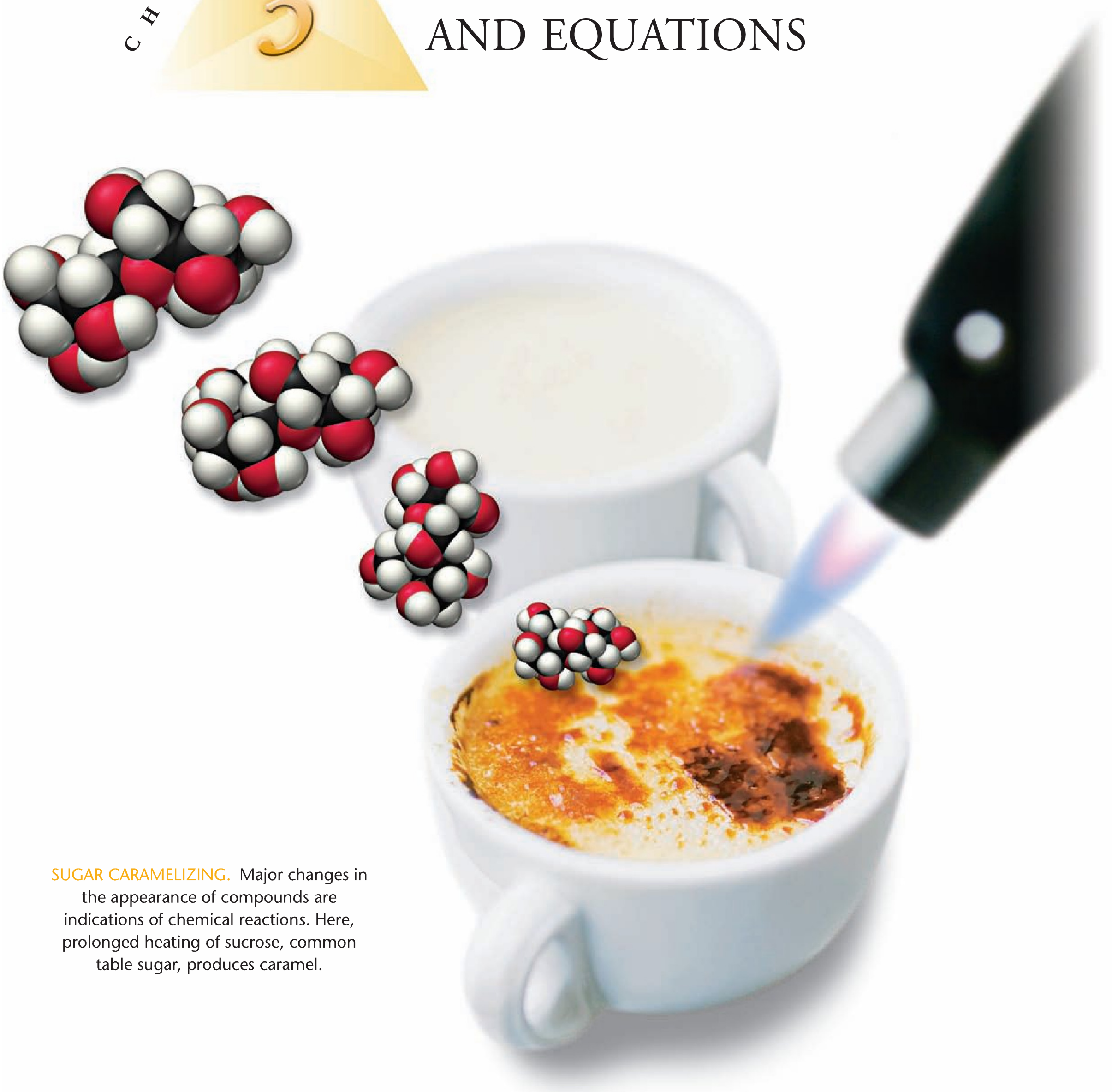
C H A P T E R 3

STOICHIOMETRY: CALCULATIONS WITH CHEMICAL FORMULAS AND EQUATIONS



SUGAR CARAMELIZING. Major changes in the appearance of compounds are indications of chemical reactions. Here, prolonged heating of sucrose, common table sugar, produces caramel.

WHAT'S AHEAD

3.1 Chemical Equations
We begin by considering how we can use chemical formulas to write equations that represent chemical reactions.

3.2 Some Simple Patterns of Chemical Reactivity
We then examine some simple chemical reactions: *combination reactions*, *decomposition reactions*, and *combustion reactions*.

3.3 Formula Weights
We can obtain quantitative information from chemical formulas by using their *formula weights*.

3.4 Avogadro's Number and the Mole
We use chemical formulas to relate the masses of substances to the numbers of atoms, molecules, or ions contained in the substances, a relationship that leads to the crucially important concept of a *mole*. A *mole* is 6.022×10^{23} objects (atoms, molecules, ions, etc.).

3.5 Empirical Formulas from Analyses
We apply the mole concept to determine chemical formulas from the masses of each element in a given quantity of a compound.

3.6 Quantitative Information from Balanced Equations
We use the quantitative information inherent in chemical formulas and equations together with the mole concept to predict the amounts of substances consumed or produced in chemical reactions.

3.7 Limiting Reactants
We recognize that one of the reactants may be used up before the others in a chemical reaction. This is the *limiting reactant*. The reaction therefore stops, leaving some of the excess starting material unreacted.

YOU POUR VINEGAR INTO A glass of water containing baking soda, and bubbles form. You strike a match and use the flame to light a candle. You heat sugar in a pan, and it turns brown (caramelizes). The bubbles, the flame, and the color change are visual evidence that something is happening.

To an experienced eye, these visual changes indicate a chemical change, or chemical reaction. The study of chemical changes is at the heart of chemistry. Some chemical changes are simple; others are complex. Some are dramatic; some are very subtle. Even as you sit reading this chapter, chemical changes are occurring within your body. Chemical changes that occur in your eyes and brain, for example, allow you to see these words and think about them. Although such chemical changes are not as obvious as some, they are nevertheless remarkable for how they allow us to function.

In this chapter we begin to explore some important aspects of chemical change. Our focus will be both on the use of chemical formulas to represent reactions and on the quantitative information we can obtain about the amounts of substances involved in reactions. **Stoichiometry** (pronounced stoy-key-OM-uh-tree) is the area of study that examines the quantities of substances consumed and produced in chemical reactions. The name is derived from the Greek *stoicheion* ("element") and *metron* ("measure"). This study of stoichiometry provides an essential set of tools that is widely used in chemistry. Aspects of stoichiometry include such diverse problems as measuring the concentration of ozone in the atmosphere, determining the potential yield of gold from an ore, and assessing different processes for converting coal into gaseous fuels.

Stoichiometry is built on an understanding of atomic masses (Section 2.4), chemical formulas, and the law of conservation of mass. (Section 2.1)

▲ **Figure 3.1 Antoine Lavoisier (1734–1794).** Lavoisier conducted many important studies on combustion reactions. Unfortunately, the French Revolution cut his career short. He was a member of the French nobility and a tax collector. He was guillotined in 1794 during the final months of the Reign of Terror. He is now generally considered to be the father of modern chemistry because he conducted carefully controlled experiments and used quantitative measurements.

The French nobleman and scientist Antoine Lavoisier (Figure 3.1◀) discovered this important chemical law during the late 1700s. In a chemistry text published in 1789, Lavoisier stated the law in this eloquent way: "We may lay it down as an incontestable axiom that, in all the operations of art and nature, nothing is created; an equal quantity of matter exists both before and after the experiment. Upon this principle, the whole art of performing chemical experiments depends." With the advent of Dalton's atomic theory, chemists understood the basis for this law: *Atoms are neither created nor destroyed during any chemical reaction*. The changes that occur during any reaction merely rearrange the atoms. The same collection of atoms is present both before and after the reaction.

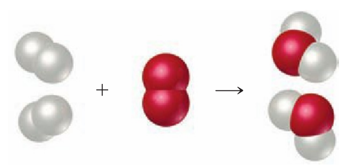
3.1 CHEMICAL EQUATIONS

Chemical reactions are represented in a concise way by **chemical equations**. When the gas hydrogen (H_2) burns, for example, it reacts with oxygen (O_2) in the air to form water (H_2O). We write the chemical equation for this reaction as follows:

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O \qquad [3.1]$$

We read the + sign as "reacts with" and the arrow as "produces." The chemical formulas to the left of the arrow represent the starting substances, called **reactants**. The chemical formulas to the right of the arrow represent substances produced in the reaction, called **products**. The numbers in front of the formulas are *coefficients*. (As in algebraic equations, the numeral 1 is usually not written.) The coefficients indicate the relative numbers of molecules of each kind involved in the reaction.

Because atoms are neither created nor destroyed in any reaction, a chemical equation must have an equal number of atoms of each element on each side of the arrow. When this condition is met, the equation is said to be *balanced*. On the right side of Equation 3.1, for example, there are two molecules of H_2O, each composed of two atoms of hydrogen and one atom of oxygen. Thus, $2\,H_2O$ (read "two molecules of water") contains $2 \times 2 = 4$ H atoms and $2 \times 1 = 2$ O atoms. Notice that the number of atoms is obtained by multiplying the coefficient and the subscripts in the chemical formula. Because there are four H atoms and two O atoms on each side of the equation, the equation is balanced. We can represent the balanced equation by the following molecular models, which illustrate that the number of atoms of each kind is the same on both sides of the arrow:



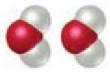
GIVE IT SOME THOUGHT

How many atoms of Mg, O, and H are represented by $3\,Mg(OH)_2$?

Balancing Equations

Once we know the formulas of the reactants and products in a reaction, we can write an unbalanced equation. We then balance the equation by determining the coefficients that provide equal numbers of each type of atom on each side of the equation. For most purposes, a balanced equation should contain the smallest possible whole-number coefficients.

In balancing an equation, you need to understand the difference between a coefficient in front of a formula and a subscript within a formula. Refer to Figure 3.2▶. Notice that changing a subscript in a formula—from H_2O to H_2O_2,

Chemical symbol	Meaning		Composition
H_2O	One molecule of water:		Two H atoms and one O atom
$2\ H_2O$	Two molecules of water:		Four H atoms and two O atoms
H_2O_2	One molecule of hydrogen peroxide:		Two H atoms and two O atoms

◀ **Figure 3.2 The difference between a subscript and a coefficient.** Notice how adding the coefficient 2 in front of the formula (line 2) has a different effect on the implied composition than adding the subscript 2 to the formula (in line 3). The number of atoms of each type (listed under composition) is obtained by multiplying the coefficient and the subscript associated with each element in the formula.

for example—changes the identity of the chemical. The substance H_2O_2, hydrogen peroxide, is quite different from the substance H_2O, water. *You should never change subscripts when balancing an equation*. In contrast, placing a coefficient in front of a formula changes only the *amount* of the substance and not its *identity*. Thus, 2 H_2O means two molecules of water, 3 H_2O means three molecules of water, and so forth.

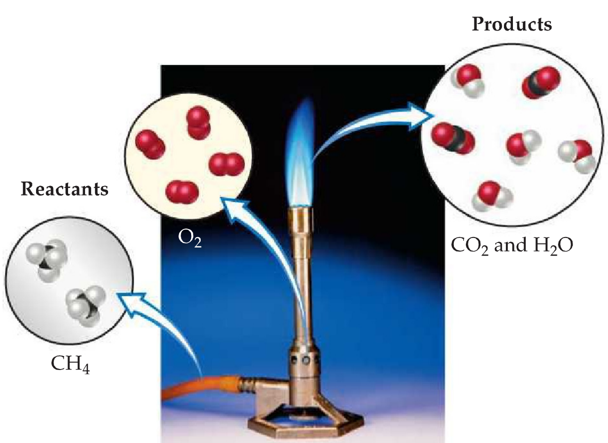
To illustrate the process of balancing an equation, consider the reaction that occurs when methane (CH_4), the principal component of natural gas, burns in air to produce carbon dioxide gas (CO_2) and water vapor (H_2O) (Figure 3.3▼). Both of these products contain oxygen atoms that come from O_2 in the air. Thus, O_2 is a reactant, and the unbalanced equation is

$$CH_4 + O_2 \longrightarrow CO_2 + H_2O \quad \text{(unbalanced)} \qquad [3.2]$$

It is usually best to balance first those elements that occur in the fewest chemical formulas on each side of the equation. In our example both C and H appear in only one reactant and, separately, in one product each. So we begin by focusing on CH_4. Let's consider first carbon and then hydrogen.

One molecule of the reactant CH_4 contains the same number of C atoms (one) as one molecule of the product CO_2. The coefficients for these substances *must* be the same, therefore, we start the balancing process by choosing the coefficient one for each. However, one molecule of CH_4 contains more H atoms (four) than one molecule of the product H_2O (two). If we place a coefficient 2 in front of H_2O, there will be four H atoms on each side of the equation:

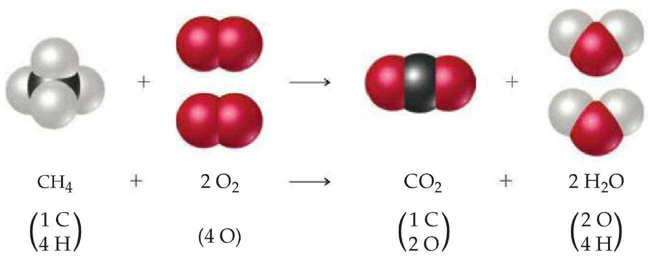
$$CH_4 + O_2 \longrightarrow CO_2 + 2\ H_2O \quad \text{(unbalanced)} \qquad [3.3]$$



◀ **Figure 3.3 Methane reacts with oxygen to produce the flame in a Bunsen burner.** The methane (CH_4) in natural gas and oxygen (O_2) from the air are the reactants in the reaction, while carbon dioxide (CO_2) and water vapor (H_2O) are the products.

At this stage the products have more O atoms (four—two from CO_2 and two from 2 H_2O) than the reactants (two). If we place the coefficient 2 in front of the reactant O_2, we balance the equation by making the number of O atoms equal on both sides of the equation:

$$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O \quad \text{(balanced)} \qquad [3.4]$$



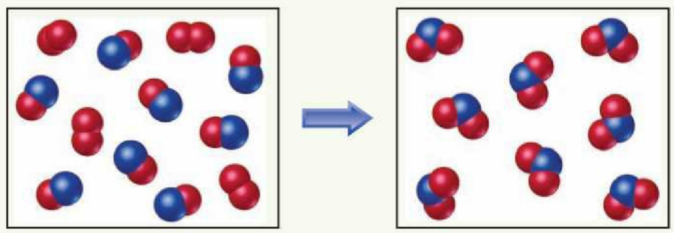
▲ **Figure 3.4 Balanced chemical equation for the combustion of CH_4.** The drawings of the molecules involved call attention to the conservation of atoms through the reaction.

The molecular view of the balanced equation is shown in Figure 3.4 ◀. We see one C, four H, and four O atoms on each side of the arrow, indicating that the equation is balanced.

The approach we have taken in arriving at balanced Equation 3.4 is largely trial and error. We balance each kind of atom in succession, adjusting coefficients as necessary. This approach works for most chemical equations.

SAMPLE EXERCISE 3.1 | Interpreting and Balancing Chemical Equations

The following diagram represents a chemical reaction in which the red spheres are oxygen atoms and the blue spheres are nitrogen atoms. (a) Write the chemical formulas for the reactants and products. (b) Write a balanced equation for the reaction. (c) Is the diagram consistent with the law of conservation of mass?



SOLUTION

(a) The left box, which represents the reactants, contains two kinds of molecules, those composed of two oxygen atoms (O_2) and those composed of one nitrogen atom and one oxygen atom (NO). The right box, which represents the products, contains only molecules composed of one nitrogen atom and two oxygen atoms (NO_2).

(b) The unbalanced chemical equation is

$$O_2 + NO \longrightarrow NO_2 \quad \text{(unbalanced)}$$

This equation has three O atoms on the left side of the arrow and two O atoms on the right side. We can increase the number of O atoms by placing a coefficient 2 on the product side:

$$O_2 + NO \longrightarrow 2\,NO_2 \quad \text{(unbalanced)}$$

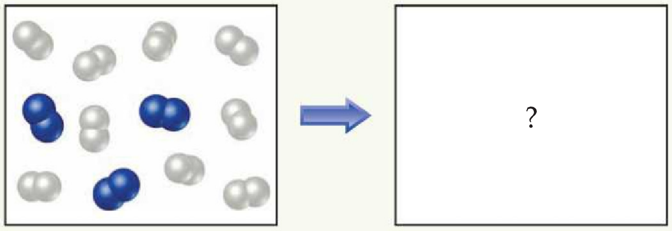
Now there are two N atoms and four O atoms on the right. Placing the coefficient 2 in front of NO balances both the number of N atoms and O atoms:

$$O_2 + 2\,NO \longrightarrow 2\,NO_2 \quad \text{(balanced)}$$

(c) The left box (reactants) contains four O_2 molecules and eight NO molecules. Thus, the molecular ratio is one O_2 for each two NO as required by the balanced equation. The right box (products) contains eight NO_2 molecules. The number of NO_2 molecules on the right equals the number of NO molecules on the left as the balanced equation requires. Counting the atoms, we find eight N atoms in the eight NO molecules in the box on the left. There are also $4 \times 2 = 8$ O atoms in the O_2 molecules and eight O atoms in the NO molecules, giving a total of 16 O atoms. In the box on the right, we find eight N atoms and $8 \times 2 = 16$ O atoms in the eight NO_2 molecules. Because there are equal numbers of both N and O atoms in the two boxes, the drawing is consistent with the law of conservation of mass.

PRACTICE EXERCISE

In the following diagram, the white spheres represent hydrogen atoms, and the blue spheres represent nitrogen atoms. To be consistent with the law of conservation of mass, how many NH_3 molecules should be shown in the right box?



Answer: Six NH_3 molecules

Indicating the States of Reactants and Products

Additional information is often added to the formulas in balanced equations to indicate the physical state of each reactant and product. We use the symbols (g), (l), (s), and (aq) for gas, liquid, solid, and aqueous (water) solution, respectively. Thus, Equation 3.4 can be written

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g) \qquad [3.5]$$

Sometimes the conditions (such as temperature or pressure) under which the reaction proceeds appear above or below the reaction arrow. The symbol Δ (the Greek uppercase letter delta) is often placed above the arrow to indicate the addition of heat.

SAMPLE EXERCISE 3.2 | Balancing Chemical Equations

Balance this equation:

$$Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$$

SOLUTION

Begin by counting each kind of atom on both sides of the arrow. The Na and O atoms are balanced—one Na and one O on each side—but there are two H atoms on the left and three H atoms on the right. Thus, we need to increase the number of H atoms on the left. To begin balancing H, let's try placing the coefficient 2 in front of H_2O:

$$Na(s) + 2\,H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$$

Beginning this way does not balance H but does increase the number of H atoms among the reactants, which we need to do. Adding the coefficient 2 causes O to be unbalanced; we will take care of that after we balance H. Now that we have 2 H_2O on the left, we can balance H by putting the coefficient 2 in front of NaOH on the right:

$$Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

Balancing H in this way fortuitously brings O into balance. But notice that Na is now unbalanced, with one Na on the left and two on the right. To rebalance Na, we put the coefficient 2 in front of the reactant:

$$2\,Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

Finally, we check the number of atoms of each element and find that we have two Na atoms, four H atoms, and two O atoms on each side of the equation. The equation is balanced.

Comment Notice that in balancing this equation, we moved back and forth placing a coefficient in front of H_2O, then NaOH, and finally Na. In balancing equations, we often find ourselves following this pattern of moving back and forth from one side of the arrow to the other, placing coefficients first in front of a formula on one side and then in front of a formula on the other side until the equation is balanced. You can always tell if you have balanced your equation correctly, no matter how you did it, by checking that the number of atoms of each element is the same on both sides of the arrow.

PRACTICE EXERCISE

Balance the following equations by providing the missing coefficients:

(a) $\_\_Fe(s) + \_\_O_2(g) \longrightarrow \_\_Fe_2O_3(s)$

(b) $\_\_C_2H_4(g) + \_\_O_2(g) \longrightarrow \_\_CO_2(g) + \_\_H_2O(g)$

(c) $\_\_Al(s) + \_\_HCl(aq) \longrightarrow \_\_AlCl_3(aq) + \_\_H_2(g)$

Answers: **(a)** 4, 3, 2; **(b)** 1, 3, 2, 2; **(c)** 2, 6, 2, 3

3.2 SOME SIMPLE PATTERNS OF CHEMICAL REACTIVITY

In this section we examine three simple kinds of reactions that we will see frequently throughout this chapter. Our first reason for examining these reactions is merely to become better acquainted with chemical reactions and their balanced equations. Our second reason is to consider how we might predict the products of some of these reactions knowing only their reactants. The key to predicting the products formed by a given combination of reactants is recognizing general patterns of chemical reactivity. Recognizing a pattern of reactivity for a class of substances gives you a broader understanding than merely memorizing a large number of unrelated reactions.

Combination and Decomposition Reactions

Table 3.1▼ summarizes two simple types of reactions: combination and decomposition reactions. In **combination reactions** two or more substances react to form one product. There are many examples of combination reactions, especially those in which elements combine to form compounds. For example, magnesium metal burns in air with a dazzling brilliance to produce magnesium oxide, as shown in Figure 3.5▶:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s) \qquad [3.6]$$

This reaction is used to produce the bright flame generated by flares and some fireworks.

When a combination reaction occurs between a metal and a nonmetal, as in Equation 3.6, the product is an ionic solid. Recall that the formula of an ionic compound can be determined from the charges of the ions involved. (Section 2.7) When magnesium reacts with oxygen, for example, the magnesium loses electrons and forms the magnesium ion, Mg^{2+}. The oxygen gains electrons and forms the oxide ion, O^{2-}. Thus, the reaction product is MgO. You should be able to recognize when a reaction is a combination reaction and to predict the products of a combination reaction in which the reactants are a metal and a nonmetal.

GIVE IT SOME THOUGHT

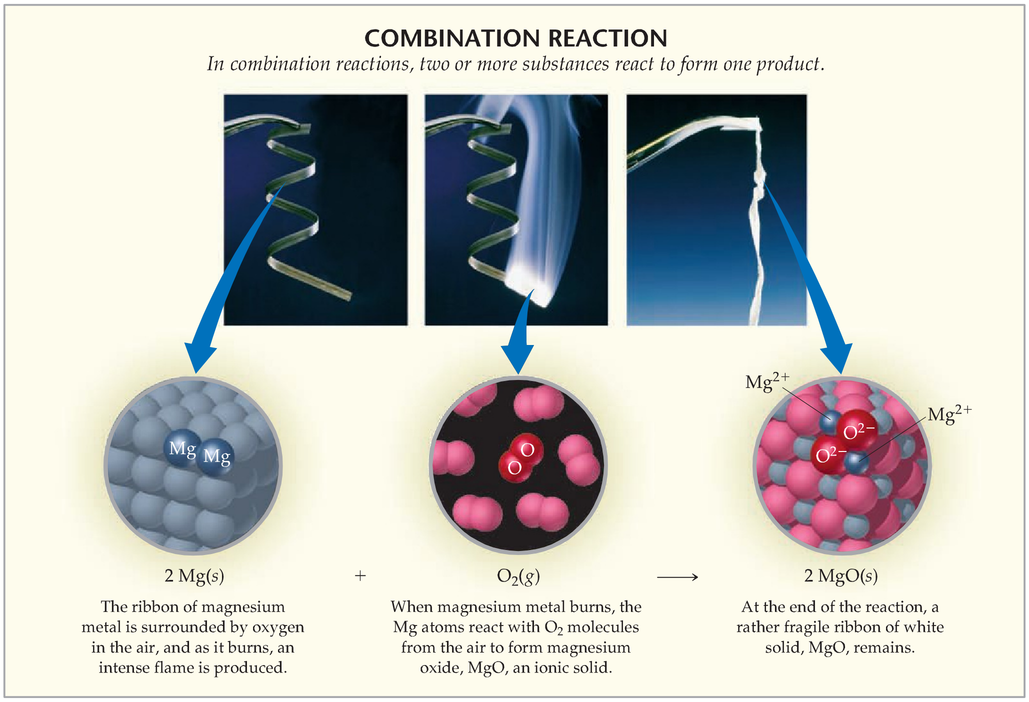
When Na and S undergo a combination reaction, what is the chemical formula of the product?

In a **decomposition reaction** one substance undergoes a reaction to produce two or more other substances. Many compounds undergo decomposition reactions when heated. For example, many metal carbonates decompose to form metal oxides and carbon dioxide when heated:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g) \qquad [3.7]$$

TABLE 3.1 ■ Combination and Decomposition Reactions

Combination Reactions	
$A + B \longrightarrow C$ $C(s) + O_2(g) \longrightarrow CO_2(g)$ $N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$ $CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(s)$	Two reactants combine to form a single product. Many elements react with one another in this fashion to form compounds.
Decomposition Reactions	
$C \longrightarrow A + B$ $2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g)$ $PbCO_3(s) \longrightarrow PbO(s) + CO_2(g)$ $Cu(OH)_2(s) \longrightarrow CuO(s) + H_2O(l)$	A single reactant breaks apart to form two or more substances. Many compounds react this way when heated.



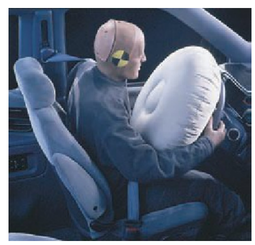
▲ **Figure 3.5 Combustion of magnesium metal in air.**

The decomposition of $CaCO_3$ is an important commercial process. Limestone or seashells, which are both primarily $CaCO_3$, are heated to prepare CaO, which is known as lime or quicklime. About 2×10^{10} kg (20 million tons) of CaO is used in the United States each year, principally in making glass, in obtaining iron from its ores, and in making mortar to bind bricks.

The decomposition of sodium azide (NaN_3) rapidly releases $N_2(g)$, so this reaction is used to inflate safety air bags in automobiles (Figure 3.6▶):

$$2\,NaN_3(s) \longrightarrow 2\,Na(s) + 3\,N_2(g) \qquad [3.8]$$

The system is designed so that an impact ignites a detonator cap, which in turn causes NaN_3 to decompose explosively. A small quantity of NaN_3 (about 100 g) forms a large quantity of gas (about 50 L). We will consider the volumes of gases produced in chemical reactions in Section 10.5.



▲ **Figure 3.6 An automobile air bag.** The decomposition of sodium azide, $NaN_3(s)$, is used to inflate automobile air bags. When properly ignited, the NaN_3 decomposes rapidly, forming nitrogen gas, $N_2(g)$, which expands the air bag.

SAMPLE EXERCISE 3.3 | Writing Balanced Equations for Combination and Decomposition Reactions

Write balanced equations for the following reactions: **(a)** The combination reaction that occurs when lithium metal and fluorine gas react. **(b)** The decomposition reaction that occurs when solid barium carbonate is heated. (Two products form: a solid and a gas.)

SOLUTION

(a) The symbol for lithium is Li. With the exception of mercury, all metals are solids at room temperature. Fluorine occurs as a diatomic molecule (see Figure 2.19). Thus, the reactants are Li(s) and $F_2(g)$. The product will be composed of a metal and a nonmetal, so we expect it to be an ionic solid. Lithium ions have a 1+ charge, Li^+, whereas fluoride ions have a 1− charge, F^-. Thus, the chemical formula for the product is LiF. The balanced chemical equation is

$$2\,Li(s) + F_2(g) \longrightarrow 2\,LiF(s)$$

(b) The chemical formula for barium carbonate is $BaCO_3$. As noted in the text, many metal carbonates decompose to form metal oxides and carbon dioxide when heated. In Equation 3.7, for example, $CaCO_3$ decomposes to form CaO and CO_2. Thus, we would expect that $BaCO_3$ decomposes to form BaO and CO_2. Barium and calcium are both in group 2A in the periodic table, which further suggests they would react in the same way:

$$BaCO_3(s) \longrightarrow BaO(s) + CO_2(g)$$

PRACTICE EXERCISE

Write balanced chemical equations for the following reactions: **(a)** Solid mercury(II) sulfide decomposes into its component elements when heated. **(b)** The surface of aluminum metal undergoes a combination reaction with oxygen in the air.
Answers: **(a)** $HgS(s) \longrightarrow Hg(l) + S(s)$; **(b)** $4\,Al(s) + 3\,O_2(g) \longrightarrow 2\,Al_2O_3(s)$

Combustion in Air

Combustion reactions are rapid reactions that produce a flame. Most of the combustion reactions we observe involve O_2 from air as a reactant. Equation 3.5 illustrates a general class of reactions that involve the burning, or combustion, of hydrocarbon compounds (compounds that contain only carbon and hydrogen, such as CH_4 and C_2H_4). (Section 2.9)

When hydrocarbons are combusted in air, they react with O_2 to form CO_2 and H_2O.* The number of molecules of O_2 required in the reaction and the number of molecules of CO_2 and H_2O formed depend on the composition of the hydrocarbon, which acts as the fuel in the reaction. For example, the combustion of propane (C_3H_8), a gas used for cooking and home heating, is described by the following equation:

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g) \qquad [3.9]$$

▲ **Figure 3.7 Propane burning in air.** The liquid propane, C_3H_8, vaporizes and mixes with air as it escapes through the nozzle. The combustion reaction of C_3H_8 and O_2 produces a blue flame.

The state of the water, $H_2O(g)$ or $H_2O(l)$, depends on the conditions of the reaction. Water vapor, $H_2O(g)$, is formed at high temperature in an open container. The blue flame produced when propane burns is shown in Figure 3.7◄.

Combustion of oxygen-containing derivatives of hydrocarbons, such as CH_3OH, also produces CO_2 and H_2O. The simple rule that hydrocarbons and related oxygen-containing derivatives of hydrocarbons form CO_2 and H_2O when they burn in air summarizes the behavior of about 3 million compounds. Many substances that our bodies use as energy sources, such as the sugar glucose ($C_6H_{12}O_6$), similarly react in our bodies with O_2 to form CO_2 and H_2O. In our bodies, however, the reactions take place in a series of intermediate steps that occur at body temperature. These reactions that involve intermediate steps are described as *oxidation reactions* instead of combustion reactions.

SAMPLE EXERCISE 3.4 | Writing Balanced Equations for Combustion Reactions

Write the balanced equation for the reaction that occurs when methanol, $CH_3OH(l)$, is burned in air.

SOLUTION

When any compound containing C, H, and O is combusted, it reacts with the $O_2(g)$ in air to produce $CO_2(g)$ and $H_2O(g)$. Thus, the unbalanced equation is

$$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

**When there is an insufficient quantity of O_2 present, carbon monoxide (CO) will be produced along with the CO_2; this is called incomplete combustion. If the amount of O_2 is severely restricted, fine particles of carbon that we call soot will be produced. Complete combustion produces only CO_2 and H_2O. Unless specifically stated to the contrary, we will always take combustion to mean complete combustion.*

In this equation the C atoms are balanced with one carbon on each side of the arrow. Because CH_3OH has four H atoms, we place the coefficient 2 in front of H_2O to balance the H atoms:

$$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Adding the coefficient balances H but gives four O atoms in the products. Because there are only three O atoms in the reactants (one in CH_3OH and two in O_2), we are not finished yet. We can place the fractional coefficient $\frac{3}{2}$ in front of O_2 to give a total of four O atoms in the reactants (there are $\frac{3}{2} \times 2 = 3$ O atoms in $\frac{3}{2}\,O_2$):

$$CH_3OH(l) + \tfrac{3}{2}\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Although the equation is now balanced, it is not in its most conventional form because it contains a fractional coefficient. If we multiply each side of the equation by 2, we will remove the fraction and achieve the following balanced equation:

$$2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

PRACTICE EXERCISE

Write the balanced equation for the reaction that occurs when ethanol, $C_2H_5OH(l)$, is burned in air.

Answer: $C_2H_5OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$

3.3 FORMULA WEIGHTS

Chemical formulas and chemical equations both have a *quantitative* significance; the subscripts in formulas and the coefficients in equations represent precise quantities. The formula H_2O indicates that a molecule of this substance (water) contains exactly two atoms of hydrogen and one atom of oxygen. Similarly, the coefficients in a balanced chemical equation indicate the relative quantities of reactants and products. But how do we relate the numbers of atoms or molecules to the amounts we measure in the laboratory? Although we cannot directly count atoms or molecules, we can indirectly determine their numbers if we know their masses. Therefore, before we can pursue the quantitative aspects of chemical formulas or equations, we must examine the masses of atoms and molecules, which we do in this section and the next.

Formula and Molecular Weights

The **formula weight** of a substance is the sum of the atomic weights of each atom in its chemical formula. Using atomic masses from a periodic table, we find, for example, that the formula weight of sulfuric acid (H_2SO_4) is 98.1 amu:*

$$\begin{aligned}\text{FW of } H_2SO_4 &= 2(\text{AW of H}) + (\text{AW of S}) + 4(\text{AW of O})\\ &= 2(1.0\text{ amu}) + 32.1\text{ amu} + 4(16.0\text{ amu})\\ &= 98.1\text{ amu}\end{aligned}$$

For convenience, we have rounded off all the atomic weights to one place beyond the decimal point. We will round off the atomic weights in this way for most problems.

If the chemical formula is merely the chemical symbol of an element, such as Na, then the formula weight equals the atomic weight of the element, in this case 23.0 amu. If the chemical formula is that of a molecule, then the formula weight is also called the **molecular weight**. The molecular weight of glucose ($C_6H_{12}O_6$), for example, is

$$\text{MW of } C_6H_{12}O_6 = 6(12.0\text{ amu}) + 12(1.0\text{ amu}) + 6(16.0\text{ amu}) = 180.0\text{ amu}$$

Because ionic substances, such as NaCl, exist as three-dimensional arrays of ions (Figure 2.23), it is inappropriate to speak of molecules of NaCl. Instead,

**The abbreviation AW is used for atomic weight, FW for formula weight, and MW for molecular weight.*

we speak of *formula units*, represented by the chemical formula of the substance. The formula unit of NaCl consists of one Na^+ ion and one Cl^- ion. Thus, the formula weight of NaCl is the mass of one formula unit:

$$\text{FW of NaCl} = 23.0 \text{ amu} + 35.5 \text{ amu} = 58.5 \text{ amu}$$

SAMPLE EXERCISE 3.5 | **Calculating Formula Weights**

Calculate the formula weight of **(a)** sucrose, $C_{12}H_{22}O_{11}$ (table sugar), and **(b)** calcium nitrate, $Ca(NO_3)_2$.

SOLUTION

(a) By adding the atomic weights of the atoms in sucrose, we find the formula weight to be 342.0 amu:

$$\begin{aligned} 12 \text{ C atoms} &= 12(12.0 \text{ amu}) = 144.0 \text{ amu} \\ 22 \text{ H atoms} &= 22(1.0 \text{ amu}) = 22.0 \text{ amu} \\ 11 \text{ O atoms} &= 11(16.0 \text{ amu}) = \underline{176.0 \text{ amu}} \\ & \qquad\qquad\qquad\quad 342.0 \text{ amu} \end{aligned}$$

(b) If a chemical formula has parentheses, the subscript outside the parentheses is a multiplier for all atoms inside. Thus, for $Ca(NO_3)_2$, we have

$$\begin{aligned} 1 \text{ Ca atom} &= 1(40.1 \text{ amu}) = 40.1 \text{ amu} \\ 2 \text{ N atoms} &= 2(14.0 \text{ amu}) = 28.0 \text{ amu} \\ 6 \text{ O atoms} &= 6(16.0 \text{ amu}) = \underline{96.0 \text{ amu}} \\ & \qquad\qquad\qquad\quad 164.1 \text{ amu} \end{aligned}$$

PRACTICE EXERCISE

Calculate the formula weight of **(a)** $Al(OH)_3$ and **(b)** CH_3OH.
Answers: **(a)** 78.0 amu, **(b)** 32.0 amu

Percentage Composition from Formulas

Occasionally we must calculate the *percentage composition* of a compound—that is, the percentage by mass contributed by each element in the substance. For example, to verify the purity of a compound, we can compare the calculated percentage composition of the substance with that found experimentally. Forensic chemists, for example, will measure the percentage composition of an unknown white powder and compare it to the percentage compositions for sugar, salt, or cocaine to identify the powder. Calculating percentage composition is a straightforward matter if the chemical formula is known. The calculation depends on the formula weight of the substance, the atomic weight of the element of interest, and the number of atoms of that element in the chemical formula:

$$\% \text{ element} = \frac{\begin{pmatrix}\text{number of atoms}\\ \text{of that element}\end{pmatrix}\begin{pmatrix}\text{atomic weight}\\ \text{of element}\end{pmatrix}}{\text{formula weight of compound}} \times 100\% \qquad [3.10]$$

SAMPLE EXERCISE 3.6 | **Calculating Percentage Composition**

Calculate the percentage of carbon, hydrogen, and oxygen (by mass) in $C_{12}H_{22}O_{11}$.

SOLUTION

Let's examine this question using the problem-solving steps in the "Strategies in Chemistry: Problem Solving" essay that appears on the next page.

Analyze We are given a chemical formula, $C_{12}H_{22}O_{11}$, and asked to calculate the percentage by mass of its component elements (C, H, and O).

Plan We can use Equation 3.10, relying on a periodic table to obtain the atomic weight of each component element. The atomic weights are first used to determine the formula weight of the compound. (The formula weight of $C_{12}H_{22}O_{11}$, 342.0 amu, was calculated in Sample Exercise 3.5.) We must then do three calculations, one for each element.

Solve Using Equation 3.10, we have

$$\%\text{C} = \frac{(12)(12.0 \text{ amu})}{342.0 \text{ amu}} \times 100\% = 42.1\%$$

$$\%\text{H} = \frac{(22)(1.0 \text{ amu})}{342.0 \text{ amu}} \times 100\% = 6.4\%$$

$$\%O = \frac{(11)(16.0 \text{ amu})}{342.0 \text{ amu}} \times 100\% = 51.5\%$$

Check The percentages of the individual elements must add up to 100%, which they do in this case. We could have used more significant figures for our atomic weights, giving more significant figures for our percentage composition, but we have adhered to our suggested guideline of rounding atomic weights to one digit beyond the decimal point.

PRACTICE EXERCISE

Calculate the percentage of nitrogen, by mass, in $Ca(NO_3)_2$.
Answer: 17.1%

3.4 AVOGADRO'S NUMBER AND THE MOLE

Even the smallest samples that we deal with in the laboratory contain enormous numbers of atoms, ions, or molecules. For example, a teaspoon of water (about 5 mL) contains 2×10^{23} water molecules, a number so large that it almost defies comprehension. Chemists, therefore, have devised a special counting unit for describing such large numbers of atoms or molecules.

In everyday life we use counting units such as a dozen (12 objects) and a gross (144 objects) to deal with modestly large quantities. In chemistry the unit for dealing with the number of atoms, ions, or molecules in a common-sized sample is the **mole**, abbreviated mol.* A mole is the amount of matter that contains as many objects (atoms, molecules, or whatever objects we are considering) as the number of atoms in exactly 12 g of isotopically pure ^{12}C. From experiments, scientists have determined this number to be 6.0221421×10^{23}. Scientists call this number **Avogadro's number**, in honor of the Italian scientist Amedeo Avogadro (1776–1856). Avogadro's number has the symbol N_A, which we will usually round to 6.02×10^{23} mol^{-1}. The unit mol^{-1} ("inverse mole" or "per mole") reminds us that there are 6.02×10^{23} objects per one mole. A mole of atoms, a mole of molecules, or a mole of anything else all contain Avogadro's number of these objects:

$$1 \text{ mol } ^{12}C \text{ atoms} = 6.02 \times 10^{23}\ ^{12}C \text{ atoms}$$

$$1 \text{ mol } H_2O \text{ molecules} = 6.02 \times 10^{23}\ H_2O \text{ molecules}$$

$$1 \text{ mol } NO_3^- \text{ ions} = 6.02 \times 10^{23}\ NO_3^- \text{ ions}$$

Strategies in Chemistry — PROBLEM SOLVING

Practice is the key to success in problem solving. As you practice, you can improve your skills by following these steps:

Step 1: Analyze the problem. Read the problem carefully for understanding. What does it say? Draw any picture or diagram that will help you to visualize the problem. Write down both the data you are given and the quantity that you need to obtain (the unknown).

Step 2: Develop a plan for solving the problem. Consider the possible paths between the given information and the unknown. What principles or equations relate the known data to the unknown? Recognize that some data may not be given explicitly in the problem; you may be expected to know certain quantities (such as Avogadro's number) or look them up in tables (such as atomic weights). Recognize also that your plan may involve either a single step or a series of steps with intermediate answers.

Step 3: Solve the problem. Use the known information and suitable equations or relationships to solve for the unknown. Dimensional analysis (Section 1.6) is a very useful tool for solving a great number of problems. Be careful with significant figures, signs, and units.

Step 4: Check the solution. Read the problem again to make sure you have found all the solutions asked for in the problem. Does your answer make sense? That is, is the answer outrageously large or small, or is it in the ballpark? Finally, are the units and significant figures correct?

**The term* mole *comes from the Latin word* moles, *meaning "a mass." The term* molecule *is the diminutive form of this word and means "a small mass."*

Avogadro's number is so large that it is difficult to imagine. Spreading 6.02×10^{23} marbles over the entire surface of Earth would produce a continuous layer about 3 miles thick. If Avogadro's number of pennies were placed side by side in a straight line, they would encircle Earth 300 trillion (3×10^{14}) times.

SAMPLE EXERCISE 3.7 | Estimating Numbers of Atoms

Without using a calculator, arrange the following samples in order of increasing numbers of carbon atoms: 12 g ^{12}C, 1 mol C_2H_2, 9×10^{23} molecules of CO_2.

SOLUTION

Analyze We are given amounts of different substances expressed in grams, moles, and number of molecules and asked to arrange the samples in order of increasing numbers of C atoms.

Plan To determine the number of C atoms in each sample, we must convert g ^{12}C, 1 mol C_2H_2, and 9×10^{23} molecules CO_2 all to numbers of C atoms. To make these conversions, we use the definition of mole and Avogadro's number.

Solve A mole is defined as the amount of matter that contains as many units of the matter as there are C atoms in exactly 12 g of ^{12}C. Thus, 12 g of ^{12}C contains 1 mol of C atoms (that is, 6.02×10^{23} C atoms). One mol of C_2H_2 contains 6×10^{23} C_2H_2 molecules. Because there are two C atoms in each C_2H_2 molecule, this sample contains 12×10^{23} C atoms. Because each CO_2 molecule contains one C atom, the sample of CO_2 contains 9×10^{23} C atoms. Hence, the order is 12 g ^{12}C (6×10^{23} C atoms) $<$ 9×10^{23} CO_2 molecules (9×10^{23} C atoms) $<$ 1 mol C_2H_2 (12×10^{23} C atoms).

Check We can check our results by comparing the number of moles of C atoms in each sample because the number of moles is proportional to the number of atoms. Thus, 12 g of ^{12}C is 1 mol C; 1 mol of C_2H_2 contains 2 mol C, and 9×10^{23} molecules of CO_2 contain 1.5 mol C, giving the same order as above: 12 g ^{12}C (1 mol C) $< 9 \times 10^{23}$ CO_2 molecules (1.5 mol C) $<$ 1 mol C_2H_2 (2 mol C).

PRACTICE EXERCISE

Without using a calculator, arrange the following samples in order of increasing number of O atoms: 1 mol H_2O, 1 mol CO_2, 3×10^{23} molecules O_3.
Answer: 1 mol H_2O (6×10^{23} O atoms) $< 3 \times 10^{23}$ molecules O_3 (9×10^{23} O atoms) $<$ 1 mol CO_2 (12×10^{23} O atoms)

SAMPLE EXERCISE 3.8 | Converting Moles to Number of Atoms

Calculate the number of H atoms in 0.350 mol of $C_6H_{12}O_6$.

SOLUTION

Analyze We are given both the amount of a substance (0.350 mol) and its chemical formula ($C_6H_{12}O_6$). The unknown is the number of H atoms in the sample.

Plan Avogadro's number provides the conversion factor between the number of moles of $C_6H_{12}O_6$ and the number of molecules of $C_6H_{12}O_6$. Once we know the number of molecules of $C_6H_{12}O_6$, we can use the chemical formula, which tells us that each molecule of $C_6H_{12}O_6$ contains 12 H atoms. Thus, we convert moles of $C_6H_{12}O_6$ to molecules of $C_6H_{12}O_6$ and then determine the number of atoms of H from the number of molecules of $C_6H_{12}O_6$:

$$\text{Moles } C_6H_{12}O_6 \longrightarrow \text{molecules } C_6H_{12}O_6 \longrightarrow \text{atoms H}$$

Solve

$$\text{H atoms} = (0.350\ \cancel{\text{mol } C_6H_{12}O_6})\left(\frac{6.02 \times 10^{23}\ \cancel{\text{molecules } C_6H_{12}O_6}}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)\left(\frac{12\text{ H atoms}}{1\ \cancel{\text{molecule } C_6H_{12}O_6}}\right)$$
$$= 2.53 \times 10^{24}\text{ H atoms}$$

Check The magnitude of our answer is reasonable. It is a large number about the magnitude of Avogadro's number. We can also make the following ballpark calculation: Multiplying $0.35 \times 6 \times 10^{23}$ gives about 2×10^{23} molecules. Multiplying this result by 12 gives $24 \times 10^{23} = 2.4 \times 10^{24}$ H atoms, which agrees with the previous, more detailed calculation. Because we were asked for the number of H atoms, the units of our answer are correct. The given data had three significant figures, so our answer has three significant figures.

PRACTICE EXERCISE

How many oxygen atoms are in **(a)** 0.25 mol $Ca(NO_3)_2$ and **(b)** 1.50 mol of sodium carbonate?
Answers: **(a)** 9.0×10^{23}, **(b)** 2.71×10^{24}

Molar Mass

A dozen (12) is the same number whether we have a dozen eggs or a dozen elephants. Clearly, however, a dozen eggs does not have the same mass as a dozen elephants. Similarly, a mole is always the *same number* (6.02×10^{23}), but 1-mole samples of different substances will have *different masses*. Compare, for example, 1 mol of ^{12}C and 1 mol of ^{24}Mg. A single ^{12}C atom has a mass of 12 amu, whereas a single ^{24}Mg atom is twice as massive, 24 amu (to two significant figures). Because a mole always has the same number of particles, a mole of ^{24}Mg must be twice as massive as a mole of ^{12}C. Because a mole of ^{12}C has a mass of 12 g (by definition), then a mole of ^{24}Mg must have a mass of 24 g. This example illustrates a general rule relating the mass of an atom to the mass of Avogadro's number (1 mol) of these atoms: *The mass of a single atom of an element (in amu) is numerically equal to the mass (in grams) of 1 mol of that element.* This statement is true regardless of the element:

$$\text{1 atom of } ^{12}\text{C has a mass of 12 amu} \Rightarrow \text{1 mol } ^{12}\text{C has a mass of 12 g}$$
$$\text{1 atom of Cl has an atomic weight of 35.5 amu} \Rightarrow \text{1 mol Cl has a mass of 35.5 g}$$
$$\text{1 atom of Au has an atomic weight of 197 amu} \Rightarrow \text{1 mol Au has a mass of 197 g}$$

Notice that when we are dealing with a particular isotope of an element, we use the mass of that isotope; otherwise we use the atomic weight (the average atomic mass) of the element. (Section 2.4)

For other kinds of substances, the same numerical relationship exists between the formula weight (in amu) and the mass (in grams) of one mole of that substance:

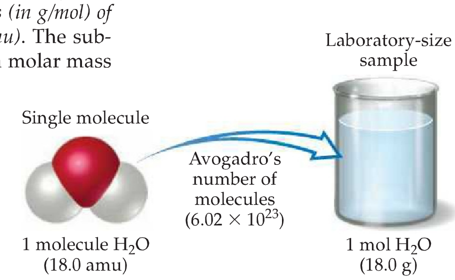
$$\text{1 H}_2\text{O molecule has a mass of 18.0 amu} \Rightarrow \text{1 mol H}_2\text{O has a mass of 18.0 g}$$
$$\text{1 NO}_3^- \text{ ion has a mass of 62.0 amu} \Rightarrow \text{1 mol NO}_3^- \text{ has a mass of 62.0 g}$$
$$\text{1 NaCl unit has a mass of 58.5 amu} \Rightarrow \text{1 mol NaCl has a mass of 58.5 g}$$

Figure 3.8 ▼ illustrates the relationship between the mass of a single molecule of H_2O and that of a mole of H_2O.

GIVE IT SOME THOUGHT

(a) Which has more mass, a mole of water (H_2O) or a mole of glucose ($C_6H_{12}O_6$)?
(b) Which contains more molecules, a mole of water or a mole of glucose?

The mass in grams of one mole of a substance (that is, the mass in grams per mol) is called the **molar mass** of the substance. *The molar mass (in g/mol) of any substance is always numerically equal to its formula weight (in amu).* The substance NaCl, for example, has a formula weight of 58.5 amu and a molar mass

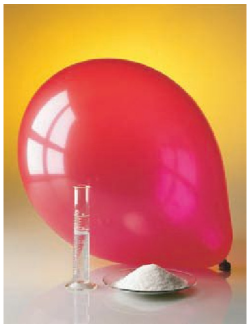


▶ **Figure 3.8 Comparing the mass of 1 molecule H_2O and 1 mol H_2O.** Notice that both masses have the same number but have different units (18.0 amu compared to 18.0 g) representing the huge difference in mass.

TABLE 3.2 ■ Mole Relationships

Name of Substance	Formula	Formula Weight (amu)	Molar Mass (g/mol)	Number and Kind of Particles in One Mole
Atomic nitrogen	N	14.0	14.0	6.02×10^{23} N atoms
Molecular nitrogen	N_2	28.0	28.0	$\begin{cases} 6.02 \times 10^{23}\ N_2 \text{ molecules} \\ 2(6.02 \times 10^{23})\ \text{N atoms} \end{cases}$
Silver	Ag	107.9	107.9	6.02×10^{23} Ag atoms
Silver ions	Ag^+	107.9[a]	107.9	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208.2	208.2	$\begin{cases} 6.02 \times 10^{23}\ BaCl_2 \text{ units} \\ 6.02 \times 10^{23}\ Ba^{2+} \text{ ions} \\ 2(6.02 \times 10^{23})\ Cl^- \text{ ions} \end{cases}$

[a] Recall that the electron has negligible mass; thus, ions and atoms have essentially the same mass.



▲ **Figure 3.9 One mole each of a solid, a liquid, and a gas.** One mole of NaCl, the solid, has a mass of 58.45 g. One mole of H_2O, the liquid, has a mass of 18.0 g and occupies a volume of 18.0 mL. One mole of O_2, the gas, has a mass of 32.0 g and occupies a balloon whose diameter is 35 cm.

of 58.5 g/mol. Further examples of mole relationships are shown in Table 3.2▲. Figure 3.9◀ shows 1-mole quantities of several common substances.

The entries in Table 3.2 for N and N_2 point out the importance of stating the exact chemical form of a substance when we use the mole concept. Suppose you read that 1 mol of nitrogen is produced in a particular reaction. You might interpret this statement to mean 1 mol of nitrogen atoms (14.0 g). Unless otherwise stated, however, what is probably meant is 1 mol of nitrogen molecules, N_2 (28.0 g), because N_2 is the most common chemical form of the element. To avoid ambiguity, it is important to state explicitly the chemical form being discussed. Using the chemical formula N_2 avoids ambiguity.

SAMPLE EXERCISE 3.9 | **Calculating Molar Mass**

What is the mass in grams of 1.000 mol of glucose, $C_6H_{12}O_6$?

SOLUTION

Analyze We are given a chemical formula and asked to determine its molar mass.

Plan The molar mass of a substance is found by adding the atomic weights of its component atoms.

Solve

$$\begin{aligned} 6\text{ C atoms} &= 6(12.0\text{ amu}) = 72.0\text{ amu} \\ 12\text{ H atoms} &= 12(1.0\text{ amu}) = 12.0\text{ amu} \\ 6\text{ O atoms} &= 6(16.0\text{ amu}) = \underline{96.0\text{ amu}} \\ & \qquad\qquad\qquad\qquad 180.0\text{ amu} \end{aligned}$$

Because glucose has a formula weight of 180.0 amu, one mole of this substance has a mass of 180.0 g. In other words, $C_6H_{12}O_6$ has a molar mass of 180.0 g/mol.

Check The magnitude of our answer seems reasonable, and g/mol is the appropriate unit for the molar mass.

Comment Glucose is sometimes called dextrose. Also known as blood sugar, glucose is found widely in nature, occurring in honey and fruits. Other types of sugars used as food are converted into glucose in the stomach or liver before the body uses them as energy sources. Because glucose requires no conversion, it is often given intravenously to patients who need immediate nourishment. People who have diabetes must carefully monitor the amount of glucose in their blood (See "Chemistry and Life" box in Section 3.6).

PRACTICE EXERCISE

Calculate the molar mass of $Ca(NO_3)_2$.
Answer: 164.1 g/mol

Interconverting Masses and Moles

Conversions of mass to moles and of moles to mass are frequently encountered in calculations using the mole concept. These calculations are simplified using dimensional analysis, as shown in Sample Exercises 3.10 and 3.11.

SAMPLE EXERCISE 3.10 | Converting Grams to Moles

Calculate the number of moles of glucose ($C_6H_{12}O_6$) in 5.380 g of $C_6H_{12}O_6$.

SOLUTION

Analyze We are given the number of grams of a substance and its chemical formula and asked to calculate the number of moles.

Plan The molar mass of a substance provides the factor for converting grams to moles. The molar mass of $C_6H_{12}O_6$ is 180.0 g/mol (Sample Exercise 3.9).

Solve Using 1 mol $C_6H_{12}O_6$ = 180.0 g $C_6H_{12}O_6$ to write the appropriate conversion factor, we have

$$\text{Moles } C_6H_{12}O_6 = (5.380\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right) = 0.02989 \text{ mol } C_6H_{12}O_6$$

Check Because 5.380 g is less than the molar mass, a reasonable answer is less than one mole. The units of our answer (mol) are appropriate. The original data had four significant figures, so our answer has four significant figures.

PRACTICE EXERCISE

How many moles of sodium bicarbonate ($NaHCO_3$) are in 508 g of $NaHCO_3$?
Answer: 6.05 mol $NaHCO_3$

SAMPLE EXERCISE 3.11 | Converting Moles to Grams

Calculate the mass, in grams, of 0.433 mol of calcium nitrate.

SOLUTION

Analyze We are given the number of moles and the name of a substance and asked to calculate the number of grams in the sample.

Plan To convert moles to grams, we need the molar mass, which we can calculate using the chemical formula and atomic weights.

Solve Because the calcium ion is Ca^{2+} and the nitrate ion is NO_3^-, calcium nitrate is $Ca(NO_3)_2$. Adding the atomic weights of the elements in the compound gives a formula weight of 164.1 amu. Using 1 mol $Ca(NO_3)_2$ = 164.1 g $Ca(NO_3)_2$ to write the appropriate conversion factor, we have

$$\text{Grams } Ca(NO_3)_2 = (0.433\ \cancel{\text{mol } Ca(NO_3)_2})\left(\frac{164.1 \text{ g } Ca(NO_3)_2}{1\ \cancel{\text{mol } Ca(NO_3)_2}}\right) = 71.1 \text{ g } Ca(NO_3)_2$$

Check The number of moles is less than 1, so the number of grams must be less than the molar mass, 164.1 g. Using rounded numbers to estimate, we have $0.5 \times 150 = 75$ g. The magnitude of our answer is reasonable. Both the units (g) and the number of significant figures (3) are correct.

PRACTICE EXERCISE

What is the mass, in grams, of **(a)** 6.33 mol of $NaHCO_3$ and **(b)** 3.0×10^{-5} mol of sulfuric acid?
Answers: **(a)** 532 g, **(b)** 2.9×10^{-3} g

▶ **Figure 3.10 Procedure for interconverting the mass and the number of formula units of a substance.** The number of moles of the substance is central to the calculation; thus, the mole concept can be thought of as the bridge between the mass of a substance in grams and the number of formula units.

Interconverting Masses and Numbers of Particles

The mole concept provides the bridge between mass and the number of particles. To illustrate how we can interconvert mass and numbers of particles, let's calculate the number of copper atoms in an old copper penny. Such a penny weighs about 3 g, and we will assume that it is 100% copper:

$$\begin{aligned}\text{Cu atoms} &= (3\ \cancel{\text{g Cu}})\left(\frac{1\ \cancel{\text{mol Cu}}}{63.5\ \cancel{\text{g Cu}}}\right)\left(\frac{6.02\times10^{23}\ \text{Cu atoms}}{1\ \cancel{\text{mol Cu}}}\right)\\ &= 3\times10^{22}\ \text{Cu atoms}\end{aligned}$$

We have rounded our answer to one significant figure, since we used only one significant figure for the mass of the penny. Notice how dimensional analysis ∞ (Section 1.6) provides a straightforward route from grams to numbers of atoms. The molar mass and Avogadro's number are used as conversion factors to convert grams → moles → atoms. Notice also that our answer is a very large number. Any time you calculate the number of atoms, molecules, or ions in an ordinary sample of matter, you can expect the answer to be very large. In contrast, the number of moles in a sample will usually be much smaller, often less than 1. The general procedure for interconverting mass and number of formula units (atoms, molecules, ions, or whatever is represented by the chemical formula) of a substance is summarized in Figure 3.10▲.

SAMPLE EXERCISE 3.12 | Calculating the Number of Molecules and Number of Atoms from Mass

(a) How many glucose molecules are in 5.23 g of $C_6H_{12}O_6$? **(b)** How many oxygen atoms are in this sample?

SOLUTION

Analyze We are given the number of grams and the chemical formula and asked to calculate (a) the number of molecules and (b) the number of O atoms in the sample.

(a) Plan The strategy for determining the number of molecules in a given quantity of a substance is summarized in Figure 3.10. We must convert 5.23 g $C_6H_{12}O_6$ to moles $C_6H_{12}O_6$, which can then be converted to molecules $C_6H_{12}O_6$. The first conversion uses the molar mass of $C_6H_{12}O_6$: 1 mol $C_6H_{12}O_6$ = 180.0 g $C_6H_{12}O_6$. The second conversion uses Avogadro's number.

Solve

Molecules $C_6H_{12}O_6$

$$\begin{aligned}&= (5.23\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6.02\times10^{23}\ \text{molecules } C_6H_{12}O_6}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)\\ &= 1.75\times10^{22}\ \text{molecules } C_6H_{12}O_6\end{aligned}$$

Check The magnitude of the answer is reasonable. Because the mass we began with is less than a mole, there should be fewer than 6.02×10^{23} molecules. We can make a ballpark estimate of the answer: $5/200 = 2.5 \times 10^{-2}$ mol; $2.5 \times 10^{-2} \times 6 \times 10^{23} = 15 \times 10^{21} = 1.5 \times 10^{22}$ molecules. The units (molecules) and significant figures (three) are appropriate.

(b) Plan To determine the number of O atoms, we use the fact that there are six O atoms in each molecule of $C_6H_{12}O_6$. Thus, multiplying the number of molecules $C_6H_{12}O_6$ by the factor (6 atoms O/1 molecule $C_6H_{12}O_6$) gives the number of O atoms.

Solve

$$\begin{aligned}\text{Atoms O} &= (1.75\times10^{22}\ \cancel{\text{molecules } C_6H_{12}O_6})\left(\frac{6\ \text{atoms O}}{1\ \cancel{\text{molecule } C_6H_{12}O_6}}\right)\\ &= 1.05\times10^{23}\ \text{atoms O}\end{aligned}$$

Check The answer is simply 6 times as large as the answer to part (a). The number of significant figures (three) and the units (atoms O) are correct.

PRACTICE EXERCISE

(a) How many nitric acid molecules are in 4.20 g of HNO_3? **(b)** How many O atoms are in this sample?
Answers: **(a)** 4.01×10^{22} molecules HNO_3, **(b)** 1.20×10^{23} atoms O

3.5 EMPIRICAL FORMULAS FROM ANALYSES

As we learned in Section 2.6, the empirical formula for a substance tells us the relative number of atoms of each element it contains. The empirical formula H_2O shows that water contains two H atoms for each O atom. This ratio also applies on the molar level: 1 mol of H_2O contains 2 mol of H atoms and 1 mol of O atoms. Conversely, *the ratio of the number of moles of each element in a compound gives the subscripts in a compound's empirical formula*. In this way, the mole concept provides a way of calculating the empirical formulas of chemical substances, as shown in the following examples.

Mercury and chlorine combine to form a compound that is 73.9% mercury and 26.1% chlorine by mass. This means that if we had a 100.0-g sample of the solid, it would contain 73.9 g of mercury (Hg) and 26.1 g of chlorine (Cl). (Any size sample can be used in problems of this type, but we will generally use 100.0 g to simplify the calculation of mass from percentage.) Using the atomic weights of the elements to give us molar masses, we can calculate the number of moles of each element in the sample:

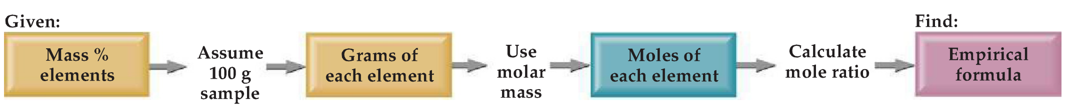
$$(73.9\ \cancel{\text{g Hg}})\left(\frac{1\ \text{mol Hg}}{200.6\ \cancel{\text{g Hg}}}\right) = 0.368\ \text{mol Hg}$$

$$(26.1\ \cancel{\text{g Cl}})\left(\frac{1\ \text{mol Cl}}{35.5\ \cancel{\text{g Cl}}}\right) = 0.735\ \text{mol Cl}$$

We then divide the larger number of moles (0.735) by the smaller (0.368) to obtain a Cl: Hg mole ratio of 1.99:1:

$$\frac{\text{moles of Cl}}{\text{moles of Hg}} = \frac{0.735\ \text{mol Cl}}{0.368\ \text{mol Hg}} = \frac{1.99\ \text{mol Cl}}{1\ \text{mol Hg}}$$

Because of experimental errors, the results may not lead to exact integers for the ratios of moles. The number 1.99 is very close to 2, so we can confidently conclude that the empirical formula for the compound is $HgCl_2$. The empirical formula is correct because its subscripts are the smallest integers that express the *ratios* of atoms present in the compound. (Section 2.6) The general procedure for determining empirical formulas is outlined in Figure 3.11 ▼.



▲ **Figure 3.11 Procedure for calculating an empirical formula from percentage composition.** The central part of the calculation is determining the number of moles of each element in the compound. The procedure is also summarized as "percent to mass, mass to mole, divide by small, multiply 'til whole."

SAMPLE EXERCISE 3.13 | Calculating an Empirical Formula

Ascorbic acid (vitamin C) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid?

SOLUTION

Analyze We are to determine an empirical formula of a compound from the mass percentages of its elements.

Plan The strategy for determining the empirical formula involves the three steps given in Figure 3.11.

Solve We *first* assume, for simplicity, that we have exactly 100 g of material (although any mass can be used). In 100 g of ascorbic acid, therefore, we have

40.92 g C, 4.58 g H, and 54.50 g O.

Second, we calculate the number of moles of each element:

$$\text{Moles C} = (40.92\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}}\right) = 3.407\ \text{mol C}$$

$$\text{Moles H} = (4.58\ \cancel{\text{g H}})\left(\frac{1\ \text{mol H}}{1.008\ \cancel{\text{gH}}}\right) = 4.54\ \text{mol H}$$

$$\text{Moles O} = (54.50\ \cancel{\text{g O}})\left(\frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}}\right) = 3.406\ \text{mol O}$$

Third, we determine the simplest whole-number ratio of moles by dividing each number of moles by the smallest number of moles, 3.406:

$$\text{C:}\ \frac{3.407}{3.406} = 1.000 \qquad \text{H:}\ \frac{4.54}{3.406} = 1.33 \qquad \text{O:}\ \frac{3.406}{3.406} = 1.000$$

The ratio for H is too far from 1 to attribute the difference to experimental error; in fact, it is quite close to $1\frac{1}{3}$. This suggests that if we multiply the ratio by 3, we will obtain whole numbers:

C:H:O = 3(1:1.33:1) = 3:4:3

The whole-number mole ratio gives us the subscripts for the empirical formula:

$C_3H_4O_3$

Check It is reassuring that the subscripts are moderately sized whole numbers. Otherwise, we have little by which to judge the reasonableness of our answer.

PRACTICE EXERCISE

A 5.325-g sample of methyl benzoate, a compound used in the manufacture of perfumes, contains 3.758 g of carbon, 0.316 g of hydrogen, and 1.251 g of oxygen. What is the empirical formula of this substance?
Answer: C_4H_4O

Molecular Formula from Empirical Formula

For any compound, the formula obtained from percentage compositions is always the empirical formula. We can obtain the molecular formula from the empirical formula if we are given the molecular weight or molar mass of the compound. *The subscripts in the molecular formula of a substance are always a whole-number multiple of the corresponding subscripts in its empirical formula.* (Section 2.6) This multiple can be found by comparing the empirical formula weight with the molecular weight:

$$\text{Whole-number multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}} \qquad [3.11]$$

In Sample Exercise 3.13, for example, the empirical formula of ascorbic acid was determined to be $C_3H_4O_3$, giving an empirical formula weight of 3(12.0 amu) + 4(1.0 amu) + 3(16.0 amu) = 88.0 amu. The experimentally determined molecular weight is 176 amu. Thus, the molecular weight is 2 times the empirical formula weight (176/88.0 = 2.00), and the molecular formula must therefore have twice as many of each kind of atom as the empirical formula. Consequently, we multiply the subscripts in the empirical formula by 2 to obtain the molecular formula: $C_6H_8O_6$.

SAMPLE EXERCISE 3.14 | Determining a Molecular Formula

Mesitylene, a hydrocarbon that occurs in small amounts in crude oil, has an empirical formula of C_3H_4. The experimentally determined molecular weight of this substance is 121 amu. What is the molecular formula of mesitylene?

SOLUTION

Analyze We are given an empirical formula and a molecular weight and asked to determine a molecular formula.

Plan The subscripts in the molecular formula of a compound are whole-number multiples of the subscripts in its empirical formula. To find the appropriate multiple, we must compare the molecular weight with the formula weight of the empirical formula.

Solve First, we calculate the formula weight of the empirical formula, C_3H_4:

$$3(12.0 \text{ amu}) + 4(1.0 \text{ amu}) = 40.0 \text{ amu}$$

Next, we divide the molecular weight by the empirical formula weight to obtain the multiple used to multiply the subscripts in C_3H_4:

$$\frac{\text{Molecular weight}}{\text{Empirical formula weight}} = \frac{121}{40.0} = 3.02$$

Only whole-number ratios make physical sense because we must be dealing with whole atoms. The 3.02 in this case could result from a small experimental error in the molecular weight. We therefore multiply each subscript in the empirical formula by 3 to give the molecular formula: C_9H_{12}.

Check We can have confidence in the result because dividing the molecular weight by the formula weight yields nearly a whole number.

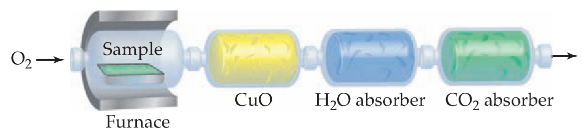
PRACTICE EXERCISE

Ethylene glycol, the substance used in automobile antifreeze, is composed of 38.7% C, 9.7% H, and 51.6% O by mass. Its molar mass is 62.1 g/mol. **(a)** What is the empirical formula of ethylene glycol? **(b)** What is its molecular formula?
Answers: **(a)** CH_3O, **(b)** $C_2H_6O_2$

Combustion Analysis

The empirical formula of a compound is based on experiments that give the number of moles of each element in a sample of the compound. The word "empirical" means "based on observation and experiment." Chemists have devised a number of experimental techniques to determine empirical formulas. One technique is combustion analysis, which is commonly used for compounds containing principally carbon and hydrogen as their component elements.

When a compound containing carbon and hydrogen is completely combusted in an apparatus such as that shown in Figure 3.12▶, the carbon in the compound is converted to CO_2, and the hydrogen is converted to H_2O. (Section 3.2) The amounts of CO_2 and H_2O produced are determined by measuring the mass increase in the CO_2 and H_2O absorbers. From the masses of CO_2 and H_2O we can calculate the number of moles of C and H in the original compound and thereby the empirical formula. If a third element is present in the compound, its mass can be determined by subtracting the masses of C and H from the compound's original mass. Sample Exercise 3.15 shows how to determine the empirical formula of a compound containing C, H, and O.



▲ **Figure 3.12 Apparatus to determine percentages of carbon and hydrogen in a compound.** The compound is combusted to form CO_2 and H_2O. Copper oxide helps to oxidize traces of carbon and carbon monoxide to carbon dioxide and to oxidize hydrogen to water.

SAMPLE EXERCISE 3.15 | **Determining Empirical Formula by Combustion Analysis**

Isopropyl alcohol, a substance sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of isopropyl alcohol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of isopropyl alcohol.

SOLUTION

Analyze We are told that isopropyl alcohol contains C, H, and O atoms and given the quantities of CO_2 and H_2O produced when a given quantity of the alcohol is combusted. We must use this information to determine the empirical formula for isopropyl alcohol, a task that requires us to calculate the number of moles of C, H, and O in the sample.

Plan We can use the mole concept to calculate the number of grams of C present in the CO_2 and the number of grams of H present in the H_2O. These amounts are the quantities of C and H present in the isopropyl alcohol before combustion. The number of grams of O in the compound equals the mass of the isopropyl alcohol minus the sum of the C and H masses. Once we have the number of grams of C, H, and O in the sample, we can then proceed as in Sample Exercise 3.13. We can calculate the number of moles of each element, and determine the mole ratio, which gives the subscripts in the empirical formula.

Solve To calculate the number of grams of C, we first use the molar mass of CO_2, 1 mol CO_2 = 44.0 g CO_2, to convert grams of CO_2 to moles of CO_2. Because each CO_2 molecule has only 1 C atom, there is 1 mol of C atoms per mole of CO_2 molecules. This fact allows us to convert the moles of CO_2 to moles of C. Finally, we use the molar mass of C, 1 mol C = 12.0 g C, to convert moles of C to grams of C. Combining the three conversion factors, we have

$$\text{Grams C} = (0.561\ \cancel{\text{g CO}_2})\left(\frac{1\ \cancel{\text{mol CO}_2}}{44.0\ \cancel{\text{g CO}_2}}\right)\left(\frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol CO}_2}}\right)\left(\frac{12.0\ \text{g C}}{1\ \cancel{\text{mol C}}}\right) = 0.153\ \text{g C}$$

The calculation of the number of grams of H from the grams of H_2O is similar, although we must remember that there are 2 mol of H atoms per 1 mol of H_2O molecules:

$$\text{Grams H} = (0.306\ \cancel{\text{g H}_2\text{O}})\left(\frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.0\ \cancel{\text{g H}_2\text{O}}}\right)\left(\frac{2\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol H}_2\text{O}}}\right)\left(\frac{1.01\ \text{g H}}{1\ \cancel{\text{mol H}}}\right) = 0.0343\ \text{g H}$$

The total mass of the sample, 0.255 g, is the sum of the masses of the C, H, and O. Thus, we can calculate the mass of O as follows:

$$\begin{aligned}\text{Mass of O} &= \text{mass of sample} - (\text{mass of C} + \text{mass of H}) \\ &= 0.255\ \text{g} - (0.153\ \text{g} + 0.0343\ \text{g}) = 0.068\ \text{g O}\end{aligned}$$

We then calculate the number of moles of C, H, and O in the sample:

$$\text{Moles C} = (0.153\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.0\ \cancel{\text{g C}}}\right) = 0.0128\ \text{mol C}$$

$$\text{Moles H} = (0.0343\ \cancel{\text{g H}})\left(\frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}}\right) = 0.0340\ \text{mol H}$$

$$\text{Moles O} = (0.068\ \cancel{\text{g O}})\left(\frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}}\right) = 0.0043\ \text{mol O}$$

To find the empirical formula, we must compare the relative number of moles of each element in the sample. The relative number of moles of each element is found by dividing each number by the smallest number, 0.0043. The mole ratio of C:H:O so obtained is 2.98:7.91:1.00. The first two numbers are very close to the whole numbers 3 and 8, giving the empirical formula C_3H_8O.

Check The subscripts work out to be moderately sized whole numbers, as expected.

PRACTICE EXERCISE

(a) Caproic acid, which is responsible for the foul odor of dirty socks, is composed of C, H, and O atoms. Combustion of a 0.225-g sample of this compound produces 0.512 g CO_2 and 0.209 g H_2O. What is the empirical formula of caproic acid? **(b)** Caproic acid has a molar mass of 116 g/mol. What is its molecular formula?

Answers: **(a)** C_3H_6O, **(b)** $C_6H_{12}O_2$

GIVE IT SOME THOUGHT

In Sample Exercise 3.15, how do you explain the fact that the ratios C:H:O are 2.98:7.91:1.00, rather than exact integers 3:8:1?

3.6 QUANTITATIVE INFORMATION FROM BALANCED EQUATIONS

The coefficients in a chemical equation represent the relative numbers of molecules in a reaction. The mole concept allows us to convert this information to the masses of the substances. Consider the following balanced equation:

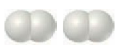
$$2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(l) \qquad [3.12]$$

The coefficients indicate that two molecules of H_2 react with each molecule of O_2 to form two molecules of H_2O. It follows that the relative numbers of moles are identical to the relative numbers of molecules:

$2\ H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2\ H_2O(l)$
2 molecules		1 molecule		2 molecules
$2(6.02 \times 10^{23}$ molecules)		$1(6.02 \times 10^{23}$ molecules)		$2(6.02 \times 10^{23}$ molecules)
2 mol		1 mol		2 mol

We can generalize this observation for all balanced chemical equations: *The coefficients in a balanced chemical equation indicate both the relative numbers of molecules (or formula units) in the reaction and the relative numbers of moles.* Table 3.3▼ further summarizes this result and shows how it corresponds to the law of conservation of mass. Notice that the total mass of the reactants (4.0 g + 32.0 g) equals the total mass of the products (36.0 g).

TABLE 3.3 ■ Information from a Balanced Equation

Equation:	**$2\ H_2(g)$**	**+**	**$O_2(g)$**	**⟶**	**$2\ H_2O(l)$**
Molecules:	2 molecules H_2	+	1 molecule O_2	⟶	2 molecules H_2O
					
Mass (amu):	4.0 amu H_2	+	32.0 amu O_2	⟶	36.0 amu H_2O
Amount (mol):	2 mol H_2	+	1 mol O_2	⟶	2 mol H_2O
Mass (g):	4.0 g H_2	+	32.0 g O_2	⟶	36.0 g H_2O

The quantities 2 mol H_2, 1 mol O_2, and 2 mol H_2O, which are given by the coefficients in Equation 3.12, are called *stoichiometrically equivalent quantities*. The relationship between these quantities can be represented as

$$2\ \text{mol}\ H_2 \simeq 1\ \text{mol}\ O_2 \simeq 2\ \text{mol}\ H_2O$$

where the $\simeq$ symbol means "is stoichiometrically equivalent to." In other words, Equation 3.12 shows 2 mol of H_2 and 1 mol of O_2 forming 2 mol of H_2O. These stoichiometric relations can be used to convert between quantities of reactants and products in a chemical reaction. For example, the number of moles of H_2O produced from 1.57 mol of O_2 can be calculated as follows:

$$\text{Moles}\ H_2O = (1.57\ \cancel{\text{mol}\ O_2})\left(\frac{2\ \text{mol}\ H_2O}{1\ \cancel{\text{mol}\ O_2}}\right) = 3.14\ \text{mol}\ H_2O$$

GIVE IT SOME THOUGHT

When 1.57 mol O_2 reacts with H_2 to form H_2O, how many moles of H_2 are consumed in the process?

As an additional example, consider the combustion of butane (C_4H_{10}), the fuel in disposable cigarette lighters:

$$2\ C_4H_{10}(l) + 13\ O_2(g) \longrightarrow 8\ CO_2(g) + 10\ H_2O(g) \qquad [3.13]$$

Let's calculate the mass of CO_2 produced when 1.00 g of C_4H_{10} is burned. The coefficients in Equation 3.13 tell how the amount of C_4H_{10} consumed is related to the amount of CO_2 produced: 2 mol $C_4H_{10} \simeq$ 8 mol CO_2. To use this relationship we must use the molar mass of C_4H_{10} to convert grams of C_4H_{10} to moles of C_4H_{10}. Because 1 mol C_4H_{10} = 58.0 g C_4H_{10}, we have

$$\text{Moles}\ C_4H_{10} = (1.00\ \cancel{\text{g}\ C_4H_{10}})\left(\frac{1\ \text{mol}\ C_4H_{10}}{58.0\ \cancel{\text{g}\ C_4H_{10}}}\right)$$
$$= 1.72 \times 10^{-2}\ \text{mol}\ C_4H_{10}$$

We can then use the stoichiometric factor from the balanced equation, 2 mol $C_4H_{10} \simeq$ 8 mol CO_2, to calculate moles of CO_2:

$$\text{Moles}\ CO_2 = (1.72 \times 10^{-2}\ \cancel{\text{mol}\ C_4H_{10}})\left(\frac{8\ \text{mol}\ CO_2}{2\ \cancel{\text{mol}\ C_4H_{10}}}\right)$$
$$= 6.88 \times 10^{-2}\ \text{mol}\ CO_2$$

Finally, we can calculate the mass of the CO_2, in grams, using the molar mass of CO_2 (1 mol CO_2 = 44.0 g CO_2):

$$\text{Grams } CO_2 = (6.88 \times 10^{-2}\ \cancel{\text{mol } CO_2})\left(\frac{44.0 \text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right)$$
$$= 3.03 \text{ g } CO_2$$

Thus, the conversion sequence is

Grams reactant → Moles reactant → Moles product → Grams product

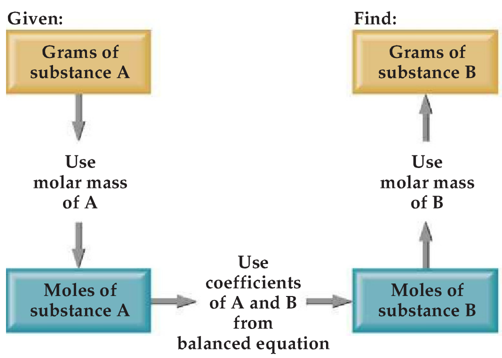
These steps can be combined in a single sequence of factors:

$$\text{Grams } CO_2 = (1.00\ \cancel{\text{g } C_4H_{10}})\left(\frac{1\ \cancel{\text{mol } C_4H_{10}}}{58.0\ \cancel{\text{g } C_4H_{10}}}\right)\left(\frac{8\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_4H_{10}}}\right)\left(\frac{44.0 \text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right)$$
$$= 3.03 \text{ g } CO_2$$

Similarly, we can calculate the amount of O_2 consumed or H_2O produced in this reaction. For example, to calculate the amount of O_2 consumed, we again rely on the coefficients in the balanced equation to give us the appropriate stoichiometric factor: 2 mol $C_4H_{10} \simeq$ 13 mol O_2:

$$\text{Grams } O_2 = (1.00\ \cancel{\text{g } C_4H_{10}})\left(\frac{1\ \cancel{\text{mol } C_4H_{10}}}{58.0\ \cancel{\text{g } C_4H_{10}}}\right)\left(\frac{13\ \cancel{\text{mol } O_2}}{2\ \cancel{\text{mol } C_4H_{10}}}\right)\left(\frac{32.0 \text{ g } O_2}{1\ \cancel{\text{mol } O_2}}\right)$$
$$= 3.59 \text{ g } O_2$$

Figure 3.13 ▼ summarizes the general approach used to calculate the quantities of substances consumed or produced in chemical reactions. The balanced chemical equation provides the relative numbers of moles of reactants and products in the reaction.



▶ **Figure 3.13 The procedure for calculating amounts of reactants or products in a reaction.** The number of grams of a reactant consumed or of a product formed in a reaction can be calculated, starting with the number of grams of one of the other reactants or products. Notice how molar masses and the coefficients in the balanced equation are used.

GIVE IT SOME THOUGHT

If 20.00 g of a compound reacts completely with 30.00 g of another compound in a combination reaction, how many grams of product were formed?

SAMPLE EXERCISE 3.16 | **Calculating Amounts of Reactants and Products**

How many grams of water are produced in the oxidation of 1.00 g of glucose, $C_6H_{12}O_6$?

$$C_6H_{12}O_6(s) + 6\ O_2(g) \longrightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

SOLUTION

Analyze We are given the mass of a reactant and are asked to determine the mass of a product in the given equation.

Plan The general strategy, as outlined in Figure 3.13, requires three steps. First, the amount of $C_6H_{12}O_6$ must be converted from grams to moles. Second, we can use the balanced equation, which relates the moles of $C_6H_{12}O_6$ to the moles of H_2O: 1 mol $C_6H_{12}O_6 \simeq 6$ mol H_2O. Third, we must convert the moles of H_2O to grams.

Solve First, use the molar mass of $C_6H_{12}O_6$ to convert from grams $C_6H_{12}O_6$ to moles $C_6H_{12}O_6$:

$$\text{Moles } C_6H_{12}O_6 = (1.00\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\text{ mol } C_6H_{12}O_6}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)$$

Second, use the balanced equation to convert moles of $C_6H_{12}O_6$ to moles of H_2O:

$$\text{Moles } H_2O = (1.00\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6\text{ mol } H_2O}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)$$

Third, use the molar mass of H_2O to convert from moles of H_2O to grams of H_2O:

$$\begin{aligned}\text{Grams } H_2O &= (1.00\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6\ \cancel{\text{mol } H_2O}}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)\left(\frac{18.0\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}}\right)\\ &= 0.600\text{ g } H_2O\end{aligned}$$

The steps can be summarized in a diagram like that in Figure 3.13:

$$\begin{array}{ccccc}
\boxed{1.00\text{ g } C_6H_{12}O_6} & \xrightarrow[\text{calculation}]{\text{no direct}} & & & \boxed{0.600\text{ g } H_2O}\\
\downarrow & & & & \uparrow\\
\times\left(\dfrac{1\text{ mol } C_6H_{12}O_6}{180.0\text{ g } C_6H_{12}O_6}\right) & & & & \times\left(\dfrac{18.0\text{ g } H_2O}{1\text{ mol } H_2O}\right)\\
\downarrow & & & & \uparrow\\
\boxed{5.56\times10^{-3}\text{ mol } C_6H_{12}O_6} & \longrightarrow & \times\left(\dfrac{6\text{ mol } H_2O}{1\text{ mol } C_6H_{12}O_6}\right) & \longrightarrow & \boxed{3.33\times10^{-2}\text{ mol } H_2O}
\end{array}$$

Check An estimate of the magnitude of our answer, $18/180 = 0.1$ and $0.1 \times 6 = 0.6$, agrees with the exact calculation. The units, grams H_2O, are correct. The initial data had three significant figures, so three significant figures for the answer is correct.

Comment An average person ingests 2 L of water daily and eliminates 2.4 L. The difference between 2 L and 2.4 L is produced in the metabolism of foodstuffs, such as in the oxidation of glucose. (*Metabolism* is a general term used to describe all the chemical processes of a living animal or plant.) The desert rat (kangaroo rat), on the other hand, apparently never drinks water. It survives on its metabolic water.

PRACTICE EXERCISE

The decomposition of $KClO_3$ is commonly used to prepare small amounts of O_2 in the laboratory: $2\ KClO_3(s) \longrightarrow 2\ KCl(s) + 3\ O_2(g)$. How many grams of O_2 can be prepared from 4.50 g of $KClO_3$?
Answer: 1.77 g

SAMPLE EXERCISE 3.17 | Calculating Amounts of Reactants and Products

Solid lithium hydroxide is used in space vehicles to remove the carbon dioxide exhaled by astronauts. The lithium hydroxide reacts with gaseous carbon dioxide to form solid lithium carbonate and liquid water. How many grams of carbon dioxide can be absorbed by 1.00 g of lithium hydroxide?

SOLUTION

Analyze We are given a verbal description of a reaction and asked to calculate the number of grams of one reactant that reacts with 1.00 g of another.

Plan The verbal description of the reaction can be used to write a balanced equation:

$$2\ LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

We are given the grams of LiOH and asked to calculate grams of CO_2. We can accomplish this task by using the following sequence of conversions:

$$\text{Grams LiOH} \longrightarrow \text{moles LiOH} \longrightarrow \text{moles } CO_2 \longrightarrow \text{grams } CO_2$$

The conversion from grams of LiOH to moles of LiOH requires the molar mass of LiOH ($6.94 + 16.00 + 1.01 = 23.95$ g/mol). The conversion of moles of LiOH to moles of CO_2 is based on the balanced chemical equation: 2 mol LiOH $\simeq$ 1 mol CO_2. To convert the number of moles of CO_2 to grams, we must use the molar mass of CO_2: $12.01 + 2(16.00) = 44.01$ g/mol.

Solve

$$(1.00\ \cancel{\text{g LiOH}})\left(\frac{1\ \cancel{\text{mol LiOH}}}{23.95\ \cancel{\text{g LiOH}}}\right)\left(\frac{1\ \cancel{\text{mol CO}_2}}{2\ \cancel{\text{mol LiOH}}}\right)\left(\frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol CO}_2}}\right) = 0.919\ \text{g CO}_2$$

Check Notice that $23.95 \approx 24$, $24 \times 2 = 48$, and $44/48$ is slightly less than 1. The magnitude of the answer is reasonable based on the amount of starting LiOH; the significant figures and units are appropriate, too.

PRACTICE EXERCISE

Propane, C_3H_8, is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?
Answer: 3.64 g

Chemistry and Life — GLUCOSE MONITORING

Over 20 million Americans have diabetes. In the world, the number approaches 172 million. Diabetes is a disorder of metabolism in which the body cannot produce or properly use the hormone insulin. One signal that a person is diabetic is that the concentration of glucose in her or his blood is higher than normal. Therefore, people who are diabetic need to measure their blood glucose concentrations regularly. Untreated diabetes can cause severe complications such as blindness and loss of limbs.

How does insulin relate to glucose? The body converts most of the food we eat into glucose. After digestion, glucose is delivered to cells via the bloodstream; cells need glucose to live. Insulin must be present for glucose to enter the cells. Normally, the body adjusts the concentration of insulin automatically, in concert with the glucose concentration after eating. However, in a diabetic person, little or no insulin is produced (Type 1 diabetes), or the cells cannot take up insulin properly (Type 2 diabetes). The result is that the blood glucose concentration is too high. People normally have a range of 70–120 mg glucose per deciliter of blood (about 4–6 mmol glucose per liter of blood). If a person has not eaten for 8 hours or more, he or she would be diagnosed as diabetic if the glucose levels were 126 mg/dL or higher. In the United States, diabetics monitor their blood glucose concentrations in mg/dL; in Europe, they use different units—millimoles glucose per liter of blood.

Glucose monitors work by the introduction of blood from a person, usually by a prick of the finger, onto a small strip of paper that contains numerous chemicals that react specifically with glucose. Insertion of the strip into a small battery-operated reader gives the glucose concentration (Figure 3.14▼). The actual mechanism of the readout varies for different devices—it may be a small electrical current or a measure of light produced in a chemical reaction. Depending on the result, a diabetic person may need to receive an injection of insulin or simply stop eating sweets for a while.

▲ **Figure 3.14 Glucose meter.** This is an example of a commercial glucose meter and its readout.

3.7 LIMITING REACTANTS

Suppose you wish to make several sandwiches using one slice of cheese and two slices of bread for each sandwich. Using Bd = bread, Ch = cheese, and Bd_2Ch = sandwich, the recipe for making a sandwich can be represented like a chemical equation:

$$2\ \text{Bd} + \text{Ch} \longrightarrow \text{Bd}_2\text{Ch}$$

If you have 10 slices of bread and 7 slices of cheese, you will be able to make only five sandwiches before you run out of bread. You will have 2 slices of cheese left over. The amount of available bread limits the number of sandwiches.

An analogous situation occurs in chemical reactions when one of the reactants is used up before the others. The reaction stops as soon as any one of the reactants is totally consumed, leaving the excess reactants as leftovers.

Suppose, for example, that we have a mixture of 10 mol H_2 and 7 mol O_2, which react to form water:

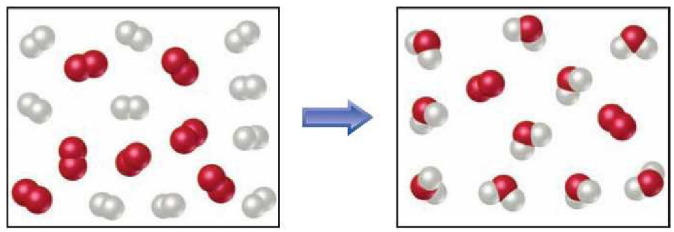
$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

Because 2 mol $H_2 \simeq$ 1 mol O_2, the number of moles of O_2 needed to react with all the H_2 is

$$\text{Moles } O_2 = (10\ \cancel{\text{mol } H_2})\left(\frac{1\text{ mol } O_2}{2\ \cancel{\text{mol } H_2}}\right) = 5\text{ mol } O_2$$

Because 7 mol O_2 was available at the start of the reaction, 7 mol O_2 − 5 mol O_2 = 2 mol O_2 will still be present when all the H_2 is consumed. The example we have considered is depicted on a molecular level in Figure 3.15 ▶.

The reactant that is completely consumed in a reaction is called either the **limiting reactant** or *limiting reagent* because it determines, or limits, the amount of product formed. The other reactants are sometimes called either *excess reactants* or *excess reagents*. In our example, H_2 is the limiting reactant, which means that once all the H_2 has been consumed, the reaction stops. O_2 is the excess reactant, and some is left over when the reaction stops.



▲ **Figure 3.15 Example illustrating a limiting reactant.** Because the H_2 is completely consumed, it is the limiting reagent in this case. Because there is a stoichiometric excess of O_2, some is left over at the end of the reaction. The amount of H_2O formed is related directly to the amount of H_2 consumed.

There are no restrictions on the starting amounts of the reactants in any reaction. Indeed, many reactions are carried out using an excess of one reagent. The quantities of reactants consumed and the quantities of products formed, however, are restricted by the quantity of the limiting reactant. When a combustion reaction takes place in the open air, oxygen is plentiful and is therefore the excess reactant. You may have had the unfortunate experience of running out of gasoline while driving. The car stops because you have run out of the limiting reactant in the combustion reaction, the fuel.

Before we leave our present example, let's summarize the data in a tabular form:

	$2\,H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2\,H_2O(g)$
Initial quantities:	10 mol		7 mol		0 mol
Change (reaction):	−10 mol		−5 mol		+10 mol
Final quantities:	0 mol		2 mol		10 mol

The initial amounts of the reactants are what we started with (10 mol H_2 and 7 mol O_2). The second line in the table (change) summarizes the amounts of the reactants consumed and the amount of the product formed in the reaction. These quantities are restricted by the quantity of the limiting reactant and depend on the coefficients in the balanced equation. The mole ratio H_2: O_2: H_2O = 10:5:10 conforms to the ratio of the coefficients in the balanced equation, 2:1:2. The changes are negative for the reactants because they are consumed during the reaction and positive for the product because it is formed during the reaction. Finally, the quantities in the third line of the table (final quantities) depend on the initial quantities and their changes, and these entries are found by adding the entries for the initial quantity and change for each column. No amount of the limiting reactant (H_2) remains at the end of the reaction. All that remains is 2 mol O_2 and 10 mol H_2O.

SAMPLE EXERCISE 3.18 | Calculating the Amount of Product Formed from a Limiting Reactant

The most important commercial process for converting N_2 from the air into nitrogen-containing compounds is based on the reaction of N_2 and H_2 to form ammonia (NH_3):

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$$

How many moles of NH_3 can be formed from 3.0 mol of N_2 and 6.0 mol of H_2?

SOLUTION

Analyze We are asked to calculate the number of moles of product, NH_3, given the quantities of each reactant, N_2 and H_2, available in a reaction. Thus, this is a limiting reactant problem.

Plan If we assume that one reactant is completely consumed, we can calculate how much of the second reactant is needed in the reaction. By comparing the calculated quantity with the available amount, we can determine which reactant is limiting. We then proceed with the calculation, using the quantity of the limiting reactant.

Solve The number of moles of H_2 needed for complete consumption of 3.0 mol of N_2 is:

$$\text{Moles } H_2 = (3.0\ \cancel{\text{mol } N_2})\left(\frac{3\text{ mol } H_2}{1\ \cancel{\text{mol } N_2}}\right) = 9.0\text{ mol } H_2$$

Because only 6.0 mol H_2 is available, we will run out of H_2 before the N_2 is gone, and H_2 will be the limiting reactant. We use the quantity of the limiting reactant, H_2, to calculate the quantity of NH_3 produced:

$$\text{Moles } NH_3 = (6.0\ \cancel{\text{mol } H_2})\left(\frac{2\text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}}\right) = 4.0\text{ mol } NH_3$$

Comment The table on the right summarizes this example:

	$N_2(g)$	+ $3\,H_2(g)$	$\longrightarrow$ $2\,NH_3(g)$
Initial quantities:	3.0 mol	6.0 mol	0 mol
Change (reaction):	−2.0 mol	−6.0 mol	+4.0 mol
Final quantities:	1.0 mol	0 mol	4.0 mol

Notice that we can calculate not only the number of moles of NH_3 formed but also the number of moles of each of the reactants remaining after the reaction. Notice also that although the number of moles of H_2 present at the beginning of the reaction is greater than the number of moles of N_2 present, the H_2 is nevertheless the limiting reactant because of its larger coefficient in the balanced equation.

Check The summarizing table shows that the mole ratio of reactants used and product formed conforms to the coefficients in the balanced equation, 1:3:2. Also, because H_2 is the limiting reactant, it is completely consumed in the reaction, leaving 0 mol at the end. Because 6.0 mol H_2 has two significant figures, our answer has two significant figures.

PRACTICE EXERCISE

Consider the reaction $2\,Al(s) + 3\,Cl_2(g) \longrightarrow 2\,AlCl_3(s)$. A mixture of 1.50 mol of Al and 3.00 mol of Cl_2 is allowed to react. **(a)** Which is the limiting reactant? **(b)** How many moles of $AlCl_3$ are formed? **(c)** How many moles of the excess reactant remain at the end of the reaction?

Answers: **(a)** Al, **(b)** 1.50 mol, **(c)** 0.75 mol Cl_2

SAMPLE EXERCISE 3.19 | Calculating the Amount of Product Formed from a Limiting Reactant

Consider the following reaction that occurs in a fuel cell:

$$2\,H_2(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(g)$$

This reaction, properly done, produces energy in the form of electricity and water. Suppose a fuel cell is set up with 150 g of hydrogen gas and 1500 grams of oxygen gas (each measurement is given with two significant figures). How many grams of water can be formed?

SOLUTION

Analyze We are asked to calculate the amount of a product, given the amounts of two reactants, so this is a limiting reactant problem.

Plan We must first identify the limiting reagent. To do so, we can calculate the number of moles of each reactant and compare their ratio with that required by the balanced equation. We then use the quantity of the limiting reagent to calculate the mass of water that forms.

Solve From the balanced equation, we have the following stoichiometric relations:

$$2 \text{ mol } H_2 \simeq 1 \text{ mol } O_2 \simeq 2 \text{ mol } H_2O$$

Using the molar mass of each substance, we can calculate the number of moles of each reactant:

$$\text{Moles } H_2 = (150 \cancel{\text{g } H_2})\left(\frac{1 \text{ mol } H_2}{2.00 \cancel{\text{g } H_2}}\right) = 75 \text{ mol } H_2$$

$$\text{Moles } O_2 = (1500 \cancel{\text{g } O_2})\left(\frac{1 \text{ mol } O_2}{32.0 \cancel{\text{g } O_2}}\right) = 47 \text{ mol } O_2$$

Thus, there are more moles of H_2 than O_2. The coefficients in the balanced equation indicate, however, that the reaction requires 2 moles of H_2 for every 1 mole of O_2. Therefore, to completely react all the O_2, we would need $2 \times 47 = 94$ moles of H_2. Since there are only 75 moles of H_2, H_2 is the limiting reagent. We therefore use the quantity of H_2 to calculate the quantity of product formed. We can begin this calculation with the grams of H_2, but we can save a step by starting with the moles of H_2 that were calculated previously in the exercise:

$$\begin{aligned}\text{Grams } H_2O &= (75 \cancel{\text{moles } H_2})\left(\frac{2 \cancel{\text{mol } H_2O}}{2 \cancel{\text{mol } H_2}}\right)\left(\frac{18.0 \text{ g } H_2O}{1 \cancel{\text{mol } H_2O}}\right)\\ &= 1400 \text{ g } H_2O \text{ (to two significant figures)}\end{aligned}$$

Check The magnitude of the answer seems reasonable. The units are correct, and the number of significant figures (two) corresponds to those in the numbers of grams of the starting materials.

Comment The quantity of the limiting reagent, H_2, can also be used to determine the quantity of O_2 used (37.5 mol = 1200 g). The number of grams of the excess oxygen remaining at the end of the reaction equals the starting amount minus the amount consumed in the reaction, 1500 g − 1200 g = 300 g.

PRACTICE EXERCISE

A strip of zinc metal with a mass of 2.00 g is placed in an aqueous solution containing 2.50 g of silver nitrate, causing the following reaction to occur:

$$Zn(s) + 2\,AgNO_3(aq) \longrightarrow 2\,Ag(s) + Zn(NO_3)_2(aq)$$

(a) Which reactant is limiting? **(b)** How many grams of Ag will form? **(c)** How many grams of $Zn(NO_3)_2$ will form? **(d)** How many grams of the excess reactant will be left at the end of the reaction?

Answers: **(a)** $AgNO_3$, **(b)** 1.59 g, **(c)** 1.39 g, **(d)** 1.52 g Zn

Theoretical Yields

The quantity of product that is calculated to form when all of the limiting reactant reacts is called the **theoretical yield**. The amount of product actually obtained in a reaction is called the *actual yield*. The actual yield is almost always less than (and can never be greater than) the theoretical yield. There are many reasons for this difference. Part of the reactants may not react, for example, or they may react in a way different from that desired (side reactions). In addition, it is not always possible to recover all of the product from the reaction mixture. The **percent yield** of a reaction relates the actual yield to the theoretical (calculated) yield:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \qquad [3.14]$$

Strategies in Chemistry HOW TO TAKE A TEST

At about this time in your study of chemistry, you are likely to face your first hour-long examination. The best way to prepare for the exam is to study and do homework diligently and to make sure you get help from the instructor on any material that is unclear or confusing. (See the advice for learning and studying chemistry presented in the preface of the book.) We present here some general guidelines for taking tests.

Depending on the nature of your course, the exam could consist of a variety of different types of questions. Let's consider some of the more common types and how they can best be addressed.

1. *Multiple-choice questions* In large-enrollment courses, the most common kind of testing device is the multiple-choice question. You are given the problem and usually presented with four or five possible answers from which you must select the correct one. The first thing to realize is that the instructor has written the question so that all of the answer choices appear at first glance to be correct. (There would be little point in offering choices you could tell were wrong even without knowing much about the concept being tested.) Thus, you should not jump to the conclusion that because one of the choices looks correct, it must be so.

   If a multiple-choice question involves a calculation, perform the calculation, quickly double-check your work, and *only then* compare your answer with the choices. If you find a match, you have probably found the correct answer. Keep in mind, though, that your instructor has anticipated the most common errors one can make in solving a given problem and has probably listed the incorrect answers resulting from those errors. Always double-check your reasoning and make sure to use dimensional analysis to arrive at the correct answer, with the correct units.

   In multiple-choice questions that do not involve calculations, if you are not sure of the correct choice, eliminate all the choices you know for sure to be incorrect. Additionally, the reasoning you used in eliminating incorrect choices will help you in reasoning about which choice is correct.

SAMPLE EXERCISE 3.20 | **Calculating the Theoretical Yield and Percent Yield for a Reaction**

Adipic acid, $H_2C_6H_8O_4$, is used to produce nylon. The acid is made commercially by a controlled reaction between cyclohexane (C_6H_{12}) and O_2:

$$2\,C_6H_{12}(l) + 5\,O_2(g) \longrightarrow 2\,H_2C_6H_8O_4(l) + 2\,H_2O(g)$$

(a) Assume that you carry out this reaction starting with 25.0 g of cyclohexane and that cyclohexane is the limiting reactant. What is the theoretical yield of adipic acid?
(b) If you obtain 33.5 g of adipic acid from your reaction, what is the percent yield of adipic acid?

SOLUTION

Analyze We are given a chemical equation and the quantity of the limiting reactant (25.0 g of C_6H_{12}). We are asked first to calculate the theoretical yield of a product ($H_2C_6H_8O_4$) and then to calculate its percent yield if only 33.5 g of the substance is actually obtained.

Plan **(a)** The theoretical yield, which is the calculated quantity of adipic acid formed in the reaction, can be calculated using the following sequence of conversions:

$$\text{g } C_6H_{12} \longrightarrow \text{mol } C_6H_{12} \longrightarrow \text{mol } H_2C_6H_8O_4 \longrightarrow \text{g } H_2C_6H_8O_4$$

(b) The percent yield is calculated by comparing the actual yield (33.5 g) to the theoretical yield using Equation 3.14.

Solve

(a) $$\text{Grams } H_2C_6H_8O_4 = (25.0\ \cancel{\text{g } C_6H_{12}})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}}}{84.0\ \cancel{\text{g } C_6H_{12}}}\right)\left(\frac{2\ \cancel{\text{mol } H_2C_6H_8O_4}}{2\ \cancel{\text{mol } C_6H_{12}}}\right)\left(\frac{146.0\text{ g } H_2C_6H_8O_4}{1\ \cancel{\text{mol } H_2C_6H_8O_4}}\right)$$
$$= 43.5\text{ g } H_2C_6H_8O_4$$

(b) $$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{33.5\text{ g}}{43.5\text{ g}} \times 100\% = 77.0\%$$

Check Our answer in (a) has the appropriate magnitude, units, and significant figures. In (b) the answer is less than 100% as necessary.

2. *Calculations in which you must show your work* Your instructor may present you with a numerical problem in which you are to show your work in arriving at a solution. In questions of this kind, you may receive partial credit even if you do not arrive at the correct answer, depending on whether the instructor can follow your line of reasoning. It is important, therefore, to be as neat and organized as you can be, given the pressures of exam taking. It is helpful in approaching such questions to take a few moments to think about the direction you are going to take in solving the problem. You may even want to write a few words or a diagram on the test paper to indicate your approach. Then write out your calculations as neatly as you can. Show the units for every number you write down, and use dimensional analysis as much as you can, showing how units cancel.
3. *Questions requiring drawings* Sometimes a test question will require you to draw a chemical structure, a diagram related to chemical bonding, or a figure showing some kind of chemical process. Questions of this kind will come later in the course, but it is useful to talk about them here. (You should review this box before each exam you take, to remind yourself of good exam-taking practices.) Be sure to label your drawing as completely as possible.
4. *Other types of questions* Other exam questions you might encounter include true-false questions and ones in which you are given a list and asked to indicate which members of the list match some criterion given in the question. Often students answer such questions incorrectly because, in their haste, they misunderstand the nature of the question. Whatever the form of the question, ask yourself this: What is the instructor testing here? What material am I supposed to know that this question covers?

Finally, if you find that you simply do not understand how to arrive at a reasoned response to a question, do not linger over the question. Put a check next to it and go on to the next one. If time permits, you can come back to the unanswered questions, but lingering over a question when nothing is coming to mind is wasting time you may need to finish the exam.

PRACTICE EXERCISE

Imagine that you are working on ways to improve the process by which iron ore containing Fe_2O_3 is converted into iron. In your tests you carry out the following reaction on a small scale:

$$Fe_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Fe(s) + 3\,CO_2(g)$$

(a) If you start with 150 g of Fe_2O_3 as the limiting reagent, what is the theoretical yield of Fe? **(b)** If the actual yield of Fe in your test was 87.9 g, what was the percent yield?
Answers: **(a)** 105 g Fe, **(b)** 83.7%

CHAPTER REVIEW

SUMMARY AND KEY TERMS

Introduction and Section 3.1 The study of the quantitative relationships between chemical formulas and chemical equations is known as **stoichiometry**. One of the important concepts of stoichiometry is the law of conservation of mass, which states that the total mass of the products of a chemical reaction is the same as the total mass of the reactants. The same numbers of atoms of each type are present before and after a chemical reaction. A balanced **chemical equation** shows equal numbers of atoms of each element on each side of the equation. Equations are balanced by placing coefficients in front of the chemical formulas for the **reactants** and **products** of a reaction, *not* by changing the subscripts in chemical formulas.

Section 3.2 Among the reaction types described in this chapter are (1) **combination reactions**, in which two reactants combine to form one product; (2) **decomposition reactions**, in which a single reactant forms two or more products; and (3) **combustion reactions** in oxygen, in which a hydrocarbon or related compound reacts with O_2 to form CO_2 and H_2O.

Section 3.3 Much quantitative information can be determined from chemical formulas and balanced chemical equations by using atomic weights. The **formula weight** of a compound equals the sum of the atomic weights of the atoms in its formula. If the formula is a molecular formula, the formula weight is also called the **molecular weight**. Atomic weights and formula weights can be used to determine the elemental composition of a compound.

Section 3.4 A **mole** of any substance is **Avogadro's number** (6.02×10^{23}) of formula units of that substance. The mass of a mole of atoms, molecules, or ions (the **molar mass**) equals the formula weight of that material expressed in grams. The mass of one molecule of H_2O, for example, is 18 amu, so the mass of 1 mol of H_2O is 18 g. That is, the molar mass of H_2O is 18 g/mol.

Section 3.5 The empirical formula of any substance can be determined from its percent composition by calculating the relative number of moles of each atom in 100 g of the substance. If the substance is molecular in nature, its molecular formula can be determined from the empirical formula if the molecular weight is also known.

Sections 3.6 and 3.7 The mole concept can be used to calculate the relative quantities of reactants and products in chemical reactions. The coefficients in a balanced equation give the relative numbers of moles of the reactants and products. To calculate the number of grams of a product from the number of grams of a reactant, first convert grams of reactant to moles of reactant. Then use the coefficients in the balanced equation to convert the number of moles of reactant to moles of product. Finally, convert moles of product to grams of product.

A **limiting reactant** is completely consumed in a reaction. When it is used up, the reaction stops, thus limiting the quantities of products formed. The **theoretical yield** of a reaction is the quantity of product calculated to form when all of the limiting reagent reacts. The actual yield of a reaction is always less than the theoretical yield. The **percent yield** compares the actual and theoretical yields.

KEY SKILLS

- Balance chemical equations.
- Calculate molecular weights.
- Convert grams to moles and moles to grams using molar masses.
- Convert number of molecules to moles and moles to number of molecules using Avogadro's number.
- Calculate the empirical and molecular formula of a compound from percentage composition and molecular weight.
- Calculate amounts, in grams or moles, of reactants and products for a reaction.
- Calculate the percent yield of a reaction.

KEY EQUATIONS

- $$\% \text{ element} = \frac{\begin{pmatrix}\text{number of atoms}\\ \text{of that element}\end{pmatrix}\begin{pmatrix}\text{atomic weight}\\ \text{of element}\end{pmatrix}}{\text{formula weight of compound}} \times 100\% \qquad [3.10]$$

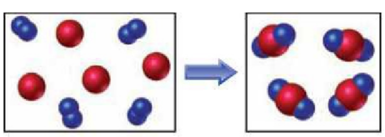
  This is the formula to calculate the mass percentage of each element in a compound. The sum of all the percentages of all the elements in a compound should add up to 100%.

- $$\% \text{ yield} = \frac{(\text{actual yield})}{(\text{theoretical yield})} \times 100\% \qquad [3.14]$$

  This is the formula to calculate the percent yield of a reaction. The percent yield can never be more than 100%.

VISUALIZING CONCEPTS

3.1 The reaction between reactant A (blue spheres) and reactant B (red spheres) is shown in the following diagram:



Based on this diagram, which equation best describes the reaction? [Section 3.1]

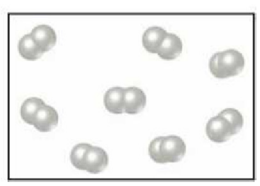
(a) $A_2 + B \longrightarrow A_2B$

(b) $A_2 + 4\,B \longrightarrow 2\,AB_2$

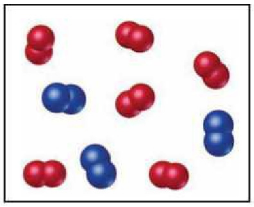
(c) $2\,A + B_4 \longrightarrow 2\,AB_2$

(d) $A + B_2 \longrightarrow AB_2$

3.2 Under appropriate experimental conditions, H_2 and CO undergo a combination reaction to form CH_3OH. The drawing below represents a sample of H_2. Make a corresponding drawing of the CO needed to react completely with the H_2. How did you arrive at the number of CO molecules in your drawing? [Section 3.2]

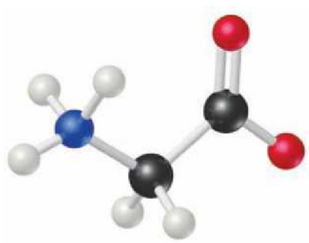


3.3 The following diagram represents the collection of elements formed by a decomposition reaction. **(a)** If the blue spheres represent N atoms and the red ones represent O atoms, what was the empirical formula of the original compound? **(b)** Could you draw a diagram representing the molecules of the compound that had been decomposed? Why or why not? [Section 3.2]

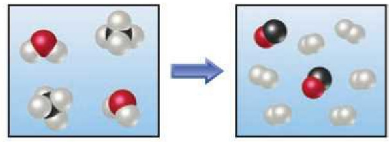


3.4 The following diagram represents the collection of CO_2 and H_2O molecules formed by complete combustion of a hydrocarbon. What is the empirical formula of the hydrocarbon? [Section 3.2]

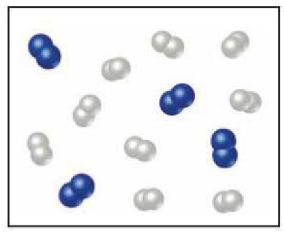
3.5 Glycine, an amino acid used by organisms to make proteins, is represented by the molecular model below.
(a) Write its molecular formula.
(b) Determine its molar mass.
(c) Calculate the mass of 3 moles of glycine.
(d) Calculate the percent nitrogen by mass in glycine. [Sections 3.3 and 3.5]



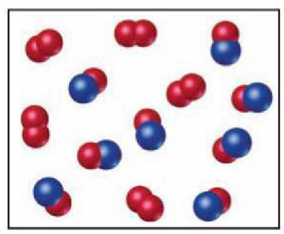
3.6 The following diagram represents a high-temperature reaction between CH_4 and H_2O. Based on this reaction, how many moles of each product can be obtained starting with 4.0 mol CH_4? [Section 3.6]



3.7 Nitrogen (N_2) and hydrogen (H_2) react to form ammonia (NH_3). Consider the mixture of N_2 and H_2 shown in the accompanying diagram. The blue spheres represent N, and the white ones represent H. Draw a representation of the product mixture, assuming that the reaction goes to completion. How did you arrive at your representation? What is the limiting reactant in this case? [Section 3.7]



3.8 Nitrogen monoxide and oxygen react to form nitrogen dioxide. Consider the mixture of NO and O_2 shown in the accompanying diagram. The blue spheres represent N, and the red ones represent O. **(a)** Draw a representation of the product mixture, assuming that the reaction goes to completion. What is the limiting reactant in this case? **(b)** How many NO_2 molecules would you draw as products if the reaction had a percent yield of 75%? [Section 3.7]



EXERCISES

Balancing Chemical Equations

3.9 **(a)** What scientific principle or law is used in the process of balancing chemical equations? **(b)** In balancing equations, why should you not change subscripts in chemical formulas? **(c)** How would one write out liquid water, water vapor, aqueous sodium chloride, and solid sodium chloride in chemical equations?

3.10 **(a)** What is the difference between adding a subscript 2 to the end of the formula for CO to give CO_2 and adding a coefficient in front of the formula to give 2 CO? **(b)** Is the following chemical equation, as written, consistent with the law of conservation of mass?

$$3\,Mg(OH)_2\,(s) + 2\,H_3PO_4(aq) \longrightarrow Mg_3(PO_4)_2(s) + 6\,H_2O(l)$$

Why or why not?

3.11 Balance the following equations:
(a) $CO(g) + O_2(g) \longrightarrow CO_2(g)$
(b) $N_2O_5(g) + H_2O(l) \longrightarrow HNO_3(aq)$
(c) $CH_4(g) + Cl_2(g) \longrightarrow CCl_4(l) + HCl(g)$

(d) $Al_4C_3(s) + H_2O(l) \longrightarrow Al(OH)_3(s) + CH_4(g)$
(e) $C_5H_{10}O_2(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
(f) $Fe(OH)_3(s) + H_2SO_4(aq) \longrightarrow Fe_2(SO_4)_3(aq) + H_2O(l)$
(g) $Mg_3N_2(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + (NH_4)_2SO_4(aq)$

3.12 Balance the following equations:
(a) $Li(s) + N_2(g) \longrightarrow Li_3N(s)$
(b) $La_2O_3(s) + H_2O(l) \longrightarrow La(OH)_3(aq)$
(c) $NH_4NO_3(s) \longrightarrow N_2(g) + O_2(g) + H_2O(g)$
(d) $Ca_3P_2(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq) + PH_3(g)$
(e) $Ca(OH)_2(aq) + H_3PO_4(aq) \longrightarrow Ca_3(PO_4)_2(s) + H_2O(l)$
(f) $AgNO_3(aq) + Na_2SO_4(aq) \longrightarrow Ag_2SO_4(s) + NaNO_3(aq)$
(g) $CH_3NH_2(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g) + N_2(g)$

3.13 Write balanced chemical equations to correspond to each of the following descriptions: **(a)** Solid calcium carbide, CaC_2, reacts with water to form an aqueous solution of calcium hydroxide and acetylene gas, C_2H_2. **(b)** When solid potassium chlorate is heated, it decomposes to form solid potassium chloride and oxygen gas. **(c)** Solid zinc metal reacts with sulfuric acid to form hydrogen gas and an aqueous solution of zinc sulfate. **(d)** When liquid phosphorus trichloride is added to water, it reacts to form aqueous phosphorous acid, $H_3PO_3(aq)$, and aqueous hydrochloric acid. **(e)** When hydrogen sulfide gas is passed over solid hot iron(III) hydroxide, the resultant reaction produces solid iron(III) sulfide and gaseous water.

3.14 Write balanced chemical equations to correspond to each of the following descriptions: **(a)** When sulfur trioxide gas reacts with water, a solution of sulfuric acid forms. **(b)** Boron sulfide, $B_2S_3(s)$, reacts violently with water to form dissolved boric acid, H_3BO_3, and hydrogen sulfide gas. **(c)** When an aqueous solution of lead(II) nitrate is mixed with an aqueous solution of sodium iodide, an aqueous solution of sodium nitrate and a yellow solid, lead iodide, are formed. **(d)** When solid mercury(II) nitrate is heated, it decomposes to form solid mercury(II) oxide, gaseous nitrogen dioxide, and oxygen. **(e)** Copper metal reacts with hot concentrated sulfuric acid solution to form aqueous copper(II) sulfate, sulfur dioxide gas, and water.

Patterns of Chemical Reactivity

3.15 **(a)** When the metallic element sodium combines with the nonmetallic element bromine, $Br_2(l)$, how can you determine the chemical formula of the product? How do you know whether the product is a solid, liquid, or gas at room temperature? Write the balanced chemical equation for the reaction. **(b)** When a hydrocarbon burns in air, what reactant besides the hydrocarbon is involved in the reaction? What products are formed? Write a balanced chemical equation for the combustion of benzene, $C_6H_6(l)$, in air.

3.16 **(a)** Determine the chemical formula of the product formed when the metallic element calcium combines with the nonmetallic element oxygen, O_2. Write the balanced chemical equation for the reaction. **(b)** What products form when a compound containing C, H, and O is completely combusted in air? Write a balanced chemical equation for the combustion of acetone, $C_3H_6O(l)$, in air.

3.17 Write a balanced chemical equation for the reaction that occurs when **(a)** $Mg(s)$ reacts with $Cl_2(g)$; **(b)** barium carbonate decomposes into barium oxide and carbon dioxide gas when heated; **(c)** the hydrocarbon styrene, $C_8H_8(l)$, is combusted in air; **(d)** dimethylether, $CH_3OCH_3(g)$, is combusted in air.

3.18 Write a balanced chemical equation for the reaction that occurs when **(a)** aluminum metal undergoes a combination reaction with $O_2(g)$; **(b)** copper(II) hydroxide decomposes into copper(II) oxide and water when heated; **(c)** heptane, $C_7H_{16}(l)$, burns in air; **(d)** the gasoline additive MTBE (methyl tert-butyl ether), $C_5H_{12}O(l)$, burns in air.

3.19 Balance the following equations, and indicate whether they are combination, decomposition, or combustion reactions:
(a) $Al(s) + Cl_2(g) \longrightarrow AlCl_3(s)$
(b) $C_2H_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
(c) $Li(s) + N_2(g) \longrightarrow Li_3N(s)$
(d) $PbCO_3(s) \longrightarrow PbO(s) + CO_2(g)$
(e) $C_7H_8O_2(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$

3.20 Balance the following equations, and indicate whether they are combination, decomposition, or combustion reactions:
(a) $C_3H_6(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
(b) $NH_4NO_3(s) \longrightarrow N_2O(g) + H_2O(g)$
(c) $C_5H_6O(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
(d) $N_2(g) + H_2(g) \longrightarrow NH_3(g)$
(e) $K_2O(s) + H_2O(l) \longrightarrow KOH(aq)$

Formula Weights

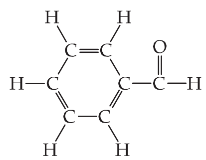
3.21 Determine the formula weights of each of the following compounds: **(a)** nitric acid, HNO_3; **(b)** $KMnO_4$; **(c)** $Ca_3(PO_4)_2$; **(d)** quartz, SiO_2; **(e)** gallium sulfide, **(f)** chromium(III) sulfate, **(g)** phosphorus trichloride.

3.22 Determine the formula weights of each of the following compounds: **(a)** nitrous oxide, N_2O, known as laughing gas and used as an anesthetic in dentistry; **(b)** benzoic acid, $HC_7H_5O_2$, a substance used as a food preservative; **(c)** $Mg(OH)_2$, the active ingredient in milk of magnesia; **(d)** urea, $(NH_2)_2CO$, a compound used as a nitrogen fertilizer; **(e)** isopentyl acetate, $CH_3CO_2C_5H_{11}$, responsible for the odor of bananas.

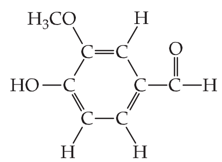
3.23 Calculate the percentage by mass of oxygen in the following compounds: **(a)** morphine, $C_{17}H_{19}NO_3$; **(b)** codeine, $C_{18}H_{21}NO_3$; **(c)** cocaine, $C_{17}H_{21}NO_4$; **(d)** tetracycline, $C_{22}H_{24}N_2O_8$; **(e)** digitoxin, $C_{41}H_{64}O_{13}$; **(f)** vancomycin, $C_{66}H_{75}Cl_2N_9O_{24}$.

3.24 Calculate the percentage by mass of the indicated element in the following compounds: **(a)** carbon in acetylene, C_2H_2, a gas used in welding; **(b)** hydrogen in ascorbic acid, $HC_6H_7O_6$, also known as vitamin C; **(c)** hydrogen in ammonium sulfate, $(NH_4)_2SO_4$, a substance used as a nitrogen fertilizer; **(d)** platinum in $PtCl_2(NH_3)_2$, a chemotherapy agent called cisplatin; **(e)** oxygen in the female sex hormone estradiol, $C_{18}H_{24}O_2$; **(f)** carbon in capsaicin, $C_{18}H_{27}NO_3$, the compound that gives the hot taste to chili peppers.

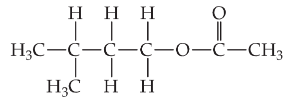
3.25 Based on the following structural formulas, calculate the percentage of carbon by mass present in each compound:

(a) 

Benzaldehyde (almond fragrance)

(b) 

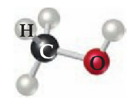
Vanillin (vanilla flavor)

(c) 

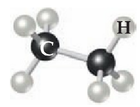
Isopentyl acetate (banana flavor)

3.26 Calculate the percentage of carbon by mass in each of the compounds represented by the following models:

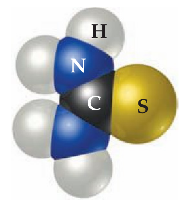
(a)



(b)



(c)



(d)

Avogadro's Number and the Mole

3.27 **(a)** What is Avogadro's number, and how is it related to the mole? **(b)** What is the relationship between the formula weight of a substance and its molar mass?

3.28 **(a)** What is the mass, in grams, of a mole of ^{12}C? **(b)** How many carbon atoms are present in a mole of ^{12}C?

3.29 Without doing any detailed calculations (but using a periodic table to give atomic weights), rank the following samples in order of increasing number of atoms: 0.50 mol H_2O, 23 g Na, 6.0×10^{23} N_2 molecules.

3.30 Without doing any detailed calculations (but using a periodic table to give atomic weights), rank the following samples in order of increasing number of atoms: 3.0×10^{23} molecules of H_2O_2, 2.0 mol CH_4, 32 g O_2.

3.31 What is the mass, in kilograms, of an Avogadro's number of people, if the average mass of a person is 160 lb? How does this compare with the mass of Earth, 5.98×10^{24} kg?

3.32 If Avogadro's number of pennies is divided equally among the 300 million men, women, and children in the United States, how many dollars would each receive? How does this compare with the gross domestic product of the United States, which was \$13.5 trillion in 2006? (The GDP is the total market value of the nation's goods and services.)

3.33 Calculate the following quantities:
(a) mass, in grams, of 0.105 moles sucrose ($C_{12}H_{22}O_{11}$)
(b) moles of $Zn(NO_3)_2$ in 143.50 g of this substance
(c) number of molecules in 1.0×10^{-6} mol CH_3CH_2OH
(d) number of N atoms in 0.410 mol NH_3

3.34 Calculate the following quantities
(a) mass, in grams, of 5.76×10^{-3} mol of CdS
(b) number of moles of NH_4Cl in 112.6 g of this substance
(c) number of molecules in 1.305×10^{-2} mol C_6H_6
(d) number of O atoms in 4.88×10^{-3} mol $Al(NO_3)_3$

3.35 **(a)** What is the mass, in grams, of 2.50×10^{-3} mol of ammonium phosphate?
(b) How many moles of chloride ions are in 0.2550 g of aluminum chloride?
(c) What is the mass, in grams, of 7.70×10^{20} molecules of caffeine, $C_8H_{10}N_4O_2$?
(d) What is the molar mass of cholesterol if 0.00105 mol weighs 0.406 g?

3.36 **(a)** What is the mass, in grams, of 0.0714 mol of iron(III) sulfate?
(b) How many moles of ammonium ions are in 8.776 g of ammonium carbonate?

(c) What is the mass, in grams, of 6.52×10^{21} molecules of aspirin, $C_9H_8O_4$?
(d) What is the molar mass of diazepam (Valium®) if 0.05570 mol weighs 15.86 g?

3.37 The molecular formula of allicin, the compound responsible for the characteristic smell of garlic, is $C_6H_{10}OS_2$. **(a)** What is the molar mass of allicin? **(b)** How many moles of allicin are present in 5.00 mg of this substance? **(c)** How many molecules of allicin are in 5.00 mg of this substance? **(d)** How many S atoms are present in 5.00 mg of allicin?

3.38 The molecular formula of aspartame, the artificial sweetener marketed as NutraSweet®, is $C_{14}H_{18}N_2O_5$. **(a)** What is the molar mass of aspartame? **(b)** How many moles of aspartame are present in 1.00 mg of aspartame? **(c)** How many molecules of aspartame are present in 1.00 mg of aspartame? **(d)** How many hydrogen atoms are present in 1.00 mg of aspartame?

3.39 A sample of glucose, $C_6H_{12}O_6$, contains 1.250×10^{21} carbon atoms. **(a)** How many atoms of hydrogen does it contain? **(b)** How many molecules of glucose does it contain? **(c)** How many moles of glucose does it contain? **(d)** What is the mass of this sample in grams?

3.40 A sample of the male sex hormone testosterone, $C_{19}H_{28}O_2$, contains 7.08×10^{20} hydrogen atoms. **(a)** How many atoms of carbon does it contain? **(b)** How many molecules of testosterone does it contain? **(c)** How many moles of testosterone does it contain? **(d)** What is the mass of this sample in grams?

3.41 The allowable concentration level of vinyl chloride, C_2H_3Cl, in the atmosphere in a chemical plant is 2.0×10^{-6} g/L. How many moles of vinyl chloride in each liter does this represent? How many molecules per liter?

3.42 At least 25 μg of tetrahydrocannabinol (THC), the active ingredient in marijuana, is required to produce intoxication. The molecular formula of THC is $C_{21}H_{30}O_2$. How many moles of THC does this 25 μg represent? How many molecules?

Empirical Formulas

3.43 Give the empirical formula of each of the following compounds if a sample contains **(a)** 0.0130 mol C, 0.0390 mol H, and 0.0065 mol O; **(b)** 11.66 g iron and 5.01 g oxygen; **(c)** 40.0% C, 6.7% H, and 53.3% O by mass.

3.44 Determine the empirical formula of each of the following compounds if a sample contains **(a)** 0.104 mol K, 0.052 mol C, and 0.156 mol O; **(b)** 5.28 g Sn and 3.37 g F; **(c)** 87.5% N and 12.5% H by mass.

3.45 Determine the empirical formulas of the compounds with the following compositions by mass:
(a) 10.4% C, 27.8% S, and 61.7% Cl
(b) 21.7% C, 9.6% O, and 68.7% F
(c) 32.79% Na, 13.02% Al, and 54.19% F

3.46 Determine the empirical formulas of the compounds with the following compositions by mass:
(a) 55.3% K, 14.6% P, and 30.1% O
(b) 24.5% Na, 14.9% Si, and 60.6% F
(c) 62.1% C, 5.21% H, 12.1% N, and 20.7% O

3.47 What is the molecular formula of each of the following compounds?
(a) empirical formula CH_2, molar mass = 84 g/mol
(b) empirical formula NH_2Cl, molar mass = 51.5 g/mol

3.48 What is the molecular formula of each of the following compounds?
(a) empirical formula HCO_2, molar mass = 90.0 g/mol
(b) empirical formula C_2H_4O, molar mass = 88 g/mol

3.49 Determine the empirical and molecular formulas of each of the following substances:
(a) Styrene, a compound substance used to make Styrofoam® cups and insulation, contains 92.3% C and 7.7% H by mass and has a molar mass of 104 g/mol.
(b) Caffeine, a stimulant found in coffee, contains 49.5% C, 5.15% H, 28.9% N, and 16.5% O by mass and has a molar mass of 195 g/mol.
(c) Monosodium glutamate (MSG), a flavor enhancer in certain foods, contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na, and has a molar mass of 169 g/mol.

3.50 Determine the empirical and molecular formulas of each of the following substances:
(a) Ibuprofen, a headache remedy, contains 75.69% C, 8.80% H, and 15.51% O by mass, and has a molar mass of 206 g/mol.
(b) Cadaverine, a foul smelling substance produced by the action of bacteria on meat, contains 58.55% C, 13.81% H, and 27.40% N by mass; its molar mass is 102.2 g/mol.
(c) Epinephrine (adrenaline), a hormone secreted into the bloodstream in times of danger or stress, contains 59.0% C, 7.1% H, 26.2% O, and 7.7% N by mass; its MW is about 180 amu.

3.51 **(a)** Combustion analysis of toluene, a common organic solvent, gives 5.86 mg of CO_2 and 1.37 mg of H_2O. If the compound contains only carbon and hydrogen, what is its empirical formula? **(b)** Menthol, the substance we can smell in mentholated cough drops, is composed of C, H, and O. A 0.1005-g sample of menthol is combusted, producing 0.2829 g of CO_2 and 0.1159 g of H_2O. What is the empirical formula for menthol? If menthol has a molar mass of 156 g/mol, what is its molecular formula?

3.52 **(a)** The characteristic odor of pineapple is due to ethyl butyrate, a compound containing carbon, hydrogen, and oxygen. Combustion of 2.78 mg of ethyl butyrate produces 6.32 mg of CO_2 and 2.58 mg of H_2O. What is the empirical formula of the compound? **(b)** Nicotine, a component of tobacco, is composed of C, H, and N. A 5.250-mg sample of nicotine was combusted, producing 14.242 mg of CO_2 and 4.083 mg of H_2O. What is the empirical formula for nicotine? If nicotine has a molar mass of 160 ± 5 g/mol, what is its molecular formula?

3.53 Washing soda, a compound used to prepare hard water for washing laundry, is a hydrate, which means that a certain number of water molecules are included in the solid structure. Its formula can be written as $Na_2CO_3 \cdot xH_2O$, where x is the number of moles of H_2O per mole of Na_2CO_3. When a 2.558-g sample of washing soda is heated at 25 °C, all the water of hydration is lost, leaving 0.948 g of Na_2CO_3. What is the value of x?

3.54 Epsom salts, a strong laxative used in veterinary medicine, is a hydrate, which means that a certain number of water molecules are included in the solid structure. The formula for Epsom salts can be written as $MgSO_4 \cdot xH_2O$, where x indicates the number of moles of H_2O per mole of $MgSO_4$. When 5.061 g of this hydrate is heated to 250 °C, all the water of hydration is lost, leaving 2.472 g of $MgSO_4$. What is the value of x?

Calculations Based on Chemical Equations

3.55 Why is it essential to use balanced chemical equations when determining the quantity of a product formed from a given quantity of a reactant?

3.56 What parts of balanced chemical equations give information about the relative numbers of moles of reactants and products involved in a reaction?

3.57 Hydrofluoric acid, HF(*aq*), cannot be stored in glass bottles because compounds called silicates in the glass are attacked by the HF(*aq*). Sodium silicate (Na_2SiO_3), for example, reacts as follows:

$$Na_2SiO_3(s) + 8\,HF(aq) \longrightarrow H_2SiF_6(aq) + 2\,NaF(aq) + 3\,H_2O(l)$$

(a) How many moles of HF are needed to react with 0.300 mol of Na_2SiO_3?
(b) How many grams of NaF form when 0.500 mol of HF reacts with excess Na_2SiO_3?
(c) How many grams of Na_2SiO_3 can react with 0.800 g of HF?

3.58 The fermentation of glucose ($C_6H_{12}O_6$) produces ethyl alcohol (C_2H_5OH) and CO_2:

$$C_6H_{12}O_6(aq) \longrightarrow 2\,C_2H_5OH(aq) + 2\,CO_2(g)$$

(a) How many moles of CO_2 are produced when 0.400 mol of $C_6H_{12}O_6$ reacts in this fashion?
(b) How many grams of $C_6H_{12}O_6$ are needed to form 7.50 g of C_2H_5OH?
(c) How many grams of CO_2 form when 7.50 g of C_2H_5OH are produced?

3.59 Several brands of antacids use $Al(OH)_3$ to react with stomach acid, which contains primarily HCl:

$$Al(OH)_3(s) + HCl(aq) \longrightarrow AlCl_3(aq) + H_2O(l)$$

(a) Balance this equation.
(b) Calculate the number of grams of HCl that can react with 0.500 g of $Al(OH)_3$.
(c) Calculate the number of grams of $AlCl_3$ and the number of grams of H_2O formed when 0.500 g of $Al(OH)_3$ reacts.
(d) Show that your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

3.60 An iron ore sample contains Fe_2O_3 together with other substances. Reaction of the ore with CO produces iron metal:

$$Fe_2O_3(s) + CO(g) \longrightarrow Fe(s) + CO_2(g)$$

(a) Balance this equation.
(b) Calculate the number of grams of CO that can react with 0.150 kg of Fe_2O_3.
(c) Calculate the number of grams of Fe and the number of grams of CO_2 formed when 0.150 kg of Fe_2O_3 reacts.
(d) Show that your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

3.61 Aluminum sulfide reacts with water to form aluminum hydroxide and hydrogen sulfide. **(a)** Write the balanced chemical equation for this reaction. **(b)** How many grams of aluminum hydroxide are obtained from 14.2 g of aluminum sulfide?

3.62 Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas. **(a)** Write a balanced chemical equation for the reaction. **(b)** How many grams of calcium hydride are needed to form 8.500 g of hydrogen?

3.63 Automotive air bags inflate when sodium azide, NaN_3, rapidly decomposes to its component elements:

$$2\,NaN_3(s) \longrightarrow 2\,Na(s) + 3\,N_2(g)$$

(a) How many moles of N_2 are produced by the decomposition of 1.50 mol of NaN_3?
(b) How many grams of NaN_3 are required to form 10.0 g of nitrogen gas?
(c) How many grams of NaN_3 are required to produce 10.0 ft^3 of nitrogen gas, about the size of an automotive air bag, if the gas has a density of 1.25 g/L?

3.64 The complete combustion of octane, C_8H_{18}, the main component of gasoline, proceeds as follows:

$$2\,C_8H_{18}(l) + 25\,O_2(g) \longrightarrow 16\,CO_2(g) + 18\,H_2O(g)$$

(a) How many moles of O_2 are needed to burn 1.25 mol of C_8H_{18}?
(b) How many grams of O_2 are needed to burn 10.0 g of C_8H_{18}?
(c) Octane has a density of 0.692 g/mL at 20 °C. How many grams of O_2 are required to burn 1.00 gal of C_8H_{18}?

3.65 A piece of aluminum foil 1.00 cm square and 0.550 mm thick is allowed to react with bromine to form aluminum bromide as shown in the accompanying photo.

(a) How many moles of aluminum were used? (The density of aluminum is 2.699 g/cm^3.) **(b)** How many grams of aluminum bromide form, assuming the aluminum reacts completely?

3.66 Detonation of nitroglycerin proceeds as follows:

$$4\,C_3H_5N_3O_9(l) \longrightarrow 12\,CO_2(g) + 6\,N_2(g) + O_2(g) + 10\,H_2O(g)$$

(a) If a sample containing 2.00 mL of nitroglycerin (density = 1.592 g/mL) is detonated, how many total moles of gas are produced? **(b)** If each mole of gas occupies 55 L under the conditions of the explosion, how many liters of gas are produced? **(c)** How many grams of N_2 are produced in the detonation?

Limiting Reactants; Theoretical Yields

3.67 **(a)** Define the terms *limiting reactant* and *excess reactant*. **(b)** Why are the amounts of products formed in a reaction determined only by the amount of the limiting reactant? **(c)** Why should you base your choice of what compound is the limiting reactant on its number of initial moles, not on its initial mass in grams?

3.68 **(a)** Define the terms *theoretical yield*, *actual yield*, and *percent yield*. **(b)** Why is the actual yield in a reaction almost always less than the theoretical yield? **(c)** Can a reaction ever have 110% actual yield?

3.69 A manufacturer of bicycles has 4815 wheels, 2305 frames, and 2255 handlebars. **(a)** How many bicycles can be manufactured using these parts? **(b)** How many parts of each kind are left over? **(c)** Which part limits the production of bicycles?

3.70 A bottling plant has 121,515 bottles with a capacity of 355 mL, 122,500 caps, and 40,875 L of beverage. **(a)** How many bottles can be filled and capped? **(b)** How much of each item is left over? **(c)** Which component limits the production?

3.71 Sodium hydroxide reacts with carbon dioxide as follows:

$$2\,NaOH(s) + CO_2(g) \longrightarrow Na_2CO_3(s) + H_2O(l)$$

Which reagent is the limiting reactant when 1.85 mol NaOH and 1.00 mol CO_2 are allowed to react? How many moles of Na_2CO_3 can be produced? How many moles of the excess reactant remain after the completion of the reaction?

3.72 Aluminum hydroxide reacts with sulfuric acid as follows:

$$2\,Al(OH)_3(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 6\,H_2O(l)$$

Which reagent is the limiting reactant when 0.500 mol $Al(OH)_3$ and 0.500 mol H_2SO_4 are allowed to react? How many moles of $Al_2(SO_4)_3$ can form under these conditions? How many moles of the excess reactant remain after the completion of the reaction?

3.73 The fizz produced when an Alka-Seltzer® tablet is dissolved in water is due to the reaction between sodium bicarbonate ($NaHCO_3$) and citric acid ($H_3C_6H_5O_7$):

$$3\,NaHCO_3(aq) + H_3C_6H_5O_7(aq) \longrightarrow 3\,CO_2(g) + 3\,H_2O(l) + Na_3C_6H_5O_7(aq)$$

In a certain experiment 1.00 g of sodium bicarbonate and 1.00 g of citric acid are allowed to react. **(a)** Which is the limiting reactant? **(b)** How many grams of carbon dioxide form? **(c)** How many grams of the excess reactant remain after the limiting reactant is completely consumed?

3.74 One of the steps in the commercial process for converting ammonia to nitric acid is the conversion of NH_3 to NO:

$$4\,NH_3(g) + 5\,O_2(g) \longrightarrow 4\,NO(g) + 6\,H_2O(g)$$

In a certain experiment, 1.50 g of NH_3 reacts with 2.75 g of O_2 **(a)** Which is the limiting reactant? **(b)** How many grams of NO and of H_2O form? **(c)** How many grams of the excess reactant remain after the limiting reactant is completely consumed? **(d)** Show that your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

3.75 Solutions of sodium carbonate and silver nitrate react to form solid silver carbonate and a solution of sodium nitrate. A solution containing 3.50 g of sodium carbonate is mixed with one containing 5.00 g of silver nitrate. How many grams of sodium carbonate, silver nitrate, silver carbonate, and sodium nitrate are present after the reaction is complete?

3.76 Solutions of sulfuric acid and lead(II) acetate react to form solid lead(II) sulfate and a solution of acetic acid. If 7.50 g of sulfuric acid and 7.50 g of lead(II) acetate are mixed, calculate the number of grams of sulfuric acid, lead(II) acetate, lead(II) sulfate, and acetic acid present in the mixture after the reaction is complete.

3.77 When benzene (C_6H_6) reacts with bromine (Br_2), bromobenzene (C_6H_5Br) is obtained:

$$C_6H_6 + Br_2 \longrightarrow C_6H_5Br + HBr$$

(a) What is the theoretical yield of bromobenzene in this reaction when 30.0 g of benzene reacts with 65.0 g of bromine? **(b)** If the actual yield of bromobenzene was 42.3 g, what was the percentage yield?

3.78 When ethane (C_2H_6) reacts with chlorine (Cl_2), the main product is C_2H_5Cl; but other products containing Cl, such as $C_2H_4Cl_2$, are also obtained in small quantities. The formation of these other products reduces the yield of C_2H_5Cl. **(a)** Calculate the theoretical yield of C_2H_5Cl when 125 g of C_2H_6 reacts with 255 g of Cl_2, assuming that C_2H_6 and Cl_2 react only to form C_2H_5Cl and HCl. **(b)** Calculate the percent yield of C_2H_5Cl if the reaction produces 206 g of C_2H_5Cl.

3.79 Hydrogen sulfide is an impurity in natural gas that must be removed. One common removal method is called the Claus process, which relies on the reaction:

$$8\,H_2S(g) + 4\,O_2(g) \longrightarrow S_8(l) + 8\,H_2O(g)$$

Under optimal conditions the Claus process gives 98% yield of S_8 from H_2S. If you started with 30.0 grams of H_2S and 50.0 grams of O_2, how many grams of S_8 would be produced, assuming 98% yield?

3.80 When hydrogen sulfide gas is bubbled into a solution of sodium hydroxide, the reaction forms sodium sulfide and water. How many grams of sodium sulfide are formed if 1.50 g of hydrogen sulfide is bubbled into a solution containing 2.00 g of sodium hydroxide, assuming that the sodium sulfide is made in 92.0% yield?

ADDITIONAL EXERCISES

3.81 Write the balanced chemical equations for **(a)** the complete combustion of acetic acid (CH_3COOH), the main active ingredient in vinegar; **(b)** the decomposition of solid calcium hydroxide into solid calcium(II) oxide (lime) and water vapor; **(c)** the combination reaction between nickel metal and chlorine gas.

3.82 The effectiveness of nitrogen fertilizers depends on both their ability to deliver nitrogen to plants and the amount of nitrogen they can deliver. Four common nitrogen-containing fertilizers are ammonia, ammonium nitrate, ammonium sulfate, and urea [$(NH_2)_2CO$]. Rank these fertilizers in terms of the mass percentage nitrogen they contain.

3.83 **(a)** Diamond is a natural form of pure carbon. How many moles of carbon are in a 1.25-carat diamond (1 carat = 0.200 g)? How many atoms are in this diamond? **(b)** The molecular formula of acetylsalicylic acid (aspirin), one of the most common pain relievers, is $C_9H_8O_4$. How many moles of $C_9H_8O_4$ are in a 0.500-g tablet of aspirin? How many molecules of $C_9H_8O_4$ are in this tablet?

3.84 **(a)** One molecule of the antibiotic known as penicillin G has a mass of 5.342×10^{-21} g. What is the molar mass of penicillin G? **(b)** Hemoglobin, the oxygen-carrying protein in red blood cells, has four iron atoms per molecule and contains 0.340% iron by mass. Calculate the molar mass of hemoglobin.

3.85 Very small crystals composed of 1000 to 100,000 atoms, called quantum dots, are being investigated for use in electronic devices.

(a) A quantum dot was made of solid silicon in the shape of a sphere, with a diameter of 4 nm. Calculate the mass of the quantum dot, using the density of silicon ($2.3\ g/cm^3$).

(b) How many silicon atoms are in the quantum dot?

(c) The density of germanium is $5.325\ g/cm^3$. If you made a 4 nm quantum dot of germanium, how many Ge atoms would it contain? Assume the dot is spherical.

3.86 Serotonin is a compound that conducts nerve impulses in the brain. It contains 68.2 mass percent C, 6.86 mass percent H, 15.9 mass percent N, and 9.08 mass percent O. Its molar mass is 176 g/mol. Determine its molecular formula.

3.87 The koala dines exclusively on eucalyptus leaves. Its digestive system detoxifies the eucalyptus oil, a poison to other animals. The chief constituent in eucalyptus oil is a substance called eucalyptol, which contains 77.87% C, 11.76% H, and the remainder O. **(a)** What is the empirical formula for this substance? **(b)** A mass spectrum of eucalyptol shows a peak at about 154 amu. What is the molecular formula of the substance?

3.88 Vanillin, the dominant flavoring in vanilla, contains C, H, and O. When 1.05 g of this substance is completely combusted, 2.43 g of CO_2 and 0.50 g of H_2O are produced. What is the empirical formula of vanillin?

[3.89] An organic compound was found to contain only C, H, and Cl. When a 1.50-g sample of the compound was completely combusted in air, 3.52 g of CO_2 was formed. In a separate experiment the chlorine in a 1.00-g sample of the compound was converted to 1.27 g of AgCl. Determine the empirical formula of the compound.

[3.90] An oxybromate compound, $KBrO_x$, where x is unknown, is analyzed and found to contain 52.92% Br. What is the value of x?

[3.91] An element X forms an iodide (XI_3) and a chloride (XCl_3). The iodide is quantitatively converted to the chloride when it is heated in a stream of chlorine:

$$2\,XI_3 + 3\,Cl_2 \longrightarrow 2\,XCl_3 + 3\,I_2$$

If 0.5000 g of XI_3 is treated, 0.2360 g of XCl_3 is obtained. **(a)** Calculate the atomic weight of the element X. **(b)** Identify the element X.

3.92 If 1.5 mol of each of the following compounds is completely combusted in oxygen, which one will produce the largest number of moles of H_2O? Which will produce the least? Explain. C_2H_5OH, C_3H_8, $CH_3CH_2COCH_3$.

3.93 A method used by the U.S. Environmental Protection Agency (EPA) for determining the concentration of ozone in air is to pass the air sample through a "bubbler" containing sodium iodide, which removes the ozone according to the following equation:

$$O_3(g) + 2\,NaI(aq) + H_2O(l) \longrightarrow O_2(g) + I_2(s) + 2\,NaOH(aq)$$

(a) How many moles of sodium iodide are needed to remove 5.95×10^{-6} mol of O_3? **(b)** How many grams of sodium iodide are needed to remove 1.3 mg of O_3?

3.94 A chemical plant uses electrical energy to decompose aqueous solutions of NaCl to give Cl_2, H_2, and NaOH:

$$2\,NaCl(aq) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g) + Cl_2(g)$$

If the plant produces 1.5×10^6 kg (1500 metric tons) of Cl_2 daily, estimate the quantities of H_2 and NaOH produced.

3.95 The fat stored in the hump of a camel is a source of both energy and water. Calculate the mass of H_2O produced by metabolism of 1.0 kg of fat, assuming the fat consists entirely of tristearin ($C_{57}H_{110}O_6$), a typical animal fat, and assuming that during metabolism, tristearin reacts with O_2 to form only CO_2 and H_2O.

[3.96] When hydrocarbons are burned in a limited amount of air, both CO and CO_2 form. When 0.450 g of a particular hydrocarbon was burned in air, 0.467 g of CO, 0.733 g of CO_2, and 0.450 g of H_2O were formed. **(a)** What is the empirical formula of the compound? **(b)** How many grams of O_2 were used in the reaction? **(c)** How many grams would have been required for complete combustion?

3.97 A mixture of $N_2(g)$ and $H_2(g)$ reacts in a closed container to form ammonia, $NH_3(g)$. The reaction ceases before either reactant has been totally consumed. At this stage 3.0 mol N_2, 3.0 mol H_2, and 3.0 mol NH_3 are present. How many moles of N_2 and H_2 were present originally?

[3.98] A mixture containing $KClO_3$, K_2CO_3, $KHCO_3$, and KCl was heated, producing CO_2, O_2, and H_2O gases according to the following equations:

$$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g)$$
$$2\,KHCO_3(s) \longrightarrow K_2O(s) + H_2O(g) + 2\,CO_2(g)$$
$$K_2CO_3(s) \longrightarrow K_2O(s) + CO_2(g)$$

The KCl does not react under the conditions of the reaction. If 100.0 g of the mixture produces 1.80 g of H_2O, 13.20 g of CO_2, and 4.00 g of O_2, what was the composition of the original mixture? (Assume complete decomposition of the mixture.)

3.99 When a mixture of 10.0 g of acetylene (C_2H_2) and 10.0 g of oxygen (O_2) is ignited, the resultant combustion reaction produces CO_2 and H_2O. **(a)** Write the balanced chemical equation for this reaction. **(b)** Which is the limiting reactant? **(c)** How many grams of C_2H_2, O_2, CO_2, and H_2O are present after the reaction is complete?

3.100 Aspirin ($C_9H_8O_4$) is produced from salicylic acid ($C_7H_6O_3$) and acetic anhydride ($C_4H_6O_3$):

$$C_7H_6O_3 + C_4H_6O_3 \longrightarrow C_9H_8O_4 + HC_2H_3O_2$$

(a) How much salicylic acid is required to produce 1.5×10^2 kg of aspirin, assuming that all of the salicylic acid is converted to aspirin? **(b)** How much salicylic acid would be required if only 80% of the salicylic acid is converted to aspirin? **(c)** What is the theoretical yield of aspirin if 185 kg of salicylic acid is allowed to react with 125 kg of acetic anhydride? **(d)** If the situation described in part (c) produces 182 kg of aspirin, what is the percentage yield?

INTEGRATIVE EXERCISES

(These exercises require skills from earlier chapters as well as skills from the present chapter.)

3.101 Consider a sample of calcium carbonate in the form of a cube measuring 2.005 in. on each edge. If the sample has a density of 2.71 g/cm^3, how many oxygen atoms does it contain?

3.102 **(a)** You are given a cube of silver metal that measures 1.000 cm on each edge. The density of silver is 10.5 g/cm^3. How many atoms are in this cube? **(b)** Because atoms are spherical, they cannot occupy all of the space of the cube. The silver atoms pack in the solid in such a way that 74% of the volume of the solid is actually filled with the silver atoms. Calculate the volume of a single silver atom. **(c)** Using the volume of a silver atom and the formula for the volume of a sphere, calculate the radius in angstroms of a silver atom.

3.103 **(a)** If an automobile travels 225 mi with a gas mileage of 20.5 mi/gal, how many kilograms of CO_2 are produced? Assume that the gasoline is composed of octane, $C_8H_{18}(l)$, whose density is 0.69 g/mL. **(b)** Repeat the calculation for a truck that has a gas mileage of 5 mi/gal.

3.104 In 1865 a chemist reported that he had reacted a weighed amount of pure silver with nitric acid and had recovered all the silver as pure silver nitrate. The mass ratio of silver to silver nitrate was found to be 0.634985. Using only this ratio and the presently accepted values for the atomic weights of silver and oxygen, calculate the atomic weight of nitrogen. Compare this calculated atomic weight with the currently accepted value.

3.105 A particular coal contains 2.5% sulfur by mass. When this coal is burned at a power plant, the sulfur is converted into sulfur dioxide gas, which is a pollutant. To reduce sulfur dioxide emissions, calcium oxide (lime) is used. The sulfur dioxide reacts with calcium oxide to form solid calcium sulfite. **(a)** Write the balanced chemical equation for the reaction. **(b)** If the coal is burned in a power plant that uses 2000 tons of coal per day, what mass of calcium oxide is required daily to eliminate the sulfur dioxide? **(c)** How many grams of calcium sulfite are produced daily by this power plant?

3.106 Copper is an excellent electrical conductor widely used in making electric circuits. In producing a printed circuit board for the electronics industry, a layer of copper is laminated on a plastic board. A circuit pattern is then printed on the board using a chemically resistant polymer. The board is then exposed to a chemical bath that reacts with the exposed copper, leaving the desired copper circuit, which has been protected by the overlaying polymer. Finally, a solvent removes the polymer. One reaction used to remove the exposed copper from the circuit board is

$$Cu(s) + Cu(NH_3)_4Cl_2(aq) + 4\,NH_3(aq) \longrightarrow 2\,Cu(NH_3)_4Cl(aq)$$

A plant needs to produce 5000 circuit boards, each with a surface area measuring 2.0 in. $\times$ 3.0 in. The boards are covered with a 0.65-mm layer of copper. In subsequent processing, 85% of the copper is removed. Copper has a density of 8.96 g/cm^3. Calculate the masses of $Cu(NH_3)_4Cl_2$ and NH_3 needed to produce the circuit boards, assuming that the reaction used gives a 97% yield.

3.107 Hydrogen cyanide, HCN, is a poisonous gas. The lethal dose is approximately 300 mg HCN per kilogram of air when inhaled. **(a)** Calculate the amount of HCN that gives the lethal dose in a small laboratory room measuring 12 × 15 × 8.0 ft. The density of air at 26 °C is 0.00118 g/cm^3. **(b)** If the HCN is formed by reaction of NaCN with an acid such as H_2SO_4, what mass of NaCN gives the lethal dose in the room?

$$2\,NaCN(s) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2\,HCN(g)$$

(c) HCN forms when synthetic fibers containing Orlon® or Acrilan® burn. Acrilan® has an empirical formula of CH_2CHCN, so HCN is 50.9% of the formula by mass. A rug measures 12 × 15 ft and contains 30 oz of Acrilan® fibers per square yard of carpet. If the rug burns, will a lethal dose of HCN be generated in the room? Assume that the yield of HCN from the fibers is 20% and that the carpet is 50% consumed.

3.108 The source of oxygen that drives the internal combustion engine in an automobile is air. Air is a mixture of gases, which are principally N_2 (~79%) and O_2 (~20%). In the cylinder of an automobile engine, nitrogen can react with oxygen to produce nitric oxide gas, NO. As NO is emitted from the tailpipe of the car, it can react with more oxygen to produce nitrogen dioxide gas. (**a**) Write balanced chemical equations for both reactions. (**b**) Both nitric oxide and nitrogen dioxide are pollutants that can lead to acid rain and global warming; collectively, they are called "NO_x" gases. In 2004, the United States emitted an estimated 19 million tons of nitrogen dioxide into the atmosphere. How many grams of nitrogen dioxide is this? **(c)** The production of NO_x gases is an unwanted side reaction of the main engine combustion process that turns octane, C_8H_{18}, into CO_2 and water. If 85% of the oxygen in an engine is used to combust octane, and the remainder used to produce nitrogen dioxide, calculate how many grams of nitrogen dioxide would be produced during the combustion of 500 grams of octane.