3

Chemical Reactions and Reaction Stoichiometry

Have you ever poured vinegar into a vessel containing baking soda? If so, you know the result is an immediate and effervescent cascade of bubbles. The bubbles contain carbon dioxide gas that is produced by the chemical reaction between sodium bicarbonate in the baking soda and acetic acid in the vinegar.

The bubbles released when baking soda reacts with an acid play an important role in baking, where the release of gaseous CO_2 causes the dough in your biscuits or the batter in your pancakes to rise. An alternative way to produce CO_2 in cooking is to use yeasts that rely on chemical reactions to convert sugar into CO_2, ethanol, and other organic compounds. These types of chemical reactions have been used for thousands of years in the baking of breads as well as in the production of alcoholic beverages like beer and wine. Chemical reactions that produce CO_2 are not limited to cooking, though—they occur in places as diverse as the cells in your body and the engine of your car.

In this chapter we explore some important aspects of chemical reactions. 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 those 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. Stoichiometry (Greek *stoicheion*, "element," and *metron*, "measure") provides an essential set of tools widely used in chemistry, including such diverse applications as measuring ozone concentrations in the atmosphere and assessing different processes for converting coal into gaseous fuels.

▶ **THE TEXTURE AND FLAVORS** of bread and beer are dependent on chemical reactions that occur when yeasts ferment sugars to produce carbon dioxide and ethanol.

WHAT'S AHEAD ▶

3.1 CHEMICAL EQUATIONS We begin by considering how we can use chemical formulas to write equations representing chemical reactions.

3.2 SIMPLE PATTERNS OF CHEMICAL REACTIVITY We then examine some simple chemical reactions: *combination reactions*, *decomposition reactions*, and *combustion reactions*.

3.3 FORMULA WEIGHTS We see how to obtain quantitative information from chemical formulas by using *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 the *mole*, defined as 6.022×10^{23} objects (atoms, molecules, ions, and so on).

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 reactant may be used up before others in a chemical reaction. This is the *limiting reactant*. When this happens the reaction stops, leaving some excess of the other starting materials.

▲ **Figure 3.1 Antoine Lavoisier (1734–1794).** The science career of Lavoisier, who conducted many important studies on combustion reactions, was cut short by the French Revolution. Guillotined in 1794 during the Reign of Terror, he is generally considered the father of modern chemistry because he conducted carefully controlled experiments and used quantitative measurements.

Stoichiometry is built on an understanding of atomic masses (Section 2.4), chemical formulas, and the **law of conservation of mass.** (Section 2.1) The French nobleman and scientist Antoine Lavoisier (◀ Figure 3.1) discovered this important chemical law during the late 1700s. 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 came to understand the basis for this law: *Atoms are neither created nor destroyed during a 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.

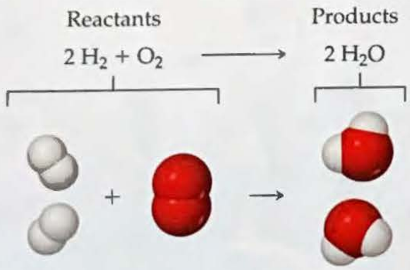
3.1 | Chemical Equations

We represent chemical reactions 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

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

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, called coefficients, indicate the relative numbers of molecules of each kind involved in the reaction. (As in algebraic equations, *the coefficient 1 is usually not written.*)

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 *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 (◀ Figure 3.2). 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 each subscript in a chemical formula by the coefficient for the formula.* Because there are four H atoms and two O atoms on each side of the equation, the equation is balanced.



▲ **Figure 3.2 A balanced chemical equation.**

Give It Some Thought

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

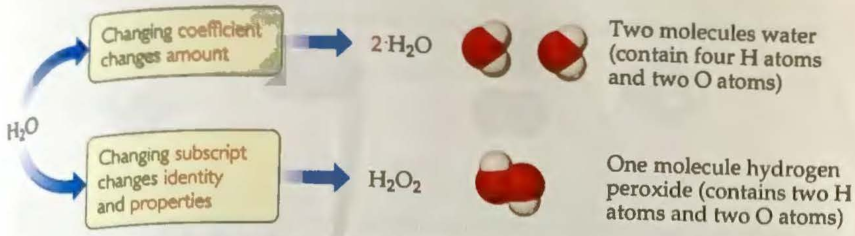
Balancing Equations

To construct a balanced chemical equation we start by writing the formulas for the reactants on the left-hand side of the arrow and the products on the right-hand side. Next we balance the equation by determining the coefficients that provide equal numbers of each type of atom on both sides 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 coefficients and subscripts. As ▶ Figure 3.3 illustrates, changing a subscript in a formula—from H_2O to H_2O_2, for example—changes the identity of the substance. The substance H_2O_2, hydrogen peroxide, is quite different from the substance H_2O, water. *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

*Lavoisier, Antoine. "Elements of Chemistry." 1790.



▲ **Figure 3.3 The difference between changing subscripts and changing coefficients in chemical equations.**

carbon dioxide gas (CO_2) and water vapor (H_2O) (▼ **Figure 3.4**). Both 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 in the equation. In our example, C appears in only one reactant (CH_4) and one product (CO_2). The same is true for H (CH_4 and H_2O). Notice, however, that O appears in one reactant (O_2) and two products (CO_2 and H_2O). So, let's begin with C. Because one molecule of CH_4 contains the same number of C atoms (one) as one molecule of CO_2, the coefficients for these substances *must* be the same in the balanced equation. Therefore, we start by choosing the coefficient 1 (unwritten) for both CH_4 and CO_2.

Next we focus on H. On the left side of the equation we have CH_4, which has four H atoms, whereas on the right side of the equation we have H_2O, containing two H atoms. To balance the H atoms in the equation we place the coefficient 2 in front of H_2O. Now there are four H atoms on each side of the equation:

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

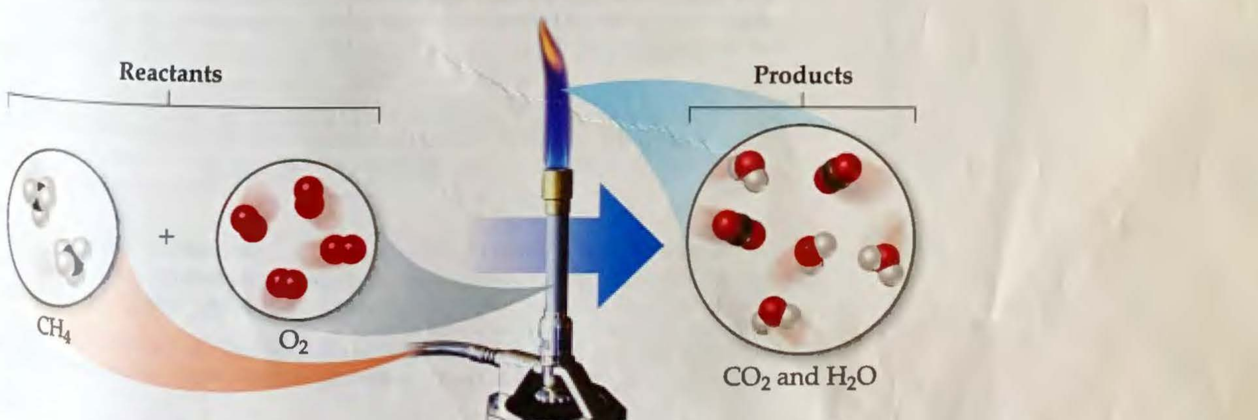
While the equation is now balanced with respect to hydrogen and carbon, it is not yet balanced for oxygen. Adding the coefficient 2 in front of O_2 balances the equation by giving four O atoms on each side (2×2 left, $2 + 2 \times 1$ right):

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

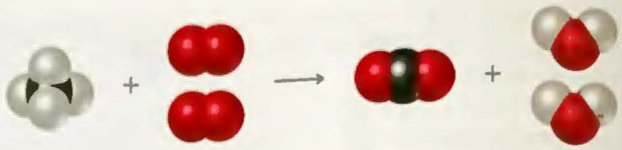
The molecular view of the balanced equation is shown in **Figure 3.5**.

GO FIGURE

In the molecular level views shown in the figure how many C, H, and O atoms are present on the reactant side? Are the same number of each type of atom present on the product side?



▲ **Figure 3.4 Methane reacts with oxygen in a Bunsen burner.**

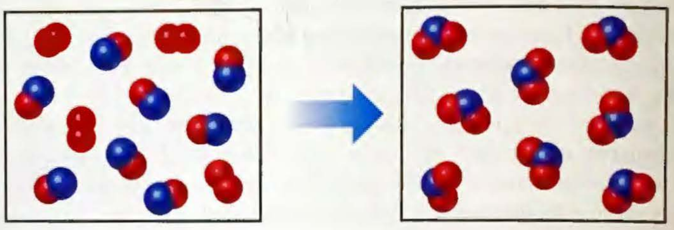


$$\underbrace{CH_4 \quad + \quad 2\,O_2}_{1\,C,\,4\,H,\,4\,O} \longrightarrow \underbrace{CO_2 \quad + \quad 2\,H_2O}_{1\,C,\,4\,H,\,4\,O}$$

▲ **Figure 3.5 Balanced chemical equation for the combustion of CH_4.**

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 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 products, contains only one kind of molecule, which is composed of one nitrogen atom and two oxygen atoms (NO_2).

(b) The unbalanced chemical equation is

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

An inventory of atoms on each side of the equation shows that there are one N and three O on the left side of the arrow and one N and two O on the right. To balance O we must increase the number of O atoms on the right while keeping the coefficients for NO and NO_2 equal. Sometimes a trial-and-error approach is required; we need to go back and forth several times from one side of an equation to the other, changing coefficients first on one side of the equation and then the other until it is balanced. In our present case, let's start by increasing the number of O atoms on the right side of the equation by placing the coefficient 2 in front of NO_2:

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

Now the equation gives two N atoms and four O atoms on the right, so we go back to the left side. Placing the coefficient 2 in front of NO balances both N and O:

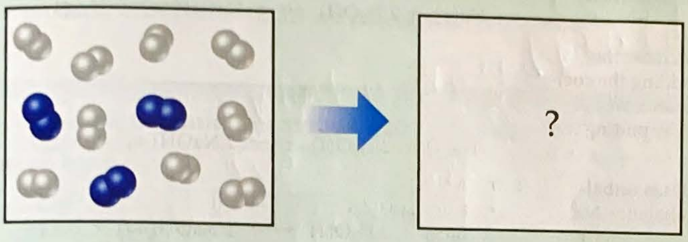
$$\underbrace{O_2 + 2\,NO}_{2\,N,\,4\,O} \longrightarrow \underbrace{2\,NO_2}_{2\,N,\,4\,O} \quad \text{(balanced)}$$

(c) The reactants box contains four O_2 and eight NO. Thus, the molecular ratio is one O_2 for each two NO, as required by the balanced equation. The products box contains eight NO_2, which means the number of NO_2 product molecules equals the number of NO reactant molecules, as the balanced equation requires.

There are eight N atoms in the eight NO molecules in the reactants box. There are also $4 \times 2 = 8$ O atoms in the O_2 molecules and 8 O atoms in the NO molecules, giving a total of 16 O atoms. In the products box, we find eight NO_2 molecules, which contain eight N atoms and $8 \times 2 = 16$ O atoms. Because there are equal numbers of N and O atoms in the two boxes, the drawing is consistent with the law of conservation of mass.

Practice Exercise 1

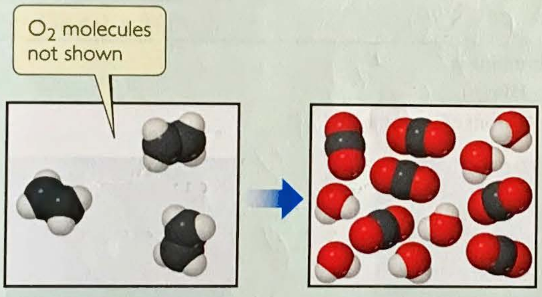
In the following diagram, the white spheres represent hydrogen atoms and the blue spheres represent nitrogen atoms.



The two reactants combine to form a single product, ammonia, NH_3, which is not shown. Write a balanced chemical equation for the reaction. Based on the equation and the contents of the left (reactants) box, find how many NH_3 molecules should be shown in the right (products) box. **(a)** 2, **(b)** 3, **(c)** 4, **(d)** 6, **(e)** 9.

Practice Exercise 2

In the following diagram, the white spheres represent hydrogen atoms, the black spheres carbon atoms, and the red spheres oxygen atoms.



In this reaction, there are two reactants, ethylene, C_2H_4, which is shown, and oxygen, O_2, which is not shown, and two products, CO_2 and H_2O, both of which are shown. **(a)** Write a balanced chemical equation for the reaction. **(b)** Determine the number of O_2 molecules that should be shown in the left (reactants) box.

Indicating the States of Reactants and Products

Symbols indicating the physical state of each reactant and product are often shown in chemical equations. We use the symbols (g), (l), (s), and (aq) for substances that are gases, liquids, solids, and dissolved in 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 symbols that represent the conditions under which the reaction proceeds appear above or below the reaction arrow. One example that we will encounter later in this chapter involves the symbol Δ (Greek uppercase delta); a cap delta above the reaction arrow indicates the addition of heat.

SAMPLE EXERCISE 3.2 Balancing Chemical Equations

Balance the equation

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

SOLUTION

Begin by counting each kind of atom on the two sides of the arrow. There are one Na, one O, and two H on the left side, and one Na, one O, and three H on the right. The Na and O atoms are balanced, but the number of H atoms is not. To increase the number of H atoms on the left, let's try placing the coefficient 2 in front of H_2O:

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

Although beginning this way does not balance H, it does increase the number of reactant H atoms, which we need to do. (Also, adding the coefficient 2 on H_2O unbalances O, but we will take care of that after we balance H.) Now that we have 2 H_2O on the left, we balance H by putting the coefficient 2 in front of NaOH:

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

Balancing H in this way brings O into balance, but now Na is 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:

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

We now have two Na atoms, four H atoms, and two O atoms on each side. The equation is balanced.

Comment Notice that 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 by checking that the number of atoms of each element is the same on the two sides of the arrow, and that you've chosen the smallest set of coefficients that balances the equation.

Practice Exercise 1

The unbalanced equation for the reaction between methane and bromine is

$$\_\,CH_4(g) + \_\,Br_2(l) \longrightarrow \_\,CBr_4(s) + \_\,HBr(g)$$

Once this equation is balanced what is the value of the coefficient in front of bromine Br_2?
(a) 1, **(b)** 2, **(c)** 3, **(d)** 4, **(e)** 6.

Practice Exercise 2

Balance these equations by providing the missing coefficients:

(a) $\_\,Fe(s) + \_\,O_2(g) \longrightarrow \_\,Fe_2O_3(s)$
(b) $\_\,Al(s) + \_\,HCl(aq) \longrightarrow \_\,AlCl_3(aq) + \_\,H_2(g)$
(c) $\_\,CaCO_3(s) + \_\,HCl(aq) \longrightarrow \_\,CaCl_2(aq) + \_\,CO_2(g) + \_\,H_2O(l)$

3.2 | Simple Patterns of Chemical Reactivity

In this section we examine three types of reactions that we see frequently throughout this chapter: combination reactions, decomposition reactions, and combustion reactions. Our first reason for examining these reactions is 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

In **combination reactions** two or more substances react to form one product (▶ **Table 3.1**). For example, magnesium metal burns brilliantly in air to produce magnesium oxide (▶ **Figure 3.6**):

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

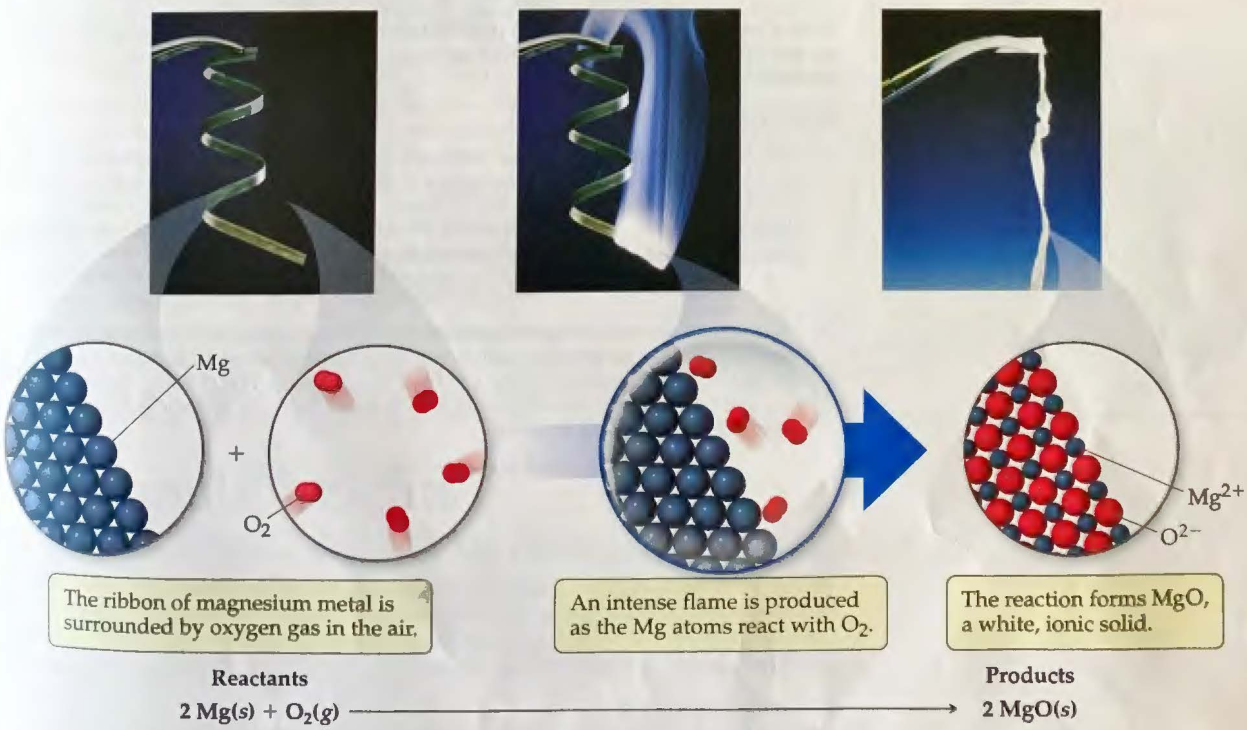
This reaction is used to produce the bright flame generated by flares and some fireworks.

A combination reaction between a metal and a nonmetal, as in Equation 3.6, produces an ionic solid. Recall that the formula of an ionic compound can be determined from the charges of its ions. (Section 2.7) When magnesium reacts with oxygen, 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 when the reactants are a metal and a nonmetal.

Table 3.1 Combination and Decomposition Reactions

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



▲ **Figure 3.6 Combustion of magnesium metal in air, a combination reaction.**

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 (Table 3.1). For example, many metal carbonates decompose to form metal oxides and carbon dioxide when heated:

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

Decomposition of $CaCO_3$ is an important commercial process. Limestone or seashells, which are both primarily $CaCO_3$, are heated to prepare CaO, known as lime or quicklime. Tens of millions of tons of CaO is used in the United States each year, in making glass, in metallurgy where it is used to isolate the metals from their ores, and in steel manufacturing where it is used to remove impurities.

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 3.7):

▲ Figure 3.7 Decomposition of sodium azide, $NaN_3(s)$, is used to inflate air bags in automobiles.

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

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).

SAMPLE EXERCISE 3.3 Writing Balanced Equations for Combination and Decomposition Reactions

Write a balanced equation for **(a)** the combination reaction between lithium metal and fluorine gas and **(b)** the decomposition reaction that occurs when solid barium carbonate is heated (two products form, a solid and a gas).

SOLUTION

(a) With the exception of mercury, all metals are solids at room temperature. Fluorine occurs as a diatomic molecule. 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 mentioned, many metal carbonates decompose to metal oxides and carbon dioxide when heated. In Equation 3.7, for example, $CaCO_3$ decomposes to form CaO and CO_2. Thus, we expect $BaCO_3$ to decompose to BaO and CO_2. Barium and calcium are both in group 2A in the periodic table, which further suggests they react in the same way:

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

Practice Exercise 1

Which of the following reactions is the balanced equation that represents the decomposition reaction that occurs when silver (I) oxide is heated? **(a)** $AgO(s) \longrightarrow Ag(s) + O(g)$; **(b)** $2\,AgO(s) \longrightarrow 2\,Ag(s) + O_2(g)$; **(c)** $Ag_2O(s) \longrightarrow 2\,Ag(s) + O(g)$; **(d)** $2\,Ag_2O(s) \longrightarrow 4\,Ag(s) + O_2(g)$; **(e)** $Ag_2O(s) \longrightarrow 2\,Ag(s) + O_2(g)$.

Practice Exercise 2

Write a balanced equation for **(a)** solid mercury (II) sulfide decomposing into its component elements when heated and **(b)** aluminum metal combining with oxygen in the air.

Combustion Reactions

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

Hydrocarbons combusted in air react with O_2 to form CO_2 and H_2O.* The number of molecules of O_2 required 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, ▶ **Figure 3.8**), a gas used for cooking and home heating, is described by the equation

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

The state of the water in this reaction, $H_2O(g)$ or $H_2O(l)$, depends on the reaction conditions. Water vapor, $H_2O(g)$, is formed at high temperature in an open container.

Combustion of oxygen-containing derivatives of hydrocarbons, such as CH_3OH, also produces CO_2 and H_2O. The rule that hydrocarbons and their oxygen-containing derivatives form CO_2 and H_2O when they burn in air summarizes the reactions of about 3 million compounds with oxygen. Many substances that our bodies use as energy sources, such as the sugar glucose ($C_6H_{12}O_6$), react 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.

GO FIGURE

Does this reaction produce or consume thermal energy (heat)?

▲ **Figure 3.8 Propane burning in air.** Liquid propane in the tank, 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.

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

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

The C atoms are balanced, one 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 this coefficient balances H but gives four O atoms in the products. Because there are only three O atoms in the reactants, we are not finished. We can place the coefficient $\frac{3}{2}$ in front of O_2 to give four O atoms in the reactants ($\frac{3}{2} \times 2 = 3$ 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 this equation is balanced, it is not in its most conventional form because it contains a fractional coefficient. However, multiplying through by 2 removes the fraction and keeps the equation balanced:

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

Practice Exercise 1

Write the balanced equation for the reaction that occurs when ethylene glycol, $C_2H_4(OH)_2$, burns in air.

(a) $C_2H_4(OH)_2(l) + 5/2\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(b) $2\,C_2H_4(OH)_2(l) + 5\,O_2(g) \longrightarrow 4\,CO_2(g) + 6\,H_2O(g)$
(c) $C_2H_4(OH)_2(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(d) $C_2H_4(OH)_2(l) + 5\,O(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(e) $4\,C_2H_4(OH)_2(l) + 10\,O_2(g) \longrightarrow 8\,CO_2(g) + 12\,H_2O(g)$

Practice Exercise 2

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

3.3 | Formula Weights

Chemical formulas and chemical equations both have a *quantitative* significance in that 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

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

we relate the numbers of atoms or molecules to the amounts we measure in the laboratory? If you wanted to react hydrogen and oxygen in exactly the right ratio to make H_2O, how would you make sure the reactants contain a 2:1 ratio of hydrogen atoms to oxygen atoms?

It is not possible to count individual atoms or molecules, but we can indirectly determine their numbers if we know their masses. So, if we are to calculate amounts of reactants needed to obtain a given amount of product, or otherwise extrapolate quantitative information from a chemical equation or formula, we need to know more about the masses of atoms and molecules.

Formula and Molecular Weights

The **formula weight** (FW) of a substance is the sum of the atomic weights (AW) of the atoms in the chemical formula of the substance. Using atomic weights, we find, for example, that the formula weight of sulfuric acid (H_2SO_4) is 98.1 amu (atomic mass units):

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

For convenience, we have rounded off the atomic weights to one decimal place, a practice we will follow in most calculations in this book.

If the chemical formula is the chemical symbol of an element, such as Na, 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, the formula weight is also called the **molecular weight** (MW). 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 exist as three-dimensional arrays of ions (see Figure 2.21), it is inappropriate to speak of molecules of these substances. Instead we use the empirical formula as the formula unit, and the formula weight of an ionic substance is determined by summing the atomic weights of the atoms in the empirical formula. For example, the formula unit of $CaCl_2$ consists of one Ca^{2+} ion and two Cl^- ions. Thus, the formula weight of $CaCl_2$ is

$$\text{FW of } CaCl_2 = 40.1\text{ amu} + 2(35.5\text{ amu}) = 111.1\text{ amu}$$

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 1

Which of the following is the correct formula weight for calcium phosphate? (**a**) 310.2 amu, (**b**) 135.1 amu, (**c**) 182.2 amu, (**d**) 278.2 amu, (**e**) 175.1 amu.

Practice Exercise 2

Calculate the formula weight of (**a**) $Al(OH)_3$, (**b**) CH_3OH, and (**c**) TaON.

Percentage Composition from Chemical Formulas

Chemists must sometimes calculate the *percentage composition* of a compound—that is, the percentage by mass contributed by each element in the substance. Forensic chemists, for example, can measure the percentage composition of an unknown powder and compare it with the percentage compositions of suspected substances (for example, sugar, salt, or cocaine) to identify the powder.

Calculating the percentage composition of any element in a substance (sometimes called the **elemental composition** of a substance) is straightforward 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:

$$\%\text{ composition of element} = \frac{\begin{pmatrix}\text{number of atoms}\\ \text{of element}\end{pmatrix}\begin{pmatrix}\text{atomic weight}\\ \text{of element}\end{pmatrix}}{\text{formula weight of substance}} \times 100\% \quad [3.10]$$

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 accompanying "Strategies in Chemistry: Problem Solving" essay.

Analyze We are given a chemical formula and asked to calculate the percentage by mass of each element.

Plan We use Equation 3.10, obtaining our atomic weights from a periodic table. We know the denominator in Equation 3.10, the formula weight of $C_{12}H_{22}O_{11}$, from Sample Exercise 3.5. We must use that value in three calculations, one for each element.

Solve

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

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

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

Check Our calculated percentages must add up to 100%, which they do. 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 1

What is the percentage of nitrogen, by mass, in calcium nitrate? **(a)** 8.54%, **(b)** 17.1%, **(c)** 13.7%, **(d)** 24.4%, **(e)** 82.9%.

Practice Exercise 2

Calculate the percentage of potassium, by mass, in K_2PtCl_6.

3.4 | Avogadro's Number and the Mole

Even the smallest samples 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 it almost defies comprehension. Chemists therefore have devised a counting unit for describing large numbers of atoms or molecules.

Strategies in Chemistry

Problem Solving

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

1. **Analyze the problem.** Read the problem carefully. What does it say? Draw a picture or diagram that will help you to visualize the problem. Write down both the data you are given and the quantity you need to obtain (the unknown).
2. **Develop a plan for solving the problem.** Consider a possible path 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.
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 useful tool for solving a great number of problems. Be careful with significant figures, signs, and units.
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?

In everyday life we use such familiar counting units as dozen (12 objects) and gross (144 objects). In chemistry the counting unit for numbers of atoms, ions, or molecules in a laboratory-size sample is the *mole*, abbreviated mol. One **mole** is the amount of matter that contains as many objects (atoms, molecules, or whatever other 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 $6.02214129 \times 10^{23}$, which we usually round to 6.02×10^{23}. Scientists call this value **Avogadro's number**, N_A, in honor of the Italian scientist Amedeo Avogadro (1776–1856), and it is often cited with units of reciprocal moles, $6.02 \times 10^{23}\ mol^{-1}$.* The unit (read as either "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 objects:

$$1\ mol\ ^{12}C\ atoms = 6.02 \times 10^{23}\ ^{12}C\ atoms$$
$$1\ mol\ H_2O\ molecules = 6.02 \times 10^{23}\ H_2O\ molecules$$
$$1\ mol\ NO_3^-\ ions = 6.02 \times 10^{23}\ NO_3^-\ ions$$

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

SAMPLE EXERCISE 3.7 Estimating Numbers of Atoms

Without using a calculator, arrange these 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 three 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 12 g ^{12}C, 1 mol C_2H_2, and 9×10^{23} molecules CO_2 to numbers of C atoms. To make these conversions, we use the definition of mole and Avogadro's number.

Solve One 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 = 6.02×10^{23} C atoms. One mol of C_2H_2 contains 6.02×10^{23} C_2H_2 molecules. Because there are two C atoms in each molecule, this sample contains 12.04×10^{23} C atoms. Because each CO_2 molecule contains one C atom, the CO_2 sample 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 numbers of moles of C atoms in the samples 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 stated previously.

Practice Exercise 1

Determine which of the following samples contains the fewest sodium atoms? (**a**) 1 mol sodium oxide, (**b**) 45 g sodium fluoride, (**c**) 50 g sodium chloride, (**d**) 1 mol sodium nitrate?

Practice Exercise 2

Without using a calculator, arrange these samples in order of increasing numbers of O atoms: 1 mol H_2O, 1 mol CO_2, 3×10^{23} molecules of O_3.

*Avogadro's number is also referred to as the Avogadro constant. The latter term is the name adopted by agencies such as the National Institute of Standards and Technology (NIST), but Avogadro's number remains in widespread usage and is used in most places in this book.

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 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 number of moles of $C_6H_{12}O_6$ and number of molecules of $C_6H_{12}O_6$: 1 mol $C_6H_{12}O_6$ = 6.02×10^{23} 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 We can do a ballpark calculation, figuring that $0.35(6 \times 10^{23})$ is about 2×10^{23} molecules of $C_6H_{12}O_6$. We know that each one of these molecules contains 12 H atoms. $12(2 \times 10^{23})$ gives $24 \times 10^{23} = 2.4 \times 10^{24}$ H atoms, which is close to our result. Because we were asked for the number of H atoms, the units of our answer are correct. We check, too, for significant figures. The given data had three significant figures, as does our answer.

Practice Exercise 1

How many sulfur atoms are in **(a)** 0.45 mol $BaSO_4$ and **(b)** 1.10 mol of aluminum sulfide?

Practice Exercise 2

How many oxygen atoms are in **(a)** 0.25 mol $Ca(NO_3)_2$ and **(b)** 1.50 mol of sodium carbonate?

Molar Mass

A dozen is the same number, 12, 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-mol samples of different substances 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 of anything always contains 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), 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 atomic weight of an element in atomic mass units is numerically equal to the mass in grams of 1 mol of that element.* For example (the symbol $\Rightarrow$ means therefore)

Cl has an atomic weight of 35.5 amu $\Rightarrow$ 1 mol Cl has a mass of 35.5 g.

Au has an atomic weight of 197 amu $\Rightarrow$ 1 mol Au has a mass of 197 g.

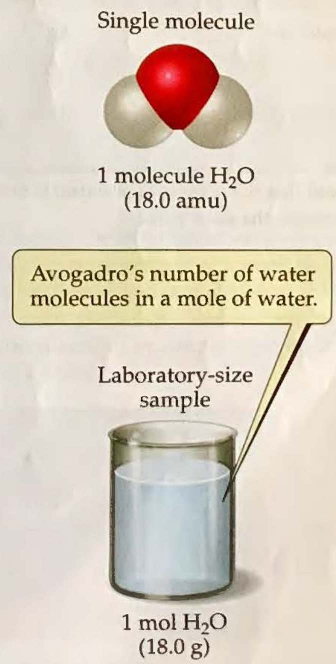
For other kinds of substances, the same numerical relationship exists between formula weight and mass of 1 mol of a substance:

H_2O has a formula weight of 18.0 amu $\Rightarrow$ 1 mol H_2O has a mass of 18.0 g (▶ Figure 3.9).

NaCl has a formula weight of 58.5 amu $\Rightarrow$ 1 mol NaCl has a mass of 58.5 g.

GO FIGURE

How many H_2O molecules are in a 9.00-g sample of water?



▲ **Figure 3.9 Comparing the mass of 1 molecule and 1 mol of H_2O.** Both masses have the same number but different units (atomic mass units and grams). Expressing both masses in grams indicates their huge difference: 1 molecule of H_2O has a mass of 2.99×10^{-23} g, whereas 1 mol H_2O has a mass of 18.0 g.

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?

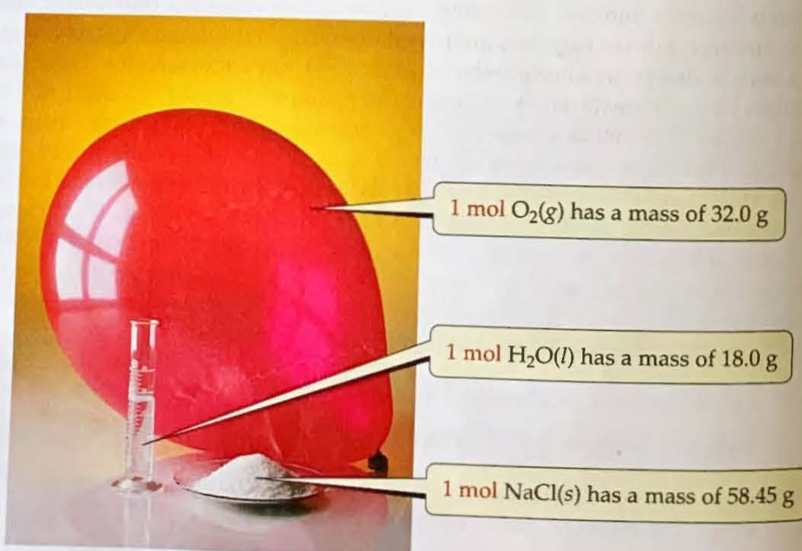
The mass in grams of one mole, often abbreviated as 1 mol, of a substance (that is, the mass in grams per mole) is called the **molar mass** of the substance. *The molar mass in grams per mole of any substance is numerically equal to its formula weight in atomic mass units.* For NaCl, for example, the formula weight is 58.5 amu and the molar mass is 58.5 g/mol. Mole relationships for several other substances are shown in ▼ **Table 3.2**, and ▼ **Figure 3.10** shows 1 mol quantities of three common substances.

The entries in Table 3.2 for N and N_2 point out the importance of stating the chemical form of a substance when using 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

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	$\begin{cases} 6.02 \times 10^{23}\ N_2 \text{ molecules} \\ 2(6.02 \times 10^{23})\ N \text{ atoms} \end{cases}$
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	$\begin{cases} 6.02 \times 10^{23}\ BaCl_2 \text{ formula units} \\ 6.02 \times 10^{23}\ Ba^{2+} \text{ ions} \\ 2(6.02 \times 10^{23})\ Cl^- \text{ ions} \end{cases}$

[a]Recall that the mass of an electron is more than 1800 times smaller than the masses of the proton and the neutron; thus, ions and atoms have essentially the same mass.



▲ **Figure 3.10** **One mole each of a solid (NaCl), a liquid (H_2O), and a gas (O_2).** In each case, the mass in grams of 1 mol—that is, the molar mass—is numerically equal to the formula weight in atomic mass units. Each of these samples contains 6.02×10^{23} formula units.

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 or N_2, for instance—avoids ambiguity.

SAMPLE EXERCISE 3.9 Calculating Molar Mass

What is the molar mass of glucose, $C_6H_{12}O_6$?

SOLUTION

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

Plan Because the molar mass of any substance is numerically equal to its formula weight, we first determine the formula weight of glucose by adding the atomic weights of its component atoms. The formula weight will have units of amu, whereas the molar mass has units of grams per mole (g/mol).

Solve Our first step is to determine the formula weight of glucose:

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

Because glucose has a formula weight of 180.0 amu, 1 mol of this substance (6.02×10^{23} molecules) 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 A molar mass below 250 seems reasonable based on the earlier examples we have encountered, and grams per mole is the appropriate unit for the molar mass.

Practice Exercise 1

A sample of an ionic compound containing iron and chlorine is analyzed and found to have a molar mass of 126.8 g/mol. What is the charge of the iron in this compound? (a) 1+, (b) 2+, (c) 3+, (d) 4+.

Practice Exercise 2

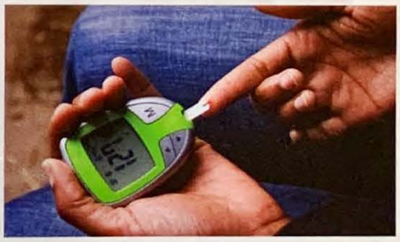
Calculate the molar mass of $Ca(NO_3)_2$.

Chemistry and Life

Glucose Monitoring

Our body converts most of the food we eat into glucose. After digestion, glucose is delivered to cells via the blood. Cells need glucose to live, and the hormone insulin must be present in order 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, either little or no insulin is produced (Type 1 diabetes) or insulin is produced but the cells cannot take it up properly (Type 2 diabetes). In either case the blood glucose levels are higher than they are in a normal person, typically 70–120 mg/dL. A person who has not eaten for 8 hours or more is diagnosed as diabetic if his or her glucose level is 126 mg/dL or higher.

Glucose meters work by the introduction of blood from a person, usually by a prick of the finger, onto a small strip of paper that contains chemicals that react with glucose. Insertion of the strip into a small battery-operated reader gives the glucose concentration (▼ Figure 3.11). The mechanism of the readout varies from one monitor to another—it may be a measurement of a small electrical current or measurement of light produced in a chemical reaction. Depending on the reading on any given day, a diabetic person may need to receive an injection of insulin or simply limit his or her intake of sugar-rich foods for a while.



▲ Figure 3.11 Glucose meter.

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 (Section 1.6), 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

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

Check Because 5.380 g is less than the molar mass, an answer less than 1 mol is reasonable. The unit mol is appropriate. The original data had four significant figures, so our answer has four significant figures.

Practice Exercise 1

How many moles of sodium bicarbonate ($NaHCO_3$) are in 508 g of $NaHCO_3$?

Practice Exercise 2

How many moles of water are in 1.00 L of water, whose density is 1.00 g/mL?

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 substance.

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^-, the chemical formula for 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

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

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, which means the magnitude of our answer is reasonable. Both the units (g) and the number of significant figures (3) are correct.

Practice Exercise 1

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

Practice Exercise 2

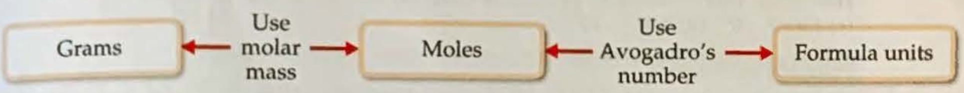
What is the mass, in grams, of **(a)** 0.50 mol of diamond (C) and **(b)** 0.155 mol of ammonium chloride?

Interconverting Masses and Numbers of Particles

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

GO FIGURE

What number would you use to convert (**a**) moles of CH_4 to grams of CH_4 and (**b**) number of molecules of CH_4 to moles of CH_4?



▲ **Figure 3.12 Procedure for interconverting mass and number of formula units.** 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 sample in grams and the number of formula units contained in the sample.

$$\text{Cu atoms} = (3\ \cancel{\text{g Cu}})\left(\frac{1\ \cancel{\text{mol Cu}}}{63.5\ \cancel{\text{g Cu}}}\right)\left(\frac{6.02 \times 10^{23}\ \text{Cu atoms}}{1\ \cancel{\text{mol Cu}}}\right)$$
$$= 3 \times 10^{22}\ \text{Cu atoms}$$

We have rounded our answer to one significant figure because we used only one significant figure for the mass of the penny. Notice how dimensional analysis 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 to moles and then moles to 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 small, often less than 1.

The general procedure for interconverting mass and number of formula units (atoms, molecules, ions, or whatever else is represented by the chemical formula) is summarized in ▲ **Figure 3.12.**

SAMPLE EXERCISE 3.12 Calculating Numbers of Molecules and 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 of a substance and asked to calculate (**a**) the number of molecules and (**b**) the number of O atoms in the substance.

Plan (**a**) The strategy for determining the number of molecules in a given quantity of a substance is summarized in Figure 3.12. We must convert 5.23 g to moles of $C_6H_{12}O_6$ and then convert moles to molecules of $C_6H_{12}O_6$. The first conversion uses the molar mass of $C_6H_{12}O_6$, 180.0 g, and the second conversion uses Avogadro's number.

Solve Molecules $C_6H_{12}O_6$

$$= (5.23\ \text{g}\ \cancel{C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}{180.0\ \text{g}\ \cancel{C_6H_{12}O_6}}\right)\left(\frac{6.02 \times 10^{23}\ \text{molecules}\ C_6H_{12}O_6}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}\right)$$
$$= 1.75 \times 10^{22}\ \text{molecules}\ C_6H_{12}O_6$$

Check Because the mass we began with is less than a mole, there should be fewer than 6.02×10^{23} molecules in the sample, which means the magnitude of our answer is reasonable. A ballpark estimate of the answer comes reasonably close to the answer we derived in this exercise: $5/200 = 2.5 \times 10^{-2}$ mol; $(2.5 \times 10^{-2})(6 \times 10^{23}) = 15 \times 10^{21} = 1.5 \times 10^{22}$ molecules. The units (molecules) and the number of significant figures (three) are appropriate.

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

Solve

$$\text{Atoms O} = (1.75 \times 10^{22}\ \cancel{\text{molecules}\ C_6H_{12}O_6})\left(\frac{6\ \text{atoms O}}{\cancel{\text{molecule}\ C_6H_{12}O_6}}\right)$$
$$= 1.05 \times 10^{23}\ \text{atoms O}$$

Check The answer is six times as large as the answer to part (**a**), exactly what it should be. The number of significant figures (three) and the units (atoms O) are correct.

Practice Exercise 1

How many chlorine atoms are in 12.2 g of CCl_4? (**a**) 4.77×10^{22}, (**b**) 7.34×10^{24}, (**c**) 1.91×10^{23}, (**d**) 2.07×10^{23}.

Practice Exercise 2

(**a**) How many nitric acid molecules are in 4.20 g of HNO_3?
(**b**) How many O atoms are in this sample?

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 in the substance. 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 numbers of moles of all elements in a compound gives the subscripts in the compound's empirical formula*. Thus, the mole concept provides a way of calculating empirical formulas.

Mercury and chlorine, for example, combine to form a compound that is measured to be 74.0% mercury and 26.0% chlorine by mass. Thus, if we had a 100.0-g sample of the compound, it would contain 74.0 g of mercury and 26.0 g of chlorine. (Samples of any size 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 atomic weights to get molar masses, we can calculate the number of moles of each element in the sample:

$$(74.0\ \cancel{\text{g Hg}})\left(\frac{1\ \text{mol Hg}}{200.6\ \cancel{\text{g Hg}}}\right) = 0.369\ \text{mol Hg}$$

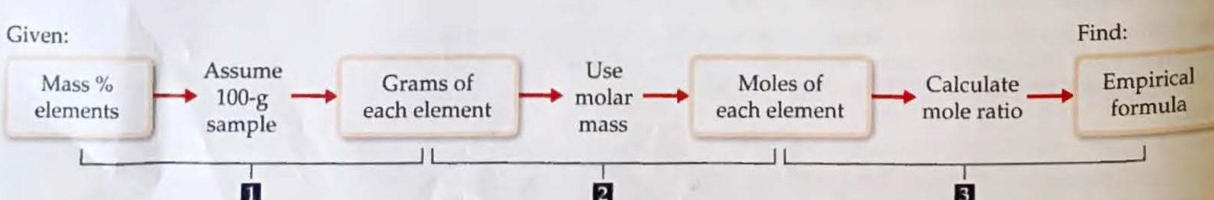
$$(26.0\ \cancel{\text{g Cl}})\left(\frac{1\ \text{mol Cl}}{35.5\ \cancel{\text{g Cl}}}\right) = 0.732\ \text{mol Cl}$$

We then divide the larger number of moles by the smaller number to obtain the Cl:Hg mole ratio:

$$\frac{\text{moles of Cl}}{\text{moles of Hg}} = \frac{0.732\ \text{mol Cl}}{0.369\ \text{mol Hg}} = \frac{1.98\ \text{mol Cl}}{1\ \text{mol Hg}}$$

Because of experimental errors, calculated values for a mole ratio may not be whole numbers, as in the calculation here. The number 1.98 is very close to 2, however, and 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 *ratio* of atoms present in the compound. (Section 2.6)

The general procedure for determining empirical formulas is outlined in ▼ **Figure 3.13.**



▲ **Figure 3.13** **Procedure for calculating an empirical formula from percentage composition.**

Give It Some Thought

Could the empirical formula determined from chemical analysis be used to tell the difference between acetylene, C_2H_2, and benzene, C_6H_6?

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?

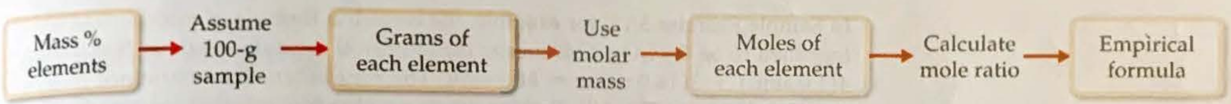
SOLUTION

Analyze We are to determine the 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 3.13.

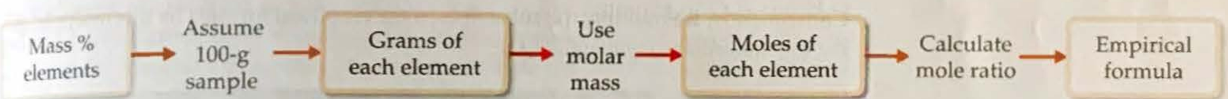
Solve

(1) For simplicity we assume we have exactly 100 g of material, although any other mass could also be used.



In 100.00 g of ascorbic acid we have 40.92 g C, 4.58 g H, and 54.50 g O.

(2) Next we calculate the number of moles of each element. We use atomic masses with four significant figures to match the precision of our experimental masses.

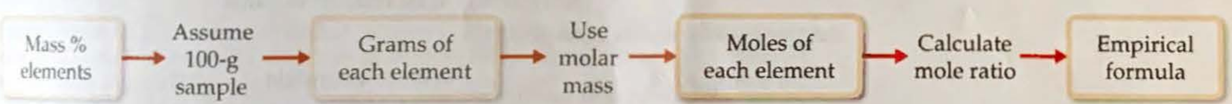


$$\text{Moles C} = (40.92\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}}\right) = 3.407\ \text{mol C}$$

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

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

(3) We determine the simplest whole-number ratio of moles by dividing each number of moles by the smallest number of moles.



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

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 we should multiply the ratios by 3 to obtain whole numbers:

$$C:H:O = (3 \times 1 : 3 \times 1.33 : 3 \times 1) = (3:4:3)$$

Thus, the empirical formula is $C_3H_4O_3$.

Check It is reassuring that the subscripts are moderate-size whole numbers. Also, calculating the percentage composition of $C_3H_4O_3$ gives values very close to the original percentages.

Practice Exercise 1

A 2.144-g sample of phosgene, a compound used as a chemical warfare agent during World War I, contains 0.260 g of carbon, 0.347 g of oxygen, and 1.537 g of chlorine. What is the empirical formula of this substance? **(a)** CO_2Cl_6, **(b)** $COCl_2$, **(c)** $C_{0.022}O_{0.022}Cl_{0.044}$, **(d)** C_2OCl_2

Practice Exercise 2

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?

Molecular Formulas from Empirical Formulas

We can obtain the molecular formula for any compound from its empirical formula if we know either the molecular weight or the molar mass of the compound. *The subscripts in the molecular formula of a substance are always whole-number multiples of the subscripts in its empirical formula.* (Section 2.6) This multiple can be found by dividing the molecular weight by the empirical formula 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 $C_3H_4O_3$. This means the empirical formula weight is $3(12.0\text{ amu}) + 4(1.0\text{ amu}) + 3(16.0\text{ amu}) = 88.0\text{ amu}$. The experimentally determined molecular weight is 176 amu. Thus, we find the whole-number multiple that converts the empirical formula to the molecular formula by dividing

$$\text{Whole-number multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}} = \frac{176\text{ amu}}{88.0\text{ amu}} = 2$$

Consequently, we multiply the subscripts in the empirical formula by this multiple, giving the molecular formula: $C_6H_8O_6$.

SAMPLE EXERCISE 3.14 Determining a Molecular Formula

Mesitylene, a hydrocarbon found in crude oil, has an empirical formula of C_3H_4 and an experimentally determined molecular weight of 121 amu. What is its molecular formula?

SOLUTION

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

Plan The subscripts in a compound's molecular formula are whole-number multiples of the subscripts in its empirical formula. We find the appropriate multiple by using Equation 3.11.

Solve The formula weight of the empirical formula C_3H_4 is

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

Next, we use this value in Equation 3.11:

$$\text{Whole-number multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}} = \frac{121}{40.0} = 3.03$$

Only whole-number ratios make physical sense because molecules contain whole atoms. The 3.03 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}.

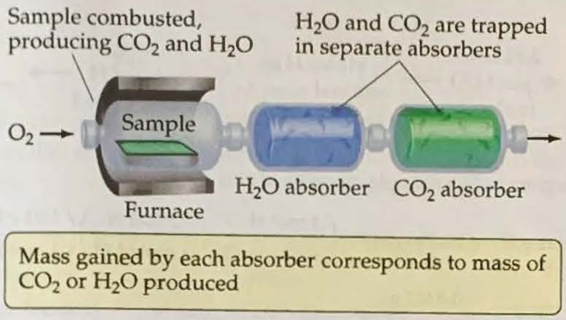
Check We can have confidence in the result because dividing molecular weight by empirical formula weight yields nearly a whole number.

Practice Exercise 1

Cyclohexane, a commonly used organic solvent, is 85.6% C and 14.4% H by mass with a molar mass of 84.2 g/mol. What is its molecular formula? (a) C_6H, (b) CH_2, (c) C_5H_{24}, (d) C_6H_{12}, (e) C_4H_8.

Practice Exercise 2

Ethylene glycol, used in automobile antifreeze, is 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?



▲ **Figure 3.14** **Apparatus for combustion analysis.**

Combustion Analysis

One technique for determining empirical formulas in the laboratory is *combustion analysis*, commonly used for compounds containing principally carbon and hydrogen.

When a compound containing carbon and hydrogen is completely combusted in an apparatus such as that shown in ▲ **Figure 3.14**, the carbon 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 sample and thereby the empirical formula. If a third element is present in the compound, its mass can be determined by subtracting the measured masses of C and H from the original sample mass.

SAMPLE EXERCISE 3.15 Determining an Empirical Formula by Combustion Analysis

Isopropyl alcohol, 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 are given the quantities of CO_2 and H_2O produced when a given quantity of the alcohol is combusted. We must 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 grams of C in the CO_2 and grams of H in the H_2O—the masses of C and H in the alcohol before combustion. The mass of O in the compound equals the mass of the original sample minus the sum of the C and H masses. Once we have the C, H, and O masses, we can proceed as in Sample Exercise 3.13.

Solve Because all of the carbon in the sample is converted to CO_2, we can use dimensional analysis and the following steps to calculate the mass C in the sample.

Mass CO_2 produced → Molar mass CO_2 44.0 g/mol → Moles CO_2 produced → 1 C atom per CO_2 molecule → Moles of C original sample → Molar mass C 12.0 g/mol → Mass C in original sample

Using the values given in this example, the mass of C is

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

Because all of the hydrogen in the sample is converted to H_2O, we can use dimensional analysis and the following steps to calculate the mass H in the sample. We use three significant figures for the atomic mass of H to match the significant figures in the mass of H_2O produced.

Mass H_2O produced → Molar mass H_2O 18.0 g/mol → Moles H_2O produced → 2 H atoms per H_2O molecule → Moles H in original sample → Molar mass H 1.01 g/mol → Mass H in original sample

Using the values given in this example, the mass of H is

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

The mass of the sample, 0.255 g, is the sum of the masses of C, H, and O. Thus, the O mass is

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

The number of moles of C, H, and O in the sample is therefore

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

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

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

To find the empirical formula, we must compare the relative number of moles of each element in the sample, as illustrated in Sample Exercise 3.13.

$$\text{C:}\frac{0.0128}{0.0043} = 3.0 \quad \text{H:}\frac{0.0340}{0.0043} = 7.9 \quad \text{O:}\frac{0.0043}{0.0043} = 1.0$$

The first two numbers are very close to the whole numbers 3 and 8, giving the empirical formula C_3H_8O.

Practice Exercise 1

The compound dioxane, which is used as a solvent in various industrial processes, is composed of C, H, and O atoms. Combustion of a 2.203-g sample of this compound produces 4.401 g CO_2 and 1.802 g H_2O. A separate experiment shows that it has a molar mass of 88.1 g/mol. Which of the following is the correct molecular formula for dioxane? **(a)** C_2H_4O, **(b)** $C_4H_4O_2$, **(c)** CH_2, **(d)** $C_4H_8O_2$.

Practice Exercise 2

(a) Caproic acid, responsible for the 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?

Give It Some Thought

In Sample Exercise 3.15, how do you explain that the values in our calculated C:H:O ratio are 3.0:7.9:1.0 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 in the reaction. For instance, the coefficients in the balanced equation

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

indicate that two molecules of H_2 react with one 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

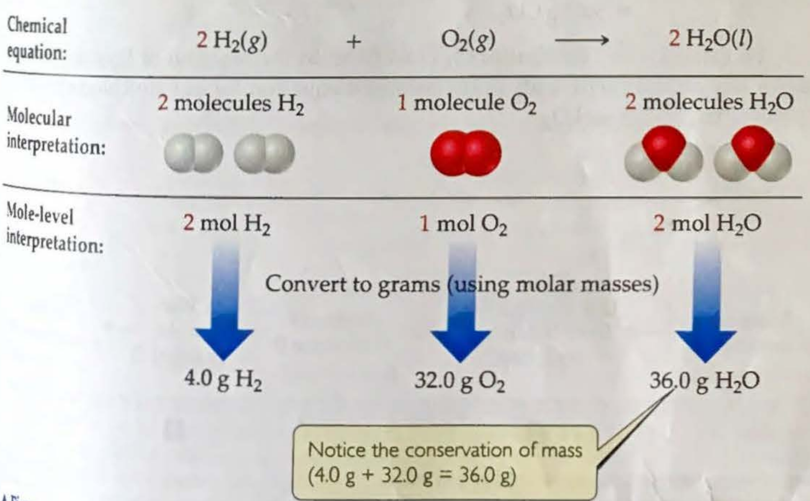
We can generalize this observation to all balanced chemical equations: *The coefficients in a balanced chemical equation indicate both the relative numbers of molecules (or formula units) in the reaction and the relative numbers of moles.* ▼ **Figure 3.15** shows how this result corresponds to the law of conservation of mass.

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

$$2 \text{ mol } H_2 \;\hat{=}\; 1 \text{ mol } O_2 \;\hat{=}\; 2 \text{ mol } H_2O$$

where the $\hat{=}$ symbol means "is stoichiometrically equivalent to." Stoichiometric relations such as these 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 is

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



▲ **Figure 3.15** Interpreting a balanced chemical equation quantitatively.

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 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 us 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 stoichiometric relationship, we must convert grams of C_4H_{10} to moles using the molar mass of C_4H_{10}, 58.0 g/mol:

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

We then use the stoichiometric factor from the balanced equation to calculate moles of CO_2:

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

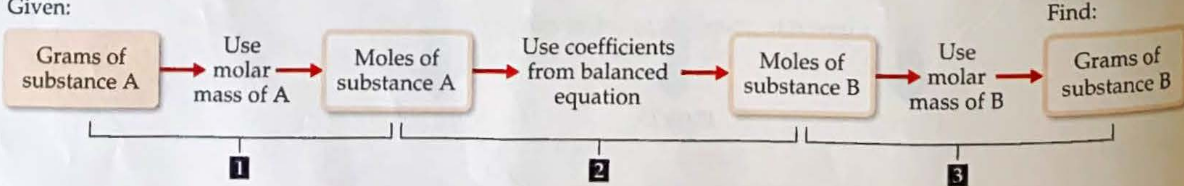
Finally, we use the molar mass of CO_2, 44.0 g/mol, to calculate the CO_2 mass in grams:

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

This conversion sequence involves three steps, as illustrated in ▼ **Figure 3.16.** These three conversions can be combined in a single equation:

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

To calculate the amount of O_2 consumed in the reaction of Equation 3.13, we again rely on the coefficients in the balanced equation for our stoichiometric factor, 2 mol $C_4H_{10} \simeq$ 13 mol O_2:



▲ **Figure 3.16 Procedure for calculating amounts of reactants consumed or products formed in a reaction.** The number of grams of a reactant consumed or product formed can be calculated in three steps, starting with the number of grams of any reactant or product.

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

Give It Some Thought

In the previous example, 1.00 g of C_4H_{10} reacts with 3.59 g of O_2 to form 3.03 g of CO_2. Using only addition and subtraction, calculate the amount of H_2O produced.

SAMPLE EXERCISE 3.16 Calculating Amounts of Reactants and Products

Determine 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 must determine the mass of a product in the given reaction.

Plan We follow the general strategy outlined in Figure 3.16:

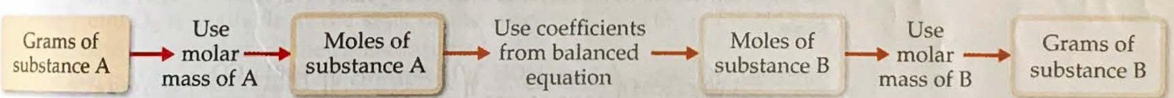
(1) Convert grams of $C_6H_{12}O_6$ to moles using the molar mass of $C_6H_{12}O_6$.

(2) Convert moles of $C_6H_{12}O_6$ to moles of H_2O using the stoichiometric relationship 1 mol $C_6H_{12}O_6 \simeq$ 6 mol H_2O.

(3) Convert moles of H_2O to grams using the molar mass of H_2O.

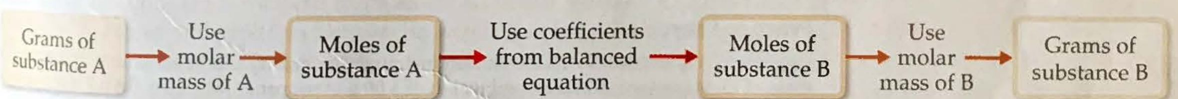
Solve

(1) First we convert grams of $C_6H_{12}O_6$ to moles using the molar mass of $C_6H_{12}O_6$.



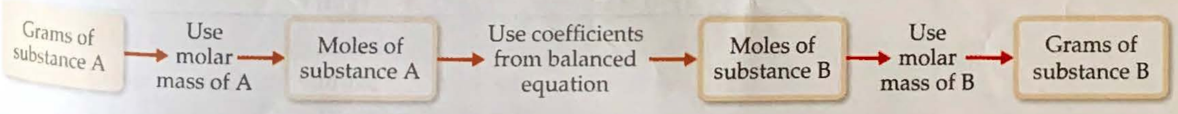
$$\text{Moles } C_6H_{12}O_6 = (1.00\ \cancel{g\ C_6H_{12}O_6})\left(\frac{1\ mol\ C_6H_{12}O_6}{180.0\ \cancel{g\ C_6H_{12}O_6}}\right)$$

(2) Next we convert moles of $C_6H_{12}O_6$ to moles of H_2O using the stoichiometric relationship 1 mol $C_6H_{12}O_6 \simeq$ 6 mol H_2O.



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

(3) Finally, we convert moles of H_2O to grams using the molar mass of H_2O.



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

Check We can check how reasonable our result is by doing a ballpark estimate of the mass of H_2O. Because the molar mass of glucose is 180 g/mol, 1 g of glucose equals 1/180 mol. Because 1 mol of glucose yields 6 mol H_2O, we would have $6/180 = 1/30$ mol H_2O. The molar mass of water is 18 g/mol, so we have $1/30 \times 18 = 6/10 = 0.6$ g of H_2O, which agrees with the full calculation. The units, grams H_2O, are correct. The initial data had three significant figures, so three significant figures for the answer is correct.

Practice Exercise 1

Sodium hydroxide reacts with carbon dioxide to form sodium carbonate and water:

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

How many grams of Na_2CO_3 can be prepared from 2.40 g of NaOH? **(a)** 3.18 g, **(b)** 6.36 g, **(c)** 1.20 g, **(d)** 0.0300 g.

Practice Exercise 2

Decomposition of $KClO_3$ is sometimes 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 $KClO_3$?

SAMPLE EXERCISE 3.17 Calculating Amounts of Reactants and Products

Solid lithium hydroxide is used in space vehicles to remove the carbon dioxide gas exhaled by astronauts. The hydroxide reacts with the 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 mass in grams of LiOH and asked to calculate the mass in grams of CO_2. We can accomplish this with the three conversion steps in Figure 3.16. The conversion of Step 1 requires the molar mass of LiOH ($6.94 + 16.00 + 1.01 = 23.95$ g/mol). The conversion of Step 2 is based on a stoichiometric relationship from the balanced chemical equation: 2 mol LiOH $\simeq$ mol CO_2. For the Step 3 conversion, we use the molar mass of CO_2 $12.01 + 2(16.00) = 44.01$ g/mol.

Solve

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

Check Notice that 23.95 g LiOH/mol $\approx$ 24 g LiOH/mol, 24 g LiOH/mol $\times$ 2 mol LiOH = 48 g LiOH, and (44 g CO_2/mol)/(48 g LiOH) is slightly less than 1. Thus, the magnitude of our answer, 0.919 g CO_2, is reasonable based on the amount of starting LiOH. The number of significant figures and units are also appropriate.

Practice Exercise 1

Propane, C_3H_8 (Figure 3.8), 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? **(a)** 5.00 g, **(b)** 0.726 g, **(c)** 2.18 g, **(d)** 3.63 g.

Practice Exercise 2

Methanol, CH_3OH, reacts with oxygen from air in a combustion reaction to form water and carbon dioxide. What mass of water is produced in the combustion of 23.6 g of methanol?

3.7 | Limiting Reactants

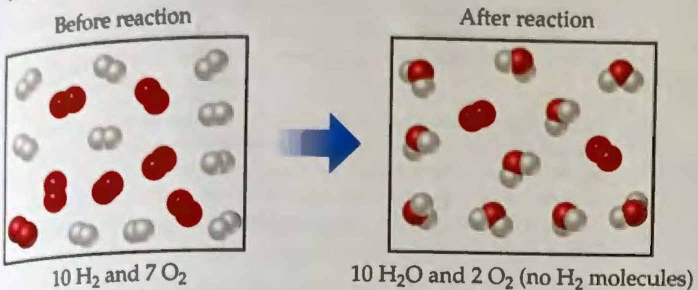
Suppose you wish to make several sandwiches using one slice of cheese and two slices of bread for each. Using Bd = bread, Ch = cheese, and Bd_2Ch = sandwich, the recipe for making a sandwich can be represented like a chemical equation:

$$2\,Bd + Ch \longrightarrow Bd_2Ch$$

If you have ten slices of bread and seven slices of cheese, you can make only five sandwiches and will have two slices of cheese left over. The amount of bread available limits the number of sandwiches.

GO FIGURE

If the amount of H_2 is doubled, how many moles of H_2O would have formed?



▲ **Figure 3.17 Limiting reactant.** Because H_2 is completely consumed, it is the limiting reactant. Because some O_2 is left over after the reaction is complete, it is the excess reactant. The amount of H_2O formed depends on the amount of limiting reactant, H_2.

An analogous situation occurs in chemical reactions when one reactant is used up before the others. The reaction stops as soon as any reactant is totally consumed, leaving the excess reactants as leftovers. Suppose, for example, 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 $H_2 \simeq$ 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 is available at the start of the reaction, 7 mol O_2 − 5 mol O_2 = 2 mol O_2 is still present when all the H_2 is consumed.

The reactant that is completely consumed in a reaction is called the **limiting reactant** because it determines, or limits, the amount of product formed. The other reactants are sometimes called *excess reactants*. In our example, shown in ▲ **Figure 3.17**, H_2 is the limiting reactant, which means that once all the H_2 has been consumed, the reaction stops. At that point some of the excess reactant O_2 is left over.

There are no restrictions on the starting amounts of reactants in any reaction. Indeed, many reactions are carried out using an excess of one reactant. The quantities of reactants consumed and products formed, however, are restricted by the quantity of the limiting reactant. For example, when a combustion reaction takes place in the open air, oxygen is plentiful and is therefore the excess reactant. If you run out of gasoline while driving, the car stops because the gasoline is the limiting reactant in the combustion reaction that moves the car.

Before we leave the example illustrated in Figure 3.17, let's summarize the data:

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

The second line in the table (Change) summarizes the amounts of reactants consumed (where this consumption is indicated by the minus signs) and the amount of the product formed (indicated by the plus sign). 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 is a multiple of the ratio of the coefficients in the balanced equation, 2:1:2. The after quantities, which depend on the before quantities and their changes, are found by adding the before quantity and change quantity for each column. The amount of the limiting reactant (H_2) must be zero at the end of the reaction. What remains is 2 mol O_2 (excess reactant) and 10 mol H_2O (product).

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. This is a limiting reactant problem.

Plan If we assume one reactant is completely consumed, we can calculate how much of the second reactant is needed. By comparing the calculated quantity of the second reactant with the amount available, 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, which tells us that H_2 is the limiting reactant. Therefore, we use the quantity of 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 It is useful to summarize the reaction data in a table:

	$N_2(g)$	+ $3\,H_2(g)$	$\longrightarrow 2\,NH_3(g)$
Before reaction:	3.0 mol	6.0 mol	0 mol
Change (reaction):	−2.0 mol	−6.0 mol	+4.0 mol
After reaction:	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 reactant remaining after the reaction. Notice also that although the initial (before) number of moles of H_2 is greater than the final (after) number of moles of N_2, H_2 is nevertheless the limiting reactant because of its larger coefficient in the balanced equation.

Check Examine the Change row of the summary table to see that the mole ratio of reactants consumed and product formed, 2:6:4, is a multiple of the coefficients in the balanced equation, 1:3:2. We confirm that H_2 is the limiting reactant because 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 1

When 24 mol of methanol and 15 mol of oxygen combine in the combustion reaction $2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$, what is the excess reactant and how many moles of it remains at the end of the reaction?
(a) 9 mol $CH_3OH(l)$, **(b)** 10 mol $CO_2(g)$,
(c) 10 mol $CH_3OH(l)$, **(d)** 14 mol $CH_3OH(l)$, **(e)** 1 mol $O_2(g)$.

Practice Exercise 2

(a) When 1.50 mol of Al and 3.00 mol of Cl_2 combine in the reaction $2\,Al(s) + 3\,Cl_2(g) \longrightarrow 2\,AlCl_3(s)$, 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?

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

The reaction

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

is used to produce electricity in a hydrogen fuel cell. Suppose a fuel cell contains 150 g of $H_2(g)$ and 1500 g of $O_2(g)$ (each measured to two significant figures). How many grams of water can form?

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 To identify the limiting reactant, we can calculate the number of moles of each reactant and compare their ratio with the ratio of coefficients in the balanced equation. We then use the quantity of the limiting reactant to calculate the mass of water that forms.

Solve From the balanced equation, we have the stoichiometric relations

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

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

$$\text{Moles } H_2 = (150\ \cancel{\text{g } H_2})\left(\frac{1\text{ mol } H_2}{2.02\ \cancel{\text{g } H_2}}\right) = 74\text{ mol } H_2$$

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

The coefficients in the balanced equation indicate that the reaction requires 2 mol of H_2 for every 1 mol of O_2. Therefore, for all the O_2 to completely react, we would need $2 \times 47 = 94$ mol of H_2. Since there are only 74 mol of H_2, all of the O_2 cannot react, so it is the excess reactant, and H_2 must be the limiting reactant. (Notice that the limiting reactant is not necessarily the one present in the lowest amount.)

We use the given quantity of H_2 (the limiting reactant) to calculate the quantity of water formed. We could begin this calculation with the given H_2 mass, 150 g, but we can save a step by starting with the moles of H_2, 74 mol, we just calculated:

$$\text{Grams } H_2O = (74\ \cancel{\text{mol } H_2})\left(\frac{2\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } H_2}}\right)\left(\frac{18.0\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}}\right)$$
$$= 1.3 \times 10^2\text{ g } H_2O$$

Check The magnitude of the answer seems reasonable based on the amounts of the reactants. The units are correct, and the number of significant figures (two) corresponds to those in the values given in the problem statement.

Comment The quantity of the limiting reactant, H_2, can also be used to determine the quantity of O_2 used:

$$\text{Grams } O_2 = (74\ \cancel{\text{mol } H_2})\left(\frac{1\ \cancel{\text{mol } O_2}}{2\ \cancel{\text{mol } H_2}}\right)\left(\frac{32.0\text{ g } O_2}{1\ \cancel{\text{mol } O_2}}\right)$$
$$= 1.2 \times 10^3\text{ g } O_2$$

The mass of O_2 remaining at the end of the reaction equals the starting amount minus the amount consumed:

$$1500\text{ g} - 1200\text{ g} = 300\text{ g}.$$

Practice Exercise 1

Molten gallium reacts with arsenic to form the semiconductor, gallium arsenide, GaAs, used in light-emitting diodes and solar cells :

$$Ga(l) + As(s) \longrightarrow GaAs(s)$$

If 4.00 g of gallium is reacted with 5.50 g of arsenic, how many grams of the excess reactant are left at the end of the reaction? **(a)** 4.94 g As, **(b)** 0.56 g As, **(c)** 8.94 g Ga, or **(d)** 1.50 g As.

Practice Exercise 2

When a 2.00-g strip of zinc metal is placed in an aqueous solution containing 2.50 g of silver nitrate, the reaction is

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

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

Theoretical and Percent Yields

The quantity of product calculated to form when all of a limiting reactant is consumed is called the **theoretical yield**. The amount of product actually obtained, called 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 actual and theoretical yields:

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

SAMPLE EXERCISE 3.20 Calculating Theoretical Yield and Percent Yield

Adipic acid, $H_2C_6H_8O_4$, used to produce nylon, is made commercially by a 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 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, what is the percent yield for the reaction?

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 to calculate the theoretical yield of a product $H_2C_6H_8O_4$ and the percent yield if only 33.5 g of product is obtained.

Plan

(a) The theoretical yield, which is the calculated quantity of adipic acid formed, can be calculated using the sequence of conversions shown in Figure 3.16.

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