Chemical Reactions and Stoichiometry

4

> *When we decode a cookbook, every one of us is a practicing chemist. Cooking is really the oldest, most basic application of physical and chemical forces to natural materials.*
>
> —Arthur E. Grosser

THE AMOUNT OF PRODUCT FORMED IN A CHEMICAL REACTION is related to the amount of reactant that is consumed. This concept makes sense intuitively, but how do we describe and understand this relationship more fully? The second half of this chapter focuses on chemical stoichiometry—the numerical relationships between the amounts of reactants and products in chemical reactions. First we will learn how to write balanced chemical equations for chemical reactions. We will also describe some general types of chemical reactions. You have probably witnessed many of these types of reactions in your daily life because they are so common. Have you ever mixed baking soda with vinegar and observed the subsequent bubbling? Or have you ever noticed the hard water deposits that form on plumbing fixtures? These reactions are the subject of the first part of this chapter.

4.1 Chemistry of Cuisine

It is pretty well-known that good chemists make good cooks. Perhaps it is the techniques that are common between cooking and chemistry. Techniques such as measuring, filtering, concentrating, and distilling that chemists should be very good at may give them an edge when it comes to cooking. Maybe it is the experimental and

creative nature by which cooking is done to produce something new and tasty that is an allure for chemists to cooking. No matter the reason, one thing is for sure: cooking *is* chemistry. When we cook, from the baking of bread to the browning of meat, a chemical reaction is occurring—we are changing the chemicals in the raw ingredients to other chemical products in the final, cooked dish.

Let's consider a few practices in preparing food that are examples of chemistry occurring before your eyes. When you heat a steak on a barbeque, the meat changes from a bright red colour to the brown colour that we recognize as a steak on our plate. What is occurring is that the proteins that compose the steak are being denatured. In this particular example, the red myoglobin—the oxygen-carrying protein—is converted to metmyoglobin, which is brown. But there are many other chemical reactions that occur during the cooking of a steak as well. When you heat food that contains both protein and carbohydrates (sugars), a complex set of reactions occur—the Maillard reaction—to form many flavourful and aromatic compounds. In fact, the Maillard reaction is extremely important when cooking and is responsible for many other flavours experienced in cuisine, such as baked and toasted bread, malted barley in whisky or beer, fried onions, and countless others.

You may have been told by a parent or grandparent to walk softly when a cake is in the oven. Have you ever wondered why? To answer this, we must know a little about what chemical reaction is causing the cake to rise. Baking powder consists of sodium bicarbonate and an agent to help acidify the mixture. The sodium bicarbonate and the acid react to form carbon dioxide. In the hot oven, this reaction occurs more quickly, and the carbon dioxide gets trapped in the cake, leavening the cake, or making it expand. This reaction is much the same as when baking soda (sodium bicarbonate) reacts with vinegar (an acid). In the early stages of baking, when a crust is forming over the top of the cake, jumping can agitate the cake, break the crust, and allow the carbon dioxide gas to escape, resulting in a "fallen cake."

A final example of the chemistry of cuisine is preparing or serving fish with lemon slices. When fish begins to decay, low molecular weight amines are produced. These compounds are volatile and produce the characteristic fishy smell. The acid from a lemon reacts with or neutralizes the amine, producing a nonvolatile and nonsmelly compound, leaving the fish smelling fresh. However, probably the best way to have fish is right out of the water, before the amines are produced; fresh fish doesn't smell.

▲ Fresh fish.

4.2 Writing and Balancing Chemical Equations

Combustion analysis (which we saw in the previous chapter) employs a **chemical reaction**, a process in which one or more substances are converted into one or more different ones. Compounds form and change through chemical reactions. Water can be made by the reaction of hydrogen with oxygen. A **combustion reaction** is a particular type of chemical reaction in which a substance combines with oxygen to form one or more oxygen-containing compounds. Combustion reactions also emit heat. The heat produced in a number of combustion reactions is critical to supplying our society's energy needs. For example, the heat from the combustion of gasoline causes the gaseous combustion products to expand in the cylinders of a car's engine, which push the pistons and propel the car. We use the heat released by the combustion of *natural gas* to cook food and to heat our homes. In many places in the world, fossil fuels, such as coal and oil, are burned to provide electricity.

A chemical reaction is represented by a **chemical equation**. The combustion of natural gas can be represented by the following, unbalanced, equation:

$$\underbrace{CH_4 + O_2}_{\text{reactants}} \longrightarrow \underbrace{CO_2 + H_2O}_{\text{products}}$$

The substances on the left side of the equation are called the **reactants** and the substances on the right side are called the **products**. We often specify the states of each reactant or product in parentheses next to the formula as follows:

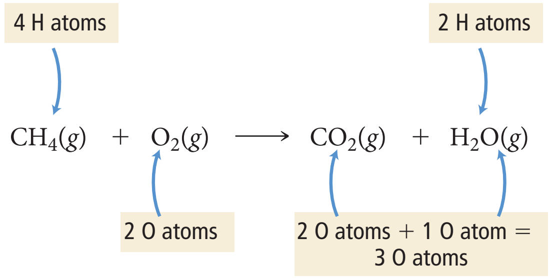
$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

The (g) indicates that these substances are gases in the reaction. The common states of reactants and products and their symbols used in chemical equations are summarized in Table 4.1.

TABLE 4.1 States of Reactants and Products in Chemical Equations

Abbreviation	State
(*g*)	Gas
(*l*)	Liquid
(*s*)	Solid
(*aq*)	Aqueous (water solution)

The equation just presented for the combustion of natural gas is not complete, however. If we look closely, we can see that the left side of the equation has two oxygen atoms and four hydrogen atoms while the right side has three oxygen atoms and two hydrogen atoms:

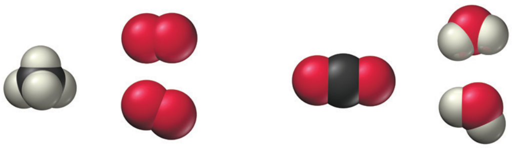


Therefore, as written, the reaction violates the law of conservation of mass because an oxygen atom formed out of nothing and hydrogen atoms have vanished. To correct these problems—that is, to write an equation that correctly represents *what actually happens*—we must balance the equation. We need to change the coefficients (those numbers *in front of* the chemical formulas), not the subscripts (those numbers within the chemical formulas), to ensure that the number of each type of atom on the left side of the equation is equal to the number on the right side. New atoms do not form during a reaction, nor do atoms vanish—matter must be conserved.

The reason that you cannot change the subscripts when balancing a chemical equation is that changing the subscripts changes the substance itself, while changing the coefficients changes the number of molecules of the substance. For example, 2 H_2O is simply two water molecules, but H_2O_2 is hydrogen peroxide, a drastically different compound.

When we add coefficients to the reactants and products to balance an equation, we change the number of molecules in the equation but not the *kind of* molecules. To balance the equation for the combustion of methane, we put the coefficient 2 before O_2 in the reactants and the coefficient 2 before H_2O in the products:

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$



The equation is now balanced because the numbers of each type of atom on either side of the equation are equal. The **balanced chemical equation** tells us that one CH_4 molecule reacts with two O_2 molecules to form one CO_2 molecule and two H_2O molecules. We verify that the equation is balanced by summing the number of each type of atom on each side of the equation:

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Reactants	Products
1 C atom (1 × $\underline{C}H_4$)	1 C atom (1 × $\underline{C}O_2$)
4 H atoms (1 × $C\underline{H_4}$)	4 H atoms (2 × $\underline{H_2}O$)
4 O atoms (2 × $\underline{O_2}$)	4 O atoms (1 × $C\underline{O_2}$ + 2 × $H_2\underline{O}$)

The numbers of each type of atom on both sides of the equation are now equal—the equation is balanced.

How to Write Balanced Chemical Equations

We can balance many chemical equations simply by trial and error. However, some guidelines are useful. For example, balancing the atoms in the most complex substances first and the atoms in the simplest substances (such as pure elements) last often makes the process shorter. The following examples illustrate how to balance chemical equations. The general guidelines are shown on the left, with two examples of how to apply them on the right. This procedure is meant only as a flexible guide, not a rigid set of steps.

PROCEDURE FOR … **Balancing Chemical Equations**	**EXAMPLE 4.1** **Balancing Chemical Equations** Write a balanced equation for the reaction between solid cobalt(III) oxide and solid carbon to produce solid cobalt and carbon dioxide gas.	**EXAMPLE 4.2** **Balancing Chemical Equations** Write a balanced equation for the combustion of gaseous butane (C_4H_{10}), a fuel used in portable stoves and grills, in which it combines with gaseous oxygen to form gaseous carbon dioxide and gaseous water.
1. Write a skeletal equation by writing chemical formulas for each of the reactants and products. (If a skeletal equation is provided, go to step 2.)	$Co_2O_3(s) + C(s) \longrightarrow Co(s) + CO_2(g)$	$C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
2. Balance atoms that occur in more complex substances first. Always balance atoms in compounds before atoms in pure elements.	**Begin with O:** $Co_2O_3(s) + C(s) \longrightarrow Co(s) + CO_2(g)$ 3 O atoms ⟶ 2 O atoms To balance O, put a 2 before $Co_2O_3(s)$ and a 3 before $CO_2(g)$: $\mathbf{2}\,Co_2O_3(s) + C(s) \longrightarrow Co(s) + \mathbf{3}\,CO_2(g)$ 6 O atoms ⟶ 6 O atoms	**Begin with C:** $C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$ 4 C atoms ⟶ 1 C atom To balance C, put a 4 before $CO_2(g)$: $C_4H_{10}(g) + O_2(g) \longrightarrow \mathbf{4}\,CO_2(g) + H_2O(g)$ 4 C atoms ⟶ 4 C atoms **Balance H:** $C_4H_{10}(g) + O_2(g) \longrightarrow 4\,CO_2(g) + H_2O(g)$ 10 H atoms ⟶ 2 H atoms To balance H, put a 5 before $H_2O(g)$: $C_4H_{10}(g) + O_2(g) \longrightarrow 4\,CO_2(g) + \mathbf{5}\,H_2O(g)$ 10 H atoms ⟶ 10 H atoms
3. Balance atoms that occur as free elements on either side of the equation last. Always balance free elements by adjusting the coefficient on the free element. If the balanced equation contains a fractional coefficient, one may choose to clear these by multiplying the entire equation by the denominator of the fraction.	**Balance Co:** $2\,Co_2O_3(s) + C(s) \longrightarrow Co(s) + 3\,CO_2(g)$ 4 Co atoms ⟶ 1 Co atom To balance Co, put a 4 before $Co(s)$: $2\,Co_2O_3(s) + C(s) \longrightarrow \mathbf{4}\,Co(s) + 3\,CO_2(g)$ 4 Co atoms ⟶ 4 Co atoms **Balance C:** $2\,Co_2O_3(s) + C(s) \longrightarrow 4\,Co(s) + 3\,CO_2(g)$ 1 C atom ⟶ 3 C atoms To balance C, put a 3 before $C(s)$: $2\,Co_2O_3(s) + \mathbf{3}\,C(s) \longrightarrow 4\,Co(s) + 3\,CO_2(g)$	**Balance O:** $C_4H_{10}(g) + O_2(g) \longrightarrow 4\,CO_2(g) + 5\,H_2O(g)$ 2 O atoms ⟶ 8 O + 5 O = 13 O atoms To balance O, put a $\frac{13}{2}$ before $O_2(g)$: $C_4H_{10}(g) + \mathbf{\frac{13}{2}}\,O_2(g) \longrightarrow 4\,CO_2(g) + 5\,H_2O(g)$ 13 O atoms ⟶ 13 O atoms $C_4H_{10}(g) + \frac{13}{2}\,O_2(g) \longrightarrow 4\,CO_2(g) + 5\,H_2O(g)$ or $2\,C_4H_{10}(g) + 13\,O_2(g) \longrightarrow 8\,CO_2(g) + 10\,H_2O(g)$

4. Check to make certain the equation is balanced by summing the total number of each type of atom on both sides of the equation.

$$2\ Co_2O_3(s) + 3\ C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$$

Left	Right
4 Co atoms	4 Co atoms
6 O atoms	6 O atoms
3 C atoms	3 C atoms

The equation is balanced.

$$C_4H_{10}(g) + \tfrac{13}{2}\ O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$$

Left	Right
4 C atoms	4 C atoms
10 H atoms	10 H atoms
13 O atoms	13 O atoms

The equation is balanced.

FOR PRACTICE 4.1

Write a balanced equation for the reaction between solid silicon dioxide and solid carbon to produce solid silicon carbide and carbon monoxide gas.

FOR PRACTICE 4.2

Write a balanced equation for the combustion of gaseous ethane (C_2H_6), a minority component of natural gas, in which it combines with gaseous oxygen to form gaseous carbon dioxide and gaseous water.

CONCEPTUAL CONNECTION 4.1

Balanced Chemical Equations

Which quantities must always be the same on both sides of a chemical equation?

(a) the number of atoms of each kind

(b) the number of molecules of each kind

(c) the number of moles of each kind of molecule

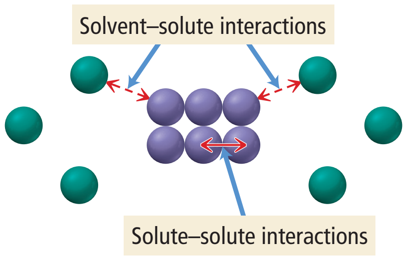
(d) the sum of the masses of all substances involved

4.3 Solutions and Solubility

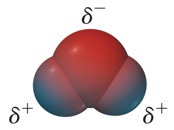
Chemical reactions involving reactants dissolved in water are among the most common and important. The reactions that occur in lakes, streams, and oceans, as well as the reactions that occur in every cell within our bodies, take place in water. A homogeneous mixture of two substances—such as salt and water—is a **solution**. The majority component of the mixture is the **solvent**, and the minority component is the **solute**. An **aqueous solution** is one in which water acts as the solvent.

Consider two familiar aqueous solutions: salt water and sugar water. Salt water is a solution of NaCl and H_2O, and sugar water is a solution of $C_{12}H_{22}O_{11}$ and H_2O. Certainly, you have made these solutions by adding table salt or sugar to water. As you stir either of these two substances into the water, it seems to disappear. However, you know that the original substance is still present because the water has a salty or a sweet taste. How do solids such as salt and sugar dissolve in water?

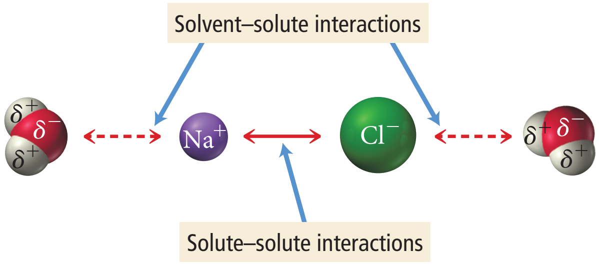
When a solid is put into a liquid solvent, the attractive forces that hold the solid together (the solute–solute interactions) come into competition with the attractive forces between the solvent molecules and the particles that compose the solid (the solvent–solute interactions), as shown in Figure 4.1 ▶. For example, when sodium chloride is put into water, there is a competition between the attraction of Na^+ cations and Cl^- anions to each other (due to their opposite charges) and the attraction of Na^+ and Cl^- to water molecules. The attraction of Na^+ and Cl^- to water is based on the *polar nature* of the water molecule. For reasons we discuss later in this text, the oxygen atom in water is electron-rich, giving it a partial negative charge (δ^-), as shown in Figure 4.2 ▶. The hydrogen atoms, in contrast, are electron-poor, giving them a partial positive charge ($\delta+$). As a result, the positively charged sodium ions are



▲ **FIGURE 4.1 Solute and Solvent Interactions** When a solid is put into a solvent, the interactions between solvent and solute particles compete with the interactions among the solute particles themselves.



▲ **FIGURE 4.2 Charge Distribution in a Water Molecule** An uneven distribution of electrons within the water molecule causes the oxygen side of the molecule to have a partial negative charge and the hydrogen side to have a partial positive charge.



▲ **FIGURE 4.3 Solute and Solvent Interactions in a Sodium Chloride Solution** When sodium chloride is put into water, the attraction of Na^+ and Cl^- ions to water molecules competes with the attraction among the oppositely charged ions themselves.

strongly attracted to the oxygen side of the water molecule, and the negatively charged chloride ions are attracted to the hydrogen side of the water molecule, as shown in Figure 4.3 ◄. In the case of NaCl, the attraction between the separated ions and many water molecules overcomes the attraction of sodium and chloride ions to each other, and the sodium chloride dissolves in the water (Figure 4.4 ▼).

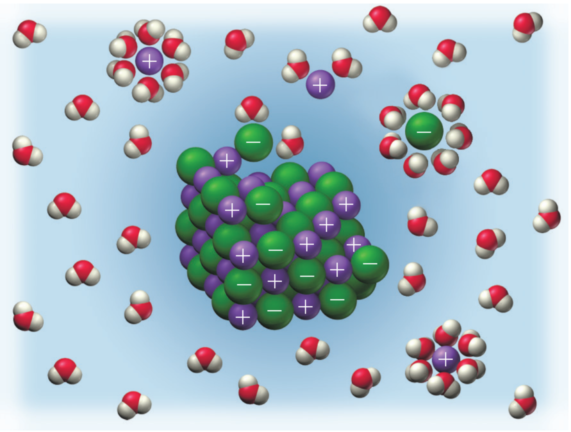
Electrolyte and Nonelectrolyte Solutions

A salt solution will conduct electricity, while a sugar solution does not. The difference between the way that salt (an ionic compound) and sugar (a molecular compound) dissolve in water illustrates a fundamental difference between types of solutions. Ionic compounds, such as sodium chloride, dissociate into their component ions when they dissolve in water. A dilute NaCl solution, represented as NaCl(*aq*), does not contain any NaCl units, but rather dissolved Na^+ and Cl^- ions. The dissolved ions act as charge carriers, allowing the solution to conduct electricity. Substances that dissolve in water to form solutions that conduct electricity are **electrolytes**. Substances such as sodium chloride that completely dissociate into ions when they dissolve in water are **strong electrolytes**, and the resulting solutions are strong electrolyte solutions.

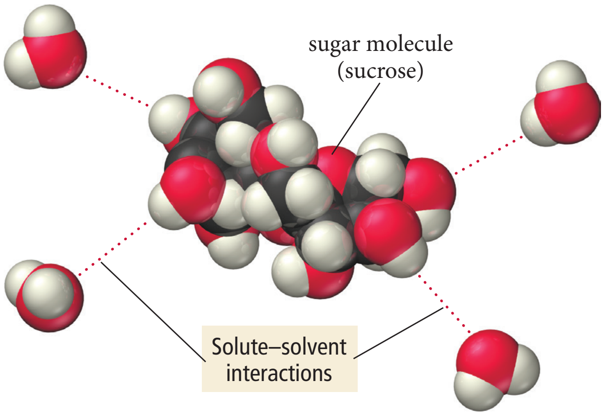
In contrast to sodium chloride, sugar is a molecular compound. Most molecular compounds—with the important exception of acids, which we discuss shortly—dissolve in water as molecules which remain intact once dissolved. Sugar dissolves because the attraction between sugar molecules and water molecules (shown in Figure 4.5 ▼) overcomes the attraction of sugar molecules to each other. So, unlike a sodium chloride solution (which is composed of dissociated ions), a sugar solution is composed of intact $C_{12}H_{22}O_{11}$ molecules homogeneously mixed with the water molecules. Compounds such as sugar that do not dissociate into ions when dissolved in water are called **nonelectrolytes**, and the resulting solutions—called *nonelectrolyte solutions*—do not conduct electricity.

Acids are molecular compounds, but they *do* ionize—form ions—when they dissolve in water. Hydrochloric acid (HCl) is a molecular compound that ionizes into H^+ and Cl^- when it dissolves in water. HCl is an example of a **strong acid**, which means that it completely ionizes in solution. Since strong acids completely ionize in solution, they

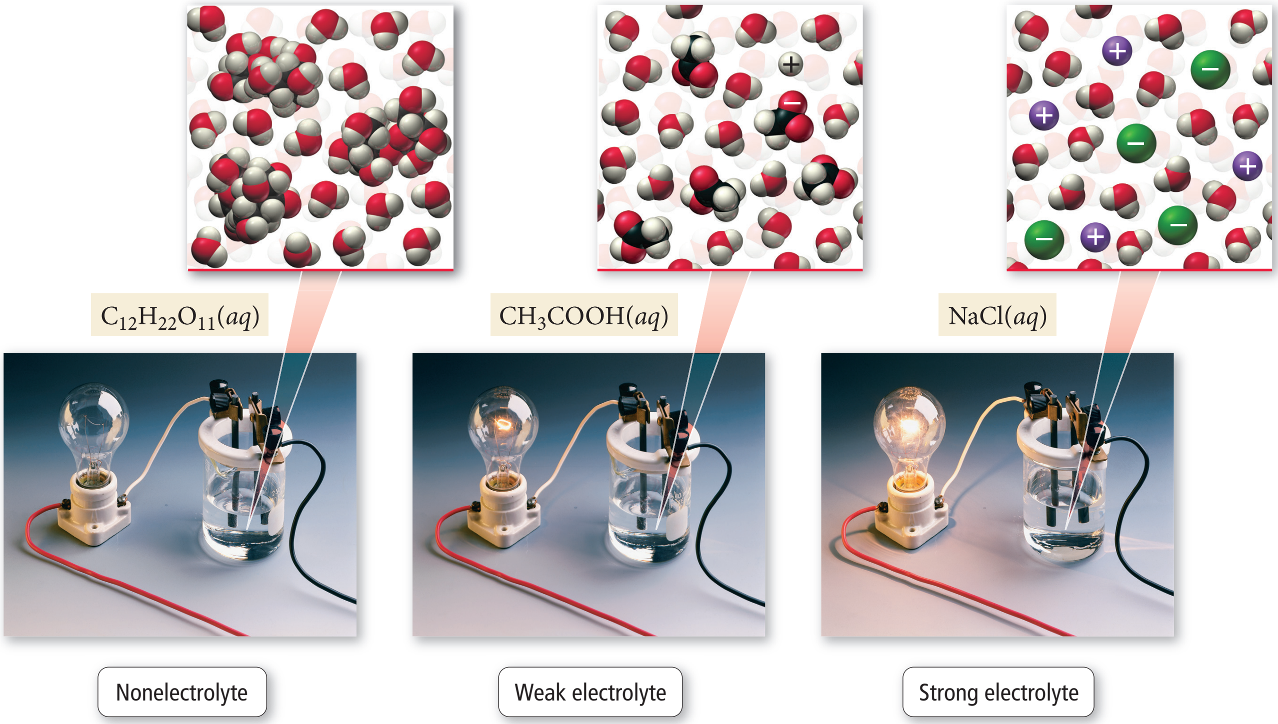
Unlike soluble ionic compounds, which contain ions and therefore *dissociate* in water, acids are molecular compounds that *ionize* in water.



▲ **FIGURE 4.4 Sodium Chloride Dissolving in Water** The attraction between water molecules and the ions of sodium chloride causes NaCl to dissolve in water.



▲ **FIGURE 4.5 Sugar and Water Interactions** Partial charges on sugar molecules and water molecules (discussed more fully in Chapter 11) result in attractions between the sugar molecules and water molecules.



▲ **FIGURE 4.6 Properties of Electrolytes in Solutions**

are also strong electrolytes. We represent the complete ionization of a strong acid with a single reaction arrow between the acid and its ionized form:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Many acids are **weak acids**; they do not completely ionize in water. For example, acetic acid CH_3COOH, the acid present in vinegar, is a weak acid. A solution of a weak acid is composed mostly of the nonionized form of the acid molecules—only a small percentage of the acid molecules ionize. We represent the partial ionization of a weak acid with opposing half arrows between the reactants and products:

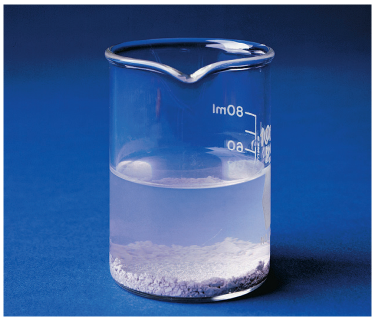
$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

Weak acids are classified as **weak electrolytes**, and the resulting solutions—called *weak electrolyte solutions*—conduct electricity only weakly. Figure 4.6 ▲ summarizes the properties of solutions containing electrolytes.

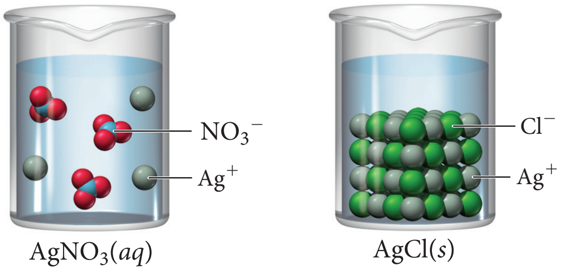
The Solubility of Ionic Compounds

We have just seen that when an ionic compound dissolves in water, the resulting solution contains the component ions dissolved in water. However, not all ionic compounds dissolve completely in water. If we add AgCl to water, for example, it remains solid and appears as a white powder at the bottom of the water.

In general, a compound is termed **soluble** if it dissolves in water and **insoluble** if it does not. However, these classifications are a bit of an oversimplification. In reality, solubility is a continuum, and even "insoluble" compounds dissolve to some extent, though usually orders of magnitude less than soluble compounds. For example, silver nitrate is soluble. If we mix solid $AgNO_3$ with water, it dissolves and forms a strong electrolyte solution. Silver chloride, on the other hand, is almost completely insoluble. If we mix solid AgCl with water, virtually all of it remains as an undissolved solid within the liquid water.



▲ AgCl does not significantly dissolve in water; it remains as a white powder at the bottom of the beaker.



There is no easy way to tell whether a particular compound is soluble or insoluble just by looking at its formula. For now, we can follow a set of empirical rules that chemists have inferred from observations on many ionic compounds. These are *solubility rules* and are listed in Table 4.2.

These rules are to be followed in order. In other words, rules with a smaller number take precedence over the rules that follow. For example, according to rule 3, all salts containing Ag^+ are insoluble, but $AgNO_3$ is soluble because according to rule 2, all compounds containing NO_3^- are soluble. As well, NaOH is quite soluble because rule 1 states that compounds containing the cations of group 1 metals are soluble, even though rule 6 states that hydroxides are insoluble.

Notice that when compounds containing polyatomic ions such as NO_3^- dissolve, the polyatomic ions dissolve as intact units.

TABLE 4.2 General Solubility Rules for Ionic Compounds in Water

1. All salts containing cations of group 1 metals (alkali metals, Li^+, Na^+, K^+, etc.) and ammonium ions (NH_4^+) are soluble.
2. All nitrates (NO_3^-), ethanoates (acetates, CH_3COO^-), chlorates (ClO_3^-), and perchlorates (ClO_4^-) are soluble.
3. Salts containing Ag^+, Pb^{2+}, and Hg_2^{2+} are insoluble.
4. Most chlorides (Cl^-), bromides (Br^-), and iodides (I^-) are soluble.
5. Sulfates (SO_4^{2-}) are soluble, except those containing Ca^{2+}, Sr^{2+}, Ba^{2+}.
6. Carbonates (CO_3^{2-}), hydroxides (OH^-), oxides (O^{2-}), phosphates (PO_4^{3-}), and sulfides (S^{2-}) are generally insoluble.

EXAMPLE 4.3 PREDICTING WHETHER AN IONIC COMPOUND IS SOLUBLE

Predict whether each compound is soluble or insoluble:

(a) $PbCl_2$ **(b)** $CuCl_2$ **(c)** $Ca(NO_3)_2$ **(d)** $BaSO_4$

SOLUTION

(a) Insoluble. Compounds containing Cl^- are normally soluble, but Pb^{2+} is an exception.

(b) Soluble. Compounds containing Cl^- are normally soluble and Cu^{2+} is not an exception.

(c) Soluble. Compounds containing NO_3^- are always soluble.

(d) Insoluble. Compounds containing SO_4^{2-} are normally soluble, but Ba^{2+} is an exception.

FOR PRACTICE 4.3

Predict whether each compound is soluble or insoluble:

(a) NiS **(b)** $Mg_3(PO_4)_2$ **(c)** Li_2CO_3 **(d)** NH_4Cl

4.4 Precipitation Reactions

▲ The reaction of ions in hard water with soap produces a grey crud you can see after you drain the water.

Have you ever taken a bath in hard water? Hard water contains dissolved ions such as Ca^{2+} and Mg^{2+} that diminish the effectiveness of soap. These ions react with soap to form a grey residue that may appear as a “bathtub ring” when you drain the tub. Hard water is particularly troublesome when washing clothes. Imagine how your white shirt would look covered with the grey crud from the bathtub and you can understand the problem. Consequently, most laundry detergents include substances designed to remove Ca^{2+} and Mg^{2+} from the laundry mixture. The most common substance used for this purpose is sodium carbonate, Na_2CO_3, which dissolves in water to form sodium cations and carbonate anions:

$$Na_2CO_3(aq) \longrightarrow 2\,Na^+(aq) + CO_3^{\ 2-}(aq)$$

Sodium carbonate is soluble (rule 1 in Table 4.2), but calcium carbonate and magnesium carbonate are not (rule 6 in Table 4.2). Consequently, the carbonate anions react with dissolved Ca^{2+} and Mg^{2+} ions in hard water to form solids that *precipitate* from (or come out of) solution:

$$\begin{aligned} Mg^{2+}(aq) + CO_3^{\ 2-}(aq) &\longrightarrow MgCO_3(s) \\ Ca^{2+}(aq) + CO_3^{\ 2-}(aq) &\longrightarrow CaCO_3(s) \end{aligned}$$

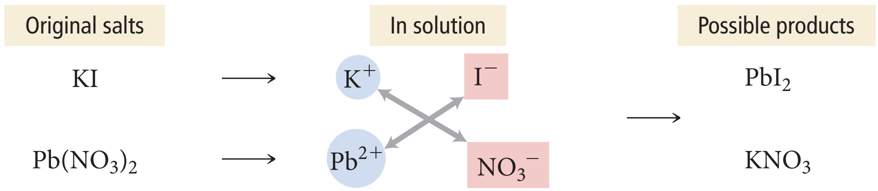
The precipitation of these ions prevents their reaction with the soap, eliminating the grey residue and preventing white shirts from turning grey.

The reactions between Ca^{2+} or Mg^{2+} with $CO_3^{\ 2-}$ are examples of **precipitation reactions**, ones in which a solid or **precipitate** forms when we mix two solutions. Precipitation reactions are common in chemistry.

As another example, consider potassium iodide and lead(II) nitrate, which each form colourless, strong electrolyte solutions when dissolved in water. When the two solutions are combined, however, a brilliant yellow precipitate forms (Figure 4.7 ▼). We can describe this precipitation reaction with the following chemical equation:

$$2\,KI(aq) + Pb(NO_3)_2(aq) \longrightarrow 2\,KNO_3(aq) + PbI_2(s) \qquad [4.1]$$

The key to predicting precipitation reactions is to understand that *only insoluble compounds form precipitates*. In a precipitation reaction, two solutions containing soluble compounds combine and an insoluble compound precipitates. In the example depicted in Figure 4.7, both KI and $Pb(NO_3)_2$ are soluble and exist in their respective solutions as K^+ and I^- and as Pb^{2+} and $NO_3^{\ -}$. Once the two solutions are mixed, two new compounds—one or both of which might be insoluble—are possible. Specifically, the cation from either compound can pair with the anion from the other to form possibly insoluble products:



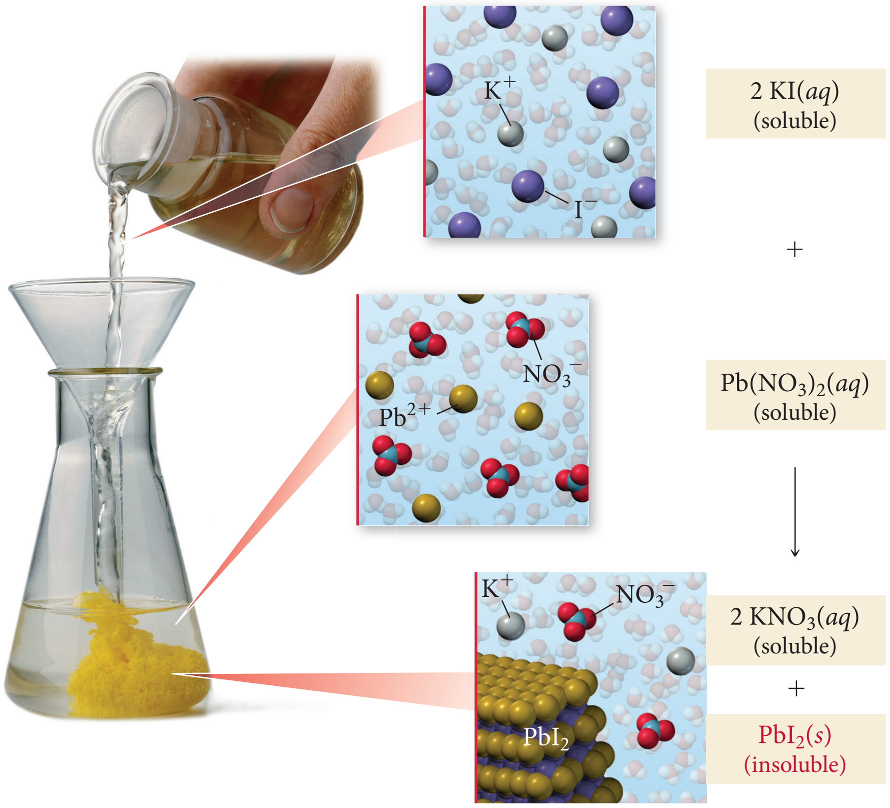
In this case, KNO_3 is soluble, but PbI_2 is essentially insoluble. Consequently, PbI_2 precipitates.

Equation 4.1 is called a **molecular equation**, an equation showing the complete neutral formulas for each compound in the reaction as if they existed as molecules. Because KI and $Pb(NO_3)_2$ are soluble, in their respective aqueous solutions, they are present only as K^+ and I^- and as Pb^{2+} and $NO_3^{\ -}$. We can write equations for reactions occurring in aqueous solution in a way that better shows the dissociated nature of dissolved ionic compounds. For example, Equation 4.1 could be rewritten as a **complete ionic equation**, which lists all of the ions present as either reactants or products in a chemical reaction:

$$\begin{aligned} &Pb^{2+}(aq) + 2\,NO_3^{\ -}(aq) + 2\,K^+(aq) + 2\,I^-(aq) \longrightarrow \\ &\qquad\qquad PbI_2(s) + 2\,NO_3^{\ -}(aq) + 2\,K^+(aq) \end{aligned}$$

FIGURE 4.7 Precipitation of Lead(II) Iodide When a potassium iodide solution is mixed with a lead(II) nitrate solution, a yellow lead(II) iodide precipitate forms.

$$\underset{\text{(soluble)}}{2\,KI(aq)} + \underset{\text{(soluble)}}{Pb(NO_3)_2(aq)} \longrightarrow \underset{\text{(soluble)}}{2\,KNO_3(aq)} + \underset{\text{(insoluble)}}{PbI_2(s)}$$



In the complete ionic equation, some ions appear on both the reactant and the product side of the equation. These ions are called **spectator ions** because they do not take part in the chemical reaction. To simplify the equation, and show more clearly what is happening, we can omit spectator ions:

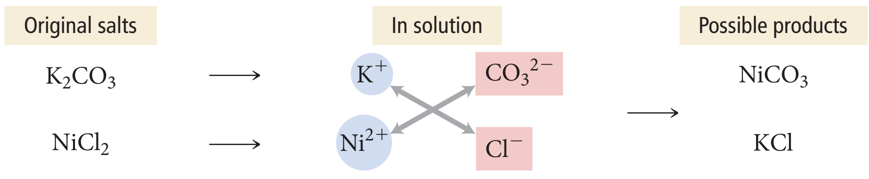
$$Pb^{2+}(aq) + 2\,I^{-}(aq) \longrightarrow PbI_2(s)$$

This **net ionic equation** is most useful because it shows only the species that take part in the reaction.

The following example shows a procedure for predicting whether a precipitation reaction occurs or not, and how to write the net ionic equation.

EXAMPLE 4.4 WRITING EQUATIONS FOR PRECIPITATION REACTIONS

Write a net ionic equation for the precipitation reaction that occurs (if any) when solutions of potassium carbonate and nickel(II) chloride are mixed.

1. Write the formula of the two compounds being mixed as reactants.	K_2CO_3 and $NiCl_2$
2. Write the ions that are present when the two ionic compounds dissolve in water, and determine the possible products that are formed.	

3. Use the solubility rules (Table 4.2) to determine whether any of the possible products are insoluble.	KCl is soluble (according to rule 1, all salts containing K^+ are soluble, and according to rule 4, all salts containing Cl^- are soluble). $NiCO_3$ is insoluble (according to rule 6, salts containing CO_3^{2-} are generally insoluble).
4. If all the possible products are soluble, there is no reaction or net ionic equation. If any of the possible products are insoluble, write the net ionic equation. It may be helpful to first write the molecular equation. Then convert the molecular equation to the complete ionic equation. Finally, cancel out the spectator ions and you will be left with the net ionic equation.	Molecular equation: $K_2CO_3(aq) + NiCl_2(aq) \longrightarrow 2\,KCl(aq) + NiCO_3(s)$ Complete ionic equation: $\cancel{2\,K^+}(aq) + CO_3^{2-}(aq) + Ni^{2+}(aq) + \cancel{2\,Cl^-}(aq) \longrightarrow \cancel{2\,K^+}(aq) + \cancel{2\,Cl^-}(aq) + NiCO_3(s)$ Net ionic equation: $Ni^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow NiCO_3(s)$

FOR PRACTICE 4.4

Write a net ionic equation for the precipitation reaction that occurs (if any) when solutions of ammonium chloride and iron(III) nitrate are mixed.

FOR MORE PRACTICE 4.4

Write a net ionic equation for the precipitation reaction that occurs (if any) when solutions of sodium hydroxide and copper(II) bromide are mixed.

4.5 Acid–Base Reactions

Another important class of reactions that occur in aqueous solution are acid–base reactions. In an **acid–base reaction** (also called a **neutralization reaction**), an acid reacts with a base and the two neutralize each other, producing water (or in some cases, a weak electrolyte). As in precipitation reactions, an acid–base reaction occurs when the anion from one reactant combines with the cation of the other.

Our stomachs contain hydrochloric acid, which acts in the digestion of food. Certain foods or stress, however, can increase the stomach's acidity to uncomfortable levels, causing what has come to be known as heartburn. Antacids are over-the-counter medicines that work by reacting with and neutralizing stomach acid. Antacids employ different *bases*—substances that produce hydroxide ions (OH^-) in water—as neutralizing agents. Milk of magnesia, for example, contains magnesium hydroxide $[Mg(OH)_2]$ and Mylanta® contains aluminum hydroxide $[Al(OH)_3]$. All antacids, regardless of the base they employ, have the same effect of neutralizing stomach acid and relieving heartburn through *acid–base reactions*.

The **Arrhenius definitions** (named after Swedish chemist Svante Arrhenius who lived from 1859 until 1927) of acids and bases are as follows:

- Acid: substance that produces H^+ ions in aqueous solution.
- Base: substance that produces OH^- ions in aqueous solution.

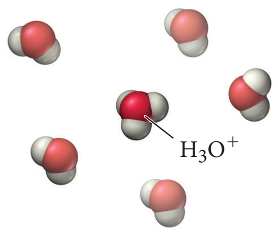
In Chapter 14, we will learn more general definitions of acids and bases, but these are sufficient to describe neutralization reactions.

According to the Arrhenius definition, HCl is an acid because it produces H^+ ions in solution:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

An H^+ ion is a bare proton. In solution, bare protons associate with water molecules to form **hydronium ions**, H_3O^+ (Figure 4.8 ▶):

$$H^+(aq) + H_2O(l) \longrightarrow H_3O^+(aq)$$



▲ **FIGURE 4.8 The Hydronium Ion** Protons associate with water molecules in solution to form H_3O^+ ions, which in turn interact with other water molecules.

▲ Lemons, limes, and vinegar contain acids. Vitamin C and aspirin are acids.

Chemists often use $H^+(aq)$ and $H_3O^+(aq)$ interchangeably. The chemical equation for the ionization of HCl is often written to show the association of the proton with a water molecule to form the hydronium ion:

$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Some acids—called **polyprotic acids**—contain more than one ionizable proton and release them sequentially. For example, sulfuric acid, H_2SO_4, is a **diprotic acid**. It is strong in its first ionizable proton, but weak in its second:

$$H_2SO_4(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + HSO_4^-(aq)$$
$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

According to the Arrhenius definition, NaOH is a base because it produces OH^- in solution:

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

In analogy to diprotic acids, some bases such as $Sr(OH)_2$ produce two moles of OH^- per mole of the base:

$$Sr(OH)_2(aq) \longrightarrow Sr^{2+}(aq) + 2\,OH^-(aq)$$

$Sr(OH)_2$, and all group 2 hydroxides, are not very soluble but all of the $Sr(OH)_2$ that does dissolve dissociates. For this reason they are considered strong bases.

In Table 4.3, some common acids and bases are listed. You can find acids and bases in many everyday substances. Foods such as citrus fruits and vinegar contain acids. Soap, baking soda, and milk of magnesia all contain bases.

When we mix an acid and a base, the H^+ from the acid—whether it is weak or strong—combines with the OH^- from the base to form H_2O (Figure 4.9 ▶). Consider the reaction between hydrochloric acid and sodium hydroxide:

$$\underset{\text{Acid}}{HCl(aq)} + \underset{\text{Base}}{NaOH(aq)} \longrightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{Salt}}{NaCl(aq)}$$

The word *salt* in this sense applies to any ionic compound and is therefore more general than the common usage, which refers only to table salt (NaCl).

Acid–base reactions generally form water and an ionic compound—called a **salt**—that usually remains dissolved in the solution. The net ionic equation for many acid–base reactions is:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

▲ Many common household products are bases.

TABLE 4.3 Common Acids and Bases

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	NaOH
Hydrobromic acid	HBr	Lithium hydroxide	LiOH
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Acetic acid	CH_3COOH (weak acid)		
Hydrofluoric acid	HF (weak acid)		

*Ammonia does not contain OH^-, but it produces OH^- in a reaction with water that occurs only to a small extent: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

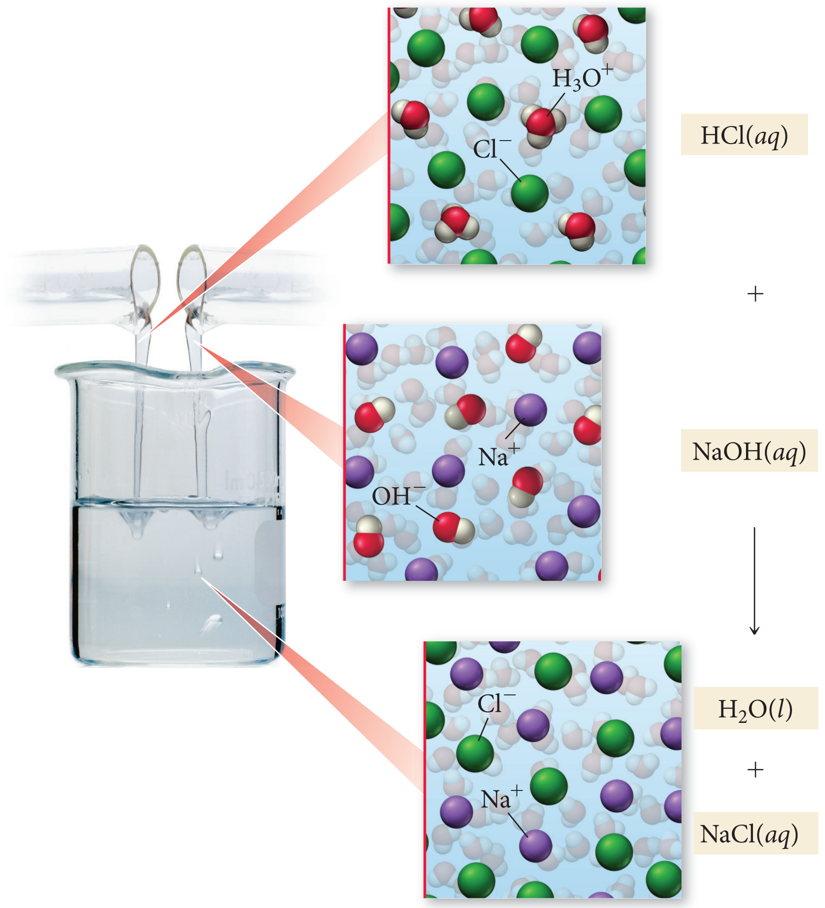


FIGURE 4.9 Acid–Base Reaction The reaction between hydrochloric acid and sodium hydroxide forms water and a salt, sodium chloride, which remains dissolved in the solution.

EXAMPLE 4.5 WRITING EQUATIONS FOR ACID–BASE REACTIONS

Write a molecular and net ionic equation for the reaction between aqueous HI and aqueous $Ba(OH)_2$.

SOLUTION

You must first identify these substances as an acid and a base. Begin by writing the skeletal reaction in which the acid and the base combine to form water and a salt.	$\underset{\text{acid}}{HI(aq)} + \underset{\text{base}}{Ba(OH)_2(aq)} \longrightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{BaI_2(aq)}$
Next, balance the equation: this is the molecular equation.	$2\ HI(aq) + Ba(OH)_2(aq) \longrightarrow 2\ H_2O(l) + BaI_2(aq)$
Write the net ionic equation by removing the spectator ions.	$2\ H^+(aq) + 2\ OH^-(aq) \longrightarrow 2\ H_2O(l)$ or simply $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

FOR PRACTICE 4.5

Write a molecular and net ionic equation for the reaction that occurs between aqueous H_2SO_4 and aqueous LiOH.

Another example of an acid–base reaction is the reaction between sulfuric acid and potassium hydroxide:

$$H_2SO_4(aq) + 2KOH(aq) \longrightarrow 2H_2O(l) + K_2SO_4(aq)$$

Again, notice the pattern of an acid and base reacting to form water and a salt:

$$\text{Acid} + \text{Base} \longrightarrow \text{Water} + \text{Salt}$$

When writing equations for acid–base reactions, write the formula of the salt using the procedure for writing formulas of ionic compounds given in Section 4.2.

Acid–Base Reactions Evolving a Gas

In some acid–base reactions, two aqueous solutions mix to form a gaseous product that bubbles out of solution. These reactions are sometimes called **gas-evolution reactions**. Some gas-evolution reactions form a gaseous product directly when the cation of one reactant combines with the anion of the other. For example, when sulfuric acid reacts with lithium sulfide, dihydrogen sulfide gas is formed:

$$H_2SO_4(aq) + Li_2S(aq) \longrightarrow H_2S(g) + Li_2SO_4(aq)$$

Similarly, when acids composed of ammonium cation react in an acid–base reaction in aqueous solution, ammonia gas is evolved:

$$NH_4Cl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NH_3(g) + NaCl(aq)$$

Other gas-evolution reactions often form an intermediate product that then decomposes (breaks down into simpler substances) to form a gas. For example, when aqueous hydrochloric acid is mixed with aqueous sodium bicarbonate, the following reactions occur:

$$HCl(aq) + NaHCO_3(aq) \longrightarrow \underset{\text{intermediate}}{H_2CO_3(aq)} + NaCl(aq)$$

$$H_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g)$$

The intermediate product, H_2CO_3, is not stable and decomposes into H_2O and gaseous CO_2. The overall reaction can be written as follows:

$$HCl(aq) + NaHCO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + NaCl(aq)$$

Like the bicarbonate example above, bases composed of carbonates, sulfites, and bisulfites also form gases in acid–base reactions occurring in aqueous solution, as is summarized in Table 4.4.

TABLE 4.4 Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H_2S	$2\,HCl(aq) + K_2S(aq) \longrightarrow H_2S(g) + 2\,KCl(aq)$
Carbonates and bicarbonates	H_2CO_3	CO_2	$2\,HCl(aq) + K_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\,KCl(aq)$
Sulfites and bisulfites	H_2SO_3	SO_2	$2\,HCl(aq) + K_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\,KCl(aq)$
Ammonium	NH_4OH	NH_3	$NH_4Cl(aq) + KOH(aq) \longrightarrow H_2O(l) + NH_3(g) + KCl(aq)$

EXAMPLE 4.6 WRITING EQUATIONS FOR GAS-EVOLUTION REACTIONS

Write a molecular equation for the gas-evolution reaction that occurs when you mix aqueous nitric acid and aqueous sodium carbonate.

Begin by writing a skeletal equation in which the cation of each reactant combines with the anion of the other.	$HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2CO_3(aq) + NaNO_3(aq)$
You must then recognize that $H_2CO_3(aq)$ decomposes into $H_2O(l)$ and $CO_2(g)$, and write these products into the equation.	$HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + NaNO_3(aq)$
Finally, balance the equation.	$2\ HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\ NaNO_3(aq)$

FOR PRACTICE 4.6

Write a molecular equation for the gas-evolution reaction that occurs when you mix aqueous hydrobromic acid and aqueous potassium sulfite.

FOR MORE PRACTICE 4.6

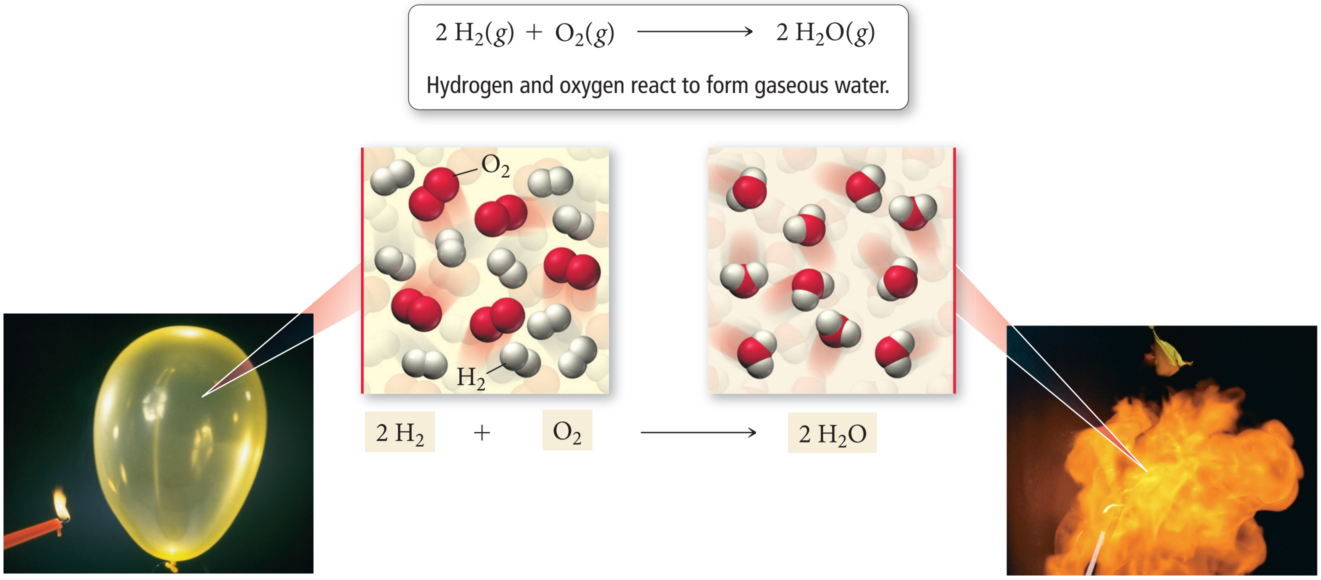
Write a net ionic equation for the reaction that occurs when you mix hydroiodic acid with calcium sulfide.

4.6 Oxidation–Reduction Reactions

Oxidation–reduction reactions or **redox reactions** are reactions in which electrons transfer from one reactant to the other. The rusting of iron, the bleaching of hair, and the production of electricity in batteries involve redox reactions. Many redox reactions (for example, combustion reactions) involve the reaction of a substance with oxygen (Figure 4.10 ▼):

Applications of oxidation–reduction reactions are covered in Chapter 18.

$$4\ Fe(s) + 3\ O_2(g) \longrightarrow 2\ Fe_2O_3(s)$$
$$C_8H_{18}(l) + \tfrac{25}{2}\ O_2(g) \longrightarrow 8\ CO_2(g) + 9\ H_2O(g)$$
$$2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(l)$$


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

Hydrogen and oxygen react to form gaseous water.

▲ **FIGURE 4.10 Oxidation–Reduction Reaction** The hydrogen in the balloon reacts with oxygen upon ignition to form gaseous water (which is dispersed in the flame).

$$2\ \text{Na}(s) + \text{Cl}_2(g) \longrightarrow 2\ \text{NaCl}(s)$$

Electrons are transferred from sodium to chlorine, forming sodium chloride. Sodium is oxidized and chlorine is reduced.

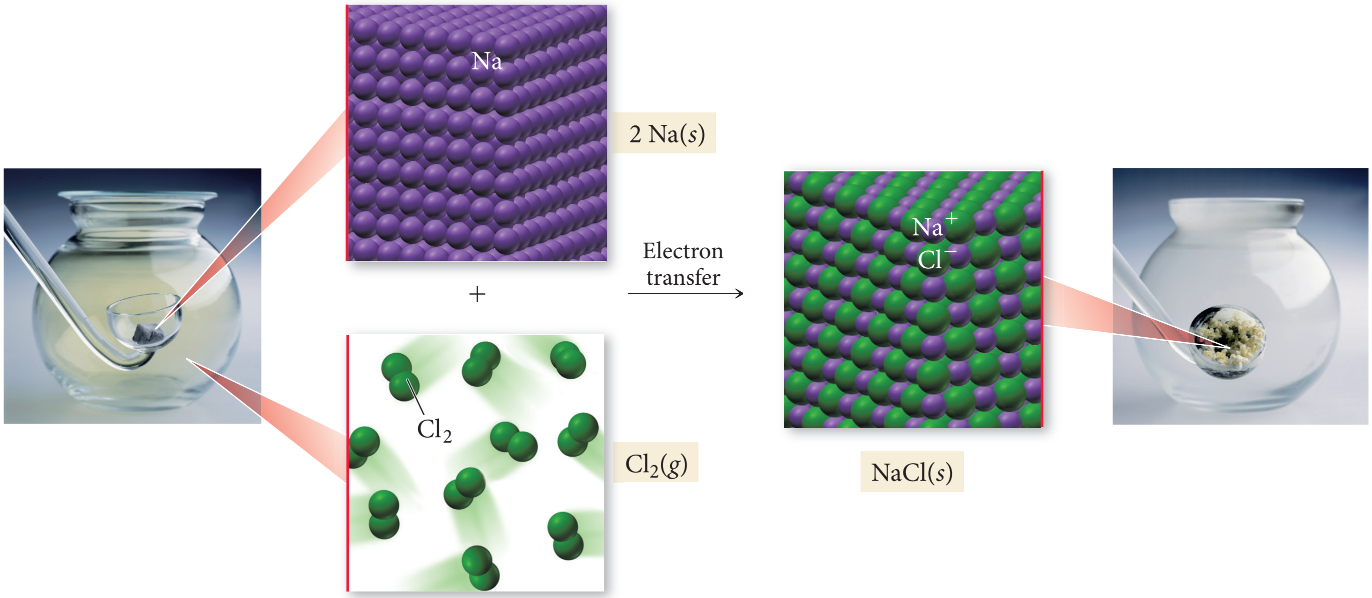


FIGURE 4.11 Oxidation–Reduction Without Oxygen When sodium reacts with chlorine, electrons transfer from the sodium to the chlorine, resulting in the formation of sodium chloride. In this redox reaction, sodium is oxidized and chlorine is reduced.

However, redox reactions need not involve oxygen. Consider, for example, the reaction between sodium and chlorine to form sodium chloride (NaCl), depicted in Figure 4.11 ▲:

$$2\ \text{Na}(s) + \text{Cl}_2(g) \longrightarrow 2\ \text{NaCl}(s)$$

This reaction is similar to the reaction between sodium and oxygen, which forms sodium oxide:

$$4\ \text{Na}(s) + \text{O}_2(g) \longrightarrow 2\ \text{Na}_2\text{O}(s)$$

In both cases, a metal (which has a tendency to lose electrons) reacts with a nonmetal (which has a tendency to gain electrons). In both cases, metal atoms lose electrons to nonmetal atoms. A fundamental definition of **oxidation** is the loss of electrons, and a fundamental definition of **reduction** is the gain of electrons.

The transfer of electrons need not be a *complete* transfer (as occurs in the formation of an ionic compound) for the reaction to qualify as oxidation–reduction. For example, consider the reaction between hydrogen gas and chlorine gas:

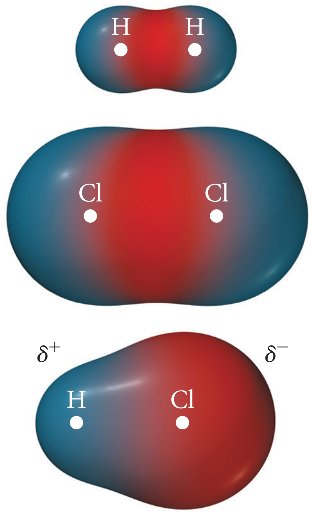
$$\text{H}_2(g) + \text{Cl}_2(g) \longrightarrow 2\ \text{HCl}(g)$$

Even though hydrogen chloride is a molecular compound with a covalent bond, and even though the hydrogen has not completely transferred its electron to chlorine during the reaction, you can see from the electron density diagrams (Figure 4.12 ◀) that hydrogen has lost some of its electron density—it has *partially* transferred its electron to chlorine. In the reaction, hydrogen is oxidized and chlorine is reduced and, therefore, this is a redox reaction.

Oxidation States

Identifying whether or not a reaction between a metal and a nonmetal is a redox reaction is fairly straightforward because of ion formation. But how do we identify redox reactions that occur between nonmetals? Chemists have devised a scheme to track electrons before and after a chemical reaction. In this scheme—which is like bookkeeping for electrons—each shared electron is assigned to the atom that attracts the electrons most strongly. Then

The ability of an element to attract electrons in a chemical bond is called electronegativity. We cover electronegativity in more detail in Section 9.7.



Hydrogen loses electron density (oxidation) and chlorine gains electron density (reduction).

FIGURE 4.12 Redox with Partial Electron Transfer When hydrogen bonds to chlorine, the electrons are unevenly shared, resulting in an increase of electron density (reduction) for chlorine and a decrease in electron density (oxidation) for hydrogen.

a number, called the **oxidation state** or oxidation number, is given to each atom based on the electron assignments. In other words, the oxidation number of an atom in a compound is the "charge" it would have if all shared electrons were assigned to the atom with the greatest attraction for those electrons.

For example, consider HCl. Since chlorine attracts electrons more strongly than hydrogen, we assign the two shared electrons in the bond to chlorine; then H (which has lost an electron in our assignment) has an oxidation state of +1, and Cl (which has gained one electron in our assignment) has an oxidation state of −1. You can use the following rules to assign oxidation states to atoms in elements and compounds.

Do not confuse oxidation state with ionic charge. Unlike ionic charge—which is a real property of an ion—the oxidation state of an atom is merely a theoretical (but useful) construct.

Rules for Assigning Oxidation States These rules are hierarchical. If any two rules conflict, follow the rule with the smaller number (i.e., rule 1 takes precedence over rule 2):

1. The oxidation state of each atom in an element is 0.
2. The oxidation state of the atom in a monoatomic ion is equal to the ion's charge.
3. The sum of the oxidation states of all atoms in:
   - a neutral molecule is always 0.
   - a polyatomic ion is always equal to the charge of the ion.
4. In their compounds, metals have positive oxidation states:
   - Group 1 metals always have an oxidation state of +1.
   - Group 2 metals always have an oxidation state of +2.
5. The oxidation state of hydrogen in a compound is usually +1.
6. In their compounds, the nonmetals typically have negative oxidation states:
   - Fluorine always has an oxidation state of −1.
   - The other group 17 elements usually have an oxidation state of −1.
   - Oxygen usually has an oxidation state of −2.
   - The other group 16 elements usually have an oxidation state of −2.
   - Group 15 elements usually have an oxidation number of −3.

Another point to keep in mind is that when assigning oxidation states to elements that are not covered by the rules, such as carbon, use rule 3 to deduce their oxidation state once all other oxidation states have been assigned. For example, when assigning oxidation states to C and H in methane, CH_4, rule 5 says that the oxidation state of hydrogen is +1. According to rule 3, the sum of all oxidation states for methane must be 0, so the oxidation state of carbon must be −4.

EXAMPLE 4.7 ASSIGNING OXIDATION STATES

Assign an oxidation state to each atom in each element, ion, or compound:

(a) Cl_2 **(b)** Na^+ **(c)** KF **(d)** CO_2 **(e)** SO_4^{2-} **(f)** K_2O_2

SOLUTION

(a) Since Cl_2 is an element, the oxidation state of both Cl atoms is 0 (rule 1).	Cl_2 $\underset{0\ \ 0}{ClCl}$
(b) Since Na^+ is a monoatomic ion, its oxidation state is equal to its charge, 1+ (rule 2).	Na^+ $\underset{+1}{Na^+}$
(c) The oxidation state of K is +1 (rule 4). Since there are only two atoms in this compound, we can use rule 3 to assign an oxidation number to F. The sum of the oxidation numbers must be 0, so the oxidation number of F is −1, which agrees with rule 6.	KF $\underset{+1\,-1}{KF}$ sum: $+1-1=0$

(continued)

EXAMPLE 4.7 (CONTINUED)

(d) The oxidation state of each oxygen is −2 (rule 6). The sum of the oxidation states must be 0 (rule 3), so the oxidation state of carbon is +4.	CO_2 (C ox state) + 2(O ox state) = 0 (C ox state) + 2(−2) = 0 C ox state = +4 CO_2 +4 −2 sum: +4+2(−2) = 0
(e) The oxidation state of each oxygen is −2 (rule 6). We might expect the oxidation state of sulfur to be −2. However, if that were the case, the sum of the oxidation states would not equal the charge on the ion. Since O is mentioned higher on the list of rules than S, it takes priority and we compute the oxidation state of sulfur by setting the sum of all the oxidation states equal to −2 (the charge of the ion, rule 3).	SO_4^{2-} (S ox state) + 4(O ox state) = −2 (S ox state) + 4(−2) = −2 S ox state = +6 SO_4^{2-} +6 −2 sum: +6+4(−2) = −2
(f) The oxidation state of potassium is +1 (rule 4). We might ordinarily expect the oxidation state of O to be −2 (rule 6), but group 1 metals are mentioned higher on the list of rules (rule 4), so we deduce the oxidation state of O by setting the sum of all the oxidation states equal to 0 (rule 3).	K_2O_2 2(K ox state) + 2(O ox state) = 0 2(+1) + 2(O ox state) = 0 O ox state = −1 K_2O_2 +1 −1 sum: 2(+1) + 2(−1) = 0

FOR PRACTICE 4.7

Assign an oxidation state to each atom in each element, ion, or compound:

(a) Cr **(b)** Cr^{3+} **(c)** CCl_4 **(d)** $SrBr_2$ **(e)** SO_3 **(f)** NO_3^-

In most cases, oxidation states are positive or negative integers; on occasion, an atom within a compound can have a fractional oxidation state. Consider KO_2. The oxidation states are assigned as follows:

KO_2
$+1\ -\frac{1}{2}$
sum: $+1+2\left(-\frac{1}{2}\right) = 0$

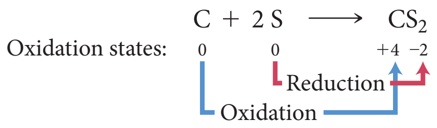
In KO_2, oxygen has a $-\frac{1}{2}$ oxidation state. Although this seems unusual, it is acceptable because oxidation states are merely an imposed electron bookkeeping scheme, not an actual physical quantity.

Identifying Redox Reactions

We can use oxidation states to identify redox reactions, even between nonmetals. For example, is the following reaction between carbon and sulfur a redox reaction?

$$C + 2S \longrightarrow CS_2$$

If so, what element is oxidized? What element is reduced? We can use the oxidation state rules to assign oxidation states to all elements on both sides of the equation.


$$\text{Oxidation states:}\quad \underset{0}{C} + 2\,\underset{0}{S} \longrightarrow \underset{+4}{C}\underset{-2}{S_2}$$

Carbon changes from an oxidation state of 0 to an oxidation state of +4. In terms of our electron bookkeeping scheme (the assigned oxidation state), carbon *loses electrons* during the conversion of reactants to products and is *oxidized.* During the conversion of reactants to products, sulfur changes from an oxidation state of 0 to an oxidation state of −2. In terms of our electron bookkeeping scheme, sulfur *gains electrons* and is *reduced.* In terms of oxidation states, oxidation and reduction are defined as follows:

- Oxidation: an increase in oxidation state.
- Reduction: a decrease in oxidation state.

Notice that oxidation and reduction must occur together. If one substance loses electrons (oxidation), then another substance must gain electrons (reduction). A substance that causes the oxidation of another substance is called an **oxidizing agent**. Oxygen, for example, is an excellent oxidizing agent because it causes the oxidation of many substances. In a redox reaction, *the oxidizing agent is always reduced*. A substance that causes the reduction of another substance is called a **reducing agent**. Hydrogen, for example, as well as the group 1 and group 2 metals (because of their tendency to lose electrons) are excellent reducing agents. In a redox reaction, *the reducing agent is always oxidized*.

Remember that a reduction is a *reduction* in oxidation state.

You will learn more about redox reactions below, including how to balance them. For now, you need to be able to identify redox reactions, as well as oxidizing and reducing agents, according to the following guidelines:

Redox reactions:

- Any reaction in which there is a change in the oxidation states of atoms in going from reactants to products.

In a redox reaction:

- The oxidizing agent oxidizes another substance (and is itself reduced).
- The reducing agent reduces another substance (and is itself oxidized).

EXAMPLE 4.8 IDENTIFYING REDOX REACTIONS, OXIDIZING AGENTS, AND REDUCING AGENTS

Determine whether each reaction is an oxidation–reduction reaction. For each oxidation–reduction reaction, identify the oxidizing agent and the reducing agent.

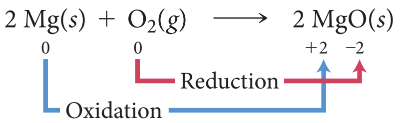
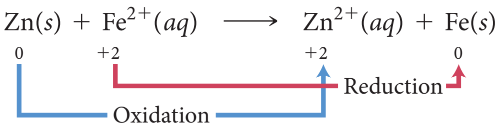
(a) $2\,\text{Mg}(s) + \text{O}_2(g) \longrightarrow 2\,\text{MgO}(s)$

(b) $2\,\text{HBr}(aq) + \text{Ca(OH)}_2(aq) \longrightarrow 2\,\text{H}_2\text{O}(l) + \text{CaBr}_2(aq)$

(c) $\text{Zn}(s) + 2\,\text{Fe}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + 2\,\text{Fe}(s)$

(d) $\text{C}_2\text{H}_6(g) + \frac{7}{2}\text{O}_2(g) \longrightarrow 2\,\text{CO}_2(g) + 3\,\text{H}_2\text{O}(l)$

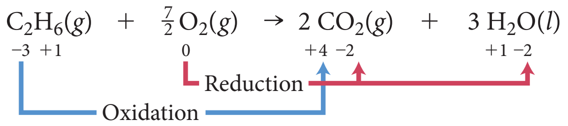
SOLUTION

(a) This is a redox reaction because magnesium increases in oxidation number (oxidation) and oxygen decreases in oxidation number (reduction).	 Oxidizing agent: O_2 Reducing agent: Mg
(b) This is not a redox reaction because none of the atoms undergoes a change in oxidation number. This is an acid–base or neutralization reaction.	$2\,\underset{+1\,-1}{\text{HBr}}(aq) + \underset{+2\,-2\,+1}{\text{Ca(OH)}_2}(aq) \longrightarrow 2\,\underset{+1\,-2}{\text{H}_2\text{O}}(l) + \underset{+2\,-1}{\text{CaBr}_2}(aq)$
(c) This is a redox reaction because zinc increases in oxidation number (oxidation) and iron decreases in oxidation number (reduction).	 Oxidizing agent: Fe^{2+} Reducing agent: Zn

(continued)

EXAMPLE 4.8 (CONTINUED)

(d) This is a redox reaction because carbon increases in oxidation number (oxidation) and oxygen decreases in oxidation number (reduction). Combustion reactions are redox reactions.



$$\underset{-3\ +1}{C_2H_6}(g) + \tfrac{7}{2}\underset{0}{O_2}(g) \rightarrow 2\,\underset{+4\ -2}{CO_2}(g) + 3\,\underset{+1\ -2}{H_2O}(l)$$

Oxidizing agent: O_2
Reducing agent: C_2H_6

FOR PRACTICE 4.8

Determine whether each reaction is an oxidation–reduction reaction. For all redox reactions, identify the oxidizing agent and the reducing agent.

(a) $2\,Li(s) + Cl_2(g) \longrightarrow 2\,LiCl(s)$
(b) $2\,Al(s) + 3\,Sn^{2+}(aq) \longrightarrow 2\,Al^{3+}(aq) + 3\,Sn(s)$
(c) $Pb(NO_3)_2(aq) + 2\,LiCl(aq) \longrightarrow PbCl_2(s) + 2\,LiNO_3(aq)$
(d) $C(s) + O_2(g) \longrightarrow CO_2(g)$

CONCEPTUAL CONNECTION 4.2

Oxidation and Reduction

Which statement is true?

(a) A redox reaction involves *either* the transfer of an electron *or* a change in the oxidation state of an element.

(b) If any of the reactants or products in a reaction contain oxygen, the reaction is a redox reaction.

(c) In a reaction, oxidation can occur independently of reduction.

(d) In a redox reaction, any increase in the oxidation state of a reactant must be accompanied by a decrease in the oxidation state of a reactant.

CHEMISTRY IN YOUR DAY Bleached Blonde

Have you ever bleached your hair? Most home kits for hair bleaching contain hydrogen peroxide (H_2O_2), an excellent oxidizing agent. When applied to hair, hydrogen peroxide oxidizes melanin, the dark pigment that gives hair its colour. Once melanin is oxidized, it no longer imparts a dark colour to hair, leaving the hair with the familiar bleached look. Hydrogen peroxide also oxidizes other components of hair. For example, protein molecules in hair contain $-SH$ groups called thiols. Hydrogen peroxide oxidizes these thiol groups to sulfonic acid groups, $-SO_3H$. The oxidation of thiol groups to sulfonic acid groups causes changes in the proteins that compose hair, making the hair more brittle and more likely to tangle. Consequently, people with heavily bleached hair generally use conditioners, which contain compounds that form thin, lubricating coatings on individual hair shafts. These coatings prevent tangling and make hair softer and more manageable.

Question

The following is a reaction of hydrogen peroxide with an alkene:

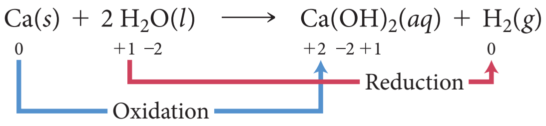
$$H_2O_2 + C_2H_4 \longrightarrow C_2H_4O + H_2O$$

Can you see why this reaction is a redox reaction? Can you identify the oxidizing and reducing agents?

▶ The bleaching of hair involves a redox reaction in which melanin—the main pigment in hair—is oxidized.

Balancing Oxidation–Reduction Equations

Consider the following reaction between calcium and water:



$$\underset{0}{Ca(s)} + 2\,\underset{+1\ -2}{H_2O(l)} \longrightarrow \underset{+2\ -2\ +1}{Ca(OH)_2(aq)} + \underset{0}{H_2(g)}$$

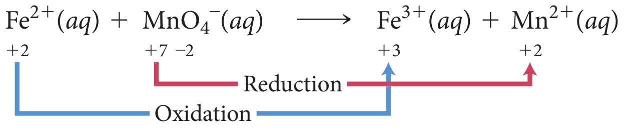
Since calcium increases in oxidation state from 0 to +2, it is oxidized. Since hydrogen decreases in oxidation state from +1 to 0, it is reduced.

Balancing redox reactions can be more complicated than balancing other types of reactions because both the mass (or number of each type of atom) and the *charge* must be balanced. Redox reactions occurring in aqueous solutions can be balanced by using a special procedure called the *half-reaction method of balancing.* In this procedure, the overall equation is broken down into two half-reactions: one for oxidation and one for reduction. The half-reactions are balanced individually and then added together. The steps differ slightly for reactions occurring in acidic and in basic solution. The following example demonstrates the method used for an acidic solution, and Example 4.10 demonstrates the method used for a basic solution.

EXAMPLE 4.9 HALF-REACTION METHOD OF BALANCING AQUEOUS REDOX REACTIONS IN ACIDIC SOLUTION

Balance the following redox reaction in acidic solution:

$$Fe^{2+}(aq) + MnO_4^-(aq) \longrightarrow Fe^{3+}(aq) + Mn^{2+}(aq)$$

1. *Assign oxidation states* to all atoms and identify the substances being oxidized and reduced. This step simply allows you to categorize, with confidence, the half-reactions as oxidation or reduction in the next step.	 $\underset{+2}{Fe^{2+}(aq)} + \underset{+7\ -2}{MnO_4^-(aq)} \longrightarrow \underset{+3}{Fe^{3+}(aq)} + \underset{+2}{Mn^{2+}(aq)}$ Fe^{2+} is being oxidized by MnO_4^- to Fe^{3+}. MnO_4^- is the oxidizing agent. The reducing agent, Fe^{2+}, reduces MnO_4^- to Mn^{2+}.
2. *Separate the overall reaction into two half-reactions*: one for oxidation and one for reduction.	**Oxidation:** $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq)$ **Reduction:** $MnO_4^-(aq) \longrightarrow Mn^{2+}(aq)$
3. *Balance each half-reaction with respect to mass* in the following order: • Balance all elements other than H and O. • Balance O by adding H_2O. • Balance H by adding H^+.	All elements other than O and H are balanced, so proceed to balance O and then H. The oxidation half-reaction does not contain O or H, so it remains the same for this step. $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq)$ Since the left side of the reaction contains 4 O, we add 4 H_2O to the right side. This results in 8 H on the right side, so we add 8 H^+ to the left side. All elements should now be balanced. $8\,H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\,H_2O(l)$
4. *Balance each half-reaction with respect to charge* by adding electrons. (Make the sum of the charges on both sides of the equation equal by adding as many electrons as necessary.)	For the oxidation reaction, there is a net positive charge on the right side, so we add one electron to the right to balance this charge: $Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$ The reduction reaction has a net 7+ on the left side and 2+ on the right side. To balance the charge, we add five electrons to the left side: $5e^- + 8\,H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\,H_2O(l)$

(*continued*)

EXAMPLE 4.9 (CONTINUED)

5. *Make the number of electrons in both half-reactions equal* by multiplying one or both half-reactions by a small whole number.

Multiply the oxidation reaction by 5 so that there are an equal number of electrons on the left side of the oxidation half-reaction as there are electrons on the right side of the reduction half-reaction.

$$5\ Fe^{2+}(aq) \longrightarrow 5\ Fe^{3+}(aq) + 5e^-$$

$$5\ e^- + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$

6. *Add the two half-reactions together,* cancelling electrons and other species as necessary.

$$5\ Fe^{2+}(aq) \longrightarrow 5\ Fe^{3+}(aq) + \cancel{5e^-}$$

$$\cancel{5e^-} + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$

$$5\ Fe^{2+}(aq) + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(l)$$

7. *Verify that the reaction is balanced* both with respect to mass and with respect to charge.

Reactants	Products
5 Fe	5 Fe
8 H	8 H
1 Mn	1 Mn
4 O	4 O
+17 charge	+17 charge

FOR PRACTICE 4.9

Balance the following redox reaction in acidic solution:

$$H^+(aq) + Cr(s) \longrightarrow H_2(g) + Cr^{2+}(aq)$$

FOR MORE PRACTICE 4.9

Balance the following redox reaction in acidic solution:

$$Cu(s) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g)$$

When a redox reaction occurs in basic solution, you can balance the reaction in exactly the same way, except that you must add an additional step to neutralize any H^+ with OH^-. The H^+ and the OH^- combine to form H_2O as shown in the following example.

EXAMPLE 4.10 BALANCING REDOX REACTIONS OCCURRING IN BASIC SOLUTION

In Example 4.9, we balanced the reaction between iron(II) and the permanganate ion in acidic solution:

$$5\ Fe^{2+}(aq) + 8\ H^+(aq) + MnO_4^-(aq) \longrightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(l)$$

Balance this reaction in basic solution.

SOLUTION

To balance redox reactions occurring in basic solution, follow the half-reaction method outlined above for acidic solutions, but add an extra step to neutralize the H^+ with OH^- as shown in step 1 below.

1. Starting with the balanced redox reaction in acidic solution, neutralize H^+ by adding the same number of OH^- ions to both sides of the equation.

$$5\ Fe^{2+}(aq) + 8\ H^+(aq) + 8\ OH^-(aq) + MnO_4^-(aq) \longrightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(l) + 8\ OH^-(aq)$$

2. The side with both H^+ and OH^- will combine to form water.	$5\ Fe^{2+}(aq) + \cancel{8}^{4}\ H_2O(l) + MnO_4^-(aq)$ $\longrightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + \cancel{4H_2O(l)} + 8\ OH^-(aq)$
Then cancel as much water as possible from both sides of the equation.	$5\ Fe^{2+}(aq) + 4\ H_2O(l) + MnO_4^-(aq)$ $\longrightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 8\ OH^-(aq)$

3. Verify that the reaction is balanced.

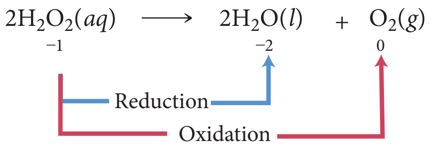
Reactants	Products
5 Fe	5 Fe
8 H	8 H
5 O	5 O
1 Mn	1 Mn
−9 charge	−9 charge

FOR PRACTICE 4.10

Balance the following redox reaction occurring in basic solution.

$$ClO^-(aq) + Cr(OH)_4^-(aq) \longrightarrow CrO_4^{2-}(aq) + Cl^-(aq)$$

Disproportionation Reactions A **disproportionation reaction** is one in which a reactant is both oxidized and reduced. For example, hydrogen peroxide (H_2O_2) decomposes to form H_2O and O_2. Some H_2O_2 is reduced, which forms H_2O, and some H_2O_2 is oxidized, which forms O_2:



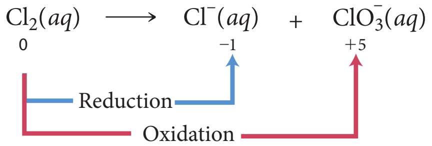
$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

To balance disproportionation reactions we follow the same method as laid out in Examples 4.9 and 4.10.

EXAMPLE 4.11 BALANCING DISPROPORTIONATION REACTIONS

Balance the disproportionation reaction of chlorine to form chloride and chlorate ions in basic solution:

$$CI_2(aq) \longrightarrow CI^-(aq) + CIO_3^-(aq)$$

1. Assign oxidation states to help categorize the half-reactions.	 $Cl_2(aq) \longrightarrow Cl^-(aq) + ClO_3^-(aq)$
2. Separate the overall reaction into half-reactions.	**Reduction:** $CI_2(aq) \longrightarrow CI^-(aq)$ **Oxidation:** $CI_2(aq) \longrightarrow CIO_3^-(aq)$
3. Balance each half-reaction with respect to mass:	
- Balance all elements other than H and O.	$CI_2(aq) \longrightarrow 2\ CI^-(aq)$ $CI_2(aq) \longrightarrow 2\ CIO_3^-(aq)$
- Balance O by adding H_2O.	$CI_2(aq) \longrightarrow 2\ CI^-(aq)$ $CI_2(aq) + 6\ H_2O(l) \longrightarrow 2\ CIO_3^-(aq)$
- Balance H by adding H^+.	$CI_2(aq) \longrightarrow 2\ CI^-(aq)$ $CI_2(aq) + 6\ H_2O(l) \longrightarrow 2\ CIO_3^-(aq) + 12\ H^+(aq)$

(continued)

EXAMPLE 4.11 **(CONTINUED)**

<table>
<tr><td>4. Balance each half-reaction with respect to charge.</td><td>$CI_2\,(aq) + 2e^- \longrightarrow 2\,CI^-(aq)$
$CI_2\,(aq) + 6\,H_2O(l) \longrightarrow 2\,CIO_3^-(aq) + 12\,H^+(aq) + 10e^-$</td></tr>
<tr><td>5. Make the number of electrons in both equations the same by multiplying each by a small whole number.</td><td>$5\,CI_2\,(aq) + 10e^- \longrightarrow 10\,CI^-(aq)$
$CI_2\,(aq) + 6\,H_2O(l) \longrightarrow 2\,CIO_3^-(aq) + 12\,H^+(aq) + 10e^-$</td></tr>
<tr><td>6. Add the half-reactions together and make necessary cancellations.</td><td>$5\,CI_2\,(aq) + \cancel{10e^-} \longrightarrow 10\,CI^-(aq)$
$CI_2\,(aq) + 6\,H_2O(l) \longrightarrow 2\,CIO_3^-(aq) + 12\,H^+(aq) + \cancel{10e^-}$
$6\,CI_2\,(aq) + 6\,H_2O(l) \longrightarrow 2\,CIO_3^-(aq) + 10\,CI^-(aq) + 12\,H^+\,(aq)$</td></tr>
<tr><td>7. Neutralize H+ by adding the same number of OH^- ions to both sides of the equation.
The side with both H^+ and OH^- will combine to form water.
Then cancel as much water as possible from both sides of the equation. In this case, we can divide each stoichiometric coefficient by a factor of two.</td><td>$6\,CI_2\,(aq) + 6\,H_2O(l) + 12\,OH^-\,(aq) \longrightarrow 2\,CIO_3^-(aq) + 10\,CI^-(aq) + 12\,H^+\,(aq) + 12\,OH^-(aq)$
$6\,CI_2\,(aq) + \cancel{6\,H_2O}(l) + 12\,OH^-\,(aq) \longrightarrow 2\,CIO_3^-(aq) + 10\,CI^-(aq) + \overset{6}{\cancel{12}}\,H_2O\,(l)$
$6\,CI_2\,(aq) + 12\,OH^-(aq) \longrightarrow 2\,CIO_3^-(aq) + 10\,CI^-(aq) + 6\,H_2O\,(l)$
$3\,CI_2\,(aq) + 6\,OH^-(aq) \longrightarrow CIO_3^-(aq) + 5\,CI^-(aq) + 3\,H_2O\,(l)$</td></tr>
<tr><td>8. Verify that the reaction is balanced with respect to mass and charge.</td><td><table><tr><th>Reactants</th><th>Products</th></tr><tr><td>6 CI</td><td>6 CI</td></tr><tr><td>6 H</td><td>6 H</td></tr><tr><td>6 O</td><td>6 O</td></tr><tr><td>−6 charge</td><td>−6 charge</td></tr></table></td></tr>
</table>

FOR PRACTICE 4.11

Balance the following disproportionation reaction of chlorine dioxide in acidic medium.

$$ClO_2 \longrightarrow ClO_2^- + ClO_3^-$$

4.7 Reaction Stoichiometry: How Much Is Produced?

A balanced chemical equation provides the exact relationship between the amount of reactant and the amount of product. For example, in a combustion reaction, the chemical equation provides a relationship between the amount of fuel burned and the amount of carbon dioxide emitted. In the following discussion, we use octane (a component of gasoline) as our fuel. The balanced equation for the combustion of octane is:

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

The balanced equation shows that 16 CO_2 molecules are produced for every 2 molecules of octane burned, or that 8 CO_2 molecules are produced for every mole of octane burned. It is equally correct to write a combustion reaction such as the preceding one to show only one molecule of the fuel by dividing all the coefficients by a factor of 2:

$$C_8H_{18}(g) + \tfrac{25}{2}O_2(g) \longrightarrow 8\,CO_2(g) + 9\,H_2O(l)$$

Nothing has changed; there are still 8 CO_2 molecules produced for every mole of octane burned. We can extend this numerical relationship between molecules to the amounts in moles as follows:

> **The coefficients in a chemical reaction specify the relative amounts in moles of each of the substances involved in the reaction.**

In other words, from the equation, we know that eight *moles* of CO_2 are produced for every *mole* of octane burned. The use of the numerical relationships between chemical amounts in a balanced chemical equation are called reaction **stoichiometries**. Stoichiometry allows us to predict the amounts of products that will form in a chemical reaction based on the amounts of reactants that react. Stoichiometry also allows us to determine the amount of reactants necessary to form a given amount of product. These calculations are central to chemistry, allowing chemists to plan and carry out chemical reactions in order to obtain products in the desired quantities.

Stoichiometry is pronounced stoy-kee-om-e-tree.

Making Molecules: Mole-to-Mole Conversions

A balanced chemical equation is simply a “recipe” for how reactants combine to form products. From our balanced equation for the combustion of octane, for example, we can write the following stoichiometric ratio:

$$1\ \text{mol}\ C_8H_{18} : 8\ \text{mol}\ CO_2$$

We can use this ratio to determine how many moles of CO_2 form when a given number of moles of C_8H_{18} burns. A typical car or small truck requires 50.0 L, which is about 300 mol of octane, to fill the tank. Suppose we burn this entire tank of gas, 300 mol of C_8H_{18}; how many moles of CO_2 form? We use the ratio from the balanced chemical equation. The ratio acts as a conversion factor between the amount in moles of the reactant C_8H_{18} and the amount in moles of the product:

$$300\ \cancel{\text{mol}\ C_8H_{18}} \times \frac{8\ \text{mol}\ CO_2}{1\ \cancel{\text{mol}\ C_8H_{18}}} = 2.40 \times 10^3\ \text{mol}\ CO_2$$

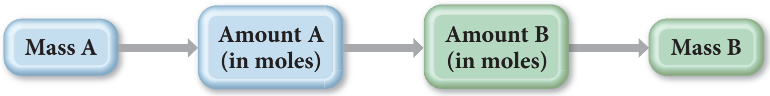
The combustion of 300 mol of C_8H_{18} (1 tank of gas) adds 2400 mol of CO_2 to the atmosphere!

Making Molecules: Mass-to-Mass Conversions

Respiration is a chemical reaction where glucose reacts with oxygen to form carbon dioxide and water:

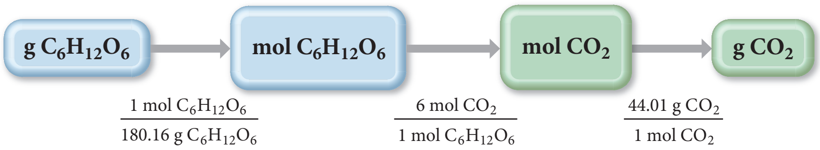
$$C_6H_{12}O_6(s) + 6\ O_2(g) \longrightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

Let’s estimate the mass of CO_2 that is produced by respiration of the glucose equivalent to the sugar in one chocolate bar, approximately 36.0 g. The calculation is similar to the one done in the previous section, but this time we are given the mass of fuel ($C_6H_{12}O_6$) instead of the number of moles. Consequently, we must first convert the mass (in grams) to the amount (in moles). The general conceptual plan for calculations where you are given the mass of a reactant or product in a chemical reaction and asked to find the mass of a different reactant or product takes the form:



where A and B are two different substances involved in the reaction. We use the molar mass of A to convert from the mass of A to the amount of A (in moles). Then, we use the appropriate ratio from the balanced chemical equation to convert from the amount of A (in moles) to the amount of B (in moles). Finally, we use the molar mass of B to convert from the amount of B (in moles) to the mass of B. To calculate the mass of CO_2 produced from the respiration of 36.0 g of glucose, we use the following conceptual plan:

Conceptual Plan



Relationships Used

$$\begin{aligned}&1 \text{ mol } C_6H_{12}O_6 : 6 \text{ mol } CO_2 \text{ (from the chemical equation)}\\&\text{molar mass of } C_6H_{12}O_6 = 180.16 \text{ g mol}^{-1}\\&\text{molar mass of } CO_2 = 44.01 \text{ g mol}^{-1}\end{aligned}$$

Solution

We follow the conceptual plan to solve the problem, beginning with the mass of $C_6H_{12}O_6$ and cancelling units to arrive at the mass of CO_2:

$$36.0 \text{ g } \cancel{C_6H_{12}O_6} \times \frac{1 \text{ mol } \cancel{C_6H_{12}O_6}}{180.16 \text{ g } \cancel{C_6H_{12}O_6}} \times \frac{6 \text{ mol } \cancel{CO_2}}{1 \text{ mol } \cancel{C_6H_{12}O_6}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } \cancel{CO_2}}$$
$$= 52.8 \text{ g } CO_2$$

The following are additional examples of stoichiometric calculations.

EXAMPLE 4.12 STOICHIOMETRY

In photosynthesis, plants convert carbon dioxide and water into glucose ($C_6H_{12}O_6$) according to the following reaction:

$$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{sunlight}} 6O_2(g) + C_6H_{12}O_6(aq)$$

Suppose you determine that a particular plant consumes 37.8 g of CO_2 in one week. Assuming that there is more than enough water present to react with all of the CO_2, what mass of glucose (in grams) can the plant synthesize from the CO_2?

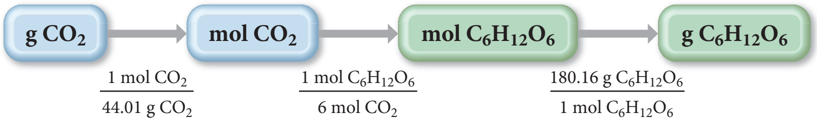
SORT The problem gives the mass of carbon dioxide and asks you to find the mass of glucose that can be produced.

GIVEN: $37.8 \text{ g } CO_2$

FIND: g $C_6H_{12}O_6$

STRATEGIZE The conceptual plan follows the general pattern of mass A $\longrightarrow$ amount A (in moles) $\longrightarrow$ amount B (in moles) $\longrightarrow$ mass B. From the chemical equation, deduce the relationship between moles of carbon dioxide and moles of glucose. Use the molar masses to convert between grams and moles.

CONCEPTUAL PLAN



RELATIONSHIPS USED

molar mass $CO_2 = 44.01 \text{ g mol}^{-1}$
$6 \text{ mol } CO_2 : 1 \text{ mol } C_6H_{12}O_6$
molar mass $C_6H_{12}O_6 = 180.16 \text{ g mol}^{-1}$

SOLVE Follow the conceptual plan to solve the problem. Begin with g CO_2 and use the conversion factors to arrive at g $C_6H_{12}O_6$.

SOLUTION

$$37.8 \text{ g } \cancel{CO_2} \times \frac{1 \text{ mol } \cancel{CO_2}}{44.01 \text{ g } \cancel{CO_2}} \times \frac{1 \text{ mol } \cancel{C_6H_{12}O_6}}{6 \text{ mol } \cancel{CO_2}} \times \frac{180.16 \text{ g } C_6H_{12}O_6}{1 \text{ mol } \cancel{C_6H_{12}O_6}} = 25.8 \text{ g } C_6H_{12}O_6$$

CHECK The units of the answer are correct. The magnitude of the answer (25.8 g) is less than the initial mass of CO_2 (37.8 g). This is reasonable because each carbon in CO_2 has two oxygen atoms associated with it, while in $C_6H_{12}O_6$, each carbon has only one oxygen atom and two hydrogen atoms (which are much lighter than oxygen) associated with it. Therefore, the mass of glucose produced should be less than the mass of carbon dioxide for this reaction.

FOR PRACTICE 4.12

Magnesium hydroxide, the active ingredient in milk of magnesia, neutralizes stomach acid, primarily HCl, according to the following reaction:

$$Mg(OH)_2(aq) + 2\,HCl(aq) \longrightarrow 2\,H_2O(l) + MgCl_2(aq)$$

What mass of HCl, in grams, can be neutralized by a dose of milk of magnesia containing 3.26 g $Mg(OH)_2$?

FOR MORE PRACTICE 4.12

One component of acid rain is nitric acid, which forms when NO_2, also a pollutant, reacts with oxygen and water according to the following simplified equation:

$$4\,NO_2(g) + O_2(g) + 2\,H_2O(l) \longrightarrow 4\,HNO_3(aq)$$

The generation of the electricity used by a medium-sized home produces about 16 kg of NO_2 per year. Assuming that there is adequate O_2 and H_2O, what mass of HNO_3, in kg, can form from this amount of NO_2 pollutant?

CONCEPTUAL CONNECTION 4.3

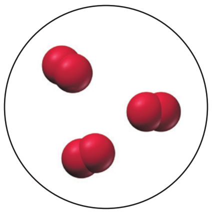
Stoichiometry

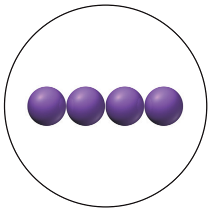
Under certain conditions, sodium can react with oxygen to form sodium oxide according to the following reaction:

$$4\,Na(s) + O_2(g) \longrightarrow 2\,Na_2O(s)$$

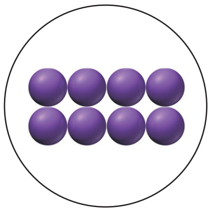
A flask contains the amount of oxygen represented by the diagram on the right ▶.

Which diagram best represents the amount of sodium required to completely react with all of the oxygen in the flask, according to the above equation?

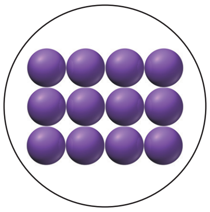




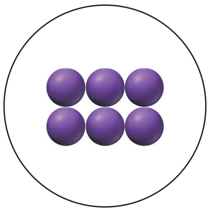
(a)



(b)



(c)



(d)

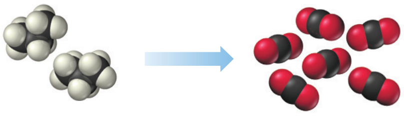
4.8 Limiting Reactant, Theoretical Yield, and Percent Yield

It is intuitive to begin with an example to understand three more important concepts in reaction stoichiometry: limiting reactant, theoretical yield, and percent yield. We will use the combustion of propane (C_3H_8), the most common barbeque fuel, as our example. The balanced equation for the combustion of propane is:

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(l)$$

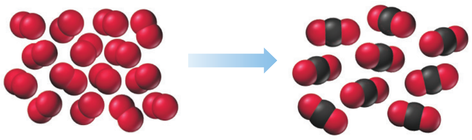
Supposing we start with 2 molecules of C_3H_8 and 15 molecules of O_2, what is the reactant that limits the amount of the products that can be formed (the **limiting reactant**)? What is the maximum amount of CO_2 that can be produced from the limiting reactant (the **theoretical yield**)? First we must calculate the number of CO_2 molecules that can be made from 2 molecules of C_3H_8:

$$2\,\cancel{C_3H_8} \times \frac{3\,CO_2}{1\,\cancel{C_3H_8}} = 6\,CO_2$$



Next, we calculate the number of CO_2 molecules that can be made from 15 O_2 molecules:

$$15\,\cancel{O_2} \times \frac{3\,CO_2}{5\,\cancel{O_2}} = 9\,CO_2$$



We have enough C_3H_8 to make 6 CO_2 molecules and enough O_2 to make 9 CO_2 molecules; for this reason, C_3H_8 is the *limiting reactant* and 6 molecules of CO_2 is the *theoretical yield* for our present starting mixture of 2 propane molecules and 15 O_2 molecules. O_2 is in excess.

An alternate way of calculating the limiting reactant is to pick any reactant and determine how much of the *other reactant* is necessary to completely react with it. For the reaction we just examined, we have 2 C_3H_8 molecules and 15 O_2 molecules. Let's see how much O_2 is need to completely react with the 2 molecules of C_3H_8:

$$2\,\cancel{C_3H_8} \times \frac{5\,O_2}{1\,\cancel{C_3H_8}} = 10\,O_2$$

Since we only need 10 O_2 molecules to completely react with the 2 C_3H_8 molecules and since we have 15 O_2 molecules, we know that C_3H_8 is the limiting reactant and that O_2 is in excess.

Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Reactant Masses

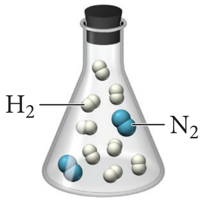
When working in the laboratory, we normally measure the initial quantities of reactants in grams, not in number of molecules. To find the limiting reactant and theoretical yield from initial masses, we must first convert the masses to amounts in moles. Consider the following reaction:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

A reaction mixture contains 42.5 g Mg and 33.8 g O_2; what is the limiting reactant and theoretical yield? To solve this problem, we must determine which of the given amounts of reactants makes the least amount of product.

CONCEPTUAL CONNECTION 4.4

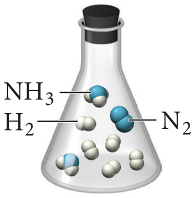
Limiting Reactant and Theoretical Yield



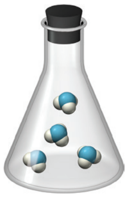
Nitrogen and hydrogen gas react to form ammonia according to the following reaction:

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

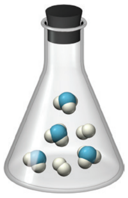
If a flask contains a mixture of reactants represented by the diagram on the right ▸, which diagram best represents the mixture in the flask after the reactants have reacted as completely as possible? What is the limiting reactant? Which reactant is in excess?



(a)

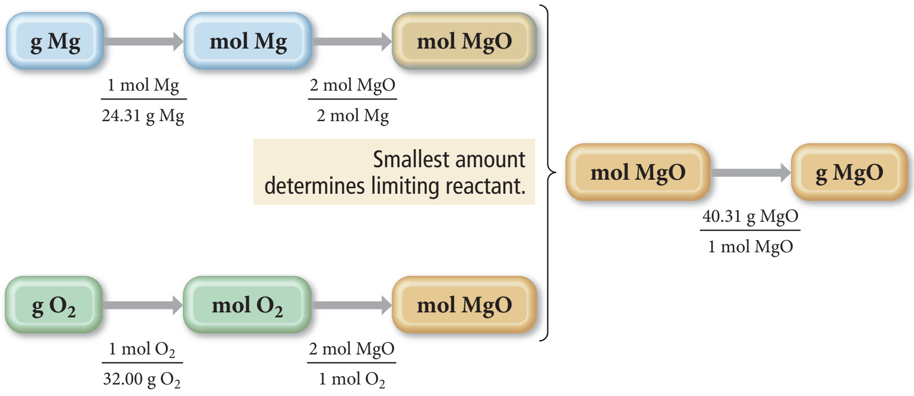


(b)



(c)

Conceptual Plan We can find the limiting reactant by calculating how much product can be made from each reactant. However, since we are given the initial quantities in grams, and stoichiometric relationships are between moles, we must first convert grams to moles. We then convert from moles of the reactant to moles of the product. The reactant that makes the *least amount of product* is the limiting reactant. The conceptual plan is as follows:



In the plan, we compare the number of moles of MgO made by each reactant and convert only the smaller amount to grams.

Relationships Used

molar mass Mg $= 24.31$ g mol^{-1}
molar mass $O_2 = 32.00$ g mol^{-1}
2 mol Mg : 2 mol MgO
1 mol O_2 : 2 mol MgO
molar mass MgO $= 40.31$ g mol^{-1}

Solution

Beginning with the masses of each reactant, we follow the conceptual plan to calculate how much product can be made from each:

$$\boxed{42.5\ \cancel{\text{g Mg}}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{2\ \text{mol MgO}}{2\ \cancel{\text{mol Mg}}} = \boxed{1.7\underline{4}83\ \text{mol MgO}}$$

Limiting reactant

Least amount of product

$$33.8\ \cancel{\text{g } O_2} \times \frac{1\ \cancel{\text{mol } O_2}}{32.00\ \cancel{\text{g } O_2}} \times \frac{2\ \text{mol MgO}}{1\ \cancel{\text{mol } O_2}} = 2.1\underline{1}25\ \text{mol MgO}$$

$$1.7\underline{4}83\ \cancel{\text{mol MgO}} \times \frac{40.31\ \text{g MgO}}{1\ \cancel{\text{mol MgO}}} = 70.5\ \text{g MgO}$$

Since Mg makes the least amount of product, it is the limiting reactant, and O_2 is in excess. Notice that the limiting reactant is not necessarily the reactant with the least mass. In this case, the mass of O_2 is less than the mass of Mg, yet Mg is the limiting reactant because it makes the least amount of MgO. The theoretical yield is 70.5 g of MgO, the mass of product possible based on the moles of limiting reactant.

Suppose that after the reaction has occurred, we determine a measured yield or an **actual yield** of MgO is 55.9 g. What is the **percent yield**? We compute the percent yield as follows:

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{55.9\ \cancel{\text{g}}}{70.5\ \cancel{\text{g}}} \times 100\% = 79.3\%$$

EXAMPLE 4.13 LIMITING REACTANT AND THEORETICAL YIELD

Ammonia, NH_3, can be synthesized by the following reaction:

$$2\ NO(g) + 5\ H_2(g) \longrightarrow 2\ NH_3(g) + 2\ H_2O(g)$$

Starting with 86.3 g NO and 25.6 g H_2, find the theoretical yield of ammonia in grams.

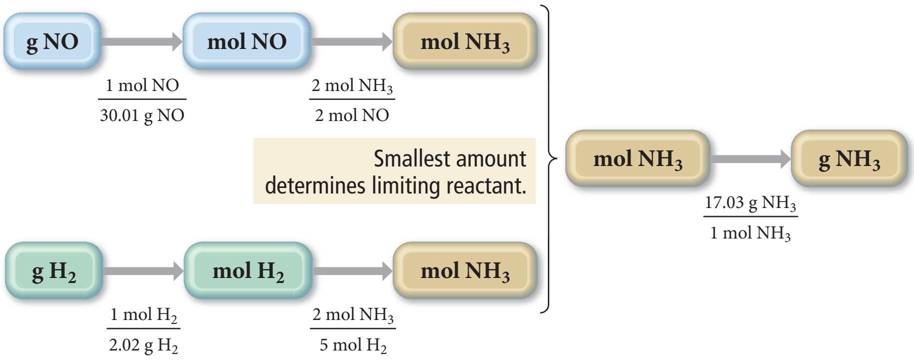
SORT You are given the mass of each reactant in grams and asked to find the theoretical yield of a product.

GIVEN: 86.3 g NO, 25.6 g H_2

FIND: theoretical yield of $NH_3(g)$

STRATEGIZE Determine which reactant makes the least amount of product by converting from grams of each reactant to moles of the reactant to moles of the product. Use molar masses to convert between grams and moles and use the stoichiometric relationships (deduced from the chemical equation) to convert between moles of reactant and moles of product. Remember that the reactant that makes *the least amount of product* is the limiting reactant. Convert the number of moles of product obtained using the limiting reactant to grams of product.

CONCEPTUAL PLAN



RELATIONSHIPS USED

molar mass NO $= 30.01\ \text{g mol}^{-1}$

molar mass $H_2 = 2.02\ \text{g mol}^{-1}$

2 mol NO : 2 mol NH_3 (from chemical equation)

5 mol H_2 : 2 mol NH_3 (from chemical equation)

molar mass $NH_3 = 17.03\ \text{g mol}^{-1}$

SOLVE Beginning with the given mass of each reactant, calculate the amount of product that can be made in moles. Convert the amount of product made by the limiting reactant to grams—this is the theoretical yield.

SOLUTION

$$\boxed{86.3\ \cancel{\text{g NO}}} \times \frac{1\ \cancel{\text{mol NO}}}{30.01\ \cancel{\text{g NO}}} \times \frac{2\ \text{mol NH}_3}{2\ \cancel{\text{mol NO}}} = \boxed{2.8\underline{7}57\ \text{mol NH}_3}$$

Limiting reactant

Least amount of product

$$25.6\ \cancel{\text{g H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{2.016\ \cancel{\text{g H}_2}} \times \frac{2\ \text{mol NH}_3}{5\ \cancel{\text{mol H}_2}} = 5.0\underline{6}93\ \text{mol NH}_3$$

$$2.8\underline{7}57\ \cancel{\text{mol NH}_3} \times \frac{17.03\ \text{g NH}_3}{1\ \cancel{\text{mol NH}_3}} = 49.0\ \text{g NH}_3$$

Since NO makes the least amount of product, it is the limiting reactant, and the theoretical yield of ammonia is 49.0 g.

CHECK The units of the answer (g NH_3) are correct. The magnitude (49.0 g) seems reasonable given that 86.3 g NO is the limiting reactant. NO contains one oxygen atom per nitrogen atom and NH_3 contains three hydrogen atoms per nitrogen atom. Since three hydrogen atoms have less mass than one oxygen atom, it is reasonable that the mass of NH_3 obtained is less than the initial mass of NO.

FOR PRACTICE 4.13

Ammonia can also be synthesized by the following reaction:

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

What is the theoretical yield of ammonia, in kg, that we can synthesize from 5.22 kg of H_2 and 31.5 kg of N_2?

EXAMPLE 4.14 LIMITING REACTANT, THEORETICAL YIELD, AND PERCENT YIELD

Titanium metal can be obtained from its oxide according to the following balanced equation:

$$TiO_2(s) + 2\ C(s) \longrightarrow Ti(s) + 2\ CO(g)$$

When 28.6 kg of C reacts with 88.2 kg of TiO_2, 42.8 kg of Ti is produced. Find the limiting reactant, theoretical yield (in kg), and percent yield.

SORT You are given the mass of each reactant and the mass of product formed. You are asked to find the limiting reactant, theoretical yield, and percent yield.

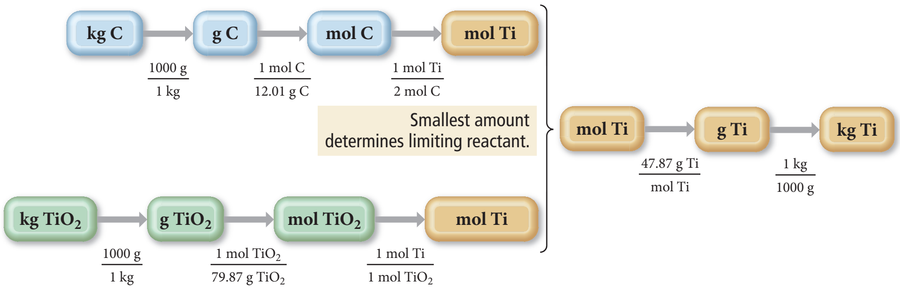
GIVEN: 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti produced

FIND: limiting reactant, theoretical yield, percent yield

STRATEGIZE Determine which of the reactants makes the least amount of product by converting from kilograms of each reactant to moles of product. Convert between grams and moles using molar mass. Convert between moles of reactant and moles of product using the stoichiometric relationships derived from the chemical equation. Remember that the reactant that makes the *least amount of product* is the limiting reactant.

Determine the theoretical yield (in kg) by converting the number of moles of product obtained with the limiting reactant to kilograms of product.

CONCEPTUAL PLAN



RELATIONSHIPS USED

$1000\ \text{g} = 1\ \text{kg}$

$\text{molar mass of C} = 12.01\ \text{g mol}^{-1}$

$\text{molar mass of TiO}_2 = 79.87\ \text{g mol}^{-1}$

$1\ \text{mol TiO}_2 : 1\ \text{mol Ti}$

$2\ \text{mol C} : 1\ \text{mol Ti}$

$\text{molar mass of Ti} = 47.87\ \text{g mol}^{-1}$

(continued)

EXAMPLE 4.14 **(CONTINUED)**

SOLVE Beginning with the actual amount of each reactant, calculate the amount of product that can be made in moles. Convert the amount of product made by the limiting reactant to kilograms—this is the theoretical yield.

SOLUTION

$$\mathbf{28.6\,\cancel{kg\ C}} \times \frac{1000\,\cancel{g}}{1\,\cancel{kg}} \times \frac{1\,\cancel{mol\ C}}{12.01\,\cancel{g\ C}} \times \frac{1\ \text{mol Ti}}{2\,\cancel{mol\ C}} = \mathbf{1.1\underline{9}07 \times 10^{3}\ mol\ Ti}$$

Limiting reactant

Least amount of product

$$\mathbf{88.2\,\cancel{kg\ TiO_2}} \times \frac{1000\,\cancel{g}}{1\,\cancel{kg}} \times \frac{1\,\cancel{mol\ TiO_2}}{79.87\,\cancel{g\ TiO_2}} \times \frac{1\ \text{mol Ti}}{1\,\cancel{mol\ TiO_2}} = \mathbf{1.1\underline{0}43 \times 10^{3}\ mol\ Ti}$$

$$\mathbf{1.1\underline{0}43 \times 10^{3}\,\cancel{mol\ Ti}} \times \frac{47.87\,\cancel{g\ Ti}}{1\,\cancel{mol\ Ti}} \times \frac{1\ \text{kg}}{1000\,\cancel{g}} = \mathbf{52.9\ kg\ Ti}$$

Since TiO_2 makes the least amount of product, it is the limiting reactant, and 52.9 kg Ti is the theoretical yield.

Calculate the percent yield by dividing the actual yield (42.8 kg Ti) by the theoretical yield.

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{42.8\,\cancel{kg}}{52.9\,\cancel{kg}} \times 100\% = 80.9\%$$

CHECK The theoretical yield has the correct units (kg Ti) and has a reasonable magnitude compared to the mass of TiO_2. Since Ti has a lower molar mass than TiO_2, the amount of Ti made from TiO_2 should have a lower mass. The percent yield is reasonable (under 100%, as it should be).

FOR PRACTICE 4.14

Mining companies use the following reaction to obtain iron from iron ore:

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

The reaction of 167 g Fe_2O_3 with 85.8 g CO produces 72.3 g Fe. Find the limiting reactant, theoretical yield, and percent yield.

CONCEPTUAL CONNECTION 4.5

Reactant in Excess

Nitrogen dioxide reacts with water to form nitric acid and nitrogen monoxide according to the following equation:

$$3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(l) + NO(g)$$

Suppose that 5 mol NO_2 and 1 mol H_2O combine and react completely. How many moles of the reactant in excess are present after the reaction has completed?

4.9 Solution Concentration and Solution Stoichiometry

In this section, we first examine how to quantify the concentration of a solution (the amount of solute relative to solvent) and then turn to applying the principles of stoichiometry, which we learned in the previous section, to reactions occurring in solution.

Solution Concentration

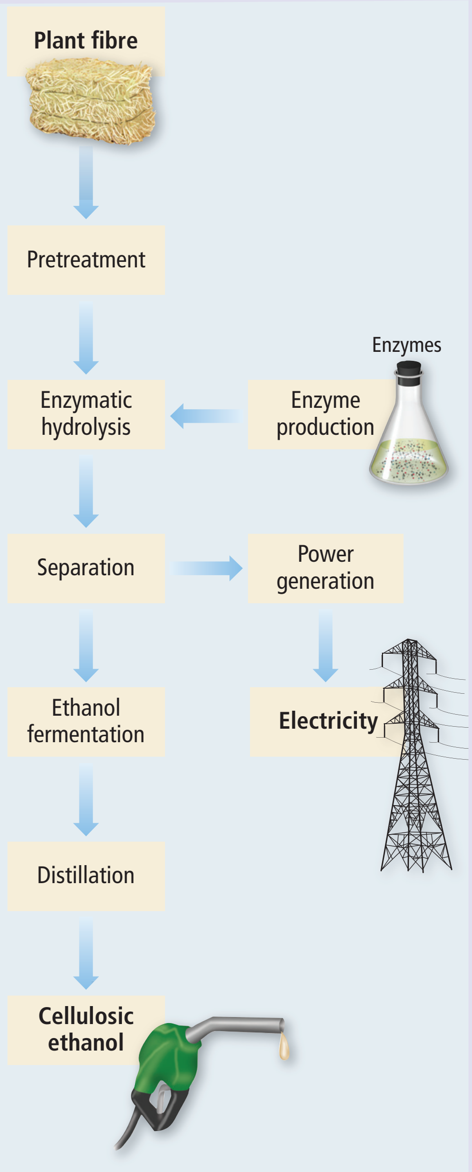
A common way to express solution concentration is **molarity** (*M*, in mol L^{-1}), the amount of solute (in moles) divided by the volume of solution (in litres):

$$\text{molarity (in mol L}^{-1}) = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

$$\text{or} \quad M = \frac{n}{V}$$

The abbreviation or symbol for molarity is *M* and the units of molarity are mol L^{-1}. You may see the units of molarity written as M. You may also see the concentration of a compound expressed as [X], where the square brackets denote the concentration of X in mol L^{-1}.

CHEMISTRY IN YOUR DAY Blended Ethanol Gasoline



We have discussed the balanced chemical equation for the combustion of octane, a component of gasoline, as follows:

$$C_8H_{18}(g) + \tfrac{25}{2}\,O_2(g) \longrightarrow 8\,CO_2(g) + 9\,H_2O(l)$$

The equation shows that 12.5 mol of O_2 are required to completely react with a mole of C_8H_{18}. What if there is not enough O_2 in an automobile cylinder to fully react with the amount of octane that is present? For many reactions, a shortage of one reactant means that less product forms—oxygen would become the limiting reactant. However, for some reactions, a shortage of one reactant causes side reactions to occur along with the desired reaction. In the case of burning octane in an automobile, those side reactions produce pollutants such as carbon monoxide and soot or carbon. (Have you ever seen black smoke coming out of a car's tailpipe?) These side reactions are called *incomplete combustion*:

$$C_8H_{18}(g) + \tfrac{17}{2}\,O_2(g) \longrightarrow 8\,CO(g) + 9\,H_2O(l)$$

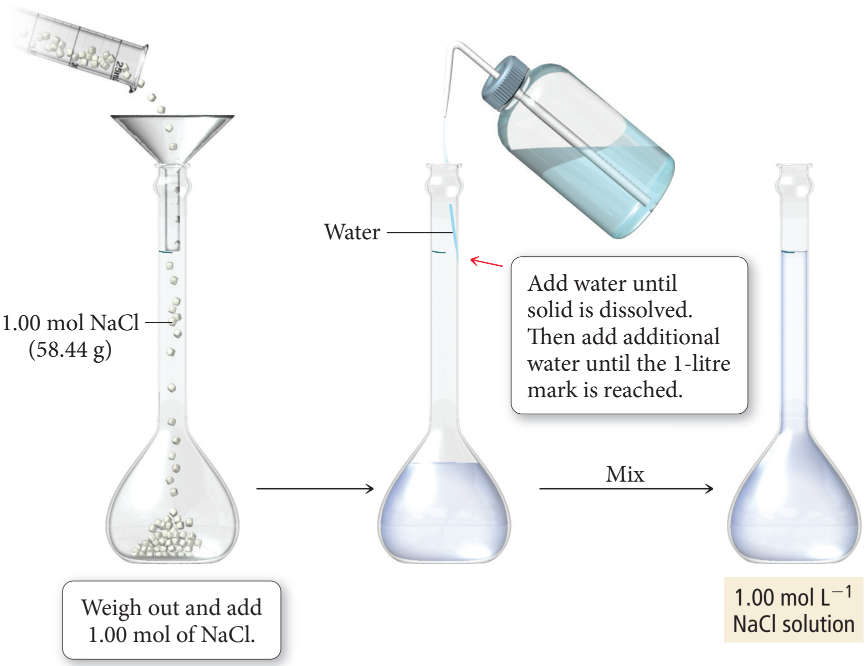
$$C_8H_{18}(g) + \tfrac{9}{2}\,O_2(g) \longrightarrow 8\,C(s) + 9\,H_2O(l)$$

Blended ethanol gasoline is a fuel where a small amount of ethanol, typically 10% by volume, is added to the gasoline. Ethanol has the molecular formula C_2H_5OH and is a partially oxidized fuel—it already contains oxygen. Blended ethanol gasoline burns cleaner—produces less emissions—as it doesn't require as much oxygen to fully oxidize to CO_2. There are other benefits of blended ethanol gasoline. Since ethanol dissolves water that might collect in the fuel tank and hoses leading to the engine, ethanol can pull it through the engine before it gums up your tank and fuel line. Ethanol is also a renewable fuel. Depending on the source of ethanol, it reduces CO_2 emissions. Ethanol from corn reduces CO_2 emissions by only about 13% but that from cellulose-based plants (sugar cane and switchgrass) can reduce CO_2 emissions by 88% compared to using fossil fuels. There are some cons associated with blended ethanol gasoline as well. Since it is partially oxidized, there is less energy content in ethanol than in a hydrocarbon such as octane. It is estimated that a 10% ethanol blend contains about 97% of the energy of gasoline, but some of the energy loss is offset by the increased combustion efficiency of the blend. In order to burn blended fuels with greater than 10% ethanol, costly modifications to the engine are required.

Finally, one of the main cons of using ethanol as fuel is that producing ethanol from corn raises the price of corn for consumption. Furthermore, if farmers decide that it is more financially beneficial to grow corn for fuel, costs of other foods will also increase. Fortunately, research into producing ethanol from cellulosic sources, such as wood and nonedible by-products from plants, is making nonfood sources of ethanol more feasible. Cellulosic ethanol was first sold to the general public in Ottawa, Ontario, from Iogen Corporation's demonstration plant on Hunt Club Road in Ottawa. The demonstration plant uses straw as the feedstock for producing ethanol. Other feedstocks for cellulosic ethanol include corn stalks and leaves (stover) left over after corn harvesting, and the fibrous material (bagasse) remaining after sugar has been extracted from sugar cane.

Notice that molarity is a ratio of the amount of solute per litre of *solution,* not per litre of solvent. To make an aqueous solution of a specified molarity, we usually put the solute into a flask and then add water to reach the desired volume of solution. For example, to make 1 L of a 1 mol L^{-1} NaCl solution, we add 1 mol of NaCl to a flask and then add enough water to make 1 L of solution (Figure 4.13 ▼). We *do not* combine 1 mol of NaCl with 1 L of water because the resulting solution would have a total volume exceeding 1 L and therefore a molarity of less than 1 mol L^{-1}. To calculate molarity, divide the amount of the solute in moles by the volume of the solution (solute *and* solvent) in litres, as shown in the following example.

FIGURE 4.13 **Preparing a One-Molar NaCl Solution**



EXAMPLE 4.15 CALCULATING SOLUTION CONCENTRATION

If 25.5 g KBr is dissolved in enough water to make 1.75 L of solution, what is the molarity of the solution?

SORT You are given the mass of KBr and the volume of a solution and asked to find its molarity.

GIVEN: 25.5 g KBr, 1.75 L of solution

FIND: molarity

STRATEGIZE When formulating the conceptual plan, think about the definition of molarity, the amount of solute *in moles* per litre of solution.

You are given the mass of KBr, so first use the molar mass of KBr to convert from g KBr to mol KBr.

Then use the number of moles of KBr and litres of solution to find the molarity.

CONCEPTUAL PLAN

$$\begin{tikzcd} \text{g KBr} \arrow[r, "\frac{1\ \text{mol}}{119.00\ \text{g}}"'] & \text{mol KBr} \end{tikzcd}$$

$$\begin{tikzcd} \text{mol KBr, L solution} \arrow[r, "\text{Molarity} = \frac{\text{amount of solute}}{\text{volume of solution}}"'] & \text{Molarity} \end{tikzcd}$$

RELATIONSHIP USED

molar mass of KBr = 119.00 g mol^{-1}

SOLVE Follow the conceptual plan. Begin with g KBr and convert to mol KBr, then use mol KBr and L solution to compute molarity.

SOLUTION

$$25.5\ \cancel{\text{g KBr}} \times \frac{1\ \text{mol KBr}}{119.00\ \cancel{\text{g KBr}}} = 0.21\underline{4}29\ \text{mol KBr}$$

$$\text{molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

$$= \frac{0.21\underline{4}29\ \text{mol}}{1.75\ \text{L}}$$

$$= 0.122\ \text{mol L}^{-1}$$

CHECK The unit of the answer (mol L^{-1}) is correct. The magnitude is reasonable since common solutions range in concentration from 0 to about 18 mol L^{-1}. Concentrations significantly above 18 mol L^{-1} are suspect and should be double-checked.

FOR PRACTICE 4.15

Calculate the molarity of a solution made by adding 45.4 g of $NaNO_3$ to a flask and dissolving it with water to create a total volume of 2.50 L.

FOR MORE PRACTICE 4.15

What mass of KBr (in grams) do you need to make 250.0 mL of a 1.50 mol L^{-1} KBr solution?

Using Molarity in Calculations

We can use the molarity of a solution as a conversion factor between moles of the solute and litres of the solution. For example, a 0.500 mol L^{-1} NaCl solution contains 0.500 mol NaCl for every litre of solution:

$$\frac{0.500\text{ mol NaCl}}{1\text{ L solution}} \quad \textit{converts} \quad \boxed{\text{L solution}} \longrightarrow \boxed{\text{mol NaCl}}$$

This conversion factor converts from L solution to mol NaCl. If we want to go the other way, we invert the conversion factor:

$$\frac{1\text{ L solution}}{0.500\text{ mol NaCl}} \quad \textit{converts} \quad \boxed{\text{mol NaCl}} \longrightarrow \boxed{\text{L solution}}$$

The following example shows how to use molarity in this way.

EXAMPLE 4.16 USING MOLARITY IN CALCULATIONS

How many litres of a 0.125 mol L^{-1} NaOH solution contains 0.255 mol of NaOH?

SORT You are given the concentration of a NaOH solution. You are asked to find the volume of the solution that contains a given amount (in moles) of NaOH.

GIVEN: 0.125 mol L^{-1} NaOH solution, 0.255 mol NaOH

FIND: volume of NaOH solution (in L)

STRATEGIZE The conceptual plan begins with mol NaOH and shows the conversion to L of solution using molarity as a conversion factor.

CONCEPTUAL PLAN

$$\boxed{\text{mol NaOH}} \xrightarrow[\frac{1\text{ L solution}}{0.125\text{ mol NaOH}}]{} \boxed{\text{L solution}}$$

RELATIONSHIP USED

$$0.125\text{ mol L}^{-1}\text{ NaOH} = \frac{0.125\text{ mol NaOH}}{1\text{ L solution}}$$

SOLVE Follow the conceptual plan. Begin with mol NaOH and convert to L solution.

SOLUTION

$$0.255\,\cancel{\text{mol NaOH}} \times \frac{1\text{ L solution}}{0.125\,\cancel{\text{mol NaOH}}} = 2.04\text{ L solution}$$

CHECK The units of the answer (L) are correct. The magnitude seems reasonable because the solution contains 0.125 mol per litre. Therefore, roughly 2 L contains the given amount of moles (0.255 mol).

FOR PRACTICE 4.16

How many grams of sucrose $(C_{12}H_{22}O_{11})$ are in 1.55 L of 0.758 mol L^{-1} sucrose solution?

FOR MORE PRACTICE 4.16

How many mL of a 0.155 mol L^{-1} KCl solution contains 2.55 g KCl?

CONCEPTUAL CONNECTION 4.6

Solutions

If we dissolve 25 g of salt in 251 g of water, what is the mass of the resulting solution?

(a) 251 g **(b)** 276 g **(c)** 226 g

Solution Dilution To save space while transporting and in stockrooms, it is convenient to have solutions in concentrated forms called **stock solutions**. For example, hydrochloric acid (HCl) is typically stored as a 12 mol L^{-1} stock solution. However, many lab procedures call for much less concentrated HCl solutions, so we must dilute the stock solution to the required concentration. The easiest way to solve dilution problems is to use the following dilution equation:

You may also see this equation written as $c_1V_1 = c_2V_2$; molarity (M) is just a specific case of concentration (c).

$$M_1V_1 = M_2V_2 \qquad [4.2]$$

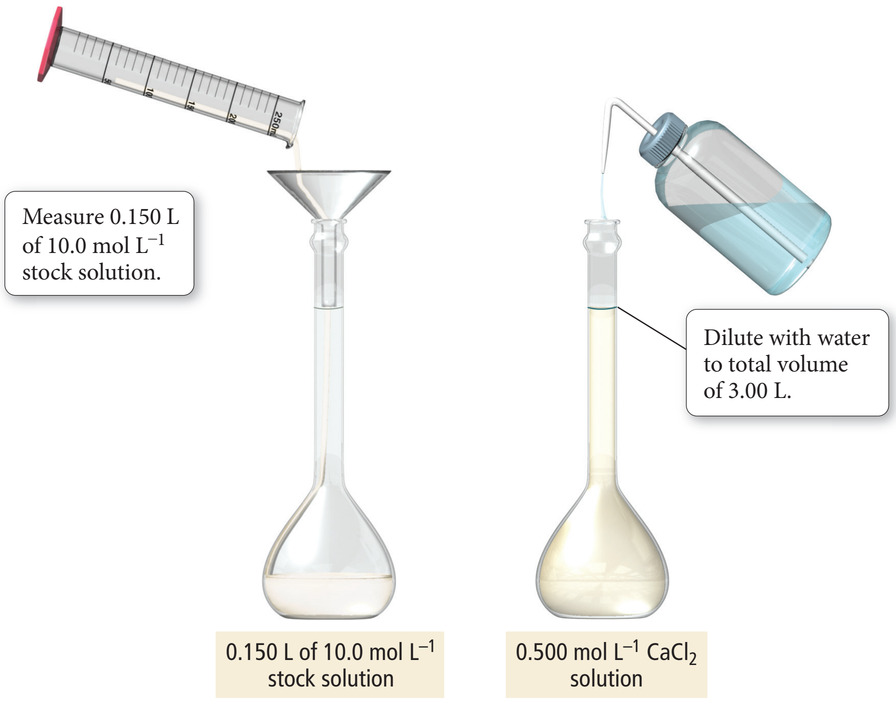
where M_1 and V_1 are the molarity and volume of the initial **concentrated solution**, and M_2 and V_2 are the molarity and volume of the final **diluted solution**. This equation works because the molarity multiplied by the volume gives the number of moles of solute, which is the same for both solutions. In other words, the number of moles of solute does not change when we dilute a solution.

For example, suppose a laboratory procedure calls for 3.00 L of a 0.500 mol L^{-1} solution of $CaCl_2$. How should we prepare this solution from a 10.0 mol L^{-1} $CaCl_2$ stock solution? We solve Equation 4.2 for V_1, the volume of the stock solution required for the dilution, and then substitute in the correct values to compute it:

$$\begin{aligned} M_1V_1 &= M_2V_2 \\ V_1 &= \frac{M_2V_2}{M_1} \\ &= \frac{0.500\ \cancel{\text{mol L}^{-1}} \times 3.00\ \text{L}}{10.0\ \cancel{\text{mol L}^{-1}}} \\ &= 0.150\ \text{L} \end{aligned}$$

We make the solution by diluting 0.150 L of the stock solution to a total volume of 3.00 L (V_2). The resulting solution will be 0.500 mol L^{-1} of $CaCl_2$ (Figure 4.14 ▼). It makes sense that we requre a small volume of the more concentrated solution to make a large volume of a more dilute solution.

▶ **FIGURE 4.14 Preparing 3.00 L of 0.500 mol L^{-1} $CaCl_2$ from a 10.0 mol L^{-1} Stock Solution**



$$M_1V_1 = M_2V_2$$

$$10.0\ \text{mol}\ \cancel{\text{L}^{-1}} \times 0.150\ \cancel{\text{L}} = 0.500\ \text{mol}\ \cancel{\text{L}^{-1}} \times 3.00\ \cancel{\text{L}}$$

$$1.50\ \text{mol} = 1.50\ \text{mol}$$

EXAMPLE 4.17 SOLUTION DILUTION

To what volume should you dilute 0.200 L of a 15.0 mol L^{-1} NaOH solution to obtain a 3.00 mol L^{-1} NaOH solution?

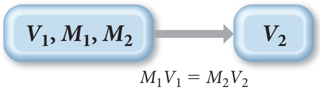
SORT You are given the initial volume, initial concentration, and final concentration of a solution, and you need to find the final volume.

GIVEN: $V_1 = 0.200\ \text{L}$
$M_1 = 15.0\ \text{mol L}^{-1}$
$M_2 = 3.00\ \text{mol L}^{-1}$

FIND: V_2

STRATEGIZE Equation 4.2 relates the initial and final volumes and concentrations for solution dilution problems. You are asked to find V_2. The other quantities (V_1, M_1, and M_2) are all given in the problem.

CONCEPTUAL PLAN



RELATIONSHIP USED

$M_1V_1 = M_2V_2$

SOLVE Begin with the solution dilution equation and solve it for V_2.

Substitute in the required quantities and compute V_2.

Make the solution by diluting 0.200 L of the stock solution to a total volume of 1.00 L (V_2). The resulting solution will have a concentration of 3.00 mol L^{-1}.

SOLUTION

$$M_1V_1 = M_2V_2$$
$$V_2 = \frac{M_1V_1}{M_2}$$
$$= \frac{15.0\ \cancel{\text{mol L}^{-1}} \times 0.200\ \text{L}}{3.00\ \cancel{\text{mol L}^{-1}}}$$
$$= 1.00\ \text{L}$$

CHECK The final unit (L) is correct. The magnitude of the answer is reasonable because the solution is diluted from 15.0 mol L^{-1} to 3.00 mol L^{-1}, a factor of five. Therefore, the volume should increase by a factor of five.

FOR PRACTICE 4.17

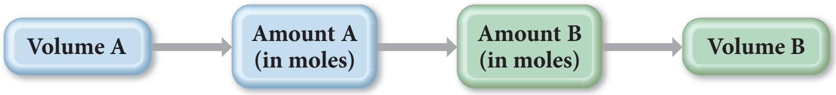
To what volume (in mL) should you dilute 100.0 mL of a 5.00 mol L^{-1} $CaCl_2$ solution to obtain a 0.750 mol L^{-1} $CaCl_2$ solution?

FOR MORE PRACTICE 4.17

What volume of a 6.00 mol L^{-1} $NaNO_3$ solution should you use to make 0.525 L of a 1.20 mol L^{-1} $NaNO_3$ solution?

Solution Stoichiometry

In Section 4.7, we learned how the coefficients in chemical equations are used as conversion factors between the amounts of reactants (in moles) and the amounts of products (in moles). In aqueous reactions, quantities of reactants and products are often specified in terms of volumes and concentrations. We can use the volume and concentration of a reactant or product to calculate its amount in moles. We can then use the stoichiometric coefficients in the chemical equation to convert to the amount of another reactant or product in moles. The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product:



We make the conversions between solution volumes and amounts of solute in moles using the molarities of the solutions. We make the conversions between amounts in moles of A and B using the stoichiometric coefficients from the balanced chemical equation. The following example demonstrates solution stoichiometry.

EXAMPLE 4.18 SOLUTION STOICHIOMETRY

What volume of a 0.150 mol L^{-1} KCl solution will completely react with 0.150 L of a 0.175 mol L^{-1} $Pb(NO_3)_2$ solution according to the following balanced chemical equation?

$$2\ \mathrm{KCl}(aq) + \mathrm{Pb(NO_3)_2}(aq) \longrightarrow \mathrm{PbCl_2}(s) + 2\ \mathrm{KNO_3}(aq)$$

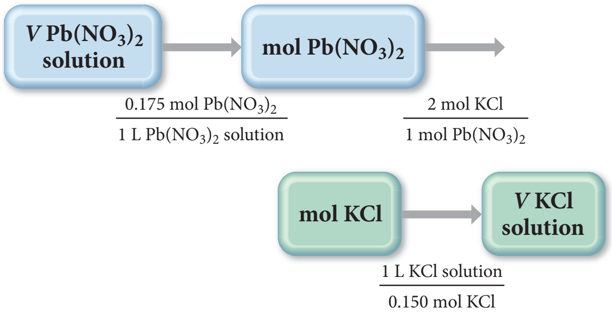
SORT You are given the volume and concentration of a $Pb(NO_3)_2$ solution. You are asked to find the volume of KCl solution (of a given concentration) required to react with it.

GIVEN: 0.150 L of $Pb(NO_3)_2$ solution, 0.175 mol L^{-1} solution, 0.150 mol L^{-1} KCl solution

FIND: volume KCl solution (in L)

STRATEGIZE The conceptual plan has the following form: volume A ⟶ amount A (in moles) ⟶ amount B (in moles) ⟶ volume B. The molar concentrations of the KCl and $Pb(NO_3)_2$ solutions can be used as conversion factors between the number of moles of reactants in these solutions and their volumes. The stoichiometric coefficients from the balanced equation are used to convert between number of moles of $Pb(NO_3)_2$ and number of moles of KCl.

CONCEPTUAL PLAN



RELATIONSHIPS USED

$$[\mathrm{Pb(NO_3)_2}] = \frac{0.175\ \mathrm{mol\ Pb(NO_3)_2}}{1\ \mathrm{L\ Pb(NO_3)_2\ solution}}$$

2 mol KCl : 1 mol $Pb(NO_3)_2$

$$[\mathrm{KCl}] = \frac{0.150\ \mathrm{mol\ KCl}}{1\ \mathrm{L\ KCl\ solution}}$$

Don't forget that [X] is a shorthand way of writing the concentration of X in units of mol L^{-1}.

SOLVE Begin with L $Pb(NO_3)_2$ solution and follow the conceptual plan to arrive at the volume of KCl solution.

SOLUTION

$$0.150\ \mathrm{L}\ \cancel{\mathrm{Pb(NO_3)_2\ solution}} \times \frac{0.175\ \cancel{\mathrm{mol\ Pb(NO_3)_2}}}{1\ \cancel{\mathrm{L\ Pb(NO_3)_2\ solution}}} \times \frac{2\ \cancel{\mathrm{mol\ KCl}}}{1\ \cancel{\mathrm{mol\ Pb(NO_3)_2}}} \times \frac{1\ \mathrm{L\ KCl\ solution}}{0.150\ \cancel{\mathrm{mol\ KCl}}} = 0.350\ \mathrm{L\ KCl\ solution}$$

CHECK The final units (L KCl solution) are correct. The magnitude (0.350 L) seems reasonable because the reaction stoichiometry requires 2 mol of KCl per mole of $Pb(NO_3)_2$. Since the concentrations of the two solutions are not very different (0.150 mol L^{-1} compared to 0.175 mol L^{-1}), the volume of KCl required should be roughly two times the 0.150 L of $Pb(NO_3)_2$ given in the problem.

FOR PRACTICE 4.18

What volume (in mL) of a 0.150 mol L^{-1} HNO_3 solution will completely react with 35.7 mL of a 0.108 mol L^{-1} Na_2CO_3 solution according to the following balanced chemical equation?

$$\mathrm{Na_2CO_3}(aq) + 2\ \mathrm{HNO_3}(aq) \longrightarrow 2\ \mathrm{NaNO_3}(aq) + \mathrm{CO_2}(g) + \mathrm{H_2O}(l)$$

FOR MORE PRACTICE 4.18

In the reaction above, what mass (in grams) of carbon dioxide forms?

CHAPTER IN REVIEW

Key Terms

Section 4.2

chemical reaction (102)
combustion reaction (102)
chemical equation (102)
reactants (103)
products (103)
balanced chemical equation (103)

Section 4.3

solution (105)
solvent (105)
solute (105)
aqueous solution (105)
electrolyte (106)
strong electrolyte (106)
nonelectrolyte (106)
strong acid (106)
weak acid (107)
weak electrolyte (107)
soluble (107)
insoluble (107)

Section 4.4

precipitation reaction (109)
precipitate (109)
molecular equation (109)
complete ionic equation (109)
spectator ion (110)
net ionic equation (110)

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acid–base reaction (111)
neutralization reaction (111)
Arrhenius definitions (111)
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polyprotic acid (112)
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oxidation–reduction reaction (115)
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oxidation (116)
reduction (116)
oxidation state (117)
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Section 4.8

limiting reactant (128)
theoretical yield (128)
actual yield (130)
percent yield (130)

Section 4.9

molarity (*M*) (132)
stock solution (136)
concentrated solution (136)
diluted solution (136)

Key Concepts

Writing and Balancing Chemical Equations (4.2)

In chemistry, we represent chemical reactions with chemical equations. The substances on the left-hand side of a chemical equation are called the reactants, and the substances on the right-hand side are called the products. Chemical equations are balanced when the number of each type of atom on the left side of the equation is equal to the number on the right side.

Aqueous Solutions and Solubility (4.3)

An aqueous solution is a homogeneous mixture of water (the solvent) with another substance (the solute). Solutes that completely dissociate (or completely ionize in the case of acids) to ions in solution are strong electrolytes and are good conductors of electricity. Solutes that only partially dissociate (or partially ionize) are weak electrolytes, and solutes that do not dissociate (or ionize) at all are nonelectrolytes. A substance that dissolves in water to form a solution is soluble. The solubility rules are an empirical set of guidelines that help predict the solubilities of ionic compounds; these rules are especially useful in Section 4.4 when determining whether or not a precipitate will form.

Precipitation Reactions (4.4)

In a precipitation reaction, we mix two aqueous solutions and a solid—or precipitate—forms. Aqueous reactions, such as precipitation reactions, can be represented with a molecular equation, which shows the complete neutral formula for each compound in the reaction. Alternatively, a reaction can be represented with a complete ionic equation, which shows the dissociated nature of the aqueous ionic compounds. Finally, a third representation is the net ionic equation, in which the spectator ions—those that do not change in the course of the reaction—are left out of the equation.

Acid–Base Reactions (4.5)

An acid is a substance which produces H^+ in solution and a base is a substance which produces OH^- in solution. In an acid–base reaction, the acid and base neutralize each other, producing water (or in some cases, a weak electrolyte). In gas-evolution reactions, the acid and base combine in solution and a gas is produced.

Oxidation–Reduction Reactions (4.6)

In oxidation–reduction reactions, one substance transfers electrons to another substance. The substance that loses electrons is oxidized, and the substance that gains them is reduced. An oxidation state is a fictitious charge given to each atom in a redox reaction by assigning all shared electrons to the atom with the greater attraction for those electrons. Oxidation states are an imposed electronic bookkeeping scheme, not an actual physical state. The oxidation state of an atom increases upon oxidation and decreases upon reduction. A combustion reaction is a specific type of oxidation–reduction reaction in which a substance reacts with oxygen—emitting heat and forming one or more oxygen-containing products.

Reaction Stoichiometry (4.7)

Reaction stoichiometry refers to the numerical relationships between the reactants and products in a balanced chemical equation. Reaction stoichiometry allows us to predict, for example, the amount of product that can be formed for a given amount of reactant, or how much of one reactant is required to react with a given amount of another.

Limiting Reactant, Theoretical Yield, and Percent Yield (4.8)

When a chemical reaction actually occurs, the reactants are usually not present in the exact stoichiometric ratios specified by the balanced chemical equation. The limiting reactant is the one that is available in the smallest stoichiometric quantity—it will be completely consumed in the reaction and it limits the amount of product that can be made. Any reactant that does not limit the amount of product is said to be in excess. The amount of product that can be made from the limiting reactant is the theoretical yield. The actual yield—always equal to or less than the theoretical yield—is the amount of product that is actually made when the reaction is carried out. The percentage of the theoretical yield that is actually produced is the percent yield.

Solution Concentration and Solution Stoichiometry (4.9)

We often express the concentration of a solution in molarity, the number of moles of solute per litre of solution. We can use the molarities and volumes of reactant solutions to predict the amount of product that will form in an aqueous reaction.

Key Equations and Relationships

Mass-to-Mass Conversion: Stoichiometry (4.7)

$$\text{mass A} \longrightarrow \text{amount A (in moles)} \longrightarrow \text{amount B (in moles)} \longrightarrow \text{mass B}$$

Percent Yield (4.8)

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

Molarity (4.9)

$$\text{molarity} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}} \quad \text{or} \quad M = \frac{n}{V}$$

Solution Dilution (4.9)

$$M_1V_1 = M_2V_2$$

Solution Stoichiometry (4.9)

$$\text{volume A} \longrightarrow \text{amount A (in moles)} \longrightarrow \text{amount B (in moles)} \longrightarrow \text{volume B}$$

Key Skills

Balancing Chemical Equations (4.2)

• Examples 4.1, 4.2 • For Practice 4.1, 4.2 • Exercises 25–34, 59, 60

Predicting Whether a Compound Is Soluble (4.3)

• Example 4.3 • For Practice 4.3 • Exercises 35–38

Writing Equations for Precipitation Reactions (4.4)

• Example 4.4 • For Practice 4.4 • For More Practice 4.4 • Exercises 39–42

Writing Equations for Acid–Base Reactions (4.5)

• Example 4.5 • For Practice 4.5 • Exercises 49–50

Writing Equations for Gas-Evolution Reactions (4.5)

• Example 4.6 • For Practice 4.6 • For More Practice 4.6 • Exercises 51–52

Assigning Oxidation States (4.6)

• Example 4.7 • For Practice 4.7 • Exercises 53–56

Identifying Redox Reactions, Oxidizing Agents, and Reducing Agents, and Using Oxidation States (4.6)

• Example 4.8 • For Practice 4.8 • Exercises 57–58

Balancing Redox Reactions in Acidic or Basic Solution; Balancing Disproportionation Reactions (4.6)

• Examples 4.9, 4.10, 4.11 • For Practice 4.9, 4.10, 4.11 • For More Practice 4.9 • Exercises 61–68

Calculations Involving the Stoichiometry of a Reaction (4.7)

• Example 4.12 • For Practice 4.12 • For More Practice 4.12 • Exercises 69–80

Determining the Limiting Reactant and Calculating Theoretical and Percent Yields (4.8)

• Examples 4.13, 4.14 • For Practice 4.13, 4.14 • Exercises 81–96

Calculating Solution Concentration and Using Molarity in Calculations (4.9)

• Examples 4.15, 4.16 • For Practice 4.15, 4.16 • For More Practice 4.15, 4.16 • Exercises 97–102

Solution Dilutions (4.9)

• Example 4.17 • For Practice 4.17 • For More Practice 4.17 • Exercises 103–106

Solution Stoichiometry (4.9)

• Example 4.18 • For Practice 4.18 • For More Practice 4.18 • Exercises 107–110

EXERCISES

Review Questions

1. What is reaction stoichiometry? What is the significance of the coefficients in a balanced chemical equation?

2. In a chemical reaction, what is the limiting reactant? The theoretical yield? The percent yield? What do we mean when we say a reactant is in excess?

3. The percent yield is normally calculated using the actual yield and theoretical yield in units of mass (g or kg). Would the percent yield be different if the actual yield and theoretical yield were in units of amount (moles)?

4. What is an aqueous solution? What is the difference between the solute and the solvent?

5. What is molarity? How is it useful?

6. Explain how a strong electrolyte, a weak electrolyte, and a nonelectrolyte differ.

7. Explain the difference between a strong acid and a weak acid.

8. What does it mean for a compound to be soluble? Insoluble?

9. What are the solubility rules? How are they useful?

10. What cations and anions have compounds that are usually soluble? What are the exceptions? What anions have compounds that are mostly insoluble? What are the exceptions?

11. What is a precipitation reaction? Give an example.

12. How can you predict whether a precipitation reaction will occur upon mixing two aqueous solutions?

13. Explain how a molecular equation, a complete ionic equation, and a net ionic equation differ.

14. What are the Arrhenius definitions of an acid and a base?

15. What is an acid–base reaction? Give an example.

16. Why can you not change the subscripts on the chemical formulas in order to balance a chemical reaction?

17. What is a gas-evolution reaction? Give an example.

18. What reactant types give rise to gas-evolution reactions?

19. What is an oxidation–reduction reaction? Give an example.

20. What are oxidation states?

21. How can oxidation states be used to identify redox reactions?

22. What happens to a substance when it becomes oxidized? Reduced?

23. In a redox reaction, which reactant is the oxidizing agent? The reducing agent?

24. What is a combustion reaction? Why are they important? Give an example.

Problems by Topic

Writing and Balancing Chemical Equations

25. Sulfuric acid is a component of acid rain formed when gaseous sulfur dioxide pollutant reacts with gaseous oxygen and liquid water to form aqueous sulfuric acid. Write a balanced chemical equation for this reaction.

26. Nitric acid is a component of acid rain that forms when gaseous nitrogen dioxide pollutant reacts with gaseous oxygen and liquid water to form aqueous nitric acid. Write a balanced chemical equation for this reaction.

27. In a popular classroom demonstration, solid sodium is added to liquid water and reacts to produce hydrogen gas and aqueous sodium hydroxide. Write a balanced chemical equation for this reaction.

28. When iron rusts, solid iron reacts with gaseous oxygen to form solid iron(III) oxide. Write a balanced chemical equation for this reaction.

29. Write a balanced chemical equation for the fermentation of sucrose ($C_{12}H_{22}O_{11}$) by yeasts in which the aqueous sugar reacts with water to form aqueous ethyl alcohol (C_2H_5OH) and carbon dioxide gas.

30. Write a balanced equation for the photosynthesis reaction in which gaseous carbon dioxide and liquid water react in the presence of chlorophyll to produce aqueous glucose ($C_6H_{12}O_6$) and oxygen gas.

31. Write a balanced chemical equation for each reaction:
  - **a.** Solid lead(II) sulfide reacts with aqueous hydrobromic acid to form solid lead(II) bromide and dihydrogen monosulfide gas.
  - **b.** Gaseous carbon monoxide reacts with hydrogen gas to form gaseous methane (CH_4) and liquid water.
  - **c.** Aqueous hydrochloric acid reacts with solid manganese(IV) oxide to form aqueous manganese(II) chloride, liquid water, and chlorine gas.
  - **d.** Liquid pentane (C_5H_{12}) reacts with gaseous oxygen to form carbon dioxide and liquid water.

32. Write a balanced chemical equation for each reaction:
  - **a.** Solid copper reacts with solid sulfur (S_8) to form solid copper(I) sulfide.
  - **b.** Solid iron(III) oxide reacts with hydrogen gas to form solid iron and liquid water.
  - **c.** Sulfur dioxide gas reacts with oxygen gas to form sulfur trioxide gas.
  - **d.** Gaseous ammonia (NH_3) reacts with gaseous oxygen to form gaseous nitrogen monoxide and gaseous water.

33. Balance each chemical equation:
  - **a.** $CO_2(g) + CaSiO_3(s) + H_2O(l) \longrightarrow SiO_2(s) + Ca(HCO_3)_2(aq)$
  - **b.** $Co(NO_3)_3(aq) + (NH_4)_2S(aq) \longrightarrow Co_2S_3(s) + NH_4NO_3(aq)$
  - **c.** $Cu_2O(s) + C(s) \longrightarrow Cu(s) + CO(g)$
  - **d.** $H_2(g) + Cl_2(g) \longrightarrow HCl(g)$

34. Balance each chemical equation:
  - **a.** $Na_2S(aq) + Cu(NO_3)_2(aq) \longrightarrow NaNO_3(aq) + CuS(s)$
  - **b.** $N_2H_4(l) \longrightarrow NH_3(g) + N_2(g)$
  - **c.** $HCl(aq) + O_2(g) \longrightarrow H_2O(l) + Cl_2(g)$
  - **d.** $FeS(s) + HCl(aq) \longrightarrow FeCl_2(aq) + H_2S(g)$

Types of Aqueous Solutions and Solubility

35. For each compound (all water soluble), would you expect the resulting aqueous solution to conduct electrical current?
a. $CsCl$ **b.** CH_3OH **c.** $Ca(NO_2)_2$ **d.** $C_6H_{12}O_6$

36. Classify each compound as a strong electrolyte or nonelectrolyte:
a. $MgBr_2$ **b.** $C_{12}H_{22}O_{11}$ **c.** Na_2CO_3 **d.** KOH

37. Determine whether each compound is soluble or insoluble. If the compound is soluble, list the ions present in solution.
a. $AgNO_3$ **b.** $Pb(C_2H_3O_2)_2$ **c.** KNO_3 **d.** $(NH_4)_2S$

38. Determine whether each compound is soluble or insoluble. For the soluble compounds, list the ions present in solution.
a. AgI **b.** $Cu_3(PO_4)_2$ **c.** $CoCO_3$ **d.** K_3PO_4

Precipitation Reactions

39. Complete and balance each equation. If no reaction occurs, write NO REACTION.
a. $LiI(aq) + BaS(aq) \longrightarrow$
b. $KCl(aq) + CaS(aq) \longrightarrow$
c. $CrBr_2(aq) + Na_2CO_3(aq) \longrightarrow$
d. $NaOH(aq) + FeCl_3(aq) \longrightarrow$

40. Complete and balance each equation. If no reaction occurs, write NO REACTION.
a. $NaNO_3(aq) + KCl(aq) \longrightarrow$
b. $NaCl(aq) + Hg_2(C_2H_3O_2)_2(aq) \longrightarrow$
c. $(NH_4)_2SO_4(aq) + SrCl_2(aq) \longrightarrow$
d. $NH_4Cl(aq) + AgNO_3(aq) \longrightarrow$

41. Write a molecular equation for the precipitation reaction that occurs (if any) when each pair of aqueous solutions is mixed. If no reaction occurs, write NO REACTION.
a. potassium carbonate and lead(II) nitrate
b. lithium sulfate and lead(II) acetate
c. copper(II) nitrate and magnesium sulfide
d. strontium nitrate and potassium iodide

42. Write a molecular equation for the precipitation reaction that occurs (if any) when each pair of aqueous solutions is mixed. If no reaction occurs, write NO REACTION.
a. sodium chloride and lead(II) acetate
b. potassium sulfate and strontium iodide
c. cesium chloride and calcium sulfide
d. chromium(III) nitrate and sodium phosphate

Ionic and Net Ionic Equations

43. Write balanced complete ionic and net ionic equations for each reaction:
a. $HCl(aq) + LiOH(aq) \longrightarrow H_2O(l) + LiCl(aq)$
b. $MgS(aq) + CuCl_2(aq) \longrightarrow CuS(s) + MgCl_2(aq)$
c. $NaOH(aq) + HNO_3(aq) \longrightarrow H_2O(l) + NaNO_3(aq)$
d. $Na_3PO_4(aq) + NiCl_2(aq) \longrightarrow Ni_3(PO_4)_2(s) + NaCl(aq)$

44. Write balanced complete ionic and net ionic equations for each reaction:
a. $K_2SO_4(aq) + CaI_2(aq) \longrightarrow CaSO_4(s) + KI(aq)$
b. $NH_4Cl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NH_3(g) + NaCl(aq)$
c. $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$
d. $HC_2H_3O_2(aq) + K_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + KC_2H_3O_2(aq)$

45. Dimercury (1+) ions (Hg_2^{2+}) can be removed from solution by precipitation with Cl^-. Suppose that a solution contains aqueous $Hg_2(NO_3)_2$. Write complete ionic and net ionic equations to show the reaction of aqueous $Hg_2(NO_3)_2$ with aqueous sodium chloride to form solid Hg_2Cl_2 and aqueous sodium nitrate.

46. Lead ions can be removed from solution by precipitation with sulfate ions. Suppose that a solution contains lead(II) nitrate. Write complete ionic and net ionic equations to show the reaction of aqueous lead(II) nitrate with aqueous potassium sulfate to form solid lead(II) sulfate and aqueous potassium nitrate.

47. Write balanced molecular and net ionic equations for the reaction between hydrobromic acid and potassium hydroxide.

48. Write balanced molecular and net ionic equations for the reaction between nitric acid and calcium hydroxide.

49. Complete and balance each acid–base equation:
a. $H_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow$
b. $HClO_4(aq) + KOH(aq) \longrightarrow$
c. $H_2SO_4(aq) + NaOH(aq) \longrightarrow$

50. Complete and balance each acid–base equation:
a. $HI(aq) + LiOH(aq) \longrightarrow$
b. $HC_2H_3O_2(aq) + Ca(OH)_2(aq) \longrightarrow$
c. $HCl(aq) + Ba(OH)_2(aq) \longrightarrow$

51. Complete and balance each gas-evolution equation:
a. $HBr(aq) + NiS(s) \longrightarrow$
b. $NH_4I(aq) + NaOH(aq) \longrightarrow$
c. $HBr(aq) + Na_2S(aq) \longrightarrow$
d. $HClO_4(aq) + Li_2CO_3(aq) \longrightarrow$

52. Complete and balance each gas-evolution equation:
a. $HNO_3(aq) + Na_2SO_3(aq) \longrightarrow$
b. $HCl(aq) + KHCO_3(aq) \longrightarrow$
c. $HC_2H_3O_2(aq) + NaHSO_3(aq) \longrightarrow$
d. $(NH_4)_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow$

Oxidation–Reduction and Combustion

53. Assign oxidation states to each atom in each element, ion, or compound:
a. Ag **b.** Ag^+ **c.** CaF_2
d. H_2S **e.** CO_3^{2-} **f.** CrO_4^{2-}

54. Assign oxidation states to each atom in each element, ion, or compound:
a. Cl_2 **b.** Fe^{3+} **c.** $CuCl_2$
d. CH_4 **e.** $Cr_2O_7^{2-}$ **f.** HSO_4^-

55. What is the oxidation state of Cr in each compound?
a. CrO **b.** CrO_3 **c.** Cr_2O_3

56. What is the oxidation state of Cl in each ion?
a. ClO^- **b.** ClO_2^- **c.** ClO_3^- **d.** ClO_4^-

57. Determine whether each reaction is a redox reaction. For each redox reaction, identify the oxidizing agent and the reducing agent.
a. $4\,Li(s) + O_2(g) \longrightarrow 2\,Li_2O(s)$
b. $Mg(s) + Fe^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Fe(s)$
c. $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow PbSO_4(s) + 2\,NaNO_3(aq)$
d. $HBr(aq) + KOH(aq) \longrightarrow H_2O(l) + KBr(aq)$

58. Determine whether each reaction is a redox reaction. For each redox reaction, identify the oxidizing agent and the reducing agent.
a. $Al(s) + 3\,Ag^+(aq) \longrightarrow Al^{3+}(aq) + 3\,Ag(s)$
b. $SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$
c. $Ba(s) + Cl_2(g) \longrightarrow BaCl_2(s)$
d. $Mg(s) + Br_2(l) \longrightarrow MgBr_2(s)$

59. Complete and balance each combustion reaction equation:

a. $S(s) + O_2(g) \longrightarrow$
b. $C_3H_6(g) + O_2(g) \longrightarrow$
c. $Ca(s) + O_2(g) \longrightarrow$
d. $C_5H_{12}S(l) + O_2(g) \longrightarrow$

60. Complete and balance each combustion reaction equation:

a. $C_4H_6(g) + O_2(g) \longrightarrow$
b. $C(s) + O_2(g) \longrightarrow$
c. $CS_2(s) + O_2(g) \longrightarrow$
d. $C_3H_8O(l) + O_2(g) \longrightarrow$

Balancing Redox Reactions

61. Balance each redox reaction occurring in acidic aqueous solution.

a. $K(s) + Cr^{3+}(aq) \longrightarrow Cr(s) + K^+(aq)$
b. $Al(s) + Fe^{2+}(aq) \longrightarrow Al^{3+}(aq) + Fe(s)$
c. $BrO_3^-(aq) + N_2H_4(g) \longrightarrow Br^-(aq) + N_2(g)$

62. Balance each redox reaction occurring in acidic aqueous solution.

a. $Zn(s) + Sn^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Sn(s)$
b. $Mg(s) + Cr^{3+}(aq) \longrightarrow Mg^{2+}(aq) + Cr(s)$
c. $MnO_4^-(aq) + Al(s) \longrightarrow Mn^{2+}(aq) + Al^{3+}(aq)$

63. Balance each redox reaction occurring in acidic aqueous solution.

a. $PbO_2(s) + I^-(aq) \longrightarrow Pb^{2+}(aq) + I_2(s)$
b. $SO_3^{2-}(aq) + MnO_4^-(aq) \longrightarrow SO_4^{2-}(aq) + Mn^{2+}(aq)$
c. $S_2O_3^{2-}(aq) + Cl_2(g) \longrightarrow SO_4^{2-}(aq) + Cl^-(aq)$

64. Balance each redox reaction occurring in acidic aqueous solution.

a. $I^-(aq) + NO_2^-(aq) \longrightarrow I_2(s) + NO(g)$
b. $ClO_4^-(aq) + Cl^-(aq) \longrightarrow ClO_3^-(aq) + Cl_2(g)$
c. $NO_3^-(aq) + Sn^{2+}(aq) \longrightarrow Sn^{4+}(aq) + NO(g)$

65. Balance each redox reaction occurring in basic aqueous solution.

a. $H_2O_2(aq) + ClO_2(aq) \longrightarrow ClO_2^-(aq) + O_2(g)$
b. $Al(s) + MnO_4^-(aq) \longrightarrow MnO_2(s) + Al(OH)_4^-(aq)$
c. $Cl_2(g) \longrightarrow Cl^-(aq) + ClO_3^-(aq)$

66. Balance each redox reaction occurring in basic aqueous solution.

a. $MnO_4^-(aq) + Br^-(aq) \longrightarrow MnO_2(s) + BrO_3^-(aq)$
b. $Ag(s) + CN^-(aq) + O_2(g) \longrightarrow Ag(CN)_2^-(aq)$
c. $NO_2^-(aq) + Al(s) \longrightarrow NH_3(g) + AlO_2^-(aq)$

67. Balance each disproportionation reaction occurring in acidic solution.

a. $ClO^-(aq) \longrightarrow Cl^-(aq) + ClO_3^-(aq)$
b. $Cu_2O(aq) \longrightarrow Cu(s) + Cu^{2+}(aq)$

68. Balance each disproportionation reaction occurring in acidic solution.

a. $Sn^{2+}(aq) \longrightarrow Sn(s) + Sn^{4+}(aq)$
b. $HNO_2(aq) \longrightarrow HNO_3(aq) + NO(g)$

Reaction Stoichiometry

69. Consider the unbalanced equation for the combustion of hexane:

$$C_6H_{14}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Balance the equation and determine how many moles of O_2 are required to react completely with 7.2 mol C_6H_{14}.

70. Consider the unbalanced equation for the neutralization of acetic acid:

$$CH_3COOH(aq) + Ba(OH)_2(aq) \longrightarrow H_2O(l) + (CH_3COO)_2Ba(aq)$$

Balance the equation and determine how many moles of $Ba(OH)_2$ are required to completely neutralize 0.461 mol CH_3COOH.

71. Calculate how many moles of NO_2 form when each quantity of reactant completely reacts:

$$2\ N_2O_5(g) \longrightarrow 4\ NO_2(g) + O_2(g)$$

a. 2.5 mol N_2O_5
b. 6.8 mol N_2O_5
c. 15.2 g N_2O_5
d. 2.87 kg N_2O_5

72. Calculate how many moles of NH_3 form when each quantity of reactant completely reacts:

$$3\ N_2H_4(l) \longrightarrow 4\ NH_3(g) + N_2(g)$$

a. 2.6 mol N_2H_4
b. 3.55 mol N_2H_4
c. 65.3 g N_2H_4
d. 4.88 kg N_2H_4

73. Consider the balanced equation:

$$SiO_2(s) + 3\ C(s) \longrightarrow SiC(s) + 2\ CO(g)$$

Complete the table showing the appropriate number of moles of reactants and products. If the number of moles of a *reactant* is provided, fill in the required amount of the other reactant, as well as the moles of each product formed. If the number of moles of a *product* is provided, fill in the required amount of each reactant to make that amount of product, as well as the amount of the other product that is made.

mol SiO_2	mol C	mol SiC	mol CO
3	______	______	______
______	6	______	______
______	______	______	10
2.8	______	______	______
______	1.55	______	______

74. Consider the balanced equation:

$$2\ N_2H_4(g) + N_2O_4(g) \longrightarrow 3\ N_2(g) + 4\ H_2O(g)$$

Complete the table showing the appropriate number of moles of reactants and products. If the number of moles of a *reactant* is provided, fill in the required amount of the other reactant, as well as the moles of each product formed. If the number of moles of a *product* is provided, fill in the required amount of each reactant to make that amount of product, as well as the amount of the other product that is made.

mol N_2H_4	mol N_2O_4	mol N_2	mol H_2O
2	______	______	______
______	5	______	______
______	______	______	10
2.5	______	______	______
______	4.2	______	______
______	______	11.8	______

75. Hydrobromic acid dissolves solid iron according to the reaction:

$$Fe(s) + 2\ HBr(aq) \longrightarrow FeBr_2(aq) + H_2(g)$$

What mass of HBr (in g) would you need to dissolve a 3.2 g pure iron bar on a padlock? What mass of H_2 would the complete reaction of the iron bar produce?

76. Sulfuric acid dissolves aluminum metal according to the reaction:

$$2\,\mathrm{Al}(s) + 3\,\mathrm{H_2SO_4}(aq) \longrightarrow \mathrm{Al_2(SO_4)_3}(aq) + 3\,\mathrm{H_2}(g)$$

Suppose you wanted to dissolve an aluminum block with a mass of 15.2 g. What minimum mass of $\mathrm{H_2SO_4}$ (in g) would you need? What mass of $\mathrm{H_2}$ gas (in g) would the complete reaction of the aluminum block produce?

77. For each of the reactions, calculate the mass (in grams) of the product formed when 3.67 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

a. $\underline{\mathrm{Ba}(s)} + \mathrm{Cl_2}(g) \longrightarrow \mathrm{BaCl_2}(s)$
b. $\underline{\mathrm{CaO}(s)} + \mathrm{CO_2}(g) \longrightarrow \mathrm{CaCO_3}(s)$
c. $2\,\underline{\mathrm{Mg}(s)} + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{MgO}(s)$
d. $4\,\underline{\mathrm{Al}(s)} + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{Al_2O_3}(s)$

78. For each of the reactions, calculate the mass (in grams) of the product formed when 15.39 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

a. $2\,\mathrm{K}(s) + \underline{\mathrm{Cl_2}(g)} \longrightarrow 2\,\mathrm{KCl}(s)$
b. $2\,\mathrm{K}(s) + \underline{\mathrm{Br_2}(l)} \longrightarrow 2\,\mathrm{KBr}(s)$
c. $4\,\mathrm{Cr}(s) + 3\,\underline{\mathrm{O_2}(g)} \longrightarrow 2\,\mathrm{Cr_2O_3}(s)$
d. $2\,\underline{\mathrm{Sr}(s)} + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{SrO}(s)$

79. For each of the acid–base reactions, calculate the mass (in grams) of each acid necessary to completely react with and neutralize 4.85 g of the base.

a. $\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \longrightarrow \mathrm{H_2O}(l) + \mathrm{NaCl}(aq)$
b. $2\,\mathrm{HNO_3}(aq) + \mathrm{Ca(OH)_2}(aq) \longrightarrow 2\,\mathrm{H_2O}(l) + \mathrm{Ca(NO_3)_2}(aq)$
c. $\mathrm{H_2SO_4}(aq) + 2\,\mathrm{KOH}(aq) \longrightarrow 2\,\mathrm{H_2O}(l) + \mathrm{K_2SO_4}(aq)$

80. For each precipitation reaction, calculate how many grams of the first reactant are necessary to completely react with 55.8 g of the second reactant.

a. $2\,\mathrm{KI}(aq) + \mathrm{Pb(NO_3)_2}(aq) \longrightarrow \mathrm{PbI_2}(s) + 2\,\mathrm{KNO_3}(aq)$
b. $\mathrm{Na_2CO_3}(aq) + \mathrm{CuCl_2}(aq) \longrightarrow \mathrm{CuCO_3}(s) + 2\,\mathrm{NaCl}(aq)$
c. $\mathrm{K_2SO_4}(aq) + \mathrm{Sr(NO_3)_2}(aq) \longrightarrow \mathrm{SrSO_4}(s) + 2\,\mathrm{KNO_3}(aq)$

Limiting Reactant, Theoretical Yield, and Percent Yield

81. For the reaction, find the limiting reactant for each of the initial amounts of reactants.

$$2\,\mathrm{Na}(s) + \mathrm{Br_2}(g) \longrightarrow 2\,\mathrm{NaBr}(s)$$

a. 2 mol Na, 2 mol $\mathrm{Br_2}$
b. 1.8 mol Na, 1.4 mol $\mathrm{Br_2}$
c. 2.5 mol Na, 1 mol $\mathrm{Br_2}$
d. 12.6 mol Na, 6.9 mol $\mathrm{Br_2}$

82. For the reaction, find the limiting reactant for each initial amount of reactants.

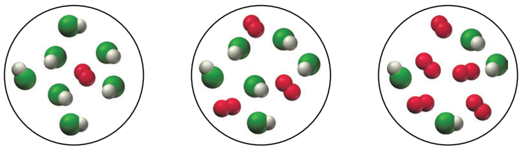
$$4\,\mathrm{Al}(s) + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{Al_2O_3}(s)$$

a. 1 mol Al, 1 mol $\mathrm{O_2}$
b. 4 mol Al, 2.6 mol $\mathrm{O_2}$
c. 16 mol Al, 13 mol $\mathrm{O_2}$
d. 7.4 mol Al, 6.5 mol $\mathrm{O_2}$

83. Consider the reaction:

$$4\,\mathrm{HCl}(g) + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{H_2O}(g) + 2\,\mathrm{Cl_2}(g)$$

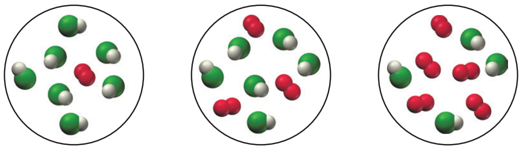
Each molecular diagram represents an initial mixture of the reactants. How many molecules of $\mathrm{Cl_2}$ would be formed from the reaction mixture that produces the greatest amount of products?


(a) **(b)** **(c)**

84. Consider the reaction:

$$2\,\mathrm{CH_3OH}(g) + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{CO_2}(g) + 4\,\mathrm{H_2O}(g)$$

Each of the molecular diagrams represents an initial mixture of the reactants. How many $\mathrm{CO_2}$ molecules would be formed from the reaction mixture that produces the greatest amount of products?


(a) **(b)** **(c)**

85. For the reaction, compute the theoretical yield of the product (in moles) for each initial amount of reactants.

$$\mathrm{Ti}(s) + 2\,\mathrm{Cl_2}(g) \longrightarrow \mathrm{TiCl_4}(s)$$

a. 4 mol Ti, 4 mol $\mathrm{Cl_2}$
b. 7 mol Ti, 17 mol $\mathrm{Cl_2}$
c. 12.4 mol Ti, 18.8 mol $\mathrm{Cl_2}$

86. For the reaction, compute the theoretical yield of product (in moles) for each initial amount of reactants.

$$2\,\mathrm{Mn}(s) + 2\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{MnO_2}(s)$$

a. 3 mol Mn, 3 mol $\mathrm{O_2}$
b. 4 mol Mn, 7 mol $\mathrm{O_2}$
c. 27.5 mol Mn, 43.8 mol $\mathrm{O_2}$

87. Zinc sulfide reacts with oxygen according to the reaction:

$$2\,\mathrm{ZnS}(s) + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{ZnO}(s) + 2\,\mathrm{SO_2}(g)$$

A reaction mixture initially contains 4.2 mol ZnS and 6.8 mol $\mathrm{O_2}$. Once the reaction has occurred as completely as possible, what amount (in moles) of the excess reactant is left?

88. Iron(II) sulfide reacts with hydrochloric acid according to the reaction:

$$\mathrm{FeS}(s) + 2\,\mathrm{HCl}(aq) \longrightarrow \mathrm{FeCl_2}(s) + \mathrm{H_2S}(g)$$

A reaction mixture initially contains 0.223 mol FeS and 0.652 mol HCl. Once the reaction has occurred as completely as possible, what amount (in moles) of the excess reactant is left?

89. For the reaction, compute the theoretical yield of product (in grams) for each initial amount of reactants.

$$2\,\mathrm{Al}(s) + 3\,\mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{AlCl_3}(s)$$

a. 2.0 g Al, 2.0 g $\mathrm{Cl_2}$
b. 7.5 g Al, 24.8 g $\mathrm{Cl_2}$
c. 0.235 g Al, 1.15 g $\mathrm{Cl_2}$

90. For the reaction, compute the theoretical yield of the product (in grams) for each initial amount of reactants.

$$\mathrm{Ti}(s) + 2\,\mathrm{F_2}(g) \longrightarrow \mathrm{TiF_4}(s)$$

a. 5.0 g Ti, 5.0 g $\mathrm{F_2}$
b. 2.4 g Ti, 1.6 g $\mathrm{F_2}$
c. 0.233 g Ti, 0.288 g $\mathrm{F_2}$

91. Iron(III) sulfide reacts with carbon monoxide according to the equation:

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

A reaction mixture initially contains 22.55 g Fe_2O_3 and 14.78 g CO. Once the reaction has occurred as completely as possible, what mass (in grams) of the excess reactant is left?

92. Elemental phosphorus reacts with chlorine gas according to the equation:

$$P_4(s) + 6\,Cl_2(g) \longrightarrow 4\,PCl_3(l)$$

A reaction mixture initially contains 45.69 g P_4 and 131.3 g Cl_2. Once the reaction has occurred as completely as possible, what mass (in grams) of the excess reactant is left?

93. Lead ions can be precipitated from solution with KCl according to the reaction:

$$Pb^{2+}(aq) + 2\,KCl(aq) \longrightarrow PbCl_2(s) + 2\,K^+(aq)$$

When 28.5 g KCl is added to a solution containing 25.7 g Pb^{2+}, a $PbCl_2$ precipitate forms. The precipitate is filtered and dried and found to have a mass of 29.4 g. Determine the limiting reactant, theoretical yield of $PbCl_2$, and percent yield for the reaction.

94. Magnesium oxide can be made by heating magnesium metal in the presence of oxygen. The balanced equation for the reaction is:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

When 10.1 g Mg is allowed to react with 10.5 g O_2, 11.9 g MgO is collected. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

95. Urea (CH_4N_2O) is a common fertilizer that can be synthesized by the reaction of ammonia (NH_3) with carbon dioxide:

$$2\,NH_3(aq) + CO_2(aq) \longrightarrow CH_4N_2O(aq) + H_2O(l)$$

In an industrial synthesis of urea, a chemist combines 136.4 kg of ammonia with 211.4 kg of carbon dioxide and obtains 168.4 kg of urea. Determine the limiting reactant, theoretical yield of urea, and percent yield for the reaction.

96. Many computer chips are manufactured from silicon, which occurs in nature as SiO_2. When SiO_2 is heated to melting, it reacts with solid carbon to form liquid silicon and carbon monoxide gas. In an industrial preparation of silicon, 155.8 kg of SiO_2 reacts with 78.3 kg of carbon to produce 66.1 kg of silicon. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

Solution Concentration and Solution Stoichiometry

97. Calculate the molarity of each solution:
a. 3.25 mol LiCl in 2.78 L of solution
b. 28.33 g $C_6H_{12}O_6$ in 1.28 L of solution
c. 32.4 mg NaCl in 122.4 mL of solution

98. Calculate the molarity of each solution:
a. 0.38 mol $LiNO_3$ in 6.14 L of solution
b. 72.8 g C_2H_6O in 2.34 L of solution
c. 12.87 mg KI in 112.4 mL of solution

99. How many moles of KCl are contained in each solution?
a. 0.556 L of a 2.3 mol L^{-1} KCl solution
b. 1.8 L of a 0.85 mol L^{-1} KCl solution
c. 114 mL of a 1.85 mol L^{-1} KCl solution

100. What volume of a 0.200 mol L^{-1} ethanol solution contains each amount in moles of ethanol?
a. 0.45 mol ethanol
b. 1.22 mol ethanol
c. 1.2×10^{-2} mol ethanol

101. A laboratory procedure calls for making 400.0 mL of a 1.1 mol L^{-1} $NaNO_3$ solution. What mass of $NaNO_3$ (in grams) is needed?

102. A chemist wants to make 5.5 L of a 0.300 mol L^{-1} $CaCl_2$ solution. What mass of $CaCl_2$ (in grams) should the chemist use?

103. If 123 mL of a 1.1 mol L^{-1} glucose solution is diluted to 500.0 mL, what is the molarity of the diluted solution?

104. If 3.5 L of a 4.8 mol L^{-1} $SrCl_2$ solution is diluted to 45 L, what is the molarity of the diluted solution?

105. To what volume should you dilute 50.0 mL of a 12 mol L^{-1} stock HNO_3 solution to obtain a 0.100 mol L^{-1} HNO_3 solution?

106. To what volume should you dilute 25 mL of a 10.0 mol L^{-1} H_2SO_4 solution to obtain a 0.150 mol L^{-1} H_2SO_4 solution?

107. Consider the precipitation reaction:

$$2\,Na_3PO_4(aq) + 3\,CuCl_2(aq) \longrightarrow Cu_3(PO_4)_2(s) + 6\,NaCl(aq)$$

What volume of a 0.175 mol L^{-1} Na_3PO_4 solution is necessary to completely react with 95.4 mL of 0.102 mol L^{-1} $CuCl_2$?

108. Consider the reaction:

$$Li_2S(aq) + Co(NO_3)_2(aq) \longrightarrow 2\,LiNO_3(aq) + CoS(s)$$

What volume of a 0.150 mol L^{-1} Li_2S solution is required to completely react with 125 mL of 0.150 mol L^{-1} $Co(NO_3)_2$?

109. What is the minimum amount of 6.0 mol L^{-1} H_2SO_4 necessary to produce 25.0 g of $H_2(g)$ according to the reaction between aluminum and sulfuric acid?

$$2\,Al(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$$

110. What is the molarity of $ZnCl_2$ that forms when 25.0 g of zinc completely reacts with $CuCl_2$ according to the following reaction? Assume a final volume of 275 mL.

$$Zn(s) + CuCl_2(aq) \longrightarrow ZnCl_2(aq) + Cu(s)$$

Cumulative Problems

111. The density of a 20.0%-by-mass ethylene glycol ($C_2H_6O_2$) solution in water is 1.03 g mL^{-1}. Find the molarity of the solution.

112. Find the percent by mass of sodium chloride in a 1.35 mol L^{-1} NaCl solution. The density of the solution is 1.05 g mL^{-1}.

113. People often use sodium bicarbonate as an antacid to neutralize excess hydrochloric acid in an upset stomach. What mass of hydrochloric acid (in grams) can 2.5 g of sodium bicarbonate neutralize? (*Hint:* Begin by writing a balanced equation for the reaction between aqueous sodium bicarbonate and aqueous hydrochloric acid.)

114. Toilet bowl cleaners often contain hydrochloric acid, which dissolves the calcium carbonate deposits that accumulate within a toilet bowl. What mass of calcium carbonate (in grams) can 3.8 g of HCl dissolve? (*Hint:* Begin by writing a balanced equation for the reaction between hydrochloric acid and calcium carbonate.)

115. The combustion of gasoline produces carbon dioxide and water. Assume gasoline to be pure octane (C_8H_{18}) and calculate the mass (in kilograms) of carbon dioxide that is added to the atmosphere per 1.0 kg of octane burned. (*Hint:* Begin by writing a balanced equation for the combustion reaction.)

116. Many home barbeques are fuelled with propane gas (C_3H_8). What mass of carbon dioxide (in kilograms) is produced upon the complete combustion of 18.9 L of propane (approximate contents of one 5-gallon tank)? Assume that the density of the liquid propane in the tank is 0.621 g mL^{-1}. (*Hint:* Begin by writing a balanced equation for the combustion reaction.)

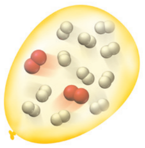
117. Aspirin can be made in the laboratory by reacting acetic anhydride ($C_4H_6O_3$) with salicylic acid ($C_7H_6O_3$) to form aspirin ($C_9H_8O_4$) and acetic acid ($C_2H_4O_2$). The balanced equation is:

$$C_4H_6O_3 + C_7H_6O_3 \longrightarrow C_9H_8O_4 + C_2H_4O_2$$

In a laboratory synthesis, a student begins with 3.00 mL of acetic anhydride (density $= 1.08$ g mL^{-1}) and 1.25 g of salicylic acid. Once the reaction is complete, the student collects 1.22 g of aspirin. Determine the limiting reactant, theoretical yield of aspirin, and percent yield for the reaction.

118. The combustion of liquid ethanol (C_2H_5OH) produces carbon dioxide and water. After 4.62 mL of ethanol (density $=$ 0.789 g mL^{-1}) was allowed to burn in the presence of 15.55 g of oxygen gas, 3.72 mL of water (density $= 1.00$ g mL^{-1}) was collected. Determine the limiting reactant, theoretical yield of H_2O, and percent yield for the reaction. (*Hint:* Write a balanced equation for the combustion of ethanol.)

119. A loud classroom demonstration involves igniting a hydrogenfilled balloon. The hydrogen within the balloon reacts explosively with oxygen in the air to form water. If the balloon is filled with a mixture of hydrogen and oxygen, the explosion is even louder than if the balloon is filled only with hydrogen—the intensity of the explosion depends on the relative amounts of oxygen and hydrogen within the balloon. Look at the molecular views representing different amounts of hydrogen and oxygen in four different balloons. Based on the balanced chemical equation, which balloon will make the loudest explosion?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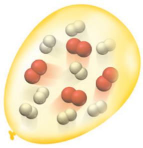
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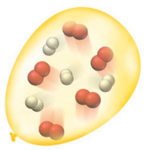
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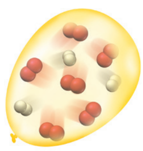
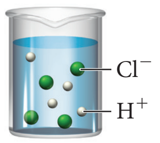
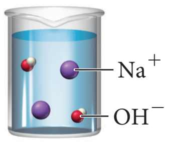

(d)

120. A hydrochloric acid solution will neutralize a sodium hydroxide solution. Look at the molecular views showing one beaker of HCl and four beakers of NaOH. Which NaOH beaker will just neutralize the HCl beaker? Begin by writing a balanced chemical equation for the neutralization reaction.





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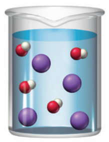
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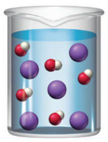
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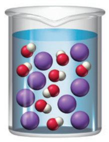

(d)

121. Predict the products and write a balanced molecular equation for each reaction. If no reaction occurs, write NO REACTION.

a. $HCl(aq) + Hg_2(NO_3)_2(aq) \longrightarrow$
b. $KHSO_3(aq) + HNO_3(aq) \longrightarrow$
c. aqueous ammonium chloride and aqueous lead(II) nitrate
d. aqueous ammonium chloride and aqueous calcium hydroxide

122. Predict the products and write a balanced molecular equation for each reaction. If no reaction occurs, write NO REACTION.

a. $H_2SO_4(aq) + HNO_3(aq) \longrightarrow$
b. $Cr(NO_3)_3(aq) + LiOH(aq) \longrightarrow$
c. liquid pentanol ($C_5H_{12}O$) and gaseous oxygen
d. aqueous strontium sulfide and aqueous copper(II) sulfate

123. Hard water often contains dissolved Ca^{2+} and Mg^{2+} ions. One way to soften water is to add phosphates. The phosphate ion forms insoluble precipitates with calcium and magnesium ions, removing them from solution. Suppose that a solution is 0.050 mol L^{-1} in calcium chloride and 0.085 mol L^{-1} in magnesium nitrate. What mass of sodium phosphate would have to be added to 1.5 L of this solution to completely eliminate the hard water ions? Assume complete reaction.

124. An acid solution is 0.100 mol L^{-1} in HCl and 0.200 mol L^{-1} in H_2SO_4. What volume of a 0.150 mol L^{-1} KOH solution would completely neutralize all the acid in 500.0 mL of this solution?

125. Find the mass of barium metal (in grams) that must react with O_2 to produce enough barium oxide to prepare 1.0 L of a 0.10 mol L^{-1} solution of OH^-.

126. A solution contains Cr^{3+} ions and Mg^{2+} ions. The addition of 1.00 L of 1.51 mol L^{-1} NaF solution is required to cause the complete precipitation of these ions as $CrF_3(s)$ and $MgF_2(s)$. The total mass of the precipitate is 49.6 g. Find the mass of Cr^{3+} in the original solution.

127. The nitrogen in sodium nitrate and in ammonium sulfate is available to plants as fertilizer. Which is the more economical source of nitrogen, a fertilizer containing 30.0% sodium nitrate by weight and costing $18 per 100 kg or one containing 20.0% ammonium sulfate by weight and costing $16.20 per 100 kg?

128. Find the volume of 0.110 mol L^{-1} of hydrochloric acid necessary to react completely with 1.52 g $Al(OH)_3$.

129. Treatment of gold metal with BrF_3 and KF produces Br_2 and $KAuF_4$, a salt of gold. Identify the oxidizing agent and the reducing agent in this reaction. Find the mass of the gold salt that forms when a 73.5 g mixture of equal masses of all three reactants is prepared.

130. We prepare a solution by mixing 0.10 L of 0.12 mol L^{-1} sodium chloride with 0.23 L of a 0.18 mol L^{-1} $MgCl_2$ solution. What volume of a 0.20 mol L^{-1} silver nitrate solution do we need to precipitate all the Cl^- ions in the solution as AgCl?

131. A solution contains one or more of the following ions: Ag^+, Ca^{2+}, and Cu^{2+}. When you add sodium chloride to the solution, no precipitate forms. When you add sodium sulfate to the solution, a white precipitate forms. You filter off the precipitate and add sodium carbonate to the remaining solution, producing another precipitate. Which ions were present in the original solution? Write net ionic equations for the formation of each of the precipitates observed.

132. A solution contains one or more of the following ions: Hg_2^{2+}, Ba^{2+}, and Fe^{2+}. When potassium chloride is added to the solution, a precipitate forms. The precipitate is filtered off, and potassium sulfate is added to the remaining solution, producing no precipitate. When potassium carbonate is added to the remaining solution, a precipitate forms. Which ions were present in the original solution? Write net ionic equations for the formation of each of the precipitates observed.

133. The reaction of NH_3 and O_2 forms NO and water. The NO can be used to convert P_4 to P_4O_6, forming N_2 in the process. The P_4O_6 can be treated with water to form H_3PO_3, which forms PH_3 and H_3PO_4 when heated. Find the mass of PH_3 that forms from the reaction of 1.00 g NH_3.

134. An important reaction that takes place in a blast furnace during the production of iron is the formation of iron metal and CO_2 from Fe_2O_3 and CO. Find the mass of Fe_2O_3 required to form 910 kg of iron. Find the amount of CO_2 that forms in this process.

135. A liquid fuel mixture contains 30.35% hexane (C_6H_{14}), 15.85% heptane (C_7H_{16}), and the rest octane (C_8H_{18}). What maximum mass of carbon dioxide is produced by the complete combustion of 10.0 kg of this fuel mixture?

136. Titanium occurs in the magnetic mineral ilmenite ($FeTiO_3$), which is often found mixed with sand. The ilmenite can be separated from the sand with magnets. The titanium can then be extracted from the ilmenite by the following set of reactions:

$$FeTiO_3(s) + 3\,Cl_2(g) + 3\,C(s) \longrightarrow 3\,CO(g) + FeCl_2(s) + TiCl_4(g)$$

$$TiCl_4(l) + 2\,Mg(s) \longrightarrow 2\,MgCl_2(l) + Ti(s)$$

Suppose that an ilmenite–sand mixture contains 22.8% ilmenite by mass and that the first reaction is carried out with a 90.8% yield. If the second reaction is carried out with an 85.9% yield, what mass of titanium can be obtained from 1.00 kg of the ilmenite–sand mixture?

Challenge Problems

137. A mixture of C_3H_8 and C_2H_2 has a mass of 2.0 g. It is burned in excess O_2 to form a mixture of water and carbon dioxide that contains 1.5 times as many moles of CO_2 as of water. Find the mass of C_2H_2 in the original mixture.

138. A mixture of 20.6 g P and 79.4 g Cl_2 reacts completely to form PCl_3 and PCl_5 as the only products. Find the mass of PCl_3 formed.

139. A solution contains Ag^+ and Hg^{2+} ions. The addition of 0.100 L of 1.22 mol L^{-1} NaI solution is just enough to precipitate all the ions as AgI and HgI_2. The total mass of the precipitate is 28.1 g. Find the mass of AgI in the precipitate.

140. Lakes that have been acidified by acid rain (HNO_3 and H_2SO_4) can be neutralized by a process called liming, in which limestone ($CaCO_3$) is added to the acidified water. What mass of limestone (in kg) would completely neutralize a 15.2 billion-litre lake that is 1.8×10^{-5} mol L^{-1} in H_2SO_4 and 8.7×10^{-6} mol L^{-1} in HNO_3?

141. We learned in Section 4.4 that sodium carbonate is often added to laundry detergents to soften hard water and make the detergent more effective. Suppose that a particular detergent mixture is designed to soften hard water that is 3.5×10^{-3} mol L^{-1} in Ca^{2+} and 1.1×10^{-3} mol L^{-1} in Mg^{2+} and that the average capacity of a washing machine is 75 L of water. If the detergent requires using 0.65 kg detergent per load of laundry, determine what percentage (by mass) of the detergent should be sodium carbonate in order to completely precipitate all of the calcium and magnesium ions in an average load of laundry water.

142. Lead poisoning is a serious condition resulting from the ingestion of lead in food, water, or other environmental sources. It affects the central nervous system, leading to a variety of symptoms such as distractibility, lethargy, and loss of motor coordination. Lead poisoning is treated with chelating agents, substances that bind to metal ions, allowing it to be eliminated in the urine. A modern chelating agent used for this purpose is succimer ($C_4H_6O_4S_2$). Suppose you are trying to determine the appropriate dose for succimer treatment of lead poisoning. What minimum mass of succimer (in milligrams) is needed to bind all of the lead in a patient's bloodstream? Assume that patient blood lead levels are 45 μg dL^{-1}, that total blood volume is 5.0 L, and that one mole of succimer binds one mole of lead.

143. A particular kind of emergency breathing apparatus—often placed in mines, caves, or other places where oxygen might become depleted or where the air might become poisoned—works via the following chemical reaction:

$$4\,KO_2(s) + 2\,CO_2(g) \longrightarrow 2\,K_2CO_3(s) + 3\,O_2(g)$$

Notice that the reaction produces O_2, which can be breathed, and absorbs CO_2, a product of respiration. Suppose you work for a company interested in producing a self-rescue breathing apparatus (based on the above reaction) that would allow the user to survive for 10 minutes in an emergency situation. What are the important chemical considerations in designing such a unit? Estimate how much KO_2 would be required for the apparatus. (Find any necessary additional information—such as human breathing rates—from appropriate sources. Assume that normal air is 20% oxygen.)

144. Metallic aluminum reacts with MnO_2 at elevated temperatures to form manganese metal and aluminum oxide. A mixture of the two reactants is 67.2% mole percent Al. Find the theoretical yield (in grams) of manganese from the reaction of 250 g of this mixture.

145. Hydrolysis of the compound B_5H_9 forms boric acid, H_3BO_3. Fusion of boric acid with sodium oxide forms a borate salt, $Na_2B_4O_7$. Without writing complete equations, find the mass (in grams) of B_5H_9 required to form 151 g of the borate salt by this reaction sequence.

146. A mixture of carbon and sulfur has a mass of 9.0 g. Complete combustion with excess O_2 gives 23.3 g of a mixture of CO_2 and SO_2. Find the mass of sulfur in the original mixture.

Conceptual Problems

147. Consider the reaction:

$$4\,K(s) + O_2(g) \longrightarrow 2\,K_2O(s)$$

The molar mass of K is 39.10 g mol^{-1} and that of O_2 is 32.00 g mol^{-1}. Without doing any extensive calculations, pick the conditions under which potassium is the limiting reactant and explain your reasoning.

a. 170 g K, 31 g O_2
b. 16 g K, 2.5 g O_2
c. 165 kg K, 28 kg O_2
d. 1.5 g K, 0.38 g O_2

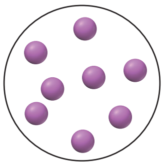
148. Consider the reaction:

$$2\,NO(g) + 5\,H_2(g) \longrightarrow 2\,NH_3(g) + 2\,H_2O(g)$$

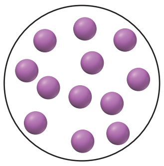
A reaction mixture initially contains 5 moles of NO and 10 moles of H_2. Without doing any calculations, determine which set of amounts best represents the mixture after the reactants have reacted as completely as possible. Explain your reasoning.

a. 1 mol NO, 0 mol H_2, 4 mol NH_3, 4 mol H_2O
b. 0 mol NO, 1 mol H_2, 5 mol NH_3, 5 mol H_2O
c. 3 mol NO, 5 mol H_2, 2 mol NH_3, 2 mol H_2O
d. 0 mol NO, 0 mol H_2, 4 mol NH_3, 4 mol H_2O

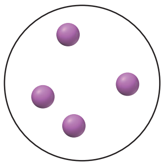
149. The circle below represents 1.0 L of a solution with a solute concentration of 1 mol L^{-1}:



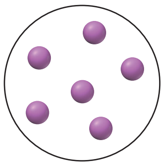
Explain what you would add (the amount of solute or volume of solvent) to the solution to obtain a solution represented by each diagram:



(a)



(b)

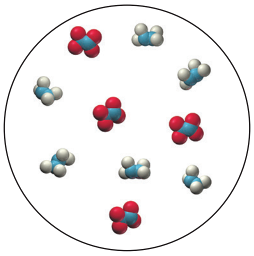


(c)

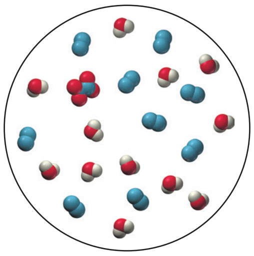
150. Consider the reaction:

$$2\,N_2H_4(g) + N_2O_4(g) \longrightarrow 3\,N_2(g) + 4\,H_2O(g)$$

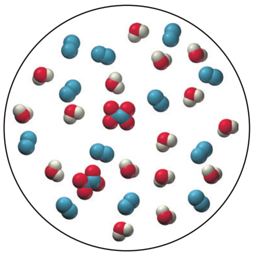
Consider also this representation of an initial mixture of N_2H_4 and N_2O_4:



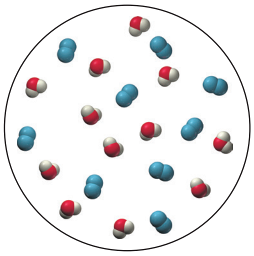
Which diagram best represents the reaction mixture after the reactants have reacted as completely as possible?



(a)



(b)



(c)