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5 CHEMICAL REACTIONS, EQUATIONS, AND STOICHIOMETRY

5.1 THE SENTENCES OF CHEMISTRY

As noted earlier, chemistry is a language. Success in the study of chemistry depends upon how well chemical language is learned. This chapter presents the last of the most basic parts of the chemical language. When it has been learned, the reader will have the essential tools needed to speak and write chemistry and to apply it in environmental and other areas.

Recall that the discussion of chemical language began by learning about the *elements*, the *atoms* composing the elements, and the *symbols* used to designate these elements and their atoms. Atoms of the elements bond together in various combinations to produce *chemical compounds*. These are designated by *chemical formulas* consisting of symbols for the kinds of atoms in the compound and subscripts indicating the relative numbers of atoms of each kind in the compound. In chemical language, the symbols of the elements are the letters of the chemical alphabet and the formulas are the words of chemistry.

Chemical Reactions and Equations: The Sentences of the Chemical Language

The formation of chemical compounds, their decomposition, and their interactions with one another fall under the category of **chemical reactions**. Chemical reactions are involved in the annual production of millions of kilograms of industrial chemicals, bacterially mediated degradation of water pollutants, the chemical analysis of the kinds and quantities of components of a sample, and practically any other operation involving chemicals. To a very large extent, chemistry is the study of chemical reactions expressed on paper as chemical equations. A **chemical equation** is a sentence of chemistry, made up of words consisting of chemical formulas. A sentence should be put together according to rules understood by all those literate in the language. The rules of the chemical language are particularly rigorous. Although a grammatically sloppy sentence in a spoken language can still convey a meaningful message, a chemical equation with even a small error is misleading and often meaningless.

Quantitative Calculations from Chemical Equations

Chemistry is a quantitative science, and it is important to know how to do some of the basic chemical calculations early in a beginning chemistry course. Among the most important of these are the calculations of the quantities of substances consumed or produced in a chemical reaction. Such calculations are classified as **stoichiometry**. Heat is normally evolved or taken up in the course of a chemical reaction. The calculation of the quantity of heat involved in a reaction falls in the branch of chemistry called **thermochemistry**.

5.2 THE INFORMATION IN A CHEMICAL EQUATION

Chemical Reactions

A chemical reaction is a process involving the breaking and/or formation of chemical bonds and a change in the chemical composition of the materials participating in the reaction. A chemical reaction might involve the combination of two elements to form a compound. An example of this is the reaction of elemental hydrogen and oxygen to produce the compound water. Passage of an electrical current through water can cause the compound to break down and produce elemental hydrogen and oxygen. When wood burns, cellulose, a compound in the wood, reacts with elemental oxygen in air to produce the two compounds, carbon dioxide and water. If the carbon dioxide produced is bubbled through a solution of the compound calcium hydroxide dissolved in water, it produces the compounds calcium carbonate (a form of limestone) and water. Energy is involved in chemical reactions; some reactions produce energy, others require it in order for them to occur.

Expressing a Chemical Reaction as a Chemical Equation

A **chemical equation** is a means of expressing what happens when a chemical reaction occurs. It tells what reacts, what is produced, and the relative quantities of each. The information provided can best be understood by examining a typical chemical equation. For example, consider the burning of propane, a gas extracted from petroleum that is widely used for heating, cooking, grain drying, and other applications in which a clean-burning fuel is needed in areas where piped natural gas is not available. When propane burns in a camp stove, it reacts with oxygen in the air. The chemical equation for this reaction and the information in it are the following:

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O \tag{5.2.1}$$

- Propane reacts with oxygen to give carbon dioxide and water.

- There are two reactants on the left side of the equation—propane, chemical formula C_3H_8, and oxygen, chemical formula O_2.
- There are two products on the right side of the equation—carbon dioxide, chemical formula CO_2, and water, chemical formula H_2O.
- For the smallest possible unit of this reaction, 1 propane molecule reacts with 5 oxygen molecules to produce 3 carbon dioxide molecules, and 4 water molecules, as shown by the respective numbers preceding the chemical formulas (there is understood to be a 1 in front of the C_3H_8).
- There are 3 C atoms altogether on the left side of the equation, all contained in the C_3H_8 molecule, and 3 C atoms on the right side contained in 3 molecules of CO_2.
- There are 8 H atoms among the reactants, all in the C_3H_8 molecule, and 8 H atoms among the 4 molecules of H_2O in the products.
- There are 10 O atoms in the 5 O_2 molecules on the left side of the equation and 10 O atoms on the right side. The 10 O atoms in the products are present in 3 CO_2 molecules and 4 H_2O molecules.

Like mathematical equations, the left side of a chemical equation must be equivalent to the right side. Chemical equations are **balanced** in terms of atoms:

A correctly written chemical equation has equal numbers of each kind of atom on both sides of the equation.

As was just seen, the chemical equation being discussed has 3 carbon atoms, 8 hydrogen atoms and 10 oxygen atoms on both the left and right sides of the equation. Therefore, the equation is balanced. Balancing a chemical equation is a very important operation that is discussed in the next section.

Symbols Used in Chemical Equations

Several symbols are used in chemical equations. The two sides of the equation may be separated by an arrow, $\rightarrow$ or a double arrow, $\leftarrow\rightarrow$. The double arrow denotes a reversible reaction, that is, one that can go in either direction. The physical state (see Section 2.5) of a reaction component is indicated by letters in parentheses immediately following the formula. Therefore, (*s*) stands for a solid, (*l*) for a liquid (*g*) for a gas, and (*aq*) for a substance dissolved in aqueous (water) solution. An arrow pointing up, ↑, immediately after the formula of a product indicates that the product is evolved as a gas, whereas ↓ shows that it is a precipitate (solid forming from a reaction in solution and settling to the bottom of the container). These two symbols are not used extensively in this book, but they are encountered in some of the older chemical literature. The symbol Δ over the arrow dividing products from reactants shows that heat is applied to the reaction. As an example of the uses of some of the symbols just defined above, consider the following reaction:

$$CaCO_3(s) + H_2SO_4(aq) \xrightarrow{\Delta} CaSO_4(s) + CO_2(g) + H_2O(l) \qquad (5.2.2)$$

This reaction shows that solid calcium carbonate reacts with a heated solution of sulfuric acid dissolved in water to form solid calcium sulfate, carbon dioxide gas, and liquid water.

5.3 BALANCING CHEMICAL EQUATIONS

As indicated in the preceding section, a correctly written chemical equation has equal numbers of atoms of each element on both sides of the equation. Balancing a chemical equation is accomplished by placing the correct number in front of each formula in the chemical equation. However, the following must be remembered:

Only the numbers in front of the chemical formulas may be changed to balance a chemical equation. The chemical formulas, themselves, (subscript numbers) may not be changed in balancing the equation.

Balancing an equation is best accomplished by considering one element at a time, balancing it by changing the numbers preceding the formulas in which it is contained, then successively balancing other elements in the formulas contained in the equation.

Balancing the Equation for the Reaction of Hydrogen Sulfide with Sulfur Dioxide

As an example of how to balance a chemical equation, consider the reaction of hydrogen sulfide gas (H_2S) with sulfur dioxide (SO_2) to yield elemental sulfur (S) and water (H_2O). This reaction is the basis of the Claus process by which commercially valuable elemental sulfur is recovered from pollutant sulfur dioxide and from toxic hydrogen sulfide in “sour” natural gas. The steps used in balancing the equation are the following:

1. Write the correct formulas of the reactants and products on either side of the equation. *These must remain the same throughout the balancing process*.

$$\underbrace{\underset{\text{Sulfur dioxide}}{SO_2} \quad + \quad \underset{\text{hydrogen sulfide}}{H_2S}}_{\text{Reactants}} \quad \rightarrow \quad \underbrace{\underset{\text{sulfur}}{S} \quad + \quad \underset{\text{water}}{H_2O}}_{\text{Products}}$$

2. Choose an element to balance initially, preferably one that is contained in only one reactant and one product. In this case, oxygen may be chosen. The 2 oxygen atoms in the SO_2 molecule on the left may be balanced by placing a 2 in front of the H_2O product.

$$SO_2 + H_2S \rightarrow S + 2H_2O$$

3. Choose another element in one of the formulas involved in the preceding operation and balance it on both sides of the equation. In this case, the H in H_2O can be balanced by placing a 2 in front of H_2S.

   $$SO_2 + 2H_2S \rightarrow S + 2H_2O$$

4. Proceed to the remaining element. So far, sulfur has not yet been considered. There are 3 sulfur atoms on the left, contained in 1 SO_2 molecule and 2 H_2S molecules. We have already considered these molecules in preceding operations and should avoid changing the numbers of either one. However, sulfur can be balanced by placing a 3 in front of the S product.

   $$SO_2 + 2H_2S \rightarrow 3S + 2H_2O$$

5. Add up the numbers of each kind of element on both sides of the equation to see if they balance. In this case, it is seen that there are 3 S atoms, 4 H atoms, and 2 O atoms on both sides, so that the equation is, in fact, balanced.

Some Other Examples of Balancing Equations

Two other examples of balancing equations are considered here. The first of these is for the combustion in a moist atmosphere of aluminum phosphide, AlP, to give aluminum oxide and phosphoric acid, H_3PO_4. The unbalanced equation for this reaction is

$$AlP + O_2 + H_2O \rightarrow Al_2O_3 + H_3PO_4$$

The steps in balancing this equation are the following, starting with Al:

$$2AlP + O_2 + H_2O \rightarrow Al_2O_3 + H_3PO_4$$

Balance P:

$$2AlP + O_2 + H_2O \rightarrow Al_2O_3 + 2H_3PO_4$$

Balance H:

$$2AlP + O_2 + 3H_2O \rightarrow Al_2O_3 + 2H_3PO_4$$

Balance O:

$$2AlP + 4O_2 + 3H_2O \rightarrow Al_2O_3 + 2H_3PO_4 \quad (5.3.1)$$

Check each element for balance:

Reactants	Products
2 Al in 2 AlP	2 Al in 1 Al_2O_3
11 O in 4 O_2 and 3 H_2O	11 O in 1 Al_2O_3 and 2 H_3PO_4
6 H in 3 H_2O	6 H in 2 H_3PO_4

The second example of balancing equations is illustrated for trimethylchlorosilane a flammable liquid used to produce high-purity silicon for semiconductor applications. Transportation accidents have resulted in spillage of this chemical and fires that produce carbon dioxide and a fog of silicon dioxide and hydrogen chloride dissolved in water droplets.. The unbalanced equation for this reaction is

$$(CH_3)_3SiCl + O_2 \rightarrow CO_2 + H_2O + SiO_2 + HCl$$

Si and Cl are balanced as the equation stands, so balance H:

$$(CH_3)_3SiCl + O_2 \rightarrow CO_2 + 4H_2O + SiO_2 + HCl$$

Balance C:

$$(CH_3)_3SiCl + O_2 \rightarrow 3CO_2 + 4H_2O + SiO_2 + HCl$$

Balance O:

$$(CH_3)_3SiCl + 6O_2 \rightarrow 3CO_2 + 4H_2O + SiO_2 + HCl \qquad (5.3.2)$$

Checking the quantities of each of the elements in the products and reactants shows that the equation is balanced.

In some cases, the presence of a diatomic species such as O_2 necessitates doubling quantities of everything else. As an example, consider the combustion of methane (natural gas, CH_4) in an oxygen-deficient atmosphere such that toxic carbon monoxide is produced. The unbalanced equation is

$$CH_4 + O_2 \rightarrow CO + H_2O$$

Balancing C and H gives

$$CH_4 + O_2 \rightarrow CO + 2H_2O$$

This puts a total of 3 O atoms on the right, so the addition of a single oxygen atom to the reactants side to give a total of 3 Os on the left would balance oxygen. This can be done by taking ${}^{3}/_{2}O_2$ to give

$$CH_4 + {}^{3}/_{2}O_2 \rightarrow CO + 2H_2O$$

Ordinarily, however, integer numbers should be used for the coefficients. The fraction can be eliminated by multiplying everything by 2 to give the balanced equation:

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O \qquad (5.3.3)$$

Exercise: Balance the following:

1. $Fe_2O_3 + CO \rightarrow Fe + CO_2$
2. $Fe(SO)_4 + O_2 + H_2O \rightarrow Fe(OH)_3 + H_2SO_4$

3. $C_2H_2 + O_2 \rightarrow CO_2 + H_2O$
4. $Mg_3N_2 + H_2O \rightarrow Mg(OH)_2 + NH_3$
5. $NaAlH_4 + H_2O \rightarrow H_2 + NaOH + Al(OH)_3$
6. $Zn(C_2H_5)_2 + O_2 \rightarrow ZnO + CO_2 + H_2O$

Answers: (1) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$, (2) $4FeSO_4 + O_2 + 10H_2O \rightarrow 4Fe(OH)_3 + 4H_2SO_4$, (3) $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$, (4) $Mg_3N_2 + 6H_2O \rightarrow 3Mg(OH)_2 + 2NH_3$ (5) $NaAlH_4 + 4H_2O \rightarrow 4H_2 + NaOH + Al(OH)_3$, (6) $2Zn(C_2H_5)_2 + 9O_2 \rightarrow 2ZnO + 4CO_2 + 10H_2O$

Summary of Steps in Balancing an Equation

Below is a summary of steps that can be followed to balance a chemical equation. Keep in mind that there is a limit to the usefulness of following a set of rules for this procedure. Ultimately, it is a matter of experience and good judgment. In general, the best sequence of steps to take is the following:

1. Express the equation in words representing the compounds, elements, and (where present) the ions participating in the reaction.
2. Write down the correct formulas of all the reactants and all the products.
3. Examine the unbalanced equation for groups of atoms, such as those in the SO_4 ion, that go through the reaction intact. Balancing is simplified by considering these atoms as a group.
4. Examine the unbalanced equation for diatomic molecules, such as O_2, whose presence may require doubling the numbers in front of the other reaction participants.
5. Choose an element, preferably one that is found in only one reactant and one product, and balance that element by placing the appropriate numbers in front of both the reactants and products involved.
6. Balance another element that appears in one of the species balanced in the preceding step.
7. Continue the balancing process, one element at a time, until all the elements have been balanced.
8. Check to make sure that the same number of atoms of each kind of element appear on both sides of the equation and that the charges from charged species (ions, whose appearance in chemical equations will be considered later) in the equation also balance on the left and right.

Exercise: The reaction of liquid hydrazine, N_2H_4, with liquid dinitrogen tetroxide, N_2O_4, to produce nitrogen gas and water is used in some rocket engines for propulsion. Balance the equation for this reaction by going through the following steps:

(a) What is the unbalanced equation for the reaction?

(b) What is the equation after balancing O?

(c) What is the equation after balancing H?

(d) What is the equation after balancing N?

(e) How many atoms of each element are on both sides of the equation after going through these steps:

Answers: (a) $N_2O_4 + N_2H_4 \rightarrow H_2O + N_2$, (b) $N_2O_4 + N_2H_4 \rightarrow 4H_2O + N_2$, (c) $N_2O_4 + 2N_2H_4 \rightarrow 4H_2O + N_2$, (d) $N_2O_4 + 2N_2H_4 \rightarrow 4H_2O + 3N_2$, (e) 6 N, 4 O, 8 H.

5.4 WILL A REACTION OCCUR?

It is possible to write chemical equations for reactions that do not occur, or which occur only to a limited extent. This can be illustrated with a couple of examples. Consider the laboratory problem faced by a technician doing studies of plant nutrient metal ions leached from soil by water. The technician was using atomic absorption analysis, a sensitive instrumental technique for the determination of metal ions in solution. While determining the concentration of zinc ion, Zn^{2+}, dissolved in the soil leachate, the technician ran out of standard zinc solution used to provide known concentrations of zinc to calibrate the instrument, so that its readings would give known values from the sample solutions. Each liter of the standard solution contained exactly 1 mg of zinc in the form of dissolved zinc chloride, $ZnC1_2$. The technician reasoned that such a solution could be prepared by weighing out 100 mg of pure zinc metal, dissolving it in a solution of hydrochloric acid (HC1), diluting the solution to a volume of 1000 milliliters (mL), and in turn diluting 10 mL of that solution to 1000 mL to give the desired solution containing 1 mg of zinc per L. After thinking a bit, the technician came up with the equation,

$$Zn(s) + 2HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq) \tag{5.4.1}$$

to describe the chemical reaction. When the 100-mg piece of zinc metal was added to some hydrochloric acid in a flask, bubbles of hydrogen gas were evolved, the zinc dissolved as zinc chloride, and the standard solution containing the desired concentration of dissolved zinc was prepared according to the plan.

Later in the investigation, the technician ran out of standard copper solution containing 1 mg/L of copper in the form of dissolved copper(II) chloride, $CuCl_2$. The same procedure that was used to prepare the standard zinc solution was tried, with a 100-mg piece of copper wire substituted for the zinc metal. However, nothing happened to the copper metal when it was placed in hydrochloric acid. No amount of heating, stirring or waiting could persuade the copper wire to dissolve. The technician wrote the chemical equation,

$$Cu(s) + 2HCl(aq) \rightarrow H_2(g) + CuCl_2(aq) \tag{5.4.2}$$

for the reaction analogous to that of zinc, but copper metal and hydrochloric acid simply do not react with each other. Even though a plausible chemical equation can be written for a reaction, it does not tell whether the chemical reaction will, in fact, occur. Consideration of whether particular reactions take place is considered in the realms of chemical thermodynamics and chemical equilibrium, covered in later chapters.

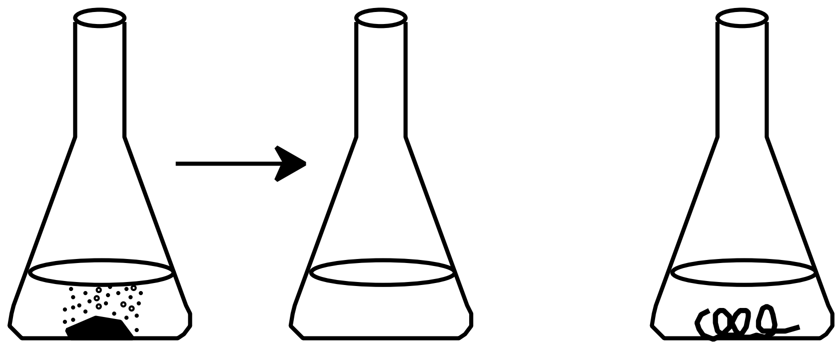

Figure 5.1 A piece of zinc metal in contact with hydrochloric acid solution reacts rapidly, giving off hydrogen gas and going into solution as $ZnCl_2$.

A piece of copper metal (wire) placed in hydrochloric acid solution does not react.

5.5 HOW FAST DOES A REACTION GO?

Consider a spoonful of table sugar, sucrose, exposed to air. Will sucrose, chemical formula $C_{12}H_{22}O_{11}$, react with oxygen in the air? The following equation can be written for the chemical reaction that might occur as follows:

$$C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(l) \qquad (5.5.1)$$

One may have an intuitive feeling that this reaction should occur from having seen sugar burn in a fire, or from knowing that the human body "burns" sugar to obtain energy. Furthermore, it is true that from the standpoint of energy, the atoms shown in the above equation are more stable when present as 12 molecules of CO_2 and 11 molecules of H_2O, rather than as one molecule of $C_{12}H_{22}O_{11}$ and 12 molecules of O_2. But anyone knows from experience that a spoonful of sugar can be exposed to dry air for a very long time without the occurrence of any visible change.

Figure 5.2: Sugar exposed to air at room temperature does not react with the oxygen in the air at a detectable rate.

The answer to the question raised above lies in the **rate of reaction**. Sucrose does, indeed, tend to react with O_2 as shown in the chemical equation above. But at room temperature, the reaction is just too slow to be significant. Of course, if the sugar were thrown into a roaring fire in a fireplace, it would burn rapidly. Special proteins known as enzymes inside of living cells can bring about the reaction of sugar and oxygen at body temperature of about 37°C, enabling the body to use the energy from the reaction. The enzymes themselves are not used up in the reaction, though they speed it up greatly; a substance that acts in such a manner is called a **catalyst**.

An important distinction must be made between reactions, such as that between copper and hydrochloric acid that will not occur under any circumstances, and others that "want to occur," but that are just too slow to be perceptible at moderate temperatures or in the absence of a catalyst. The latter type of reaction often does take place under the proper conditions or with a catalyst. Rates of reactions are quite important in chemistry.

5.6 CLASSIFICATION OF CHEMICAL REACTIONS

Chemical reactions may involve several kinds of processes. Reactions may consist of putting elements together to make compounds, or taking compounds apart to produce the component elements. Reactions may occur between compounds, between compounds and elements, or between ions. Many reactions involve the transfer of electrons, whereas others do not. Given these possibilities, plus others, it is helpful to categorize chemical reactions in several major classes. These are defined below.

A **combination reaction** is one in which two reactants bond together to form a single product. An example of such a reaction is provided by the burning of elemental phosphorus as one of the steps in the manufacture of phosphoric acid, a widely used industrial chemical and fertilizer ingredient. The reaction is,

$$P_4 \quad + \quad 5O_2(g) \quad \rightarrow \quad P_4O_{10} \tag{5.6.1}$$

Elemental phosphorus, which occurs as the molecule with 4 P atoms. — Tetraphosphorus decoxide

This is one example of the many combination reactions in which two elements combine together to form a compound. The general classification, however, may be applied to combinations of two compounds or of a compound and an element to form a compound. For example, the P_4O_{10} produced in the preceding reaction is combined with water to yield phosphoric acid:

$$6H_2O + P_4O_{10} \rightarrow 4H_3PO_4 \tag{5.6.2}$$

Phosphoric acid

A **decomposition reaction** is the opposite of a combination reaction. An example of a decomposition reaction in which a compound decomposes to form the elements in it is provided by the manufacture of carbon black. This material is a finely divided form of pure carbon, C, and is used as a filler in rubber tire manufacture and as an ingredient in the paste used to fill electrical dry cells. It is made by heating methane (natural gas) to temperatures in the range of 1260–1425°C in a special furnace, causing the following reaction to occur:

$$CH_4(g) \xrightarrow{\Delta} C(s) + 2H_2(g) \tag{5.6.3}$$

The finely divided carbon black product is collected in a special device called a cyclone collector, shown in Figure 5.3, and the hydrogen gas by-product is recycled as a fuel to the furnace that heats the methane. As illustrated by the preceding reaction, a reaction in which a compound is broken down into its component elements is a decomposition reaction.

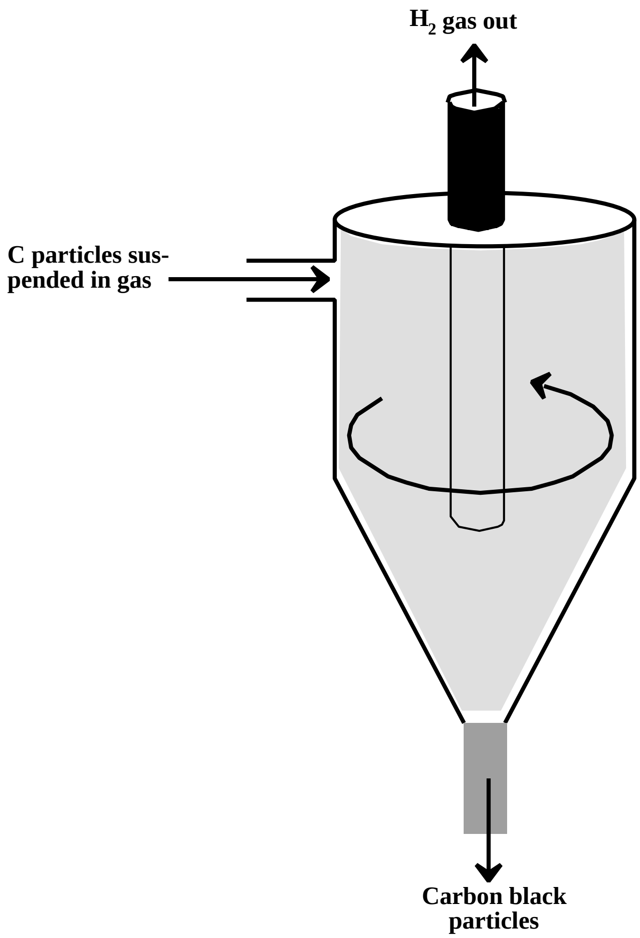

Figure 5.3. A cyclone collector is used to collect carbon black particles from a gas stream.

Decomposition reactions may also involve the breakdown of a compound to another compound and an element, or to two or more compounds. As an example of the latter, consider the reaction for the manufacture of calcium oxide, CaO, commonly called quicklime:

$$\underset{\substack{\text{Calcium carbonate}\\ \text{(limestone)}}}{CaCO_3(s)} \xrightarrow{\Delta} \underset{\substack{\text{Calcium oxide}\\ \text{(quicklime)}}}{CaO(s)} + CO_2(g) \tag{5.6.4}$$

In this reaction, high temperatures are used to decompose limestone ($CaCO_3$) to CaO and carbon dioxide gas. This is an important reaction because quicklime, CaO, ranks second only to sulfuric acid in annual chemical production. It is used in the manufacture of cement, for water treatment, and in many other applications.

Substitution or **replacement** reactions occur when a component of a chemical compound is replaced by something else. For example, zinc displaces hydrogen in Reaction 5.4.1. A **double replacement** or **metathesis** reaction occurs when there is a two-way exchange of ions between compounds. This happens, for example, when a solution of sulfuric acid reacts with a solution of barium hydroxide,

$$H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2H_2O(l) \qquad (5.6.5)$$

to yield solid barium sulfate and water. This reaction also falls into two other categories. Because it involves the combination of H^+ ions (from H_2SO_4) and OH^- ions (from $Ba(OH)_2$) to produce water, it is a **neutralization** reaction. This particular neutralization reaction is also a **precipitation** reaction because of the formation of a solid, $BaSO_4$. Such a solid formed by the reaction of two dissolved chemicals is called a precipitate.

The reaction of H^+ from any acid with OH^- to produce water is a neutralization reaction.

Evolution of a gas can also be used as a basis for classifying reactions. An example is provided by the treatment of industrial wastewater containing dissolved ammonium chloride, NH_4Cl. This compound is composed of the ammonium ion, NH_4^+, and the chloride ion, Cl^-. Commercially valuable byproduct ammonia gas can be recovered from such water by the addition of calcium hydroxide,

$$Ca(OH)_2(aq) + 2NH_4Cl(aq) \rightarrow 2NH_3(g) + H_2O(l) + CaCl_2(aq) \qquad (5.6.6)$$

resulting in the evolution of ammonia gas, NH_3, which can be recovered.

<u>Exercise</u>: Classify each of the following reactions as combination, decomposition, substitution, metathesis, neutralization, precipitation, or evolution of a gas. In some cases, a reaction will fit into more than one category.

(a) $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

(b) $2KClO_3(s) \xrightarrow{\Delta} 2KCl(s) + 3O_2(g)$

(c) $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

(d) $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

(e) $Fe(s) + CuCl_2(aq) \rightarrow Cu(s) + FeCl_2(aq)$

(f) $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

(g) $MgCl_2(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + 2NaCl(a)$

Answers: (a) Combination, (b) decomposition, evolution of a gas, (c) combination, (d) metathesis, evolution of a gas, (e) substitution, (f) neutralization, metathesis, (g) precipitation, metathesis.

5.7 QUANTITATIVE INFORMATION FROM CHEMICAL REACTIONS

Review of Quantitative Chemical Terms

So far, chemical equations have been described largely in terms of individual atoms and molecules. Chemistry deals with much larger quantities, of course. On an industrial level, kilograms, tons, or even thousands of tons are commonly used. It is easy to scale up to such large quantities, because the relative quantities of materials involved remain the same, whether one is dealing with just a few atoms and molecules, or train-car loads of material. Before proceeding with the discussion of quantitative calculations with chemical equations, it will be helpful to consider some terms that have been defined previously:

Formula mass: The sum of the atomic masses of all the atoms in a formula unit of a compound. Although the average masses of atoms and molecules may be expressed in atomic mass units (amu or u), formula mass is generally viewed as being relative and without units.

Molar mass: Where X is the formula mass, the molar mass is X grams of an element or compound, that is, the mass in grams of 1 mole of the element or compound.

Mole: The fundamental unit for quantity of material. Each mole contains Avogadro's number (6.022×10^{23}) of formula units of the element or compound.

Formula	Formula mass	Molar mass	Number of formula units per mole
H	1.01	1.0l g/mol	6.022×10^{23} H atoms/mol
N	14.01	14.01 g/mol	6.022×10^{23} N atoms/mol
N_2	2 x 14.01 = 28.02	28.02 g/moL	6.022×10^{23} N_2 molecules/mol
NH_3	14.01 + 3 x 1.01=17.04	17.04 g/mol	6.022×10^{23} NH_3 molecules/ mol
CaO	40.08 + 16.00 =56.08	56.08 g/mol	6.022×10^{23} $\frac{\text{formula units CaO}}{\text{mol}}$*

* Since CaO consists of Ca^{2+} and O^{2-} ions, there are not really individual CaO molecules, so it is more correct to refer to a formula unit of CaO consisting of 1 Ca^{2+} ion and 1 O^{2-} ion.

Calcination of Limestone

To illustrate some of the quantitative information that may be obtained from chemical equations, consider the calcination of limestone to make quicklime for water treatment:

$$\underset{\text{Limestone, calcium carbonate}}{CaCO_3(s)} \xrightarrow{\Delta} \underset{\text{Quicklime, calcium oxide}}{CaO(s)} + CO_2(g) \qquad (5.7.1)$$

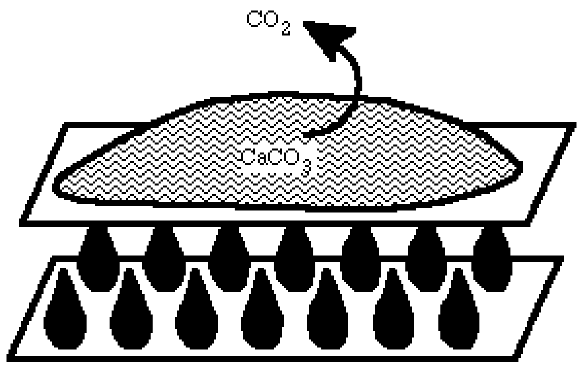

Figure 5.4 Heating calcium carbonate to a high temperature results in the production of quicklime, CaO, and the evolution of carbon dioxide. The reaction is a decomposition reaction and the process is called calcination.

The quantitative information contained in this equation can be summarized as follows:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

At the formula unit (molecular) level

1 formula unit	1 formula unit	1 molecule
1 Ca atom (atomic mass 40.1)	1 Ca atom	1 C atom
1 C atom (atomic mass 12.0)	1 O atom	2 O atoms
3 O atoms (atomic mass 16.0)		
100. 1 u	56.1 u	44.0 u

At the mole level

1 mole of $CaCO_3$	1 mole of CO	1 mole of CO_2
100. 1 g of $CaCO_3$	56.1 g of CO	44.0 g of CO_2

From the quantities above, it is seen that the equation can be viewed in terms as small as the smallest number of molecules and formula units. In this case that involves simply 1 formula unit of $CaCO_3(s)$, 1 formula unit of $CaO(s)$, and 1 molecule of CO_2. This would involve a total of 1 Ca atom, 1 C atom, and 3 O atoms. From such a small scale it is possible to expand to moles by scaling up by 6.022×10^{23} (Avogadro's number), giving 100.1 g of $CaCO_3$, 56.1 g of CaO, and 44.0 g of CO_2. Actually, these quantitative relationships are applicable to any amount of matter and they enable the calculation of the amounts of material reacting and produced in a chemical reaction. Next, it is shown how these kinds of calculations are performed.

5.8 WHAT IS STOICHIOMETRY AND WHY IS IT IMPORTANT?

Stoichiometry is the calculation of the quantities of reactants or products involved in a chemical reaction. The importance of stoichiometry can be appreciated by visualizing industrial operations that process hundreds or thousands of tons of chemicals per day. The economics of many chemical manufacturing processes are such that an unnecessary excess of only a percent or so of a reacting chemical can lead to waste that can make the operation unprofitable. Obtaining accurate values in chemical analysis, which may need to be known to within about a part per thousand, often involves highly exacting stoichiometric calculations.

Stoichiometric calculations are based upon **the law of conservation of mass**, which states that:

The total mass of reactants in a chemical reaction equals the total mass of products; Matter is neither created nor destroyed in chemical reactions.

The key to doing stoichiometric calculations correctly is the following:

The relative masses (or number of moles, number of atoms, or molecules) of the participants in a designated chemical reaction remain the same, regardless of the overall quantities of reaction participants.

The Mole Ratio Method of Stoichiometric Calculations

In a chemical reaction, there is a definite ratio between the number of moles of a particular reactant or product and the number of moles of any other reactant or product. These ratios are readily seen by simply examining the coefficients in front of the reaction species in the chemical equation. Normally, a stoichiometric calculation is made to relate the quantities of only two of the reaction participants. The objective may be to figure out how much of one reactant will react with a given quantity of another reactant. Or, a particular quantity of a product may be desired, so that it is necessary to calculate the quantity of a specific reactant needed to give the amount of product. To perform stoichiometric calculations involving only two reaction participants, it is necessary only to know the relative number of moles of each and their molar masses. The most straightforward type of stoichiometric calculation is the **mole ratio method** defined below:

The ***mole ratio method*** is *a means of performing stoichiometric calculations based upon the constant ratios of the numbers of moles of various reactants and products regardless of the overall quantity of reaction taking place.*

The mole ratio method greatly simplifies stoichiometric calculations. It can even be used to relate relative quantities of reaction participants in a series of reactions. For example, if a particular quantity of reactant is involved in a reaction followed by one or more additional reactions, the amount of a product in the final reaction is readily calculated by the mole ratio method.

To illustrate the mole ratio method, consider a typical reaction, that of hydrogen gas and carbon monoxide gas to produce methane:

$3H_2$	+	CO	→	CH_4	+	H_2O		(5.8.1)
3 moles		1 mole		1 mole		1 mole		
6.0 g		28.0 g		16.0 g		18.0 g		

This reaction is called the **methanation reaction**, and is used in the petroleum and synthetic fuels industry for the manufacture of non-polluting synthetic natural gas (CH_4). From examination of the chemical equation it is easy to get the ratio of the number of moles of any reaction participant to the number of moles of any other reaction participant as shown in Table 5.1.

Table 5.1 Mole Ratios Used in Calculations with the Methanation Reaction

Equality of number of moles	Mole ratios	
$3\text{ mol } H_2 = 1\text{ mol } CO$	$\frac{3\text{ mol } H_2}{1\text{ mol } CO}$	$\frac{1\text{ mol } CO}{3\text{ mol } H_2}$
$3\text{ mol } H_2 = 1\text{ mol } CH_4$	$\frac{3\text{ mol } H_2}{1\text{ mol } CH_4}$	$\frac{1\text{ mol } CH_4}{3\text{ mol } H_2}$
$3\text{ mol } H_2 = 1\text{ mol } H_2O$	$\frac{3\text{ mol } H_2}{1\text{ mol } H_2O}$	$\frac{1\text{ mol } H_2O}{3\text{ mol } H_2}$
$1\text{ mol } CO = 1\text{ mol } CH_4$	$\frac{1\text{ mol } CO}{1\text{ mol } CH_4}$	$\frac{1\text{ mol } CH_4}{1\text{ mol } CO}$
$1\text{ mol } CO = 1\text{ mol } H_2O$	$\frac{1\text{ mol } CO}{1\text{ mol } H_2O}$	$\frac{1\text{ mol } H_2O}{1\text{ mol } CO}$
$1\text{ mol } CH_4 = 1\text{ mol } H_2O$	$\frac{1\text{ mol } CH_4}{1\text{ mol } H_2O}$	$\frac{1\text{ mol } H_2O}{1\text{ mol } CH_4}$

These ratios enable calculation of the number of moles of any reaction participant, if the number of moles of any other participant is known. For example, if it is known that 1.00 mole of H_2 reacts, the calculation of the number of moles of CH_4 produced is simply the following:

$$1.00\text{ mol } H_2 \times \frac{1\text{ mol } CH_4}{3\text{ mol } H_2} = 0.333\text{ mol } CH_4 \tag{5.8.2}$$

Calculation of the mass of a substance requires conversion between moles and mass. Suppose that one needs to know the mass of H_2 required to produce 4.00 g of CH_4. The first step is to convert the mass of CH_4 to moles:

$$4.00 \text{ g } CH_4 \times \underset{\substack{\text{The molar mass of} \\ CH_4 \text{ is 16.0 g/mol}}}{\frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4}} = 0.250 \text{ mol } CH_4 \tag{5.8.3}$$

The next step is to multiply by the mole ratio of H_2 to CH_4:

$$0.250 \text{ mol } CH_4 \times \frac{3 \text{ mol } H_2}{1 \text{ mol } CH_4} = 0.750 \text{ mol } H_2 \tag{5.8.4}$$

The last step is to multiply by the molar mass of H_2:

$$0.750 \text{ mol } H_2 \times \frac{2.00 \text{ g } H_2}{1 \text{ mol } H_2} = 1.50 \text{ g } H_2 \tag{5.8.5}$$

All of these steps can be performed at once, as shown below:

$$4.00 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4} \times \frac{3 \text{ mol } H_2}{1 \text{ mol } CH_4} \times \frac{2.00 \text{ g } H_2}{1 \text{ mol } H_2} = 1.50 \text{ g } H_2 \tag{5.8.6}$$

Several problems will be shown that illustrate the mole ratio method. First, however, it will be helpful to learn the following steps used in solving a stoichiometric problem by this method:

1. Write the balanced chemical equation for the reaction involved.

2. Identify the reactant or product whose quantity is known (known substance) and the one whose quantity is being calculated (desired substance).

3. Express the number of moles of the known substance, which usually must be calculated from its mass.

4. Multiply the number of moles of known substance times the mole ratio of desired substance to obtain the number of moles of desired substance.

$$\begin{matrix}\text{Moles desired} \\ \text{substance}\end{matrix} = \begin{matrix}\text{moles of known} \\ \text{substance}\end{matrix} \times \begin{matrix}\text{mole ratio of desired substance} \\ \text{to known substance}\end{matrix}$$

5. Calculate the number of grams of desired substance by multiplying its molar mass times the number of moles.

$$\begin{matrix}\text{Mass in grams of} \\ \text{desired substance}\end{matrix} = \begin{matrix}\text{molar mass of} \\ \text{desired substance}\end{matrix} \times \begin{matrix}\text{moles of desired} \\ \text{substance}\end{matrix}$$

To illustrate these steps, consider the preparation of ammonia, NH_3, from hydrogen gas and atmospheric nitrogen. The reaction in words is

Hydrogen plus nitrogen yields ammonia

Insertion of the correct formulas gives,

$$H_2 \quad + \quad N_2 \quad \rightarrow \quad NH_3 \qquad (5.8.7)$$

and the equation is balanced by placing the correct coefficients in front of each formula:

$$\underset{3\text{ mol}}{3H_2} \quad + \quad \underset{1\text{ mol}}{N_2} \quad \rightarrow \quad \underset{2\text{ mol}}{2NH_3} \qquad (5.8.8)$$

As an example, calculate the number of grams of H_2 required for the synthesis of 4.25 g of NH_3 using the following steps:

1.Calculate the number of moles of NH_3 (molar mass 17.0 g/mole).

$$\text{Mol } NH_3 = 4.25\text{ g } NH_3 \times \frac{1\text{ mol } NH_3}{17.0\text{ g } NH_3} = 0.250\text{ mol } NH_3$$

2.Express the mole ratio of H_2 to NH_3 from examination of Equation 5.8.8.

$$\frac{3\text{ mol } H_2}{2\text{ mol } NH_3}$$

3.Calculate the number of moles of H_2.

$$\text{Mol } H_2 = 0.250\text{ mol } NH_3 \times \frac{3\text{ mol } H_2}{2\text{ mol } NH_3} = 0.375\text{ mol } H_2$$

4.Calculate the mass of H_2.

$$0.375\text{ mol } H_2 \times \frac{2.00\text{ g } H_2}{1\text{ mol } H_2} = 0.750\text{ g } H_2$$

Once the individual steps involved are understood, it is easy to combine them all into a single calculation as follows:

$$4.25\text{ g } H_2 \times \frac{1\text{ mol } NH_3}{17.0\text{ g } NH_3} \times \frac{3\text{ mol } H_2}{2\text{ mol } NH_3} \times \frac{2\text{ g } H_2}{1\text{ mol } H_2} = 0.750\text{ g } H_2 \qquad (5.8.9)$$

As a second example of a stoichiometric calculation by the mole ratio method, consider the reaction of iron(III) sulfate, $Fe(SO_4)_3$, with calcium hydroxide, $Ca(OH)_2$. This reaction is used in water treatment processes for the preparation of gelatinous iron(III) hydroxide, $Fe(OH)_3$, which settles in the water carrying solid particles with it. The iron(III) hydroxide acts to remove suspended matter (turbidity) from water. The $Ca(OH)_2$ (slaked lime) is added as a base (source of OH^- ion) to react with iron(III) sulfate. The reaction is

$$Fe_2(SO_4)_3(aq) + 3Ca(OH)_2(aq) \rightarrow 2Fe(OH)_3(s) + 3CaSO_4(s) \tag{5.8.10}$$

Suppose that a mass of 1000 g of iron(III) sulfate is to be used to treat a tankful of water. What mass of calcium hydroxide is required to react with the iron(III) sulfate? The steps required to solve this problem are the following:

1. $$\text{Formula mass } Fe_2(SO_4)_3 = \underbrace{2 \times 55.8}_{\text{2 Fe atoms}} + \underbrace{3 \times 32.0}_{\text{3 S atoms}} + \underbrace{12 \times 16.0}_{\text{12 O atoms}} = 399.6$$

   $$\text{Formula mass } Ca(OH)_2 = \underbrace{1 \times 40.1}_{\text{1 Ca atom}} + \underbrace{2 \times 16.0}_{\text{2 O atoms}} + \underbrace{2 \times 1.0}_{\text{2 H atoms}} = 74.1$$

2. $$\text{Mol } Fe_2(SO_4)_3 = 1000 \text{ g } Fe_2(SO_4)_3 \times \frac{1 \text{ mol } Fe_2(SO_4)_3}{399.6 \text{ g } Fe_2(SO_4)_3}$$

3. $$\text{Mole ratio} = \frac{3 \text{ mol } Ca(OH)_2}{1 \text{ mol } Fe_2(SO_4)_3}$$

4. $$\text{Mass of } Ca(OH)_2 = 1000 \text{ g } Fe_2(SO_4)_3 \times \frac{1 \text{ mol } Fe_2(SO_4)_3}{400 \text{ g } Fe_2(SO_4)_3} \times \frac{3 \text{ mol } Ca(OH)_2}{1 \text{ mol } Fe_2(SO_4)_3} \times$$

   $$\frac{74.1 \text{ g } Ca(OH)_2}{1 \text{ mol } Ca(OH)_2} = 556 \text{ g } Ca(OH)_2$$

<u>Exercise</u>: Gelatinous aluminum hydroxide for the treatment of water can be generated by the following reaction with sodium bicarbonate:

$$Al_2(SO_4)_3(aq) + 6NaHCO_3(aq) \rightarrow 2Al(OH)_3(s) + 3Na_2SO_4(aq) + 6CO_2(g)$$

Calculate the mass in kg of $NaHCO_3$ required to react with and precipitate 80.0 g of $Al_2(SO_4)_3$.

Answer: The molar mass of $NaHCO_3$ is 72.0 and that of $Al_2(SO_4)_3$ is 342.3. From the above reaction it is seen that 6 mol of $NaHCO_3$ are required per mol of $Al_2(SO_4)_3$. Therefore, the calculation is

$$\text{Mass of } NaHCO_3 = 80 \text{ g } Al_2(SO_4)_3 \times \frac{1 \text{ mol } Al_2(SO_4)_3}{342.3 \text{ g } Al_2(SO_4)_3} \times \frac{6 \text{ mol } NaHCO_3}{1 \text{ mol } Al_2(SO_4)_3} \times$$

$$\frac{72.0 \text{ g } NaHCO_3}{1 \text{ mol } NaHCO_3} = 101.0 \text{ g } Ca(OH)_2$$

CHAPTER SUMMARY

The chapter summary below is presented in a programmed format to review the main points covered in this chapter. It is used most effectively by filling in the blanks, referring back to the chapter as necessary. The correct answers are given at the end of the summary.

The formation of chemical compounds, their decomposition, and their interactions with one another fall under the category of [1]____________________________. A chemical equation may be viewed as a [2]__________________ in chemical language. Calculations of the quantities of substances consumed or produced in a chemical reaction are classified as [3]_______________. The calculation of the quantity of heat involved in a reaction falls in the branch of chemistry called [4]___________________. Substances on the left side of a chemical equation are called [5]___________________ whereas those on the right are [6]___________________. A correctly written chemical equation has equal numbers of [7]____ on both sides of the equation. When used in a chemical equation, the symbols $\rightarrow$, $\leftarrow\rightarrow$, (*s*), (*l*), (*g*), (*aq*), $\uparrow$, $\downarrow$, and Δ stand for [8]______________________________
____
____
____
____________________________________, respectively. Chemical formulas [9]_____
____ in balancing chemical equations. After expressing an equation in words, the next step in balancing it is to [10]____
_________________________________. Balancing an equation may be simplified by considering groups of atoms that [11]_____________________________________. The presence of [12]____ may require doubling the numbers in front of the other reaction participants. The final step in balancing an equation is [13]____
____
____
____.

Simply because it is possible to write a chemical equation for a reaction does not necessarily mean that [14]____. Rates of reactions [15]_______________________ but may be increased by a [16]________
________________. The reaction $C + O_2 \rightarrow CO_2$ is an example of a [17]_____________ reaction. The reaction $2H_2O \rightarrow 2H_2 + O_2$ is an example of a [18]___________________ reaction. The reaction $H_2SO_4 + Fe \rightarrow FeSO_4(s) + H_2$ is a [19]____________________ reaction. A neutralization reaction is [20]____
____ and a precipitation reaction is [21]____.
Formula mass is [22]____
____. Molar mass is [23]____________
____. Each mole of an element or compound contains [24]____

____.

The molar mass of NH_3 is [25]______________ a mass that contains [26]_____________ individual H atoms. For the smallest whole numbers of moles, the equation

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

states that [27]_______________ g of $CaCO_3$ are heated to produce [28]__________ g of CaO and [29]________ g of CO_2. Stoichiometry is [30]_____________________________
____.

Stoichiometric calculations are based upon the law of conservation of mass, which states that [31]____
____
____.

The "key to doing stoichiometric calculations correctly" is the fact that [32]_________
____
____
____
____.

In a chemical reaction there is a [33]______________________________ between the number of moles of a particular reactant or product and the number of moles of any other reactant or product. Using the mole ratio method for stoichiometric calculations, the steps involved after writing the balanced chemical equation for the reaction are [34]____
____
____
____
____
____
____
____
____
____
____.

Answers to Chapter Summary

1. chemical reactions
2. sentence
3. stoichiometry
4. thermochemistry
5. reactants
6. products
7. each kind of atom
8. division between reactants and products, reversible reaction, solid, liquid, gas, substance dissolved in water, evolution of a gas, formation of a precipitate, and application of heat
9. cannot be changed

10. write down the correct formulas of all the reactants and all the products
11. go through the reaction intact
12. diatomic molecules, such as O_2
13. to check to make sure that the same number of atoms of each kind of element appear on both sides of the equation and that the charges from ions also balance on the left and right
14. the reaction will occur
15. vary greatly
16. catalyst
17. combination
18. decomposition
19. displacement
20. the reaction of H^+ from any acid with OH^- to produce water
21. one in which a solid forms from a reaction in solution
22. the sum of the atomic masses of all the atoms in a formula unit of a compound
23. the mass in grams of 1 mole of an element or compound
24. Avogadro's number (6.022×10^{23}) of formula units of the element or compound
25. 17.0 g/mol
26. $3 \times 6.022 \times 10^{23}$
27. 100. 1
28. 56.1
29. 44.0
30. the calculation of the quantities of reactants or products involved in a chemical reaction
31. the total mass of reactants in a chemical reaction equals the total mass of products, that is, matter is neither created nor destroyed in chemical reactions
32. The relative masses (or number of moles, number of atoms, or molecules) of the participants in a designated chemical reaction remain the same, regardless of the overall quantities of reaction participants.
33. definite ratio
34. (1) identify the reactant or product whose quantity is known and the one whose quantity is being calculated, (2) express the number of moles of the known substance, (3) multiply the number of moles of known substance times the mole ratio of desired substance to known substance to obtain the number of moles of desired substance, (4) calculate the number of grams of desired substance by multiplying its molar mass times its number of moles.

QUESTIONS AND PROBLEMS

Section 5.1. The Sentences of Chemistry

1. Describe chemical reactions and chemical equations and distinguish between them.

Section 5.2. The Information in a Chemical Reaction

2. Briefly summarize the information in the chemical equation

   $CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$

3. Summarize the information in the chemical equation, $H_2O(l) + Na(s) \rightarrow NaOH(aq) + H_2(g)$, including the meanings of the terms in italics.

4. What are the meanings of the following symbols in a chemical equation: $\rightarrow$, $\leftarrow\rightarrow$, $\uparrow$, $\downarrow$, Δ, (*s*), (*l*), (*g*), (*aq*)?

Section 5.3. Balancing Chemical Equations

5. A typical word statement of a chemical reaction is, "Ammonia reacts with sulfuric acid to yield ammonium sulfate." The compounds or the ions involved in them have been given earlier in this book. Using correct chemical formulas, show the unbalanced and balanced chemical equations for this reaction.

6. What is wrong with balancing $Ca + O_2 \rightarrow CaO$ as $Ca + O_2 \rightarrow CaO_2$?

7. Iron (II) sulfate dissolved in acid mine water, a pollutant from coal mines, reacts with oxygen from air according to the unbalanced equation $FeSO_4 + H_2SO_4 + O_2 \rightarrow Fe_3(SO_4)_3 + H_2O$. Examination of this equation reveals two things about groups of atoms that simplify the balancing of the equation. What are these two factors, and how can awareness of them help to balance the equation?

8. Balance the equation from Question 6.

9. Balance each of the following: (a) $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$, (b) $KClO_3 \rightarrow KCl + O_2$, (c) $Ag_2SO_4 + BaI_2 \rightarrow AgI + BaSO_4$, (d) $KClO_4 + C_6H_{12}O_6 \rightarrow KCl + CO_2 + H_2O$, (e) $Fe + O_2 \rightarrow Fe_2O_3 + H_2O$, (f) $P + Cl_2 \rightarrow PCl_3$

Section 5.4. Will a Reaction Occur?

10. Having studied Section 5.4, and knowing something about silver metal and its uses, suggest what would happen if a small item of silver jewelry were placed in a solution of hydrochloric acid.

11. From the information given about Reactions 5.4.1 and 5.4.2, suggest a reaction that might occur if a piece of zinc metal were placed in a solution of $CuCl_2$. Explain. If a chemical reaction does occur, write an equation describing the reaction.

Section 5.5 How Fast Does a Reaction Go?

12. Steel wool heated in a flame and thrust into a bottle of oxygen gas burns vigorously, implying that a reaction such as $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ does in fact occur. Why does a steel girder, such as one used for bridge construction, not burn when exposed to air?

13. What does a catalyst do?

Section 5.6. Classification of Chemical Reactions

14. Classify each of the following reactions according to the categories that are given in Section 5.6:

(a) $NaCl(l, \textit{melted liquid}) \xrightarrow[\text{current}]{\text{Electrical}} 2Na(l) + Cl_2(g)$
(b) $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$
(c) $Zn(s) + CuCl_2(aq) \rightarrow Cu(s) + ZnCl_2(aq)$
(d) $2NaOH(aq) + H_3PO_4(aq) \rightarrow Na_2HPO_4(aq) + 2H_2O(l)$
(e) $BaCl_2(aq) + Ag_2SO_4(aq) \rightarrow BaSO_4(s) + 2AgCl(s)$
(f) $(NH_4)_2SO_4(s) + Ca(OH)_2(s) \xrightarrow{\Delta} 2NH_3(g) + 2H_2O(g) + CaSO_4(s)$

Section 5.7. Quantitative Information from Chemical Reactions

15. Define (a) molecular mass (formula mass when the formula unit is a molecule), (b) mole, (c) molar mass.

16. What are the three main entities that can compose a formula unit of a substance?

17. A total of 7.52 g of $A1Cl_3$ contains (a) __________ moles and (b) ____________ formula units of the compound.

18. A total of 336 g of methane, CH_4, contains (a) __________________ moles and (b) _________________ molecules of the compound.

Section 5.8. What is Stoichiometry and Why is it Important?

19. Ammonium nitrate, NH_4NO_3, can be made by a reaction between NH_3 and HNO_3. A “recipe” for the manufacture of ammonium nitrate specifies mixing 270 kg of NH_3 with 1000 kg of HNO_3. What is the significance of the relative quantities of these two ingredients?

20. What are some reasons that stoichiometric ratios of reactants are used for many industrial processes? In what cases are stoichiometric ratios not used?

21. What is the relationship constituting the basis of stoichiometry between the masses of the reactants and products in a chemical reaction?

22. In addition to the law of conservation of mass, another important stoichiometric relationship involves the proportions in mass of the reaction participants. What is this relationship?

23. Briefly define and explain the mole ratio method of stoichiometric calculations.

24. Carbon disulfide, CS_2, burns rapidly in air because of the reaction with oxygen. What is the mole ratio of O_2 to CS_2 in this reaction?

25. What is the first step required to solve a problem by the mole ratio method?

26. What is the mass (g) of NO_2 produced by the reaction of 10.44 g of O_2 with NO, yielding NO_2?

27. Plants utilize light energy in the photosynthesis process to synthesize glucose, $C_6H_{12}O_6$, from CO_2 and H_2O by way of the reaction $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$. How many grams of CO_2 are consumed in the production of 90 g of glucose?

28. Bacteria in water utilize organic material as a food source, consuming oxygen in a process called respiration. If glucose sugar is the energy source, the reaction is $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ (chemically, the reverse of the photosynthesis reaction in the preceding problem). At 25°C the maximum amount of oxygen in 1.00 liter of water due to dissolved air is 8.32 mg. What is the mass of glucose in mg that will cause all the oxygen in 1.50 L of water to be used up by bacterial respiration?

29. In a blast furnace, the overall reaction by which carbon in coke is used to produce iron metal from iron ore is $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$. How many tons of C are required to produce 100 tons of iron metal?

30. The gravimetric chemical analysis of NaC1 may be carried out by precipitating dissolved chloride from solution by the following reaction with silver nitrate solution: $AgNO_3(aq) + NaCl \rightarrow AgCl(s) + NaNO_3(aq)$ If this reaction produced 1.225 g of AgCl, what was the mass of NaCl?

31. For the reaction of 100.0 g of NaOH with Cl_2, $2NaOH + Cl_2 \rightarrow NaClO + NaCl + H_2O$, give the masses of each of the products.

32. Hydrochloric acid (HCl gas dissolved in water) reacts with calcium carbonate in a piece of limestone as follows: $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$. If 14.6 g of CO_2 are produced in this reaction, what is the total mass of reactants and the total mass of the products?.

33. Given the reaction in the preceding problem, how many moles of HC1 are required to react with 0.618 moles of $CaCO_3$?

34. Given the reaction in Problem 30, how many moles of $CaCO_3$ are required to produce 100.0 g of $CaCl_2$?

35. Silicon tetrachloride, $SiCl_4$, is used to make organosilicon compounds (silicones) and produces an excellent smokescreen for military operations. In the latter application, the $SiCl_4$ reacts with atmospheric moisture (water), $SiCl_4(g) + 2H_2O(g) \rightarrow SiO_2(s) + 4HCl(g)$ to form particles of silicon dioxide and hydrogen chloride gas. The HC1 extracts additional moisture from the atmosphere to produce droplets of hydrochloric acid which, along with small particles of SiO_2, constitute the "smoke" in the smokescreen. Air in a smokescreen was sampled by drawing 100 m^3 of the air through water to collect HCl and to cause any unreacted $SiCl_4$ to react according to the above reaction. After the sampling was completed, the water was found to contain 1.85 g of HC1. What was the original concentration of $SiCl_4$ in the atmosphere (before any of the above reactions occurred) in units of milligrams of $SiCl_4$ per cubic meter?

36. Match each reaction below with the type of reaction

A. Combination (addition)	1. $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
B. Decomposition	2. $Pb(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow PbSO_4(s) + 2HNO_3(aq)$
C. Metathetical (double displacement)	3. $CuSO_4 \cdot 5H_2O(s) \xrightarrow{\Delta} CuSO_4(s) + 5H_2O(g)$
D. Neutralization	4. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

37. Given atomic masses of Al 27 and O 16.0, the mass of O_2 consumed and the mass of Al_2O_3 produced when 29.5 g of Al undergo the reaction $4Al + 3O_2 \rightarrow 2Al_2O_3$ are

A. 26.2 g O_2 and 55.7 g Al_2O_3
B. 23.4 g O_2 and 49.7 g Al_2O_3
C. 28.7 g O_2 and 64.7 g Al_2O_3
D. 30.6 g O_2 and 65.1 g Al_2O_3
E. 20.3 g O_2 and49.8 g Al_2O_3

38. Given atomic masses H 1.0, S, 32.0, O 16.0, the mass of SO_2 that must react to yield 46.8 g of S by the reaction $2H_2S + SO_2 \rightarrow 3S + 2H_2O$ is

A. 35.1 g SO_2
B. 19.1 g SO_2
C. 31.2 g SO_2
D. 40.6 g SO_2
E. 28.3 g SO_2

39. Given atomic masses H 1.0, S, 32.0, O 16.0, the number of moles of H^+ ion that can be obtained from 24.0 g of H_2SO_4 is

A. 0.669 mol H^+
B. 0.490 mol H^+
C. 1.22 mol H^+
D. 0.555 mol H^+
E. 0.380 mol H^+

40. Of the following, mark the statement that is **incorrect** regarding the equation below:

$$CaCO_3(s) + H_2SO_4(aq) \xrightarrow{\Delta} CaSO_4(s) + CO_2(g) + H_2O(l)$$

A. $CaSO_4$ is dissolved in water.
B. $CaCO_3$ is a reactant.
C. CO_2 is a gas product.
D. The reaction mixture was heated.
E. H_2SO_4 was dissolved in water.

41. Given atomic masses of H 1.0, Fe 55.8, and O 16.0, the number of *moles* of CO_2 produced when 79.8 g of Fe_2O_3 undergo the reaction $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$ are

A. 2.56 mol of CO_2
B. 1.36 mol of CO_2
C. 1.50 mol of CO_2
D. 1.14 mol of CO_2
E. 1.62 mol of CO_2

42. Consider the combustion of tetraethyllead, $Pb(C_2H_5)_4$, formerly used as a gasoline octane booster, that burns in the presence of O_2 to give PbO,H_2O, and CO_2. What is the balanced chemical equation for this reaction? Assume that 1.00 g of tetraethyllead was burned. What mole ratio would be used to calculate the mass of H_2O produced? What mass of H_2O was produced?