WHAT'S AHEAD

3 CHEMICAL REACTIONS AND STOICHIOMETRY

3.1 The Conservation of Mass, Chemical Equations, and Stoichiometry

Much of the energy that we use in our daily lives, including transportation, comes from chemical reactions. In the internal combustion engine, fuel and air react to produce energy, along with gaseous byproducts (mostly carbon dioxide and water) that exit the exhaust pipes of vehicles, an issue that is central to discussions surrounding climate change. Suppose we wanted to know how many molecules of CO_2 are produced by burning a certain amount of fuel. The tools of chemistry we will learn in this section enable us to accurately determine how many molecules of hydrocarbon and oxygen are consumed and how many molecules of byproducts are generated.

When you finish this section, you should be able to:

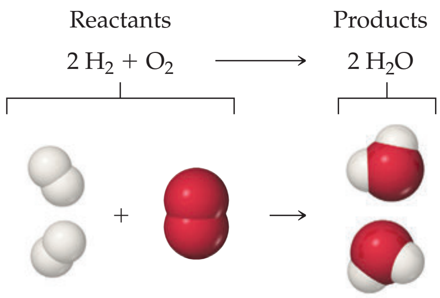
- Explain the law of conservation of mass in terms of reactants and products in a chemical equation
- Balance chemical equations by writing out the chemical formulas (and appropriate coefficients) of the reactants and products in chemical reactions

Stoichiometry (pronounced stoy-key-OM-uh-tree, roughly meaning "element measure" in Greek) is the area of study that examines the quantities of substances consumed and produced in chemical reactions. Chemists and chemical engineers use stoichiometry every day in running the reactions of the worldwide chemical industry.

Stoichiometry is built on an understanding of atomic weights (Section 2.4), chemical formulas, and the **law of conservation of mass** (Section 2.1). This important principle tells us that: *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.

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

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



▲ **Figure 3.1 A balanced chemical equation.**

We read the + sign as "reacts with" and the arrow as "produces." The chemical formulas on the left side of the arrow represent the starting substances, called **reactants**. The chemical formulas on the right side 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*).

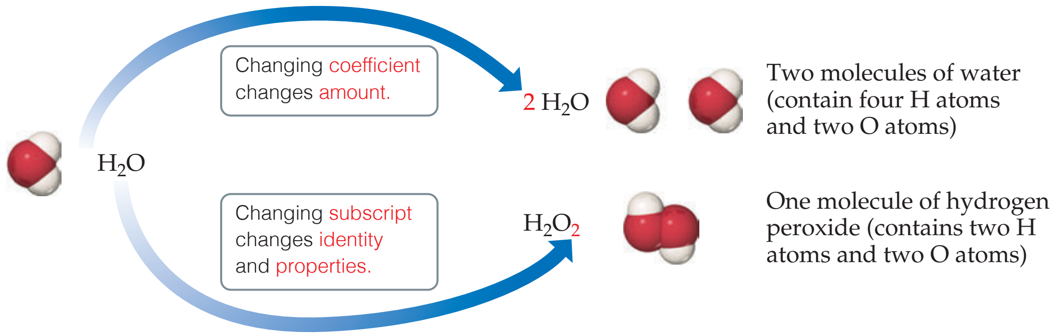
Because atoms are neither created nor destroyed in any chemical reaction, a balanced chemical equation must have an equal number of atoms of each element on each side of the arrow. On the right side of Equation 3.1, for example, there are two molecules of H_2O, each composed of two atoms of hydrogen and one atom of oxygen (**Figure 3.1**). 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.

How to Balance Chemical Equations

Chemists write chemical equations to identify the reactants and products in a reaction. To determine the amount of product that can be made, or the amount of a reactant that is required, the equation needs to be *balanced* using stoichiometry.

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. 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 3.2** 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.



▶ **Figure 3.2 The difference between changing subscripts and changing coefficients in chemical equations.**

A Step-by-Step Example of Balancing a Chemical Equation

To illustrate the process of balancing an equation, consider the reaction that occurs when methane (CH_4), the principal component of natural gas, burns in air to produce carbon dioxide gas (CO_2) and water vapor (H_2O) (**Figure 3.3**). Both 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 first balance 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 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: there are 2 O atoms on the left-hand side, and a total of 4 O atoms on the right-hand side. Adding the coefficient 2 in front of O_2 balances the equation by giving four O atoms on each side:

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

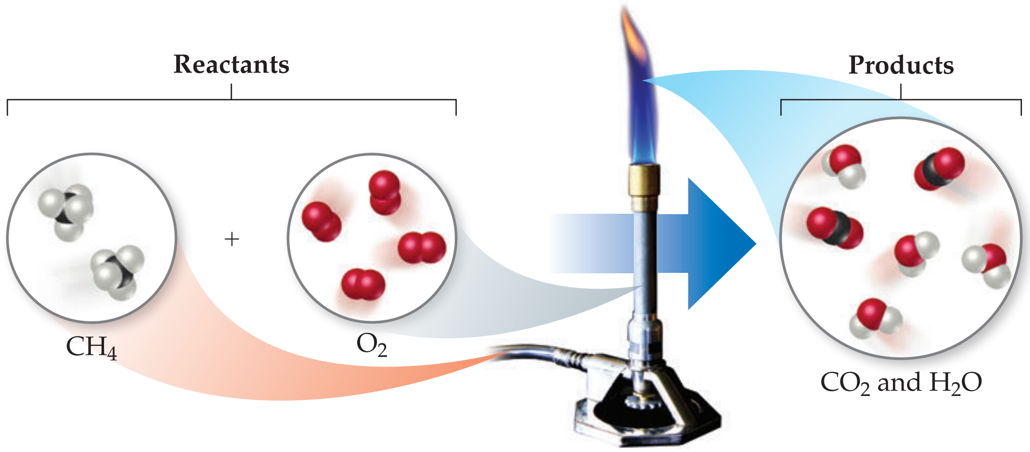
We can provide even more information in a chemical equation: the physical state of the reactants, products, and more details about what conditions are required for the reaction to proceed. We use the symbols (*g*), (*l*), (*s*), and (*aq*) for substances that are gases, liquids, solids, and dissolved in aqueous (water) solution, respectively. Thus, Equation 3.4 is fully written as:

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

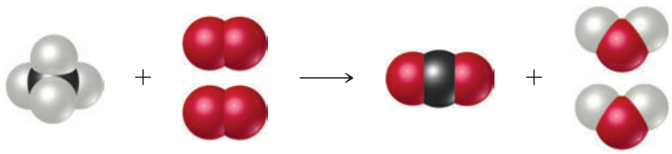
Symbols that represent the conditions under which the reaction proceeds can appear above or below the reaction arrow. One example that we will encounter later in this chapter involves the symbol Δ (Greek uppercase delta); a delta above the reaction arrow indicates the addition of heat.

For Equation 3.5, Figures 3.3 and 3.4 provide molecular views of the reaction as it would happen in a Bunsen burner and the balanced reaction, respectively.

Go Figure In the molecular level views shown in the figure, how many C, H, and O atoms are present as reactants? Are the same number of each type of atom present as products?



▲ **Figure 3.3** **Methane reacts with oxygen in a Bunsen burner.**

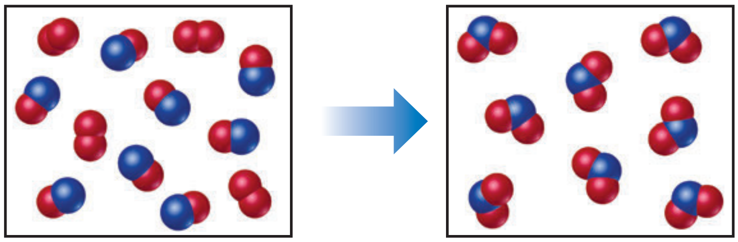


$$\underbrace{CH_4 + 2\,O_2}_{(1\,C,\,4\,H,\,4\,O)} \longrightarrow \underbrace{CO_2 + 2\,H_2O}_{(1\,C,\,4\,H,\,4\,O)}$$

◀ **Figure 3.4** **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:

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

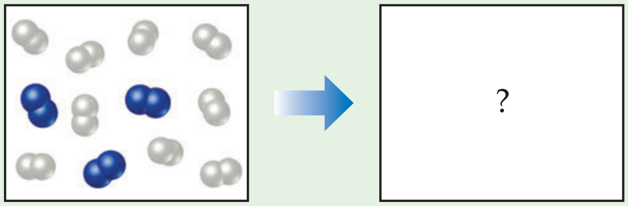
Continued

(**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

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, how many NH_3 molecules should be shown in the right (products) box?
(**a**) 2 (**b**) 3 (**c**) 4 (**d**) 6 (**e**) 9

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

Balance these equations by providing the missing coefficients:

(**a**) $\_\_\,Fe(s) + \_\_\,O_2(g) \longrightarrow \_\_\,Fe_2O_3(s)$

(**b**) $\_\_\,Al(s) + \_\_\,HCl(aq) \longrightarrow \_\_\,AlCl_3(aq) + \_\_\,H_2(g)$

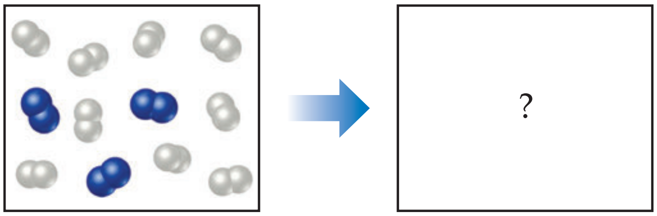
(**c**) $\_\_\,CaCO_3(s) + \_\_\,HCl(aq) \longrightarrow \_\_\,CaCl_2(aq) + \_\_\,CO_2(g) + \_\_\,H_2O(l)$

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.1 How many atoms of oxygen are represented by the notation $3\,Mg(OH)_2$?
(a) 1 **(b)** 2 **(c)** 3 **(d)** 5 **(e)** 6

3.2 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, how many NH_3 molecules should be shown in the right (products) box?

(a) 2 **(b)** 3 **(c)** 4 **(d)** 6 **(e)** 9

3.3 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 with the smallest possible integers, what is the value of the coefficient in front of bromine, Br_2?

(a) 1 **(b)** 2 **(c)** 3 **(d)** 4 **(e)** 6

Exercises

3.4 Write "true" or "false" for each statement. (**a**) We balance chemical equations as we do because energy must be conserved. (**b**) If the reaction $2\,O_3(g) \rightarrow 3\,O_2(g)$ goes to completion and all O_3 is converted to O_2, then the mass of O_3 at the beginning of the reaction must be the same as the mass of O_2 at the end of the reaction. (**c**) You can balance the "water-splitting" reaction $H_2O(l) \rightarrow H_2(g) + O_2(g)$ by writing it this way: $H_2O_2(l) \rightarrow H_2(g) + O_2(g)$.

3.5 Balance the following equations:

(a) $SiCl_4(l) + H_2O(l) \longrightarrow Si(OH)_4(s) + HCl(aq)$

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

(c) $Al(OH)_3(s) + H_2SO_4(l) \longrightarrow Al_2(SO_4)_3(s) + H_2O(l)$

(d) $H_3PO_4(aq) \longrightarrow H_4P_2O_7(aq) + H_2O(l)$

3.6 Balance the following equations:

(a) $CaS(s) + H_2O(l) \longrightarrow Ca(HS)_2(aq) + Ca(OH)_2(aq)$

(b) $NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$

(c) $FeCl_3(s) + Na_2CO_3(aq) \longrightarrow Fe_2(CO_3)_3(s) + NaCl(aq)$

(d) $FeS_2(s) + O_2(g) \longrightarrow Fe_2O_3(s) + SO_2(g)$

3.7 Write balanced chemical equations corresponding to each of the following descriptions: (**a**) Potassium cyanide reacts with an aqueous solution of sulfuric acid to form hydrogen cyanide gas. (**b**) When an aqueous solution of ammonium nitrite (NH_4NO_2) reacts with an aqueous solution of potassium hydroxide, ammonia gas, water and metal nitrate is formed. (**c**) When hydrogen gas is passed over solid hot iron(III) oxide, the resulting reaction produces iron and gaseous water. (**d**) When liquid ethanoic acid (CH_3COOH) is combusted, carbon dioxide and water are formed.

Answers to Self-Assessment Exercises

3.1 (e) **3.2** (d) **3.3** (d)

3.2 | Simple Patterns of Chemical Reactivity: Combination, Decomposition, and Combustion

One of the great triumphs of chemistry over the last hundred years is the development of fertilizers that enable us to feed the world. Ammonia, NH_3, is one of the principle chemicals farmers use to increase crop yield. The industrial process that are used to convert the elements nitrogen and hydrogen into ammonia is one of the most important chemical reactions in the world.

In this section, we will learn about broad classes of chemical reactions. When you finish this section, you should be able to:

- Recognize chemical reactions that are combination, decomposition, or combustion reactions
- Predict the products of these reactions
- Balance chemical equations for these reactions

Combination and Decomposition Reactions

In **combination reactions**, two or more substances react to form one product, according to

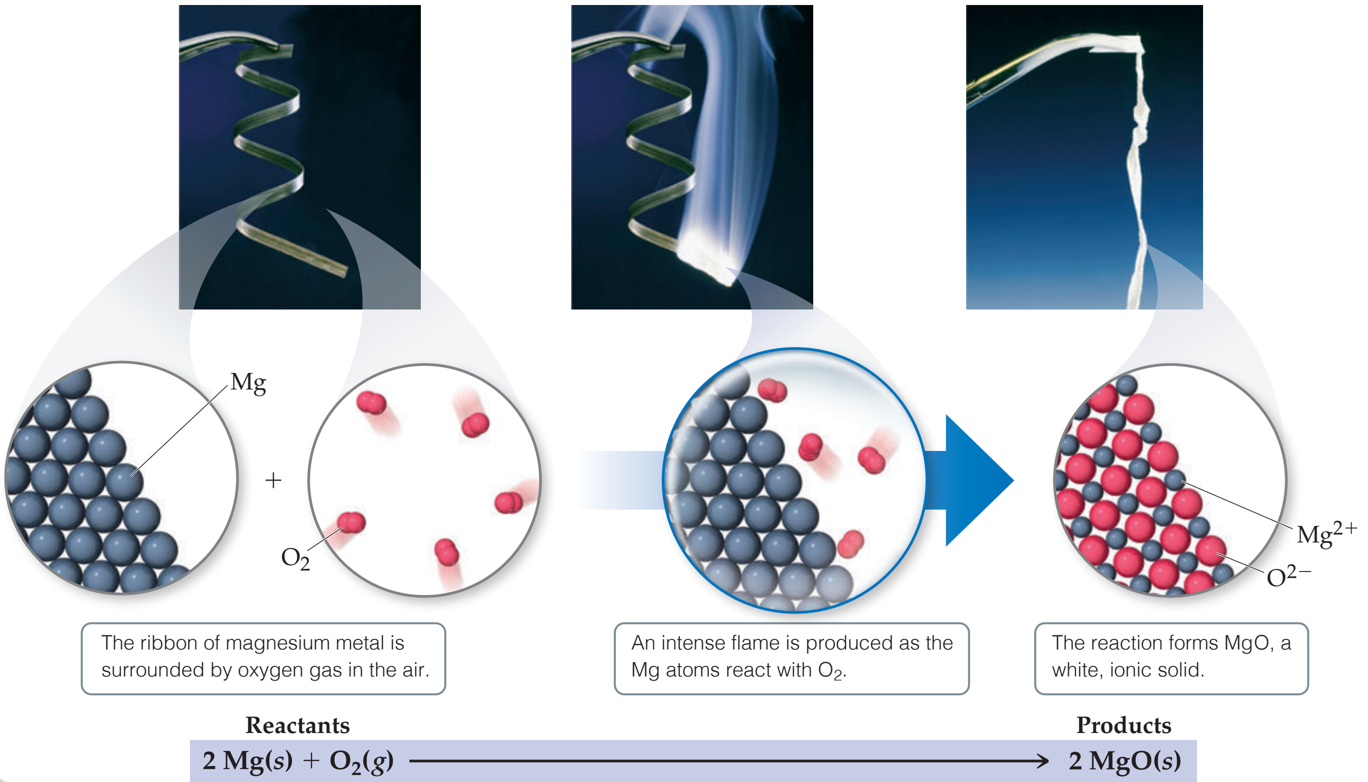
$$A + B \longrightarrow C \qquad [3.6]$$

For example, magnesium metal burns brilliantly in air to produce magnesium oxide. This reaction is used to produce the bright white flame generated by flares and some fireworks (**Figure 3.5**):

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s) \qquad [3.7]$$

A combination reaction between a metal and a nonmetal, as in Equation 3.7, 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.



▲ **Figure 3.5 Combustion of magnesium metal in air, a combination reaction.**

Other examples of combination reactions include the formation of small gaseous molecules such as CO_2 and NH_3 from their elements; and the reaction of calcium oxide with water to produce calcium hydroxide:

$$C(s) + O_2(g) \longrightarrow CO_2(g) \quad [3.8]$$

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g) \quad [3.9]$$

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq) \quad [3.10]$$

In a **decomposition reaction**, a single substance undergoes a reaction to produce two or more products:

$$C \longrightarrow A + B \quad [3.11]$$

An example of a decomposition reaction is when a metal carbonate decomposes to a metal oxide and carbon dioxide upon heating:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g) \quad [3.12]$$

Industrially, the decomposition of calcium carbonate at high temperatures is quite important. Limestone and seashells, natural sources of calcium carbonate, are heated to prepare calcium oxide, known as “quicklime” or “lime.” Tens of millions of tons of CaO are used each year in making glass, in metallurgy to isolate metals from ores, in cement, and in steel manufacturing to remove impurities.

Another important example of a decomposition reaction is the decomposition of sodium azide, NaN_3, to form sodium metal and nitrogen gas:

$$2\,NaN_3(s) \longrightarrow 2\,Na(s) + 3\,N_2(g) \quad [3.13]$$

This reaction is what you find in automobile airbags (**Figure 3.6**). Approximately 100 g of NaN_3, upon physical impact, will explosively produce about 50 L of nitrogen gas.

▲ **Figure 3.6 Decomposition of sodium azide, $NaN_3(s)$, produces $N_2(g)$ that inflates air bags in automobiles.**

Combustion Reactions

Combustion reactions are rapid reactions that produce a flame. Most combustion reactions involve O_2 from air as a reactant. The combustion of hydrocarbons (compounds that contain only carbon and hydrogen) in air, illustrated in Equation 3.5, is a major energy-producing process in our world.

Hydrocarbons combusted in air react with O_2 to form CO_2 and H_2O. The number of molecules of O_2 required, as well as the number of product molecules 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 3.7**), a gas used for cooking and home heating, is described by the chemical equation

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g) \quad [3.14]$$

Go Figure

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

▲ **Figure 3.7 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.

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.
$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g)$	
$PbCO_3(s) \longrightarrow PbO(s) + CO_2(g)$	
$Cu(OH)_2(s) \longrightarrow CuO(s) + H_2O(g)$	

The state of the water in this reaction is listed as gas here, since the propane flame burns at a high temperature; but depending on conditions, the water molecules that are produced could be in the gas or liquid phase.

Millions of compounds are made only of carbon, hydrogen, and oxygen. Notable classes of such molecules are sugars and alcohols. Combustion of these oxygen-containing derivatives of hydrocarbons in air also produces CO_2, H_2O and energy. Many of the substances that function as energy sources in metabolism, such as the sugar glucose ($C_6H_{12}O_6$), react with O_2 to ultimately 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 called *oxidation reactions* rather than combustion reactions.

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 2 in the periodic table, which further suggests they react in the same way:

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

Practice Exercise

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**) $Ag_2O(s) \longrightarrow 2\,Ag(s) + O_2(g)$

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

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

(**a**) $C_2H_4(OH)_2(l) + 5\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(**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**) $4\,C_2H_4(OH)_2(l) + 10\,O_2(g) \longrightarrow 8\,CO_2(g) + 12\,H_2O(g)$

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

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

(a) NaS **(c)** NaS_2 **(e)** Na_3S_2
(b) Na_2S **(d)** Na_2S_3

3.9 Which of the following reactions is the balanced equation for the decomposition of silver(I) oxide?

(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) $2Ag_2O(s) \longrightarrow 4\,Ag(s) + O_2(g)$
(e) $Ag_2O(s) \longrightarrow 2\,Ag(s) + O_2(g)$

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

(a) $C_2H_4(OH)_2(l) + 5\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$
(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) $4\,C_2H_4(OH)_2(l) + 10\,O_2(g) \longrightarrow 8\,CO_2(g) + 12\,H_2O(g)$

Exercises

3.11 (**a**) When the metallic element lithium combines with the nonmetallic element chlorine, $Cl_2(g)$, what is the chemical formula of the product? (**b**) Is the product a solid, liquid, or gas at room temperature? (**c**) In the balanced chemical equation for this reaction, what is the coefficient in front of the product if the coefficient in front of $Cl_2(g)$ is 1?

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

3.13 Balance the following equations and indicate whether they are combination, decomposition, or combustion reactions:

(a) $C_7H_{16}(s) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$
(b) $Li_3N(s) + BN(s) \longrightarrow Li_3BN_2(s)$
(c) $Zn(OH)_2(s) \longrightarrow ZnO(s) + H_2O(l)$
(d) $Ag_2O(s) \longrightarrow Ag(s) + O_2(g)$

Answers to Self-Assessment Exercises

3.8 (b) **3.9** (d) **3.10** (b)

3.3 Formula Weights and Elemental Compositions of Substances

Sulfuric acid, $H_2SO_4(l)$, is a common laboratory chemical that is used by the metric ton in many reactions in the chemical industry. We can see from the molecular model on the left that one molecule of H_2SO_4 contains one sulfur atom (yellow), four oxygen atoms (red), and two hydrogen atoms (white). But in the lab, we dispense milliliters of liquid H_2SO_4, or more commonly, milliliters of its aqueous solution, $H_2SO_4(aq)$. How can we connect the chemical equations that we write, which represent individual molecules, to the reactions we do in the lab, where quantities are measured in grams or milliliters? That is the topic we explore in this section.

When you finish this section, you should be able to:

- Calculate the formula weight of a substance from its empirical formula or it's molecular weight from its molecular formula
- Calculate the elemental composition of a substance from the mass percentages of the elements that make up the substance

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 the atomic weights from the periodic table, we find, for example, that the formula weight of sulfuric acid, H_2SO_4, is 98.1 u (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 weight 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 (for Na, this would be 23.1 amu). If the chemical formula is that of a molecule, like H_2SO_4, the formula weight can also be called the **molecular weight** (MW).

Not all substances, though, are molecules. For instance, ionic substances such as calcium chloride exist as three-dimensional arrays of ions (see Figure 2.18). In these cases, the empirical formula is used as the formula unit, and the formula weight is the sum of 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

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

Elemental Compositions of Substances

Let's say you are a forensic chemist, working in a crime lab. Your colleagues find a mysterious white powder at a crime scene. Is it salt, sugar, methamphetamine, cocaine or something else?

One way to determine the identity of a substance is to measure its **elemental composition** and compare it to the calculated elemental compositions of possible candidate substances. We do these calculations based on the masses of each element in the compound:

$$\begin{matrix}\text{\% mass composition} \\ \text{of element}\end{matrix} = \frac{\begin{pmatrix}\text{number of atoms} \\ \text{of element}\end{pmatrix}\begin{pmatrix}\text{AW} \\ \text{of element}\end{pmatrix}}{\text{FW of substance}} \times 100\% \qquad [3.15]$$

The sum of all the mass percentages of each element in the compound must add up to 100%.

As an example, let's calculate the mass percentage of sulfur in sulfuric acid. Based on atoms we can see that for each H_2SO_4 molecule, one atom out of seven is sulfur. But that does not mean that 1/7 of the mass of the compound is sulfur, since the atoms weigh different amounts:

$$\%\,S \text{ in } H_2SO_4 = \frac{(1)(32.1 \text{ amu})}{98.1 \text{ amu}} \times 100\% = 32.7\%$$

So we find that almost a third of the mass of a given quantity of pure H_2SO_4 is due to sulfur.

Sample Exercise 3.5

Calculating Formula Weights

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

SOLUTION

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

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

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

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

Practice Exercise

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

STRATEGIES FOR SUCCESS 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 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 1.7) 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?

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

We'll use the steps outlined in the Strategies For Success: Problem Solving feature to answer the question.

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

$$\%C = \frac{(12)(12.0\,\cancel{u})}{342.0\,\cancel{u}} \times 100\% = 42.1\%$$

$$\%H = \frac{(22)(1.0\,\cancel{u})}{342.0\,\cancel{u}} \times 100\% = 6.4\%$$

$$\%O = \frac{(11)(16.0\,\cancel{u})}{342.0\,\cancel{u}} \times 100\% = 51.5\%$$

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

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%

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.14 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

3.15 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%

3.16 A mysterious white powder found at a crime scene is analyzed and contains 66.8 ± 0.5% carbon by mass. One of the investigators hypothesizes that the substance is cocaine ($C_{17}H_{21}NO_4$). What percent carbon, by mass, is in cocaine?

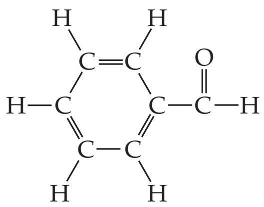
(a) 39.5% **(b)** 64.3%
(c) 67.3% **(d)** 70.6%
(e) 72.5%

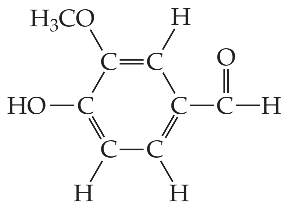
Exercises

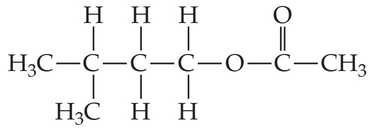
3.17 Determine the formula weights of each of the following compounds: (**a**) lead (IV) chloride; (**b**) copper(II) oxide; (**c**) iodic acid, HIO_3; (**d**) sodium perchlorate, $NaClO_4$; (**e**) indium nitride, (**f**) phosphorus pentoxide, P_4O_{10}; (**g**) boron trichloride.

3.18 Calculate the percentage by mass of oxygen in the following compounds: (**a**) vanillin, $C_8H_8O_3$; (**b**) isopropyl alcohol, C_3H_8O; (**c**) acetaminophen, $C_8H_9NO_2$; (**d**) cyclopropanone, C_3H_4O; (**e**) dioxin, $C_{12}H_4Cl_4O_2$; (**f**) penicillin, $C_{16}H_{18}N_2O_4S$.

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

(a) 
Benzaldehyde (almond fragrance)

(b) 
Vanillin (vanilla flavor)

(c) 
Isopentyl acetate (banana flavor)

Answers to Self-Assessment Exercises

3.14 (a) **3.15** (b) **3.16** (c)

3.4 Avogadro's Number and the Mole; Molar Mass

Even the smallest samples in the laboratory contain enormous numbers of atoms, ions or molecules.For example, a teaspoon (about 5 mL) of water contains 2×10^{23} water molecules, a number so large it defies comprehension. Chemists, therefore, have devised a way to conveniently count such enormous numbers.

When you finish this section, you should be able to:

- Explain the *mole* and the origin of *Avogadro's number*
- Calculate the *molar mass* for a compound and relate this to its formula weight
- Convert between grams, molecules, and moles of a substance

The Mole and Avogadro's Number

In everyday life, we use counting units such as dozen (12 objects), score (20 objects), gross (144 objects) or ream (500 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 as the number of atoms in exactly 12 g of isotopically pure ^{12}C. From experiments, scientists have determined this number to be 6.0221415×10^{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, $\mathbf{6.02 \times 10^{23}\ mol^{-1}}$.* 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:

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

Avogadro's number is so large that it is difficult to imagine.Spreading 6.02×10^{23} marbles over the 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 the Earth 300 trillion(3×10^{14}) times.

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 1 mol of ^{12}C has a mass of 12 g (by definition), 1 mol of ^{24}Mg must have a mass of 24 g.This example illustrates a general rule relating the mass of an atom 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, Cl has an atomic weight of 35.5 amu; therefore, 1 mol of Cl atoms has a mass of 35.5 g.

Au has an atomic weight of 197 amu; therefore, 1 mol of Au atoms has a mass of 197 g.

The same relationship holds for the formula weight or molecular weight of a substance and the mass of 1 mol of that substance:

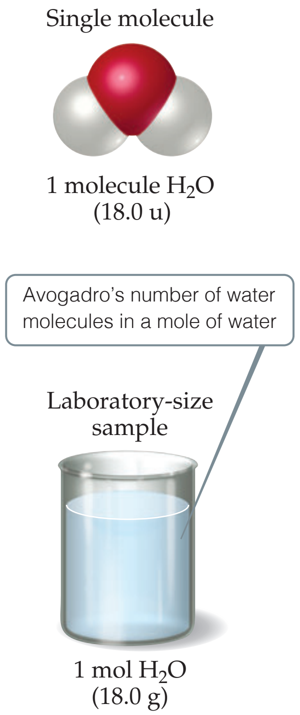
H_2O has a molecular weight of 18.0 amu; therefore, 1 mol of H_2O has a mass of 18.0 g

NaCl has a formula weight of 58.5 amu; therefore 1 mol of NaCl has a mass of 58.5 g.

The mass in grams of one mole of a substance is called the **molar mass** of the substance.The units of molar mass are g/mol, or g mol^{-1}.*The molar mass in grams per mole of any substance is numerically equal to its formula weight in atomic mass units*.For NaCl, for

Go Figure

What number do you get if you divide the mass of 1 mol of water by the mass of 1 molecule of water?

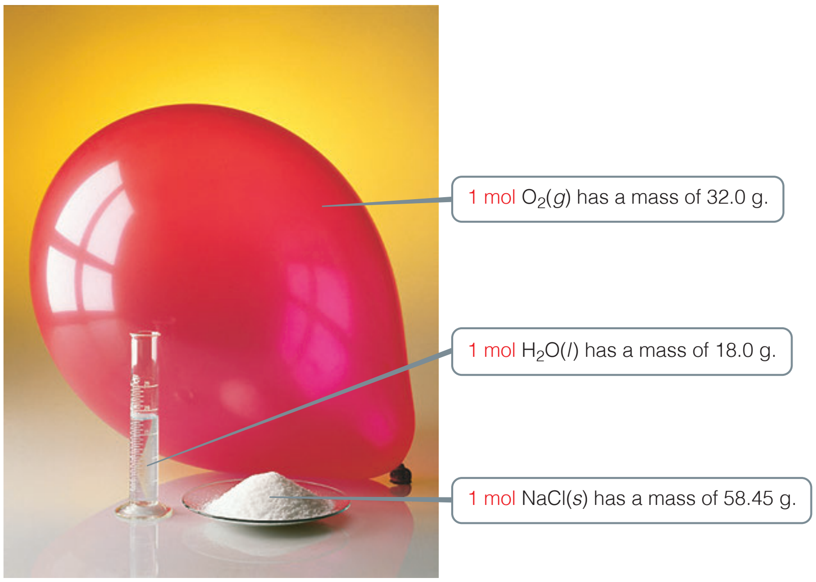


▲ **Figure 3.8 Comparing the mass of 1 molecule and 1 mol of H_2O.** Both masses have the same number but different units (atomic mass units and grams). Expressing both masses in grams indicates their huge difference: 1 molecule of H_2O has a mass of 2.99×10^{-23} g, whereas 1 mol H_2O has a mass of 18.0 g.

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 or "dinitrogen"	N_2	28.0	28.0	$\left\{ \begin{array}{l} 6.02 \times 10^{23}\ N_2 \text{ molecules} \\ 2(6.02 \times 10^{23})\ N \text{ atoms} \end{array} \right.$
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	$\left\{ \begin{array}{l} 6.02 \times 10^{23}\ BaCl_2 \text{ formula units} \\ 6.02 \times 10^{23}\ Ba^{2+} \text{ ions} \\ 2(6.02 \times 10^{23})\ Cl^- \text{ ions} \end{array} \right.$

[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 3.9 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.

instance, the formula weight is 58.5 amu and therefore its molar mass is 58.5 g/mol.Mole relationships for several substances are shown in Table 3.1, and **Figures 3.8** and **3.9** illustrate 1 mol quantities of 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. For instance, suppose you read that 1 mol of nitrogen is used to make ammonia.You might interpret this statement to mean that 1 mol of nitrogen atoms was used (14.0 g). Unless otherwise stated, however, what is meant is 1 mol of nitrogen molecules, N_2 (28.0 g), was used, because N_2 is the naturally occurring form of the element. To avoid ambiguity, it is important to explicitly state the chemical form being discussed.Using the chemical formula—N or N_2, for instance—avoids any confusion.

Converting Between Masses, Moles, and Atoms/Molecules/Ions

Now we are ready to learn how to relate chemical equations to the amounts of chemicals used and produced in reactions.We will use the mole concept with dimensional analysis (Section 1.7) to help us (**Figure 3.10**).

Go Figure What units would you put under "molar mass" and "Avogadro's number" on this diagram?

Grams ⟷ Use molar mass ⟷ Moles ⟷ Use Avogadro's number ⟷ Formula units

▲ **Figure 3.10 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.

For example, let's calculate how many copper atoms are in an old copper penny. Such a penny has a mass of about 3 grams, and let's assume for simplicity that the penny is pure copper:

$$\begin{aligned}\text{Number of Cu atoms} &= (3\text{ g Cu})\left(\frac{1\text{ mol Cu}}{63.5\text{ g Cu}}\right)\left(\frac{6.02\times 10^{23}\text{ Cu atoms}}{1\text{ mol Cu}}\right)\\ &= 3\times 10^{22}\text{ Cu atoms}\end{aligned}$$

We have rounded the answer to one significant figure because we only used one significant figure for the mass of the penny. Notice how dimensional analysis provides a straightforward route from grams to number of atoms.The molar mass and Avogadro's number are used as conversion factors to convert grams to moles and then moles to atoms.Also notice that our answer is a very large number: this makes sense, since there are enormous numbers of atoms in a macroscopic sample we can pick up with our hands.Any time that you calculate the number of atoms, ions, or molecules in a laboratory-scale sample, you should expect the number to be very large.However, if you were to calculate number of moles in a laboratory sample, the number may not be that large, and in fact might be less than 1.That is true for the moles of copper in our old penny:

$$\begin{aligned}\text{Moles of Cu} &= (3\text{ g Cu})\left(\frac{1\text{ mol Cu}}{63.5\text{ g Cu}}\right)\\ &= 5\times 10^{-2}\text{ mol Cu}\end{aligned}$$

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, so that normal blood glucose levels are 70–120 mg/dL. 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. 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 a 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.**

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 $\times$ 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 $\times$ 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 $= 6.02 \times 10^{23}$ 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

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

Sample Exercise 3.8

Converting Grams to Moles

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

SOLUTION

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

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

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

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

Check Because 5.380 g is less than the molar mass, 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

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

Sample Exercise 3.9

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 u. Using 1 mol $Ca(NO_3)_2$ = 164.1 g $Ca(NO_3)_2$ to write the appropriate conversion factor, we have

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

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

Practice Exercise

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

Sample Exercise 3.10

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/mol, and the second conversion uses Avogadro's number.

(**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

(**a**) Convert grams $C_6H_{12}O_6$ to molecules $C_6H_{12}O_6$.

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

(**b**) Convert molecules $C_6H_{12}O_6$ to atoms O.

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

Check

(**a**) 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.

(**b**) 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

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}

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.20 Which of the following samples contains the fewest sodium atoms?
(a) 1.0 mol sodium oxide
(b) 45 g sodium fluoride
(c) 50 g sodium chloride
(d) 1.0 mol sodium nitrate

3.21 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+

3.22 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}

Exercises

3.23 (**a**) Write "true" or "false" for each statement. (**a**) A mole of ducks contain a mole of feathers. (**b**) A mole of ammonia gas has a mass of 17.0 g. (**c**) The mass of 1 ammonia molecule is 17.0 g. (**d**) A mole of $MgSO_4(s)$ contains 4 moles of oxygen atoms.

3.24 Without doing any detailed calculations (but using a periodic table to give atomic weights), rank the following samples in order of increasing numbers of atoms: 0.5 mol BCl_3 molecules, 197 g gold, 6.0×10^{23} CCl_4 molecules.

3.25 Calculate the following quantities:

(**a**) mass, in grams, of 0.105 mol sucrose ($C_{12}H_{22}O_{11}$)

(**b**) moles of $Zn(NO_3)_2$ in 143.50 g of this substance

(**c**) number of molecules in 1.0×10^{-6} mol CH_3CH_2OH

(**d**) number of N atoms in 0.410 mol NH_3

3.26 (**a**) What is the mass, in grams, of 2.50×10^{-3} mol of ammonium phosphate?

(**b**) How many moles of chloride ions are in 0.2550 g of aluminum chloride?

(**c**) What is the mass, in grams, of 7.70×10^{20} molecules of caffeine, $C_8H_{10}N_4O_2$?

(**d**) What is the molar mass of cholesterol if 0.00105 mol has a mass of 0.406 g?

3.27 The molecular formula of saccharin, an artificial sweetener, is $C_7H_5NO_3S$. (**a**) What is the molar mass of saccharin? (**b**) How many moles of sachharin are in 2.00 mg of this substance? (**c**) How many molecules are in 2.00 mg of this substance? (**d**) How many C atoms are present in 2.00 mg of saccharin?

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

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

Answers to Self-Assessment Exercises

3.20 (c) **3.21** (b) **3.22** (c)

3.5 Formula Weights and Elemental Compositions of Substances

The empirical formula for a substance tells us the relative number of atoms of each element in the substance. (Section 2.6) 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 from experimental data.

When you finish this section, you should be able to:

- Calculate the empirical formula of a compound from the mass percentages of the elements that make up the compound.
- Calculate the molecular formula of a compound from its molar mass and empirical formula.

Mercury and chlorine 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 calculate the number of moles of each element in the sample:

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

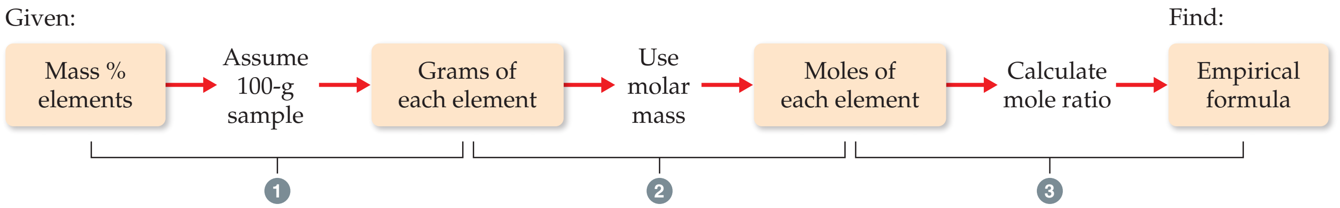
$$(26.0\ \cancel{\text{g Cl}})\left(\frac{1\ \text{mol Cl}}{35.5\ \cancel{\text{g Cl}}}\right) = 0.732\ \text{mol Cl}$$

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.

The general procedure for determining empirical formulas is outlined in **Figure 3.12**



▲ **Figure 3.12 Procedure for calculating an empirical formula from percentage composition.**

Sample Exercise 3.11

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.

Continued

Solve

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

Mass % elements → Assume 100-g sample → Grams of each element → Use molar mass → Moles of each element → Calculate mole ratio → Empirical formula

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.

Mass % elements → Assume 100-g sample → Grams of each element → Use molar mass → Moles of each element → Calculate mole ratio → Empirical formula

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

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

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

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

Mass % elements → Assume 100-g sample → Grams of each element → Use molar mass → Moles of each element → Calculate mole ratio → Empirical formula

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

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:

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

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

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

Molecular Formulas from Empirical Formulas

For molecular substances, the empirical formula and the molecular formula are often different. For example, benzene has a molecular formula of C_6H_6, but it's empirical formula CH is the same as that of the gas acetylene, whose molecular formula is C_2H_2. Knowledge of the empirical formula is not sufficient to differentiate these two very different compounds. Fortunately, we can obtain the molecular formula for any compound from its empirical formula if we know either the molecular weight of the compound, which can be measured by a variety of methods, including mass spectrometry (link to Section 2.4). *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:

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

For example, the empirical formula of ascorbic acid (vitamin C) is $C_3H_4O_3$.. This means the empirical formula weight is 3(12.0 amu) + 4(1.0 amu) + 3(16.0 amu) = 88.0 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 3.12

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.

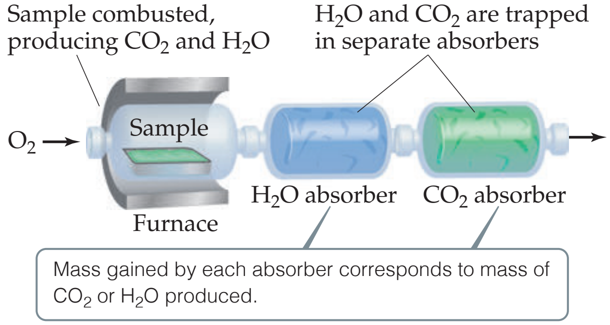
Practice Exercise

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

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 3.13**, the carbon is converted to CO_2 and the hydrogen is converted to H_2O. (Section 3.2). 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.



◀ **Figure 3.13 Apparatus for combustion analysis.**

Sample Exercise 3.13

Determining an Empirical Formula by Combustion Analysis

2-propanol, sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of 2-propanol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of 2-propanol.

SOLUTION

Analyze We are told that 2-propanol 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 2-propanol, 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, we find that 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

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

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

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)** $C_4H_8O_2$

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.30 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_2

(b) $COCl_2$

(c) $C_{0.022}O_{0.022}Cl_{0.044}$

(d) C_2OCl_2

3.31 Cyclohexane, a commonly used organic solvent, is 85.6% carbon and 14.4% hydrogen 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

3.32 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) $C_4H_8O_2$

Exercises

3.33 Give the empirical formula of each of the following compounds if a sample contains (**a**) 0.052 mol C, 0.103 mol H, and 0.017 mol O; (**b**) 2.10 g nickel and 0.58 g oxygen; (**c**) 26.56% K, 35.41% Cr, and 38.03% O by mass.

3.34 Determine the empirical formulas of the compounds with the following compositions by mass:

(**a**) 74.0% C, 8.7% H, and 17.3% N

(**b**) 57.5% Na, 40.0% O, and 2.5% H

(**c**) 41.1% N, 11.8% H, and the remainder S

3.35 A compound whose empirical formula is XF_3 consists of 65% F by mass. What is the atomic mass of X?

3.36 What is the molecular formula of each of the following compounds?

(**a**) empirical formula CH, molar mass = 78.0 g/mol

(**b**) empirical formula OH, molar mass = 34.0 g/mol

3.37 Determine the empirical and molecular formulas of each of the following substances:

(**a**) Styrene, a compound used to make Styrofoam® cups and insulation, contains 92.3% C and 7.7% H by mass and has a molar mass of 104 g/mol.

(**b**) Caffeine, a stimulant found in coffee, contains 49.5% C, 5.15% H, 28.9% N, and 16.5% O by mass and has a molar mass of 195 g/mol.

(**c**) Monosodium glutamate (MSG), a flavor enhancer in certain foods, contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na, and has a molar mass of 169 g/mol.

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

3.39 Valproic acid, used to treat seizures and bipolar disorder, is composed of C, H, and O. A 0.165-g sample is combusted to produce 0.166 g of water and 0.403 g of carbon dioxide. What is the empirical formula for valproic acid? If the molar mass is 144 g/mol, what is the molecular formula?

3.40 Washing soda, a compound used to prepare hard water for washing laundry, is a hydrate, which means that a certain number of water molecules are included in the solid structure. Its formula can be written as $Na_2CO_3 \cdot xH_2O$, where x is the number of moles of H_2O per mole of Na_2CO_3. When a 2.558-g sample of washing soda is heated at 125 °C, all the water of hydration is lost, leaving 0.948 g of Na_2CO_3. What is the value of x?

Answers to Self-Assessment Exercises

3.30 (b) 3.31 (d) 3.32 (d)

3.6 | Reaction Stoichiometry

When a chemical reaction is carried out, it's vital to understand how much of each product will be produced and how much of each reactant will be consumed. To carry out chemical reactions without this knowledge can lead to unintended consequences. Maybe an expensive reactant will be wasted because much more of it was added than was needed. A reaction might generate more gas than the reaction container can hold, leading to an explosion. In some reactions, particularly those involving solids, it can be difficult to separate the desired product from excess reactants. In this section, you will learn how to calculate the quantities of reactants consumed and products produced, given a balanced chemical equation representing the reaction.

When you finish this section, you should be able to:

- Determine the number of grams (or moles) of a product formed in a chemical reaction given the number of grams (or moles) of the reactants, and vice versa

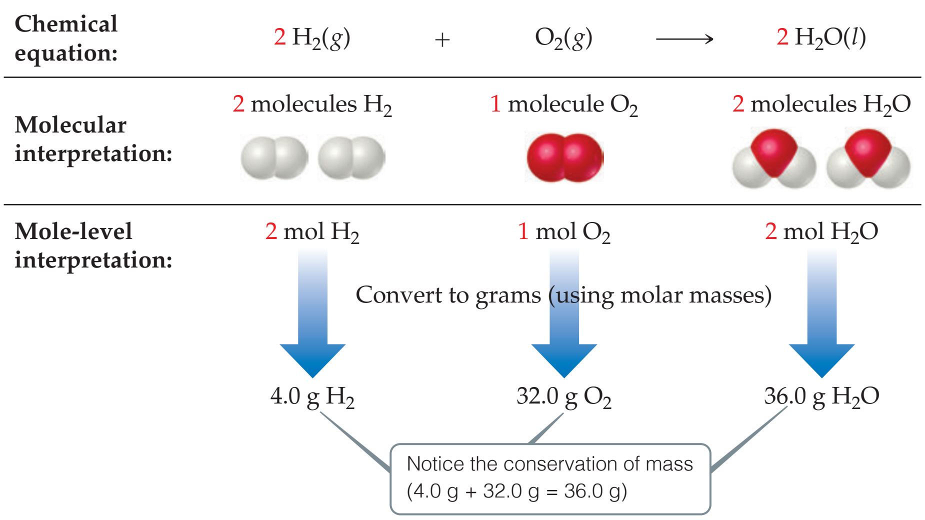
The coefficients in a chemical equation represent the relative numbers of molecules in a reaction. The mole concept allows us to convert this information to the masses of the substances in the reaction. For instance, the coefficients in the balanced equation:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l) \qquad [3.16]$$

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 3.14** shows how this result corresponds to the law of conservation of mass.



▲ **Figure 3.14 Interpreting a balanced chemical equation quantitatively.**

The quantities 2 mol H_2, 1 mol O_2, and 2 mol H_2O given by the coefficients in Equation 3.16 are called stoichiometrically equivalent quantities. The relationship between these quantities can be represented as

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

where the $\simeq$ symbol means "is stoichiometrically equivalent to." 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

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

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

$$2\ C_4H_{10}(l) + 13\ O_2(g) \longrightarrow 8\ CO_2(g) + 10\ H_2O(g) \qquad [3.17]$$

Let's calculate the mass of CO_2 produced when 1.00 g of C_4H_{10} is burned. The coefficients in Equation 3.17 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:

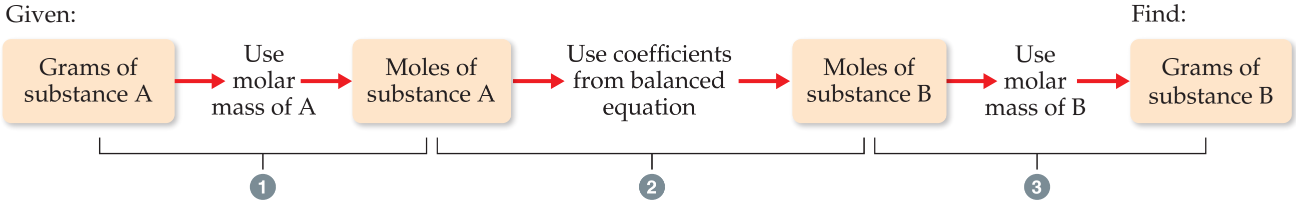
$$\text{Moles}\ CO_2 = (1.72 \times 10^{-2}\ \cancel{\text{mol}\ C_4H_{10}})\left(\frac{8\ \text{mol}\ CO_2}{2\ \cancel{\text{mol}\ C_4H_{10}}}\right)$$
$$= 6.88 \times 10^{-2}\ \text{mol}\ CO_2$$

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 3.15**. These three conversions can be combined in a single equation:

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



▲ **Figure 3.15 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.

To calculate the amount of O_2 consumed in the reaction of Equation 3.17, we again rely on the coefficients in the balanced equation for our stoichiometric factor, $2 \text{ mol } C_4H_{10} \simeq 13 \text{ mol } O_2$:

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

Many chemical reactions either consume or produce heat (Figure 3.7). This heat is also a stoichiometric quantity. For instance, if a reaction of a given number of reactant moles produces 100 J of energy in the form of heat, performing the reaction with twice the number of reactant moles will produce 200 J of heat. We will explore these ideas further in Chapter 5.

Sample Exercise 3.14
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:

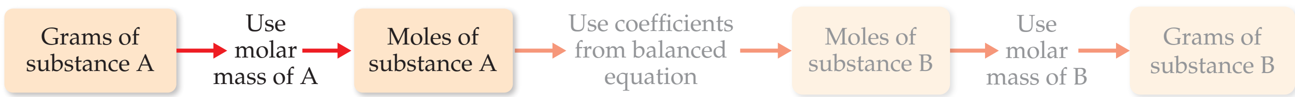
(1) Convert grams of $C_6H_{12}O_6$ to moles using the molar mass of $C_6H_{12}O_6$.

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

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

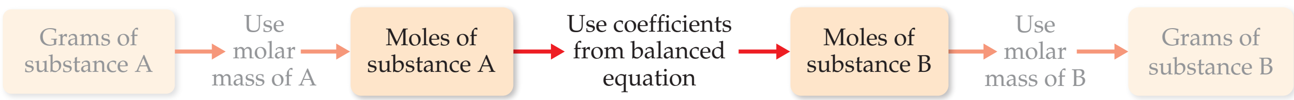
Solve

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



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

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



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

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

Grams of substance A → Use molar mass of A → Moles of substance A → Use coefficients from balanced equation → Moles of substance B → Use molar mass of B → Grams of substance B

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

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

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

Sample Exercise 3.15

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 3.15. The conversion of Step 1 requires the molar mass of LiOH $(6.94 + 16.00 + 1.01 = 23.95\text{ g/mol})$. The conversion of Step 2 is based on a stoichiometric relationship from the balanced chemical equation: 2 mol LiOH $\simeq$ 1 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

$$(1.00\text{ g } \cancel{\text{LiOH}})\left(\frac{1\ \cancel{\text{mol LiOH}}}{23.95\text{ g } \cancel{\text{LiOH}}}\right)\left(\frac{1\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol LiOH}}}\right)\left(\frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right)$$
$$= 0.919\text{ g } CO_2$$

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

Propane, C_3H_8 (Figure 3.7), 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

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.41 In this section, we learned that when 1.00 g of butane (C_4H_{10}) reacts with 3.59 g of oxygen (O_2) it produces 3.03 g of carbon dioxide (CO_2). True or False: Is it possible using only addition and/or subtraction, to calculate the number of grams H_2O produced.

3.42 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

3.43 Propane, C_3H_8, is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?
(a) 5.00 g
(b) 0.726 g
(c) 2.18 g
(d) 3.63 g

3.44 Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas. How many grams of calcium hydride are needed to form 4.50 g of hydrogen?
(a) 1.11 g
(b) 2.25 g
(c) 46.9 g
(d) 93.8 g

Exercises

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

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

(a) How many moles of HF are needed to react with 0.300 mol of Na_2SiO_3?

(b) How many grams of NaF form when 0.500 mol of HF reacts with excess Na_2SiO_3?

(c) How many grams of Na_2SiO_3 can react with 0.800 g of HF?

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

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

(a) Balance this equation.

(b) Calculate the number of grams of HCl that can react with 0.500 g of $Al(OH)_3$.

(c) Calculate the number of grams of $AlCl_3$ and the number of grams of H_2O formed when 0.500 g of $Al(OH)_3$ reacts.

(d) Show that your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

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

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

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

(a) How many moles of N_2 are produced by the decomposition of 1.50 mol of NaN_3?

(b) How many grams of NaN_3 are required to form 10.0 g of nitrogen gas?

(c) How many grams of NaN_3 are required to produce 10.0 ft^3 of nitrogen gas, about the size of an automotive air bag, if the gas has a density of 1.25 g/L?

3.49 A piece of aluminum foil 1.00 cm^2 and 0.550-mm thick is allowed to react with bromine to form aluminum bromide.

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

3.50 The complete combustion of octane, C_8H_{18}, produces 5470 kJ of heat. Calculate how many grams of octane is required to produce 20,000 kJ of heat.

Answers to Self-Assessment Exercises

3.41 True **3.42** (a) **3.43** (d) **3.44** (c)

3.7 Limiting Reactants

Often, the reactants used in a chemical reaction are not present in precise stoichiometric amounts. For example, a natural gas-fired power plant generates electricity by producing hot gases that drive turbines, predominantly through the following combustion reaction:

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g) \quad [3.18]$$

Power plants typically operate with an excess of $O_2(g)$ to achieve the maximum energy from the hydrocarbon fuel and minimize production of harmful byproducts, like carbon monoxide, that result from incomplete combustion. Consequently, the amount of CH_4 introduced determines how much CO_2 and H_2O water are produced, as well as the amount of energy released. In this section, we will learn how to do quantitative calculations for reactions where the reactants are not present in stoichiometrically equivalent quantities.

When you have completed this section, you should be able to:

- Identify limiting reactants and calculate amounts, in grams or moles, of reactants consumed and products formed in chemical reactions.
- Calculate the percent yield of a chemical reaction from the actual yield and the quantities of each reactant.

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, we can represent the recipe for making a sandwich like a chemical equation:

$$2\,Bd + Ch \longrightarrow Bd_2Ch \quad [3.19]$$

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.

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:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g) \quad [3.20]$$

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

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

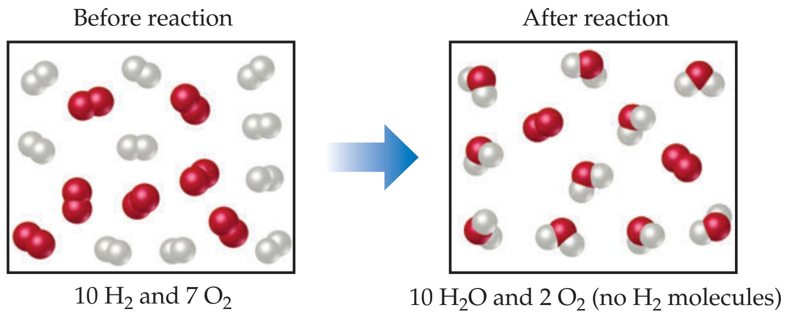
Because 7 mol O_2 is available at the start of the reaction, 7 mol O_2 − 5 mol O_2 = 2 mol O_2 is still present when all the H_2 is consumed.

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 3.16, 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 fuel while driving, the car stops because the fuel is the limiting reactant in the combustion reaction that moves the car. Before we leave the example illustrated in **Figure 3.16**, 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

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



▲ **Figure 3.16 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.

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 are found by adding the before and change quantities 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 3.16

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:

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

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.

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

▶ **Practice Exercise**

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 $O_2(g)$

Sample Exercise 3.17

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

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

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**

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) 1.20 g As **(b)** 1.50 g As **(c)** 4.30 g As **(d)** 8.30 g Ga

Theoretical and Percent Yields

The quantity of product calculated to form when all the 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. Some 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 is the actual yield divided by the theoretical yield, multiplied by 100 to convert to percent:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Sample Exercise 3.18

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 3.15.

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

Solve

(**a**) The theoretical yield is:

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

(**b**) The percent yield is:

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

Check We can check our answer in (**a**) by doing a ballpark calculation. From the balanced equation we know that each mole of cyclohexane gives 1 mol adipic acid. We have $25/84 \approx 25/75 = 0.3$ mol hexane, so we expect 0.3 mol adipic acid, which equals about $0.3 \times 150 = 45$ g, about the same magnitude as the 43.5 g obtained in the more detailed calculation given previously. In addition, our answer has the appropriate units and number of significant figures. In (**b**) the answer is less than 100%, as it must be from the definition of percent yield.

Practice Exercise

If 3.00 g of titanium metal is reacted with 6.00 g of chlorine gas, Cl_2, to form 7.7 g of titanium(IV) chloride in a combination reaction, what is the percent yield of the product?
(**a**) 65% (**b**) 96% (**c**) 48% (**d**) 86%

STRATEGIES FOR SUCCESS Design an Experiment

One of the most important skills you can learn in school is how to think like a scientist. Questions such as: "What experiment might test this hypothesis?", "How do I interpret these data?", and "Do these data support the hypothesis?" are asked every day by chemists and other scientists as they go about their work.

We want you to become a good critical thinker as well as an active, logical, and curious learner. For this purpose, starting in this chapter, we include at the end of each chapter a special exercise called "Design an Experiment." Here is an example:

Is milk a pure liquid or a mixture of chemical components in water? Design an experiment to distinguish between these two possibilities.

You might already know the answer—milk is indeed a mixture of components in water—but the goal is to think of how to demonstrate this in practice. Upon thinking about it, you will likely realize that the key idea for this experiment is separation: You can prove that milk is a mixture of chemical components if you can figure out how to separate these components.

Testing a hypothesis is a creative endeavor. While some experiments may be more efficient than others, there is often more than one good way to test a hypothesis. Our question about milk, for example, might be explored by an experiment in which you boil a known quantity of milk until it is dry. Does a solid residue form in the bottom of the pan? If so, you could weigh it and calculate the percentage of solids in milk, which would offer good evidence that milk is a mixture. If there is no residue after boiling, then you still cannot distinguish between the two possibilities.

What other experiments might you do to demonstrate that milk is a mixture? You could put a sample of milk in a centrifuge, which you might have used in a biology lab, spin your sample, and observe if any solids collect at the bottom of the centrifuge tube; large molecules can be separated in this way from a mixture. Measurement of the mass of the solid at the bottom of the tube is a way to obtain a value for the % solids in milk and also tells you that milk is indeed a mixture.

Keep an open mind: Lacking a centrifuge, how else might you separate solids in the milk? You could consider using a filter with really tiny holes in it or perhaps even a fine strainer. You could propose that if milk were poured through this filter, some (large) solid components should stay on the top of the filter, while water (and really small molecules or ions) would pass through the filter. That result would be evidence that milk is a mixture. Does such a filter exist? Yes! But for our purposes, the existence of such a filter is not the point: The point is, can you use your imagination and your knowledge of chemistry to design a reasonable experiment? Don't worry too much about the exact apparatus you need for the Design an Experiment exercises. The goal is to imagine what you would need to do, or what kind of data you would need to collect, in order to answer the question. If your instructor allows it, you can collaborate with others in your class to develop ideas. Scientists discuss their ideas with other scientists all the time. We find that discussing ideas, and refining them, makes us better scientists and helps us collectively answer important questions.

The design and interpretation of scientific experiments is at the heart of the scientific method. Think of the Design an Experiment exercises as puzzles that can be solved in various ways, and enjoy your explorations!

Self-Assessment Exercises

Here are a few problems designed to test your understanding of the material.

3.51 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 $O_2(g)$

3.52 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 reacted with 5.50 g of arsenic, how many grams of the excess reactant are left at the end of the reaction?

(a) 1.20 g As

(b) 1.50 g As

(c) 4.30 g As

(d) 8.30 g Ga

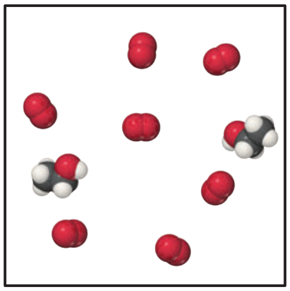
3.53 If 3.00 g of titanium metal is reacted with 6.00 g of chlorine gas, Cl_2, to form 7.7 g of titanium(IV) chloride in a combination reaction, what is the percent yield of the product?

(a) 65%
(b) 96%
(c) 48%
(d) 86%

Exercises

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

3.55 Consider the mixture of ethanol, C_2H_5OH, and O_2 shown in the accompanying diagram. (**a**) Write a balanced equation for the combustion reaction that occurs between ethanol and oxygen. (**b**) Which reactant is the limiting reactant? (**c**) How many molecules of CO_2, H_2O, C_2H_5OH, and O_2 will be present if the reaction goes to completion?



3.56 Sodium hydroxide reacts with carbon dioxide as follows:

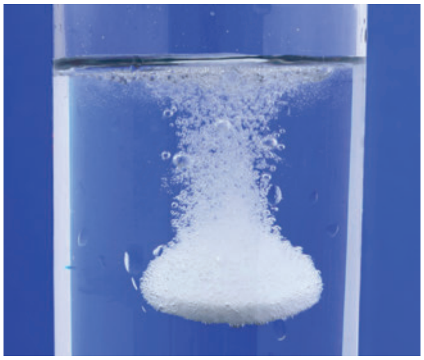
$$2\,NaOH(s) + CO_2(g) \longrightarrow Na_2CO_3(s) + H_2O(l)$$

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

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

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

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



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

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

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

(**a**) When 30.0 g of benzene reacts with 65.0 g of bromine, what is the theoretical yield of bromobenzene? (**b**) If the actual yield of bromobenzene is 42.3 g, what is the percentage yield?

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

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

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

Answers to Self-Assessment Exercises

3.51 (d) 3.52 (a) 3.53 (b)

Chapter Summary and Key Terms

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

SIMPLE PATTERNS OF CHEMICAL REACTIVITY (SECTION 3.2) Among the reaction types described in this chapter are (1) **combination reactions**, in which two reactants combine to form one product; (2) **decomposition reactions**, in which a single reactant forms two or more products; and (3) **combustion reactions** in oxygen, in which a substance, typically a hydrocarbon, reacts rapidly with O_2 to form CO_2 and H_2O.

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

AVOGADRO'S NUMBER AND THE MOLE (SECTION 3.4) A mole of any substance contains **Avogadro's number** (6.02×10^{23}) of formula units of that substance. The mass of a **mole** of atoms, molecules, or ions (the **molar mass**) equals the formula weight of that material expressed in grams. The mass of 1 molecule of H_2O, for example, is 18.0 u, so the mass of 1 mol of H_2O is 18.0 g. That is, the molar mass of H_2O is 18.0 g/mol.

EMPIRICAL FORMULAS FROM ANALYSIS (SECTION 3.5) The empirical formula of any substance can be determined from its percent composition by calculating the relative number of moles of each atom in 100 g of the substance. If the substance is molecular in nature, its molecular formula can be determined from the empirical formula if the molecular weight is also known. Combustion analysis is a special technique for determining the empirical formulas of compounds containing only carbon, hydrogen, and/or oxygen.

QUANTITATIVE INFORMATION FROM BALANCED EQUATIONS AND LIMITING REACTANTS (SECTIONS 3.6 AND 3.7) The mole concept can be used to calculate the relative quantities of reactants and products in chemical reactions. The coefficients in a balanced equation give the relative numbers of moles of the reactants and products. To calculate the number of grams of a product from the number of grams of a reactant, first convert grams of reactant to moles of reactant. Then use the coefficients in the balanced equation to convert the number of moles of reactant to moles of product. Finally, convert moles of product to grams of product.

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

Learning Outcomes After studying this chapter, you should be able to:

- Balance chemical equations. (Section 3.1) *Related Exercises: 3.5, 3.70*
- Predict the products of simple combination, decomposition, and combustion reactions. (Section 3.2) *Related Exercises: 3.12, 3.74*
- Calculate formula weights. (Section 3.3) *Related Exercises: 3.17, 3.76*
- Convert grams to moles and vice versa using molar masses. (Section 3.4) *Related Exercises: 3.25, 3.81*
- Convert number of molecules to moles and vice versa using Avogadro's number. (Section 3.4) *Related Exercises: 3.27, 3.83*
- Calculate the empirical and molecular formulas of a compound from percentage composition and molecular weight. (Section 3.5) *Related Exercises: 3.37, 3.90*
- Identify limiting reactants and calculate amounts, in grams or moles, of reactants consumed and products formed for a reaction. (Section 3.6) *Related Exercises: 3.47, 3.96*
- Calculate the percent yield of a reaction. (Section 3.7) *Related Exercises: 3.59, 3.105*

Key Equations

- $$\begin{matrix}\text{\% mass composition}\\ \text{of element}\end{matrix} = \frac{\begin{pmatrix}\text{number of atoms}\\ \text{of element}\end{pmatrix}\begin{pmatrix}\text{atomic weight}\\ \text{of element}\end{pmatrix}}{\text{formula weight of compound}} \times 100\% \quad [3.15]$$

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

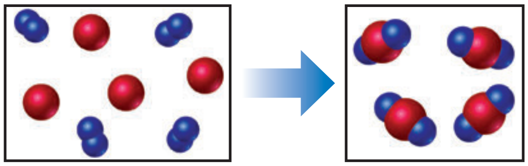
- $$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

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

Exercises

Visualizing Concepts

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



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

(**a**) $A_2 + B \longrightarrow A_2B$

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

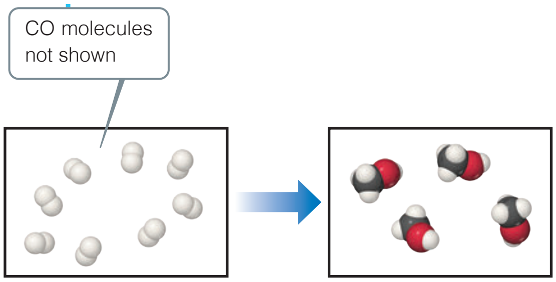
(**c**) $2\,A + B_4 \longrightarrow 2\,AB_2$

(**d**) $A + B_2 \longrightarrow AB_2$

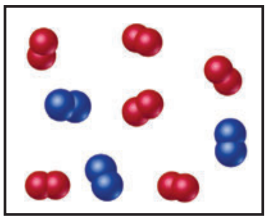
3.62 The following diagram shows the combination reaction between hydrogen, H_2, and carbon monoxide, CO, to produce methanol, CH_3OH (white spheres are H, black spheres are C, red spheres are O). The correct number of CO molecules involved in this reaction is not shown. [Section 3.1]

(**a**) Determine the number of CO molecules that should be shown in the left (reactants) box.

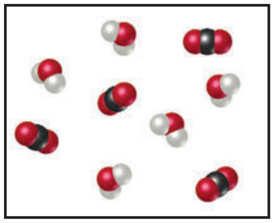
(**b**) Write a balanced chemical equation for the reaction.



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



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



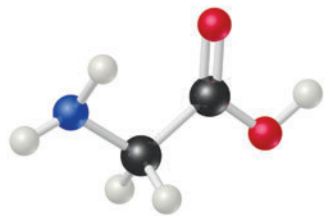
3.65 Glycine, an amino acid used by organisms to make proteins, is represented by the following molecular model.

(**a**) Write its molecular formula.

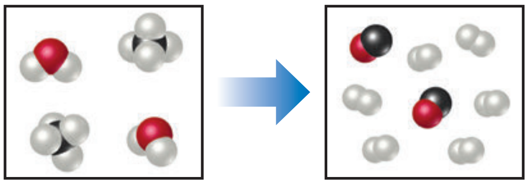
(**b**) Determine its molar mass.

(**c**) Calculate how many moles of glycine are in a 100.0-g sample of glycine.

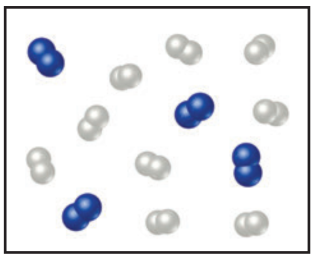
(**d**) Calculate the percent nitrogen by mass in glycine. [Sections 3.3 and 3.5]



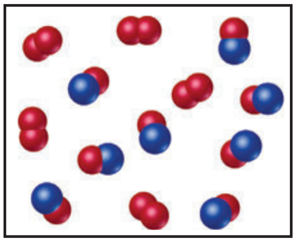
3.66 The following diagram represents a high-temperature reaction between CH_4 and H_2O. Based on this reaction, find how many moles of each product can be obtained starting with 4.0 mol CH_4. [Section 3.6]



3.67 Nitrogen (N_2) and hydrogen (H_2) react to form ammonia (NH_3). Consider the mixture of N_2 and H_2 shown in the accompanying diagram. The blue spheres represent N, and the white ones represent H. (**a**) Write the balanced chemical equation for the reaction. (**b**) What is the limiting reactant? (**c**) How many molecules of ammonia can be made, assuming the reaction goes to completion, based on the diagram? (**d**) Are any reactant molecules left over, based on the diagram? If so, how many of which type are left over? [Section 3.7]



3.68 Nitrogen monoxide and oxygen react to form nitrogen dioxide. Consider the mixture of NO and O_2 shown in the accompanying diagram. The blue spheres represent N, and the red ones represent O. (**a**) How many molecules of NO_2 can be formed, assuming the reaction goes to completion? (**b**) What is the limiting reactant? (**c**) If the actual yield of the reaction was 75% instead of 100%, how many molecules of each kind would be present after the reaction was over?



Chemical Equations and Simple Patterns of Chemical Reactivity (Section 3.1)

3.69 A key step in balancing chemical equations is correctly identifying the formulas of the reactants and products. For example, consider the reaction between calcium oxide, $CaO(s)$, and $H_2O(l)$ to form aqueous calcium hydroxide. (**a**) Write a balanced chemical equation for this combination reaction, having correctly identified the product as $Ca(OH)_2(aq)$. (**b**) Is it possible to balance the equation if you incorrectly identify the product as $CaOH(aq)$, and if so, what is the equation?

3.70 Balance the following equations:

(**a**) $HClO_4(aq) + P_4O_{10}(s) \longrightarrow HPO_3(aq) + Cl_2O_7(l)$

(**b**) $Au_2S_3(s) + H_2(g) \longrightarrow Au(s) + H_2S(g)$

(**c**) $Ba_3N_2(s) + H_2O(aq) \longrightarrow Ba(OH)_2(aq) + NH_3(g)$

(**d**) $Na_2CO_3(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

3.71 Balance the following equations:

(**a**) $CF_4(l) + Br_2(g) \longrightarrow CBr_4(l) + F_2(g)$

(**b**) $Cu(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$

(**c**) $MnO_2(s) + HCl(aq) \longrightarrow MnCl_2(s) + H_2O(l) + Cl_2(g)$

(**d**) $KOH(aq) + H_3PO_4(aq) \longrightarrow K_3PO_4(aq) + H_2O(l)$

3.72 Write balanced chemical equations to correspond to each of the following descriptions: (**a**) When sulfur trioxide gas reacts with water, a solution of sulfuric acid forms.(**b**) Boron sulfide, $B_2S_3(s)$, reacts violently with water to form dissolved boric acid, H_3BO_3, and hydrogen sulfide gas. (**c**) Phosphine, $PH_3(g)$, combusts in oxygen gas to form water vapor and solid tetraphosphorus decaoxide. (**d**) When solid mercury(II) nitrate is heated, it decomposes to form solid mercury(II) oxide, gaseous nitrogen dioxide, and oxygen. (**e**) Copper metal reacts with hot concentrated sulfuric acid solution to form aqueous copper(II) sulfate, sulfur dioxide gas, and water.

Patterns of Chemical Reactivity (Section 3.2)

3.73 (**a**) When a compound containing C, H, and O is completely combusted in air, what reactant besides the hydrocarbon is involved in the reaction? (**b**) What products form in this reaction? (**c**) What is the sum of the coefficients in the balanced chemical equation for the combustion of one mole of acetone, $C_3H_6O(l)$, in air?

3.74 Write a balanced chemical equation for the reaction that occurs when (**a**) titanium metal reacts with $O_2(g)$; (**b**) silver(I) oxide decomposes into silver metal and oxygen gas when heated; (**c**) propanol, $C_3H_7OH(l)$ burns in air; (**d**) methyl *tert*-butyl ether, $C_5H_{12}O(l)$, burns in air.

3.75 Balance the following equations and indicate whether they are combination, decomposition, or combustion reactions:

(**a**) $NaClO_3(s) \longrightarrow NaCl(s) + O_2(g)$

(**b**) $NH_4OH(aq) \longrightarrow H_2O(l) + NH_3(g)$

(**c**) $K(s) + Cl_2(g) \longrightarrow KCl(s)$

(**d**) $C_2H_5OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$

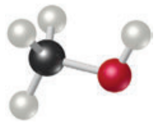
Formula Weights (Section 3.3)

3.76 Determine the formula weights of each of the following compounds: (**a**) Butyric acid, $CH_3CH_2CH_2COOH$, which is responsible for the rotten smell of spoiled food; (**b**) sodium perborate, $NaBO_3$, a substance used as bleach; (**c**) calcium carbonate, $CaCO_3$, a substance found in marble. (c) CF_2Cl_2, a refrigerant known as Freon; (**d**) $NaHCO_3$, known as baking soda and used in bread and pastry baking; (**e**) iron pyrite, FeS_2, which has a golden appearance and is known as "Fool's Gold."

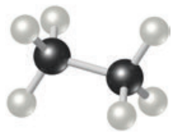
3.77 Calculate the percentage by mass of the indicated element in the following compounds: (**a**) hydrogen in methane, CH_4, the major hydrocarbon in natural gas; (**b**) oxygen in vitamin E, $C_{29}H_{50}O_2$; (**c**) sulphur in magnesium sulphate, $MgSO_4$, a substance used as a drying agent; (**d**) nitrogen in epinephrine, $C_9H_{13}NO_3$, also known as adrenalin, a hormone that is important for the fight-or-flight response; (**e**) oxygen in the insect pheromone sulcatol, $C_8H_{16}O$; (**f**) carbon in sucrose, $C_{12}H_{22}O_{11}$, the compound that is responsible for the sweet taste of table sugar.

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

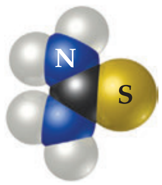
(**a**)



(**b**)



(**c**)



(**d**)

Avogadro's Number and the Mole (Section 3.4)

3.79 (**a**) What is the mass, in grams, of one mole of ^{79}Br? (b) How many bromine atoms are present in one mole of ^{79}Br?

3.80 Without doing any detailed calculations (but using a periodic table to give atomic weights), rank the following samples in order of increasing numbers of atoms: 0.2 mol PCl_5 molecules, 80 g Fe_2O_3, 3.0×10^{23} CO molecules.

3.81 Calculate the following quantities:

(**a**) mass, in grams, of 1.50×10^{-2} mol CdS

(**b**) number of moles of NH_4Cl in 86.6 g of this substance

(**c**) number of molecules in 8.447×10^{-2} mol C_6H_6

(**d**) number of O atoms in 6.25×10^{-3} mol $Al(NO_3)_3$

3.82 (**a**) What is the mass, in grams, of 1.223 mol of iron(III) sulfate?

(**b**) How many moles of ammonium ions are in 6.955 g of ammonium carbonate?

(**c**) What is the mass, in grams, of 1.50×10^{21} molecules of aspirin, $C_9H_8O_4$?

(**d**) What is the molar mass of diazepam (Valium®) if 0.05570 mol has a mass of 15.86 g?

3.83 The molecular formula of salicylic acid, a compound commonly found in facial cleanser, is $C_7H_6O_3$. (**a**) What is the molar mass of salicylic acid? (**b**) How many moles of salicylic acid are present in 0.5 mg of this substance? (**c**) How many molecules of salicylic acid are in 0.5 mg of this substance? (**d**) How many oxygen atoms are present in 0.5 mg of salicylic acid?

3.84 A sample of the male sex hormone testosterone, $C_{19}H_{28}O_2$, contains 3.88×10^{21} hydrogen atoms. (**a**) How many atoms of carbon does it contain? (**b**) How many molecules of testosterone does it contain? (**c**) How many moles of testosterone does it contain? (**d**) What is the mass of this sample in grams?

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

Empirical Formulas from Analyses (Section 3.5)

3.86 Determine the empirical formula of each of the following compounds if a sample contains (**a**) 3.92 mol C, 5.99 mol H, and 2.94 mol O; (**b**) 12.0 g calcium and 2.8 g nitrogen; (**c**) 89.14% Au and 10.86% O by mass.

3.87 Determine the empirical formulas of the compounds with the following compositions by mass:

(**a**) 42.1% Na, 18.9% P, and 39.0% O

(**b**) 18.7% Li, 16.3% C, and 65.0% O

(**c**) 60.0% C, 4.4% H, and the remainder O

3.88 The compound XCl_4 contains 75.0% Cl by mass. What is the element X?

3.89 What is the molecular formula of each of the following compounds?

(**a**) empirical formula CH_3O, molar mass = 62.0 g/mol

(**b**) empirical formula NH_2, molar mass = 32.0 g/mol

3.90 Determine the empirical and molecular formulas of each of the following substances:

(**a**) Ibuprofen, a headache remedy, contains 75.69% C, 8.80% H, and 15.51% O by mass and has a molar mass of 206 g/mol.

(**b**) Cadaverine, a foul-smelling substance produced by the action of bacteria on meat, contains 58.55% C, 13.81% H, and 27.40% N by mass; its molar mass is 102.2 g/mol.

(**c**) Epinephrine (adrenaline), a hormone secreted into the bloodstream in times of danger or stress, contains 59.0% C, 7.1% H, 26.2% O, and 7.7% N by mass; its molar mass is about 180 u.

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

3.92 Propenoic acid, $C_3H_4O_2$, is a reactive organic liquid that is used in the manufacturing of plastics, coatings, and adhesives. An unlabeled container is thought to contain this liquid. A 0.275-g sample of the liquid is combusted to produce 0.102 g of water and 0.374 g carbon dioxide. Is the unknown liquid propenoic acid? Support your reasoning with calculations.

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

Reaction Stoichiometry (Section 3.6)

3.94 The reaction between potassium superoxide, KO_2, and CO_2,

$$4\,KO_2 + 2\,CO_2 \longrightarrow 2K_2CO_3 + 3\,O_2$$

is used as a source of O_2 and absorber of CO_2 in self-contained breathing equipment used by rescue workers.

(**a**) How many moles of O_2 are produced when 0.400 mol of KO_2 reacts in this fashion?

(**b**) How many grams of KO_2 are needed to form 7.50 g of O_2?

(**c**) How many grams of CO_2 are used when 7.50 g of O_2 are produced?

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

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

(**a**) Balance this equation.

(**b**) Calculate the number of grams of CO that can react with 0.350 kg of Fe_2O_3.

(**c**) Calculate the number of grams of Fe and the number of grams of CO_2 formed when 0.350 kg of Fe_2O_3 reacts.

(**d**) Show that your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

3.96 Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas. (**a**) Write a balanced chemical equation for the reaction. (**b**) How many grams of calcium hydride are needed to form 4.500 g of hydrogen?

3.97 The complete combustion of octane, C_8H_{18}, a component of gasoline, proceeds as follows:

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

(**a**) How many moles of O_2 are needed to burn 1.50 mol of C_8H_{18}?

(**b**) How many grams of O_2 are needed to burn 10.0 g of C_8H_{18}?

(**c**) Octane has a density of 0.692 g/mL at 20 °C. How many grams of O_2 are required to burn 15.0 gal of C_8H_{18} (the capacity of an average fuel tank)?

(**d**) How many grams of CO_2 are produced when 15.0 gal of C_8H_{18} are combusted?

3.98 Detonation of nitroglycerin proceeds as follows:

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

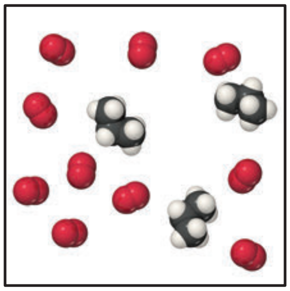
(**a**) If a sample containing 2.00 mL of nitroglycerin (density = 1.592 g/mL) is detonated, how many moles of gas are produced? (**b**) If each mole of gas occupies 55 L under the conditions of the explosion, how many liters of gas are produced? (**c**) How many grams of N_2 are produced in the detonation?

3.99 The combustion of one mole of liquid octane, $CH_3(CH_2)_6CH_3$, produces 5470 kJ of heat. Calculate how much heat is produced if 1.000 gallon of octane is combusted. See Exercise 3.97 for necessary information about octane.

Limiting Reactants (Section 3.7)

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

3.101 Consider the mixture of propane, C_3H_8, and O_2 shown here. (**a**) Write a balanced equation for the combustion reaction that occurs between propane and oxygen. (**b**) Which reactant is the limiting reactant? (**c**) How many molecules of CO_2, H_2O, C_3H_8, and O_2 will be present if the reaction goes to completion?



3.102 Aluminum hydroxide reacts with sulfuric acid as follows:

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

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

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

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

In a certain experiment, 2.00 g of NH_3 reacts with 2.50 g of O_2. (**a**) Which is the limiting reactant? (**b**) How many grams of NO and H_2O form? (**c**) How many grams of the excess reactant remain after the limiting reactant is completely consumed? (**d**) Show that your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

3.104 Solutions of sulfuric acid and lead(II) acetate react to form solid lead(II) sulfate and a solution of acetic acid. If 5.00 g of sulfuric acid and 5.00 g of lead(II) acetate are mixed, calculate the number of grams of sulfuric acid, lead(II) acetate, lead(II) sulfate, and acetic acid present in the mixture after the reaction is complete.

3.105 When ethane (C_2H_6) reacts with chlorine (Cl_2), the main product is C_2H_5Cl, but other products containing Cl, such as $C_2H_4Cl_2$, are also obtained in small quantities. The formation of these other products reduces the yield of C_2H_5Cl. (**a**) Calculate the theoretical yield of C_2H_5Cl when 125 g of C_2H_6 reacts with 255 g of Cl_2, assuming that C_2H_6 and Cl_2 react only to form C_2H_2Cl and HCl. (**b**) Calculate the percent yield of C_2H_5Cl if the reaction produces 206 g of C_2H_5Cl.

3.106 When hydrogen sulfide gas is bubbled into a solution of sodium hydroxide, the reaction forms sodium sulfide and water. How many grams of sodium sulfide are formed if 1.25 g of hydrogen sulfide is bubbled into a solution containing 2.00 g of sodium hydroxide, assuming that the sodium sulfide is made in 92.0% yield?

Additional Exercises

3.107 Write balanced chemical equations for (**a**) the complete combustion of acetone (CH_3COCH_3), a common organic solvent; (**b**) the decomposition of solid mercury (I) carbonate into carbon dioxide gas, mercury, and solid mercury oxide; (**c**) the combination reaction between sulphur dioxide gas and liquid water to produce sulfurous acid.

3.108 If 2.0 mol $CH_3CH_2CH_2COOH$, 2.0 mol C_4H_{10}, and 2.0 mol C_6H_6 are completely combusted in oxygen, which one produces the largest number of moles of H_2O? Which one produces the least? Explain.

3.109 Calcium is an essential nutrient in our body. It is important for bone health. Four common calcium-containing supplements are calcium carbonate ($CaCO_3$), calcium citrate ($Ca_3C_{12}H_{10}O_{14}$), calcium gluconate ($CaC_{12}H_{22}O_{14}$), and calcium lactate ($CaC_6H_{10}O_6$). Rank these calcium supplements in terms of the mass percentage of calcium they contain.

3.110 (**a**) Ibuprofen is a common over-the-counter analgesic with the formula $C_{13}H_{18}O_2$. How many moles of $C_{13}H_{18}O_2$ are in a 500-mg tablet of ibuprofen? Assume the tablet is composed entirely of ibuprofen. (**b**) How many molecules of $C_{13}H_{18}O_2$ are in this tablet? (**c**) How many oxygen atoms are in the tablet?

3.111 Very small semiconductor crystals, composed of approximately 1000 to 10,000 atoms, are called quantum dots. Quantum dots made of the semiconductor CdSe are now being used in electronic reader and tablet displays because they emit light efficiently and in multiple colors, depending on dot size. The density of CdSe is 5.82 g/cm^3.

(**a**) What is the mass of one 2.5-nm CdSe quantum dot?

(**b**) CdSe quantum dots that are 2.5 nm in diameter emit blue light upon stimulation. Assuming that the dot is a perfect sphere and that the empty space in the dot can be neglected, calculate how many Cd atoms are in one quantum dot of this size.

(**c**) What is the mass of one 6.5-nm CdSe quantum dot?

(**d**) CdSe quantum dots that are 6.5 nm in diameter emit red light upon stimulation. Assuming that the dot is a perfect sphere, calculate how many Cd atoms are in one quantum dot of this size.

(**e**) If you wanted to make one 6.5-nm dot from multiple 2.5-nm dots, how many 2.5-nm dots would you need, and how many CdSe formula units would be left over, if any?

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

3.113 Cinnamaldehyde is a compound that is responsible for the characteristic aroma of cinnamon. It contains 81.79% C, 6.10% H, and the remaining is oxygen. Its molar mass is 132 g/mol. Determine its molecular formula.

3.114 Fructose, commonly called fruit sugar, is a monosaccharide found in many plants. It contains 40% C, 6.71% H, and the remainder O. (**a**) What is the empirical formula for fructose? (**b**) A mass spectrum of fructose shows a peak at about 180 u. What is the molecular formula of the substance?

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

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

3.117 A compound, $Na_2Cr_2O_x$, where x is unknown, is analyzed and found to contain 39.70% Cr. What is the value of x?

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

Integrative Exercises

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

3.126 Boron nitride, BN, is an electrical insulator with remarkable thermal and chemical stability. Its density is 2.1 g/cm^3. It can be made by reacting boric acid, H_3BO_3, with ammonia. The other product of the reaction is water. (**a**) Write a balanced chemical equation for the synthesis of BN. (**b**) If you made 225 g of boric acid react with 150 g ammonia, what mass of BN could you make? (**c**) Which reactant, if any, would be left over, and how many moles of leftover reactant would remain? (d) One application of BN is as thin film for electrical insulation. If you take the mass of BN from part (a) and make a 0.4 mm thin film from it, what area, in cm^2, would it cover?

3.127 Viridicatumtoxin B, $C_{30}H_{31}NO_{10}$, is a natural antibiotic compound. It requires a synthesis of 12 steps in the laboratory. Assuming all steps have equivalent yields of 85%, which is the final percent yield of the total synthesis?

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

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

3.130 (**a**) If an automobile travels 350 km with a gas mileage of 9.0 km/L, how many kilograms of CO_2 are produced? Assume that the gasoline is composed of octane, $C_8H_{18}(l)$, whose density is 0.692 g/mL. (**b**) Repeat the calculation for a truck that has a gas mileage of 2 km/L.

3.131 Section 2.9 introduced the idea of structural isomerism, with 1-propanol and 2-propanol as examples. Determine which of these properties would distinguish these two substances: (**a**) boiling point, (**b**) combustion analysis results, (**c**) molecular weight, (**d**) density at a given temperature and pressure. You can check on the properties of these two compounds in

Wolfram Alpha (http://www.wolframalpha.com/) or the *CRC Handbook of Chemistry and Physics.*

3.132 NO_x is a generic term for the nitrogen oxides, NO and NO_2. NO_x gases are air pollutants that react to form smog and acid rain. In order to reduce NO_x emission from vehicle, catalytic converters are installed in car exhausts to decompose NO and NO_2 respectively into N_2 and O_2 (**a**) Write the balanced chemical equations for the decomposition of NO and NO_2 respectively. (**b**) If the car produces 100 g NO_x a day, with equal mole ratio of NO and NO_2, how many grams of NO and NO_2 are produced respectively?

3.133 Hydrogen cyanide, HCN, is a poisonous gas. The lethal dose is approximately 300 mg HCN per kilogram of air when inhaled. (**a**) Calculate the amount of HCN that gives the lethal dose in a small laboratory room measuring $3.5 \times 4.5 \times 2.5$ m. The density of air at 26 °C is 0.00118 g/cm^3. (**b**) If the HCN is formed by reaction of NaCN with an acid such as H_2SO_4, what mass of NaCN gives the lethal dose in the room?

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

(**c**) HCN forms when synthetic fibers containing Orlon® or Acrilan® burn. Acrilan® has an empirical formula of CH_2CHCN, so HCN is 50.9% of the formula by mass. A rug measures 3.5×4.5 m and contains 850 g of Acrilan® fibers per square yard of carpet. If the rug burns, will a lethal dose of HCN be generated in the room? Assume that the yield of HCN from the fibers is 20% and that the carpet is 50% consumed.

3.134 The source of oxygen that drives the internal combustion engine in an automobile is air. Air is a mixture of gases, principally N_2(~79%) and O_2(~20%). In the cylinder of an automobile engine, nitrogen can react with oxygen to produce nitric oxide gas, NO. As NO is emitted from the tailpipe of the car, it can react with more oxygen to produce nitrogen dioxide gas. (**a**) Write balanced chemical equations for both reactions. (**b**) Both nitric oxide and nitrogen dioxide are pollutants that can lead to acid rain and global warming; collectively, they are called "NO_x" gases. In 2009, the United States emitted an estimated 19 million tons of nitrogen dioxide into the atmosphere. How many grams of nitrogen dioxide is this? (**c**) The production of NO_x gases is an unwanted side reaction of the main engine combustion process that turns octane, C_8H_{18}, into CO_2 and water. If 85% of the oxygen in an engine is used to combust octane and the remainder used to produce nitrogen dioxide, calculate how many grams of nitrogen dioxide would be produced during the combustion of 500 g of octane.

3.135 The thermite reaction,

$$Fe_2O_3 + Al \rightarrow Al_2O_3 + Fe$$

produces so much heat that the Fe product melts. This reaction is used industrially to weld metal parts under water, where a torch cannot be employed. It is also a favorite chemical demonstration in the lecture hall (on a small scale).

(**a**) Balance the chemical equation for the thermite reaction, and include the proper states of matter.

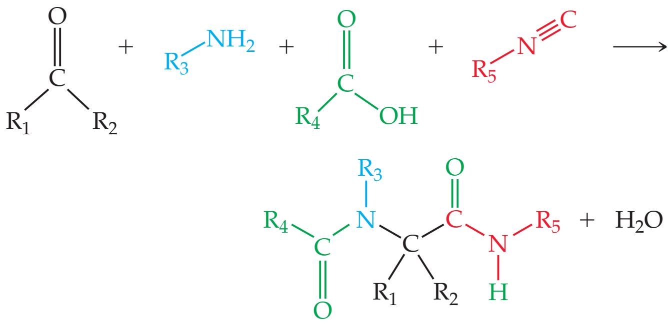
(**b**) Calculate how many grams of aluminum are needed to completely react with 500.0 g of Fe_2O_3 in this reaction.

(**c**) This reaction produces 852 kJ of heat per mole of Fe_2O_3 reacted. How many grams of Fe_2O_3 are needed to produce 1.00×10^4 kJ of heat?

(**d**) If you performed the reverse reaction— aluminum oxide plus iron makes iron oxide plus aluminum—would that reaction have heat as a reactant or a product?

3.136 One of the most bizarre reactions in chemistry is called the Ugi reaction:

$R_1C(=O)R_2 + R_3 - NH_2 + R_4COOH + R_5NC \rightarrow$
$R_4C(=O)N(R_3)C(R_1R_2)C=ONHR_5 + H_2O$



(**a**) Write out the balanced chemical equation for the Ugi reaction, for the case where R = $CH_3CH_2CH_2CH_2CH_2CH_2$— (this is called the *hexyl* group) for all compounds. (**b**) What mass of the "hexyl Ugi product" would you form if 435.0 mg of $CH_3\,CH_2CH_2CH_2CH_2CH_2NH_2$ was the limiting reactant?

Design an Experiment

You will learn later in this book that sulfur is capable of forming two common oxides, SO_2 and SO_3. One question that we might ask is whether the direct reaction between sulfur and oxygen leads to the formation of SO_2, SO_3, or a mixture of the two. This question has practical significance because SO_3 can go on to react with water to form sulfuric acid, H_2SO_4, which is produced industrially on a very large scale. Consider also that the answer to this question may depend on the relative amount of each element that is present and the temperature at which the reaction is carried out. For example, on the one hand, carbon and oxygen normally react to form CO_2, but when not enough oxygen is present, CO can form. On the other hand, under normal reaction conditions H_2 and O_2 react to form water, H_2O (rather than hydrogen peroxide H_2O_2), regardless of the starting ratio of hydrogen to oxygen.

Suppose you are given a bottle of sulfur, which is a yellow solid, a cylinder of O_2, a transparent reaction vessel that can be evacuated and sealed so that only sulfur, oxygen, and the product(s) of the reaction between the two are present, an analytical balance so that you can determine the masses of the reactants and/or products, and a furnace that can be used to heat the reaction vessel to 200 °C where the two elements react. (**a**) If you start with 0.10 mol of sulfur in the reaction vessel how many moles of oxygen would need to be added to form SO_2, assuming SO_2 forms exclusively? (**b**) How many moles of oxygen would be needed to form SO_3, assuming SO_3 forms exclusively? (**c**) Given the available equipment, how would you determine you had added the correct number of moles of each reactant to the reaction vessel? (**d**) What observation or experimental technique would you use to determine the identity of the reaction product(s)? Could differences in the physical properties of SO_2 and SO_3 be used to help identify the product(s)? Have any instruments been described in Chapters 1–3 that would allow you to identify the product(s)? (**e**) What experiments would you conduct to determine if the product(s) of this reaction (either SO_2 or SO_3 or a mixture of the two) can be controlled by varying the ratio of sulfur and oxygen that are added to the reaction vessel? What ratio(s) of S to O_2 would you test to answer this question?