoms or ions of each element on each side of the arrow is the same. You will learn how to balance equations in the sections that follow.

When reading an equation, each part represents a word or a statement. Consider the following example:

$$AgNO_3(aq) + KCl(aq) \longrightarrow AgCl(s) + KNO_3(aq)$$

The equation reads as follows: 'silver nitrate solution reacts with potassium chloride solution to produce solid silver chloride and potassium nitrate solution'.

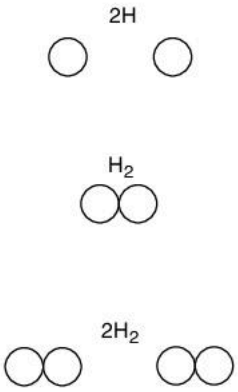
Balancing chemical equations

You should have noticed that in the three chemical equations above, the atoms are not changed during the reactions. The way in which the atoms are bonded to each other changes, but there is the same number of atoms of each element on both sides of each equation. In other words, atoms are neither created nor destroyed in chemical reactions, they are only rearranged.

When writing a chemical equation, you must ensure that the number of atoms of each element on the right side of the equation is the same as the number on the left side. A chemical equation that conforms to this is known as a **balanced chemical equation**.

You learnt about writing chemical formulae in Unit 5. To help you to understand how to balance chemical equations, consider the following:

- 2H represents two atoms of hydrogen, not bonded together
- H_2 represents one molecule of hydrogen, consisting of two atoms of hydrogen bonded together
- $2H_2$ represents two molecules of hydrogen and, in total, four atoms of hydrogen
- $3CO_2$ represents three molecules of carbon dioxide. In total, the number of atoms of each element is 3C and 6O atoms
- $Ca(NO_3)_2$ represents one formula unit of calcium nitrate, which consists of one Ca^{2+} ion and two NO_3^- ions; therefore, the number of atoms of each element is 1Ca, 2N and 6O.
- $2Ca(NO_3)_2$ represents two formula units of calcium nitrate. In total, the number of atoms of each element is 2Ca, 4N and 12O.



▲ **Figure 6.1.1** Different representations of atoms and molecules

The number *in front* of each formula in the list above is called the **coefficient**. A coefficient of one is not written. Coefficients are used when balancing equations. Consider the reaction between hydrogen gas and chlorine gas to produce hydrogen chloride gas. We can summarise this reaction using formulae as follows:

$$H_2(g) + Cl_2(g) \longrightarrow HCl(g)$$

Studying this equation we can see that there are two atoms of hydrogen on the left side, but only one atom on the right side. Also, there are two atoms of chlorine on the left side, but only one atom on the right side. The equation is, therefore, not balanced. If a coefficient, 2, is placed in front of the formula for hydrogen chloride, this now gives two atoms of hydrogen and two atoms of chlorine on the right side:

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

The equation is now balanced.

It is very important to note that when balancing equations the formulae cannot be changed. The number of atoms of the elements can only be changed by changing the coefficient in front of a formula. The subscripts after the elements must always remain the same.

How to write balanced chemical equations

When writing and balancing chemical equations, the guidelines outlined below should be followed.

- Write the correct chemical formula for each reactant and product, taking note of the following:
  - Seven of the common elements exist as diatomic molecules when in their free state. These are H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2. In reactions involving any of these elements, write the formula for the diatomic molecule.
  - In reactions involving other elements in their free state, write the atomic symbol of the element, e.g. if a reaction involves magnesium, write Mg.
- Separate the reactants and products by an arrow.
- Separate each reactant and each product by a plus sign.
- Write the state symbol after each reactant and product.
- Start balancing the elements in the product immediately after the arrow first.
- Any polyatomic ion which appears unchanged from one side to the other can be balanced as a unit, e.g. if SO_4^{2-} appears at both sides, consider it as a single unit (see Example 4).
- If any compounds contain hydrogen or oxygen, balance hydrogen second from last and balance oxygen last (see Examples 2 and 3).
- If any elements occur in their free state, leave these until the very last to balance, e.g. Fe, Ca, Cl_2, O_2.
- Use coefficients in front of each formula to balance the number of atoms of each element on the left and right sides. The formulae must not be changed.
- Check that the coefficients are in the lowest possible ratio.

Exam tip

It is essential that you can write **balanced chemical equations** for all the reactions you will encounter. Always check that you have the same type and number of atom of each element on both sides of the arrow, and that you have given the correct state symbol for each reactant and product.

Examples

1 Write a balanced chemical equation for the reaction in which magnesium burns in oxygen to form magnesium oxide as the only product.

Write the chemical formulae of the reactants and products together with their state symbols. After doing this, determine the number of atoms of each element on each side of the arrow:

$$Mg(s) + O_2(g) \longrightarrow MgO(s)$$

Reactants	Products
Mg = 1	Mg = 1
O = 2	O = 1

O does not balance. Balance by placing a **2** in front of the **MgO**.

$$Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Reactants	Products
Mg = 1	Mg = 2
O = 2	O = 2

Now **Mg** does not balance. Balance by placing a **2** in front of the **Mg**.

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Reactants	Products
Mg = 2	Mg = 2
O = 2	O = 2

The equation is now balanced.

2 Write a balanced chemical equation for the reaction between propane gas (C_3H_8) and oxygen to form carbon dioxide and steam.

Write the chemical formulae of the reactants and products together with their state symbols and determine the number of atoms of each element on each side of the arrow:

$$C_3H_8(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Reactants	Products
C = 3	C = 1
H = 8	H = 2
O = 2	O = 2 + 1 = 3

None of the elements balance. Balance the carbon and hydrogen first, leaving the oxygen in its free state until last. Balance the **C** by placing a **3** in front of the $\mathbf{CO_2}$. Balance the **H** by placing a **4** in front of the $\mathbf{H_2O}$.

$$C_3H_8(g) + O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

Reactants	Products
C = 3	C = 3
H = 8	H = 8
O = 2	O = 6 + 4 = 10

O still does not balance. Balance by placing a **5** in front of the $\mathbf{O_2}$.

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

Reactants	Products
C = 3	C = 3
H = 8	H = 8
O = 10	O = 6 + 4 = 10

The equation is now balanced.

3 Balance the following equation:

$$Ca(HCO_3)_2(aq) + HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Determine the number of atoms of each element on each side of the arrow:

$$Ca(HCO_3)_2(aq) + HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Reactants	Products
Ca = 1	Ca = 1
H = 2 + 1 = 3	H = 2
C = 2	C = 1
O = 6	O = 2 + 1 = 3
Cl = 1	Cl = 2

Start by balancing the elements in the product immediately after the arrow. **Ca** balances, but **Cl** does not balance. Balance by placing a **2** in front of the **HCl**.

$$Ca(HCO_3)_2(aq) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Reactants	Products
Ca = 1	Ca = 1
H = 2 + 2 = 4	H = 2
C = 2	C = 1
O = 6	O = 2 + 1 = 3
Cl = 2	Cl = 2

Balance the carbon and hydrogen next. Balance the **C** by placing a **2** in front of the **CO_2**. Balance the **H** by placing a **2** in front of the **H_2O**.

$$Ca(HCO_3)_2(aq) + 2HCl(aq) \longrightarrow CaCl_2(aq) + 2CO_2(g) + 2H_2O(l)$$

Reactants	Products
Ca = 1	Ca = 1
H = 2 + 2 = 4	H = 4
C = 2	C = 2
O = 6	O = 4 + 2 = 6
Cl = 2	Cl = 2

The equation is now balanced.

4 Balance the following equation

$$Al(NO_3)_3(aq) + NaOH(aq) \longrightarrow Al(OH)_3(s) + NaNO_3(aq)$$

In this reaction, the NO_3^- and the OH^- ions are polyatomic ions which remain unchanged from one side to the other. They are, therefore, treated as single units when determining the number of atoms of each element at each side of the arrow.

$$Al(NO_3)_3(aq) + NaOH(aq) \longrightarrow Al(OH)_3(s) + NaNO_3(aq)$$

Reactants	Products
Al = 1	Al = 1
NO_3 = 3	NO_3 = 1
Na = 1	Na = 1
OH = 1	OH = 3

Start with the product immediately after the arrow. **OH** does not balance. Balance by placing a **3** in front of the **NaOH**. Move on to the next product. **NO_3** does not balance. Balance by placing a **3** in front of the **$NaNO_3$**.

$$Al(NO_3)_3(aq) + 3NaOH(aq) \longrightarrow Al(OH)_3(s) + 3NaNO_3(aq)$$

Reactants	Products
Al = 1	Al = 1
NO_3 = 3	NO_3 = 3
Na = 3	Na = 3
OH = 3	OH = 3

The equation is now balanced.

State symbols of ionic compounds

In Unit 5.5 you learnt that most ionic compounds are soluble in water. Some ionic compounds are, however, insoluble in water. When writing equations involving ionic compounds, it is very important to know whether the ionic compound is soluble or insoluble in water. If it is *insoluble*, then it would always be given the state symbol (s). If it is *soluble* it would usually be given the state symbol (aq). Table 6.1.1 gives the solubility rules for determining the solubility of ionic compounds in water.

Table 6.1.1 Solubility of ionic compounds in water

Solubility rule	Exceptions
All compounds of Group I metals are *soluble*, e.g. compounds of lithium, sodium and potassium.	None.
All ammonium compounds are *soluble*.	None.
Most nitrates and ethanoates are *soluble*.	Silver ethanoate (CH_3COOAg) is only slightly soluble.
Most chlorides, bromides and iodides are *soluble*.	Those of lead and silver are insoluble. Lead(II) chloride ($PbCl_2$) and lead(II) bromide ($PbBr_2$) are moderately soluble in hot water.
Most sulfates are *soluble*.	Those of lead and barium are insoluble. Calcium sulfate ($CaSO_4$) and silver sulfate (Ag_2SO_4) are slightly soluble.
All hydrogencarbonates are *soluble*.	None.
Most carbonates and phosphates are *insoluble*.	Those of Group I metals and ammonium are soluble.
Most metal hydroxides are *insoluble*.	Those of Group I metals and ammonium are soluble. Barium hydroxide ($Ba(OH)_2$) is moderately soluble, calcium hydroxide ($Ca(OH)_2$) is slightly soluble and magnesium hydroxide ($Mg(OH)_2$) is very slightly soluble.
Most metal oxides are *insoluble*.	Those of Group I metals, calcium and barium, react with water to form the equivalent hydroxides.

Exam tip

It is essential that you learn the **solubility rules** so that you can give the correct state symbols when writing chemical equations. This is particularly important when writing ionic equations.

Writing ionic equations

Ionic equations show only the atoms or ions which actually take part in a reaction and, as a result, end up in a different state from the one in which they started. For example,

- two ions in solution may join to form an insoluble compound within the solution, known as a **precipitate**
- two ions may form a covalent compound composed of molecules
- an ion may be converted to an atom, i.e. it may be discharged
- an element in its free state may be converted to an ion, i.e. it may be ionised.

Figure 6.1.2 Example of a precipitate formed in a solution

The following steps should be followed when writing ionic equations.

1) Write the balanced chemical equation for the reaction.
2) Rewrite the equation, but show any ions that are present in solution as individual ions.
3) Delete any ions which remain unchanged, i.e. which are in solution on both sides of the equation. These are called **spectator ions** and they do not take part in the reaction.
4) Rewrite the ionic equation showing only the atoms or ions which change, i.e. those which actually take part in the reaction.
5) Cancel the coefficients to the lowest possible ratio if necessary.

It is important to note that when **acids**, such as hydrochloric acid ($HCl(aq)$), nitric acid ($HNO_3(aq)$) and sulfuric acid ($H_2SO_4(aq)$), are in aqueous solution they ionise to form hydrogen ions, or H^+ ions, and negative anions.

Examples

1 Write the ionic equation for the reaction between lead nitrate solution and sodium chloride solution to produce insoluble lead chloride and sodium nitrate solution.

1) Write the balanced chemical equation:

$Pb(NO_3)_2(aq) + 2NaCl(aq) \longrightarrow PbCl_2(s) + 2NaNO_3(aq)$

2) Rewrite the equation showing ions present in solution as individual ions:

$Pb^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + 2Cl^-(aq) \longrightarrow PbCl_2(s) + 2Na^+(aq) + 2NO_3^-(aq)$

3) Delete the $Na^+(aq)$ and $NO_3^-(aq)$ ions since they remain unchanged:

$Pb^{2+}(aq) + \cancel{2NO_3^-(aq)} + \cancel{2Na^+(aq)} + 2Cl^-(aq) \longrightarrow PbCl_2(s) + \cancel{2Na^+(aq)} + \cancel{2NO_3^-(aq)}$

4) Rewrite the ionic equation: $Pb^{2+}(aq) + 2Cl^-(aq) \longrightarrow PbCl_2(s)$

2 Write the ionic equation for the reaction between potassium hydroxide and sulfuric acid ($H_2SO_4(aq)$) to produce potassium sulfate and water.

1) $2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$

2) $2K^+(aq) + 2OH^-(aq) + 2H^+(aq) + SO_4^{2-}(aq) \longrightarrow 2K^+(aq) + SO_4^{2-}(aq) + 2H_2O(l)$

3) $\cancel{2K^+(aq)} + 2OH^-(aq) + 2H^+(aq) + \cancel{SO_4^{2-}(aq)} \longrightarrow \cancel{2K^+(aq)} + \cancel{SO_4^{2-}(aq)} + 2H_2O(l)$

4) $2OH^-(aq) + 2H^+(aq) \longrightarrow 2H_2O(l)$

5) $OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$

3 Write the ionic equation for the reaction between magnesium and hydrochloric acid ($HCl(aq)$) to produce magnesium chloride and hydrogen gas.

1) $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$

2) $Mg(s) + 2H^+(aq) + 2Cl^-(aq) \longrightarrow Mg^{2+}(aq) + 2Cl^-(aq) + H_2(g)$

3) $Mg(s) + 2H^+(aq) + \cancel{2Cl^-(aq)} \longrightarrow Mg^{2+}(aq) + \cancel{2Cl^-(aq)} + H_2(g)$

4) $Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$

! Key fact

It is extremely important that you give the correct **state symbols** when writing ionic equations and that the equations **balance**.

Summary questions

1 Write balanced equations for the following reactions:
- **a** Aluminium + oxygen ⟶ aluminium oxide
- **b** $N_2(g) + H_2(g) \rightleftharpoons NH_3(g)$
- **c** $CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
- **d** The reaction between potassium and water to produce potassium hydroxide and hydrogen gas.
- **e** The reaction between aluminium hydrogencarbonate and hydrochloric acid ($HCl(aq)$) to produce aluminium chloride, carbon dioxide and water.
- **f** The reaction between hydrogen and oxygen to produce steam.

2 State whether the following compounds are soluble or insoluble in water:

AgI $Fe(NO_3)_3$ Na_2CO_3 $KMnO_4$ $(NH_4)_3PO_4$

3 Balance the following equations and then write the ionic equation for each reaction:
- **a** $Na_2CO_3(aq) + HCl(aq) \longrightarrow NaCl(aq) + CO_2(g) + H_2O(l)$
- **b** $KOH(aq) + (NH_4)_2SO_4(aq) \longrightarrow K_2SO_4(aq) + NH_3(g) + H_2O(l)$
- **c** $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + NaNO_3(aq)$

A6.2 Types of chemical reactions

Objectives

By the end of this topic you will be able to:

- identify the seven main types of chemical reactions
- give examples of the main types of chemical reactions
- write balanced equations for the different types of chemical reactions.

Chemical reactions form an integral part of our lives, from reactions that take place in our bodies, such as respiration which changes glucose into energy, to complex reactions, which produce pharmaceutical products. There are *seven* main types of chemical reactions which are described below.

Synthesis reactions

Synthesis reactions are reactions in which two or more substances combine chemically to form a single product.

A synthesis reaction is generally symbolised by:

$$A + B \longrightarrow AB$$

Examples of synthesis reactions are as follows:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

$$2FeCl_2(s) + Cl_2(g) \longrightarrow 2FeCl_3(s)$$

Decomposition reactions

Decomposition reactions are reactions in which a single reactant is broken down into two or more products. A decomposition reaction will occur if the compound is unstable, if the compound is heated (thermal decomposition) or if an electric current is passed through the compound when in the liquid state or dissolved in aqueous solution (electrolysis).

A decomposition reaction is generally symbolised by:

$$AB \longrightarrow A + B$$

Examples of decomposition reactions are as follows:

$$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

$$2Pb(NO_3)_2(s) \xrightarrow{\text{heat}} 2PbO(s) + 4NO_2(g) + O_2(g)$$

Did you know?

Many islands of the Caribbean, being coral islands, are composed largely of **calcium carbonate**. When the sugar industry first started in Barbados, a coral island, calcium carbonate was quarried and roasted in lime kilns to produce calcium oxide. The calcium oxide was then used to clarify the cane juice during the extraction of sucrose from sugar cane, which at one time was a major industry within the island (Unit 2.5).

Single displacement reactions

A **single displacement reaction** involves the reaction between an element in its free state and a compound. The element in its free state takes the place of (displaces) another element in the compound. A more reactive element will always displace a less reactive element from a compound. Displacement reactions can be divided into two types.

- A **metal** may displace another metal from a compound or it may displace the hydrogen from an acid.

  This type of single displacement reaction can generally be symbolised by:

  $$A + BX \longrightarrow AX + B$$

  Examples of this type of single displacement reaction are as follows:

  $$Mg(s) + CuSO_4(aq) \longrightarrow MgSO_4(aq) + Cu(s)$$

  $$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

Writing these as ionic equations:

$$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$$
$$Zn(s) + 2H^{+}(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$$

In the above examples, magnesium is more reactive than copper and zinc appears above hydrogen in the reactivity series (see Unit 18.1).

- A **non-metal** may displace another non-metal from a compound.

  This type of single displacement reaction can be generally symbolised by:

  $$X + AY \longrightarrow AX + Y$$

  An example of this type of single displacement reaction is as follows:

  $$Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$$

  Writing this as an ionic equation:

  $$Cl_2(g) + 2Br^{-}(aq) \longrightarrow 2Cl^{-}(aq) + Br_2(aq)$$

  In this example, chlorine is more reactive than bromine.

Ionic precipitation reactions

Ionic precipitation reactions are sometimes referred to as **double displacement reactions**. They usually involve two compounds which are in solution, where the compounds exchange ions.

An ionic precipitation reaction is generally symbolised by:

$$AX + BY \longrightarrow AY + BX$$

To have a successful ionic precipitation reaction, one or both of the products must be incapable of undergoing the reverse reaction. This is generally because they form one product which is insoluble, known as a **precipitate**.

Key fact

A **precipitate** is an insoluble solid that forms in a solution, often as a result of a chemical reaction.

Examples of ionic precipitation reactions are as follows:

$$AgNO_3(aq) + KBr(aq) \longrightarrow AgBr(s) + KNO_3(aq)$$
$$BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NaCl(aq)$$

Writing these as ionic equations:

$$Ag^{+}(aq) + Br^{-}(aq) \longrightarrow AgBr(s)$$
$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

Neutralisation reactions

Neutralisation reactions are reactions between a base (or an alkali) and an acid. The acid is neutralised by the base (or alkali) and the products formed are a salt and water.

An example of a neutralisation reaction is:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Writing this as an ionic equation:

$$OH^{-}(aq) + H^{+}(aq) \longrightarrow H_2O(l)$$

Redox reactions

Redox reactions, or oxidation–reduction reactions, are chemical reactions in which one reactant is **reduced** and the other reactant is **oxidised**. Redox reactions will be covered in more detail in Unit 9.

Reversible reactions

Reversible reactions occur when the direction of a chemical change can be reversed easily, i.e. the products can react to produce the original reactants again. Most reactions are not reversible; they can only proceed in one direction.

If a reaction is reversible, a double arrow ($\rightleftharpoons$) is used. In a reversible reaction, we refer to the forward reaction as proceeding from left to right and the reverse or backward reaction as proceeding from right to left.

A reversible reaction is generally symbolised by:

$$A + B \rightleftharpoons C + D$$

For example, when ammonium chloride is heated, it decomposes into ammonia and hydrogen chloride gases:

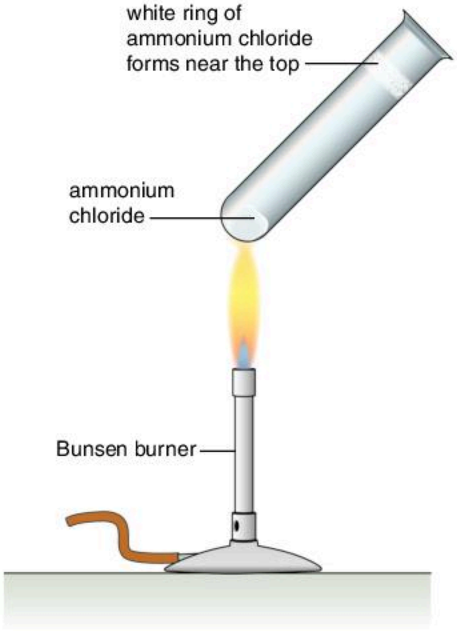
$$NH_4Cl(s) \longrightarrow NH_3(g) + HCl(g)$$

If the ammonia and hydrogen chloride gases are cooled (or mixed at room temperature) they react to form ammonium chloride:

$$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$$

The reaction is, therefore, reversible and the equation can be written as follows:

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$



▲ **Figure 6.2.1** Heating ammonium chloride

In Figure 6.2.1, as the ammonium chloride is heated it forms ammonia and hydrogen chloride gases, which diffuse up the test tube. The top of the tube is much cooler and when the gases reach here, they react and form a ring of ammonium chloride around the inside of the tube.

In many reversible reactions, the reaction can proceed in both directions at the same time such that the reaction always contains a mixture of reactants and products. If the reaction occurs in a sealed system where no substances can enter or leave, then a stage is reached where the rates of the forward and backward reactions are equal. The reaction does not stop and is said to have reached **dynamic equilibrium**.

Summary questions

1 Balance the following equations, then classify the reactions as synthesis, decomposition, single displacement, ionic precipitation or neutralisation reactions:

- a $Mg(NO_3)_2(s) \longrightarrow MgO(s) + NO_2(g) + O_2(g)$
- b $Fe(s) + Cl_2(g) \longrightarrow FeCl_3(s)$
- c $Zn(s) + HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
- d $NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(l)$
- e $AgNO_3(aq) + MgBr_2(aq) \longrightarrow AgBr(s) + Mg(NO_3)_2(aq)$.

Key concepts

- **Chemical equations** are representations of chemical reactions using symbols and formulae.
- The reactants are shown at the left of an equation and the products on the right.
- The state of each reactant and product is given after its formula using (s) for solid, (l) for liquid, (g) for gas and (aq) for aqueous solution.
- In a balanced chemical equation there must be the same number of atoms of each element on both sides of the equation.
- **Coefficients** are numbers placed in front of formulae. They are used when balancing equations. Formulae must never be changed.
- There are solubility rules which can be used to determine the solubility of ionic compounds in water. These are useful in determining state symbols of ionic compounds.
- **Ionic equations** show only the atoms or ions which actually take part in a reaction and, as a result, end up in a different situation from the one in which they started.
- Writing ionic equations involves writing the full equation first, but showing any ions that are present in solution as individual ions to determine which ions do not change in any way.
- **Spectator ions** are the ions which do not take part in a reaction and, therefore, remain unchanged. These are left out of the final ionic equation.
- There are seven main types of chemical reactions: synthesis, decomposition, single displacement, ionic precipitation, neutralisation, redox and reversible reactions.
- A **synthesis reaction** is any reaction in which two or more reactants combine to form a single product.
- A **decomposition reaction** is any reaction where a single reactant is broken down into two or more products.
- A **single displacement reaction** is any reaction between an element in its free state and a compound where the element displaces another element from the compound.
- An **ionic precipitation reaction** usually involves two compounds in aqueous solution reacting to produce a precipitate.
- A **precipitate** is an insoluble solid that forms in a solution, often as a result of a chemical reaction.
- A **neutralisation reaction** occurs when an acid reacts with a base (or alkali) to form a salt and water.
- A **redox reaction** occurs when one reactant is reduced and the other is oxidised.
- A **reversible reaction** occurs when the direction of a chemical change can be reversed easily, i.e. the products can react to produce the original reactants again.

Practice exam-style questions

Multiple-choice questions

1 Which of the following correctly represents the chemical reaction between sodium hydroxide and sulfuric acid to produce sodium sulfate and water?

A $NaOH(aq) + H_2SO_4(aq) \longrightarrow NaSO_4(aq) + H_2O(l)$

B $2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(s) + 2H_2O(l)$

C $2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

D $Na(OH)_2(aq) + H_2SO_4(aq) \longrightarrow NaSO_4(aq) + 2H_2O(l)$

2 Which of the following compounds is soluble in water?

A copper(II) sulfate

B magnesium carbonate

C zinc hydroxide

D calcium phosphate

3 Aqueous solutions of which of the following compounds would react to form a precipitate when mixed?

A sodium chloride and copper(II) nitrate

B zinc sulfate and magnesium chloride

C potassium hydroxide and sodium nitrate

D lead nitrate and ammonium carbonate

4 Consider the following reaction:

$$Mg(HCO_3)_2(aq) + 2HCl(aq) \longrightarrow MgCl_2(aq) + 2CO_2(g) + 2H_2O(l)$$

The best ionic equation for this reaction is:

A $Mg^{2+}(aq) + 2HCO_3^-(aq) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + 2CO_2(g) + 2H_2O(l)$

B $2HCO_3^-(aq) + 2H^+(aq) \longrightarrow 2CO_2(g) + 2H_2O(l)$

C $HCO_3^-(aq) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$

D $HCO_3^-(aq) + H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$

5 What type of reaction would you consider the following to be?

$$2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$$

A a neutralisation reaction

B a single displacement reaction

C a double displacement reaction

D a synthesis reaction

Equations to write

6 Write a balanced chemical equation for each of the following reactions.

a The reaction between aluminium and chlorine to produce aluminium chloride as the only product.

b The reaction between aluminium oxide and sulfuric acid ($H_2SO_4(aq)$), which produces aluminium sulfate and water.

c The reaction between potassium carbonate solution and hydrochloric acid ($HCl(aq)$) to produce potassium chloride, carbon dioxide and water.

d The reaction between ethane ($C_2H_6(g)$) and oxygen to make carbon dioxide and steam.

e The reaction between aqueous solutions of lead(II) nitrate and potassium sulfate to make lead(II) sulfate and potassium nitrate.

f The reaction between potassium hydroxide solution and ammonium sulfate solution to form potassium sulfate, ammonia gas and water.

g The reaction between calcium and hydrochloric acid ($HCl(aq)$), which makes calcium chloride and hydrogen.

h The reaction between sodium and water to produce sodium hydroxide and hydrogen.

i The reaction between magnesium hydrogencarbonate solution and nitric acid ($HNO_3(aq)$), which produces magnesium nitrate, carbon dioxide and water.

j The reaction between sodium hydroxide solution and sulfuric acid ($H_2SO_4(aq)$) to make sodium sulfate and water.

7 Write balanced ionic equations for the reactions in **c**, **e**, **f**, **g**, **i** and **j** in question **6** above.

A7 The mole concept

Chemistry often involves determining the amount of a substance required to produce a product. For example, how much ammonia and nitric acid are required to produce 1 kg of the fertiliser ammonium nitrate? The mass of an atom is so small that chemists have had to devise another method of measuring the amount of substances; this method uses the mole.

Objectives

By the end of this topic you will be able to:

- define the terms relative atomic mass, relative molecular mass and relative formula mass
- calculate relative molecular and formula masses
- define the terms mole and molar mass
- give Avogadro's constant
- perform calculations involving moles, masses and number of particles.

A7.1 The mole and mass

Relative atomic, molecular and formula masses

You learnt in Unit 3.3 that because the mass of atoms is extremely small, scientists use a much smaller unit than a gram to compare their masses, known as relative atomic mass. They designed a system which compares the mass of an atom with one-twelfth the mass of a carbon-12 atom. A carbon-12 atom was assigned a mass of 12.00 atomic mass units or amu, therefore, one-twelfth the mass of a carbon-12 atom has a mass of 1.00 amu.

Relative atomic mass, therefore, compares the mass of an atom of an element to one-twelfth the mass of a carbon-12 atom. Because it is a comparative value, relative atomic mass has *no* units. Relative atomic mass is given the symbol $\boldsymbol{A_r}$.

In the definition of relative atomic mass given in the Key fact box, we use the word 'average' because relative atomic mass takes into account the relative abundance of each isotope. The relative atomic mass of an element, to the nearest whole number, can be found in the periodic table on page 360. For example, the relative atomic mass of calcium is 40 and of fluorine is 19.

When calculating the masses of covalent and ionic compounds we use the same logic by comparing their masses to the carbon-12 atom.

Molecular elements and compounds formed by covalent bonding are composed of molecules. We use the term relative molecular mass to compare the mass of a molecule of a molecular element or compound with one-twelfth the mass of a carbon-12 atom. Relative molecular mass is given the symbol $\boldsymbol{M_r}$.

Compounds formed by ionic bonding are represented by formula units. We use the term relative formula mass to compare the mass of a formula unit of an ionic compound with one-twelfth the mass of a carbon-12 atom.

To determine the relative molecular mass or relative formula mass, we *add* together the relative atomic masses of all the elements present in the compound. Since relative molecular and formula masses are comparative values, they have no units.

! Key fact

Relative atomic mass, A_r, is the average mass of one atom of an element compared to one-twelfth the mass of an atom of carbon-12.

! Key fact

Relative molecular mass, M_r, is the average mass of one molecule of an element or compound compared to one-twelfth the mass of an atom of carbon-12.

! Key fact

Relative formula mass is the average mass of one formula unit of a compound compared to one-twelfth the mass of an atom of carbon-12.

Examples

1 Calculate the relative molecular mass of nitrogen, N_2.

N_2 consists of 2 atoms of N.

∴ relative molecular mass of $N_2 = (2 \times 14) = \mathbf{28}$.

2 Calculate the relative molecular mass of glucose, $C_6H_{12}O_6$.

$C_6H_{12}O_6$ consists of 6 atoms of H, 12 atoms of H and 6 atoms of O.

∴ relative molecular mass of
$C_6H_{12}O_6 = (6 \times 12) + (12 \times 1) + (6 \times 16) = \mathbf{180}$.

3 Calculate the relative formula mass of magnesium carbonate, $MgCO_3$.

$MgCO_3$ consists of 1 atom of Mg, 1 atom of C and 3 atoms of O.

∴ relative formula mass of
$MgCO_3 = (1 \times 24) + (1 \times 12) + (3 \times 16) = \mathbf{84}$.

The mole

We use different terms to represent a fixed amount of something, e.g. a dozen apples represents 12 apples and a pair of gloves represents two gloves. The same concept is used in chemistry. The mole represents a fixed number. It can be used to describe anything, e.g. a mole of apples, a mole of gloves, a mole of atoms or a mole of molecules. However, a mole represents a very large number and is, therefore, used mainly as a measurement of atomic sized particles.

Key fact

A **mole** is the amount of a substance that contains the same number of particles as there are atoms in 12.00 g of carbon-12.

Like relative atomic mass, the mole has been defined using the carbon-12 isotope because of its high abundance and stability. It was found that 12.00 g of carbon-12 contains 6.0×10^{23} atoms of carbon-12. One mole represents the number of atoms in 12.00 g of carbon-12, i.e. 6.0×10^{23}. This number is known as Avogadro's constant, or $\boldsymbol{N_A}$.

In other words, a mole is the amount of a substance that contains 6.0×10^{23} particles of the substance.

Looking at the definition of a mole in the Key fact box:

- 'amount' can refer to the mass of a substance or the volume of a substance if the substance is a gas
- 'particles' can refer to atoms, molecules, formula units or ions.

In the same way that it was found that 12.00 g of carbon-12 contains 6.0×10^{23} carbon-12 atoms, it was found that:

- 40 g of calcium (Ca) contains 6.0×10^{23} Ca atoms
- 28 g of nitrogen (N_2) contains 6.0×10^{23} N_2 molecules
- 180 g of glucose ($C_6H_{12}O_6$) contains 6.0×10^{23} $C_6H_{12}O_6$ molecules
- 84 g of magnesium carbonate ($MgCO_3$) contains 6.0×10^{23} $MgCO_3$ formula units.

If we look at the masses given above we see that each mass has the same numerical value as the relative mass of the element or compound and each is the mass of *one mole* of particles of the substance.

NB When writing the mole as a unit, it is usually abbreviated to **mol**.

Molar mass

The mass of one mole of a substance is known as its molar mass, or ***M***.

The molar mass of an element or a compound is given the unit grams per mole or **g mol^{-1}**. For example, because 12.00 g of carbon contains 6.0×10^{23} carbon atoms, or 1 mol of carbon atoms, the molar mass of carbon is 12 g mol^{-1}.

The molar mass of different elements and compounds will be different even though they are all the mass of one mole, just as the mass of one dozen apples will be different from the mass of a dozen grapes even though they are both a dozen. This is because the mass of the atoms of each element is different, just as the mass of an apple is different from the mass of a grape.

Key fact

Molar mass, *M*, is the mass, in grams, of one mole of a substance.

Calculating molar mass

Looking at the findings above, it can be seen that:

- because 40 g of calcium contains 6.0×10^{23} Ca atoms, or 1 mol of Ca atoms, the molar mass of calcium is **40 g mol^{-1}**
- because 28 g of nitrogen contains 6.0×10^{23} N_2 molecules, or 1 mol of N_2 molecules, the molar mass of nitrogen is **28 g mol^{-1}**
- because 180 g of glucose contains 6.0×10^{23} $C_6H_{12}O_6$ molecules, or 1 mol of $C_6H_{12}O_6$ molecules, the molar mass of glucose is **180 g mol^{-1}**
- because 84 g of magnesium carbonate contains 6.0×10^{23} $MgCO_3$ formula units, or 1 mol of $MgCO_3$ formula units, the molar mass of magnesium carbonate is **84 g mol^{-1}**.

Therefore, the molar mass of an element or compound is the relative atomic, molecular or formula mass amount expressed in grams per mole.

More examples

$$\begin{aligned}
M(H_2O) &= (2 \times 1) + 16 = \mathbf{18\,g\,mol^{-1}} \\
M(CuSO_4) &= 64 + 32 + (4 \times 16) = \mathbf{160\,g\,mol^{-1}} \\
M(Ca_3(PO_4)_2) &= (3 \times 40) + (2 \times 31) + (8 \times 16) = \mathbf{310\,g\,mol^{-1}} \\
M((NH_4)_2CO_3) &= (2 \times 14) + (8 \times 1) + 12 + (3 \times 16) = \mathbf{96\,g\,mol^{-1}}
\end{aligned}$$

Moles and mass

Molar mass gives us the relationship between the number of moles and the mass of an element or compound. We can use the molar mass to calculate the mass or number of moles of an element or compound.

Examples

1 Calculate the number of moles in 10 g of calcium carbonate, $CaCO_3$.

First calculate the molar mass of $CaCO_3$.

$M(CaCO_3) = 40 + 12 + (3 \times 16) = 100\,g\,mol^{-1}$

i.e. mass of 1 mol $CaCO_3$ = 100 g

$\therefore$ number of moles in 10 g $= \frac{10}{100}$ mol $=$ **0.1 mol**

Exam tip

When performing calculations in chemistry, including those involving the mole, it is essential that you show each step of your working clearly by making a **clear statement**. You must not just put down a set of numbers without stating what you are calculating.

2 Calculate the mass of 0.2 mol of sulfuric acid, H_2SO_4.

First calculate the molar mass of H_2SO_4.

$M(H_2SO_4) = (2 \times 1) + 32 + (4 \times 16) = 98\ \text{g mol}^{-1}$

i.e. mass of 1 mol H_2SO_4 = 98 g

$\therefore$ mass of 0.2 mol of $H_2SO_4 = 0.2 \times 98\ \text{g} = \mathbf{19.6\ g}$

Moles and number of particles

The number of particles in one mole is always 6.0×10^{23}. Using this fact we can calculate the number of particles in any number of moles. Remember that the type of particle in a substance depends on the type of bonding in the substance.

- If a substance is an element, e.g. metals and the noble gases, then the particles are individual *atoms*.
- If a substance is a molecular element, e.g. oxygen, or a covalent compound, e.g. water, then the particles are *molecules* made up of *atoms*.
- If a substance is an ionic compound, e.g. sodium chloride, then the particles are *formula units* made up of *ions*.

Examples

1 Calculate the number of atoms in 0.5 mole of copper, Cu.

1 mol Cu contains 6.0×10^{23} atoms.

$\therefore$ 0.5 mol Cu contains $0.5 \times 6.0 \times 10^{23}$ atoms = $\mathbf{3.0 \times 10^{23}}$ **atoms**

2 Calculate the number of moles of carbon dioxide, CO_2, that contain 2.0×10^{23} carbon dioxide molecules.

1 mol CO_2 contains 6.0×10^{23} molecules.

$$\therefore \text{ number of moles in } 2.0 \times 10^{23} \text{ molecules} = \frac{2.0 \times 10^{23}}{6.0 \times 10^{23}}\ \text{mol} = \mathbf{0.33\ mol}$$

We can extend our calculations of moles and particles to look at the number of moles or particles inside a compound. For example:

- 1 mol of water, H_2O, consists of: 1 mol of H_2O molecules, or 2 mol of H atoms and 1 mol of O atoms
- 2 mol of potassium carbonate, K_2CO_3 consists of: 2 mol of K_2CO_3 formula units, or 4 mol of K^+ ions and 2 mol of CO_3^{2-} ions.

Example

Calculate the number of hydrogen atoms in 1.5 mol of methane, CH_4.

1 mol CH_4 contains 4 mol H atoms.

$\therefore$ 1.5 mol CH_4 contains 1.5×4 mol H atoms = **6 mol H atoms**.

1 mol H atoms contains 6.0×10^{23} atoms.

$\therefore$ 6 mol H atoms contains $6 \times 6.0 \times 10^{23}$ atoms = $\mathbf{3.6 \times 10^{24}}$ **atoms**.

Exam tip

It is essential that you place the **correct unit** after each value given in your calculations, especially your final answers. The only values that have no units are relative atomic, molecular and formula masses. Your answer is incorrect if it has the wrong unit or no unit, unless it is the value for relative mass.

Moles, mass and number of particles

We can now combine our calculation of moles and mass, and moles and number of particles.

Examples

1 Calculate the number of molecules in 1.8 g of water, H_2O.

$M(H_2O) = (2 \times 1) + 16 = 18\,g\,mol^{-1}$

i.e. mass of 1 mol $H_2O = 18\,g$

$\therefore$ number of moles in $1.8\,g = \frac{1.8}{18}\,mol = \mathbf{0.1\,mol}$

1 mol H_2O contains 6.0×10^{23} molecules

$\therefore$ 0.1 mol H_2O contains $0.1 \times 6.0 \times 10^{23}$ molecules

$= \mathbf{6.0 \times 10^{22}}$ **molecules**

2 Calculate the mass of 1.5×10^{23} molecules of hydrogen chloride, HCl.

1 mol HCl contains 6.0×10^{23} molecules

$\therefore$ number of moles in 1.5×10^{23} molecules $= \frac{1.5 \times 10^{23}}{6.0 \times 10^{23}}\,mol = \mathbf{0.25\,mol}$

$M(HCl) = 1 + 35.5 = 36.5\,g\,mol^{-1}$

i.e. mass of 1 mol HCl $= 36.5\,g$

$\therefore$ mass of 0.25 mol HCl $= 0.25 \times 36.5\,g = \mathbf{9.13\,g}$

3 Calculate the number of aluminium ions, Al^{3+}, in 40.8 g of aluminium oxide, Al_2O_3.

$M(Al_2O_3) = (2 \times 27) + (3 \times 16) = 102\,g\,mol^{-1}$

i.e. mass of 1 mol $Al_2O_3 = 102\,g$

$\therefore$ number of moles in $40.8\,g = \frac{40.8}{102}\,mol = \mathbf{0.4\,mol}$

1 mol Al_2O_3 contains 2 mol Al^{3+} ions.

$\therefore$ 0.4 mol Al_2O_3 contains 0.4×2 mol Al^{3+} ions

$=$ **0.8 mol Al^{3+} ions**.

1 mol Al^{3+} ions contains 6.0×10^{23} ions

$\therefore$ 0.8 mol Al^{3+} ions contains $0.8 \times 6.0 \times 10^{23}$ ions

$= \mathbf{4.8 \times 10^{23}}$ **ions**.

Summary questions

1 Calculate the mass of:
  a 0.25 mol of sodium sulfate, Na_2SO_4 **b** 1.2 mol of glucose, $C_6H_{12}O_6$.

2 Calculate the number of moles in:
  a 20 g of neon **b** 63.0 g of magnesium carbonate.

3 Calculate the number of molecules in 1.5 mol of hydrogen chloride, HCl.

4 Calculate the mass of 2.5×10^{23} formula units of calcium phosphate, $Ca_3(PO_4)_2$.

5 Calculate the number of molecules in 34 g of ammonia, NH_3.

6 Calculate the number of chloride ions in 6.8 g of zinc chloride, $ZnCl_2$.

Objectives

By the end of this topic you will be able to:

- state Avogadro's law
- define molar volume
- perform calculations involving moles and volumes of gases.

A7.2 The mole and gas volumes

Gases have very small masses because of their low densities, so chemists have investigated the relationship between the number of moles of a gas and the *volume* occupied by the gas.

In 1811, Amedeo Avogadro noticed that all gases under the same conditions of temperature and pressure, and occupying the same volume, contain the same number of molecules. He proposed **Avogadro's law**.

As an example, using Avogadro's law, we can say that 100 cm^3 of oxygen gas at a temperature of 25 °C and pressure of 101.3 kPa has the same number of molecules as 100 cm^3 of nitrogen gas at 25 °C and 101.3 kPa.

If the number of molecules in each gas is 6.0×10^{23}, i.e. 1 mol, it follows that 1 mol of all gases under the same conditions of temperature and pressure must occupy the same volume.

Key fact

Avogadro's law states that equal volumes of all gases, under the same conditions of temperature and pressure, contain the same number of molecules.

Molar volume

The volume of one mole of a gas is known as its **molar volume**, or $\boldsymbol{V_m}$.

Molar volume is, therefore, the volume of a gas which contains 6.0×10^{23} molecules of the gas.

Since temperature and pressure both affect the volume of a gas, molar volume will depend on both of these. Chemists mostly work with two sets of conditions.

- **Standard temperature and pressure**, or **stp**, which equates to a temperature of 0 °C (273 K) and a pressure of 101.3 kPa (1 atmosphere).

  Molar volume at stp is $\mathbf{22.4\ dm^3}$ or $\mathbf{22\,400\ cm^3}$.

- **Room temperature and pressure**, or **rtp**, which equates to a temperature of 25 °C (298 K) and a pressure of 101.3 kPa (1 atmosphere).

  Molar volume at rtp is $\mathbf{24.0\ dm^3}$ or $\mathbf{24\,000\ cm^3}$.

Molar volume gives us the relationship between the number of moles and the volume of a gas. We can use the molar volume to calculate the volume or number of moles of a gas.

Key fact

Molar volume, V_m, is the volume occupied by 1 mol of a gas.

Examples

1 Calculate the volume occupied by 0.25 mol nitrogen, N_2, at stp.

Volume of 1 mol of N_2 at stp = 22.4 dm^3

$\therefore$ volume of 0.25 mol of N_2 at stp = $0.25 \times 22.4\ dm^3 = \mathbf{5.6\ dm^3}$

2 Calculate the number of moles in 2.4 dm^3 of oxygen gas, O_2, at rtp.

Volume of 1 mol of O_2 at rtp = 24.0 dm^3

$\therefore$ number of moles in 2.4 dm^3 of O_2 at rtp = $\frac{2.4}{24.0}$ mol = **0.1 mol**

We can now combine our calculations with those in Unit 7.1 to involve moles, mass, volume and number of particles.

Examples

1 Calculate the volume occupied by 6.4 g of oxygen, O_2 at stp.

Calculate the molar mass of O_2 gas.

$M(O_2) = 2 \times 16 = 32\,g\,mol^{-1}$

i.e. mass of 1 mol of $O_2 = 32\,g$

$\therefore$ number of moles in 6.4 g $= \frac{6.4}{32}$ mol $=$ **0.2 mol**

volume of 1 mol O_2 at stp $= 22.4\,dm^3$

$\therefore$ volume of 0.2 mol O_2 gas at stp $= 0.2 \times 22.4\,dm^3 =$ **$4.48\,dm^3$**

2 Calculate the mass of $600\,cm^3$ of carbon dioxide, CO_2, at rtp.

Volume of 1 mol CO_2 at rtp $= 24\,000\,cm^3$

$\therefore$ number of moles in $600\,cm^3 = \frac{600}{24\,000}$ mol

$=$ **0.025 mol**

$M(CO_2) = 12 + (2 \times 16) = 44\,g\,mol^{-1}$

i.e. mass of 1 mol $CO_2 = 44\,g$

$\therefore$ mass of 0.025 mol $CO_2 = 0.025 \times 44\,g$

$=$ **1.1 g**

3 Calculate the number of molecules in $6.72\,dm^3$ of ammonia, NH_3, at stp.

Volume of 1 mol NH_3 at stp $= 22.4\,dm^3$

$\therefore$ number of moles in $6.72\,dm^3 = \frac{6.72}{22.4}$ mol

$=$ **0.3 mol**

1 mol NH_3 contains 6.0×10^{23} molecules

$\therefore$ 0.3 mol NH_3 contains $0.3 \times 6.0 \times 10^{23}$ molecules

$=$ **1.8×10^{23} molecules**

4 Calculate the volume occupied by 4.5×10^{22} molecules of hydrogen chloride gas, HCl at rtp.

1 mol HCl contains 6.0×10^{23} molecules.

$\therefore$ number of moles in 4.5×10^{22} molecules $= \frac{4.5 \times 10^{22}}{6.0 \times 10^{23}}$ mol

$=$ **0.075 mol**

volume of 1 mol HCl at rtp $= 24\,dm^3$

$\therefore$ volume of 0.075 mol HCl at rtp $= 0.075 \times 24.0\,dm^3$

$=$ **$1.8\,dm^3$**

Summary questions

1 Calculate the mass, volume and number of molecules in 2.5 mol of ammonia, NH_3.

2 Calculate the number of atoms in $4.8\,dm^3$ of neon gas at rtp.

3 Calculate the volume of 16 g of sulfur dioxide, SO_2, at stp.

4 Determine the mass of $840\,cm^3$ of oxygen gas, O_2, at rtp.

A7.3 The mole and concentration of solutions

Objectives

By the end of this topic you will be able to:

- explain what is meant by concentration of a solution
- explain the terms mass concentration and molar concentration
- define the term standard solution
- perform calculations involving concentrations of solutions.

You have already come across ways of describing solutions such as dilute, concentrated and saturated. However, we can express the concentration of a solution in more precise ways, which allow us to perfume calculations involving solutions. The concentration of a solution is a measure of the quantity of a solute dissolved in a fixed volume of the solution. The volume of solution usually used is 1 dm^3, i.e. 1000 cm^3.

Concentration of a solution can be expressed in two ways.

- Mass concentration, which gives the mass of solute dissolved in 1 dm^3 of solution. The unit for mass concentration is grams of solute per cubic decimetre of solution or $\mathbf{g\,dm^{-3}}$.
- Molar concentration, which gives the number of moles of solute dissolved in 1 dm^3 of solution. The unit for molar concentration is moles of solute per cubic decimetre of solution or $\mathbf{mol\,dm^{-3}}$.

It is very important to note that concentration is expressed as the quantity of solute in 1 dm^3 of solution and not in 1 dm^3 of solvent.

Example

A solution of sodium hydroxide contains 10.0 g of sodium hydroxide dissolved in 1 dm^3 of solution. Express the concentration of the solution in $g\,dm^{-3}$ and $mol\,dm^{-3}$.

The solution contains 10.0 g of NaOH in 1 dm^3.

Therefore, mass concentration = $\mathbf{10.0\,g\,dm^{-3}}$.

To determine the molar concentration, find the number of moles of sodium hydroxide present:

$$M(\text{NaOH}) = 23 + 16 + 1 = 40\,\text{g mol}^{-1}$$

i.e. mass of 1 mole NaOH = 40 g

$$\therefore \quad \text{number of moles in 10.0 g NaOH} = \frac{10}{40}\,\text{mol} = \mathbf{0.25\,mol}$$

Therefore, molar concentration = $\mathbf{0.25\,mol\,dm^{-3}}$

? Did you know?

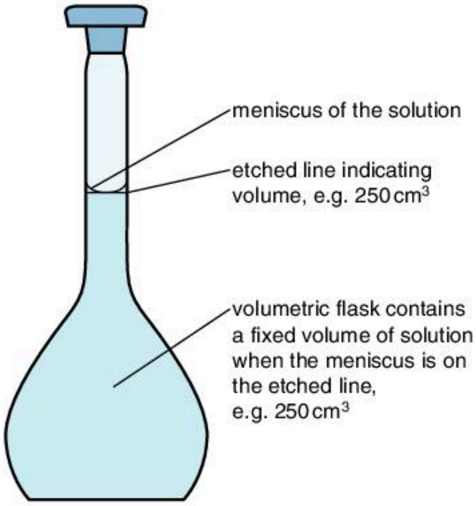
Molar concentration may also be called **molarity**. The unit for molarity is M. A solution with a concentration of 1 $mol\,dm^{-3}$ may be described as a '1 molar solution' or '1 M solution'.

! Key fact

A **standard solution** is one whose concentration is known accurately.

A standard solution

A solution whose concentration is known accurately is referred to as a standard solution. A standard solution is made in a volumetric flask as shown in Figure 7.3.1. These flasks come in various sizes, e.g. 100 cm^3, 250 cm^3, 500 cm^3 and 1 dm^3. A known mass of solute is added to the flask and distilled water is then added until the solution reaches the mark on the neck of the flask.



▲ **Figure 7.3.1** A volumetric flask

To make a standard solution which does not have a volume of 1 dm^3, the mass of solute required must be calculated first.

Example

Determine the mass of potassium carbonate which must be dissolved in distilled water to make 250 cm^3 of potassium carbonate solution with a concentration of 0.2 $mol\,dm^{-3}$.

1000 cm^3 (1 dm^3) of the required solution contains 0.2 mol K_2CO_3.

$\therefore$ 1 cm^3 of the required solution contains $\frac{0.2}{1000}$ mol K_2CO_3

and 250 cm^3 of the required solution contains

$$\frac{0.2}{1000} \times 250 \text{ mol } K_2CO_3 = \mathbf{0.05\ mol\ K_2CO_3}$$

Therefore 0.05 mol K_2CO_3 is required to make the solution.

$M(K_2CO_3) = (2 \times 39) + 12 + (3 \times 16) = 138 \text{ g mol}^{-1}$

i.e. mass of 1 mol $K_2CO_3 = 138$ g

$\therefore$ mass of 0.05 mol $K_2CO_3 = 0.05 \times 138 \text{ g} = \mathbf{6.9\ g}$

Therefore the mass of K_2CO_3 required is **6.9 g**.

Practical activity

Making a standard solution

Your teacher may use this activity to assess:

- manipulation and measurement.

You are going to make 250 cm^3 of sodium carbonate solution with a concentration of 0.1 mol dm^{-3}.

You will be supplied with sodium carbonate, distilled water, a balance, a sheet of weighing paper, a 100 cm^3 beaker, a glass rod and a 250 cm^3 volumetric flask.

Method

1. Calculate the mass of sodium carbonate required to make the solution. (Use the calculation given in the example above to help you.)
2. Weigh a sheet of weighing paper on a balance.
3. Weigh the mass of sodium carbonate you determined in your calculation on the sheet of weighing paper.
4. Carefully transfer the sodium carbonate to the beaker, making sure it is all transferred.
5. Add some distilled water and stir the solution with the glass rod until all the sodium carbonate has dissolved completely.
6. Transfer the solution to the volumetric flask.
7. Rinse the glass rod with distilled water over the beaker and transfer the washings to the flask. Wash the beaker with distilled water three times and transfer these washings to the flask each time.
8. Add distilled water to just below the line in the volumetric flask. Add the final quantity of distilled water drop by drop until the bottom of the meniscus is on the line.
9. Place the stopper on the flask. Holding the stopper in place, invert the flask three times to mix the solution.

Calculations

We can extend our calculations involving solutions as shown in the examples below.

Examples

1 Calculate the molar concentration of a solution of hydrochloric acid, HCl, which has a mass concentration of 9.125 g dm^{-3}.

$M(HCl) = 1 + 35.5 = 36.5 \text{ g mol}^{-1}$

i.e. mass of 1 mol HCl = 36.5 g

$\therefore$ number of moles HCl in 9.125 g $= \frac{9.125}{36.5} \text{ mol} = \mathbf{0.25\ mol}$

i.e. number of moles HCl in 1 dm^3 = 0.25 mol

Therefore, molar concentration = **0.25 mol dm^{-3}**

2 Calculate the molar concentration of a solution of sodium hydroxide which contains 6 g of sodium hydroxide dissolved in distilled water to make 200 cm^3 of solution.

Calculate the molar mass of sodium hydroxide, NaOH.

$M(\text{NaOH}) = 40\ \text{g mol}^{-1}$

i.e. mass of 1 mol of NaOH = 40 g

$\therefore$ number of moles in 6 g NaOH = $\frac{6}{40}$ mol = **0.15 mol**

200 cm^3 of solution contains 0.15 mol NaOH

$\therefore$ 1 cm^3 of solution contains $\frac{0.15}{200}$ mol NaOH

and 1000 cm^3(1 dm^3) of solution contains $\frac{0.15}{200} \times 1000$ mol NaOH

= **0.75 mol NaOH**

Therefore the molar concentration = **0.75 mol dm^{-3}**

3 Determine how many moles of sodium sulfate are present in 250 cm^3 of a sodium sulfate solution which has a molar concentration of 0.2 mol dm^{-3}.

1 dm^3 of solution contains 0.2 mol Na_2SO_4,

i.e. 1000 cm^3 of solution contains 0.2 mol Na_2SO_4.

$\therefore$ 1 cm^3 of solution contains $\frac{0.2}{1000}$ mol Na_2SO_4

and 250 cm^3 of solution contains $\frac{0.2}{1000} \times 250$ mol Na_2SO_4

= **0.05 mol Na_2SO_4**

Therefore, number of moles of sodium sulfate present = **0.05 mol**

4 You wish to make a sodium chloride solution with a molar concentration of 0.25 mol dm^{-3}, but you only have 11.7 g of sodium chloride. What is the maximum volume of solution you can make?

Calculate the molar mass of sodium chloride, NaCl.

$M(\text{NaCl}) = 23 + 35.5 = 58.5\ \text{g mol}^{-1}$

i.e. mass of 1 mol = 58.5 g.

$\therefore$ number of moles in 11.7 g = $\frac{11.7}{58.5}$ mol

= **0.2 mol**

Required solution has a molar concentration of 0.25 mol dm^{-3},

i.e. 0.25 mol NaCl is contained in 1 dm^3 of solution

or 0.25 mol NaCl is contained in 1000 cm^3 of solution.

$\therefore$ 1 mol NaCl is contained in $\frac{1000}{0.25}$ cm^3 of solution

and 0.2 mol NaCl is contained in $\frac{1000}{0.25} \times 0.2$ cm^3 of solution

= **800 cm^3 of solution**

Therefore, the maximum volume of solution which can be made is **800 cm^3**.

Summary questions

1. Define the terms ‘molar concentration’ and ‘standard solution’.
2. Calculate the mass concentration and molar concentration if 2.8 g of potassium hydroxide is dissolved in distilled water to make up 500 cm^3 of solution.
3. What mass of calcium chloride, $CaCl_2$, needs to be weighed to make a solution of volume 200 cm^3 and concentration 0.1 $mol\,dm^{-3}$?
4. How many moles of sodium carbonate are present in 400 cm^3 of sodium carbonate solution, which has a concentration of 0.25 $mol\,dm^{-3}$?

A7.4 The mole and chemical formulae

Objectives

By the end of this topic you will be able to:

- explain the terms empirical formula, molecular formula and percentage composition
- determine the empirical and molecular formulae of a compound
- determine the percentage composition of a compound.

Empirical and molecular formulae

In terms of moles, a chemical formula shows how many moles of each element combine to form one mole of the compound. For example, H_2O represents 2 mol of hydrogen atoms combined with 1 mol of oxygen atoms. In Unit 5.1 you learnt that chemical formulae can be written in three main ways. We will be considering *two* of these ways here.

- The **empirical formula**, which gives the simplest whole number mole ratio between the atoms or ions in the compound.
- The **molecular formula**, which gives the actual number of moles of atoms of each element present in one mole of the compound.

Ionic compounds are always represented by *empirical formulae,* which give the ratio of the ions within the compound.

Covalent compounds are represented by *molecular formulae*. The empirical and molecular formulae of covalent compounds may be the same, e.g. water (H_2O), ammonia (NH_3) and methane (CH_4), or the molecular formula may be a simple whole number multiple of the empirical formula, e.g. the molecular formula of butene is C_4H_8 and its empirical formula is CH_2.

If we know the proportions of the elements, by mass, in a compound, then we can use this to work out the empirical formula of the compound. If the molecular formula of a compound is different from the empirical formula and we know the molar mass of the compound, we can use this to work out its molecular formula.

To determine the empirical formula, we need to calculate the number of moles of each element present and then determine the simplest mole ratio of these elements.

Examples

1 A sample of a solid is decomposed and found to contain 6.52 g of potassium, 4.34 g of chromium and 5.34 g of oxygen. What is the empirical formula of the compound?

	K	**Cr**	**O**
mass (g)	6.52	4.34	5.34
molar mass ($g\,mol^{-1}$)	39	52	16
number of moles	$\frac{6.52}{39} = 0.167$	$\frac{4.34}{52} = 0.083$	$\frac{5.34}{16} = 0.333$
simplest mole ratio	**2**	**1**	**4**

The empirical formula is $\mathbf{K_2CrO_4}$.

Note: To determine the simplest mole ratio, divide the number of moles of each element by the smallest number, i.e. 0.083.

2 On analysis, a sample of glucose was found to contain 40.0% carbon, 6.7% hydrogen and 53.3% oxygen. The molar mass of the compound is 180 g mol^{-1}. Determine the molecular formula of glucose.

If the mass of the sample is assumed to be 100 g, then the percentage of each element can be expresses in grams.

	C	H	O
mass (g)	40	6.7	53.3
molar mass (g mol^{-1})	12	1	16
number of moles	$\frac{40}{12} = 3.3$	$\frac{6.7}{1} = 6.7$	$\frac{53.3}{16} = 3.3$
simplest mole ratio	**1**	**2**	**1**

The empirical formula of glucose is $\mathbf{CH_2O}$.

To determine the molecular formula, calculate the molar mass of CH_2O:

$M(CH_2O) = 12 + (2 \times 1) + 16 = 30\ \text{g mol}^{-1}$

$M(\text{glucose}) = 180\ \text{g mol}^{-1}$

Ratio between $M(\text{glucose})$ and $M(CH_2O) = \frac{180}{30} = \mathbf{6}$

$\therefore$ the molecular formula is 6 × the empirical formula.

The molecular formula of glucose is $\mathbf{C_6H_{12}O_6}$.

Percentage composition

The **percentage composition** of a compound indicates the percentage, by mass, of each element in the compound. For example, the percentage composition of water shows what percentage of the mass of a water molecule is made up of hydrogen and what percentage is made up of oxygen.

If we know the formula of a compound, the percentage composition, by mass, of each element present can be calculated by calculating the mass of 1 mol of the compound and the mass which each element contributes to 1 mol of the compound.

Examples

1 Calculate the percentage composition of hydrogen and oxygen in water.

$M(H_2O) = (2 \times 1) + 16 = 18\ \text{g mol}^{-1}$

i.e. mass of 1 mol H_2O = 18 g

Mass of hydrogen in 1 mol $H_2O = 2 \times 1 = 2$ g

$$\therefore \text{ percentage hydrogen} = \frac{2}{18} \times 100\% = \mathbf{11.11\%}$$

Mass of oxygen in 1 mol H_2O = 16 g

$$\therefore \text{ percentage oxygen} = \frac{16}{18} \times 100\% = \mathbf{88.89\%}$$

2 Calculate the percentage, by mass, of nitrogen in ammonium phosphate, $(NH_4)_3PO_4$.

$M((NH_4)_3PO_4) = (3 \times 14) + (12 \times 1) + 31 + (4 \times 16) = 149\,g\,mol^{-1}$

i.e. mass of 1 mol $(NH_4)_3PO_4 = 149\,g$

$$\text{Mass of nitrogen in 1 mol } (NH_4)_3PO_4 = 3 \times 14\,g = 42\,g$$

$$\therefore \text{ percentage nitrogen} = \frac{42}{149} \times 100\% = \mathbf{28.19\%}$$

Summary questions

1 On analysis a compound was found to contain 11.64 g of sodium, 16.20 g of sulfur and 12.15 g of oxygen. Determine the empirical formula of the compound.

2 A liquid used in rocket fuel is found to consist of 3.04 g of nitrogen and 6.94 g of oxygen. The molar mass of the compound is 92 g mol^{-1}. Determine the molecular formula of the compound.

3 Calculate the percentage composition of hydrogen, sulfur and oxygen in sulfuric acid, H_2SO_4.

4 Determine the percentage, by mass, of carbon in aluminium carbonate, $Al_2(CO_3)_3$.

A7.5 The mole and chemical reactions

The Law of Conservation of Matter

In Unit 6.1 you learnt that atoms are neither created nor destroyed in chemical reactions, they are only rearranged, and that when writing a chemical equation the number of atoms of each element on the right side of the equation must be the same as the number on the left side. This is summed up in the **Law of Conservation of Matter** given in the Key fact box.

Since all the atoms that were present at the beginning of the reaction are present at the end of the reaction, it follows that the total mass of the products must be the same as the total mass of the original reactants.

We can prove this using the reaction between hydrogen and oxygen to make water. The balanced equation for the reaction is:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

This means:

$$2 \text{ molecules of } H_2 + 1 \text{ molecule of } O_2 \longrightarrow 2 \text{ molecules of } H_2O$$

Or, on a bigger scale, the coefficients show the number of moles of each reactant and product. The equation, therefore, means:

$$2\,mol \text{ of } H_2 + 1\,mol \text{ of } O_2 \longrightarrow 2\,mol \text{ of } H_2O$$

If we now take this a step further and use our mole/mass relationship, it means:

$$2(2 \times 1)\,g \text{ of } H_2 + 1(2 \times 16)\,g \text{ of } O_2 \longrightarrow 2[(2 \times 1) + 16)]\,g \text{ of } H_2O$$
$$\mathbf{4\,g} \text{ of } H_2 + \mathbf{32\,g} \text{ of } O_2 \longrightarrow \mathbf{36\,g} \text{ of } H_2O$$

Objectives

By the end of this topic you will be able to:

- state the Law of Conservation of Matter
- apply the mole concept to balanced chemical equations using masses
- apply the mole concept to ionic equations
- apply the mole concept to balanced chemical equations using volumes of gases
- apply the mole concept to balanced chemical equations using concentration of solutions.

Key fact

The **Law of Conservation of Matter** states that matter can neither be created nor destroyed during a chemical reaction.

We can clearly see that the mass of the water produced is the same as the total mass of the original reactants, hydrogen and oxygen, i.e. 36 g.

We can use balanced chemical equations to determine the quantities of unknown reactants and products in a reaction. If the quantity of one reactant or product is known, it is possible to calculate the quantities of any of the other reactants and products.

Mass and chemical reactions

When answering questions involving masses of reactants and products, you will always be given the mass of one reactant or product and be asked to determine the mass of another reactant or product. To do this we use our mole/mass relationship and the mole ratio from the balanced equation for the reaction.

The steps involved in solving mass–mass problems are as follows:

1) Calculate the number of **moles** of the known reactant or product using its given mass and its molar mass. The known reactant or product is the one whose mass you have been given in the question.
2) Use the balanced chemical equation to determine the **mole ratio** between the known and the unknown reactants and/or products. The unknown reactant or product is the one whose mass you are calculating.
3) Use the number of moles of the known reactant or product found in step 1), and the mole ratio found in step 2), to calculate the number of **moles** of the unknown reactant or product.
4) Use the number of moles of the unknown reactant or product found in step 3), and its molar mass, to determine its **mass**.

Note: If you have not been given the equation for the reaction, you must write the balanced chemical equation before you begin your calculation.

Examples

1 Calculate the mass of magnesium oxide formed by burning 12 g of magnesium in excess oxygen.

The balanced equation for the reaction is:

$$2\text{Mg(s)} + \text{O}_2\text{(g)} \longrightarrow 2\text{MgO(s)}$$

The mass of the reactant, magnesium is known. The mass of the product, magnesium oxide, is to be determined.

1) Find the number of **moles** of Mg using its mass and molar mass:

   Calculate the molar mass of Mg.

   $M(\text{Mg}) = 24\ \text{g mol}^{-1}$

   i.e. mass 1 mol of Mg = 24 g

   $\therefore$ number of moles in 12 g of Mg $= \frac{12}{24}$ mol = **0.5 mol**

2) Use the balanced chemical equation to determine the **mole ratio** between Mg and MgO:

   2 mol Mg form 2 mol MgO.

   $\therefore$ 1 mol Mg forms **1 mol MgO**.

3) Use the number of moles of Mg from 1), and the mole ratio from 2), to calculate the number of **moles** of MgO produced:

   0.5 mol Mg forms **0.5 mol MgO**.

4) Use the number of moles of MgO from 3), and its molar mass, to determine the **mass** of MgO produced.

$M(\text{MgO}) = 24 + 16 = 40\ \text{g mol}^{-1}$

i.e. mass of 1 mol of MgO = 40 g

$\therefore$ mass of 0.5 mol of MgO = 0.5 × 40 g = **20 g**

2 Determine the mass of sodium carbonate required to produce 23.4 g of sodium chloride, when the sodium carbonate reacts with excess hydrochloric acid.

The balanced equation for the reaction is:

$$\text{Na}_2\text{CO}_3(\text{s}) + 2\text{HCl(aq)} \longrightarrow 2\text{NaCl(aq)} + \text{CO}_2(\text{g}) + \text{H}_2\text{O(l)}$$

The mass of the product, sodium chloride, is known. The mass of the reactant, sodium carbonate, is to be determined.

1) $M(\text{NaCl}) = 23 + 35.5 = 58.5\ \text{g mol}^{-1}$

i.e. mass of 1 mol NaCl = 58.5 g

$\therefore$ number of moles in 23.4 g $= \frac{23.4}{58.5}$ mol = **0.4 mol**

2) 1 mol Na_2CO_3 forms 2 mol NaCl.

3) $\frac{1}{2} \times 0.4$ mol Na_2CO_3 forms 0.4 mol NaCl = **0.2 mol $\mathbf{Na_2CO_3}$**

4) $M(\text{Na}_2\text{CO}_3) = (2 \times 23) + 12 + (3 \times 16) = 106\ \text{g mol}^{-1}$

i.e. mass of 1 mol Na_2CO_3 = 106 g

$\therefore$ mass of 0.2 mol = 0.2 × 106 g = **21.2 g**

3 Determine the decrease in mass that would occur when 21.0 g of magnesium carbonate is heated until its mass remains constant.

The balanced equation for the reaction is:

$$\text{MgCO}_3(\text{s}) \longrightarrow \text{MgO(s)} + \text{CO}_2(\text{g})$$

The decrease in mass is due to the loss of carbon dioxide, therefore, to find the decrease in mass, the mass of carbon dioxide produced must be determined.

The mass of the reactant, magnesium carbonate, is known. The mass of the product, carbon dioxide, is to be determined.

1) $M(\text{MgCO}_3) = 24 + 12 + (3 \times 16) = 84\ \text{g mol}^{-1}$

i.e. mass of 1 mol MgCO_3 = 84 g

$\therefore$ number of moles in 21.0 g $= \frac{21.0}{84}$ mol = **0.25 mol**

2) 1 mol MgCO_3 forms 1 mol CO_2.

3) 0.25 mol MgCO_3 forms **0.25 mol $\mathbf{CO_2}$**.

4) $M(\text{CO}_2) = 12 + (2 \times 16) = 44\ \text{g mol}^{-1}$

i.e. mass of 1 mol CO_2 = 44 g

$\therefore$ mass of 0.25 mol CO_2 = 0.25 × 44 g = **11.0 g**

Therefore, the decrease in mass is **11.0 g**.

We can also apply the mole concept to ionic equations using the same four steps.

Example

Calculate the mass of lead(II) hydroxide that could be produced when a solution containing 3.4 g of hydroxide ions reacts with a solution containing excess lead ions.

The ionic equation for the reaction is:

$$Pb^{2+}(aq) + 2OH^-(aq) \longrightarrow Pb(OH)_2(s)$$

The mass of the hydroxide ions present in the solution is known. The mass of the product, lead(II) hydroxide, is to be determined.

1) $M(OH^-) = 16 + 1 = 17\,g\,mol^{-1}$

   i.e. mass of 1 mol OH^- ions = 17 g

   $\therefore$ number of moles in 3.4 g = $\frac{3.4}{17}$ mol = **0.2 mol**

2) 2 mol OH^- ions form 1 mol $Pb(OH)_2$.

3) 0.2 mol OH^- ions form **0.1 mol $Pb(OH)_2$**.

4) $M(Pb(OH)_2) = 207 + (2 \times 16) + (2 \times 1) = 241\,g\,mol^{-1}$

   i.e. mass of 1 mol $Pb(OH)_2$ = 241 g

   $\therefore$ mass of 0.1 mol $Pb(OH)_2 = 0.1 \times 241$ g = **24.1 g**

Gas volumes and chemical reactions

Exam tip

You must pay very careful attention to the **conditions** under which reactions involving gases occur, i.e. stp or rtp, to ensure that you use the correct value for molar volume.

A balanced chemical equation for a reaction which involves gases can also be used to determine the unknown volume of a reactant or product using the mole/volume relationship and the mole ratio from the balanced equation for the reaction.

When answering questions involving volumes of gases, use the same four steps as are used in solving mass–mass problems.

Example

Nitrogen reacts with hydrogen to produce ammonia. Calculate the minimum volume of hydrogen that would be required to react completely with $6.0\,dm^3$ of nitrogen at rtp.

The balanced equation for the reaction is:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

The volume of the reactant, nitrogen, is known. The volume of the other reactant, hydrogen, is to be determined.

1) Volume of 1 mol N_2 at rtp = $24\,dm^3$

   $\therefore$ number of moles in $6.0\,dm^3$ = $\frac{6.0}{24}$ mol = **0.25 mol**

2) 1 mol N_2 reacts with 3 mol H_2.

3) 0.25 mol N_2 reacts with 3×0.25 mol H_2 = **0.75 mol H_2**.

4) Volume of 1 mol H_2 at rtp = $24.0\,dm^3$

   $\therefore$ volume of 0.75 mol H_2 at rtp = $0.75 \times 24.0\,dm^3$ = **$18.0\,dm^3$**

Problems can also be solved involving both mass and volume using the same four steps.

Example

Calculate the volume of sulfur trioxide gas formed at rtp when 9.6 g of oxygen reacts with excess sulfur dioxide.

The balanced equation for the reaction is:

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

The mass of the reactant, oxygen, is known. The volume of the product, sulfur dioxide, is to be determined.

1) $M(O_2) = 32\,\text{g mol}^{-1}$

   i.e. mass of 1 mol $O_2 = 32\,\text{g}$

   $\therefore$ number of moles in 9.6 g $= \frac{9.6}{32}$ mol $=$ **0.3 mol**

2) 1 mol O_2 forms 2 mol SO_3.

3) 0.3 mol O_2 forms 2×0.3 mol $SO_3 =$ **0.6 mol SO_3**.

4) Volume of 1 mol SO_3 at rtp $= 24.0\,\text{dm}^3$

   $\therefore$ volume of 0.6 mol SO_3 at rtp $= 0.6 \times 24.0\,\text{dm}^3 =$ **14.4 dm³**

Molar concentration and chemical reactions

A balanced chemical equation for a reaction which involves a solution whose volume and molar concentration are known can also be used to determine the mass of the other reactant, or the mass or volume of a product, using the same four steps as for solving mass and volume problems.

Examples

1 To prepare the salt, sodium sulfate, a student adds $50\,\text{cm}^3$ of sodium hydroxide solution with a concentration of $2.0\,\text{mol dm}^{-3}$ to excess sulfuric acid. Calculate the mass of sodium sulfate that would form.

The balanced equation for the reaction is:

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The volume and molar concentration of the reactant, sodium hydroxide, are known. The mass of the product, sodium sulfate, is to be determined.

1) $1000\,\text{cm}^3$ ($1\,\text{dm}^3$) NaOH(aq) contains 2 mol NaOH

   $\therefore$ $1\,\text{cm}^3$ NaOH(aq) contains $\frac{2}{1000}$ mol NaOH

   and $50\,\text{cm}^3$ NaOH(aq) contains $\frac{2}{1000} \times 50$ mol NaOH $=$ **0.1 mol NaOH**

2) 2 mol NaOH forms 1 mol Na_2SO_4.

3) 0.1 mol NaOH forms $\frac{1}{2} \times 0.1$ mol $Na_2SO_4 =$ **0.05 mol Na_2SO_4**

4) $M(Na_2SO_4) = (2 \times 23) + 32 + (4 \times 16) = 142\,\text{g mol}^{-1}$

   i.e. mass of 1 mol $Na_2SO_4 = 142\,\text{g}$

   $\therefore$ mass of 0.05 mol $Na_2SO_4 = 0.05 \times 142\,\text{g} =$ **7.1 g**

2 What volume of carbon dioxide, measured at stp, would be produced when 20 cm^3 of hydrochloric acid with a concentration of 1.5 $mol\,dm^{-3}$ reacts with excess aluminium carbonate?

The balanced equation for the reaction is:

$$Al_2(CO_3)_3(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3CO_2(g) + 3H_2O(l)$$

The volume and molar concentration of the reactant, hydrochloric acid, is known. The volume of the product, carbon dioxide is to be determined.

1) 1000 cm^3 HCl(aq) contains 1.5 mol HCl.

$\therefore$ 1 cm^3 contains $\frac{1.5}{1000}$ mol HCl

and 20 cm^3 contains $\frac{1.5}{1000} \times 20$ mol HCl = **0.03 mol HCl**.

2) 6 mol HCl form 3 mol CO_2.

$\therefore$ 2 mol HCl form 1 mol CO_2.

3) 0.03 mol HCl forms $\frac{1}{2} \times 0.03$ mol CO_2 = **0.015 mol CO_2**

4) Volume of 1 mol CO_2 at stp = 22.4 dm^3

$\therefore$ volume of 0.015 mol CO_2 at stp = 0.015 × 22.4 dm^3
= **0.336 dm^3**

Summary questions

1 Ethene (C_2H_4) burns in oxygen to form carbon dioxide and steam according to the following equation:

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

If 2.8 g of ethene burn in excess oxygen,

a calculate the number of moles in 2.8 g of ethene

b calculate the number of moles of carbon dioxide formed

c calculate the mass of carbon dioxide formed

d calculate the volume of carbon dioxide formed at stp.

2 Calcium carbonate decomposes into calcium oxide and carbon dioxide when heated. Calculate the mass of carbon dioxide released when 300 g of calcium carbonate is heated.

3 Carbon monoxide gas reacts with oxygen gas to form carbon dioxide gas. Calculate the volume and mass of carbon monoxide gas required to produce 4 dm^3 of carbon dioxide gas at rtp.

4 Potassium carbonate reacts with nitric acid (HNO_3) to produce potassium nitrate, carbon dioxide and water. Calculate the mass of potassium chloride formed when 40 cm^3 of potassium carbonate solution with a concentration of 0.5 $mol\,dm^{-3}$ reacts with excess nitric acid.

Key concepts

- Relative atomic mass, molecular mass and formula mass compare the masses of atoms, molecules and formula units with one-twelfth the mass of a carbon-12 atom.
- Relative atomic mass, molecular mass and formula mass have no units.
- A **mole** is the amount of a substance that contains the same number of particles as there are atoms in 12.00 g of carbon-12.
- The number of particles in a mole is 6.0×10^{23}, which is also known as **Avogadro's number** or $\boldsymbol{N_A}$.
- **Molar mass** is the mass, in grams, of one mole of a substance.
- The molar mass of an element or compound is the relative atomic mass, molecular mass or formula mass amount expressed in grams per mole. Molar mass is given the unit $\mathbf{g\,mol^{-1}}$.
- **Avogadro's law** states that equal volumes of all gases, under the same conditions of temperature and pressure, contain the same number of molecules.
- **Molar volume** is the volume occupied by one mole of a gas. The molar volume is the same for all gases under the same conditions of temperature and pressure.
- Molar volume at standard temperature and pressure (stp) is $\mathbf{22.4\,dm^3}$.
- Molar volume at room temperature and pressure (rtp) is $\mathbf{24\,dm^3}$.
- The **concentration of a solution** is a measure of the quantity of a solute dissolved in a fixed volume of the solution.
- Mass concentration gives the mass of solute dissolved in 1 dm^3 of solution. Mass concentration is given the unit $\mathbf{g\,dm^{-3}}$.
- Molar concentration gives the number of moles of solute dissolved in 1 dm^3 of solution. Molar concentration is given the unit $\mathbf{mol\,dm^{-3}}$.
- A **standard solution** is one whose concentration is known accurately. A standard solution is made using a volumetric flask.
- The **empirical formula** of a compound gives the simplest mole ratio between the atoms or ions in the compound.
- The **molecular formula** of a compound gives the actual number of moles of atoms of each element present in one mole of the compound.
- If the proportions of the elements, by mass, in a compound are known, then its empirical formula can be determined. If the molecular formula of a compound is different from the empirical formula and the molar mass of the compound is known, its molecular formula can be determined.
- The **percentage composition** of a compound indicates the percentage, by mass, of each element in the compound.
- The **Law of Conservation of Matter** states that matter can neither be created nor destroyed during a chemical reaction.
- The coefficients in a balanced chemical equation give the mole ratios between the reactants and the products.
- Balanced chemical equations can be used to determine quantities of unknown reactants and products in a reaction. If the quantity of one reactant or product is known, it is possible to calculate the quantities of any of the other reactants and products.

Practice exam-style questions

Multiple-choice questions

1 The mass of 0.25 mol of potassium carbonate is:

A 24.75 g
B 34.5 g
C 38.5 g
D 552.0 g

2 A mass of 32 g of oxygen:

A contains 6.0×10^{23} oxygen atoms
B contains 2 mol of oxygen
C contains 6.0×10^{23} molecules of oxygen
D contains the same number of molecules of oxygen as 35.5 g of chlorine

3 What is the mass of 448 cm^3 of carbon dioxide at stp?

A 0.88 g
B 224 g
C 448 g
D 880 g

4 You are required to make 250 cm^3 of a solution of sodium hydroxide which has a concentration of 0.1 mol dm^{-3}. What mass of sodium hydroxide would you need to weigh out?

A 0.1 g
B 1.0 g
C 4.0 g
D 25.0 g

5 A compound was found to consist of 30 g of element W combined with 7 g of Z. If the relative atomic mass of W is 40 and the relative atomic mass of Z is 14, which of the following is the correct empirical formula for the compound?

A WZ
B W_2Z
C W_2Z_3
D W_3Z_2

6 Hydrogen reacts with oxygen according to the following equation:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

What volume of steam could be produced at stp if 12.8 g of oxygen reacts with excess hydrogen?

A 8.96 dm^3
B 12.8 dm^3
C 17.92 dm^3
D 25.6 dm^3

Structured question

7 This question is concerned with the mole and its application.

a Give a definition for the term ‘mole’. *(1 mark)*

b Calcium hydrogencarbonate reacts with nitric acid according to the following equation:

$$Ca(HCO_3)_2(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + 2CO_2(g) + 2H_2O(l)$$

During an investigation in the laboratory, a student added 8.1 g of calcium hydrogencarbonate to excess nitric acid.

i) Determine the number of moles of calcium hydrogencarbonate used by the student. *(2 marks)*

ii) Determine the number of moles of carbon dioxide made. *(2 marks)*

iii) Determine the volume of carbon dioxide made at rtp. *(1 mark)*

c Magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen. In another investigation, a student reacted 20 cm^3 of hydrochloric acid with a concentration of 1.5 mol dm^{-3} with excess magnesium.

i) Write a balanced chemical equation for the reaction. *(2 marks)*

ii) Calculate the number of moles of hydrochloric acid used. *(1 mark)*

iii) Determine the number of moles of magnesium chloride produced. *(2 marks)*

iv) Calculate the mass of magnesium chloride produced. *(2 marks)*

d On analysis a compound was found to contain 24.45 g of iron and 10.55 g of oxygen. Determine the empirical formula of the compound. *(2 marks)*

Total 15 marks

Extended response question

8 a i) What do you understand by the term ‘standard solution’? *(1 mark)*

ii) You wish to prepare 250 cm^3 of potassium carbonate solution of concentration 0.8 mol dm^{-3} in the laboratory. Determine the mass of potassium carbonate required. *(3 marks)*

iii) Name the piece of apparatus you would use to make your solution in. *(1 mark)*

b A student wishes to make a sample of the insoluble salt, lead(II) chloride, in the laboratory. Determine the mass of lead(II) chloride which could be made by reacting a solution containing excess lead(II) nitrate with a solution containing 23.4 g of sodium chloride. *(6 marks)*

c Ammonium nitrate and ammonium sulfate are used extensively by gardeners as fertilisers to supply plants with the nitrogen they need to make protein for use in growth. By calculating the percentage, by mass, of nitrogen in each of the fertilisers, determine which would be the better value for money if they both cost the same per kilogram. *(4 marks)*

Total 15 marks

A8 Acids, bases and salts

The concept of acids, bases and salts has been around for a very long time; as early as the 17th century. As our knowledge of chemistry has developed, the classification of a compound as an acid, a base or a salt has been refined. Acids, bases and salts are used in all aspects of our lives.

Objectives

By the end of this topic you will be able to:

- explain the ionisation of acids when dissolved in water
- define an acid in terms of protons
- give the general properties of aqueous acids
- describe the reactions of acids with reactive metals, carbonates, hydrogencarbonates and bases
- describe the ways in which acids can be classified
- define acid anhydride and give examples of acid anhydrides
- give examples of acids in living systems.

A8.1 Properties and reactions of acids

People have always known that vinegar and lemon juice taste sour, but it was not until a few hundred years ago that it was discovered that these have something in common, they both contain **acids**. An acid can be defined as a substance that reacts with a base to form a salt and water.

Acids in their anhydrous form, i.e. not dissolved in water, are composed of covalent molecules and they may be solid, liquid or gas at room temperature. *Solids* include citric acid, tartaric acid and ascorbic acid (vitamin C). *Liquids* include sulfuric acid and nitric acid. *Gases* include hydrogen chloride gas.

All acids have hydrogen in their formulae, e.g. nitric acid, HNO_3 and sulfuric acid, H_2SO_4. Table 8.1.1 gives the names and formulae of some common acids.

▼ **Table 8.1.1** Some common acids

Name	Formula
hydrochloric acid	HCl
sulfuric acid	H_2SO_4
nitric acid	HNO_3
phosphoric acid	H_3PO_4
methanoic acid	$HCOOH$
ethanoic acid	CH_3COOH

▲ **Figure 8.1.1** **(a)** Some common laboratory acids, **(b)** some common household acids

Ionisation of acids in water

When an acid is added to water, the acid molecules **ionise**, i.e. they form ions. Acids always ionise to form **hydrogen ions**, or **H^+ ions**, and negative anions. The H^+ ions then become attached to the polar water molecules forming **hydronium ions** or **H_3O^+ ions** (sometimes called hydroxonium ions).

Using the ionisation of hydrogen chloride gas to form hydrochloric acid as an example:

$$\mathrm{HCl(g)} + \text{water} \longrightarrow \mathrm{H^+(aq)} + \mathrm{Cl^-(aq)}$$

$$\mathrm{H^+(aq)} + \mathrm{H_2O(l)} \longrightarrow \underset{\text{hydronium ion}}{\mathrm{H_3O^+(aq)}}$$

the overall reaction can be summarised as:

$$\mathrm{HCl(g)} + \mathrm{H_2O(l)} \longrightarrow \mathrm{H_3O^+(aq)} + \mathrm{Cl^-(aq)}$$

For simplicity, it is usual to represent the reaction as follows:

$$\mathrm{HCl(aq)} \longrightarrow \mathrm{H^+(aq)} + \mathrm{Cl^-(aq)}$$

Other examples that show the ionisation of acids in water are:

$$\mathrm{HNO_3(aq)} \longrightarrow \mathrm{H^+(aq)} + \mathrm{NO_3{}^-(aq)}$$

$$\mathrm{H_2SO_4(aq)} \longrightarrow \mathrm{2H^+(aq)} + \mathrm{SO_4{}^{2-}(aq)}$$

$$\mathrm{H_3PO_4(aq)} \longrightarrow \mathrm{3H^+(aq)} + \mathrm{PO_4{}^{3-}(aq)}$$

$$\mathrm{CH_3COOH(aq)} \rightleftharpoons \mathrm{CH_3COO^-(aq)} + \mathrm{H^+(aq)}$$

Key fact

An **acid** is a proton donor.

A hydrogen atom, $^1_1\mathrm{H}$, has one proton in its nucleus and one electron spinning around the nucleus in the first energy shell. A hydrogen ion is, therefore, a **single proton** since it is formed by a hydrogen atom losing its one valence electron, leaving only the nucleus containing the one proton. The ability of an acid to donate its hydrogen ions, or protons, to another reactant can be used to define an acid as being a proton donor.

When an acid reacts, the hydrogen ions, or protons, formed when it dissolves in water, can be *given* to the other reactant. For example, when an aqueous solution of hydrochloric acid reacts with sodium hydroxide, the hydrochloric acid gives its H^+ ions, or protons, to the OH^- ions of the sodium hydroxide, forming water. This can be summarised by the following ionic equation:

$$\mathrm{OH^-(aq)} + \mathrm{H^+(aq)} \longrightarrow \mathrm{H_2O(l)}$$

General properties of aqueous acids

The presence of *H^+ ions* in aqueous solution gives acids their characteristic properties. Aqueous solutions of acids are described as being **acidic** and they have the following common properties:

- they have a sour taste
- they change blue litmus to red
- they have a pH value of less than 7

- they are corrosive
- they are electrolytes, i.e. they conduct an electric current.

Chemical reactions of aqueous acids

Since all acids form hydrogen ions when in aqueous solution, acids have common reactions with certain other reactants. In all of these reactions, the hydrogen ions in the acid are replaced by metal ions or ammonium ions to form a salt. A salt is, therefore, a compound formed when the hydrogen ion in an acid is replaced by a metal ion or an ammonium ion.

The reactions of acids are discussed in detail below.

Acids react with reactive metals

Acids, except nitric acid, react with reactive metals to produce a salt and hydrogen. Reactive metals are those above hydrogen in the reactivity series of metals (see Unit 18.1), i.e. potassium, sodium, calcium, magnesium, aluminium, zinc, iron and lead.

This reaction can be summarised by the following general word equation:

reactive metal + acid → salt + hydrogen

When reactive metals react with acids, effervescence (bubbling) is seen as hydrogen gas is evolved, and the reactions become hotter because they produce heat, i.e. the reactions are described as being exothermic.

It is important to remember that nitric acid does not react in this way. When nitric acid reacts with metals it produces oxides of nitrogen, e.g. nitrogen dioxide (NO_2), and not hydrogen.

Example

Magnesium reacts with sulfuric acid to produce magnesium sulfate and hydrogen:

$$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$$

The ionic equation for this reaction is:

$$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$$

Acids react with metal carbonates and hydrogencarbonates

Acids react with metal carbonates and metal hydrogencarbonates to produce a salt, carbon dioxide and water.

These reactions can be summarised by the following general word equations:

metal carbonate + acid → salt + carbon dioxide + water

metal hydrogencarbonate + acid → salt + carbon dioxide + water

When metal carbonates and hydrogencarbonates react with acids, effervescence is seen as carbon dioxide gas is evolved.

Examples

1 Potassium carbonate reacts with nitric acid to produce potassium nitrate, carbon dioxide and water:

$$K_2CO_3(aq) + 2HNO_3(aq) \longrightarrow 2KNO_3(aq) + CO_2(g) + H_2O(l)$$

The ionic equation for this reaction is:

$$CO_3^{2-}(aq) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$$

2 Calcium hydrogencarbonate reacts with hydrochloric acid to produce calcium chloride, carbon dioxide and water:

$$Ca(HCO_3)_2(aq) + 2HCl(aq) \longrightarrow CaCl_2(aq) + 2CO_2(g) + 2H_2O(l)$$

The ionic equation for this reaction is:

$$HCO_3^-(aq) + H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$$

Did you know?

Any reactive metals, insoluble carbonates and insoluble bases that react with acids to form soluble salts appear to dissolve because the salt produced dissolves.

Acids react with bases

Acids react with bases to produce a salt and water. These are known as **neutralisation reactions.** It is important to note that a **base** is a metal oxide or a metal hydroxide.

These reactions can be summarised by the following general word equations:

metal oxide + acid ⟶ salt + water

metal hydroxide + acid ⟶ salt + water

When bases react with acids, the reactions become warmer because they are **exothermic** reactions.

Examples

1 Copper(II) oxide reacts with sulfuric acid to produce copper(II) sulfate and water:

$$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

Ionic equation:

$$CuO(s) + 2H^+(aq) \longrightarrow Cu^{2+}(aq) + H_2O(l)$$

In this reaction, the *black* copper(II) oxide appears to dissolve slowly and a blue solution forms. This is because soluble copper(II) sulfate is produced and the solution contains *blue* Cu^{2+} ions.

2 Sodium hydroxide reacts with hydrochloric acid to produce sodium chloride and water:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Ionic equation:

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$$

Exam tip

It is extremely important that you can write **balanced chemical equations** for the reactions of acids. To help you to do this, it is a good idea to memorise the general word equations given in this section so that you know the general products for each type of reaction. You must also learn the formulae of the common acids.

Practical activity

Investigating reactions of acids

Your teacher may use this activity to assess:

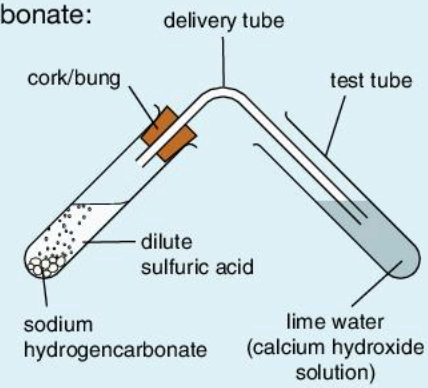
- observation, recording and reporting
- analysis and interpretation.

You will be supplied with magnesium ribbon, zinc, sodium hydrogencarbonate, copper(II) carbonate, calcium hydroxide, copper(II) oxide, hydrochloric acid, sulfuric acid, nitric acid, lime water, test tubes and a wooden splint.

Method

1 Reactions between an acid and metals:
   a Place $2\,cm^2$ of hydrochloric acid in a test tube and add a piece of magnesium ribbon. Observe the reaction. Feel the bottom of the tube to see if it becomes hotter.
   b Place a burning splint at the mouth of the test tube to test for the presence of hydrogen gas. If hydrogen is being produced, the splint will make a squeaky pop and the flame will be extinguished.
   c Repeat the experiment using a small piece of zinc.

2 Reaction between an acid and a hydrogencarbonate:
   a Place $2\,cm^3$ of sulfuric acid in a test tube and add a spatula of sodium hydrogencarbonate. Observe the reaction.
   b Cork the test tube with a delivery tube running through the cork into a test tube of colourless lime water, as shown in Figure 8.1.2, to test for the presence of carbon dioxide. If carbon dioxide is being produced, a white precipitate will form in the lime water.



▲ **Figure 8.1.2** Reaction between sulfuric acid and sodium hydrogencarbonate

3 Reaction between an acid and a carbonate:
   a Place $2\,cm^3$ of sulfuric acid in a test tube and add a spatula of copper(II) carbonate. Observe the reaction.
   b Test for the presence of carbon dioxide as in step **2**.
   c Observe any colour changes occurring during the reaction.

4 Reactions between an acid and bases:
   a Place $2\,cm^3$ of nitric acid in a test tube and add a spatula of calcium hydroxide. Shake the tube and observe the reaction. Feel the bottom of the tube to see if it becomes hotter.
   b Repeat the experiment using a spatula of copper(II) oxide instead of calcium hydroxide. This time shake the tube periodically until no further changes occur. Observe any colour changes.

5 Record all your observations for each reaction, including the results of the tests for hydrogen and carbon dioxide.

6 Write a balanced chemical equation for each reaction.

7 Explain any colour changes you observed in steps **3c** and **4b**, and any heat changes you felt in steps **1a** and **4a**.

During the experiment, you learnt how to test for the presence of hydrogen and carbon dioxide gases.

If **hydrogen** is produced in a reaction, it causes a burning splint to make a *squeaky pop* and to be extinguished. The 'pop' is the sound of a small explosion as the hydrogen reacts explosively with oxygen in the air to form steam:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

If **carbon dioxide** is produced in a reaction, it forms a *white precipitate* when bubbled into lime water. This is because the lime water is a solution of calcium hydroxide, and the carbon dioxide reacts with this forming white, insoluble calcium carbonate and water. The calcium carbonate forms the white precipitate:

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

Classification of acids

Acids can be classified in a variety of ways.

Inorganic and organic acids

Acids can be classified as inorganic acids or organic acids.

- An **inorganic acid** contains a non-metallic element or polyatomic group as well as hydrogen. Examples of inorganic acids are given in Table 8.1.2.
- Most **organic acids** contain the carboxyl group, i.e. COOH. In these acids it is the hydrogen at the end of the carboxyl group which forms the H^+ ion when the acid dissolves in water. Examples of organic acids are given in Table 8.1.3.

Did you know?

Inorganic compounds are of geological origin. Most do not contain carbon, the exceptions being carbonates, hydrogencarbonates and carbon dioxide. **Organic compounds** are mainly of biological origin, e.g. from plants and animals. They always contain carbon and most contain hydrogen.

▼ **Table 8.1.2** Common inorganic acids

Acid	Formula
hydrochloric acid	HCl
nitric acid	HNO_3
sulfuric acid	H_2SO_4
phosphoric acid	H_3PO_4
nitrous acid	HNO_2
sulfurous acid	H_2SO_3
carbonic acid	H_2CO_3

▼ **Table 8.1.3** Common organic acids

Acid	Formula
methanoic acid	$HCOOH$
ethanoic acid	CH_3COOH
lactic acid	$CH_3CH(OH)COOH$
ascorbic acid	$C_6H_8O_6$
citric acid	$C_6H_8O_7$

Monobasic, dibasic and tribasic acids

Acids can also be classified according to the number of H^+ ions they produce per molecule when dissolved in water, known as their **basicity**.

- A **monobasic acid** produces one H^+ ion per molecule when it dissolves in water. Hydrochloric acid (HCl), nitric acid (HNO_3) and ethanoic acid (CH_3COOH) are examples of monobasic acids.
- A **dibasic acid** produces two H^+ ions per molecule when it dissolves in water. Sulfuric acid (H_2SO_4) is an example of a dibasic acid.
- A **tribasic acid** produces three H^+ ions per molecule when it dissolves in water. Phosphoric acid (H_3PO_4) is an example of a tribasic acid.

Key fact

Basicity is the number of H^+ ions produced per molecule of acid when it dissolves in water.

Dilute and concentrated acids

Acids can also be classified based on the quantity of water present.

- A **dilute acid** is one that contains a lot of water, e.g. hydrochloric acid, which has a concentration of 0.1 mol dm^{-3}, is dilute.
- A **concentrated acid** is one that contains very little water, e.g. hydrochloric acid, which has a concentration of 12 mol dm^{-3}, is concentrated.

Strong and weak acids

Acids can also be classified based on the degree of ionisation that occurs when the acid is dissolved in water.

- A **strong acid** is fully ionised when dissolved in water, e.g. hydrochloric acid and sulfuric acid.
- A **weak acid** is only partially ionised when dissolved in water, e.g. carbonic acid and ethanoic acid.

We will be looking at strong and weak acids in more detail in Unit 8.3.

Did you know?

A concentrated acid is always **diluted** by adding the acid slowly to the distilled water. When an acid ionises on being added to water, the reaction gives off a lot of heat energy, i.e. the reaction is exothermic. This can cause spattering and can be dangerous. If the acid is added slowly to the water, the larger volume of water can absorb the heat energy produced.

Acid anhydrides

Certain compounds react with water to form an acid. These are known as **acid anhydrides**.

Key fact

An **acid anhydride** is a compound which reacts with water to form an acid.

Acid anhydride literally means 'an acid without water'. Many acid anhydrides are acidic oxides of non-metals. Carbon dioxide (CO_2), sulfur dioxide (SO_2), sulfur trioxide (SO_3) and nitrogen dioxide (NO_2) are examples of acid anhydrides:

$$CO_2(g) + H_2O(l) \rightleftharpoons \underset{\text{carbonic acid}}{H_2CO_3(aq)}$$

$$SO_2(g) + H_2O(l) \rightleftharpoons \underset{\text{sulfurous acid}}{H_2SO_3(aq)}$$

$$SO_3(g) + H_2O(l) \longrightarrow \underset{\text{sulfuric acid}}{H_2SO_4(aq)}$$

$$2NO_2(g) + H_2O(l) \longrightarrow \underset{\text{nitrous acid}}{HNO_2(aq)} + \underset{\text{nitric acid}}{HNO_3(aq)}$$

Did you know?

Nitrogen dioxide is made naturally in the atmosphere when lightning occurs and it is also present in the exhaust fumes of motor vehicles. When it reacts with water vapour in the air it forms **acid rain**. This acid rain can then cause statues and buildings made of calcium carbonate to slowly dissolve as it reacts with them to make soluble calcium nitrate. It can also make water too acidic for certain aquatic organisms to survive.

Examples of acids in living systems

We come across a variety of organic acids in living organisms. We also use some of these acids in our everyday activities. These are summarised in Table 8.1.4.

▼ **Table 8.1.4** Acids in living systems

Acid	Occurrence	Important points to note
Ascorbic acid or vitamin C ($C_6H_8O_6$)	In many foods, e.g. citrus fruits, West Indian cherries, sweet peppers, tomatoes, green leafy vegetables	A shortage of vitamin C in the diet can lead to scurvy. On exposure to heat, e.g. during cooking, vitamin C is destroyed by being oxidised. Sodium hydrogencarbonate is sometimes added to fruits and vegetables to improve their appearance and texture. This neutralises any vitamin C present, reducing the vitamin C content.
Methanoic acid (HCOOH)	In the venom of bee and ant stings	Causes itching, redness, swelling and pain around the sting. Stings can be treated by applying a paste of sodium hydrogencarbonate or calamine lotion which contains zinc hydrogencarbonate. These both neutralise the acid.
Lactic acid ($C_3H_6O_3$)	Produced in muscle cells during strenuous activity	If too much lactic acid builds up it prevents muscles from contracting and the person collapses.
Ethanoic acid (CH_3COOH)	In vinegar	Vinegar is used to preserve food items. Being acidic, it has a low pH, which denatures (destroys) the enzymes that cause decay and inhibits the growth of bacteria and fungi.

▼ **Table 8.1.4** (continued)

Acid	Occurrence	Important points to note
Citric acid ($C_6H_8O_7$)	In citrus fruits, e.g. limes	Lime juice is used to remove rust stains from clothes. The acid in the lime juice reacts with the iron(III) oxide (Fe_2O_3) in the rust stains making a soluble compound which can be washed out of the clothes: $\underset{\text{iron(III) oxide (rust)}}{Fe_2O_3(s)} + \underset{\text{from the acid}}{6H^+(aq)} \longrightarrow \underset{\text{stain removed}}{2Fe^{3+}(aq) + 3H_2O(l)}$

Summary questions

1 Define the terms 'acid' and 'acid anhydride'.

2 Write equations which show how the following acids ionise when placed in water: sulfuric acid, hydrochloric acid and nitric acid.

3 Give three properties, other than their reactions, which are typical of aqueous acids.

4 Give the products that are formed when acids react with:
  a metal oxides
  b carbonates.

5 Write balanced chemical equations for the following:
  a the reaction between zinc hydroxide and hydrochloric acid
  b the reaction between magnesium hydrogencarbonate and nitric acid
  c the reaction between aluminium and sulfuric acid.

6 How does vinegar work to preserve some of the foods we eat?

7 Why is sodium hydrogencarbonate used as a treatment for ant stings?

Objectives

By the end of this topic you will be able to:

- define a base in terms of protons
- define an alkali
- give the general properties of aqueous alkalis
- describe the reactions of bases with acids and ammonium salts
- define an amphoteric substance
- classify oxides into acidic, basic, amphoteric and neutral.

A8.2 Properties and reactions of bases

A **base** can be defined as a substance which reacts with an acid to form a salt and water. Bases are normally metal oxides and metal hydroxides. However, ammonia is also classified as a base. Examples of bases include magnesium oxide (MgO), copper(II) oxide (CuO), magnesium hydroxide ($Mg(OH)_2$) and copper(II) hydroxide ($Cu(OH)_2$).

You learnt in Unit 8.1 that when a base reacts with an acid, the acid donates **protons** to the base, forming water. The base has *accepted* the protons from the acid and this fact can be used to define a base as a proton acceptor.

Using the same example as in Unit 8.1, when hydrochloric acid reacts with sodium hydroxide, the OH^- ions of the sodium hydroxide accept the H^+ ions, or protons, from the hydrochloric acid forming water, as summarised by the following ionic equation:

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$$

Alkalis

Some bases are soluble in water, others are insoluble. Soluble bases are known as **alkalis**. When they dissolve in water, they form hydroxide ions, or OH^- ions.

The common alkalis include potassium hydroxide (KOH) and sodium hydroxide (NaOH), which are fully soluble in water, and calcium hydroxide ($Ca(OH)_2$), which is moderately soluble. For example:

$$NaOH(s) + \text{water} \longrightarrow Na^+(aq) + OH^-(aq)$$

Key fact

A **base** is a proton acceptor.

Key fact

An **alkali** is a base which dissolves in water to form a solution containing OH^- ions.

When ammonia gas is added to water it reacts with water to form ammonium hydroxide (NH_4OH), which is also an alkali:

$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Potassium oxide (K_2O), sodium oxide (Na_2O) and calcium oxide (CaO) also react with water to form the equivalent hydroxide. For example:

$$Na_2O(s) + H_2O(l) \longrightarrow 2Na^+(aq) + 2OH^-(aq)$$

Like acids, alkalis can be classified into strong alkalis and weak alkalis based on the degree of ionisation that occurs when the alkali is dissolved in water.

- A **strong alkali** is fully ionised when dissolved in water, e.g. potassium hydroxide and sodium hydroxide.
- A **weak alkali** is only partially ionised when dissolved in water, e.g. ammonia.

We will be looking at strong and weak alkalis in more detail in Unit 8.3.

▲ **Figure 8.2.1** These products contain bases

General properties of aqueous alkalis

The presence of *OH^- ions* in aqueous solution gives alkalis their characteristic properties. Aqueous solutions of alkalis are described as being **alkaline** and they have the following common properties:

- they have a bitter taste
- they change red litmus to blue
- they have a pH value of more than 7
- they are corrosive
- they are electrolytes, i.e. they conduct an electric current
- they feel soapy when touched.

Chemical reactions of bases

Since bases contain oxide or hydroxide ions, bases have common reactions with certain other reactants. The reactions of bases are discussed in detail below.

Bases react with acids

Bases react with acids to produce a salt and water.

These reactions, which you looked at in Unit 8.1, can be summarised by the following general word equation:

$$\textbf{base + acid} \longrightarrow \textbf{salt + water}$$

Remember that when bases react with acids, the reactions become warmer because they are exothermic reactions. Also, any insoluble bases which react to produce soluble salts, appear to dissolve.

Examples

1 Magnesium oxide reacts with hydrochloric acid to produce magnesium chloride and water:

$$MgO(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$$

Ionic equation:

$$MgO(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2O(l)$$

2 Potassium hydroxide reacts with sulfuric acid to produce potassium sulfate and water:

$$2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

Ionic equation:

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$$

Bases react with ammonium salts

Bases react with ammonium salts to produce a salt, ammonia and water:

base + ammonium salt ⟶ salt + ammonia + water

In order to react, the reactants need to be heated.

Examples

1 Copper(II) oxide reacts with ammonium chloride to produce copper(II) chloride, ammonia and water:

$$CuO(s) + 2NH_4Cl(s) \longrightarrow CuCl_2(s) + 2NH_3(g) + H_2O(l)$$

2 Calcium hydroxide reacts with ammonium sulfate to produce calcium sulfate, ammonia and water:

$$Ca(OH)_2(s) + (NH_4)_2SO_4(s) \longrightarrow CaSO_4(s) + 2NH_3(g) + 2H_2O(l)$$

Practical activity

Investigating the reaction between calcium hydroxide and ammonium chloride

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with calcium hydroxide, ammonium chloride, red litmus paper, a dry test tube and tongs.

Method

1 Place two small spatulas of calcium hydroxide in the test tube.
2 Add two small spatulas of ammonium chloride and shake the tube to mix the solids.
3 Carefully heat the mixture and while heating hold a piece of moist red litmus paper across the mouth of the tube to test for the presence of ammonia gas. If ammonia gas is produced the litmus paper will turn blue.
4 Look at the sides of the test tube to see if any water droplets form.
5 Write a balance chemical equation for the reaction.
6 Explain the formation of the water droplets.

Did you know?

Ammonia gas causes moist red litmus paper to turn blue because it reacts with the water present to form ammonium hydroxide which is alkaline:

$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4OH(aq)$$

Amphoteric oxides and hydroxides

Some metal oxides and hydroxides can behave as bases because they can react with acids and also behave as acids because they can react with strong alkalis. These substances are referred to as being amphoteric.

The following general word equations summarise how amphoteric oxides and hydroxides react.

- The amphoteric substance reacting as a base:

amphoteric oxide or hydroxide + acid ⟶ salt + water

- The amphoteric substance reacting as an acid:

strong alkali + amphoteric oxide or hydroxide ⟶ salt + water

The amphoteric oxides and hydroxides are listed in Table 8.2.1.

Key fact

An **amphoteric substance** is a substance which can react with both acids and with strong alkalis.

▼ **Table 8.2.1** Amphoteric oxides and hydroxides

Amphoteric oxide	Formula	Amphoteric hydroxide	Formula
aluminium oxide	Al_2O_3	aluminium hydroxide	$Al(OH)_3$
zinc oxide	ZnO	zinc hydroxide	$Zn(OH)_2$
lead(II) oxide	PbO	lead(II) hydroxide	$Pb(OH)_2$

Examples

1 Aluminium hydroxide reacts with hydrochloric acid to produce aluminium chloride and water:

$$Al(OH)_3(s) + 3HCl(aq) \longrightarrow AlCl_3(aq) + 3H_2O(l)$$

2 Aluminium hydroxide reacts with the strong alkali, sodium hydroxide, to produce sodium aluminate and water:

$$NaOH(aq) + Al(OH)_3(s) \longrightarrow \underset{\text{sodium aluminate}}{NaAlO_2(aq)} + 2H_2O(l)$$

3 Zinc oxide reacts with hydrochloric acid to produce zinc chloride and water:

$$ZnO(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2O(l)$$

4 Zinc oxide reacts with the strong alkali, sodium hydroxide, to produce sodium zincate and water:

$$2NaOH(aq) + ZnO(s) \longrightarrow \underset{\text{sodium zincate}}{Na_2ZnO_2(aq)} + H_2O(l)$$

Classification of oxides

So far in this unit you have come across acidic, basic and amphoteric oxides. Another type of oxide also exists, known as a neutral oxide. Oxides can, therefore, be classified as acidic, basic, amphoteric and neutral.

Acidic oxides

Acidic oxides are oxides of certain non-metals which react with alkalis to form a salt and water, for example,

$$2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

Acidic oxides also react with water to form an acid, for example,

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

Examples of acidic oxides include carbon dioxide (CO_2), sulfur dioxide (SO_2), sulfur trioxide (SO_3) and nitrogen dioxide (NO_2).

! Key fact

Acidic oxides are oxides of certain non-metals which react with alkalis to form a salt and water.

Basic oxides

Basic oxides are oxides of metals which react with acids to form a salt and water, for example,

$$CuO(s) + 2HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + H_2O(l)$$

A few basic oxides also react with water to form a hydroxide, i.e. they are alkalis, for example,

$$K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$$

Examples of basic oxides include magnesium oxide (MgO), iron(III) oxide (Fe_2O_3) and copper(II) oxide (CuO).

! Key fact

Basic oxides are oxides of metals which react with acids to form a salt and water.

Basic oxides which are alkalis include potassium oxide (K_2O), sodium oxide (Na_2O) and calcium oxide (CaO).

Amphoteric oxides

Amphoteric oxides are the oxides of certain metals which react with both acids and strong alkalis to form a salt and water, for example,

$$PbO(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l)$$
$$2NaOH(aq) + PbO(s) \longrightarrow \underset{\text{sodium plumbate}}{Na_2PbO_2(aq)} + H_2O(l)$$

The three amphoteric oxides are aluminium oxide (Al_3O_3), zinc oxide (ZnO) and lead(II) oxide (PbO).

Key fact

Amphoteric oxides are the oxides of certain metals which react with both acids and strong alkalis to form a salt and water.

Neutral oxides

Neutral oxides are oxides of certain non-metals which do not react with acids or alkalis. These include carbon monoxide (CO), nitrogen monoxide (NO) and dinitrogen monoxide (N_2O).

Key fact

Neutral oxides are oxides of certain non-metals which do not react with acids or alkalis.

Summary questions

1 Define the terms 'base', 'alkali' and 'amphoteric substance'.
2 Give three properties, other than their reactions, which are typical of aqueous alkalis.
3 Write balanced chemical equations for the following:
  a the reaction between magnesium hydroxide and nitric acid
  b the reaction between calcium oxide and ammonium nitrate
  c the reaction between sodium hydroxide and zinc hydroxide.
4 Calcium oxide is a basic oxide. Explain, using a chemical equation, why calcium oxide is basic.

A8.3 Strength of acids and alkalis

Objectives

By the end of this topic you will be able to:

- explain the difference between a strong acid and a weak acid
- give examples of strong and weak acids
- explain the difference between a strong alkali and a weak alkali
- give examples of strong and weak alkalis
- explain the relationship between acidity, alkalinity and the pH scale
- give the pH values for strong acids, weak acids, strong alkalis and weak alkalis
- explain how to determine the pH of a solution
- give the colours of certain indicators in acidic and alkaline solutions.

Aqueous solutions of both acids and alkalis can be classified as strong or weak based on the **degree of ionisation** which occurs when they are dissolved in water. You must not confuse this with concentration. The concentration of an acid or alkali is based on the quantity of water present in the solution (see Unit 8.1).

Strong and weak acids

When an acid dissolves in water, its molecules ionise to form hydrogen ions, or H^+ ions, and negative anions.

Strong acids

A **strong acid** is fully ionised when it dissolves in water. This means that all of the acid molecules ionise and there is a high concentration of H^+ ions in the solution. Hydrochloric acid is an example of a strong acid:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Other strong acids include nitric acid (HNO_3) and sulfuric acid (H_2SO_4).

Weak acids

A **weak acid** is only partially ionised when it dissolves in water. This means that the solution contains a mixture of acid molecules which have not ionised and ions produced from the molecules which have ionised. The solution has a low concentration of H^+ ions. Ethanoic acid is an example of a weak acid:

$$CH_3COOH(aq) \rightleftharpoons \underset{\text{ethanoate ion}}{CH_3COO^-(aq)} + H^+(aq)$$

At any one time, only about 1% of ethanoic acid molecules are ionised, i.e. the solution contains 1% ethanoate ions (CH_3COO^-) and hydrogen ions (H^+) and 99% ethanoic acid molecules (CH_3COOH).

Most organic acids are weak. Weak inorganic acids include nitrous acid (HNO_2), sulfurous acid (H_2SO_3) and carbonic acid (H_2CO_3).

Strong and weak alkalis

When an alkali is added to water, it ionises to form positive cations and hydroxide ions, or OH^- ions.

Strong alkalis

A **strong alkali** is fully ionised when it dissolves in water. The solution contains a high concentration of OH^- ions. Sodium hydroxide is an example of a strong alkali:

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

Potassium hydroxide (KOH) is another example of a strong alkali.

Weak alkalis

A **weak alkali** is only partially ionised when it dissolves in water. The solution contains a low concentration of OH^- ions. Ammonia is an example of a weak alkali:

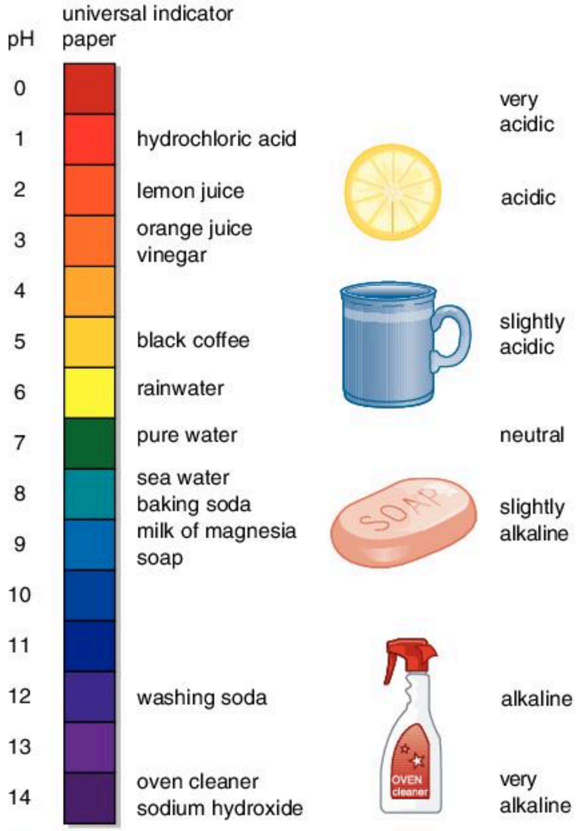
$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4{}^+(aq) + OH^-(aq)$$

Measuring the strength of acids and alkalis

The strength of an aqueous acid or alkali is measured using the **pH scale**. The pH scale is a number scale with numbers ranging from 0 to 14. Aqueous solutions with a pH less than 7 are considered to be *acidic* and those with a pH greater than 7 are considered to be *alkaline*. An aqueous solution with a pH of 7 is defined as a *neutral* solution. Distilled water is neutral.

The pH scale also tells us how strong the acidic or alkaline solution is. In general, the *lower* the pH, the *stronger* the acid, and the *higher* the pH, the *stronger* the alkali.

- A strong acid has a pH close to 1, e.g. hydrochloric acid (HCl), nitric acid (HNO_3) and sulfuric acid (H_2SO_4).
- A weak acid has a pH of about 4 or 5, e.g. ethanoic acid (CH_3COOH).
- A neutral substance has a pH of 7, e.g. distilled water (H_2O).
- A weak alkali has a pH of about 9 or 10, e.g. aqueous ammonia (NH_4OH).
- A strong alkali has a pH close to 14, e.g. sodium hydroxide (NaOH) and potassium hydroxide (KOH).



▲ **Figure 8.3.1** The pH scale

Did you know?

Many plant parts contain coloured compounds called **anthocyanins** which are red in acidic conditions and purple-blue in alkaline conditions. They are responsible for the colour of many flowers and fruit skins. Anthocyanins can be extracted easily from these plant parts, e.g. the leaves of red cabbage, and used as a crude acid–alkali indicator. Litmus is a naturally occurring indicator which is made by extracting dyes from certain species of lichens.

There are two main methods that are used in the laboratory to measure the pH of a solution. One is universal indicator and the other is a pH meter.

- **Universal indicator** is a chemical substance which indicates the pH of a solution by colour, i.e. its colour changes depending on the pH of the solution. Universal indicator can be in paper form or solution form. When using the paper, a small piece of paper is dipped into the solution and its colour is compared with a pH colour chart. Figure 8.3.1 shows the colour of universal indicator paper at different pH values.
- A **pH meter** is an electronic instrument. It consists of a measuring probe connected to an electronic meter, which displays the pH reading. A pH meter is a more accurate method for determining the pH of a solution than universal indicator paper.

Other acid–alkali indicators

There are other **indicators** which change to one colour in an acidic solution and another colour in an alkaline solution. It is important to note that these indicators do not measure the pH of a solution, they only indicate if it is acidic or alkaline.

Litmus is an example of this type of indicator. Other indicators which you may come across in the laboratory include methyl orange, screened methyl orange, phenolphthalein and bromothymol blue. These are used in neutralisation experiments to indicate the neutralisation point (see Unit 8.5). The common indicators and their colours in acidic and alkaline solutions are given in Table 8.3.1.

▼ **Table 8.3.1** Indicators used in the laboratory

Indicator	Colour in acidic solution	Colour in alkaline solution
litmus	red	blue
methyl orange	red	yellow
screened methyl orange	red	green
phenolphthalein	colourless	pink
bromothymol blue	yellow	blue

Practical activity

Investigating the acidity and alkalinity of various solutions

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

You will be supplied with samples of solutions to be tested, test tubes, universal indicator paper, pH colour chart, red and blue litmus paper, methyl orange and phenolphthalein.

Method

1. You are going to measure the pH of each of the solutions and investigate the colour of each indicator in each solution. Only use methyl orange and phenolphthalein when the solution you are testing is colourless.
   - **a** Place 2 cm^3 of the solution to be tested in each of four test tubes.
   - **b** Dip a piece of universal indicator paper into the solution in the first tube and observe its colour. Compare the colour with the pH colour chart to determine the pH of the solution.
   - **c** Dip a piece of red and a piece of blue litmus paper into the solution in the second tube. Observe their colours.
   - **d** Place two drops of the methyl orange and phenolphthalein, respectively, into each of the other two tubes. Observe their colours.
2. If there is a pH meter available, use the meter to determine the pH of each of the solutions. Ensure that you rinse the measuring probe of the meter in distilled water after each reading.
3. Record all your results in a table.
4. Classify each substance as a strong acid, a weak acid, a strong alkali, a weak alkali or a neutral substance.

Summary questions

1. What is the difference between a strong acid and a weak acid?
2. Name three strong acids and three weak acids.
3. Describe the pH scale, giving the range of the scale and the pH values of acidic, alkaline and neutral solutions.
4. Give the colours of methyl orange and phenolphthalein and the pH values for each of the following:
   - **a** aqueous ethanoic acid
   - **b** sodium hydroxide solution
   - **c** aqueous ammonia
   - **d** aqueous hydrochloric acid.

A8.4 Salts

You have already learnt that all acids form at least one hydrogen ion per molecule when they dissolve in water, e.g. hydrochloric acid (HCl) forms one hydrogen ion per molecule and sulfuric acid (H_2SO_4) forms two. These hydrogen ions can be replaced by other positive ions when acids react and they are known as **replaceable hydrogen ions**. When acids react and the hydrogen ions are replaced by metal or ammonium ions from the other reactant, a **salt** is formed.

Classification of salts

Salts can be classified into two groups, normal salts and acid salts, based on whether all or only some of the hydrogen ions are replaced:

Objectives

By the end of this topic you will be able to:

- define the term salt
- distinguish between a normal salt and an acid salt
- explain water of crystallisation
- distinguish between a hydrated salt and an anhydrous salt
- decide on the most appropriate method of preparing a salt
- describe the different methods of preparing salts
- describe the uses of salts in everyday life
- state some of the dangers of salts.

Key fact

A **salt** is a compound formed when some or all of the hydrogen ions in an acid are replaced by metal or ammonium ions.

Normal salts

Normal salts are formed when *all* of the hydrogen ions in an acid are replaced by metal or ammonium ions. For example, when sodium hydroxide (NaOH) reacts with sulfuric acid (H_2SO_4), if all of the H^+ ions in the acid are replaced by Na^+ ions, the normal salt, sodium sulfate (Na_2SO_4), is formed:

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

All acids can form normal salts.

Acid salts

Acid salts are formed when the hydrogen ions in an acid are only *partially* replaced by metal or ammonium ions. For example, when sodium hydroxide reacts with sulfuric acid, if only one of the H^+ ions in the acid is replaced by a Na^+ ion, the acid salt, sodium hydrogensulfate ($NaHSO_4$), is formed:

$$NaOH(aq) + H_2SO_4(aq) \longrightarrow NaHSO_4(aq) + H_2O(l)$$

Only *dibasic* and *tribasic* acids can form acid salts.

The type of salt formed by dibasic and tribasic acids depends on the relative quantity of each reactant. For example, in the two reactions above, when 2 mol of sodium hydroxide and 1 mol of sulfuric acid reacted, a normal salt was produced. However, when only 1 mol of sodium hydroxide reacted with 1 mol of sulfuric acid, an acid salt was produced.

The tribasic acid, phosphoric acid (H_3PO_4), can produce *one* normal salt and *two* different acid salts. For example, in the reaction with sodium hydroxide, it can form the normal salt, sodium phosphate (Na_3PO_4), when all three H^+ ions are replaced. It can also form the two acid salts, sodium hydrogenphosphate (Na_2HPO_4) when two H^+ ions are replaced and sodium dihydrogenphosphate (NaH_2PO_4) when one H^+ ion is replaced:

$$3NaOH(aq) + H_3PO_4(aq) \longrightarrow Na_3PO_4(aq) + 3H_2O(l)$$
$$2NaOH(aq) + H_3PO_4(aq) \longrightarrow Na_2HPO_4(aq) + 2H_2O(l)$$
$$NaOH(aq) + H_3PO_4(aq) \longrightarrow NaH_2PO_4(aq) + H_2O(l)$$

Again, you can see that the relative quantity of the two reactants determines which salt is produced.

Table 8.4.1 shows the salts formed by the common acids.

▼ **Table 8.4.1** Salts formed by common acids

Acid	Formula	Salt(s) formed	Anion present	Type of salt	Example
hydrochloric acid	HCl	chlorides	Cl^-	normal salt	NaCl
nitric acid	HNO_3	nitrates	NO_3^-	normal salt	$NaNO_3$
ethanoic acid	CH_3COOH	ethanoates	CH_3COO^-	normal salt	CH_3COONa
sulfuric acid	H_2SO_4	sulfates hydrogensulfates	SO_4^{2-} HSO_4^-	normal salt acid salt	Na_2SO_4 $NaHSO_4$
carbonic acid	H_2CO_3	carbonates hydrogencarbonates	CO_3^{2-} HCO_3^-	normal salt acid salt	Na_2CO_3 $NaHCO_3$
phosphoric acid	H_3PO_4	phosphates hydrogenphosphates dihydrogenphosphates	PO_4^{3-} HPO_4^{2-} $H_2PO_4^-$	normal salt acid salt acid salt	Na_3PO_4 Na_2HPO_4 NaH_2PO_4

Water of crystallisation

Some salts contain a fixed number of water molecules between the ions in their crystal lattice, known as **water of crystallisation**. Salts containing water of crystallisation are referred to as **hydrated salts**. Water of crystallisation can be given in the formula of the salt as shown in the following examples.

- $CuSO_4.5H_2O$ represents hydrated copper(II) sulfate. The formula shows that for every mole of Cu^{2+} and SO_4^{2-} ions within the crystal lattice, there are *five* moles of water molecules.
- $CoCl_2.6H_2O$ represents hydrated cobalt(II) chloride.
- $MgSO_4.7H_2O$ represents hydrated magnesium sulfate.
- $FeSO_4.7H_2O$ represents hydrated iron(II) sulfate.

Water of crystallisation is essential to develop the crystalline structure, and sometimes the colour, of the crystals containing it. If this water of crystallisation is removed by heating, the solid loses its crystalline structure, it becomes powdery and its colour may also change. For example, hydrated copper(II) sulfate crystals are bright blue and have a regular shape. When heated they become white and powdery:

$$\underset{\text{blue crystals}}{CuSO_4.5H_2O(s)} \xrightarrow{\text{heat}} \underset{\substack{\text{white}\\\text{powder}}}{CuSO_4(s)} + 5H_2O(g)$$

Salts without water of crystallisation are known as **anhydrous salts**.

Preparation of salts

We use salts in all aspects of our lives. Some of these uses will be discussed later in this unit. There are many different methods by which these salts can be produced, some of which we will be discussing in this section.

When deciding on a preparation method, it is important to know the *solubility* of the salt that is to be produced and the *solubility* of the compounds which may be used to prepare the salt. Methods of preparing insoluble salts differ from those used to prepare soluble salts.

You can find a table summarising the solubility rules in Unit 6.1.

Preparation of insoluble salts

Insoluble salts can be prepared by reacting two solutions, one containing the cations of the salt being prepared and the other containing the anions of the salt. This is known as **ionic precipitation**. In an ionic precipitation reaction, the two soluble salts in solution react to form an insoluble salt, i.e. a precipitate, and a soluble salt. For example, barium nitrate solution reacts with sodium sulfate solution to form insoluble barium sulfate and sodium nitrate solution:

$$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NaNO_3(aq)$$

Ionic equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

Method

The following general method is used for preparing insoluble salts.

1) Choose two soluble salts, one containing the *cations* required to make the salt and the other containing the *anions* required.
2) Dissolve the two salts in water to make solutions.
3) Mix the two solutions to form the insoluble salt as a precipitate.
4) Filter the mixture and collect the precipitate as the *residue*.
5) Wash the residue with distilled water while it is still in the filter funnel and leave it to dry.

When choosing the two salts to use in your preparation it is important to remember that all nitrates are soluble, as are all sodium salts. Therefore, the best salts to choose would be a *nitrate* to supply the cations, i.e. barium nitrate in the above preparation, and a *sodium salt* to supply the anions, e.g. sodium sulfate.

Practical activity

To prepare an insoluble salt by ionic precipitation

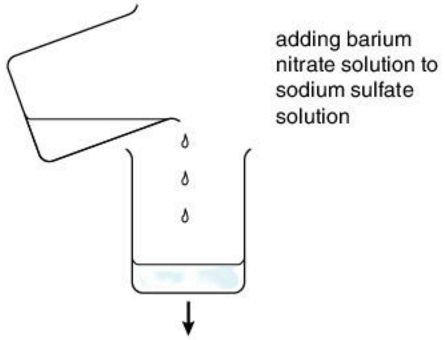
Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

You will be supplied with solutions of barium nitrate and sodium sulfate, two beakers, a filter funnel and filter paper.

Method

1. Place 10 cm^3 of sodium sulfate solution into a small beaker.
2. Add 10 cm^3 of barium nitrate solution and swirl to mix.
3. Using the filter funnel and filter paper, filter the mixture.
4. Wash the residue with distilled water while it is still in the filter funnel and allow it to dry.
5. Record all your observations as you prepare the sample of the salt.
6. Write both a balanced chemical equation and an ionic equation for the reaction and name the insoluble salt that you have prepared.



▲ **Figure 8.4.1** Preparation of the insoluble salt, barium sulfate

▲ **Figure 8.4.2** A precipitate of barium sulfate

Preparation of soluble salts

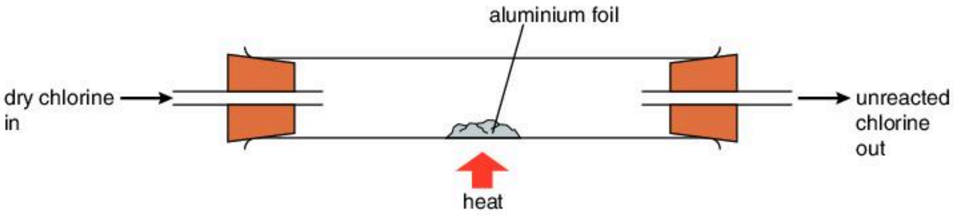
There are *three* main methods for preparing soluble salts:

- direct combination
- the reaction between a reactive metal, an insoluble carbonate or an insoluble base and an acid
- the reaction between a soluble base, i.e. an alkali, and an acid, known as a titration.

Direct combination

Salts composed of two simple ions, such as metal chlorides, can be prepared by reacting two elements, a metal and a non-metal, directly with each other. The metal supplies the cations of the salt and the non-metal supplies the anions.

For example, aluminium chloride can be prepared in a fume cupboard by passing a stream of dry chlorine gas over heated aluminium foil in a long glass tube as shown in Figure 8.4.3.



▲ **Figure 8.4.3** Preparation of aluminium chloride

The equation for the reaction is:

$$2Al(s) + 3Cl_2(g) \longrightarrow 2AlCl_3(s)$$

Since the reaction occurs in the absence of any moisture, the salt formed is an **anhydrous** salt. This is particularly suitable for preparing anhydrous aluminium chloride and anhydrous iron(III) chloride.

Reactions with acids

Many soluble salts can be prepared by reacting a reactive metal, insoluble carbonate or insoluble base with an acid as shown in the examples below.

- Reacting a **reactive metal** and an acid, e.g. to prepare magnesium chloride:

  $$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

  This method cannot be used to prepare salts of very reactive metals, e.g. potassium or sodium salts, because the reaction of these metals with an acid is too violent.
- Reacting an **insoluble carbonate** and an acid, e.g. to prepare calcium nitrate:

  $$CaCO_3(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$
- Reacting an **insoluble base** and an acid, e.g. to prepare copper(II) sulfate:

  $$CuO(S) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

In these reactions, the final product will only be a pure solution of the required salt if the reaction has reached completion and no acid remains. One way to achieve this is by using a reactant which is insoluble. This reactant is added to the acid until some remains. This indicates there is no more acid left for the reactant to react with and, since it is insoluble, the excess will remain undissolved in the solution of the salt.

Method

The following general method is used to prepare soluble salts using reactions with acids.

1) Choose the appropriate reactive metal, insoluble carbonate or insoluble base to provide the *cations* and the appropriate acid to provide the *anions*.
2) Place the acid into a beaker and add the metal, carbonate or base until excess solid remains and effervescence stops if a metal or carbonate is used. This indicates that all the acid has reacted. Heating may be required to speed up the reaction when using a metal or a base.
3) Dip a piece of blue litmus paper into the solution to ensure all the acid has reacted. The litmus should remain blue.
4) Remove the excess, unreacted solid by filtration.

Exam tip

Being able to describe the **methods** used to prepare salts is very important for exams. If you are asked to describe how to prepare a soluble salt using a reaction with an acid and you can choose the reactants, it is a good idea to choose the appropriate insoluble carbonate to supply the cations since the reaction between a carbonate and an acid is rapid, it doesn't need heating and visible effervescence stops when the reaction has reached completion.

5) Collect the *filtrate* and evaporate the water over a beaker of boiling water. If a hydrated salt is required, evaporate some of the water to concentrate the solution and leave the concentrated solution to crystallise.

Practical activity

To prepare a soluble salt by reacting a base with an acid

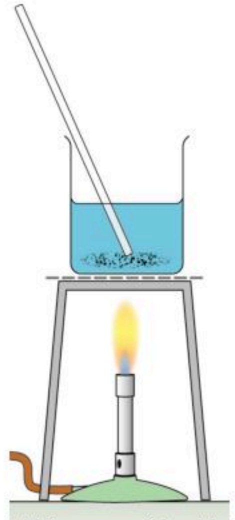
Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

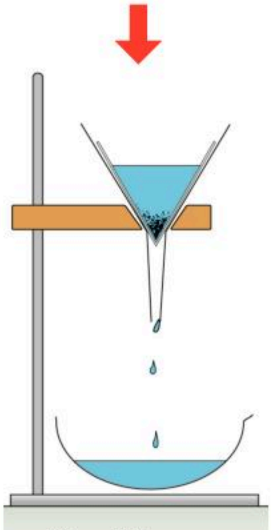
You will be supplied with samples of copper(II) oxide and sulfuric acid, a beaker, an evaporating dish, a filter funnel and filter paper.

Method

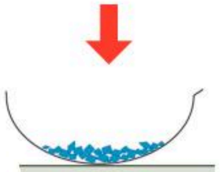
1. Place 20 cm^3 of sulfuric acid into the beaker.
2. Add copper(II) oxide to the acid a spatula at a time, heating and stirring gently, until no more will react, i.e. until you can see excess copper(II) oxide in the beaker.
3. Using the filter funnel and filter paper, filter the mixture to remove the excess copper(II) oxide and collect the filtrate.
4. Pour the filtrate into the evaporating dish. Place the evaporating dish over a beaker of boiling water and heat until the solution becomes concentrated.
5. Place a piece of filter paper over the evaporating dish containing the concentrated solution and leave it to crystallise and form hydrated copper(II) sulfate crystals.
6. Record all your observations as you prepare the salt.
7. Write a balanced chemical equation for the reaction.
8. To produce anhydrous copper(II) sulfate, you can heat the filtrate in the evaporating dish over a beaker of boiling water until all the water has evaporated.



Add the copper(II) oxide to the sulfuric acid while heating gently, until no more will dissolve



Filter off the excess copper(II)oxide



To produce hydrated copper(II) sulfate allow the crystals to form over a few days. To produce anhydrous copper(II) sulfate evaporate over a beaker of boiling water

▲ **Figure 8.4.4** Steps involved in the preparation of hydrated copper(II) sulfate crystals

Titration

Potassium, sodium and ammonium salts cannot be prepared by the method which you have just learnt for two reasons. Firstly, potassium and sodium react in a violent way with acids making the reactions too dangerous. Secondly, potassium, sodium and ammonium carbonates and hydroxides are all soluble, so when the reaction has reached completion and no more acid remains, the excess solid *dissolves* and the solution of the salt will then not be pure.

Potassium, sodium and ammonium salts are prepared by adding an acid to an aqueous alkali until the solution is just neutral, indicating that the reaction has reached completion. For example, to prepare potassium sulfate, sulfuric acid is added to potassium hydroxide solution until the resulting solution is just neutral:

$$2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

The colour change of an indicator is used to determine the **neutralisation point**. The technique used to determine the exact volume of acid needed to neutralise a fixed volume of aqueous alkali is known as a **titration**. A carbonate solution may also be used in place of the aqueous alkali.

Method

The general method for performing a titration is as follows.

1) Choose an appropriate alkali or soluble carbonate to provide the *cations* and an appropriate acid to supply the *anions*.
2) Measure a fixed volume of the aqueous alkali or carbonate using a pipette. Run it into a conical flask and add a few drops of indicator solution, e.g. phenolphthalein.
3) Place the acid in a burette and take an initial burette reading.
4) Add the acid to the aqueous alkali or carbonate until the neutralisation point is reached.
5) Take a final burette reading and determine the volume of acid added.
6) Repeat the titration until you have three volumes of acid within 0.1 cm^3 of each other. Average these volumes to determine the volume of acid needed.
7) Add this volume of acid to the fixed volume of aqueous alkali or carbonate without the indicator.
8) Evaporate the water from the solution.

Practical activity

To prepare a soluble salt by titration

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

You will be supplied with 0.1 $mol\,dm^{-3}$ sodium hydroxide solution, 0.1 $mol\,dm^{-3}$ hydrochloric acid, phenolphthalein solution, a 25 cm^3 pipette, a conical flask and a burette.

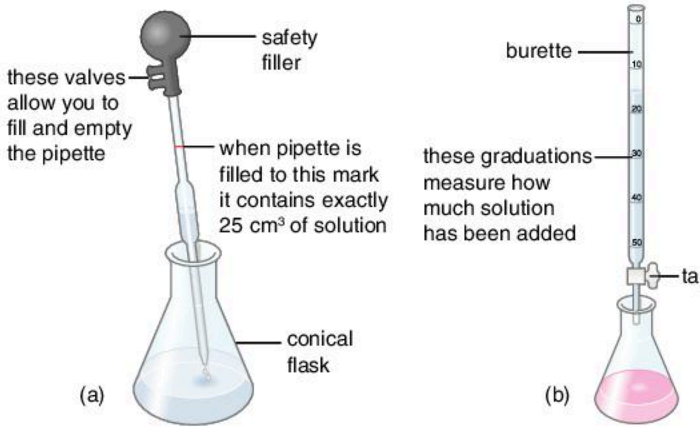
Method

1. Rinse the burette with some hydrochloric acid and fill it with the acid.
2. Rinse the pipette with some sodium hydroxide solution. Measure 25 cm^3 of sodium hydroxide solution in the pipette and run it into the conical flask.
3. Add a few drops of phenolphthalein solution to the conical flask and swirl the contents.
4. Take an initial burette reading, remembering to take the reading from the bottom of the meniscus.
5. Slowly add the hydrochloric acid to the conical flask, swirling constantly, until the solution just turns from pink to colourless. The neutralisation point is the point where the colour changes after the addition of just one drop of the acid.
6. Take a final burette reading and determine the volume of hydrochloric acid added. This is your rough titration.
7. Discard the contents of the conical flask and rinse the flask thoroughly with distilled water.
8. Repeat your titration until you have three volumes of acid which are within 0.1 cm^3. Each time you can add the acid rapidly until you have added about 1 cm^3 less than you added in your rough titration, then almost close the tap and add the acid drop by drop until the neutralisation point is reached.
9. Average the three volumes of acid, which are within 0.1 cm^3. This is the accurate volume of acid needed to neutralise 25 cm^3 of the sodium hydroxide solution. You can now prepare the salt.
10. Repeat your titration, but this time do not add the phenolphthalein solution. When you approach the volume determined in step **9**, add the acid drop by drop until you have added the correct volume.
11. Place the solution into an evaporating dish and carefully evaporate all the water to obtain sodium chloride.

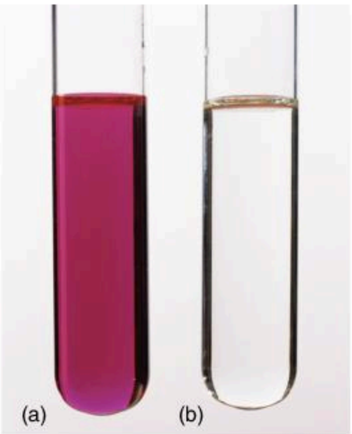
12 Record your results in a table like the one opposite.

13 Write a chemical equation for the reaction.

Burette reading	**Titration number**			
	Rough	**1**	**2**	**3**
Final reading/cm^3				
Initial reading/cm^3				
Volume of acid added/cm^3				



▲ **Figure 8.4.5** **(a)** Using a pipette to measure a fixed volume of sodium hydroxide solution and **(b)** using a burette to add a varying volume of hydrochloric acid



▲ **Figure 8.4.6** Phenolphthalein changes from **(a)** pink to **(b)** colourless when hydrochloric acid is added to sodium hydroxide solution

Preparation of acid salts

You learnt earlier in this unit that when sodium hydroxide (NaOH) reacts with sulfuric acid (H_2SO_4) there are two possibilities:

The formation of a normal salt:

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The formation of an acid salt:

$$NaOH(aq) + H_2SO_4(aq) \longrightarrow NaHSO_4(aq) + H_2O(l)$$

Looking at the balanced equations for these two reactions:

- to produce a normal salt, the ratio of reactants is 2 mol of sodium hydroxide to 1 mol of sulfuric acid
- to produce an acid salt, the ratio of reactants is 1 mol of sodium hydroxide to 1 mol of sulfuric acid.

If the concentrations of the sodium hydroxide solution and the sulfuric acid are the same, we can apply the ratio to volumes. For example, if we use 0.1 mol dm^{-3} sodium hydroxide solution and 0.1 mol dm^{-3} sulfuric acid:

- to produce a normal salt, we would react 50 cm^3 of sodium hydroxide solution with 25 cm^3 of sulfuric acid
- to produce an acid salt, we would react 50 cm^3 of sodium hydroxide solution with 50 cm^3 of sulfuric acid.

Comparing the two, to prepare the acid salt we would need *double* the volume of acid that we used to prepare the normal salt.

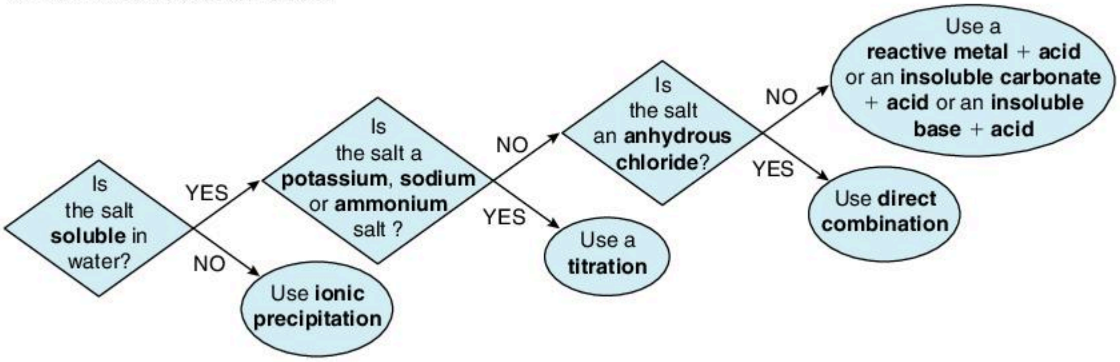
Method

The general method for preparing the acid salt, sodium hydrogensulfate, is as follows.

- Determine the volume of sulfuric acid needed to neutralise a fixed volume of sodium hydroxide solution using a titration.
- Repeat the titration without the indicator and adding *twice* the volume of sulfuric acid determined.
- Evaporate the water from the solution.

Summary of salt preparation

You have just learnt that there are several methods that can be used to prepare salts and it is important that you choose the correct method. The flow chart in Figure 8.4.7 outlines a series of questions to ask which will help you decide which method to choose.



▲ **Figure 8.4.7** Flow chart for the preparation of salts

Uses of salts in everyday life

Salts play an extremely important part in our everyday lives, from the sodium chloride that we use to flavour our food to the sodium carbonate that we use as washing soda. Table 8.4.2 gives some of the uses of salts.

▲ **Figure 8.4.8** Washing soda is a salt which some of us have in our homes

▼ **Table 8.4.2** Uses of salts

Salt	Use	Explanation
Sodium hydrogencarbonate ($NaHCO_3$)	An ingredient in baking powder	Baking powder is used to make cakes rise. It contains sodium hydrogencarbonate and an acid. When mixed with the liquid in the cake mixture, the two compounds react forming carbon dioxide: $HCO_3^-(aq) + H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$ The carbon dioxide forms bubbles in the cake which expand on heating causing the cake to rise.
Sodium carbonate (washing soda) ($NaCO_3$)	To soften hard water	Hard water does not lather with soap. It is caused by dissolved calcium and magnesium salts. When sodium carbonate is added, the Ca^{2+} and Mg^{2+} ions precipitate out as insoluble calcium and magnesium carbonate: $Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$ The water becomes soft and will lather with soap.
Calcium carbonate (limestone) ($CaCO_3$)	Manufacture of cement for use in the construction industry	To make cement, calcium carbonate is heated with small quantities of other materials such as silicon dioxide (sand) in a kiln to about 1400 °C. The calcium carbonate decomposes forming calcium oxide (quicklime): $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$ The calcium oxide is then blended with the other materials inside the kiln to form clinker. The clinker is then ground with a small amount of calcium sulfate (gypsum) to make cement.

Did you know?

Epsom salt derives its name from the town of Epsom in Surrey, England where the salt was first discovered in a bitter saline spring. The healing power of the waters from the spring were realised in about 1618 and news of these powers spread rapidly. Epsom quickly developed into a spa town and people travelled from far and wide to visit Epsom and 'take the waters'.

Did you know?

The origins of **gypsum plaster** date back more than 9000 years and it was used through the centuries in many parts of the world. By the 1700s, Paris had become the centre for its use and it became known as plaster of Paris. At the time, the houses in Paris were made of wood and, by law, plaster of Paris was used to cover the walls to make them fireproof so that Paris did not suffer the same fate as London had in the Great Fire of 1666.

▼ **Table 8.4.2** (continued)

Salt	Use	Explanation
Magnesium sulfate (Epsom salt) ($MgSO_4.7H_2O$)	Various medicinal uses	Added to bath water, Epsom salt is used to ease stress and relax the body, soothe back pain and aching limbs, ease muscle strain, help reduce inflammation, help cure skin problems and help heal cuts. Taken orally, Epsom salt is used as a saline laxative and to help eliminate toxins from the body.
Sodium chloride ($NaCl$)	Food preservation	Sodium chloride is used to preserve food items, such as meat and fish, by withdrawing water from their cells by osmosis so that the water is unavailable for the reactions that cause decay. It also withdraws water from the microorganisms which cause decay, preventing them from growing.
Sodium nitrate ($NaNO_3$) and sodium nitrite ($NaNO_2$)	Food preservation	Sodium nitrate and sodium nitrite are often used with sodium chloride to preserve meat. They destroy certain bacteria that cause severe food poisoning and retard the development of rancidity. Sodium nitrite gives an attractive red colour to the meat, e.g. ham and bacon, and adds flavour.
Sodium benzoate (C_6H_5COONa)	Food preservation	Sodium benzoate is used to preserve foods with a low pH, e.g. fizzy drinks and fruit juices. At a pH of below 4.5, sodium benzoate is converted to benzoic acid (C_6H_5COOH) which inhibits the growth of microorganisms.
Calcium sulfate (gypsum) ($CaSO_4.2H_2O$)	Manufacture of plaster of Paris, which is used when setting broken bones and as a building material	Plaster of Paris (gypsum plaster) is made of calcium sulfate which has been heated to about 300 °C to remove the water of crystallisation. Bandages are then impregnated with the dry plaster of Paris powder. When water is added to the bandage, a paste forms and heat is given off. The bandage, with its paste, is then wrapped around the damaged limb, where it hardens forming an orthopaedic cast. Plaster of Paris is also used as a building material. Water is added to the dry plaster to make a paste which is used to coat walls and ceilings.

Dangers of salts

While salts play such an important role in our everyday lives, some have also been implicated in causing various health problems. Table 8.4.3 summarises some of these dangers.

▼ **Table 8.4.3** Dangers of salts

Salt	Dangers
Sodium chloride ($NaCl$)	Excessive consumption can lead to hypertension (high blood pressure).
Sodium nitrate ($NaNO_3$) and sodium nitrite ($NaNO_2$)	May increase a person's risk of developing cancer. Have been implicated in causing brain damage in children.
Sodium benzoate (C_6H_5COONa)	Has been implicated in increasing hyperactivity and asthma in children. May increase a person's risk of developing cancer.

Summary questions

1 Explain the difference between a normal salt and an acid salt.

2 What is water of crystallisation?

3 Name the most appropriate method for preparing each of the following salts. In each case name the two compounds you would choose to prepare the salt and write a balanced chemical equation for the reaction:
   - a magnesium chloride
   - b lead(II) sulfate
   - c anhydrous iron(III) chloride
   - d potassium nitrate.

4 Describe, giving full details, how you would prepare magnesium carbonate starting with magnesium nitrate.

5 Describe how we use the following salts in our everyday lives:
   - a calcium sulfate
   - b sodium hydrogencarbonate
   - c magnesium sulfate.

A8.5 Neutralisation reactions

In Unit 6.2 you learnt that a **neutralisation reaction** is any reaction between a base and an acid. In a neutralisation reaction a salt and water are always formed.

When a strong alkali reacts with a strong acid, the reaction is complete when neither alkali nor acid is present in excess. At this point the products have a pH of 7, i.e. they are neutral. This point is known as the **neutralisation point** or **end point**.

In a neutralisation reaction between an aqueous alkali and an acid, the actual reaction occurring is between the hydroxide ions of the alkali and the hydrogen ions of the acid. These ions react to form water and the other ions remain in solution as spectator ions. For example, the reaction between sodium hydroxide solution and hydrochloric acid can be summarised by the following chemical equation:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

The OH^- ions from the sodium hydroxide and the H^+ ions from the hydrochloric acid react to form water, as shown in the following ionic equation:

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$$

The Na^+ and Cl^- ions remain in solution as spectator ions.

The reaction between the hydroxide and hydrogen ions is described as being an **exothermic reaction** (see Unit 12.1) because it produces heat energy. We can make use of this fact when determining the neutralisation point of such a reaction.

Objectives

By the end of this topic you will be able to:

- explain a neutralisation reaction
- describe how to carry out neutralisation reactions using an indicator, pH meter and temperature change
- describe how we use neutralisation reactions in our daily lives.

! Key fact

A **neutralisation reaction** is a reaction between a base and an acid to form a salt and water.

! Key fact

The **neutralisation point** is the point in an acid–alkali reaction where the reaction is complete and neither acid nor alkali is present in excess.

Determining the neutralisation point in an acid–alkali reaction

The neutralisation point of a reaction between an aqueous alkali and an aqueous acid is determined by performing a titration. During the titration, the volume of one solution required to neutralise a fixed volume of another solution is determined, and the neutralisation point is identified in one of *three* ways: using an indicator, using a pH meter or using temperature change. We will now look at two of these ways, using an indicator and using temperature change.

Using an indicator

In Unit 8.3 you learnt that an **indicator** is a substance which has one colour in an acidic solution and another colour in an alkaline solution.

When performing a titration using an indicator, a fixed volume of one solution, usually the alkali, is placed in a conical flask and a few drops of the indicator are added. The other solution, usually the acid, is added from the burette and the indicator will change colour at the neutralisation point. This point is determined when the colour changes on the addition of a single drop of solution from the burette.

Phenolphthalein and methyl orange are the most commonly used indicators to determine the neutralisation point in a titration.

? Did you know?

The neutralisation point can be determined using a **pH meter** instead of an indicator. One solution is added in small quantities, e.g. 2 cm^3, to a fixed volume of the other solution and the pH of the solution changes sharply to 7 at the neutralisation point.

Using temperature change

When using a **temperature change** to determine the neutralisation point, a fixed volume of one solution is placed in an insulated container, e.g. a polystyrene cup, and its temperature is recorded. The other solution is added in small quantities, e.g. 2 cm^3, from the burette. Using the thermometer, the solution is quickly stirred after each addition and the temperature is recorded. The procedure is continued until several successive drops in temperature have been recorded. This is known as a **thermometric titration**.

A graph is then drawn showing temperature against the volume of solution added from the burette.

To draw the graph, the points are plotted and then two straight lines of best fit are drawn, one through the points where the temperature is increasing and one through the points where the temperature is decreasing. The point of *intersection* of the two lines is the neutralisation point.

An example of a graph which might be obtained when hydrochloric acid is added to 25 cm^3 of sodium hydroxide solution is shown in Figure 8.5.1. The graph shows that 11.0 cm^3 of hydrochloric acid is needed to neutralise 25 cm^3 of sodium hydroxide solution.

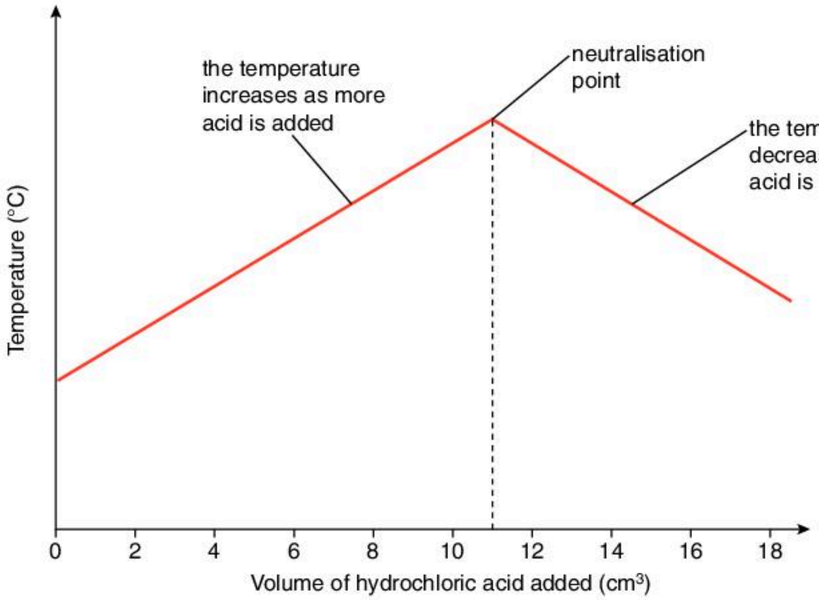


Figure 8.5.1 Graph showing temperature against volume of hydrochloric acid added to 25 cm^3 of sodium hydroxide solution

The reaction between the OH^- ions in the sodium hydroxide and the H^+ ions in the hydrochloric acid produces heat:

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$$

The temperature increases as more acid is added because more H^+ ions are being added to react with the OH^- ions. The temperature stops increasing when all the OH^- ions have reacted and there are none left in the solution. This is the neutralisation point. The temperature then decreases for three reasons; no more heat is being produced because the reaction has reached completion, acid at a lower temperature is being added which cools the solution and heat is being lost to the environment.

Practical activity

To determine the neutralisation point by using temperature change

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation.

You will be supplied with 1.0 mol dm^{-3} sodium hydroxide solution, 1.0 mol dm^{-3} sulfuric acid, a 25 cm^3 pipette, a polystyrene cup, a burette and a thermometer.

Method

1. Rinse the burette with some sulfuric acid and fill it with the acid.
2. Rinse the pipette with some sodium hydroxide solution. Measure 25 cm^3 of sodium hydroxide solution in the pipette and run it into the polystyrene cup.
3. Using the thermometer, measure and record the temperature of the sodium hydroxide solution.
4. Keeping the thermometer in the sodium hydroxide solution, add 2 cm^3 of acid from the burette. Quickly stir the solution with the thermometer and record its maximum temperature.
5. Immediately add another 2 cm^3 of hydrochloric acid, quickly stir the solution and record its maximum temperature.
6. Repeat step **5** until you have recorded four successive drops in temperature. It is essential that you do this as quickly as possible.
7. Record your results in a table.
8. Use your results to draw a graph showing the temperature of the solution against the volume of sulfuric acid added. Draw two straight lines of best fit.
9. Use your graph to determine the volume of sulfuric acid needed to neutralise 25 cm^3 of sodium hydroxide solution.
10. Explain:
    - a why the temperature of the solution increased (use an ionic equation to help you)
    - b why the temperature of the solution stopped increasing
    - c why the temperature of the solution decreased.
11. Suggest reasons why this method of determining the neutralisation point is less accurate than using an indicator.

Using neutralisation reactions

We have already studied how neutralisation reactions can be used to treat ant and bee stings. Table 8.5.1 summarises three other ways we use neutralisation reactions.

Table 8.5.1 Using neutralisation reactions

Where used	Action
Antacids	Antacids are taken to treat indigestion and acid reflux. They may contain sodium hydrogencarbonate ($NaHCO_3$), magnesium hydroxide [$Mg(OH)_2$], aluminium hydroxide [$Al(OH)_3$], magnesium carbonate ($MgCO_3$) or calcium carbonate ($CaCO_3$). They work by neutralising the excess hydrochloric acid in the stomach.
Toothpaste	Acid, produced by bacteria in the mouth, causes tooth decay by reacting with the calcium hydroxyapatite [$Ca_{10}(PO_4)_6(OH)_2$] in tooth enamel. Two active ingredients in toothpaste, sodium hydrogencarbonate and sodium monofluorophosphate (Na_2FPO_3), help reduce decay. The sodium hydrogencarbonate neutralises any acid in the mouth. The fluoride ions in the sodium monofluorophosphate displace the hydroxide ions in the calcium hydroxyapatite forming calcium fluoroapatite [$Ca_{10}(PO_4)_6F_2$]: $Ca_{10}(PO_4)_6(OH)_2(s) + 2F^-(aq) \longrightarrow Ca_{10}(PO_4)_6F_2(s) + 2OH^-(aq)$ Acids do not react with calcium fluoroapatite hence the tooth enamel does not decay.
Soil treatment	Most plants grow best if the soil is neutral. Finely ground calcium carbonate (limestone, $CaCO_3$) or lime in the form of calcium oxide (quicklime, CaO) or calcium hydroxide (slaked lime, $Ca(OH)_2$) can be added to soil to neutralise any acids present. However, lime cannot be added at the same time as an ammonium fertiliser because the two react to make a salt, ammonia gas and water which eliminates the beneficial effects of both: $CaO(s) + 2NH_4^+(aq) \longrightarrow Ca^{2+}(aq) + 2NH_3(g) + H_2O(l)$

Summary questions

1 What is a neutralisation reaction?

2 Outline how you would perform a neutralisation reaction using:
  a an indicator
  b temperature change.

3 Give three uses of neutralisation reactions in our daily lives.

4 Why should lime and an ammonium fertiliser not be added at the same time to soil?

Objectives

By the end of this topic you will be able to:

- explain volumetric analysis
- determine the mole ratio in which reactants combine
- determine the mass concentration and molar concentration of one reactant used in a titration.

A8.6 Volumetric analysis

Volumetric analysis involves performing a titration to determine the exact volume of one solution, usually an acid, required to neutralise a fixed volume of another solution, usually an aqueous alkali or carbonate. The results of the titration can then be used quantitatively in one of two ways:

- to determine the **mole ratio** in which the two reactants combine
- to determine the **molar concentration** and **mass concentration** of one of the reactants.

Determining mole ratios

The **mole ratio** in which reactants combine can be determined using a titration if the mass or molar concentration of both reactants is known, i.e. if they are both standard solutions.

Example

During a titration it was found that 25 cm^3 of potassium hydroxide of molar concentration 0.6 $mol\,dm^{-3}$ was neutralised by 10 cm^3 of sulfuric acid of molar concentration 0.75 $mol\,dm^{-3}$. Determine the mole ratio in which the reactants combine.

Calculate the number of **moles** of potassium hydroxide that reacted:

Molar concentration of KOH(aq) = 0.6 $mol\,dm^{-3}$

i.e. 1000 cm^3 (1 dm^3) KOH(aq) contains 0.6 mol KOH.

$\therefore$ 1 cm^3 KOH(aq) contains $\frac{0.6}{1000}$ mol KOH

and 25 cm^3 KOH(aq) contains $\frac{0.6}{1000} \times 25$ mol KOH

= **0.015 mol KOH**

Calculate the number of **moles** of sulfuric acid that reacted:

Molar concentration of H_2SO_4(aq) = 0.75 $mol\,dm^{-3}$

i.e. 1000 cm^3 (1 dm^3) H_2SO_4(aq) contains 0.75 mol H_2SO_4.

$\therefore$ 1 cm^3 H_2SO_4(aq) contains $\frac{0.75}{1000}$ mol H_2SO_4

and 10 cm^3 H_2SO_4(aq) contains $\frac{0.75}{1000} \times 10$ mol H_2SO_4

= **0.0075 mol $\mathbf{H_2SO_4}$**

Use the number of moles of each reactant to determine the mole ratio:

0.015 mol KOH reacts with 0.0075 mol H_2SO_4

∴ **2 mol KOH** react with **1 mol H_2SO_4**.

The mole ratio in which the reactants combine is 2 mol potassium hydroxide to 1 mol sulfuric acid.

Determining concentration

The concentration of one reactant used in a titration can be determined if the mass or molar concentration of the other reactant is known, i.e. if it is a standard solution. This is done by following the same four steps, i.e. 1) to 4), that you followed in your calculations in Unit 7.5.

Examples

1 A student performed a titration to determine the concentration of a solution of hydrochloric acid. The student used a standard solution of potassium hydroxide with a concentration of 0.1 mol dm^{-3} and added the hydrochloric acid to 25 cm^3 of the potassium hydroxide solution. The results of the experiment are shown in the table.

Burette reading	**Titration number**			
	Rough	**1**	**2**	**3**
Final reading/cm^3	21.5	20.5	41.0	20.4
Initial reading/cm^3	0.0	0.0	20.5	0.0
Volume of acid added/cm^3	21.5	20.5	20.5	20.4

Determine the molar concentration of the hydrochloric acid.

Calculate the average volume of hydrochloric acid used:

$$\text{average volume of HCl(aq)} = \left(\frac{20.5 + 20.5 + 20.4}{3}\right)\text{cm}^3$$
$$= \mathbf{20.5\ cm^3}$$

i.e. it takes 20.5 cm^3 of hydrochloric acid to neutralise 25 cm^3 of potassium hydroxide solution, which has a concentration of 0.1 mol dm^{-3}.

The balanced chemical equation for the reaction is:

$$KOH(aq) + HCl(aq) \longrightarrow KCl(aq) + H_2O(l)$$

1) Calculate the number of **moles** of potassium hydroxide that reacted:

   Molar concentration of KOH(aq) = 0.1 mol dm^{-3}

   i.e. 1000 cm^3 (1 dm^3) KOH(aq) contains 0.1 mol KOH.

   $$\therefore\ 1\ \text{cm}^3\ \text{KOH(aq) contains } \frac{0.1}{1000}\ \text{mol KOH}$$

   $$\text{and } 25\ \text{cm}^3\ \text{KOH(aq) contains } \frac{0.1}{1000} \times 25\ \text{mol KOH}$$
   $$= \mathbf{0.0025\ mol\ KOH}.$$

2) Use the balanced equation to determine the **mole ratio** between KOH and HCl:

   1 mol KOH reacts with 1 mol HCl.

3) Use the number of moles of KOH from 1) and the mole ratio from 2) to calculate the number of **moles** of HCl reacting:

   0.0025 mol KOH reacts with **0.0025 mol HCl**.

4) Use the number of moles of HCl from 3) and the volume used to calculate the number of **moles** of HCl in 1 dm^3:

   Volume of HCl(aq) used = 20.5 cm^3

   i.e. 20.5 cm^3 HCl(aq) contains 0.0025 mol HCl.

   $\therefore$ 1 cm^3 HCl(aq) contains $\frac{0.0025}{20.5}$ mol HCl

   and 1000 cm^3 (1 dm^3) HCl(aq) contains $\frac{0.025}{20.5} \times 1000$ mol HCl

   $= \mathbf{0.122\ mol\ HCl}$

   Therefore, molar concentration of HCl(aq) = **0.122 mol dm^{-3}**.

2 A student found that 20.0 cm^3 of hydrochloric acid of unknown concentration were required to neutralise 25 cm^3 of sodium carbonate solution with a concentration of 0.12 mol dm^{-3}. Calculate the mass concentration of the hydrochloric acid.

The balanced chemical equation for the reaction is:

$$Na_2CO_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

1) Molar concentration of Na_2CO_3(aq) = 0.12 mol dm^{-3}

   i.e. 1000 cm^3 (1 dm^3) Na_2CO_3(aq) contains 0.12 mol Na_2CO_3.

   $\therefore$ 1 cm^3 Na_2CO_3(aq) contains $\frac{0.12}{1000}$ mol Na_2CO_3

   and 25 cm^3 Na_2CO_3(aq) contains $\frac{0.12}{1000} \times 25$ mol Na_2CO_3

   $= \mathbf{0.003\ mol\ Na_2CO_3}$.

2) 1 mol Na_2CO_3 reacts with 2 mol HCl.

3) 0.003 mol Na_2CO_3 reacts with 2 × 0.003 mol HCl

   $= \mathbf{0.006\ mol\ HCl}$.

4) Volume of HCl(aq) used = 20.0 cm^3

   i.e. 20.0 cm^3 HCl(aq) contains 0.006 mol HCl.

   $\therefore$ 1 cm^3 HCl(aq) contains $\frac{0.006}{20.0}$ mol HCl

   and 1000 cm^3 (1 dm^3) HCl(aq) contains $\frac{0.006}{20.0} \times 1000$ mol HCl

   $= \mathbf{0.3\ mol\ HCl}$.

   Therefore, molar concentration of HCl(aq) = 0.3 mol dm^{-3}

   $M(HCl) = 1 + 35.5 = 36.5\ g\ mol^{-1}$

   i.e. mass of 1 mol HCl = 36.5 g

   $\therefore$ mass of 0.3 mol HCl = 0.3 × 36.5 g

   $= \mathbf{10.95\ g}$.

Therefore, mass concentration of HCl(aq) = **10.95 g dm^{-3}**.

Practical activity

To determine the concentration of a solution of hydrochloric acid

Your teacher may use this activity to assess:

- manipulation and measurement
- analysis and interpretation

You will be supplied with 0.2 mol dm^{-3} sodium hydroxide solution, hydrochloric acid of unknown concentration, methyl orange indicator, a 25 cm^3 pipette, a conical flask and a burette.

Method

1. Rinse the burette with some hydrochloric acid and fill it with the acid.
2. Rinse the pipette with some sodium hydroxide solution. Measure 25 cm^3 of sodium hydroxide solution in the pipette and run it into the conical flask.
3. Add a few drops of methyl orange to the conical flask and swirl to mix the contents.
4. Carry out a titration as described in the practical activity on page 141. This time add the acid until the solution just turns from yellow to orange on adding one drop of acid. If it turns red you have added too much acid.
5. Record your results in a table as outlined in the practical activity on page 142 and average the three volumes of acid which were within 0.1 cm^3.
6. Use this volume to calculate the molar concentration and mass concentration of the hydrochloric acid. Use the examples above to help you.

Determining vitamin C content

The vitamin C (ascorbic acid) content of vitamin C tablets, fresh fruit juices or packaged fruit juices can be determined by performing a titration using iodine solution. Iodine acts as an **oxidising agent** (see Unit 9.3) during the titration and oxidises the ascorbic acid to dehydroascorbic acid. The ascorbic acid acts as a **reducing agent** and reduces the iodine to I$^-$ ions. The reaction is summarised in the following equation:

$$\underset{\text{ascorbic acid}}{C_6H_8O_6(aq)} + I_2(aq) \longrightarrow \underset{\text{dehydroascorbic acid}}{C_6H_6O_6(aq)} + 2HI(aq)$$

During the titration, *brown* iodine solution is added from the burette to a fixed volume of fruit juice or a solution of a vitamin C tablet. As the iodine is added it is immediately reduced by the ascorbic acid to *colourless* I$^-$ ions. At the point where all of the ascorbic acid has just reacted, iodine will no longer be reduced. The iodine remains in the solution and reacts with the starch to produce a blue-black colour. The end point of the titration is when a permanent *blue-black* colour just forms.

If the concentration of the iodine solution is known, the concentration of ascorbic acid in the fruit juice or solution of a vitamin C tablet can be calculated since 1 mol of ascorbic acid reacts with 1 mol of iodine.

Practical activity

To determine the vitamin C content of a fruit juice

Your teacher may use this activity to assess:

- manipulation and measurement
- analysis and interpretation.

You will be supplied with a sample of a fruit juice without pulp, e.g. grapefruit juice, iodine solution of concentration 0.005 mol dm^{-3}, 1% starch indicator solution, a 25 cm^3 pipette, a conical flask and a burette.

Method

1 Rinse the burette with some iodine solution and fill it with the iodine solution.

2 Rinse the pipette with some of the fruit juice. Measure 25 cm^3 of the fruit juice in the pipette and run it into the conical flask.

3 Add ten drops of starch indicator to the conical flask and swirl to mix the contents.

4 Carry out a titration as described in the practical activity on page 141. This time add the iodine solution until the first permanent trace of a blue-black colour is obtained.

5 Record your results in a table as outlined in the same practical activity on page 142 and average the three volumes of iodine solution which were within 0.1 cm^3.

6 Use this volume to calculate the molar concentration and mass concentration of ascorbic acid in the grapefruit juice. Use the examples on pages 149–50 to help you.

You could adapt the experiment which you have just performed to determine the effect of heat on the vitamin C content of a fruit juice. **Planning and designing** experiments is discussed in more detail in the School-Based Assessment section on the CD.

Summary questions

1 A volume of 40 cm^3 of sodium carbonate solution of concentration 0.2 mol dm^{-3} reacts with 20 cm^3 of nitric acid of concentration 0.8 mol dm^{-3}. Determine the mole ratio in which the reactants combine.

2 A volume of 30 cm^3 of sodium hydroxide solution of concentration 0.4 mol dm^{-3} were neutralised by 20 cm^3 of sulfuric acid. Determine the concentration of the sulfuric acid in:
  a mol dm^{-3}
  b g dm^{-3}.

Key concepts

- An acid is a substance which forms hydrogen ions (H^+) when dissolved in water.
- An **acid** is defined as a proton donor.
- Aqueous solutions of acids have a sour taste, change the colour of litmus to red, have a pH value of less than 7 and conduct an electric current.
- Aqueous acids react with reactive metals to form a salt and hydrogen; react with carbonates and hydrogencarbonates to form a salt, carbon dioxide and water; and react with bases to form a salt and water.
- Aqueous acids may be classified as organic or inorganic; monobasic, dibasic or tribasic; dilute or concentrated; and strong or weak.
- An **acid anhydride** is a compound that reacts with water to form an acid.
- We come across a variety of organic acids in living organisms and we make use of some of these organic acids in our everyday activities.

- A base is defined as a proton donor.
- An alkali is a base which dissolves in water to form a solution containing hydroxide ions (OH^-).
- Aqueous solutions of alkalis have a bitter taste, change the colour of litmus to blue, have a pH value greater than 7, conduct an electric current and feel soapy.
- Bases react with acids to form a salt and water and react with ammonium salts to form a salt, ammonia and water.
- Amphoteric substances can react with acids and strong alkalis.
- Oxides can be classified as acidic, basic, amphoteric and neutral.
- Strong acids and alkalis are fully ionised when dissolved in water.
- Weak acids and alkalis are partially ionised when dissolved in water.
- The strength of an acid or alkali is measured using the pH scale which ranges from 0 to 14.
- The stronger the acid, the lower the pH. The stronger the alkali, the higher the pH.
- Neutral substances have a pH of 7.
- Universal indicator or a pH meter is used to measure pH.
- Other acid–alkali indicators have one colour in an acidic solution and another colour in an alkaline solution.
- A salt is a compound formed when some or all of the replaceable hydrogen ions in an acid are replaced by metal or ammonium ions.
- A normal salt is formed when all the hydrogen ions in an acid are replaced. An acid salt is formed when only some of the hydrogen ions are replaced.
- Some salts contain water molecules trapped between their ions in the crystal lattice. This is known as water of crystallisation.
- Insoluble salts can be prepared by ionic precipitation.
- Anhydrous binary salts can be prepared by direct combination.
- Potassium, sodium and ammonium salts can be prepared using a titration.
- Other soluble salts can be prepared by reacting a reactive metal, an insoluble carbonate or an insoluble base with an acid.
- Salts play an extremely important part in our everyday lives, but can also be dangerous.
- A neutralisation reaction is a reaction between a base and an acid to form a salt and water.
- The neutralisation point is the point in an acid–alkali reaction where the reaction is complete and neither acid nor alkali is present in excess.
- The neutralisation point of an acid–alkali reaction can be determined during a titration by using an indicator, a pH meter or temperature change.
- Neutralisation reactions are used in various aspects of our daily lives.
- Volumetric analysis involves performing a titration to determine the exact volume of one solution required to neutralise a fixed volume of another solution.
- Results of a titration can be used to determine the mole ratio in which the two reactants combine or to determine the concentration of one reactant.
- The vitamin C content of a vitamin C tablet or fruit juice can be determined using a titration.

Practice exam-style questions

Multiple-choice questions

1 Which of the following is not a property of hydrochloric acid?

A It turns blue litmus red.
B It is a strong electrolyte.
C It reacts with magnesium to form magnesium chloride and water.
D It reacts with sodium carbonate to form sodium chloride, carbon dioxide and water.

2 Which of the following is/are properties of aqueous alkalis?

I They react with acids to form a salt and hydrogen.
II They have a pH less than 7.
III They react with ammonium salts to form a salt, ammonia and water.

A I and III only
B I and II only
C I, II and III
D III only

3 A weak acid is one which:

A is dilute
B ionises fully in aqueous solution
C has a low concentration of H^+ ions
D does not ionise in aqueous solution.

4 A solution which has a pH of 11:

A is more alkaline than a solution with a pH of 12
B is less alkaline than a solution with a pH of 9
C is more acidic than water
D is less alkaline than a sodium hydroxide solution of the same concentration.

5 The actual change which occurs when sodium hydroxide solution reacts with sulfuric acid is:

A $2Na^+(aq) + SO_4^{2-}(aq) \longrightarrow Na_2SO_4(aq)$
B $OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$
C $2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(s) + 2H_2O(l)$
D $NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(l)$

6 Which of the following acids is found in an ant sting?

A Methanoic acid
B Ethanoic acid
C Lactic acid
D Hydrochloric acid

7 Which of the following is not an amphoteric substance?

A Lead(II) oxide
B Aluminium hydroxide
C Iron(III) oxide
D Zinc hydroxide

8 A pure sample of lead(II) chloride can be prepared by reacting:

A lead with dilute hydrochloric acid
B lead(II) hydroxide with dilute hydrochloric acid
C lead(II) nitrate solution with sodium chloride solution
D lead(II) carbonate with hydrochloric acid.

9 Magnesium sulfate can be prepared by:

I reacting magnesium carbonate with sulfuric acid
II reacting magnesium nitrate with sulfuric acid
III reacting magnesium with sulfuric acid.

A I and II only
B I and III only
C I, II and III
D III only

10 Which of the following cannot be used to determine the neutralisation point in a reaction between an acid and an alkali?

A An anemometer
B A thermometer
C A pH meter
D An indicator

11 A volume of 25 cm^3 of a carbonate solution of concentration 0.2 $mol\,dm^{-3}$ was neutralised by 20 cm^3 of an acid of concentration 0.5 $mol\,dm^{-3}$. The mole ratio in which the reactants combine is:

A 1 mol of carbonate to 1 mol of acid
B 1 mol of carbonate to 2 mol of acid
C 2 mol of carbonate to 1 mol of acid
D 3 mol of carbonate to 1 mol of acid.

12 A volume of 10 cm^3 of sulfuric acid of unknown concentration neutralised 25 cm^3 of sodium hydroxide solution of concentration 0.4 $mol\,dm^{-3}$. The concentration of the acid is:

A 0.1 $mol\,dm^{-3}$
B 0.25 $mol\,dm^{-3}$
C 0.4 $mol\,dm^{-3}$
D 0.5 $mol\,dm^{-3}$.

Structured question

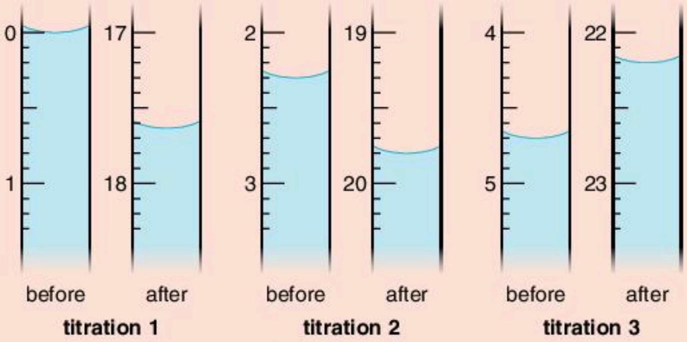
13 A titration is used to determine the volume of one solution needed to neutralise a fixed volume of another solution. Titrations can be used as a means of preparing certain soluble salts and in volumetric analysis to determine the concentration of one of the solutions used. During a titration, the neutralisation point is usually determined by using an indicator.

a What is meant by the neutralisation point of a titration? *(2 marks)*

b In order to determine the concentration of a sodium hydroxide solution, a student is provided with two solutions labelled X and Y.

- X is aqueous sulfuric acid with a mass concentration of 19.6 $g\,dm^{-3}$
- Y is a sodium hydroxide solution of unknown concentration.

To determine the concentration of Y, the student titrates X against 25.0 cm^3 of Y. The figure shows the readings from the burette before and after each titration using 25.0 cm^3 of solution Y.



Burette readings

i) Using the information from the figure, complete the table below.

	Titration number		
	1	2	3
Final burette reading/cm^3			
Initial burette reading/cm^3			
Volume of solution X used/cm^3			

(3 marks)

ii) Calculate the volume of X needed to neutralise 25 cm^3 of Y. *(1 mark)*

iii) Calculate the concentration of sulfuric acid in solution X in mol dm^{-3}. *(1 mark)*

iv) Calculate the number of moles of sulfuric acid used in the titration. *(1 mark)*

v) Write a balanced equation for the reaction. *(2 marks)*

vi) Determine the number of moles of sodium hydroxide in the 25.0 cm^3 of solution Y used. *(1 mark)*

vii) Calculate the concentration of sodium hydroxide in solution Y in mol dm^{-3}. *(1 mark)*

c How would the student use the titration carried out in **b** above:

i) to prepare a pure dry sample of sodium sulfate *(2 marks)*

ii) to prepare a pure sodium hydrogensulfate solution? *(1 mark)*

Total 15 marks

Extended response question

14 a An acid can be defined as a proton donor and a base as a proton acceptor. Explain the reason for EACH definition. *(4 marks)*

b Carbonic acid and sulfuric acid are both acidic. However, they have different pH values. You are supplied with a 1.0 mol dm^{-3} solution of each acid.

i) Suggest a pH value for EACH solution. *(2 marks)*

ii) Explain the reason for their different pH values. *(2 marks)*

iii) Give one way you could use to find out if your suggested pH values are correct. *(1 mark)*

c Salts can be prepared in the laboratory by a variety of different methods.

i) Describe how you would prepare a pure sample of anhydrous copper(II) sulfate in the laboratory starting with copper(II) carbonate. Include in your answer a balanced chemical equation for the reaction. *(5 marks)*

ii) How would your method be different if you wanted to prepare hydrated copper(II) sulfate crystals? *(1 mark)*

Total 15 marks

A9 Oxidation–reduction reactions

Objectives

By the end of this topic you will be able to:

- describe the action of common oxidising and reducing substances in everyday activities
- define oxidation and reduction in terms of electron transfer
- identify the reactant being oxidised and the reactant being reduced using electron transfer.

Oxidation and reduction reactions form a very important part of our everyday lives, for example, in batteries, photography, photosynthesis and the rusting of cars. In all these reactions electrons are being transferred. It was only after the discovery of oxygen that the chemistry behind oxidation–reduction reactions was understood.

A9.1 Oxidation and reduction – an introduction

The discovery and laboratory preparation of oxygen was a very important stepping stone in chemistry. Antoine Lavoisier (1743–94) discovered that when metals are burned, they combine with oxygen from the air to form oxides. The term **oxidation** was used later to define any reaction in which a reactant gains oxygen. An example, according to this definition of oxidation, is the reaction between magnesium and oxygen where the magnesium has gained oxygen, forming magnesium oxide:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

The removal of oxygen from a metal oxide was described as a reduction reaction because the metal oxide was reduced to the pure metal. The term **reduction** was used to define any reaction in which a reactant loses oxygen. An example, according to this definition of reduction, is the reaction between iron(III) oxide and carbon where the iron(III) oxide has lost oxygen to form iron:

$$2Fe_2O_3(s) + 3C(s) \longrightarrow 4Fe(s) + 3CO_2(g)$$

These definitions were later extended to include reactions involving hydrogen. Hydrogen is chemically opposite to oxygen, therefore, oxidation was defined as a loss of hydrogen and reduction was defined as a gain of hydrogen.

With the discovery of protons, neutrons and, most importantly, **electrons**, scientists refined their definitions of oxidation and reduction, and the definitions were refined again using the concept of **oxidation number** or **oxidation state**. You will be studying these definitions in detail later in this unit.

▲ **Figure 9.1.1** The copper from which the statue of Marcus Garvey is made has started to turn green because of oxidation

Oxidation and reduction in our everyday activities

We encounter, and make use of, many oxidation and reduction reactions in our everyday lives. We will look at some of these below.

Action of bleaches

Bleach is added to clothes to remove coloured stains. The most commonly used bleaches are chlorine bleaches containing sodium chlorate(I) ($NaClO$) and oxygen bleaches containing hydrogen peroxide (H_2O_2). Both types of bleach remove stains by *oxidising* the coloured chemicals, or dyes, in the stain to their colourless form, hence the stain disappears. The equation below summarises how the chlorate(I) ion (ClO^-) in a chlorine bleach works:

$$ClO^-(aq) + \text{coloured dye} \longrightarrow Cl^-(aq) + \text{colourless dye}$$

Rusting

When iron and its alloy, steel, come into contact with oxygen and moisture the iron is *oxidised* by the oxygen to form hydrated iron(III) oxide ($Fe_2O_3.xH_2O$), otherwise known as **rust**. Given sufficient time and exposure to oxygen and moisture, any iron object will eventually oxidise entirely to rust and disintegrate.

▲ **Figure 9.1.2** A cut apple turns brown on exposure to oxygen in the air

Browning of cut fruits and vegetables

As soon as certain fruits such as apples and bananas, or certain vegetables such as potatoes, are peeled or cut, the cut surface is exposed to oxygen in the air. Enzymes in the plant cells on the surface of the fruit or vegetable begin to *oxidise* certain chemicals in the cells to *brown* compounds known as **melanins**. These melanins cause the cut or damaged surface to turn brown.

In many cases, browning is undesirable because it causes changes in appearance, flavour, quality and taste. The melanins formed are not toxic and in some cases their production is useful since it contributes to the colour and flavour of items such as raisins, prunes, coffee, tea and cocoa.

? Did you know?

Melanins are dark brown and black pigments responsible for the colour of skin, hair, scales, feathers and eyes of animals. Exposure to ultraviolet light stimulates the production of melanin pigments in human skin, which is why skin darkens in sunlight.

Food preservation

Sodium sulfite and sulfur dioxide are used as **food preservatives** to enhance flavour, preserve freshness and reduce or prevent spoilage and discolouration of foods such as wine, fruit juices, dried fruits, dried potatoes and shrimp. They do this by preventing oxidation. They prevent bacteria oxidising wine to vinegar and prevent the vitamin C in fruit juice being destroyed by oxidation. They also prevent browning of dried fruits, dried potato products and shrimp by *reducing* chemicals produced during the browning process back to their colourless form.

The breathalyser test

The **breathalyser test** is used to test the alcohol level in the breath of drivers. In the test, the driver blows into a sample of orange potassium dichromate(VI) crystals ($K_2Cr_2O_7$) which have been acidified with sulfuric acid. Any ethanol vapours in the driver's breath will *reduce* the *orange* dichromate(VI) ion ($Cr_2O_7^{2-}$) to the *green* chromium(III) ion (Cr^{3+}), turning the crystals green (see Unit 9.3).

Oxidation–reduction and electrons

By the turn of the 20th century, chemists realised that oxidation reactions always seemed to involve a loss of electrons. They developed a model for oxidation reactions that involved a transfer of electrons. A substance that *lost* electrons was said to have been **oxidised** and a substance that *gained* electrons was said to have been **reduced**.

Oxidation and reduction can be defined in terms of electron transfer, oxidation being the *loss* of electrons and reduction being the *gain* of electrons.

! Key fact

Oxidation is the loss of electrons by an element in its free state or an element in a compound.

! Key fact

Reduction is the gain of electrons by an element in its free state or an element in a compound.

Exam tip

You must be able to define oxidation and reduction in terms of electron transfer. An easy way to remember these definitions is by remembering two words: **OIL RIG**.

OIL **O**xidation **I**s **L**oss
RIG **R**eduction **I**s **G**ain

Key fact

A **redox reaction** is a chemical reaction in which one reactant is reduced and the other is oxidised.

Exam tip

When writing balanced oxidation and reduction half equations you must always include **state symbols** and it is important to ensure that the number of **atom**s and **charges balance** at each side of the equation.

A substance will not lose electrons unless there is another substance available to gain the electrons, i.e. for every oxidation reaction there will always be a reduction reaction. These oxidation–reduction reactions are known as **redox reactions**. The equation that only shows the loss of electrons is known as the **oxidation half equation** and the equation that only shows the gain of electrons is known as the **reduction half equation**.

Examples of redox reactions

1 Burning magnesium in oxygen:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

The magnesium oxide (MgO) produced in the reaction is an ionic compound consisting of Mg^{2+} ions and O^{2-} ions. During the reaction:

- Each magnesium atom loses two electrons to form a magnesium ion. Magnesium (Mg) is **oxidised**:

  $$Mg - 2e^- \longrightarrow Mg^{2+}$$

  This is more correctly written as:

  $$Mg \longrightarrow Mg^{2+} + 2e^-$$

  The overall oxidation half equation is:

  $$2Mg(s) \longrightarrow 2Mg^{2+}(s) + 4e^-$$

- Each oxygen atom in the oxygen molecule gains two electrons to form an oxide ion. Oxygen (O_2) is **reduced**:

  $$O + 2e^- \longrightarrow O^{2-}$$

  The overall reduction half equation is:

  $$O_2(g) + 4e^- \longrightarrow 2O^{2-}(s)$$

2 Adding zinc to copper(II) sulfate solution:

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$

The zinc displaces the copper in the copper(II) sulfate to form zinc sulfate and copper. Writing the ionic equation for the reaction:

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

During the reaction:

- Each zinc atom loses two electrons to form a zinc ion. Zinc (Zn) is **oxidised**:

  $$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

- Each copper(II) ion gains two electrons to form a copper atom. The copper(II) ions (Cu^{2+}) are **reduced**:

  $$Cu^{2+}(aq) + 2e^- \longrightarrow 2Cu(s)$$

3 Bubbling chlorine gas through potassium bromide solution:

$$Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$$

The chlorine gas displaces the bromine in the potassium bromide to form potassium chloride and bromine. Writing the ionic equation for the reaction:

$$Cl_2(g) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$$

During the reaction:

- Each bromide ion loses one electron to form a bromine atom. The bromide ions (Br^-) are **oxidised**:

  $$2Br^-(aq) \longrightarrow Br_2(aq) + 2e^-$$

- Each chlorine atom in the chlorine molecule gains one electron to form a chloride ion. Chlorine (Cl_2) is **reduced**:

  $$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$$

Ionic bonding always involves a transfer of electrons from the metal atom to the non-metal atom. Any reaction in which ionic bonds are formed is a redox reaction. Using the reaction between sodium and chlorine as an example:

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

- Each sodium atom loses one electron to form a sodium ion. Sodium (Na) is **oxidised**:

  $$2Na(s) \longrightarrow 2Na^+(s) + 2e^-$$

- Each chlorine atom in the chlorine molecule gains one electron to form a chloride ion. Chlorine (Cl_2) is **reduced**:

  $$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(s)$$

Summary questions

1. Define oxidation and reduction in terms of a transfer of electrons.
2. State whether each of the following half equations shows oxidation or reduction:

   a $Fe^{3+}(aq) + 3e^- \longrightarrow Fe(s)$ b $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$

   c $Br_2(aq) \longrightarrow 2Br^-(aq) + 2e^-$ d $Cu^{2+} + 2e^- \longrightarrow Cu(s)$.
3. In the following redox reaction, state what has been oxidised and what has been reduced:

   $$Br_2(aq) + 2KI(aq) \longrightarrow 2KBr(aq) + I_2(aq)$$
4. Why does an apple turn brown when it is cut in half?

A9.2 Oxidation numbers

Redox chemistry developed further when chemists recognised that oxidation–reduction reactions did not always involve the transfer of electrons. For example, in the reaction between hydrogen and oxygen to form water, there is no change in the number of valence electrons on any of the atoms, i.e. there is no transfer of valence electrons. However, the reaction is definitely a redox reaction because the reaction involves both hydrogen and oxygen:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

The model of oxidation–reduction reactions was expanded by the development of the concept of oxidation number or oxidation state. Oxidation number indicates the number of electrons lost, gained or shared as a result of chemical bonding.

Objectives

By the end of this topic you will be able to:

- define oxidation number
- determine oxidation numbers from formulae
- define oxidation and reduction in terms of oxidation number
- identify redox reactions using a change in oxidation number
- identify the reactant being oxidised and the reactant being reduced using a change in oxidation number.

! Key fact

The **oxidation number** of an element is the *theoretical* charge that an atom of the element would have *if* all the bonds between the atoms in the compound containing it were ionic and the compound was composed entirely of ions.

Oxidation numbers can be positive, negative or zero. When writing the oxidation number of an element, unless the number is zero, a plus or minus sign is placed before the number to indicate if it is positive or negative, e.g. $+1$, -2, and there is always a number after the sign, i.e. the number 1 is always written.

Oxidation numbers play an important role in the naming of certain compounds as you will learn in sections that follow.

Rules for determining oxidation numbers

Before you can learn how to recognise redox reactions using the concept of oxidation number, you must learn how to determine the oxidation numbers of elements. To do this, a set of rules need to be followed.

- The oxidation number of each *atom* of an element in its free state is **zero**. For example:

  oxidation number of an Al atom in Al $= 0$
  oxidation number of each N atom in $N_2 = 0$.

- The oxidation number of each *simple ion* in an ionic compound is equal to the **charge** on the ion. For example:

  in Al_2O_3: oxidation number of each Al^{3+} ion $= +3$
  oxidation number of each O^{2-} ion $= -2$.

- The oxidation number of *hydrogen* when present in a compound is always **+1** except when bonded to a metal in a metal hydride where its oxidation number is -1, e.g. in sodium hydride (NaH) and calcium hydride (CaH_2).
- The oxidation number of *oxygen* when present in a compound is always **−2** except in peroxides where it is -1, e.g. in hydrogen peroxide (H_2O_2).
- The sum of the oxidation numbers of all elements present in a compound is **zero**. For example, in H_2O, the oxidation number of each atom must add up to zero:

$$\begin{array}{cccc} 2(\text{oxidation number of H}) & + & (\text{oxidation number of O}) & = 0 \\ 2(+1) & + & (-2) & = 0 \\ (+2) & + & (-2) & = 0 \end{array}$$

- The sum of the oxidation numbers of all elements present in a polyatomic ion is equal to the **charge** on the ion. For example, in the OH^- ion, the oxidation number of each atom must add up to -1:

$$\begin{array}{cccc} (\text{oxidation number of O}) & + & (\text{oxidation number of H}) & = -1 \\ (-2) & + & (+1) & = -1 \end{array}$$

Following these rules, the oxidation number of any element can be determined from the formula of the compound or polyatomic ion in which the element occurs.

Examples

1 Determine the oxidation number of sulfur in sulfur dioxide (SO_2).

$$\begin{array}{lccc} (\text{oxidation number of S}) & + & 2(\text{oxidation number of O}) & = 0 \\ (\text{oxidation number of S}) & + & 2(-2) & = 0 \\ (\text{oxidation number of S}) & + & (-4) & = 0 \\ & & \therefore\ \text{oxidation number of S} & = \mathbf{+4} \end{array}$$

The compound sulfur dioxide (SO_2) can be called **sulfur(IV) oxide** because the oxidation number of sulfur in the compound is $+4$.

2 Determine the oxidation number of nitrogen in the nitrate ion (NO_3^-).

$$\begin{aligned} (\text{oxidation number of N}) + 3(\text{oxidation number of O}) &= -1 \\ (\text{oxidation number of N}) + 3(-2) &= -1 \\ (\text{oxidation number of N}) + (-6) &= -1 \\ \therefore \text{oxidation number of N} &= \mathbf{+5} \end{aligned}$$

The nitrate ion (NO_3^-) is more correctly called the **nitrate(V) ion** because the oxidation number of nitrogen in the ion is +5.

3 Determine the oxidation number of carbon in methane (CH_4).

$$\begin{aligned} C + 4(+1) &= 0 \\ C + (+4) &= 0 \\ \therefore C &= -4 \end{aligned}$$

The oxidation number of carbon is **−4**.

4 Determine the oxidation number of manganese in the MnO_4^- ion and name the ion.

$$\begin{aligned} Mn + 4(-2) &= -1 \\ Mn + (-8) &= -1 \\ \therefore Mn &= +7 \end{aligned}$$

The oxidation number of manganese is **+7**. The ion is called the **manganate(VII) ion** because of the oxidation number of manganese in the ion.

5 Determine the oxidation number of chromium in $K_2Cr_2O_7$ and name the compound.

$$\begin{aligned} 2(+1) + 2Cr + 7(-2) &= 0 \\ (+2) + 2Cr + (-14) &= 0 \\ 2Cr &= +12 \\ \therefore Cr &= +6 \end{aligned}$$

The oxidation number of chromium is **+6**.
The compound is called **potassium dichromate(VI)** because of the oxidation number of chromium in the compound.

6 Determine the oxidation number of sulfur in the sulfite ion (SO_3^{2-}) and suggest an alternative name for the ion.

$$\begin{aligned} S + 3(-2) &= -2 \\ S + (-6) &= -2 \\ \therefore S &= +4 \end{aligned}$$

The oxidation number of sulfur is **+4**. The ion can also be called the **sulfate(IV) ion**.

Oxidation number and naming compounds

Certain elements have more than one possible oxidation number. For this reason, when naming a compound or ion that contains any of these elements, it is important to include the oxidation number as part of the name. This is done by placing the oxidation number of the element, without its charge, in Roman numerals in brackets after its name. When naming ions, the name always ends in '**ate**'. Tables 9.2.1, 9.2.2 and 9.2.3 give examples.

▼ **Table 9.2.1** Names of compounds containing manganese

Compound	Oxidation number of Mn	Name
MnO_2	+4	manganese(IV) oxide
$KMnO_4$	+7	potassium manganate(VII)

▼ **Table 9.2.2** Names of compounds and ions containing nitrogen

Compound	Oxidation number of N	Name
NO	+2	nitrogen(II) oxide
NO_2	+4	nitrogen(IV) oxide
NO_2^-	+3	nitrate(III) ion
NO_3^-	+5	nitrate(V) ion

▼ **Table 9.2.3** Names of compounds and ions containing sulfur

Compound	Oxidation number of S	Name
SO_2	+4	sulfur(IV) oxide
SO_3	+6	sulfur(VI) oxide
SO_3^{2-}	+4	sulfate(IV) ion
SO_4^{2-}	+6	sulfate(VI) ion

Redox reactions and oxidation number

If we consider the loss or gain of electrons and a change in oxidation number, we can see a clear link between the two.

- Whenever an atom or ion *loses* electrons it has been *oxidised*. At the same time, its oxidation number *increases*. For example, when a sodium atom forms a sodium ion it loses one electron and its oxidation number increases from 0 to +1. The sodium atom has been **oxidised**:

$$\begin{array}{lll} & Na - e^- & \longrightarrow Na^+ \\ \text{oxidation numbers:} & (0) & \qquad (+1) \end{array}$$

- Whenever an atom or ion *gains* electrons it has been *reduced*. At the same time, its oxidation number *decreases*. For example, when a chlorine atom forms a chloride ion it gains one electron and its oxidation number decreases from 0 to −1. The chlorine atom has been **reduced**:

$$\begin{array}{lll} & Cl + e^- & \longrightarrow Cl^- \\ \text{oxidation numbers:} & (0) & \qquad (-1) \end{array}$$

By considering the *change* in oxidation number of each element in a reaction it is possible to determine if it is a redox reaction. If it is a redox reaction, the oxidation number of one element will increase and the oxidation number of another element will decrease.

Consider the following reaction again:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

We can write the half equation for the conversion of magnesium to magnesium ions and show the oxidation number of each below:

$$\begin{array}{lll} 2Mg(s) & \longrightarrow & 2Mg^{2+}(s) + 4e^- \\ (0) & & (+2) \end{array}$$

Looking at this, we can see that the oxidation number of each magnesium atom has increased from 0 to +2. Magnesium (Mg) has been **oxidised**.

We can write the half equation for the conversion of oxygen to oxide ions and show the oxidation numbers below:

$$\begin{array}{lll} O_2(g) + 4e^- & \longrightarrow & 2O^{2-}(s) \\ (0) & & (-2) \end{array}$$

Looking at this, we can see that the oxidation number of each oxygen atom in the oxygen molecule has decreased from 0 to −2. Oxygen (O_2) has been **reduced**.

It is important to note that in writing the oxidation number of each element we disregard any subscripts or coefficients. We are writing the oxidation number of a single atom or ion of the element only.

Key fact

Oxidation is the increase in the oxidation number of an element in its free state or an element in a compound.

Key fact

Reduction is the decrease in the oxidation number of an element in its free state or an element in a compound.

Exam tip

It is easy to remember that reduction is a decrease in oxidation number, because when you reduce anything, you make it **less**.

Recognising redox reactions using oxidation numbers

Changes in oxidation number are not always brought about by a transfer of electrons. The oxidation numbers of elements in covalent substances can also change when they react to form other covalent substances. For example, when hydrogen and oxygen react to form water:

$$\begin{array}{lllll} & 2H_2(g) & + O_2(g) & \longrightarrow & 2H_2O(g) \\ \text{oxidation numbers:} & (0) & (0) & & (+1)(-2) \end{array}$$

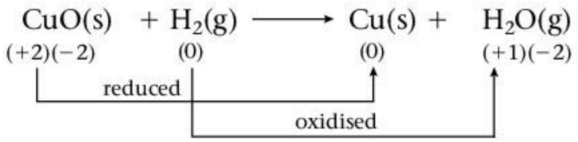
- The oxidation number of each hydrogen atom in the hydrogen molecule has increased from 0 to +1. Hydrogen (H_2) has been **oxidised**.
- The oxidation number of each oxygen atom in the oxygen molecule has decreased from 0 to −2. Oxygen (O_2) has been **reduced**.

Redox reactions which involve a transfer of electrons and also those which don't involve a transfer of electrons can be recognised using **oxidation number** by the steps outlined below.

1) Write the balanced chemical equation for the reaction.
2) Write the oxidation number of each element in brackets below it. You do not need to determine the oxidation numbers of elements in polyatomic ions which remain unchanged during a reaction.
3) Determine which element shows an *increase* in its oxidation number. This is the element which has been *oxidised*.
4) Determine which element shows a *decrease* in its oxidation number. This is the element which has been *reduced*.
5) If there is no increase in oxidation number of one element and a decrease in oxidation number of another element, then the reaction is *not* a redox reaction.

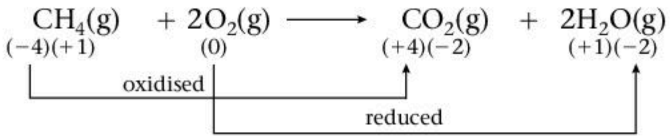
Examples

1 Determine which element has been oxidised and which element has been reduced in the reaction between copper(II) oxide and hydrogen.



- The oxidation number of each hydrogen atom in the hydrogen molecule has increased from 0 to +1. Hydrogen (H_2) has been **oxidised**.
- The oxidation number of the Cu^{2+} ion in the copper(II) oxide has decreased from +2 to 0. Copper(II) oxide (CuO) has been **reduced**.

2 Determine which element has been oxidised and which has been reduced in the reaction between methane and oxygen.



- The oxidation number of the carbon atom in the methane molecule has increased from −4 to +4. Methane (CH_4) has been **oxidised**.
- The oxidation number of each oxygen atom in the oxygen molecule has decreased from 0 to −2. Oxygen (O_2) has been **reduced**.

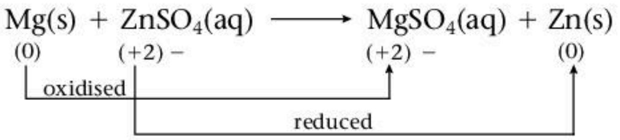
3 Determine whether the following reaction is a redox reaction or not.

$$\underset{(+1)\quad -}{2AgNO_3(aq)} + \underset{(+2)(-1)}{MgCl_2(aq)} \longrightarrow \underset{(+1)(-1)}{2AgCl(s)} + \underset{(+2)\quad -}{Mg(NO_3)_2(aq)}$$

Because the NO_3^- ion remains unchanged in the reaction, i.e. it is one of the spectator ions, the oxidation numbers of the nitrogen and oxygen in the ion do not need to be considered.

None of the elements has undergone a change in oxidation number. This is *not* a redox reaction.

4 Determine which element has been oxidised and which has been reduced in the reaction between magnesium and zinc sulfate.



$$\underset{(0)}{Mg(s)} + \underset{(+2)\ -}{ZnSO_4(aq)} \longrightarrow \underset{(+2)\ -}{MgSO_4(aq)} + \underset{(0)}{Zn(s)}$$

Because the SO_4^{2-} ion remains unchanged in the reaction, the oxidation numbers of the sulfur and oxygen in the ion do not need to be considered.

- The oxidation number of the magnesium has increased from 0 to +2. Magnesium (Mg) has been **oxidised**.
- The oxidation number of each Zn^{2+} ion in the zinc sulfate has decreased from +2 to 0. Zinc sulfate ($ZnSO_4$) has been **reduced**.

Summary questions

1 Define oxidation and reduction in terms of oxidation number.

2 Determine the oxidation number of:

a copper in CuO
b sulfur in H_2S
c carbon in each of CO, CO_2, H_2CO_3 and C_2H_4
d nitrogen in each of NH_3, Mg_3N_2 and NH_4Cl
e chlorine in each of $AlCl_3$, NaClO and $KClO_3$.

3 Use oxidation number to determine whether the following are redox reactions. If they are redox reactions give the reactant which has been oxidised and the reactant which has been reduced:

a $Cl_2(g) + 2KI(aq) \longrightarrow 2KCl(aq) + I_2(aq)$
b $4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$
c $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
d $BaCl_2(aq) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + 2HCl(aq)$.

Objectives

By the end of this topic you will be able to:

- define the terms oxidising and reducing agent in terms of electrons
- define the terms oxidising and reducing agent in terms of oxidation number
- recognise oxidising and reducing agents
- give examples of compounds which can act as both an oxidising agent and a reducing agent
- describe tests to identify oxidising and reducing agents.

A9.3 Oxidising and reducing agents

In any redox reaction, oxidation and reduction occur simultaneously.

- One reactant brings about the oxidation of the other reactant. This is known as the **oxidising agent**.
- One reactant brings about the reduction of the other reactant. This is known as the **reducing agent**.

This can generally be symbolised as follows, where A is the reactant that has been oxidised and B is the reactant that has been reduced:

$$\underset{\text{oxidised}}{A} + \underset{\text{reduced}}{B} \longrightarrow C + D$$

In the above reaction:

- A must have been oxidised by B. B is the oxidising agent.
- B must have been reduced by A. A is the reducing agent.

Oxidising and reducing agents in terms of electrons

Consider the example of aluminium reacting with chlorine gas:

$$2Al(s) + 3Cl_2(g) \longrightarrow 2AlCl_3(s)$$

- The aluminium has been oxidised because each atom has lost three electrons to form an aluminium ion (Al^{3+}):

  $$2Al(s) \longrightarrow 2Al^{3+}(s) + 6e^-$$

  The chlorine atoms brought about this loss by taking these electrons. Chlorine (Cl_2) is the **oxidising agent**.
- The chlorine has been reduced because each atom in the molecule gained one electron to form a chloride ion (Cl^-):

  $$3Cl_2(g) + 6e^- \longrightarrow 6Cl^-(s)$$

  The aluminium atoms brought about this gain by donating these electrons. Aluminium (Al) is the **reducing agent**.

An oxidising agent, therefore, causes another reactant to *lose* electrons and a reducing agent causes another reactant to *gain* electrons.

Key fact

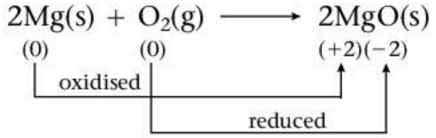
An **oxidising agent** brings about the oxidation of another reactant by causing an atom or ion in that reactant to **lose** electrons.

Key fact

A **reducing agent** brings about the reduction of another reactant by causing an atom or ion in that reactant to **gain** electrons.

Oxidising and reducing agents in terms of oxidation number

Consider the example of magnesium burning in oxygen again:



$$\underset{(0)}{2Mg(s)} + \underset{(0)}{O_2(g)} \longrightarrow \underset{(+2)(-2)}{2MgO(s)}$$

- The magnesium has been oxidised because the oxidation number of each magnesium atom has increased from 0 to +2. Oxygen (O_2) is the **oxidising agent** because it is the reactant that caused the increase in oxidation number.
- The oxygen has been reduced because the oxidation number of each oxygen atom has decreased from 0 to −2. Magnesium (Mg) is the **reducing agent** because it is the reactant that caused the decrease in oxidation number.

An oxidising agent, therefore, causes the oxidation number of an atom or ion in another reactant to *increase*. A reducing agent causes the oxidation number of an atom or ion in another reactant to *decrease*.

Exam tip

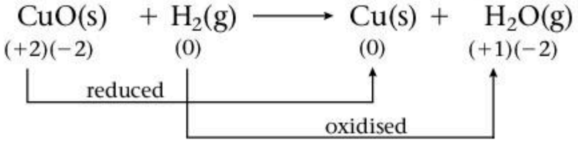
It can be very confusing to try to learn the definitions of oxidation and reduction as well as those of oxidising agent and reducing agent. It is better to **work them out** when needed.

You know that **O**xidation **I**s the **L**oss of electrons (**OIL**). An oxidising agent must cause this loss by taking these electrons. The oxidising agent itself gains electrons.

You know that **R**eduction **I**s the **G**ain of electrons (**RIG**). A reducing agent must cause this gain by giving electrons. The reducing agent itself loses electrons.

Examples

1 Determine which reactant is the oxdising agent and which is the reducing agent in the reaction between copper(II) oxide and hydrogen.



$$\underset{(+2)(-2)}{CuO(s)} + \underset{(0)}{H_2(g)} \longrightarrow \underset{(0)}{Cu(s)} + \underset{(+1)(-2)}{H_2O(g)}$$

- Copper(II) oxide (CuO) is the **oxidising agent** because it oxidised the hydrogen by causing the oxidation number of each hydrogen atom to increase from 0 to +1.
- Hydrogen (H_2) is the **reducing agent** because it reduced the copper(II) oxide by causing the oxidation number of the Cu^{2+} ion to decrease from +2 to 0.

Key fact

An **oxidising agent** brings about the oxidation of another reactant by causing the oxidation number of an atom or ion in that reactant to **increase**.

Key fact

A **reducing agent** brings about the reduction of another reactant by causing the oxidation number of an atom or ion in that reactant to **decrease**.

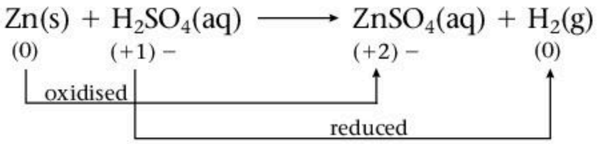
Exam tip

It is better to **work out** the definitions of oxidising agent and reducing agent using the concept of oxidation number when you need them.

If you remember that oxidation involves an increase in oxidation number, an oxidising agent must cause this increase.

If you remember that reduction involves a decrease in oxidation number, a reducing agent must cause this decrease.

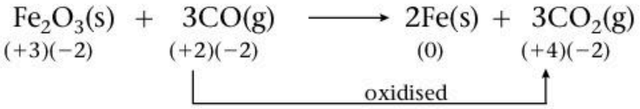
2 Determine which reactant is the oxdising agent and which is the reducing agent in the reaction between zinc and sulfuric acid.



$$\underset{(0)}{Zn(s)} + \underset{(+1)\,-}{H_2SO_4(aq)} \longrightarrow \underset{(+2)\,-}{ZnSO_4(aq)} + \underset{(0)}{H_2(g)}$$

- Sulfuric acid (H_2SO_4) is the **oxidising agent** because it oxidised the zinc by causing the oxidation number of the zinc atom to increase from 0 to +2.
- Zinc (Zn) is the **reducing agent** because it reduced the sulfuric acid by causing the oxidation number of each H^+ ion to decrease from +1 to 0.

3 Is iron(III) oxide behaving as an oxdising agent or a reducing agent when it reacts with carbon monoxide?



$$\underset{(+3)(-2)}{Fe_2O_3(s)} + \underset{(+2)(-2)}{3CO(g)} \longrightarrow \underset{(0)}{2Fe(s)} + \underset{(+4)(-2)}{3CO_2(g)}$$

To determine the behaviour of iron(III) oxide we need to look to see if the carbon monoxide has been oxidised or reduced.

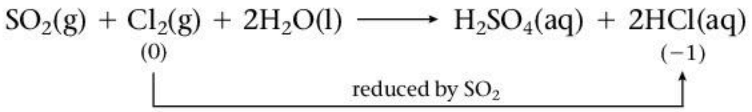
The carbon monoxide has been oxidised. The iron(III) oxide (Fe_2O_3) is behaving as an **oxidising agent** because it caused the oxidation number of the carbon atom in the carbon monoxide (CO) to increase from +2 to +4.

Substances that can behave as both oxidising and reducing agents

There are some compounds which can act as both oxidising and reducing agents. Their behaviour depends on the other reactant. Examples of two compounds which can behave in this way are given below.

Sulfur dioxide, SO_2

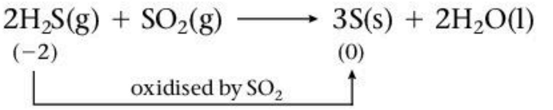
Sulfur dioxide usually acts as a *reducing agent*. This is seen, for example, when it reacts with chlorine:



$$SO_2(g) + \underset{(0)}{Cl_2(g)} + 2H_2O(l) \longrightarrow H_2SO_4(aq) + \underset{(-1)}{2HCl(aq)}$$

The sulfur dioxide reduced the chlorine by causing the oxidation number of each chlorine atom in the chlorine molecule to decrease from 0 to −1.

When it reacts with a stronger reducing agent than itself, such as hydrogen sulfide, sulfur dioxide behaves as an *oxidising agent*:



$$\underset{(-2)}{2H_2S(g)} + SO_2(g) \longrightarrow \underset{(0)}{3S(s)} + 2H_2O(l)$$

The sulfur dioxide oxidised the hydrogen sulfide by causing the oxidation number of the sulfur atom in the hydrogen sulfide molecule to increase from −2 to 0.

Acidified hydrogen peroxide, H^+/H_2O_2

Acidified hydrogen peroxide is hydrogen peroxide solution to which sulfuric acid has been added. It usually behaves as an *oxidising agent*. This is seen, for example, when it reacts with potassium iodide:

$$\underset{(-1)}{2\text{KI(aq)}} + \text{H}_2\text{O}_2\text{(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \longrightarrow \underset{(0)}{\text{I}_2\text{(g)}} + \text{K}_2\text{SO}_4\text{(aq)} + 2\text{H}_2\text{O(l)}$$

oxidised by H_2O_2

The hydrogen peroxide oxidised the potassium iodide by causing the oxidation number of the I^- ion to increase from −1 to 0.

When it reacts with a stronger oxidising agent than itself, such as acidified potassium manganate(VII), acidified hydrogen peroxide behaves as a *reducing agent*:

$$\underset{(+7)}{2\text{KMnO}_4\text{(aq)}} + 5\text{H}_2\text{O}_2\text{(aq)} + 3\text{H}_2\text{SO}_4\text{(aq)} \longrightarrow \text{K}_2\text{SO}_4\text{(aq)} + 5\text{O}_2\text{(g)} + \underset{(+2)}{2\text{MnSO}_4\text{(aq)}} + 8\text{H}_2\text{O(l)}$$

reduced by H_2O_2

The hydrogen peroxide reduced the potassium manganate(VII) by causing the oxidation number of the manganese to decrease from +7 to +2.

Tests for oxidising agents and reducing agents

There are certain chemicals which undergo oxidation or reduction in a chemical reaction in a very distinct way, e.g. they may change to a distinct colour, they may form a particular precipitate or they may give off a particular gas. Chemists use these chemicals to determine whether unknown substances are oxidising agents or reducing agents. They do this by reacting the unknown substance with these chemicals and observing the products. These chemicals form the basis of tests for oxidising agents and reducing agents.

Examples

Potassium iodide solution, KI(aq)

Potassium iodide solution is used in the laboratory to test for an oxidising agent. There is a distinct colour change in the solution from *colourless* to *brown* when the potassium iodide is *oxidised*. The colour change occurs because the oxidising agent oxidises the colourless iodide ion (I^-) to iodine (I_2) which dissolves forming a brown solution.

$$\underset{(-1)}{2\text{I}^-\text{(aq)}} \longrightarrow \underset{(0)}{\text{I}_2\text{(aq)}} + 2e^-$$

oxidised

The potassium iodide is, itself, a reducing agent.

Acidified potassium manganate(VII) solution, $H^+/KMnO_4(aq)$

Acidified potassium manganate(VII) solution is used in the laboratory to test for a reducing agent. There is a distinct colour change in the solution from *purple* to *colourless* when the potassium manganate(VII) is *reduced*. The colour change occurs because the reducing agent reduces the purple manganate(VII) ion (MnO_4^-) to the colourless manganese(II) ion (Mn^{2+}).

$$\underset{(+7)}{\text{MnO}_4^-\text{(aq)}} \longrightarrow \underset{(+2)}{\text{Mn}^{2+}\text{(aq)}}$$

reduced

The acidified potassium manganate(VII) is, itself, an oxidising agent.

▲ **Figure 9.3.1** Potassium iodide turns from colourless to brown when oxidised

Tests for oxidising agents

Table 9.3.1 summarises the names and colour changes of the main two chemicals which can be used as the reactants to test for oxidising agents. These chemicals are, themselves, reducing agents.

▼ **Table 9.3.1** Reducing agents used to test for the presence of an oxidising agent

Reducing agent used for test	Visible change when agent is oxidised	Explanation
Potassium iodide solution (KI(aq))	Colourless to brown	The colourless iodide ion (I^-) is oxidised to iodine (I_2) which dissolves forming a brown solution
Iron(II) sulfate solution ($FeSO_4(aq)$)	Pale green to yellow-brown	The pale green iron(II) ion (Fe^{2+}) is oxidised to the yellow-brown iron(III) ion (Fe^{3+})

There are several other chemicals which behave as reducing agents in reactions, although they are not usually used to test for the presence of an oxidising agent. These are given in Table 9.3.2.

▼ **Table 9.3.2** Other common reducing agents

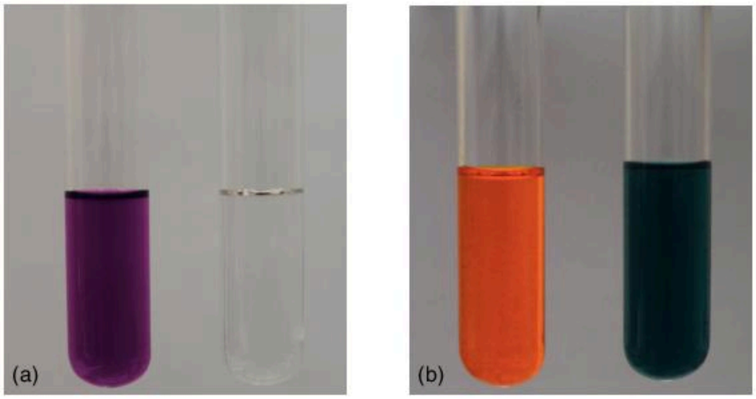
Reducing agent	Visible change when agent is oxidised	Explanation
Hydrogen sulfide gas ($H_2S(g)$)	A yellow precipitate forms	Yellow insoluble sulfur (S) is produced
Concentrated hydrochloric acid (HCl(aq))	A yellow-green gas is evolved	Yellow-green chlorine gas (Cl_2) is produced
Others: hydrogen gas (H_2); carbon (C); carbon monoxide (CO); reactive metals		

Tests for reducing agents

Table 9.3.3 summarises the names and colour changes of the main two chemicals which can be used as the reactants to test for reducing agents. These chemicals are, themselves, oxidising agents.

▼ **Table 9.3.3** Oxidising agents used to test for the presence of a reducing agent

Oxidising agent used for test	Visible change when agent is reduced	Explanation
Acidified potassium manganate(VII) solution ($H^+/KMnO_4(aq)$)	Purple to colourless	The purple manganate(VII) ion (MnO_4^-) is reduced to the colourless manganese(II) ion (Mn^{2+})
Acidified potassium dichromate(VI) solution ($H^+/K_2Cr_2O_7(aq)$)	Orange to green	The orange dichromate(VI) ion ($Cr_2O_7^{2-}$) is reduced to the green chromium(III) ion (Cr^{3+})



▶ **Figure 9.3.2 (a)** Acidified potassium manganate(VII) turns from purple to colourless when reduced. **(b)** Acidified potassium dichromate(VI) turns from orange to green when reduced

There are several other chemicals which behave as oxidising agents, although they are not usually used to test for the presence of a reducing agent. These are given in Table 9.3.4.

▼ **Table 9.3.4** Other common oxidising agents

Oxidising agent	Visible change when agent is reduced	Explanation
Iron(III) sulfate solution ($Fe_2(SO_4)_3(aq)$)	Yellow-brown to pale green	The yellow-brown iron(III) ion (Fe^{3+}) is reduced to the pale green iron(II) ion (Fe^{2+})
Dilute or concentrated nitric acid ($HNO_3(aq)$)	A brown gas is evolved	Brown nitrogen dioxide gas (NO_2) is produced
Hot concentrated sulfuric acid $H_2SO_4(l)$)	A pungent gas is evolved which turns acidified potassium manganate(VII) colourless	Sulfur dioxide gas (SO_2) is produced
Others: sodium chlorate(I) solution (NaClO); oxygen gas (O_2); chlorine gas (Cl_2); manganese(IV) oxide (MnO_2)		

Practical activity

To investigate the reactions of oxidising agents and reducing agents

Your teacher may use this activity to assess:

- observation, recording and reporting
- analysis and interpretation.

You will be supplied with aqueous solutions of acidified potassium manganate(VII), acidified potassium dichromate(VI), potassium iodide, iron(II) sulfate and acidified hydrogen peroxide and five test tubes.

Method

1. Place 2 cm^3 of acidified potassium manganate(VII) solution into the first test tube. Slowly add potassium iodide solution, shaking as you add the solution, until you see no further colour change.
2. Place 2 cm^3 of acidified potassium manganate(VII) solution into the second test tube, 2 cm^3 of acidified potassium dichromate(VI) solution into the third tube, 2 cm^3 of potassium iodide solution into the fourth tube and 2 cm^3 of iron(II) sulfate solution into the fifth tube. Slowly add acidified hydrogen peroxide solution to each tube, shaking as you add the solution, until you see no further colour change.
3. Record in a table the original colour of each solution and the final colour after mixing.
4. Explain the reason for the colour change that you observed in each reaction. Use the information given in Tables 9.3.1 and 9.3.3 above to help you.
5. Comment on the behaviour of the acidified hydrogen peroxide.

Summary questions

1. Define the terms oxidising agent and reducing agent
   a in terms of electrons
   b in terms of oxidation number.
2. In each of the following reactions identify the oxidising agent and the reducing agent:
   a $2Fe_2O_3(s) + 3C(s) \longrightarrow 3CO_2(g) + 4Fe(s)$
   b $2NO(g) + 2CO(g) \longrightarrow N_2(g) + 2CO_2(g)$
   c $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
3. Give an example of a compound which can act both as an oxidising agent and as a reducing agent.

Key concepts

- We encounter, and make use of, oxidation and reduction reactions in our everyday lives. We encounter them in rusting and the browning of fruits, and make use of them in bleaches, food preservation and the breathalyser test.
- Oxidation–reduction reactions are known as **redox reactions**.
- Many redox reactions involve a transfer of electrons.
- **Oxidation** is defined as the loss of electrons by an element in its free state or an element in a compound.
- **Reduction** is defined as the gain of electrons by an element in its free state or an element in a compound.
- Oxidation and reduction can also be defined using the concept of oxidation number.
- **Oxidation number** is the theoretical charge that an atom of an element would have if all the bonds between the atoms in the compound containing it were ionic and the compound was composed entirely of ions.
- In any redox reaction there will always be a change in the oxidation numbers of two elements.
- **Oxidation** is defined as an increase in oxidation number of an element in its free state or an element in a compound.
- **Reduction** is defined as a decrease in oxidation number of an element in its free state or an element in a compound.
- An **oxidising agent** is the reactant that brings about the oxidation of another reactant.
- A **reducing agent** is the reactant that brings about the reduction of another reactant.
- An oxidising agent causes the other reactant to lose electrons.
- A reducing agent causes the other reactant to gain electrons.
- An oxidising agent causes the oxidation number of another element to increase.
- A reducing agent causes the oxidation number of another element to decrease.
- The distinctive colour changes that accompany the oxidation or reduction of certain compounds allow these compounds to be used when testing for the presence of oxidising or reducing agents.
- Potassium iodide and iron(II) sulfate are reducing agents which can be used to test for the presence of an oxidising agent.
- Acidified potassium manganate(VII) and acidified potassium dichromate(VI) are oxidising agents which can be used to test for the presence of a reducing agent.

Practice exam-style questions

Multiple-choice questions

1 Oxidation is:

I a gain of oxygen
II a gain of electrons
III a decrease in oxidation number

A I only
B I and II only
C I, II and III
D III only

2 The oxidation number of sulfur in the $S_2O_3^{2-}$ ion is:

A +2
B +4
C +6
D +8

3 The name of the BrO_3^- ion is:

A bromite(I)
B bromate(I)
C bromite(V)
D bromate(V)

4 Consider the following:

$$M(s) \longrightarrow M^{2+}(aq)$$

In the reaction represented above, metal M:

I has been oxidised
II has increased in oxidation number
III has gained electrons

A I only
B I and II only
C II and III only
D I, II and III

5 Which of the following is true about a reducing agent?

A It causes an increase in oxidation number.
B It causes a loss of electrons.
C Its oxidation number increases.
D It gains electrons.

6 In which of the following reactions does the oxidation number of nitrogen show the greatest increase?

A $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$
B $4NO_2(g) + O_2(g) + 2H_2O(l) \longrightarrow 4HNO_3(aq)$
C $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
D $2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$

7 Which of the following reactions is not a redox reaction?

A $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$
B $Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$
C $MgO(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$
D $C(s) + O_2(g) \longrightarrow CO_2(g)$

8 In the following reaction:

$$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

which of the following statements about copper is true?

A Copper undergoes reduction and its oxidation number decreases.
B Copper undergoes oxidation and its oxidation number decreases.
C Copper undergoes reduction and its oxidation number increases.
D Copper undergoes oxidation and its oxidation number increases.

9 In separate experiments, sulfur dioxide was bubbled into acidified solutions of potassium dichromate(VI) and acidified potassium manganate(VII). Which of the following pairs correctly describes the colour changes observed in the experiment?

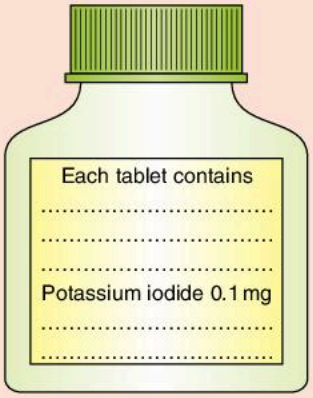
	Colour of potassium dichromate(VI)	**Colour of potassium manganate(VII)**
A	orange to green	colourless to purple
B	orange to green	purple to colourless
C	green to orange	colourless to purple
D	green to orange	purple to colourless

10 When acidified hydrogen peroxide solution is added to iron(II) sulfate solution:

A a pale green solution is formed
B the iron(II) ion is oxidised to the iron(III) ion
C the oxidation number of the iron(II) ion decreases
D the acidified hydrogen peroxide is acting as a reducing agent

Structured question

11 **a** The figure below shows the label on a multivitamin and mineral supplement bottle. Much of the information has been left out.



A multivitamin and mineral supplement bottle

A solution is made of one of the tablets and acidified hydrogen peroxide is added to the solution.

i) State what you would expect to observe. *(1 mark)*
ii) Write an ionic half equation to represent the reaction. *(2 marks)*
iii) With a reason, state whether the reaction represents oxidation or reduction. *(1 mark)*

b Three redox reactions are represented by the equations given below. For EACH reaction, name the substance which has been oxidised. In EACH case, give the reason for your choice based on oxidation number.

i) $Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$ ***(2 marks)***

ii) $2Fe^{2+}(aq) + Cl_2(g) \longrightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$ ***(2 marks)***

iii) $2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$ ***(2 marks)***

c One type of the breathalyser test uses acidified potassium dichromate(VI) crystals.

i) For what purpose is the breathalyser test used? ***(1 mark)***

ii) Describe what you would observe if the result of the test is positive. ***(1 mark)***

iii) Explain the chemistry involved in the breathalyser test. ***(2 marks)***

d Using oxidation number, suggest a name for the ClO_2^- ion. ***(1 mark)***

Total 15 marks

Extended response question

12 a Define oxidation in terms of:

i) electrons

ii) oxidation number. ***(2 marks)***

b Support EACH definition given in **a** above by reference to the following reaction:

$$Cl_2(g) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(g)$$

(4 marks)

c 'Metals are reducing agents and non-metals are oxidising agents.'
Use EACH of the reactions below to determine whether the above statement is true. Give reasons for your answers based on oxidation number.

i) $2PbO(s) + C(s) \longrightarrow Pb(s) + CO_2(g)$ ***(2 marks)***

ii) The reaction between magnesium and copper(II) sulfate solution. ***(3 marks)***

d You are provided with two bottles labelled X and Y containing an oxidising agent and a reducing agent, respectively. Both agents are colourless liquids. Describe TWO tests you could carry out to confirm what you are told about X and Y. ***(4 marks)***

Total 15 marks

A10 Electrochemistry

Electrochemistry is the study of the relationship between chemical reactions and electrical energy. An electrochemical reaction is one which either produces electrical energy or requires electrical energy in order to proceed. From the reactions that occur in a battery to produce an electric current, to the electrolysis of aluminium oxide to produce aluminium, electrochemistry is important in our lives.

A10.1 The electrochemical series

The electrochemical series of metals

In Unit 5.2 you learnt that when atoms bond ionically, the metal atom always loses electrons to form a positive metal cation. Some metals lose electrons much more easily than others. When metals are placed in order of ease with which they lose electrons, i.e. ease with which they ionise, a series known as the **electrochemical series of metals** is created.

A section of the electrochemical series of the common metals is given in Table 10.1.1. In this table, the ease of ionisation increases going upwards.

The more easily a metal atom ionises, the more stable are the ions that it forms. Metals at the top of the series form very *stable* ions which are hard to convert back to atoms. Metals at the bottom of the series form *unstable* ions which are easy to convert back to atoms.

Since metals lose electrons when they ionise, they act as **reducing agents** when they react with other substances, i.e. they give electrons to the other reactant. The more easily they give electrons, the stronger their reducing power. The strength as a reducing agent, therefore, *increases* going *up* the series.

Potassium is highest in the series, which means that it ionises the most easily, forms the most stable ions and is the *strongest* reducing agent. Silver is the lowest in the series, which means that it ionises the least easily, forms the least stable ions and is the *weakest* reducing agent.

We can make use of the electrochemical series of metals to predict certain chemical reactions.

Displacement of metals

A metal will displace a metal that is *below* it in the electrochemical series from a compound containing the lower metal. This is because the higher metal is a stronger reducing agent, therefore, readily gives electrons to the ions of the lower metal. As a result, the higher metal **ionises** and the ions of the lower metal are converted to atoms, i.e. the ions are said to have been **discharged**.

Objectives

By the end of this topic you will be able to:

- give the electrochemical series of the common metals
- determine if a displacement reaction will occur between metals and their compounds based on their relative positions in the electrochemical series
- determine if a metal will displace hydrogen from an acid based on its position in the electrochemical series
- give the electrochemical series of some common non-metals
- determine if a displacement reaction will occur between certain non-metals and their compounds based on their relative positions in the electrochemical series.

▼ **Table 10.1.1** The electrochemical series of the common metals

Metal	Cation
potassium	K^+
calcium	Ca^{2+}
sodium	Na^+
magnesium	Mg^{2+}
aluminium	Al^{3+}
zinc	Zn^{2+}
iron	Fe^{2+}
lead	Pb^{2+}
hydrogen	H^+
copper	Cu^{2+}
silver	Ag^+

Examples

1 Will a reaction occur if a strip of magnesium ribbon is placed in copper(II) sulfate solution?

Magnesium is *higher* than copper in the electrochemical series, therefore, magnesium can displace the Cu^{2+} ions from the copper(II) sulfate. Magnesium is a stronger reducing agent, therefore, it gives electrons to the Cu^{2+} ions. The magnesium **ionises** forming Mg^{2+} ions and the Cu^{2+} ions are **discharged** forming copper atoms:

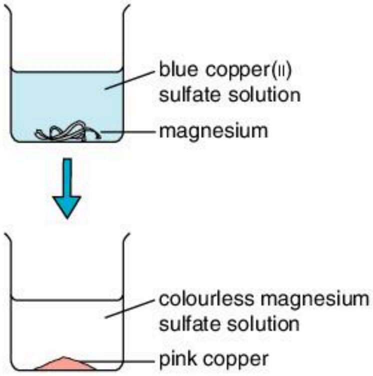
$$Mg(s) + CuSO_4(aq) \longrightarrow MgSO_4(aq) + Cu(s)$$

The ionic equation is:

$$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$$

The transfer of electrons is shown in the ionic half equations:

$$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$$
$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$



▲ **Figure 10.1.1** Magnesium is added to copper(II) sulfate solution

As the reaction proceeds the following are observed:

- The magnesium ribbon gradually gets smaller as it ionises and the Mg^{2+} ions dissolve in the solution.
- A pink solid builds up in the solution as copper is formed.
- The blue colour of the copper(II) sulfate solution gradually fades as the blue Cu^{2+} ions are discharged from the solution, forming copper.

2 Will a reaction occur if some iron filings are placed in a solution of zinc nitrate?

Iron is *lower* than zinc in the electrochemical series. *No* reaction will occur.

Displacement of hydrogen

Metals above hydrogen in the electrochemical series will displace the H^+ ions in an acid, forming hydrogen gas (H_2). Metals below hydrogen will not displace the H^+ ions. Metals above hydrogen are stronger reducing agents and will readily give electrons to the H^+ ions of the acid. Metals below hydrogen are weaker reducing agents, so will not reduce the H^+ ions.

Examples

1 Will a reaction occur if zinc is placed in hydrochloric acid?

Zinc is *higher* than hydrogen in the electrochemical series, therefore, zinc can displace the H^+ ions from the hydrochloric acid. Zinc is a stronger reducing agent, therefore, it gives electrons to the H^+ ions. The zinc **ionises** forming Zn^{2+} ions and the H^+ ions are **discharged** forming hydrogen gas (H_2):

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

The ionic equation is:

$$Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$$

The transfer of electrons is shown in the ionic half equations:

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$
$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

2 Will a reaction occur if some copper turnings are placed in sulfuric acid?

Copper is *lower* than hydrogen in the electrochemical series. *No* reaction will occur.

The electrochemical series of non-metals

When non-metal atoms bond ionically, the non-metal atom always gains electrons to form a negative non-metal anion. Some non-metals gain electrons much more easily than others. When non-metals are placed in order of ease with which they gain electrons, i.e. ease with which they ionise, a series known as the **electrochemical series of non-metals** is created.

A section of the electrochemical series of some non-metals is given in Table 10.1.2. In this table, the ease of ionisation increases going upwards.

▼ **Table 10.1.2** The electrochemical series of some non-metals

Non-metal	Anion
fluorine	F^-
chlorine	Cl^-
bromine	Br^-
iodine	I^-

The easier a non-metal atom ionises, the more stable are the ions that it forms. Non-metals at the top of the series form very *stable* ions, which are hard to convert back to atoms. Non-metals at the bottom of the series form *unstable* ions, which are easy to convert back to atoms.

Since non-metals gain electrons when they ionise, they act as **oxidising agents** when they react with other substances, i.e. they remove electrons from the other reactant. The more easily they remove electrons, the stronger their oxidising power. The strength as an oxidising agent, therefore, *increases* going *up* the series.

Fluorine is highest in the series, which means that it ionises the most easily, forms the most stable ions and is the *strongest* oxidising agent. Iodine is the lowest in the series, which means that it ionises the least easily, forms the least stable ions and is the *weakest* oxidising agent.

We can make use of the electrochemical series of non-metals to predict certain chemical reactions.

Displacement of non-metals

A non-metal will displace a non-metal which is *below* it in the electrochemical series from a compound containing the lower non-metal. This is because the higher non-metal is a stronger oxidising agent, therefore, readily removes electrons from the ions of the lower non-metal. As a result, the higher non-metal ionises, and the ions of the lower non-metal are discharged forming atoms.

Examples

1 Will a reaction occur if chlorine gas is bubbled into sodium iodide solution?

Chlorine is *higher* than iodine in the electrochemical series of non-metals, therefore, chlorine can displace the I^- ions from the sodium iodide. Chlorine is a stronger oxidising agent, therefore, it removes electrons from the I^- ions. The chlorine **ionises** forming Cl^- ions and the I^- ions are **discharged** forming iodine (I_2):

$$Cl_2(g) + 2NaI(aq) \longrightarrow 2NaCl(aq) + I_2(aq)$$

The ionic equation is:

$$Cl_2(g) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(aq)$$

The transfer of electrons is shown in the ionic half equations:

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$$

$$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$$

2 Will a reaction occur if bromine solution is added to a solution of potassium chloride?

Bromine is *lower* than chlorine in the electrochemical series. *No* reaction will occur.

Summary questions

1 If zinc is placed in a solution of copper(II) nitrate, a reaction occurs.
   - a Explain why a reaction occurs.
   - b Give a balanced chemical equation and a balanced ionic equation for the reaction.
   - c What happens to the colour of the copper(II) nitrate solution? Explain why this occurs.

2 For each of the following, predict whether or not a displacement reaction will occur:
   - a aluminium + lead(II) nitrate solution
   - b silver + calcium nitrate solution
   - c iodine + sodium chloride solution
   - d bromine + potassium iodide solution.

3 Which of the metals in the electrochemical series will release hydrogen gas when added to hydrochloric acid?

4 A metal X was found to displace zinc from zinc nitrate solution but no displacement occurred when it was added to magnesium nitrate solution. Arrange the three metals, X, zinc and magnesium in decreasing order of ease of ionisation.

A10.2 Electrical conduction

Objectives

By the end of this topic you will be able to:

- distinguish between a conductor and a non-conductor
- distinguish between metallic and electrolytic conduction
- distinguish between a strong electrolyte, a weak electrolyte and a non-electrolyte
- give examples of strong electrolytes, weak electrolytes and non-electrolytes.

In Unit 5 you learnt that in order for a substance to conduct an electric current, it must contain charged particles which are able to move through the substance. These charged particles can be either electrons or ions. Materials can be classified into two groups based on their ability to conduct an electric current.

- **Conductors** are substances which allow electricity to pass through. They can be solids, liquids or solutions. Examples of conductors include metals, graphite, molten ionic compounds, solutions of ionic compounds and aqueous acids and alkalis.
- **Non-conductors** are substances which do not allow an electric current to pass through. Non-conductors can be solids, liquids, solutions and gases. Examples of non-conductors include non-metals (except graphite), plastics, covalent substances and solid ionic compounds.

Practical activity

To investigate the electrical conductivity of various substances

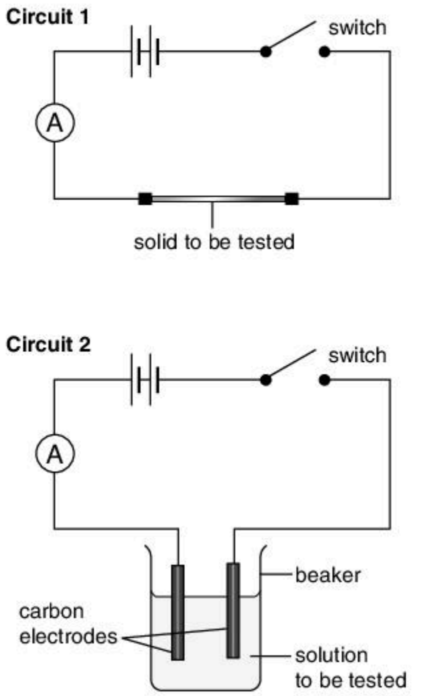
Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement.

You will be supplied with circuit components, samples of various substances, e.g. copper wire, magnesium ribbon, aluminium foil, an iodine crystal, a graphite rod, a plastic ruler, a sodium chloride crystal, a copper(II) sulfate crystal, sodium chloride solution, copper(II) sulfate solution, dilute hydrochloric acid, ethanol and kerosene.

Method

1. Draw up a table with three columns headed 'Test substance', 'Prediction' and 'Results'. In the first column write down the name of each test substance that you have been supplied with. In the second column fill in your prediction of whether the substance is a conductor or a non-conductor.
2. Connect the circuit as shown in Figure 10.2.1. If the test substance is a solid use Circuit 1 and if it is a solution or a liquid use Circuit 2.
3. Record the results of your experiment in the third column of your table.
4. Compare your predictions with the results.



▲ **Figure 10.2.1** Two circuits to test the conductivity of a solid and a liquid

Metallic and electrolytic conduction

When ionic compounds melt or dissolve in water, the liquid or solution created contains ions and is known as an **electrolyte**. Electrolytes are conductors.

There are certain differences between conduction in a metal, known as **metallic conduction**, and conduction in an electrolyte, known as **electrolytic conduction**.

Key fact

An **electrolyte** is a compound that forms ions when molten or in aqueous solution.

Metallic conduction

In a metal, the valence electrons of the atoms are delocalised. These delocalised electrons, known as mobile electrons, are able to move throughout the metal. Metals are able to conduct an electric current because these **mobile electrons** can move throughout the metal. Although graphite is a non-metal, it also conducts an electric current because of the presence of delocalised, mobile electrons (see Unit 5.5).

During metallic conduction, the metal remains chemically *unchanged*.

Electrolytic conduction

When an ionic compound melts, the ions are no longer held together by ionic bonds and they become free to move throughout the liquid:

$$\mathrm{NaCl(s)} \xrightarrow{\text{heat}} \mathrm{Na^+(l) + Cl^-(l)}$$

Similarly, when an ionic compound dissolves in water, the ionic bonds break and the ions are free to move throughout the solution:

$$\mathrm{NaCl(s)} \xrightarrow{\text{water}} \mathrm{Na^+(aq)} + \mathrm{Cl^-(aq)}$$

These electrolytes conduct an electric current because of the presence of **mobile ions** which are able to move throughout the liquid or aqueous solution.

During electrolytic conduction, the electrolyte is *decomposed*, i.e. it is chemically changed.

▼ **Table 10.2.1** Metallic and electrolytic conduction compared

Metallic conduction	Electrolytic conduction
Mobile electrons carry the electric current through the metal.	**Mobile ions** carry the electric current through the electrolyte
The metal remains chemically **unchanged**	The electrolyte **decomposes**

Strength of an electrolyte

We can distinguish between strong and weak electrolytes based on the concentration of ions in the electrolyte.

Strong electrolytes

Strong electrolytes are **fully ionised** when dissolved in water, e.g. strong acids, strong alkalis and soluble ionic compounds. These have a high concentration of ions in solution. For example:

$$\mathrm{HCl(aq)} \longrightarrow \mathrm{H^+(aq)} + \mathrm{Cl^-(aq)}$$

Molten ionic compounds are also classified as strong electrolytes.

Weak electrolytes

Weak electrolytes are only **partially ionised** when dissolved in water, e.g. weak acids and weak alkalis. These have a low concentration of ions in solution. For example:

$$\mathrm{CH_3COOH(aq)} \rightleftharpoons \mathrm{CH_3COO^-(aq)} + \mathrm{H^+(aq)}$$

▼ **Table 10.2.2** Examples of electrolytes

	Strong electrolytes	Weak electrolytes
Acids	Hydrochloric acid ($HCl(aq)$) Nitric acid ($HNO_3(aq)$) Sulfuric acid ($H_2SO_4(aq)$)	Carbonic acid ($H_2CO_3(aq)$) Ethanoic acid ($CH_3COOH(aq)$)
Alkalis	Potassium hydroxide solution ($KOH(aq)$) Sodium hydroxide solution ($NaOH(aq)$)	Aqueous ammonia ($NH_3(aq)$)
Salts	Molten sodium chloride ($NaCl(l)$) Sodium chloride solution ($NaCl(aq)$) Molten potassium bromide ($KBr(l)$) Potassium bromide solution ($KBr(aq)$)	

Strong electrolytes are much better electrical conductors than weak electrolytes. The strength of an electrolyte can be tested with a simple circuit which has a light bulb (or light emitting diode (LED)) connected to a battery and two electrodes which are dipped in the electrolyte. The greater the strength of the electrolyte, the more brightly the bulb will glow.

A **non-electrolyte** is a substance which remains as molecules when in the liquid state or dissolved in water. The liquid or solution does not contain any ions. Non-electrolytes are non-conductors. Examples of non-electrolytes include liquids such as kerosene and gasoline, molten covalent substances such as wax and solutions of covalent substances such as ethanol and glucose.

Pure water

Pure water is an extremely **weak electrolyte**. Water undergoes spontaneous ionisation into H^+ ions and OH^- ions. At any one time, approximately one in every 5.56×10^8 water molecules is ionised:

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

The concentration of H^+ ions and OH^- ions in pure water is 1×10^{-7} mol dm^{-3}. As you will learn in the next unit, this is of particular significance when an electric current is passed through an aqueous electrolyte.

Summary questions

1 Distinguish between:
   - a a conductor and a non-conductor
   - b metallic and electrolytic conduction
   - c a strong and a weak electrolyte.

2 Classify each of the following as a strong electrolyte, a weak electrolyte or a non-electrolyte:
   - a potassium chloride solution
   - b a solution of sucrose in water
   - c aqueous ethanoic acid
   - d kerosene
   - e aqueous nitric acid.

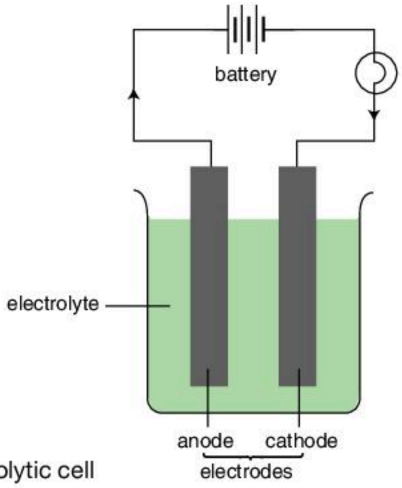
3 Why is pure water an extremely weak electrolyte?

A10.3 Electrolysis

When an electric current is passed through an electrolyte, the electrolyte is decomposed into simpler substances, i.e. it undergoes a chemical change known as **electrolysis**.

Electrolysis takes place in a piece of apparatus known as an **electrolytic cell**. An electrolytic cell has *three* main components.

- A battery or other d.c. **power supply** that provides the electric current.
- Two **electrodes** connected, via wires, to the battery or power supply, which take the electric current into and out of the electrolyte. They must be able to conduct an electric current and are usually made of an inert material such as graphite (carbon) or platinum. The **anode** is connected to the positive terminal of the battery and the **cathode** is connected to the negative terminal.
- The **electrolyte**, which is a molten ionic compound or a solution containing mobile ions.



▶ **Figure 10.3.1** An electrolytic cell

Objectives

By the end of this topic you will be able to:

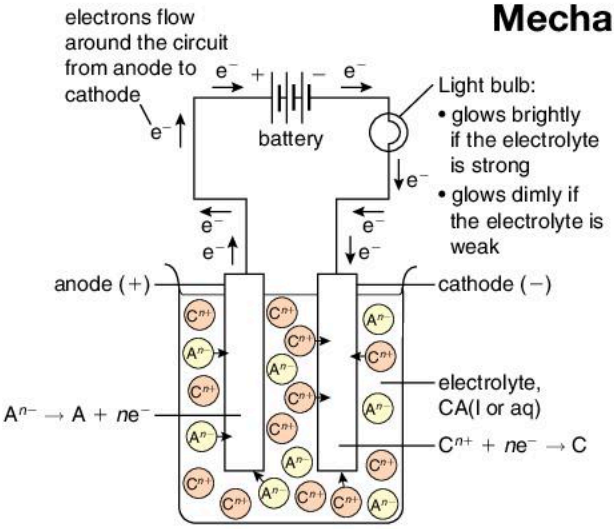
- define the terms electrolysis, anode, cathode, anion and cation
- identify ions present in electrolytes
- predict the electrode to which an ion will drift during electrolysis
- discuss the electrolysis of molten electrolytes
- explain the preferential discharge of ions
- discuss the electrolysis of certain aqueous electrolytes.

Key fact

Electrolysis is the chemical change which occurs when an electric current is passed through an electrolyte.

Key fact

The **anode** is the positive electrode. The **cathode** is the negative electrode.



▲ **Figure 10.3.2** Mechanism of electrolysis

Exam tip

It can be confusing to try to learn which electrode is positive and which is negative. It is better to work it out, remembering that the electrodes are named after the ions which are attracted to them.

- You know that **anions** have a negative charge. Anions must be attracted to the positive electrode. The positive electrode must be called the **anode**.
- You know that **cations** have a positive charge. Cations must be attracted to the negative electrode. The negative electrode must be called the **cathode**.

Mechanism of electrolysis

During electrolysis both anions and cations are **discharged**, i.e. they lose or gain electrons to form neutral atoms. This occurs as follows.

- The **anions** (negative ions) are attracted to the anode (positive electrode). The anions lose electrons to the anode and form atoms, i.e. they are discharged:

  $$A^{n-} \longrightarrow A + ne^-$$

  Oxidation occurs at the anode (OIL). The anode acts as the **oxidising agent**.

- The **electrons**, lost by the anions at the anode, travel through the circuit to the positive terminal of the battery. They then re-enter the circuit from the negative terminal of the battery and travel to the cathode.

- The **cations** (positive ions) are attracted to the cathode (negative electrode). The cations gain the electrons from the cathode forming atoms, i.e. they are discharged:

  $$C^{n+} + ne^- \longrightarrow C$$

  Reduction occurs at the cathode (RIG). The cathode acts as the **reducing agent**.

Electrolysis of molten electrolytes

Molten electrolytes contain only *two* different ions, one cation and one anion. Both are discharged during electrolysis.

Example

Electrolysis of molten (fused) lead(II) bromide using inert graphite electrodes

The ions present in the molten lead(II) bromide are $Pb^{2+}(l)$ and $Br^-(l)$.

At the anode:

The Br^- ions move towards the anode where they are discharged, i.e. they lose electrons to the anode forming bromine atoms. Pairs of bromine atoms immediately bond covalently forming bromine molecules:

$$2Br^-(l) \longrightarrow Br_2(g) + 2e^-$$

Brown fumes of **bromine** are evolved at the anode.

At the cathode:

The Pb^{2+} ions move towards the cathode where they are discharged, i.e. they gain electrons to form lead atoms:

$$Pb^{2+}(l) + 2e^- \longrightarrow Pb(l)$$

Molten **lead** forms around the cathode and drips off.

Practical activity

To electrolyse molten lead(II) bromide using graphite electrodes

(Teacher demonstration)

Your teacher may use this activity to assess:

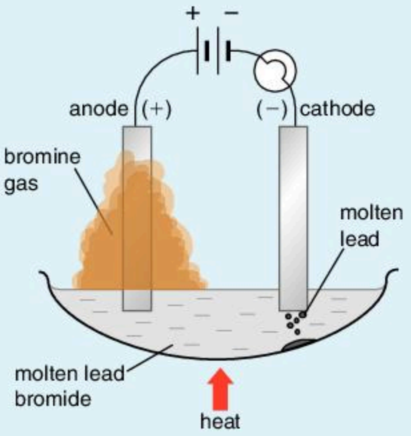
- observation, recording and reporting.

Your teacher will perform the following experiment.

The experiment must be performed in a fume cupboard.

Method

1. Place some lead(II) bromide into a crucible.
2. Heat the lead(II) bromide until it melts.
3. Connect the circuit as shown in Figure 10.3.3.
4. Observe the product formed at each electrode.
5. Write an ionic half equation for the reaction occurring at each electrode.



▲ **Figure 10.3.3** Electrolysis of molten lead(II) bromide

Electrolysis of aqueous solutions

During the electrolysis of an aqueous solution, the ions present are not only those of the solute. Because water undergoes spontaneous ionisation, an aqueous solution always contains H^+ ions and OH^- ions from the ionisation of water molecules as well as the ions from the solute. As a result, aqueous solutions always contain at least *two* different cations and *two* different anions.

In the electrolysis of an aqueous solution, one type of ion of each charge will be discharged in preference to the other. This is known as **preferential discharge**.

Preferential discharge of anions

There are *three* main factors which influence the preferential discharge of the anions that you will be studying:

The position of the ion in the electrochemical series

Considering the electrochemical series of anions given in Table 10.3.1, ions at the top of the series are the most stable and the hardest to discharge. Ions at the bottom of the series are the least stable and the easiest to discharge, i.e. the ease of discharge increases downwards.

The *lower* the ion in the electrochemical series, the more likely it is to be discharged.

This is seen, for example, when we look at the electrolysis of dilute sulfuric acid, dilute sodium chloride solution and copper(II) sulfate solution using inert electrodes; the OH^- ion is preferentially discharged from each solution.

▼ **Table 10.3.1** The electrochemical series of anions

Anion
F^-
SO_4^{2-}
NO_3^-
Cl^-
Br^-
I^-
OH^-

The concentration of the electrolyte

The concentration of the electrolyte has an impact on which ions are preferentially discharged. There is a tendency for the *more concentrated* ion to be preferentially discharged. However, this rule really only applies to solutions containing *halide ions*, i.e. Cl^- ions, Br^- ions and I^- ions.

This is seen when we compare the electrolysis of dilute sodium chloride solution with the electrolysis of concentrated sodium chloride solution, both using inert electrodes. In the dilute solution, the OH^- ion is preferentially discharged but in the concentrated solution, the Cl^- ion is preferentially discharged.

The type of electrode

The type of anode chosen for the electrolysis can affect the reaction occurring at the anode. If the anode is *inert*, for example graphite or platinum, it does not affect the reaction occurring. However, if the anode is *active*, for example copper, it can take part in the electrolysis process and this affects what happens at the anode.

This is seen when we compare the electrolysis of copper(II) sulfate solution using an inert anode with its electrolysis using an active copper anode. When an inert anode is used, the OH^- ion is preferentially discharged. With an active copper anode, the anode ionises.

Preferential discharge of cations

There is *one* main factor which influences the preferential discharge of the cations that you will be studying:

▼ **Table 10.3.2** The electrochemical series of cations

Cation
K^+
Ca^{2+}
Na^+
Mg^{2+}
Al^{3+}
Zn^{2+}
Fe^{2+}
Pb^{2+}
H^+
Cu^{2+}
Ag^+

The position of the ion in the electrochemical series

Considering the electrochemical series of cations given in Table 10.3.2, ions at the top of the series are the most stable and the hardest to discharge. Ions at the bottom of the series are the least stable and the easiest to discharge, i.e. the ease of discharge increases downwards.

The *lower* the ion in the electrochemical series, the more likely it is to be discharged.

This is seen, for example, when we look at the electrolysis of dilute and concentrated sodium chloride solution; the H^+ ion is preferentially discharged in both. Also when we look at the electrolysis of copper(II) sulfate solution; the Cu^{2+} ion is preferentially discharged. We will now look at the examples referred to above in greater detail.

Examples

1 Electrolysis of dilute sulfuric acid using inert electrodes

Ions present:

From the H_2SO_4: $H^+(aq)$ and $SO_4^{2-}(aq)$.

From the H_2O: $H^+(aq)$ and $OH^-(aq)$.

At the anode:

The SO_4^{2-} ions and the OH^- ions move towards the anode. The OH^- ions are preferentially discharged because they are lower in the electrochemical series of anions:

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

Effervescence occurs as oxygen gas is evolved at the anode.

At the cathode:

The H^+ ions move towards the cathode where they are discharged:

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

Effervescence occurs as **hydrogen** gas is evolved at the cathode.

Relative volumes of gases produced:

Comparing the two equations, for every 4 mol of electrons transferred from the anode to the cathode, 1 mol of oxygen and 2 mol of hydrogen are produced:

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

$$4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$$

1 volume of oxygen is produced for every *2* volumes of hydrogen.

Changes in the electrolyte:

The sulfuric acid becomes *more concentrated* because H^+ ions and OH^- ions are discharged from it, i.e. water is being removed from the electrolyte. Because of the discharge of the H^+ ions and OH^- ions, the electrolysis of sulfuric acid is referred to as the **electrolysis of water**.

Practical activity

To electrolyse dilute sulfuric acid using inert electrodes

(Teacher demonstration)

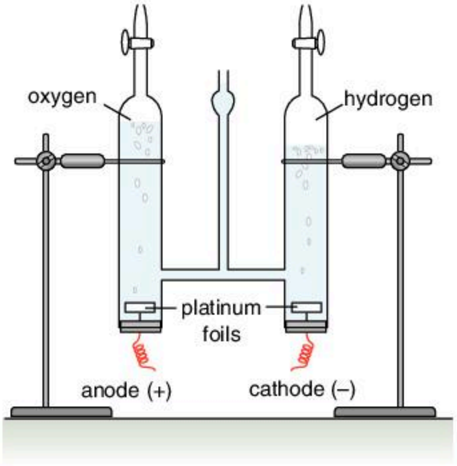
Your teacher may use this activity to assess:

- observation, recording and reporting
- analysis and interpretation.

Your teacher will perform the following experiment.

Method

1. Set up the Hoffmann's voltameter as shown in Figure 10.3.4, using dilute sulfuric acid as the electrolyte.
2. Observe the relative volume of gases given off at each electrode.
3. Test the gas given off at the anode to see if it is oxygen by using a glowing splint. The splint should glow brighter or re-light if the gas is oxygen.
4. Test the gas given off at the cathode to see if it is hydrogen by using a burning splint. The splint should make a squeaky pop and be extinguished if the gas is hydrogen.
5. Explain:
   - **a** why oxygen was produced at the anode (include an equation)
   - **b** why hydrogen was produced at the cathode (include an equation)
   - **c** why the volume of hydrogen was greater than the volume of oxygen.
6. What would you expect to happen to the electrolyte as the reaction proceeds?



▲ **Figure 10.3.4** Hoffmann's voltameter showing the electrolysis of dilute sulfuric acid

Did you know?

Oxygen gas causes the glowing splint to glow brighter or relight because it supports the combustion, i.e. burning, of the splint.

2 Electrolysis of dilute sodium chloride solution using inert electrodes

Ions present:

From the NaCl: $Na^+(aq)$ and $Cl^-(aq)$.

From the H_2O: $H^+(aq)$ and $OH^-(aq)$.

At the anode:

The Cl^- ions and the OH^- ions move towards the anode. The OH^- ions are preferentially discharged because they are in a dilute solution and are lower in the electrochemical series of anions:

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

Effervescence occurs as **oxygen** gas is evolved at the anode.

At the cathode:

The Na^+ ions and the H^+ ions move towards the cathode. The H^+ ions are preferentially discharged because they are lower in the electrochemical series of cations:

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

Effervescence occurs as **hydrogen** gas is evolved at the cathode.

Relative volumes of gases produced:

Comparing the two equations, for every 4 mol of electrons transferred from the anode to the cathode, 1 mol of oxygen and 2 mol of hydrogen are produced, i.e. *1* volume of oxygen is produced for every *2* volumes of hydrogen.

Changes in the electrolyte:

The sodium chloride solution becomes *more concentrated* because H^+ ions and OH^- ions are discharged from it.

3 Electrolysis of concentrated sodium chloride solution (brine) using inert electrodes

Ions present:

From the NaCl: $Na^+(aq)$ and $Cl^-(aq)$.

From the H_2O: $H^+(aq)$ and $OH^-(aq)$.

At the anode:

The Cl^- ions and the OH^- ions move towards the anode. The Cl^- ions are preferentially discharged because they are halide ions in a concentrated solution:

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

Effervescence occurs as *yellow-green* **chlorine** gas is evolved at the anode.

At the cathode:

The Na^+ ions and the H^+ ions move towards the cathode. The H^+ ions are preferentially discharged because they are lower in the electrochemical series of cations:

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

Effervescence occurs as **hydrogen** gas is evolved at the cathode.

Relative volumes of gases produced:

Comparing the two equations, for every 2 mol of electrons transferred from the anode to the cathode, 1 mol of chlorine and 1 mol of hydrogen are produced, i.e. *equal* volumes of chlorine and hydrogen are produced.

Changes in the electrolyte:

The electrolyte becomes *alkaline* because H^+ ions and Cl^- ions are discharged from it leaving an excess of Na^+ ions and OH^- ions, which form sodium hydroxide.

Practical activity

To investigate the effect of the concentration of the electrolyte on the preferential discharge of ions

(Teacher demonstration)

Your teacher may use this activity to assess:

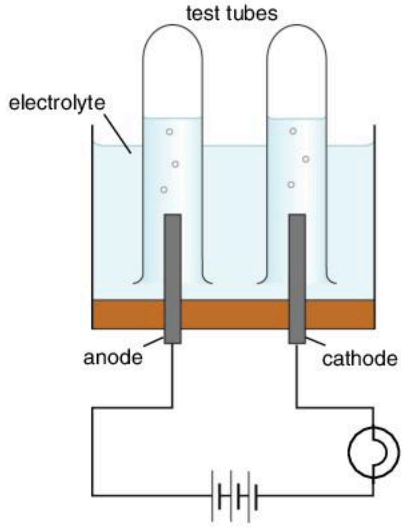
- observation, recording and reporting
- analysis and interpretation.

Your teacher will perform the following experiment.

Method

1. Set up an electrolytic cell as shown in Figure 10.3.5, using inert graphite electrodes and dilute sodium chloride solution as the electrolyte.
2. Place an inverted test tube filled with dilute sodium chloride solution over each electrode to collect the gas evolved at each.
3. Test the gas in the tube over the anode for oxygen using a glowing splint.
4. Test the gas in the tube over the cathode for hydrogen using a burning splint.
5. Repeat the experiment using concentrated sodium chloride solution.
6. Test the gas in the tube over the anode for chlorine using a piece of moist blue litmus paper. If the gas is chlorine the paper should turn red and then white.
7. Test the gas in the tube over the cathode for hydrogen using a burning splint.
8. Dip a piece of red litmus paper into the electrolyte.
9. How were your observations in the two experiments:
   a. similar
   b. different?
10. Which ions were preferentially discharged at the anode:
    a. in the dilute sodium chloride solution
    b. in the concentrated sodium chloride solution?

    In each case, explain why and include the relevant equation.
11. Which ions were preferentially discharged at the cathode in both experiments? Give a reason and a relevant equation.
12. Explain why the litmus paper changed colour when dipped into the electrolyte in step 8.



▲ **Figure 10.3.5** Electrolysing concentrated sodium chloride solution (brine) in the laboratory

Did you know?

Chlorine gas turns moist blue litmus paper red and then white because it reacts with the water to form hydrochloric acid (HCl) and chloric(I) acid (HClO):

$$Cl_2(g) + H_2O(l) \longrightarrow HCl(aq) + HClO(aq)$$

The acids turn litmus red and the chloric(I) acid oxidises it to colourless.

4 Electrolysis of copper(II) sulfate solution using inert electrodes

Ions present:

From the $CuSO_4$: $Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$.

From the H_2O: $H^+(aq)$ and $OH^-(aq)$.

At the anode:

The SO_4^{2-} ions and the OH^- ions move towards the anode. The OH^- ions are preferentially discharged because they are lower in the electrochemical series of anions:

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

Effervescence occurs as **oxygen** gas is evolved at the anode.

At the cathode:

The Cu^{2+} ions and the H^+ ions move towards the cathode. The Cu^{2+} ions are preferentially discharged because they are lower in the electrochemical series of cations:

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

A *pink* deposit of **copper** builds up around the cathode and the cathode increases in size.

Changes in the electrolyte:

The electrolyte becomes *paler blue* because the blue Cu^{2+} ions are discharged from it. It also becomes *acidic* because Cu^{2+} ions and OH^- ions are discharged leaving an excess of H^+ ions and SO_4^{2-} ions, which form sulfuric acid.

5 Electrolysis of copper(II) sulfate solution using active copper electrodes

Ions present:

From the $CuSO_4$: $Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$.

From the H_2O: $H^+(aq)$ and $OH^-(aq)$.

At the anode:

The SO_4^{2-} ions and the OH^- ions move towards the anode. However, the anode, being made of copper, is active and the copper atoms ionise, forming Cu^{2+} ions, this requires less energy than discharging the OH^- ions.

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

The Cu^{2+} ions go into the electrolyte and the anode gradually *decreases* in size.

At the cathode:

The Cu^{2+} ions and the H^+ ions move towards the cathode. The Cu^{2+} ions are preferentially discharged because they are lower in the electrochemical series of cations:

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

A *pink* deposit of **copper** builds up around the cathode and the cathode increases in size.

Exam tip

When you are writing equations for the reactions occurring at the anode and cathode, remember to always include the correct **state symbols**.

Changes in the electrolyte:

The electrolyte remains *unchanged*. For every 2 mol of electrons transferred from anode to cathode, 1 mol of Cu^{2+} ions enter the electrolyte at the anode and 1 mol of Cu^{2+} ions is discharged at the cathode.

Practical activity

To investigate the effect of using inert and active electrodes on the electrolysis of copper(II) sulfate solution

(Teacher demonstration)

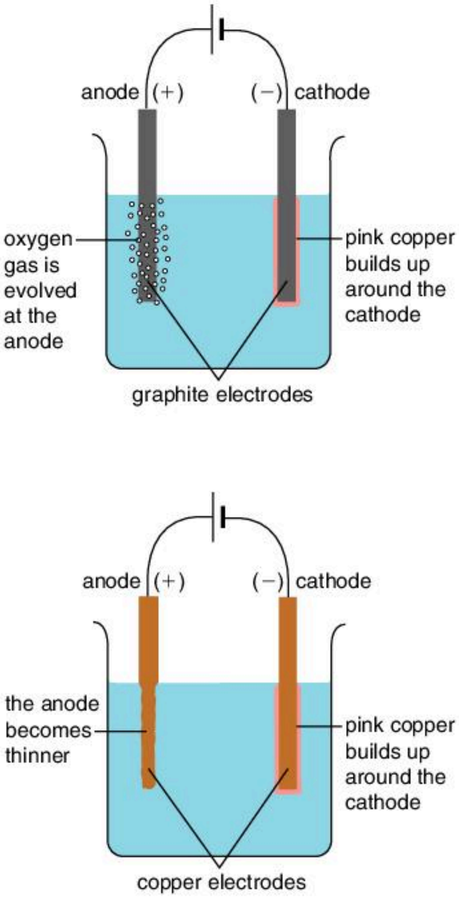
Your teacher may use this activity to assess:

- observation, recording and reporting
- analysis and interpretation.

Your teacher will perform the following experiment.

Method

1. Set up an electrolytic cell as shown in Figure 10.3.6, using graphite electrodes and copper(II) sulfate solution as the electrolyte.
2. Observe the anode and the cathode, taking note of what occurs at each electrode. Also observe the colour of the electrolyte and dip a piece of blue litmus paper into the electrolyte.
3. Repeat the experiment using copper electrodes.
4. Observe the anode and the cathode and the colour of the electrolyte.
5. Classify the electrodes as inert or active.
6. How were your observations in the two experiments:
   - a similar
   - b different?
7. Why was the reaction occurring at the anode different in the two experiments? Include the relevant equations.
8. Why did the litmus paper change colour in step **2**?
9. Why was there a change in colour of the electrolyte in the first experiment but not in the second?



▲ **Figure 10.3.6** Electrolysis of copper(II) sulfate solution using graphite and copper electrodes

Summary questions

1. Define the term electrolysis.
2. Give the formula and state symbol of the ions present in each of the following:
   - a molten sodium chloride
   - b dilute sulfuric acid
   - c copper(II) sulfate solution
   - d potassium bromide solution
3. Draw a labelled diagram of the apparatus used in the laboratory for the electrolysis of copper(II) sulfate solution using inert graphite electrodes.
4. Describe the differences in the electrolysis of dilute sodium chloride solution and the electrolysis of concentrated sodium chloride solution using inert graphite electrodes.
5. By means of an equation in each case, show the reaction occurring at the anode when copper(II) sulfate solution is electrolysed using:
   - a a graphite anode
   - b a copper anode.

Exam tip

It is important that you can draw line diagrams to show **electrolytic cells**.

Objectives

By the end of this topic you will be able to:

- define the Faraday constant
- calculate the volumes of gases and masses of substances produced during electrolysis.

Did you know?

The electrical charge on one electron is **1.6×10^{-19} C**. This means that 1 C of electrical charge is equivalent to $\frac{1}{1.6 \times 10^{-19}}$ electrons.

A10.4 Quantitative electrolysis

The volumes of gases and masses of substances produced during electrolysis can be calculated using the mole concept that you studied in Unit 7. Michael Faraday (1791–1867) was the first person to propose that the mass of a substance produced at an electrode during electrolysis is directly proportional to the **quantity of electricity** passing through the electrolytic cell.

The quantity of electricity, or quantity of electrical charge, is given the symbol $\boldsymbol{Q}$ and is measured in units called coulombs, **C**. During electrolysis, electrons flow from the anode to the cathode through electrical wires. Each electron flowing through these wires has an electrical charge which is extremely small. In fact, the electrical charge on one electron is so small that when 1 C of electrical charge flows through a circuit, 6.25×10^{18} electrons have flowed.

The quantity of electricity or electrical charge flowing through an electrolytic cell during electrolysis depends on *two* factors.

- The rate of flow of the electrical charge, i.e. the current, which is given the symbol $\boldsymbol{I}$ and is measured in units called amperes (also known as amps), **A**. A current of 1 amp is equivalent to 1 C, or 6.25×10^{18} electrons, flowing for 1 s.
- The length of time, $\boldsymbol{t}$, that the current flows for, measured in seconds, **s**.

The quantity of electricity or electrical charge can then be calculated using the following formula:

$$\text{quantity of electrical charge (C)} = \text{current (A)} \times \text{time (s)}$$

i.e. $$Q = I \times t$$

Example

Calculate the quantity of electrical charge that flows when a current of 2.0 A flows in a circuit for 20 min.

Current = 2.0 A

Time in seconds = $20 \times 60 = 1200$ s

$\therefore$ quantity of electrical charge flowing = 2×1200 C = **2400 C**

Key fact

The **Faraday constant** is the size of the electrical charge on one mole of electrons, i.e. **96 500 C mol^{-1}**.

We can relate this to **moles** if we consider the electrical charge which 1 mol of electrons has. You will recall that 1 mol of electrons is equivalent to 6.0×10^{23} electrons. The charge on one electron is 1.6×10^{-19} C, therefore, if we have 1 mol of electrons, these electrons would have a total charge of 96 500 C. This value is known as the **Faraday constant**.

During electrolysis, a cation with a single positive charge, i.e. M^+, is discharged at the cathode according to the following equation:

$$M^+ + e^- \longrightarrow M$$

This shows that 1 mol of electrons is required to discharge 1 mol of M^+ ions and form 1 mol of M or, applying the Faraday constant, 96 500 C is required to discharge 1 mol of M^+ ions and form 1 mol of M. This also applies when discharging 1 mole of anions with a single negative charge at the anode, i.e. 96 500 C is required.

The following examples will show how we can use the Faraday constant to calculate the volumes of gases and masses of substances formed during electrolysis. You will recall the following quantities that relate moles to mass and volumes of gases.

Molar mass, M = mass of 1 mol of a substance.

Molar volume, V_m = volume of 1 mol of a gas.

At stp, $V_m = 22.4\,dm^3$, at rtp, $V_m = 24.0\,dm^3$.

Examples

1 Calculate the mass of lead produced at the cathode when a current of 5.0 A flows through molten lead(II) bromide for 16 minutes and 5 seconds.

Determine the quantity of electricity that flows:

Current = 5.0 A

Time in seconds = (16 × 60) + 5 = 965 s

$\therefore$ quantity of electricity = (5.0 × 965) C = **4825 C**

Write the equation for the reaction at the cathode:

$$Pb^{2+}(l) + 2e^- \longrightarrow Pb(l)$$

From the equation:

2 mol of electrons are required to form 1 mol Pb.

$\therefore$ 2 × 96 500 C form 1 mol Pb

i.e. 193 000 C form 1 mol Pb

$\therefore$ 1 C forms $\frac{1}{193\,000}$ mol Pb

$$\therefore\ 4825\text{ C forms } \frac{1}{193\,000} \times 4825 \text{ mol Pb} = \mathbf{0.025\ mol\ Pb}$$

$M(Pb) = 207\,g\,mol^{-1}$

i.e. mass of 1 mol Pb = 207 g

$\therefore$ mass of 0.025 mol Pb = 0.025 × 207 g = **5.175 g**

2 If an electric current of 2.5 A is passed through dilute sulfuric acid for 2 hours 34 minutes and 24 seconds, calculate the volume of oxygen produced at the anode at rtp.

Determine the quantity of electricity that flows:

Current = 2.5 A

Time in seconds = (2 × 60 × 60) + (34 × 60) + 24 = 9264 s

$\therefore$ quantity of electricity = (2.5 × 9264) C = **23 160 C**

Equation for the reaction at the anode:

$$4OH^-(aq) \longrightarrow 2H_2O(aq) + O_2(g) + 4e^-$$

From the equation:

4 mol electrons are lost in forming 1 mol O_2.

$\therefore$ $4 \times 96\,500$ C form 1 mol O_2

i.e. $386\,000$ C form 1 mol O_2

$\therefore$ 1 C forms $\frac{1}{386\,000}$ mol O_2

$\therefore$ $23\,160$ C forms $\frac{1}{386\,000} \times 23\,160$ mol O_2

$= \mathbf{0.06\,mol\,O_2}$

Volume of 1 mol O_2 at rtp $= 24.0\,\text{dm}^3$

$\therefore$ volume of 0.06 mol $O_2 = 0.06 \times 24.0\,\text{dm}^3 = \mathbf{1.44\,dm^3}$

3 A solution of copper(II) sulfate is electrolysed using copper electrodes. Determine the decrease in mass of the anode if a current of 4.0 A passes through the solution for 32 minutes and 10 seconds.

Determine the quantity of electricity that flows:

Current = 4.0 A

Time in seconds $= (32 \times 60) + 10 = 1930$ s

$\therefore$ quantity of electricity $= 4.0 \times 1930\,\text{C} = \mathbf{7720\,C}$

Equation for the reaction at the anode:

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

From the equation:

2 mol electrons are lost in forming 1 mol Cu^{2+} ions.

$\therefore$ $2 \times 96\,500$ C form 1 mol Cu^{2+} ions

i.e. $193\,000$ C form 1 mol Cu^{2+} ions

$\therefore$ 1 C forms $\frac{1}{193\,000}$ mol Cu^{2+} ions

$\therefore$ 7720 C forms $\frac{1}{193\,000} \times 7720$ mol Cu^{2+} ions

$= \mathbf{0.04\,mol\,Cu^{21}\,ions}$

$M(Cu^{2+}) = 64\,\text{g mol}^{-1}$

i.e. mass of 1 mol Cu^{2+} ions = 64 g

$\therefore$ mass of 0.04 mol Cu^{2+} ions $= 0.04 \times 64\,\text{g} = 2.56\,\text{g}$

Therefore, decrease in mass of the anode = **2.56 g**.

4 A steel spoon needs to be electroplated with 13 g of chromium. Calculate the length of time that a current of 15 A would have to flow to electroplate the spoon when placed as the cathode of the electrolytic cell.

Calculate the number of moles of chromium which need to be produced:

$M(Cr) = 52\,\text{g mol}^{-1}$

i.e. mass of 1 mol Cr = 52 g

$\therefore$ number of moles in 13.0 g $= \frac{13}{52}$ mol $= 0.25$ mol

Equation for the reaction at the cathode:

$$Cr^{3+}(aq) + 3e^{-} \longrightarrow Cr(s)$$

From the equation:

3 mol electrons form 1 mol Cr.

$\therefore$ 3 × 96 500 C form 1 mol Cr

i.e. 289 500 C form 1 mol Cr.

$\therefore$ 0.25 × 289 500 C form 0.25 mol Cr = 72 375 C

Therefore, the quantity of electricity required = **72 375 C**

Calculate the time taken for 72 375 C to flow using a current of 15 A:

quantity of electricity (C) = current (A) × time (s)

$$72\,375\text{ C} = 15\text{ A} \times \text{time (s)}$$

$$\therefore \quad \text{time} = \frac{72\,375}{15}\text{s}$$

$$= \mathbf{4825\ s}$$

$$= 1\text{ hour }20\text{ minutes }25\text{ seconds}$$

It would take **1 hour 20 minutes 25 seconds** to electroplate the spoon.

Summary questions

1 Define the Faraday constant.

2 Calculate the mass of magnesium that would be deposited on the cathode when a current of 5.0 A flows through molten magnesium chloride for 2 hours 40 minutes and 50 seconds.

3 Determine the volume of chlorine that would be produced at the anode at rtp when a current of 2.5 A flows through concentrated sodium chloride solution for 51 minutes and 28 seconds.

4 How long must a current of 10.0 A flow through a solution of copper(II) sulfate to produce 25.6 g of copper at the cathode?

A10.5 Industrial applications of electrolysis

Electrolysis is used in many ways in industry including:

- extracting metals from their ores
- purifying metals, i.e. electrorefining
- plating one metal with another, i.e. electroplating
- coating a metal with its oxide, i.e. anodising.

We will now look at each of these in more detail.

Extraction of metals

Very few metals are found in their free state in the Earth's crust. Most are found bonded to other elements in ionic compounds known as minerals. Minerals from which metals can be extracted are known as ores. The process

Objectives

By the end of this topic you will be able to:

- describe how electrolysis can be used to extract metals from their ores
- describe how electrolysis can be used to purify metals
- describe the process of electroplating
- describe the process of anodising.

of extracting metals from their ores involves converting the metal ions to atoms, i.e. it is a reduction process:

$$M^{n+} + ne^- \longrightarrow M$$

Metals high in the electrochemical series ionise easily and form very stable ions. This means that their ions are very difficult to reduce to atoms. For this reason these metals need a powerful method of reduction to extract them from their ores. **Electrolysis of the molten ore** is a powerful method of reduction and is used to extract aluminium and metals above aluminium in the electrochemical series from their ores.

During electrolysis of the molten ore, the metal ions move towards the cathode where they are discharged.

An example is the extraction of aluminium from its molten ore, aluminium oxide (Al_2O_3), using inert graphite electrodes.

The aluminium ions move towards the cathode where they are discharged to form molten aluminium:

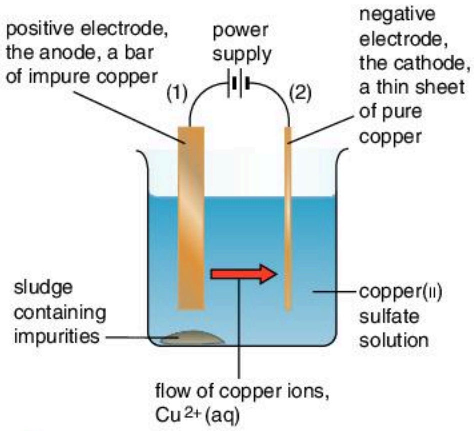
$$Al^{3+}(l) + 3e^- \longrightarrow Al(l)$$

The molten aluminium which forms is denser than the molten electrolyte and it sinks to the bottom of the electrolytic cell and is tapped off. This will be discussed in more detail in Unit 18.2.

Purification of metals

Electrolysis is very useful in the production of pure metals from impure samples of the metal. This process is known as **electrorefining**. The principles of electrorefining are as follows.

- The impure metal is made the anode, i.e. it is connected to the positive terminal of the battery.
- The cathode is a very thin sample of the pure metal.
- The electrolyte is an aqueous solution which contains ions of the metal being purified.
- The anode, being active, ionises, and ions of the metal enter the electrolyte where they join the other metal ions.
- The metal ions are discharged at the cathode where they form the pure metal which builds up around the cathode.
- The electrolyte remains unchanged since the metal ions entering it at the anode are removed as they are discharged at the cathode.
- Any insoluble impurities present in the anode fall to the bottom of the cell forming a sludge which can be removed.



▲ **Figure 10.5.1** The electrorefining of copper

An example of electrorefining is the **purification of copper**. This is illustrated in Figure 10.5.1. The impure copper is the anode (1) and the cathode is a thin sheet of pure copper (2). The electrolyte contains Cu^{2+} ions, i.e. copper(II) sulfate solution.

The copper anode ionises:

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

The Cu^{2+} ions enter the electrolyte and are discharged at the cathode, forming pure copper:

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

Electrorefining is only suitable for purifying metals whose ions are *below hydrogen* in the electrochemical series. Any metal ions above hydrogen in the series would not be discharged at the cathode since the hydrogen ions, being lower in the series, would be preferentially discharged, forming hydrogen gas.

Electroplating

Electroplating involves depositing a thin layer of one metal on top of another metal. Electroplating is carried out to either protect the original metal from corrosion, to improve its appearance or to make a cheap metal object appear more valuable. The object that is being electroplated can be made of any metal, but it is usually made of brass, copper or steel. The principles of electroplating are very similar to those of electrorefining.

- The object to be electroplated is made the cathode, i.e. it is connected to the negative terminal of the battery.
- The anode is a pure sample of the metal which is being used for plating.
- The electrolyte is an aqueous solution which contains ions of the metal being used for plating.
- The anode, being active, ionises, and ions of the metal enter the electrolyte where they join the other metal ions.
- The metal ions are discharged at the cathode where they form the pure metal which builds up as a coating around the object.
- The electrolyte remains unchanged since the metal ions entering it at the anode are removed as they are discharged at the cathode.

As with electrorefining, only metals whose ions are *below hydrogen* in the electrochemical series can be used for plating.

Examples of electroplating are chromium plating, silver plating, gold plating and nickel plating.

Chromium plating, or chrome plating, gives steel objects a shiny, silver look as well as protecting them from corrosion. The steel object is placed as the cathode, the anode is pure chromium and the electrolyte is usually a solution of chromium(III) sulfate [$Cr_2(SO_4)_3$].

▲ **Figure 10.5.2** Chrome plated taps

The chromium anode ionises:

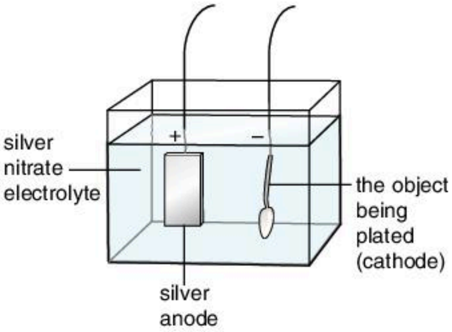
$$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^-$$

The Cr^{3+} ions enter the electrolyte and are discharged at the cathode forming a layer of chromium on the object:

$$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$$

The electrolytes used in **nickel plating** and **silver plating** are usually solutions of nickel(II) sulfate ($NiSO_4$) and silver nitrate ($AgNO_3$). During the process of silver plating illustrated in Figure 10.5.3, the silver anode ionises:

$$Ag(s) \longrightarrow Ag^+(aq) + e^-$$



▲ **Figure 10.5.3** Electroplating with silver

The Ag^+ ions enter the electrolyte and are discharged at the cathode forming a layer of silver on the object:

$$Ag^+(aq) + e^- \longrightarrow Ag(s)$$

Anodising

On exposure to air, aluminium immediately reacts with the oxygen to form a layer of aluminium oxide (Al_2O_3) on its surface. This layer is relatively unreactive and it adheres firmly to the metal below and does not flake off. As a result it protects the metal from attack by oxygen and moisture in the air.

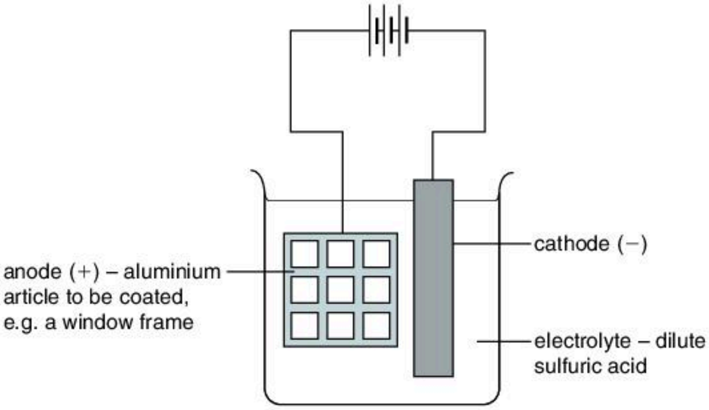
The thickness of the aluminium oxide layer on the surface of aluminium objects can be artificially increased by a process known as **anodising**. This makes the objects, such as window frames and saucepans, even more resistant to corrosion. The anodising process involves the following steps.

- The aluminium article to be anodised is pre-treated using sodium hydroxide solution to clean and etch the surface.
- The cleaned aluminium article is made the anode, i.e. it is connected to the positive terminal of the battery.
- The electrolyte is usually a dilute solution of sulfuric acid.
- As soon as the electrolytic cell is connected up, the aluminium anode begins to ionise, forming Al^{3+} ions, and at the same time the SO_4^{2-} ions and the OH^- ions in the electrolyte move towards the anode. The OH^- ions react with the Al^{3+} ions forming a layer of aluminium oxide on the surface of the aluminium:

$$Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$$

$$2Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al_2O_3(s) + 3H^+(aq)$$

The silver-grey aluminium oxide coating readily absorbs dyes so it can be attractively coloured.



▲ **Figure 10.5.4** The anodising of aluminium

Summary questions

1 Explain why electrolysis is used to extract metals high in the electrochemical series from their ores.

2 Describe how you would purify a sample of silver. Include the equations for the reactions occurring at the anode and the cathode.

3 Give the equations for the reactions that occur at the anode and cathode during:

a chromium plating

b nickel plating.

4 What is anodising?

Key concepts

- The **electrochemical series of metals** is a list of metals in order of how easily their atoms lose electrons.
- The higher the metal is in the electrochemical series, the more easily it ionises and the stronger are its reducing powers.
- A metal can displace a metal below it in the electrochemical series from a compound containing the lower metal.
- Metals above hydrogen in the electrochemical series displace the hydrogen ions from an acid, forming hydrogen gas; metals below hydrogen will not displace the hydrogen ions from an acid.
- The **electrochemical series of non-metals** is a list of non-metals in order of how easily their atoms gain electrons.
- The higher the non-metal is in the electrochemical series, the more easily it ionises and the stronger are its oxidising powers.
- A non-metal can displace a non-metal below it in the electrochemical series from a compound containing the lower non-metal.
- **Conductors** allow electricity to pass through. **Non-conductors** do not allow electricity to pass through.
- **Electrolytes** are compounds that form ions when molten or in aqueous solution.
- Electrolytes conduct electricity because they contain **mobile ions**.
- When electricity passes through an electrolyte, the electrolyte decomposes.
- Metals conduct electricity because they contain **mobile electrons**.
- When electricity passes through a metal, the metal remains unchanged.
- **Strong electrolytes** are fully ionised; they contain a high concentration of ions. **Weak electrolytes** are partially ionised; they contain a low concentration of ions.
- Pure water is a very weak electrolyte due to the very low concentration of H^+ and OH^- ions.
- **Electrolysis** is the chemical change which occurs when an electric current is passed through an electrolyte.
- **Electrodes** are used during electrolysis to conduct electricity in to and out of the electrolyte.
- The **anode** is the positive electrode. Anions move towards the anode and lose electrons. Oxidation occurs at the anode.
- The **cathode** is the negative electrode. Cations move towards the cathode and gain electrons. Reduction occurs at the cathode.
- Molten electrolytes contain one type of cation and one type of anion. Both are discharged during electrolysis.
- Aqueous electrolytes contain solute ions as well as H^+ ions and OH^- ions from the water. One type of ion of each charge will be discharged in preference to the other.
- Factors which influence the **preferential discharge** of anions are the position of the anion in the electrochemical series, the concentration of the aqueous solution and whether the anode is active or inert.
- The position of the cation in the electrochemical series determines the preferential discharge of cations.

- When dilute sulfuric acid is electrolysed using inert electrodes, oxygen is produced at the anode and hydrogen at the cathode.
- When sodium chloride solution is electrolysed using inert electrodes, if the solution is dilute, oxygen is produced at the anode, if it is concentrated, chlorine is produced at the anode. Hydrogen is produced at the cathode in both cases.
- When copper(II) sulfate solution is electrolysed, if the anode is inert, oxygen is produced at it, if the anode is made of copper, it ionises. Copper is produced at the cathode in both cases.
- The volume of gas or mass of a substance produced at an electrode during electrolysis is directly proportional to the quantity of electricity passing through the electrolytic cell.
- The **quantity of electricity**, Q, is measured in coulombs, C.
- The quantity of electricity is dependent on the rate of flow of electrical charge, i.e. the current, I, and the time the current flows for, t, i.e. $Q = I \times t$.
- The **Faraday constant** is the size of the electrical charge on one mole of electrons, i.e. $96\,500\,\mathrm{C\,mol^{-1}}$.
- Industrial uses of electrolysis include the extraction of metals from their ores, the purification of metals, electroplating and anodising.

Practice exam-style questions

Multiple-choice questions

1 Metal X was found to displace metal Y from a solution containing the sulfate of Y. However, X did not displace metal Z from a solution containing Z sulfate. Which of the following is correct?

A X is higher in the electrochemical series than Z.
B Z is a stronger reducing agent than X.
C Y is higher in the electrochemical series than X.
D Y is a stronger reducing agent than Z.

2 Which of the following lists consists of conductors?

A sodium chloride, magnesium, graphite
B hydrochloric acid, ethanol, aluminium
C aqueous sodium hydroxide, copper, ethanoic acid
D kerosene, molten potassium chloride, tap water

3 During electrolysis, the reaction which takes place at the anode is:

A displacement
B reduction
C oxidation
D decomposition

4 If a current of 5.0 A flows for 16 minutes and 5 seconds through a silver nitrate solution, what mass of silver would be deposited at the anode?

A 2.7 g
B 5.4 g
C 10.8 g
D 27.0 g

5 What quantity of electricity is needed to produce 0.75 mol of chlorine during the electrolysis of molten sodium chloride?

A 72 375 C
B 96 500 C
C 144 750 C
D 193 000 C

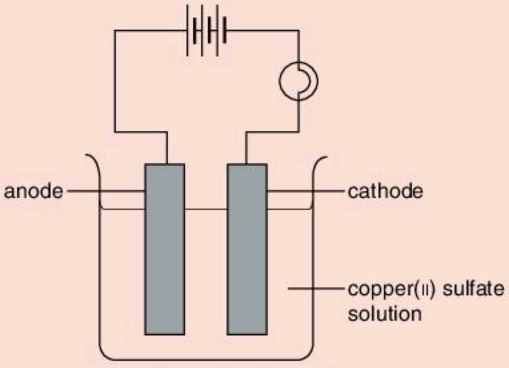
6 Which of the following statements is incorrect?

A Iron can be purified by electrolysis.
B Anodising is used to protect aluminium items from corrosion.
C Electroplating can be used to coat steel items with a thin layer of silver.
D Magnesium can be extracted from its ore by electrolysis.

Structured question

7 a Define the term 'electrolysis'. ***(1 mark)***

b An electrolytic cell was set up, as shown in the figure, with copper(II) sulfate solution as the electrolyte, a graphite anode and a copper cathode.



Electrolytic cell

i) State, giving a reason, what you would observe at the anode. Include a relevant equation. ***(3 marks)***
ii) State, giving a reason, what you would observe in the electrolyte. *(2 marks)*
iii) If a current of 5.0 A flows for 2 hours 8 minutes and 40 seconds through the electrolytic cell, determine the increase in mass of the cathode. *(4 marks)*
iv) Give TWO differences that you would observe in the electrolytic cell if the copper electrode was made the anode and the graphite electrode was made the cathode. *(2 marks)*

c i) Give TWO differences between electroplating and anodising. *(2 marks)*
ii) Give ONE use of anodising. *(1 mark)*

Total 15 marks

Extended response question

8 a Electrolysis is a chemical process. Use the electrolysis of molten potassium bromide to show how electrolysis is a chemical process. Support your answer with relevant ionic equations. *(4 marks)*

b Using the electrolysis of dilute and concentrated sodium chloride solutions with inert electrodes, discuss the effect of EACH of the following on the products of electrolysis.

i) The position of ions in the electrochemical series. *(2 marks)*
ii) The concentration of the electrolyte. *(2 marks)*

c The process of electroplating can be used to coat metal items with a layer of another metal to make them more resistant to corrosion, improve their appearance and make them appear more valuable. Cutlery made of steel is often electroplated with a layer of nickel followed by a layer of silver.

i) Draw a labelled diagram to show the apparatus you could use to electroplate a steel spoon with nickel using nickel sulfate solution ($NiSO_4(aq)$) as the electrolyte. Clearly indicate what you would use as the anode and the cathode. *(3 marks)*
ii) How long must a steady current of 4.0 A flow through the electrolyte to coat the spoon with 9.44 g of nickel? *(4 marks)*

Total 15 marks

A11 Rates of reaction

Objectives

By the end of this topic you will be able to:

- define what is meant by rate of reaction
- explain how the rate of a reaction can be measured
- explain the collision theory for reactions
- interpret rate curves for reactions.

Chemical reactions proceed at different speeds. Some reactions can take years, such as the rusting of cars, while others happen in less than a second, such as an explosion. Other reactions take place at varying speeds between these two extremes. Knowing the speed at which reactions occur is extremely important in industry to maximise productivity while minimising costs and the risk of accidents.

A11.1 Measuring rates of reaction

The study of the speed at which a chemical reaction occurs, known as the **rate of reaction**, is called **chemical kinetics**. During a reaction, the concentration of the reactants decreases as the concentration of products increases. If we can measure the decrease in concentration of a reactant or the increase in concentration of a product, and know how long it took for the decrease or increase to occur, we can determine the rate of a reaction.

The rate of a reaction can, therefore, be determined by:

$$\text{rate of reaction} = \frac{\text{decrease in concentration of a reactant}}{\text{time taken for the decrease}}$$

or $$\text{rate of reaction} = \frac{\text{increase in concentration of a product}}{\text{time taken for the increase}}.$$

! Key fact

The **rate of reaction** is a measured change in the concentration of a reactant or product with time at a stated temperature.

Changes in concentration are not always easy to measure. However, there are other *property changes*, which depend on the nature of the reactants and products, that can be measured to determine the rate of a reaction. Some of the changes that can be measured are:

- if one of the products is a gas, the volume of gas produced over time can be measured using a gas syringe, e.g. a reaction between a reactive metal and an acid
- if one of the products is a gas, if the gas is allowed to escape, the decrease in mass of the reaction can be measured over time, e.g. a reaction between a carbonate and an acid
- if a precipitate forms, the appearance of the precipitate can be measured, e.g. the reaction between sodium thiosulfate and hydrochloric acid
- if there is a change in colour intensity, this can be measured, e.g. the reaction between potassium iodide and acidified hydrogen peroxide
- if there is a distinct change in pressure, temperature or pH as the reaction proceeds, any of these can be measured over time.

Collision theory for reactions

Chemical reactions involve the breaking of original bonds in the reactants and the formation of new bonds in the products. In order for a reaction to take place, all of the following must occur:

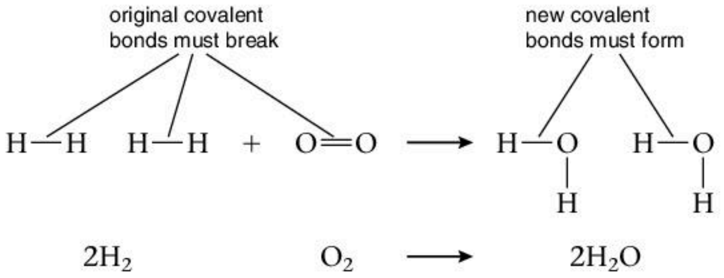
- the reactant particles must *collide* with each other in order to break the original bonds in the reactants so that new bonds can form in the products
- the reactant particles must collide with *enough energy*, known as **activation energy** (see Unit 12.1), to break the original bonds in the reactants and enable the particles to have enough energy to form products (see Unit 12.1)
- the reactant particles must collide with the *correct orientation* so that the energy produced by the collision can be passed on to the original bonds in order to break them.

Not all collisions occur with the required activation energy or with the correct orientation of particles. Therefore, not all collisions result in a reaction occurring. Collisions which result in a reaction are called **effective collisions**. The rate of a reaction depends on the number of effective collisions.

For example, in the reaction between hydrogen and oxygen to form water:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

The hydrogen (H_2) and oxygen (O_2) molecules must collide with each other to break the covalent bonds between the atoms within the molecules. This will then allow new covalent bonds to form between the hydrogen and oxygen atoms, forming water molecules.



▲ **Figure 11.1.1** The reaction between hydrogen and oxygen

To be an effective collision, each collision must generate enough activation energy to break the covalent bonds within the hydrogen and oxygen molecules and the molecules must be correctly orientated to each other.

Rate curves

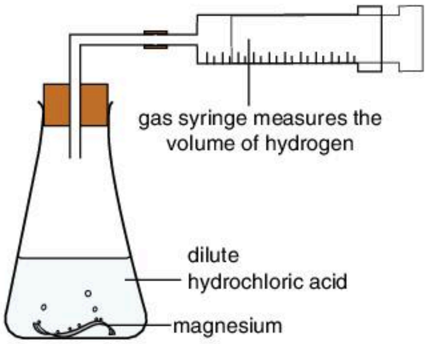
If a measured property is plotted on a graph against time as the reaction proceeds and the property changes, a **rate curve** is obtained.

Examples

1 The reaction between magnesium and hydrochloric acid:

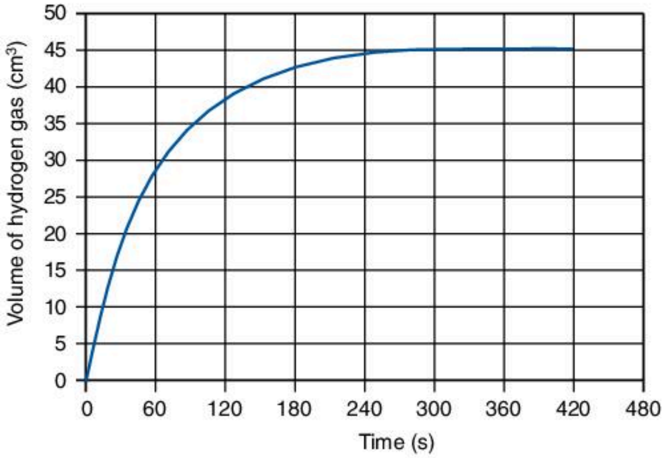
$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

If magnesium ribbon is added to hydrochloric acid, the volume of hydrogen gas given off can be measured in a gas syringe at regular intervals using the apparatus shown in Figure 11.1.2. If the volume of gas is then plotted against time, the curve in Figure 11.1.3 may be obtained.



▲ **Figure 11.1.2** Investigating the rate of reaction between magnesium and hydrochloric acid.

▶ **Figure 11.1.3** Graph of volume of hydrogen gas released against time



We can use the values on the graph to determine the average rate of reaction at different stages in the reaction by using the general formula:

$$\text{rate of reaction} = \frac{\text{increase in volume of hydrogen}}{\text{time taken for the increase}}$$

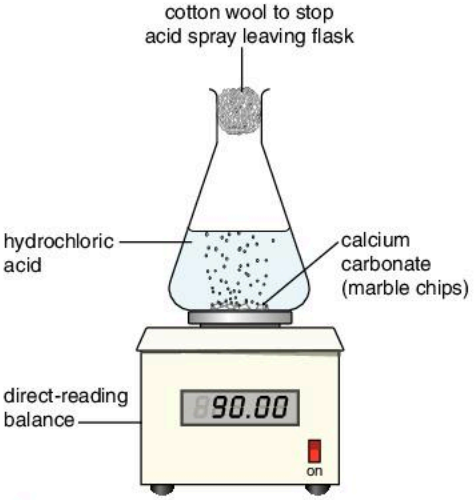
- To determine the average rate of the reaction in the first minute (60 s):
  volume of gas produced in the first minute = **28 cm³**
  $\therefore$ average rate of reaction in the first minute $= \frac{28\,\text{cm}^3}{60\,\text{s}} = \mathbf{0.47\,cm^3\,s^{-1}}$
- To determine the average rate of the reaction in the second minute:
  Volume of gas produced in the second minute $= (38 - 28)\,\text{cm}^3 = \mathbf{10\,cm^3}$
  $\therefore$ average rate of reaction in the second minute $= \frac{10\,\text{cm}^3}{60\,\text{s}}$
  $= \mathbf{0.17\,cm^3\,s^{-1}}$

This clearly shows that as the reaction proceeds, the rate of the reaction *decreases*. This can also be seen by looking at the gradient or slope of the graph. The gradient is steep at the beginning of the graph indicating the reaction is rapid. However, as the reaction progresses, the gradient becomes shallower indicating the reaction is slowing down. We will discuss this in more detail in the next section.

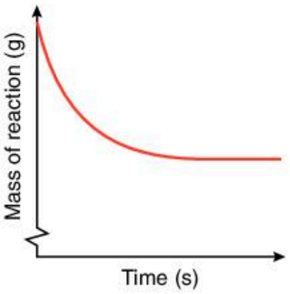
2 The reaction between calcium carbonate and hydrochloric acid:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

If some hydrochloric acid is poured into a conical flask, which is on a balance, and calcium carbonate is added, the mass of the reaction can be measured at regular intervals. If the mass is then plotted against time, the curve in Figure 11.1.5 may be obtained.



▲ **Figure 11.1.4** Investigating the rate of reaction between calcium carbonate and hydrochloric acid



▲ **Figure 11.1.5** Graph of mass of reaction against time

Looking at the gradient of the graph we can also clearly see that as the reaction proceeds, the rate of the reaction decreases.

Explanation of the shape of rate curves

We can see that both rate curves have a very similar shape. The gradient is steep at the beginning of the reaction and gradually becomes shallower until the curve eventually becomes horizontal. This shows that the rate of a reaction changes as the reaction proceeds. We will now look at the reasons for this.

- The gradient is steepest at the beginning of the reaction indicating that the rate of the reaction is at its *highest*. This is because the concentration of the reactant particles is at its highest at the beginning of the reaction. The frequency of collision between the particles is at its highest at the beginning.
- As the reaction proceeds, the gradient becomes shallower indicating that the rate of the reaction is *decreasing*. This is because the concentration of the reactant particles is decreasing as the reaction proceeds. The frequency of collision between the particles is decreasing as the reaction proceeds.
- After a period of time the curve becomes horizontal, i.e. the gradient becomes zero, indicating that the reaction has reached completion and has *stopped*. One of the reactants has been used up and there are no more of its particles left to collide. The reactant which is used up in a reaction is known as the **limiting reactant**. The quantity of the limiting reactant determines the quantity of the products produced.

Practical activity

To investigate the change in rate of a reaction as the reaction proceeds

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation.

You will be supplied with 4.0 g of calcium carbonate crystals, 0.3 mol dm^{-3} hydrochloric acid, a conical flask, a cork with a delivery tube running through it to a gas syringe, a measuring cylinder and a stopwatch.

Method

1. Place the calcium carbonate crystals in the conical flask.
2. Measure 30 cm^3 of hydrochloric acid in the measuring cylinder.
3. As quickly as possible pour the acid into the conical flask and cork the conical flask with the cork which has a delivery tube running to the gas syringe, as in Figure 11.1.2. At the same time, start the stopwatch.
4. Record the volume of gas in the gas syringe every 15 seconds until the volume remains constant for three consecutive readings.
5. Record your results in a table and use them to plot a graph of volume of gas against time. Draw a curve of best fit between the points.
6. Using your graph, calculate the average rate of the reaction for:
   a the first minute **b** the second minute.
7. Explain why the average rate of the reaction is higher in the first minute than in the second minute.
8. Explain why the curve of your graph eventually becomes horizontal.

Exam tip

When drawing graphs using data obtained from experiments, it is important that you draw a curve or straight line of best fit. You do not necessarily join all the points. A curve or line of **best fit** must be smooth or straight and must pass between the points so that it is as close as possible to all the points.

Summary questions

1 Define what is meant by rate of reaction.
2 Give two ways of calculating the rate of a reaction.
3 Give two methods that a student could use to determine how the rate of reaction between magnesium ribbon and dilute hydrochloric acid changes with time.
4 What does the collision theory for a reaction state?
5 Explain how the rate of a reaction changes as the reaction proceeds.

Objectives

By the end of this topic you will be able to:

- identify the factors which affect the rates of reaction
- explain how different factors affect the rates of reaction
- describe experiments to determine the effect of changing a factor on the rate of a reaction
- interpret graphical representations of data obtained in studying rates of reaction.

? Did you know?

Photosynthesis is possibly one of the most important reactions in the world. It is the process by which green plants make food by combining carbon dioxide and water. The plants use the energy from sunlight to bring about the reaction. Without photosynthesis, animals would have no food to eat and both plants and animals would have no oxygen for respiration.

A11.2 Factors affecting rates of reaction

The rate of a reaction is dependent on several factors. There are four main factors that we will be studying which can affect the rate of a reaction. These are:

- concentration
- temperature
- surface area (particle size)
- presence or absence of a catalyst.

Light and **pressure** also affect the rate of certain reactions. For example, the reaction between methane and chlorine (see Unit 14.1) and photosynthesis are both initiated by light, and as the light intensity increases, the rate of the reaction increases. Chemical reactions which are initiated or speeded up by light are known as **photochemical reactions**.

Pressure affects reactions between reactants in the gaseous state. If the pressure on the gaseous reactants is increased, their volume decreases, meaning that there are more reactant particles per unit volume. As a result, the particles collide more frequently and this increases the chances of effective collisions, thereby *increasing* the rate of the reaction. This is seen in the reaction between nitrogen and hydrogen to form ammonia:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

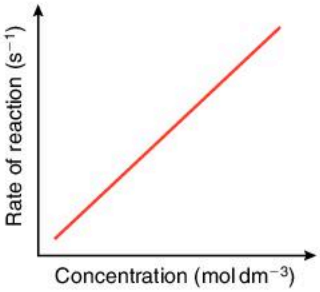
We will now look in detail at how each of the other *four* factors affects the rate of a reaction. When considering a particular factor, it is assumed that all other factors are kept constant.

Concentration

The general rule is that the *higher* the concentration of a reactant, the *higher* the rate of the reaction.

This applies particularly to solutions. If the concentration of a reactant is increased, this means that there are more reactant particles per unit volume. As a result, the particles collide more frequently which increases the chances of effective collisions, therefore, the rate of the reaction increases.

If we measure the rate of a reaction at different concentrations of one of the reactants, and plot the rate against concentration, we get a graph similar to the one in Figure 11.2.1. The graph clearly shows that as the concentration increases the rate increases. It also shows that, for this reaction, the rate of the reaction is *directly proportional* to the concentration.



▶ **Figure 11.2.1** Rate of reaction against concentration

Practical activity

To investigate the effect of concentration on the rate of a reaction

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation.

You will be supplied with 0.005 mol dm^{-3} potassium iodide solution, 0.01 mol dm^{-3} sodium thiosulfate solution, a solution of acidified hydrogen peroxide and starch with a hydrogen peroxide concentration of 0.085 mol dm^{-3}, two 25 cm^3 measuring cylinders, one 10 cm^3 measuring cylinder, a burette, a small beaker and a stopwatch.

Method

1. Place the sodium thiosulfate solution into the burette.
2. Measure 25 cm^3 of potassium iodide solution in a measuring cylinder and pour it into the beaker. Add 1 cm^3 of sodium thiosulfate solution from the burette and swirl to mix the solutions.
3. Measure 25 cm^3 of acidified hydrogen peroxide and starch mixture in another measuring cylinder.
4. Add the mixture to the beaker and at the same time start the stopwatch. Swirl the beaker and place it on a piece of white paper.
5. Observe the solution in the beaker from above and as soon as a blue-black colour begins to appear stop the stopwatch.
6. Record the time taken for the blue-black colour to start to appear.
7. Repeat the experiment four more times using the volumes of potassium iodide solution and water given in the table below.
8. Calculate the concentration of potassium iodide solution for each experiment using the following formula:

$$\text{concentration of KI(aq)} = \frac{\text{volume of KI(aq)}}{25} \times 0.005\,\text{mol dm}^{-3}$$

9. Calculate the rate of reaction for each experiment using the following formula:

$$\text{rate of reaction} = \frac{1}{\text{time taken for blue-black colour to appear}}\,\text{s}^{-1}$$

10. Record your results in a table similar to the one outlined below and plot a graph to show the rate of the reaction against the concentration of the potassium iodide solution. Draw a straight line of best fit between the points.

Volume of potassium iodide solution/cm^3	Volume of distilled water/cm^3	Concentration of potassium iodide solution/mol dm^{-3}	Volume of sodium thiosulfate solution/cm^3	Volume of acidified hydrogen peroxide and starch mixture/cm^3	Time taken for blue/black colour to appear/s	Rate of reaction/ s^{-1}
25	0		1	25		
20	5		1	25		
15	10		1	25		
10	15		1	25		
5	20		1	25		

11 What can you deduce about the effect of concentration on the rate of the reaction?

12 Explain the effect that increasing the concentration has on the rate of the reaction.

13 Use your graph to determine the rate of the reaction at potassium iodide concentrations of 0.0022 mol dm^{-3} and 0.0044 mol dm^{-3}. In theory, if the concentration of a reactant doubles, the rate of a reaction should also double. Did the rate of your reaction double? If it didn't double, can you suggest reasons why? Use the information given below to help you.

During the experiment, the hydrogen peroxide oxidises the I$^-$ ions in the potassium iodide to iodine molecules (I_2):

$$2I^-(aq) \longrightarrow I_2(aq) + 2e^-$$

As soon as an iodine molecule is formed the sodium thiosulfate reduces it back to I$^-$ ions:

$$I_2(aq) + 2e^- \longrightarrow 2I^-(aq)$$

When all the sodium thiosulfate has reacted, iodine molecules remain in the solution and react with the starch to produce a *blue-black* colour. Since the concentration and volume of sodium thiosulfate solution used each time is the same, the blue-black colour starts to appear when the same quantity of iodine has been made each time. In your experiment you determined the time taken for the *same quantity* of iodine to be produced using different concentrations of potassium iodide solution.

Temperature

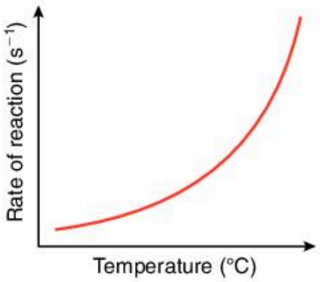
The general rule is that the *higher* the temperature at which a reaction occurs, the *higher* the rate of the reaction.

If the temperature at which a reaction occurs is increased, the reactant particles gain **kinetic energy** which has the following effects:

- the particles move faster, therefore they collide more frequently
- the particles collide with more energy, therefore more collisions occur with sufficient activation energy for the particles to react.

Combining these two effects, as the temperature increases, the chances of effective collisions increase, therefore the rate of the reaction increases. For some chemical reactions which occur at room temperature, if the temperature increases by 10 °C, the rate of the reaction approximately *doubles*.

If we measure the rate of a reaction at different temperatures and plot the rate against temperature we get a graph similar to the one shown in Figure 11.2.2. The graph clearly shows that as the temperature increases, the rate of the reaction increases, but it shows that the rate is not directly proportional to temperature.



▲ **Figure 11.2.2** Rate of reaction against temperature

Practical activity

To investigate the effect of temperature on the rate of a reaction

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation

You will be supplied with 1.0 mol dm^{-3} hydrochloric acid, 0.05 mol dm^{-3} sodium thiosulfate solution ($Na_2S_2O_3(aq)$), a small conical flask, a 50 cm^3 measuring cylinder, a burette, a thermometer and a stopwatch.

Method

1. Using a black pen, draw a cross on a piece of white paper. Do not make the cross too dark.
2. Measure 50 cm^3 of sodium thiosulfate solution using the measuring cylinder and pour it into the conical flask.
3. Pour the hydrochloric acid into the burette.
4. Add 1 cm^3 of hydrochloric acid to the conical flask and at the same time start the stopwatch. Swirl the flask to mix the solutions and place the flask on top of the cross.
5. When the two reactants are mixed, a precipitate of sulfur slowly forms which gradually obscures the cross. Look down into the conical flask and stop the stopwatch as soon as the cross is no longer visible.
6. Immediately measure and record the temperature of the mixture in the conical flask and record the time taken for the cross to disappear.
7. Repeat the experiment, but this time heat the sodium thiosulfate solution in the conical flask to about 40 °C before adding the hydrochloric acid.
8. Repeat the experiment three more times, heating the sodium thiosulfate solution to about 50 °C, 60 °C and then 70 °C.
9. Record your results in a table and calculate the rate of the reaction at each temperature using the following formula:
$$\text{rate of reaction} = \frac{1}{\text{time taken for cross to disappear}}\ \text{s}^{-1}$$
10. Plot a graph to show the rate of the reaction against the final temperature of the mixture. Draw a curve of best fit between the points.
11. What can you deduce about the effect of temperature on the rate of the reaction?
12. Explain the effect that increasing the temperature has on the rate of the reaction.
13. Use your graph to determine the rate of the reaction at 40 °C and 50 °C. In theory, the rate of a reaction should approximately double if the temperature is increased by 10 °C. Did the rate of your reaction double? If it didn't double, can you suggest reasons why? Use the information given below to help you.
14. Given that the products of the reaction are sodium chloride, sulfur, water and sulfur dioxide, write a balanced chemical equation for the reaction.

During the experiment, the sodium thiosulfate reacted with the hydrochloric acid to form insoluble sulfur, which slowly built up and obscured the cross:

$$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + S(s) + H_2O(l) + SO_2(g)$$

Since the volume and concentration of the two solutions used was the same each time, the cross disappeared when the same quantity of sulfur had been made each time. In your experiment you were determining the time taken for the *same quantity* of sulfur to be produced at different temperatures.

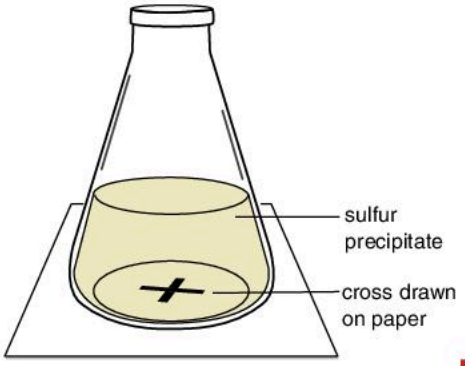


Figure 11.2.3 Following the reaction between sodium thiosulfate solution and hydrochloric acid

Surface area (particle size)

The general rule is that the *smaller* the particles of a reactant, the *higher* the rate of a reaction.

This only applies to reactants in the solid state. When solids react, the reaction occurs on the surface of the solid. Small solid particles have a larger total **surface area** than the same mass of large solid particles. If the particle size is decreased, the smaller particles have a greater surface area exposed to particles of the other reactant. As a result, the particles collide more frequently which increases the chances of effective collisions, therefore, the rate of the reaction increases.

In flour mills and coal mines, the finely divided flour and coal are highly flammable. Any spark, e.g. from a cigarette, can initiate a reaction with the oxygen in the air which could be explosive because of the large surface area of the fine particles of flour and coal.

Practical activity

To investigate the effect of surface area on the rate of a reaction

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation.

You will be supplied with 5.0 g of small calcium carbonate crystals, 5.0 g of large calcium carbonate crystals, 0.3 mol dm^{-3} hydrochloric acid, a conical flask, a cork with a delivery tube running through it to a gas syringe, a measuring cylinder and a stopwatch.

Method

1. Place the small calcium carbonate crystals in the conical flask.
2. Measure 30 cm^3 of hydrochloric acid in the measuring cylinder.
3. As quickly as possible pour the acid into the conical flask and cork the conical flask with the cork which has a delivery tube running to the gas syringe, as in Figure 11.1.2. At the same time, start the stopwatch.
4. Record the volume of gas in the gas syringe every 30 seconds until the volume remains constant for three consecutive readings.
5. Repeat the experiment using the large crystals.
6. Record your results in a table and use them to plot a graph of volume of gas against time for each set of crystals. You can plot both curves on the same set of axes.
7. Using your graph, calculate the average rate of the reaction for:
   a. the small crystals
   b. the large crystals.
8. Explain why the average rate of the reaction is higher for the small crystals.
9. Explain why both sets of crystals produced almost the same volume of gas.

Presence or absence of a catalyst

Certain chemical substances can be added to reactions to alter the rate of the reaction. These substances are known as catalysts and they are used extensively in industry to increase the rate of reactions. Catalysts are chemically unchanged at the end of the reaction.

Key fact

A **catalyst** is a substance which alters the rate of a chemical reaction without itself undergoing any permanent chemical change.

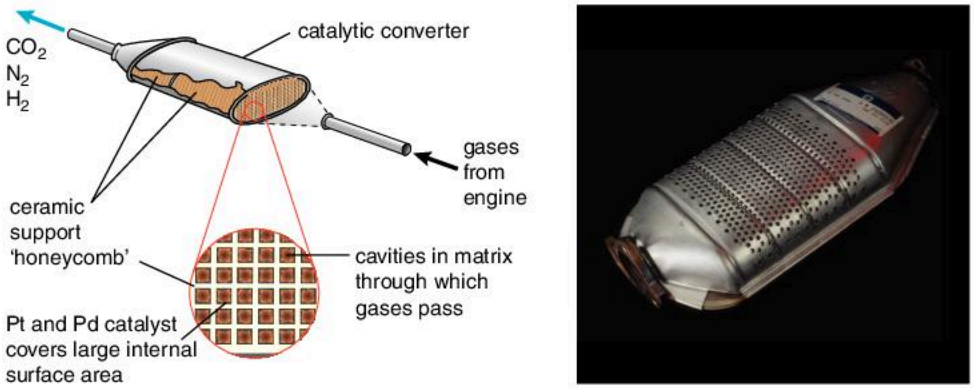
Most catalysts are added to *speed up* the rate of a reaction. They work by providing an alternative pathway for the reaction which requires *less* activation energy than the normal pathway (see Unit 12.1). As a result, more of the collisions occur with sufficient energy for the particles to react, which increases the number of effective collisions in a given length of time.

For example, hydrogen peroxide (H_2O_2) decomposes very slowly at room temperature into water and oxygen:

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

The reaction can be speeded up considerably by adding manganese(IV) oxide (MnO_2). The manganese(IV) oxide acts as a catalyst and it is chemically unchanged at the end of the reaction. When the reaction has finished, the mass of manganese(IV) oxide is exactly the same as it was at the beginning of the reaction.

Enzymes are biological catalysts made of protein molecules. They are present in the cells of all living organisms where they speed up chemical reactions that would otherwise occur too slowly for life to exist.



▲ **Figure 11.2.4** A catalytic converter in a car exhaust allows harmful pollutant gases to react together forming harmless products

Did you know?

Tetraethyl lead(IV) was first mixed with gasoline in the 1920s as an inexpensive way to increase engine performance and fuel economy by preventing uncontrolled combustion known as 'engine knocking'. Hence it became known as 'antiknock'. Because of the toxic nature of lead compounds, the use of leaded fuel has been phased out almost worldwide. This phasing out started in the USA in the mid-1970s.

A few catalysts are added to *slow down* the rate of a reaction. These are known as negative catalysts or inhibitors. They work by providing an alternative pathway for the reaction which requires *more* activation energy than the normal pathway. This decreases the number of effective collisions in a given length of time.

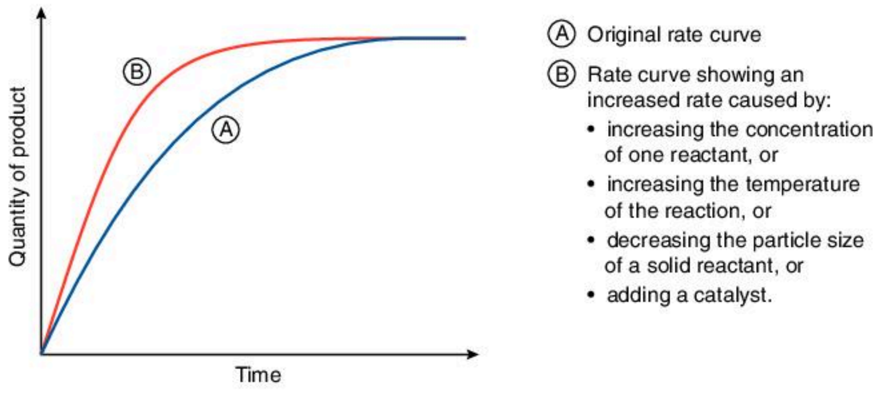
An example of a negative catalyst is tetraethyl lead(IV) ($(C_2H_5)_4Pb$), which used to be added to gasoline (petrol) for use in internal combustion engines. It prevented uncontrolled combustion of the gasoline so that the engine ran more smoothly.

Exam tip

If you are asked to explain how a particular factor affects the rate of a chemical reaction, your answer must be based on the **collision theory** for reactions.

Effect of changing factors on rate curves

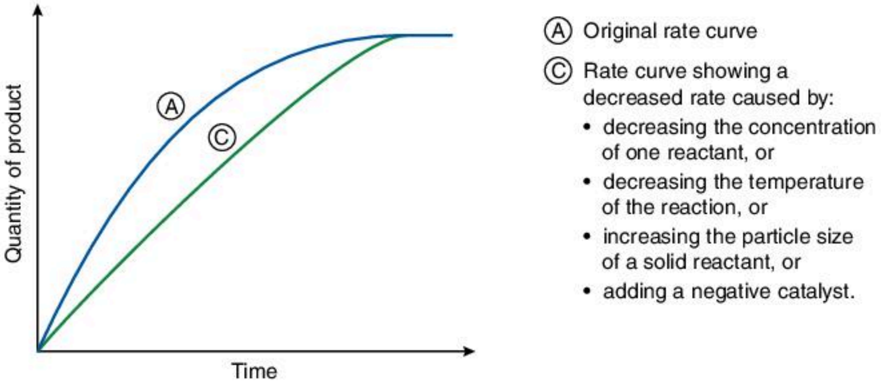
The effect of changing any factor which alters the rate of a reaction can be shown on the rate curve for the reaction. In each of the graphs in Figures 11.2.5 and 11.2.6, the number of moles of the limiting reactant remains unchanged. Remember that the limiting reactant is the one which is used up in a reaction and determines the quantity of products produced.



▲ **Figure 11.2.5** A rate curve showing the effect of any factor which increases the rate of a reaction

In Figure 11.2.5, curve Ⓐ is the original curve and curve Ⓑ shows the effect of changing any one single factor as indicated. Looking carefully at what the graph shows:

- curve Ⓑ has a *steeper gradient* – this is because the reaction is occurring at a higher rate
- curve Ⓑ *becomes horizontal sooner* – this is because the reaction takes less time to reach completion
- both curves become horizontal when *the same quantity of product has been made* – this is because the original number of moles of the limiting reactant was not changed.



▲ **Figure 11.2.6** A rate curve showing the effect of any factor which decreases the rate of a reaction

In Figure 11.2.6, curve Ⓐ is the original curve and curve Ⓒ shows the effect of changing any one single factor as indicated. Looking carefully at what the graph shows:

- curve Ⓒ has *a shallower gradient* – this is because the reaction is occurring at a slower rate
- curve Ⓒ becomes *horizontal later* – this is because the reaction takes more time to reach completion
- both curves become horizontal when *the same quantity of product has been made* – this is because the original number of moles of the limiting reactant was not changed.

Exam tip

It is extremely important that you can interpret **graphs** which show rates of reaction.

Summary questions

1 Explain how each of the following increases the rate of a reaction:
   a increasing the temperature
   b decreasing the particle size of a reactant
   c increasing the concentration of a reactant.

2 **a** What is a catalyst?
   b Explain how a catalyst alters the rate of a reaction.

3 Describe an experiment you could perform to determine the effect of particle size on the rate of a reaction.

Key concepts

- The **rate of reaction** is a measured change in the concentration of a reactant or product with time at a stated temperature.
- The collision theory states that to react, particles of the reactants must collide, the collisions must occur with the required **activation energy** and the particles must be correctly orientated.
- An **effective collision** is one that results in the formation of the product.
- A **rate curve** shows a measured property plotted against time.
- All rate curves have the same basic shape.
- The rate of a reaction changes as the reaction proceeds. It is fastest at the beginning and the rate decreases with time until the reaction reaches completion when the **limiting reactant** is used up.
- Four main factors affect the rate of a chemical reaction: concentration, temperature, surface area (particle size) and the presence or absence of a catalyst.
- The higher the concentration of a reactant, the higher the rate of a reaction.
- The higher the temperature, the higher the rate of a reaction.
- The smaller the particle size, the higher the rate of a reaction.
- A **catalyst** is a substance which alters the rate of a chemical reaction without itself undergoing any permanent chemical change.
- Most catalysts, if present, speed up the rate of a reaction.
- A negative catalyst, if present, slows down the rate of a reaction.

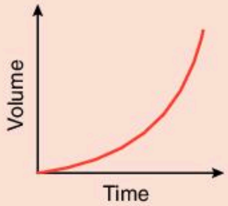
Practice exam-style questions

Multiple-choice questions

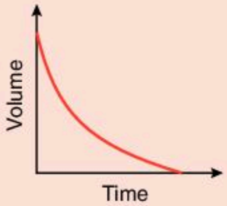
1 Which of the following would not result in an increase in the rate of reaction between 25 cm^3 of 0.1 mol dm^{-3} sulfuric acid and excess magnesium carbonate crystals at 28 °C?

A Increasing the temperature of the acid to 38 °C.
B Increasing the volume of the acid to 40 cm^3.
C Using the same mass of magnesium carbonate powder.
D Increasing the concentration of the acid to 0.2 mol dm^{-3}.

2 An aqueous solution of hydrogen peroxide was decomposed and the oxygen produced was collected in a gas syringe. Which of the following graphs represents the total volume of oxygen as measured at various time intervals?



A



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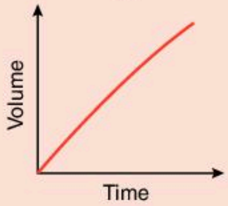

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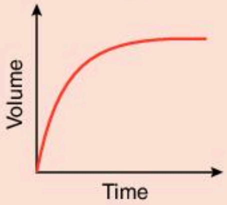


D

3 A catalyst is best described as a substance which:

A speeds up the rate of a chemical reaction
B slows down the rate of a chemical reaction
C takes no part in a chemical reaction
D alters the rate of a chemical reaction

4 The rate of a chemical reaction is:

A independent of the size of the particles of a solid reactant
B independent of the concentration of the reactants
C dependent on the temperature
D unaffected by a catalyst

5 If the temperature at which hydrochloric acid reacts with sodium thiosulfate is raised by 10 °C, which of the following occurs?

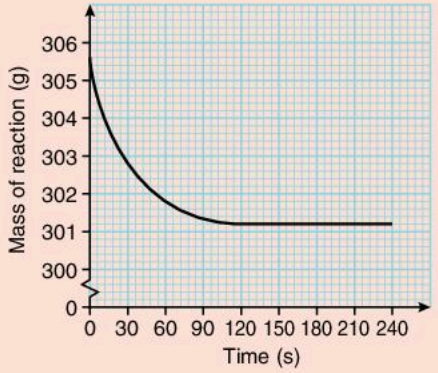
A The rate of the reaction is increased by a factor of 10.
B The rate of the reaction is decreased by a factor of 10.
C The rate of the reaction is doubled.
D The rate of the reaction is halved.

6 A mass of 5.0 g of powdered zinc will react the slowest with:

A 30 cm^3 of 0.1 mol dm^{-3} sulfuric acid
B 20 cm^3 of 0.2 mol dm^{-3} sulfuric acid
C 20 cm^3 of 0.3 mol dm^{-3} sulfuric acid
D 10 cm^3 of 0.4 mol dm^{-3} sulfuric acid

Structured question

7 a An experiment was set up to investigate what effect changing various factors has on the rate of reaction between calcium carbonate and hydrochloric acid. The reaction was carried out at 30 °C in a conical flask that was placed on a balance. A mass of 10.0 g of calcium carbonate was placed in the flask and excess hydrochloric acid of concentration 1.0 mol dm^{-3} was added. The flask and its contents were weighed at regular intervals and their mass plotted against time, as shown in the figure.



Mass of reaction against time

i) The rate of a reaction decreases as the reaction proceeds. Using the data from the graph, determine the average rate of the reaction in g s^{-1}:
– in the first minute
– in the second minute. *(4 marks)*

ii) Why is the rate faster in the first minute than the second minute? *(2 marks)*

iii) What effect would repeating the reaction at 40 °C have on the rate of the reaction? *(1 mark)*

iv) By referring to the collision theory for reactions, explain your answer to part **iii)** above. *(3 marks)*

v) What effect, if any, would using excess hydrochloric acid of concentration 0.5 mol dm^{-3} have on the rate of the reaction? Give a reason for your answer. *(2 marks)*

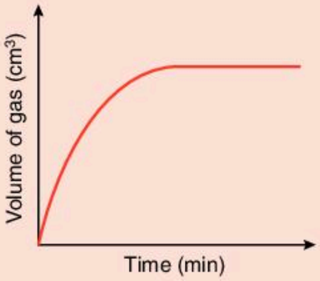
b **i)** What is a catalyst? *(1 mark)*

ii) Explain how a catalyst works. *(2 marks)*

Total 15 marks

Extended response question

8 The graph below shows the results of an experiment which was carried out to determine how the rate of a reaction between magnesium ribbon and hydrochloric acid varies over time. In the experiment, excess magnesium ribbon was added to 25 cm^3 of hydrochloric acid at 35 °C and the volume of gas was measured at regular intervals.



a **i)** Write a balanced equation for the reaction. *(2 marks)*

ii) Account for the shape of the curve. *(4 marks)*

b Draw a labelled diagram to show how the apparatus could be set up to carry out the experiment. *(3 marks)*

c **i)** Copy the graph onto your answer paper and on your graph draw a second curve to show the expected results if the experiment was repeated at 25 °C. Make sure that you distinguish between your curves. *(2 marks)*

ii) If the experiment was repeated again with magnesium powder instead of ribbon at 35 °C, what effect, if any, would this have on:

– the maximum volume of gas evolved

– the rate of evolution of the gas?

Explain your answer in each case. *(4 marks)*

Total 15 marks

A12 Energetics

Objectives

By the end of this topic you will be able to:

- distinguish between an exothermic reaction and an endothermic reaction
- give examples of exothermic and endothermic reactions
- explain exothermic and endothermic reactions based on bond breaking and bond forming
- explain what is meant by enthalpy change, ΔH
- explain exothermic and endothermic reactions based on enthalpy change
- draw energy profile diagrams for exothermic and endothermic reactions
- show the effect of adding a catalyst using energy profile diagrams.

Chemical energetics is the branch of chemistry that deals with energy changes during physical processes and chemical reactions. Energy changes usually accompany all chemical reactions. In fact, in some chemical reactions the energy produced is more important than the products of the chemical reaction, e.g. the burning of fuels. Energy changes can take many forms, e.g. changes in heat, light and electricity.

A12.1 Energy changes during reactions

Exothermic and endothermic reactions

Whenever there is a chemical reaction, there tends to be a change in energy of the system. The energy changes in chemical reactions are usually in the form of heat energy, however, they may also be in the form of light, electrical or nuclear energy. We can distinguish between two types of reaction based on the energy changes which occur.

- If the reaction produces heat, causing the reaction mixture and its surroundings to get hotter, the reaction is said to be an **exothermic reaction**. Exothermic reactions transfer energy to their surroundings.
- If the reaction absorbs heat, causing the causing the reaction mixture and its surroundings to get colder, the reaction is said to be an **endothermic reaction**. Endothermic reactions absorb energy from their surroundings.

Key fact

An **exothermic reaction** releases energy to its surroundings, causing the surroundings to get hotter.

Key fact

An **endothermic reaction** absorbs energy from its surroundings, causing the surroundings to get colder.

Practical activity

To investigate exothermic and endothermic reactions

Your teacher may use this activity to assess:

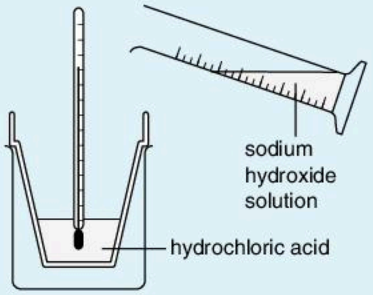
- observation, recording and reporting
- manipulation and measurement.

You will be supplied with 2.0 mol dm^{-3} hydrochloric acid, 2.0 mol dm^{-3} sodium hydroxide solution, zinc metal, potassium nitrate, sodium hydroxide pellets, ammonium chloride, distilled water, a polystyrene cup supported in a beaker, two measuring cylinders and a thermometer.

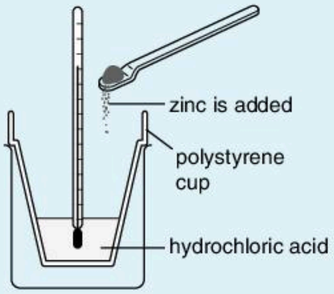
Method

1. Measure 25 cm^3 of hydrochloric acid using a measuring cylinder and pour it into the polystyrene cup. Record the temperature of the acid.
2. Measure 25 cm^3 of sodium hydroxide solution in the other measuring cylinder and add it to the acid. Stir the solutions with the thermometer and record the maximum or minimum temperature.
3. Measure another 25 cm^3 of hydrochloric acid and pour it into the cleaned polystyrene cup. Record the temperature of the acid and add a small amount of zinc. Stir and record the maximum or minimum temperature.

4. Measure 25 cm^3 of distilled water in a cleaned measuring cylinder and pour it into the cleaned polystyrene cup. Record the temperature of the water and add a small amount of potassium nitrate. Stir and record the maximum or minimum temperature.
5. Repeat step **4** using sodium hydroxide pellets and again using ammonium chloride in place of the potassium nitrate in each case. Make sure you wash the polystyrene cup between each experiment.
6. Record your results in a table.
7. Classify each reaction as exothermic or endothermic.



▲ **Figure 12.1.1** Sodium hydroxide solution is added to hydrochloric acid



▲ **Figure 12.1.2** Zinc is added to hydrochloric acid

Examples of exothermic reactions include:

- burning of fuels
- neutralisation reactions between acids and bases, e.g. the reaction between sodium hydroxide and hydrochloric acid
- reactions between reactive metals and acids
- respiration, which releases energy from food in the cells of all living organisms
- dissolving certain substances in water, e.g. sodium hydroxide and sulfuric acid.

Examples of endothermic reactions include:

- the decomposition of compounds when heated, known as **thermal decomposition**, e.g. calcium carbonate must absorb heat in order to decompose into calcium oxide and carbon dioxide
- photosynthesis in plants, which absorbs sunlight energy
- dissolving certain salts in water, e.g. ammonium chloride and potassium nitrate.

Bond breaking and bond forming

In Unit 11.1 you learnt that a chemical reaction involves breaking the original bonds in the reactants and making new bonds in the products.

- When the original bonds are broken in the reactants, energy is *absorbed*.
- When new bonds are formed in the products, energy is *released*.

reactants	⟶	**products**
bonds are broken		bonds are formed
energy is *absorbed*		energy is *released*

- In an exothermic reaction, the energy absorbed to break the existing bonds in the reactants is less than the energy released in forming the new bonds in the products. The extra energy is *released* to the surroundings and the temperature of the surroundings increases.
- In an endothermic reaction, the energy absorbed to break the existing bonds in the reactants is greater than the energy released in forming the new bonds in the products. The extra energy is *absorbed* from the surroundings and the temperature of the surroundings decreases.

Every chemical bond has a specific amount of energy, no matter what compound it is in, known as its bond energy. The amount of energy which has to be absorbed to break that bond is the same as the amount of energy released when that bond is formed.

Did you know?

Average **bond energies** can be found in chemical data books. The amount of energy released or absorbed during a chemical reaction can be calculated by adding up the bond energies for all the bonds in the reactants and for all the bonds in the products and applying the following formula:

energy change = sum of all bond energies in the reactant − sum of all bond energies in the products

Enthalpy change

The energy content of a substance is called its enthalpy and is given the symbol ***H***. The energy content of a substance cannot be measured directly, but it is possible to measure the change in enthalpy that occurs during a reaction. This is known as enthalpy change and is given the symbol $\boldsymbol{\Delta H}$ and is usually expressed in kilojoules, **kJ**, or kilojoules per mol, $\mathbf{kJ\,mol^{-1}}$.

The enthalpy change of a reaction is summarised by the following formula:

$$\text{enthalpy change of a reaction} = (\text{total enthalpy of products}) - (\text{total enthalpy of reactants})$$

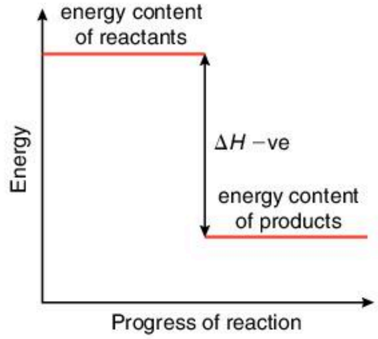
i.e. $\Delta H_{\text{reaction}} = H_{\text{products}} - H_{\text{reactants}}$

Exothermic reactions

In an exothermic reaction the enthalpy of the products is *less* than the enthalpy of the reactants:

i.e $H_{\text{products}} < H_{\text{reactants}}$

Therefore, ΔH has a value which is less than zero, i.e. ΔH is *negative*. The extra energy is released to the surroundings and the temperature of the surroundings increases. Another way to think of this is that since the reaction *loses* energy, ΔH is negative, i.e. ΔH −ve.



▲ **Figure 12.1.3** An exothermic reaction

An example of an exothermic reaction is the combustion of methane (the value of ΔH is given after the equation):

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) \qquad \Delta H = -891\ \text{kJ mol}^{-1}$$

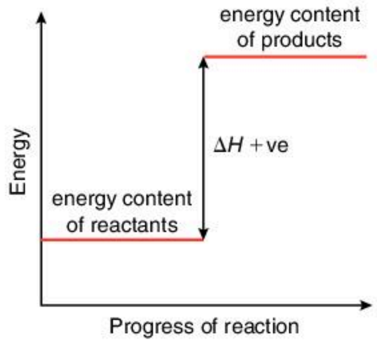
The value for ΔH tells us that 891 kJ of energy is lost when 1 mol of methane is completely burned in oxygen.

Endothermic reactions

In an **endothermic reaction** the enthalpy of the products is greater than the enthalpy of the reactants:

i.e. $H_{\text{products}} > H_{\text{reactants}}$

Therefore, ΔH has a value which is greater than zero, i.e. ΔH is *positive*. Energy is absorbed from the surroundings and the temperature of the surroundings decreases. Another way to think of this is that since the reaction *gains* energy, ΔH is positive, i.e. ΔH +ve.



▲ **Figure 12.1.4** An endothermic reaction

An example of an endothermic reaction is the reaction between hydrogen and iodine:

$$H_2(g) + I_2(s) \longrightarrow 2HI(g) \qquad \Delta H = +26.5\ \text{kJ mol}^{-1}$$

The value for ΔH tells us that 26.5 kJ of energy is gained when 1 mol of hydrogen iodide is made during the reaction. The overall energy absorbed by the reaction represented in the equation would be double this value, i.e. 53 kJ, since it shows 2 mol of hydrogen iodide being made.

Reversible reactions

In a **reversible reaction**, if the forward reaction is exothermic, then the reverse reaction is endothermic and vice versa. The ΔH value given for a reversible reaction is for the forward reaction, i.e. the reaction which proceeds from left to right. For example, the reaction between nitrogen and hydrogen to form ammonia is:

$$N_2(g) + 3H_2(s) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -46.1\ \text{kJ mol}^{-1}$$

This shows that the forward reaction which produces ammonia is exothermic. Therefore, the reverse reaction is endothermic. It also shows us that when 1 mol of ammonia is made during the reaction, 46.1 kJ of energy is released.

Energy profile diagrams

The energy change in a chemical reaction can be illustrated with an **energy profile diagram**. The diagram shows the energy content or enthalpy of the reactants and products, ΔH and the activation energy. Most reactions need some energy to get them started. This minimum amount of energy is called the **activation energy**. This activation energy can be thought of as the energy barrier of a reaction.

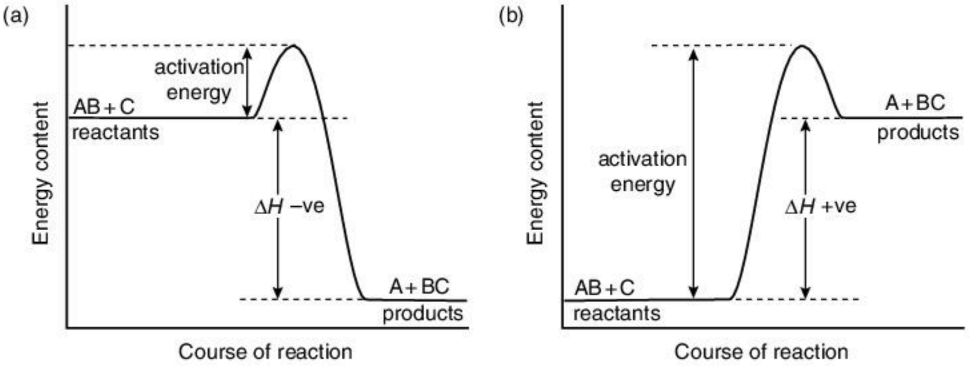
The amount of activation energy does not affect the overall energy change for the reaction and is not included in the calculation of the value of ΔH. The general energy profile diagrams of an exothermic and an endothermic reaction are shown in Figure 12.1.5.

Exam tip

It may be very confusing to think of a reaction which causes an increase in temperature of its surroundings as having a negative value for ΔH and vice versa. Try to remember that an **exothermic reaction** is losing energy to its surroundings and if you lose anything you are minus it, i.e. ΔH is –ve. Similarly, the **endothermic reaction** is gaining energy from its surroundings and if you gain anything you are plus it, i.e. ΔH is +ve.

Key fact

Activation energy is the minimum amount of energy that reactants must be given, in excess of what they normally possess, in order for bonds to start breaking in the reactants and products to start forming.



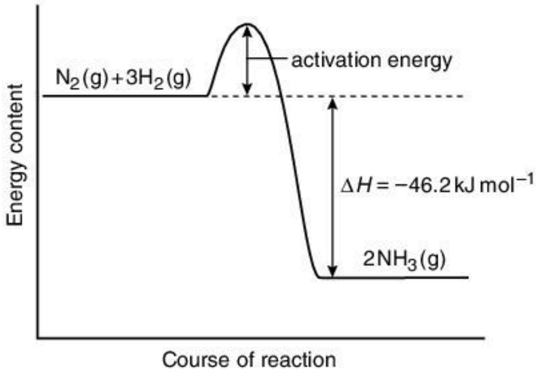
▲ **Figure 12.1.5** Energy profile diagrams.
(a) An exothermic reaction, **(b)** an endothermic reaction.

Drawing energy profile diagrams

When you are drawing an energy profile diagram for a specific reaction it is important that you include the following information on your diagram:

- the formulae of the reactants
- the formulae of the products
- arrows indicating the activation energy and ΔH
- the value of ΔH.

An example of how this can be done for the reaction between nitrogen and hydrogen is shown in Figure 12.1.6.

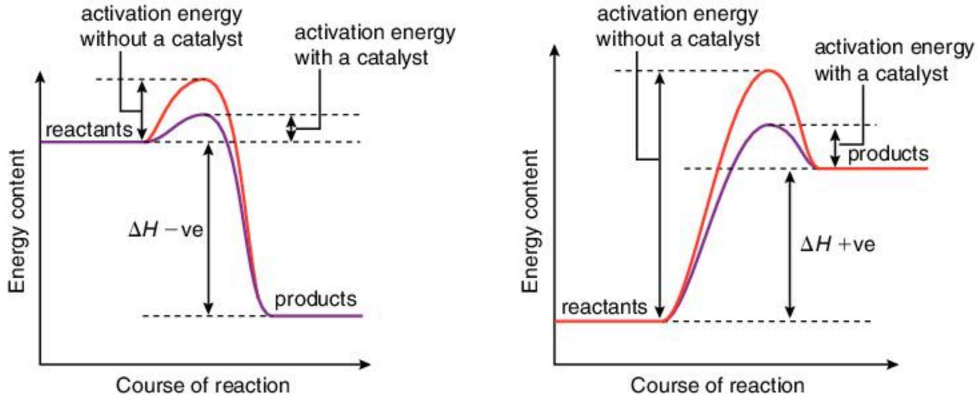


▲ **Figure 12.1.6** Energy profile diagram for the formation of ammonia

Catalysts and energy profile diagrams

In Unit 11.2 you learnt that a **catalyst** is a substance that can increase the rate of a reaction. In terms of energy, a reaction in which a catalyst is used has a *lower activation energy* than the same reaction without a catalyst. This is because a catalyst provides an alternative pathway for the reaction which requires *less* activation energy than the normal pathway. As a result, more reactant particles have the required energy to react when they collide, and the reaction speeds up because there are more effective collisions (collisions that result in the formation of the products) in the same length of time.

The effect of using a catalyst can be shown on energy profile diagrams. It is important to note that although the reaction has a lower activation energy, as shown in the diagrams in Figure 12.1.7, the energy values for the reactants and products remain unchanged, therefore, the value for ΔH remains the same.



▲ **Figure 12.1.7** Energy profile diagrams with and without a catalyst. **(a)** an exothermic reaction, **(b)** an endothermic reaction.

Summary questions

1 Explain what is meant by an exothermic reaction and an endothermic reaction.

2 Explain exothermic and endothermic reactions by reference to:
  a bond breaking and bond forming
  b enthalpy change.

3 Give the formula used to calculate enthalpy change.

4 Draw a fully labelled energy profile diagram for the following reaction. Indicate on your diagram what effect adding a catalyst would have.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \qquad \Delta H = -241.8\,\text{kJ mol}^{-1}$$

A12.2 Calculating energy changes

In the laboratory, the energy change, or enthalpy change, in a reaction can be determined by measuring the heat change that occurs during the reaction, known as the **heat of reaction**. We cannot measure the energy content, or enthalpy, of the reactants and products directly. However, we can calculate the heat of reaction by measuring the change in temperature which occurs during a reaction.

In this unit you will learn how to determine heats of reaction such as the heat of neutralisation and the heat of solution experimentally. In these experiments you will measure the change in temperature which accompanies the neutralisation reaction between a strong acid and a strong alkali and the dissolving of a solid in water. You can then use what is known as the **specific heat capacity** to determine the change in enthalpy for the reaction.

The specific heat capacity of a substance, c, is the quantity of heat required to raise the temperature of a unit mass of the substance by 1 °C or 1 K. We usually take the unit mass of the substance as 1 g and measure temperature in degrees Celsius. This being the case, specific heat capacity is given the units joules per gram per degree Celsius or $\mathbf{J\,g^{-1}\,°C^{-1}}$. Using water as an example, the specific heat capacity of water is $4.2\,\text{J}\,\text{g}^{-1}\,°\text{C}^{-1}$. This means that it takes 4.2 J of heat energy to raise the temperature of 1 g of water by 1 °C.

Objectives

By the end of this topic you will be able to:

- define the terms specific heat capacity, heat of neutralisation and heat of solution
- give the formula to calculate the heat change of a reaction
- calculate heat changes from experimental data
- explain why the heat of neutralisation for the reaction between a strong acid and strong alkali is always the same
- describe experiments to determine the heat of neutralisation and heat of solution
- give the assumptions made when calculating the heat of neutralisation and heat of solution from experimental data.

Key fact

The **specific heat capacity** of a substance is the quantity of heat energy required to raise the temperature of a unit mass of the substance by 1 °C or 1 K.

If we know the mass of the substances reacting, $\boldsymbol{m}$, we can determine the heat of reaction, $\boldsymbol{q}$, from the specific heat capacity of the substances, $\boldsymbol{c}$, and the temperature change, $\boldsymbol{\Delta T}$, using the following formula:

$$q = m \times c \times \Delta T.$$

Measuring heats of reaction

To measure the heat of neutralisation and heat of solution a piece of apparatus known as a **calorimeter** is used. A calorimeter is an insulated container which prevents heat released by a reaction escaping into the surroundings if the reaction is exothermic, or heat from the surroundings being absorbed by the reaction if the reaction is endothermic. A simple calorimeter can be made out of a polystyrene cup because polystyrene is a good insulator.

The general method for determining heats of reaction involves measuring a fixed volume of solution or water to be used in a measuring cylinder, pouring it into the calorimeter and measuring its initial temperature using a thermometer. The reactants are then mixed in the calorimeter and the maximum or minimum temperature reached in the reaction is measured. The temperature change can then be calculated and used in the formula given above to calculate the heat of the reaction.

Assumptions

In the experiments you will be carrying out to determine the heat of neutralisation and heat of solution, you will make three assumptions in your calculations.

- The density of a dilute aqueous solution is the same as water, i.e. $1\,\text{g}\,\text{cm}^{-3}$. This means that the mass of $1\,\text{cm}^3$ of solution is 1 g.
- The specific heat capacity of a dilute aqueous solution is the same as water, i.e. $4.2\,\text{J}\,\text{g}^{-1}\,{}^\circ\text{C}^{-1}$.
- Negligible heat is lost to the surroundings or absorbed from the surroundings during the reaction.

Key fact

The **heat of neutralisation** is the heat change which occurs when 1 mol of water is produced in a reaction between an acid and an alkali.

Heat of neutralisation

The neutralisation reaction between a strong acid and a strong alkali is an exothermic reaction. The amount of heat produced per mole of water formed in the reaction is known as the **heat of neutralisation**.

In calculating the heat of neutralisation, the initial temperature of both solutions must be measured and averaged to determine the initial temperature, and the number of moles of water produced in the reaction must be calculated.

Example

A volume of $50\,\text{cm}^3$ of sodium hydroxide solution of concentration $1.0\,\text{mol}\,\text{dm}^{-3}$ and temperature 26.0 °C is added to $50\,\text{cm}^3$ of hydrochloric acid of concentration $1.0\,\text{mol}\,\text{dm}^{-3}$ and temperature 27.0 °C. After mixing, the maximum temperature of the solution is 33.2 °C. Determine the heat of neutralisation and draw an energy profile diagram for the reaction.

The balanced equation for the reaction is:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Determine the number of moles of water made in the reaction:

1000 cm^3 of NaOH(aq) contain 1.0 mol NaOH

$\therefore$ 50 cm^3 of NaOH(aq) contain $\frac{1.0}{1000} \times 50$ mol NaOH = **0.05 mol NaOH**

And 1000 cm^3 of HCl contains 1.0 mol HCl

$\therefore$ 50 cm^3 of HCl(aq) contain $\frac{1.0}{1000} \times 50$ mol HCl = **0.05 mol HCl**

From the equation:

1 mol NaOH reacts with 1 mol HCl forming 1 mol H_2O

$\therefore$ 0.05 mol NaOH reacts with 0.05 mol HCl forming 0.05 mol H_2O

i.e. **0.05 mol H_2O** is produced in the reaction.

Determine the heat of neutralisation:

Total volume of solution = 50 + 50 = 100 cm^3

$\therefore$ mass of solution, m = **100 g**

Average initial temperature, $T_1 = \frac{26.0 + 27.0}{2}\,°C = 26.5\,°C$

Final temperature, $T_2 = 33.2\,°C$

$\therefore$ temperature increase, $\Delta T = T_2 - T_1 = (33.2 - 26.5)\,°C = \mathbf{6.7\,°C}$

Specific heat capacity of the solution, $c = \mathbf{4.2\,J\,g^{-1}\,°C^{-1}}$

Using $q = m \times c \times \Delta T$,

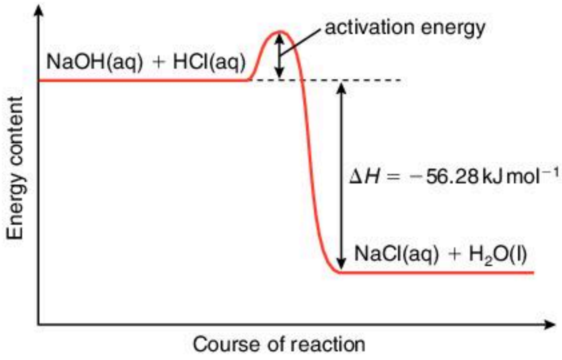
$$\begin{aligned}\text{heat evolved in forming 0.05 mol } H_2O &= (100 \times 4.2 \times 6.7)\,J \\ &= 2814\,J \\ &= \mathbf{2.814\,kJ}\end{aligned}$$

$$\begin{aligned}\therefore \text{ heat evolved in forming 1 mol } H_2O &= \frac{2.814}{0.05}\,kJ \\ &= \mathbf{56.28\,kJ}\end{aligned}$$

i.e. heat of neutralisation, $\Delta H = \mathbf{-56.28\,kJ\,mol^{-1}}$

The heat of neutralisation is negative because the temperature increased, therefore, the reaction was exothermic.

The energy profile diagram is shown below:



Practical activity

To determine the heat of neutralisation

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation.

You will be supplied with 1.0 mol dm^{-3} sulfuric acid, 2.0 mol dm^{-3} potassium hydroxide solution, a calorimeter (polystyrene cup), a measuring cylinder and a thermometer.

Method

1. Measure 50 cm^3 of sulfuric acid in the measuring cylinder and pour it into the dry calorimeter.
2. Measure and record the temperature of the acid using the thermometer.
3. Measure 50 cm^3 of potassium hydroxide solution in the cleaned measuring cylinder. Measure and record its temperature.
4. Pour the potassium hydroxide solution into the calorimeter and stir the solution with the thermometer while watching the temperature change.
5. Record the maximum temperature reached.
6. Calculate the average initial temperature of the two solutions and use this to calculate the temperature increase.
7. Assuming that the density and specific heat capacity of the solution are both the same as water, calculate the heat of neutralisation for the reaction between sulfuric acid and potassium hydroxide. Use the calculation on the previous page to help you.
8. Draw an energy profile diagram for the reaction.
9. The heat of neutralisation for the reaction should be approximately -56.3 kJ mol^{-1}. If your value is lower than this, can you suggest reasons why?

The heat of neutralisation for the reaction between any strong acid and any strong alkali is always about -56.3 kJ mol^{-1}. This is because strong acids and strong alkalis are fully ionised in aqueous solution and the common reaction that releases heat in all cases is:

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l) \qquad \Delta H = -56.3\ \text{kJ mol}^{-1}$$

The equation for the neutralisation reaction between sodium hydroxide and sulfuric acid is:

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

In this reaction, 2 mol of water are produced. The heat of neutralisation is still about -56.3 kJ mol^{-1}, however, the overall heat change for the reaction shown in the equation is about -112.6 kJ, i.e. double the heat of neutralisation.

The heat of neutralisation can also be determined from the results of a **thermometric titration** explained in Unit 8.5. In this method, the initial temperatures of the acid and alkali must be measured and the point of intersection of the two lines of best fit on the graph represents the maximum temperature recorded during the reaction.

Heat of solution

When a solid, liquid or gas dissolves in a solvent, there is usually a change in enthalpy. This enthalpy change which occurs when 1 mol of solute dissolves in a solvent is known as heat of solution.

In calculating the heat of solution, the initial temperature of the water must be measured and the exact quantity of solute dissolving must be known so that the number of moles of solute can be calculated.

Key fact

The **heat of solution** is the heat change which occurs when 1 mol of solute dissolves in such a volume of solvent that further dilution by the solvent produces no further heat change.

Example

A student dissolves 5.35 g of ammonium chloride in 100 cm^3 of distilled water. The initial temperature of the water is 22.5 °C. After mixing, a minimum temperature of 16.4 °C is recorded by the student. Determine the heat of solution of ammonium chloride and draw an energy profile diagram for the solution process.

Determine the number of moles of NH_4Cl dissolved:

$M(NH_4Cl) = 53.5\ \text{g mol}^{-1}$

i.e. mass of 1 mol $NH_4Cl = 53.5$ g

$\therefore$ number of moles in 5.35 g $= \dfrac{5.35}{53.5}$ mol $= \mathbf{0.1\ mol}$

Determine the heat of solution:

Mass of water, $m = \mathbf{100\ g}$

Initial temperature, $T_1 = 22.5\ °\text{C}$

Final temperature, $T_2 = 16.4\ °\text{C}$

$\therefore$ temperature decrease $(\Delta T) = T_1 - T_2 = 22.5 - 16.4\ °\text{C} = \mathbf{6.1\ °C}$

Specific heat capacity of the solution, $c = \mathbf{4.2\ J\ g^{-1}\ °C^{-1}}$

Using $q = m \times c \times \Delta T$,

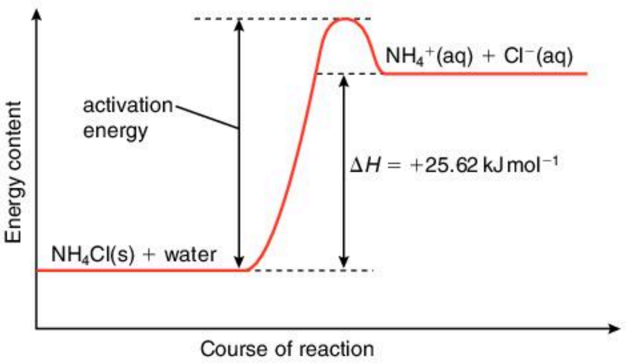
$$\text{heat absorbed in dissolving 0.1 mol } NH_4Cl = 100 \times 4.2 \times 6.1\ \text{J} = 2562\ \text{J} = \mathbf{2.562\ kJ}$$

$\therefore$ heat absorbed in dissolving 1 mol $NH_4Cl = \dfrac{2.562}{0.1}$ kJ $= \mathbf{25.62\ kJ}$

i.e. heat of solution, $\Delta H = \mathbf{+25.62\ kJ\ mol^{-1}}$.

The heat of solution is positive because the temperature decreased, therefore, the reaction was endothermic.

The energy profile diagram is shown below:



Exam tip

It is essential that you give the correct unit for the heat of neutralisation or solution. Since you are calculating the heat change for producing 1 mol of water or dissolving 1 mol of solute, the unit will always be **J mol^{-1}** or **kJ mol^{-1}**. It is also essential that you indicate whether the reaction is exothermic or endothermic by putting a negative or positive sign before your value.

Practical activity

To determine the heat of solution

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation.

You will be supplied with potassium nitrate, distilled water, a balance, a calorimeter (polystyrene cup), a measuring cylinder and a thermometer.

Method

1. Measure 100 cm^3 of distilled water in the measuring cylinder and pour it into the dry calorimeter.
2. Measure and record the temperature of the water.
3. Using the balance, weigh 12.1 g of potassium nitrate and add it to the water in the calorimeter.
4. Stir the solution with the thermometer while watching the temperature change.
5. Record the minimum temperature reached.
6. Calculate the temperature decrease.
7. Assuming that the specific heat capacity of the solution is the same as water, calculate the heat of solution of potassium nitrate. Use the calculation on the previous page to help you.
8. Is dissolving potassium nitrate an exothermic or endothermic reaction? Give a reason for your answer.
9. Draw an energy profile diagram for the solution process.
10. By reference to bonds breaking and bonds forming, explain the reason for the enthalpy change which occurred when potassium nitrate dissolved in water.

When a solute dissolves in a solvent the following events occur.

- Bonds break between the particles of the solute; this *absorbs* energy from the surroundings. In the case of ionic compounds, it is the ionic bonds which break, e.g. when sodium chloride dissolves in water it is the ionic bonds between the Na^+ ions and the Cl^- ions which break. In the case of covalent substances, it is the intermolecular forces between the molecules that break, e.g. when ethanol dissolves in water it is the intermolecular forces between the ethanol molecules that break.
- The intermolecular forces between the solvent molecules also break; this also *absorbs* energy from the surroundings.
- Attractions are formed between the ions or molecules of the solute and the molecules of the solvent. This process is called solvation and it *releases* energy to the surroundings.

The reaction is exothermic if more energy is released during solvation than is absorbed to break the bonds in the solute and solvent. The reaction is endothermic if more energy is absorbed to break bonds in the solute and solvent than is released during solvation.

Summary questions

1. Define the following:
   - **a** heat of neutralisation
   - **b** heat of solution
   - **c** specific heat capacity.
2. What are the assumptions made in an experiment to determine the heat of neutralisation?
3. Why is the heat of neutralisation for the reaction between sodium hydroxide and hydrochloric acid the same as the heat of neutralisation for the reaction between potassium hydroxide and sulfuric acid?
4. It was found that when 4.0 g of ammonium nitrate dissolved in 50 cm^3 of distilled water, the temperature decreased from 27.4 °C to 21.6 °C. Determine the heat of solution for ammonium nitrate and draw an energy profile diagram for the process.

Key concepts

- **Exothermic reactions** produce heat energy causing the reaction mixture and its surroundings to get hotter, i.e. they transfer energy to their surroundings.
- **Endothermic reactions** absorb heat energy causing the reaction mixture and its surroundings to get colder, i.e. they absorb energy from their surroundings.
- When bonds are broken in reactants energy is absorbed. When new bonds are formed in products energy is released.
- In an exothermic reaction, the energy absorbed to break bonds is less than the energy released in forming new bonds.
- In an endothermic reaction, the energy absorbed to break bonds is greater than the energy released in forming new bonds.
- The energy content of a substance is called **enthalpy** and is given the symbol H.
- The enthalpy change in a reaction can be measured and is given the symbol ΔH.
- $\Delta H_{\text{reaction}} = H_{\text{products}} - H_{\text{reactants}}$
- In an exothermic reaction $H_{\text{products}} < H_{\text{reactants}}$, therefore, ΔH is less than zero, i.e. ΔH is negative.
- In an endothermic reaction, $H_{\text{products}} > H_{\text{reactants}}$, therefore, ΔH is greater than zero, i.e. ΔH is positive.
- The enthalpy change in a chemical reaction can be illustrated by an **energy profile diagram**.
- The **activation energy** is the minimum amount of energy required for a reaction to occur.
- A **catalyst** reduces the amount of activation energy required for a reaction to occur by providing an alternative route for the reaction to proceed which requires less energy.
- The specific heat capacity of a substance enables us to determine enthalpy change from a change in temperature.
- The **specific heat capacity** is the quantity of heat energy required to raise the temperature of a unit mass of the substance by 1 °C or 1 K.
- The **heat of reaction**, q, can be calculated from the masses of the substances reacting, m, the specific heat capacity of the substances, c, and the change in temperature, ΔT, using the formula:

  $$q = m \times c \times \Delta T.$$
- The **heat of neutralisation** is the heat change which occurs when 1 mol of water is produced in a reaction between an acid and an alkali.
- The **heat of solution** is the heat change which occurs when 1 mol of solute dissolves in such a volume of solvent that further dilution by the solvent produces no further heat change.
- To determine the heat of neutralisation or heat of solution we assume that 1 cm^3 of a dilute aqueous solution has a mass of 1 g and a specific heat capacity which is the same as water, i.e. $4.2\ \text{J g}^{-1}\,°\text{C}^{-1}$, and that negligible heat is lost to, or absorbed from, the surroundings.
- The heat of neutralisation for the reaction between a strong acid and a strong alkali is always the same, about $-56.3\ \text{kJ mol}^{-1}$.

Practice exam-style questions

Multiple-choice questions

1 Which of the following statements best describes an endothermic reaction?

A The enthalpy of the reactants is greater than the enthalpy of the products.
B Energy is released to the surroundings.
C The enthalpy change is negative.
D The temperature of the surroundings decreases.

2 When sodium hydroxide is added to water at 28 °C, the temperature of the solution increases to 40 °C because:

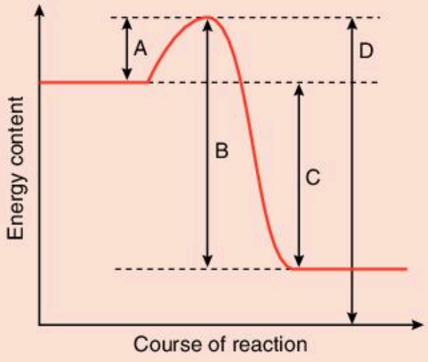
A the enthalpy change is positive
B energy is released to the surroundings
C the reaction is endothermic
D more energy is absorbed to break the ionic bonds between the Na^+ ion and OH^- ions than is released during solvation

3 Which of the following processes releases energy?

I breaking of a bond
II solvation of an ion
III formation of a bond

A I only
B I and II only
C II and III only
D III only

4 The diagram below shows the energy profile diagram of a reaction. Which letter in the diagram represents the energy necessary for the reaction to take place?



5 Consider the following equation for the reaction between sodium hydroxide and phosphoric acid

$$3NaOH(aq) + H_3PO_4(aq) \longrightarrow Na_3PO_4(aq) + 3H_2O(l) \quad \Delta H = -168.9\,kJ$$

Which of the following statements is incorrect?

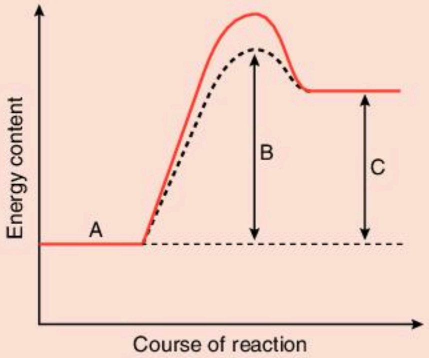
A The reaction is exothermic.
B The heat of neutralisation is $-168.9\,kJ\,mol^{-1}$.
C The reactants have more energy than the products.
D The temperature of the surroundings would increase.

6 When 0.2 mol sodium nitrate dissolved in $100\,cm^3$ of water, the temperature decreased by 7.2 °C. Assuming that the specific heat capacity of the solution is $4.2\,J\,g^{-1}\,°C^{-1}$, the heat of solution of sodium nitrate is:

A $604.8\,J\,mol^{-1}$
B $3024\,J\,mol^{-1}$
C $15.12\,kJ\,mol^{-1}$
D $151.2\,kJ\,mol^{-1}$

Structured question

7 **a** The figure below shows an energy profile diagram for a reaction and the effect of a catalyst on the reaction pathway.



Energy profile diagram

i) State, with a reason, if the reaction is exothermic or endothermic. *(2 marks)*

ii) Identify the parts labelled A, B, and C on the diagram. *(3 marks)*

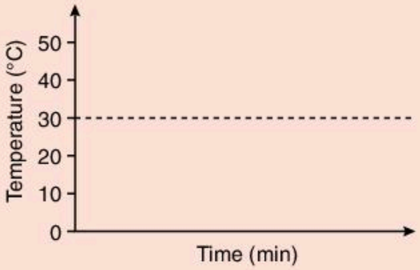
iii) How does a catalyst work to speed up the rate of a reaction? *(2 marks)*

b The dissolving of ammonium nitrate in water may be represented by the following equation:

$$NH_4NO_3(s) + water \longrightarrow NH_4^+(aq) + NO_3^-(aq) \quad \Delta H\ +ve$$

In order to determine the value of ΔH, a student dissolved 16.0 g of ammonium nitrate in $75\,cm^3$ of distilled water at room temperature (30.0 °C). He stirred the mixture and recorded its temperature at regular intervals until all the solid had dissolved and the temperature of the mixture had returned to room temperature. The largest difference he observed between the initial temperature and the solution temperature was 15 °C.

i) On the axes below, sketch a graph to show the variation in solution temperature with time.



(2 marks)

ii) Calculate the energy change for the reaction, assuming that the specific heat capacity of the solution is $4.2\,J\,g^{-1}\,°C^{-1}$. *(2 marks)*

iii) Determine the number of moles of ammonium nitrate that dissolved. *(2 marks)*

iv) Determine the heat of solution of ammonium nitrate. *(2 marks)*

Total 15 marks

Extended response question

8 a i) Define 'heat of neutralisation'. *(2 marks)*

ii) Account for the fact that the heat of neutralisation for the reaction between any strong acid and any strong alkali is the same. Give a relevant ionic equation to support your answer. *(2 marks)*

b A student was required to carry out an experiment to determine the heat of neutralisation in the reaction between potassium hydroxide and sulfuric acid. She was provided with the following materials and equipment:

- 50 cm^3 of potassium hydroxide solution
- 50 cm^3 of sulfuric acid of concentration 0.6 $mol\,dm^{-3}$
- a thermometer
- a polystyrene cup.

She recorded the initial temperature of the two solutions, mixed them in the polystyrene cup and recorded the maximum temperature reached. Her temperature recordings are given below:

Initial temperature of potassium hydroxide solution = 29.5 °C
Initial temperature of sulfuric acid = 29.5 °C
Maximum temperature after mixing = 37.1 °C

i) Determine the heat change for the reaction, assuming that the specific heat capacity of the solution is $4.2\,J\,g^{-1}\,°C^{-1}$. *(3 marks)*

ii) State TWO other assumptions that you have made in your calculation in **i)** above. *(2 marks)*

iii) From your answer to **i)** above, calculate the heat of neutralisation given that the sulfuric acid is the limiting reactant. *(4 marks)*

iv) Draw a fully labelled energy profile diagram for the reaction. *(2 marks)*

Total 15 marks

Section B

B13 Introduction to organic chemistry

Objectives

By the end of this topic you will be able to:

- give examples of organic compounds
- give the main compounds containing carbon
- give the electron structure of a carbon atom
- understand and illustrate that carbon can form single bonds, double bonds, unbranched chains, branched chains and ringed structures
- explain the term functional group
- identify certain functional groups
- represent organic compounds using molecular, fully displayed and condensed structural formulae.

Organic chemistry is the branch of chemistry in which compounds containing carbon atoms, known as organic compounds, are studied. Studying organic compounds is very important since they are all around us and form an integral part of our daily lives, for example plastics, oil, gasoline, diesel and alcohols. Organic compounds also form a major part of the bodies of living organisms and of the food we eat. Proteins, carbohydrates and fats are all organic compounds.

B13.1 Organic compounds

All **organic compounds** contain carbon. Most also contain hydrogen and many contain oxygen. Other elements may also be present such as nitrogen, sulfur and phosphorus. Organic compounds which contain only carbon and hydrogen are known as **hydrocarbons**.

A vast number of organic compounds exist and they are used in all aspects of our lives. Some examples of organic compounds that we use in our everyday lives are given below.

- Methane, CH_4, is the main component of natural gas which is used as a fuel to generate electricity, as a fuel for domestic heating and cooking purposes and as a fuel for vehicles in some countries.
- Ethanoic acid, CH_3COOH, is the main component in vinegar which is used to preserve and flavour food.
- Ethanol, C_2H_5OH, forms the basis of wines and spirits.
- Glucose, $C_6H_{12}O_6$, is used by both plants and animals as one of their main sources of energy.
- Protein is an example of a natural polymer (see Unit 16) which is used by plants and animals to enable them to grow.
- Polyethene is an example of a man-made polymer which is widely used as a plastic, for example, in sandwich bags and cling wrap.

Formation of carbon compounds

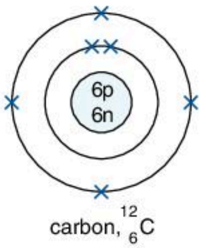
The great abundance of organic compounds in the world is because of the chemistry of the **carbon atom**. Carbon is found in:

- carbon dioxide and carbon monoxide
- metal carbonates and hydrogencarbonates
- fossil fuels, such as crude oil, natural gas and coal
- carbohydrates, proteins and lipids (fats and oils)
- organic acids, such as ascorbic acid and citric acid
- diamond and graphite which are pure carbon.

The atomic number of carbon is 6. A carbon atom therefore has six electrons and has an electronic configuration of C (2,4). The shell diagram of a carbon atom is shown in Figure 13.1.1.

Looking at the electron structure, a carbon atom has *four* valence electrons. This means that carbon can share four electrons with other atoms, i.e. it can form four covalent bonds with other atoms. These other atoms include hydrogen, oxygen, nitrogen and the halogens.

It is important to note that not all carbon-containing compounds are organic compounds. Carbon dioxide, carbon monoxide, metal carbonates and metal hydrogencarbonates are classified as inorganic compounds.



▲ **Figure 13.1.1** Electron structure of carbon

Bonding in organic compounds

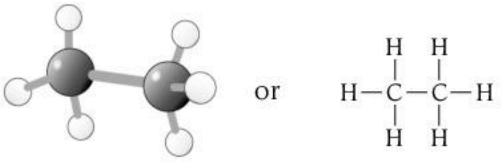
As a result of carbon having four valence electrons, carbon atoms can bond with other carbon atoms in an almost unlimited way. Because of this, a vast number of different organic compounds exist compared to inorganic compounds. The simplest of these is methane (CH_4). If we look at the methane molecule in the ball and stick model in Figure 13.1.2, we can see that carbon is bonded to four hydrogen atoms arranged in a tetrahedron around it. The bond angle between covalent bonds is 109.5°.

Organic molecules, therefore, have very distinctive three-dimensional structures.

When carbon atoms bond with each other a variety of structures can form.

Single bonds

Single bonds can form between adjacent carbon atoms. In a single bond, a carbon atom shares *one* pair of electrons with another carbon atom. This leaves three other electrons per carbon atom available for bonding with other atoms. For example, ethane (C_2H_6) has a single bond between the two carbon atoms, represented by the single line:



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

ethane (C_2H_6)

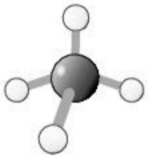
Organic compounds containing only single bonds between adjacent carbon atoms are known as saturated compounds.

Double bonds

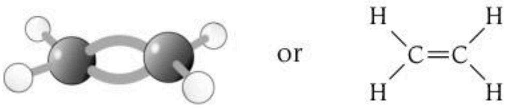
Double bonds can form between adjacent carbon atoms. In a double bond, a carbon atom shares *two* pairs of electrons with another carbon atom. This leaves two other electrons per carbon atom available for bonding with other atoms. For example, ethene (C_2H_4) has a double bond between two carbon atoms, represented by two lines:

? **Did you know?**

Carbon is the sixteenth most abundant element in the Earth's crust and the sixth most abundant element by mass in the Universe after hydrogen, helium, oxygen, neon and nitrogen. It is present in all forms of life and is the second most abundant element by mass in the human body after oxygen.



▲ **Figure 13.1.2** Ball and stick model of methane

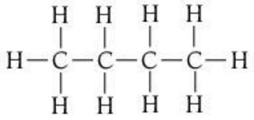


ethene (C_2H_4)

Organic compounds containing one or more double bonds between adjacent carbon atoms are known as **unsaturated compounds**.

Unbranched chains

Unbranched chains of carbon atoms of different lengths can form. For example, butane (C_4H_{10}) is an unbranched chain of four carbon atoms with single bonds between adjacent atoms:

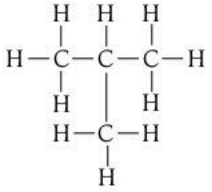


$$\mathrm{H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H}$$

butane (C_4H_{10})

Branched chains

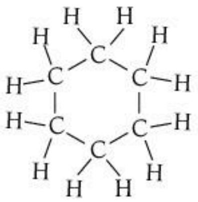
Branched chains of carbon atoms of different lengths can form. For example, methylpropane (C_4H_{10}) is a chain of three carbon atoms with the fourth carbon atom branching off from the middle atom of the chain:



methylpropane (C_4H_{10})

Rings

Rings of carbon atoms can form. For example, cyclohexane (C_6H_{12}) is a ring of six carbon atoms:



cyclohexane (C_6H_{12})

The ability of carbon atoms to form covalent bonds with other carbon atoms forming long chains and rings of carbon atoms is known as **catenation**.

It is important to note that, with the exception of the ball and stick models, the drawings above show each molecule in two-dimensions only.

Functional groups

Organic compounds contain one or more **functional groups**. A functional group is a particular atom, a particular group of atoms, or a particular bond between two carbon atoms within the molecule. The functional group determines the chemical properties of the compound containing it. Organic compounds are classified into groups known as **homologous series** based on the functional group that they contain (see Unit 13.2). Some of the functional groups we will be studying are given in Table 13.1.1.

The simplest organic compounds that you will be studying can be thought of as being composed of two parts:

- the **hydrocarbon part** which is composed of carbon and hydrogen atoms only.
- the **functional group**.

Formulae of organic compounds

The formulae of organic compounds can be written in a variety of different ways. Using butanoic acid as an example, the most common methods are:

- The **molecular formula**. This shows the actual number of atoms of each element in one molecule of the compound. The molecular formula for butanoic acid is $C_4H_8O_2$. Molecular formulae have a limited use in organic chemistry since they don't give any information about how the atoms are joined within the molecule.
- The **fully displayed structural formula**. This shows in two-dimensional diagrammatic form how the atoms are arranged in one molecule of the compound (Figure 13.1.3).

  The fully displayed formula has the limitation that it is a two-dimensional drawing of a three-dimensional molecule, i.e. the molecule is straightened out and flattened. The **ball and stick model** in Figure 13.1.4 represents the actual molecule in three-dimensions. However, you would not be required to draw this.
- The **condensed structural formula**. This shows the sequence and arrangement of atoms in one molecule of the compound in such a way that the nature and position of attachment of the functional group is shown without actually drawing the molecule. Each carbon atom is written separately and in sequence, followed by the other atoms that are bonded to it. The condensed structural formula for butanoic acid is:

  $$CH_3CH_2CH_2COOH$$

  This may be shortened further by showing the total number of carbon atoms and total number of hydrogen atoms in the hydrocarbon part of the molecule, i.e.

  $$C_3H_7COOH$$

Summary questions

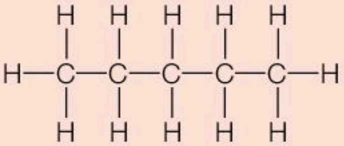
1 How many bonds can a carbon atom form with other atoms?

2 What is a hydrocarbon?

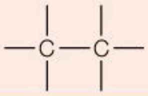
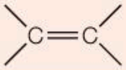
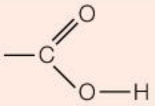
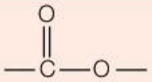
3 What is the difference between a saturated and an unsaturated organic compound?

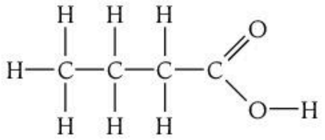
4 What is a functional group?

5 Give the condensed structural formula and molecular formula of this compound:

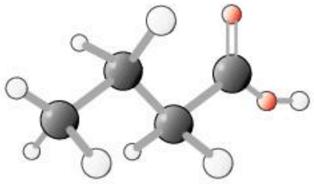


▼ **Table 13.1.1** Functional groups

Name of functional group	Structure of functional group
carbon–carbon single bond	
carbon–carbon double bond	
hydroxyl group	—O—H or —OH
carboxyl group	 or —COOH
ester group	 $-\overset{O}{\overset{\Vert}{C}}-O-$ or —COO—



▲ **Figure 13.1.3** Fully displayed structural formula of butanoic acid



▲ **Figure 13.1.4** Ball and stick model of butanoic acid

B13.2 Homologous series

We can *classify* compounds into groups based on the functional group which they contain. These groups are known as **homologous series**.

Characteristics of a homologous series

The characteristics of a homologous series are given below.

- All members of the series have the same functional group.
- All members of the series can be represented by the same general formula.
- Each member of the series differs from the member before or after it by a $-CH_2-$ group and by relative molecular mass of 14.
- All members of a series possess similar chemical properties. Reactivity decreases as molar mass increases, i.e. as the number of carbon atoms per molecule increases.
- All members of a series show a gradual change in their physical properties as their molar mass increases. In general, as molar mass increases, melting point, boiling point and density increase.
- All members of a homologous series can be prepared by the same general method.

Naming straight chain members of a homologous series

Because of the great variety of organic compounds, a set of rules for naming these compounds has been developed by the International Union of Pure and Applied Chemistry (IUPAC). The process of naming chemical compounds is known as **chemical nomenclature**. The main function of chemical nomenclature is to ensure that when a person reads a chemical name there is no ambiguity as to which chemical compound it refers.

The names of straight chain members of a homologous series consist of two parts:

- The first part, or *prefix*, is related to the *total number of carbon atoms* present in one molecule. Table 13.2.1 gives the prefixes for organic compounds up to ten carbon atoms.
- The second part which is based on the homologous series to which the compound belongs and relates to the *functional group* present. For example, if the compound belongs to the alkane series, its name ends in '-ane'.

Total number of carbon atoms	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-
9	non-
10	dec-

▶ **Table 13.2.1** Prefixes for naming organic compounds

Table 13.2.2 gives the main homologous series you will be studying, together with an example of a member of each series containing two carbon atoms.

Objectives

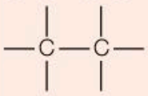
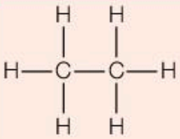
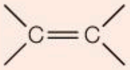
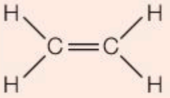
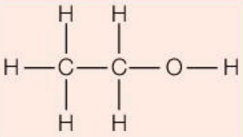
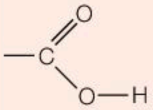
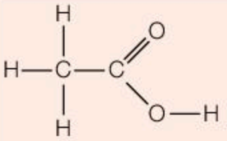
By the end of this topic you will be able to:

- define the term homologous series
- list the general characteristics of a homologous series
- write the general formula of a given homologous series
- write the formulae for members of a given homologous series
- name straight chain members of the common homologous series
- determine the homologous series to which a compound belongs from its name or formula
- define structural isomer and structural isomerism
- explain how structural isomers can be formed
- name branched chain isomers.

! Key fact

A **homologous series** is a group of organic compounds that all possess the same functional group.

▼ **Table 13.2.2** The main homologous series

Name of homologous series	General formula	Functional group present	Naming	Example containing two carbon atoms	Condensed formula	Structural formula
Alkane	C_nH_{2n+2}	carbon–carbon single bond 	prefix + ane	ethane	CH_3CH_3 or C_2H_6	
Alkene	C_nH_{2n}	carbon–carbon double bond 	prefix + ene	ethene	$CH_2{=}CH_2$ or C_2H_4	
Alcohol or alkanol	$C_nH_{2n+1}OH$	hydroxyl group —O—H (—OH)	prefix + anol	ethanol	CH_3CH_2OH or C_2H_5OH	
Alkanoic acid or carboxylic acid	$C_nH_{2n+1}COOH$	carboxyl group  (—COOH)	prefix + anoic acid	ethanoic acid	CH_3COOH	

Determining homologous series

The homologous series that a compound belongs to can be determined if you are given the name of the compound or its formula. You can also work out the formula of a compound if you have been given its name or you can deduce its name if you have been given its formula.

Examples

- To determine the *homologous series* that a compound belongs to from the name, look at the second part of the name. For example, if the second part of the name is '-ene' then the compound belongs to the alkene series.
- To work out the *formula* of the compound from the name, first determine the homologous series to which it belongs and then look at the prefix of the name. For example, if the name of the compound is *butane* then it belongs to the alkane series. Its prefix is 'but-' so it has four carbon atoms. The general formula of an alkane is C_nH_{2n+2}. The formula of butane must be C_4H_{10}.
- To determine the *homologous series* that a compound belongs to from the formula, look and see what functional group is present. For example, if the functional group present is the —O—H group, then the compound belongs to the alcohol series.
- To determine the *name* of the compound from the formula, first determine the homologous series to which the compound belongs and then look at the total number of carbon atoms in the molecule. For example, if the formula of the compound is C_2H_5COOH then it belongs to the alkanoic acid series so the second part of the name must be '-anoic acid'. It has a total of three carbon atoms in the molecule so the prefix must be 'prop-'. The name of the compound is propanoic acid.

Exam tip

It is very important that you can work out the formula of a compound if you have been given its name or that you can deduce the name of a compound if you have been given the formula. To do either of these, you must be able to identify the **homologous series** to which the compound belongs.

Structural isomers

Many organic compounds have the same molecular formula but different structural formulae, i.e. their molecules contain the same number of atoms of each element, but these atoms are bonded together differently. These compounds are called structural isomers and this phenomenon is called structural isomerism.

Key fact

Structural isomers are compounds which have the same molecular formula but different structural formulae.

Key fact

Structural isomerism is the occurrence of two or more organic compounds with the same molecular formula but different structural formulae.

Structural isomers can belong to the same homologous series or to different homologous series.

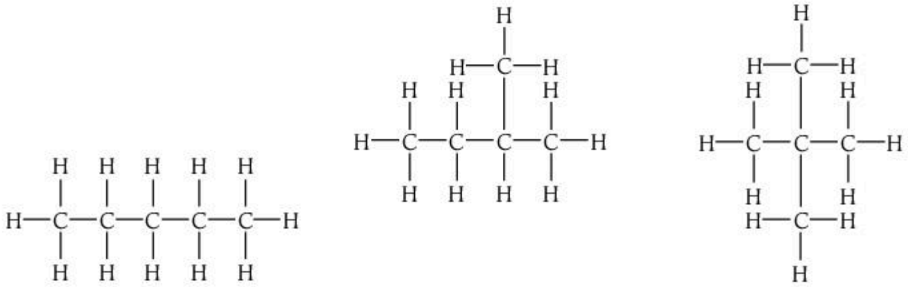
- If they contain the *same* functional group they must belong to the same homologous series. Because they have the same functional group, they have very similar chemical properties, however, because their atoms are bonded differently they have different physical properties.
- If they contain *different* functional groups they belong to different homologous series. Because of this they have different chemical and physical properties.

Each different structural isomer has a different name.

Structural isomers of straight chain molecules can be formed in two ways, by the chain of carbon atoms becoming branched or by the position of the functional group changing.

Forming isomers by branching

Carbon chains can have side branches of one or more carbon atoms. For example, pentane (C_5H_{12}) has three isomers, one of which is the straight chain isomer and two of which are formed by branching of the carbon chain as shown in Figure 13.2.1. The condensed structural formula of each isomer is given below the isomer.



$CH_3CH_2CH_2CH_2CH_3$

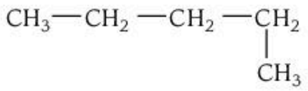
Straight chain isomer

$CH_3CH_2CH(CH_3)CH_3$ $CH_3C(CH_3)_2CH_3$

Branched chain isomers

▶ **Figure 13.2.1** Isomers of pentane

It is important to note that when drawing the structural formula of an organic compound, the longest continuous chain of carbon atoms is always drawn *horizontally*. Care must also be taken not to draw bent or twisted versions of the straight chain isomer. For example, the following is a bent version of the straight chain isomer of pentane and, therefore, is incorrect:



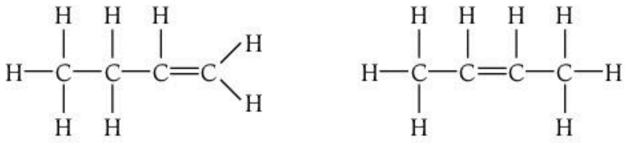
$$CH_3—CH_2—CH_2—\underset{\displaystyle CH_3}{\underset{|}{CH_2}}$$

Forming isomers by changing the position of the functional group

The functional group is usually situated at the right-hand end of the molecule. However, the position of the functional group can change.

Examples

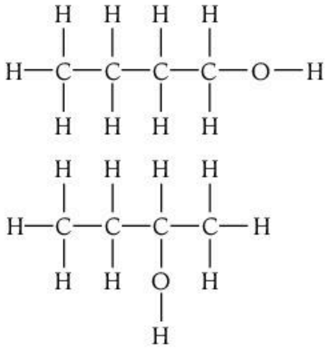
Butene (C_4H_8) has two isomers because the position of the carbon–carbon double bond can change as shown in Figure 13.2.2.



◀ **Figure 13.2.2** Isomers of butene

When drawing the structural formulae of alkenes, the carbon–carbon double bond is always drawn *horizontally*. You will learn more about the isomers of alkenes in Unit 14.2.

Butanol (C_4H_9OH) has two isomers because the position of the hydroxyl group (—OH) can change as shown in Figure 13.2.3.



▲ **Figure 13.2.3** Isomers of butanol

Naming branched chain isomers

The side chains which branch off from the longest chain of carbon atoms in a molecule are known as alkyl groups. These groups have the general formula C_nH_{2n+1}. Alkyl groups are named by adding the suffix **'-yl'** to the prefix.

▼ **Table 13.2.3** Naming alkyl groups

Formula of the alkyl group	Name
$-CH_3$	methyl
$-C_2H_5$	ethyl
$-C_3H_7$	propyl

The name of a branched chain molecule is made up of three parts:

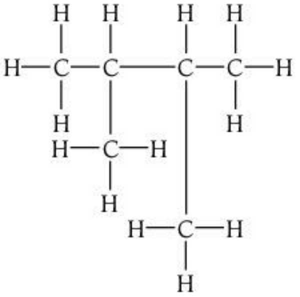
- the first part indicates the number of the carbon atom to which the side chain is attached
- the second part is the name of the side chain, i.e. the alkyl group
- the third part is derived from the homologous series to which the compound belongs and the longest continuous chain of carbon atoms.

To name branched chain molecules, therefore, the following four steps should be used.

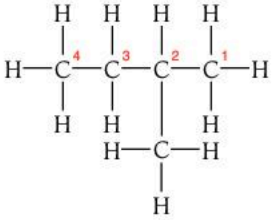
1) Find the **number** of the carbon atom to which the side chain is attached.
2) Determine the **name** of the side chain.
3) Determine the **homologous series** to which the compound belongs and the **total number** of carbon atoms in the longest continuous carbon chain.
4) Put the information from steps 1) to 3) together to name the compound, separating the number from the name by a hyphen.

Examples

1 Determine the name of the following branched chain isomer of C_5H_{12}.



1) Find the longest continuous chain of carbon atoms in the structure. Number the carbon atoms from the end closest to the side chain so that the side chain is at the *lowest* number:



The side chain is attached to carbon atom number **2**.

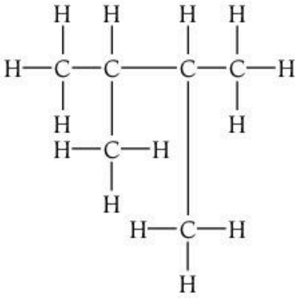
2) The side chain is an alkyl group with *one* carbon atom. It is the **methyl** group.

3) The compound has only carbon–carbon single bonds. Therefore, it belongs to the *alkane* series.

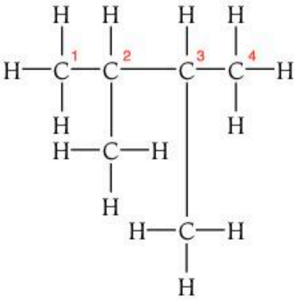
The longest continuous carbon chain has *four* carbon atoms in it. Its name is, therefore, **butane**.

4) The name of the compound is **2-methylbutane**.

2 Determine the name of the following branched chain isomer of C_6H_{14}.



1) Number the carbon atoms in the longest continuous carbon chain so that the side chains are at the lowest numbers:



There are two side chains. One is attached to carbon atom number 2 and one is attached to carbon atom number 3. When writing the name we write **2,3-** because both side chains are the same, i.e. CH_3.

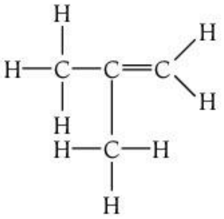
2) Both side chains are alkyl groups with *one* carbon atom. They are both the methyl group and because there are two of them we use the prefix 'di-', i.e. we write **dimethyl**.

3) The compound has only carbon–carbon single bonds. Therefore, it belongs to the *alkane* series.

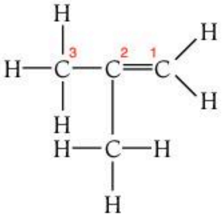
   The longest continuous carbon chain has *four* carbon atoms in it. Its name is, therefore, **butane**.

4) The name of the compound is **2,3-dimethylbutane**.

3 Determine the name of the following branched chain isomer of C_4H_8.



1) Number the carbon atoms in the longest continuous carbon chain so that the side chain is at the lowest number:



   The side chain is attached to carbon atom number **2**.

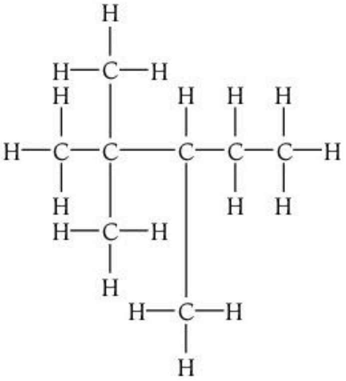
2) The side chain is an alkyl group with *one* carbon atom. It is the **methyl** group.

3) The compound has a carbon–carbon double bond; therefore, it belongs to the *alkene* series.

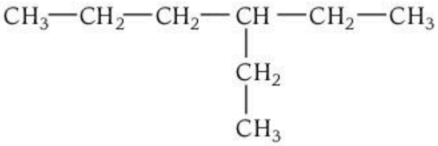
   The longest continuous carbon chain has *three* carbon atoms in it. Its name is, therefore, **propene**.

4) The name of the compound is **2-methylpropene**.

Other examples are given below:



2,2,3-trimethylpentane



3-ethylhexane

Summary questions

1 Give four characteristics of a homologous series.

2 Give the name and condensed formula of each of the following straight chain isomers:

a an alcohol with one carbon atom

b an alkene with four carbon atoms

c an alkanoic acid with three carbon atoms

d an alkane with two carbon atoms.

3 Name the following compounds:

a C_3H_6

b C_4H_9OH

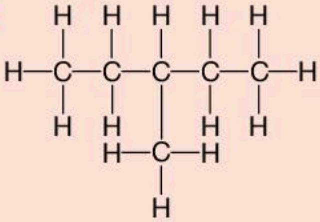
c CH_4

d C_2H_5COOH

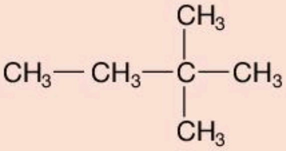
4 Define structural isomerism.

5 Name the following isomers:

a



b



6 Draw the structural formula of 2,3-dimethylhexane.

Key concepts

- Organic chemistry is the study of organic compounds.
- All organic compounds contain carbon. Most also contain hydrogen and many contain oxygen.
- A carbon atom can form four covalent bonds with other atoms.
- Carbon atoms have the ability to form single and double bonds with other carbon atoms. They can also form straight chains, branched chains and ringed structures.
- A functional group is a particular atom, a particular group of atoms or a particular bond between two carbon atoms within the molecule.
- The functional group determines the chemical properties of the compound containing it.
- An organic compound can be represented by a molecular formula, a fully displayed structural formula or a condensed structural formula.
- Organic compounds are classified into groups known as homologous series based on the functional group that they contain.
- Members of a homologous series all possess the same functional group, can all be represented by the same general formula and have certain other common characteristics.
- The names of straight chain members of a homologous series consist of two parts, the first part, or prefix, which is related to the total number of carbon atoms in the molecule and the second part which is related to the functional group.
- Structural isomerism is the occurrence of two or more organic compounds with the same molecular formula but different structural formulae. The compounds are known as structural isomers.
- Structural isomers of straight chain molecules can be formed in two ways, by the chain of carbon atoms becoming branched or by the position of the functional group changing.
- Each different structural isomer has a different name.

Practice exam-style questions

Multiple-choice questions

1 Members of a homologous series:

I have similar chemical properties
II have the same molecular formula
III differ from succeeding members by CH_2

A I and II only
B I and III only
C II and III only
D I, II and III

2 To which homologous series does C_4H_8 belong?

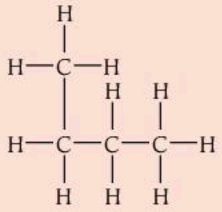
A alkanes
B alkenes
C alcohols
D alkanoic acid

3 Isomers always:

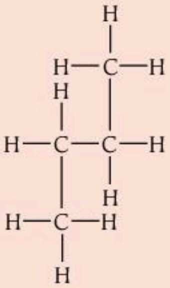
A have the same functional group
B have similar chemical properties
C have similar physical properties
D have different structural formulae

4 Which of the following formulae do not represent isomers?

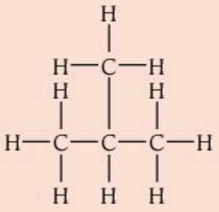
I



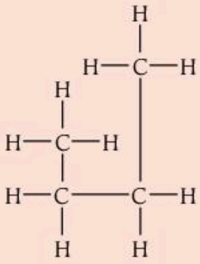
II



III



IV



A I and II
B I and III
C II and III
D III and IV

5 The general formula of an alkane is:

A C_nH_{2n}
B C_nH_{2n+1}
C C_nH_{2n+2}
D C_nH_{2n+3}

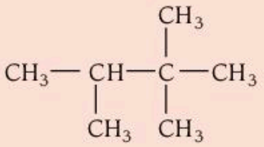
6 A compound has the formula C_3H_7COOH. The name of the compound is:

A ethanoic acid
B propanoic acid
C propanol
D butanoic acid

7 Which of the following is the correct formula of pentanol?

A $C_5H_{10}OH$
B $C_5H_{11}OH$
C $C_5H_{12}OH$
D $C_5H_{13}OH$

8 What is the name of the following isomer?

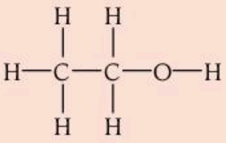


$$CH_3-\underset{CH_3}{CH}-\overset{CH_3}{\underset{CH_3}{C}}-CH_3$$

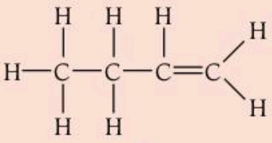
A 2,2,3-trimethylbutane
B 2,3,3-trimethylbutane
C 2,3,3-trimethylheptane
D 2,2,3-trimethylheptane

Structured question

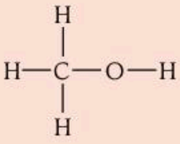
9 a The fully displayed structural formulae of five organic compounds are given below:



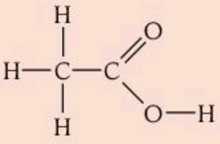
A



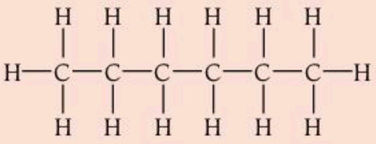
B



C



D



E

i) Which compounds belong to the same homologous series? *(1 mark)*

ii) Name the homologous series to which the compounds you named in **a i)** belong and give a reason for placing them in the same series. *(2 marks)*

iii) Give the names of compounds **B** and **D**. *(2 marks)*

iv) Give the name and fully displayed structural formula of the compound which has FOUR carbon atoms in its molecule and belongs to the same homologous series as compound **D**. *(2 marks)*

v) Draw the fully displayed structures of any TWO isomers of compound **E**. Give the name of EACH isomer you have drawn. *(4 marks)*

b Name the homologous series to which EACH of the following organic compounds belongs and draw the fully displayed structural formula of EACH compound:

i) propene

ii) 2,2-dimethylpentane. *(4 marks)*

Total 15 marks

Extended response question

10 a Explain fully how the electronic configuration of the carbon atom, ${}^{12}_{6}C$, can account for the fact that carbon can form such a wide variety of organic compounds. *(5 marks)*

b Organic compounds can be classified into homologous series.

i) List FOUR characteristics of a homologous series. *(4 marks)*

ii) A compound has the formula C_2H_5COOH

- To which homologous series does the compound belong?
- Give the name of the compound. *(2 marks)*

c **i)** What do you understand by the term 'structural isomerism'? *(2 marks)*

ii) Explain why structural isomers of members of the same homologous series have very similar chemical properties, but they have different physical properties. *(2 marks)*

Total 15 marks

B14 Hydrocarbons – alkanes and alkenes

Objectives

By the end of this topic you will be able to:

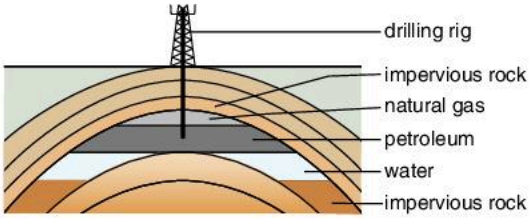
- identify natural gas and petroleum as natural sources of hydrocarbons
- describe the fractional distillation of petroleum
- list the main uses of the fractions obtained from the fractional distillation of petroleum
- describe the thermal and catalytic cracking of alkanes.

Hydrocarbons are organic compounds which consist entirely of carbon and hydrogen. The majority of hydrocarbons are naturally formed by the decomposition of dead marine organisms under the effects of high pressures, heat and bacteria deep under the sea, converting them into oily liquids or gases. Hydrocarbons are one of the most important energy sources on Earth.

B14.1 Sources and extraction of hydrocarbons

Sources of hydrocarbons

Hydrocarbons are organic compounds containing carbon and hydrogen atoms only. They include the alkanes, the alkenes and the ringed hydrocarbons. Naturally occurring hydrocarbons are extracted from deep under the Earth's surface where they exist in both the liquid form, known as petroleum, or crude oil, and the gaseous form, known as natural gas.



▲ **Figure 14.1.1** Diagram showing the occurrence of petroleum and natural gas under the surface of the Earth

Natural gas

Natural gas is a mixture of four alkanes: methane (CH_4), ethane (C_2H_6), propane (C_3H_8) and butane (C_4H_{10}). Methane makes up about 80% of natural gas and the other three hydrocarbons make up 20%. The propane and butane are removed before natural gas is sold commercially. Commercial natural gas, therefore, is a mixture of methane and ethane. This natural gas can then be used to generate electricity, it can be piped to homes and used for cooking and heating, or it can be liquefied under pressure to produce liquefied natural gas, or LNG. LNG is used in place of petrol in vehicles in some countries, e.g. Germany.

After removal, the propane and butane are also liquefied under pressure and used to produce liquefied petroleum gas, or LPG. LPG is also known as 'bottled gas' and many of you use it for cooking.

? Did you know?

Petroleum is a finite, non-renewable resource and once all the existing supplies within the Earth's crust have been used they will not be replaced.

Petroleum

Petroleum is an oily liquid which is a complex mixture of many hydrocarbons, mainly alkanes and ringed hydrocarbons. The smallest gaseous hydrocarbons and the largest solid hydrocarbons are dissolved in the oily liquid. The mixture can contain hydrocarbons with as many as 70 or more carbon atoms in their molecules. To be useful, petroleum must be separated into its different components, or fractions, as they are called. This is done by **fractional distillation** at an oil refinery.

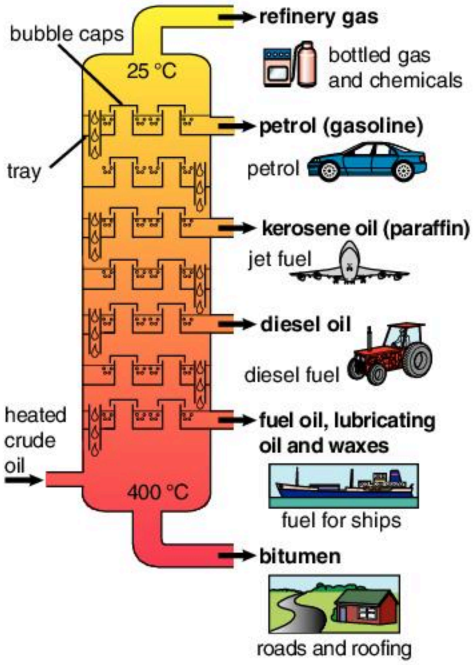
Fractional distillation of petroleum

When it is first removed from under the ground, petroleum contains many impurities. Before it can be separated into its components, these impurities have to be removed. Once the impurities have been removed, the petroleum is heated to about 400 °C and the liquid and vapour produced are piped into the lower part of a **fractionating tower**. The fraction which remains as a viscous liquid at 400 °C sinks to the bottom of the tower and is tapped off forming **bitumen** or asphalt.

The vapours rise up the tower and pass through a series of bubble caps and trays. The temperature of the tower *decreases* upwards. The vapours of the larger hydrocarbons with the higher boiling points condense on the bubble caps in the lower part of the tower, and those of the smaller hydrocarbons with the lower boiling points rise further and condense at higher levels. The liquids produced are tapped off from the trays at the different levels in the tower. Gases that do not condense at the temperatures within the tower are removed as **refinery gas** at the top of the tower.

Each fraction is a mixture of hydrocarbons of similar molecular size and boiling points within a specific temperature range as shown in Table 14.1.1. Petrol is tapped off from the top trays, kerosene and diesel oil are tapped off from the middle trays and fuel oil, lubricating oil and waxes are tapped off from the lower trays.

Table 14.1.1 summarises the fractions which are obtained by fractional distillation of petroleum and some of their uses.



◀ **Figure 14.1.2** Fractional distillation of petroleum

▼ **Table 14.1.1** Uses of the different fractions obtained from petroleum

Name of fraction	Boiling point/°C	Number of carbon atoms	Uses
Refinery gas	below 25	1–4	Fuel for domestic use, e.g. cooking and heating. To manufacture a variety of petrochemicals.
Petrol (gasoline)	25–170	5–10	Fuel for internal combustion engines, e.g. cars and light aircraft. As solvents.
Kerosene oil (paraffin oil)	170–250	10–14	Fuel for cooking, heating, kerosene lamps and jet engines. Cracked into smaller hydrocarbons.
Diesel oil	250–350	14–20	Fuel for diesel engines, e.g. cars, lorries, buses, trains, trucks and generators. Cracked into smaller hydrocarbons.
Fuel oil, lubricating oil and waxes	350–400	18–30	Fuel oils are used as fuel for factory boilers, ships and power stations. Lubricating oils are used to lubricate mechanical parts in machinery and vehicles. Waxes are used to make polishing waxes, wax paper, petroleum jelly and candles.
Bitumen	above 400	more than 30	Road surfacing. Roofing.

Cracking hydrocarbons

When petroleum is fractionally distilled into its different fractions, two problems arise.

- The fractions which contain the smaller hydrocarbon molecules, e.g. petrol, are in much greater demand in today's world than the fractions containing the larger hydrocarbon molecules, e.g. diesel oil. However, the distillation of petroleum produces an excess of the larger hydrocarbons and insufficient of the smaller ones to meet modern demands.
- Most of the hydrocarbons obtained from petroleum have only carbon–carbon single bonds in their molecules and, apart from burning in air to release energy, they are fairly unreactive. To be used in the petrochemical industry to manufacture the great variety of other organic compounds that we use every day, they need to be converted into more reactive compounds which have carbon–carbon double bonds.

Both these problems can be overcome by breaking down the larger hydrocarbon molecules into smaller, more useful molecules by a process known as **cracking**.

Cracking can be carried out in one of two ways, **thermal cracking** and **catalytic cracking**.

? Did you know?

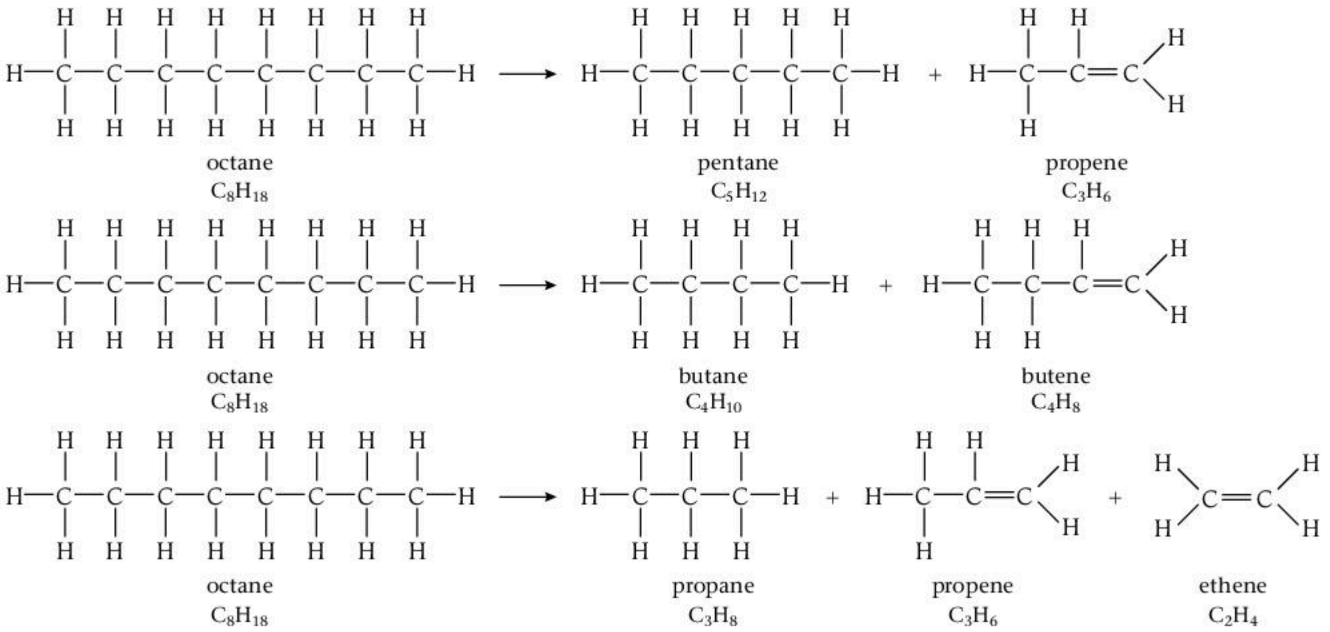
The **Pitch Lake** in Trinidad is the largest natural deposit of asphalt in the world. Located in south-west Trinidad, the lake covers about 40 hectares and is over 75 m deep. The lake serves as a tourist attraction and is also mined for its asphalt. Sir Walter Raleigh was introduced to the lake by Amerindian guides in 1595 and he immediately recognised its potential and started caulking his ships with the tar.

! Key fact

Cracking is the process by which long-chain hydrocarbons are broken down into shorter-chain hydrocarbons.

Thermal cracking

During **thermal cracking** temperatures of above 700 °C and pressures up to 70 atmospheres are used to break the larger alkanes into smaller ones. Under these conditions, the long-chain alkane molecules vibrate strongly enough to break. Thermal cracking of an alkane molecule always forms a shorter-chain alkane and at least one alkene. The molecules are broken up in a fairly random way so thermal cracking always gives a mixture of products containing a fairly high proportion of alkene molecules. For example, octane (C_8H_{18}) can be cracked in a variety of different ways, three of these are illustrated in Figure 14.1.3.



▲ **Figure 14.1.3** Thermal cracking of octane

Catalytic cracking

During **catalytic cracking** the long-chain alkanes are heated to temperatures of about 500 °C at fairly low pressures in the presence of a catalyst. The catalysts are normally synthetic mixtures of aluminium oxide and silicon dioxide known as zeolites. Catalytic cracking is the most important source of petrol and raw materials for the petrochemical industry.

Natural gas and oil in the Caribbean

The Republic of Trinidad and Tobago is the leading Caribbean producer of oil and natural gas, with these accounting for about 40% of gross domestic product and a large percentage of exports.

The oil refinery in Trinidad is found at Pointe-à-Pierre and accounts for most of the prosperity of the region. All the natural gas produced is used domestically. More than two-thirds of it is piped to methanol, urea and ammonia plants and to the electricity sector.

▲ **Figure 14.1.4** Pointe-à-Pierre oil refinery

Did you know?

The Pointe-à-Pierre oil refinery is the only oil refinery in the world to co-exist with a **wildlife reserve**. The Pointe-à-Pierre Wild Fowl Trust is a not-for-profit environmental organisation dedicated to environmental education and conservation of wetlands and waterfowl. It consists of two lakes and about 25 hectares of land within the oil refinery.

Summary questions

1. Name two natural sources of hydrocarbons.
2. List the main uses of three of the fractions obtained from the fractional distillation of petroleum.
3. Describe what is meant by 'cracking' and distinguish between thermal cracking and catalytic cracking of alkanes.
4. Give the formulae and names of two products which could be obtained by cracking pentane (C_5H_{12}).

B14.2 Alkanes: C_nH_{2n+2}

Objectives

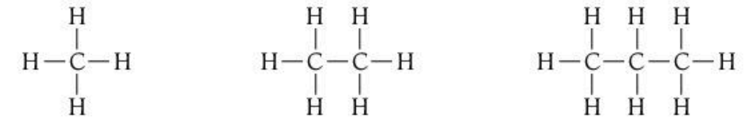
By the end of this topic you will be able to:

- give the names and structural formulae of unbranched and branched alkanes up to six carbon atoms
- relate the reactivity of alkanes to the presence of single bonds
- describe the combustion reaction of alkanes
- describe substitution reactions of alkanes with halogens
- explain why alkanes undergo substitution reactions
- relate the uses of alkanes to their properties
- describe the production of biogas.

Alkanes form a homologous series with the general formula $\mathbf{C_nH_{2n+2}}$.

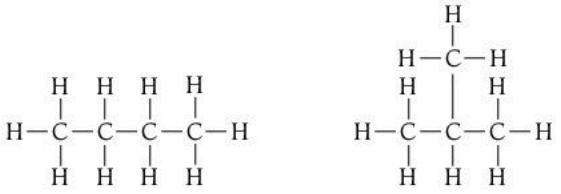
Alkanes are saturated hydrocarbons. This means that they contain only single bonds between carbon atoms. Alkanes are relatively unreactive because of the strong carbon–carbon single bonds and carbon–hydrogen bonds in the molecules.

Alkanes are named by using the prefix which indicates the total number of carbon atoms in the molecule and the suffix **'-ane'**. The first three alkanes in the series do not show structural isomerism. Their names and formulae are given in Figure 14.2.1.

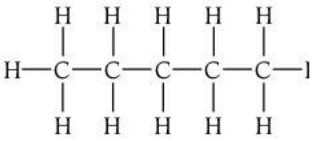

methane (CH_4) ethane (C_2H_6) propane (C_3H_8)

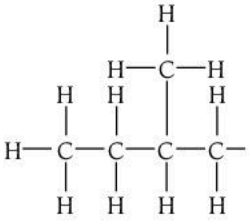
▲ **Figure 14.2.1** Structural formulae of the first three alkanes

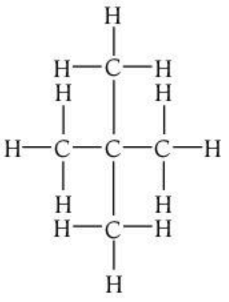
Alkanes with *four or more* carbon atoms show structural isomerism resulting from their ability to form branched chains. Following steps 1) to 4) on page 233 for naming branched chain isomers, Figures 14.2.2, 14.2.3 and 14.2.4 show the names and structural formulae of the isomers of butane (C_4H_{10}), pentane (C_5H_{12}) and hexane (C_6H_{14}).


butane 2-methylpropane

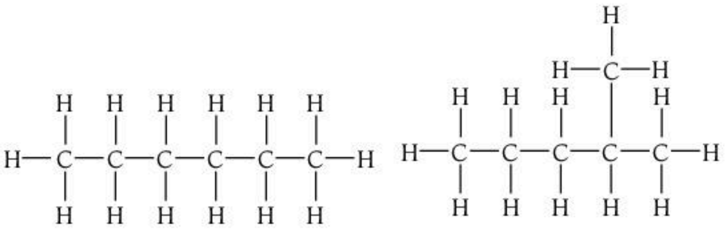
▲ **Figure 14.2.2** The structural isomers of butane (C_4H_{10})

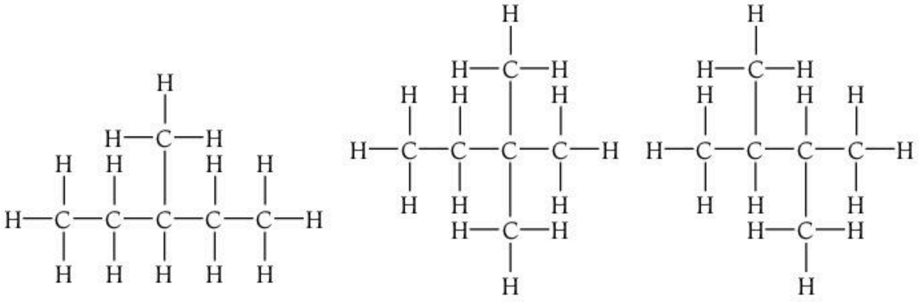

pentane


2-methylbutane


2,2-dimethylpropane

▲ **Figure 14.2.3** The structural isomers of pentane (C_5H_{12})


hexane 2-methylpentane


3-methylpentane 2,2-dimethylbutane 2,3-dimethylbutane

▶ **Figure 14.2.4** The structural isomers of hexane (C_6H_{14})

The melting and boiling points of the alkanes *increase* with an increase in the size of the molecules, i.e. with an increase in molar mass. The first four alkanes, methane, ethane, propane and butane, are gases at room temperature.

Alkanes with 5 to 16 carbon atoms in their molecules are liquids, while those with 17 or more carbon atoms are solids. Alkanes become less volatile as their molecular size increases.

Alkanes are *relatively unreactive* and only undergo a few types of reaction, namely combustion, substitution and cracking. The cracking of alkanes was described in Unit 14.1. We will now look at the other two reactions.

Combustion reactions of alkanes

Alkanes burn very easily in air or oxygen. The reaction is exothermic producing large amounts of heat energy, which is why alkanes are used extensively as fuels. The products of the **combustion reaction** vary depending on the oxygen supply. If the oxygen supply is plentiful, they undergo complete combustion, if it is limited, they undergo incomplete combustion.

When alkanes burn they do so with a *clear blue flame* because the ratio of carbon to hydrogen atoms in the molecules is low. All the carbon is therefore converted to carbon dioxide or carbon monoxide and no unreacted carbon remains to give the flame a yellow, smoky appearance. The flame, therefore, has a very low soot content. Alkanes are regarded as clean burning fuels.

Complete combustion of alkanes

When the oxygen supply is plentiful, alkanes burn completely, i.e. they undergo **complete combustion**. The products of the complete combustion of alkanes are always carbon dioxide and water as steam. The complete combustion of methane and ethane is shown in the following equations:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) \qquad \Delta H \ -\text{ve}$$
$$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g) \qquad \Delta H \ -\text{ve}$$

Incomplete combustion of alkanes

When the oxygen supply is limited, alkanes undergo **incomplete combustion**. The products of the incomplete combustion of alkanes are carbon monoxide and water as steam. The incomplete combustion of methane is shown in the following equation:

$$2CH_4(g) + 3O_2(g) \longrightarrow 2CO(g) + 4H_2O(g) \qquad \Delta H \ -\text{ve}$$

Substitution reactions of alkanes

Under the right conditions, alkanes can react with halogens. In alkane molecules, the carbon atoms are bonded to the maximum number of hydrogen atoms possible. When a reaction occurs between an alkane and a halogen, the hydrogen atoms are replaced by halogen atoms. These reactions are known as **substitution reactions** because halogen atoms take the place of the hydrogen atoms in the alkane molecules, one hydrogen atom at a time. They can also be called **halogenation**.

Energy in the form of *light* is required for the reactions to occur. Ultraviolet light works the best, but often sunlight is sufficient. A substitution reaction takes place in stages and the speed of the reaction depends on *three* factors.

- Light intensity – the brighter the light, the faster the reaction.
- The reactivity of the halogen – the more reactive the halogen, the faster the reaction. Fluorine, being the most reactive halogen, has the most rapid reaction. Iodine does not react with alkanes to any extent.
- The reactivity of the alkane – the smaller the alkane molecule, the faster the

? Did you know?

The prefix '**normal**' or '**n-**' is sometimes used before the name when naming the straight chain isomers. The straight chain isomer of butane is sometimes called normal butane or n-butane. Similarly, the straight chain isomer of pentane can be called normal pentane or n-pentane.

✓ Exam tip

It is important that you can give the fully displayed structural formulae and names of unbranched and branched chain isomers of any alkane containing up to six carbon atoms if you are given its molecular formula.

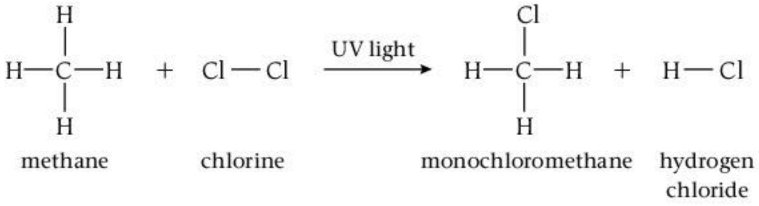
✓ Exam tip

Balancing the combustion reactions of alkanes can be difficult. Begin by balancing the carbon atoms followed by the hydrogen atoms. Leave the oxygen atoms until last. After balancing the carbon and the hydrogen, if you find that you have an odd number of oxygen atoms on the right, multiply each coefficient (including the unwritten 1) by 2, before balancing the oxygen.

reaction. Methane, being the smallest molecule, has the most rapid reaction.

This is seen in the reaction between methane and chlorine.

No reaction occurs in the dark. In bright light the reaction is rapid. In dim light a *slow substitution* reaction occurs in stages, one hydrogen atom being replaced by one chlorine atom at a time. The first stage of the reaction produces chloromethane and hydrogen chloride gas:



or

$$CH_4(g) + Cl_2(g) \xrightarrow{\text{UV light}} CH_3Cl(g) + HCl(g)$$

As the reaction continues, the hydrogen atoms are replaced one by one until tetrachloromethane (CCl_4) is produced:

$$CH_3Cl(g) + Cl_2(g) \xrightarrow{\text{UV light}} \underset{\text{dichloromethane}}{CH_2Cl_2(l)} + HCl(g)$$

$$CH_2Cl_2(l) + Cl_2(g) \xrightarrow{\text{UV light}} \underset{\text{trichloromethane}}{CHCl_3(l)} + HCl(g)$$

$$CHCl_3(l) + Cl_2(g) \xrightarrow{\text{UV light}} \underset{\text{tetrachloromethane}}{CCl_4(l)} + HCl(g)$$

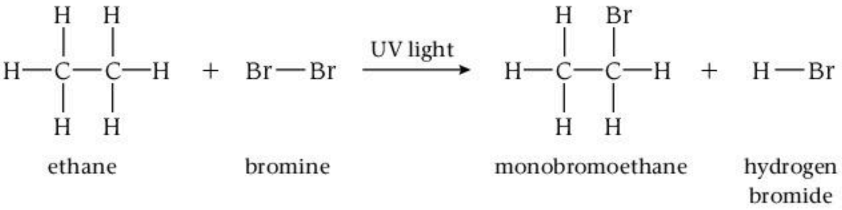
The overall reaction can be summarised by the following equation:

$$CH_4(g) + 4Cl_2(g) \xrightarrow{\text{UV light}} CCl_4(l) + 4HCl(g)$$

The products of the halogenations of alkanes are called **haloalkanes** or **alkyl halides**.

The reaction of bromine vapour or bromine solution with methane is similar, but since bromine is less reactive than chlorine, the reaction proceeds more slowly. Once again, energy in the form of ultraviolet light is required for the reaction to proceed.

During the reaction between bromine and any alkane, the colour of the bromine *slowly* fades from *red-brown* to *colourless*. The first step in the reaction between ethane and bromine is shown in the equation below:



? Did you know?

Trichloromethane (or chloroform) and **tetrachloromethane** (or carbon tetrachloride) are both volatile, 'sweet' smelling liquids. At one time, trichloromethane was used as a general anaesthetic because the vapours depress the nervous system, and tetrachloromethane was used in fire extinguishers and as a solvent in dry cleaning. Their uses were discontinued when it was discovered how toxic they both were.

Uses of alkanes

Because of their properties, alkanes have a variety of uses. The main uses of alkanes, related to their properties, are given below.

- Alkanes are used as *fuels* for several reasons:
  - they burn easily when ignited
  - they release a large amount of energy when they undergo combustion, i.e. the combustion of an alkane is an exothermic reaction
  - they are clean burning fuels, i.e. they burn with a clear blue flame which has a very low soot content
  - they are stored easily as liquids in containers, which makes them easy to transport.
- Alkanes are used as *solvents*. Alkanes are non-polar molecules therefore they dissolve a variety of other non-polar solutes. For example, hexane and heptane are used as solvents in the manufacture of fast-drying lacquers, glues and other adhesives, petroleum jelly is oil dissolved in paraffin wax, a solid alkane.

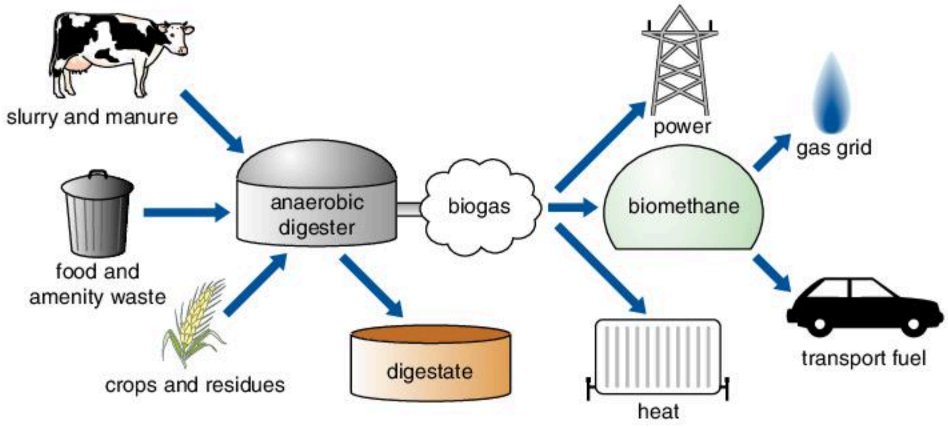
Biogas production

Biogas refers to a gas produced by the breakdown of organic matter in the absence of oxygen. It is a renewable energy source which can be produced from any organic matter including manure and other farmyard animal waste, crop residues, sewage, garden waste, food waste and organic waste from food industries.

The organic waste is put into a sealed anaerobic digester where naturally occurring anaerobic microorganisms, e.g. bacteria, digest it in the absence of oxygen into biogas. Biogas is a mixture of approximately 60% methane, 40% carbon dioxide and traces of other contaminant gases such as hydrogen sulfide. The exact composition depends on the nature of the original organic matter.

Biogas can be used directly as a fuel to provide heat and generate electricity or it can be converted to almost pure methane, often called **biomethane**, by removal of the other gases. This biomethane can then be piped to homes with natural gas or liquefied under pressure to produce liquefied natural gas (LNG) for use in vehicles designed to run on LNG.

The leftover indigestible material and dead microorganisms, known as **digestate**, can then be used as a fertiliser. The production of biogas reduces demands on non-renewable fossil fuels, recycles soil nutrients and reduces the amount of waste going to garbage dumps and landfills.



▲ **Figure 14.2.5** Production of biogas

Summary questions

1. Give the names and fully displayed structural formulae of three isomers of hexane.
2. Write a balanced equation to show the complete combustion of butane.
3. Propane reacts with chlorine gas.
   - a What condition is necessary for the reaction to occur?
   - b What is the name given to this type of reaction?
   - c Use structural formulae to show the first step in the reaction.
   - d Name the products of the first step in the reaction.
4. Explain why the main use of alkanes is as fuels.

Objectives

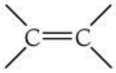
By the end of this topic you will be able to:

- recognise alkenes by the presence of a carbon–carbon double bond
- give the names and structural formulae of unbranched alkenes up to six carbon atoms
- relate the reactivity of alkenes to the presence of a carbon–carbon double bond
- describe the combustion reaction of alkenes
- explain why alkenes undergo addition reactions
- describe addition reactions of alkenes with hydrogen, the halogens, the hydrogen halides and water
- describe reactions which can be used to distinguish between an alkane and an alkene
- relate the uses of alkenes to their properties.

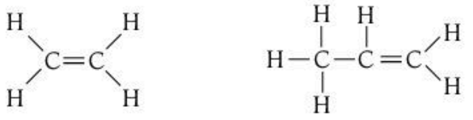
B14.3 Alkenes: C_nH_{2n}

Alkenes form a homologous series with the general formula $\mathbf{C_nH_{2n}}$.

Alkenes are unsaturated hydrocarbons. They contain one double bond between two carbon atoms, i.e. their functional group is the carbon–carbon double bond:



Alkenes are named by using the prefix which indicates the total number of carbon atoms in the molecule and the suffix '**-ene**'. The first two alkenes in the series do not show structural isomerism. Their names and formulae are given in Figure 14.3.1.



ethene (C_2H_4)

propene (C_3H_6)

▲ **Figure 14.3.1** Structural formulae of the first two alkenes

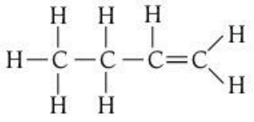
Alkenes with *four or more* carbon atoms show structural isomerism. Structural isomers of alkenes result from the following:

- the change in position of the functional group, i.e. the change in position of the carbon–carbon double bond (it is important to note, however, that the bond must always be horizontal)
- branching of the molecule.

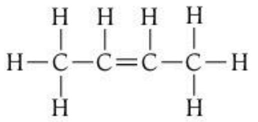
When naming isomers of alkenes:

- number the carbon atoms in the longest chain from the end closest to the double bond so that the double bond is closest to the lowest number and indicate the position of the double bond in the name using the lowest possible number
- follow steps 1) to 4) on page 233 for naming branched chain isomers.

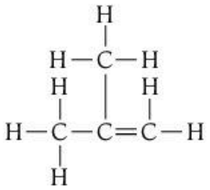
Using the rules above, the structures and names of the unbranched and branched isomers of butene (C_4H_8) and pentene (C_5H_{10}) and the unbranched isomers of hexene (C_6H_{12}) are given in Figures 14.3.2, 14.3.3 and 14.3.4.



but-1-ene
(or 1-butene)

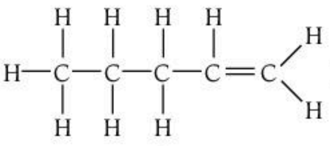


but-2-ene
(or 2-butene)

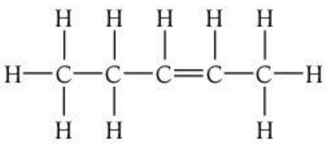


2-methylpropene

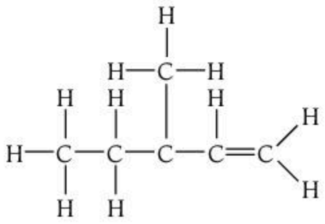
▲ **Figure 14.3.2** The structural isomers of butene (C_4H_8)



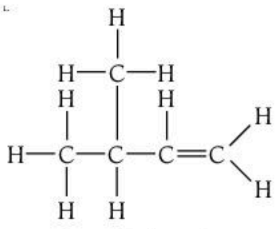
pent-1-ene
(or 1-pentane)



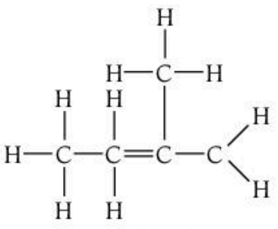
pent-2-ene
(or 2-pentene)



2-methylbut-1-ene
(or 2-methyl-1-butene)

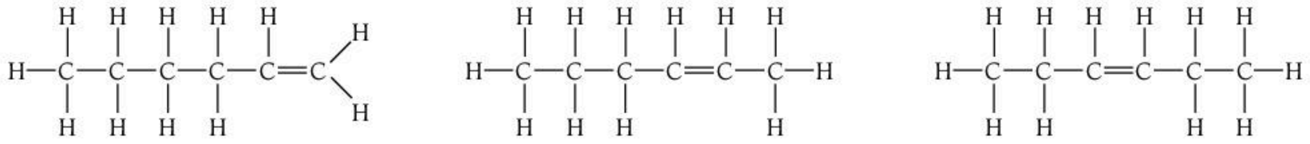


3-methylbut-1-ene
(or 3-methyl-1-butene)



2-methylbut-2-ene
(or 2-methyl-2-butene)

▶ **Figure 14.3.3** The structural isomers of pentene (C_5H_{10})


hex-1-ene (or 1-hexene)

hex-2-ene (or 2-hexene)

hex-3-ene (or 3-hexene)

▲ **Figure 14.3.4** The straight chain isomers of hexene (C_6H_{12})

The presence of the double bond makes alkenes *more reactive* than alkanes. Alkenes undergo a variety of different reactions. The ones that we will be looking at are combustion and addition reactions.

Combustion reactions of alkenes

Alkenes burn very easily in air or oxygen. The **complete combustion** of an alkene produces carbon dioxide and water as steam. The reaction is exothermic, producing a large amount of heat energy. The complete combustion of ethene is shown in the equation below:

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g) \qquad \Delta H\ -\text{ve}$$

Alkenes require more oxygen for complete combustion than the corresponding alkane with the same number of carbon atoms because alkenes have a higher ratio of carbon atoms to hydrogen atoms in their molecules. Because of this higher ratio of carbon to hydrogen atoms, alkenes do not burn with clean flames. When they burn they release unreacted carbon as soot and sometimes carbon monoxide as well as carbon dioxide and steam. As a result of the presence of the soot in the flame, alkenes burn with a *smoky flame*.

Addition reactions of alkenes

Addition reactions of alkenes involve adding extra atoms to the molecules, with one bond in the double bond being broken in the process. They are called addition reactions because *two* molecules react to form *one* molecule. Unlike substitution reactions, the alkene does not lose any hydrogen atoms.

The addition reaction which an alkene undergoes is summarised in Figure 14.3.5, where A–B represents a molecule with a single covalent bond between two atoms or small groups of atoms in a molecule. During the reaction, the covalent bond between A and B breaks and one of the bonds in the carbon–carbon double bond also breaks. The double bond is converted to a single bond forming a **saturated compound** and A becomes bonded to the carbon atom at one side of the original double bond and B becomes bonded to the carbon atom at the other side.

Addition reaction with hydrogen

Alkenes undergo a type of addition reaction with hydrogen gas known as **hydrogenation**. During hydrogenation, the single bond between the hydrogen atoms in the hydrogen molecule breaks and one atom is bonded to each carbon atom at either side of the original double bond in the alkene. This always produces an **alkane**.

The following conditions are required for the reaction:

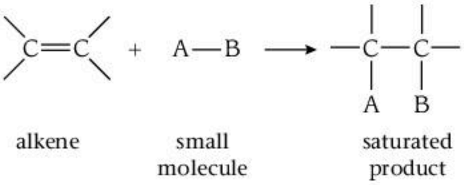
- a nickel catalyst
- a pressure of about 5 atmospheres
- a temperature of about 150 °C.

Exam tip

It is important that you can name the unbranched isomers of any alkene containing up to *six* carbon atoms. It is also important that you can draw the fully displayed structural formulae of unbranched and branched isomers of any alkene containing up to *six* carbon atoms if you are given its molecular formula. It is not essential that you can name the branched isomers of alkenes.

Did you know?

Hexene (C_6H_{12}) has thirteen structural isomers, the three unbranched chain isomers shown in Figure 14.3.4 and ten branched chain isomers. For practice, you could see how many of these you can draw.

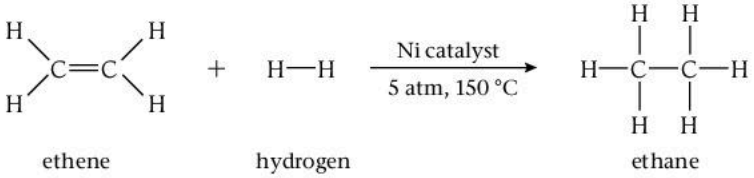


▲ **Figure 14.3.5** Summary of addition reactions of alkenes

Did you know?

Hydrogenation is used in industry to ‘harden’ fats and oils, i.e. increase their melting points. For example, margarine can be made by hydrogenating the carbon–carbon double bonds in vegetable oils. This hardens the oils making the margarine more spreadable and it also has possible health benefits. If margarine is used in place of butter, it reduces the amount of saturated animal fats in the diet.

For example, ethene reacts with hydrogen to form ethane:



or

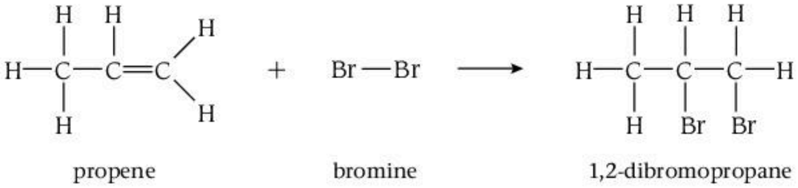
$$C_2H_4(g) + H_2(g) \xrightarrow[\text{5 atm, 150 °C}]{\text{Ni catalyst}} C_2H_6(g)$$

Addition reaction with halogens

Alkenes undergo a type of addition reaction with halogens known as **halogenation**. During halogenation, the single bond between the halogen atoms in the molecule breaks and one atom is bonded to each carbon atom at either side of the original double bond in the alkene. This always produces a **haloalkane**.

In contrast to the halogenation reactions which alkanes undergo, this reaction does not need light energy; it can take place in the dark at room temperature. The reaction is also a *rapid* reaction in contrast to the slow reaction that occurs with alkanes.

For example, propene reacts with bromine vapour to form 1,2-dibromopropane:



or

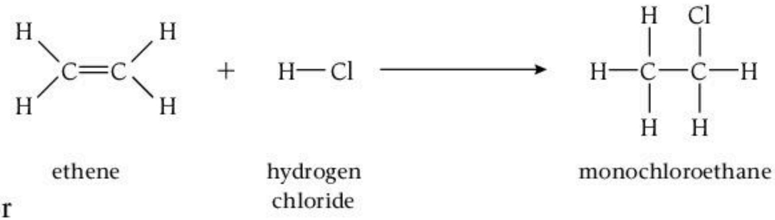
$$C_3H_6(g) + Br_2(g) \longrightarrow C_3H_6Br_2(l)$$

In this reaction, bromine can be in the form of a vapour or a solution. In both cases, the *red-brown* bromine is *rapidly decolourised* as the alkene reacts with it.

Addition reactions with hydrogen halides

Alkenes undergo addition reactions with hydrogen halides, for example, hydrogen chloride (HCl) and hydrogen bromide (HBr), at room temperature to produce **haloalkanes**.

For example, ethene reacts with hydrogen chloride to form monochloroethane:



or

$$C_2H_4(g) + HCl(g) \longrightarrow C_2H_5Cl(l)$$

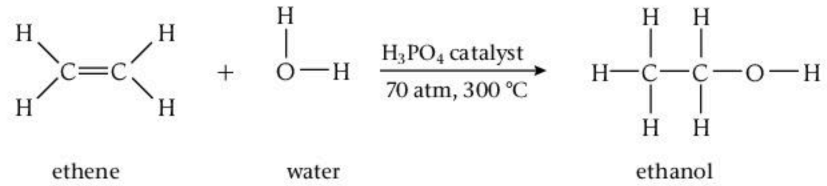
Addition reactions with water

Alkenes undergo a type of addition reaction with water in the form of *steam* known as **hydration**. During hydration, one of the bonds between a hydrogen atom and the oxygen atom in the water molecule breaks and the hydrogen atom is bonded to the carbon atom at one side of the original double bond in the alkene and the remaining OH group is bonded to the carbon atom at the other side. This always produces an **alcohol**.

The following conditions are required for the reaction:

- a catalyst of phosphoric acid in sand (silica)
- a pressure of 60–70 atmospheres
- a temperature of about 300 °C.

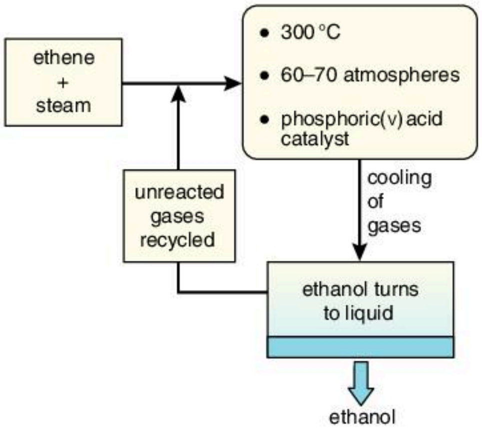
For example, ethene reacts with water, in the form of steam, to form ethanol:



or

$$C_2H_4(g) + H_2O(g) \xrightarrow[70\ \text{atm},\ 300\ °C]{H_3PO_4\ \text{catalyst}} C_2H_5OH(l)$$

Ethanol is produced industrially by the addition reaction between ethene and steam using the conditions listed above. The yield of the reaction is very low so the unreacted ethene is recycled back into the reaction vessel to increase the yield. The flow diagram in Figure 14.3.6 shows the industrial preparation of ethanol.

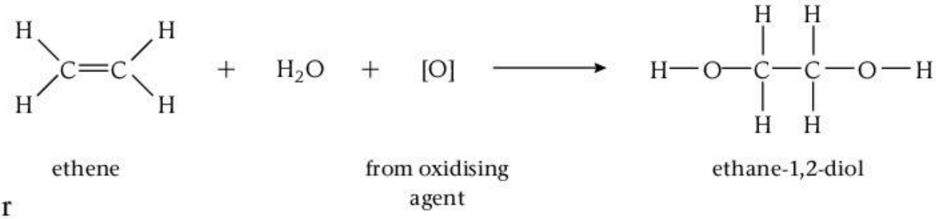


▲ **Figure 14.3.6** The industrial preparation of ethanol

Addition reaction with acidified potassium manganate(VII)

Alkenes undergo an addition reaction with acidified potassium manganate(VII) solution. The reaction is a redox reaction in which the alkene is oxidised by the acidified potassium manganate(VII) to form a **dialcohol**, i.e. an alcohol with two –OH groups.

For example, ethene reacts with acidified potassium manganate(VII) to form ethane-1,2-diol:



or

$$C_2H_4(g) + H_2O(l) + [O] \longrightarrow C_2H_4(OH)_2(l)$$

During the reaction, the potassium manganate(VII) solution rapidly changes colour from *purple* to *colourless*. The alkene acts as a reducing agent and reduces the purple manganate(VII) ion (MnO_4^-) to the colourless manganese(II) ion (Mn^{2+}).

Exam tip

It is very important that you can write equations for the addition reactions of alkenes and that you can state the **conditions** required for the reactions to occur.

Did you know?

Alkenes can also add to themselves in **addition polymerisation** reactions. These reactions make very large molecules known as polymers. We will be looking at this in more detail in Unit 16.

Reactions to distinguish between an alkane and an alkene

You have already learnt that both alkanes and alkenes are hydrocarbons, but alkanes are saturated whereas alkenes are unsaturated. Alkanes undergo slow substitution reactions because they contain only carbon–carbon single bonds whereas alkenes undergo rapid addition reactions because of the presence of a carbon–carbon double bond.

The difference between the reactions of alkanes and alkenes with bromine solution or acidified potassium manganate(VII) solution is used in the laboratory to distinguish between the two. Both reactions are testing to see if the compound contains a carbon–carbon double bond, i.e. if it is **unsaturated**.

The tests used to distinguish between an alkane and an alkene are summarised in Table 14.3.1.

▼ **Table 14.3.1** Distinguishing between an alkane and an alkene

Test	**Observations and explanation**	
	Alkane	**Alkene**
Add acidified potassium manganate(VII) solution	The solution remains purple. Alkanes do not react with acidified potassium manganate(VII) solution.	The solution changes colour *rapidly* from purple to colourless. Alkenes rapidly reduce the purple MnO_4^- ion to the colourless Mn^{2+} ion.
Add a solution of bromine in water (bromine water) or in tetrachloromethane under standard laboratory conditions	The solution remains red-brown. Alkanes do not react with bromine solution under standard laboratory conditions.	The solution changes colour *rapidly* from red-brown to colourless. Alkenes undergo a rapid addition reaction with bromine solution under any conditions, including darkness.

It is important to note the following.

- In the presence of ultraviolet light, alkanes cause a solution of bromine in water or tetrachloromethane to slowly change colour from red-brown to colourless due to the slow substitution reaction occurring.
- Both alkanes and alkenes react with bromine vapour, changing it from red-brown to colourless. Alkanes slowly decolourise the vapour in ultraviolet light due to the slow substitution reaction occurring. Alkenes rapidly decolourise the vapour due to the rapid addition reaction and this reaction does not need light, therefore it will occur in darkness.

These reactions are, however, not usually used as laboratory tests to distinguish an alkane and an alkene.

Practical activity

To distinguish between an alkane and an alkene

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with samples of two hydrocarbons cyclohexane and cyclohexene, acidified potassium manganate(VII) solution, bromine solution and test tubes.

Method

1. Place 2 cm^3 of cyclohexane into each of two test tubes. Label the tubes A and B.
2. Place 2 cm^3 of cyclohexene into each of two test tubes. Label the tubes C and D.
3. Add a few drops of acidified potassium manganate(VII) solution to tubes A and C.
4. Shake and observe any colour changes which occur.
5. Add a few drops of bromine solution to test tubes B and D.
6. Shake and observe any colour changes.
7. Which hydrocarbon is saturated and which is unsaturated?

Uses of alkenes

Alkenes are extremely important *starting materials* for the manufacture of a wide range of chemicals because of their ability to undergo addition reactions due to the presence of the carbon–carbon double bond. For example, they are used to manufacture ethanol and other alcohols because they undergo addition reactions with steam and to manufacture various haloalkanes because of their addition reactions with halogens and hydrogen halides. They are also used to manufacture antifreezes, such as ethane-1,2-diol and synthetic rubbers. Because they can undergo addition reactions with themselves they are used to manufacture a wide variety of polymers, or plastics. You will learn more about plastics in Unit 16.

? Did you know?

The combustion of alkenes is an exothermic reaction, therefore, alkenes could be used as fuels. However, they are far too important as starting materials for making other organic products to waste them by using them as fuels.

Summary questions

1. Give the names and fully displayed structural formulae of the three straight chain isomers of hexene.
2. a Describe what you would observe when bromine solution is added to ethene.
   b Give a balanced equation for the reaction using structural formulae.
3. Give the conditions under which the following reactions occur:
   a hydrogenation of ethene to produce ethane
   b hydration of ethane to produce ethanol.
4. Describe an experiment that you could perform in the laboratory to distinguish between an alkane and an alkene.

Key concepts

- **Hydrocarbons** are organic compounds composed of carbon and hydrogen atoms only.
- **Natural gas** and **petroleum** are natural sources of hydrocarbons.
- Natural gas is composed mainly of methane with small amounts of ethane, propane and butane.
- Petroleum is a complex mixture mainly of alkanes and ringed hydrocarbons.

- Separation of petroleum into its components is done by **fractional distillation** at an oil refinery.
- The main uses of the fractions obtained from the fractional distillation of petroleum are as fuels, in the manufacture of petrochemicals, as lubricants and in surfacing roads.
- The larger molecules produced during fractional distillation of petroleum can be broken down into smaller, more useful molecules by **cracking**.
- **Thermal cracking** uses high temperature and pressures to break the molecules down.
- **Catalytic cracking** uses fairly high temperatures, low pressures and a catalyst to break the molecules down.
- Cracking of an alkane produces at least one alkene.
- **Alkanes** are **saturated hydrocarbons** with the general formula $\mathbf{C_nH_{2n+2}}$.
- Alkanes with four or more carbon atoms display structural isomerism.
- Alkanes burn in air or oxygen with a clear blue flame.
- The complete combustion of alkanes is an exothermic reaction and produces carbon dioxide and water as steam. The incomplete combustion produces carbon monoxide and steam.
- Alkanes undergo **substitution reactions** with halogens in the presence of light.
- Substitution reactions take place in stages, one hydrogen atom being replaced by one halogen atom at a time.
- Alkanes are used mainly as fuels and organic solvents.
- **Biogas** is a renewable source of methane.
- **Alkenes** are **unsaturated hydrocarbons** with the general formula $\mathbf{C_nH_{2n}}$.
- Alkenes have the carbon–carbon double bond as their functional group.
- Alkenes with four or more carbon atoms display structural isomerism.
- Alkenes burn in air or oxygen with a smoky flame.
- The complete combustion of alkenes is an exothermic reaction and produces carbon dioxide and water as steam.
- Alkenes undergo **addition reactions** in which other atoms are added to the molecule. In the process one bond in the double bond breaks.
- The addition reaction with hydrogen is known as **hydrogenation** and it forms an alkane. It requires a nickel catalyst, a pressure of 5 atm and a temperature of 150 °C.
- The addition reaction with the halogens is known as **halogenation** and it forms a haloalkane.
- The addition reaction with the hydrogen halides also forms a haloalkane.
- The addition reaction with water in the form of steam is known as **hydration** and it forms an alcohol. It requires a phosphoric acid in sand catalyst, a pressure of 60–70 atm and a temperature of 300 °C.
- Alkenes change the colour of bromine solution from red-brown to colourless under standard laboratory conditions, whereas alkanes do not.
- Alkenes change the colour of acidified potassium manganate(VII) solution from purple to colourless, whereas alkanes do not.
- Alkenes are used mainly as starting materials for the manufacture of a wide range of other chemicals because of their ability to undergo addition reactions.

Practice exam-style questions

Multiple-choice questions

1 Natural sources of hydrocarbons include:

I petroleum
II natural gas
III coal

A I and II only
B I and III only
C II only
D I, II and III

2 Which of the following could not be obtained by cracking hexane?

A methane and pentene
B propene and propane
C butane and ethane
D methane, propene and ethene

3 What are the main uses of bitumen?

A As a fuel and for surfacing roads.
B To manufacture petrochemicals and as a fuel.
C As a roofing material and to surface roads.
D As a lubricant and as a roofing material.

4 The different fractions which make up crude oil are separated based on:

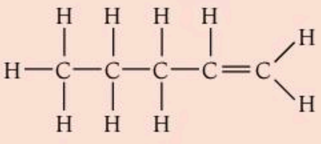
A their different melting points
B their different boiling points
C their different densities
D their different solubilities in water

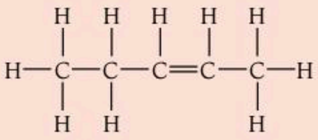
5 When ethane reacts with bromine vapour which of the following is/are true?

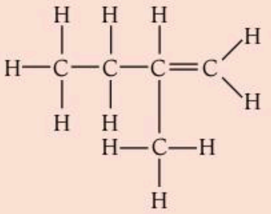
I a substitution reaction occurs
II the bromine is rapidly decolourised
III only one product is produced

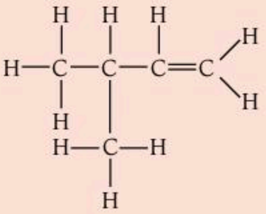
A I only
B I and II only
C I and III only
D II and III only

6 The following are all structural isomers of pentene except:


A


B


C


D

7 The process of converting ethene to ethane can be described as:

I hydrogenation
II hydration
III addition

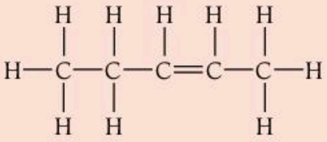
A I and II only
B I and III only
C II and III only
D I, II and III

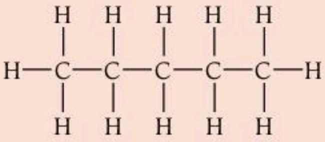
8 Which of the following could best be used to distinguish propane from propene?

A Burn both in air.
B React both with bromine vapour.
C React both with hydrogen gas.
D React both with acidified potassium manganate(VII) solution.

Structured question

9 The fully displayed structural formulae of two hydrocarbons, **A** and **B** are given below:


A


B

a To which homologous series does EACH compound belong? *(2 marks)*

b Draw THREE structures to represent isomers of compound **A**. *(3 marks)*

c **A** and **B** both react with chlorine but under different conditions.

- **i)** Give ONE condition necessary for **B** to react with chlorine. *(1 mark)*
- **ii)** Draw the fully displayed structure of the compound formed when **A** reacts with chlorine. *(1 mark)*
- **iii)** Draw the fully displayed structure of the organic product produced during the first step of the reaction between **B** and chlorine. *(1 marks)*

d Describe a chemical test that could be used to distinguish between **A** and **B**. *(1 marks)*

e Give the reagent and the reaction conditions necessary to convert **A** to **B**. *(2 marks)*

f Both **A** and **B** burn in air or oxygen.

- **i)** Write the balanced chemical equation for the complete combustion of **B**. *(2 marks)*
- **ii)** If equal volumes of **A** or **B** are burnt separately, which would you expect to produce the more soot? Give a reason for your choice. *(2 marks)*

Total 15 marks

Extended response question

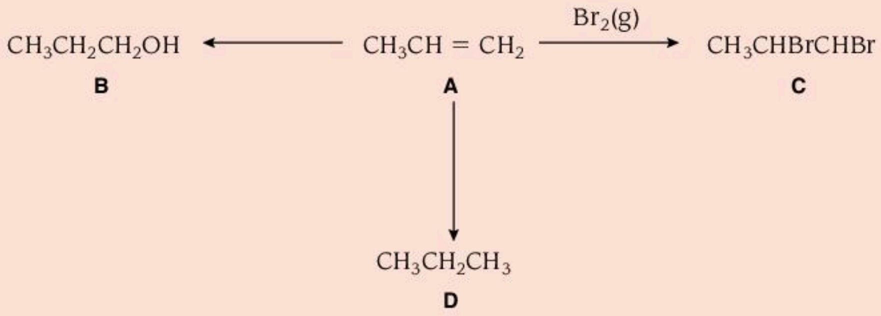
10 a One of the main sources of hydrocarbons used by man is petroleum.

i) What is a hydrocarbon? *(1 mark)*

ii) Name TWO fractions obtained from the fractional distillation of petroleum and give the main uses of each. *(4 marks)*

iii) Fractional distillation of petroleum produces too few of the smaller hydrocarbons to meet the demands of today's world. To increase the supplies of these smaller hydrocarbons, larger alkanes are cracked. Write an equation to represent one possible way of cracking hexane. *(2 marks)*

b Carefully study the reaction scheme below and answer the questions which follow.



i) Identify suitable reagents and reaction conditions which can be used to convert:

- compound **A** to compound **B**
- compound **A** to compound **D**. *(4 marks)*

ii) Describe what would be observed during the formation of compound **C** from compound **A**. *(1 mark)*

iii) Under the right conditions, compound **D** reacts with chlorine.

- Name the type of reaction taking place.
- Write a chemical equation to represent the first step in the reaction. *(3 marks)*

Total 15 marks

B15 Alcohols, alkanoic acids and esters

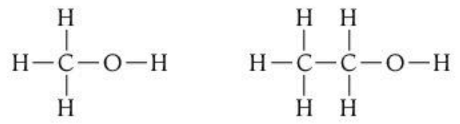
Alcohols, alkanoic acids and esters are organic compounds that all contain oxygen. Although they all contain oxygen, their properties and reactions differ because of the way the oxygen atoms are bonded in the molecules. Because of their varied properties and reactions, members of these three homologous series have widespread uses in industry and our daily lives.

B15.1 Alcohols: $C_nH_{2n+1}OH$

Alcohols are also known as **alkanols**. They are organic compounds which have the **hydroxyl group**, –OH, as their functional group.

The general formula of alcohols which contain one hydroxyl group is $\mathbf{C_nH_{2n+1}OH}$. This can also be written R–OH, where R represents C_nH_{2n+1}, i.e. an alkyl group.

Alcohols are named by using the prefix which indicates the total number of carbon atoms in the molecule plus the ending '**-anol**'. The first two alcohols in the series do not show structural isomerism. Their names and formulae are given in Figure 15.1.1.


methanol (CH_3OH)

ethanol (C_2H_5OH)

▲ **Figure 15.1.1** Structural formulae of the first two alcohols

Alcohols with *three or more* carbon atoms show structural isomerism. Structural isomers of alcohols result from the following:

- the change in position of the functional group, i.e. the change in position of the hydroxyl group, –OH
- branching of the molecule.

When naming isomers of alcohols:

- number the carbon atoms in the longest chain from the end closest to the –OH functional group so that the group is at the lowest number and indicate the position of the hydroxyl group in the name
- follow steps 1) to 4) on page 233 for naming branched chain isomers.

Using the rules above, the structures and names of the isomers of propanol (C_3H_7OH) and butanol (C_4H_9OH) are given in Figures 15.1.2 and 15.1.3.

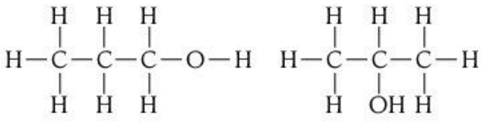
Objectives

By the end of this topic you will be able to:

- identify alcohols by the presence of the hydroxyl group, –OH
- give the names and structural formulae of unbranched alcohols with up to six carbon atoms
- relate the volatility and solubility of alcohols to the presence of the polar –OH group
- describe the reactions of ethanol
- explain the principles of the breathalyser test
- describe the fermentation process by which ethanol is produced from carbohydrates
- describe the processes involved in wine making and rum manufacture.

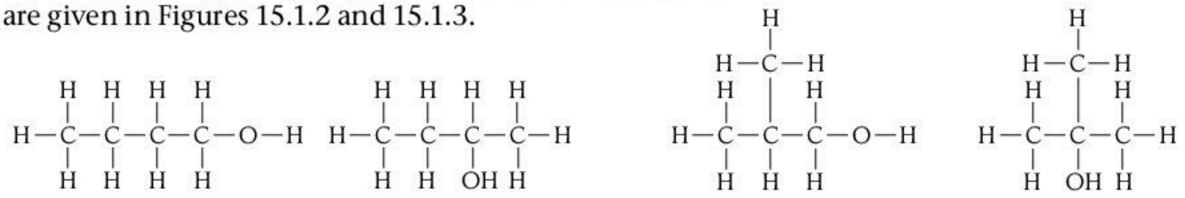
Exam tip

You must not confuse the –OH group in alcohols with the OH^- ion that is present in hydroxides.


propan-1-ol
(or 1-propanol)

propan-2-ol
(or 2-propanol)

▲ **Figure 15.1.2** The structural isomers of propanol (C_3H_7OH)

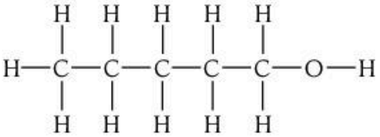

butan-1-ol
(or 1-butanol)

butan-2-ol
(or 2-butanol)

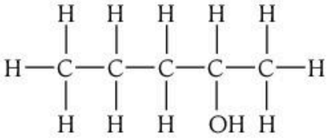
2-methylpropan-1-ol
(or 2-methyl-1-propanol)

2-methylpropan-2-ol
(or 2-methyl-2-propanol)

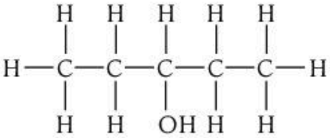
◀ **Figure 15.1.3** The structural isomers of butanol (C_4H_9OH)



pentan-1-ol
(or 1-pentanol)



pentan-2-ol
(or 2-pentanol)



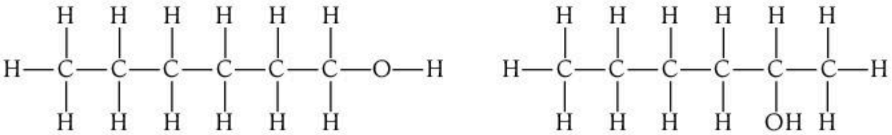
pentan-3-ol
(or 3-pentanol)

▲ **Figure 15.1.4** The straight chain isomers of pentanol ($C_5H_{11}OH$)

Exam tip

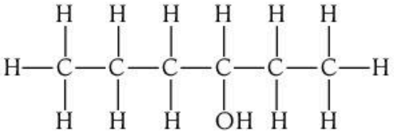
It is important that you can name the unbranched isomers of any alcohol containing up to *six* carbon atoms. It is also important that you can draw the fully displayed structural formulae of unbranched isomers of any alcohol containing up to *six* carbon atoms. It is not essential that you can name the branched isomers of alcohols.

The straight chain isomers of pentanol ($C_5H_{11}OH$) and hexanol ($C_6H_{13}OH$) are given in Figures 15.1.4 and 15.1.5.



hexan-1-ol
(or 1-hexanol)

hexan-2-ol
(or 2-hexanol)



hexan-3-ol
(or 3-hexanol)

▲ **Figure 15.1.5** The straight chain isomers of hexanol ($C_6H_{13}OH$)

General properties of alcohols

Alcohol molecules are **polar** due to the presence of the polar –OH group. Recall that a polar molecule is a covalent molecule in which one part of the molecule has a slightly positive charge and another part has a slightly negative charge. The volatility of alcohols and their solubility in water are both related to the polar nature of their molecules.

Volatility

Alcohols are **less volatile** than their corresponding alkanes with the same number of carbon atoms, i.e. the boiling point of an alcohol is always much higher than that of the corresponding alkane. For example, the boiling point of ethanol (C_2H_5OH) is 78 °C and the boiling point of ethane (C_2H_6) is −89 °C. This is because the polar –OH groups cause the forces of attraction between the alcohol molecules to be stronger than between the non-polar alkane molecules. As a result, all alcohols are either liquids or solids at room temperature.

The boiling point of alcohols *increases* as the number of carbon atoms in the molecules increases. This is because the intermolecular forces between the molecules increase as the size of the molecules increases.

Solubility in water

The general rule on solubility states that polar solutes dissolve in polar solvents. Water is a polar solvent, therefore alcohols, being polar, are **soluble** in water. The small alcohols, such as methanol, ethanol and propanol, are completely soluble in water, however, the solubility *decreases* as the number of carbon atoms in the molecules increases.

Reactions of ethanol

Alcohols undergo a variety of reactions due to the presence of the hydroxyl (–OH) functional group. In this section we will be looking specifically at the reactions of ethanol, however, the other alcohols undergo similar reactions since they all contain the hydroxyl group. The strength of the reactions *decreases* as the number of carbon atoms in the alcohol molecules increases.

Combustion of ethanol

Ethanol burns very easily in air or oxygen producing carbon dioxide and water as steam. The reaction is exothermic, producing a large amount of heat energy. The **complete combustion** of ethanol is shown in the equation below:

$$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g) \qquad \Delta H \text{ −ve}$$

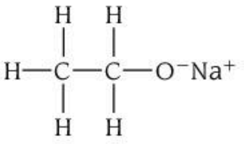
Ethanol burns with a *clear blue flame* because the ratio of carbon to hydrogen atoms in the molecule is low. All the carbon is therefore converted to carbon dioxide and no unreacted carbon remains to produce soot.

Reaction of ethanol with sodium

Ethanol reacts with sodium to produce sodium ethoxide (C_2H_5ONa) and hydrogen. When a piece of sodium metal is dropped into ethanol steady effervescence occurs as the hydrogen gas is evolved and a colourless solution of sodium ethoxide in alcohol is formed. The equation for the reaction between ethanol and sodium is given below:

$$2C_2H_5OH(l) + 2Na(s) \longrightarrow \underset{\text{sodium ethoxide}}{2C_2H_5ONa(\text{alc sol})} + H_2(g)$$

Sodium ethoxide is an ionic compound containing the ethoxide ion, $C_2H_5O^-$. The fully displayed structure of sodium ethoxide is represented in Figure 15.1.6.



▲ **Figure 15.1.6** The fully displayed structure of sodium ethoxide

The reaction between sodium and ethanol is similar to the reaction between sodium and water. In both reactions, the sodium **displaces** hydrogen from the ethanol or the water, forming hydrogen gas. However, the strength of the reaction between sodium and water is much more vigorous (see Unit 17.2).

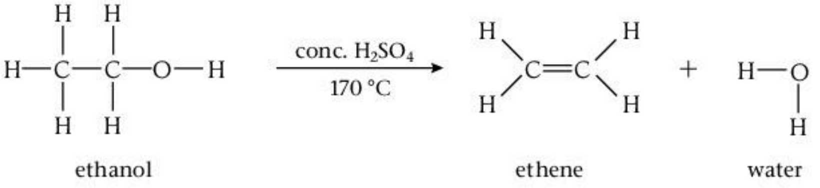
Dehydration of ethanol

When ethanol is mixed with concentrated sulfuric acid and the mixture is heated to 170 °C, the ethanol is **dehydrated** to ethene. This can be achieved in one of two ways:

- heating the ethanol with excess concentrated sulfuric acid at a temperature of 170 °C, where the acid acts as a catalyst in the reaction
- passing the ethanol vapour over heated aluminium oxide, where the aluminium oxide acts as a catalyst.

In both methods, a water molecule is removed from each ethanol molecule, i.e. the hydroxyl functional group is removed from one carbon atom and a hydrogen atom is removed from the other carbon atom. This results in the formation of a carbon–carbon double bond between the two adjacent carbon atoms.

The equation for the dehydration of ethanol using concentrated sulfuric acid is given below:



or

$$C_2H_5OH(l) \xrightarrow[170\ °C]{\text{conc. } H_2SO_4} C_2H_4(g) + H_2O(g)$$

Oxidation of ethanol

When ethanol is heated with a powerful oxidising agent, such as acidified potassium manganate(VII) solution or acidified potassium dichromate(VI) solution, the ethanol is **oxidised** to ethanoic acid. The oxidation of ethanol can be represented by the following equation:

$$\underset{\text{ethanol}}{C_2H_5OH(l)} + \underset{\substack{\text{from}\\ \text{oxidising}\\ \text{agent}}}{2[O]} \longrightarrow \underset{\text{ethanoic acid}}{CH_3COOH(aq)} + H_2O(l)$$

In the above reaction, ethanol is behaving as a **reducing agent**. If the reaction occurs with acidified potassium manganate(VII) solution, the solution changes from purple to colourless because the ethanol reduces the *purple* manganate(VII) ion (MnO_4^-) to the *colourless* manganese(II) ion (Mn^{2+}). If the reaction occurs with acidified potassium dichromate(VI) solution, the solution changes from orange to green because the ethanol reduces the *orange* dichromate(VI) ($Cr_2O_7^{2-}$) to the *green* chromium(III) ion (Cr^{3+}).

Reaction of ethanol with alkanoic acids

Ethanol reacts with alkanoic acids to form an ester and water. We will be discussing this in greater detail in Unit 15.3.

Practical activity

To investigate the oxidation of ethanol

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with ethanol, acidified potassium dichromate(VI) solution and a test tube.

Method

1. Place about 2 cm^3 of ethanol into the test tube.
2. Add an equal volume of acidified potassium dichromate(VI) solution and shake to mix.
3. Observe the colour of the solution periodically for about 15 minutes.

The breathalyser test

The reduction of acidified potassium dichromate(VI) and the accompanying colour change is used in the **breathalyser test** to test the alcohol content of a driver's breath. The driver blows over a sample of orange acidified potassium dichromate(VI) crystals. If the driver's breath contains ethanol vapour, the ethanol will reduce the orange dichromate(VI) ion ($Cr_2O_7^{2-}$) to the green chromium(III) ion (Cr^{3+}). As a result, the crystals turn *green*.

Production of ethanol by fermentation

Fermentation is the chemical reaction in which carbohydrates are converted into ethanol by yeast under anaerobic conditions, i.e. in the absence of oxygen. Yeast is a unicellular fungus which produces enzymes that break down any complex carbohydrates into simple sugars, mainly glucose. The yeast then produces another enzyme called zymase which converts the simple sugars into ethanol and carbon dioxide. This process, also known as **anaerobic respiration**, releases energy which the yeast cells use.

The equation for the fermentation of glucose is given below:

$$\underset{\text{glucose}}{C_6H_{12}O_6(aq)} \xrightarrow{\text{zymase in yeast}} \underset{\text{ethanol}}{2C_2H_5OH(aq)} + 2CO_2(g) \qquad \Delta H\ -\text{ve}$$

When the concentration of ethanol in the fermentation mixture reaches about 14%, it begins to destroy the zymase enzyme and the fermentation reaction stops. The ethanol can then be separated from the mixture by **fractional distillation** (Unit 2.4). During the distillation process, the ethanol is collected at 78 °C and is about 96% pure.

Practical activity

To prepare a sample of ethanol

(Teacher demonstration)

Your teacher may use this activity to assess:

- observation, recording and reporting.

Your teacher will perform the following experiment.

Method

1. Place 10 g of glucose into a 250 cm³ conical flask. Add 100 cm³ of water and mix to make a glucose solution.
2. Make a paste with 1 g of dried yeast and a small quantity of water. Add the paste to the glucose solution.
3. Place a cork on the flask with a delivery tube leading to a beaker containing lime water. Ensure that all the connections are properly sealed. The lime water acts as an airlock to prevent the entry of aerobic bacteria which can cause the formation of ethanoic acid (vinegar). The lime water also shows when the production of carbon dioxide starts by the formation of a white precipitate.
4. Leave the reaction until fermentation has stopped.
5. Pour the mixture into a distillation flask and fractionally distil it, collecting the fraction which distils at 78 °C, the boiling point of ethanol.

Fermentation is used in the production of different **alcoholic beverages** such as wine and rum. It is also used in baking where the carbon dioxide produced during the process makes the dough rise. We will look at wine and rum making in more detail in the next sections.

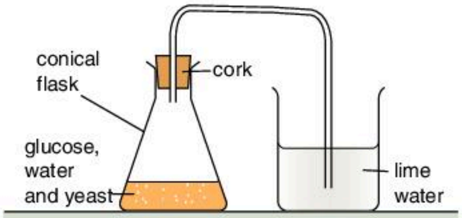
Winemaking

Wine is made primarily from grapes in a winery. The grapes are first crushed to form a mixture of pulp, skins and juice known as 'must'. To make white wine the skins are removed. The **vintner** (winemaker) then adds yeast to the must in fermentation tanks and the yeast ferments the sugars present forming ethanol and carbon dioxide. When fermentation is complete any skins are removed and the wine is filtered to clarify it. The wine is then aged in storage tanks or oak barrels before bottling. The ethanol content of most wines is between 10% and 14%.

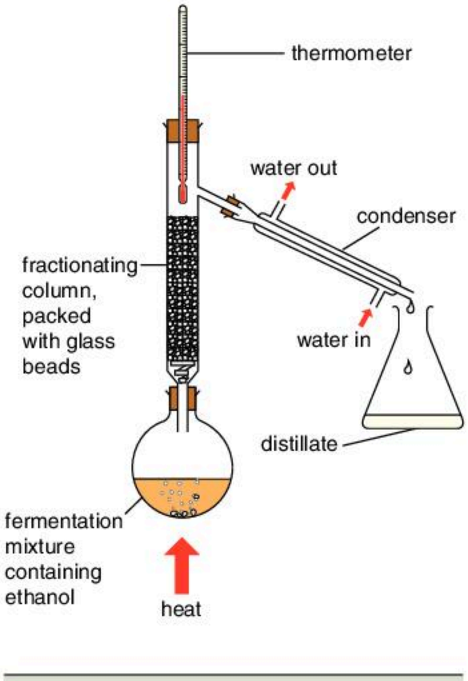
Great care must be taken throughout the winemaking process to minimise the exposure of the wine to air. This is because certain aerobic bacteria may

Exam tip

It is very important that you can draw a line diagram to show the apparatus used to carry out **fractional distillation** in the laboratory.



▲ **Figure 15.1.7** Production of ethanol by fermentation



▲ **Figure 15.1.8** Separation of ethanol by fractional distillation

Did you know?

Red or black grapes are used to make **red wine** and the grape skins are present throughout fermentation to give the wine its colour. **White wine** is made mainly from 'white' grapes, which actually vary in colour from green to pale yellow.

be present in the wine from the original grapes and, on exposure to oxygen in the air, these bacteria **oxidise** the ethanol to ethanoic acid, or vinegar:

$$C_2H_5OH(aq) + O_2(g) \xrightarrow{\text{aerobic bacteria}} \underset{\text{ethanoic acid}}{CH_3COOH(aq)} + H_2O(l)$$

The ethanoic acid then causes the wine to become sour and undrinkable.

Rum manufacture

▲ **Figure 15.1.9** A master blender in front of the oak barrels

Rum is made from the black treacle-like substance known as **molasses**, which is produced during the extraction of sugar from sugar cane (Unit 2.5). The molasses is first diluted with water and yeast is added. The yeast ferments the sugar in the molasses forming ethanol and carbon dioxide. The mixture produced is then sent to the still where it is **fractionally distilled** to produce a distillate with an ethanol content of between 70% and 95%.

The distillate is then diluted and transferred to oak barrels to be aged. This process can take 20 years or more. The oak barrels are charred on the inside to impart a variety of flavours to the rum, give it a smoother taste and add colour. More water is then added to reduce the ethanol content to about 40% and a master blender is responsible for blending various rums of different ages and/or from different barrels to produce the unique brands of the different rum manufacturers.

Summary questions

1 Give the names and fully displayed structural formulae of three straight chain isomers of pentanol.

2 Explain how the functional group of the alcohols affects their solubility and boiling points.

3 Ethanol burns in oxygen with a blue flame. Give a balanced equation for the reaction.

4 Ethanol can be dehydrated.

- a Give the conditions required for the reaction.
- b Draw the fully displayed structural formulae of the products formed in the reaction.

5 Give a balance chemical equation for the fermentation of glucose and give two ways in which the reaction may be used.

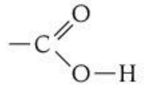
B15.2 Alkanoic acids: $C_nH_{2n+1}COOH$

Objectives

By the end of this topic you will be able to:

- identify alkanoic acids by the presence of the carboxyl group, –COOH
- give the names and structural formulae of unbranched alkanoic acids with up to six carbon atoms
- relate the volatility and solubility of alkanoic acids to the presence of the polar –OH part of the functional group
- describe the reactions of ethanoic acid.

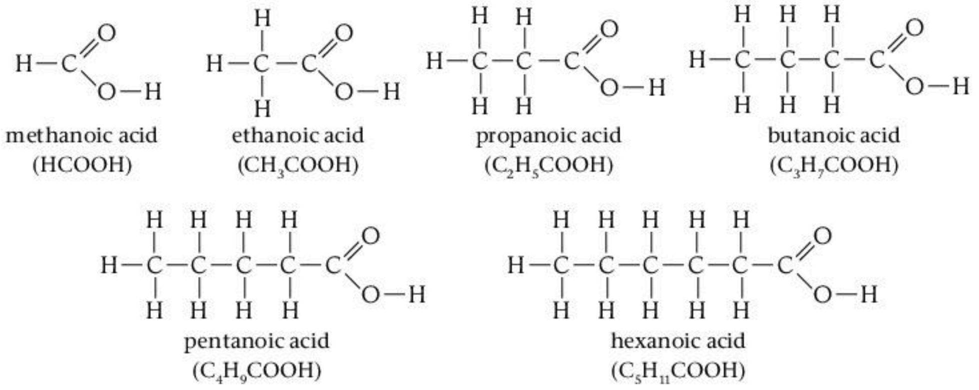
Alkanoic acids are also known as **carboxylic acids**. They are organic compounds which have the **carboxyl group**, –COOH, as their functional group. The carboxyl group can also be represented as:



The general formula of alkanoic acids which contain one carboxyl group is $\mathbf{C_nH_{2n+1}COOH}$. This can also be written R–COOH, where R represents C_nH_{2n+1}, i.e. an alkyl group.

Alkanoic acids are named by using the prefix which indicates the total number of carbon atoms in the molecule plus the suffix '-anoic acid'. It is important to remember that the functional group contains one carbon atom which must be taken into account when naming alkanoic acids, i.e. the prefix is not derived directly from '*n*' in the general formula, it is derived by adding 1 to the value of '*n*'. For example, the name of the acid with the formula CH_3COOH is ethanoic acid and not methanoic acid.

The names and formulae of the straight chain isomers of the first six alkanoic acids are shown in Figure 15.2.1.



▲ **Figure 15.2.1** Structural formulae of the first six alkanoic acids

Did you know?

Organic acids containing one or more carboxyl functional groups can be found throughout the natural world. Methanoic acid is found in the venom of ant and bee stings while ethanoic acid gives vinegar its sour taste. Citric acid, found in citrus fruits, has three carboxyl groups and tartaric acid, found in tamarinds, has two functional groups. Aspirin, or acetylsalicylic acid, which is used to reduce fever and relieve aches and pains, has one carboxyl group.

Properties of alkanoic acids

Alkanoic acid molecules are **polar** due to the presence of the polar –OH part of the carboxyl functional group. Like alcohols, the volatility of alkanoic acids and their solubility in water are both related to the polar nature of their molecules.

Volatility

Alkanoic acids are **less volatile** than their corresponding alkanes with the same number of carbon atoms, i.e. the boiling point of an alkanoic acid is always much higher than that of the corresponding alkane. For example, the boiling point of ethanoic acid (CH_3COOH) is 112 °C and the boiling point of ethane (C_2H_6) is -89 °C. This is because the forces of attraction between the polar alkanoic acid molecules are stronger than between the non-polar alkane molecules. All alkanoic acids are either liquids or solids at room temperature.

The boiling point of alkanoic acids *increases* as the number of carbon atoms in the molecules increases because the intermolecular forces between the molecules increase as the size of the molecules increases.

Solubility in water

Alkanoic acids, being polar, are **soluble** in water. The small alkanoic acids such as methanoic acid, ethanoic acid and propanoic acid are completely soluble in water, however, the solubility *decreases* as the number of carbon atoms in the molecules increases.

When alkanoic acids dissolve in water, they also ionise to a very small degree to form hydrogen (H^+) ions, hence they are **weak acids**. For example, when ethanoic acid dissolves in water about 1% of the ethanoic acid molecules are ionised at any one time, i.e. the solution contains 1% ethanoate ions (CH_3COO^-) and hydrogen ions (H^+) and 99% ethanoic acid molecules (CH_3COOH).

The ionisation of ethanoic acid is shown in the following equation:

$$CH_3COOH(aq) \rightleftharpoons \underset{\text{ethanoate ion}}{CH_3COO^-(aq)} + H^+(aq)$$

Reactions of aqueous ethanoic acid

Aqueous solutions of alkanoic acids react in a very similar way to inorganic acids such as hydrochloric acid when they react with reactive metals, metal oxides, metal hydroxides and metal carbonates. The reactions are *slower* than they would be with inorganic acids because the alkanoic acids are weak acids.

In this section we will be looking specifically at the reactions of ethanoic acid. However, the other alkanoic acids undergo similar reactions since they all contain the carboxyl group. The strength of the reactions *decreases* as the number of carbon atoms in the molecules increases.

Salts formed by ethanoic acid are known as **ethanoates**.

Reaction with reactive metals

Reactive metals react with ethanoic acid to produce a salt and hydrogen gas. Using the reaction between magnesium and ethanoic acid as an example:

$$\mathrm{Mg(s) + 2CH_3COOH(aq) \longrightarrow \underset{\text{magnesium ethanoate}}{(CH_3COO)_2Mg(aq)} + H_2(g)}$$

It is important to note that since the magnesium has displaced the H^+ ion from the –COOH functional group, it is written in the place of this H^+ ion in the formula of the salt. However, in the name of the salt, the metal still comes first.

Reaction with metal oxides

Metal oxides react with ethanoic acid to produce a salt and water. These reactions are **neutralisation reactions**. Using the reaction between zinc oxide and ethanoic acid as an example:

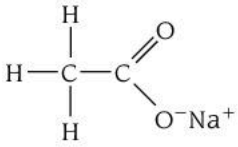
$$\mathrm{ZnO(s) + 2CH_3COOH(aq) \longrightarrow \underset{\text{zinc ethanoate}}{(CH_3COO)_2Zn(aq)} + H_2O(l)}$$

Reaction with metal hydroxides

Metal hydroxides react with ethanoic acid to produce a salt and water. These are also **neutralisation reactions**. Using the reaction between sodium hydroxide and ethanoic acid as an example:

$$\mathrm{NaOH(aq) + CH_3COOH(aq) \longrightarrow \underset{\text{sodium ethanoate}}{CH_3COONa(aq)} + H_2O(l)}$$

The fully displayed structure of sodium ethanoate is represented in Figure 15.2.2.



▲ **Figure 15.2.2** The fully displayed structure of sodium ethanoate

Reaction with metal carbonates

Metal carbonates react with ethanoic acid to produce a salt, carbon dioxide and water. Using the reaction between calcium carbonate and ethanoic acid as an example:

$$\mathrm{CaCO_3(s) + 2CH_3COOH(aq) \longrightarrow \underset{\text{calcium ethanoate}}{(CH_3COO)_2Ca(aq)} + CO_2(g) + H_2O(l)}$$

Esterification

Alkanoic acids react with alcohols to produce an ester and water. The reaction is known as **esterification** and it requires a catalyst of concentrated sulfuric acid and heat in order to take place. Since two molecules are reacting together with the loss of a water molecule from between them, the reaction may also be referred to as a **condensation reaction**. The alkanoic acid and alcohol must be in the pure liquid form and not in aqueous solution for the reaction to occur.

Using the reaction between ethanoic acid and propanol as an example:

$$\mathrm{\underset{\text{ethanoic acid}}{CH_3COOH(l)} + \underset{\text{propanol}}{C_3H_7OH(l)} \underset{\text{heat}}{\overset{\text{conc. }H_2SO_4}{\rightleftharpoons}} \underset{\text{propyl ethanoate}}{CH_3COOC_3H_7(l)} + \underset{\text{water}}{H_2O(l)}}$$

Propyl ethanoate is an **ester**. We will be looking at esters in more detail in the next unit.

Summary questions

1 Give the name of each of the following compounds:
   a C_3H_7COOH
   b $C_5H_{11}COOH$

2 Explain why methanoic acid is an acid. As part of your answer write a balanced chemical equation which shows how methanoic acid ionises in water.

3 Why is ethanoic acid less volatile than ethane?

4 Give a balanced chemical equation for the reaction between ethanoic acid and:
   a calcium oxide
   b potassium carbonate.

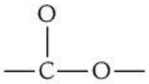
5 Give the conditions necessary for the reaction between ethanoic acid and ethanol.

B15.3 Esters: RCOOR′

An **ester** is a compound formed when an alkanoic acid reacts with an alcohol. The smaller esters are volatile liquids which usually have very distinctive sweet fruity smells. Esters are widespread in nature where they occur in both animals and plants, especially in flowers and fruits, giving them their distinctive smells and flavours. Naturally occurring animal fats and vegetable oils are esters of long-chain alkanoic acids and an alcohol called glycerol.

Esters are used extensively as artificial perfumes, food flavourings and essences, e.g. artificial flavourings such as banana, strawberry and apple are esters. They are also used in the manufacture of soap and as industrial solvents for making paints and varnishes. **Biodiesel** is an ester produced from the reaction of fats with methanol.

Esters have the **ester group**, –COO–, as their functional group. The ester group can also be represented as:



$$-\overset{\overset{\displaystyle O}{|}}{C}-O-$$

The general formula of a simple ester is $\mathbf{C_nH_{2n+1}COOC_xH_{2x+1}}$ where 'n' and 'x' may be the same or different numbers. This can also be written R–COO–R′, where R and R′ represent alkyl groups which may be the same or different.

Formation of esters

Esters are made by a condensation reaction known as **esterification**. During esterification, an alkanoic acid is heated with an alcohol in the presence of concentrated sulfuric acid. The reaction is reversible and produces an ester and water.

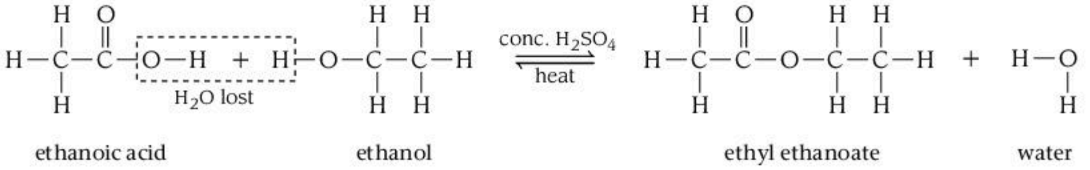
The reaction can be summarised by the following general word equation:

$$\textbf{alkanoic acid + alcohol} \xrightleftharpoons[\textbf{heat}]{\textbf{conc. }\mathbf{H_2SO_4}} \textbf{ester + water}$$

The concentrated sulfuric acid is added for two reasons:

- it acts as a catalyst to speed up the reaction
- it is a powerful dehydrating agent, therefore, it removes the water produced during the reaction which favours the forward reaction and increases the yield of ester produced.

Using structural formulae to represent the reaction between ethanoic acid and ethanol to produce ethyl ethanoate, we can see the structure of the ethyl ethanoate that is formed.



The chemical equation for the reaction is given below:

$$CH_3COOH(l) + C_2H_5OH(l) \xrightleftharpoons[\text{heat}]{\text{conc. } H_2SO_4} CH_3COOC_2H_5(l) + H_2O(l)$$

ethanoic acid ethanol ethyl ethanoate water

Objectives

By the end of this topic you will be able to:

- identify esters by the presence of the ester functional group, –COO–
- describe how esters are formed
- write the names and formulae of esters
- explain the hydrolysis of esters
- explain the saponification of fats and oils
- distinguish between soapy and soapless detergents.

Did you know?

Some esters play an important role in animal communication. Animals produce chemical signals called **pheromones** which trigger natural responses in members of the same species. Some of these are esters. For example, 3-methylbutyl ethanoate, which is the main component of the banana aroma, is also the alarm pheromone of the honeybee. It is released when a bee is attacked by a predator to trigger aggression in the other bees which provokes them to sting.

Formulae and names of esters

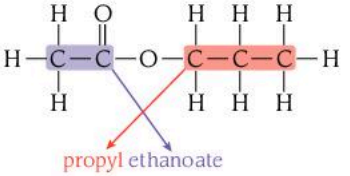
When writing the structural formula of an ester, the first part of the formula comes from the alkanoic acid and the second part comes from the alcohol that formed the ester.

The name of an ester is also derived from the alkanoic acid and the alcohol. However, the first part of the name is derived from the alcohol and the second part of the name is derived from the alkanoic acid.

To summarise:

Formula: *acid* part first; *alcohol* part second.

Name: *alcohol* part first with the ending 'yl'; *acid* part second with the ending 'anoate'.



▲ **Figure 15.3.1** Structural formula of propyl ethanoate

The structural formula of propyl ethanoate is shown in Figure 15.3.1 as an example of this.

Some more examples of the formulae and names of esters are given in Figure 15.3.2.

▲ **Figure 15.3.2** Other examples of esters

Exam tip

When writing the equation for the reaction between an alkanoic acid and an alcohol, always write the formula of the **acid first**. In this way, when you write the formula of the ester, you will always write the acid part first.

Practical activity

To prepare a sample of ethyl ethanoate

(Teacher demonstration)

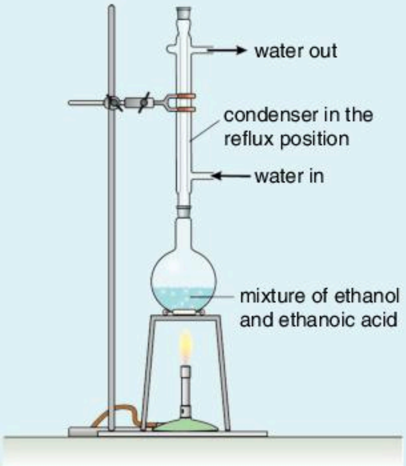
Your teacher may use this activity to assess:

- observation, recording and reporting.

Your teacher will perform the following experiment.

Method

1. Place 10 cm^3 of ethanoic acid into a round-bottomed flask. Add 10 cm^3 of ethanol and 2 cm^3 of concentrated sulfuric acid.
2. Attach a condenser to the flask in the reflux position as shown in Figure 15.3.3.
3. Gently boil the mixture for 5 minutes.
4. Pour the mixture into a small beaker half filled with cold water and identify the ester formed by its characteristic odour and oily appearance on the surface of the water.



▲ **Figure 15.3.3** Preparing a sample of ethyl ethanoate

The reaction between ethanoic acid and ethanol to produce ethyl ethanoate proceeds relatively slowly. As the reaction is heated, both reactants and any ester produced boil forming vapours which must be retained. The condenser is placed in the **reflux position**, i.e. vertically, so that the ethanoic acid, ethanol and ethyl ethanoate vapours enter the condenser where they condense and return to the flask. In this way, the ethanoic acid and ethanol are returned for further reaction and no ethyl ethanoate is lost. When the resulting mixture is poured into the water, any unreacted acid and alcohol dissolve in the water. Ethyl ethanoate is only slightly soluble and, being less dense than water, it floats forming an oily layer on the surface of the water.

Hydrolysis of esters

During a **hydrolysis reaction** molecules are broken down into smaller molecules by reacting with water.

Esters can be hydrolysed by heating with a dilute acid or an aqueous alkali. The products of hydrolysis depend on whether an acid or an alkali is used.

Acid hydrolysis

During **acid hydrolysis**, the ester is heated with an acid catalyst such as dilute hydrochloric or sulfuric acid. The products of acid hydrolysis are an alkanoic acid and an alcohol.

Using the acid hydrolysis of ethyl ethanoate as an example:

$$\underset{\text{ethyl ethanoate}}{CH_3COO_2H_5(l)} + H_2O(l) \xrightleftharpoons{H^+\text{ ions}} \underset{\text{ethanoic acid}}{CH_3COOH(aq)} + \underset{\text{ethanol}}{C_2H_5OH(aq)}$$

Alkaline hydrolysis

During **alkaline hydrolysis**, the ester is heated with sodium or potassium hydroxide solution. The products of alkaline hydrolysis are the sodium or potassium salt of an alkanoic acid and an alcohol. The reaction occurs in *two* stages.

Using the alkaline hydrolysis of ethyl ethanoate by reacting it with sodium hydroxide solution as an example:

Stage 1: ethanoic acid and ethanol are formed first as in acid hydrolysis:

$$\underset{\text{ethyl ethanoate}}{CH_3COOC_2H_5(l)} + H_2O(l) \xrightleftharpoons{OH^-\text{ ions}} \underset{\text{ethanoic acid}}{CH_3COOH(aq)} + \underset{\text{ethanol}}{C_2H_5OH(aq)}$$

Stage 2: the ethanoic acid is then neutralised by the sodium hydroxide solution to form sodium ethanoate:

$$\underset{\text{ethanoic acid}}{CH_3COOH(aq)} + NaOH(aq) \longrightarrow \underset{\text{sodium ethanoate}}{CH_3COONa(aq)} + H_2O(l)$$

The overall reaction is shown in the following equation:

$$\underset{\text{ethyl ethanoate}}{CH_3COOC_2H_5(l)} + NaOH(aq) \longrightarrow \underset{\text{sodium ethanoate}}{CH_3COONa(aq)} + \underset{\text{ethanol}}{C_2H_5OH(aq)}$$

Saponification of esters

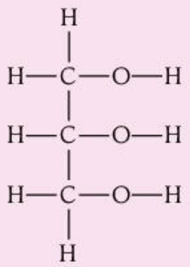
Saponification refers to the alkaline hydrolysis of large esters found in animal fats and vegetable oils. Saponification is used in the manufacture of **soap**. During the manufacturing process, fats and oils are boiled with concentrated potassium or sodium hydroxide solution to produce the sodium salts of long-chain alkanoic acids, also known as **fatty acids**. The other product of the reaction is the alcohol, glycerol, $C_3H_5(OH)_3$. As the formula shows, glycerol is an alcohol with three hydroxyl functional groups.

A simplified equation for this reaction is:

$$\text{fat or oil} + NaOH \xrightarrow{\text{heat}} \text{sodium salt of a long-chain alkanoic acid} + \text{glycerol}$$

Did you know?

Glycerol is the common name for propane-1,2,3-triol or 1,2,3-propanetriol. The structural formula of glycerol is drawn for you below:

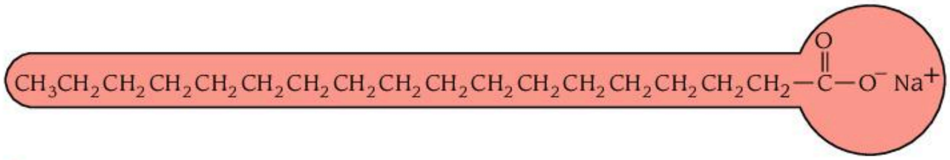


Example

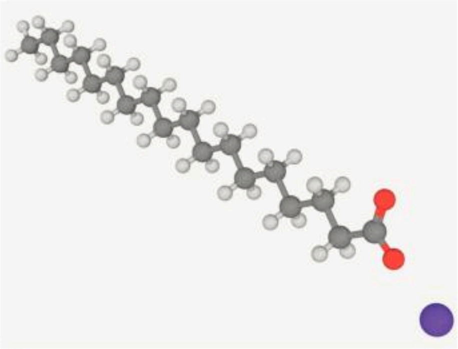
One example of a fat is glyceryl octadecanoate ($(C_{17}H_{35}COO)_3C_3H_5$). Glyceryl octadecanoate is an ester of the fatty acid known as octadecanoic acid ($C_{17}H_{35}COOH$) and glycerol ($C_3H_5(OH)_3$). When glyceryl octadecanoate is hydrolysed by boiling with concentrated sodium hydroxide solution, sodium octadecanoate ($C_{17}H_{35}COONa$) and glycerol are produced. Sodium octadecanoate is one form of soap.

$$\underset{\substack{\text{glyceryl octadecanoate}\\ \text{(fat)}}}{(C_{17}H_{35}COO)_3C_3H_5(l)} + 3NaOH(aq) \xrightarrow{\text{heat}} \underset{\substack{\text{sodium octadecanoate}\\ \text{(soap)}}}{3C_{17}H_{35}COONa(aq)} + \underset{\text{glycerol}}{C_3H_5(OH)_3(aq)}$$

When the soap has been produced, the mixture is added to saturated sodium chloride solution to precipitate out and solidify the soap.



▲ **Figure 15.3.4** A soap molecule



▲ **Figure 15.3.5** Model of a soap molecule

Practical activity

To prepare a sample of soap

(Teacher demonstration)

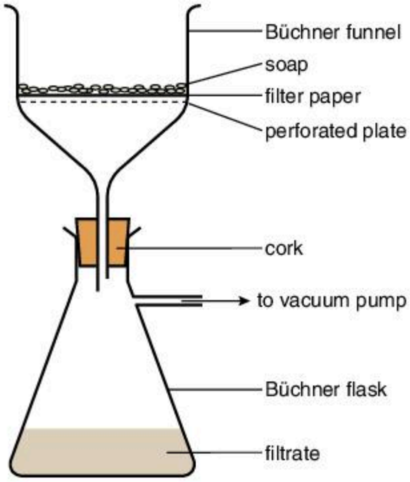
Your teacher may use this activity to assess:

- observation, recording and reporting.

Your teacher will perform the following experiment.

Method

1. Place 10 cm³ of cooking oil into a beaker and carefully add 15 cm³ of ethanol and 15 cm³ of 5.0 mol dm⁻³ sodium hydroxide solution. Stir the mixture thoroughly.
2. Continue stirring the mixture while heating it gently on a hot plate for about 30 minutes – the mixture must no longer separate into two layers and it must be fairly thick. During heating, make sure that the mixture doesn't overheat or foam and that the volume doesn't decrease too much. Top the mixture up with a solution containing equal volumes of ethanol and water if necessary.
3. When thick, let the mixture cool for a while, but do not let it cool completely.
4. Add the mixture to 100 cm³ of saturated sodium chloride solution and stir to precipitate out the soap.
5. Filter the mixture using suction filtration and collect the soap in the filter paper as the residue.
6. Rinse the soap with two separate 10 cm³ portions of iced distilled water during filtration and continue the suction filtration to dry the soap.
7. Place a small portion of the soap in a test tube and add about 4 cm³ of distilled water. Place a cork on the tube, shake thoroughly and observe.



▲ **Figure 15.3.6** Suction filtration to separate soap from the reaction mixture

Soapy and soapless detergents

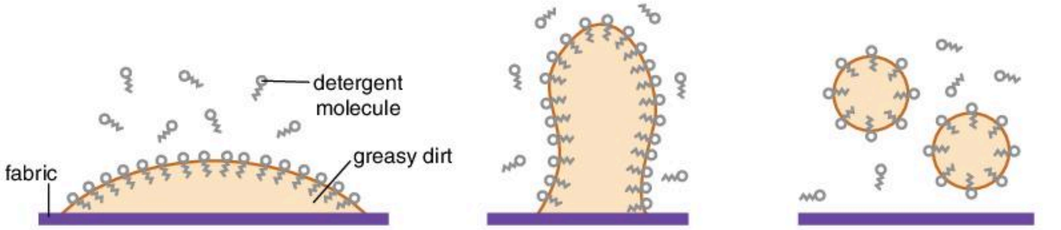
A **detergent** is a chemical added to water to remove dirt from clothing, surfaces and floors. Detergent molecules are composed of long molecules with two ends which have very different structures and properties.

- One end of the molecule, the 'head', is ionic and readily dissolves in water. We say that this end is **hydrophilic** or water-loving and oil-hating. This end is often referred to as the **ionic head**.
- The opposite end of the molecule, the 'tail', is covalent and is composed of carbon and hydrogen atoms. This end readily dissolves in organic substances such as oils and grease. We say that this end is **hydrophobic** or water-hating and oil-loving. This end is referred to as the **hydrocarbon tail**.

Dirt often consists of oily or greasy substances that do not dissolve in water. This makes them difficult to remove. Detergents help to remove this dirt by lowering the surface tension of water, loosening the dirt molecules and dispersing them in the water, i.e. they help to make the greasy dirt mix with the water.

When a detergent is added to oily or greasy dirt on fabrics or the skin, the hydrophobic (oil-loving) tail of the detergent molecule dissolves in the greasy dirt and the hydrophilic (water-loving) head sticks out. The head dissolves in the water and the dirt is lifted from the fabric or skin and is washed away. This is helped by agitation such as occurs in a washing machine.

Figure 15.3.7 shows how greasy dirt can be removed from fabric.



The hydrophobic tails of the detergent molecules dissolve in the greasy dirt. The hydrophilic heads remain dissolved in the water.

The greasy dirt is lifted off the fabric.

The greasy dirt is surrounded by detergent molecules and broken into smaller droplets that can mix with water. This forms an emulsion which can be washed away.

Figure 15.3.7 How detergent molecules remove greasy dirt

Detergent molecules can be classified into soapy or soapless depending on their structure.

Soapy detergents

Soapy detergents are often just known as **soap**. They are made by boiling animal fats or vegetable oils with concentrated potassium or sodium hydroxide solution, as described in the previous section. The structure of a soap molecule is shown in Figure 15.3.4.

Soapy detergents do not lather easily in hard water. Hard water contains dissolved calcium and magnesium salts such as calcium and magnesium hydrogencarbonate (Unit 21.1). The calcium and magnesium ions in the hard water displace the sodium ions in the soap forming an insoluble precipitate of calcium or magnesium octadecanoate. This precipitate is known as **scum**.

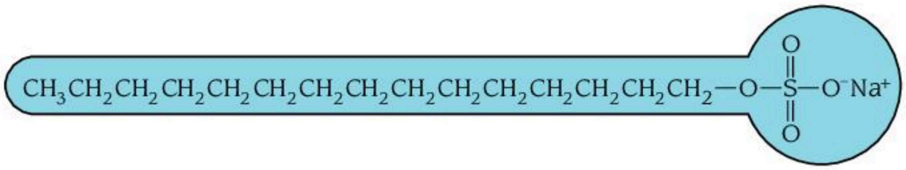
$$\underset{\text{soap}}{2C_{17}H_{35}COONa(aq)} + Ca^{2+}(aq) \longrightarrow \underset{\text{scum}}{(C_{17}H_{35}COO)_2Ca(s)} + 2Na^{+}(aq)$$

Figure 15.3.8 Scum forms when soap and hard water react

Scum removes the soap from the solution which prevents it from doing the job of removing greasy dirt. It is only when all of the calcium and magnesium ions have been precipitated out of the water as scum that any remaining soap molecules will form a lather and do their job of removing the dirt.

Soapless detergents

Soapless detergents are also known as synthetic detergents and often just called **detergents**. They are formed from petroleum. The structure of a soapless detergent molecule is shown in Figure 15.3.9.



$$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2-O-\overset{\overset{\displaystyle O}{\|}}{\underset{\underset{\displaystyle O}{\|}}{S}}-O^-Na^+$$

▲ **Figure 15.3.9** A soapless detergent molecule

Soapless detergents lather much more easily with hard water than soapy detergents. This is because their calcium and magnesium salts are soluble, therefore, they do not form scum.

Apart from their different lathering abilities in hard water, soapy and soapless detergents have other differences which are summarised in Table 15.3.1.

▼ **Table 15.3.1** Soapy and soapless detergents compared

Soapy detergents	**Soapless detergents**
They do not lather easily in hard water. Instead they form unpleasant **scum** which wastes soap, forms a grey, greasy layer around sinks, showers, etc. and discolours clothes.	They lather in hard water. They do not form scum since their calcium and magnesium salts are soluble.
They are **biodegradable**, i.e. they are broken down by bacteria in the environment, therefore, they do not cause foaming in sewerage systems or on waterways, e.g. lakes and rivers.	Some are **non-biodegradable**. These cause foam to build up in sewerage systems and on waterways. This foam makes sewage treatment difficult and prevents oxygen dissolving for aquatic organisms resulting in their death. Most of the modern soapless detergents are, however, biodegradable.
They do not contain phosphates, therefore, they do not cause pollution.	They contain phosphates which are added to improve their cleaning ability. Phosphates pollute aquatic environments by causing **eutrophication**, i.e. the rapid growth of green algae (Unit 20.5). Eutrophication turns waterways green and eventually makes them stagnant.
They are manufactured from a renewable resource, i.e. fats and oils. The supply of fats and oils will not run out.	They are manufactured from a non-renewable resource, i.e. petroleum. The supplies of petroleum will eventually run out.

▲ **Figure 15.3.10** Some soapless detergents are non-biodegradable and can damage the environment

Practical activity

To compare the effects of a soapy and a soapless detergent on hard water

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with a soapy detergent, a soapless detergent, a sample of soft water, a sample of hard water, four boiling tubes and two dropping pipettes.

Method

1. Label the boiling tubes A, B, C and D. Place 20 cm^3 of soft water into each of A and B, and 20 cm^3 of hard water into each of C and D.
2. Using a dropping pipette, add one drop of soapy detergent to the water in tube A. Place a cork on the tube and shake vigorously for 10 seconds.
3. Allow the test tube to stand for 1 minute. If there is no permanent lather, repeat step **2**. Continue repeating step **2** until a permanent lather forms.
4. Observe the appearance of the mixture and note how many drops of soapy detergent were required to form a permanent lather.
5. Repeat steps **2** to **4** adding soapy detergent to C and soapless detergent to B and D.
6. Record your results in a table.
7. Compare the appearance of the contents of each tube when a permanent lather formed and the number of drops of detergent needed to form a permanent lather in each tube.

Summary questions

1. Give the conditions necessary for the formation of an ester.
2. Name the following esters:
   a $C_3H_7COOCH_3$ **b** $C_2H_5COOC_3H_7$
3. Give a balanced equation for:
   a the formation of butyl ethanoate from ethanoic acid and butanol
   b the hydrolysis of ethyl methanoate using sodium hydroxide solution.
4. What is saponification?
5. Compare a soapy detergent and a soapless detergent by referring to their effects on:
   a hard water **b** the environment.

Key concepts

- **Alcohols** have the general formula $\mathbf{C_nH_{2n+1}OH}$.
- The functional group of alcohols is the **hydroxyl group**, –OH.
- Alcohols with three or more carbon atoms display structural isomerism.
- Alcohol molecules are **polar**.
- Alcohols are less volatile than their corresponding alkanes because of the polar nature of the molecules.
- The smaller alcohols are completely soluble in water because of the polar nature of the molecules. Solubility decreases as the molecular size increases.
- **Ethanol** burns in air or oxygen with a clean blue flame to form carbon dioxide and water as steam. The reaction is exothermic.
- Ethanol reacts with sodium to form sodium ethoxide and hydrogen.
- Ethanol can be dehydrated to ethene by being heated to 170 °C with concentrated sulfuric acid.

- Ethanol can be oxidised by reacting it with acidified potassium manganate(VII) solution or acidified potassium dichromate(VI) solution. The reaction produces ethanoic acid and water.
- The reduction of acidified potassium dichromate(VII) crystals by ethanol and the accompanying colour change from orange to green is used in the breathalyser test.
- Ethanol reacts with alkanoic acids to produce an ester and water.
- Ethanol can be produced by fermentation.
- **Fermentation** is the chemical reaction in which carbohydrates are converted into ethanol by yeast under anaerobic conditions.
- Wine and rum are made by fermentation.
- **Alkanoic acids** have the general formula $\mathbf{C_nH_{2n+1}COOH}$.
- The functional group of alkanoic acids is the **carboxyl group**, –COOH.
- Alkanoic acid molecules are **polar**.
- Alkanoic acids are less volatile than their corresponding alkanes because of the polar nature of the molecules.
- The smaller alkanoic acids are completely soluble in water because of the polar nature of the molecules. Solubility decreases as the molecular size increases.
- Alkanoic acids partially ionise in water making them weak acids.
- **Aqueous ethanoic acid** reacts with reactive metals, metal oxides, metal hydroxide and metal carbonates in the same way as inorganic acids.
- Ethanoic acid reacts with alcohols to produce an ester and water.
- An **ester** is a compound formed when an alkanoic acid reacts with an alcohol.
- The reaction which makes an ester is a **condensation reaction** known as **esterification**. It requires a catalyst of concentrated sulfuric acid and heat.
- Esters often have distinctive sweet fruity smells. Animal fats and vegetable oils are esters of long-chain alkanoic acids and glycerol.
- The functional group of esters is the **ester group**, –COO–.
- **Acid hydrolysis** of an ester forms an alkanoic acid and an alcohol.
- **Alkaline hydrolysis** of an ester forms the salt of an alkanoic acid and an alcohol.
- **Saponification** is the process by which soap is made by the alkaline hydrolysis of animal fats or vegetable oils.
- **Soapless detergents** are produced from petroleum.
- Soapy detergents do not lather easily in hard water, instead they form scum. Soapless detergents do lather in water.
- Soapy detergents do not pollute because they are biodegradable and do not contain phosphates. Soapless detergents can cause pollution because some are non-biodegradable and they contain phosphates.

Practice exam-style questions

Multiple-choice questions

Questions 1 and 2 refer to the following types of reaction. When answering the question, select the most appropriate type of reaction.

A fermentation
B reduction
C substitution
D oxidation

1 The process by which ethanol is produced from carbohydrates.

2 The process by which ethanol is converted to ethanoic acid.

3 During the breathalyser test to detect the presence of ethanol on a driver's breath, the ethanol:
A oxidises the green Cr^{3+} ion to the orange $Cr_2O_7^{2-}$ ion
B reduces the orange $Cr_2O_7^{2-}$ ion to the green Cr^{3+} ion
C reduces the green Cr^{3+} ion to the orange $Cr_2O_7^{2-}$ ion
D oxidises the orange $Cr_2O_7^{2-}$ ion the green Cr^{3+} ion

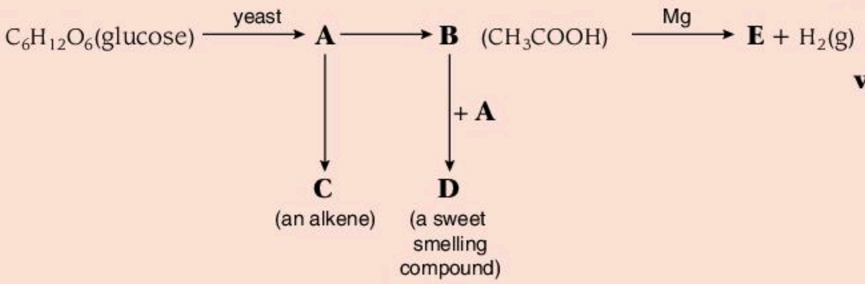
4 Concentrated sulfuric acid is used in the reaction between ethanoic acid and ethanol as:
A a reducing agent
B an oxidising agent
C a catalyst
D a neutralising agent

5 Acid hydrolysis of methyl ethanoate will produce:
A methanoic acid and ethanol
B methanoic acid, ethanol and carbon dioxide
C ethanoic acid and methanol
D ethanoic acid, methanol and water

6 Which of the following will not react with aqueous sodium hydroxide?
A ethanol
B propanoic acid
C ethyl methanoate
D olive oil

Structured question

7 Study the information given in the following reaction scheme and answer the questions which follow. Compound **B** has been identified for you.



a Draw the fully displayed structures of each of compounds **A**, **C** and **D** shown in the reaction scheme above. *(4 marks)*

b Name the type of reaction involved in the conversion of:
- **i)** Glucose to **A**. *(1 mark)*
- **ii)** **A** to **C**. *(1 mark)*

c State the reaction conditions and any reagents necessary to convert:
- **i)** **A** to **B** *(2 marks)*
- **ii)** **C** to **A** *(2 marks)*
- **iii)** **D** to **A** and **B** *(2 marks)*

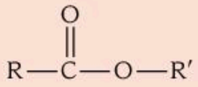
d Write a balanced chemical equation for the reaction between **B** and magnesium and name compound **E**. *(3 marks)*

Total 15 marks

Extended response question

8 a Describe how you could prepare a pure sample of ethanol from a suitable carbohydrate in the laboratory. Include an equation for the reaction and name the most suitable method of obtaining a sample of ethanol from the fermentation mixture. *(5 marks)*

b Fats and oils are used to manufacture soaps. Fats and oils are esters which consist of long-chain fatty acids and glycerol.
- **i)** If the structure of a fat is represented by the following general formula,



$$R-\overset{\overset{\displaystyle O}{\|}}{C}-O-R'$$

write a balanced chemical equation to represent the formation of a soap from this fat. *(2 marks)*
- **ii)** Calculate the number of hydrogen atoms that would be present in the R group of the fat represented in **i)** above if that group contains 18 carbon atoms. *(1 mark)*
- **iii)** Glycerol is a by-product in the manufacture of soap. The formula of glycerol is $C_3H_5(OH)_3$. Would you expect glycerol to be soluble or insoluble in water? Give a reason for your answer. *(2 marks)*
- **iv)** Describe what you would expect to observe when glycerol reacts with sodium. Give a balanced equation to represent the reaction. *(3 marks)*
- **v)** State TWO ways in which a soap prepared from a fat or oil differs from a soapless detergent. *(2 marks)*

Total 15 marks

B16 Polymers

Objectives

By the end of this topic you will be able to:

- define the term polymer
- explain the process of addition polymerisation
- show how alkenes are linked in the structure of a polyalkene
- give examples of polyalkenes
- state the uses of different polyalkenes.

Polymers are very large molecules known as macromolecules. They can be natural or synthetic, i.e. man-made. Proteins, starch and cellulose are examples of natural polymers, whereas polystyrene, PVC, terylene and nylon are examples of synthetic polymers. Synthetic polymers are collectively called plastics. Understanding the chemistry of polymers is important as it allows us to use and recycle them efficiently.

B16.1 Addition polymers

The carbon atom is unique among all elements that are found in nature. It can form long-chain molecules known as macromolecules. Some of these macromolecules are known as **polymers**. Polymers are made from small molecules known as **monomers** which are linked together, usually in chains. A polymer molecule may contain thousands of monomers.

The reaction which produces a polymer from monomers is known as **polymerisation**.

Addition polymers are synthetic, i.e. man-made, polymers which are made by **addition polymerisation**. Addition polymerisation occurs when monomers containing a carbon–carbon double bond, i.e. alkenes, are linked together to form a polymer. For this to occur, one of the bonds in the double bond of each alkene molecule breaks. The molecules then bond to one another by single covalent bonds between adjoining carbon atoms to form a polymer.

Key fact

A **polymer** is a macromolecule formed by linking together 50 or more small molecules, known as monomers, usually in chains.

Polyalkenes

Addition polymers produced from alkenes by addition polymerisation are known as **polyalkenes**. Polyalkenes are saturated compounds because they only have single bonds between carbon atoms.

The carbon–carbon single bond (–C–C–) which links each monomer unit together is known as an **alkane linkage**. The specific name of a polymer is derived from the name of the monomer from which it was formed. Some examples of polyalkenes are given in the Table 16.1.1.

▼ **Table 16.1.1** Names of some polyalkenes

Monomer	ethene	propene	monochloroethene
Polymer	polyethene	polypropene	polychloroethene

We will now look at the formation of polyethene and polypropene in more detail.

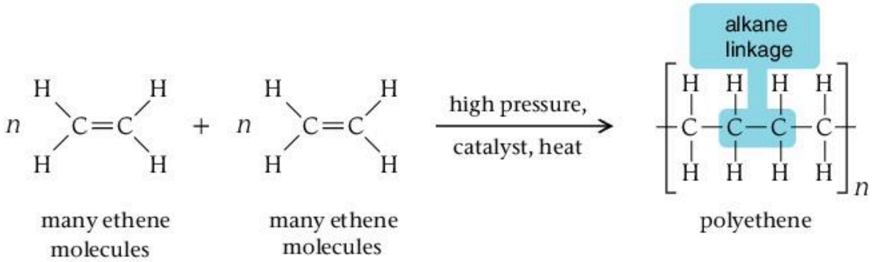
Formation of polyethene

Polyethene is formed by placing ethene (C_2H_4) in a pressurised container and heating it with a catalyst. One of the bonds in the double bond of each ethene molecule breaks and the saturated polymer, polyethene, is formed.

The addition polymerisation of ethene can be represented by the following:

$$\text{ethene} \xrightarrow[\text{catalyst, heat}]{\text{high pressure}} \text{polyethene}$$

In the formation of polyethene, thousands of ethene molecules join together. We can represent the reaction using structural formulae as in Figure 16.1.1.

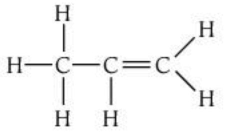


▲ **Figure 16.1.1** The formation of polyethene

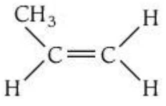
Formation of polypropene

Polypropene is formed by placing propene (C_3H_6) in a pressurised container and heating it with a catalyst. Thousands of propene molecules then join together to form polypropene.

In order to show the polymerisation of propene using structural formulae, we must redraw the propene molecule shown in **A** below so that only the two carbon atoms joined by the double bond are drawn horizontally. This is shown in **B** below.

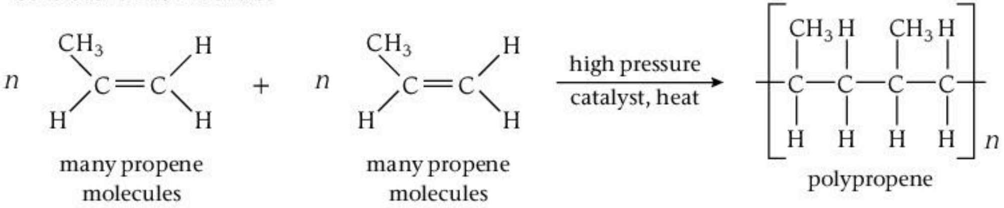


A



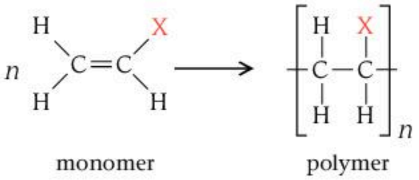
B

We can now show how propene can polymerise to form polypropene using structural formulae:



▲ **Figure 16.1.2** The formation of polypropene

If the monomer contains other atoms, groups of atoms or functional groups, e.g. chlorine, fluorine or the ethyl group, then the monomer and the resulting polymer can be represented as in Figure 16.1.3.



where X is Cl, F, C_2H_5 etc.

▲ **Figure 16.1.3** Formation of a polyalkenes

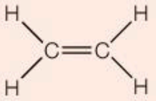
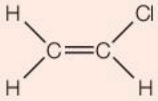
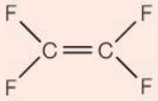
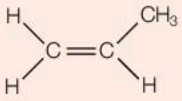
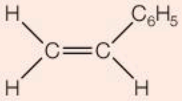
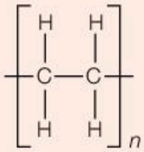
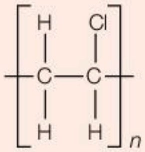
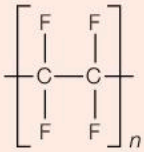
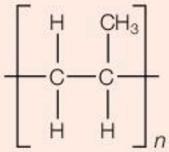
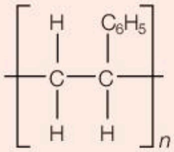
In addition polymerisation:

- the polymer is the only product – the breaking of the double bond makes it possible for the monomers to link together without losing any atoms from between them

- the monomers which make the polymer are all the same
- the empirical formula of the polymer is the same as the monomer that formed it since no atoms are lost during its formation, therefore, the ratio of atoms in the two is the same.

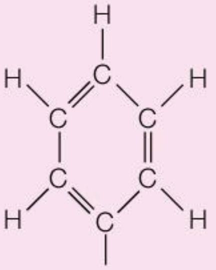
Examples of monomers and their addition polymers are given in the Table 16.1.2.

▼ **Table 16.1.2** Examples of monomers and their addition polymers

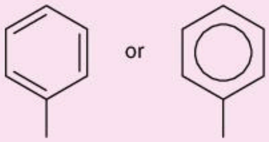
Monomer	 ethene	 chloroethene	 tetrafluoroethene	 propene	 styrene
Polymer	 polyethene	 polychloroethene	 polytetrafluoroethene	 polypropene	 polystyrene

? **Did you know?**

In **polystyrene**, C_6H_5 is a ringed structure of six carbon atoms as shown below.



This can also be represented as:



Some addition polymers are more commonly known by other names, for example, polychloroethene is commonly known as polyvinyl chloride or PVC, polytetrafluoroethene is commonly know°n as PTFE and polystyrene is often called styrofoam.

Uses of polyalkenes

Because polyalkenes can be moulded easily into different shapes, are light in weight, are coloured easily, can be made to be flexible or rigid, are good thermal and electrical insulators and are durable, i.e. they do not corrode or decay, they have a great many uses. Table 16.1.3 summarises some of these uses.

▼ **Table 16.1.3** Uses of some polyalkenes

Name of polymer	Uses
Polyethene	To make plastic bags, plastic bottles, sandwich bags, cling wrap, packaging materials and moisture barriers used in construction.
Polychloroethene (PVC)	To make pipe fittings, water pipes, guttering, plastic windows and insulation for electrical wires and cables.
Polytetrafluoroethene (PTFE)	As a non-stick coating on kitchen and garden utensils and to make low-friction bearings.
Polypropene	To make ropes, plastic toys, carpets and plastic food containers.
Polystyrene (styrofoam)	To make fast food containers and vending cups for drinks. As a protective packaging material and an insulation material in walls and roofs.

▲ **Figure 16.1.4** Polyvinyl chloride is used to make guttering

◀ **Figure 16.1.5** Polypropene is used to make ropes

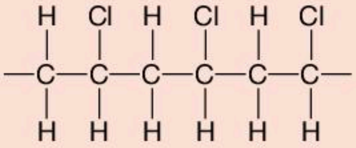
Summary questions

1 What is a polymer?

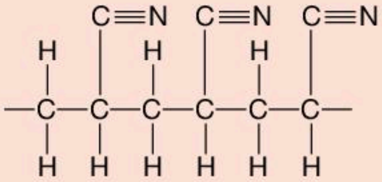
2 Explain what happens during addition polymerisation.

3 For each of the following polymers, draw the structural formula of the monomer that formed the polymer:

a



b



4 Give three examples of addition polymers and give two uses of each of them.

B16.2 Condensation polymers

Condensation polymers are made by **condensation polymerisation**. Condensation polymerisation involves linking monomers together in long chains by eliminating a small molecule from between adjacent monomer units. The molecule is often water (H_2O), where one hydrogen atom (H) comes from one monomer unit and an oxygen and a hydrogen atom (OH) come from the adjacent monomer. Instead of water, other small molecules can also be eliminated such as hydrogen chloride (HCl) or ammonia (NH_3). The elimination of a small molecule makes it possible for the monomers to link up in chains.

In contrast to addition polymerisation, during condensation polymerisation:

- the polymer is not the only product made – condensation polymerisation always forms *two* products, the polymer and another compound made of small molecules
- the polymer is often made from more than one different type of monomer
- the empirical formula of the polymer is different from the monomer since atoms are lost during the formation of the polymer.

In order for monomers to form condensation polymers, the monomer molecules must have *two* functional groups, e.g. hydroxyl groups (–OH) and/or carboxyl groups (–COOH). We will be studying three different types of condensation polymers, polyesters, polyamides and polysaccharides.

Objectives

By the end of this topic you will be able to:

- explain the process of condensation polymerisation
- show how monomers are linked in the structure of a polyester, a polyamide and a polysaccharide
- give examples of polyesters, synthetic and natural polyamides and polysaccharides
- state the uses of different condensation polymers
- discuss the harmful effects of synthetic polymers on the environment.

Polyesters

Polyesters are synthetic condensation polymers that are formed from the condensation reaction between an alkanoic acid and an alcohol. The acid and the alcohol must both have *two* functional groups in their molecules. The acid is, therefore, known as a diacid and the alcohol is known as a dialcohol.

The number of carbon atoms in the acid and the alcohol can vary, so we can represent the carbon chain to which the functional groups are attached by X or Y. The structural formulae in Figure 16.2.1 show the two monomers that are involved in the formation of a polyester.

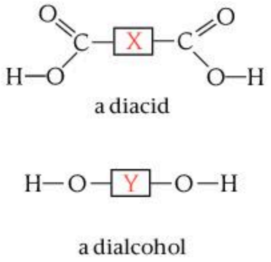


Figure 16.2.1 Structural formulae of a diacid and a dialcohol

When these two molecules react with each other, a water molecule is eliminated from between them and an **ester functional group** (–COO– or –OOC–) forms which links the two together. The structure of the ester functional group is shown in Figure 16.2.2.

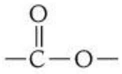

$$-\overset{\overset{\displaystyle O}{\|}}{C}-O-$$

Figure 16.2.2 The ester functional group or ester linkage

A polyester forms by linking thousands of these monomers together alternately in chains, with an ester functional group between adjacent monomers. Therefore, we say that polyesters are polymers where the linkages between monomers are **ester linkages**.

Figure 16.2.3 uses a diacid molecule and a dialcohol molecule to give a simple representation of how they can be condensed to form a polyester molecule.

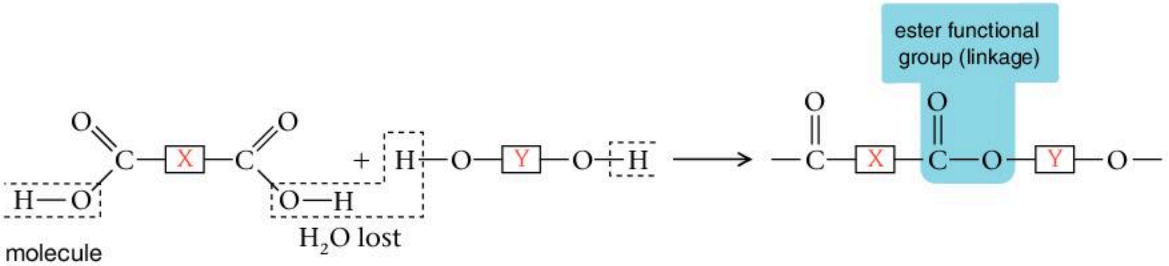


Figure 16.2.3
A section of a polyester molecule

An example of a polyester is polyethylene terephthalate, also known as PET, or by brand names such as Terylene or Dacron. Polyethylene terephthalate is formed from a diacid in which X has the formula C_6H_4 and a dialcohol in which Y has the formula C_2H_4.

Polyamides

Polyamides all contain the **amide functional group** (–CONH– or –NHCO–). The structure of the amide functional group is shown in Figure 16.2.4.

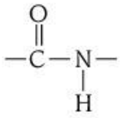

$$-\overset{\overset{\displaystyle O}{\|}}{C}-\underset{\underset{\displaystyle H}{|}}{N}-$$

Figure 16.2.4 The amide functional group or amide linkage

Some polyamides are synthetic while others occur naturally. We will be looking at both types in this section.

Synthetic polyamides

Synthetic polyamides are made from a diacid and a diamine. A diamine has two amino functional groups as shown in Figure 16.2.5.

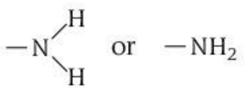


Figure 16.2.5 The amino functional group

The structural formulae in Figure 16.2.6 show the two monomers that are involved in the formation of a synthetic polyamide.

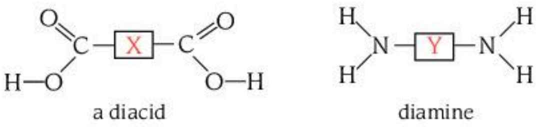
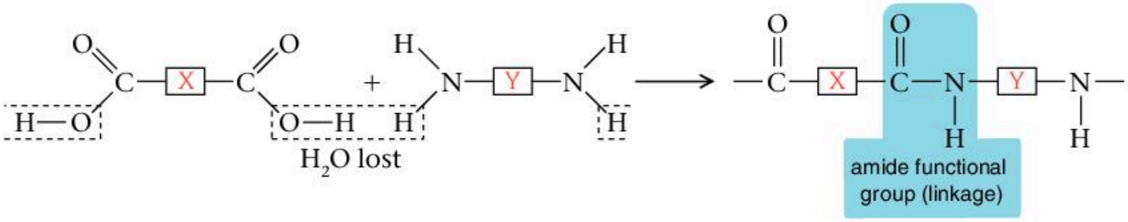


Figure 16.2.6 Structural formulae of a diacid and a diamine

When these two molecules react with each other, a water molecule is eliminated from between them and an amide functional group forms which links the two together. A polyamide forms by linking thousands of these monomers together alternately in chains, with an amide functional group between adjacent monomers. Therefore, we say that polyamides are polymers where the linkages between monomers are **amide linkages**.

Figure 16.2.7 uses a diacid molecule and a diamine molecule to give a simple representation of how they can be condensed to form a synthetic polyamide molecule.

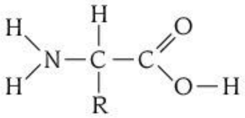


▲ **Figure 16.2.7** A section of a polyamide molecule

An example of a synthetic polyamide is nylon 6,6, which is formed from a diacid in which X has the formula C_4H_8 and a dialcohol in which Y has the formula C_6H_{12}.

Natural polyamides

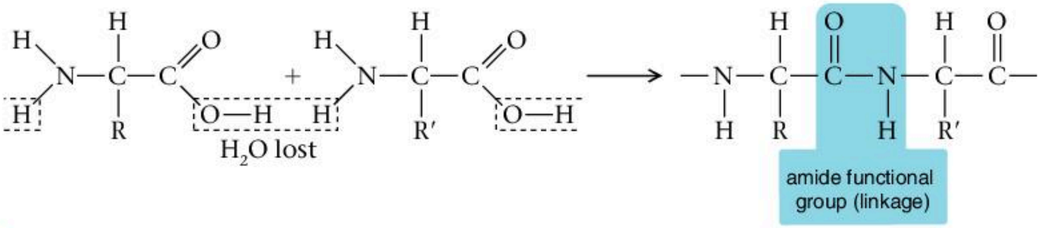
Proteins are **natural polyamides** formed by linking monomers known as **amino acids** together. Amino acids are molecules which have one amino functional group and one carboxyl functional group. The structural formula of an amino acid is given in Figure 16.2.8.



◀ **Figure 16.2.8** Structural formula of an amino acid molecule

In the amino acid molecule, R represents a variable hydrocarbon group which may contain other atoms such as nitrogen, oxygen and sulfur, as well as carbon and hydrogen. There are 20 different groups, therefore there are 20 different amino acids which make up protein molecules.

Figure 16.2.9 uses two amino acid molecules to give a simple representation of the condensation of amino acid molecules to form a protein molecule.



▲ **Figure 16.2.9** Formation of a protein molecule

Like synthetic polyamides, proteins contain the amide functional group or amide linkage. In proteins, this amide linkage is also known as the **peptide linkage**.

Examples of proteins include keratin which is found in hair and nails and collagen which is found in the connective tissue that makes up structures such as ligaments and tendons. Haemoglobin, enzymes, antibodies and many hormones are also made of proteins.

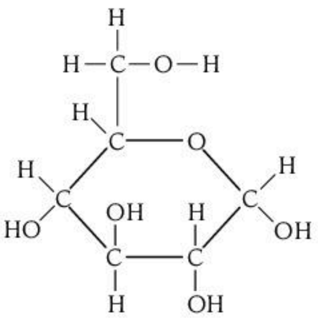
Polysaccharides

Polysaccharides are natural polymers formed by linking monomers known as **monosaccharides** together. Examples of monosaccharides include glucose, fructose and galactose. These are all isomers with the molecular formula $C_6H_{12}O_6$.

Starch is a polysaccharide formed when glucose undergoes condensation polymerisation. The simplified chemical equation for this reaction can be written as follows, where *n* is the number of repeating glucose units in the polysaccharide:

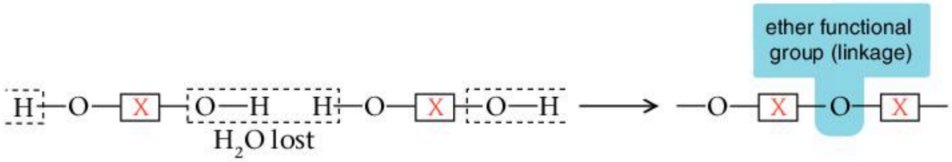
$$\underset{\text{glucose}}{nC_6H_{12}O_6} \longrightarrow \underset{\text{starch}}{(C_5H_{10}O_5)_n} + nH_2O$$

The structural formula of a glucose molecule is given in Figure 16.2.10.



▲ **Figure 16.2.10** Structural formula of a glucose molecule

If we think of *two* of the –OH groups of a glucose molecule as being functional groups, then we can represent the glucose molecule as H–O–X–O–H, where X represents the rest of the glucose molecule, i.e. $C_6H_{10}O_4$. Figure 16.2.11 uses two glucose molecules to give a simple representation of the condensation of glucose molecules to form a starch molecule.



▲ **Figure 16.2.11** Formation of a starch molecule

Polysaccharides contain the **ether linkage** (– O –) between the monomers. Other polysaccharides include glycogen or animal starch and cellulose which is found in plant cell walls.

Uses of condensation polymers

Like polyalkenes, synthetic condensation polymers have a variety of uses. Their molecules are extremely long and one of their main uses is to make fibres. These fibres are strong, light in weight and durable, which makes them very useful in the production of a variety of fabrics. Table 16.2.1 summarises some of the uses of both synthetic and natural condensation polymers.

▼ **Table 16.2.1** Uses of some condensation polymers

Type of polymer	Name of polymer	Uses
Polyester	Polyethylene terephthalate (PET, Terylene or Dacron)	To make clothing, especially outdoor clothing, bed sheets, towels, boat sails and carpets. To make PET bottles for soft drinks.
Polyamide	Nylon	To make fishing lines, seatbelts, carpets and clothing, especially where stretch is required, e.g. nylon stockings and activity wear.
Polyamide	Protein	To build body cells, hair and nails. To make enzymes, antibodies and some hormones.
Polysaccharide	Starch	Used to produce energy during respiration in living organisms after first being digested to glucose. Stored as a food reserve in living organisms.
Polysaccharide	Cellulose	Used to build cell walls of plant cells.

Exam tip

It is very important that you know the **uses** of the different types of polymers.

Environmental impact of synthetic polymers

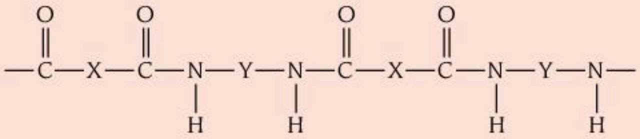
Synthetic polymers, otherwise known as **plastics**, can have a variety of harmful effects on living systems and the environment. Some of these are given below.

- During the manufacture of plastics, a variety of toxic chemicals are released into the environment. Some of these continue to be released from the plastic items themselves during use and when disposed of. Many of these chemicals are persistent, i.e. they remain in the environment doing damage for a great many years.
- Most plastics are non-biodegradable. They build up in the environment leading to land and water pollution. They are particularly harmful to aquatic organisms, often leading to their death by being ingested or by causing suffocation.
- Plastics produce dense smoke and poisonous gases when burnt. These can lead to air pollution and serious health problems, such as asthma, immune system damage, cancer and damage to the nervous and reproductive systems.
- Many plastics are flammable. As a result they pose fire hazards.
- Plastics are made from petroleum which is a non-renewable resource. Their manufacture is contributing to the world wide depletion of petroleum.

▲ **Figure 16.2.12** Polyethylene terephthalate is used to make **(a)** boat sails and **(b)** PET bottles

Summary questions

1. Explain what happens during condensation polymerisation.
2. Give two differences between addition polymerisation and condensation polymerisation.
3. Draw the partial structure of a polyester which is made from four monomer units.
4. Draw the structural formulae of the two monomer units in the following polymer:



$$-\overset{\overset{\displaystyle O}{\|}}{C}-X-\overset{\overset{\displaystyle O}{\|}}{C}-\underset{\underset{\displaystyle H}{|}}{N}-Y-\underset{\underset{\displaystyle H}{|}}{N}-\overset{\overset{\displaystyle O}{\|}}{C}-X-\overset{\overset{\displaystyle O}{\|}}{C}-\underset{\underset{\displaystyle H}{|}}{N}-Y-\underset{\underset{\displaystyle H}{|}}{N}-$$

5. Give two examples of synthetic condensation polymers and one example of a natural condensation polymer. Give two uses of each polymer named.

Key concepts

- A **polymer** is a macromolecule formed by linking together 50 or more small molecules, known as monomers, usually in chains.
- Polymers can be synthetic, i.e. man-made, or natural. Synthetic polymers are collectively known as **plastics**.
- The reaction which forms a polymer is known as **polymerisation**.
- **Addition polymers** are formed by **addition polymerisation**. Alkenes form addition polymers when one bond in the carbon–carbon double bond of each alkene molecule breaks. The molecules then bond to one another by single covalent bonds between adjoining carbon atoms.
- In addition polymerisation the only product is the polymer, all the monomers are identical and the empirical formulae of the polymer and the monomer are the same.
- Alkenes form polymers known as **polyalkenes**.
- Polyalkenes contain the **alkane linkage** (–C–C–) between monomers.
- Polyethene, polypropene, polychloroethene (PVC) and polystyrene are examples of polyalkenes.
- Polyalkenes have a variety of uses.
- **Condensation polymers** are formed by **condensation polymerisation**. Condensation polymerisation involves linking monomers together in long chains by eliminating a small molecule from between adjacent monomer units.
- In condensation polymerisation the polymer and another compound composed of small molecules are produced, the monomers are not necessarily identical and the empirical formulae of the polymer and the monomer are different.
- **Polyesters** are synthetic condensation polymers formed by the condensation of a diacid and a dialcohol.
- Polyesters contain the **ester linkage** (–COO–) between monomers.
- Polyethylene terephthalate (PET), also known as Terylene or Dacron, is an example of a polyester.
- **Synthetic polyamides** are condensation polymers formed by the condensation of a diacid and a diamine.
- Nylon is an example of a synthetic polyamide.
- **Natural polyamides** are condensation polymers formed by the condensation of amino acids.
- Proteins are natural polyamides.
- All polyamides contain the **amide linkage** (–CONH–) between monomers.
- **Polysaccharides** are natural condensation polymers formed by the condensation of monosaccharides.
- Polysaccharides contain the **ether linkage** (–O–) between monomers.
- Starch, glycogen and cellulose are examples of polysaccharides.
- Condensation polymers have a variety of uses.
- Synthetic polymers can have a variety of harmful effects on living systems and the environment.

Practice exam-style questions

Multiple-choice questions

1 Which of the following is not an addition polymer?
- **A** polystyrene
- **B** terylene
- **C** polyethene
- **D** polypropene

2 Amide linkages are found in:
- I proteins
- II terylene
- III nylon

- **A** I and II only
- **B** II only
- **C** I and III only
- **D** I, II and III

3 Amino acids are so called because they:
- **A** contain the NH_2 group
- **B** contain the COOH group
- **C** turn blue litmus red
- **D** contain the NH_2 group and the COOH group

4 In a condensation polymerisation reaction,
- **A** the polymer has a different empirical formula to the monomer
- **B** one product is formed
- **C** the polymer is always formed from one type of monomer only
- **D** water is always produced

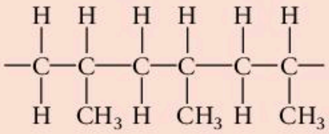
5 The monomer of starch is:
- **A** fructose
- **B** sucrose
- **C** galactose
- **D** glucose

6 Which of the following is not true about plastics?
- **A** They are flammable.
- **B** They are easily biodegraded.
- **C** They are good insulators.
- **D** They are easily moulded.

Structured question

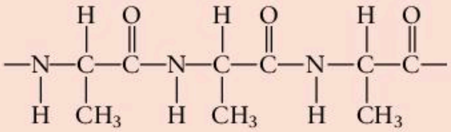
7 **a** The structures **A**, **B** and **C** represent three types of polymers.

A



$$-\underset{H}{\overset{H}{C}}-\underset{CH_3}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{CH_3}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{CH_3}{\overset{H}{C}}-$$

B

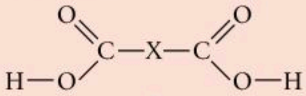


$$-\underset{H}{N}-\underset{CH_3}{\overset{H}{C}}-\overset{O}{\overset{\|}{C}}-\underset{H}{N}-\underset{CH_3}{\overset{H}{C}}-\overset{O}{\overset{\|}{C}}-\underset{H}{N}-\underset{CH_3}{\overset{H}{C}}-\overset{O}{\overset{\|}{C}}-$$

C —O—X—O—X—O—X—

Use the information to answer the following questions.
- **i)** What type of polymerisation is involved in the formation of EACH polymer? *(3 marks)*
- **ii)** Draw the structural formula of EACH monomer unit of **A** and of **B**. *(2 marks)*
- **iii)** Name the monomer unit of **B**. *(1 mark)*
- **iv)** Name ONE naturally occurring polymer that could be represented by **C** and name the monomer which makes up the polymer you have named. *(2 marks)*

b The two compounds, **E** and **F**, illustrated below, undergo polymerisation.



E

HO—Y—OH

F

- **i)** Using TWO molecules of **E** and TWO molecules of **F**, show how the compounds are linked together to form a polymer. *(2 marks)*
- **ii)** Name the type of linkage formed between the monomers in the polymer you have drawn in part **i)** above and name the type of polymer. *(2 marks)*
- **iii)** Give a named example of the type of polymer drawn in **i)** above and state TWO of its uses. *(3 marks)*

Total 15 marks

Extended response question

8 Polymers are found in every aspect of our daily lives.

a What are polymers? *(1 mark)*

b Polymers can be classified into two groups, addition polymers and condensation polymers. Briefly discuss ONE named addition polymer and ONE named synthetic condensation polymer. You should include in your answer:
- **i)** The structure of the monomer units of EACH polymer. *(2 marks)*
- **ii)** Appropriate structures to show how the monomers are linked to form EACH polymer. *(4 marks)*
- **iii)** The name of the type of linkage formed between the monomer units in EACH polymer. *(2 marks)*
- **iv)** TWO uses of EACH polymer. *(2 marks)*

c Addition polymers are formed by a process called addition polymerisation and condensation polymers are formed by condensation polymerisation. Give TWO differences between the two types of polymerisation. *(2 marks)*

d Give TWO disadvantages of the extensive use of synthetic polymers (plastics) in today's world. *(2 marks)*

Total 15 marks

Section C

C17 Characteristics of metals

Objectives

By the end of this topic you will be able to:

- describe the physical properties of metals
- relate the physical properties of metals to the way in which metal atoms are bonded in the metallic lattice.

Most elements in the periodic table are metals. In fact, metals account for about 80% of all the elements in the table. Metallurgy, i.e. the study of metals, is one of the oldest applied sciences, dating back to 6000BC with the discovery of gold. Metals have a variety of physical and chemical properties which enable them to play an extremely important part in our daily lives.

C17.1 Physical properties of metals

Metals are elements whose atoms have a small number of valence electrons, usually 1, 2 or 3. Most metals are found in Groups I, II and III of the periodic table and between Groups II and III as the transition metals.

In Unit 5.4 you learnt that the type of bonding between metal atoms in a metal is known as **metallic bonding** and that metals have very distinct physical properties as a result of this bonding. Recall that a metal lattice is composed of metal cations surrounded by a 'sea' of delocalised, mobile electrons and that the lattice is held together by the strong electrostatic force of attraction between the positive cations and the delocalised, negative electrons known as the **metallic bond**.

The way in which the metal atoms are bonded in metal lattices helps us to understand the physical properties of metals. We will consider some of these in detail below.

▲ **Figure 17.1.1** Because metals are good conductors of heat they are used to make frying pans

Metals have high melting points and boiling points

The electrostatic forces of attraction between the positive cations and negative, delocalised electrons are relatively *strong*. As a result, large amounts of heat energy are required to separate the atoms in order to melt or boil the metals. All metals are found in the *solid* state at room temperature, with the exception of mercury which is a *liquid*.

Metals are good conductors of electricity and heat

The **delocalised electrons** in the metal lattice are not held in any fixed position and are free to move. These moving electrons act as charge carriers allowing an electric current to be carried through the metal. They also act as heat carriers allowing heat to be carried through the metal when, for example, one end of a piece of metal is placed in a hot environment, such as a flame, the other end of the metal quickly heats up.

Metals have a shiny lustre

Lustre is the shininess of a metal. Metals generally have a very shiny, metallic lustre. This is because the mobile electrons reflect the photons of light back from the metal surface before they are allowed to penetrate far into the surface.

▶ **Figure 17.1.2** This silver cup has a shiny metallic lustre

Metals are hard, malleable and ductile

Metals are considered to be hard because they are not easily damaged when a force is applied to them. They are, however, **malleable** meaning that they can be bent and hammered into different shapes. They are also **ductile**, meaning that they can be drawn out into wires. This is because the atoms in the metal are all of the same type, therefore they are all the same size. If force is applied to the metal, the atoms can roll over each other into new positions without breaking the metallic bond.

▲ **Figure 17.1.3** This gold ring demonstrates the malleability of metals

Metals have high densities

The density of any substance is calculated by dividing the mass of the substance by its volume. Metals have *high* densities because their atoms are packed very closely together. This means that there are as many metal atoms as possible in a given volume of metal.

Summary questions

1 Explain why metals have high melting points and boiling points.
2 Why are metals good conductors of electricity and heat?
3 Explain the terms 'malleable' and 'ductile'. Explain why metals are malleable and ductile.
4 Why do metals have a high density?

C17.2 Chemical properties and reactions of metals

Because metals are elements whose atoms have a small number of valence electrons, usually 1, 2 or 3, they form positive cations in chemical reactions by *losing* these valence electrons. This can be summarised as follows:

$$M \longrightarrow M^{n+} + ne^-$$

As a result, metals form **ionic compounds** when they react and the metal behaves as a **reducing agent** since it gives electrons to the other reactant, i.e. it causes the other reactant to gain electrons (RIG).

Some metals are relatively stable while others react violently when exposed to other substances, e.g. oxygen, water and acids. The way in which a particular metal reacts gives an indication of its reactivity. The reactivity of a metal is related to the metal's position in the periodic table.

We will now look at the reactions of certain metals with oxygen, water and dilute acids.

Objectives

By the end of this topic you will be able to:

- describe the reactions of various metals with oxygen, water and dilute acids
- name the products formed by the reactions
- write balanced chemical equations for the reactions of various metals with oxygen, water and dilute acids.

▲ **Figure 17.2.1** Sodium has to be stored under paraffin to prevent it from reacting with oxygen in the air

The reactions of metals with oxygen

Some metals can be exposed to air without reacting to any great extent with the oxygen in the air. These metals include the ones that we use regularly and see around us. The surface of other metals **oxidises** immediately on exposure to air and some metals react so vigorously with oxygen they have to be stored under paraffin, e.g. potassium and sodium. The paraffin acts as a barrier, preventing the metals from coming into contact with the air.

When metals react with oxygen, they form ionic compounds known as **metal oxides**. This can be summarised by the following general word equation:

$$\textbf{metal + oxygen} \longrightarrow \textbf{metal oxide}$$

Table 17.2.1 shows what happens when certain metals react with oxygen in the air.

▼ **Table 17.2.1** Reactions of metals with oxygen

Metal	Description of the reaction when exposed to dry air	Description of the reaction when heated in air	Products formed when heated and equation
Potassium	Reacts very readily forming potassium oxide	Burns vigorously with a lilac flame	Forms a white powdery solid, potassium oxide: $4K(s) + O_2(g) \longrightarrow 2K_2O(s)$
Sodium	Reacts very readily forming sodium oxide	Burns vigorously with an orange flame	Forms a white powdery solid, sodium oxide: $4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$
Calcium	Reacts readily to form coating of calcium oxide	Burns very easily with a brick red flame	Forms a white powdery solid called calcium oxide: $2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$
Magnesium	Reacts slowly to form a coating of magnesium oxide	Burns easily with a bright white flame	Forms a white powdery solid, magnesium oxide: $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
Aluminium	Reacts slowly to form a thin coating of aluminium oxide	Burns when heated strongly, especially if powdered	Forms a white powdery solid, aluminium oxide: $4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$
Zinc	Reacts very slowly to form a thin coating of zinc oxide	Burns when heated strongly, especially if powdered	Forms a white powdery solid, zinc oxide, which is yellow when hot: $2Zn(s) + O_2(g) \longrightarrow 2ZnO(s)$
Iron	Does not react with dry air	Burns when heated strongly, especially if powdered	Forms a black powdery solid, iron(II, III) oxide: $3Fe(s) + 2O_2(g) \longrightarrow Fe_3O_4(s)$
Copper	Does not react with dry air	Does not burn when heated, but does form an oxide coating if heated very strongly	Forms a black solid, copper(II) oxide: $2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$
Silver	Does not react with dry air	Does not react, even when heated very strongly	–

If the metal oxides produced in these reactions can dissolve in water, the resulting solution is **alkaline**. For example, potassium oxide reacts with water to form soluble potassium hydroxide, sodium oxide reacts with water to form soluble sodium hydroxide and calcium oxide reacts with water to form calcium hydroxide which is slightly soluble. Potassium hydroxide, sodium hydroxide and calcium hydroxide are alkalis. All the other oxides in Table 17.2.1 are insoluble.

Looking at the reactions of metals with oxygen in the air we see that potassium, sodium, calcium and magnesium are the most reactive metals. These metals are found in Groups I and II of the periodic table, i.e. to the far left of the table. In general, the reactivity of metals *decreases* from left to right across the periodic table.

The reactions of metals with water

Metals can also be added to water to determine their reactivity. In general, metals that are very reactive with oxygen are highly reactive when exposed to water.

When a metal reacts with water, the metal hydroxide and hydrogen are produced. The reaction can be summarised by the following general word equation:

metal + water ⟶ metal hydroxide + hydrogen

When a metal reacts with steam, the metal oxide and hydrogen are produced. The reaction can be summarised by the following general word equation:

metal + steam ⟶ metal oxide + hydrogen

Table 17.2.2 shows what happens when certain metals are allowed to come into contact with water or steam.

Table 17.2.2 Reactions of metals with water or steam

Metal	Description of the reaction	Products formed and equation
Potassium	Reacts very vigorously with cold water producing a lilac flame	Potassium hydroxide and hydrogen: $2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$
Sodium	Reacts vigorously with cold water, producing an orange flame	Sodium hydroxide and hydrogen: $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$
Calcium	Reacts moderately with cold water	Slightly soluble calcium hydroxide and hydrogen: $Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$
Magnesium	Reacts very slowly with cold water and slowly with hot water	With water: very slightly soluble magnesium hydroxide and hydrogen: $Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$
	Reacts vigorously with steam	With steam: magnesium oxide and hydrogen: $Mg(s) + H_2O(g) \longrightarrow MgO(s) + 2H_2(g)$
Aluminium	Does not react with cold or hot water. Reacts with steam	Aluminium oxide and hydrogen: $2Al(s) + 3H_2O(g) \longrightarrow Al_2O_3(s) + 3H_2(g)$
Zinc	Does not react with cold or hot water. Reacts with steam	Zinc oxide and hydrogen: $Zn(s) + H_2O(g) \longrightarrow ZnO(s) + H_2(g)$
Iron	Does not react with cold or hot water. Reacts with steam	Iron(II,III) oxide and hydrogen: $3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$
Copper	Does not react with water or steam	–
Silver	Does not react with water or steam	–

It can be seen from Table 17.2.2 that some of the metals do not react with water. They do, however, react with steam. The fact that they require steam to react means that they are *less* reactive than the ones that react with cold water. Once again, potassium, sodium and calcium are metals that react most vigorously with water.

When potassium and sodium react with cold water, they both roll into a ball and move rapidly around on the surface of the water as they decrease in size. As they move, a lilac flame forms around the ball of potassium and an orange flame forms around the ball of sodium. The lilac flame around the potassium is shown in Figure 17.2.2.

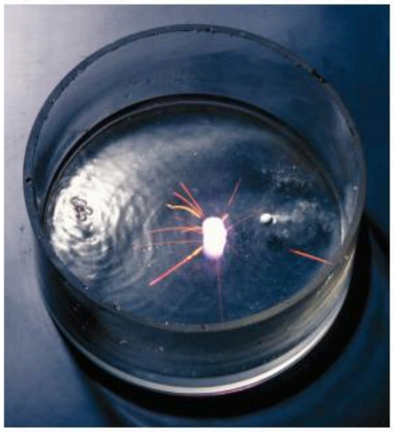


Figure 17.2.2 A lilac flame forms around potassium when it reacts with cold water

Practical activity

Investigating reactions of metals with water

(Teacher demonstration)

Your teacher may use this activity to assess:

- observation, recording and reporting.

Your teacher will perform the following experiment.

Great care must be taken when reacting potassium and sodium with water.

Method

1. Place about 2 cm^3 of distilled water into a test tube and add a piece of clean magnesium ribbon. Leave undisturbed for one hour.
2. Look for signs of effervescence and dip a piece of red litmus into the solution.
3. Place about 2 cm^3 of distilled water into another test tube and add a piece of calcium.
4. Look for effervescence and test the gas for the presence of hydrogen by placing a burning splint at the mouth of the tube. Dip a piece of red litmus paper into the solution.
5. Half fill a large trough with water and place a very small piece of potassium onto the surface of the water.
6. Look for a trail of effervescence as the potassium moves around on the surface of the water. Note the colour of the flame that forms around the potassium. When the potassium disappears, dip a piece of red litmus paper into the solution.
7. Repeat step **6** with a small piece of sodium. If a flame does not form around the ball of sodium, test for the presence of hydrogen by touching the sodium with a burning splint.
8. Which metal reacted the most vigorously with water and which reacted the least vigorously?
9. Write balanced chemical equations for the four reactions.

The reactions of metals with dilute acids

In general, reactive metals react with acids, except nitric acid, to form a salt and hydrogen. We can again see that the metals which are very reactive with oxygen and water are also very reactive with acids. The reaction can be summarised by the following general word equation:

$$\textbf{reactive metal + acid} \longrightarrow \textbf{salt + hydrogen}$$

The salt that is produced depends on the metal used and the acid.

Table 17.2.3 shows what happens when certain metals react with dilute hydrochloric acid. When a metal reacts with hydrochloric acid the salt formed is a chloride.

▼ **Table 17.2.3** Reactions of metals with hydrochloric acid

Metal	Description of the reaction	Products formed and equation	
Potassium	An extremely violent reaction occurs	Potassium chloride and hydrogen:	$2K(s) + 2HCl(aq) \longrightarrow 2KCl(aq) + H_2(g)$
Sodium	A violent reaction occurs	Sodium chloride and hydrogen:	$2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$
Calcium	A fairly violent reaction occurs	Calcium chloride and hydrogen:	$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$
Magnesium	A very vigorous reaction occurs	Magnesium chloride and hydrogen:	$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
Aluminium	A vigorous reaction occurs	Aluminium chloride and hydrogen:	$2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2(g)$
Zinc	A fairly vigorous reaction occurs	Zinc chloride and hydrogen:	$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
Iron	A very slow reaction occurs	Iron(II) chloride and hydrogen:	$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$
Copper	Does not react with dilute acids	–	
Silver	Does not react with dilute acids	–	

Looking at the reactions of metals with hydrochloric acid we can again see that potassium, sodium, calcium and magnesium are the most reactive metals.

The reactions between the metals and dilute sulfuric acid are very similar to the reactions which occur with dilute hydrochloric acid. However, when a metal reacts with sulfuric acid, the salt formed is a sulfate.

Did you know?

It is unadvisable to perform the reactions of potassium, sodium and calcium with an acid because these three metals react so **violently** that the reactions are dangerous.

Practical activity

Investigating reactions of metals with dilute acids

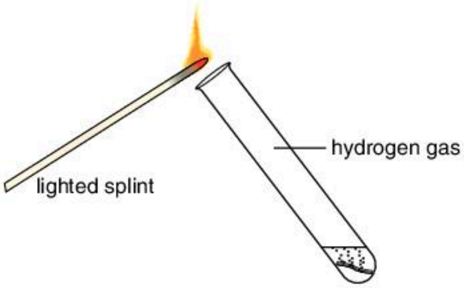
Your teacher may use this activity to assess:

- observation, recording and reporting
- analysis and interpretation.

You will be supplied with samples of the metals magnesium, aluminium, zinc, iron and copper, dilute hydrochloric acid, dilute sulfuric acid, five test tubes and a wooden splint.

Method

1. Carefully place about 2 cm^3 of dilute hydrochloric acid into each of the five test tubes.
2. Label each tube with the symbol of the particular metal that will be added.
3. Add a small piece of metal to each tube so that each tube contains a different metal as indicated by its label.
4. Observe each tube for signs of effervescence. If effervescence is seen, test the gas for the presence of hydrogen by placing a lighted splint at the mouth of the tube and feel the bottom of the tube to see if it becomes warm.
5. If no effervescence is seen, heat the tube gently. If effervescence is seen on heating, test the gas for the presence of hydrogen. Make sure that you do not confuse effervescence with bubbling that occurs if the acid boils. Effervescence continues after the tube is removed from the source of heat, boiling does not.



▲ **Figure 17.2.3** The production of hydrogen gas can be confirmed by holding a lighted splint near the mouth of the test tube. A 'squeaky pop' sound will identify the gas as hydrogen.

6 Repeat the method using dilute sulfuric acid in place of dilute hydrochloric acid.

7 Record your observations in a table, noting if the reaction needed heating.

8 Did the metals differ in their reactivity with the acids? If they did, which metal reacted the most vigorously and which showed no reaction at all, even when heated?

9 In what ways were the reactions between the metals and the two acids similar? Were there any differences?

10 Aluminium did not react as described in Table 17.2.3 until the test tubes were heated. Can you suggest why?

11 Write balanced chemical equations for the reactions of the metals with dilute sulfuric acid.

During the experiment, you possibly discovered that aluminium did not react with the acids until the test tubes were heated. This is because aluminium is coated with a thin film of **aluminium oxide** which prevents it from reacting with acids when cold. When heated, the acids react with the aluminium oxide forming soluble aluminium chloride or aluminium sulfate. These salts dissolve and the aluminium is then able to react with the acids.

Summary questions

1 Write a balanced chemical equation for each of the following reactions:
   - a calcium reacting with oxygen
   - b zinc reacting with hydrochloric acid
   - c iron reacting with sulfuric acid.

2 How do the products of the reactions differ when magnesium reacts with water and when it reacts with steam? Write a balanced chemical equation for each reaction.

3 Which seems to be more reactive, iron or copper?

4 Which seem to be the least reactive, Group I metals, Group II metals or the transition metals?

C17.3 Reactions of metal compounds

Objectives

By the end of this topic you will be able to:

- describe the reactions of metal oxides, hydroxides and carbonates with acids
- describe the decomposition reactions of metal nitrates, carbonates and hydroxides.

You have already come across some of the reactions of metal compounds. We will review these and look at some new reactions in this unit. The compounds we will be looking at include metal oxides, metal hydroxides, metal carbonates and metal nitrates.

Reactions of metal oxides

Metal oxides react with acids to form a salt and water. The reaction can be summarised by the following general word equation:

$$\textbf{metal oxide + acid} \longrightarrow \textbf{salt + water}$$

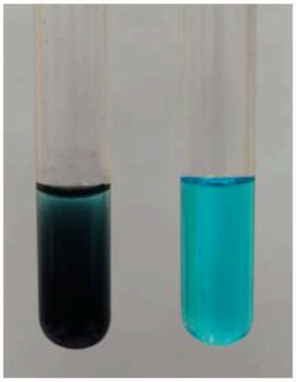
This reaction is known as a **neutralisation reaction** since the metal oxide is a base that neutralises the acid to form a salt and water. When a metal oxide reacts with an acid the reaction becomes warmer because it is exothermic.

Examples

1 Copper(II) oxide reacts with sulfuric acid to produce soluble copper(II) sulfate and water:

$$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

During the reaction, the black, powdery copper(II) oxide disappears slowly and a blue solution forms. This is because soluble copper(II) sulfate is produced which contains blue Cu^{2+} ions. The reaction can be speeded up by heating it.



▲ **Figure 17.3.1** Black powdery copper(II) oxide slowly disappears in sulfuric acid as it reacts to form soluble, blue copper(II) sulfate

2 Magnesium oxide reacts with hydrochloric acid to produce soluble magnesium chloride and water:

$$MgO(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$$

The white, powdery magnesium oxide disappears as it reacts with the acid to form a colourless solution of magnesium chloride.

Reactions of metal hydroxides

Metal hydroxides also react with acids to form a salt and water. The reaction can be summarised by the following general word equation:

metal hydroxide + acid ⟶ salt + water

This reaction is also a **neutralisation reaction** since the metal hydroxide is a base and it neutralises the acid to form a salt and water. The reaction is exothermic, therefore it becomes warmer.

Examples

1 Copper(II) hydroxide reacts with nitric acid to produce soluble copper(II) nitrate and water:

$$Cu(OH)_2(s) + 2HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2H_2O(l)$$

During the reaction, the blue copper(II) hydroxide disappears and a blue solution forms due to the production of soluble copper(II) nitrate which contains blue Cu^{2+} ions.

2 Zinc hydroxide reacts with hydrochloric acid to produce soluble zinc chloride and water:

$$Zn(OH)_2(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + 2H_2O(l)$$

The white, powdery zinc hydroxide disappears at it reacts with the acid to form a colourless solution of zinc chloride.

Reactions of metal carbonates

Metal carbonates react with acids to form a salt, carbon dioxide and water. The reaction can be summarised by the following general word equation:

metal carbonate + acid ⟶ salt + carbon dioxide + water

When a metal carbonate reacts with an acid, effervescence is seen as the carbon dioxide is evolved.

Examples

1 Calcium carbonate reacts with nitric acid to produce soluble calcium nitrate, carbon dioxide and water:

$$CaCO_3(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

During the reaction, fairly vigorous effervescence occurs, and the white calcium carbonate quickly disappears as it reacts with the acid to form a colourless solution of calcium nitrate.

2 Iron(II) carbonate reacts with sulfuric acid to produce iron(II) sulfate, carbon dioxide and water:

$$FeCO_3(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + CO_2 + H_2O(l)$$

Fairly vigorous effervescence occurs. The green iron(II) carbonate quickly disappears and a green solution forms due to the production of soluble iron(II) sulfate which contains green Fe^{2+} ions.

Decomposition of metal compounds

Many metal compounds decompose when they are heated. Table 17.3.1 summarises the effects of heat on metal nitrates, metal carbonates and metal hydroxides.

▼ **Table 17.3.1** The effect of heat on metal nitrates, metal carbonates and metal hydroxides

Metal	**Metal compound**		
	Nitrates	**Carbonates**	**Hydroxides**
Potassium Sodium	Decompose to form the metal nitrite and oxygen: $2NaNO_3(s) \longrightarrow 2NaNO_2(s) + O_2(g)$	Do not decompose. The carbonates are stable.	Do not decompose. The hydroxides are stable.
Calcium Magnesium Aluminium Zinc Iron Lead Copper	Decompose to form the metal oxide, nitrogen dioxide and oxygen: $2Mg(NO_3)_2(s) \longrightarrow 2MgO(s) + 4NO_2(g) + O_2(g)$ $2Cu(NO_3)_2(s) \longrightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$	Decompose to form the metal oxide and carbon dioxide: $MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$ $CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$	Decompose to form the metal oxide and water (as steam): $Mg(OH)_2(s) \longrightarrow MgO(s) + H_2O(g)$ $Cu(OH)_2(s) \longrightarrow CuO(s) + H_2O(g)$
Sliver	Decomposes to form the metal, silver, nitrogen dioxide and oxygen: $2AgNO_3(s) \longrightarrow 2Ag(s) + 2NO_2(g) + O_2(g)$	Silver carbonate does not exist since it is too unstable.	Silver hydroxide does not exist since it is too unstable

When some metal compounds are heated and they decompose, *colour changes* are observed. For example, when the compounds of copper are heated they change colour from blue or blue-green to black because black copper(II) oxide is formed in each reaction. When compounds of lead are heated, they change colour from white to yellow because yellow lead(II) oxide is produced. Tests given in Table 22.3.1 on page 354 can be used to identify the oxygen, nitrogen dioxide, carbon dioxide and water vapour produced when metal compounds decompose.

Looking at the decomposition of metal compounds we can see that potassium and sodium form *stable* compounds which do not decompose on heating or, in the case of the nitrates, only decompose slightly. However, silver forms compounds which are too unstable to exist or, in the case of silver nitrate, decomposes very easily and completely when heated.

Practical activity

Investigating the decomposition of metal compounds

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with samples of sodium nitrate, calcium nitrate, sodium carbonate, copper(II) carbonate, magnesium hydroxide, lime water, a wooden splint, a piece of blue litmus paper, a piece of cobalt chloride paper and dry test tubes, one fitted with a cork which has a delivery tube running through it.

Method

1 Place a spatula of sodium nitrate into a test tube and heat strongly until you see vigorous effervescence. While still heating test for the presence of oxygen by inserting a glowing splint slowly into the tube.

2 Repeat step **1** using calcium nitrate in place of sodium nitrate. This time test for oxygen when a brown gas begins to be produced. Place a piece of moist blue litmus paper into the brown gas. Take care not to inhale the fumes. If you have a fume cupboard available, you should do this in the fume cupboard.

3 Place a spatula of copper(II) carbonate into a test tube. Place the cork into the mouth of the tube with the other end of the delivery tube leading into a test tube containing lime water (see Figure 8.1.2). Heat the test tube strongly and observe the lime water.

4 Repeat step **3** using sodium carbonate in place of copper(II) carbonate.

5 Place a spatula of magnesium hydroxide in a test tube and heat strongly. Place a piece of dry cobalt chloride paper across the mouth of the tube.

6 Describe what you saw when each compound was heated.

7 Which compound did not decompose? Explain why.

8 Explain the colour change you saw when you heated copper(II) carbonate.

9 Write balanced chemical equations for the decomposition reactions.

Summary questions

1 Write a balanced chemical equation for each of the following:
  - a the reaction between magnesium hydroxide and nitric acid
  - b the reaction between copper(II) carbonate and hydrochloric acid
  - c the reaction between zinc oxide and sulfuric acid.

2 Describe and explain the colour change which occurs when copper(II) oxide is put into sulfuric acid and heated gently.

3 Write balanced chemical equations to show the effect of heat on the following metal compounds:
  - a calcium carbonate
  - b potassium nitrate
  - c zinc hydroxide
  - d lead(II) nitrate.

Key concepts

- **Metals** are elements whose atoms have a small number of valence electrons, usually 1, 2 or 3, and they form positive cations in chemical reactions by losing these valence electrons.
- A metal lattice is composed of metal cations surrounded by a 'sea' of delocalised electrons.
- The electrostatic attraction between positive cations and delocalised negative electrons, known as the metallic bond, holds the metal lattice together.
- In general, metals have high melting points and boiling points, conduct electricity and heat, have a shiny lustre, are malleable and ductile and have high densities.
- The properties of metals can be explained by relating them to the structure of the metallic lattice.
- Metals form positive **cations** in chemical reactions by losing valence electrons. As a result they form **ionic compounds**.
- Metals behave as **reducing agents** in reactions.
- Metals react with oxygen to form metal oxides.
- If metal oxides can dissolve in water they form alkaline solutions.
- Potassium, sodium, calcium and magnesium react with water to form the metal hydroxide and hydrogen.
- Other metals, except copper, react with steam to form the metal oxide and hydrogen.
- Reactive metals react with hydrochloric acid and sulfuric acid to form a salt and hydrogen.
- Metal oxides react with acids to form a salt and water. The reaction is a **neutralisation** reaction.
- Metal hydroxides react with acids to form a salt and water. The reaction is also a neutralisation reaction.
- Metal carbonates react with acids to form a salt, carbon dioxide and water.
- The nitrates of potassium and sodium decompose when heated to form the metal nitrite and oxygen.
- The nitrates of other metals, except silver, decompose when heated to form the metal oxide, nitrogen dioxide and oxygen.
- Silver nitrate decomposes when heated to form silver, nitrogen dioxide and oxygen.
- The carbonates and hydroxides of potassium and sodium are not decomposed when heated.
- The carbonates of other metals, except silver, decompose when heated to form the metal oxide and carbon dioxide. Silver carbonate does not exist.
- The hydroxides of metals, except silver, decompose when heated to form the metal oxide and water as steam. Silver hydroxide does not exist.

Practice exam-style questions

Multiple-choice questions

1 All metals:

I are solid at room temperature
II lose electrons when they react
III form soluble basic oxides when they react with oxygen

A II only
B I and II only
C I and III only
D II and III only

2 The following statements are all true except:

A When calcium reacts with cold water it forms calcium hydroxide and hydrogen.
B Iron reacts more vigorously than zinc with hydrochloric acid.
C Magnesium burns readily in air.
D When silver nitrate decomposes it forms silver.

3 Metals can conduct an electric current because:

A the atoms in the metal vibrate
B the cations in the metal lattice are mobile
C there are small spaces between the atoms in the metal lattice which allow electricity to pass through
D some of the electrons in the metal lattice are mobile

4 Which of the following is a true statement?

A Metals react with oxygen in the air to form acidic oxides.
B All metal hydroxides decompose when heated.
C Magnesium carbonate decomposes into magnesium oxide and carbon dioxide when heated.
D Copper reacts readily with hydrochloric acid.

5 Zinc liberates hydrogen gas from:

I cold water
II steam
III sulfuric acid

A I and II only
B II and III only
C I, II and III
D III only

6 When sodium nitrate crystals are heated:

A they melt but do not decompose
B they form a brown, acidic gas
C they form a gas which relights a glowing splint
D they evolve a gas which causes a lighted splint to be extinguished with a squeaky pop

Structured question

7 a The table below gives some information about two metals, A and B, and their reactions.

Metal	Charge on the metal ion	Action of water on the metal	Action of dilute hydrochloric acid on the metal	Action of air on the metal
A	2+	Does not react with cold or hot water but does react with steam		Burns when heated strongly, especially if powdered
B	+		Reacts violently	

i) Complete the table by writing in the missing information. *(3 marks)*

ii) Using the symbols A and B, as shown in the table above, write balanced equations for the following:

– the reaction between A and steam
– the reaction between B and hydrochloric acid
– the reaction between A and oxygen. *(5 marks)*

b i) When most nitrates are heated, one or two gases may be given off. From the nitrates of the metals calcium, sodium, zinc and copper, select ONE nitrate that:

– gives off ONE gas
– gives off TWO gases. *(2 marks)*

ii) Write the equation for the reaction which gives off TWO gases above. *(2 marks)*

c An unknown metal, Z, was found to react in a very similar way to magnesium. How would the action of heat on the carbonate of Z differ from the action of heat on sodium carbonate?
Give the relevant equation(s) given that Z forms the Z^{2+} ion. *(3 marks)*

Total 15 marks

Extended response question

8 a Metals have certain physical properties in common. They are good conductors of electricity, they are malleable and ductile and they have high melting points. By reference to the bonding in metals, explain the reason for EACH of the properties listed. *(6 marks)*

b Zinc carbonate, copper(II) oxide and magnesium hydroxide are each treated with excess dilute nitric acid in three separate experiments. Discuss EACH reaction using the following sub-titles:

i) observations on adding dilute nitric acid
ii) explanation of the observations
iii) equation for the reaction. *(9 marks)*

Total 15 marks

C18 Reactivity, extraction and uses of metals

Objectives

By the end of this topic you will be able to:

- discuss the reactivity of metals
- describe displacement reactions between metals and metal compounds
- explain the basis of the reactivity series of metals
- deduce the order of reactivity of metals based on experimental results.

Metals vary considerably in their reactivity. Some are so reactive that they must be stored out of contact with oxygen or moisture in the air. Others are extremely unreactive. The reactive metals occur in the Earth's crust in ionic compounds from which they are extracted. These compounds are known as ores. The unreactive ones occur in the Earth's crust in their elemental state. Metals are used extensively by man to make many of the things we see around us and this makes their extraction extremely important.

C18.1 The reactivity of metals

It is possible to determine the **reactivity series** of metals by looking at their reactions. The reactivity series of metals is a list of metals arranged in order from the most reactive to the least reactive based on the following:

- the strength of their reactions with oxygen, water and dilute hydrochloric acid or dilute sulfuric acid
- how easily their compounds are decomposed when heated
- displacement reactions between the metals and metal compounds.

We will now look at each of these.

Strength of reaction with oxygen, water and dilute acids

You studied the reactions of a variety of metals with oxygen, water and dilute acids in Unit 17.2. Looking at these we can see certain trends.

- Potassium, sodium, calcium and magnesium appear to be the most reactive metals with oxygen, water and hydrochloric and sulfuric acid.
- Copper is the least reactive metal with oxygen and it does not react with steam or dilute acid. Silver does not react with oxygen, water or dilute acids.
- Starting from potassium, the strength of the reaction of the metals with oxygen, water and dilute acid *decreases* in the following order: potassium, sodium, calcium, magnesium, zinc, iron, copper and finally silver.

In Unit 4 you learnt that the reactivity of a metal is determined by how easily its atoms ionise, known as **ease of ionisation**. The more easily a metal ionises, the more reactive it is. Of the metals studied, potassium is the most reactive metal because it ionises the most easily and silver is the least reactive because it ionises the least easily.

Decomposition of metal compounds

We studied the **decomposition** of metal compounds in Unit 17.3. Looking at these reactions we can again see certain trends.

- The very reactive metals, potassium and sodium, form fairly stable compounds. Their carbonates and hydroxides are so stable that they are not decomposed when heated and their nitrates are decomposed only very slightly.
- The compounds of silver are extremely unstable such that silver carbonate and silver hydroxide are too unstable to exist and silver nitrate decomposes very easily when heated.
- Compounds of the other metals decompose when heated. Starting from the compounds of calcium, the ease of decomposition of the compounds *increases* in the following order: calcium, magnesium, zinc, aluminium, iron, lead and finally copper.

From this it can be seen that the more reactive the metal is, the more stable are its compounds. The less reactive the metal is, the less stable are its compounds and the more easily they decompose. This is because the reactive metals ionise easily and the ions they form are very *stable*. This makes their compounds very stable. The less reactive metals do not ionise easily and the ions they form are *unstable*. This means that their compounds are also unstable.

Displacement reactions

A **displacement reaction** occurs when a metal in its free state (chemically uncombined) takes the place of (displaces) another metal from a compound. A *more* reactive metal will always displace a *less* reactive metal from its compounds.

If a more reactive metal is placed in a solution containing ions of a less reactive metal, the more reactive metal will **ionise** and the ions of the less reactive metal will be **discharged** to form atoms. As a result, a precipitate of the less reactive metal will form on the surface of the more reactive metal. This often drops off into the solution when disturbed and the more reactive metal gradually decreases in size. Depending on the nature of the solution, it may also undergo a colour change as the reaction proceeds.

Displacement reactions are **exothermic reactions**. The greater the difference in the reactivity of the two metals, the more heat is produced in the reaction.

A displacement reaction can be summarised by the following general equation:

$$\underset{\substack{\vdots\\\text{more}\\\text{reactive}\\\text{than B}}}{\mathrm{A}} \quad + \quad \underset{\substack{\vdots\\\text{less}\\\text{reactive}\\\text{than A}}}{\mathrm{BX}} \quad \longrightarrow \quad \mathrm{AX} \quad + \quad \mathrm{B}$$

- A is more reactive than B. A ionises to form A^{n+} ions:

  $$\mathrm{A} \longrightarrow \mathrm{A}^{n+} + ne^{-}$$

- B is less reactive than A. The B^{n+} ions are discharged to form B atoms:

  $$\mathrm{B}^{n+} + ne^{-} \longrightarrow \mathrm{B}$$

Displacement reactions are also **redox reactions**. The more reactive metal acts as a reducing agent because it gives electrons to the ions of the less reactive metal. The more reactive the metal is, the more powerful it is as a reducing agent.

Example

The reaction between magnesium and copper(II) sulfate solution:

$$\mathrm{Mg(s)} + \mathrm{CuSO_4(aq)} \longrightarrow \mathrm{MgSO_4(aq)} + \mathrm{Cu(s)}$$

Magnesium is *more* reactive than copper, therefore, the magnesium ionises and the Mg^{2+} ions enter the solution.

$$\mathrm{Mg(s)} \longrightarrow \mathrm{Mg^{2+}(aq)} + 2e^-$$

The Cu^{2+} ions in the copper(II) sulfate solution are discharged forming copper atoms:

$$\mathrm{Cu^{2+}(aq)} + 2e^- \longrightarrow \mathrm{Cu(s)}$$

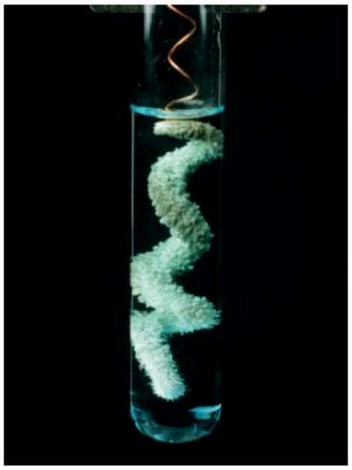
A pink precipitate of copper builds up on the magnesium and drops off into the solution when disturbed. The magnesium becomes smaller and the blue colour of the solution gradually fades as the blue Cu^{2+} ions are discharged.

The overall ionic equation for the reaction is:

$$\mathrm{Mg(s)} + \mathrm{Cu^{2+}(aq)} \longrightarrow \mathrm{Mg^{2+}(aq)} + \mathrm{Cu(s)}$$

Did you know?

It is possible to make a 'silver tree' by placing a coil of copper wire into a solution of silver nitrate. Copper is more reactive than silver, therefore, it displaces the silver from the silver nitrate and after a while silver is deposited on the copper wire forming a beautiful 'tree'.



▲ **Figure 18.1.1** When copper is placed into a solution containing silver ions, silver is deposited on the copper

Practical activity

Investigating displacement reactions

Your teacher may use this activity to assess:

- observation, recording and reporting
- analysis and interpretation.

You will be supplied with clean samples of the following metals of approximately the same size: magnesium, zinc and copper, magnesium sulfate solution, zinc sulfate solution, copper(II) sulfate solution and six test tubes.

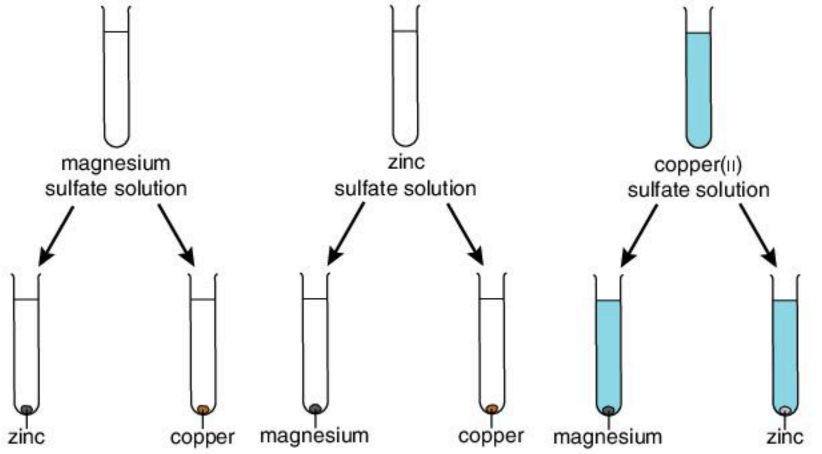
Method

1. Place about 5 cm^3 magnesium sulfate solution into each of two test tubes, 5 cm^3 of zinc sulfate solution into another two tubes and 5 cm^3 of copper(II) sulfate solution into the last two tubes. Measure and record the temperature of each solution.
2. Add a piece of magnesium to the zinc sulfate and the copper(II) sulfate solutions, a piece of zinc to the magnesium sulfate and the copper(II) sulfate solutions and a piece of copper to the magnesium sulfate and the zinc sulfate solutions.
3. Gently shake each tube periodically and observe what occurs in each. If a displacement reaction occurs, it should be possible to see the displaced metal deposited on the metal you added and possibly dropping off from the metal into the solution when shaken.
4. After 10 minutes, measure the temperature of the contents of each tube. If the temperature of any tube increases, calculate the increase.

5. Observe the colour of the copper(II) sulfate solution in each tube after about 30 minutes.
6. Record your observations and any temperature increases in a table similar to the one outlined below. The dash indicates that the reaction was not carried out. If no reaction occurred, place a cross in the appropriate place.

	Observations on adding		
	Magnesium	**Zinc**	**Copper**
Magnesium sulfate solution	–		
Zinc sulfate solution		–	
Copper(II) sulfate solution			–

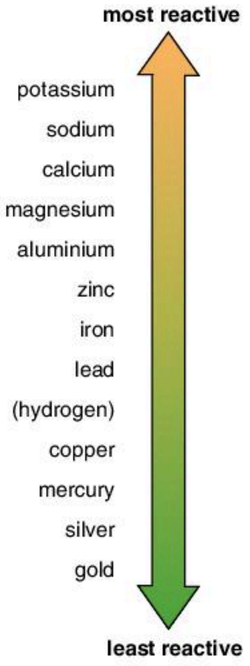
7. Using the information from the table, draw up a list of the three metals so that the most reactive metal occurs at the top and the least reactive metal appears at the bottom.
8. Why was the temperature increase greater when magnesium was placed in copper(II) sulfate solution than when it was placed in zinc sulfate solution?
9. Write balanced chemical equations for each reaction that occurred.
10. Explain the following:
    - a why a reaction occurred between magnesium and copper(II) sulfate solution
    - b why no reaction occurred between copper and zinc sulfate solution
    - c why the blue colour of the solution faded in the reaction between zinc and copper(II) sulfate solution.



▲ **Figure 18.1.2** The displacement reactions of various metals with metal salts

The reactivity series of metals

The **reactivity series** of some of the common metals, drawn up from the strength of the reactions of the metals with oxygen, water and dilute acids, the ease of decomposition of the metal compounds and displacement reactions, is given in Figure 18.1.3.



▲ **Figure 18.1.3** The reactivity series of metals

Exam tip

It is really important that you learn the **order of reactivity** of the metals in the reactivity series.

Potassium, at the top, is the *most* reactive. It ionises the most easily and forms the most stable ions which are very hard to discharge. Gold, at the bottom, is the *least* reactive. It ionises with great difficulty and forms very unstable ions which are very easy to discharge.

We can use the reactivity series of metals to make certain *predictions*.

- We can predict how any metal will *react* if its position in the reactivity series is known. For example, if a metal occurs between calcium and magnesium it will react less vigorously than calcium with oxygen, water and dilute acids but more vigorously than magnesium. The metal will also not displace calcium but will displace magnesium from their compounds.
- We can predict what *position* a metal will occupy in the reactivity series if the way it reacts is known. For example, if a metal is found to react more vigorously with dilute hydrochloric acid than zinc and is also found to displace zinc from zinc sulfate solution but not magnesium from magnesium sulfate solution, then the metal must occur above zinc but below magnesium.

If two metals are very similar in the way they react, then it is easy to determine which is higher in the series by carrying out displacement reactions between the two metals and their compounds. The *higher* metal will always displace the *lower* metal from its compounds.

Hydrogen is included in the reactivity series even though it is a non-metal. We can use the position of hydrogen to predict how metals will react with acids since acids contain the hydrogen ion, H^+:

- if a metal is above hydrogen in the reactivity series, then it will react with the acid because it displaces the hydrogen in the acid
- if a metal is below hydrogen in the reactivity series, it will not react with the acid because it will not displace the hydrogen in the acid
- the higher up the metal is in the series, i.e. the further it is away from hydrogen, the more vigorous the reaction will be with an acid.

Exam tip

It is important that you know how to deduce the **order of reactivity** of metals based on experimental results or data supplied.

Summary questions

1 Why is sodium more reactive than copper?

2 Why don't potassium carbonate and sodium carbonate decompose when heated?

3 Would a reaction occur if:

a a piece of clean magnesium ribbon was placed into copper(II) sulfate solution

b a clean piece of copper was placed into sodium sulfate solution?

If a reaction occurs, explain why and give a balanced equation for the reaction.

4 Explain the benefit of including hydrogen in the reactivity series of metals even though it is a non-metal.

5 Metal A will react with a solution containing ions of metal B but will not react with a solution containing ions of metal C. Write metals A, B and C in their correct order of reactivity from the most reactive to the least reactive.

C18.2 The extraction of metals

Objectives

By the end of this topic you will be able to:

- explain how the position of a metal in the reactivity series determines its extraction method
- describe the extraction of aluminium
- describe the extraction of iron.

The very unreactive metals occur in the Earth's crust in their free elemental state, e.g. silver and gold, and can, therefore, be mined directly from the Earth. Most metals, however, occur combined with other elements in impure ionic compounds and they then have to be extracted from these compounds. Compounds from which metals can be extracted are known as **ores**. The most important ores are metal oxides, sulfides and carbonates.

The extraction of a metal from its ore is a **reduction process** because the positive metal cations have to be discharged to form atoms by gaining electrons:

$$M^{n+} + ne^- \longrightarrow M$$

The actual method used depends on the position of the metal in the reactivity series.

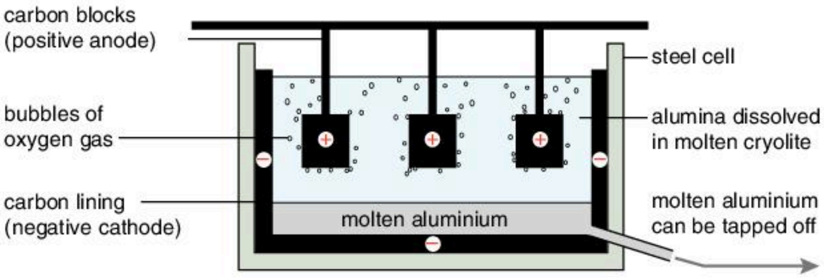
- Metals *high* in the reactivity series, i.e. potassium, sodium, calcium, magnesium and aluminium, form very stable ions which are difficult to reduce. They need a powerful method of reduction. The most powerful method is **electrolysis of the molten ore**. This method uses the most energy and is the most expensive.
- Metals *lower* down in the reactivity series, i.e. zinc and below, form less stable ions which are easier to reduce than those high in the series. They need a less powerful method of reduction which involves **heating the ore with a reducing agent** such as carbon, carbon monoxide or hydrogen. This method uses less energy than electrolysis and is less expensive.

Extraction of aluminium

Aluminium is the third most abundant in the Earth's crust after oxygen and silicon, and the most abundant metal. The main ore from which aluminium is extracted is known as **bauxite**. Bauxite is impure, hydrated aluminium oxide, $Al_2O_3.xH_2O$. There are large bauxite reserves in both Jamaica and Guyana. After the bauxite has been mined, it is purified to form pure, anhydrous aluminium oxide, also known as **alumina**, Al_2O_3. The alumina is then exported, mainly to Canada and the USA, where the aluminium is extracted by **electrolysis**.

The electrolysis of alumina takes place in a large steel tank lined with carbon (graphite), which forms an **electrolytic cell**. The carbon lining acts as the negative cathode in the electrolysis process. In the middle of the tank there are huge blocks of carbon, hung from above, which act as the positive anode.

An electrolytic cell used for the extraction of aluminium is shown in Figure 18.2.1.



▲ **Figure 18.2.1** Extracting aluminium by electrolysis

The extraction process involves the following steps:

1. The pure alumina (aluminium oxide) is dissolved in molten **cryolite** at about 950 °C. Cryolite is sodium aluminium fluoride, Na_3AlF_6. The alumina is dissolved in the cryolite because the melting point of alumina is extremely high, i.e. 2050 °C, and the liquid formed when it melts is a poor conductor. By dissolving it in molten cryolite, its melting point is reduced meaning less energy is required and the solution produced is a much better conductor of electricity. Once the alumina has dissolved in the molten cryolite, its ions are free to move.
2. The molten cryolite/alumina solution is electrolysed in the electrolytic cell:
    - The aluminium ions move towards the cathode where they are **reduced** to form aluminium atoms:

      $$Al^{3+}(l) + 3e^- \longrightarrow Al(l)$$

      The molten aluminium that forms collects at the bottom of the cell and is tapped off. It is then made into sheets or blocks.
    - The oxide ions move towards the anode where they are oxidised to form oxygen gas:

      $$2O^{2-}(l) \longrightarrow O_2(g) + 4e^-$$

      The oxygen gas that forms is released as bubbles. It forms carbon dioxide by reacting with the carbon anodes, and over time the anodes need to be replaced since they disintegrate as a result of the reaction.

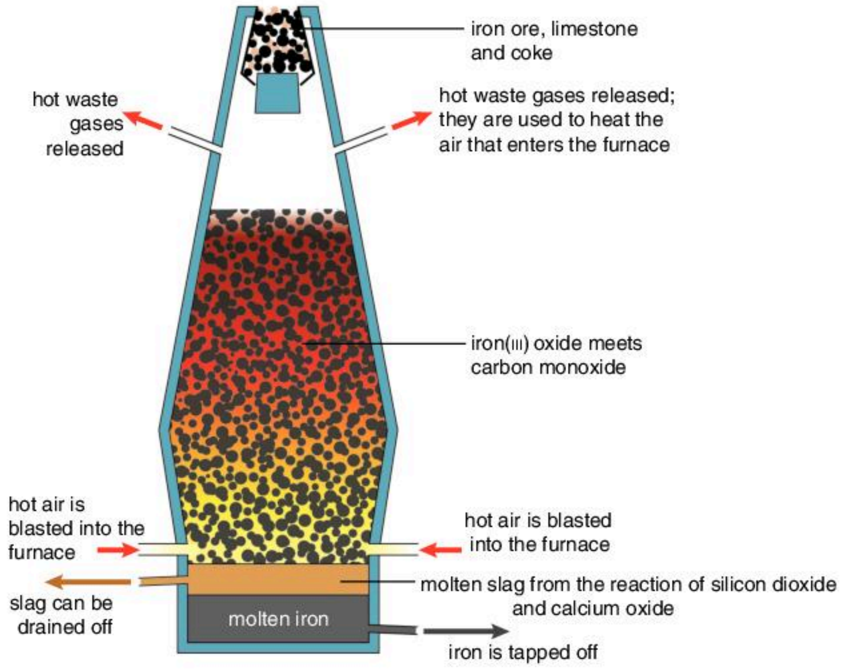
? Did you know?

The extraction of aluminium by electrolysis is an expensive process since it requires energy to melt the cryolite, it uses huge amounts of electricity and replacing the anodes is costly.

Extraction of iron

The main ores from which iron is extracted are **haematite** and **magnetite**. Haematite is impure iron(III) oxide, Fe_2O_3 and magnetite is impure iron(II, III) oxide Fe_3O_4. The extraction of iron from its ores involves reducing the ores using the reducing agent, **carbon monoxide**, in a **blast furnace**.

A blast furnace for the extraction of iron is shown in Figure 18.2.2. It is a chimney shaped tower that can be up to 30 m tall.



▲ **Figure 18.2.2** A blast furnace used for extracting iron

The extraction process

The following steps are involved in the extraction of iron from its ores.

1 A mixture of iron ore, limestone and coke is added through the top of the furnace. The limestone is calcium carbonate and coke is the name given to a substance that is made from coal which is mainly carbon. The mixture is known as 'the charge'.

2 Hot air is blown in through the bottom of the furnace.

3 At the *bottom* of the furnace, the coke burns in the hot air forming carbon dioxide:

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H \ -\text{ve}$$

The reaction is exothermic and keeps the temperature at the bottom of the furnace at about 1900 °C. The carbon dioxide rises further up the furnace.

4 In the *middle* of the furnace, the carbon dioxide reacts with more coke to form carbon monoxide:

$$CO_2(g) + C(s) \longrightarrow 2CO(g)$$

The carbon monoxide then rises further up the furnace.

5 In the *upper part* of the furnace, the carbon monoxide then **reduces** the iron ores to iron:

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

or

$$Fe_3O_4(s) + 4CO(g) \longrightarrow 3Fe(l) + 4CO_2(g)$$

The **molten iron** runs to the bottom of the furnace where it is tapped off. This iron contains a fairly high percentage of carbon (about 4%), as well as other impurities such as silicon and phosphorus, and is known as 'pig iron'. Most of the pig iron is then purified and converted into **steel**, an alloy of iron.

? Did you know?

The iron that runs out of the bottom of the blast furnace flows into a series of moulds that branch off from the main channel. These moulds resemble a litter of suckling piglets in the way that the piglets lie next to each other when they are feeding from a sow. This is why the iron that is tapped off from the blast furnace is known as '**pig iron**'.

▲ **Figure 18.2.3** 'Pig iron' gets its name from the moulds in which liquid iron collects. These moulds resemble a litter of suckling piglets.

The function of the limestone

Iron ore contains a lot of impurities, mainly silicon dioxide (sand), which would eventually build up and 'clog up' the blast furnace. Limestone is added to remove the silicon dioxide. The steps involved are as follows:

- The heat in the *upper part* of the blast furnace causes the calcium carbonate to decompose forming calcium oxide and carbon dioxide:

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

- The calcium oxide, being basic, then reacts with the acidic silicon dioxide to form calcium silicate, $CaSiO_3$. Calcium silicate is also known as **slag**:

$$CaO(s) + SiO_2(s) \longrightarrow \underset{\text{slag}}{CaSiO_3(l)}$$

  The slag melts and runs down to the bottom of the blast furnace where it forms a layer above the molten iron and is tapped off separately. It is then used in road building or can be finely ground and mixed with cement.

In the extraction of iron it is possible to use **carbon** or **hydrogen** in place of carbon monoxide as the reducing agent to reduce the iron ore, for example, the Ispat steel plant in Port Lisas, Trinidad, uses hydrogen. Hydrogen is ideal since it produces water vapour as a product and no other harmful gases. Pollution problems associated with the use of coke, such as the release of carbon dioxide and other waste gases into the atmosphere, are avoided with the use of hydrogen.

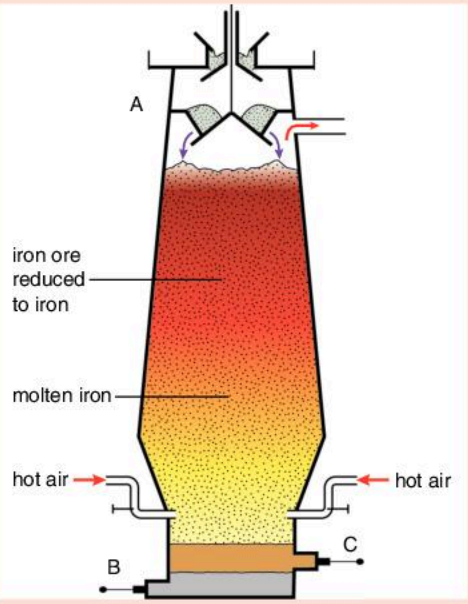
Summary questions

1 Why are different methods used to extract aluminium and iron from their ores?

2 **a** Name the ore from which aluminium is extracted.
   b Outline the process by which aluminium is extracted from the ore named in **a** above.

3 The diagram shows a blast furnace:
   a What are the three materials that are added to the top of the blast furnace at A in the extraction of iron?
   b What is removed at B and C?
   c Explain why air is blown into the blast furnace.
   d The carbon monoxide acts as a reducing agent when it reacts with the iron(III) oxide. What does this mean?
   e Why is it easy to remove the slag?

4 Write the balanced chemical equation to show how iron(III) oxide is converted to iron in the blast furnace.



Objectives

By the end of this topic you will be able to:

- describe the uses of aluminium, lead and iron
- relate the uses of aluminium, lead and iron to their properties
- explain what an alloy is
- explain why metal alloys are often used in place of the metals
- describe the uses of aluminium, iron and lead alloys
- relate the uses of aluminium, iron and lead alloys to their properties.

C18.3 Uses of metals and their alloys

Metals are used extensively by man. The use of a metal is directly linked to its particular properties, especially its physical properties. Metals are used because they are strong and last a relatively long time. They are also shiny, malleable and ductile, have high densities and can conduct electricity and heat. All of these properties make metals particularly useful. Metals can also be mixed with other metals to make **alloys**. This is done to improve or modify the properties of the metals.

Uses of metals

In this unit we will look at the uses of three metals, aluminium, lead and iron.

Aluminium

Aluminium is a particularly useful metal because it has a thin layer of aluminium oxide on its surface. This layer adheres to the metal and makes it very resistant to corrosion (Unit 19). Because of this, aluminium is used to produce louvred windows and window frames, which can resist corrosion when they are exposed to oxygen and moisture in the air. Another major use of aluminium is to make aluminium cans used as containers to store food and drink.

▲ **Figure 18.3.1** Aluminium is used to make louvred windows

▲ **Figure 18.3.2** Aluminium cans are used to store drinks

▲ **Figure 18.3.3** Aluminium can be used to make saucepans

Aluminium is a good conductor of electricity which enables it to be used in overhead power cables. These cables have a steel core down the centre to strengthen them. Aluminium is also a good conductor of heat which makes it useful to make saucepans. Aluminium foil is used for cooking since it is protected by the thin, unreactive layer of aluminium oxide and it has a highly reflective surface that allows it to reflect heat back into the food, thus keeping it warm for a longer period of time.

Aluminium has the disadvantage that it is not the strongest of metals. For this reason it is usually alloyed with other metals to make it stronger. We will be discussing some of the uses of aluminium alloys in the next section.

Did you know?

There was a time when **aluminium** was considered more precious than gold. Napoleon III was believed to have held a banquet where the honoured guests were given aluminium utensils to eat with while the other guests were only given gold utensils.

Lead

As one of the first metals to be discovered, **lead** has been used for several thousand years. The largest use of lead today is in lead–acid batteries because it is highly resistant to corrosion and conducts electricity. A typical 12-volt battery used in motor vehicles consists of six cells. Each cell has an anode made of lead alloyed with a small quantity of antimony, a cathode made of lead(IV) oxide (PbO_2) and an electrolyte of sulfuric acid. Chemical reactions occurring in each cell produce electricity. Since the reactions are reversible, the battery can be recharged.

▲ **Figure 18.3.4** A lead–acid battery

Another important use of lead is as a radiation shield around X-ray equipment and nuclear reactors because X-rays and nuclear radiation do not penetrate through it. Whenever you have an X-ray taken you are given a lead apron to wear which shields you from the harmful effects of the X-rays.

Iron

Today, **iron** is the most widely used of all metals because of its high abundance, low cost and high strength. However, because pure iron is quite soft it is usually hardened by alloying it with carbon to form **steel**. Pure iron, commercially known as wrought iron, has been used for centuries to make a variety of ornamental structures, such as furniture, lighting fixtures, gates and railings. Most items described as wrought iron today are actually made of mild steel because it is easier and cheaper to produce than wrought iron. The description wrought iron is still used because these items have the bent and shaped appearance of items formerly made of wrought iron.

▲ **Figure 18.3.5** Wrought iron gates

The uses of metals and the properties that make them particularly suitable for each use are summarised in Table 18.3.1.

▼ **Table 18.3.1** Properties and uses of metals

Metal	Use	Properties
Aluminium	• to make overhead electrical cables (with a steel core)	• good conductor of electricity, resistant to corrosion, light in weight, ductile
	• to make saucepans	• good conductor of heat, resistant to corrosion, non-toxic, light in weight, attractive appearance
	• to make aluminium cans to store food and drink	• resistant to corrosion, non-toxic, light in weight, malleable
	• to make window frames	• resistant to corrosion
	• to make foil for cooking	• unreactive (due to the aluminium oxide coating), therefore does not react with food, non-toxic, highly reflective
	• manufacture of aluminium alloys	• easily alloyed with other metals
Lead	• to make lead–acid batteries, e.g. car batteries	• good conductor of electricity, resistant to corrosion
	• as a radiation shield	• high density prevents radiation from penetrating through
	• to make weights, e.g. fishing weights, keel weights for sailboats	• high density, malleable
	• manufacture of lead alloys	• easily alloyed with other metals
Wrought iron (rarely used)	• to produce ornamental iron work	• malleable and ductile, therefore easily shaped, easily welded, strong therefore resistant to stress
	• manufacture of steel	• easily alloyed with carbon and other metals

Did you know?

The word **plumbing** was derived from the Latin *plumbum*, meaning lead. Plumbing was originated by ancient civilisations such as Roman and Greek as they developed public baths, piped drinking water and drainage systems. Some historians have speculated that the decline of the Roman Empire was due to lead from water pipes getting into the water supply and causing widespread lead poisoning.

Exam tip

It is important that you can relate the uses of metals and their alloys to their **physical properties**.

Alloys and their uses

There are many situations where metals are not used in their pure form. More than one metal may be combined to produce a mixture of metals known as an alloy. Sometimes a non-metal may also be added.

To make an alloy, the metals are heated, mixed thoroughly and cooled so that they solidify again. Alloys are solid metallic solutions. The purpose of making an alloy is to improve the properties of the metals or to modify their properties. Alloys are usually harder, stronger and more resistant to corrosion than the original metals.

The atoms of the metals that are mixed are usually of *different sizes*. Even if they are only slightly different in size, the regular packing of atoms is changed and it makes it more difficult for the atoms to slide over each other when a force is applied. This is why alloys are harder and stronger than the metals alone.

We will now look at the alloys of aluminium, lead and iron.

Aluminium alloys

Aluminium is often mixed with copper, magnesium, manganese, silicon and zinc to produce different alloys. Two of the most common alloys of aluminium are duralumin and magnalium. The composition, uses and properties of duralumin and magnalium are given in Table 18.3.2.

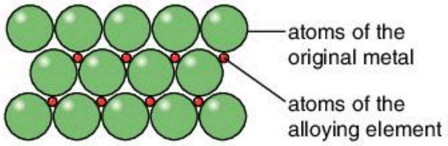
▼ **Table 18.3.2** Composition, uses and properties of aluminium alloys

Aluminium alloy	Composition	Uses	Properties
Duralumin	About 94% aluminium, 4% copper with small amounts of magnesium and manganese	In the construction of aircraft and to make precision tools	Stronger than aluminium so it can withstand greater stress, as light in weight as aluminium
Magnalium	About 95% aluminium and 5% magnesium	Construction of aircraft and automobile parts, to make metallic mirrors and scientific instruments	Stronger, harder and more resistant to corrosion than aluminium, lighter in weight than aluminium

▲ **Figure 18.3.7** Duralumin is strong and lightweight and is used to construct aircraft

Iron alloys

The building industry makes extensive use of iron, usually in the form of steel. Steel is an alloy of iron with carbon or various other metals. There are quite a few different types of steel. Steels formed by alloying iron with carbon are known as carbon steels. Generally the amount of carbon in carbon steels varies between 0.1% and 1.5%. As the carbon content increases, the steel becomes harder but less malleable and ductile and therefore becomes more brittle and harder to weld.



▲ **Figure 18.3.6** Structure of an alloy

Did you know?

Brass is a copper alloy consisting of about 70% copper and 30% zinc. It is used to manufacture electrical pins for plugs since it is stronger than copper and to make musical instruments due to being malleable and having very good acoustic properties.

Did you know?

Duralumin was developed and patented by a German metallurgist called Alfred Wilm in 1909. Originally it was only made by a German company called Dürener Metallwerke. The name duralumin is a contraction of 'Dürener' and 'Aluminium'. In 1917 during World War I, Hugo Junkers, a German aircraft designer, pioneered the use of duralumin in the construction of the J 7 fighter plane.

Did you know?

Iron can be alloyed with a variety of metals to form what are known as **specialist steels**. For example, titanium steel is an alloy of iron and titanium, manganese steel is an alloy of iron and manganese and cobalt steel is an alloy of iron and cobalt. Each steel has specialist uses; titanium steel can withstand very high temperatures so is used to make spacecrafts, cobalt steel is magnetised easily so is used to make magnets.

▲ **Figure 18.3.8** Steel is widely used in the construction industry

The composition, uses and properties of some of the common alloys of iron are given in Table 18.3.3.

▼ **Table 18.3.3** Composition, uses and properties of some iron alloys

Iron alloy	Composition	Uses	Properties
Mild steel	Contains less than 0.25% carbon	In the construction of buildings, bridges, ships and motor vehicles. To make wire, nails and 'tin cans' to store food and other items, i.e. 'tin cans' are made of steel coated in a thin layer of tin to prevent corrosion.	Harder and stronger than iron, malleable and ductile so easily shaped, easy to weld
High carbon steel	Contains between 0.25% and 1.5% carbon	To make cutting tools, drill bits, knives and masonry nails	Harder than mild steel, but more brittle
Stainless steel	Typical values are about 70% iron, 20% chromium and 10% nickel	To make cutlery, cooking utensils, kitchen sinks, hospital equipment, catering equipment	Hard, strong and much more resistant to corrosion (rusting) than carbon steels. Has a shiny, attractive appearance, but is more expensive to produce than carbon steels.
Cast iron	Contains about 4% carbon	To make small castings, e.g. manhole covers, railings, grates, cylinder blocks in engines and saucepans	Inexpensive and easy to cast into exact shapes. Hard, but more brittle than steel, therefore it tends to shatter rather than bend when hit.

Lead alloys

Lead is often mixed with tin, antimony or copper to make alloys. One common alloy of lead is **lead solder**. Lead solder is composed of about 60% lead and 40% tin. It is used to join metallic items together. To do this, the solder is melted so that it flows into the joint between the items. On cooling the solder solidifies, joining the items together. The melting points of the metals being joined must be higher than that of the lead solder.

By alloying lead with tin, the lead solder produced has a lower melting point than lead and is also harder and stronger than lead.

Summary questions

1. What properties of aluminium make it useful for:
   - a making overhead electrical cables
   - b making saucepans?
2. Other than making alloys, give two uses of lead and in each case explain why lead is suitable for the use given.
3. What is an alloy?
4. What are the advantages of using an alloy rather than a pure metal?
5. Discuss two named alloys of aluminium by referring to the composition and the uses of each.

Key concepts

- It is possible to determine the **reactivity series of metals** by looking at the strength of their reactions with oxygen, water and dilute acids, how easily their compounds are decomposed when heated, and displacement reactions between the metals and metal compounds.
- The more reactive the metal, the more stable its compounds.
- In a **displacement reaction**, a more reactive metal displaces a less reactive metal from its compounds.
- The reactivity series of the common metals from the most reactive to the least reactive is: potassium, sodium, calcium, magnesium, aluminium, zinc, iron, lead (hydrogen), copper, mercury, silver, gold.
- The position of hydrogen in the reactivity series can be used to compare how certain metals will react with dilute acids.
- The very unreactive metals occur in the Earth's crust in their free **elemental state**, e.g. silver and gold, and can, therefore, be mined directly from the Earth.
- Most metals occur in the Earth's crust combined with other elements in impure ionic compounds and they have to be extracted from these compounds.
- Compounds from which metals can be extracted are known as **ores**. The most important ores are metal oxides, sulfides and carbonates.
- The extraction of a metal from its ore is a **reduction** process.
- Metals high in the reactivity series, i.e. potassium, sodium, calcium, magnesium and aluminium, form very stable ions which are difficult to reduce. They need a powerful method of reduction, i.e. **electrolysis of the molten ore**.
- Metals lower down in the reactivity series, i.e. zinc and below, form less stable ions which are easier to reduce. They need a less powerful method of reduction, i.e. **heating the ore with a reducing agent**.
- Aluminium is extracted from its ore, **bauxite**, by electrolysis.
- Bauxite is impure, hydrated aluminium oxide, $Al_2O_3.xH_2O$.
- During electrolysis, aluminium forms at the cathode and oxygen forms at the anode.
- Iron is extracted from its ores, **haematite** and **magnetite** by reducing the ores using carbon monoxide as the reducing agent in a **blast furnace**.
- Haematite is impure iron(III) oxide, Fe_2O_3 and magnetite is impure iron(II, III) oxide Fe_3O_4.
- Molten iron and molten slag, calcium silicate ($CaSiO_3$), are produced in the blast furnace.
- The physical properties of metals make them very useful to man.
- The main uses of **aluminium** are to make overhead electrical cables, saucepans, window frames, aluminium storage cans and foil for cooking.
- The main use of **lead** is to make lead–acid batteries. It is also used to make radiation shields and weights.
- The main use of **iron** is to manufacture steel.
- An **alloy** is a mixture of metals although sometimes a non-metal can be added.
- Alloys are often more useful than the pure metal because they are harder, stronger and more resistant to corrosion than the pure metal
- **Duralumin** and **magnalium** are alloys of aluminium. Their main use is in the construction of aircraft.
- **Steel** is an alloy of iron with carbon or other elements. Different types of steel have different properties and uses.
- **Solder** is an alloy of lead and tin. It is used to join metal items.

Practice exam-style questions

Multiple-choice questions

1 Which of the following reactions is likely to take place most readily?

A $Mg(s) + 2AgNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + 2Ag(s)$
B $Mg(s) + Pb(NO_3)_2(aq) \longrightarrow Mg(NO_3)_2(aq) + Pb(s)$
C $Mg(s) + Cu(NO_3)_2(aq) \longrightarrow Mg(NO_3)_2(aq) + Cu(s)$
D $Mg(s) + Zn(NO_3)_2(aq) \longrightarrow Mg(NO_3)_2(aq) + Zn(s)$

2 The method used to extract a metal from its ore is determined by:

A the melting point of the metal
B the position of the metal in the reactivity series
C the solubility of the ore
D the electrical conductivity of the molten ore

3 Aluminium is extracted from alumina by:

A reduction of molten alumina using carbon
B reduction of alumina dissolved in cryolite using carbon monoxide
C electrolysis of molten alumina
D electrolysis of alumina dissolved in molten cryolite

4 Which of the following is a list of the raw materials used in the extraction of iron?

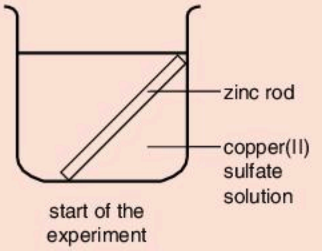
A haematite, limestone, air, carbon monoxide
B limestone, carbon dioxide, magnetite, coke
C haematite, air, limestone, coke
D coke, air, haematite, silicon dioxide

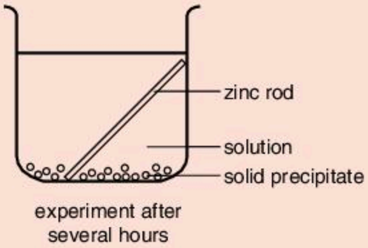
5 Which of the following is most widely used in the construction industry?

A stainless steel
B mild steel
C duralumin
D solder

Structured question

6 a A student places a zinc rod in a beaker containing copper(II) sulfate solution and leaves it for several hours, gently shaking the solution occasionally. The diagrams below show what the student saw at the start of the experiment and then again after several hours.





i) Explain why the zinc rod decreased in size during the experiment. Give a relevant ionic equation to support your answer. *(3 marks)*

ii) What change would you expect to observe in the solution? *(1 mark)*

iii) Explain your answer to **a ii)**. *(1 mark)*

iv) What would be the colour of the solid precipitate that was seen in the bottom of the beaker after several hours? *(1 mark)*

v) Name the type of reaction occurring in the beaker. *(1 mark)*

vi) How would the rate of the reaction be affected if the zinc rod was replaced with a piece of magnesium of the same dimensions? *(1 mark)*

vii) Explain your answer to **a vi)**. *(2 marks)*

b Another student carried out experiments 1 and 2 below in order to determine the relative reactivity of lead, iron and an unknown metal X, which has a valency of 2.

Experiment 1: A piece of metal X was placed into a solution of lead(II) nitrate. Metal X became coated with lead.

Experiment 2: A piece of metal X was placed into a solution of iron(II) sulfate. There was no change in the appearance of metal X.

i) Use the information given to place lead, iron and metal X in order of reactivity, starting with the most reactive. *(2 marks)*

ii) Write an ionic equation for the reaction between metal X and the lead(II) nitrate solution. *(2 marks)*

iii) How would you expect metal X to react with hydrochloric acid? *(1 mark)*

Total 15 marks

Extended response question

7 a Chemical and electrolytic reduction are the two main methods used to extract metals from their ores. Identify ONE factor that determines the method that should be employed in the extraction of metals from their ores. Explain how this factor determines the method to be used. *(3 marks)*

b Outline the main steps involved in the extraction of iron from its ore, iron(III) oxide. Your answer must include THREE relevant equations. Details of how the impurity, silicon dioxide, is removed are not required. *(7 marks)*

c Stainless steel, an alloy of iron, is used to make cooking utensils.

i) What is an alloy? *(1 mark)*

ii) Suggest TWO reasons why stainless steel is used to make cooking utensils instead of iron. *(2 marks)*

d Aluminium is another metal used to make alloys. Name ONE alloy of aluminium and give its main use. *(2 marks)*

Total 15 marks

C19 Metals in living systems and the environment

Some metals are extremely important in the lives of plants and animals, while others can cause severe poisoning if consumed and can have harmful effects on the environment. Conversely, the environment can have devastating effects on metals and can lead to their corrosion.

Objectives

By the end of this topic you will be able to:

- explain what happens when metals corrode
- give the conditions necessary for metals to corrode
- explain what happens when aluminium corrodes
- explain what happens when iron rusts.

C19.1 The corrosion of metals

Metals, especially those used outdoors, tend to undergo **corrosion**. Corrosion occurs when the surface of the metal gradually wears away due to its reaction with chemicals in the environment, mainly oxygen and moisture, and it is speeded up by the presence of certain pollutants. In some cases it may lead to the total destruction of the metal object. When metals corrode, they are **oxidised** to their oxides. They sometimes also form salts, e.g. carbonates are formed by the reaction with carbon dioxide.

The rate at which most metals corrode is linked to the metal's reactivity. In general, the *more* reactive the metal, the *faster* it will corrode. Metals that are used close to the sea will tend to corrode at an even faster rate than normal because of the presence of sodium chloride in the atmosphere.

Aluminium and iron corrode with quite different results; the corrosion of aluminium is mostly beneficial while the corrosion of iron is detrimental. We will now look at the corrosion of both these metals in more detail.

The corrosion of aluminium

On exposure to air, a fresh piece of **aluminium** immediately forms a layer of **aluminium oxide**, Al_2O_3. This layer adheres to the metal below, does not peel off and is relatively unreactive. As a result, it protects the aluminium from further corrosion. It is because of this aluminium oxide layer that aluminium appears less reactive than is predicted by its position in the reactivity series.

If a piece of aluminium is scratched and the oxide layer is removed, it is possible to see the shiny metal underneath. It does not take very long for that shiny section to become dull as the aluminium reacts with the oxygen in the air again and renews its oxide film.

Aluminium can be made even more resistant to corrosion by **anodising**. This is the process by which the thickness of the oxide layer on the aluminium is artificially *thickened* by electrolysis (Unit 10.5). The oxide layer gives added advantages to the aluminium; it makes the surface harder and it readily absorbs dyes, so anodised objects can be coloured attractively. Because of these added advantages, anodised aluminium has a great many uses, from making cookware such as pots and pans to protecting satellites from space debris.

▲ **Figure 19.1.1** Rust formation on a padlock

The corrosion of iron

On exposure to oxygen and moisture in the air **iron** and **steel** objects immediately begin to corrode to form mainly **hydrated iron(III) oxide**, $Fe_2O_3.xH_2O$, which has an orange-brown colour. Hydrated iron(III) oxide is commonly known as **rust**. Corrosion of iron and steel is known as **rusting**.

In order for iron to rust, both oxygen and water or moisture must be present. It does not occur if only one is present. The rusting of iron can be both damaging and costly. Unlike the adherent coating of aluminium oxide that forms on the surface of aluminium, rust does not adhere to the iron below. Rust flakes off and exposes fresh iron to oxygen and moisture. The newly exposed iron then rusts and the process continues and gradually wears away the iron.

Practical activity

To investigate what conditions are needed for rusting to occur

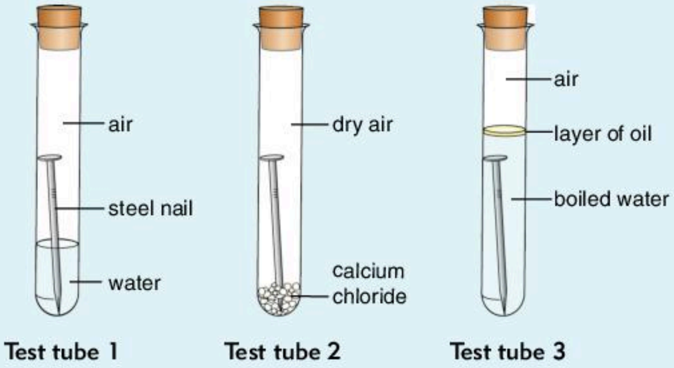
Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with three steel nails, some calcium chloride, some oil and three test tubes with corks.

Method

1. Place one nail into the first test tube, add enough water to half cover the nail and cork the tube.
2. Place the second nail into a dry test tube. Add a small amount of calcium chloride and cork the tube. The calcium chloride absorbs any moisture from the air in the tube.
3. Boil some water to remove the dissolved oxygen and allow it to cool quickly. Place the third nail into another test tube and add enough of this boiled water to completely cover the nail. Gently pour some oil down the side of the tube so that it forms a layer on top of the water. Cork the tube.
4. Leave the nails in the three test tubes for a few days and carefully observe any changes in the nails.
5. Record your observations in a suitable table.
6. What was the purpose of adding the oil layer to the third tube?
7. Which nail rusted the most? What conditions were present in this case?
8. Did any nails show no signs of rusting? If so, what condition was absent in each case?
9. What conclusion can you draw about the conditions necessary for rusting?
10. Explain how you could test whether a steel nail rusts more quickly in sea water or tap water.



▲ **Figure 19.1.2** Investigating the conditions needed for rusting

Summary questions

1. What happens when a metal corrodes?
2. You are given the four metals copper, iron, zinc and aluminium.
   - a Which metal tends to appear less reactive than its position in the reactivity series would suggest? Why is this?
   - b Which metal rusts?
3. What conditions must be present for rusting to occur?
4. Explain why aluminium and iron corrode differently.
5. Suggest reasons for the following.
   - a Metal railings made of iron tend to rust more at the bottom of the railing than at the top.
   - b Pieces of iron on shipwrecks tend to rust very slowly when found deep below the surface of the water.

C19.2 The impact of metals on living systems and the environment

Objectives

By the end of this topic you will be able to:

- explain the importance of metals and their compounds in living organisms
- discuss the harmful effects of metals and their compounds on living organisms and the environment.

Many metals and their compounds are extremely important to living organisms. However, certain metals and their compounds can have harmful effects on the same living organisms and also on the environment.

The importance of metals and their compounds

Metal ions play a vital role in the functioning of living organisms. Some of these ions are required by the human body in amounts in excess of 100 mg per day, e.g. calcium, potassium, sodium and magnesium. These are sometimes referred to as **macro-minerals**. Others are recommended in smaller amounts for good health, these include manganese, iron, zinc, cobalt, copper, molybdenum, selenium and chromium. These are referred to as **micro-minerals** or **trace minerals**. Plants also require certain metal ions for healthy growth. Some of the ions required by plants and animals are used to make **organometallic compounds**, for example, chlorophyll and haemoglobin.

Table 19.2.1 summarises the importance of certain metal ions in living organisms.

▼ **Table 19.2.1** The importance of metals and their compounds in living organisms

Metal ion	Importance in living organisms
Magnesium	Magnesium ions are necessary for the formation of **chlorophyll** in green plants. Chlorophyll is the green pigment found in the chloroplasts of plant cells. It gives green plant parts their colour. Chlorophyll is essential in plants to trap the sunlight energy so that the plant can manufacture its own food by **photosynthesis**. The equation for photosynthesis is given below: $6CO_2(g) + 6H_2O(l) \xrightarrow[\text{absorbed by chlorophyll}]{\text{sunlight energy}} \underset{\text{glucose}}{C_6H_{12}O_6(aq)} + 6O_2(g)$ A shortage of magnesium causes the leaves of plants to become yellow, a condition known as **chlorosis**. Magnesium ions are also required for more than 300 biochemical reactions in the human body since they are essential for the functioning of many enzymes. The ions play a vital role in the production of energy in cells, in the formation of strong bones and teeth, in the transmission of nerve impulses and in the contraction of muscles.

▶ **Table 19.2.1** (continued)

Metal ion	Importance in living organisms
Iron	Iron ions are necessary for the formation of **haemoglobin** in animals. Haemoglobin is the red pigment found in the red blood cells. It is essential to carry oxygen around the body. The oxygen is then used by the body cells in **respiration** to provide the body with energy. The equation for respiration is given below: $\underset{\text{glucose}}{C_6H_{12}O_6(aq)} + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l) + \text{energy}$ A shortage of iron leads to a reduction in the number of red blood cells in the body, tiredness and a lack of energy, a condition known as **anaemia**.
Calcium	Calcium ions are necessary for the production of calcium phosphate in the body. Calcium phosphate is needed to build strong bones and teeth. Calcium ions are also necessary for the blood to clot at a cut. A shortage of calcium leads to bones and teeth becoming soft and weak. It may also lead to **rickets**, a condition in which the legs become bowed.
Zinc	Zinc ions play an important part in the functioning of the immune system, in the healing of wounds, and in cell and tissue growth and repair.
Sodium and potassium	Sodium and potassium ions play an important role in the transmission of nerve impulses and the contraction of muscles.

! Key fact

Pollution is the contamination of the natural environment by the release of unpleasant and harmful substances.

The harmful effects of metals and their compounds

Some of the same metals required by living organisms to maintain their health become **toxic** to the same living organisms when their levels become too high. Other metals, not required by living organisms, can also be toxic to them. Many of these are in a group known as **heavy metals**, which mainly includes the transition metals and some of the metalloids. All of these metals occur naturally in ionic compounds, however, their concentrations within the environment are increasing due to **pollution** caused by man's activities.

Heavy metal ions become particularly toxic when they form organometallic compounds by combining with organic compounds, e.g. tetraethyl lead and methyl mercury. Table 19.2.2 summarises some metal ions that can have harmful effects on the human body.

? Did you know?

The phrase 'mad as a hatter' was coined during the 1800s when **mercury** was used in the production of felt, which was then used by felters to make hats. The felters were exposed daily to trace levels of mercury. This gradually accumulated in their bodies causing them to develop symptoms of dementia (madness).

? Did you know?

Minamata disease was first discovered in the city of Minamata, Japan, in 1956. It was caused by the release of methyl mercury in industrial wastewater from a chemical factory into Minamata Bay. This methyl mercury gradually built up in shellfish and fish in the bay, which, when eaten by the local population, resulted in severe **mercury poisoning** and over 1700 deaths.

▼ **Table 19.2.2** Harmful effects of metals

Metal ion	Sources in the environment	Toxic effects
Lead	Disused car batteries, car exhaust fumes when using leaded petrol, lead based paints.	Lead damages a variety of body tissues and organs, in particular the kidneys, bones and the nervous system, especially the brain. It also interferes with the normal formation of red blood cells leading to anaemia. It is especially harmful to young children causing reduced IQ and potentially permanent learning disorders.
Mercury	Disused thermometers from hospitals and laboratories, disused fluorescent lights, industry, e.g. combustion of coal, extraction and refining of metals such as gold, manufacture of sodium hydroxide.	Mercury damages the central nervous system impairing muscular coordination, hearing, sight, speech and the sense of touch known as **Minamata disease**.
Cadmium	Disused nickel–cadmium batteries, cigarette smoke, industry, e.g. combustion of fossil fuels, extraction and refining of metals, manufacture of paint and certain plastics, e.g. PVC.	Cadmium causes damage to the respiratory system, kidneys and liver. It can also cause osteoporosis where the bones become weakened and fragile.
Arsenic	Mining of certain metals, especially gold, extraction and purification of metals, burning of fossil fuels, especially coal.	Arsenic can cause cancer and skin problems. It also damages the nervous system, heart and blood vessels.

Heavy metal ions are particularly harmful in the environment because they are **persistent**, i.e. they remain in the environment and do not decay like organic pollutants. They also have the potential to build up in the tissues of living organisms, especially fatty tissue, and they concentrate up food chains, harming the top consumers such as humans.

Consumption of fish, especially large fish, is the major source of ingested **mercury** in humans due to mercury concentrating up the food chain. Small fish eat plankton. If this plankton contains small amounts of mercury, this mercury gradually builds up in the bodies of the fish throughout their lifetime. Larger fish then eat the smaller fish, constantly building up the mercury in their bodies throughout their lifespan. This continues up the food chain to the very large, long-lived fish such as sharks, marlin, tuna and swordfish.

Disposal of **solid waste** containing heavy metals poses a serious problem since these metals should not be allowed to enter the environment. This waste includes:

- lead–acid batteries from motor vehicles
- thermometers containing mercury from hospitals and laboratories
- compact fluorescent light bulbs and other fluorescent bulbs containing mercury vapour
- nickel–cadmium batteries.

These items should not be disposed of in landfill sites because of the possibility of contaminating groundwater and nearby soil. They should also not be incinerated because of the possibility of releasing harmful gases containing the metal ions into the air. The problem of their disposal can be solved to a large extent by **recycling** all items containing heavy metals.

Summary questions

1 What is haemoglobin? Why is it important in our bodies?

2 What is chlorophyll? Why is it important to plants?

3 Why are zinc and calcium important to humans?

4 a Name three different metals whose ions have negative effects on the human body.

b Describe some of the negative effects of each metal named in a.

5 Explain why it may be dangerous to consume large fish such as swordfish or marlin.

? Did you know?

In general women need more iron than men since they lose iron from their bodies each month in haemoglobin during **menstruation**. Also women must ensure that they get adequate supplies of zinc during pregnancy since it is essential for growth of the cells and tissues of the foetus.

▲ **Figure 19.2.1** Compact fluorescent light bulbs contain mercury vapour

Key concepts

- Metals, especially those used outdoors, tend to undergo **corrosion**.
- Corrosion occurs when the surface of the metal gradually wears away due to its reaction with chemicals in the environment, mainly oxygen and moisture, and it is speeded up by the presence of certain pollutants.
- When metals corrode, they are oxidised to their oxides and sometimes they form salts.
- Reactive metals tend to corrode much faster than those that are lower in the reactivity series.
- When aluminium corrodes it forms a layer of aluminium oxide which is relatively unreactive and adheres to the metal below, protecting it from further corrosion.
- The layer of aluminium oxide can be thickened by **anodising**.
- When iron corrodes it forms hydrated iron(III) oxide, also known as **rust**.
- In order for iron to rust, both oxygen and water or moisture must be present.
- Rust flakes off over time exposing fresh iron to oxygen and moisture. The newly exposed iron then rusts and the process continues and gradually wears away the iron.
- Metal ions play a vital role in the functioning of living organisms. Some of these are combined with organic compounds to form **organometallic compounds**.
- **Macro-minerals** are needed by the human body in amounts in excess of 100 mg per day. Calcium, potassium, sodium and magnesium are the macro-minerals needed.
- **Micro-minerals** are required in small amounts by the human body for good health. Two of the most important micro-minerals are iron and zinc.
- One important metal ion required by plants is magnesium.
- Metals can have a negative effect on living organisms. Many of these are in a group known as **heavy metals**, which mainly includes the transition metals and some of the metalloids.
- The concentration of these harmful heavy metals within the environment is increasing due to pollution caused by man's activities.
- The most harmful metals include lead, mercury, cadmium and arsenic.
- Heavy metal ions are harmful in the environment because they do not decay and they concentrate up food chains, harming the top consumers such as humans.
- Consumption of fish, especially large fish, is the major source of ingested mercury in humans due to mercury concentrating up the food chain.
- Disposal of solid waste containing heavy metals poses a serious problem since these metals should not be allowed to enter the environment. This waste includes lead–acid batteries, mercury thermometers, fluorescent light bulbs and nickel–cadmium batteries.
- Waste containing heavy metals should be recycled.

Practice exam-style questions

Multiple-choice questions

1 Which of the following must be present for iron to rust?

I oxygen
II carbon dioxide
III moisture

A I only
B I and III only
C II and III only
D I, II and III

2 Which of the following is not a benefit to be gained from the corrosion of aluminium?

A It causes aluminium to be less reactive than its position in the reactivity series predicts.
B It forms an oxide layer on the aluminium, which can be coloured attractively.
C The oxide layer formed flakes off easily.
D It forms a relatively unreactive oxide layer on the aluminium.

3 Which of the following metals is a constituent of chlorophyll?

A magnesium
B zinc
C calcium
D iron

4 Which of the following is a trace mineral needed by the human body for good health?

A calcium
B potassium
C zinc
D magnesium

5 Which of the following is a list of heavy metal ions that are harmful to the human body?

A arsenic, iron, mercury, lead
B arsenic, cadmium, lead, mercury
C cadmium, calcium, mercury, lead
D arsenic, cadmium, magnesium, mercury

6 Consumption of large fish is most likely to cause:

A rickets
B anaemia
C cancer
D Minamata disease

Extended response question

7 a Most metals undergo corrosion. In the case of aluminium this can be beneficial to the metal while in the case of iron it can be harmful and costly.

i) What happens when a metal corrodes? *(1 mark)*

ii) Explain why the corrosion of aluminium is beneficial while the corrosion of iron is not. Your answer should include reference to the products of corrosion and their properties. *(6 marks)*

b 'Metal ions are extremely important to living organisms.' By reference to magnesium, iron and calcium, provide evidence to support this statement. *(6 marks)*

c Certain metals are harmful to living organisms. Explain why these metals are so harmful when they are released into the environment by human activities. *(2 marks)*

Total 15 marks

C20 Non-metals

Objectives

By the end of this topic you will be able to:

- give the general physical properties of non-metals
- give the specific physical properties of hydrogen, oxygen, nitrogen, chlorine, carbon and sulfur
- state where hydrogen, oxygen, nitrogen, chlorine, carbon and sulfur occur naturally.

Most elements in the periodic table are metals, the rest make up a group of elements known as the non-metals. The non-metals are found on the right-hand side of the table. Many non-metals are found naturally in their elemental state, mostly as gases or solids. They are also found combined with other non-metals in covalent compounds and combined with metals in ionic compounds. Non-metals are used extensively by man. However, because the properties of the different non-metals vary significantly, their uses also vary quite considerably.

C20.1 Physical properties of non-metals

Non-metals are elements whose atoms usually have a large number of valence electrons, usually 5, 6, 7 or 8. Most non-metals are found in Groups V, VI, VII and 0 of the periodic table.

The bonding between non-metal atoms in a non-metal varies. In Unit 5.5 you learnt about the bonding in the two allotropes of carbon, i.e. diamond and graphite. Several of the other non-metals are composed of **diatomic molecules**. Each diatomic molecule is made up of two atoms of the non-metal bonded covalently. Weak intermolecular forces then exist between the diatomic molecules. Non-metals composed of diatomic molecules include hydrogen (H_2), oxygen (O_2), nitrogen (N_2) and all the halogens, i.e. those in Group VII of the periodic table including fluorine (F_2), chlorine (Cl_2), bromine (Br_2) and iodine (I_2). The non-metals in Group 0, i.e. the noble gases, exist as individual atoms due to having full valence electron shells. Bonding between the atoms of other non-metals varies.

Because the way in which the non-metal atoms are bonded differs, the physical properties of the different non-metals also vary to some extent. However, non-metals have the following general physical properties:

- they have low melting points and boiling points
- they are poor conductors of heat and electricity
- they are dull in appearance when in the solid state
- they are weak and brittle when in the solid state
- they have low densities.

Many of the non-metals are *gases* at room temperature because of the weak forces of attraction between the diatomic molecules or atoms, e.g. hydrogen, oxygen, nitrogen, fluorine, chlorine and all the noble gases. Bromine is found as a *liquid* at room temperature and the other non-metals are *solids*.

The specific properties and characteristics of some of the non-metals are given in Table 20.1.1.

Exam tip

It is important that you can give the differences between the **physical properties** of metals and non-metals.

▼ **Table 20.1.1** Properties and characteristics of some non-metals

Non-metal	Properties and characteristics	Occurrence
Hydrogen (H_2)	A colourless, tasteless, odourless gas The lightest of all elements Acts as a reducing agent	Occurs in the compound water (H_2O), in hydrocarbons and in most other organic compounds
Oxygen (O_2)	A colourless, tasteless, odourless gas Acts as an oxidising agent	Occurs in the elemental state as O_2 molecules in the air. It makes up about 21% of the air. Occurs in water, in oxides of metals and non-metals, e.g. metal ores, silicon dioxide (sand) and carbon dioxide, in metal carbonates and metal sulfates, in many organic compounds and in very many other compounds
Nitrogen (N_2)	A colourless, tasteless, odourless gas One of the least reactive non-metals; it is considered relatively inert	Occurs in the elemental state as N_2 molecules in the air. It makes up about 78% of the air. Occurs in proteins, metal nitrates and ammonium compounds
Chlorine (Cl_2)	A yellow green gas Has a strong odour Poisonous to humans Acts as an oxidising agent	Occurs combined in metal chlorides, e.g. sodium chloride (table salt)
Carbon	Has two main allotropes, diamond and graphite (Unit 5.5) **Diamond** is an extremely hard, transparent, colourless, sparkling solid. It has a very high melting point. **Graphite** is a soft, flaky, opaque, dark grey solid. It has a very high melting point and conducts electricity. Acts as a reducing agent	Occurs in the elemental state as diamond and graphite Occurs in all organic compounds, in metal carbonates and metal hydrogencarbonates and in carbon dioxide
Sulfur	A light yellow solid	Occurs in the elemental state, e.g. near hot springs and in volcanic regions Occurs in metal sulfides and metal sulfates

Practical activity

Comparing the properties of two metals and two non-metals

Your teacher may use this activity to assess:

- observation, recording and reporting.

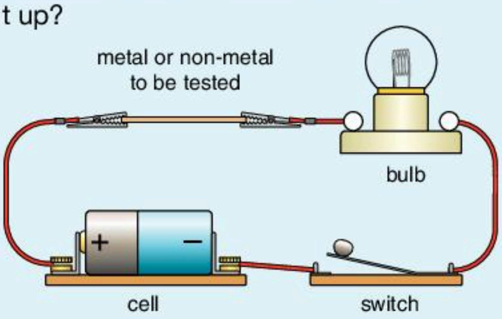
You will be supplied with samples of two metals, e.g. aluminium foil and copper, samples of two non-metals, e.g. graphite and roll sulfur, a cell, a bulb, a switch, conducting wire and two crocodile clips.

Method

1 Set up a circuit using the cell, bulb, switch, connecting wire and crocodile clips as shown in Figure 20.1.1. Attach each metal and non-metal in turn between the crocodile clips and close the switch. In each case, does the bulb light up?

2 Classify each metal and non-metal according to the following criteria:
  - Is it shiny or dull?
  - Is it malleable or brittle?
  - Is it a conductor or a non-conductor?

3 Record your findings in a table.



▲ **Figure 20.1.1** A circuit to test electrical conductivity

Summary questions

1 List five properties of hydrogen.

2 Name seven diatomic elements.

3 Where would you find oxygen in nature?

4 Compare diamond and graphite, giving the similarities and the differences between them.

5 By means of a table, compare the properties of metals with those of non-metals.

C20.2 Chemical properties and reactions of non-metals

Objectives

By the end of this topic you will be able to:

- describe the reactions of non-metals with oxygen and metals
- state the products formed by the reactions of non-metals with oxygen and metals
- explain the oxidising and reducing properties of non-metals.

Because non-metals are elements whose atoms usually have a large number of valence electrons, usually 5, 6, 7 or 8, they form negative anions in chemical reactions with metals by gaining electrons into their valence electron shell.

$$\mathrm{N} + n\mathrm{e}^- \longrightarrow \mathrm{N}^{n-}$$

As a result they form **ionic compounds** when they react with metals. The non-metal behaves as an oxidising agent since it gains electrons from the metal, i.e. it causes the metal to lose electrons (OIL).

Non-metals can also react with each other. In these reactions they share valence electrons forming **covalent compounds** and may behave as oxidising and reducing agents.

We will now look at the reactions of certain non-metals with oxygen and metals, and at the oxidising and reducing properties of non-metals.

Reactions of non-metals with oxygen

Non-metals react with oxygen to form compounds known as **non-metal oxides**. These non-metal oxides are usually gases at room temperature. This is in contrast with metal oxides which are solids at room temperature.

Most non-metal oxides are *acidic*, e.g. carbon dioxide (CO_2), sulfur dioxide (SO_2), sulfur trioxide (SO_3) and nitrogen dioxide (NO_2). These oxides are also known as **acid anhydrides** (Unit 8.1) since they react with water to form acidic solutions. A few non-metal oxides are *neutral*, e.g. water (H_2O), carbon monoxide (CO) and nitrogen monoxide (NO). They don't react with acids or alkalis.

Table 20.2.1 summarises the reactions of certain non-metals with oxygen.

▼ **Table 20.2.1** Reactions of some non-metals with oxygen

Non-metal	Reaction with oxygen
Hydrogen	Hydrogen burns with a very pale blue, almost colourless, flame forming water as steam: $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$ A mixture of hydrogen and air or oxygen explodes when lit.
Nitrogen	Nitrogen reacts with oxygen if the temperature is high enough, forming nitrogen monoxide (nitrogen(II) oxide): $N_2(g) + O_2(g) \longrightarrow 2NO(g)$ This reaction occurs during lightning storms and in the engines of motor vehicles.
Carbon	Carbon burns forming either carbon monoxide or carbon dioxide. If the oxygen supply is limited, carbon monoxide is produced: $2C(s) + O_2(g) \longrightarrow 2CO(g)$ If the oxygen supply is plentiful, carbon dioxide is produced: $C(s) + O_2(g) \longrightarrow CO_2(g)$
Sulfur	Sulfur burns with a blue flame forming sulfur dioxide (sulfur(IV) oxide): $S(s) + O_2(g) \longrightarrow SO_2(g)$

? Did you know?

The destruction of the Hindenburg airship killing 35 passengers is an infamous example of the combustion of **hydrogen**. The Hindenburg was a luxurious airship measuring 804 feet long and 135 feet tall. Inside were two decks for passengers and crew and sixteen gas cells filled with hydrogen to make it lighter than air. As it was coming to land in New Jersey on 6 May 1937, a flame appeared at the rear of the ship and within 34 seconds the entire ship was consumed in flames.

Practical activity

Investigating reactions of non-metals with oxygen

(Teacher demonstration)

Your teacher may use this activity to assess:

- observation, recording and reporting.

Your teacher will perform the following experiment.

Method

1. Place a small piece of carbon on to a deflagrating spoon.
2. Carefully heat the spoon in a Bunsen burner flame until the carbon begins to burn. Once burning, place it into a gas jar full of oxygen and observe what happens.
3. Repeat steps **1** and **2** using sulfur in place of carbon.
4. Compare the violence of the reaction of carbon with oxygen and sulfur with oxygen. Which seems more reactive?
5. What products were formed in each reaction?

Reactions of non-metals with metals

Non-metals react with metals to form **ionic compounds**. A reaction between a metal and a non-metal is a redox reaction because the metal loses electrons, i.e. it is oxidised, and the non-metal gains these electrons, i.e. it is reduced. The non-metal acts as an **oxidising agent** in all of these reactions.

- **Hydrogen** reacts with metals to form **metal hydrides**,

  e.g. $Ca(s) + H_2(g) \longrightarrow \underset{\text{calcium hydride}}{CaH_2(s)}$

- **Oxygen** reacts with metals to form **metal oxides**,

  e.g. $2Mg(s) + O_2(g) \longrightarrow \underset{\text{magnesium oxide}}{2MgO(s)}$

- **Nitrogen** reacts with metals to form **metal nitrides**,

  e.g. $3Mg(s) + N_2(g) \longrightarrow \underset{\text{magnesium nitride}}{Mg_3N_2(s)}$

- **Chlorine** reacts with metals to form **metal chlorides**,

  e.g. $2Al(s) + 3Cl_2(s) \longrightarrow \underset{\text{aluminium chloride}}{2AlCl_3(s)}$

- **Sulfur** reacts with metals to form **metal sulfides**,

  e.g. $Fe(s) + S(s) \longrightarrow \underset{\text{iron(II) sulfide}}{FeS(s)}$

Non-metals as oxidising and reducing agents

Non-metals can behave as both oxidising agents and reducing agents. They generally behave as oxidising agents, however, some non-metals can also behave as reducing agents, for example, hydrogen, carbon and sulfur.

Non-metals as oxidising agents

- *All* non-metals behave as **oxidising agents** when they react with *metals* to form ionic compounds. They remove electrons from the metal atoms,

  e.g. $2Zn(s) + O_2(g) \longrightarrow 2ZnO(s)$

  $2Fe(s) + 3Cl_2(s) \longrightarrow 2FeCl_3(s)$

  $Mg(s) + S(s) \longrightarrow MgS(s)$

- **Chlorine** and **oxygen** are particularly powerful **oxidising agents**. Whenever they react with another element or compound they act as oxidising agents,

  e.g. $S(s)+O_2(g) \longrightarrow SO_2(g)$

  $2KBr(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + Br_2(aq)$

  Oxygen caused the oxidation number of sulfur to increase from 0 to +4 in the first reaction. Chlorine removed electrons from the Br^- ions causing their oxidation number to increase from −1 to 0 in the second reaction.

Non-metals as reducing agents

- **Hydrogen**, **carbon** and **sulfur** act as **reducing agents** when they react with *oxygen*, a very powerful oxidising agent,

  e.g. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

  $C(s) + O_2(g) \longrightarrow CO_2(g)$

  $S(s) + O_2(g) \longrightarrow SO_2(g)$

  In each these reactions, the non-metal caused the oxidation number of each oxygen atom in the oxygen to decrease from 0 to −2.

- **Hydrogen** and **carbon** also act as **reducing agents** when they react with *metal oxides*. They reduce the metal ions to atoms,

  e.g. $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$

  $2Fe_2O_3(s) + 3C(s) \longrightarrow 4Fe(s) + 3CO_2(g)$

  Hydrogen caused the oxidation number of the Cu^{2+} ion in the copper(II) oxide to decrease from +2 to 0 in the first reaction. Carbon caused the oxidation number of each Fe^{3+} ion in the iron(III) oxide to decrease from +3 to 0 in the second reaction.

Summary questions

1. What is produced when a non-metal reacts with oxygen?
2. When carbon reacts with oxygen it can form one of two different oxides.
   - **a** What determines which oxide is produced?
   - **b** Write balanced chemical equations to show the formation of the two different oxides.
3. Write balanced chemical equations for the following reactions:
   - **a** the reaction between aluminium and chlorine
   - **b** the reaction between sodium and hydrogen.
4. When a non-metal reacts with a metal, the non-metal always acts as an oxidising agent. Explain the reason for this.
5. Certain non-metals can act as both oxidising agents and reducing agents. Name one non-metal that can behave as both and support your answer by reference to two different reactions of the named non-metal. Include balanced equations for the reactions.

C20.3 Laboratory preparation and uses of gases

Objectives

By the end of this topic you will be able to:

- describe how oxygen, carbon dioxide and ammonia can be prepared in the laboratory
- relate the methods of drying and collecting the gases to their properties
- explain the uses of oxygen and carbon dioxide based on their properties.

Oxygen, carbon dioxide and ammonia are all gases which, if required, can be made with relative ease in the laboratory. These same gases also have some very important uses in today's world.

When considering the method to be used to prepare a gas in the laboratory, the following properties of the gas need to be considered:

- The *solubility* of the gas in water to determine if the gas can be collected by bubbling it through water.
- The reactivity of the gas with different *drying agents* to determine which drying agents can be used.
- The *density* of the gas relative to the density of air to determine the method used to collect the dry gas.

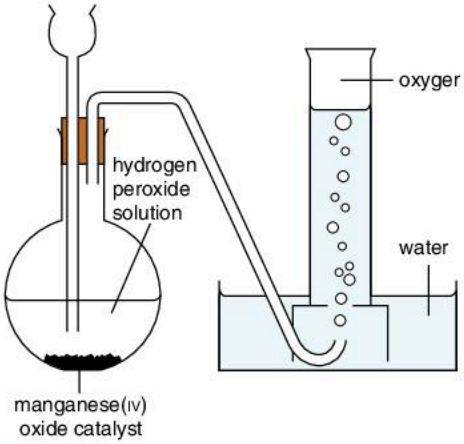
Laboratory preparation of oxygen

Oxygen is prepared in the laboratory by the decomposition of hydrogen peroxide solution using a catalyst of manganese(IV) oxide. The hydrogen peroxide decomposes according to the following equation:

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

The oxygen that is produced can then be collected in one of two ways which depends on whether dry oxygen is required or not.

- If dry oxygen is *not* required, then the gas can be collected by bubbling it through water into an upside down gas jar, as shown in Figure 20.3.1, since oxygen is not very soluble in water. As the oxygen fills the gas jar, the water level in the jar drops. This is known as **downward displacement of water** and it is easy to see when the jar is full. The oxygen collected in this way contains water vapour due to it having been bubbled through water.
- If dry oxygen *is* required, then it can be dried by passing it through a U-tube containing anhydrous calcium chloride or calcium oxide, or by bubbling it through a wash bottle containing concentrated sulfuric acid, as shown in Figure 20.3.2. All these compounds absorb the water vapour in the oxygen. The dry gas is then collected directly into a gas jar by **upward displacement of air** since it is slightly denser than air and sinks in the jar, displacing the air upwards. The problem with this method is that it cannot be determined when the jar is full.



▲ **Figure 20.3.1** The laboratory preparation of oxygen

Laboratory preparation of carbon dioxide

Carbon dioxide can be prepared by reacting a carbonate with an acid. The usual method is to react calcium carbonate with hydrochloric acid as shown in the following equation:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

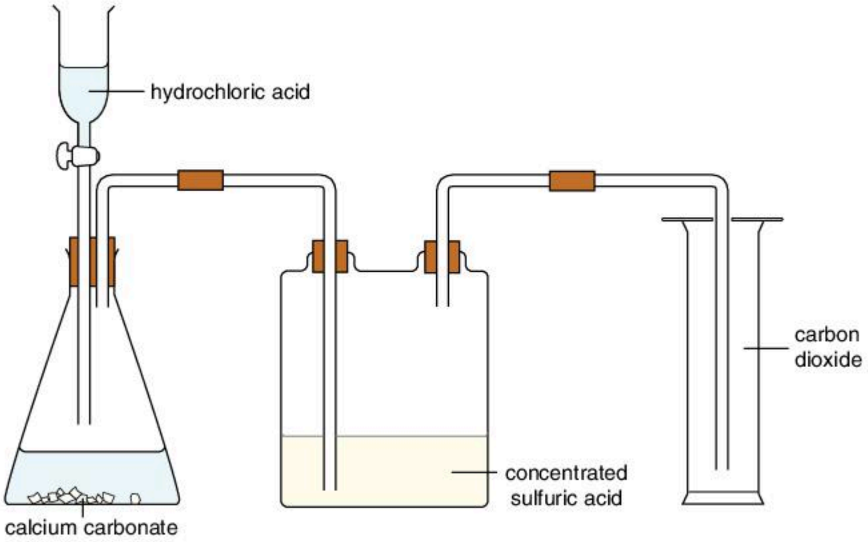
As with the preparation of oxygen, the method of collection depends on whether dry carbon dioxide is required or not.

- If dry carbon dioxide is *not* required, then the gas can be collected by bubbling it through water into an upside down gas jar since it is not very soluble in water. However, the carbon dioxide collected will contain water vapour.

Exam tip

It is very important that you can draw line drawings of the **apparatus** used to prepare oxygen, carbon dioxide and ammonia in the laboratory.

- If dry carbon dioxide *is* required, then it can be dried by passing it through a U-tube containing anhydrous calcium chloride or by bubbling it through a wash bottle containing concentrated sulfuric acid. Calcium oxide cannot be used to dry the gas because carbon dioxide reacts with it. The dry gas is then collected directly into a gas jar by **upward displacement of air** since it is denser than air.



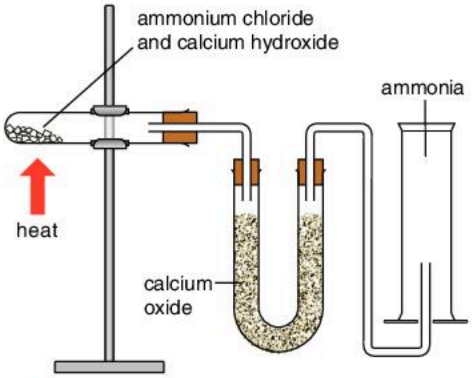
▲ **Figure 20.3.2** The laboratory preparation of dry carbon dioxide

It is important to note that calcium carbonate and sulfuric acid cannot be used to produce carbon dioxide since the reaction produces insoluble calcium sulfate. This forms a layer around the calcium carbonate crystals, which prevents them from continuing to react with the acid and the reaction quickly stops.

Laboratory preparation of ammonia

Ammonia can be prepared in the laboratory by reacting an alkali with an ammonium salt, for example, calcium hydroxide and ammonium chloride as shown in the following equation:

$$Ca(OH)_2(s) + 2NH_4Cl(s) \longrightarrow CaCl_2(s) + 2NH_3(g) + 2H_2O(g)$$



▲ **Figure 20.3.3** The laboratory preparation of dry ammonia

The solid calcium hydroxide and ammonium chloride are mixed together in a boiling tube. They are then heated using a Bunsen burner with the boiling tube placed horizontally, as shown in Figure 20.3.3. Since ammonia is very soluble in water, it cannot be collected over water. The gas is dried by passing it through a U-tube or up a drying tower containing calcium oxide. Neither calcium chloride nor concentrated sulfuric acid can be used to dry the gas since ammonia reacts with both of them.

The ammonia is collected in a gas jar by **downward displacement of air** since it less dense than air and rises to the top of the jar, displacing the air downwards. It is possible to determine when the gas jar is full of ammonia by placing a piece of moist red litmus paper at the mouth of the jar. If the litmus turns blue, the jar is full.

Table 20.3.1 summarises the methods used to prepare oxygen, carbon dioxide and ammonia in the laboratory.

Table 20.3.1 Summary of the methods used to prepare gases in the laboratory

Gas	Usual reagents	Solubility of gas in water	Drying agent(s)	Method of collecting the dry gas
Oxygen	Hydrogen peroxide and a catalyst of manganese(VI) oxide	Slightly soluble. Can be collected over water.	Anhydrous calcium chloride Calcium oxide Concentrated sulfuric acid	Upward displacement of air. Oxygen is slightly denser than air.
Carbon dioxide	Calcium carbonate and hydrochloric acid	Slightly soluble. Can be collected over water.	Anhydrous calcium chloride Concentrated sulfuric acid	Upward displacement of air. Carbon dioxide is slightly denser than air.
Ammonia	Calcium hydroxide and ammonium chloride	Very soluble. Ammonia reacts with water to form ammonium hydroxide, therefore, cannot be collected over water.	Calcium oxide	Downward displacement of air. Ammonia is less dense than air.

Uses of gases

The gases which we can make in the laboratory also have many important uses. We will look at some of the uses of oxygen and carbon dioxide based on their properties.

Uses of oxygen

Oxygen is essential for all living organisms to make energy during the process of **aerobic respiration**. Humans obtain this oxygen from the air when we breathe. Oxygen is also essential for **combustion** to occur. Because of these two properties, humans have found a variety of uses for oxygen, some of which are given below.

- Oxygen is used in hospitals for patients who have difficulty breathing and also to ease certain medical disorders, e.g. emphysema, asthma, chronic bronchitis and heart disease.
- Oxygen is carried on aeroplanes and submarines for breathing purposes in emergencies.
- Nearly pure oxygen is used in spacesuits for astronauts to breathe.
- Oxygen is used in oxyacetylene and oxyhydrogen torches to burn the acetylene or hydrogen and produce extremely high-temperature flames that are used to weld and cut metals.
- Liquid oxygen is used to burn the fuel that generates the lift in spaceships.

Uses of carbon dioxide

Carbon dioxide has certain properties that make it useful to humans.

- Pressurised liquid carbon dioxide is used in fire extinguishers. When the pressure is reduced by opening the valve of the extinguisher, the carbon dioxide is released. The carbon dioxide is non-flammable and denser than air, so it smothers the flames, keeping out the oxygen in the air.
- Solid carbon dioxide ('dry ice') is used as a refrigerant in the food industry. It sublimes at −78.5 °C, therefore, it can be used to keep food items extremely cold and leaves no liquid when it sublimes to a gas.
- Carbon dioxide is used to make carbonated soft drinks. The solubility of the gas increases as the pressure increases. The carbon dioxide is dissolved in the drink under pressure. When the pressure is released by opening the bottle or can, its solubility decreases and bubbles form causing the drink to 'fizz'. The carbon dioxide also adds a pleasant taste to the drink.
- Carbon dioxide is used as an aerosol propellant in cans containing certain foodstuffs, such as whipped cream, because it is relatively inert.

Figure 20.3.4 A fire extinguisher

Summary questions

1 Describe how you could prepare dry oxygen in the laboratory.

2 Ammonia can be prepared in the laboratory by heating a mixture of calcium hydroxide and ammonium chloride.
  a Write a balanced equation for the reaction.
  b Name a suitable drying agent for the gas.
  c What method would be used to collect the dry gas and why is this method used?
  d Why is ammonia not collected over water?

3 a What method is used to collect dry carbon dioxide in the laboratory?
  b Why is this method used?

4 Explain why oxygen is used:
  a in hospitals b in welding.

5 Explain why carbon dioxide is used in fire extinguishers.

Objectives

By the end of this topic you will be able to:

- list the uses of various non-metals and their compounds.

C20.4 The uses of non-metals

Although only a small number of elements in the periodic table are non-metals, these elements are used in a wide range of applications. Some of them play a particularly important role in industry. The uses of some non-metals and their compounds are given in Tables 20.4.1–20.4.6.

▼ **Table 20.4.1** The uses of carbon and its compounds

Non-metal/compound	Uses
Carbon (diamond)	• in jewellery • to cut and engrave glass and other diamonds • in the tips of drills • diamond powder is used as an abrasive for polishing and fine grinding
Carbon (graphite)	• as the 'lead' in pencils • to make electrodes for use in industry • as a solid lubricant • to make carbon fibres to strengthen plastics
Sodium carbonate (Na_2CO_3)	• to soften hard water • in the manufacture of glass
Sodium hydrogencarbonate ($NaHCO_3$)	• as a raising agent in baking powder • as an antacid • in fire extinguishers

▲ **Figure 20.4.1** This tennis racquet has been reinforced with carbon fibre

▲ **Figure 20.4.2** Carbon in the form of diamond is used to make jewellery

? Did you know?

Vulcanisation is the process of hardening rubber. The rubber is heated with sulfur to make it stronger, more elastic and more resistant to chemical attack. This is achieved by the sulfur forming cross-links between the polymer chains of the rubber. Charles Goodyear is generally credited with its discovery in 1839, however, the history of stabilising rubber dates back to 1600 BC.

▼ **Table 20.4.2** The uses of sulfur and its compounds

Non-metal/compound	Uses
Sulfur	• to vulcanise (harden) rubber, e.g. for car tyres • in the manufacture of medicinal drugs and ointments to treat fungal infections • in the manufacture of fungicides to prevent fungal attacks on crops and garden plants • to make the heads of matches • to manufacture sulfur dioxide and sulfuric acid
Sulfur dioxide (SO_2)	• as a food preservative, e.g. in jams and fruit juices • as a bleaching agent in the manufacture of paper
Sulfuric acid (H_2SO_4)	• in the manufacture of fertilisers, e.g. ammonium sulfate • in lead–acid batteries, e.g. car batteries

▼ **Table 20.4.3** The uses of phosphorus and its compounds

Non-metal/compound	Uses
Phosphorus	• to make the striking surface of safety matches • to make flares, fireworks and explosives • to make pesticides • in the manufacture of phosphoric acid
Phosphoric acid (H_3PO_4)	• to manufacture phosphate fertilisers, e.g. ammonium phosphate [$(NH_4)_3PO_4$] • as a rust remover
Phosphorus sulfides, e.g. P_4S_3	• to make the heads of matches

▲ **Figure 20.4.3** The heads of these matches contain a phosphorus sulfide

▼ **Table 20.4.4** The uses of chlorine and its compounds

Non-metal/compound	Uses
Chlorine	• to treat drinking water to destroy bacteria • to make solvents and dry-cleaning fluids, e.g. tetrachloroethene (C_2Cl_4) • to make pesticides, e.g. DDT • to make antiseptics, e.g. TCP and Dettol • to make monochloroethene, which is used to make PVC for pipes, window frames and guttering • to manufacture sodium chlorate(I) and calcium chlorate(I)
Sodium chlorate(I) (NaClO) and calcium chlorate(I) [$Ca(ClO)_2$]	• as bleaching agents • to sterilise drinking water and swimming pool water – the compound releases 'free chlorine' which is toxic to bacteria • as disinfectants

▼ **Table 20.4.5** The uses of nitrogen and its compounds

Non-metal/compound	Uses
Nitrogen	• liquid nitrogen is used to quick freeze foods • to fill light bulbs together with argon • in food packaging to provide an inert atmosphere which prevents aerobic decay • in the manufacture of ammonia
Ammonia (NH_3)	• in household cleaners, e.g. oven and window cleaners • to manufacture fertilisers, e.g. ammonium nitrate (NH_4NO_3), ammonium sulfate [$(NH_4)_2SO_4$] and ammonium phosphate [$(NH_4)_3PO_4$] • in the manufacture of nitric acid
Nitric acid (HNO_3)	• to manufacture fertilisers, e.g. ammonium nitrate (NH_4NO_3) • to manufacture explosives and dyes

? Did you know?

Fertilisers are added to the soil to supply plants with one or more of the elements that they require for healthy growth. Three of the major elements needed by plants are nitrogen (N), phosphorus (P) and potassium (K). Fertilisers which supply mainly these elements are known as NPK fertilisers.

▲ **Figure 20.4.4** Chlorine is used to manufacture chlorine bleaches

? Did you know?

Chlorine gas was used as a poison gas during World War I in the trenches in Europe.

▲ **Figure 20.4.5** Ammonia is often used in cleaning agents

▼ **Table 20.4.6** The uses of silicon and its compounds

Non-metal/compound	Uses
Silicon	• in the manufacture of silicon chips used in electronic devices from pocket calculators to computers • to make silicone implants for plastic and reconstructive surgery • to make a silicon steel alloy in the steel industry
Silicon dioxide (quartz) (SiO_2)	• in jewellery, e.g. amethyst is purple quartz • in the manufacture of glass – silicon dioxide is the main ingredient of glass • in the manufacture of cement • mixed with cement and aggregate to make concrete
Metal silicates	• to make ceramic products, such as pottery and porcelain products, roof tiles, floor tiles, building bricks and sewer pipes

Summary questions

1. Which four compounds of non-metals are used to make fertilisers?
2. What is mean by 'vulcanisation of rubber'? Name the non-metal that is used to vulcanise rubber and explain what happens during the process of vulcanisation.
3. Discuss some of the uses of the following non-metals and their compounds:
   a carbon b chlorine c silicon.

Objectives

By the end of this topic you will be able to:

- discuss the harmful effects of non-metals and their compounds on living systems and the environment.

C20.5 Harmful effects of non-metals and their compounds

While non-metals and their compounds are extremely useful to humans, some compounds of the same non-metals can have harmful effects on living systems and the environment. Like certain metals, the concentrations of these non-metal compounds within the environment are increasing due to pollution caused by man's activities. These harmful substances are called **pollutants**.

We will now look at some of the sources of some of these pollutants and their negative effects on the environment in Table 20.5.1. Some of these effects will also be discussed in more detail after the table.

▶ **Figure 20.5.1** Pollution harms the environment

▼ **Table 20.5.1** Harmful effects of some compounds of non-metals

Non-metal compound/ pollutant	Sources in the environment	Harmful effects
Sulfur dioxide (SO_2)	Combustion of fossil fuels, especially coal and heavy oils, e.g. in industry and power stations	Sulfur dioxide causes respiratory problems in humans, e.g. bronchitis. It reduces the growth of plants and dissolves in rainwater forming an acidic solution known as **acid rain** (see later in this unit).
Carbon monoxide (CO)	Incomplete combustion of fossil fuels, mainly in motor vehicles	Carbon monoxide binds with haemoglobin much more readily than oxygen causing a reduction in the amount of oxygen reaching body tissues. This reduces respiration and mental awareness, causes visual impairment, dizziness, headaches, unconsciousness and even death.

▼ **Table 20.5.1** (continued)

Non-metal compound/ pollutant	Sources in the environment	Harmful effects
Carbon dioxide (CO_2)	Complete combustion of fossil fuels, e.g. in motor vehicles, industry, power stations and aeroplanes	Carbon dioxide, together with other gases, builds up in the upper atmosphere contributing to the **greenhouse effect** and **global warming** (see later in this unit). About 25% of the carbon dioxide released is absorbed by the oceans leading to **ocean acidification**. This increase in acidity is expected to affect the ability of shellfish such as oysters, sea urchins and corals to produce and maintain their shells.
Hydrogen sulfide (H_2S)	Decay of organic matter in garbage dumps, landfills and farmyards. Released from petroleum refineries.	Hydrogen sulfide is extremely toxic. Like carbon monoxide, it combines readily with haemoglobin, which reduces cellular respiration and causes dizziness, headaches, unconsciousness and even death. Low concentrations irritate the eyes and respiratory system.
Oxides of nitrogen (NO and NO_2)	Combustion at high temperatures in motor vehicle engines and power stations causing nitrogen and oxygen in the air to react	Oxides of nitrogen are extremely toxic causing lung damage. Low concentrations irritate the skin, eyes and respiratory system. They also cause leaves to die, reduce the growth of plants, play a major role in the development of photochemical smog and dissolve in rainwater forming **acid rain** (see later in this unit).
Chlorofluorocarbons (CFCs)	Used as a refrigerant in air-conditioners and refrigerators and as a propellant in some aerosol sprays	CFCs break down the ozone layer in the upper atmosphere allowing more ultraviolet light to reach the Earth's surface (see later in this unit).
Carbon particles (C)	Combustion of fossil fuels, e.g. in industry. Bush fires and cigarettes.	Carbon particles blacken trees reducing photosynthesis and blacken buildings. They combine with water vapour and sulfur dioxide to form **smog**, which causes poor visibility for drivers and aircraft pilots and respiratory problems, e.g. bronchitis, asthma and lung disease.
Nitrate ions (NO_3^-) and phosphate ions (PO_4^{3-})	Fertilisers used in agriculture, especially in intensive farming. Synthetic detergents.	These ions cause **eutrophication** (see later in this unit).
Pesticides, e.g. insecticides, fungicides and herbicides	Used in agriculture for pest control, disease control and control of weeds. Used to control vectors of disease, e.g. mosquitoes.	Pesticides concentrate up food chains harming top consumers (see later in this unit). They can also harm useful as well as harmful organisms, for example, bees which are essential for pollination are being killed by insecticides.

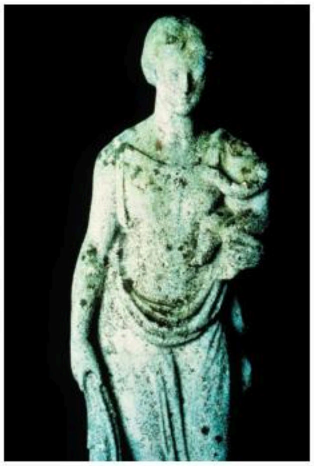
Acid rain

Acid rain is formed when pollutant gases such as sulfur dioxide and oxides of nitrogen are released into the atmosphere. These oxides are acidic and when they react with rainwater they form acidic solutions, for example:

$$SO_2(g) + H_2O(l) \rightleftharpoons \underbrace{H_2SO_3(aq)}_{\text{sulfurous acid}}$$

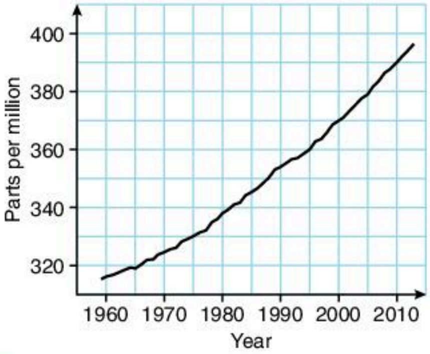
$$2NO_2(g) + H_2O(l) \longrightarrow \underbrace{HNO_2(aq)}_{\text{nitrous acid}} + \underbrace{HNO_3(aq)}_{\text{nitric acid}}$$

These acidic solutions then fall to Earth as acid rain which can have a serious effect on plants, animals, the soil and water. Acid rain decreases the pH of the soil, damages crops, trees and other plants and harms animals. It also corrodes buildings, cars, machinery and historical monuments and it causes lakes, streams and rivers to become acidic and unsuitable for aquatic organisms, such as fish.



▲ **Figure 20.5.2** A limestone statue corroded by acid rain

▲ **Figure 20.5.3** The effects of acid rain on trees



▲ **Figure 20.5.5** Average carbon dioxide levels in the atmosphere from 1960 to 2012

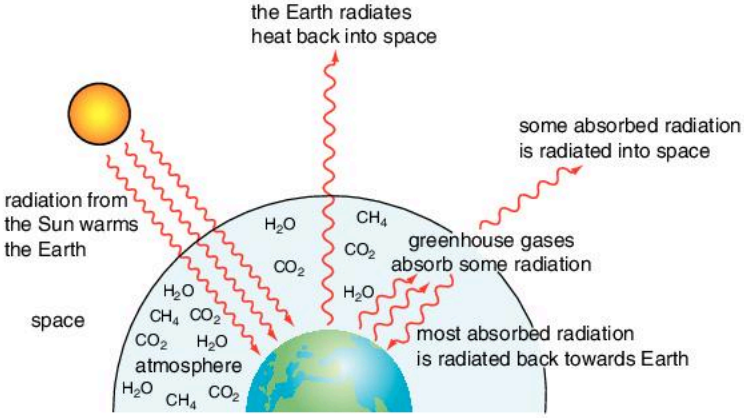
Did you know?

In 1800 the average carbon dioxide concentration in the atmosphere was about 280 parts per million (ppm) or 0.028%. In 1960 the figure was about 316 ppm and in 2012 it had risen to about 394 ppm. The concentration is currently increasing at about 2 ppm per year.

The greenhouse effect and global warming

When fossil fuels are burnt a number of gases are produced and released into the atmosphere, including carbon dioxide. Carbon dioxide is known as a **greenhouse gas**. Other greenhouse gases include water vapour, dinitrogen monoxide (N_2O) and methane (CH_4).

Greenhouse gases form a layer around the Earth. Radiation from the Sun passes through this layer, reaches the Earth's surface and warms the Earth. The Earth's surface then radiates this heat back into the atmosphere as infrared radiation. Some of this radiation goes back into space and some is absorbed by the greenhouse gases around the Earth. The greenhouse gases then re-emit this radiation. Some of it is re-emitted into space, however, most of it is re-emitted back towards the Earth as heat. The heat causes a warming known as **the greenhouse effect**.



▲ **Figure 20.5.4** The greenhouse effect

An increase in greenhouse gases, especially carbon dioxide caused by the combustion of fossil fuels, is resulting in the Earth getting warmer. The warming of the Earth's surface as a result of increasing concentrations of greenhouse gases in the atmosphere is called **global warming**. This global warming is causing polar ice caps and glaciers to melt, sea levels to rise, flooding of low-lying coastal areas, global climate changes and more severe weather patterns.

Ozone layer depletion

Ozone (O_3) is a gas found in the upper atmosphere which shields the Earth from dangerous ultraviolet (UV) radiation. This **ozone layer** is becoming depleted by chlorofluorocarbons (CFCs) which, until recently, were used extensively in aerosol sprays, refrigerators and air-conditioners. These CFCs are inert and insoluble in water causing them to accumulate in the atmosphere where they are damaging the ozone layer. Depletion of the ozone layer means that more UV light is reaching the Earth and this is leading to more and more people developing skin cancer, cataracts and depressed immune systems.

Eutrophication

Eutrophication is the rapid growth of green plants and algae in waterways such as lakes, ponds and rivers caused by nitrates and phosphates entering the water. Nitrate and phosphate ions are needed by green plants to make proteins so they can grow. When their concentrations increase, it promotes the growth of plant life, especially algae. This rapid growth causes the water to turn green. Underwater plants then begin to die due to reduced light penetration and the algae start to die due to overcrowding. The dead plants and algae are decomposed by aerobic bacteria which multiply and use up all the dissolved oxygen. This leads to the death of other aquatic organisms such as fish.

▲ **Figure 20.5.6** Eutrophication occurs when there is uncontrolled growth of green algae in water

Effects of pesticides

Pesticides are chemicals used to destroy pests. They are used especially in agriculture and in the control of vectors of disease such as mosquitoes. There are different types of pesticides: **herbicides** are used to kill weeds, **insecticides** are used to kill insects and **fungicides** are used to destroy fungi.

Some pesticides are particularly harmful in the environment because they are **persistent**, i.e. they do not break down in the environment. They also have the potential for bioaccumulation which occurs when organisms absorb toxic substances at a faster rate than they excrete them. For example, if fat soluble pesticides such as DDT are absorbed by living organisms they are stored in fatty tissues of the organisms where they gradually accumulate. The concentration of the pesticides in the fatty tissues of animals then increases moving up food chains. As a result, consumers at the top of the food chain are very often harmed, especially birds of prey.

Disposal of solid waste

Plastics are organic compounds composed of non-metals. Even though they are organic compounds, most plastics are **non-biodegradable**, i.e. they cannot be broken down by microorganisms and therefore they remain in the environment for very long periods of time. This makes the disposal of solid waste containing plastics a serious problem.

- Plastics make up about 25% of all solid waste going to landfills, meaning that more and more land is being used up to extend and build landfills to accept all these plastics.
- Some plastics release toxic chemicals which are persistent within the environment. When plastics are disposed of in landfills there is the danger of these chemicals contaminating groundwater and nearby soil.
- Plastics produce toxic gases when burnt which can lead to air pollution and serious health problems if they are disposed of in incinerators.
- When not disposed of properly, plastics very often land up in waterways and oceans where they are particularly harmful to aquatic organisms, often leading to their death by being ingested or by causing suffocation.

The problem of the disposal of plastics can be solved to a large extent by **recycling** all items made of plastic.

Summary questions

1 Explain how carbon dioxide contributes to global warming and suggest some of the consequences of global warming.

2 Name two gases which can form acid rain and discuss some of the effects of acid rain.

3 What are pesticides and why are they harmful to organisms at the top of food chains?

4 a What is eutrophication and what is it caused by?

b Why is eutrophication a problem in aquatic environments?

5 What are some of the problems of disposing of plastics?

Objectives

By the end of this topic you will be able to:

- define green chemistry
- outline the principles of green chemistry.

Key fact

Green chemistry is the utilisation of a set of principles that reduces or eliminates the use or generation of hazardous substances in the design, manufacture and application of chemical products and processes.

C20.6 Green chemistry

Green chemistry is also known as **sustainable chemistry**. Green chemistry involves designing chemical products and processes that reduce or eliminate the use or generation of substances that are hazardous to living organisms and the environment.

The principles of green chemistry provide a framework for chemists to use when they are designing new materials, products, processes and systems. The principles focus on the development of sustainable designs which allow for economic development at the same time as protecting human health and the environment.

Traditional chemical processes usually take place in several steps and use a variety of reagents, including solvents. They often require elevated temperatures and pressures and purification has to take place at each step. They usually create yields below 100%, generate substantial waste, incur heavy costs associated with raw materials, energy usage and waste disposal and have a negative impact on the environment.

Green chemistry is doing chemistry the way **nature** does chemistry; using renewable, biodegradable materials which do not persist in the environment or harm the environment.

Using the technologies of green chemistry provides a number of benefits including:

- reduced wastage
- reduced pollution
- safer products
- reduced use of energy and natural resources
- improved competitiveness of chemical manufacturers.

The twelve principles of green chemistry

There are *twelve* principles of green chemistry, which provide a road map for chemists to follow when designing new materials, products, processes and systems. The term ‘green chemistry’ was coined by Paul Anastas in 1991 while working at the United States Environmental Protection Agency (EPA) and, together with John Warner, he developed the twelve principles. These twelve principles are stated below with a short explanation of each.

1 Prevent waste

‘It’s better to prevent waste than to treat or clean up waste after it has been created.’

Chemical processes should be designed to produce as little hazardous waste as possible.

2 Maximise atom economy

‘Synthetic methods should be designed to maximise the incorporation of all materials used in the process into the final product.’

Chemical processes which incorporate most or all of the starting materials into the final products should be used. This would then result in few, if any, wasted atoms.

3 Design less hazardous chemical syntheses

'Wherever practicable, synthetic methods should be designed which use and generate substances that possess little or no toxicity to human health and the environment.'

Chemical processes should be designed using reactants which pose the least risk to human health and the environment and which produce harmless by-products.

4 Design safer chemicals and products

'Chemical products should be designed which perform their desired functions while minimising their toxicity.'

All new products should be designed to be safer. These products should still be effective for their specific uses, but have minimum toxicity.

5 Use safer solvents and auxiliaries

'The use of auxiliary substances should be minimised wherever possible and they should be made innocuous when used.'

Many chemical processes use auxiliary substances such as solvents and separating agents. Many of these are hazardous and lead to considerable waste. The use of these substances should be reduced to a minimum, eliminated totally or they should be replaced with less hazardous alternatives.

6 Increase energy efficiency

'The energy requirements of chemical processes should be minimised and synthetic methods should be conducted at ambient temperature and pressure if possible.'

Chemical processes should be designed to use the minimum amounts of energy. If possible, processes should be carried out at the ambient (surrounding) temperature and pressure to avoid the need for large amounts of energy associated with using higher temperatures and pressures.

7 Use renewable feedstocks

'Raw materials or feedstocks should be renewable rather than depleting whenever practicable.'

The term feedstock refers to any raw material used in an industrial process. Whenever possible, chemical processes should use raw materials which are **renewable**. These include agricultural products such as purpose-grown crops, wood chips, processed solid waste and waste from other processes.

The use of non-renewable raw materials should be reduced to a minimum since they are rapidly running out. Non-renewable raw materials include those obtained from fossil fuels, such as petroleum, natural gas and coal, and those which are mined from the Earth, such as ores.

8 Reduce derivatives

'Unnecessary derivatisation should be minimised or avoided if possible.'

Derivatisation involves temporarily modifying physical and chemical processes. Since this requires additional reagents and can generate waste, derivatisation should be avoided or minimised.

9 Use catalysts

'Catalytic reagents are superior to stoichiometric reagents.'

Catalysts are not used up in reactions since they do not take part in the reaction itself. Stoichiometric reagents take part in the reaction so they are used up and the reaction often forms unwanted or hazardous by-products.

Catalysts should be used in chemical processes for several reasons. The reaction occurs faster, the catalyst reduces the temperature needed for the reaction which saves energy, it improves the conversion of reactants to products reducing wastage of reagents and it reduces the production of by-products.

10 Design for degradation

'Chemical products should be designed so they break down into innocuous products that do not persist in the environment.'

Chemical products should be designed so that at the end of their functional life they break down into harmless products which do not persist in the environment.

11 Analyse in real-time to prevent pollution

'Analytical methods should be developed to allow for real-time, in-process monitoring and control to minimise the formation of hazardous substances.'

The progress of any chemical process should always be monitored to know when the reaction is complete and to detect the formation of any unwanted or hazardous by-products.

12 Minimise the potential for accidents

'Choose reagents to be used in chemical processes which minimise the potential for chemical accidents, including explosions, fires and toxic releases into the atmosphere.'

Reagents should always be chosen to reduce the risk of chemical accidents to a minimum.

Green chemistry in action

The best way to understand the potential of green chemistry is to see it in action. Each year, the US Environmental Protection Agency (EPA) sponsors various awards including The Presidential Green Chemistry Challenge Awards. The following are a few examples of award-wining innovations which show how the **effective application** of the principles of green chemistry by small businesses and larger corporations is beginning to solve our environmental challenges.

- Virent Energy Systems use a water-based, catalytic method to make gasoline, diesel and jet fuel from the **sugar**, **starch** and **cellulose** of plants. The process requires little external energy and can compete economically with fuels produced from petroleum. Plants are a renewable resource so their use helps reduce dependence on petroleum.
- Eastman Chemical Company has developed a new method to make esters for use in cosmetics and personal care products. Their method uses **enzymes** to make the esters which saves energy, avoids using strong acids and organic solvents, minimises the production of undesirable by-products and takes place at lower temperatures than traditional methods.
- Battelle and its partners have developed a toner for laser printers and photocopiers made from **soya oil** and **protein** together with **carbohydrates** from corn. This toner is much easier to remove from paper than traditional toners which are made from petroleum-based starting materials, allowing the paper to be easily recycled. Its production also saves considerable amounts of energy and reduces the use of petroleum.
- Buckman International has carefully selected and designed **enzymes**, derived from natural sources, to modify the cellulose fibres in wood in order to make paper. Paper made using these modified fibres is much stronger and of higher quality than paper made by traditional methods. This means that the paper can be made using less wood fibre and more recycled paper, thus conserving natural resources. The process also saves significant amounts of energy and reduces the use of additional chemicals.
- Seventh Generation produces green cleaning, paper and personal care products, such as laundry and dishwasher detergents, paper towels, baby diapers, body washes and lotions. The company uses **biodegradable** and **phosphate-** and **chlorine-free ingredients** in its products and **recycled** materials in its packaging.

Figure 20.6.1 Seventh Generation products use biodegradable and phosphate- and chlorine-free ingredients

Summary questions

1 Define 'green chemistry'.

2 Suggest some benefits of using the technologies of green chemistry.

3 Outline *six* principles of green chemistry.

Key concepts

- **Non-metals** are elements whose atoms usually have a large number of valence electrons, usually 5, 6, 7 or 8.
- In general, non-metals have low melting points and boiling points, do not conduct electricity and heat, are dull, weak and brittle in the solid state and have low densities.
- Many non-metals are gases at room temperature. Bromine is a liquid and a few are solids.
- Hydrogen, oxygen and nitrogen are colourless, tasteless and odourless gases, chlorine is a yellow green gas and sulfur is a yellow solid.
- Carbon has two allotropes, diamond and graphite.
- Diamond is a transparent, colourless, sparkling solid.
- Graphite is a soft, flaky, opaque, dark grey solid.
- Non-metals form negative anions in chemical reactions with metals by gaining electrons into their valence electron shell. As a result they form **ionic compounds**.
- Non-metals share valence electrons when they react with each other forming **covalent compounds**.
- Non-metals react with oxygen to form **non-metal oxides**.
- Non-metals react with metals to form **ionic compounds**.
- All non-metals behave as **oxidising agents** when they react with metals.
- Oxygen and chlorine always act as oxidising agents when they react with other non-metals or compounds.
- Hydrogen, carbon and sulfur act as **reducing agents** when they react with oxygen.
- Hydrogen and carbon act as reducing agents when they react with metal oxides.
- Oxygen can be prepared in the laboratory by the decomposition of hydrogen peroxide in the presence of a manganese(IV) oxide catalyst.
- Carbon dioxide can be prepared in the laboratory by reacting a carbonate with an acid, usually calcium carbonate and hydrochloric acid.
- Ammonia can be prepared in the laboratory by reacting an alkali with an ammonium salt, e.g. calcium hydroxide and ammonium chloride.
- Calcium chloride, calcium oxide and concentrated sulfuric acid may be used to dry gases made in the laboratory.
- Dry oxygen and carbon dioxide are collected by upward displacement of air, dry ammonia is collected by downward displacement of air.
- Two important uses of oxygen are in hospitals and for welding and cutting metals.
- Carbon dioxide is used in fire extinguishers, as a refrigerant, to make carbonated soft drinks and as an aerosol propellant.
- Non-metals and their compounds have a great many uses including making fibres to strengthen plastics, to strengthen rubber used in tyres, in jewellery and to make insecticides, matches, fertilisers, bleaches, glass and ceramics.
- Non-metals and their compounds can have a negative effect on living organisms and the environment.

- The concentration of these harmful non-metals and their compounds within the environment is increasing due to pollution caused by man's activities.
- The most harmful non-metal compounds include sulfur dioxide, carbon monoxide, carbon dioxide, hydrogen sulfide, oxides of nitrogen, chlorofluorocarbons, carbon particles, nitrate and phosphate ions and compounds in pesticides.
- **Acid rain** forms when sulfur dioxide and oxides of nitrogen react with rainwater.
- Carbon dioxide is contributing to **the greenhouse effect** and **global warming**.
- Chlorofluorocarbons are breaking down the **ozone layer** in the upper atmosphere allowing more ultraviolet light to reach the Earth's surface.
- Nitrate and phosphate ions are causing **eutrophication**.
- **Pesticides** concentrate up food chains, harming the top consumers.
- Disposal of solid waste containing plastics poses a serious problem. Waste containing plastics should be **recycled**.
- **Green chemistry** is the utilisation of a set of principles that reduces or eliminates the use or generation of hazardous substances in the design, manufacture and application of chemical products and processes.
- Using the technologies of green chemistry provides a number of benefits including reduced wastage, reduced pollution, safer products, reduced use of energy and natural resources and improved competitiveness of chemical manufacturers.
- There are twelve principles of green chemistry.

Practice exam-style questions

Multiple-choice questions

1 Which of the following is true of non-metals?
A Most have high densities.
B They are all non-conductors of electricity.
C They are usually solid at room temperature.
D Most have low melting points.

2 Which of the following is a list of gaseous diatomic non-metals at room temperature?
A chlorine, oxygen, bromine, nitrogen
B hydrogen, oxygen, neon, nitrogen
C fluorine, oxygen, hydrogen, chlorine
D oxygen, nitrogen, hydrogen, helium

3 Most oxides of non-metals are:
A acidic
B amphoteric
C basic
D neutral

4 In which of the following reactions is the non-metal behaving as a reducing agent?
A $Zn(s) + S(s) \longrightarrow ZnS(s)$
B $2KI(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + I_2(aq)$
C $PbO(s) + H_2(g) \longrightarrow Pb(s) + H_2O(l)$
D $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

5 Oxygen can be prepared in the laboratory by the decomposition of hydrogen peroxide. Manganese(IV) oxide is also usually used in the preparation as:
A a source of the oxygen
B a catalyst
C an oxidising agent
D a reducing agent

6 Which of the following is used as the drying agent in the preparation of dry ammonia in the laboratory?
A calcium oxide
B calcium chloride
C calcium carbonate
D concentrated sulfuric acid

7 Which of the following pairs of reactants would be the least suitable to prepare a sample of carbon dioxide in the laboratory?
A calcium carbonate and dilute nitric acid
B magnesium carbonate and dilute sulfuric acid
C copper(II) carbonate and dilute hydrochloric acid
D calcium carbonate and dilute sulfuric acid

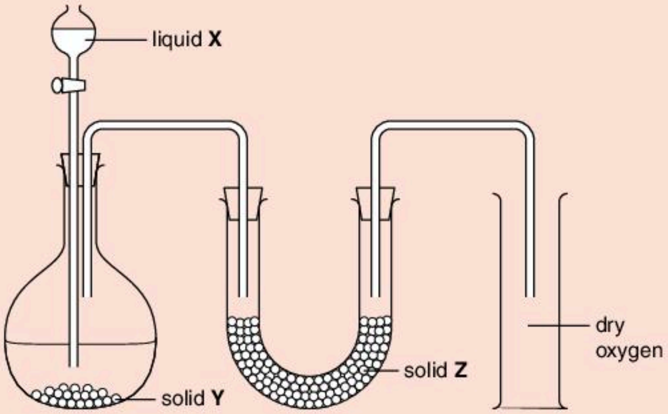
8 Carbon dioxide is used in all of the following except:
A fire extinguishers
B soft drinks
C hospitals
D aerosol sprays

9 Acid rain is caused by:
I sulfur dioxide
II nitrogen dioxide
III carbon monoxide
A I only
B I and II only
C II and III only
D I, II and III

10 The increase in carbon dioxide levels in the atmosphere are predicted to cause all of the following except:
A depletion of the ozone layer
B changes in global weather patterns
C melting of polar ice caps
D flooding of low-lying coastal areas

Structured question

11 a The diagram below shows the apparatus that could be used to prepare a dry sample of oxygen in the laboratory.



i) Identify the liquid **X** and the solid **Y**. *(2 marks)*
ii) State the role of the solid **Y**. *(1 mark)*
iii) Write an equation to show how oxygen is produced by **X** and **Y**. *(2 marks)*
iv) Name ONE compound that could be **Z** and state its role. *(2 marks)*
v) Give ONE reason for collecting the oxygen by the method shown and name the method. *(2 marks)*

b Oxygen can react with both metals and non-metals to form oxides.
i) Write a balanced equation for the reaction between zinc and oxygen. *(2 marks)*
ii) If water, followed by a few drops of blue litmus solution, is added to the oxide formed when oxygen reacts with sulfur, what would you expect to observe? Give a reason for your answer. *(2 marks)*

c The non-metal hydrogen can behave as both an oxidising agent and a reducing agent. Write TWO equations, one to show hydrogen behaving as an oxidising agent and one to show it behaving as a reducing agent. *(2 marks)*

Total 15 marks

Extended response question

12 **a** The Fourth Annual Assessment Report of the United Nations Intergovernmental Panel on Climate Change published in 2007 predicts that overall global warming is likely to be about 0.2 °C per decade for the next two decades. This global warming has been partially attributed to the increase in greenhouse gases in the atmosphere.

i) Name the main greenhouse gas. *(1 mark)*

ii) Explain how the greenhouse effect works. *(4 marks)*

iii) Discuss the possible effects of global warming on the environment. *(3 marks)*

b Over the years studies have shown that there has been a significant depletion of the ozone layer. This has been linked with the extensive use of chlorofluorocarbons for many years.
Discuss some of the consequences of this depletion. Your answer must include reference to the function of the ozone layer. *(3 marks)*

c 'Non-metals and their compounds are used extensively by man.' By reference to TWO different non-metals, provide evidence to support this statement. *(4 marks)*

Total 15 marks

C21 Water

Objectives

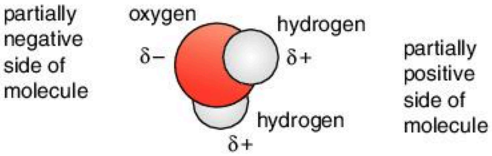
By the end of this topic you will be able to:

- describe some of the unique properties of water
- explain why the properties of water can have useful and harmful effects on living systems
- discuss the consequences of the solvent properties of water
- distinguish between the two types of water hardness.

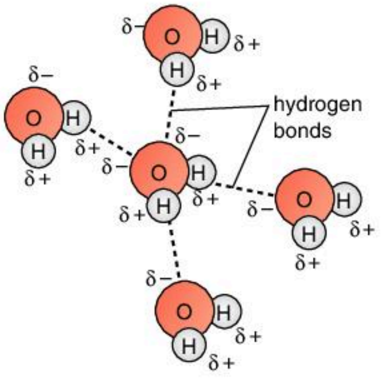
Water makes up between 60% and 70% of the human body and it covers 71% of the Earth's surface. Water is a liquid at room temperature and pressure, however, it exists on Earth in all three states, solid (ice), liquid and gas (water vapour or steam). Water has certain unique properties that make it essential for life on Earth. These properties help us to understand such things as how water becomes polluted and why life can exist at the bottom of lakes when the top is frozen.

C21.1 The unique properties of water

Pure water, i.e. water that does not contain any impurities, is a colourless, tasteless and odourless liquid that boils at 100 °C and freezes at 0 °C at standard pressure. Water molecules are **polar**, meaning that each molecule has a side that has a partial positive charge and a side that has a partial negative charge. The two hydrogen atoms in each molecule have partial positive charges and the oxygen atom has a partial negative charge. This is shown in Figure 21.1.1.



▲ **Figure 21.1.1** A polar water molecule



▲ **Figure 21.1.2** Hydrogen bonding in water

The partial positive hydrogen atoms and partial negative oxygen atoms of the water molecules are attracted to each other. This attraction forms what is known as the **hydrogen bond**. Hydrogen bonds are usually stronger than other intermolecular forces which exist between molecules and this gives water several **unique properties**. We will now look at some of these properties and at their significance to living organisms.

The maximum density of water occurs at 4 °C

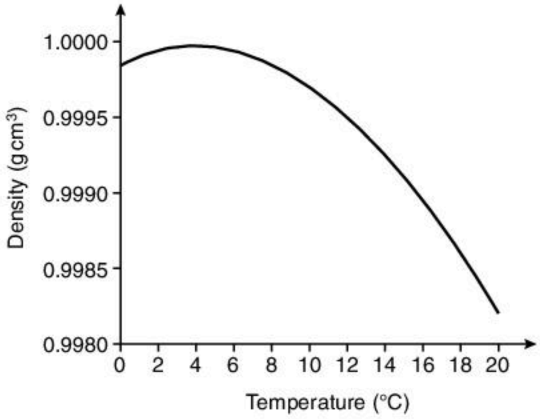
Density is a measure of how compact a substance is. It is defined as the mass of a substance divided by its volume, i.e.

$$\text{density} = \frac{\text{mass}}{\text{volume}}.$$

If the volume of a fixed mass of a substance *decreases*, its density *increases*. If its volume *increases*, its density *decreases*.

When most liquids are cooled and become solids they contract. The volume of the solid is less than the volume of the liquid from which it forms. The solid is, therefore, denser than the liquid and if it is placed in more of the liquid it sinks.

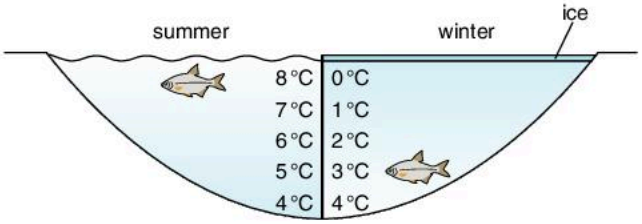
Something slightly different happens in the case of water. When water is cooled down to 4 °C it contracts and becomes denser, just like other liquids. However, if it is cooled *below* 4 °C it starts to *expand* and continues to expand until it freezes at 0 °C. This means that the volume of ice is *greater* than the volume of the water from which it forms. Solid ice is, therefore, *less dense* than the liquid water from which it formed and it floats on the liquid as it forms. This is why ice cubes float in a cold drink.



▲ **Figure 21.1.3** The maximum density of water occurs at 4 °C

▲ **Figure 21.1.4** Ice floats in water

When a pond or lake freezes, ice forms at the surface and the denser, warmer water remains below the ice. This enables aquatic organisms to survive under the ice.



▲ **Figure 21.1.5** The temperature of water in a lake in summer and in winter

Water has a high specific heat capacity

In Unit 12.2 you learnt that the **specific heat capacity** of a substance is the amount of heat needed to raise the temperature of a unit mass of the substance, e.g. 1 g, by 1 °C. Because water has a *high* specific heat capacity, it requires a lot of heat energy to increase its temperature by 1 °C. This means that water can absorb a lot of heat energy without its temperature changing very much or, in other words, when the temperature of the surroundings changes, the temperature of water does not change very much. This is of significance to living organisms for two main reasons.

- The bodies of living organisms contain between 60% and 70% water. Because of this, living organisms can absorb a lot of heat energy without their body temperature changing very much. This means they can survive in extremes of temperature.
- As environmental temperatures change, e.g. from winter to summer, the temperature of large bodies of water such as lakes and seas does not change very much. This means that organisms living in aquatic environments do not experience extreme fluctuations in temperature.

Water has a high heat of vaporisation

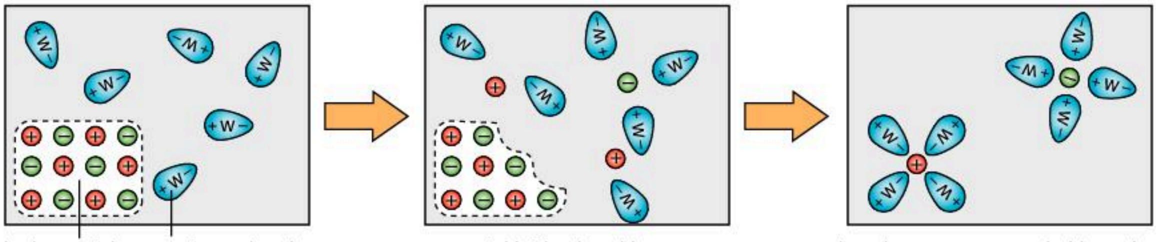
The **heat of vaporisation** is the amount of heat energy required to change a liquid to a gas. A lot of heat energy is required to change liquid water to a gas due to the hydrogen bonds between the water molecules. This means that water is not very volatile. Because water has a high heat of vaporisation, when water evaporates from the surface of a living organism it removes a large amount of heat energy from the organism. This makes sweating and transpiration very effective methods of cooling organisms.

Water has a relatively high melting point and boiling point

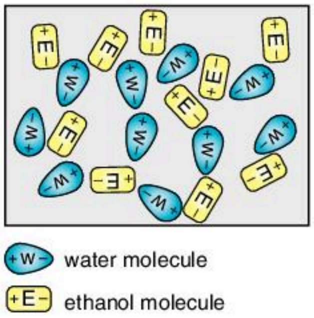
The **melting point** of ice and the **boiling point** of water are much higher than other molecules of a similar size due to the hydrogen bonds between the molecules. Water exists as a liquid between 0 °C and 100 °C, which means that at the temperatures experienced on Earth, most water is in the liquid state. Because of this, lakes, rivers and seas exist and provide an environment in which aquatic organisms can live.

Water dissolves a large number of substances

Water is sometimes referred to as a '**universal solvent**' because it dissolves a large number of substances. Because water molecules are polar, water can dissolve both ionic and polar covalent substances.



(a) Dissolving ionic substances



(b) Dissolving covalent substances

▲ **Figure 21.1.6** Water can dissolve both ionic (**a**) and covalent (**b**) substances

The fact that water can dissolve so many substances is of great significance to living organisms.

- Water dissolves chemicals in cells so that chemical reactions can take place, e.g. respiration.
- Water dissolves many useful substances so that they can be absorbed and transported around the bodies of organisms, e.g. food and mineral salts.
- Water dissolves waste products so that they can be excreted, e.g. urea.

However, the solvent properties of water can also cause *problems*.

- Water can become hard.
- Water can become polluted.
- Mineral salts can be leached out of the soil.

Consequences of the solvent properties of water

We will now look at some of the consequences of the solvent properties of water in detail.

Water hardness

Hard water is water that does not lather easily with soap, whereas **soft water** lathers easily when soap is added. Water hardness is caused by dissolved calcium and magnesium salts. When soap is put into hard water containing Ca^{2+} or Mg^{2+} ions, it forms unpleasant **scum**. This can be shown by the following equation using the soap sodium octadecanoate ($C_{17}H_{35}COONa$):

$$\underset{\text{soap}}{2C_{17}H_{35}COONa(aq)} + Ca^{2+}(aq) \longrightarrow \underset{\text{scum}}{(C_{17}H_{35}COO)_2Ca(s)} + 2Na^{+}(aq)$$

Soap only lathers when all of the calcium and magnesium ions have been precipitated out as scum, therefore, hard water wastes soap. Also scum forms a grey, greasy layer around sinks and showers and discolours clothes.

There are two types of water hardness, **temporary hardness** and **permanent hardness**.

Temporary water hardness

Temporary hardness of water is hardness that *can* be removed by boiling.

Temporary hardness is caused by dissolved calcium hydrogencarbonate ($Ca(HCO_3)_2$) and dissolved magnesium hydrogencarbonate ($Mg(HCO_3)_2$). It is found in areas rich in limestone, i.e. calcium carbonate ($CaCO_3$).

Rainwater contains dissolved carbon dioxide. When this rainwater passes through rocks containing calcium carbonate the dissolved carbon dioxide reacts with the calcium carbonate forming soluble calcium hydrogencarbonate. This is shown in the following equation:

$$CaCO_3(s) + H_2O(l) + CO_2(g) \longrightarrow Ca(HCO_3)_2(aq)$$

The soluble calcium hydrogencarbonate then dissolves in the water making it hard. This process is responsible for forming caves in limestone regions.

Permanent water hardness

Permanent hardness of water is hardness that *cannot* be removed by boiling.

Permanent hardness is caused by dissolved calcium sulfate ($CaSO_4$) and dissolved magnesium sulfate ($MgSO_4$). Calcium sulfate and magnesium sulfate are only slightly soluble and they dissolve in rainwater passing through rocks containing them.

Water pollution

When water dissolves harmful substances in the environment it becomes **polluted** (Units 19.2 and 20.5). The main water pollutants are summarised in the list below.

- Heavy metal ions, mainly from industry.
- Sulfur dioxide and oxides of nitrogen which are produced during combustion dissolve in water to form acid rain.
- Nitrate and phosphate ions present in fertilisers and detergents.
- Pesticides used in agriculture and the control of vectors of disease.
- Organic waste from untreated sewage and farmyards.

Leaching

Leaching is the loss of water soluble substances from the soil as water passes through it. When rainwater or irrigation water washes through the soil, it dissolves water soluble mineral salts and takes the salts with it into deeper layers of the soil. This often takes the salts out of the reach of plant roots, making the soil less fertile.

Summary questions

1 Explain why water has unique physical properties.
2 Explain why ice floats in water.
3 Explain how each of the following properties of water is useful to living organisms.
  a Water has a high specific heat capacity
  b Water has a high heat of vaporisation
4 Discuss:
  a the advantages of water having good solvent properties
  b the disadvantages of water having good solvent properties.
5 What is water hardness and how is it caused?

C21.2 The treatment of water for domestic purposes

Objectives

By the end of this topic you will be able to:

- describe the stages in the large-scale treatment of water
- describe ways in which water can be treated at home
- describe how hard water can be softened.

? Did you know?

In many parts of the world, humans have inadequate access to potable water and use sources **contaminated** with pathogens, toxic chemicals and suspended solids. Drinking this water or using it in food preparation is a cause of death in these countries.

The water that is piped to our homes comes from sources such as reservoirs, lakes, rivers and aquifers. It is sometimes cloudy, may have an unpleasant taste and may contain harmful microorganisms, so it must be treated to reduce or remove all the contaminants and make it safe to use. Water, which is safe enough to be consumed by humans, is known as drinking water or potable water.

Various methods are used to treat water to make it safe for use.

Large-scale water treatment

Large-scale water treatment is used to treat water before it is piped to homes. It involves the following stages.

Flocculation and sedimentation

Flocculation is the first step in the treatment process. Certain chemicals, e.g. alum, are added to the water to cause any fine suspended solid particles to clump together to form larger particles called floc. This floc is then allowed to settle within the water supply, a process known as sedimentation.

Filtration

Once the floc has settled, the clear water above it is passed through filters in order to remove any remaining particles, including some bacteria and viruses. The filters, which are usually beds of gravel, sand and charcoal, have varying compositions and pore sizes.

Chlorination

Chlorination is carried out after filtration to destroy any bacteria and viruses remaining in the water supply. The water is treated with chlorine gas, which is highly toxic to bacteria and viruses. Since chlorine is so toxic, only very small amounts are added (5 parts per million). One disadvantage of this is that it can cause the water to have an unpleasant taste if too much is added. Monochloroamine (NH_2Cl) is now being used as an alternative to chlorine since it is more stable than chlorine so remains effective for longer periods and it lacks the distinctive odour and taste of gaseous chlorine.

Water treatment at home

We can also treat the water in our homes to ensure it is safe to drink using the following methods.

Filtration

Suspended sediment can be removed from water in homes by using fibre filters made from spun cellulose or rayon fibres. Dissolved organic compounds, odours and unpleasant tastes can be removed by using carbon filters containing activated charcoal.

▲ **Figure 21.2.1** A home water filter

Chlorination

Water can be chlorinated in the home by adding sodium chlorate(I) solution (NaClO) or calcium chlorate(I) tablets [$Ca(ClO)_2$] to the water. Sodium chlorate(I) is found in chlorine bleaches. Ten drops of chlorine bleach added to 5 dm^3 of water, stirred and left for 30 minutes, is sufficient to destroy any microorganisms in the water, but not leave an unpleasant taste.

Boiling

Boiling water for 15 minutes is sufficient to kill any microorganisms in the water. After boiling, the water should be cooled before drinking.

Softening hard water

Water softening removes dissolved calcium and magnesium ions. It therefore converts hard water to soft water. With the exception of boiling, it removes both temporary and permanent hardness. There are five methods of softening hard water.

Boiling

Boiling water only removes temporary hardness. Boiling water causes the dissolved calcium hydrogencarbonate and magnesium hydrogencarbonate to decompose into insoluble calcium carbonate and magnesium carbonate, water and carbon dioxide. The insoluble carbonates precipitate out of the water thus removing the calcium and magnesium ions.

$$\text{e.g.}\quad Ca(HCO_3)_2(aq) \xrightarrow{\text{heat}} CaCO_3(s) + H_2O(l) + CO_2(g)$$

Calcium carbonate is also known as limescale. You often see it accumulating inside kettles and around shower heads.

▲ **Figure 21.2.2** An ion-exchange water softener. The black cylinders contain the ion-exchange resin and the grey tank contains the sodium chloride used during regeneration.

Exam tip

It is important that you can write **equations** to show the different methods to soften hard water.

Addition of sodium carbonate

Sodium carbonate is also known as **washing soda**. When added to hard water it precipitates out the dissolved calcium and magnesium ions as insoluble calcium carbonate and magnesium carbonate. Adding sodium carbonate removes both temporary and permanent hardness, as shown in the following equations.

Removal of temporary hardness:

$$\text{e.g. } Ca(HCO_3)_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaHCO_3(aq)$$

Removal of permanent hardness:

$$\text{e.g. } CaSO_4(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + Na_2SO_4(aq)$$

These reactions can be represented by the common ionic equation:

$$Ca^{+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

Ion-exchange

In an **ion-exchange** water softening device, water is slowly passed through an ion-exchange column containing an ion-exchange resin known as **zeolite**. Zeolite is a complex compound containing sodium ions. It can be represented by the formula Na_2Z. As the water passes through the zeolite, the calcium and magnesium ions displace the sodium ions and are absorbed into the zeolite. The calcium and magnesium ions are, therefore, removed from the water and the sodium ions enter the water in their place. Sodium ions do not cause water to be hard.

$$\text{E.g. } Ca^{2+}(aq) + \underset{\substack{\text{ion-}\\ \text{exchange}\\ \text{resin}}}{Na_2Z(s)} \longrightarrow CaZ(s) + 2Na^{+}(aq)$$

The calcium and magnesium ions are periodically removed from the zeolite by passing a highly concentrated solution of sodium chloride (brine) through it, a process known as regeneration.

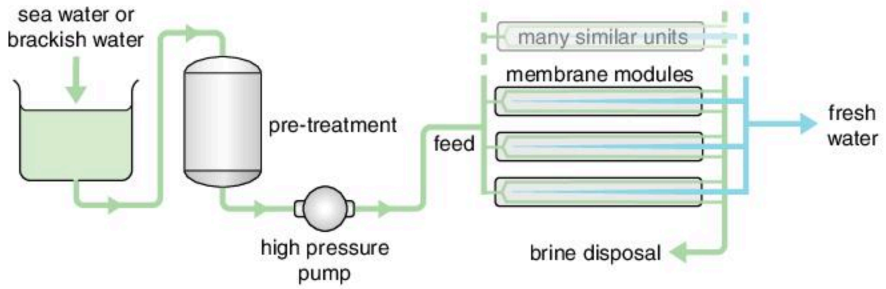
Distillation

During distillation, water is boiled and the steam produced is condensed to form pure **distilled water**. This leaves any dissolved salts and microorganisms behind. This method can be used to desalinate sea water in areas where water is in short supply. However, it is not an economical method since it uses a lot of energy to heat the water and a residue of solid builds up in the heating vessel. It can be used on a small scale in areas with a lot of sunshine where solar energy is used to heat the water in a **solar still**.

Reverse osmosis

During **reverse osmosis**, water is forced through a semi-permeable membrane under pressure. Only water molecules are forced through the membrane and all dissolved substances remain behind on the pressurised side. This has the advantage over other methods of softening water since it removes the ions which cause water hardness and it also removes other ions and any dissolved organic matter.

Most **desalination plants** make use of reverse osmosis to desalinate sea water or brackish water to produce fresh water. Brackish water contains more dissolved salts than fresh water, but less than sea water.



▲ **Figure 21.2.3** Reverse osmosis obtains fresh water from sea water

Key concepts

- Water molecules are **polar**.
- The attraction between the partial positive and partial negative parts of water molecules forms the **hydrogen bond** which has a significant effect on the physical properties of water, giving it several unique properties.
- The unique properties of water make it extremely valuable in living systems.
- The maximum density of water occurs at 4 °C. When a pond or lake freezes, ice forms at the surface and the denser, warmer water remains below the ice enabling aquatic organisms to survive under the ice.
- Water has a high **specific heat capacity**. Living organisms and large bodies of water can absorb a lot of heat energy without their temperatures changing very much.
- Water has a high **heat of vaporisation**. This means water is not very volatile. When water evaporates from the surface of a living organism it removes heat energy cooling the organism.
- Water has relatively high melting and boiling points. At temperatures experienced on Earth most water is in the liquid state making the existence of aquatic environments possible.
- Water dissolves a very large number of substances. Water dissolves chemicals in cells so they can react, food and minerals so they can be transported around organisms' bodies and waste products so they can be excreted.
- Water hardness, water pollution and leaching all result from water's solvent properties.
- **Temporary water hardness** is caused by dissolved calcium hydrogencarbonate and dissolved magnesium hydrogencarbonate.
- **Permanent water hardness** is caused by dissolved calcium sulfate and dissolved magnesium sulfate.
- Large-scale treatment of water for domestic purposes involves flocculation, sedimentation, filtration and chlorination.
- Water can be treated in the home by filtration, chlorination and boiling.
- Temporary water hardness can be removed by boiling.
- Temporary and permanent water hardness can be removed by adding sodium carbonate to the water, passing the water through an ion-exchange resin, distillation and reverse osmosis.

Summary questions

1 Describe what happens at each stage in the large-scale treatment of water.

2 **a** List three ways that water can be treated in a home.
  b What, in terms of palatability, do you think could be a disadvantage of using a chlorine bleach to treat water?

3 Explain how water can be softened by:
  a boiling
  b using an ion-exchange resin
  c using sodium carbonate.

Give the relevant equations for each of your answers.

4 What is reverse osmosis and how is it used to obtain fresh water from sea water?

Practice exam-style questions

Multiple-choice questions

1 Which of the following properties of water is important in the control of body temperature?
- **A** its boiling point
- **B** its density
- **C** its heat of vaporisation
- **D** its solvent properties

2 Aquatic organisms can live on the bottom of a lake in winter even though the surface is frozen because:
- **A** water has a relatively low melting point
- **B** the maximum density of water occurs at 4 °C
- **C** water is not very volatile
- **D** water has a high specific heat capacity

3 All of the following result from the solvent properties of water except:
- **A** acid rain
- **B** water hardness
- **C** chemical reactions in cells
- **D** soil erosion

4 The compound mainly responsible for temporary hardness of water is:
- **A** calcium sulfate
- **B** calcium hydrogencarbonate
- **C** calcium carbonate
- **D** calcium chloride

5 Permanent hardness can be removed by:

I adding sodium carbonate to the water
II passing the water through an ion-exchange column
III boiling the water

- **A** II only
- **B** I and II only
- **C** I and III only
- **D** I, II and III

6 Which of the following lists the stages involved in the large-scale treatment of water in the correct order?
- **A** flocculation, sedimentation, filtration, chlorination
- **B** sedimentation, flocculation, filtration, chlorination
- **C** chlorination, flocculation, sedimentation, filtration
- **D** flocculation, chlorination, sedimentation, filtration

Extended response question

7 a Water possesses a number of unique properties which make it essential for life on Earth. Outline THREE of these properties and explain how EACH is essential to living organisms. *(6 marks)*

b The presence of certain ions in water can cause temporary and permanent hardness.

i) What is meant by the term 'hard water'? *(1 mark)*

ii) Sodium carbonate is used to remove both temporary and permanent hardness, while temporary hardness can also be removed by boiling. Write an equation to show the removal of EACH of the following:
- permanent hardness using sodium carbonate
- temporary hardness by boiling. *(4 marks)*

iii) Suggest TWO reasons why it is useful to remove hardness from water. *(2 marks)*

c Briefly outline how water is treated before it is piped to homes. *(2 marks)*

Total 15 marks

C22 Qualitative analysis

Ionic compounds are composed of positive cations and negative anions. Being able to identify the cations and anions present in unknown ionic compounds is extremely important. This is known as qualitative analysis. Simple observations of a compound may indicate what it contains, e.g. the copper(II) ion is blue. Positively identifying ions makes use of their reactions with other chemical compounds. Being able to identify different gases is also very important. This makes use of the chemical properties of the gases.

Objectives

By the end of this topic you will be able to:

- describe tests to identify cations using sodium hydroxide solution, aqueous ammonia and potassium iodide solution
- give results of tests to identify cations
- write ionic equations for the reactions occurring in tests to identify cations.

C22.1 Identification of cations

Cations are positively charged ions. The cation present in a compound can be identified by carrying out various tests on the compound.

Identification of cations using sodium hydroxide solution

All hydroxides are *insoluble* except ammonium hydroxide and those of the alkali metals, i.e. the metals in Group I of the periodic table including sodium and potassium. This means that if sodium hydroxide solution is added to a solution containing any metal ion, except ions of the alkali metals, a precipitate will form since the metal ion reacts with the hydroxide ion (OH^-) forming an insoluble metal hydroxide. This can be summarised by the following general ionic equation:

$$M^{n+}(aq) + nOH^-(aq) \longrightarrow M(OH)_n(s)$$

The *colour* of the precipitate and its *solubility* in excess sodium hydroxide solution can give us some idea of the identity of the metal cation present in a solution.

The test is carried out by making a solution of a salt containing the cation to be identified in a test tube and adding a few drops of sodium hydroxide solution. The colour of the precipitate is noted. Sodium hydroxide solution is then added until it is in excess to see if the precipitate dissolves. If no precipitate forms, then the solution is warmed and tested for the presence of ammonia by placing a piece of moist red litmus paper across the mouth of the tube.

Table 22.1.1 shows the colour of the precipitate formed on adding a few drops of sodium hydroxide solution to salt solutions containing different cations, and the solubility of the precipitate formed when excess sodium hydroxide solution is added.

▼ **Table 22.1.1** Effect of sodium hydroxide solution on cations

Cation	**Colour of precipitate after dropwise addition of sodium hydroxide solution**	**Equation for the reaction**	**Effect of adding excess sodium hydroxide solution on the precipitate**
Ca^{2+}	White	$Ca^{2+}(aq) + 2OH^-(aq) \longrightarrow Ca(OH)_2(s)$	Precipitate is insoluble in excess, therefore it remains.
Al^{3+}	White	$Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al(OH)_3(s)$	Precipitate is soluble in excess; it dissolves and forms a colourless solution.
Zn^{2+}	White	$Zn^{2+}(aq) + 2OH^-(aq) \longrightarrow Zn(OH)_2(s)$	Precipitate is soluble in excess; it dissolves and forms a colourless solution.
Pb^{2+}	White	$Pb^{2+}(aq) + 2OH^-(aq) \longrightarrow Pb(OH)_2(s)$	Precipitate is soluble in excess; it dissolves and forms a colourless solution.
Fe^{2+}	Green, turns brown on standing	$Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$	Precipitate is insoluble in excess, therefore it remains
Fe^{3+}	Red-brown	$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$	Precipitate is insoluble in excess, therefore it remains
Cu^{2+}	Blue	$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$	Precipitate is insoluble in excess, therefore it remains
NH_4^+	No precipitate. Ammonia is evolved on warming, which turns moist red litmus paper blue.	$NH_4^+(aq) + OH^-(aq) \longrightarrow NH_3(g) + H_2O(l)$	–

▲ **Figure 22.1.1** A blue precipitate forms when sodium hydroxide solution is added to a solution containing Cu^{2+} ions

Calcium hydroxide, **iron(II) hydroxide**, **iron(III) hydroxide** and **copper(II) hydroxide** are **basic hydroxides** which do not react with sodium hydroxide. When all the metal ions have been precipitated out and excess sodium hydroxide solution is added, these precipitates *remain*.

Aluminium hydroxide, **zinc hydroxide** and **lead(II) hydroxide** are **amphoteric hydroxides** which react with the strong alkali, sodium hydroxide. When all the metal ions have been precipitated out and excess sodium hydroxide solution is added, these precipitates react with the sodium hydroxide forming soluble salts that dissolve, hence the precipitate *dissolves* and a solution is formed.

Ammonium hydroxide is soluble, therefore, *no* precipitate forms on adding sodium hydroxide to an ammonium salt. Warming the solution causes the ammonium ion and the hydroxide ion to react and form ammonia gas as shown in the equation in Table 22.1.1.

If an ammonium salt is heated directly in a dry test tube, ammonia gas is also produced:

e.g. $NH_4Cl(s) \longrightarrow NH_3(g) + HCl(g)$

Identification of cations using aqueous ammonia (ammonium hydroxide solution)

Aqueous ammonia (NH_4OH) also contains the hydroxide ion. If this is added to a solution containing any metal ion, except ions of the alkali metals, a precipitate of an **insoluble metal hydroxide** forms. The *colour* of the precipitate and its *solubility* in excess ammonium hydroxide solution can also give us some idea of the identity of the metal cation present in a solution.

The method used is the same as the one used with sodium hydroxide solution. Table 22.1.2 shows the colour of the precipitate formed on adding a few drops of aqueous ammonia and its solubility in excess aqueous ammonia.

Exam tip

You must be able to write the **ionic equations** for all the reactions occurring in the tests to identify cations and anions. When writing these equations it is essential that you remember to put the charges on the ions, that you balance the equations and that you give the correct state symbols.

▼ **Table 22.1.2** Effect of aqueous ammonia on cations

Cation	Colour of precipitate after dropwise addition of aqueous ammonia	Equation for the reaction	Effect of adding excess aqueous ammonia on the precipitate
Ca^{2+}	No precipitate	–	–
Al^{3+}	White	$Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al(OH)_3(s)$	Precipitate is insoluble in excess, therefore it remains.
Zn^{2+}	White	$Zn^{2+}(aq) + 2OH^-(aq) \longrightarrow Zn(OH)_2(s)$	Precipitate is soluble in excess; it dissolves and forms a colourless solution.
Pb^{2+}	White	$Pb^{2+}(aq) + 2OH^-(aq) \longrightarrow Pb(OH)_2(s)$	Precipitate is insoluble in excess, therefore it remains.
Fe^{2+}	Green, turns brown on standing	$Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$	Precipitate is insoluble in excess, therefore it remains.
Fe^{3+}	Red-brown	$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$	Precipitate is insoluble in excess, therefore it remains.
Cu^{2+}	Blue	$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$	Precipitate is soluble in excess; it dissolves and forms a deep blue solution.

Aluminium hydroxide, **lead(II) hydroxide**, **iron(II) hydroxide** and **iron(III) hydroxide** do not react with the weak alkali, ammonium hydroxide. When all the metal ions have been precipitated out and excess ammonium hydroxide solution is added, these precipitates *remain*.

Zinc hydroxide and **copper(II) hydroxide** react with ammonium hydroxide. When all the metal ions have been precipitated out and excess ammonium hydroxide solution is added, these precipitates react with the ammonium hydroxide forming complex soluble salts that dissolve, hence the precipitate *dissolves* and a solution is formed.

Distinguishing between the Al^{3+} ion and the Pb^{2+} ion

Combining the results of the tests using sodium hydroxide solution and aqueous ammonia, the only cations which cannot be distinguished are the Al^{3+} ion and the Pb^{2+} ion. Both ions form a *white* precipitate with sodium hydroxide solution that is *soluble* in excess and a *white* precipitate with aqueous ammonia that is *insoluble* in excess. Therefore, it is necessary to perform a further test to distinguish between these two ions.

The test can be performed by adding a few drops of potassium iodide solution to a solution containing the metal cation. The Al^{3+} ion does not produce a precipitate, however, the Pb^{2+} ion reacts with the iodide ion (I^-) to produce a *bright yellow* precipitate of lead(II) iodide:

$$Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$$

▶ **Figure 22.1.2** A bright yellow precipitate forms when potassium iodide solution is added to a solution containing Pb^{2+} ions

Practical activity

To identify cations

Your teacher may use this activity to assess:

- observation, recording and reporting.

You will be supplied with samples of salts containing calcium, aluminium, zinc, lead(II), iron(II), iron(III), copper(II) and ammonium ions, sodium hydroxide solution, aqueous ammonia, potassium iodide solution, a piece of red litmus paper and test tubes.

Method

1. Carefully place a very small spatula of each salt into a separate test tube. Add 2 cm^3 of distilled water to each tube and shake to dissolve the salt.
2. Label each tube with the symbol of the cation it contains.
3. Add a few drops of sodium hydroxide solution to each tube. Observe the colour of the precipitate.
4. Add excess sodium hydroxide solution to each tube and see if the precipitate dissolves.
5. If no precipitate forms, gently heat the contents of the test tube and test for ammonia gas by placing a piece of moist red litmus paper across the mouth of the tube.
6. Record your results in a table using the following headings for the columns: cation, colour of the precipitate on adding sodium hydroxide dropwise, ionic equation for the reaction occurring, solubility of the precipitate on adding excess sodium hydroxide solution.
7. Explain why the precipitates dissolved when excess sodium hydroxide was added to the test tubes containing aluminium, zinc and lead(II) ions, but did not dissolve when excess solution was added to the test tubes containing calcium, iron(II), iron(III) and copper(II) ions.
8. Repeat steps **1** to **6** replacing sodium hydroxide solution with aqueous ammonia.
9. Explain why the precipitates dissolved when excess aqueous ammonia was added to the test tubes containing zinc and copper(II) ions.
10. Place small spatulas of the salts containing aluminium and lead(II) ions into two separate test tubes. Add 2 cm^3 of distilled water to each and shake to make solutions.
11. Add a few drops of potassium iodide solution to each and look for a precipitate forming. Record the colour of the precipitate.

Summary questions

1. When a few drops of sodium hydroxide solution were added to an unknown solution, a blue precipitate formed. When excess sodium hydroxide solution was added, the precipitate remained.
   - **a** Suggest an identity for the cation present in the solution.
   - **b** Explain why the precipitate formed.
   - **c** Write an ionic equation for the formation of the precipitate.
   - **d** Explain why the precipitate remained when excess sodium hydroxide was added.
2. When a few drops of sodium hydroxide solution were added to a solution containing lead(II) ions, a white precipitate formed. When excess sodium hydroxide solution was added, the precipitate dissolved.
   - **a** Explain why the precipitate formed.
   - **b** Write an ionic equation for the formation of the precipitate.
   - **c** Explain why the precipitate dissolved when excess sodium hydroxide was added.
3. Explain how you could distinguish between a solution containing aluminium ions and one containing zinc ions.
4. Ammonium chloride solution does not form a precipitate when sodium hydroxide solution is added, however, on heating, ammonia gas is evolved.
   - **a** Explain why ammonia is evolved on heating.
   - **b** Write an ionic equation for the reaction.

C22.2 Identification of anions

Objectives

By the end of this topic you will be able to:

- describe tests to identify anions
- give results of the tests to identify anions
- write ionic equations for the reactions occurring in the tests to identify anions.

Anions are negatively charged ions. The anion present in a compound can be identified by carrying out various tests on the compound.

Identification of anions by heating the solid

Carbonate (CO_3^{2-}) and nitrate (NO_3^-) ions can be identified by heating a sample of the solid in a dry test tube and testing the gas evolved. Heating causes the ions to decompose as shown in Table 22.2.1.

▼ **Table 22.2.1** Effect of heat on a sample of the solid

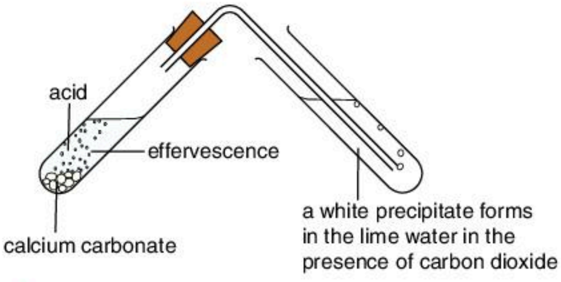
Anion	Observations on heating	Equation for the reaction
CO_3^{2-}	**Carbon dioxide** is evolved, which forms a white precipitate in lime water.	$CO_3^{2-}(s) \xrightarrow{heat} O^{2-}(s) + CO_2(g)$
NO_3^- of potassium or sodium	**Oxygen** is evolved, which relights a glowing splint.	$2NO_3^-(s) \longrightarrow 2NO_2^-(s) + O_2(g)$
NO_3^- of calcium and below	**Nitrogen dioxide** is evolved, which is a brown gas. **Oxygen** is evolved which relights a glowing splint.	$4NO_3^-(s) \xrightarrow{heat} 2O^{2-}(s) + 4NO_2(g) + O_2(g)$

Identification of anions using dilute acid

Carbonate and sulfite (SO_3^{2-}) ions can be identified by adding dilute nitric (or hydrochloric) acid to a sample of the solid, heating if necessary, and testing the gas evolved. The anions react with the hydrogen ions (H^+) in the acid as shown in Table 22.2.2.

▼ **Table 22.2.2** Effect of dilute acid on a sample of the solid

Anion	Observations on adding dilute acid	Equation for the reaction
CO_3^{2-}	Effervescence occurs. **Carbon dioxide** is evolved, which forms a white precipitate in lime water.	$CO_3^{2-}(s) + 2H^+(aq) \longrightarrow H_2O(l) + CO_2(g)$
SO_3^{2-}	Effervescence occurs. **Sulfur dioxide** is evolved on heating, which turns acidified potassium manganate(VII) solution from purple to colourless.	$SO_3^{2-}(s) + 2H^+(aq) \longrightarrow H_2O(l) + SO_2(g)$



▲ **Figure 22.2.1** Identification of a carbonate

Identification of anions using concentrated sulfuric acid

Carbonate, sulfite, chloride (Cl^-), bromide (Br^-) and iodide (I^-) ions can be identified by adding concentrated sulfuric acid to a sample of the solid in a dry test tube and testing the gas evolved. The reactions are shown in Table 22.2.3.

Table 22.2.3 Effect of concentrated sulfuric acid on the solid

Anion	Observation on adding concentrated sulfuric acid
CO_3^{2-}	**Carbon dioxide** is evolved, which forms a white precipitate in lime water
SO_3^{2-}	**Sulfur dioxide** is evolved, which turns acidified potassium manganate(VII) solution from purple to colourless
Cl^-	**Hydrogen** chloride is evolved, which forms white fumes with ammonia gas
Br^-	**Bromine** vapour is evolved, which is red-brown
I^-	**Iodine** is formed, which is a grey-black solid. The solid sublimes if heated to form a purple vapour.

Identification of anions using silver nitrate solution

Chloride, **bromide** and **iodide** ions can be identified by making a solution of the solid in dilute nitric acid, adding a few drops of silver nitrate solution, observing the colour of the precipitate and then adding aqueous ammonia and looking to see if the precipitate dissolves. The reactions are shown in Table 22.2.4.

Table 22.2.4 Effect of silver nitrate solution followed by aqueous ammonia on the solid

Anion	Observations on adding silver nitrate solution	Equation for the reaction	Observation on adding aqueous ammonia
Cl^-	A **white precipitate** of silver chloride forms, which turns slightly purple-grey in sunlight	$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$	Precipitate dissolves
Br^-	A **cream precipitate** of silver bromide forms, which turns slightly green in sunlight	$Ag^+(aq) + Br^-(aq) \longrightarrow AgBr(s)$	Precipitate partially dissolves
I^-	A pale **yellow precipitate** of silver iodide forms	$Ag^+(aq) + I^-(aq) \longrightarrow AgI(s)$	Precipitate remains

Identification of anions using barium nitrate (or chloride) solution

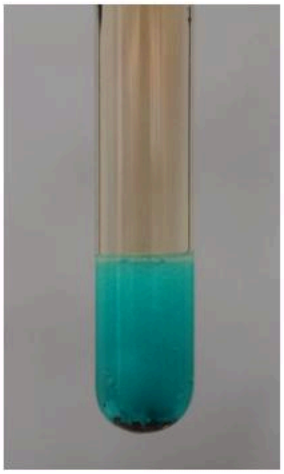
The **sulfate** (SO_4^{2-}) ion can be identified by making a solution of the solid in distilled water, adding a few drops of barium nitrate (or chloride) solution, observing the precipitate and then adding dilute nitric (or hydrochloric) acid. This test will also detect carbonate and sulfite ions. The reactions are shown in Table 22.2.5.

Table 22.2.5 Effect of barium nitrate followed by dilute acid on the solid

Anion	Observations on adding barium nitrate solution	Equation for the reaction	Observation on adding dilute acid
SO_4^{2-}	A **white precipitate** of barium sulfate forms	$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$	**White precipitate** remains
CO_3^{2-}	A **white precipitate** of barium carbonate forms	$Ba^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow BaCO_3(s)$	Precipitate dissolves releasing **carbon dioxide**
SO_3^{2-}	A **white precipitate** of barium sulfite forms	$Ba^{2+}(aq) + SO_3^{2-}(aq) \longrightarrow BaSO_3(s)$	Precipitate dissolves releasing **sulfur dioxide** on heating

Further test to identify the nitrate ion

To confirm the presence of the **nitrate** ion, concentrated sulfuric acid and a few copper turnings are added to a sample of the solid in a test tube and the tube is heated. A *blue* solution is formed and *brown* nitrogen dioxide gas is evolved.



▲ **Figure 22.2.2** Testing for the nitrate ion

Practical activity

To identify anions

Your teacher may use this activity to assess:

- observation, recording and reporting
- analysis and interpretation.

You will be supplied with samples of four salts labelled A to D (A is sodium sulfate, B is sodium chloride, C is sodium nitrate and D is sodium carbonate), test tubes and any other reagents that you have learnt about in this unit.

Method

1 You are required to carry out appropriate tests to confirm the identity of each salt. Use the information in this unit to decide on the tests you will use. You may make use of any of the reagents you have learnt about in this unit and may carry out more than one test on each salt.

2 Once you have decided on the tests to use, proceed with these tests to confirm the identity of each salt.

3 For each salt:
   a record the method used for each test you carried out and the results of the test
   b explain the results of each test
   c write ionic equations where appropriate.

Exam tip

You may be given the results of various tests on an unknown salt and be asked to **identify** the ions present in the salt. Alternatively, the names of the ions present will be given to you and you will be asked to record the expected **observations** when various tests are performed on the salt.

Summary questions

1 Barium chloride is added to a solution that contains either a sulfate or a sulfite and a white precipitate forms. What additional test should be done to determine if it is a sulfate or a sulfite?

2 Describe how you would distinguish between the following pairs of compounds:
   a sodium carbonate and sodium sulfite
   b sodium chloride and sodium iodide
   c magnesium nitrate and magnesium carbonate.

3 An unknown salt was heated and brown fumes were produced.
   a Which anion was present in the salt?
   b What other products would have been produced?
   c Write an ionic equation to show the effect of heat on the anion.

4 An unknown salt does not effervesce when dilute nitric acid is added. When barium nitrate solution is added to a solution of the salt a white precipitate forms that is insoluble in nitric acid.
   a What anion is present in the salt?
   b Write an ionic equation to show the formation of the white precipitate.

C22.3 Identification of gases

Objectives

By the end of this topic you will be able to:

- describe tests to identify various gases
- give results of the tests to identify gases
- explain the chemical principles of the tests used to identify gases.

Many gases that are produced in chemical reactions in the laboratory or are used in the laboratory are colourless. Some of them, however, do have characteristic smells. It is important to know how to test for the different gases as this can help to identify the products formed in various reactions.

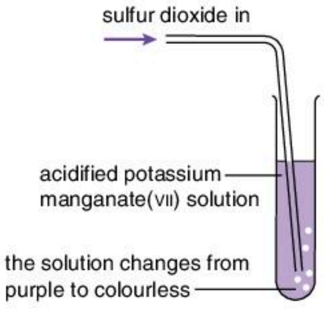
Table 22.3.1 summarises the different tests for gases produced in chemical reactions and explains the chemical principles of the tests.

▼ **Table 22.3.1** Identification of gases

Gas	Properties	Test for the presence of the gas
Hydrogen (H_2)	A colourless gas. Has no smell.	Place a **lighted splint** in the gas. A '*squeaky pop*' sound is heard and the splint is extinguished. The pop is the sound of a small explosion as the hydrogen reacts explosively with oxygen in the air to form steam: $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$
Oxygen (O_2)	A colourless gas. Has no smell.	Place a **glowing splint** in the gas. The glowing splint glows *brighter* or *re-lights*.
Carbon dioxide (CO_2)	A colourless gas. Has no smell.	Bubble the gas into **lime water** (a solution of calcium hydroxide). A *white precipitate* of calcium carbonate forms: $Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$ The precipitate redissolves on continued bubbling due to the formation of soluble calcium hydrogencarbonate: $CaCO_3(s) + CO_2(g) + H_2O(l) \longrightarrow Ca(HCO_3)_2(aq)$
Ammonia (NH_3)	A colourless gas. Has a pungent odour.	Hold a piece of **moist red litmus paper** in the gas. The litmus paper turns *blue* in colour. Ammonia reacts with water on the paper forming ammonium hydroxide, which is alkaline. $NH_3(g) + H_2O(l) \longrightarrow NH_4OH(aq)$ Place a drop of **concentrated hydrochloric acid** on a glass rod near the gas. The acid gives off hydrogen chloride fumes, which react with the ammonia and form *white fumes* of ammonium chloride: $NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$
Hydrogen chloride (HCl)	A colourless gas. Has a sharp acid smell.	Place a drop of **concentrated ammonia solution** on a glass rod near the gas. The ammonia solution gives off ammonia fumes, which react with the hydrogen chloride forming *white fumes* of ammonium chloride, as above.
Sulfur dioxide (SO_2)	A colourless gas. Has a choking smell.	Bubble the gas into **acidified potassium manganate(VII)** solution. The solution turns from *purple* to *colourless*. The sulfur dioxide reduces the purple MnO_4^- ion to the colourless Mn^{2+} ion. Bubble the gas into **acidified potassium dichromate(VI) solution**. The solution turns from *orange* to *green*. The sulfur dioxide reduces the orange $Cr_2O_7^{2-}$ ion to the green Cr^{3+} ion.
Chlorine (Cl_2)	A yellow-green gas. Has a sharp odour. Is poisonous.	Hold a piece of **moist blue litmus paper** in the gas. The litmus turns *red* and is then bleached *white*. Chlorine reacts with water on the paper, forming hydrochloric acid (HCl) and chloric(I) acid (HClO). $Cl_2(g) + H_2O(l) \longrightarrow HCl(aq) + HClO(aq)$ The chloric(I) acid oxidises the coloured litmus to colourless.
Nitrogen dioxide (NO_2)	A brown gas. Has a sharp, irritating odour. Is toxic.	Its **brown** colour makes it fairly easy to identify. Place a piece of **moist blue litmus paper** in the gas. The paper turns *red* but is not bleached. Nitrogen dioxide reacts with the water on the paper forming an acidic solution containing nitrous acid (HNO_2) and nitric acid (HNO_3): $2NO_2(g) + H_2O(l) \longrightarrow HNO_2(aq) + HNO_3(aq)$

Did you know?

The explosive reaction which occurs when hydrogen combines with oxygen produces so much energy that it is used to propel **space rockets**.



▲ **Figure 22.3.1** The test for sulfur dioxide

Did you know?

Silica gel is used to absorb moisture in anything from shoe boxes to boxes containing electronic items. Silica gel beads are often impregnated with cobalt(II) chloride, which becomes pink when the gel is saturated with water and is no longer effective. In the past, weather-predicting novelty items were coated with cobalt(II) chloride. The item was supposed to predict good weather when it turned blue and rain when it turned pink.

▼ **Table 22.3.1** (continued)

Gas	Properties	Test for the presence of the gas
Water vapour (H_2O)	A colourless gas. Has no smell.	Place a piece of **dry cobalt(II) chloride paper** in the gas. The paper changes from *blue* to *pink*. The water vapour causes the blue, anhydrous cobalt(II) chloride ($CoCl_2$) to change to pink, hydrated cobalt(II) chloride ($CoCl_2.6H_2O$). Bring the gas into contact with **anhydrous copper(II) sulfate**. The copper(II) sulfate turns from *white* to *blue*. The water vapour causes the white, anhydrous copper(II) sulfate ($CuSO_4$) to change to blue, hydrated copper(II) sulfate ($CuSO_4.5H_2O$).

Exam tip

If you are asked to describe how to test for a particular gas, you must give both the **method** you would use and the **result** you would expect to get.

Practical activity

To identify gases

Your teacher may use this activity to assess:

- observation, recording and reporting
- manipulation and measurement
- analysis and interpretation.

You will be supplied with a sample of ammonium carbonate, a sample of magnesium nitrate, some lime water, pieces of red litmus paper, blue litmus paper and dry cobalt chloride paper, tweezers, a wooden splint and test tubes, one of which is fitted with a cork with a delivery tube running through it.

Method

1 Place a spatula of ammonium carbonate into a dry test tube and place the cork with the delivery tube into the mouth of the test tube.
Place about 2 cm^3 of lime water into another tube.

2 Heat the test tube gently using a Bunsen flame. While heating, hold a piece of moist red litmus paper across the end of the delivery tube using tweezers. Observe the change in colour of the paper.

3 While still heating the tube, place a piece of dry cobalt chloride paper across the end of the delivery tube. Observe the change in colour of the paper.

4 While still heating the tube, bubble the gas into the lime water.
Observe what happens in the lime water.

5 As you are heating, observe what happens to the ammonium carbonate in the test tube.

6 Identify the three gases produced when ammonium carbonate is heated.
Hence write an equation for the decomposition of ammonium carbonate.

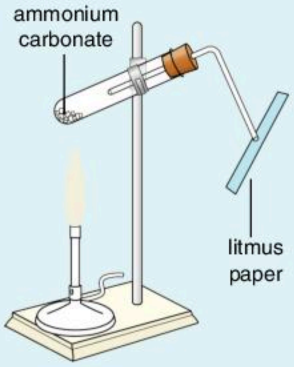
7 Explain why the quantity of ammonium carbonate decreased as it was heated and possibly disappeared completely.

8 Place a small spatula of magnesium nitrate into another dry test tube.

9 Heat the test tube gently using a Bunsen flame. As soon as a brown gas is seen, slowly insert a glowing splint into the tube. Observe what happens to the splint.

10 While still heating, place a piece of moist blue litmus paper into the brown gas.
Observe the change in colour of the paper.

11 Identify the two gases produced when magnesium nitrate is heated.
Given that the solid remaining in the tube is magnesium oxide, write a balanced equation for the reaction.



▲ **Figure 22.3.2** Identification of gases produced from the decomposition of ammonium carbonate

Summary questions

1 How would you test for the presence of oxygen gas?

2 A lighted splint is held near the mouth of a test tube and a popping sound is heard.
  a What gas is present in the tube?
  b Why did the splint make a popping sound?
  c Write a balanced equation for the reaction occurring.

3 An unknown solid was added to some hydrochloric acid in a test tube and the tube was heated. During heating, a piece of paper with acidified potassium dichromate(VI) solution on it changed colour from orange to green.
  a What gas was produced?
  b Explain why the colour of the paper changed.

4 How would you test for the presence of chlorine gas?

Key concepts

- If sodium hydroxide solution or ammonium hydroxide solution (aqueous ammonia) are added to a solution containing any metal ion, except ions of the alkali metals, a precipitate of an **insoluble metal hydroxide** will form.
- The colour of the precipitate and its *solubility* in excess sodium hydroxide or ammonium hydroxide solution can indicate the identity of a metal cation present in a solution.
- The Ca^{2+} forms a *white* precipitate with sodium hydroxide which remains in excess and does not form a precipitate with aqueous ammonia.
- Al^{3+} and Pb^{2+} ions all form *white* precipitate which are soluble in excess sodium hydroxide solution and insoluble in excess aqueous ammonia.
- The Zn^{2+} ion forms a *white* precipitate which is soluble in excess sodium hydroxide solution and excess aqueous ammonia.
- The Fe^{2+} ion forms a *green* precipitate which is insoluble in excess sodium hydroxide solution and excess aqueous ammonia.
- The Fe^{3+} ion forms a *red-brown* precipitate which is insoluble in excess sodium hydroxide solution and excess aqueous ammonia.
- The Cu^{2+} ion forms a *blue* precipitate which is insoluble in excess sodium hydroxide solution and dissolves in excess aqueous ammonia forming a deep blue solution.
- The NH_4^+ ion does not form an insoluble hydroxide but, on heating, *ammonia* gas is evolved.
- Al^{3+} and Pb^{2+} ions can be distinguished by using potassium iodide solution; Pb^{2+} ions form a *yellow* precipitate.
- When a sample of a solid is heated, the CO_3^{2-} ion produces *carbon dioxide* and the NO_3^- ion produces *brown nitrogen dioxide* and *oxygen*.
- When dilute nitric (or hydrochloric) acid is added to a solid, the CO_3^{2-} ion produces *carbon dioxide* and the SO_3^{2-} ion produces *sulfur dioxide* on heating.

- When concentrated sulfuric acid is added to a solid, the CO_3^{2-} ion produces *carbon dioxide*, the SO_3^{2-} ion produces *sulfur dioxide*, the Cl^- ion produces hydrogen chloride, the Br^- ion produces *red-brown bromine* and the I^- ion produces *grey-black iodine*.
- When a solution of a solid is made in nitric acid and silver nitrate solution is added followed by aqueous ammonia, the Cl^- ion forms a *white* precipitate which is soluble in aqueous ammonia, the Br^- ion forms a *cream* precipitate which is partially soluble and the I^- ion forms a *pale yellow* precipitate which is insoluble.
- If barium nitrate (or chloride) solution is added to a solution containing the SO_4^{2-} ion, a *white* precipitate forms which remains when dilute nitric (or hydrochloric) acid is added.
- If concentrated sulfuric acid and a few copper turnings are added to a solid containing the NO_3^- ion and heated, a *blue* solution and *brown nitrogen dioxide* form.
- Hydrogen causes a lighted splint to be extinguished with a *squeaky pop*.
- Oxygen causes a glowing splint to glow *brighter* or *re-light*.
- Carbon dioxide forms a *white precipitate* in lime water.
- Hydrogen chloride forms *white fumes* with ammonia.
- Ammonia causes moist red litmus to turn *blue* and forms *white fumes* with hydrogen chloride.
- Sulfur dioxide causes acidified potassium manganate(VII) to turn from *purple* to *colourless* and acidified potassium dichromate(VI) to turn from *orange* to *green*.
- Chlorine causes moist blue litmus paper to turn *red* and then *white*.
- Nitrogen dioxide is a *brown* gas which causes moist blue litmus paper to turn *red*.
- Water vapour causes dry cobalt(II) chloride paper to turn from *blue* to *pink* and anhydrous copper(II) sulfate to turn from *white* to *blue*.

Practice exam-style questions

Multiple-choice questions

Questions 1 and 2 refer to the following table.

	Reaction with excess sodium hydroxide solution	Reaction with potassium iodide solution	Reaction with excess aqueous ammonia
A	No precipitate	No precipitate	No precipitate
B	No precipitate	No precipitate	White precipitate
C	No precipitate	Yellow precipitate	White precipitate
D	White precipitate	No precipitate	No precipitate

Match each of the ions below to the letter that best describes its characteristics. Each letter may be used more than once, once or not at all.

1 Pb^{2+} ion.

2 Zn^{2+} ion.

3 When excess dilute nitric acid is added to salt Z, a green solution forms. On adding barium nitrate solution to half of the solution, a white precipitate forms. On adding excess aqueous ammonia to the other half of the solution, a green precipitate forms. Salt Z is:

A iron(II) carbonate
B iron(II) sulfate
C copper(II) carbonate
D copper(II) sulfate

4 A metal oxide dissolved in dilute hydrochloric acid. On addition of sodium hydroxide to the solution a white precipitate formed which remained on adding excess sodium hydroxide solution. The metal oxide was:

A zinc oxide
B lead(II) oxide
C calcium oxide
D aluminium oxide

5 A solid X was heated in a dry test tube and produced a brown gas. When a glowing splint was inserted into the tube it re-lit. X was most likely to be:

A potassium bromide
B potassium nitrate
C magnesium bromide
D magnesium nitrate

6 Aqueous silver nitrate followed by aqueous ammonia is added to a solution containing chloride ions. Which of the following observations was made?

A A white precipitate formed which was soluble in aqueous ammonia.
B A white precipitate formed which was partially soluble in aqueous ammonia.
C A white precipitate formed which was insoluble in aqueous ammonia.
D A cream precipitate formed which was soluble in aqueous ammonia.

7 Which of the following is a list of colourless, odourless gases?

A hydrogen, oxygen, sulfur dioxide
B oxygen, water vapour, ammonia
C carbon dioxide, hydrogen, oxygen
D water vapour, hydrogen, sulfur dioxide

8 Sulfur dioxide will:

I form white fumes with ammonia gas
II cause acidified potassium dichromate(VI) solution to change from green to orange
III decolourise acidified potassium manganate(VII) solution

A I and III only
B III only
C II and III only
D I and II only

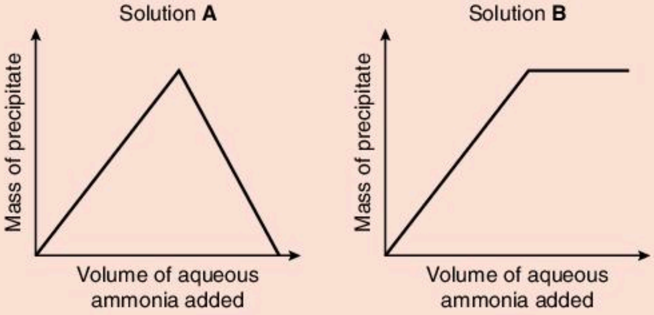
Structured question

9 a A student conducts a number of tests on a solution in order to identify the ions present. Some of the observations and inferences made from the tests are recorded in a copy of the table below. You are required to complete the missing observations and inferences in the table, including the identity of any ions present and any gases evolved, and ionic equations where required.

Test	Observation	Inference
A few drops of silver nitrate solution are added to a sample of the solution	•	• Cl^- ions present The ionic equation is:
A few drops of barium nitrate solution are added to a sample of the solution followed by dilute nitric acid	• White precipitate forms • Precipitate is insoluble in dilute nitric acid	• •
Sodium hydroxide solution is added to a sample of the solution until in excess. The mixture is warmed gently and the gas tested with moist red litmus paper.	• Green precipitate forms which is insoluble in excess sodium hydroxide solution • Litmus paper turns blue	• The ionic equation is: •

(8 marks)

b Another student was provided with two colourless solutions labelled **A** and **B**, respectively. One of the solutions contained aluminium nitrate and the other contained zinc nitrate. In order to identify the TWO solutions, the following experiment was performed: To 25.0 cm^3 portions of EACH solution, aqueous ammonia was gradually added and the mass of the precipitate obtained was determined at regular intervals. The results obtained were used to plot the following graphs:



i) Identify solutions **A** and **B**. *(2 marks)*

ii) Would there be any difference in the shapes of either of the graphs if aqueous sodium hydroxide had been used as the precipitating agent? Explain your answer. *(2 marks)*

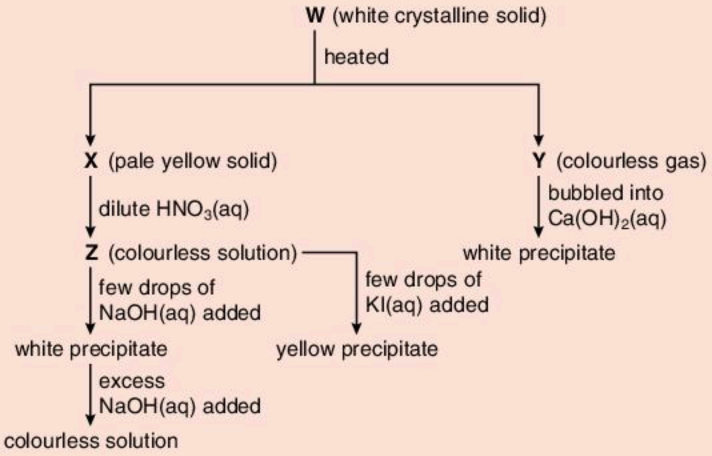
iii) Write the ionic equation for the reaction between aluminium ions and aqueous ammonia. *(2 marks)*

iv) Describe what you would expect to observe if aqueous sodium hydroxide was added, until in excess, to a solution containing copper(II) ions. *(1 mark)*

Total 15 marks

Extended response question

10 a i) Identify substances **W** to **Z** in the following reaction scheme. *(4 marks)*



ii) Write chemical equations for EACH of the following reactions:

- the effect of heat on solid **W**
- the reaction of gas **Y** with calcium hydroxide solution. *(2 marks)*

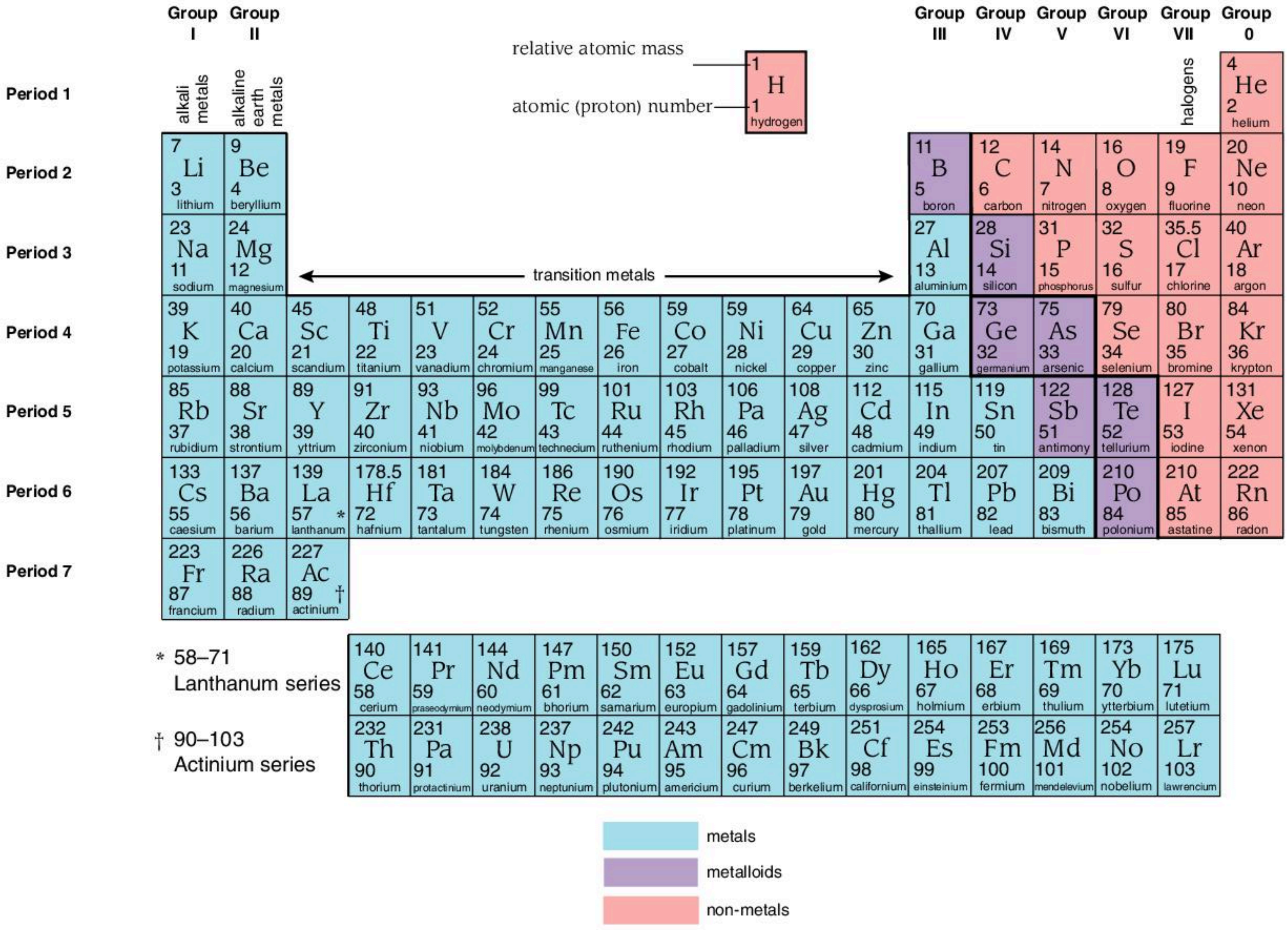
iii) Write the ionic equation for the formation of the yellow precipitate on adding a few drops of potassium iodide solution to solution **Z**. *(2 marks)*

iv) On continued bubbling of gas **Y** into the calcium hydroxide solution, the white precipitate disappeared and a colourless solution formed. Explain the reason for this observation. *(2 marks)*

b A student was given two white solids labelled **T** and **U**. She was told that **T** was potassium sulfite and **U** was calcium nitrate. Describe ONE test she could perform to confirm the identity of the anion in **T** and ONE test she could perform to confirm the identity of the cation in **U**. Your answer must include an ionic equation for ONE of the tests she used. *(5 marks)*

Total 15 marks

The periodic table



relative atomic mass — 1
H
atomic (proton) number — 1
hydrogen

	Group I alkali metals	Group II alkaline earth metals											Group III	Group IV	Group V	Group VI	Group VII halogens	Group 0 noble gases
Period 1																		4 He 2 helium
Period 2	7 Li 3 lithium	9 Be 4 beryllium											11 B 5 boron	12 C 6 carbon	14 N 7 nitrogen	16 O 8 oxygen	19 F 9 fluorine	20 Ne 10 neon
Period 3	23 Na 11 sodium	24 Mg 12 magnesium	← transition metals →										27 Al 13 aluminium	28 Si 14 silicon	31 P 15 phosphorus	32 S 16 sulfur	35.5 Cl 17 chlorine	40 Ar 18 argon
Period 4	39 K 19 potassium	40 Ca 20 calcium	45 Sc 21 scandium	48 Ti 22 titanium	51 V 23 vanadium	52 Cr 24 chromium	55 Mn 25 manganese	56 Fe 26 iron	59 Co 27 cobalt	59 Ni 28 nickel	64 Cu 29 copper	65 Zn 30 zinc	70 Ga 31 gallium	73 Ge 32 germanium	75 As 33 arsenic	79 Se 34 selenium	80 Br 35 bromine	84 Kr 36 krypton
Period 5	85 Rb 37 rubidium	88 Sr 38 strontium	89 Y 39 yttrium	91 Zr 40 zirconium	93 Nb 41 niobium	96 Mo 42 molybdenum	99 Tc 43 technecium	101 Ru 44 ruthenium	103 Rh 45 rhodium	106 Pa 46 palladium	108 Ag 47 silver	112 Cd 48 cadmium	115 In 49 indium	119 Sn 50 tin	122 Sb 51 antimony	128 Te 52 tellurium	127 I 53 iodine	131 Xe 54 xenon
Period 6	133 Cs 55 caesium	137 Ba 56 barium	139 La 57 * lanthanum	178.5 Hf 72 hafnium	181 Ta 73 tantalum	184 W 74 tungsten	186 Re 75 rhenium	190 Os 76 osmium	192 Ir 77 iridium	195 Pt 78 platinum	197 Au 79 gold	201 Hg 80 mercury	204 Tl 81 thallium	207 Pb 82 lead	209 Bi 83 bismuth	210 Po 84 polonium	210 At 85 astatine	222 Rn 86 radon
Period 7	223 Fr 87 francium	226 Ra 88 radium	227 Ac 89 † actinium															

* 58–71 Lanthanum series	140 Ce 58 cerium	141 Pr 59 praseodymium	144 Nd 60 neodymium	147 Pm 61 bhorium	150 Sm 62 samarium	152 Eu 63 europium	157 Gd 64 gadolinium	159 Tb 65 terbium	162 Dy 66 dysprosium	165 Ho 67 holmium	167 Er 68 erbium	169 Tm 69 thulium	173 Yb 70 ytterbium	175 Lu 71 lutetium
† 90–103 Actinium series	232 Th 90 thorium	231 Pa 91 protactinium	238 U 92 uranium	237 Np 93 neptunium	242 Pu 94 plutonium	243 Am 95 americium	247 Cm 96 curium	249 Bk 97 berkelium	251 Cf 98 californium	254 Es 99 einsteinium	253 Fm 100 fermium	256 Md 101 mendelevium	254 No 102 nobelium	257 Lr 103 lawrencium

🟦 metals
🟪 metalloids
🟥 non-metals

Index

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