

CHAPTER
4

AQUEOUS REACTIONS AND SOLUTION STOICHIOMETRY



A VIEW OF THE PACIFIC OCEAN along the California coastline.

4.1 General Properties of Aqueous Solutions

We begin by examining the nature of the substances dissolved in water, whether they exist in water as ions, molecules, or some mixture of the two. This information is necessary to understand the nature of reactants in aqueous solutions.

4.2 Precipitation Reactions

We will identify reactions in which soluble reactants yield an insoluble product.

4.3 Acid-Base Reactions

We will explore reactions in which protons, H^+ ions, are transferred between reactants.

4.4 Oxidation-Reduction Reactions

We will examine reactions in which electrons are transferred between reactants.

4.5 Concentrations of Solutions

We will recognize that the amount of a compound in a given volume of a solution is called its *concentration*. Concentration can be expressed in a number of ways, the most usual of which is moles of compound per liter of solution (*molarity*).

4.6 Solution Stoichiometry and Chemical Analysis

We will use what we have learned about stoichiometry and concentration to calculate the amounts or concentrations of substances in solution.

THE WATERS OF THE PACIFIC OCEAN, seen in this chapter-opening photograph of the California coast, are part of the World Ocean that covers almost two-thirds of our planet. Water has been the key to much of Earth's evolutionary history. Life itself almost certainly originated in water, and the

need for water by all forms of life has helped determine diverse biological structures. Your own body is about 60% water by mass. We will see repeatedly throughout this text that water possesses many unusual properties essential to supporting life on Earth.

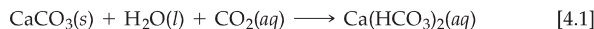
The waters of the World Ocean may not appear to be any different from those of Lake Tahoe or the water that flows from your kitchen faucet, but a taste of seawater is all it takes to demonstrate that there is an important difference. Water has an exceptional ability to dissolve a wide variety of substances. Water on Earth—whether it is drinking water from the tap, water from a clear mountain stream, or seawater—invariably contains a variety of dissolved substances. A solution in which water is the dissolving medium is called an **aqueous solution**. Seawater is different from what we call “freshwater” because it has a much higher total concentration of dissolved ionic substances.

Water is the medium for most of the chemical reactions that take place within us and around us. Nutrients dissolved in blood are carried to our cells, where they enter into reactions that help keep us alive. Automobile parts rust



▲ **Figure 4.1 Limestone cave.** When CO_2 dissolves in water, the resulting solution is slightly acidic. Limestone caves are formed by the dissolving action of this acidic solution acting on CaCO_3 in the limestone.

when they come into frequent contact with aqueous solutions that contain various dissolved substances. Spectacular limestone caves (Figure 4.1 ◀) are formed by the dissolving action of underground water containing carbon dioxide, $\text{CO}_2(aq)$:



In Chapter 3 we saw a few simple types of chemical reactions and their descriptions. In this chapter we continue to examine chemical reactions by focusing on aqueous solutions. A great deal of important chemistry occurs in aqueous solutions, so we need to learn the vocabulary and concepts used to describe and understand this chemistry. In addition, we will extend the concepts of stoichiometry that we learned in Chapter 3 by considering how solution concentrations are expressed and used.

4.1 GENERAL PROPERTIES OF AQUEOUS SOLUTIONS

Recall that a *solution* is a homogeneous mixture of two or more substances. ∞∞ (Section 1.2) The substance present in the greatest quantity is usually called the **solvent**. The other substances in the solution are called the **solutes**; they are said to be dissolved in the solvent. When a small amount of sodium chloride (NaCl) is dissolved in a large quantity of water, for example, the water is the solvent and the sodium chloride is the solute.

Electrolytic Properties

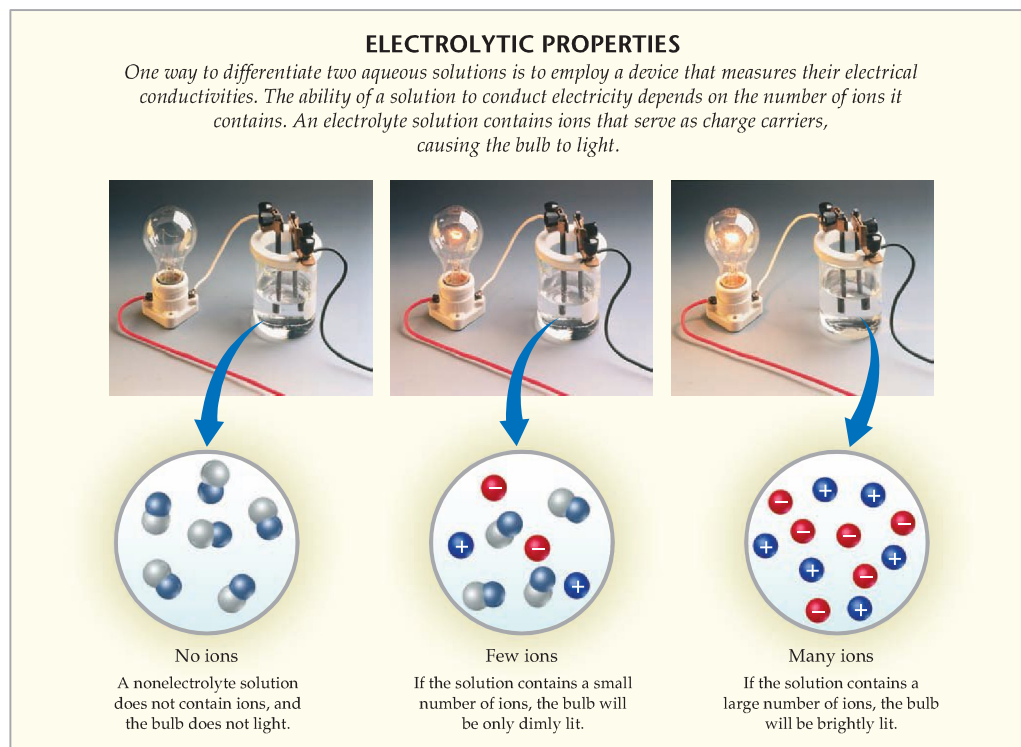
Imagine preparing two aqueous solutions—one by dissolving a teaspoon of table salt (sodium chloride) in a cup of water and the other by dissolving a teaspoon of table sugar (sucrose) in a cup of water. Both solutions are clear and colorless. How do they differ? One way, which might not be immediately obvious, is in their electrical conductivity: The salt solution is a good conductor of electricity; the sugar solution is not.

Whether a solution conducts electricity can be determined by using a device such as that shown in Figure 4.2 ▶. To light the bulb, an electric current must flow between the two electrodes that are immersed in the solution. Although water itself is a poor conductor of electricity, the presence of ions causes aqueous solutions to become good conductors. Ions carry electrical charge from one electrode to the other, completing the electrical circuit. Thus, the conductivity of NaCl solutions indicates the presence of ions in the solution. The lack of conductivity of sucrose solutions indicates the absence of ions. When NaCl dissolves in water, the solution contains Na^+ and Cl^- ions, each surrounded by water molecules. When sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) dissolves in water, the solution contains only neutral sucrose molecules surrounded by water molecules.

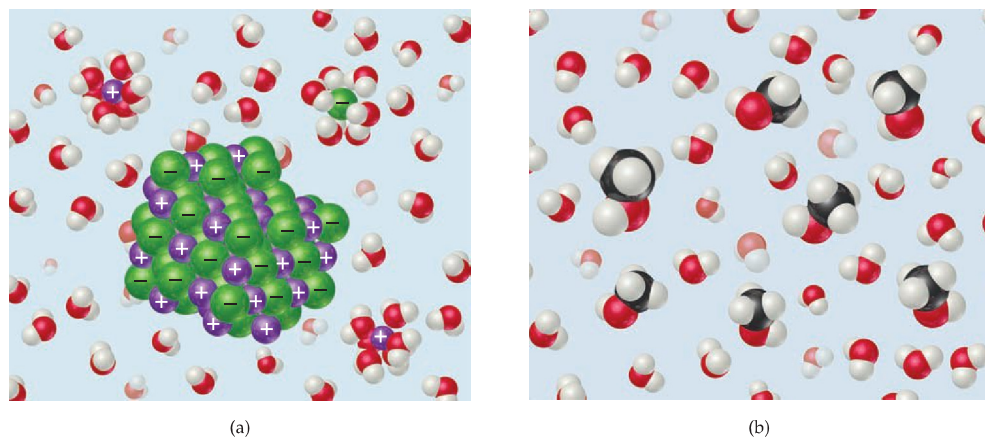
A substance (such as NaCl) whose aqueous solutions contain ions is called an **electrolyte**. A substance (such as $\text{C}_{12}\text{H}_{22}\text{O}_{11}$) that does not form ions in solution is called a **nonelectrolyte**. The difference between NaCl and $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ arises largely because NaCl is ionic, whereas $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ is molecular.

Ionic Compounds in Water

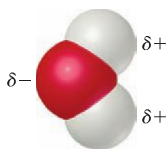
Recall from Section 2.7 and especially Figure 2.23 that solid NaCl consists of an orderly arrangement of Na^+ and Cl^- ions. When NaCl dissolves in water, each ion separates from the solid structure and disperses throughout the solution, as shown in Figure 4.3(a) ▶. The ionic solid *dissociates* into its component ions as it dissolves.



▲ **Figure 4.2** Measuring ion concentrations using conductivity.



▲ **Figure 4.3** **Dissolution in water.** (a) When an ionic compound dissolves in water, H_2O molecules separate, surround, and disperse the ions into the liquid. (b) Methanol, CH_3OH , a molecular compound, dissolves without forming ions. The methanol molecules contain black spheres, which represent carbon atoms. In both parts (a) and (b) the water molecules have been moved apart so that the solute particles can be seen more clearly.



Water is a very effective solvent for ionic compounds. Although water is an electrically neutral molecule, one end of the molecule (the O atom) is rich in electrons and has a partial negative charge, denoted by δ^- . The other end (the H atoms) has a partial positive charge, denoted by δ^+ as shown in the margin. Positive ions (cations) are attracted by the negative end of H_2O , and negative ions (anions) are attracted by the positive end.

As an ionic compound dissolves, the ions become surrounded by H_2O molecules, as shown in Figure 4.3(a). The ions are said to be solvated. We denote these ions in chemical equations by writing them as $\text{Na}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$, where “aq” is an abbreviation for “aqueous.” ∞ (Section 3.1) The **solvation** process helps stabilize the ions in solution and prevents cations and anions from recombining. Furthermore, because the ions and their shells of surrounding water molecules are free to move about, the ions become dispersed uniformly throughout the solution.

We can usually predict the nature of the ions present in a solution of an ionic compound from the chemical name of the substance. Sodium sulfate (Na_2SO_4), for example, dissociates into sodium ions (Na^+) and sulfate ions (SO_4^{2-}). You must remember the formulas and charges of common ions (Tables 2.4 and 2.5) to understand the forms in which ionic compounds exist in aqueous solution.



GIVE IT SOME THOUGHT

What dissolved species are present in a solution of (a) KCN, (b) NaClO_4 ?

Molecular Compounds in Water

When a molecular compound dissolves in water, the solution usually consists of intact molecules dispersed throughout the solution. Consequently, most molecular compounds are nonelectrolytes. As we have seen, table sugar (sucrose) is an example of a nonelectrolyte. As another example, a solution of methanol (CH_3OH) in water consists entirely of CH_3OH molecules dispersed throughout the water [Figure 4.3(b)].

A few molecular substances, however, have aqueous solutions that contain ions. Acids are the most important of these solutions. For example, when $\text{HCl}(\text{g})$ dissolves in water to form hydrochloric acid, $\text{HCl}(\text{aq})$, it *ionizes*; that is, it dissociates into $\text{H}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions.

Strong and Weak Electrolytes

Two categories of electrolytes—strong and weak—differ in the extent to which they conduct electricity. **Strong electrolytes** are those solutes that exist in solution completely or nearly completely as ions. Essentially all soluble ionic compounds (such as NaCl) and a few molecular compounds (such as HCl) are strong electrolytes. **Weak electrolytes** are those solutes that exist in solution mostly in the form of molecules with only a small fraction in the form of ions. For example, in a solution of acetic acid (CH_3COOH) most of the solute is present as $\text{CH}_3\text{COOH}(\text{aq})$ molecules. Only a small fraction (about 1%) of the CH_3COOH is present as $\text{H}^+(\text{aq})$ and $\text{CH}_3\text{COO}^-(\text{aq})$ ions.*

We must be careful not to confuse the extent to which an electrolyte dissolves with whether it is strong or weak. For example, CH_3COOH is extremely soluble in water but is a weak electrolyte. $\text{Ba}(\text{OH})_2$, on the other hand, is not very soluble, but the amount of the substance that does dissolve dissociates almost completely. Thus, $\text{Ba}(\text{OH})_2$ is a strong electrolyte.

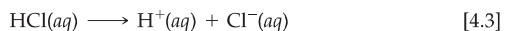
*The chemical formula of acetic acid is sometimes written as $\text{HC}_2\text{H}_3\text{O}_2$ with the acidic H written in front of the chemical formula, so the formula looks like that of other common acids such as HCl . The formula CH_3COOH conforms to the molecular structure of acetic acid, with the acidic H on the O atom at the end of the formula.

When a weak electrolyte such as acetic acid ionizes in solution, we write the reaction in the following manner:



The half-arrows in both directions mean that the reaction is significant in both directions. At any given moment some CH_3COOH molecules are ionizing to form H^+ and CH_3COO^- ions. At the same time, H^+ and CH_3COO^- ions are recombining to form CH_3COOH . The balance between these opposing processes determines the relative numbers of ions and neutral molecules. This balance produces a state of **chemical equilibrium** in which the relative numbers of each type of ion or molecule in the reaction are constant over time. This equilibrium condition varies from one weak electrolyte to another. Chemical equilibria are extremely important. We will devote Chapters 15–17 to examining them in detail.

Chemists use half-arrows in both directions to represent the ionization of weak electrolytes and a single arrow to represent the ionization of strong electrolytes. Because HCl is a strong electrolyte, we write the equation for the ionization of HCl as follows:



The absence of a reverse arrow indicates that the H^+ and Cl^- ions have no tendency to recombine in water to form HCl molecules.

In the following sections we will look more closely at how we can use the composition of a compound to predict whether it is a strong electrolyte, weak electrolyte, or nonelectrolyte. For the moment, you need only to remember that *soluble ionic compounds are strong electrolytes*. We identify ionic compounds as those composed of metals and nonmetals—such as NaCl, FeSO_4 , and $\text{Al}(\text{NO}_3)_3$ —or compounds containing the ammonium ion, NH_4^+ —such as NH_4Br and $(\text{NH}_4)_2\text{CO}_3$.

GIVE IT SOME THOUGHT

Which solute will cause the lightbulb in the experiment shown in Figure 4.2 to glow more brightly, CH_3OH or MgBr_2 ?

SAMPLE EXERCISE 4.1 Relating Relative Numbers of Anions and Cations to Chemical Formulas

The diagram on the right represents an aqueous solution of one of the following compounds: MgCl_2 , KCl, or K_2SO_4 . Which solution does the drawing best represent?

SOLUTION

Analyze: We are asked to associate the charged spheres in the diagram with ions present in a solution of an ionic substance.

Plan: We examine the ionic substances given in the problem to determine the relative numbers and charges of the ions that each contains. We then correlate these charged ionic species with the ones shown in the diagram.

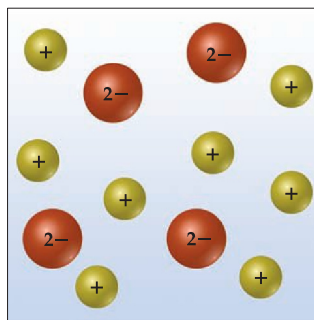
Solve: The diagram shows twice as many cations as anions, consistent with the formulation K_2SO_4 .

Check: Notice that the total net charge in the diagram is zero, as it must be if it is to represent an ionic substance.

PRACTICE EXERCISE

If you were to draw diagrams (such as that shown on the right) representing aqueous solutions of each of the following ionic compounds, how many anions would you show if the diagram contained six cations? (a) NiSO_4 , (b) $\text{Ca}(\text{NO}_3)_2$, (c) Na_3PO_4 , (d) $\text{Al}_2(\text{SO}_4)_3$

Answers: (a) 6, (b) 12, (c) 2, (d) 9



4.2 PRECIPITATION REACTIONS

Figure 4.4 ▼ shows two clear solutions being mixed. One solution contains lead nitrate, $\text{Pb}(\text{NO}_3)_2$, and the other contains potassium iodide (KI). The reaction between these two solutes produces an insoluble yellow product. Reactions that result in the formation of an insoluble product are called **precipitation reactions**. A **precipitate** is an insoluble solid formed by a reaction in solution. In Figure 4.4 the precipitate is lead iodide (PbI_2), a compound that has a very low solubility in water:



The other product of this reaction, potassium nitrate (KNO_3), remains in solution.

Precipitation reactions occur when certain pairs of oppositely charged ions attract each other so strongly that they form an insoluble ionic solid. To predict whether certain combinations of ions form insoluble compounds, we must consider some guidelines concerning the solubilities of common ionic compounds.

Solubility Guidelines for Ionic Compounds

The **solubility** of a substance at a given temperature is the amount of the substance that can be dissolved in a given quantity of solvent at the given temperature.

▼ Figure 4.4 A precipitation reaction.

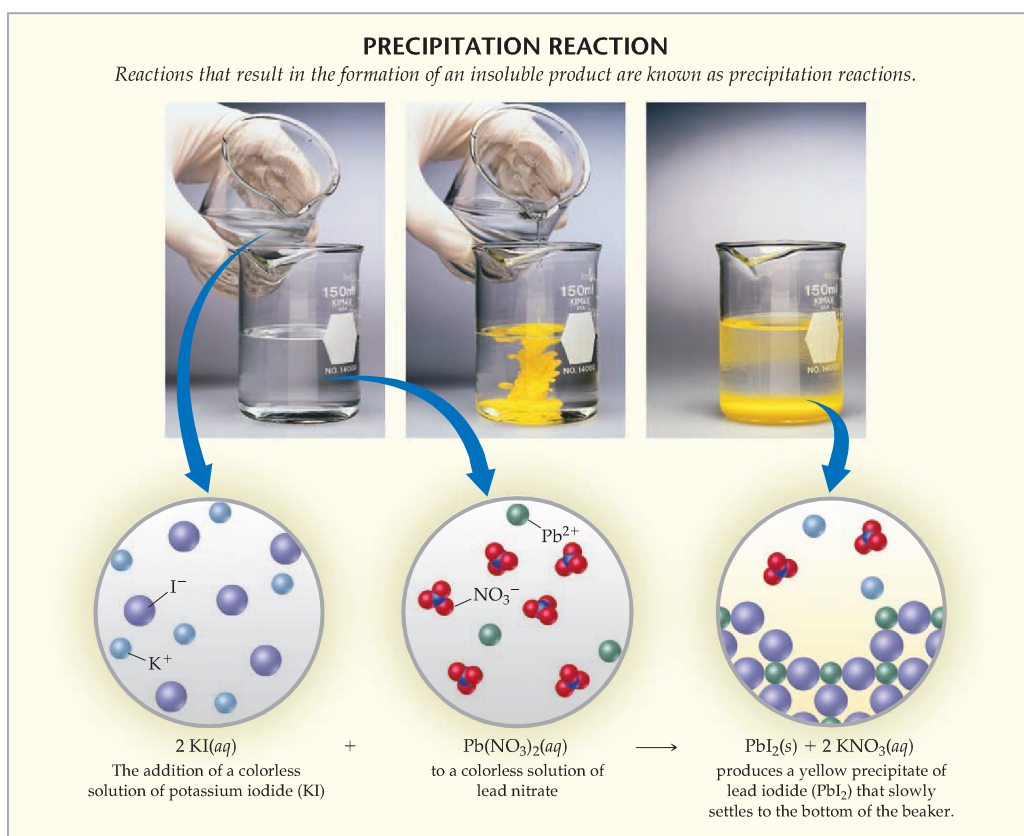


TABLE 4.1 ■ Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+}
	Br^-	Compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+}
	I^-	Compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+} , Ba^{2+} , Hg_2^{2+} , and Pb^{2+}
Insoluble Ionic Compounds		Important Exceptions
Compounds containing	S^{2-}	Compounds of NH_4^+ , the alkali metal cations, and Ca^{2+} , Sr^{2+} , and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of the alkali metal cations, and NH_4^+ , Ca^{2+} , Sr^{2+} , and Ba^{2+}

For instance, only 1.2×10^{-3} mol of PbI_2 dissolves in a liter of water at 25 °C. In our discussions, any substance with a solubility less than 0.01 mol/L will be referred to as *insoluble*. In those cases the attraction between the oppositely charged ions in the solid is too great for the water molecules to separate the ions to any significant extent; the substance remains largely undissolved.

Unfortunately, there are no rules based on simple physical properties such as ionic charge to guide us in predicting whether a particular ionic compound will be soluble. Experimental observations, however, have led to guidelines for predicting solubility for ionic compounds. For example, experiments show that all common ionic compounds that contain the nitrate anion, NO_3^- , are soluble in water. Table 4.1▲ summarizes the solubility guidelines for common ionic compounds. The table is organized according to the anion in the compound, but it also reveals many important facts about cations. Note that *all common ionic compounds of the alkali metal ions (group 1A of the periodic table) and of the ammonium ion (NH_4^+) are soluble in water.*

■ SAMPLE EXERCISE 4.2 | Using Solubility Rules

Classify the following ionic compounds as soluble or insoluble in water: (a) sodium carbonate (Na_2CO_3), (b) lead sulfate (PbSO_4).

SOLUTION

Analyze: We are given the names and formulas of two ionic compounds and asked to predict whether they are soluble or insoluble in water.

Plan: We can use Table 4.1 to answer the question. Thus, we need to focus on the anion in each compound because the table is organized by anions.

Solve:

(a) According to Table 4.1, most carbonates are insoluble. But carbonates of the alkali metal cations (such as sodium ion) are an exception to this rule and are soluble. Thus, Na_2CO_3 is soluble in water.

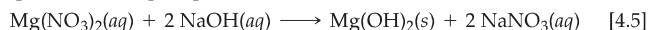
(b) Table 4.1 indicates that although most sulfates are water soluble, the sulfate of Pb^{2+} is an exception. Thus, PbSO_4 is insoluble in water.

■ PRACTICE EXERCISE

Classify the following compounds as soluble or insoluble in water: (a) cobalt(II) hydroxide, (b) barium nitrate, (c) ammonium phosphate.

Answers: (a) insoluble, (b) soluble, (c) soluble

To predict whether a precipitate forms when we mix aqueous solutions of two strong electrolytes, we must (1) note the ions present in the reactants, (2) consider the possible combinations of the cations and anions, and (3) use Table 4.1 to determine if any of these combinations is insoluble. For example, will a precipitate form when solutions of $\text{Mg}(\text{NO}_3)_2$ and NaOH are mixed? Both $\text{Mg}(\text{NO}_3)_2$ and NaOH are soluble ionic compounds and strong electrolytes. Mixing $\text{Mg}(\text{NO}_3)_2(aq)$ and $\text{NaOH}(aq)$ first produces a solution containing Mg^{2+} , NO_3^- , Na^+ , and OH^- ions. Will either of the cations interact with either of the anions to form an insoluble compound? In addition to the reactants, the other possible interactions are Mg^{2+} with OH^- and Na^+ with NO_3^- . From Table 4.1 we see that hydroxides are generally insoluble. Because Mg^{2+} is not an exception, $\text{Mg}(\text{OH})_2$ is insoluble and will thus form a precipitate. NaNO_3 , however, is soluble, so Na^+ and NO_3^- will remain in solution. The balanced equation for the precipitation reaction is



Exchange (Metathesis) Reactions

Notice in Equation 4.5 that the cations in the two reactants exchange anions— Mg^{2+} ends up with OH^- , and Na^+ ends up with NO_3^- . The chemical formulas of the products are based on the charges of the ions—two OH^- ions are needed to give a neutral compound with Mg^{2+} , and one NO_3^- ion is needed to give a neutral compound with Na^+ . ∞ (Section 2.7) The equation can be balanced only after the chemical formulas of the products have been determined.

Reactions in which positive ions and negative ions appear to exchange partners conform to the following general equation:



Such reactions are called **exchange reactions**, or **metathesis reactions** (meh-TATH-eh-sis, which is the Greek word for “to transpose”). Precipitation reactions conform to this pattern, as do many acid–base reactions, as we will see in Section 4.3.

To complete and balance a metathesis equation, we follow these steps:

1. Use the chemical formulas of the reactants to determine the ions that are present.
2. Write the chemical formulas of the products by combining the cation from one reactant with the anion of the other. (Use the charges of the ions to determine the subscripts in the chemical formulas.)
3. Finally, balance the equation.

SAMPLE EXERCISE 4.3 | Predicting a Metathesis Reaction

- (a) Predict the identity of the precipitate that forms when solutions of BaCl_2 and K_2SO_4 are mixed.
- (b) Write the balanced chemical equation for the reaction.

SOLUTION

Analyze: We are given two ionic reactants and asked to predict the insoluble product that they form.

Plan: We need to write down the ions present in the reactants and to exchange the anions between the two cations. Once we have written the chemical formulas for these products, we can use Table 4.1 to determine which is insoluble in water. Knowing the products also allows us to write the equation for the reaction.

Solve:

- (a) The reactants contain Ba^{2+} , Cl^- , K^+ , and SO_4^{2-} ions. If we exchange the anions, we will have BaSO_4 and KCl . According to Table 4.1, most compounds of SO_4^{2-} are soluble but those of Ba^{2+} are not. Thus, BaSO_4 is insoluble and will precipitate from solution. KCl , on the other hand, is soluble.

(b) From part (a) we know the chemical formulas of the products, BaSO_4 and KCl . The balanced equation with phase labels shown is



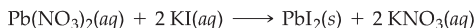
■ PRACTICE EXERCISE

(a) What compound precipitates when solutions of $\text{Fe}_2(\text{SO}_4)_3$ and LiOH are mixed? (b) Write a balanced equation for the reaction. (c) Will a precipitate form when solutions of $\text{Ba}(\text{NO}_3)_2$ and KOH are mixed?

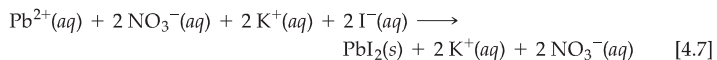
Answers: (a) $\text{Fe}(\text{OH})_3$; (b) $\text{Fe}_2(\text{SO}_4)_3(aq) + 6 \text{LiOH}(aq) \longrightarrow 2 \text{Fe}(\text{OH})_3(s) + 3 \text{Li}_2\text{SO}_4(aq)$; (c) no (both possible products are water soluble)

Ionic Equations

In writing chemical equations for reactions in aqueous solution, it is often useful to indicate explicitly whether the dissolved substances are present predominantly as ions or as molecules. Let's reconsider the precipitation reaction between $\text{Pb}(\text{NO}_3)_2$ and 2KI , shown previously in Figure 4.4:



An equation written in this fashion, showing the complete chemical formulas of the reactants and products, is called a **molecular equation** because it shows the chemical formulas of the reactants and products without indicating their ionic character. Because $\text{Pb}(\text{NO}_3)_2$, KI , and KNO_3 are all soluble ionic compounds and therefore strong electrolytes, we can write the chemical equation to indicate explicitly the ions that are in the solution:



An equation written in this form, with all soluble strong electrolytes shown as ions, is called a **complete ionic equation**.

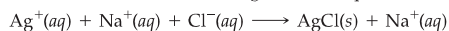
Notice that $\text{K}^+(aq)$ and $\text{NO}_3^-(aq)$ appear on both sides of Equation 4.7. Ions that appear in identical forms among both the reactants and products of a complete ionic equation are called **spectator ions**. These ions are present but play no direct role in the reaction. When spectator ions are omitted from the equation (they cancel out like algebraic quantities), we are left with the **net ionic equation**:



A net ionic equation includes only the ions and molecules directly involved in the reaction. Charge is conserved in reactions, so the sum of the charges of the ions must be the same on both sides of a balanced net ionic equation. In this case the $2+$ charge of the cation and the two $1-$ charges of the anions add to give zero, the charge of the electrically neutral product. *If every ion in a complete ionic equation is a spectator, then no reaction occurs.*

▲ GIVE IT SOME THOUGHT

Are any spectator ions shown in the following chemical equation?



Net ionic equations are widely used to illustrate the similarities between large numbers of reactions involving electrolytes. For example, Equation 4.8 expresses the essential feature of the precipitation reaction between any strong electrolyte containing Pb^{2+} and any strong electrolyte containing I^- : The $\text{Pb}^{2+}(aq)$ and $\text{I}^-(aq)$ ions combine to form a precipitate of PbI_2 . Thus, a net ionic equation demonstrates that more than one set of reactants can lead to the same net reaction. For example, aqueous solutions of KI and MgI_2 share many chemical similarities because both contain I^- ions. The complete equation, on the other hand, identifies the actual reactants that participate in a reaction.

The following steps summarize the procedure for writing net ionic equations:

1. Write a balanced molecular equation for the reaction.
2. Rewrite the equation to show the ions that form in solution when each soluble strong electrolyte dissociates into its component ions. *Only strong electrolytes dissolved in aqueous solution are written in ionic form.*
3. Identify and cancel **spectator ions**.

■ SAMPLE EXERCISE 4.4 | Writing a Net Ionic Equation

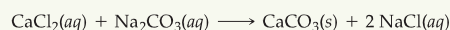
Write the net ionic equation for the precipitation reaction that occurs when solutions of calcium chloride and sodium carbonate are mixed.

SOLUTION

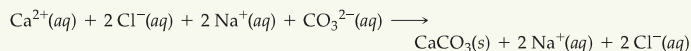
Analyze: Our task is to write a net ionic equation for a precipitation reaction, given the names of the reactants present in solution.

Plan: We first need to write the chemical formulas of the reactants and products and then determine which product is insoluble. We then write and balance the molecular equation. Next, we write each soluble strong electrolyte as separated ions to obtain the complete ionic equation. Finally, we eliminate the spectator ions to obtain the net ionic equation.

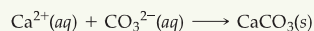
Solve: Calcium chloride is composed of calcium ions, Ca^{2+} , and chloride ions, Cl^- ; hence an aqueous solution of the substance is $\text{CaCl}_2(aq)$. Sodium carbonate is composed of Na^+ ions and CO_3^{2-} ions; hence an aqueous solution of the compound is $\text{Na}_2\text{CO}_3(aq)$. In the molecular equations for precipitation reactions, the anions and cations appear to exchange partners. Thus, we put Ca^{2+} and CO_3^{2-} together to give CaCO_3 and Na^+ and Cl^- together to give NaCl . According to the solubility guidelines in Table 4.1, CaCO_3 is insoluble and NaCl is soluble. The balanced molecular equation is



In a complete ionic equation, *only* dissolved strong electrolytes (such as soluble ionic compounds) are written as separate ions. As the *(aq)* designations remind us, CaCl_2 , Na_2CO_3 , and NaCl are all dissolved in the solution. Furthermore, they are all strong electrolytes. CaCO_3 is an ionic compound, but it is not soluble. We do not write the formula of any insoluble compound as its component ions. Thus, the complete ionic equation is



Cl^- and Na^+ are spectator ions. Canceling them gives the following net ionic equation:



Check: We can check our result by confirming that both the elements and the electric charge are balanced. Each side has one Ca, one C, and three O, and the net charge on each side equals 0.

Comment: If none of the ions in an ionic equation is removed from solution or changed in some way, then they all are spectator ions and a reaction does not occur.

■ PRACTICE EXERCISE

Write the net ionic equation for the precipitation reaction that occurs when aqueous solutions of silver nitrate and potassium phosphate are mixed.

Answers: $3 \text{Ag}^+(aq) + \text{PO}_4^{3-}(aq) \longrightarrow \text{Ag}_3\text{PO}_4(s)$



▲ Figure 4.5 Some common household acids (left) and bases (right).

4.3 ACID–BASE REACTIONS

Many acids and bases are industrial and household substances (Figure 4.5 ◀), and some are important components of biological fluids. Hydrochloric acid, for example, is an important industrial chemical and the main constituent of gastric juice in your stomach. Acids and bases are also common electrolytes.

Acids

Acids are substances that ionize in aqueous solutions to form hydrogen ions, thereby increasing the concentration of $\text{H}^+(\text{aq})$ ions. Because a hydrogen atom consists of a proton and an electron, H^+ is simply a proton. Thus, acids are often called proton donors. Molecular models of three common acids, HCl , HNO_3 and CH_3COOH , are shown in the margin.

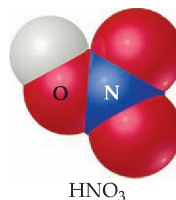
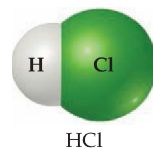
Just as cations are surrounded and bound by water molecules (see Figure 4.3[a]), the proton is also solvated by water molecules. The nature of the proton in water is discussed in detail in Section 16.2. In writing chemical equations involving the proton in water, we represent it simply as $\text{H}^+(\text{aq})$.

Molecules of different acids can ionize to form different numbers of H^+ ions. Both HCl and HNO_3 are *monoprotic* acids, which yield one H^+ per molecule of acid. Sulfuric acid, H_2SO_4 , is a *diprotic* acid, one that yields two H^+ per molecule of acid. The ionization of H_2SO_4 and other diprotic acids occurs in two steps:



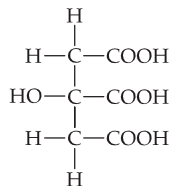
Although H_2SO_4 is a strong electrolyte, only the first ionization is complete. Thus, aqueous solutions of sulfuric acid contain a mixture of $\text{H}^+(\text{aq})$, $\text{HSO}_4^-(\text{aq})$, and $\text{SO}_4^{2-}(\text{aq})$.

The molecule CH_3COOH (acetic acid) that we have mentioned frequently is the primary component in vinegar. Acetic acid has four hydrogens, but only one of them is capable of being ionized in water. Only the hydrogen that is bound to oxygen in the COOH group will ionize in water; the other hydrogens are bound to carbon and do not break their $\text{C}-\text{H}$ bonds in water. We will discuss acids much more in Chapter 16.



GIVE IT SOME THOUGHT

The structural formula of citric acid, a main component of citrus fruits, is shown here:

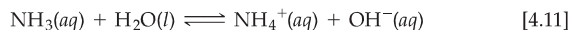


How many $\text{H}^+(\text{aq})$ can be generated by each citric acid molecule when citric acid is dissolved in water?

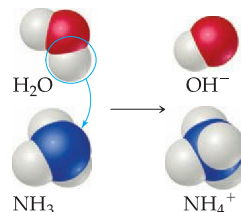
Bases

Bases are substances that accept (react with) H^+ ions. Bases produce hydroxide ions (OH^-) when they dissolve in water. Ionic hydroxide compounds such as NaOH , KOH , and $\text{Ca}(\text{OH})_2$ are among the most common bases. When dissolved in water, they dissociate into their component ions, introducing OH^- ions into the solution.

Compounds that do not contain OH^- ions can also be bases. For example, ammonia (NH_3) is a common base. When added to water, it accepts an H^+ ion from the water molecule and thereby produces an OH^- ion (Figure 4.6):



Ammonia is a weak electrolyte because only a small fraction of the NH_3 (about 1%) forms NH_4^+ and OH^- ions.



▲ Figure 4.6 Hydrogen ion transfer. An H_2O molecule acts as a proton donor (acid), and NH_3 as a proton acceptor (base). Only a fraction of the NH_3 reacts with H_2O ; NH_3 is a weak electrolyte.

TABLE 4.2 ■ Common Strong Acids and Bases

Strong Acids	Strong Bases
Hydrochloric, HCl	Group 1A metal hydroxides (LiOH, NaOH, KOH, RbOH, CsOH)
Hydrobromic, HBr	Heavy group 2A metal hydroxides [Ca(OH) ₂ , Sr(OH) ₂ , Ba(OH) ₂]
Hydroiodic, HI	
Chloric, HClO ₃	
Perchloric, HClO ₄	
Nitric, HNO ₃	
Sulfuric, H ₂ SO ₄	

Strong and Weak Acids and Bases

Acids and bases that are strong electrolytes (completely ionized in solution) are called **strong acids** and **strong bases**. Those that are weak electrolytes (partly ionized) are called **weak acids** and **weak bases**. Strong acids are more reactive than weak acids when the reactivity depends only on the concentration of $\text{H}^+(\text{aq})$. The reactivity of an acid, however, can depend on the anion as well as on $\text{H}^+(\text{aq})$. For example, hydrofluoric acid (HF) is a weak acid (only partly ionized in aqueous solution), but it is very reactive and vigorously attacks many substances, including glass. This reactivity is due to the combined action of $\text{H}^+(\text{aq})$ and $\text{F}^-(\text{aq})$.

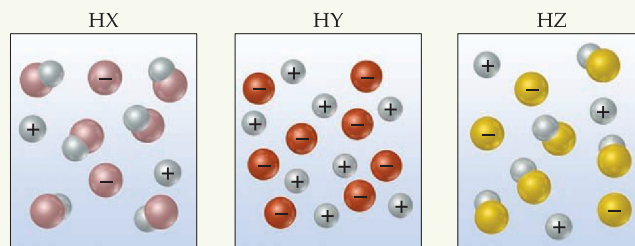
Table 4.2 ▲ lists the common strong acids and bases. You should commit these to memory. As you examine this table, notice that some of the most common acids, such as HCl, HNO₃, and H₂SO₄, are strong. (For H₂SO₄, as we noted earlier, only the first proton completely ionizes.) Three of the strong acids are the hydrogen compounds of the halogen family. (HF, however, is a weak acid.) The list of strong acids is very short. Most acids are weak. The only common strong bases are the hydroxides of Li⁺, Na⁺, K⁺, Rb⁺, and Cs⁺ (the alkali metals, group 1A) and the hydroxides of Ca²⁺, Sr²⁺, and Ba²⁺ (the heavy alkaline earths, group 2A). These are the common soluble metal hydroxides. Most other metal hydroxides are insoluble in water. The most common weak base is NH₃, which reacts with water to form OH⁻ ions (Equation 4.11).

GIVE IT SOME THOUGHT

Which of the following is a strong acid: H₂SO₃, HBr, CH₃COOH?

SAMPLE EXERCISE 4.5 | Comparing Acid Strengths

The following diagrams represent aqueous solutions of three acids (HX, HY, and HZ) with water molecules omitted for clarity. Rank them from strongest to weakest.



SOLUTION

Analyze: We are asked to rank three acids from strongest to weakest, based on schematic drawings of their solutions.

Plan: We can examine the drawings to determine the relative numbers of uncharged molecular species present. The strongest acid is the one with the most H^+ ions and fewest undissociated acid molecules in solution. The weakest acid is the one with the largest number of undissociated molecules.

Solve: The order is $\text{HY} > \text{HZ} > \text{HX}$. HY is a strong acid because it is totally ionized (no HY molecules in solution), whereas both HX and HZ are weak acids, whose solutions consist of a mixture of molecules and ions. Because HZ contains more H^+ ions and fewer molecules than HX, it is a stronger acid.

PRACTICE EXERCISE

Imagine a diagram showing 10 Na^+ ions and 10 OH^- ions. If this solution were mixed with the one pictured on the previous page for HY, what would the diagram look like that represents the solution after any possible reaction? (H^+ ions will react with OH^- ions to form H_2O .)

Answer: The final diagram would show 10 Na^+ ions, 2 OH^- ions, 8 Y^- ions, and 8 H_2O molecules.

Identifying Strong and Weak Electrolytes

If we remember the common strong acids and bases (Table 4.2) and also remember that NH_3 is a weak base, we can make reasonable predictions about the electrolytic strength of a great number of water-soluble substances. Table 4.3 summarizes our observations about electrolytes. To classify a soluble substance as a strong electrolyte, weak electrolyte, or nonelectrolyte, we simply work our way down and across this table. We first ask ourselves whether the substance is ionic or molecular. If it is ionic, it is a strong electrolyte. The second column of Table 4.3 tells us that all ionic compounds are strong electrolytes. If the substance we want to classify is molecular, we ask whether it is an acid or a base. (Does it have H first in the chemical formula or contain a COOH group?) If it is an acid, we rely on the memorized list from Table 4.2 to determine whether it is a strong or weak electrolyte: All strong acids are strong electrolytes, and all weak acids are weak electrolytes. If an acid is not listed in Table 4.2, it is probably a weak acid and therefore a weak electrolyte. For example, H_3PO_4 , H_2SO_3 , and $\text{HC}_7\text{H}_5\text{O}_2$ are not listed in Table 4.2 and are weak acids. If the substance we want to classify is a base, we again turn to Table 4.2 to determine whether it is one of the listed strong bases. NH_3 is the only molecular base that we consider in this chapter, and Table 4.3 tells us it is a weak electrolyte. (There are compounds called amines that are related to NH_3 and are also molecular bases, but we will not consider them until Chapter 16.) Finally, any molecular substance that we encounter in this chapter that is not an acid or NH_3 is probably a nonelectrolyte.

TABLE 4.3 Summary of the Electrolytic Behavior of Common Soluble Ionic and Molecular Compounds

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids Weak bases	All other compounds

SAMPLE EXERCISE 4.6 | Identifying Strong, Weak, and Nonelectrolytes

Classify each of the following dissolved substances as a strong electrolyte, weak electrolyte, or nonelectrolyte: CaCl_2 , HNO_3 , $\text{C}_2\text{H}_5\text{OH}$ (ethanol), HCOOH (formic acid), KOH .

SOLUTION

Analyze: We are given several chemical formulas and asked to classify each substance as a strong electrolyte, weak electrolyte, or nonelectrolyte.

Plan: The approach we take is outlined in Table 4.3. We can predict whether a substance is ionic or molecular based on its composition. As we saw in Section 2.7, most ionic compounds we encounter in this text are composed of a metal and a nonmetal, whereas most molecular compounds are composed only of nonmetals.

Solve: Two compounds fit the criteria for ionic compounds: CaCl_2 and KOH . Because Table 4.3 tells us that all ionic compounds are strong electrolytes, that is how we classify these two substances. The three remaining compounds are molecular. Two, HNO_3 and HCOOH , are acids. Nitric acid, HNO_3 , is a common strong acid, as shown in Table 4.2, and therefore is a strong electrolyte. Because most acids are weak acids, our best guess would be that HCOOH is a weak acid (weak electrolyte). This is correct. The remaining molecular compound, $\text{C}_2\text{H}_5\text{OH}$, is neither an acid nor a base, so it is a nonelectrolyte.

Comment: Although $\text{C}_2\text{H}_5\text{OH}$ has an OH group, it is not a metal hydroxide; thus, it is not a base. Rather, it is a member of a class of organic compounds that have C—OH bonds, which are known as alcohols. ∞ (Section 2.9). The COOH group is called the “carboxylic acid group” (Chapter 16). Molecules that have this group are all weak acids.

PRACTICE EXERCISE

Consider solutions in which 0.1 mol of each of the following compounds is dissolved in 1 L of water: $\text{Ca}(\text{NO}_3)_2$ (calcium nitrate), $\text{C}_6\text{H}_{12}\text{O}_6$ (glucose), CH_3COONa (sodium acetate), and CH_3COOH (acetic acid). Rank the solutions in order of increasing electrical conductivity, based on the fact that the greater the number of ions in solution, the greater the conductivity.

Answers: $\text{C}_6\text{H}_{12}\text{O}_6$ (nonelectrolyte) < CH_3COOH (weak electrolyte, existing mainly in the form of molecules with few ions) < CH_3COONa (strong electrolyte that provides two ions, Na^+ and CH_3COO^-) < $\text{Ca}(\text{NO}_3)_2$ (strong electrolyte that provides three ions, Ca^{2+} and 2NO_3^-)



▲ Figure 4.7 The acid–base indicator bromthymol blue. The indicator is blue in basic solution and yellow in acidic solution. The left flask shows the indicator in the presence of a base, aqueous ammonia (labeled as ammonium hydroxide). The right flask shows the indicator in the presence of hydrochloric acid, HCl.

Neutralization Reactions and Salts

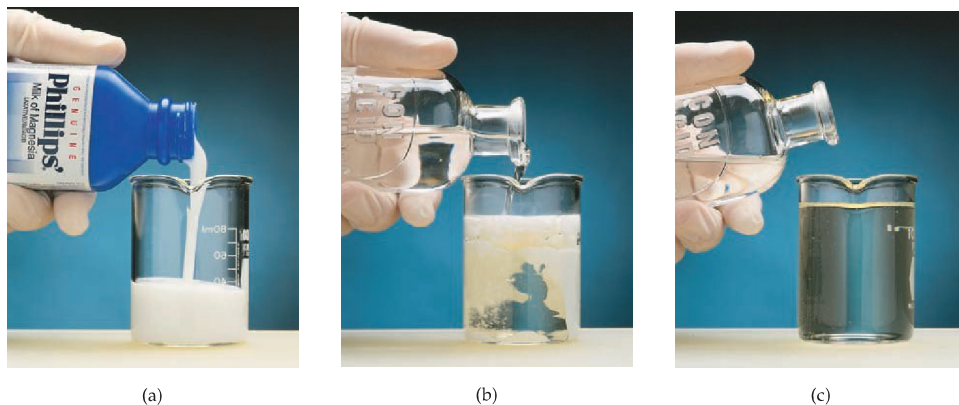
The properties of acidic solutions are quite different from those of basic solutions. Acids have a sour taste, whereas bases have a bitter taste.* Acids can change the colors of certain dyes in a specific way that differs from the effect of a base (Figure 4.7 ◀). The dye known as litmus, for example, is changed from blue to red by an acid and from red to blue by a base. In addition, acidic and basic solutions differ in chemical properties in several important ways that we will explore in this chapter and in later chapters.

When a solution of an acid and a solution of a base are mixed, a **neutralization reaction** occurs. The products of the reaction have none of the characteristic properties of either the acidic solution or the basic solution. For example, when hydrochloric acid is mixed with a solution of sodium hydroxide, the following reaction occurs:



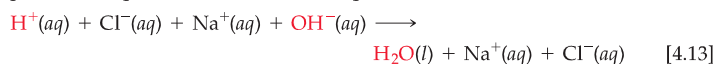
Water and table salt, NaCl, are the products of the reaction. By analogy to this reaction, the term **salt** has come to mean any ionic compound whose cation comes from a base (for example, Na^+ from NaOH) and whose anion comes from an acid (for example, Cl^- from HCl). In general, a *neutralization reaction between an acid and a metal hydroxide produces water and a salt.*

**Tasting chemical solutions is not a good practice. However, we have all had acids such as ascorbic acid (vitamin C), acetylsalicylic acid (aspirin), and citric acid (in citrus fruits) in our mouths, and we are familiar with their characteristic sour taste. Soaps, which are basic, have the characteristic bitter taste of bases.*



▲ **Figure 4.8 Reaction of $\text{Mg}(\text{OH})_2(\text{s})$ with acid.** (a) Milk of magnesia is a suspension of magnesium hydroxide, $\text{Mg}(\text{OH})_2(\text{s})$, in water. (b) The magnesium hydroxide dissolves upon the addition of hydrochloric acid, $\text{HCl}(\text{aq})$. (c) The final clear solution contains soluble $\text{MgCl}_2(\text{aq})$, as shown in Equation 4.15.

Because HCl , NaOH , and NaCl are all soluble strong electrolytes, the complete ionic equation associated with Equation 4.12 is



Therefore, the net ionic equation is



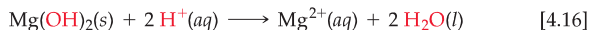
Equation 4.14 summarizes the essential feature of the neutralization reaction between any strong acid and any strong base: $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ ions combine to form H_2O .

Figure 4.8▲ shows the reaction between hydrochloric acid and the base $\text{Mg}(\text{OH})_2$, which is insoluble in water. A milky white suspension of $\text{Mg}(\text{OH})_2$ called milk of magnesia is seen dissolving as the neutralization reaction occurs:

Molecular equation:



Net ionic equation:



Notice that the OH^- ions (this time in a solid reactant) and H^+ ions combine to form H_2O . Because the ions exchange partners, neutralization reactions between acids and metal hydroxides are also metathesis reactions.

■ SAMPLE EXERCISE 4.7 | Writing Chemical Equations for a Neutralization Reaction

(a) Write a balanced molecular equation for the reaction between aqueous solutions of acetic acid (CH_3COOH) and barium hydroxide, $\text{Ba}(\text{OH})_2$. (b) Write the net ionic equation for this reaction.

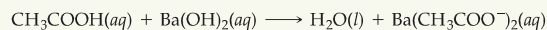
SOLUTION

Analyze: We are given the chemical formulas for an acid and a base and asked to write a balanced molecular equation and then a net ionic equation for their neutralization reaction.

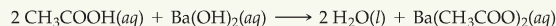
Plan: As Equation 4.12 and the italicized statement that follows it indicate, neutralization reactions form two products, H_2O and a salt. We examine the cation of the base and the anion of the acid to determine the composition of the salt.

Solve:

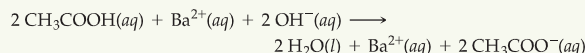
(a) The salt will contain the cation of the base (Ba^{2+}) and the anion of the acid (CH_3COO^-). Thus, the formula of the salt is $\text{Ba}(\text{CH}_3\text{COO})_2$. According to the solubility guidelines in Table 4.1, this compound is soluble. The unbalanced molecular equation for the neutralization reaction is



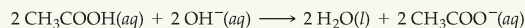
To balance this molecular equation, we must provide two molecules of CH_3COOH to furnish the two CH_3COO^- ions and to supply the two H^+ ions needed to combine with the two OH^- ions of the base. The balanced molecular equation is



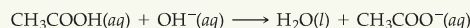
(b) To write the net ionic equation, we must determine whether each compound in aqueous solution is a strong electrolyte. CH_3COOH is a weak electrolyte (weak acid), $\text{Ba}(\text{OH})_2$ is a strong electrolyte, and $\text{Ba}(\text{CH}_3\text{COO})_2$ is also a strong electrolyte (ionic compound). Thus, the complete ionic equation is



Eliminating the spectator ions gives



Simplifying the coefficients gives the net ionic equation:



Check: We can determine whether the molecular equation is correctly balanced by counting the number of atoms of each kind on both sides of the arrow. (There are 10 H, 6 O, 4 C, and 1 Ba on each side.) However, it is often easier to check equations by counting groups: There are 2 CH_3COO groups, as well as 1 Ba, and 4 additional H atoms and 2 additional O atoms on each side of the equation. The net ionic equation checks out because the numbers of each kind of element and the net charge are the same on both sides of the equation.

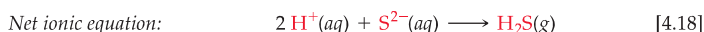
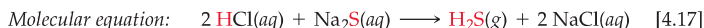
PRACTICE EXERCISE

(a) Write a balanced molecular equation for the reaction of carbonic acid (H_2CO_3) and potassium hydroxide (KOH). (b) Write the net ionic equation for this reaction.

Answers: (a) $\text{H}_2\text{CO}_3(aq) + 2 \text{KOH}(aq) \longrightarrow 2 \text{H}_2\text{O}(l) + \text{K}_2\text{CO}_3(aq)$; (b) $\text{H}_2\text{CO}_3(aq) + 2 \text{OH}^-(aq) \longrightarrow 2 \text{H}_2\text{O}(l) + \text{CO}_3^{2-}(aq)$. (H_2CO_3 is a weak acid and therefore a weak electrolyte, whereas KOH, a strong base, and K_2CO_3 , an ionic compound, are strong electrolytes.)

Acid–Base Reactions with Gas Formation

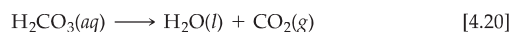
Many bases besides OH^- react with H^+ to form molecular compounds. Two of these that you might encounter in the laboratory are the sulfide ion and the carbonate ion. Both of these anions react with acids to form gases that have low solubilities in water. Hydrogen sulfide (H_2S), the substance that gives rotten eggs their foul odor, forms when an acid such as $\text{HCl}(aq)$ reacts with a metal sulfide such as Na_2S :



Carbonates and bicarbonates react with acids to form CO_2 gas. Reaction of CO_3^{2-} or HCO_3^- with an acid first gives carbonic acid (H_2CO_3). For example, when hydrochloric acid is added to sodium bicarbonate, the following reaction occurs:

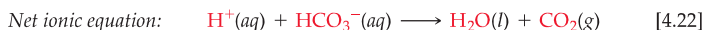


Carbonic acid is unstable. If carbonic acid is present in solution in sufficient concentrations, it decomposes to form H_2O and CO_2 , which escapes from the solution as a gas.



The decomposition of H_2CO_3 produces bubbles of CO_2 gas, as shown in Figure 4.9. The overall reaction is summarized by the following equations:

Molecular equation:



▲ **Figure 4.9** Carbonates react with acids to form carbon dioxide gas.

Here NaHCO_3 (white solid) reacts with hydrochloric acid. The bubbles contain CO_2 .

Chemistry Put to Work ANTACIDS

Your stomach secretes acids to help digest foods. These acids, which include hydrochloric acid, contain about 0.1 mol of H^+ per liter of solution. The stomach and digestive tract are normally protected from the corrosive effects of stomach acid by a mucosal lining. Holes can develop in this lining, however, allowing the acid to attack the underlying tissue, causing painful damage. These holes, known as ulcers, can be caused by the secretion of excess acids or by a weakness in the digestive lining. Studies indicate, however, that many ulcers are caused by bacterial infection. Between 10 and 20% of Americans suffer from ulcers at some point in their lives. Many others experience occasional indigestion or heartburn that is due to digestive acids entering the esophagus.

We can address the problem of excess stomach acid in two simple ways: (1) removing the excess acid, or (2) decreasing the production of acid. Those substances that remove ex-

cess acid are called *antacids*, whereas those that decrease the production of acid are called *acid inhibitors*. Figure 4.10 shows several common, over-the-counter antacids.

Antacids are simple bases that neutralize digestive acids. They are able to neutralize acids because they contain hydroxide, carbonate, or bicarbonate ions. Table 4.4 lists the active ingredients in some antacids.

The newer generation of antiulcer drugs, such as Tagamet[®] and Zantac[®], are acid inhibitors. They act on acid-producing cells in the lining of the stomach. Formulations that control acid in this way are now available as over-the-counter drugs.

Related Exercise: 4.95

TABLE 4.4 ■ Some Common Antacids

Commercial Name	Acid-Neutralizing Agents
Alka-Seltzer [®]	$NaHCO_3$
Amphojel [®]	$Al(OH)_3$
Di-Gel [®]	$Mg(OH)_2$ and $CaCO_3$
Milk of Magnesia	$Mg(OH)_2$
Maalox [®]	$Mg(OH)_2$ and $Al(OH)_3$
Mylanta [®]	$Mg(OH)_2$ and $Al(OH)_3$
Rolaids [®]	$NaAl(OH)_2CO_3$
Tums [®]	$CaCO_3$



▲ **Figure 4.10 Antacids.** These products all serve as acid-neutralizing agents in the stomach.

Both $NaHCO_3$ and Na_2CO_3 are used as acid neutralizers in acid spills. The bicarbonate or carbonate salt is added until the fizzing due to the formation of $CO_2(g)$ stops. Sometimes sodium bicarbonate is used as an antacid to soothe an upset stomach. In that case the HCO_3^- reacts with stomach acid to form $CO_2(g)$. The fizz when Alka-Seltzer[®] tablets are added to water arises from the reaction of sodium bicarbonate and citric acid.

GIVE IT SOME THOUGHT

By analogy to examples already given in the text, predict what gas forms when $Na_2SO_3(s)$ is treated with $HCl(aq)$.

4.4 OXIDATION-REDUCTION REACTIONS

In precipitation reactions, cations and anions come together to form an insoluble ionic compound. In neutralization reactions H^+ ions and OH^- ions come together to form H_2O molecules. Now let's consider a third important kind of reaction, one in which electrons are transferred between reactants. Such reactions are called **oxidation-reduction**, or *redox*, reactions.

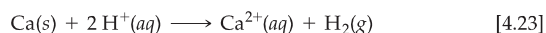
► **Figure 4.11 Corrosion of iron.**

Corrosion of iron is caused by chemical attack of oxygen and water on exposed metal surfaces. Corrosion is even more rapid in salt water.

**Oxidation and Reduction**

The corrosion of iron (rusting) and of other metals, such as the corrosion of the terminals of an automobile battery, are familiar processes. What we call *corrosion* is the conversion of a metal into a metal compound by a reaction between the metal and some substance in its environment. Rusting, as shown in Figure 4.11 ▲, involves the reaction of oxygen with iron in the presence of water.

When a metal corrodes, it loses electrons and forms cations. Calcium, for example, is vigorously attacked by acids to form calcium ions:

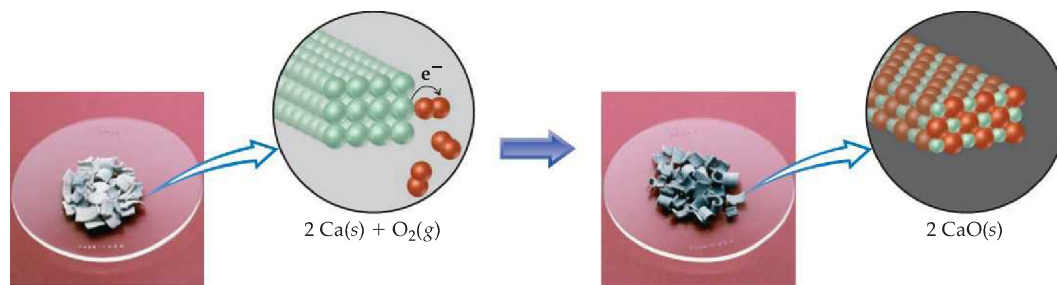


When an atom, ion, or molecule has become more positively charged (that is, when it has lost electrons), we say that it has been oxidized. *Loss of electrons by a substance is called oxidation.* Thus, Ca, which has no net charge, is *oxidized* (undergoes oxidation) in Equation 4.23, forming Ca^{2+} .

The term oxidation is used because the first reactions of this sort to be studied thoroughly were reactions with oxygen. Many metals react directly with O_2 in air to form metal oxides. In these reactions the metal loses electrons to oxygen, forming an ionic compound of the metal ion and oxide ion. For example, when calcium metal is exposed to air, the bright metallic surface of the metal tarnishes as CaO forms:

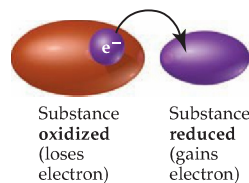


As Ca is oxidized in Equation 4.24, oxygen is transformed from neutral O_2 to two O^{2-} ions (Figure 4.12 ▼). When an atom, ion, or molecule has become more negatively charged (gained electrons), we say that it is *reduced*.



▲ **Figure 4.12 Oxidation of calcium metal by molecular oxygen.** The oxidation involves transfer of electrons from the metal to O_2 , eventually leading to formation of CaO .

The gain of electrons by a substance is called **reduction**. When one reactant loses electrons (that is, when it is *oxidized*), another reactant must gain them. The oxidation of one substance is always accompanied by the reduction of another as electrons are transferred between them.



Oxidation Numbers

Before we can properly identify an oxidation-reduction reaction, we must have a kind of bookkeeping system—a way of keeping track of the electrons gained by the substance reduced and those lost by the substance oxidized. The concept of oxidation numbers (also called *oxidation states*) was devised as a way of doing this. Each atom in a neutral molecule or charged species is assigned an **oxidation number**, which is the actual charge for a monatomic ion. Otherwise, the oxidation number is the hypothetical charge assigned to the atom, assuming that the electrons are *completely* held by one atom or the other. The oxidation numbers of certain atoms change in an oxidation-reduction reaction. Oxidation occurs when the oxidation number increases, whereas reduction occurs when the oxidation number decreases.

We use the following rules for assigning oxidation numbers:

1. For an atom in its **elemental form**, the oxidation number is always zero.
Thus, each H atom in the H_2 molecule has an oxidation number of 0, and each P atom in the P_4 molecule has an oxidation number of 0.
2. For any **monatomic ion** the oxidation number equals the charge on the ion.
Thus, K^+ has an oxidation number of +1, S^{2-} has an oxidation number of -2 , and so forth. The alkali metal ions (group 1A) always have a $1+$ charge, and therefore the alkali metals always have an oxidation number of $+1$ in their compounds. Similarly, the alkaline earth metals (group 2A) are always $+2$, and aluminum (group 3A) is always $+3$ in its compounds. (In writing oxidation numbers, we will write the sign before the number to distinguish them from the actual electronic charges, which we write with the number first.)
3. **Nonmetals** usually have negative oxidation numbers, although they can sometimes be positive:
 - (a) The oxidation number of **oxygen** is usually -2 in both ionic and molecular compounds. The major exception is in compounds called peroxides, which contain the O_2^{2-} ion, giving each oxygen an oxidation number of -1 .
 - (b) The oxidation number of **hydrogen** is usually $+1$ when bonded to nonmetals and -1 when bonded to metals.
 - (c) The oxidation number of **fluorine** is -1 in all compounds. The other **halogens** have an oxidation number of -1 in most binary compounds. When combined with oxygen, as in oxyanions, however, they have positive oxidation states.
4. **The sum of the oxidation numbers of all atoms in a neutral compound is zero.** The sum of the oxidation numbers in a polyatomic ion equals the charge of the ion.
For example, in the hydronium ion, H_3O^+ , the oxidation number of each hydrogen is $+1$ and that of oxygen is -2 . Thus, the sum of the oxidation numbers is $3(+1) + (-2) = +1$, which equals the net charge of the ion. This rule is very useful in obtaining the oxidation number of one atom in a compound or ion if you know the oxidation numbers of the other atoms, as illustrated in Sample Exercise 4.8.

GIVE IT SOME THOUGHT

- (a) What noble gas element has the same number of electrons as the fluoride ion?
- (b) What is the oxidation number of that noble gas?

SAMPLE EXERCISE 4.8 | Determining Oxidation Numbers

Determine the oxidation number of sulfur in each of the following: (a) H_2S , (b) S_8 , (c) SCl_2 , (d) Na_2SO_3 , (e) SO_4^{2-} .

SOLUTION

Analyze: We are asked to determine the oxidation number of sulfur in two molecular species, in the elemental form, and in two ionic substances.

Plan: In each species the sum of oxidation numbers of all the atoms must equal the charge on the species. We will use the rules outlined above to assign oxidation numbers.

Solve:

(a) When bonded to a nonmetal, hydrogen has an oxidation number of +1 (rule 3b). Because the H_2S molecule is neutral, the sum of the oxidation numbers must equal zero (rule 4). Letting x equal the oxidation number of S, we have $2(+1) + x = 0$. Thus, S has an oxidation number of -2 .

(b) Because this is an elemental form of sulfur, the oxidation number of S is 0 (rule 1).

(c) Because this is a binary compound, we expect chlorine to have an oxidation number of -1 (rule 3c). The sum of the oxidation numbers must equal zero (rule 4). Letting x equal the oxidation number of S, we have $x + 2(-1) = 0$. Consequently, the oxidation number of S must be $+2$.

(d) Sodium, an alkali metal, always has an oxidation number of $+1$ in its compounds (rule 2). Oxygen has a common oxidation state of -2 (rule 3a). Letting x equal the oxidation number of S, we have $2(+1) + x + 3(-2) = 0$. Therefore, the oxidation number of S in this compound is $+4$.

(e) The oxidation state of O is -2 (rule 3a). The sum of the oxidation numbers equals -2 , the net charge of the SO_4^{2-} ion (rule 4). Thus, we have $x + 4(-2) = -2$. From this relation we conclude that the oxidation number of S in this ion is $+6$.

Comment: These examples illustrate that the oxidation number of a given element depends on the compound in which it occurs. The oxidation numbers of sulfur, as seen in these examples, range from -2 to $+6$.

PRACTICE EXERCISE

What is the oxidation state of the boldfaced element in each of the following: (a) P_2O_5 , (b) NaH , (c) $\text{Cr}_2\text{O}_7^{2-}$, (d) SnBr_4 , (e) BaO_2 ?

Answers: (a) $+5$, (b) -1 , (c) $+6$, (d) $+4$, (e) -1

Oxidation of Metals by Acids and Salts

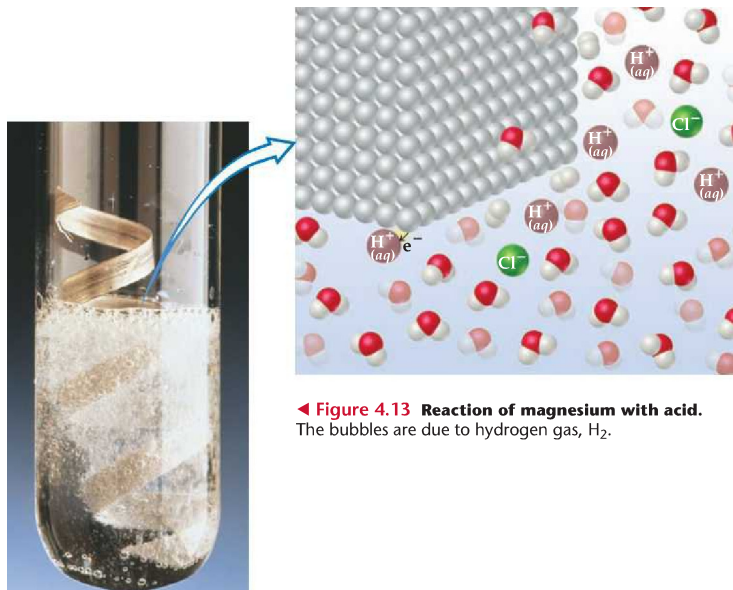
There are many kinds of redox reactions. For example, combustion reactions are redox reactions because elemental oxygen is converted to compounds of oxygen. [Section 3.2](#) In this chapter we consider the redox reactions between metals and either acids or salts. In Chapter 20 we will examine more complex kinds of redox reactions.

The reaction of a metal with either an acid or a metal salt conforms to the following general pattern:



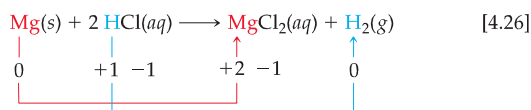
These reactions are called **displacement reactions** because the ion in solution is displaced or replaced through oxidation of an element.

Many metals undergo displacement reactions with acids, producing salts and hydrogen gas. For example, magnesium metal reacts with hydrochloric acid to form magnesium chloride and hydrogen gas (Figure 4.13▶).

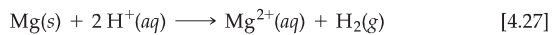


◀ **Figure 4.13** Reaction of magnesium with acid. The bubbles are due to hydrogen gas, H_2 .

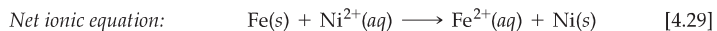
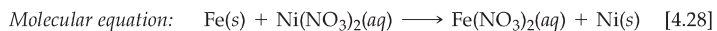
To show that oxidation and reduction have occurred, the oxidation number for each atom is shown below the chemical equation for this reaction:



Notice that the oxidation number of Mg changes from 0 to +2. The increase in the oxidation number indicates that the atom has lost electrons and has therefore been oxidized. The H^+ ion of the acid decreases in oxidation number from +1 to 0, indicating that this ion has gained electrons and has therefore been reduced. The oxidation number of the Cl^- ion remains -1 , and it is a spectator ion in the reaction. The net ionic equation is as follows:



Metals can also be oxidized by aqueous solutions of various salts. Iron metal, for example, is oxidized to Fe^{2+} by aqueous solutions of Ni^{2+} such as $\text{Ni}(\text{NO}_3)_2(aq)$:



The oxidation of Fe to form Fe^{2+} in this reaction is accompanied by the reduction of Ni^{2+} to Ni. Remember: *Whenever one substance is oxidized, some other substance must be reduced.*

SAMPLE EXERCISE 4.9 Writing Molecular and Net Ionic Equations for Oxidation-Reduction Reactions

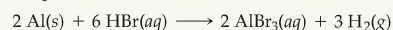
Write the balanced molecular and net ionic equations for the reaction of aluminum with hydrobromic acid.

SOLUTION

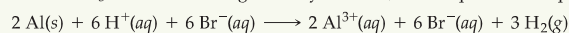
Analyze: We must write two equations—molecular and net ionic—for the redox reaction between a metal and an acid.

Plan: Metals react with acids to form salts and H_2 gas. To write the balanced equations, we must write the chemical formulas for the two reactants and then determine the formula of the salt. The salt is composed of the cation formed by the metal and the anion of the acid.

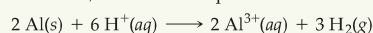
Solve: The formulas of the given reactants are Al and HBr. The cation formed by Al is Al^{3+} , and the anion from hydrobromic acid is Br^- . Thus, the salt formed in the reaction is AlBr_3 . Writing the reactants and products and then balancing the equation gives this molecular equation:



Both HBr and AlBr_3 are soluble strong electrolytes. Thus, the complete ionic equation is



Because Br^- is a spectator ion, the net ionic equation is



Comment: The substance oxidized is the aluminum metal because its oxidation state changes from 0 in the metal to +3 in the cation, thereby increasing in oxidation number. The H^+ is reduced because its oxidation state changes from +1 in the acid to 0 in H_2 .

PRACTICE EXERCISE

(a) Write the balanced molecular and net ionic equations for the reaction between magnesium and cobalt(II) sulfate. (b) What is oxidized and what is reduced in the reaction?

Answers: (a) $\text{Mg}(s) + \text{CoSO}_4(aq) \longrightarrow \text{MgSO}_4(aq) + \text{Co}(s)$;

$\text{Mg}(s) + \text{Co}^{2+}(aq) \longrightarrow \text{Mg}^{2+}(aq) + \text{Co}(s)$; (b) Mg is oxidized and Co^{2+} is reduced.

The Activity Series

Can we predict whether a certain metal will be oxidized either by an acid or by a particular salt? This question is of practical importance as well as chemical interest. According to Equation 4.28, for example, it would be unwise to store a solution of nickel nitrate in an iron container because the solution would dissolve the container. When a metal is oxidized, it reacts to form various compounds. Extensive oxidation can lead to the failure of metal machinery parts or the deterioration of metal structures.

Different metals vary in the ease with which they are oxidized. Zn is oxidized by aqueous solutions of Cu^{2+} , for example, but Ag is not. Zn, therefore, loses electrons more readily than Ag; that is, Zn is easier to oxidize than Ag.

A list of metals arranged in order of decreasing ease of oxidation is called an **activity series**. Table 4.5 gives the activity series in aqueous solution for many of the most common metals. Hydrogen is also included in the table. The metals at the top of the table, such as the alkali metals and the alkaline earth metals, are most easily oxidized; that is, they react most readily to form compounds. They are called the *active metals*. The metals at the bottom of the activity series, such as the transition elements from groups 8B and 1B, are very stable and form compounds less readily. These metals, which are used to make coins and jewelry, are called *noble metals* because of their low reactivity.

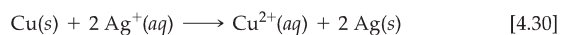
The activity series can be used to predict the outcome of reactions between metals and either metal salts or acids. *Any metal on the list can be oxidized by the ions of elements below it.* For example, copper is above silver in the series.

TABLE 4.5 Activity Series of Metals in Aqueous Solution

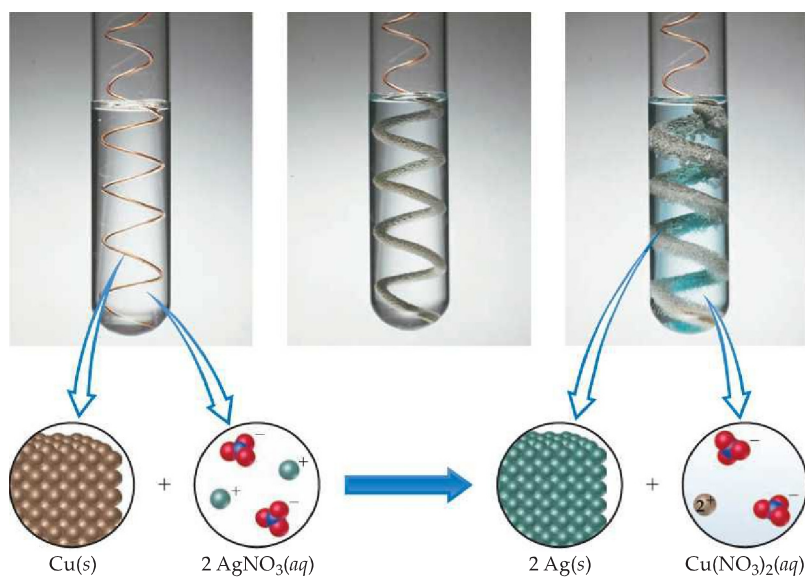
Metal	Oxidation Reaction
Lithium	$\text{Li}(s) \longrightarrow \text{Li}^+(aq) + e^-$
Potassium	$\text{K}(s) \longrightarrow \text{K}^+(aq) + e^-$
Barium	$\text{Ba}(s) \longrightarrow \text{Ba}^{2+}(aq) + 2e^-$
Calcium	$\text{Ca}(s) \longrightarrow \text{Ca}^{2+}(aq) + 2e^-$
Sodium	$\text{Na}(s) \longrightarrow \text{Na}^+(aq) + e^-$
Magnesium	$\text{Mg}(s) \longrightarrow \text{Mg}^{2+}(aq) + 2e^-$
Aluminum	$\text{Al}(s) \longrightarrow \text{Al}^{3+}(aq) + 3e^-$
Manganese	$\text{Mn}(s) \longrightarrow \text{Mn}^{2+}(aq) + 2e^-$
Zinc	$\text{Zn}(s) \longrightarrow \text{Zn}^{2+}(aq) + 2e^-$
Chromium	$\text{Cr}(s) \longrightarrow \text{Cr}^{3+}(aq) + 3e^-$
Iron	$\text{Fe}(s) \longrightarrow \text{Fe}^{2+}(aq) + 2e^-$
Cobalt	$\text{Co}(s) \longrightarrow \text{Co}^{2+}(aq) + 2e^-$
Nickel	$\text{Ni}(s) \longrightarrow \text{Ni}^{2+}(aq) + 2e^-$
Tin	$\text{Sn}(s) \longrightarrow \text{Sn}^{2+}(aq) + 2e^-$
Lead	$\text{Pb}(s) \longrightarrow \text{Pb}^{2+}(aq) + 2e^-$
Hydrogen	$\text{H}_2(g) \longrightarrow 2\text{H}^+(aq) + 2e^-$
Copper	$\text{Cu}(s) \longrightarrow \text{Cu}^{2+}(aq) + 2e^-$
Silver	$\text{Ag}(s) \longrightarrow \text{Ag}^+(aq) + e^-$
Mercury	$\text{Hg}(l) \longrightarrow \text{Hg}^{2+}(aq) + 2e^-$
Platinum	$\text{Pt}(s) \longrightarrow \text{Pt}^{2+}(aq) + 2e^-$
Gold	$\text{Au}(s) \longrightarrow \text{Au}^{3+}(aq) + 3e^-$



Thus, copper metal will be oxidized by silver ions, as pictured in Figure 4.14 ▼:



The oxidation of copper to copper ions is accompanied by the reduction of silver ions to silver metal. The silver metal is evident on the surface of the copper wires in Figure 4.14(b) and 4.14(c). The copper(II) nitrate produces a blue color in the solution, which is most evident in part (c).



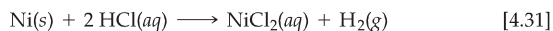
◀ **Figure 4.14 Reaction of copper with silver ion.**

When copper metal is placed in a solution of silver nitrate, a redox reaction occurs, forming silver metal and a blue solution of copper(II) nitrate.

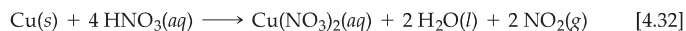
GIVE IT SOME THOUGHT

Which is the more easily reduced, $\text{Mg}^{2+}(\text{aq})$ or $\text{Ni}^{2+}(\text{aq})$?

Only those metals above hydrogen in the activity series are able to react with acids to form H_2 . For example, Ni reacts with $\text{HCl}(\text{aq})$ to form H_2 :



Because elements below hydrogen in the activity series are not oxidized by H^+ , Cu does not react with $\text{HCl}(\text{aq})$. Interestingly, copper does react with nitric acid, as shown previously in Figure 1.11. This reaction, however, is not a simple oxidation of Cu by the H^+ ions of the acid. Instead, the metal is oxidized to Cu^{2+} by the nitrate ion of the acid, accompanied by the formation of brown nitrogen dioxide, $\text{NO}_2(\text{g})$:



What substance is reduced as copper is oxidized in Equation 4.32? In this case the NO_2 results from the reduction of NO_3^- . We will examine reactions of this type in more detail in Chapter 20.

SAMPLE EXERCISE 4.10 | Determining When an Oxidation-Reduction Reaction Can Occur

Will an aqueous solution of iron(II) chloride oxidize magnesium metal? If so, write the balanced molecular and net ionic equations for the reaction.

SOLUTION

Analyze: We are given two substances—an aqueous salt, FeCl_2 , and a metal, Mg—and asked if they react with each other.

Plan: A reaction will occur if Mg is above Fe in the activity series, Table 4.5. If the reaction occurs, the Fe^{2+} ion in FeCl_2 will be reduced to Fe, and the elemental Mg will be oxidized to Mg^{2+} .

Solve: Because Mg is above Fe in the table, the reaction will occur. To write the formula for the salt that is produced in the reaction, we must remember the charges on common ions. Magnesium is always present in compounds as Mg^{2+} ; the chloride ion is Cl^- . The magnesium salt formed in the reaction is MgCl_2 , meaning the balanced molecular equation is



Both FeCl_2 and MgCl_2 are soluble strong electrolytes and can be written in ionic form. Cl^- then, is a spectator ion in the reaction. The net ionic equation is



The net ionic equation shows that Mg is oxidized and Fe^{2+} is reduced in this reaction.

Check: Note that the net ionic equation is balanced with respect to both charge and mass.

PRACTICE EXERCISE

Which of the following metals will be oxidized by $\text{Pb}(\text{NO}_3)_2$: Zn, Cu, Fe?

Answer: Zn and Fe

4.5 CONCENTRATIONS OF SOLUTIONS

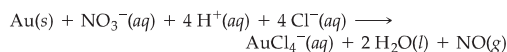
The behavior of solutions often depends on the nature of the solutes and their concentrations. Scientists use the term **concentration** to designate the amount of solute dissolved in a given quantity of solvent or quantity of solution. The concept of concentration is intuitive: The greater the amount of solute dissolved in a certain amount of solvent, the more concentrated the resulting solution. In chemistry we often need to express the concentrations of solutions quantitatively.

A Closer Look THE AURA OF GOLD

Gold has been known since the earliest records of human existence. Throughout history people have cherished gold, have fought for it, and have died for it.

The physical and chemical properties of gold serve to make it a special metal. First, its intrinsic beauty and rarity make it precious. Second, gold is soft and can be easily formed into artistic objects, jewelry, and coins (Figure 4.15 ▶). Third, gold is one of the least active metals (Table 4.5). It is not oxidized in air and does not react with water. It is unreactive toward basic solutions and nearly all acidic solutions. As a result, gold can be found in nature as a pure element rather than combined with oxygen or other elements, which accounts for its early discovery.

Many of the early studies of the reactions of gold arose from the practice of alchemy, in which people attempted to turn cheap metals, such as lead, into gold. Alchemists discovered that gold can be dissolved in a 3:1 mixture of concentrated hydrochloric and nitric acids, known as aqua regia (“royal water”). The action of nitric acid on gold is similar to that on copper (Equation 4.32) in that the nitrate ion, rather than H^+ , oxidizes the metal to Au^{3+} . The Cl^- ions interact with Au^{3+} to form highly stable $AuCl_4^-$ ions. The net ionic equation for the reaction of gold with aqua regia is



All the gold ever mined would easily fit in a cube 19 m on a side and weighing about 1.1×10^8 kg (125,000 tons). More than 90% of this amount has been produced since the beginning

of the California gold rush of 1848. Each year, worldwide production of gold amounts to about 1.8×10^6 kg (2000 tons). By contrast, over 1.5×10^{10} kg (16 million tons) of aluminum are produced annually. Gold is used mainly in jewelry (73%), coins (10%), and electronics (9%). Its use in electronics relies on its excellent conductivity and its corrosion resistance. Gold is used, for example, to plate contacts in electrical switches, relays, and connections. A typical touch-tone telephone contains 33 gold-plated contacts. Gold is also used in computers and other microelectronic devices where fine gold wire is used to link components.

Because of its resistance to corrosion by acids and other substances found in saliva, gold is an ideal metal for dental crowns and caps, which accounts for about 3% of the annual use of the element. The pure metal is too soft to use in dentistry, so it is combined with other metals to form alloys.

Related Exercise: 4.97



◀ **Figure 4.15 Portrait of Pharaoh Tutankhamun (1346–1337 bc) made of gold and precious stones.** This highly prized article is from the inner coffin of the tomb of Tutankhamun.

Strategies in Chemistry ANALYZING CHEMICAL REACTIONS

In this chapter you have been introduced to a great number of chemical reactions. A major difficulty that students face in trying to master material of this sort is gaining a “feel” for what happens when chemicals are allowed to react. In fact, you might marvel at the ease with which your professor or teaching assistant can figure out the results of a chemical reaction. One of our goals in this textbook is to help you become more adept at predicting the outcomes of reactions. The key to gaining this “chemical intuition” is understanding how to categorize reactions.

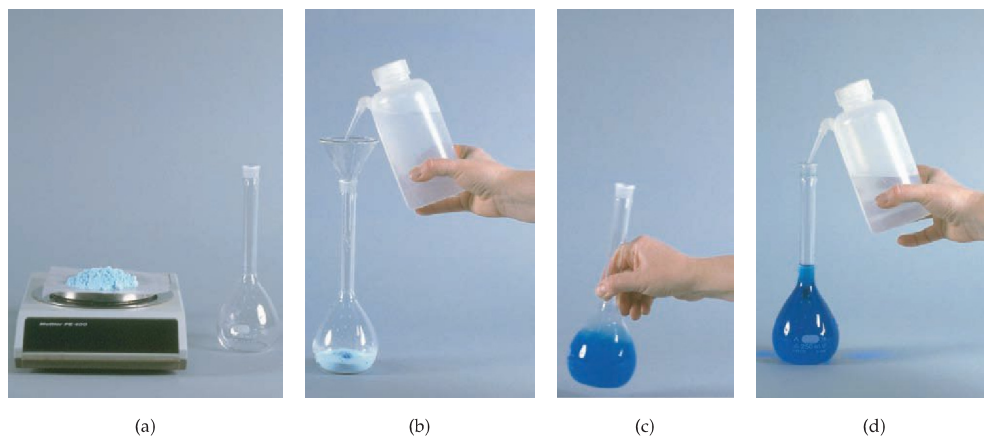
Attempting to memorize the many individual reactions in chemistry would be a futile task. It is far more fruitful to try to recognize patterns to determine the general category of a reaction, such as metathesis or oxidation-reduction. Thus, when you are faced with the challenge of predicting the outcome of a chemical reaction, ask yourself the following pertinent questions:

- What are the reactants in the reaction?
- Are they electrolytes or nonelectrolytes?
- Are they acids and bases?
- If the reactants are electrolytes, will metathesis produce a precipitate? Water? A gas?

- If metathesis cannot occur, can the reactants possibly engage in an oxidation-reduction reaction? This requires that there be both a reactant that can be oxidized and one that can be reduced.

By asking questions such as these, you should be able to predict what might happen during the reaction. You might not always be entirely correct, but if you keep your wits about you, you will not be far off. As you gain experience with chemical reactions, you will begin to look for reactants that might not be immediately obvious, such as water from the solution or oxygen from the atmosphere.

One of the greatest tools available to us in chemistry is experimentation. If you perform an experiment in which two solutions are mixed, you can make observations that help you understand what is happening. For example, using the information in Table 4.1 to predict whether a precipitate will form is not nearly as exciting as actually seeing the precipitate form, as in Figure 4.4. Careful observations in the laboratory portion of the course will make your lecture material easier to master.



▲ Figure 4.16 Procedure for preparation of 0.250 L of 1.00 M solution of CuSO_4 . (a) Weigh out 0.250 mol (39.9 g) of CuSO_4 (formula weight = 159.6 amu). (b) Put the CuSO_4 (solute) into a 250-mL volumetric flask, and add a small quantity of water. (c) Dissolve the solute by swirling the flask. (d) Add more water until the solution just reaches the calibration mark etched on the neck of the flask. Shake the stoppered flask to ensure complete mixing.

Molarity

Molarity (symbol M) expresses the concentration of a solution as the number of moles of solute in a liter of solution (soln):

$$\text{Molarity} = \frac{\text{moles solute}}{\text{volume of solution in liters}} \quad [4.33]$$

A 1.00 molar solution (written 1.00 M) contains 1.00 mol of solute in every liter of solution. Figure 4.16 ▲ shows the preparation of 250.0 mL of a 1.00 M solution of CuSO_4 by using a volumetric flask that is calibrated to hold exactly 250.0 mL. First, 0.250 mol of CuSO_4 (39.9 g) is weighed out and placed in the volumetric flask. Water is added to dissolve the salt, and the resultant solution is diluted to a total volume of 250.0 mL. The molarity of the solution is $(0.250 \text{ mol } \text{CuSO}_4)/(0.250 \text{ L soln}) = 1.00 M$.

GIVE IT SOME THOUGHT

Which is more concentrated, a $1.00 \times 10^{-2} M$ solution of sucrose or a $1.00 \times 10^{-4} M$ solution of sucrose?

SAMPLE EXERCISE 4.11 | Calculating Molarity

Calculate the molarity of a solution made by dissolving 23.4 g of sodium sulfate (Na_2SO_4) in enough water to form 125 mL of solution.

SOLUTION

Analyze: We are given the number of grams of solute (23.4 g), its chemical formula (Na_2SO_4), and the volume of the solution (125 mL). We are asked to calculate the molarity of the solution.

Plan: We can calculate molarity using Equation 4.33. To do so, we must convert the number of grams of solute to moles and the volume of the solution from milliliters to liters.

Solve: The number of moles of Na_2SO_4 is obtained by using its molar mass:

$$\text{Moles } \text{Na}_2\text{SO}_4 = (23.4 \text{ g } \text{Na}_2\text{SO}_4) \left(\frac{1 \text{ mol } \text{Na}_2\text{SO}_4}{142 \text{ g } \text{Na}_2\text{SO}_4} \right) = 0.165 \text{ mol } \text{Na}_2\text{SO}_4$$

Converting the volume of the solution to liters:

$$\text{Liters soln} = (125 \text{ mL}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.125 \text{ L}$$

Thus, the molarity is

$$\text{Molarity} = \frac{0.165 \text{ mol } \text{Na}_2\text{SO}_4}{0.125 \text{ L soln}} = 1.32 \frac{\text{mol } \text{Na}_2\text{SO}_4}{\text{L soln}} = 1.32 M$$

Check: Because the numerator is only slightly larger than the denominator, it is reasonable for the answer to be a little over 1 M . The units (mol/L) are appropriate for molarity, and three significant figures are appropriate for the answer because each of the initial pieces of data had three significant figures.

PRACTICE EXERCISE

Calculate the molarity of a solution made by dissolving 5.00 g of glucose ($C_6H_{12}O_6$) in sufficient water to form exactly 100 mL of solution.

Answer: 0.278 M

Expressing the Concentration of an Electrolyte

When an ionic compound dissolves, the relative concentrations of the ions introduced into the solution depend on the chemical formula of the compound. For example, a 1.0 M solution of NaCl is 1.0 M in Na^+ ions and 1.0 M in Cl^- ions. Similarly, a 1.0 M solution of Na_2SO_4 is 2.0 M in Na^+ ions and 1.0 M in SO_4^{2-} ions. Thus, the concentration of an electrolyte solution can be specified either in terms of the compound used to make the solution (1.0 M Na_2SO_4) or in terms of the ions that the solution contains (2.0 M Na^+ and 1.0 M SO_4^{2-}).

SAMPLE EXERCISE 4.12 | Calculating Molar Concentrations of Ions

What are the molar concentrations of each of the ions present in a 0.025 M aqueous solution of calcium nitrate?

SOLUTION

Analyze: We are given the concentration of the ionic compound used to make the solution and asked to determine the concentrations of the ions in the solution.

Plan: We can use the subscripts in the chemical formula of the compound to determine the relative concentrations of the ions.

Solve: Calcium nitrate is composed of calcium ions (Ca^{2+}) and nitrate ions (NO_3^-), so its chemical formula is $Ca(NO_3)_2$. Because there are two NO_3^- ions for each Ca^{2+} ion in the compound, each mole of $Ca(NO_3)_2$ that dissolves dissociates into 1 mol of Ca^{2+} and 2 mol of NO_3^- . Thus, a solution that is 0.025 M in $Ca(NO_3)_2$ is 0.025 M in Ca^{2+} and $2 \times 0.025 M = 0.050 M$ in NO_3^- :

$$\frac{\text{mol } NO_3^-}{L} = \left(\frac{0.025 \text{ mol } Ca(NO_3)_2}{L} \right) \left(\frac{2 \text{ mol } NO_3^-}{1 \text{ mol } Ca(NO_3)_2} \right) = 0.050 M$$

Check: The concentration of NO_3^- ions is twice that of Ca^{2+} ions, as the subscript 2 after the NO_3^- in the chemical formula $Ca(NO_3)_2$ suggests it should be.

PRACTICE EXERCISE

What is the molar concentration of K^+ ions in a 0.015 M solution of potassium carbonate?

Answer: 0.030 M K^+

Interconverting Molarity, Moles, and Volume

The definition of molarity (Equation 4.33) contains three quantities—molarity, moles solute, and liters of solution. If we know any two of these, we can calculate the third. For example, if we know the molarity of a solution, we can calculate the number of moles of solute in a given volume. Molarity, therefore, is a conversion factor between volume of solution and moles of solute. Calculation of the number of moles of HNO_3 in 2.0 L of 0.200 M HNO_3 solution illustrates the conversion of volume to moles:

$$\begin{aligned} \text{moles } HNO_3 &= (2.0 \text{ L soln}) \left(\frac{0.200 \text{ mol } HNO_3}{1 \text{ L soln}} \right) \\ &= 0.40 \text{ mol } HNO_3 \end{aligned}$$

Dimensional analysis can be used in this conversion if we express molarity as moles/liter soln. To obtain moles, therefore, we multiply liters and molarity: moles = liters \times molarity = ~~liters~~ \times moles/~~liter~~.

To illustrate the conversion of moles to volume, let's calculate the volume of 0.30 M HNO₃ solution required to supply 2.0 mol of HNO₃:

$$\text{Liters soln} = (2.0 \text{ mol HNO}_3) \left(\frac{1 \text{ L soln}}{0.30 \text{ mol HNO}_3} \right) = 6.7 \text{ L soln}$$

In this case we must use the reciprocal of molarity in the conversion: liters = moles \times 1/M = ~~moles~~ \times liters/~~mole~~.

■ SAMPLE EXERCISE 4.13 | Using Molarity to Calculate Grams of Solute

How many grams of Na₂SO₄ are required to make 0.350 L of 0.500 M Na₂SO₄?

SOLUTION

Analyze: We are given the volume of the solution (0.350 L), its concentration (0.500 M), and the identity of the solute Na₂SO₄ and asked to calculate the number of grams of the solute in the solution.

Plan: We can use the definition of molarity (Equation 4.33) to determine the number of moles of solute, and then convert moles to grams using the molar mass of the solute.

$$M_{\text{Na}_2\text{SO}_4} = \frac{\text{moles Na}_2\text{SO}_4}{\text{liters soln}}$$

Solve: Calculating the moles of Na₂SO₄ using the molarity and volume of solution gives

$$\begin{aligned} M_{\text{Na}_2\text{SO}_4} &= \frac{\text{moles Na}_2\text{SO}_4}{\text{liters soln}} \\ \text{moles Na}_2\text{SO}_4 &= \text{liters soln} \times M_{\text{Na}_2\text{SO}_4} \\ &= (0.350 \text{ L soln}) \left(\frac{0.500 \text{ mol Na}_2\text{SO}_4}{1 \text{ L soln}} \right) \\ &= 0.175 \text{ mol Na}_2\text{SO}_4 \end{aligned}$$

Because each mole of Na₂SO₄ weighs 142 g, the required number of grams of Na₂SO₄ is

$$\text{grams Na}_2\text{SO}_4 = (0.175 \text{ mol Na}_2\text{SO}_4) \left(\frac{142 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} \right) = 24.9 \text{ g Na}_2\text{SO}_4$$

Check: The magnitude of the answer, the units, and the number of significant figures are all appropriate.

■ PRACTICE EXERCISE

(a) How many grams of Na₂SO₄ are there in 15 mL of 0.50 M Na₂SO₄? (b) How many milliliters of 0.50 M Na₂SO₄ solution are needed to provide 0.038 mol of this salt?

Answers: (a) 1.1 g, (b) 76 mL

Dilution

Solutions that are used routinely in the laboratory are often purchased or prepared in concentrated form (called *stock solutions*). Hydrochloric acid, for example, is purchased as a 12 M solution (concentrated HCl). Solutions of lower concentrations can then be obtained by adding water, a process called **dilution**.*

Let's look at how to prepare a dilute solution from a concentrated one. Suppose we wanted to prepare 250.0 mL (that is, 0.2500 L) of 0.100 M CuSO₄ solution by diluting a stock solution containing 1.00 M CuSO₄. When solvent is added to dilute a solution, the number of moles of solute remains unchanged.

$$\text{Moles solute before dilution} = \text{moles solute after dilution} \quad [4.34]$$

*In diluting a concentrated acid or base, the acid or base should be added to water and then further diluted by adding more water. Adding water directly to concentrated acid or base can cause spattering because of the intense heat generated.

For a long time it was thought that dehydration was a potential danger for people engaged in extended vigorous activity. Thus, athletes were encouraged to drink lots of water while engaged in active sport. The trend toward extensive hydration has spread throughout society; many people carry water bottles everywhere and dutifully keep well hydrated.

It turns out, though, that in some circumstances, drinking too much water is a greater danger than not drinking enough. Excess water consumption can lead to *hyponatremia*, a condition in which the concentration of sodium ion in the blood is too low. In the past decade at least four marathon runners have died from hyponatremia-related trauma, and dozens more have become seriously ill. For example, a first-time marathoner named Hillary Bellamy, running in the Marine Corps marathon in 2003, collapsed near mile 22 and died the next day. One physician who treated her said that she died from hyponatremia-induced brain swelling, the result of drinking too much water before and during the race.

The normal blood sodium level is 135 to 145 mM (*millimolar*). When that level drops to as low as 125 mM, dizziness and confusion set in. A concentration below 120 mM can be critical. Low sodium level in the blood causes brain tissue to swell. Dangerously low levels can occur in a marathon runner or other active athlete who is sweating out salt at the same time that excessive salt-free water is being drunk to compensate for water loss. The condition affects women more than men because of their generally different body composition and patterns of metabolism. Drinking a sport drink, such as Gatorade, which contains some electrolytes, helps to prevent hyponatremia (Figure 4.17▶).

Contrary to popular belief, dehydration is not as likely as overhydration to present a life-threatening situation, though it can contribute to heat stroke when the temperature is high. Athletes frequently lose several pounds in the course of extreme workouts, all in the form of water loss, with no lasting adverse effects. When, for instance, Amby Burfoot ran in the Boston Marathon in 1968, his body weight went from 138 to 129 pounds during the race. He lost 6.5% of his body weight while winning the men's competition that year. Weight losses of this magnitude are typical of elite marathon runners, who produce tremendous amounts of heat and sweat and cannot afford to slow down for much drinking.

Related Exercises: 4.63, 4.64



▲ **Figure 4.17 Water stations.** To help prevent overhydration, the number of water stations such as this one has been reduced in many marathon events.

Because we know both the volume and concentration of the dilute solution, we can calculate the number of moles of CuSO_4 it contains.

$$\begin{aligned} \text{moles CuSO}_4 \text{ in dil soln} &= (0.2500 \text{ L soln}) \left(0.100 \frac{\text{mol CuSO}_4}{\text{L soln}} \right) \\ &= 0.0250 \text{ mol CuSO}_4 \end{aligned}$$

Now we can calculate the volume of the concentrated solution needed to provide 0.0250 mol CuSO_4 :

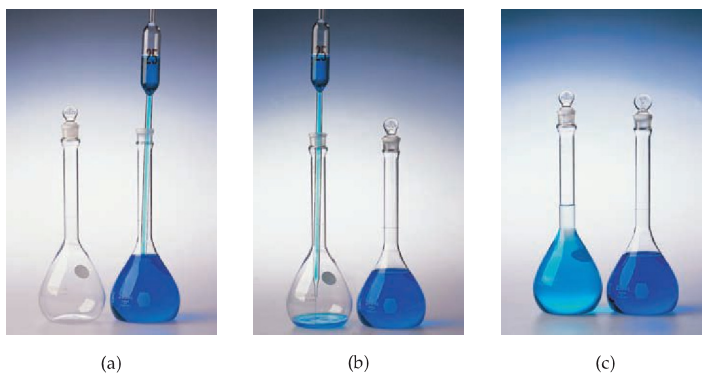
$$\text{Liters of conc soln} = (0.0250 \text{ mol CuSO}_4) \left(\frac{1 \text{ L soln}}{1.00 \text{ mol CuSO}_4} \right) = 0.0250 \text{ L}$$

Thus, this dilution is achieved by withdrawing 0.0250 L (that is, 25.0 mL) of the 1.00 M solution using a pipet, adding it to a 250-mL volumetric flask, and then diluting it to a final volume of 250.0 mL, as shown in Figure 4.18▼. Notice that the diluted solution is less intensely colored than the concentrated one.

In laboratory situations, calculations of this sort are often made very quickly with a simple equation that can be derived by remembering that the number of moles of solute is the same in both the concentrated and dilute solutions and that moles = molarity \times liters:

$$\begin{aligned} \text{moles solute in conc soln} &= \text{moles solute in dil soln} \\ M_{\text{conc}} \times V_{\text{conc}} &= M_{\text{dil}} \times V_{\text{dil}} \end{aligned} \quad [4.35]$$

► **Figure 4.18 Procedure for preparing 250 mL of 0.100 M CuSO₄ by dilution of 1.00 M CuSO₄.** (a) Draw 25.0 mL of the 1.00 M solution into a pipet. (b) Add this to a 250-mL volumetric flask. (c) Add water to dilute the solution to a total volume of 250 mL.



GIVE IT SOME THOUGHT

How is the molarity of a 0.50 M KBr solution changed when water is added to double its volume?

The molarity of the more concentrated stock solution (M_{conc}) is always larger than the molarity of the dilute solution (M_{dil}). Because the volume of the solution increases upon dilution, V_{dil} is always larger than V_{conc} . Although Equation 4.35 is derived in terms of liters, any volume unit can be used as long as that same unit is used on both sides of the equation. For example, in the calculation we did for the CuSO₄ solution, we have

$$(1.00 \text{ M})(V_{\text{conc}}) = (0.100 \text{ M})(250. \text{ mL})$$

Solving for V_{conc} gives $V_{\text{conc}} = 25.0 \text{ mL}$ as before.

SAMPLE EXERCISE 4.14 | Preparing a Solution by Dilution

How many milliliters of 3.0 M H₂SO₄ are needed to make 450 mL of 0.10 M H₂SO₄?

SOLUTION

Analyze: We need to dilute a concentrated solution. We are given the molarity of a more concentrated solution (3.0 M) and the volume and molarity of a more dilute one containing the same solute (450 mL of 0.10 M solution). We must calculate the volume of the concentrated solution needed to prepare the dilute solution.

Plan: We can calculate the number of moles of solute, H₂SO₄, in the dilute solution and then calculate the volume of the concentrated solution needed to supply this amount of solute. Alternatively, we can directly apply Equation 4.35. Let's compare the two methods.

Solve: Calculating the moles of H₂SO₄ in the dilute solution:

$$\begin{aligned} \text{moles H}_2\text{SO}_4 \text{ in dilute solution} &= (0.450 \text{ L soln}) \left(\frac{0.10 \text{ mol H}_2\text{SO}_4}{1 \text{ L soln}} \right) \\ &= 0.045 \text{ mol H}_2\text{SO}_4 \end{aligned}$$

Calculating the volume of the concentrated solution that contains 0.045 mol H₂SO₄:

$$\text{L conc soln} = (0.045 \text{ mol H}_2\text{SO}_4) \left(\frac{1 \text{ L soln}}{3.0 \text{ mol H}_2\text{SO}_4} \right) = 0.015 \text{ L soln}$$

Converting liters to milliliters gives 15 mL.

If we apply Equation 4.35, we get the same result:

$$(3.0 M)(V_{\text{conc}}) = (0.10 M)(450 \text{ mL})$$

$$(V_{\text{conc}}) = \frac{(0.10 M)(450 \text{ mL})}{3.0 M} = 15 \text{ mL}$$

Either way, we see that if we start with 15 mL of 3.0 M H₂SO₄ and dilute it to a total volume of 450 mL, the desired 0.10 M solution will be obtained.

Check: The calculated volume seems reasonable because a small volume of concentrated solution is used to prepare a large volume of dilute solution.

■ PRACTICE EXERCISE

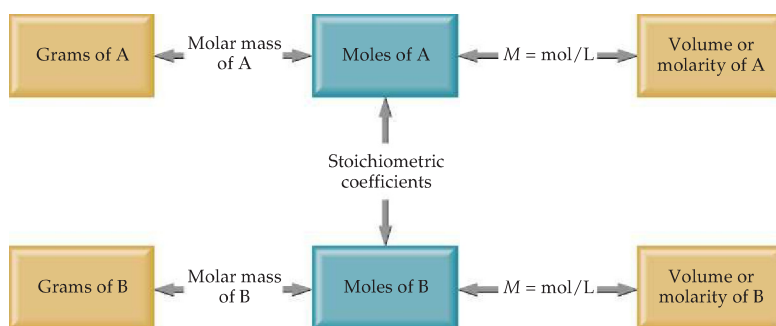
(a) What volume of 2.50 M lead(II) nitrate solution contains 0.0500 mol of Pb²⁺?
 (b) How many milliliters of 5.0 M K₂Cr₂O₇ solution must be diluted to prepare 250 mL of 0.10 M solution? (c) If 10.0 mL of a 10.0 M stock solution of NaOH is diluted to 250 mL, what is the concentration of the resulting stock solution?

Answers: (a) 0.0200 L = 20.0 mL, (b) 5.0 mL, (c) 0.40 M

4.6 SOLUTION STOICHIOMETRY AND CHEMICAL ANALYSIS

Imagine that you have to determine the concentrations of several ions in a sample of lake water. Although many instrumental methods have been developed for such analyses, chemical reactions such as those discussed in this chapter continue to be used. In Chapter 3 we learned that if you know the chemical equation and the amount of one reactant consumed in the reaction, you can calculate the quantities of other reactants and products. In this section we briefly explore such analyses of solutions.

Recall that the coefficients in a balanced equation give the relative number of moles of reactants and products. ∞ (Section 3.6) To use this information, we must convert the quantities of substances involved in a reaction into moles. When we are dealing with grams of substances, as we were in Chapter 3, we use the molar mass to achieve this conversion. When we are working with solutions of known molarity, however, we use molarity and volume to determine the number of moles (moles solute = $M \times L$). Figure 4.19 ▼ summarizes this approach to using stoichiometry.



▲ **Figure 4.19 Problem-solving procedure.** Outline of the procedure used to solve stoichiometry problems that involve measured (laboratory) units of mass, solution concentration (molarity), or volume.

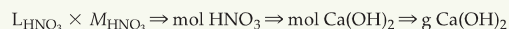
SAMPLE EXERCISE 4.15 | Using Mass Relations in a Neutralization Reaction

How many grams of $\text{Ca}(\text{OH})_2$ are needed to neutralize 25.0 mL of 0.100 M HNO_3 ?

SOLUTION

Analyze: The reactants are an acid, HNO_3 , and a base, $\text{Ca}(\text{OH})_2$. The volume and molarity of HNO_3 are given, and we are asked how many grams of $\text{Ca}(\text{OH})_2$ are needed to neutralize this quantity of HNO_3 .

Plan: We can use the molarity and volume of the HNO_3 solution to calculate the number of moles of HNO_3 . We then use the balanced equation to relate the moles of HNO_3 to moles of $\text{Ca}(\text{OH})_2$. Finally, we can convert moles of $\text{Ca}(\text{OH})_2$ to grams. These steps can be summarized as follows:



Solve: The product of the molar concentration of a solution and its volume in liters gives the number of moles of solute:

$$\begin{aligned} \text{moles HNO}_3 &= L_{\text{HNO}_3} \times M_{\text{HNO}_3} = (0.0250 \text{ L}) \left(0.100 \frac{\text{mol HNO}_3}{\text{L}} \right) \\ &= 2.50 \times 10^{-3} \text{ mol HNO}_3 \end{aligned}$$

Because this is an acid–base neutralization reaction, HNO_3 and $\text{Ca}(\text{OH})_2$ react to form H_2O and the salt containing Ca^{2+} and NO_3^- :



Thus, 2 mol $\text{HNO}_3 \approx$ 1 mol $\text{Ca}(\text{OH})_2$. Therefore,

$$\begin{aligned} \text{grams Ca}(\text{OH})_2 &= (2.50 \times 10^{-3} \text{ mol HNO}_3) \left(\frac{1 \text{ mol Ca}(\text{OH})_2}{2 \text{ mol HNO}_3} \right) \left(\frac{74.1 \text{ g Ca}(\text{OH})_2}{1 \text{ mol Ca}(\text{OH})_2} \right) \\ &= 0.0926 \text{ g Ca}(\text{OH})_2 \end{aligned}$$

Check: The size of the answer is reasonable. A small volume of dilute acid will require only a small amount of base to neutralize it.

PRACTICE EXERCISE

(a) How many grams of NaOH are needed to neutralize 20.0 mL of 0.150 M H_2SO_4 solution? (b) How many liters of 0.500 M $\text{HCl}(aq)$ are needed to react completely with 0.100 mol of $\text{Pb}(\text{NO}_3)_2(aq)$, forming a precipitate of $\text{PbCl}_2(s)$?

Answers: (a) 0.240 g, (b) 0.400 L

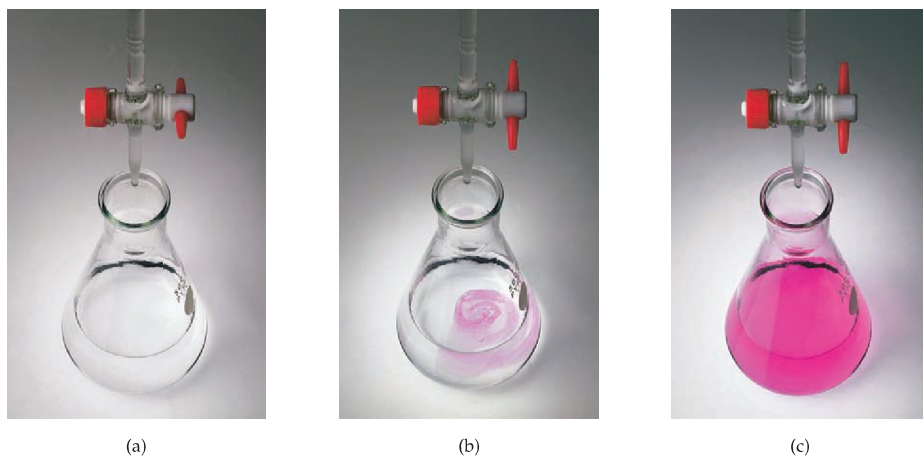
Titration

To determine the concentration of a particular solute in a solution, chemists often carry out a **titration**, which involves combining a sample of the solution with a reagent solution of known concentration, called a **standard solution**. Titrations can be conducted using acid–base, precipitation, or oxidation–reduction reactions. Suppose we have an HCl solution of unknown concentration and an NaOH solution we know to be 0.100 M. To determine the concentration of the HCl solution, we take a specific volume of that solution, say 20.00 mL. We then slowly add the standard NaOH solution to it until the neutralization reaction between the HCl and NaOH is complete. The point at which stoichiometrically equivalent quantities are brought together is known as the **equivalence point** of the titration.

GIVE IT SOME THOUGHT

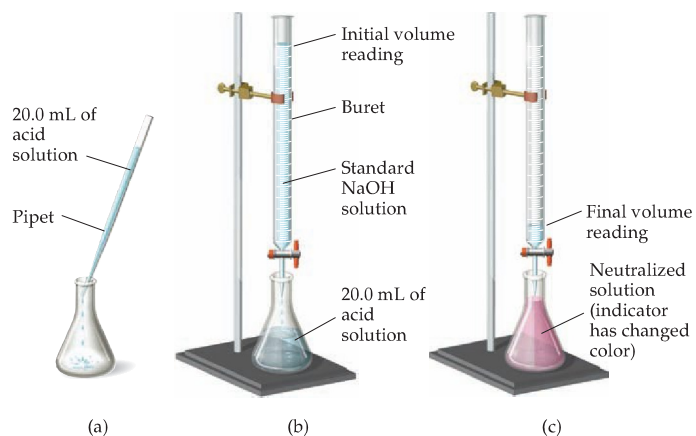
25.00 mL of a 0.100 M HBr solution is titrated with a 0.200 M NaOH solution. How many mL of the NaOH solution are required to reach the equivalence point?

To titrate an unknown with a standard solution, there must be some way to determine when the equivalence point of the titration has been reached. In acid–base titrations, dyes known as acid–base **indicators** are used for this purpose.



▲ **Figure 4.20** Change in appearance of a solution containing phenolphthalein indicator as base is added. Before the end point, the solution is colorless (a). As the end point is approached, a pale pink color forms where the base is added (b). At the end point, this pale pink color extends throughout the solution after mixing. As even more base is added, the intensity of the pink color increases (c).

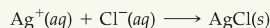
For example, the dye known as phenolphthalein is colorless in acidic solution but is pink in basic solution. If we add phenolphthalein to an unknown solution of acid, the solution will be colorless, as seen in Figure 4.20(a)▲. We can then add standard base from a buret until the solution barely turns from colorless to pink, as seen in Figure 4.20(b). This color change indicates that the acid has been neutralized and the drop of base that caused the solution to become colored has no acid to react with. The solution therefore becomes basic, and the dye turns pink. The color change signals the *end point* of the titration, which usually coincides very nearly with the equivalence point. Care must be taken to choose indicators whose end points correspond to the equivalence point of the titration. We will consider this matter in Chapter 17. The titration procedure is summarized in Figure 4.21▼.



◀ **Figure 4.21** Procedure for titrating an acid against a standardized solution of NaOH. (a) A known quantity of acid is added to a flask. (b) An acid–base indicator is added, and standardized NaOH is added from a buret. (c) The equivalence point is signaled by a color change in the indicator.

SAMPLE EXERCISE 4.16 | Determining the Quantity of Solute by Titration

The quantity of Cl^- in a municipal water supply is determined by titrating the sample with Ag^+ . The reaction taking place during the titration is



The end point in this type of titration is marked by a change in color of a special type of indicator. (a) How many grams of chloride ion are in a sample of the water if 20.2 mL of 0.100 M Ag^+ is needed to react with all the chloride in the sample? (b) If the sample has a mass of 10.0 g, what percent Cl^- does it contain?

SOLUTION

Analyze: We are given the volume (20.2 mL) and molarity (0.100 M) of a solution of Ag^+ and the chemical equation for reaction of this ion with Cl^- . We are asked first to calculate the number of grams of Cl^- in the sample and, second, to calculate the mass percent of Cl^- in the sample.

(a) Plan: We begin by using the volume and molarity of Ag^+ to calculate the number of moles of Ag^+ used in the titration. We can then use the balanced equation to determine the moles of Cl^- in the sample and from that the grams of Cl^- .

Solve:

$$\begin{aligned} \text{moles Ag}^+ &= (20.2 \text{ mL soln}) \left(\frac{1 \text{ L soln}}{1000 \text{ mL soln}} \right) \left(0.100 \frac{\text{mol Ag}^+}{\text{L soln}} \right) \\ &= 2.02 \times 10^{-3} \text{ mol Ag}^+ \end{aligned}$$

From the balanced equation we see that 1 mol $\text{Ag}^+ \approx 1$ mol Cl^- . Using this information and the molar mass of Cl, we have

$$\begin{aligned} \text{grams Cl}^- &= (2.02 \times 10^{-3} \text{ mol Ag}^+) \left(\frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+} \right) \left(\frac{35.5 \text{ g Cl}^-}{1 \text{ mol Cl}^-} \right) \\ &= 7.17 \times 10^{-2} \text{ g Cl}^- \end{aligned}$$

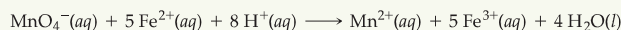
(b) Plan: To calculate the percentage of Cl^- in the sample, we compare the number of grams of Cl^- in the sample, 7.17×10^{-2} g, with the original mass of the sample, 10.0 g.

$$\text{Solve: Percent Cl}^- = \frac{7.17 \times 10^{-2} \text{ g}}{10.0 \text{ g}} \times 100\% = 0.717\% \text{ Cl}^-$$

Comment: Chloride ion is one of the most common ions in water and sewage. Ocean water contains 1.92% Cl^- . Whether water containing Cl^- tastes salty depends on the other ions present. If the only accompanying ions are Na^+ , a salty taste may be detected with as little as 0.03% Cl^- .

PRACTICE EXERCISE

A sample of an iron ore is dissolved in acid, and the iron is converted to Fe^{2+} . The sample is then titrated with 47.20 mL of 0.02240 M MnO_4^- solution. The oxidation-reduction reaction that occurs during titration is as follows:



(a) How many moles of MnO_4^- were added to the solution? (b) How many moles of Fe^{2+} were in the sample? (c) How many grams of iron were in the sample? (d) If the sample had a mass of 0.8890 g, what is the percentage of iron in the sample?

Answers: (a) 1.057×10^{-3} mol MnO_4^- (b) 5.286×10^{-3} mol Fe^{2+} , (c) 0.2952 g, (d) 33.21%

SAMPLE EXERCISE 4.17 | Determining Solution Concentration Via an Acid-Base Titration

One commercial method used to peel potatoes is to soak them in a solution of NaOH for a short time, remove them from the NaOH, and spray off the peel. The concentration of NaOH is normally in the range of 3 to 6 M. The NaOH is analyzed periodically. In one such analysis, 45.7 mL of 0.500 M H_2SO_4 is required to neutralize a 20.0-mL sample of NaOH solution. What is the concentration of the NaOH solution?

SOLUTION

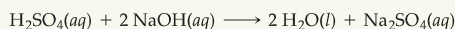
Analyze: We are given the volume (45.7 mL) and molarity (0.500 M) of an H₂SO₄ solution that reacts completely with a 20.0-mL sample of NaOH. We are asked to calculate the molarity of the NaOH solution.

Plan: We can use the volume and molarity of the H₂SO₄ to calculate the number of moles of this substance. Then, we can use this quantity and the balanced equation for the reaction to calculate the number of moles of NaOH. Finally, we can use the moles of NaOH and the volume of this solution to calculate molarity.

Solve: The number of moles of H₂SO₄ is given by the product of the volume and molarity of this solution:

$$\begin{aligned}\text{moles H}_2\text{SO}_4 &= (45.7 \text{ mL soln}) \left(\frac{1 \text{ L soln}}{1000 \text{ mL soln}} \right) \left(0.500 \frac{\text{mol H}_2\text{SO}_4}{\text{L soln}} \right) \\ &= 2.28 \times 10^{-2} \text{ mol H}_2\text{SO}_4\end{aligned}$$

Acids react with metal hydroxides to form water and a salt. Thus, the balanced equation for the neutralization reaction is



According to the balanced equation, 1 mol H₂SO₄ \approx 2 mol NaOH. Therefore,

$$\begin{aligned}\text{moles NaOH} &= (2.28 \times 10^{-2} \text{ mol H}_2\text{SO}_4) \left(\frac{2 \text{ mol NaOH}}{1 \text{ mol H}_2\text{SO}_4} \right) \\ &= 4.56 \times 10^{-2} \text{ mol NaOH}\end{aligned}$$

Knowing the number of moles of NaOH present in 20.0 mL of solution allows us to calculate the molarity of this solution:

$$\begin{aligned}\text{Molarity NaOH} &= \frac{\text{mol NaOH}}{\text{L soln}} = \left(\frac{4.56 \times 10^{-2} \text{ mol NaOH}}{20.0 \text{ mL soln}} \right) \left(\frac{1000 \text{ mL soln}}{1 \text{ L soln}} \right) \\ &= 2.28 \frac{\text{mol NaOH}}{\text{L soln}} = 2.28 \text{ M}\end{aligned}$$

PRACTICE EXERCISE

What is the molarity of an NaOH solution if 48.0 mL is needed to neutralize 35.0 mL of 0.144 M H₂SO₄?

Answers: 0.210 M

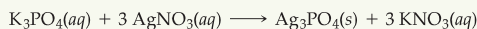
SAMPLE INTEGRATIVE EXERCISE | Putting Concepts Together

Note: Integrative exercises require skills from earlier chapters as well as ones from the present chapter.

A sample of 70.5 mg of potassium phosphate is added to 15.0 mL of 0.050 M silver nitrate, resulting in the formation of a precipitate. (a) Write the molecular equation for the reaction. (b) What is the limiting reactant in the reaction? (c) Calculate the theoretical yield, in grams, of the precipitate that forms.

SOLUTION

(a) Potassium phosphate and silver nitrate are both ionic compounds. Potassium phosphate contains K⁺ and PO₄³⁻ ions, so its chemical formula is K₃PO₄. Silver nitrate contains Ag⁺ and NO₃⁻ ions, so its chemical formula is AgNO₃. Because both reactants are strong electrolytes, the solution contains K⁺, PO₄³⁻, Ag⁺, and NO₃⁻ ions before the reaction occurs. According to the solubility guidelines in Table 4.1, Ag⁺ and PO₄³⁻ form an insoluble compound, so Ag₃PO₄ will precipitate from the solution. In contrast, K⁺ and NO₃⁻ will remain in solution because KNO₃ is water soluble. Thus, the balanced molecular equation for the reaction is



(b) To determine the limiting reactant, we must examine the number of moles of each reactant. ∞ (Section 3.7) The number of moles of K₃PO₄ is calculated from the mass of the sample using the molar mass as a conversion factor. ∞ (Section 3.4)

The molar mass of K_3PO_4 is $3(39.1) + 31.0 + 4(16.0) = 212.3 \text{ g/mol}$. Converting milligrams to grams and then to moles, we have

$$(70.5 \text{ mg } \text{K}_3\text{PO}_4) \left(\frac{10^{-3} \text{ g } \text{K}_3\text{PO}_4}{1 \text{ mg } \text{K}_3\text{PO}_4} \right) \left(\frac{1 \text{ mol } \text{K}_3\text{PO}_4}{212.3 \text{ g } \text{K}_3\text{PO}_4} \right) = 3.32 \times 10^{-4} \text{ mol } \text{K}_3\text{PO}_4$$

We determine the number of moles of AgNO_3 from the volume and molarity of the solution. ∞ (Section 4.5) Converting milliliters to liters and then to moles, we have

$$(15.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.050 \text{ mol } \text{AgNO}_3}{\text{L}} \right) = 7.5 \times 10^{-4} \text{ mol } \text{AgNO}_3$$

Comparing the amounts of the two reactants, we find that there are $(7.5 \times 10^{-4}) / (3.32 \times 10^{-4}) = 2.3$ times as many moles of AgNO_3 as there are moles of K_3PO_4 . According to the balanced equation, however, 1 mol K_3PO_4 requires 3 mol of AgNO_3 . Thus, there is insufficient AgNO_3 to consume the K_3PO_4 , and AgNO_3 is the limiting reactant.

(c) The precipitate is Ag_3PO_4 , whose molar mass is $3(107.9) + 31.0 + 4(16.0) = 418.7 \text{ g/mol}$. To calculate the number of grams of Ag_3PO_4 that could be produced in this reaction (the theoretical yield), we use the number of moles of the limiting reactant, converting $\text{mol } \text{AgNO}_3 \Rightarrow \text{mol } \text{Ag}_3\text{PO}_4 \Rightarrow \text{g } \text{Ag}_3\text{PO}_4$. We use the coefficients in the balanced equation to convert moles of AgNO_3 to moles Ag_3PO_4 , and we use the molar mass of Ag_3PO_4 to convert the number of moles of this substance to grams.

$$(7.5 \times 10^{-4} \text{ mol } \text{AgNO}_3) \left(\frac{1 \text{ mol } \text{Ag}_3\text{PO}_4}{3 \text{ mol } \text{AgNO}_3} \right) \left(\frac{418.7 \text{ g } \text{Ag}_3\text{PO}_4}{1 \text{ mol } \text{Ag}_3\text{PO}_4} \right) = 0.10 \text{ g } \text{Ag}_3\text{PO}_4$$

The answer has only two significant figures because the quantity of AgNO_3 is given to only two significant figures.

CHAPTER REVIEW

SUMMARY AND KEY TERMS

Introduction and Section 4.1 Solutions in which water is the dissolving medium are called **aqueous solutions**. The component of the solution that is in the greater quantity is the **solvent**. The other components are **solutes**.

Any substance whose aqueous solution contains ions is called an **electrolyte**. Any substance that forms a solution containing no ions is a **nonelectrolyte**. Electrolytes that are present in solution entirely as ions are **strong electrolytes**, whereas those that are present partly as ions and partly as molecules are **weak electrolytes**. Ionic compounds dissociate into ions when they dissolve, and they are strong electrolytes. The solubility of ionic substances is made possible by **solvation**, the interaction of ions with polar solvent molecules. Most molecular compounds are nonelectrolytes, although some are weak electrolytes, and a few are strong electrolytes. When representing the ionization of a weak electrolyte in solution, half-arrows in both directions are used, indicating that the forward and reverse reactions can achieve a chemical balance called a **chemical equilibrium**.

Section 4.2 **Precipitation reactions** are those in which an insoluble product, called a **precipitate**, forms. Solubility guidelines help determine whether or not an ionic compound will be soluble in water. (The **solubility** of a substance is the amount that dissolves in a given quantity of solvent.) Reactions such as precipitation reactions, in which cations and anions appear to exchange partners, are called **exchange reactions**, or **metathesis reactions**.

Chemical equations can be written to show whether dissolved substances are present in solution predominantly as ions or molecules. When the complete chemical formulas of all reactants and products are used, the equation is called a **molecular equation**. A **complete ionic equation** shows all dissolved strong electrolytes as their component ions. In a **net ionic equation**, those ions that go through the reaction unchanged (**spectator ions**) are omitted.

Section 4.3 Acids and bases are important electrolytes. **Acids** are proton donors; they increase the concentration of $\text{H}^+(\text{aq})$ in aqueous solutions to which they are added.

Bases are proton acceptors; they increase the concentration of $\text{OH}^-(aq)$ in aqueous solutions. Those acids and bases that are strong electrolytes are called **strong acids** and **strong bases**, respectively. Those that are weak electrolytes are **weak acids** and **weak bases**. When solutions of acids and bases are mixed, a **neutralization reaction** results. The neutralization reaction between an acid and a metal hydroxide produces water and a **salt**. Gases can also be formed as a result of acid–base reactions. The reaction of a sulfide with an acid forms $\text{H}_2\text{S}(g)$; the reaction between a carbonate and an acid forms $\text{CO}_2(g)$.

Section 4.4 Oxidation is the loss of electrons by a substance, whereas **reduction** is the gain of electrons by a substance. **Oxidation numbers** keep track of electrons during chemical reactions and are assigned to atoms using specific rules. The oxidation of an element results in an increase in its oxidation number, whereas reduction is accompanied by a decrease in oxidation number. Oxidation is always accompanied by reduction, giving **oxidation-reduction**, or redox, **reactions**.

Many metals are oxidized by O_2 , acids, and salts. The redox reactions between metals and acids and between metals and salts are called **displacement reactions**. The products of these displacement reactions are always an element (H_2 or a metal) and a salt. Comparing such reactions allows us to rank metals according to their ease of oxidation. A list of metals arranged in order of decreasing

ease of oxidation is called an **activity series**. Any metal on the list can be oxidized by ions of metals (or H^+) below it in the series.

Section 4.5 The composition of a solution expresses the relative quantities of solvent and solutes that it contains. One of the common ways to express the **concentration** of a solute in a solution is in terms of molarity. The **molarity** of a solution is the number of moles of solute per liter of solution. Molarity makes it possible to interconvert solution volume and number of moles of solute. Solutions of known molarity can be formed either by weighing out the solute and diluting it to a known volume or by the **dilution** of a more concentrated solution of known concentration (a stock solution). Adding solvent to the solution (the process of dilution) decreases the concentration of the solute without changing the number of moles of solute in the solution ($M_{\text{conc}} \times V_{\text{conc}} = M_{\text{dil}} \times V_{\text{dil}}$).

Section 4.6 In the process called **titration**, we combine a solution of known concentration (a **standard solution**) with a solution of unknown concentration to determine the unknown concentration or the quantity of solute in the unknown. The point in the titration at which stoichiometrically equivalent quantities of reactants are brought together is called the **equivalence point**. An **indicator** can be used to show the end point of the titration, which coincides closely with the equivalence point.

KEY SKILLS

- Recognize compounds as acids or bases, and as strong, weak, or nonelectrolytes.
- Recognize reactions as acid–base, precipitation, metathesis, or redox.
- Be able to calculate moles or grams of substances in solution using molarity.
- Understand how to carry out a dilution to achieve a desired solution concentration.
- Understand how to perform and interpret the results of a titration.

KEY EQUATIONS

$$\bullet \text{ Molarity} = \frac{\text{moles solute}}{\text{volume of solution in liters}} \quad [4.33]$$

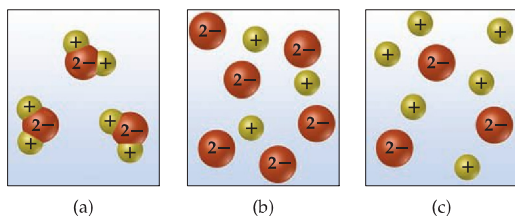
Molarity is the most commonly used unit of concentration in chemistry.

$$\bullet M_{\text{conc}} \times V_{\text{conc}} = M_{\text{dil}} \times V_{\text{dil}} \quad [4.35]$$

When adding solvent to a concentrated solution to make a dilute solution, molarities and volumes of both concentrated and dilute solutions can be calculated if three of the quantities are known.

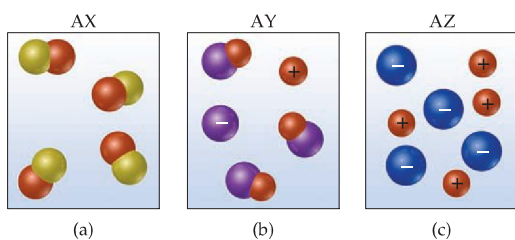
VISUALIZING CONCEPTS

- 4.1 Which of the following schematic drawings best describes a solution of Li_2SO_4 in water (water molecules not shown for simplicity)? [Section 4.1]



- 4.2 Methanol, CH_3OH , and hydrogen chloride, HCl , are both molecular substances, yet an aqueous solution of methanol does not conduct an electrical current, whereas a solution of HCl does conduct. Account for this difference. [Section 4.1]

- 4.3 Aqueous solutions of three different substances, AX, AY, and AZ, are represented by the three diagrams below. Identify each substance as a strong electrolyte, weak electrolyte, or nonelectrolyte. [Section 4.1]



- 4.4 A 0.1 M solution of acetic acid, CH_3COOH , causes the lightbulb in the apparatus of Figure 4.2 to glow about as brightly as a 0.001 M solution of HBr . How do you account for this fact? [Section 4.1]

- 4.5 You are presented with three white solids, A, B, and C, which are glucose (a sugar substance), NaOH , and AgBr . Solid A dissolves in water to form a conducting solution. B is not soluble in water. C dissolves in water to form a nonconducting solution. Identify A, B, and C. [Section 4.2]

- 4.6 We have seen that ions in aqueous solution are stabilized by the attractions between the ions and the water molecules. Why then do some pairs of ions in solution form precipitates? [Section 4.2]

- 4.7 Which of the following ions will *always* be a spectator ion in a precipitation reaction? (a) Cl^- , (b) NO_3^- , (c) NH_4^+ , (d) S^{2-} , (e) SO_4^{2-} . Explain briefly. [Section 4.2]

- 4.8 The labels have fallen off two bottles, one containing $\text{Mg}(\text{NO}_3)_2$ and the other containing $\text{Pb}(\text{NO}_3)_2$. You have a bottle of dilute H_2SO_4 . How could you use it to test a portion of each solution to identify which solution is which? [Section 4.2]

- 4.9 Explain how a redox reaction involves electrons in the same way that an acid–base reaction involves protons. [Sections 4.3 and 4.4]

- 4.10 If you want to double the concentration of a solution, how could you do it? [Section 4.5]

EXERCISES

Electrolytes

- 4.11 When asked what causes electrolyte solutions to conduct electricity, a student responds that it is due to the movement of electrons through the solution. Is the student correct? If not, what is the correct response?
- 4.12 When methanol, CH_3OH , is dissolved in water, a nonconducting solution results. When acetic acid, CH_3COOH , dissolves in water, the solution is weakly conducting and acidic in nature. Describe what happens upon dissolution in the two cases, and account for the different results.
- 4.13 We have learned in this chapter that many ionic solids dissolve in water as strong electrolytes, that is, as separated ions in solution. What properties of water facilitate this process?
- 4.14 What does it mean to say that ions are hydrated when an ionic substance dissolves in water?
- 4.15 Specify what ions are present in solution upon dissolving each of the following substances in water: (a) ZnCl_2 , (b) HNO_3 , (c) $(\text{NH}_4)_2\text{SO}_4$, (d) $\text{Ca}(\text{OH})_2$.
- 4.16 Specify what ions are present upon dissolving each of the following substances in water: (a) MgI_2 , (b) $\text{Al}(\text{NO}_3)_3$, (c) HClO_4 , (d) NaCH_3COO .
- 4.17 Formic acid, HCOOH , is a weak electrolyte. What solute particles are present in an aqueous solution of this compound? Write the chemical equation for the ionization of HCOOH .
- 4.18 Acetone, CH_3COCH_3 , is a nonelectrolyte; hypochlorous acid, HClO , is a weak electrolyte; and ammonium chloride, NH_4Cl , is a strong electrolyte. (a) What are the solute particles present in aqueous solutions of each compound? (b) If 0.1 mol of each compound is dissolved in solution, which one contains 0.2 mol of solute particles, which contains 0.1 mol of solute particles, and which contains somewhere between 0.1 and 0.2 mol of solute particles?

Precipitation Reactions and Net Ionic Equations

- 4.19 Using solubility guidelines, predict whether each of the following compounds is soluble or insoluble in water: (a) NiCl_2 , (b) Ag_2S , (c) Cs_3PO_4 , (d) SrCO_3 , (e) PbSO_4 .
- 4.20 Predict whether each of the following compounds is soluble in water: (a) $\text{Ni}(\text{OH})_2$, (b) PbBr_2 , (c) $\text{Ba}(\text{NO}_3)_2$, (d) AlPO_4 , (e) AgCH_3COO .
- 4.21 Will precipitation occur when the following solutions are mixed? If so, write a balanced chemical equation for the reaction. (a) Na_2CO_3 and AgNO_3 , (b) NaNO_3 and NiSO_4 , (c) FeSO_4 and $\text{Pb}(\text{NO}_3)_2$.
- 4.22 Identify the precipitate (if any) that forms when the following solutions are mixed, and write a balanced equation for each reaction. (a) $\text{Ni}(\text{NO}_3)_2$ and NaOH , (b) NaOH and K_2SO_4 , (c) Na_2S and $\text{Cu}(\text{CH}_3\text{COO})_2$.
- 4.23 Name the spectator ions in any reactions that may be involved when each of the following pairs of solutions are mixed.
 (a) $\text{Na}_2\text{CO}_3(\text{aq})$ and $\text{MgSO}_4(\text{aq})$
 (b) $\text{Pb}(\text{NO}_3)_2(\text{aq})$ and $\text{Na}_2\text{S}(\text{aq})$
 (c) $(\text{NH}_4)_3\text{PO}_4(\text{aq})$ and $\text{CaCl}_2(\text{aq})$
- 4.24 Write balanced net ionic equations for the reactions that occur in each of the following cases. Identify the spectator ion or ions in each reaction.
- (a) $\text{Cr}_2(\text{SO}_4)_3(\text{aq}) + (\text{NH}_4)_2\text{CO}_3(\text{aq}) \longrightarrow$
 (b) $\text{Ba}(\text{NO}_3)_2(\text{aq}) + \text{K}_2\text{SO}_4(\text{aq}) \longrightarrow$
 (c) $\text{Fe}(\text{NO}_3)_2(\text{aq}) + \text{KOH}(\text{aq}) \longrightarrow$
- 4.25 Separate samples of a solution of an unknown salt are treated with dilute solutions of HBr , H_2SO_4 , and NaOH . A precipitate forms in all three cases. Which of the following cations could the solution contain: K^+ ; Pb^{2+} ; Ba^{2+} ?
- 4.26 Separate samples of a solution of an unknown ionic compound are treated with dilute AgNO_3 , $\text{Pb}(\text{NO}_3)_2$, and BaCl_2 . Precipitates form in all three cases. Which of the following could be the anion of the unknown salt: Br^- ; CO_3^{2-} ; NO_3^- ?
- 4.27 You know that an unlabeled bottle contains a solution of one of the following: AgNO_3 , CaCl_2 , or $\text{Al}_2(\text{SO}_4)_3$. A friend suggests that you test a portion of the solution with $\text{Ba}(\text{NO}_3)_2$ and then with NaCl solutions. Explain how these two tests together would be sufficient to determine which salt is present in the solution.
- 4.28 Three solutions are mixed together to form a single solution. One contains 0.2 mol $\text{Pb}(\text{CH}_3\text{COO})_2$, the second contains 0.1 mol Na_2S , and the third contains 0.1 mol CaCl_2 . (a) Write the net ionic equations for the precipitation reaction or reactions that occur. (b) What are the spectator ions in the solution?

Acid–Base Reactions

- 4.29 Which of the following solutions has the largest concentration of solvated protons: (a) 0.2 M LiOH , (b) 0.2 M HI , (c) 1.0 M methyl alcohol (CH_3OH)? Explain.
- 4.30 Which of the following solutions is the most basic? (a) 0.6 M NH_3 , (b) 0.150 M KOH , (c) 0.100 M $\text{Ba}(\text{OH})_2$. Explain.
- 4.31 What is the difference between (a) a monoprotic acid and a diprotic acid, (b) a weak acid and a strong acid, (c) an acid and a base?
- 4.32 Explain the following observations: (a) NH_3 contains no OH^- ions, and yet its aqueous solutions are basic; (b) HF is called a weak acid, and yet it is very reactive; (c) although sulfuric acid is a strong electrolyte, an aqueous solution of H_2SO_4 contains more HSO_4^- ions than SO_4^{2-} ions.
- 4.33 HCl , HBr , and HI are strong acids, yet HF is a weak acid. What does this mean in terms of the extent to which these substances are ionized in solution?
- 4.34 What is the relationship between the solubility rules in Table 4.1 and the list of strong bases in Table 4.2? Another way of asking this question is, why is $\text{Cd}(\text{OH})_2$, for example, not listed as a strong base in Table 4.2?
- 4.35 Label each of the following substances as an acid, base, salt, or none of the above. Indicate whether the substance exists in aqueous solution entirely in molecular form, entirely as ions, or as a mixture of molecules and ions. (a) HF ; (b) acetonitrile, CH_3CN ; (c) NaClO_4 ; (d) $\text{Ba}(\text{OH})_2$.
- 4.36 An aqueous solution of an unknown solute is tested with litmus paper and found to be acidic. The solution is weakly conducting compared with a solution of NaCl of the same concentration. Which of the following substances could the unknown be: KOH , NH_3 , HNO_3 , KClO_2 , H_3PO_3 , CH_3COCH_3 (acetone)?
- 4.37 Classify each of the following substances as a nonelectrolyte, weak electrolyte, or strong electrolyte in water: (a) H_2SO_3 , (b) $\text{C}_2\text{H}_5\text{OH}$ (ethanol), (c) NH_3 , (d) KClO_3 , (e) $\text{Cu}(\text{NO}_3)_2$.
- 4.38 Classify each of the following aqueous solutions as a nonelectrolyte, weak electrolyte, or strong electrolyte: (a) HClO_4 , (b) HNO_3 , (c) NH_4Cl , (d) CH_3COCH_3 (acetone), (e) CoSO_4 , (f) $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ (sucrose).
- 4.39 Complete and balance the following molecular equations, and then write the net ionic equation for each:
 (a) $\text{HBr}(\text{aq}) + \text{Ca}(\text{OH})_2(\text{aq}) \longrightarrow$
 (b) $\text{Cu}(\text{OH})_2(\text{s}) + \text{HClO}_4(\text{aq}) \longrightarrow$
 (c) $\text{Al}(\text{OH})_3(\text{s}) + \text{HNO}_3(\text{aq}) \longrightarrow$
- 4.40 Write the balanced molecular and net ionic equations for each of the following neutralization reactions:
 (a) Aqueous acetic acid is neutralized by aqueous potassium hydroxide.
 (b) Solid chromium(III) hydroxide reacts with nitric acid.
 (c) Aqueous hypochlorous acid and aqueous calcium hydroxide react.

- 4.41 Write balanced molecular and net ionic equations for the following reactions, and identify the gas formed in each: (a) solid cadmium sulfide reacts with an aqueous solution of sulfuric acid; (b) solid magnesium carbonate reacts with an aqueous solution of perchloric acid.
- 4.42 Because the oxide ion is basic, metal oxides react readily with acids. (a) Write the net ionic equation for the following reaction:
- $$\text{FeO}(s) + 2 \text{HClO}_4(aq) \longrightarrow \text{Fe}(\text{ClO}_4)_2(aq) + \text{H}_2\text{O}(l)$$
- (b) Based on the equation in part (a), write the net ionic equation for the reaction that occurs between NiO(s) and an aqueous solution of nitric acid.

- 4.43 Write a balanced molecular equation and a net ionic equation for the reaction that occurs when (a) solid CaCO_3 reacts with an aqueous solution of nitric acid; (b) solid iron(II) sulfide reacts with an aqueous solution of hydrobromic acid.
- 4.44 As K_2O dissolves in water, the oxide ion reacts with water molecules to form hydroxide ions. Write the molecular and net ionic equations for this reaction. Based on the definitions of acid and base, what ion is the base in this reaction? What is the acid? What is the spectator ion in the reaction?

Oxidation-Reduction Reactions

- 4.45 Define oxidation and reduction in terms of (a) electron transfer and (b) oxidation numbers.
- 4.46 Can oxidation occur without accompanying reduction? Explain.
- 4.47 Which circled region of the periodic table shown here contains the most readily oxidized elements? Which contains the least readily oxidized?

The periodic table is color-coded and has four regions circled: Region A is the alkali metals (Group 1); Region B is the transition metals (Groups 3-10); Region C is the metalloids (Groups 13-15); Region D is the halogens (Group 17). The noble gases (Group 18) are not circled.

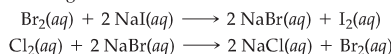
- 4.48 From the elements listed in Table 4.5, select an element that lies in region A of the periodic table shown above and an element that lies in region C. Write a balanced oxidation-reduction equation that shows the oxidation of one metal and reduction of an ion of the other. You will need to decide which element is oxidized and which is reduced.
- 4.49 Determine the oxidation number for the indicated element in each of the following substances: (a) S in SO_2 , (b) C in COCl_2 , (c) Mn in MnO_4^- , (d) Br in HBrO , (e) As in As_4 , (f) O in K_2O_2 .
- 4.50 Determine the oxidation number for the indicated element in each of the following compounds: (a) Ti in TiO_2 , (b) Sn in SnCl_3^- , (c) C in $\text{C}_2\text{O}_4^{2-}$, (d) N in N_2H_4 , (e) N in HNO_2 , (f) Cr in $\text{Cr}_2\text{O}_7^{2-}$.
- 4.51 Which element is oxidized and which is reduced in the following reactions?
- (a) $\text{N}_2(g) + 3 \text{H}_2(g) \longrightarrow 2 \text{NH}_3(g)$
- (b) $3 \text{Fe}(\text{NO}_3)_2(aq) + 2 \text{Al}(s) \longrightarrow 3 \text{Fe}(s) + 2 \text{Al}(\text{NO}_3)_3(aq)$
- (c) $\text{Cl}_2(aq) + 2 \text{NaI}(aq) \longrightarrow \text{I}_2(aq) + 2 \text{NaCl}(aq)$
- (d) $\text{PbS}(s) + 4 \text{H}_2\text{O}_2(aq) \longrightarrow \text{PbSO}_4(s) + 4 \text{H}_2\text{O}(l)$

- 4.52 Which of the following are redox reactions? For those that are, indicate which element is oxidized and which is reduced. For those that are not, indicate whether they are precipitation or acid-base reactions.
- (a) $\text{Cu}(\text{OH})_2(s) + 2 \text{HNO}_3(aq) \longrightarrow \text{Cu}(\text{NO}_3)_2(aq) + 2 \text{H}_2\text{O}(l)$
- (b) $\text{Fe}_2\text{O}_3(s) + 3 \text{CO}(g) \longrightarrow 2 \text{Fe}(s) + 3 \text{CO}_2(g)$
- (c) $\text{Sr}(\text{NO}_3)_2(aq) + \text{H}_2\text{SO}_4(aq) \longrightarrow \text{SrSO}_4(s) + 2 \text{HNO}_3(aq)$
- (d) $4 \text{Zn}(s) + 10 \text{H}^+(aq) + 2 \text{NO}_3^-(aq) \longrightarrow 4 \text{Zn}^{2+}(aq) + \text{N}_2\text{O}(g) + 5 \text{H}_2\text{O}(l)$

- 4.53 Write balanced molecular and net ionic equations for the reactions of (a) manganese with dilute sulfuric acid; (b) chromium with hydrobromic acid; (c) tin with hydrochloric acid; (d) aluminum with formic acid, HCOOH .
- 4.54 Write balanced molecular and net ionic equations for the reactions of (a) hydrochloric acid with nickel; (b) dilute sulfuric acid with iron; (c) hydrobromic acid with magnesium; (d) acetic acid, CH_3COOH , with zinc.
- 4.55 Using the activity series (Table 4.5), write balanced chemical equations for the following reactions. If no reaction occurs, simply write NR. (a) Iron metal is added to a solution of copper(II) nitrate; (b) zinc metal is added to a solution of magnesium sulfate; (c) hydrobromic acid is added to tin metal; (d) hydrogen gas is bubbled through an aqueous solution of nickel(II) chloride; (e) aluminum metal is added to a solution of cobalt(II) sulfate.
- 4.56 Based on the activity series (Table 4.5), what is the outcome (if any) of each of the following reactions?
- (a) $\text{Mn}(s) + \text{NiCl}_2(aq) \longrightarrow$
- (b) $\text{Cu}(s) + \text{Cr}(\text{CH}_3\text{COO})_3(aq) \longrightarrow$
- (c) $\text{Cr}(s) + \text{NiSO}_4(aq) \longrightarrow$
- (d) $\text{Pt}(s) + \text{HBr}(aq) \longrightarrow$
- (e) $\text{H}_2(g) + \text{CuCl}_2(aq) \longrightarrow$
- 4.57 The metal cadmium tends to form Cd^{2+} ions. The following observations are made: (i) When a strip of zinc metal is placed in $\text{CdCl}_2(aq)$, cadmium metal is deposited on

the strip. (ii) When a strip of cadmium metal is placed in $\text{Ni}(\text{NO}_3)_2(\text{aq})$, nickel metal is deposited on the strip. (a) Write net ionic equations to explain each of the observations made above. (b) What can you conclude about the position of cadmium in the activity series? (c) What experiments would you need to perform to locate more precisely the position of cadmium in the activity series?

4.58 (a) Use the following reactions to prepare an activity series for the halogens:



(b) Relate the positions of the halogens in the periodic table with their locations in this activity series. (c) Predict whether a reaction occurs when the following reagents are mixed: $\text{Cl}_2(\text{aq})$ and $\text{KI}(\text{aq})$; $\text{Br}_2(\text{aq})$ and $\text{LiCl}(\text{aq})$.

Solution Composition; Molarity

- 4.59 (a) Is the concentration of a solution an intensive or an extensive property? (b) What is the difference between 0.50 mol HCl and 0.50 M HCl?
- 4.60 (a) Suppose you prepare 500 mL of a 0.10 M solution of some salt and then spill some of it. What happens to the concentration of the solution left in the container? (b) Suppose you prepare 500 mL of a 0.10 M aqueous solution of some salt and let it sit out, uncovered, for a long time, and some water evaporates. What happens to the concentration of the solution left in the container? (c) A certain volume of a 0.50 M solution contains 4.5 g of a salt. What mass of the salt is present in the same volume of a 2.50 M solution?
- 4.61 (a) Calculate the molarity of a solution that contains 0.0250 mol NH_4Cl in exactly 500 mL of solution. (b) How many moles of HNO_3 are present in 50.0 mL of a 2.50 M solution of nitric acid? (c) How many milliliters of 1.50 M KOH solution are needed to provide 0.275 mol of KOH?
- 4.62 (a) Calculate the molarity of a solution made by dissolving 0.750 grams of Na_2SO_4 in enough water to form exactly 850 mL of solution. (b) How many moles of KMnO_4 are present in 250 mL of a 0.0475 M solution? (c) How many milliliters of 11.6 M HCl solution are needed to obtain 0.250 mol of HCl?
- 4.63 The average adult human male has a total blood volume of 5.0 L. If the concentration of sodium ion in this average individual is 0.135 M, what is the mass of sodium ion circulating in the blood?
- 4.64 A person suffering from hyponatremia has a sodium ion concentration in the blood of 0.118 M and a total blood volume of 4.6 L. What mass of sodium chloride would need to be added to the blood to bring the sodium ion concentration up to 0.138 M, assuming no change in blood volume?
- 4.65 The concentration of alcohol ($\text{CH}_3\text{CH}_2\text{OH}$) in blood, called the "blood alcohol concentration" or BAC, is given in units of grams of alcohol per 100 mL of blood. The legal definition of intoxication, in many states of the United States, is that the BAC is 0.08 or higher. What is the concentration of alcohol, in terms of molarity, in blood if the BAC is 0.08?
- 4.66 The average adult male has a total blood volume of 5.0 L. After drinking a few beers, he has a BAC of 0.10 (see Exercise 4.65). What mass of alcohol is circulating in his blood?
- 4.67 Calculate (a) the number of grams of solute in 0.250 L of 0.175 M KBr, (b) the molar concentration of a solution containing 14.75 g of $\text{Ca}(\text{NO}_3)_2$ in 1.375 L, (c) the volume of 1.50 M Na_3PO_4 in milliliters that contains 2.50 g of solute.
- 4.68 (a) How many grams of solute are present in 50.0 mL of 0.488 M $\text{K}_2\text{Cr}_2\text{O}_7$? (b) If 4.00 g of $(\text{NH}_4)_2\text{SO}_4$ is dissolved in enough water to form 400 mL of solution, what is the molarity of the solution? (c) How many milliliters of 0.0250 M CuSO_4 contain 1.75 g of solute?
- 4.69 (a) Which will have the highest concentration of potassium ion: 0.20 M KCl, 0.15 M K_2CrO_4 , or 0.080 M K_3PO_4 ? (b) Which will contain the greater number of moles of potassium ion: 30.0 mL of 0.15 M K_2CrO_4 or 25.0 mL of 0.080 M K_3PO_4 ?
- 4.70 In each of the following pairs, indicate which has the higher concentration of Cl^- ion: (a) 0.10 M CaCl_2 or 0.15 M KCl solution, (b) 100 mL of 0.10 M KCl solution or 400 mL of 0.080 M LiCl solution, (c) 0.050 M HCl solution or 0.020 M CdCl_2 solution.
- 4.71 Indicate the concentration of each ion or molecule present in the following solutions: (a) 0.25 M NaNO_3 , (b) 1.3×10^{-2} M MgSO_4 , (c) 0.0150 M $\text{C}_6\text{H}_{12}\text{O}_6$, (d) a mixture of 45.0 mL of 0.272 M NaCl and 65.0 mL of 0.0247 M $(\text{NH}_4)_2\text{CO}_3$. Assume that the volumes are additive.
- 4.72 Indicate the concentration of each ion present in the solution formed by mixing (a) 42.0 mL of 0.170 M NaOH and 37.6 mL of 0.400 M NaOH, (b) 44.0 mL of 0.100 M Na_2SO_4 and 25.0 mL of 0.150 M KCl, (c) 3.60 g KCl in 75.0 mL of 0.250 M CaCl_2 solution. Assume that the volumes are additive.
- 4.73 (a) You have a stock solution of 14.8 M NH_3 . How many milliliters of this solution should you dilute to make 1000.0 mL of 0.250 M NH_3 ? (b) If you take a 10.0-mL portion of the stock solution and dilute it to a total volume of 0.500 L, what will be the concentration of the final solution?
- 4.74 (a) How many milliliters of a stock solution of 10.0 M HNO_3 would you have to use to prepare 0.450 L of 0.500 M HNO_3 ? (b) If you dilute 25.0 mL of the stock solution to a final volume of 0.500 L, what will be the concentration of the diluted solution?
- 4.75 (a) Starting with solid sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, describe how you would prepare 250 mL of a 0.250 M sucrose solution. (b) Describe how you would prepare 350.0 mL of 0.100 M $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ starting with 3.00 L of 1.50 M $\text{C}_{12}\text{H}_{22}\text{O}_{11}$.

- 4.76 (a) How would you prepare 175.0 mL of 0.150 M AgNO_3 solution starting with pure solute? (b) An experiment calls for you to use 100 mL of 0.50 M HNO_3 solution. All you have available is a bottle of 3.6 M HNO_3 . How would you prepare the desired solution?
- [4.77] Pure acetic acid, known as glacial acetic acid, is a liquid with a density of 1.049 g/mL at 25 °C. Calculate the molarity of a solution of acetic acid made by dissolving 20.00 mL of glacial acetic acid at 25 °C in enough water to make 250.0 mL of solution.
- [4.78] Glycerol, $\text{C}_3\text{H}_8\text{O}_3$, is a substance used extensively in the manufacture of cosmetics, foodstuffs, antifreeze, and plastics. Glycerol is a water-soluble liquid with a density of 1.2656 g/L at 15 °C. Calculate the molarity of a solution of glycerol made by dissolving 50.000 mL glycerol at 15 °C in enough water to make 250.00 mL of solution.

Solution Stoichiometry; Titrations

- 4.79 What mass of KCl is needed to precipitate the silver ions from 15.0 mL of 0.200 M AgNO_3 solution?
- 4.80 What mass of NaOH is needed to precipitate the Cd^{2+} ions from 35.0 mL of 0.500 M $\text{Cd}(\text{NO}_3)_2$ solution?
- 4.81 (a) What volume of 0.115 M HClO_4 solution is needed to neutralize 50.00 mL of 0.0875 M NaOH? (b) What volume of 0.128 M HCl is needed to neutralize 2.87 g of $\text{Mg}(\text{OH})_2$? (c) If 25.8 mL of AgNO_3 is needed to precipitate all the Cl^- ions in a 785-mg sample of KCl (forming AgCl), what is the molarity of the AgNO_3 solution? (d) If 45.3 mL of 0.108 M HCl solution is needed to neutralize a solution of KOH, how many grams of KOH must be present in the solution?
- 4.82 (a) How many milliliters of 0.120 M HCl are needed to completely neutralize 50.0 mL of 0.101 M $\text{Ba}(\text{OH})_2$ solution? (b) How many milliliters of 0.125 M H_2SO_4 are needed to neutralize 0.200 g of NaOH? (c) If 55.8 mL of BaCl_2 solution is needed to precipitate all the sulfate ion in a 752-mg sample of Na_2SO_4 , what is the molarity of the solution? (d) If 42.7 mL of 0.208 M HCl solution is needed to neutralize a solution of $\text{Ca}(\text{OH})_2$, how many grams of $\text{Ca}(\text{OH})_2$ must be in the solution?
- 4.83 Some sulfuric acid is spilled on a lab bench. You can neutralize the acid by sprinkling sodium bicarbonate on it and then mopping up the resultant solution. The sodium bicarbonate reacts with sulfuric acid as follows:
- $$2 \text{NaHCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \longrightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) + 2 \text{CO}_2(\text{g})$$
- Sodium bicarbonate is added until the fizzing due to the formation of $\text{CO}_2(\text{g})$ stops. If 27 mL of 6.0 M H_2SO_4 was spilled, what is the minimum mass of NaHCO_3 that must be added to the spill to neutralize the acid?
- 4.84 The distinctive odor of vinegar is due to acetic acid, CH_3COOH , which reacts with sodium hydroxide in the following fashion:
- $$\text{CH}_3\text{COOH}(\text{aq}) + \text{NaOH}(\text{aq}) \longrightarrow \text{H}_2\text{O}(\text{l}) + \text{NaC}_2\text{H}_3\text{O}_2(\text{aq})$$
- If 3.45 mL of vinegar needs 42.5 mL of 0.115 M NaOH to reach the equivalence point in a titration, how many grams of acetic acid are in a 1.00-qt sample of this vinegar?
- 4.85 A sample of solid $\text{Ca}(\text{OH})_2$ is stirred in water at 30 °C until the solution contains as much dissolved $\text{Ca}(\text{OH})_2$ as it can hold. A 100-mL sample of this solution is withdrawn and titrated with 5.00×10^{-2} M HBr. It requires 48.8 mL of the acid solution for neutralization. What is the molarity of the $\text{Ca}(\text{OH})_2$ solution? What is the solubility of $\text{Ca}(\text{OH})_2$ in water, at 30 °C, in grams of $\text{Ca}(\text{OH})_2$ per 100 mL of solution?
- 4.86 In a laboratory, 6.82 g of $\text{Sr}(\text{NO}_3)_2$ is dissolved in enough water to form 0.500 L of solution. A 0.100-L sample is withdrawn from this stock solution and titrated with a 0.0245 M solution of Na_2CrO_4 . What volume of Na_2CrO_4 solution is needed to precipitate all the $\text{Sr}^{2+}(\text{aq})$ as SrCrO_4 ?
- 4.87 A solution of 100.0 mL of 0.200 M KOH is mixed with a solution of 200.0 mL of 0.150 M NiSO_4 . (a) Write the balanced chemical equation for the reaction that occurs. (b) What precipitate forms? (c) What is the limiting reactant? (d) How many grams of this precipitate form? (e) What is the concentration of each ion that remains in solution?
- 4.88 A solution is made by mixing 12.0 g of NaOH and 75.0 mL of 0.200 M HNO_3 . (a) Write a balanced equation for the reaction that occurs between the solutes. (b) Calculate the concentration of each ion remaining in solution. (c) Is the resultant solution acidic or basic?
- [4.89] A 0.5895-g sample of impure magnesium hydroxide is dissolved in 100.0 mL of 0.2050 M HCl solution. The excess acid then needs 19.85 mL of 0.1020 M NaOH for neutralization. Calculate the percent by mass of magnesium hydroxide in the sample, assuming that it is the only substance reacting with the HCl solution.
- [4.90] A 1.248-g sample of limestone rock is pulverized and then treated with 30.00 mL of 1.035 M HCl solution. The excess acid then requires 11.56 mL of 1.010 M NaOH for neutralization. Calculate the percent by mass of calcium carbonate in the rock, assuming that it is the only substance reacting with the HCl solution.

ADDITIONAL EXERCISES

- 4.91 Explain why a titration experiment is a good way to measure the unknown concentration of a compound in solution.
- 4.92 The accompanying photo shows the reaction between a solution of $\text{Cd}(\text{NO}_3)_2$ and one of Na_2S . What is the identity of the precipitate? What ions remain in solution? Write the net ionic equation for the reaction.



- 4.93 Suppose you have a solution that might contain any or all of the following cations: Ni^{2+} , Ag^+ , Sr^{2+} , and Mn^{2+} . Addition of HCl solution causes a precipitate to form. After filtering off the precipitate, H_2SO_4 solution is added to the resultant solution and another precipitate forms. This is filtered off, and a solution of NaOH is added to the resulting solution. No precipitate is observed. Which ions are present in each of the precipitates? Which of the four ions listed above must be absent from the original solution?
- 4.94 You choose to investigate some of the solubility guidelines for two ions not listed in Table 4.1, the chromate ion (CrO_4^{2-}) and the oxalate ion ($\text{C}_2\text{O}_4^{2-}$). You are given 0.01 M solutions (A, B, C, D) of four water-soluble salts:

Solution	Solute	Color of Solution
A	Na_2CrO_4	Yellow
B	$(\text{NH}_4)_2\text{C}_2\text{O}_4$	Colorless
C	AgNO_3	Colorless
D	CaCl_2	Colorless

When these solutions are mixed, the following observations are made:

Expt Number	Solutions	
	Mixed	Result
1	A + B	No precipitate, yellow solution
2	A + C	Red precipitate forms
3	A + D	No precipitate, yellow solution
4	B + C	White precipitate forms
5	B + D	White precipitate forms
6	C + D	White precipitate forms

- (a) Write a net ionic equation for the reaction that occurs in each of the experiments. (b) Identify the precipitate formed, if any, in each of the experiments. (c) Based on these limited observations, which ion tends to form the more soluble salts, chromate or oxalate?
- 4.95 Antacids are often used to relieve pain and promote healing in the treatment of mild ulcers. Write balanced net ionic equations for the reactions between the $\text{HCl}(aq)$ in the stomach and each of the following substances used in various antacids: (a) $\text{Al}(\text{OH})_3(s)$, (b) $\text{Mg}(\text{OH})_2(s)$, (c) $\text{MgCO}_3(s)$, (d) $\text{NaAl}(\text{CO}_3)(\text{OH})_2(s)$, (e) $\text{CaCO}_3(s)$.
- [4.96] Salts of the sulfite ion, SO_3^{2-} , react with acids in a way similar to that of carbonates. (a) Predict the chemical formula, and name the weak acid that forms when the sulfite ion reacts with acids. (b) The acid formed in part (a) decomposes to form water and a gas. Predict the molecular formula, and name the gas formed. (c) Use a source book such as the *CRC Handbook of Chemistry and Physics* to confirm that the substance in part (b) is a gas under normal room-temperature conditions. (d) Write balanced net ionic equations of the reaction of $\text{HCl}(aq)$ with (i) $\text{Na}_2\text{SO}_3(aq)$, (ii) $\text{Ag}_2\text{SO}_3(s)$, (iii) $\text{KHSO}_3(s)$, and (iv) $\text{ZnSO}_3(aq)$.
- [4.97] The commercial production of nitric acid involves the following chemical reactions:
- $$4 \text{NH}_3(g) + 5 \text{O}_2(g) \longrightarrow 4 \text{NO}(g) + 6 \text{H}_2\text{O}(g)$$
- $$2 \text{NO}(g) + \text{O}_2(g) \longrightarrow 2 \text{NO}_2(g)$$
- $$3 \text{NO}_2(g) + \text{H}_2\text{O}(l) \longrightarrow 2 \text{HNO}_3(aq) + \text{NO}(g)$$
- (a) Which of these reactions are redox reactions? (b) In each redox reaction identify the element undergoing oxidation and the element undergoing reduction.
- 4.98 Use Table 4.5 to predict which of the following ions can be reduced to their metal forms by reacting with zinc: (a) $\text{Na}^+(aq)$, (b) $\text{Pb}^{2+}(aq)$, (c) $\text{Mg}^{2+}(aq)$, (d) $\text{Fe}^{2+}(aq)$, (e) $\text{Cu}^{2+}(aq)$, (f) $\text{Al}^{3+}(aq)$. Write the balanced net ionic equation for each reaction that occurs.
- [4.99] Lanthanum metal forms cations with a charge of 3+. Consider the following observations about the chemistry of lanthanum: When lanthanum metal is exposed to air, a white solid (compound A) is formed that contains lanthanum and one other element. When lanthanum metal is added to water, gas bubbles are observed and a different white solid (compound B) is formed. Both A and B dissolve in hydrochloric acid to give a clear solution. When either of these solutions is evaporated, a soluble white solid (compound C) remains. If compound C is dissolved in water and sulfuric acid is added, a white precipitate (compound D) forms. (a) Propose identities for the substances A, B, C, and D. (b) Write net ionic equations for all the reactions described. (c) Based on the preceding observations, what can be said about the position of lanthanum in the activity series (Table 4.5)?
- 4.100 A 35.0-mL sample of 1.00 M KBr and a 60.0-mL sample of 0.600 M KBr are mixed. The solution is then heated to evaporate water until the total volume is 50.0 mL. What is the molarity of the KBr in the final solution?

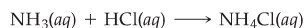
- 4.101** Using modern analytical techniques, it is possible to detect sodium ions in concentrations as low as 50 pg/mL. What is this detection limit expressed in (a) molarity of Na^+ , (b) Na^+ ions per cubic centimeter?
- 4.102** Hard water contains Ca^{2+} , Mg^{2+} , and Fe^{2+} , which interfere with the action of soap and leave an insoluble coating on the insides of containers and pipes when heated. Water softeners replace these ions with Na^+ . If 1500 L of hard water contains 0.020 M Ca^{2+} and 0.0040 M Mg^{2+} , how many moles of Na^+ are needed to replace these ions?
- 4.103** Tartaric acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$, has two acidic hydrogens. The acid is often present in wines and precipitates from solution as the wine ages. A solution containing an unknown concentration of the acid is titrated with NaOH . It requires 24.65 mL of 0.2500 M NaOH solution to titrate both acidic protons in 50.00 mL of the tartaric acid solution. Write a balanced net ionic equation for the neutralization reaction, and calculate the molarity of the tartaric acid solution.
- 4.104** The concentration of hydrogen peroxide in a solution is determined by titrating a 10.0-mL sample of the solution with permanganate ion.
- $$2 \text{MnO}_4^-(aq) + 5 \text{H}_2\text{O}_2(aq) + 6 \text{H}^+(aq) \longrightarrow 2 \text{Mn}^{2+}(aq) + 5 \text{O}_2(g) + 8 \text{H}_2\text{O}(l)$$
- If it takes 14.8 mL of 0.134 M MnO_4^- solution to reach the equivalence point, what is the molarity of the hydrogen peroxide solution?
- 4.105** A solid sample of $\text{Zn}(\text{OH})_2$ is added to 0.350 L of 0.500 M aqueous HBr . The solution that remains is still acidic. It is then titrated with 0.500 M NaOH solution, and it takes 88.5 mL of the NaOH solution to reach the equivalence point. What mass of $\text{Zn}(\text{OH})_2$ was added to the HBr solution?

INTEGRATIVE EXERCISES

- 4.106** (a) By titration, 15.0 mL of 0.1008 M sodium hydroxide is needed to neutralize a 0.2053-g sample of an organic acid. What is the molar mass of the acid if it is monoprotic? (b) An elemental analysis of the acid indicates that it is composed of 5.89% H, 70.6% C, and 23.5% O by mass. What is its molecular formula?
- 4.107** A 3.455-g sample of a mixture was analyzed for barium ion by adding a small excess of sulfuric acid to an aqueous solution of the sample. The resultant reaction produced a precipitate of barium sulfate, which was collected by filtration, washed, dried, and weighed. If 0.2815 g of barium sulfate was obtained, what was the mass percentage of barium in the sample?
- 4.108** A tanker truck carrying 5.0×10^3 kg of concentrated sulfuric acid solution tips over and spills its load. If the sulfuric acid is 95.0% H_2SO_4 by mass and has a density of 1.84 g/mL, how many kilograms of sodium carbonate must be added to neutralize the acid?
- 4.109** A sample of 5.53 g of $\text{Mg}(\text{OH})_2$ is added to 25.0 mL of 0.200 M HNO_3 . (a) Write the chemical equation for the reaction that occurs. (b) Which is the limiting reactant in the reaction? (c) How many moles of $\text{Mg}(\text{OH})_2$, HNO_3 , and $\text{Mg}(\text{NO}_3)_2$ are present after the reaction is complete?
- 4.110** A sample of 1.50 g of lead(II) nitrate is mixed with 125 mL of 0.100 M sodium sulfate solution. (a) Write the chemical equation for the reaction that occurs. (b) Which is the limiting reactant in the reaction? (c) What are the concentrations of all ions that remain in solution after the reaction is complete?
- 4.111** A mixture contains 76.5% NaCl , 6.5% MgCl_2 , and 17.0% Na_2SO_4 by mass. What is the molarity of Cl^- ions in a solution formed by dissolving 7.50 g of the mixture in enough water to form 500.0 mL of solution?
- 4.112** The average concentration of bromide ion in seawater is 65 mg of bromide ion per kg of seawater. What is the molarity of the bromide ion if the density of the seawater is 1.025 g/mL?
- 4.113** The mass percentage of chloride ion in a 25.00-mL sample of seawater was determined by titrating the sample with silver nitrate, precipitating silver chloride. It took 42.58 mL of 0.2997 M silver nitrate solution to reach the equivalence point in the titration. What is the mass percentage of chloride ion in the seawater if its density is 1.025 g/mL?
- 4.114** The arsenic in a 1.22-g sample of a pesticide was converted to AsO_4^{3-} by suitable chemical treatment. It was then titrated using Ag^+ to form Ag_3AsO_4 as a precipitate. (a) What is the oxidation state of As in AsO_4^{3-} ? (b) Name Ag_3AsO_4 by analogy to the corresponding compound containing phosphorus in place of arsenic. (c) If it took 25.0 mL of 0.102 M Ag^+ to reach the equivalence point in this titration, what is the mass percentage of arsenic in the pesticide?
- 4.115** The newest U.S. standard for arsenate in drinking water, mandated by the Safe Drinking Water Act, required that by January 2006, public water supplies must contain no greater than 10 parts per billion (ppb) arsenic. If this arsenic is present as arsenate, AsO_4^{3-} , what mass of sodium arsenate would be present in a 1.00-L sample of drinking water that just meets the standard?
- 4.116** The safe drinking water standard for arsenic (which is usually found as arsenate, see 4.115) is 50 parts per billion (ppb) in most developing countries. (a) How many grams of sodium arsenate are in 55 gallons of water, if the concentration of arsenate is 50 ppb? (b) In 1993,

naturally occurring arsenic was discovered as a major contaminant in the drinking water across the country of Bangladesh. Approximately 12 million people in Bangladesh still drink water from wells that have higher concentrations of arsenic than the standard. Recently, a chemistry professor from George Mason University was awarded a \$1 million Grainger Challenge Prize for Sustainability for his development of a simple, inexpensive system for filtering naturally occurring arsenic from drinking water. The system uses buckets of sand, cast iron, activated carbon, and wood chips for trapping arsenic-containing minerals. Assuming the efficiency of such a bucket system is 90% (meaning, 90% of the arsenic that comes in is retained in the bucket and 10% passes out of the bucket), how many times should water that is 500 ppb in arsenic be passed through to meet the 50 ppb standard?

- [4.117] Federal regulations set an upper limit of 50 parts per million (ppm) of NH_3 in the air in a work environment [that is, 50 molecules of $\text{NH}_3(g)$ for every million molecules in the air]. Air from a manufacturing operation was drawn through a solution containing 1.00×10^2 mL of 0.0105 M HCl. The NH_3 reacts with HCl as follows:



After drawing air through the acid solution for 10.0 min at a rate of 10.0 L/min, the acid was titrated. The remaining acid needed 13.1 mL of 0.0588 M NaOH to reach the equivalence point. (a) How many grams of NH_3 were drawn into the acid solution? (b) How many ppm of NH_3 were in the air? (Air has a density of 1.20 g/L and an average molar mass of 29.0 g/mol under the conditions of the experiment.) (c) Is this manufacturer in compliance with regulations?