



Science may be described as the art of systemic oversimplification—the art of discerning what we may with advantage omit.

—KARL POPPER (1902–1994)

CHAPTER

5

Introduction to Solutions and Aqueous Reactions

In this chapter, we turn to describing chemical reactions that occur in water. You have probably witnessed many of these types of reactions in your daily life because they are so common. Have you ever mixed baking soda with vinegar and observed the subsequent bubbling? Have you noticed the hard water deposits that form on plumbing fixtures? These reactions—and many others, including those that occur within the watery environment of living cells—are aqueous chemical reactions, the subject of this chapter.



The spherified cherry, a product of a cooking technique known as molecular gastronomy.

- 5.1** Molecular Gastronomy and the Spherified Cherry 167
- 5.2** Solution Concentration 168
- 5.3** Solution Stoichiometry 173
- 5.4** Types of Aqueous Solutions and Solubility 175
- 5.5** Precipitation Reactions 179

- 5.6** Representing Aqueous Reactions: Molecular, Ionic, and Net Ionic Equations 183
- 5.7** Acid–Base Reactions 185
- 5.8** Gas-Evolution Reactions 191
- 5.9** Oxidation–Reduction Reactions 193

LEARNING OUTCOMES 203

5.1

Molecular Gastronomy and the Spherified Cherry

One of my favorite chefs is the Spaniard José Andrés, who owns and operates restaurants all over the world including *The Bazaar* in Los Angeles and Miami, and *Jaleo* in Washington, D.C. Andrés was born in the province of Asturias, Spain, the same province as my grandfather (no wonder I like him so much). Andrés is not only an amazing chef, who is credited with bringing small plates (or *tapas*) to America, but he also dabbles in *molecular gastronomy*, a way of preparing food that involves chemistry.



► **FIGURE 5.1** The Spherified Cherry

Cherry The spherified cherry is made by precipitating an encapsulating layer around cherry juice.

A common chemical reaction in molecular gastronomy is precipitation. In a *precipitation reaction*, two *solutions*—homogeneous mixtures often containing a solid dissolved in a liquid—are mixed. Upon mixing, a solid (or *precipitate*) forms. For example, when we mix solutions of lead(II) nitrate and potassium iodide, a brilliant yellow solid forms. The solid is lead(II) iodide.

In molecular gastronomy, chefs use a similar precipitation reaction—called spherification—to encapsulate liquids. Among the most popular molecular gastronomy creations is the spherified cherry (Figure 5.1▲). To make a spherified cherry, chefs take juice from real cherries and mix it with a calcium salt (such as calcium chloride), which dissolves in the cherry juice. They then carefully pour the cherry juice into a bath of sodium alginate. Sodium alginate is a sodium salt that dissolves into water, resulting in the presence of alginate ions. When the calcium ions in the cherry juice encounter the alginate ions in the bath, a precipitation reaction occurs. In this case, the precipitation reaction forms in the area immediately surrounding the cherry juice, forming an encapsulating sphere around the juice. The result is a spherical, edible “cherry” that ruptures in the mouth and releases its juice.

In this chapter, we explore solutions, focusing especially on *aqueous* solutions (solutions in which one component is water). The cherry juice and calcium chloride mixture just discussed is an example of an aqueous solution. Other common aqueous solutions include seawater, vinegar, and the watery environment within biological cells. We will also explore the chemical reactions that occur within solutions, such as precipitation reactions, which have many common applications.

WATCH NOW!

KEY CONCEPT VIDEO 5.2

Solution Concentration

5.2

Solution Concentration

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

Solution Concentration

The amount of solute in a solution is variable. For example, we can add just a little salt to water to make a **dilute solution**, one that contains a small amount of solute relative to the solvent, or we can add a lot of salt to water to make a **concentrated solution**, one that contains a large amount of solute relative to the solvent (Figure 5.2►). A common way to express solution concentration is **molarity (M)**, the amount of solute (in moles) divided by the volume of solution (in liters):

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

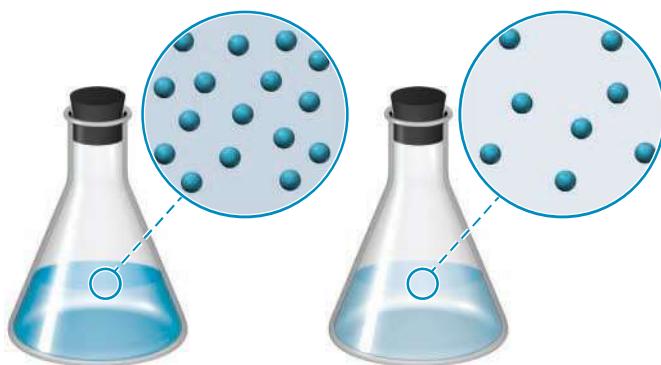
Concentrated and Dilute Solutions

Concentrated solution:

Relatively large amount of solute.

Diluted solution:

Relatively small amount of solute.

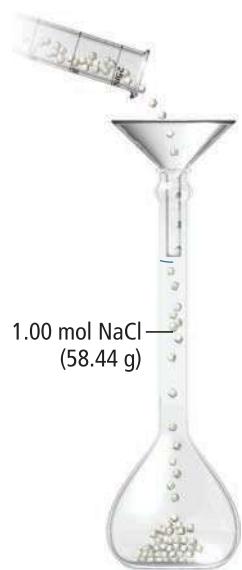


◀ FIGURE 5.2 Concentrated and Dilute Solutions

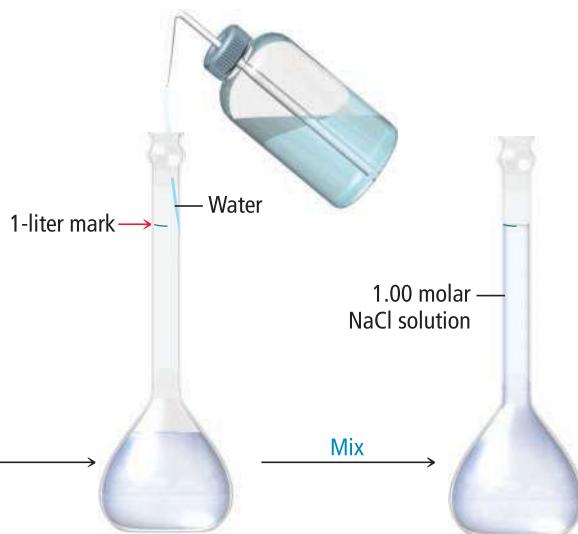
Notice that molarity is a ratio of the amount of solute per liter of *solution*, not per liter of solvent. To make an aqueous solution of a specified molarity, we usually put the solute into a flask and then add water to reach the desired volume of solution. For example, to make 1 L of a 1 M NaCl solution, we add 1 mol of NaCl to a flask and then add enough water to make 1 L of solution (Figure 5.3▼). We *do not* combine 1 mol of NaCl with 1 L of water because the resulting solution would have a total volume different from 1 L and therefore a molarity different than 1 M. To calculate molarity, divide the amount of the solute in moles by the volume of the solution (solute *and* solvent) in liters, as shown in Example 5.1.

Preparing a Solution of Specified Concentration

Weigh out and add 1.00 mol of NaCl.



Add water until solid is dissolved. Then add additional water until the 1-liter mark is reached.



The result is a 1.00 molar NaCl solution.

◀ FIGURE 5.3 Preparing a 1 Molar NaCl Solution

MOLARITY How many moles of solute are required to make 3.0 L of a 2.0 M solution?

- (a) 2.0 mol solute
- (b) 3.0 mol solute
- (c) 4.0 mol solute
- (d) 6.0 mol solute

5.1

Cc
Conceptual Connection

ANSWER NOW!



EXAMPLE 5.1 Calculating Solution Concentration

What is the molarity of a solution containing 25.5 g KBr dissolved in enough water to make 1.75 L of solution?



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

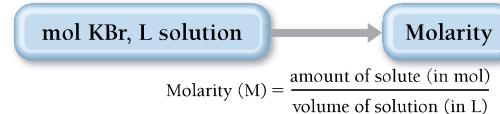
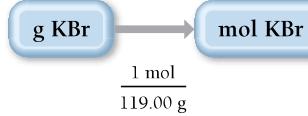
GIVEN: 25.5 g KBr, 1.75 L of solution

FIND: molarity (M)

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

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

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

CONCEPTUAL PLAN**RELATIONSHIPS USED**

molar mass of KBr = 119.00 g/mol

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

SOLUTION

$$25.5 \text{ g KBr} \times \frac{1 \text{ mol KBr}}{119.00 \text{ g KBr}} = 0.21429 \text{ mol KBr}$$

$$\begin{aligned} \text{molarity (M)} &= \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}} \\ &= \frac{0.21429 \text{ mol KBr}}{1.75 \text{ L solution}} \\ &= 0.122 \text{ M} \end{aligned}$$

CHECK The units of the answer (M) are correct. The magnitude is reasonable since common solutions range in concentration from 0 to about 18 M. Concentrations significantly above 18 M are suspect and should be double-checked.

FOR PRACTICE 5.1 Calculate the molarity of a solution made by adding 45.4 g of NaNO₃ to a flask and dissolving it with water to create a total volume of 2.50 L.

FOR MORE PRACTICE 5.1 What mass of KBr (in grams) do you need to make 250.0 mL of a 1.50 M KBr solution?

Using Molarity in Calculations

We can use the molarity of a solution as a conversion factor between moles of the solute and liters of the solution. For example, a 0.500 M NaCl solution contains 0.500 moles NaCl for every liter of solution:



This conversion factor converts from liters solution to mole NaCl. If we want to convert in the other direction, we invert the conversion factor:



Example 5.2 illustrates how to use molarity in this way.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 5.2

EXAMPLE 5.2 Using Molarity in Calculations

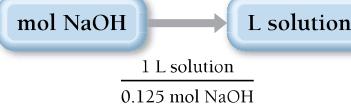
How many liters of a 0.125 M NaOH solution contain 0.255 mol of NaOH?

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

GIVEN: 0.125 M NaOH solution, 0.255 mol NaOH

FIND: volume of NaOH solution (in L)

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

CONCEPTUAL PLAN**RELATIONSHIPS USED**

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

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

SOLUTION

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

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

FOR PRACTICE 5.2 How many grams of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) are in 1.55 L of a 0.758 M sucrose solution?

FOR MORE PRACTICE 5.2 How many mL of a 0.155 M KCl solution contain 2.55 g KCl?

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

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



ANSWER NOW!

**Solution Dilution**

To save space in storerooms, laboratories often store solutions in concentrated forms called **stock solutions**. For example, hydrochloric acid is frequently stored as a 12 M stock solution. However, many lab procedures call for much less concentrated hydrochloric acid solutions, so we must dilute the stock solution to the required concentration. How do we know how much of the stock solution to use? The easiest way to solve dilution problems is to use the following dilution equation:

$$M_1 V_1 = M_2 V_2 \quad [5.1]$$

where M_1 and V_1 are the molarity and volume of the initial concentrated solution, and M_2 and V_2 are the molarity and volume of the final diluted solution. This equation works because the molarity multiplied by the volume gives the number of moles of solute, which is the same in both solutions.

$$\begin{aligned} M_1 V_1 &= M_2 V_2 \\ \text{mol}_1 &= \text{mol}_2 \end{aligned}$$

In other words, the number of moles of solute does not change when we dilute a solution.

For example, suppose a laboratory procedure calls for 3.00 L of a 0.500 M CaCl_2 solution. How should we prepare this solution from a 10.0 M stock solution? We solve

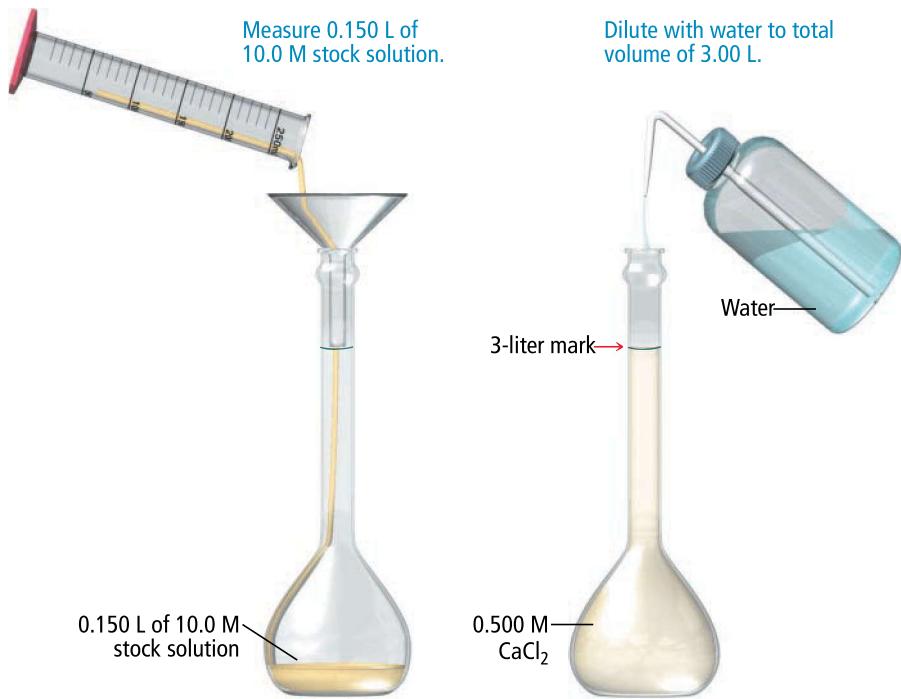
When diluting acids, always add the concentrated acid to the water. Never add water to concentrated acid solutions, as the heat generated may cause the concentrated acid to splatter and burn your skin.

Equation 5.1 for V_1 , the volume of the stock solution required for the dilution, and then substitute in the correct values to calculate it.

$$\begin{aligned} M_1 V_1 &= M_2 V_2 \\ V_1 &= \frac{M_2 V_2}{M_1} \\ &= \frac{0.500 \text{ mol/L} \times 3.00 \text{ L}}{10.0 \text{ mol/L}} \\ &= 0.150 \text{ L} \end{aligned}$$

Consequently, we make the solution by diluting 0.150 L of the stock solution to a total volume of 3.00 L (V_2). The resulting solution will be 0.500 M in CaCl_2 (Figure 5.4▼).

Diluting a Solution



$$\begin{aligned} M_1 V_1 &= M_2 V_2 \\ \frac{10.0 \text{ mol}}{x} \times 0.150 \text{ L} &= \frac{0.500 \text{ mol}}{x} \times 3.00 \text{ L} \\ 1.50 \text{ mol} &= 1.50 \text{ mol} \end{aligned}$$

► FIGURE 5.4 Preparing 3.00 L of 0.500 M CaCl_2 from a 10.0 M Stock Solution

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 5.3

EXAMPLE 5.3 Solution Dilution

To what volume should you dilute 0.200 L of a 15.0 M NaOH solution to obtain a 3.00 M NaOH solution?

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

GIVEN: $V_1 = 0.200 \text{ L}$

$M_1 = 15.0 \text{ M}$

$M_2 = 3.00 \text{ M}$

FIND: V_2

STRATEGIZE Equation 5.1 relates the initial and final volumes and concentrations for solution dilution problems. You are asked to find V_2 . The other quantities (V_1 , M_1 , and M_2) are all given in the problem.

CONCEPTUAL PLAN

$$\begin{array}{ccc} V_1, M_1, M_2 & \longrightarrow & V_2 \\ M_1 V_1 = M_2 V_2 \end{array}$$

RELATIONSHIPS USED

$$M_1 V_1 = M_2 V_2$$

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

Substitute in the required quantities and calculate V_2 .

Make the solution by diluting 0.200 L of the stock solution to a total volume of 1.00 L (V_2). The resulting solution will have a concentration of 3.00 M.

SOLUTION $M_1V_1 = M_2V_2$

$$V_2 = \frac{M_1V_1}{M_2}$$

$$= \frac{15.0 \text{ mol/L} \times 0.200 \text{ L}}{3.00 \text{ mol/L}}$$

$$= 1.00 \text{ L}$$

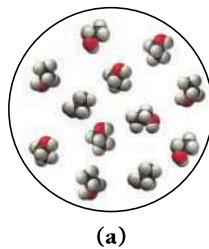
CHECK The final units (L) are correct. The magnitude of the answer is reasonable because the solution is diluted from 15.0 M to 3.00 M, a factor of five. Therefore, the volume should increase by a factor of five.

FOR PRACTICE 5.3 To what volume (in mL) should you dilute 100.0 mL of a 5.00 M CaCl_2 solution to obtain a 0.750 M CaCl_2 solution?

FOR MORE PRACTICE 5.3 What volume of a 6.00 M NaNO_3 solution should you use to make 0.525 L of a 1.20 M NaNO_3 solution?

SOLUTION DILUTION The image shown at the far right represents a small volume within 500 mL of aqueous ethanol ($\text{CH}_3\text{CH}_2\text{OH}$) solution. (The water molecules have been omitted for clarity.)

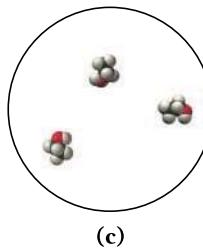
Which of the following images best represents the same volume of the solution after we add an additional 500 mL of water?



(a)



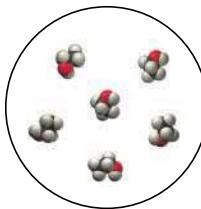
(b)



(c)



ANSWER NOW!



5.3

Solution Stoichiometry

In Section 4.3, we discussed how we can use the coefficients in chemical equations as conversion factors between the amounts of reactants (in moles) and the amounts of products (in moles). In aqueous reactions, quantities of reactants and products are often specified in terms of volumes and concentrations. We can use the volume and concentration of a reactant or product to calculate its amount in moles. We can then use the stoichiometric coefficients in the chemical equation to convert to the amount of another reactant or product in moles. The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product:



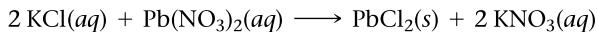
We make the conversions between solution volumes and amounts of solute in moles using the molarities of the solutions. We make the conversions between amounts in moles of A and B using the stoichiometric coefficients from the balanced chemical equation. Example 5.4 demonstrates solution stoichiometry.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 5.4

EXAMPLE 5.4 Solution Stoichiometry

What volume (in L) of a 0.150 M KCl solution will completely react with 0.150 L of a 0.175 M Pb(NO₃)₂ solution according to the following balanced chemical equation?

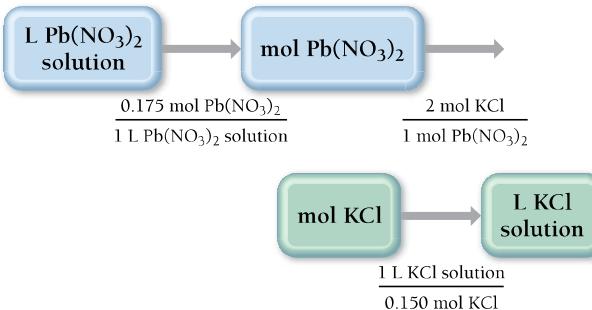


SORT You are given the volume and concentration of a Pb(NO₃)₂ solution. You are asked to find the volume of KCl solution (of a given concentration) required to react with it.

STRATEGIZE The conceptual plan has the form: volume A → amount A (in moles) → amount B (in moles) → volume B. Use the molar concentrations of the KCl and Pb(NO₃)₂ solutions as conversion factors between the number of moles of reactants in these solutions and their volumes. Use the stoichiometric coefficients from the balanced equation to convert between number of moles of Pb(NO₃)₂ and number of moles of KCl.

GIVEN: 0.150 L of Pb(NO₃)₂ solution, 0.175 M Pb(NO₃)₂ solution, 0.150 M KCl solution

FIND: volume KCl solution (in L)

CONCEPTUAL PLAN**RELATIONSHIPS USED**

$$\text{M Pb}(\text{NO}_3)_2 = \frac{0.175 \text{ mol Pb}(\text{NO}_3)_2}{1 \text{ L Pb}(\text{NO}_3)_2 \text{ solution}}$$

$$2 \text{ mol KCl} : 1 \text{ mol Pb}(\text{NO}_3)_2$$

$$\text{M KCl} = \frac{0.150 \text{ mol KCl}}{1 \text{ L KCl solution}}$$

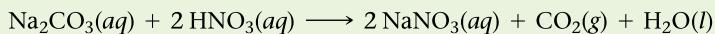
SOLVE Begin with L Pb(NO₃)₂ solution and follow the conceptual plan to arrive at L KCl solution.

SOLUTION

$$\begin{aligned}
 0.150 \text{ L Pb}(\text{NO}_3)_2 \text{ solution} &\times \frac{0.175 \text{ mol Pb}(\text{NO}_3)_2}{1 \text{ L Pb}(\text{NO}_3)_2 \text{ solution}} \\
 &\times \frac{2 \text{ mol KCl}}{1 \text{ mol Pb}(\text{NO}_3)_2} \times \frac{1 \text{ L KCl solution}}{0.150 \text{ mol KCl}} = 0.350 \text{ L KCl solution}
 \end{aligned}$$

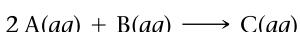
CHECK The final units (L KCl solution) are correct. The magnitude (0.350 L) is reasonable because the reaction stoichiometry requires 2 mol of KCl per mole of Pb(NO₃)₂. Since the concentrations of the two solutions are not very different (0.150 M compared to 0.175 M), the volume of KCl required is roughly two times the 0.150 L of Pb(NO₃)₂ given in the problem.

FOR PRACTICE 5.4 What volume (in mL) of a 0.150 M HNO₃ solution will completely react with 35.7 mL of a 0.108 M Na₂CO₃ solution according to the following balanced chemical equation?



FOR MORE PRACTICE 5.4 In the previous reaction, what mass (in grams) of carbon dioxide forms?

ANSWER NOW!

**SOLUTION STOICHIOMETRY** Consider the reaction:

What is the limiting reactant if you mix equal volumes of a 1 M solution of A and a 1 M solution of B?

- (a) A (b) B

5.4

Types of Aqueous Solutions and Solubility

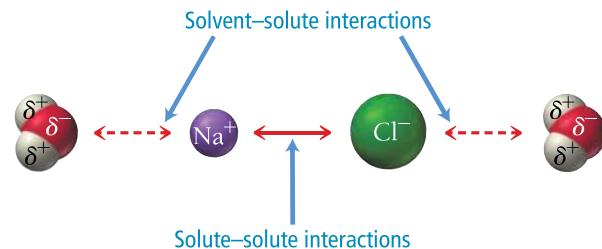
Consider two familiar aqueous solutions: salt water and sugar water. Salt water is a homogeneous mixture of NaCl and H₂O, and sugar water is a homogeneous mixture of C₁₂H₂₂O₁₁ and H₂O. You may have made these solutions yourself by adding table salt or sugar to water. As you stir either of these two substances into the water, the substance seems to disappear. However, you know that the original substance is still present because the mixture tastes salty or sweet. How do solids such as salt and sugar dissolve in water?

When a solid is put into a liquid solvent, the attractive forces that hold the solid together (the solute–solute interactions) compete with the attractive forces between the solvent molecules and the particles that compose the solid (the solvent–solute interactions), as shown in Figure 5.5►. For example, when sodium chloride is put into water, there is a competition between the attraction of Na⁺ cations and Cl⁻ anions to each other (due to their opposite charges) and the attraction of Na⁺ and Cl⁻ to water molecules. The attraction of Na⁺ and Cl⁻ to water is based on the *polar nature* of the water molecule. For reasons we discuss later in this book (Section 10.6), the oxygen atom in water is electron-rich, giving it a partial negative charge (δ^-), as shown in Figure 5.6►. The hydrogen atoms, in contrast, are electron-poor, giving them a partial positive charge (δ^+). As a result, the positively charged sodium ions are strongly attracted to the oxygen side of the water molecule (which has a partial negative charge), and the negatively charged chloride ions are attracted to the hydrogen side of the water molecule (which has a partial positive charge), as shown in Figure 5.7▼. In the case of NaCl, the attraction between the separated ions and the water molecules overcomes the attraction of sodium and chloride ions to each other, and the sodium chloride dissolves in the water (Figure 5.8▼).

Electrolyte and Nonelectrolyte Solutions

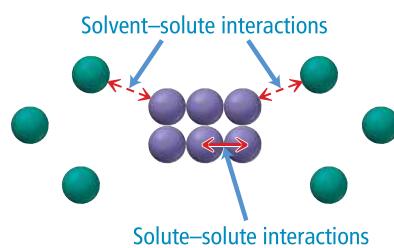
As Figure 5.9► illustrates, a salt solution conducts electricity while a sugar solution does not. The difference between the ways that salt (an ionic compound) and sugar (a molecular compound) dissolve in water illustrates a fundamental difference between types of solutions. Ionic compounds such as the sodium chloride in the previous example dissociate into their component ions when they dissolve in water. An NaCl solution, represented as NaCl(*aq*), does not contain any NaCl units, but rather dissolved Na⁺ ions and Cl⁻ ions. The dissolved ions act as charge carriers, allowing the solution to conduct electricity. Substances that dissolve in water to form solutions that conduct electricity are **electrolytes**. Substances such

Interactions in a Sodium Chloride Solution

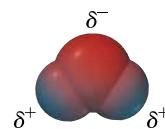


▲ FIGURE 5.7 Solute and Solvent Interactions in a Sodium Chloride Solution When sodium chloride is put into water, the attraction of Na⁺ and Cl⁻ ions to water molecules competes with the attraction between the oppositely charged ions themselves.

Solute and Solvent Interactions

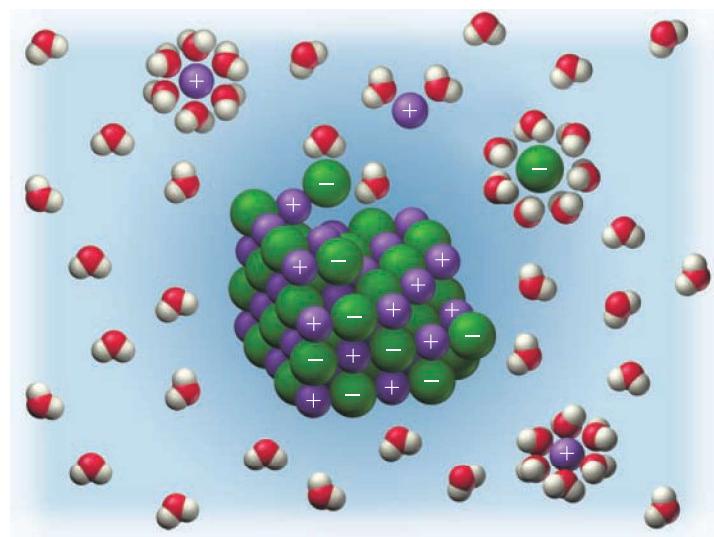


▲ FIGURE 5.5 Solute and Solvent Interactions



▲ FIGURE 5.6 Charge Distribution in Water An uneven distribution of electrons causes the oxygen side of the water molecule to have a partial negative charge and the hydrogen side to have a partial positive charge.

Dissolution of an Ionic Compound

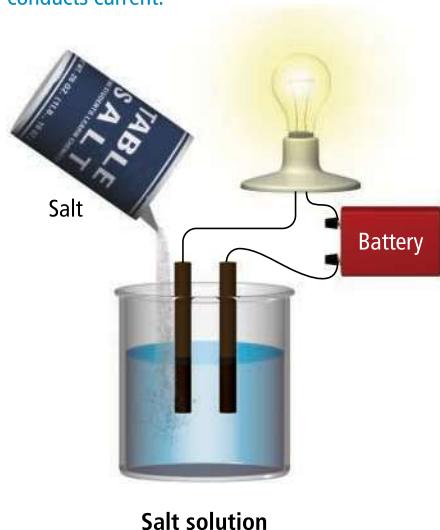


▲ FIGURE 5.8 Sodium Chloride Dissolving in Water The attraction between water molecules and the ions of sodium chloride causes NaCl to dissolve in the water.

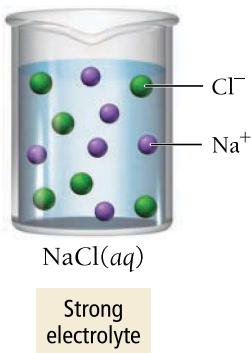
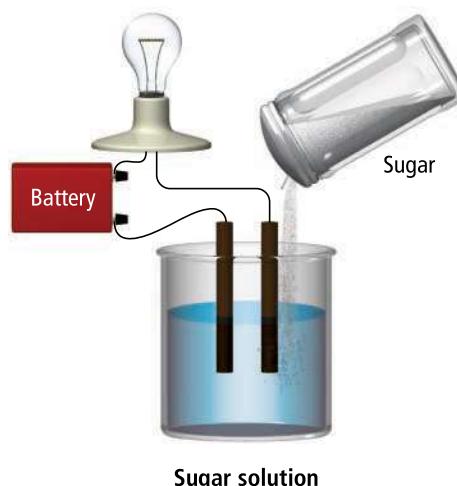
► FIGURE 5.9 Electrolyte and Nonelectrolyte Solutions

Electrolyte and Nonelectrolyte Solutions

An electrolyte solution conducts current.



A nonelectrolyte solution does not conduct current.

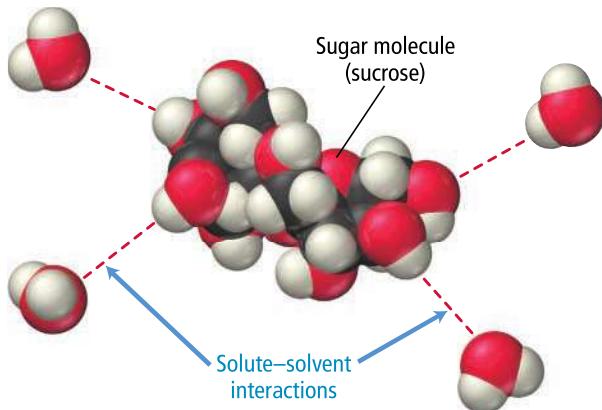


as sodium chloride that completely dissociate into ions when they dissolve in water are **strong electrolytes**, and the resulting solutions are *strong electrolyte solutions*.

In contrast to sodium chloride, sugar is a molecular compound. Most molecular compounds—with the important exception of acids, which we discuss shortly—dissolve in water as intact molecules. Sugar dissolves because the attraction between sugar molecules and water molecules (shown in Figure 5.10▼) overcomes the attraction of sugar molecules to each other (Figure 5.11▼). So unlike a sodium chloride solution (which is composed of dissociated ions), a sugar solution is composed of intact $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ molecules homogeneously mixed with the water molecules. Compounds such as sugar that do not dissociate into ions when dissolved in water are called **nonelectrolytes**, and the resulting solutions—called *nonelectrolyte solutions*—do not conduct electricity.

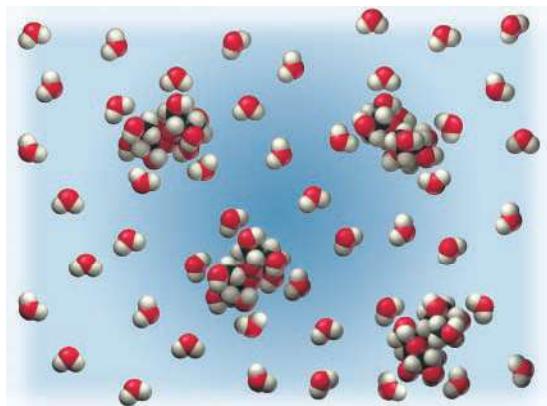
Acids, which we introduced in Section 3.6, are molecular compounds, but they ionize—form ions—when they dissolve in water. Hydrochloric acid (HCl) is a molecular compound that ionizes into H^+ and Cl^- when it dissolves in water. HCl is an example of a **strong acid**, one that completely ionizes in solution. Since strong acids completely

Interactions between Sugar and Water Molecules



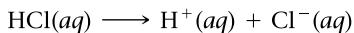
▲ FIGURE 5.10 Sugar and Water Interactions Partial charges on sugar molecules and water molecules (which we will discuss more fully in Chapter 12) result in attractions between the sugar molecules and water molecules.

Sugar Solution

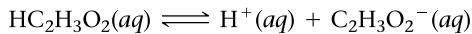


▲ FIGURE 5.11 A Sugar Solution Sugar dissolves because the attractions between sugar molecules and water molecules, which both contain a distribution of electrons that results in partial positive and partial negative charges, overcome the attractions between sugar molecules to each other.

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

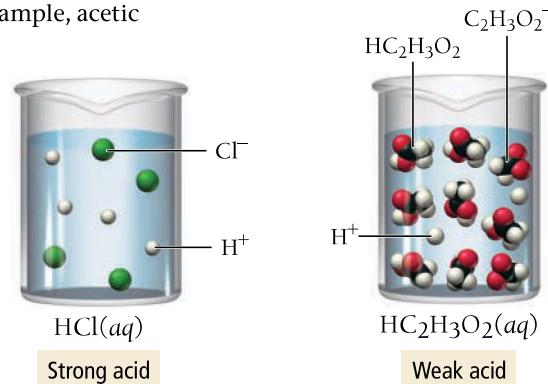


Many acids are **weak acids**; they do not completely ionize in water. For example, acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$), the acid in vinegar, is a weak acid. A solution of a weak acid is composed mostly of the nonionized acid—only a small percentage of the acid molecules ionize. We represent the partial ionization of a weak acid with opposing half arrows between the reactants and products:

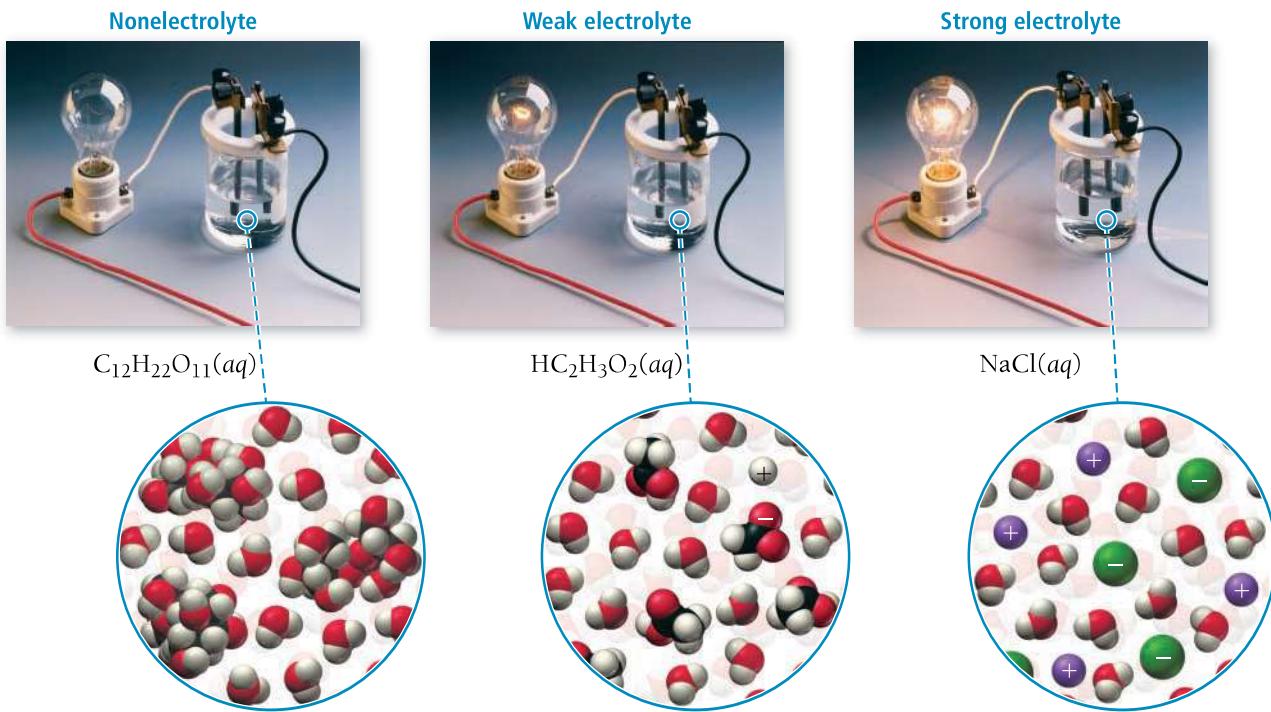


Weak acids are **weak electrolytes**, and the resulting solutions—called *weak electrolyte solutions*—conduct electricity only weakly. Figure 5.12▼ summarizes the electrolytic properties of solutions.

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



Electrolytic Properties of Solutions



▲ FIGURE 5.12 Electrolytic Properties of Solutions

ELECTROLYTE SOLUTIONS Which aqueous solution conducts electricity?

- (a) 1.0 M KBr (b) 1.0 M $\text{C}_6\text{H}_{12}\text{O}_6$ (c) 1.0 M CH_3OH

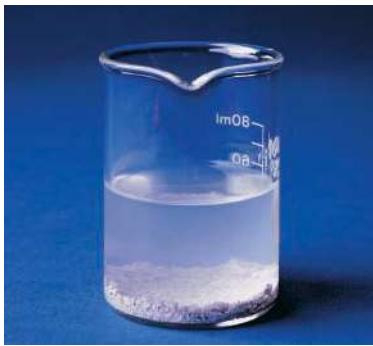
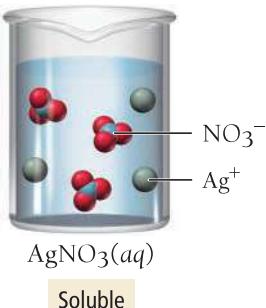
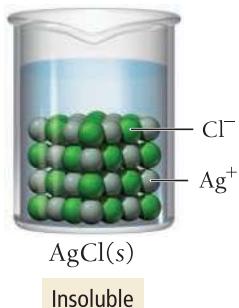
5.5
Cc
Conceptual Connection

ANSWER NOW!



The Solubility of Ionic Compounds

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



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

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

Whether a particular compound is soluble or insoluble depends on several factors. In Section 13.3, we will examine more closely the energy changes associated with solution formation. For now, we can follow a set of empirical rules that chemists have inferred from observations on many ionic compounds. Table 5.1 summarizes these *solubility rules*.

The solubility rules state that compounds containing the sodium ion are soluble. That means that compounds such as NaBr , NaNO_3 , Na_2SO_4 , NaOH , and Na_2CO_3 all dissolve in water to form strong electrolyte solutions. Similarly, the solubility rules state that compounds containing the NO_3^- ion are soluble. That means that compounds such as AgNO_3 , $\text{Pb}(\text{NO}_3)_2$, NaNO_3 , $\text{Ca}(\text{NO}_3)_2$, and $\text{Sr}(\text{NO}_3)_2$ all dissolve in water to form strong electrolyte solutions.

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

The solubility rules also state that, with some exceptions, compounds containing the CO_3^{2-} ion are insoluble. Therefore, compounds such as CuCO_3 , CaCO_3 , SrCO_3 , and FeCO_3 do not dissolve in water. Note that the solubility rules contain many exceptions. For example, compounds containing CO_3^{2-} are soluble when paired with Li^+ , Na^+ , K^+ , or NH_4^+ . Thus Li_2CO_3 , Na_2CO_3 , K_2CO_3 , and $(\text{NH}_4)_2\text{CO}_3$ are all soluble.

TABLE 5.1 ■ Solubility Rules for Ionic Compounds in Water

Compounds Containing the Following Ions Are Generally Soluble	Exceptions
Li^+ , Na^+ , K^+ , and NH_4^+	None
NO_3^- and $\text{C}_2\text{H}_3\text{O}_2^-$	None
Cl^- , Br^- , and I^-	When these ions pair with Ag^+ , Hg_2^{2+} , or Pb^{2+} , the resulting compounds are insoluble.
SO_4^{2-}	When SO_4^{2-} pairs with Sr^{2+} , Ba^{2+} , Pb^{2+} , Ag^+ , or Ca^{2+} , the resulting compound is insoluble.
Compounds Containing the Following Ions Are Generally Insoluble	Exceptions
OH^- and S^{2-}	When these ions pair with Li^+ , Na^+ , K^+ , or NH_4^+ , the resulting compounds are soluble.
	When S^{2-} pairs with Ca^{2+} , Sr^{2+} , or Ba^{2+} , the resulting compound is soluble.
	When OH^- pairs with Ca^{2+} , Sr^{2+} , or Ba^{2+} , the resulting compound is slightly soluble.
CO_3^{2-} and PO_4^{3-}	When these ions pair with Li^+ , Na^+ , K^+ , or NH_4^+ , the resulting compounds are soluble.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 5.5

EXAMPLE 5.5 Predicting Whether an Ionic Compound Is Soluble

Predict whether each compound is soluble or insoluble.

- (a) PbCl₂ (b) CuCl₂ (c) Ca(NO₃)₂ (d) BaSO₄

SOLUTION

- (a) Insoluble. Compounds containing Cl⁻ are normally soluble, but Pb²⁺ is an exception.
 (b) Soluble. Compounds containing Cl⁻ are normally soluble, and Cu²⁺ is not an exception.
 (c) Soluble. Compounds containing NO₃⁻ are always soluble.
 (d) Insoluble. Compounds containing SO₄²⁻ are normally soluble, but Ba²⁺ is an exception.

FOR PRACTICE 5.5 Predict whether each compound is soluble or insoluble.

- (a) NiS (b) Mg₃(PO₄)₂ (c) Li₂CO₃ (d) NH₄Cl

SOLUBILITY OF IONIC COMPOUNDS The presence of one of the following ions within a compound indicates that a compound is soluble with no exceptions. Which ion?

- (a) OH⁻ (b) SO₄²⁻ (c) NO₃⁻



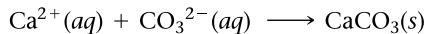
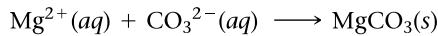
ANSWER NOW!

**5.5****Precipitation Reactions**

Have you ever taken a bath in hard water? Hard water contains dissolved ions such as Ca²⁺ and Mg²⁺ that diminish the effectiveness of soap. These ions react with soap to form a gray soap scum that may appear as a “bathtub ring” after you drain the tub. Hard water is particularly troublesome when washing clothes. Consider how your white shirt would look covered with the soap scum from the bathtub and you can understand the problem. Consequently, most laundry detergents include substances designed to remove Ca²⁺ and Mg²⁺ from the laundry mixture. The most common substance used for this purpose is sodium carbonate, which dissolves in water to form sodium cations (Na⁺) and carbonate (CO₃²⁻) anions:

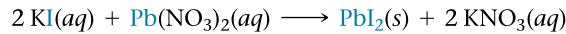


Sodium carbonate is soluble, but calcium carbonate and magnesium carbonate are not (see the solubility rules in Table 5.1). Consequently, the carbonate anions react with dissolved Mg²⁺ and Ca²⁺ ions in hard water to form solids that *precipitate* from (or come out of) solution:



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

The reactions between CO₃²⁻ and Mg²⁺ and Ca²⁺ are examples of **precipitation reactions**, reactions in which a solid or **precipitate** forms when we mix two solutions. Precipitation reactions are common in chemistry. As another example, consider potassium iodide and lead(II) nitrate, which form colorless, strong electrolyte solutions when dissolved in water separately. When the two solutions are combined, however, a brilliant yellow precipitate forms (Figure 5.13►). We describe this precipitation reaction with the following chemical equation:



WATCH NOW!

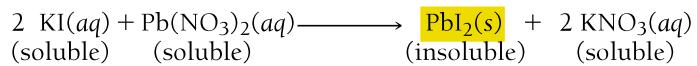
KEY CONCEPT VIDEO 5.5

Reactions in Solution



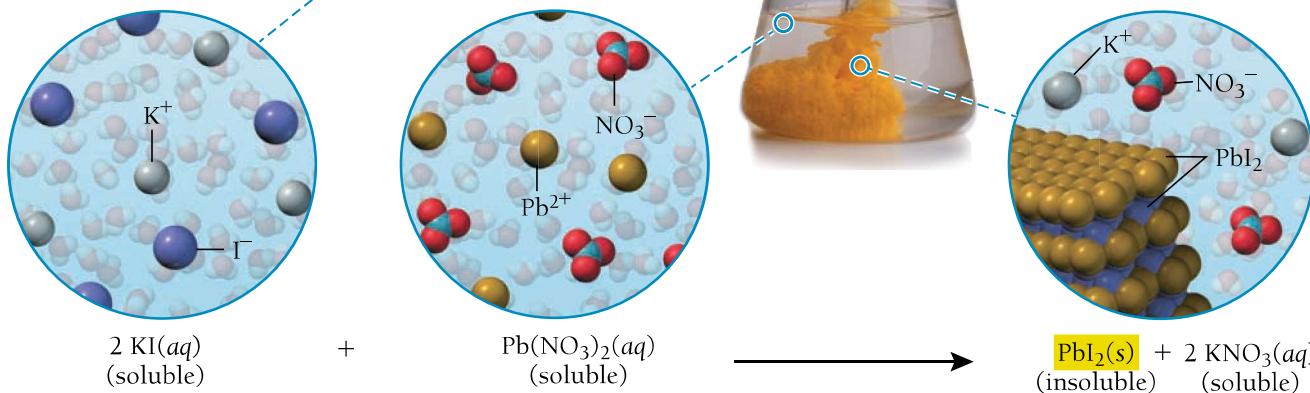
▲ The reaction of ions in hard water with soap produces a gray soap scum that is visible after you drain the bathwater.

Precipitation Reaction



◀ FIGURE 5.13 Precipitation of Lead(II) Iodide

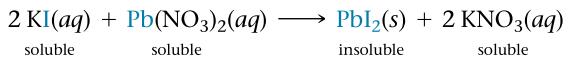
When a potassium iodide solution is mixed with a lead(II) nitrate solution, a yellow lead(II) iodide precipitate forms.



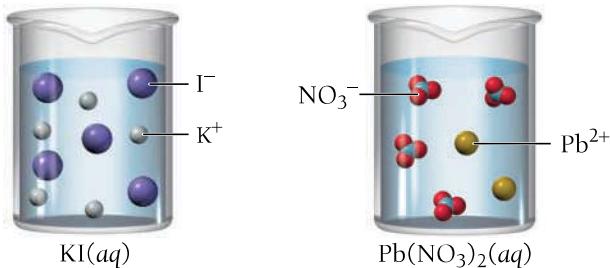
Precipitation reactions do not always occur when two aqueous solutions are mixed. For example, if we combine solutions of $\text{KI}(aq)$ and $\text{NaCl}(aq)$, nothing happens (Figure 5.14►):



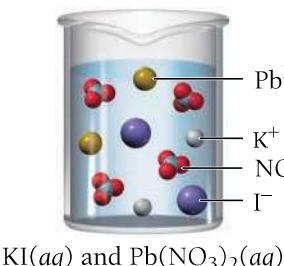
The key to predicting precipitation reactions is to understand that *only insoluble compounds form precipitates*. In a precipitation reaction, two solutions containing soluble compounds combine and an insoluble compound precipitates. Consider the precipitation reaction described previously:



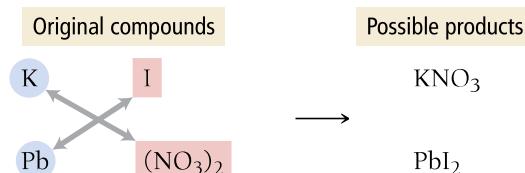
KI and Pb(NO₃)₂ are both soluble, but the precipitate, PbI₂, is insoluble. Before mixing, KI(*aq*) and Pb(NO₃)₂(*aq*) are both dissociated in their respective solutions:



The instant that the solutions come into contact, all four ions are present:



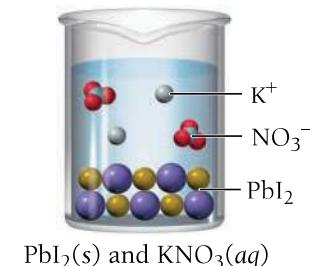
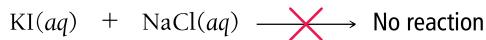
Now, new compounds—one or both of which might be insoluble—are possible. Specifically, the cation from either compound can pair with the anion from the other to form possibly insoluble products:



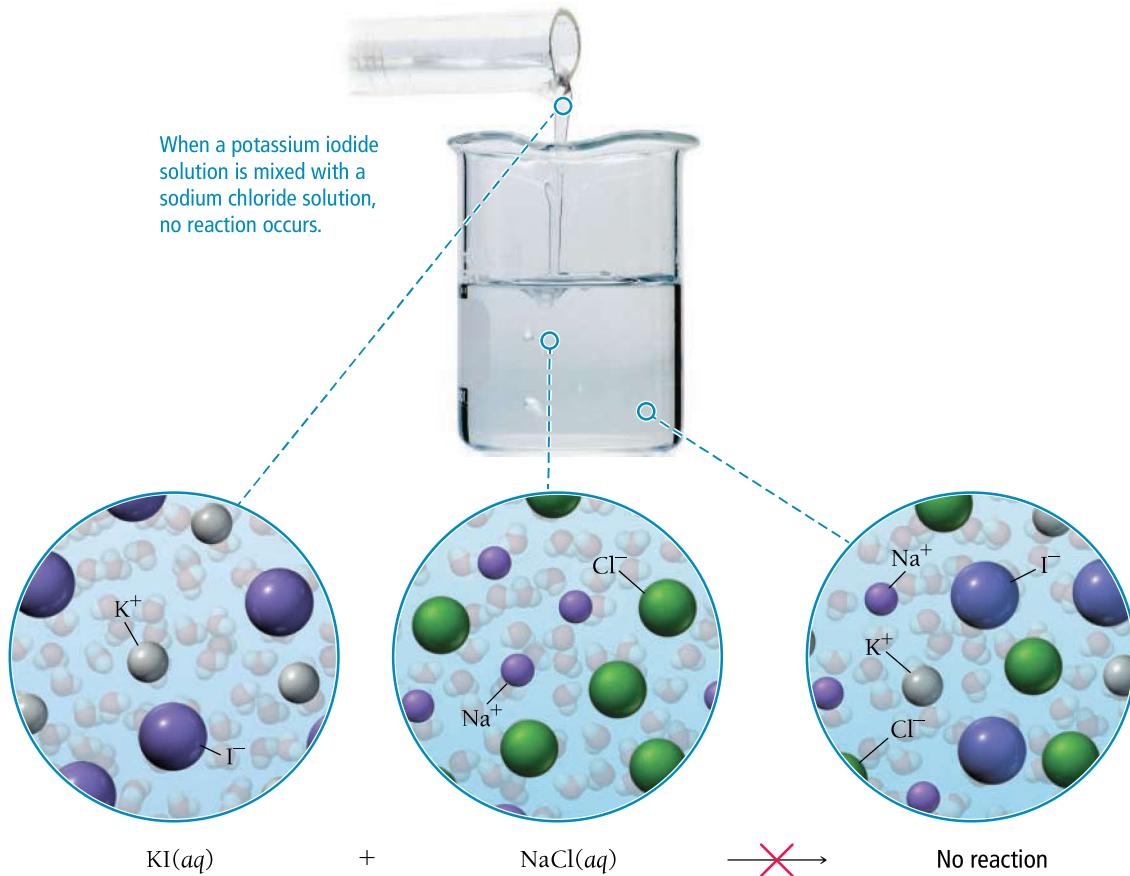
If the possible products are both soluble, no reaction occurs and no precipitate forms. If one or both of the possible products are insoluble, a precipitation reaction occurs. In this case, KNO_3 is soluble, but PbI_2 is insoluble. Consequently, PbI_2 precipitates.

To predict whether a precipitation reaction will occur when two solutions are mixed and to write an equation for the reaction, we use the procedure that follows. The steps are outlined in the left column, and two examples illustrating how to apply the procedure are shown in the center and right columns.

No Reaction



◀ FIGURE 5.14 No Precipitation



WATCH NOW!


**INTERACTIVE WORKED
EXAMPLE VIDEO 5.6**
HOW TO: Write Equations for Precipitation Reactions
EXAMPLE 5.6
Writing Equations for Precipitation Reactions

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

1. Write the formulas of the two compounds being mixed as reactants in a chemical equation.
2. Below the equation, write the formulas of the products that could form from the reactants. Obtain these by combining the cation from each reactant with the anion from the other. Make sure to write correct formulas for these ionic compounds using the procedure demonstrated in Section 3.5.



Possible products



3. Refer to the solubility rules to determine whether any of the possible products are insoluble.

KCl is soluble. (Compounds containing Cl^- are usually soluble, and K^+ is not an exception.)

NiCO_3 is insoluble. (Compounds containing CO_3^{2-} are usually insoluble, and Ni^{2+} is not an exception.)

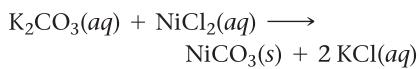
4. If all of the possible products are soluble, there is no precipitate. Write "NO REACTION" after the arrow.

Since this example has an insoluble product, we proceed to the next step.

5. If any of the possible products are insoluble, write their formulas as the products of the reaction, using (s) to indicate solid. Include an (aq) to indicate aqueous after any soluble products.



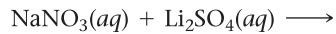
6. Balance the equation. Remember to adjust only coefficients, not subscripts.



FOR PRACTICE 5.6 Write an equation for the precipitation reaction that occurs (if any) when solutions of ammonium chloride and iron(III) nitrate mix.

EXAMPLE 5.7
Writing Equations for Precipitation Reactions

Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium nitrate and lithium sulfate are mixed.



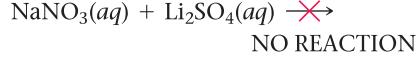
Possible products



LiNO_3 is soluble. (Compounds containing NO_3^- are soluble, and Li^+ is not an exception.)

Na_2SO_4 is soluble. (Compounds containing SO_4^{2-} are generally soluble, and Na^+ is not an exception.)

Since this example has no insoluble product, there is no reaction.



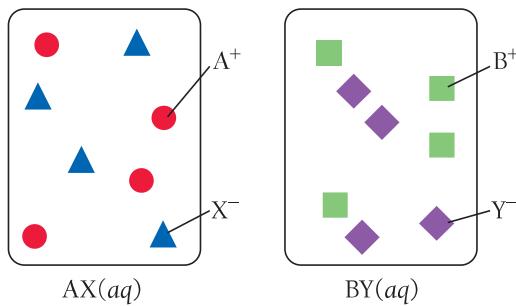
FOR PRACTICE 5.7 Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium hydroxide and copper(II) bromide mix.

PRECIPITATION REACTIONS

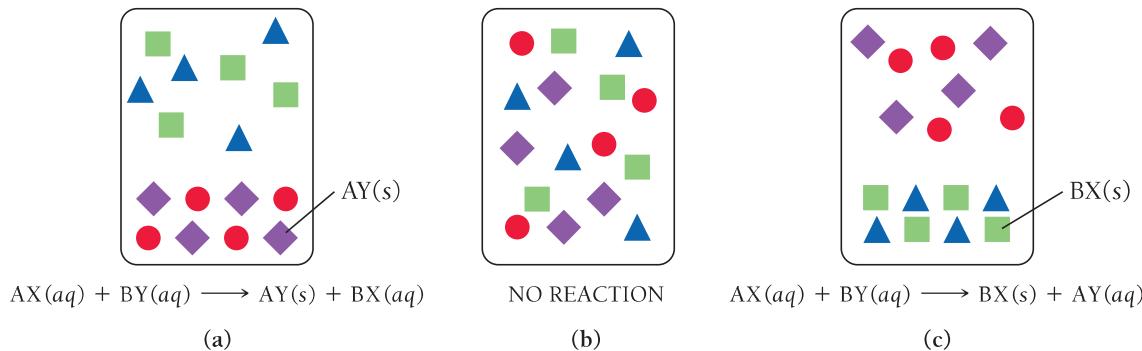
Consider the generic ionic compounds with the formulas AX and BY and the following solubility rules:

AX soluble; BY soluble; AY soluble; BX insoluble

Let circles represent A^+ ions; squares represent B^+ ions; triangles represent X^- ions; and diamonds represent Y^- ions. We represent solutions of the two compounds (AX and BY) as follows:



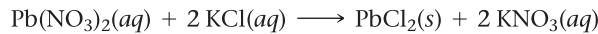
From the answer options, choose the representation that correctly shows the result of mixing the two solutions (AX and BY) and the correct equation to represent the reaction.



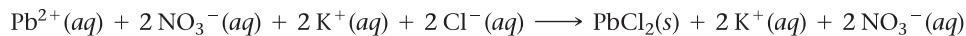
5.6

Representing Aqueous Reactions: Molecular, Ionic, and Net Ionic Equations

Consider the following equation for a precipitation reaction:



This equation is a **molecular equation**, an equation showing the complete neutral formulas for each compound in the reaction as if they existed as molecules. In actual solutions of soluble ionic compounds, dissolved substances are present as ions. We can write equations for reactions occurring in aqueous solution in a way that better shows the dissociated nature of dissolved ionic compounds. For example, we can rewrite the previous equation as:



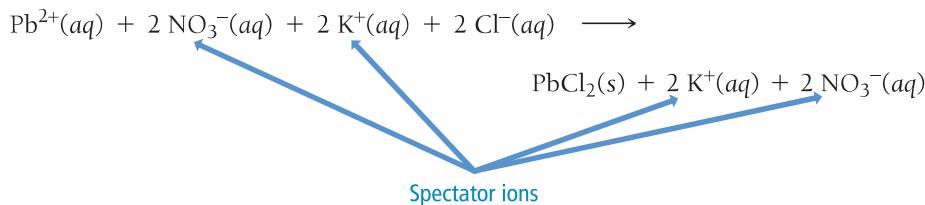
Equations such as this, which list all of the ions present as either reactants or products in a chemical reaction, are **complete ionic equations**. Strong electrolytes are always represented as their component ions in ionic equations—weak electrolytes are not.



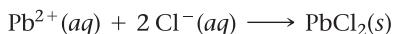
ANSWER NOW!



Notice that in the complete ionic equation, some of the ions in solution appear unchanged on both sides of the equation. These ions are called **spectator ions** because they do not participate in the reaction.



To simplify the equation and to show more clearly what is happening, we can omit spectator ions:

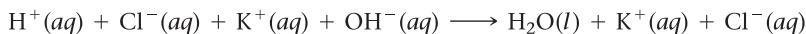


Equations that show only the species that actually change during the reaction are **net ionic equations**.

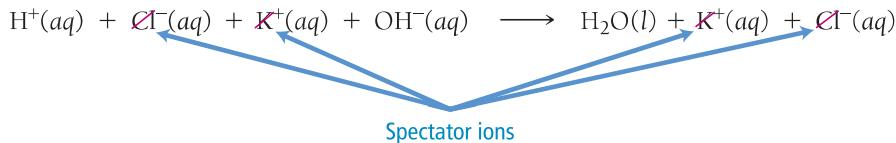
As another example, consider the reaction between $\text{HCl}(aq)$ and $\text{KOH}(aq)$:



Since HCl , KOH , and KCl all exist in solution primarily as independent ions, the complete ionic equation is:



To write the net ionic equation, we remove the spectator ions, those that are unchanged on both sides of the equation:



The net ionic equation is $\text{H}^+(aq) + \text{OH}^-(aq) \longrightarrow \text{H}_2\text{O}(l)$.

Summarizing Aqueous Equations

- A **molecular equation** is a chemical equation showing the complete, neutral formulas for every compound in a reaction.
- A **complete ionic equation** is a chemical equation showing all of the species as they are actually present in solution: strong electrolytes are therefore represented as their component ions.
- A **net ionic equation** is an equation showing only the species that actually change during the reaction.

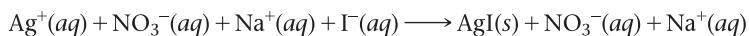
ANSWER NOW!



5.8

Conceptual Connection

SPECTATOR IONS Which of the ions listed below is a spectator ion in the complete ionic equation shown here?



- (a) $\text{Ag}^+(aq)$ (b) $\text{NO}_3^-(aq)$ (c) $\text{I}^-(aq)$

EXAMPLE 5.8 Writing Complete Ionic and Net Ionic Equations

Write complete ionic and net ionic equations for each reaction.

- (a) $3 \text{SrCl}_2(aq) + 2 \text{Li}_3\text{PO}_4(aq) \longrightarrow \text{Sr}_3(\text{PO}_4)_2(s) + 6 \text{LiCl}(aq)$
 (b) $\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{KOH}(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{KC}_2\text{H}_3\text{O}_2(aq)$

SOLUTION

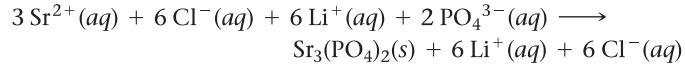
(a) Write the complete ionic equation by separating strong electrolytes into their constituent ions. The $\text{Sr}_3(\text{PO}_4)_2(s)$, precipitating as a solid, remains as one unit.

Write the net ionic equation by eliminating the spectator ions, those that do not change from one side of the reaction to the other.

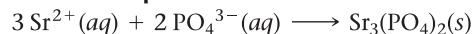
(b) Write the complete ionic equation by separating strong electrolytes into their constituent ions. Do not separate $\text{HC}_2\text{H}_3\text{O}_2(aq)$ because it is a weak electrolyte.

Write the net ionic equation by eliminating the spectator ions.

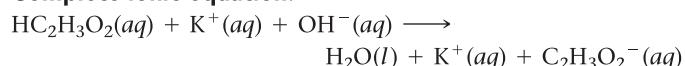
Complete ionic equation:



Net ionic equation:



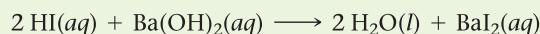
Complete ionic equation:



Net ionic equation:



FOR PRACTICE 5.8 Write the complete ionic equation and net ionic equation for the following reaction:



5.7

Acid–Base Reactions

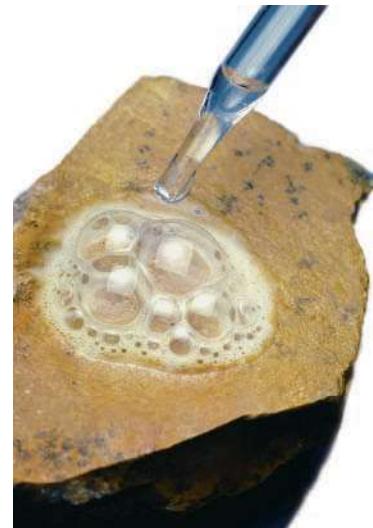
Two other important classes of reactions that occur in aqueous solution are acid–base reactions and gas-evolution reactions. In an **acid–base reaction** (also called a **neutralization reaction**), an acid reacts with a base and the two neutralize each other, producing water (or in some cases a weak electrolyte). In a **gas-evolution reaction**, a gas forms, resulting in bubbling. In both cases, as in precipitation reactions, the reactions occur when the anion from one reactant combines with the cation of the other. Many gas-evolution reactions are also acid–base reactions.

Acid–Base Reactions

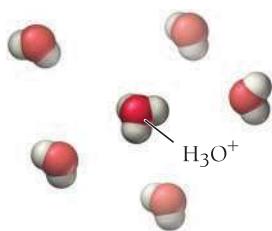
Our stomachs contain hydrochloric acid (HCl), which acts in the digestion of food. Certain foods or stress, however, can increase the stomach's acidity to uncomfortable levels, causing acid stomach or heartburn. Antacids are over-the-counter medicines that work by reacting with and neutralizing stomach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

Recall from Chapter 3 that an acid forms H^+ ions in solution, and we just saw that a base is a substance that produces OH^- ions in solution. More formally:

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



▲ Gas-evolution reactions, such as the reaction of hydrochloric acid (HCl) with limestone (CaCO_3), typically produce CO_2 ; bubbling occurs as the gas is released.



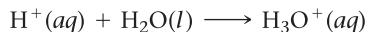
▲ FIGURE 5.15 The Hydronium Ion Protons normally associate with water molecules in solution to form H_3O^+ ions, which in turn interact with other water molecules.

These definitions of acids and bases, called the **Arrhenius definitions**, are named after Swedish chemist Svante Arrhenius (1859–1927). In Chapter 16, we will learn more general definitions of acids and bases, but these definitions are sufficient to describe neutralization reactions.

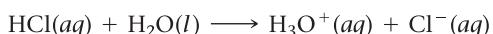
According to the Arrhenius definition, HCl is an acid because it produces H^+ ions in solution:



An H^+ ion is a bare proton. In solution, bare protons normally associate with water molecules to form **hydronium ions** (Figure 5.15◀):



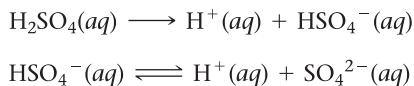
Chemists use $\text{H}^+(aq)$ and $\text{H}_3\text{O}^+(aq)$ interchangeably to mean the same thing—a hydronium ion. The chemical equation for the ionization of HCl and other acids is often written to show the association of the proton with a water molecule to form the hydronium ion:



As we discussed in Section 5.4, some acids are weak acids—they do not completely ionize in solution. We represent the partial ionization of a weak acid with opposing half arrows.



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



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



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

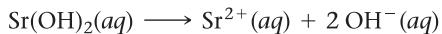
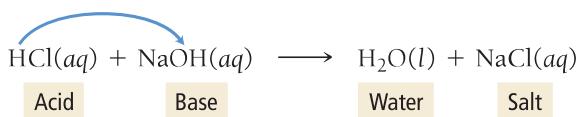
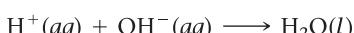


Table 5.2 lists common acids and bases. You can find acids and bases in many everyday substances. Foods such as citrus fruits and vinegar contain acids. Soap, baking soda, and milk of magnesia all contain bases.

When we mix an acid and a base, the $\text{H}^+(aq)$ from the acid—whether it is weak or strong—combines with the $\text{OH}^-(aq)$ from the base to form $\text{H}_2\text{O}(l)$ (Figure 5.16▶). Consider the reaction between hydrochloric acid and sodium hydroxide:



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



▲ These household substances all contain acids.

TABLE 5.2 ■ Some Common Acids and Bases

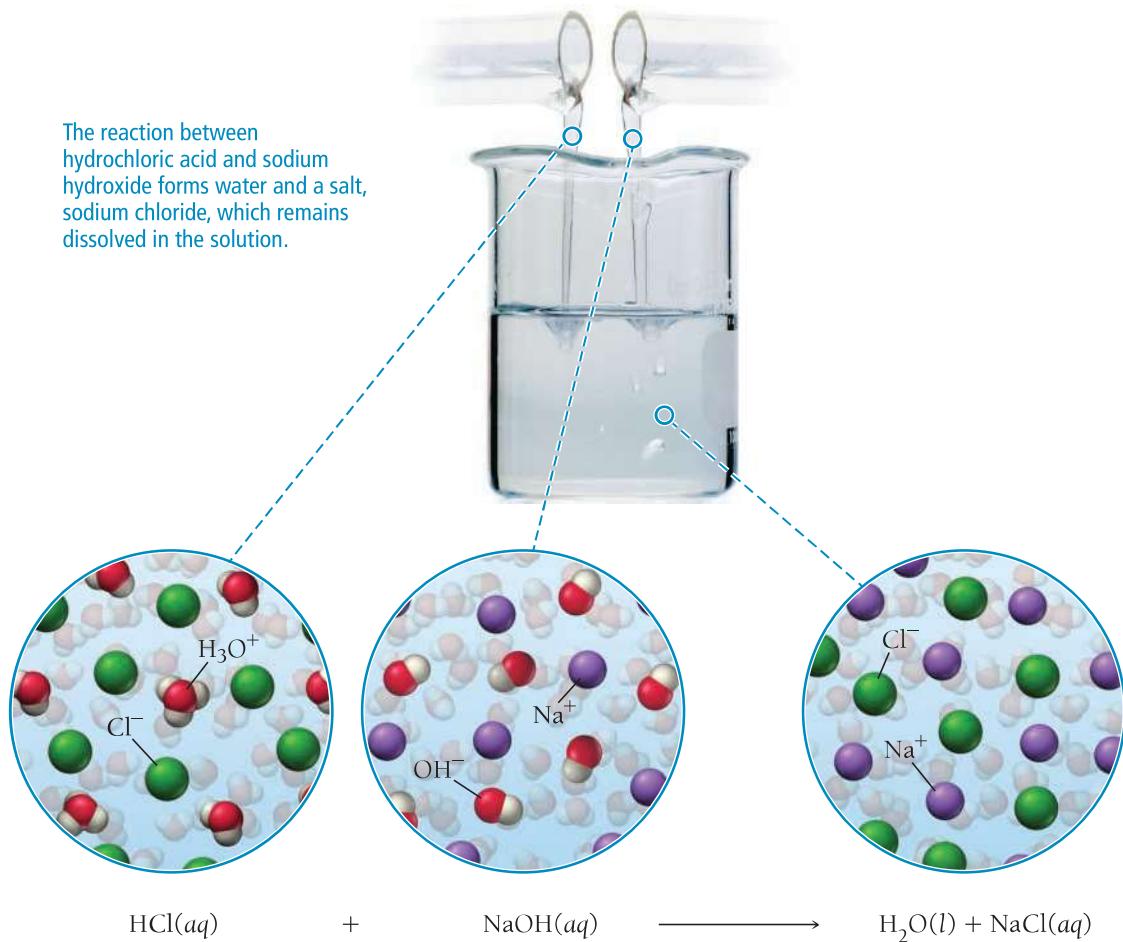
Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	NaOH
Hydrobromic acid	HBr	Lithium hydroxide	LiOH
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO ₃	Calcium hydroxide	Ca(OH) ₂
Sulfuric acid	H ₂ SO ₄	Barium hydroxide	Ba(OH) ₂
Perchloric acid	HClO ₄	Ammonia*	NH ₃ (weak base)
Formic acid	HCHO ₂ (weak acid)		
Acetic acid	HC ₂ H ₃ O ₂ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

*Ammonia does not contain OH⁻, but it produces OH⁻ in a reaction with water that occurs only to a small extent: NH₃(aq) + H₂O(l) ⇌ NH₄⁺(aq) + OH⁻(aq).



▲ Many common household products contain bases.

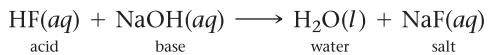
Acid–Base Reaction



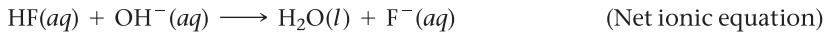
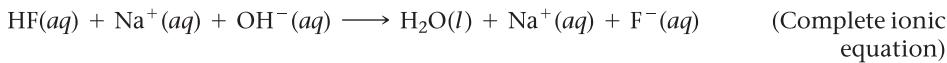
▲ **FIGURE 5.16** Acid–Base Reaction

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

However, if the acid is a weak acid, the net ionic equation is slightly different. For example, consider the acid–base equation between hydrofluoric acid and sodium hydroxide:

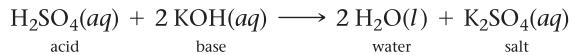


The complete ionic equation and the net ionic equation for this reaction are:



Notice that, since HF is a weak acid, we do not show it as ionized in the ionic equations.

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



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



When writing equations for acid–base reactions, write the formula of the salt using the procedure for writing formulas of ionic compounds demonstrated in Section 3.5.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 5.9

EXAMPLE 5.9 Writing Equations for Acid–Base Reactions Involving a Strong Acid

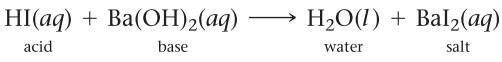
Write a molecular and net ionic equation for the reaction between aqueous HI and aqueous Ba(OH)2.



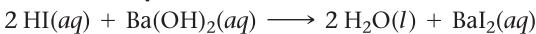
SOLUTION First identify these substances as an acid and a base. Begin by writing the unbalanced equation in which the acid and the base combine to form water and a salt.

Next, balance the equation; this is the molecular equation.

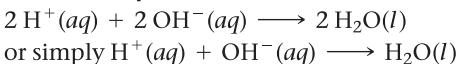
Write the net ionic equation by removing the spectator ions.



Molecular equation:



Net ionic equation:



FOR PRACTICE 5.9 Write a molecular and a net ionic equation for the reaction that occurs between aqueous HBr and aqueous LiOH.

EXAMPLE 5.10 Writing Equations for Acid–Base Reactions Involving a Weak Acid

Write a molecular equation, ionic equation, and net ionic equation for the reaction between aqueous acetic acid (HC2H3O2) and aqueous potassium hydroxide (KOH).

SOLUTION Begin by writing the molecular equation in which the acid and the base combine to form water and a salt. (The equation is already balanced.)

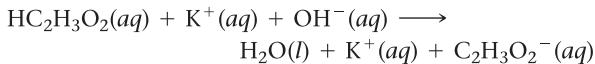
Write the complete ionic equation by separating aqueous ionic compounds into their constituent ions. Do not separate HC2H3O2(aq) because it is a weak acid (and a weak electrolyte).

Write the net ionic equation by eliminating the spectator ions.

Molecular equation:



Complete ionic equation:



Net ionic equation:



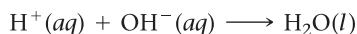
FOR PRACTICE 5.10 Write the net ionic equation for the reaction between HCHO2 (a weak acid) and NaOH.

Acid–Base Titrations

We can apply the principles of acid–base neutralization and stoichiometry to a common laboratory procedure called a *titration*. In a **titration**, a substance in a solution of known concentration is reacted with another substance in a solution of unknown concentration. For example, consider the following acid–base reaction:



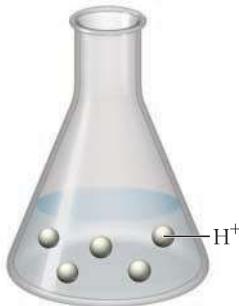
The net ionic equation for this reaction eliminates the spectator ions:



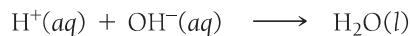
Suppose we have an HCl solution represented by the molecular diagram shown here (for purposes of clarity, we have omitted the Cl^- ions and the H_2O molecules not involved in the reaction from this representation).

In titrating this sample, we slowly add a solution of known OH^- concentration, as shown in the molecular diagrams in Figure 5.17▼. As we add the OH^- , it reacts with and neutralizes the H^+ , forming water. At the **equivalence point**—the point in the titration when the number of moles of OH^- added equals the number of moles of H^+ initially in solution—the titration is complete. The equivalence point is typically signaled by an **indicator**, a dye whose color depends on the acidity or basicity of the solution (Figure 5.18►).

We cover acid–base titrations and indicators in more detail in Chapter 18. In most laboratory titrations, the concentration of one of the reactant solutions is unknown, and the concentration of the other is precisely known. By carefully measuring the volume of each solution required to reach the equivalence point, we can determine the concentration of the unknown solution, as demonstrated in Example 5.11.



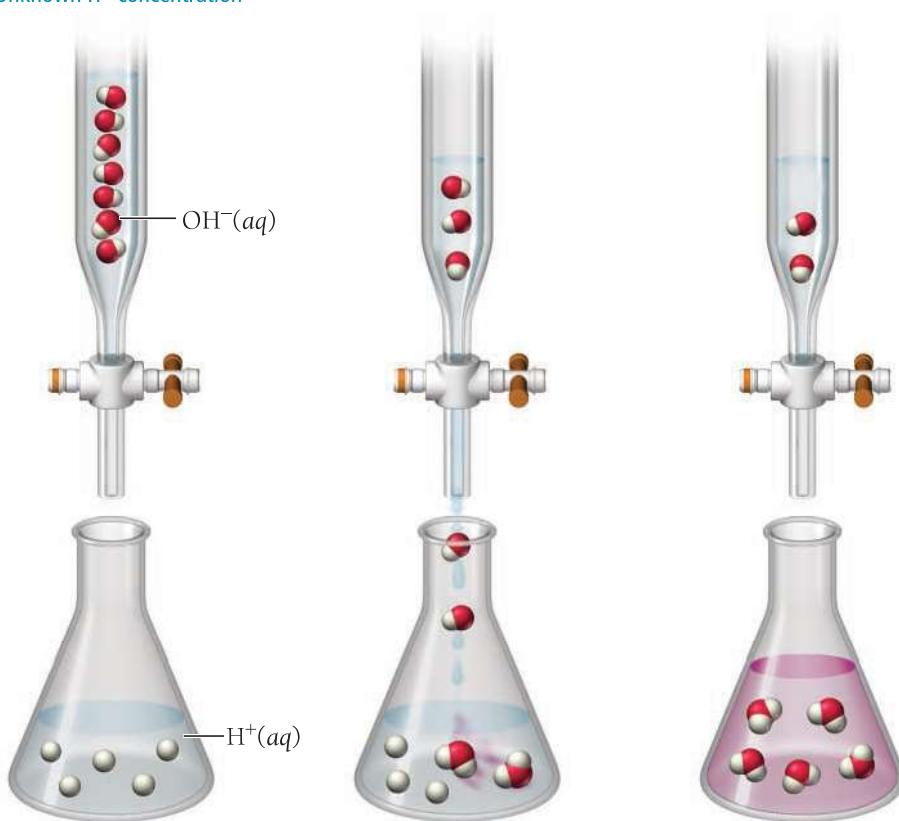
Acid–Base Titration



Beginning of titration:
Known OH^- concentration
Unknown H^+ concentration

OH^- solution is slowly added to acid solution

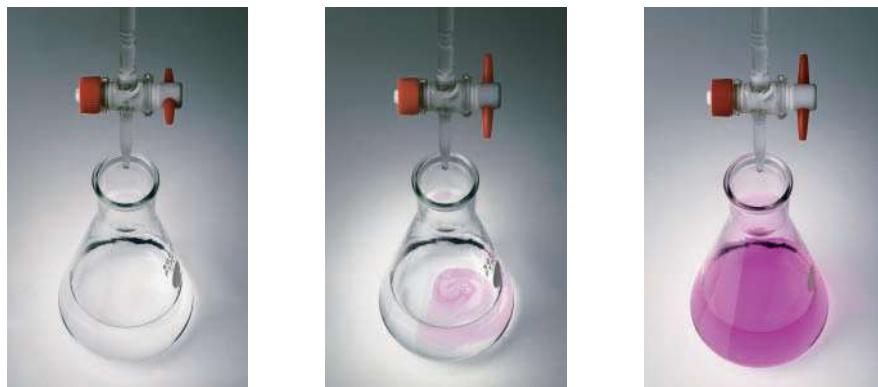
Equivalence point:
Moles of OH^- = moles of H^+



◀ FIGURE 5.17 Acid–Base Titration

► FIGURE 5.18 Titration In this titration, NaOH is added to a dilute HCl solution. When the NaOH and HCl reach stoichiometric proportions (the equivalence point), the phenolphthalein indicator changes color to pink.

Indicator in Titration



ANSWER NOW!



5.9 Cc Conceptual Connection

ACID-BASE TITRATION A 10.0 mL sample of 0.20 M HBr solution is titrated with 0.10 M NaOH. What volume of NaOH is required to reach the equivalence point?

- (a) 10.0 mL (b) 20.0 mL (c) 40.0 mL

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 5.11

EXAMPLE 5.11 Acid-Base Titration

The titration of 10.00 mL of HCl solution of unknown concentration requires 12.54 mL of a 0.100 M NaOH solution to reach the equivalence point. What is the concentration of the unknown HCl solution in M?

SORT You are given the volume and concentration of NaOH solution required to titrate a given volume of HCl solution. You are asked to find the concentration of the HCl solution.

STRATEGIZE Since this problem involves an acid-base neutralization reaction between HCl and NaOH, start by writing the balanced equation, using the techniques covered earlier in this section.

The first part of the conceptual plan has the form volume A \longrightarrow moles A \longrightarrow moles B. The concentration of the NaOH solution is a conversion factor between moles and volume of NaOH. The balanced equation provides the relationship between number of moles of NaOH and number of moles of HCl.

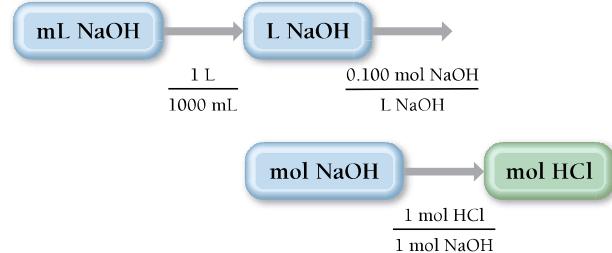
In the second part of the conceptual plan, use the number of moles of HCl (from the first part) and the volume of HCl solution (given) to calculate the molarity of the HCl solution.

GIVEN: 12.54 mL of NaOH solution, 0.100 M NaOH solution, 10.00 mL of HCl solution

FIND: concentration of HCl solution



CONCEPTUAL PLAN



$$\text{mol HCl, L HCl solution} \longrightarrow \text{Molarity}$$

$$M = \frac{\text{mol}}{\text{L}}$$

RELATIONSHIPS USED

$$1 \text{ L} = 1000 \text{ mL}$$

$$M(\text{NaOH}) = \frac{0.100 \text{ mol NaOH}}{\text{L NaOH}}$$

$$1 \text{ mol HCl} : 1 \text{ mol NaOH}$$

$$\text{Molarity (M)} = \frac{\text{moles of solute (mol)}}{\text{volume of solution (L)}}$$

SOLVE In the first part of the solution, determine the number of moles of HCl in the unknown solution.

In the second part of the solution, divide the number of moles of HCl by the volume of the HCl solution in L. 10.00 mL is equivalent to 0.01000 L.

SOLUTION

$$12.54 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.100 \text{ mol NaOH}}{1 \text{ L NaOH}} \\ \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} = 1.25 \times 10^{-3} \text{ mol HCl}$$

$$\text{Molarity} = \frac{1.25 \times 10^{-3} \text{ mol HCl}}{0.01000 \text{ L}} = 0.125 \text{ M HCl}$$

CHECK The units of the answer (M HCl) are correct. The magnitude of the answer (0.125 M) is reasonable because it is similar to the molarity of the NaOH solution, as expected from the reaction stoichiometry (1 mol HCl reacts with 1 mol NaOH) and the similar volumes of NaOH and HCl.

FOR PRACTICE 5.11 The titration of a 20.0-mL sample of an H_2SO_4 solution of unknown concentration requires 22.87 mL of a 0.158 M KOH solution to reach the equivalence point. What is the concentration of the unknown H_2SO_4 solution?

FOR MORE PRACTICE 5.11 What volume (in mL) of 0.200 M NaOH do we need to titrate 35.00 mL of 0.140 M HBr to the equivalence point?

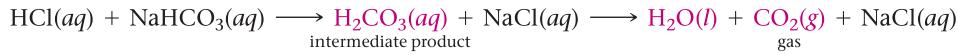
5.8

Gas-Evolution Reactions

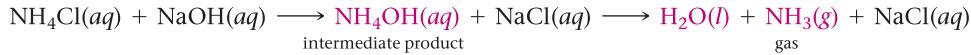
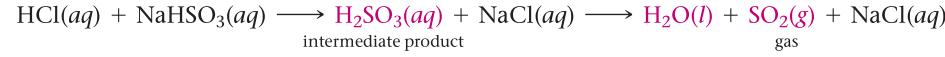
In a *gas-evolution reaction*, two aqueous solutions mix to form a gaseous product that bubbles out of solution. Some gas-evolution reactions form a gaseous product directly when the cation of one reactant combines with the anion of the other. For example, when sulfuric acid reacts with lithium sulfide, dihydrogen sulfide gas forms:



Other gas-evolution reactions often form an intermediate product that then decomposes (breaks down into simpler substances) to form a gas. For example, when aqueous hydrochloric acid is mixed with aqueous sodium bicarbonate, the following reaction occurs (Figure 5.19►):



The intermediate product, H_2CO_3 , is not stable and decomposes into H_2O and gaseous CO_2 . Other important gas-evolution reactions form either H_2SO_3 or NH_4OH as intermediate products:



Many gas-evolution reactions such as this one are also acid–base reactions. In Chapter 17 we will learn how ions such as CO_3^{2-} act as bases in aqueous solution.

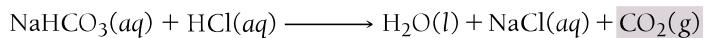
The intermediate product NH_4OH provides a convenient way to think about this reaction, but the extent to which it actually forms is debatable.

Table 5.3 lists the main types of compounds that form gases in aqueous reactions, as well as the gases formed.

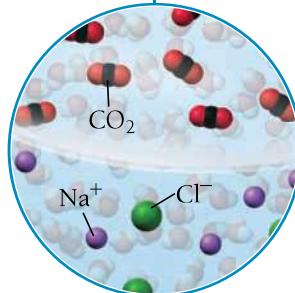
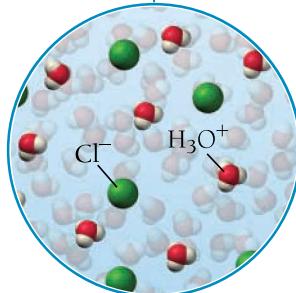
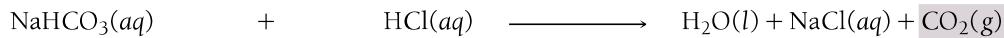
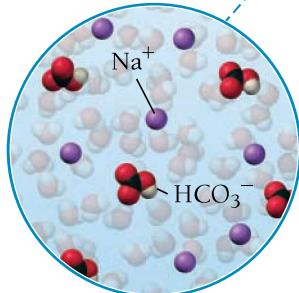
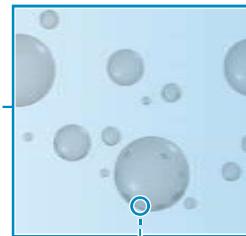
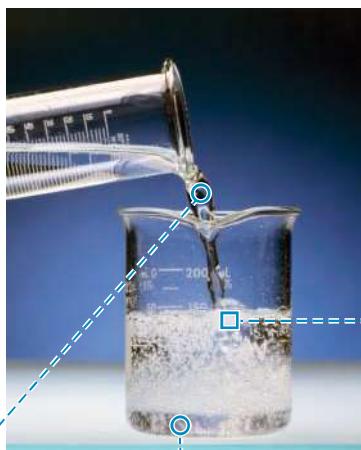
TABLE 5.3 ■ Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H_2S	$2 \text{ HCl}(aq) + \text{K}_2\text{S}(aq) \longrightarrow \text{H}_2\text{S}(g) + 2 \text{ KCl}(aq)$
Carbonates and bicarbonates	H_2CO_3	CO_2	$2 \text{ HCl}(aq) + \text{K}_2\text{CO}_3(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + 2 \text{ KCl}(aq)$
Sulfites and bisulfites	H_2SO_3	SO_2	$2 \text{ HCl}(aq) + \text{K}_2\text{SO}_3(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{SO}_2(g) + 2 \text{ KCl}(aq)$
Ammonium	NH_4OH	NH_3	$\text{NH}_4\text{Cl}(aq) + \text{KOH}(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{NH}_3(g) + \text{KCl}(aq)$

Gas-Evolution Reaction



When aqueous sodium bicarbonate is mixed with aqueous hydrochloric acid, gaseous CO_2 bubbles are the result of the reaction.

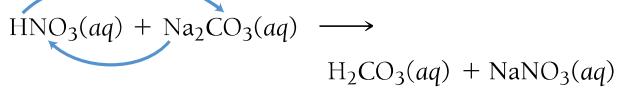


▲ FIGURE 5.19 Gas-Evolution Reaction

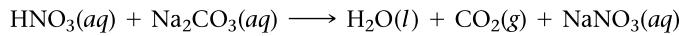
EXAMPLE 5.12 Writing Equations for Gas-Evolution Reactions

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

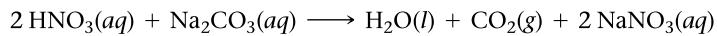
Begin by writing an unbalanced equation in which the cation of each reactant combines with the anion of the other.



You must then recognize that $\text{H}_2\text{CO}_3(aq)$ decomposes into $\text{H}_2\text{O}(l)$ and $\text{CO}_2(g)$ and write these products into the equation.



Finally, balance the equation.



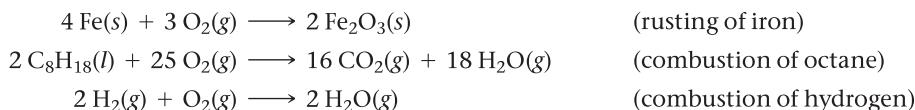
FOR PRACTICE 5.12 Write a molecular equation for the gas-evolution reaction that occurs when you mix aqueous hydrobromic acid and aqueous potassium sulfite.

FOR MORE PRACTICE 5.12 Write a net ionic equation for the reaction that occurs when you mix hydroiodic acid with calcium sulfide.

5.9

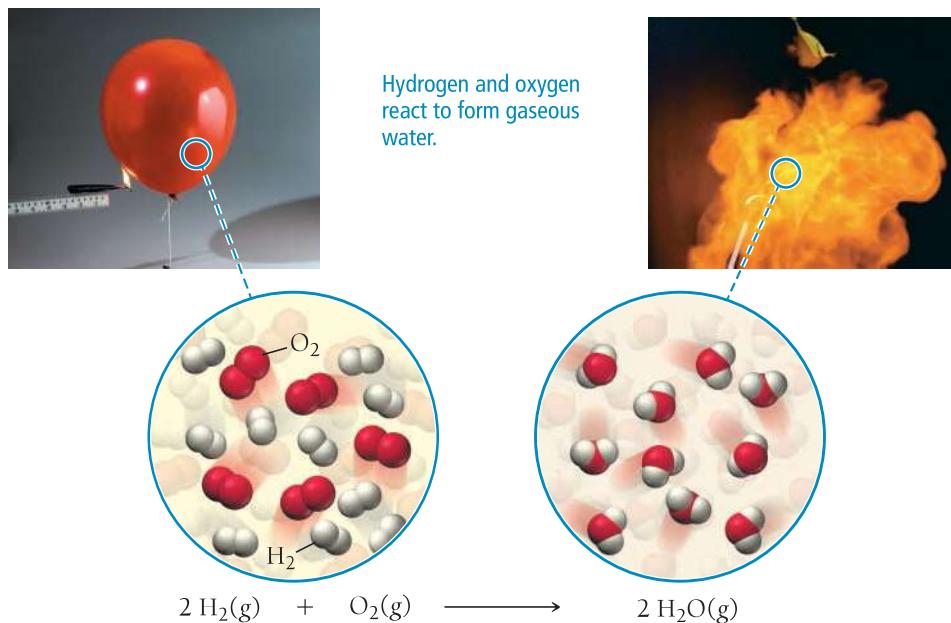
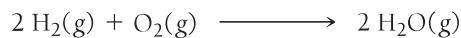
Oxidation–Reduction Reactions

Oxidation–reduction reactions or **redox reactions** are reactions in which electrons transfer from one reactant to the other. The rusting of iron, the bleaching of hair, and the production of electricity in batteries involve redox reactions. Many redox reactions involve the reaction of a substance with oxygen (Figure 5.20▼).



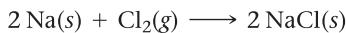
Oxidation–reduction reactions are covered in more detail in Chapter 20.

Oxidation–Reduction Reaction

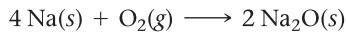


▲ FIGURE 5.20 Oxidation–Reduction Reaction

However, redox reactions need not involve oxygen. Consider, for example, the reaction between sodium and chlorine to form sodium chloride (NaCl), depicted in Figure 5.21►:



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

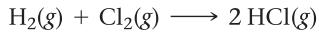


The reaction between sodium and oxygen forms other oxides as well.

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

Helpful Mnemonic: O I L R I G
Oxidation **I**s **L**oss; **R**eduction **I**s **G**ain.

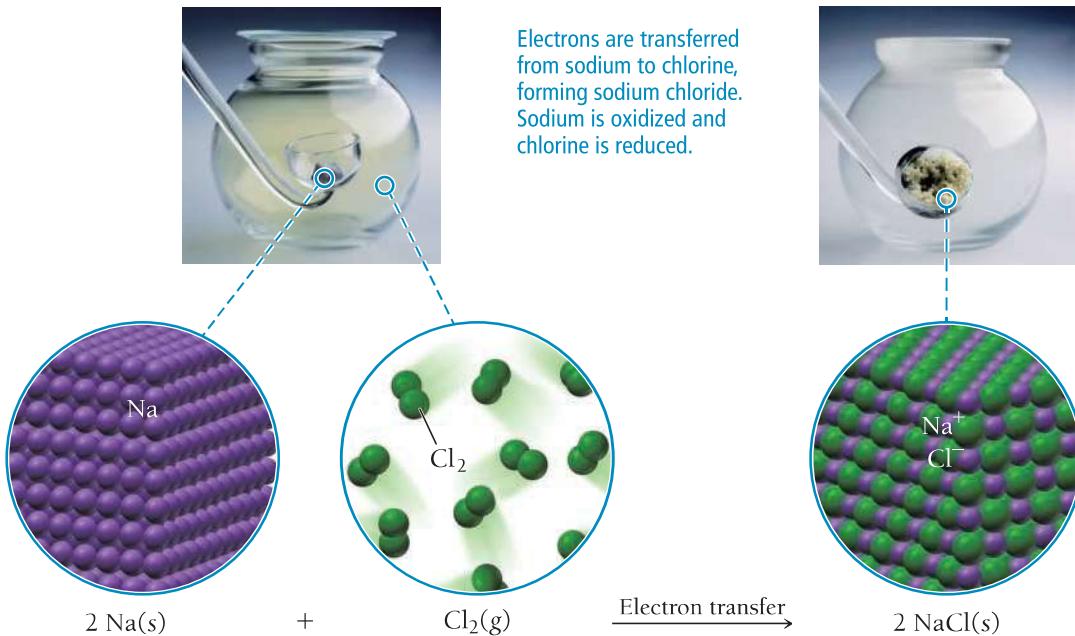
The transfer of electrons does not need to be a *complete* transfer (as occurs in the formation of an ionic compound) for the reaction to qualify as oxidation–reduction. For example, consider the reaction between hydrogen gas and chlorine gas:



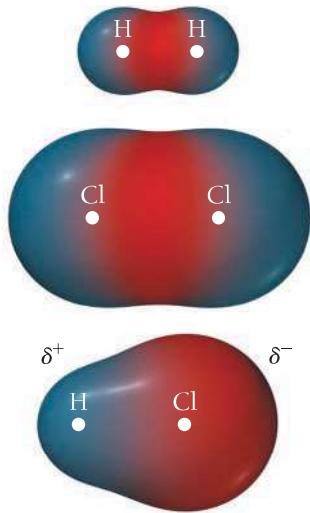
The ability of an element to attract electrons in a chemical bond is called electronegativity. We cover electronegativity in more detail in Section 10.6.

Even though hydrogen chloride is a molecular compound with a covalent bond, and even though the hydrogen has not completely transferred its electron to chlorine

Oxidation–Reduction Reaction without Oxygen



▲ FIGURE 5.21 Oxidation–Reduction without Oxygen



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

▲ FIGURE 5.22 Redox with Partial Electron Transfer When hydrogen bonds to chlorine, the electrons are unevenly shared, resulting in an increase of electron density (reduction) for chlorine and a decrease in electron density (oxidation) for hydrogen.

during the reaction, we can see from the electron density diagrams (Figure 5.22◀) that hydrogen has lost some of its electron density—it has *partially* transferred its electron to chlorine. In the reaction, hydrogen is oxidized and chlorine is reduced and, therefore, this is a redox reaction.

Oxidation States

Identifying whether or not a reaction between a metal and a nonmetal is a redox reaction is fairly straightforward because of ion formation. But how do we identify redox reactions that occur between nonmetals? Chemists have devised a scheme to track electrons before and after a chemical reaction. In this scheme—which is like bookkeeping for electrons—each shared electron is assigned to the atom that attracts the electrons most strongly. Then a number, called the **oxidation state** or **oxidation number**, is given to each atom based on the electron assignments. In other words, the oxidation number of an atom in a compound is the “charge” it would have if all shared electrons were assigned to the atom with the greatest attraction for those electrons.

For example, consider HCl. Since chlorine attracts electrons more strongly than hydrogen, we assign the two shared electrons in the bond to chlorine; then H (which has lost an electron in our assignment) has an oxidation state of +1, and Cl (which has gained one electron in our assignment) has an oxidation state of -1. Notice that in contrast to ionic charges, which are usually written with the sign of the charge *after* the magnitude (1+ and 1-, for example), oxidation states are written with the sign of the charge *before* the magnitude (+1 and -1, for example). We use the following rules to assign oxidation states to atoms in elements, ions, and compounds.

Rules for Assigning Oxidation States

(These rules are hierarchical. If any two rules conflict, follow the rule that is higher on the list.)

- The oxidation state of an atom in a free element is 0.
- The oxidation state of a monoatomic ion is equal to its charge.
- The sum of the oxidation states of all atoms in:
 - A neutral molecule or formula unit is 0.
 - An ion is equal to the charge of the ion.

- In their compounds, metals have positive oxidation states.

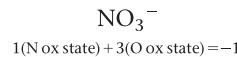
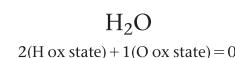
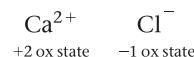
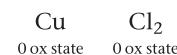
- Group 1A metals *always* have an oxidation state of +1.
- Group 2A metals *always* have an oxidation state of +2.

- In their compounds, nonmetals are assigned oxidation states according to the table shown here. Entries at the top of the table take precedence over entries at the bottom of the table.

When assigning oxidation states, keep these points in mind:

- The oxidation state of any given element generally depends on what other elements are present in the compound. (The exceptions are the group 1A and 2A metals, which are *always* +1 and +2, respectively.)
- Rule 3 must always be followed. Therefore, when following the hierarchy shown in rule 5, give priority to the element(s) highest on the list and then assign the oxidation state of the element lowest on the list using rule 3.
- When assigning oxidation states to elements that are not covered by rules 4 and 5 (such as carbon), use rule 3 to deduce their oxidation state once all other oxidation states have been assigned.

Examples



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

Nonmetal	Oxidation State	Example
Fluorine	-1	MgF_2 -1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	-2	CO_2 -2 ox state
Group 7A	-1	CCl_4 -1 ox state
Group 6A	-2	H_2S -2 ox state
Group 5A	-3	NH_3 -3 ox state

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 5.13

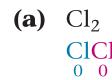
EXAMPLE 5.13 Assigning Oxidation States

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

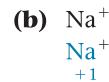
- (a) Cl_2 (b) Na^+ (c) KF (d) CO_2 (e) SO_4^{2-} (f) K_2O_2

SOLUTION

Since Cl_2 is a free element, the oxidation state of both Cl atoms is 0 (rule 1).



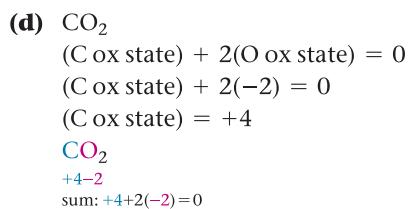
Since Na^+ is a monoatomic ion, the oxidation state of the Na^+ ion is +1 (rule 2).



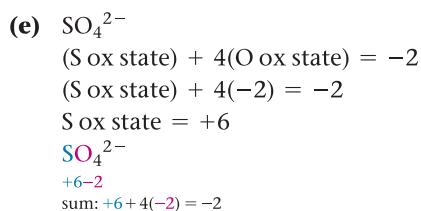
The oxidation state of K is +1 (rule 4). The oxidation state of F is -1 (rule 5). Since this is a neutral compound, the sum of the oxidation states is 0.



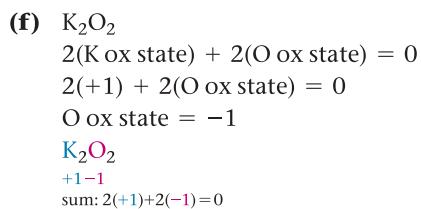
The oxidation state of oxygen is -2 (rule 5). The oxidation state of carbon must be deduced using rule 3, which says that the sum of the oxidation states of all the atoms must be 0 .



The oxidation state of oxygen is -2 (rule 5). You would ordinarily expect the oxidation state of S to be -2 (rule 5). However, if that were the case, the sum of the oxidation states would not equal the charge of the ion. Since O is higher on the list than S, it takes priority, and you calculate the oxidation state of sulfur by setting the *sum* of all of the oxidation states equal to -2 (the charge of the ion).



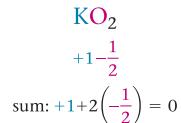
The oxidation state of potassium is $+1$ (rule 4). You would ordinarily expect the oxidation state of O to be -2 (rule 5), but rule 4 takes priority, and you deduce the oxidation state of O by setting the sum of all of the oxidation states equal to 0 .



FOR PRACTICE 5.13 Assign an oxidation state to each atom in each element, ion, or compound.

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

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



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

ANSWER NOW!



5.10 Cc

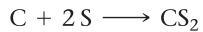
Conceptual Connection

OXIDATION NUMBERS IN POLYATOMIC IONS Which statement best describes the *difference* between the *charge* of a polyatomic ion and the *oxidation states* of its constituent atoms? (For example, the charge of NO_3^- is $1-$, and the oxidation states of its atoms are $+5$ for the nitrogen atom and -2 for each oxygen atom.)

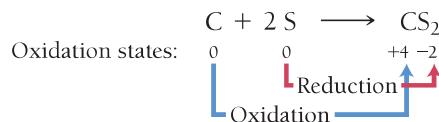
- (a) The charge of a polyatomic ion is a property of the entire ion, while the oxidation states are assigned to each individual atom.
- (b) The oxidation state of the ion is the same as its charge.
- (c) The charge of a polyatomic ion is not a real physical property, while the oxidation states of atoms are actual physical properties.

Identifying Redox Reactions

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



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



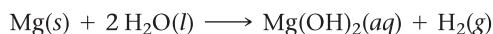
Carbon changes from an oxidation state of 0 to an oxidation state of +4. In terms of our electron bookkeeping scheme (the assigned oxidation state), carbon *loses electrons* and is *oxidized*. Sulfur changes from an oxidation state of 0 to an oxidation state of -2. In terms of our electron bookkeeping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

- Oxidation: An increase in oxidation state
- Reduction: A decrease in oxidation state

Remember that a reduction is a reduction in oxidation state.

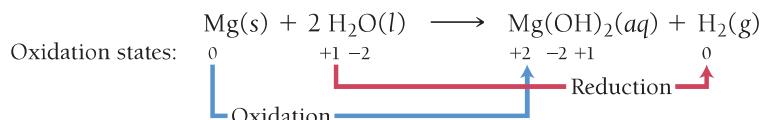
EXAMPLE 5.14 Using Oxidation States to Identify Oxidation and Reduction

Use oxidation states to identify the element that is oxidized and the element that is reduced in the following redox reaction:



SOLUTION

Begin by assigning oxidation states to each atom in the reaction.



Since Mg increased in oxidation state, it was oxidized. Since H decreased in oxidation state, it was reduced.

FOR PRACTICE 5.14 Use oxidation states to identify the element that is oxidized and the element that is reduced in the following redox reaction:



FOR MORE PRACTICE 5.14 Determine whether or not each reaction is a redox reaction. If the reaction is a redox reaction, identify which element is oxidized and which is reduced.

- (a) $\text{Hg}_2(\text{NO}_3)_2(aq) + 2 \text{KBr}(aq) \longrightarrow \text{Hg}_2\text{Br}_2(s) + 2 \text{KNO}_3(aq)$
- (b) $4 \text{Al}(s) + 3 \text{O}_2(g) \longrightarrow 2 \text{Al}_2\text{O}_3(s)$
- (c) $\text{CaO}(s) + \text{CO}_2(g) \longrightarrow \text{CaCO}_3(s)$

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

In Section 20.2, you will learn more about redox reactions, including how to balance them. For now, be able to identify redox reactions, as well as oxidizing and reducing agents, according to the following guidelines.

Redox reactions:

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

In a redox reaction:

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

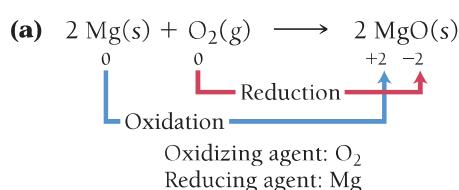
EXAMPLE 5.15 Identifying Redox Reactions, Oxidizing Agents, and Reducing Agents

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

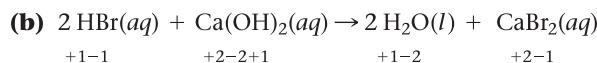
- (a) $2 \text{Mg}(s) + \text{O}_2(g) \longrightarrow 2 \text{MgO}(s)$
 (b) $2 \text{HBr}(aq) + \text{Ca}(\text{OH})_2(aq) \longrightarrow 2 \text{H}_2\text{O}(l) + \text{CaBr}_2(aq)$
 (c) $\text{Zn}(s) + \text{Fe}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + \text{Fe}(s)$

SOLUTION

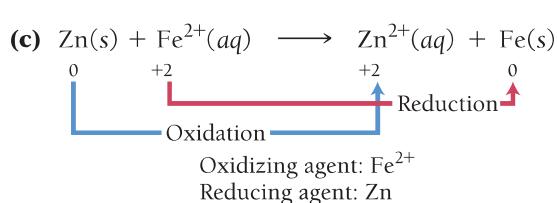
This is a redox reaction because magnesium increases in oxidation number (oxidation) and oxygen decreases in oxidation number (reduction).



This is not a redox reaction because none of the atoms undergo a change in oxidation number.



This is a redox reaction because zinc increases in oxidation number (oxidation) and iron decreases in oxidation number (reduction).



FOR PRACTICE 5.15 Determine whether or not each reaction is a redox reaction. For each redox reaction, identify the oxidizing agent and the reducing agent.

- (a) $2 \text{Li}(s) + \text{Cl}_2(g) \longrightarrow 2 \text{LiCl}(s)$
 (b) $2 \text{Al}(s) + 3 \text{Sn}^{2+}(aq) \longrightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Sn}(s)$
 (c) $\text{Pb}(\text{NO}_3)_2(aq) + 2 \text{LiCl}(aq) \longrightarrow \text{PbCl}_2(s) + 2 \text{LiNO}_3(aq)$
 (d) $\text{C}(s) + \text{O}_2(g) \longrightarrow \text{CO}_2(g)$

ANSWER NOW!

5.11
Cc
 Conceptual Connection

OXIDATION AND REDUCTION Which statement is true?

- (a) A redox reaction involves either the transfer of an electron or a change in the oxidation state of an element.
- (b) If any of the reactants or products in a reaction contain oxygen, the reaction is a redox reaction.
- (c) In a reaction, oxidation can occur independently of reduction.
- (d) In a redox reaction, any increase in the oxidation state of a reactant must be accompanied by a decrease in the oxidation state of another reactant.

The Activity Series: Predicting Whether a Redox Reaction Is Spontaneous

Is there a way to predict whether a particular redox reaction will spontaneously occur? Suppose we knew that substance A has a greater tendency to lose electrons than substance B. (In other words, A is more easily oxidized than B.) Knowing this, we could predict that mixing A with cations of B would result in a redox reaction in which A loses its electrons (A is oxidized) to the cations of B (B is reduced). For example, Mg has a greater tendency to lose electrons than Cu. Consequently, if we put solid Mg into a solution containing Cu^{2+} ions, Mg is oxidized and Cu^{2+} is reduced.





CHEMISTRY IN YOUR DAY | Bleached Blonde

Have you ever bleached your hair? Most home kits for hair bleaching contain hydrogen peroxide (H_2O_2), an excellent oxidizing agent. When applied to hair, hydrogen peroxide oxidizes melanin, the dark pigment that gives hair its color. Once melanin is oxidized, it no longer imparts a dark color, leaving the hair with the familiar bleached look. Hydrogen peroxide also oxidizes other components of hair. For example, protein molecules in hair contain $-\text{SH}$ groups called thiols. Hydrogen peroxide oxidizes these thiol groups to sulfonic acid groups, $-\text{SO}_3\text{H}$. The oxidation of thiol groups to sulfonic acid groups causes changes in the proteins that compose hair, making the hair more brittle and more likely to tangle. Consequently, people with heavily bleached hair generally use conditioners, which contain compounds that form thin, lubricating coatings on individual hair shafts. These coatings prevent tangling and make hair softer and more manageable.

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

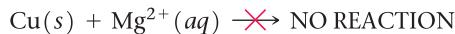


Can you see why this reaction is a redox reaction? Can you identify the oxidizing and reducing agents?

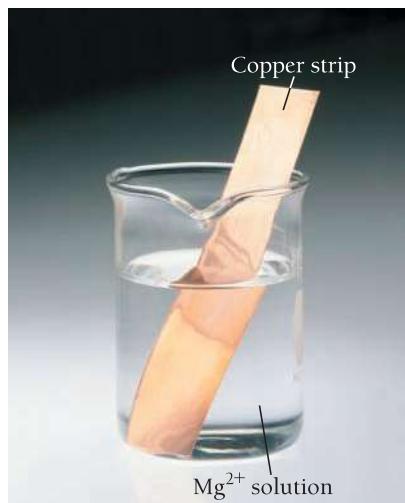


▲ The bleaching of hair involves a redox reaction in which melanin—the main pigment in hair—is oxidized.

We see this as the fading of blue (the color of the Cu^{2+} ions in solution), the dissolving of the solid magnesium, and the appearance of solid copper on the remaining magnesium surface (Figure 5.23▼). This reaction is spontaneous—it occurs on its own when Mg(s) and $\text{Cu}^{2+}(aq)$ come into contact. If, however, we put Cu(s) in a solution containing $\text{Mg}^{2+}(aq)$ ions, no reaction occurs (Figure 5.24▼).



No reaction occurs because, as noted previously, Mg atoms have a greater tendency to lose electrons than do Cu atoms; Cu atoms will therefore not lose electrons to Mg^{2+} ions.



▲ **FIGURE 5.23** Cu^{2+} Oxidizes Magnesium When we put a magnesium strip into a Cu^{2+} solution, the magnesium is oxidized to Mg^{2+} and the copper ion is reduced to Cu(s) . Notice the fading of the blue color (due to Cu^{2+} ions) in solution and the appearance of solid copper on the magnesium strip.

▲ **FIGURE 5.24** Mg^{2+} Does Not Oxidize Copper When we place solid copper in a solution containing Mg^{2+} ions, no reaction occurs.

Table 5.4 presents the **activity series of metals**. This table lists metals in order of decreasing tendency to lose electrons. The metals at the top of the list have the greatest tendency to lose electrons—they are most easily oxidized and therefore the most reactive. The metals at the bottom of the list have the lowest tendency to lose electrons—they are the most difficult to oxidize and therefore the least reactive. It is not a coincidence that the metals used for jewelry, such as silver and gold, are near the bottom of the list. They are among the least reactive metals and therefore do not form compounds easily. Instead, they tend to remain as solid silver and solid gold rather than being oxidized to silver and gold cations by elements in the environment (such as the oxygen in the air).

Each reaction in the activity series is an oxidation *half-reaction*. The half-reactions at the top are most likely to occur in the *forward* direction, and the half-reactions at the bottom are most likely to occur in the *reverse* direction. Consequently, if we pair a half-reaction from the top of the list with the reverse of a half-reaction from the bottom of the list, we get a spontaneous reaction.

More specifically,

Any half-reaction on the list is spontaneous when paired with the reverse of any half-reaction below it on the list.

TABLE 5.4 ■ Activity Series of Metals



▲ Gold is very low on the activity series. Because it is so difficult to oxidize, it resists the tarnishing and corrosion that more active metals undergo.

$\text{Li(s)} \longrightarrow \text{Li}^+(\text{aq}) + \text{e}^-$	Most reactive
$\text{K(s)} \longrightarrow \text{K}^+(\text{aq}) + \text{e}^-$	Most easily oxidized
$\text{Ca(s)} \longrightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{e}^-$	Strongest tendency to lose electrons
$\text{Na(s)} \longrightarrow \text{Na}^+(\text{aq}) + \text{e}^-$	
$\text{Mg(s)} \longrightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^-$	
$\text{Al(s)} \longrightarrow \text{Al}^{3+}(\text{aq}) + 3\text{e}^-$	
$\text{Mn(s)} \longrightarrow \text{Mn}^{2+}(\text{aq}) + 2\text{e}^-$	
$\text{Zn(s)} \longrightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$	
$\text{Cr(s)} \longrightarrow \text{Cr}^{3+}(\text{aq}) + 3\text{e}^-$	
$\text{Fe(s)} \longrightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$	
$\text{Ni(s)} \longrightarrow \text{Ni}^{2+}(\text{aq}) + 2\text{e}^-$	
$\text{Sn(s)} \longrightarrow \text{Sn}^{2+}(\text{aq}) + 2\text{e}^-$	
$\text{Pb(s)} \longrightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{e}^-$	
$\text{H}_2(\text{g}) \longrightarrow 2\text{H}^+(\text{aq}) + 2\text{e}^-$	Least reactive
$\text{Cu(s)} \longrightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$	Most difficult to oxidize
$\text{Ag(s)} \longrightarrow \text{Ag}^+(\text{aq}) + \text{e}^-$	
$\text{Au(s)} \longrightarrow \text{Au}^{3+}(\text{aq}) + 3\text{e}^-$	Least tendency to lose electrons

ANSWER NOW!



5.12 CC

Conceptual Connection

ACTIVITY SERIES

Which metal is most easily oxidized?

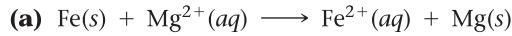
- (a) Na
- (b) Cr
- (c) Au

EXAMPLE 5.16 Predicting Spontaneous Redox Reactions

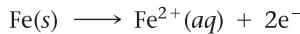
Determine whether each redox reaction is spontaneous.

- (a) $\text{Fe(s)} + \text{Mg}^{2+}(aq) \longrightarrow \text{Fe}^{2+}(aq) + \text{Mg(s)}$
 (b) $\text{Fe(s)} + \text{Pb}^{2+}(aq) \longrightarrow \text{Fe}^{2+}(aq) + \text{Pb(s)}$

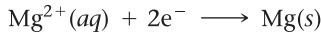
SOLUTION



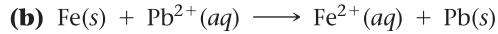
This reaction involves the oxidation of Fe:



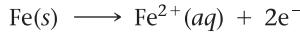
with the reverse of a half-reaction *above it* in the activity series:



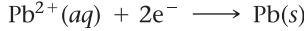
Therefore, the reaction *is not* spontaneous.



This reaction involves the oxidation of Fe:



with the reverse of a half-reaction *below it* in the activity series:



Therefore, the reaction *is* spontaneous.

FOR PRACTICE 5.16 Predicting Spontaneous Redox Reactions

Determine whether each redox reaction is spontaneous.

- (a) $\text{Zn(s)} + \text{Ni}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + \text{Ni(s)}$
 (b) $\text{Zn(s)} + \text{Ca}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + \text{Ca(s)}$

QUIZ YOURSELF NOW!

Self-Assessment Quiz



- Q1.** What is the molarity of a solution containing 55.8 g of MgCl_2 dissolved in 1.00 L of solution?

MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.1

- a) 55.8 M b) 1.71 M c) 0.586 M d) 0.558 M

- Q2.** What mass (in grams) of $\text{Mg}(\text{NO}_3)_2$ is present in 145 mL of a 0.150 M solution of $\text{Mg}(\text{NO}_3)_2$?

MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.2

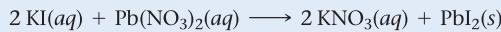
- a) 3.23 g b) 0.022 g c) 1.88 g d) 143 g

- Q3.** What volume of a 1.50 M HCl solution should you use to prepare 2.00 L of a 0.100 M HCl solution?

MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.3

- a) 0.300 L b) 0.133 L c) 30.0 L d) 2.00 L

- Q4.** Potassium iodide reacts with lead(II) nitrate in the following precipitation reaction:



What minimum volume of 0.200 M potassium iodide solution is required to completely precipitate all of the lead in 155.0 mL of a 0.112 M lead(II) nitrate solution?

MISSED THIS? Read Section 5.3; Watch IWE 5.4

- a) 348 mL b) 86.8 mL c) 174 mL d) 43.4 mL

- Q5.** Which solution forms a precipitate when mixed with a solution of aqueous Na_2CO_3 ?

MISSED THIS? Read Section 5.5; Watch KCV 5.5, IWE 5.6

- a) $\text{KNO}_3(aq)$ b) $\text{NaBr}(aq)$
 c) $\text{NH}_4\text{Cl}(aq)$ d) $\text{CuCl}_2(aq)$

- Q6.** What is the net ionic equation for the reaction that occurs when aqueous solutions of KOH and SrCl_2 are mixed?

MISSED THIS? Read Section 5.6

- a) $\text{K}^+(aq) + \text{Cl}^-(aq) \longrightarrow \text{KCl}(s)$
 b) $\text{Sr}^{2+}(aq) + 2\text{OH}^-(aq) \longrightarrow \text{Sr}(\text{OH})_2(s)$
 c) $\text{H}^+(aq) + \text{OH}^-(aq) \longrightarrow \text{H}_2\text{O}(l)$
 d) None of the above because no reaction occurs.

- Q7.** What is the net ionic equation for the reaction that occurs when aqueous solutions of KOH and HNO_3 are mixed?

MISSED THIS? Read Section 5.7; Watch IWE 5.9

- a) $\text{K}^+(aq) + \text{NO}_3^-(aq) \longrightarrow \text{KNO}_3(s)$
 b) $\text{NO}_3^-(aq) + \text{OH}^-(aq) \longrightarrow \text{NO}_3\text{OH}(s)$
 c) $\text{H}^+(aq) + \text{OH}^-(aq) \longrightarrow \text{H}_2\text{O}(l)$
 d) None of the above because no reaction occurs.

- Q8.** The titration of 15.00 mL of HBr solution of unknown concentration requires 18.44 mL of a 0.100 M KOH solution to reach the equivalence point. What is the concentration of the unknown HBr solution in M?

MISSED THIS? Read Section 5.7; Watch IWE 5.11

- a) 0.001844 M b) 0.813 M c) 44 M d) 0.123 M

—Continued on the next page

Continued—

- Q9.** What is the net ionic equation for the reaction that occurs when aqueous solutions of KHCO_3 and HBr are mixed?

MISSED THIS? Read Sections 5.6, 5.8

- a) $\text{K}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) \longrightarrow \text{KC}_2\text{H}_3\text{O}_2(s)$
 b) $\text{H}^+(aq) + \text{HCO}_3^-(aq) \longrightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l)$
 c) $\text{H}^+(aq) + \text{OH}^-(aq) \longrightarrow \text{H}_2\text{O}(l)$
 d) None of the above because no reaction occurs.

- Q10.** What is the oxidation state of carbon in CO_3^{2-} ?

MISSED THIS? Read Section 5.9; Watch KCV 5.13

- a) +4 b) +3 c) -3 d) -2

Answers: 1. (c) 2. (a) 3. (b) 4. (c) 5. (d) 6. (b) 7. (c) 8. (p) 9. (p) 10. (a) 11. (b) 12. (a)

- Q11.** Sodium reacts with water according to the reaction:



Identify the oxidizing agent.

MISSED THIS? Read Section 5.9; Watch KCV 5.13

- a) $\text{Na}(s)$ b) $\text{H}_2\text{O}(l)$ c) $\text{NaOH}(aq)$ d) $\text{H}_2(aq)$
 a) $\text{Cu}^{2+}(aq)$ b) $\text{Zn}^{2+}(aq)$
 c) $\text{Mg}^{2+}(aq)$ d) $\text{K}^+(aq)$

- Q12.** Which of these ions will spontaneously react with $\text{Ni}(s)$ in solution? **MISSED THIS? Read Section 5.9**

CHAPTER 5 IN REVIEW

TERMS

Section 5.2

- solution (168)
 solvent (168)
 solute (168)
 aqueous solution (168)
 dilute solution (168)
 concentrated solution (168)
 molarity (M) (168)
 stock solution (171)

Section 5.4

- electrolyte (175)
 strong electrolyte (176)

nonelectrolyte (176)

- strong acid (176)
 weak acid (177)
 weak electrolyte (177)
 soluble (178)
 insoluble (178)

Section 5.5

- precipitation reaction (179)
 precipitate (179)

Section 5.6

- molecular equation (183)
 complete ionic equation (183)

spectator ion (184)

net ionic equation (184)

equivalence point (189)

indicator (189)

Section 5.7

- acid–base reaction
 (neutralization reaction)
 (185)
 gas-evolution reaction (185)
 Arrhenius definitions (186)
 hydronium ion (186)
 polyprotic acid (186)
 diprotic acid (186)
 salt (186)
 titration (189)

Section 5.9

- oxidation–reduction (redox)
 reaction (193)
 oxidation (193)
 reduction (193)
 oxidation state (oxidation
 number) (194)
 oxidizing agent (197)
 reducing agent (197)
 activity series of metals (200)

CONCEPTS

Solution Concentration and Stoichiometry (5.2, 5.3)

- An aqueous solution is a homogeneous mixture of water (the solvent) with another substance (the solute).
- We express the concentration of a solution in molarity, the number of moles of solute per liter of solution.
- We can use the molarities and volumes of reactant solutions to predict the amount of product that forms in an aqueous reaction.

Aqueous Solutions and Precipitation Reactions (5.4, 5.5)

- Solutes that completely dissociate (or completely ionize in the case of the strong acids) to ions in solution are strong electrolytes, and their solutions are good conductors of electricity.
- Solutes that only partially dissociate (or partially ionize) are weak electrolytes.
- Solutes that do not dissociate (or ionize) are nonelectrolytes.
- A substance that dissolves in water to form a solution is soluble.
- In a precipitation reaction, we mix two aqueous solutions and a solid (precipitate) forms.

- The solubility rules are an empirical set of guidelines that help predict the solubilities of ionic compounds; these rules are especially useful when determining whether or not a precipitate will form.

Equations for Aqueous Reactions (5.6)

- We can represent an aqueous reaction with a molecular equation, which shows the complete neutral formula for each compound in the reaction.
- We can also represent an aqueous reaction with a complete ionic equation, which shows the dissociated nature of strong electrolytes.
- A third representation of an aqueous reaction is the net ionic equation, in which the spectator ions—those that do not change in the course of the reaction—are left out of the equation.

Acid–Base and Gas-Evolution Reactions (5.7, 5.8)

- In an acid–base reaction, an acid, a substance that produces H^+ in solution, reacts with a base, a substance that produces OH^- in solution, and the two neutralize each other, producing water (or in some cases a weak electrolyte).

- An acid-base titration is a laboratory procedure in which a reaction is carried to its equivalence point—the point at which the reactants are in exact stoichiometric proportions; titrations are useful in determining the concentrations of unknown solutions.
- In gas-evolution reactions, two aqueous solutions combine, and a gas is produced.

Oxidation–Reduction Reactions (5.9)

- In oxidation–reduction reactions, one substance transfers electrons to another substance.
- In oxidation–reduction reactions, the substance that loses electrons is oxidized, and the substance that gains them is reduced.

- An oxidation state is a fictitious charge given to each atom in an oxidation–reduction reaction by assigning all shared electrons to the atom with the greater attraction for those electrons. An oxidation state is an imposed electronic bookkeeping scheme, not an actual physical state.
- The oxidation state of an atom increases upon oxidation and decreases upon reduction.
- The activity series of metals can be used to predict spontaneous redox reaction. Any half-reaction in the series is spontaneous when paired with any reverse half-reaction below it.

EQUATIONS AND RELATIONSHIPS

Molarity (M): Solution Concentration (5.2)

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

Solution Dilution (5.2)

$$M_1 V_1 = M_2 V_2$$

Solution Stoichiometry (5.3)

$$\begin{array}{l} \text{volume A} \longrightarrow \text{amount A (in moles)} \longrightarrow \\ \qquad\qquad\qquad \text{amount B (in moles)} \longrightarrow \text{volume B} \end{array}$$

LEARNING OUTCOMES

Chapter Objectives	Assessment
Calculate molarity and apply it to conversion and dilution problems (5.2)	Examples 5.1, 5.2, 5.3 For Practice 5.1, 5.2, 5.3 For More Practice 5.1, 5.2, 5.3 Exercises 21–32
Calculate the amounts of reactants and products involved in aqueous reactions (5.3)	Example 5.4 For Practice 5.4 For More Practice 5.4 Exercises 33–38
Classify compounds as soluble or insoluble, electrolyte or nonelectrolyte (5.4)	Example 5.5 For Practice 5.5 Exercises 39–42
Write chemical equations for precipitation reactions between two or more aqueous solutions (5.5)	Examples 5.6, 5.7 For Practice 5.6, 5.7 Exercises 43–46
Express molecular equations as complete ionic and net ionic equations (5.6)	Example 5.8 For Practice 5.8 Exercises 47–50
Write molecular, complete ionic, and net ionic equations for neutralization reactions (5.7)	Examples 5.9, 5.10 For Practice 5.9, 5.10 Exercises 51–56
Perform calculations involving titration reactions (5.7)	Example 5.11 For Practice 5.11 For More Practice 5.11 Exercises 57–58
Write equations for gas-evolution reactions (5.8)	Example 5.12 For Practice 5.12 For More Practice 5.12 Exercises 59–60
Determine the oxidation state of elements in compounds (5.9)	Example 5.13 For Practice 5.13 Exercises 61–64
Determine if a reaction is a redox reaction and if so, identify the oxidizing and reducing agents (5.9)	Examples 5.14, 5.15 For Practice 5.14, 5.15 For More Practice 5.14 Exercises 65–68
Predict the spontaneity of redox reactions (5.9)	Example 5.16 For Practice 5.16 Exercises 69–72

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

- What is an aqueous solution? What is the difference between the solute and the solvent?
- What is molarity? How is it useful?
- Explain how a strong electrolyte, a weak electrolyte, and a non-electrolyte differ.
- Explain the difference between a strong acid and a weak acid.
- What does it mean for a compound to be soluble? Insoluble?
- What are the solubility rules? How are they useful?
- What are the cations and anions whose compounds are usually soluble? What are the exceptions? What are the anions whose compounds are mostly insoluble? What are the exceptions?
- What is a precipitation reaction? Give an example.
- How can you predict whether a precipitation reaction will occur upon mixing two aqueous solutions?
- Explain how a molecular equation, a complete ionic equation, and a net ionic equation differ.
- What is the Arrhenius definition of an acid? A base?
- What is an acid–base reaction? Give an example.
- Explain the principles behind an acid–base titration. What is an indicator?
- What is a gas-evolution reaction? Give an example.
- What reactant types give rise to gas-evolution reactions?
- What is an oxidation–reduction reaction? Give an example.
- What are oxidation states?
- How can oxidation states be used to identify redox reactions?
- What happens to a substance when it becomes oxidized? Reduced?
- In a redox reaction, which reactant is the oxidizing agent? The reducing agent?

PROBLEMS BY TOPIC

Solution Concentration and Solution Stoichiometry

- Calculate the molarity of each solution.
MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.1
 - 3.25 mol of LiCl in 2.78 L solution
 - 28.33 g C₆H₁₂O₆ in 1.28 L of solution
 - 32.4 mg NaCl in 122.4 mL of solution
- Calculate the molarity of each solution.
 - 0.38 mol of LiNO₃ in 6.14 L of solution
 - 72.8 g C₂H₆O in 2.34 L of solution
 - 12.87 mg KI in 112.4 mL of solution
- What is the molarity of NO₃⁻ in each solution?
MISSED THIS? Read Sections 5.2, 5.4; Watch KCV 5.2, IWE 5.1
 - 0.150 M KNO₃
 - 0.150 M Ca(NO₃)₂
 - 0.150 M Al(NO₃)₃
- What is the molarity of Cl⁻ in each solution?
 - 0.200 M NaCl
 - 0.150 M SrCl₂
 - 0.100 M AlCl₃
- How many moles of KCl are contained in each solution?
MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.2
 - 0.556 L of a 2.3 M KCl solution
 - 1.8 L of a 0.85 M KCl solution
 - 114 mL of a 1.85 M KCl solution
- What volume of 0.200 M ethanol solution contains each amount in moles of ethanol?
 - 0.45 mol ethanol
 - 1.22 mol ethanol
 - 1.2 × 10⁻² mol ethanol
- A laboratory procedure calls for making 400.0 mL of a 1.1 M NaNO₃ solution. What mass of NaNO₃ (in g) is needed?
MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.2
- A chemist wants to make 5.5 L of a 0.300 M CaCl₂ solution. What mass of CaCl₂ (in g) should the chemist use?
- If 123 mL of a 1.1 M glucose solution is diluted to 500.0 mL, what is the molarity of the diluted solution?
MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.3
- If 3.5 L of a 4.8 M SrCl₂ solution is diluted to 45 L, what is the molarity of the diluted solution?
- To what volume should you dilute 50.0 mL of a 12 M stock HNO₃ solution to obtain a 0.100 M HNO₃ solution?
MISSED THIS? Read Section 5.2; Watch KCV 5.2, IWE 5.3
- To what volume should you dilute 25 mL of a 10.0 M H₂SO₄ solution to obtain a 0.150 M H₂SO₄ solution?
- Consider the precipitation reaction:

$$2 \text{Na}_3\text{PO}_4(\text{aq}) + 3 \text{CuCl}_2(\text{aq}) \longrightarrow \text{Cu}_3(\text{PO}_4)_2(\text{s}) + 6 \text{NaCl}(\text{aq})$$
 What volume of 0.175 M Na₃PO₄ solution is necessary to completely react with 95.4 mL of 0.102 M CuCl₂?
- Consider the reaction:

$$\text{Li}_2\text{S}(\text{aq}) + \text{Co}(\text{NO}_3)_2(\text{aq}) \longrightarrow 2 \text{LiNO}_3(\text{aq}) + \text{CoS}(\text{s})$$
 What volume of 0.150 M Li₂S solution is required to completely react with 125 mL of 0.150 M Co(NO₃)₂?
- What is the minimum amount of 6.0 M H₂SO₄ necessary to produce 25.0 g of H₂(g) according to the reaction between aluminum and sulfuric acid?
MISSED THIS? Read Section 5.3; Watch IWE 5.4

$$2 \text{Al}(\text{s}) + 3 \text{H}_2\text{SO}_4(\text{aq}) \longrightarrow \text{Al}_2(\text{SO}_4)_3(\text{aq}) + 3 \text{H}_2(\text{g})$$
- What is the molarity of ZnCl₂ that forms when 25.0 g of zinc completely reacts with CuCl₂ according to the following reaction? Assume a final volume of 275 mL.

$$\text{Zn}(\text{s}) + \text{CuCl}_2(\text{aq}) \longrightarrow \text{ZnCl}_2(\text{aq}) + \text{Cu}(\text{s})$$

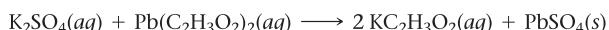
37. A 25.0-mL sample of a 1.20 M potassium chloride solution is mixed with 15.0 mL of a 0.900 M lead(II) nitrate solution, and this precipitation reaction occurs:

MISSED THIS? Read Section 4.3, 5.3; Watch KCV 4.3, IWE 4.6, 5.4



The solid PbCl_2 is collected, dried, and found to have a mass of 2.45 g. Determine the limiting reactant, the theoretical yield, and the percent yield.

38. A 55.0-mL sample of a 0.102 M potassium sulfate solution is mixed with 35.0 mL of a 0.114 M lead(II) acetate solution and this precipitation reaction occurs:



The solid PbSO_4 is collected, dried, and found to have a mass of 1.01 g. Determine the limiting reactant, the theoretical yield, and the percent yield.

Types of Aqueous Solutions and Solubility

39. For each compound (all water soluble), would you expect the resulting aqueous solution to conduct electrical current?

MISSED THIS? Read Section 5.4

- a. CsCl b. CH_3OH c. $\text{Ca}(\text{NO}_3)_2$ d. $\text{C}_6\text{H}_{12}\text{O}_6$

40. Classify each compound as a strong electrolyte or nonelectrolyte.

- a. MgBr_2 b. $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ c. Na_2CO_3 d. KOH

41. Determine whether each compound is soluble or insoluble. If the compound is soluble, list the ions present in solution.

MISSED THIS? Read Section 5.4; Watch IWE 5.5

- a. AgNO_3 b. $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$
c. KNO_3 d. $(\text{NH}_4)_2\text{S}$

42. Determine whether each compound is soluble or insoluble. If the compound is soluble, list the ions present in solution.

- a. AgI b. $\text{Cu}_3(\text{PO}_4)_2$
c. CoCO_3 d. K_3PO_4

Precipitation Reactions

43. Complete and balance each equation. If no reaction occurs, write "NO REACTION."

MISSED THIS? Read Section 5.5; Watch KCV 5.5, IWE 5.6

- a. $\text{LiI}(aq) + \text{BaS}(aq) \longrightarrow$
b. $\text{KCl}(aq) + \text{CaS}(aq) \longrightarrow$
c. $\text{CrBr}_2(aq) + \text{Na}_2\text{CO}_3(aq) \longrightarrow$
d. $\text{NaOH}(aq) + \text{FeCl}_3(aq) \longrightarrow$

44. Complete and balance each equation. If no reaction occurs, write "NO REACTION."

- a. $\text{NaNO}_3(aq) + \text{KCl}(aq) \longrightarrow$
b. $\text{NaCl}(aq) + \text{Hg}_2(\text{C}_2\text{H}_3\text{O}_2)_2(aq) \longrightarrow$
c. $(\text{NH}_4)_2\text{SO}_4(aq) + \text{SrCl}_2(aq) \longrightarrow$
d. $\text{NH}_4\text{Cl}(aq) + \text{AgNO}_3(aq) \longrightarrow$

45. Write a molecular equation for the precipitation reaction that occurs (if any) when each pair of aqueous solutions is mixed. If no reaction occurs, write "NO REACTION."

MISSED THIS? Read Section 5.5; Watch KCV 5.5, IWE 5.6

- a. potassium carbonate and lead(II) nitrate
b. lithium sulfate and lead(II) acetate
c. copper(II) nitrate and magnesium sulfide
d. strontium nitrate and potassium iodide

46. Write a molecular equation for the precipitation reaction that occurs (if any) when each pair of aqueous solutions is mixed. If no reaction occurs, write "NO REACTION."

- a. sodium chloride and lead(II) acetate
b. potassium sulfate and strontium iodide
c. cesium chloride and calcium sulfide
d. chromium(III) nitrate and sodium phosphate

Ionic and Net Ionic Equations

47. Write balanced complete ionic and net ionic equations for each reaction. **MISSED THIS? Read Section 5.6**

- a. $\text{HCl}(aq) + \text{LiOH}(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{LiCl}(aq)$
b. $\text{CaS}(aq) + \text{CuCl}_2(aq) \longrightarrow \text{CuS}(s) + \text{CaCl}_2(aq)$
c. $\text{NaOH}(aq) + \text{HC}_2\text{H}_3\text{O}_2(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{NaC}_2\text{H}_3\text{O}_2(aq)$
d. $\text{Na}_3\text{PO}_4(aq) + \text{NiCl}_2(aq) \longrightarrow \text{Ni}_3(\text{PO}_4)_2(s) + \text{NaCl}(aq)$

48. Write balanced complete ionic and net ionic equations for each reaction.

- a. $\text{K}_2\text{SO}_4(aq) + \text{CaI}_2(aq) \longrightarrow \text{CaSO}_4(s) + \text{KI}(aq)$
b. $\text{NH}_4\text{Cl}(aq) + \text{NaOH}(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{NH}_3(g) + \text{NaCl}(aq)$
c. $\text{AgNO}_3(aq) + \text{NaCl}(aq) \longrightarrow \text{AgCl}(s) + \text{NaNO}_3(aq)$
d. $\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{K}_2\text{CO}_3(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + \text{KC}_2\text{H}_3\text{O}_2(aq)$

49. Mercury(I) ions (Hg_2^{2+}) can be removed from solution by precipitation with Cl^- . Suppose that a solution contains aqueous $\text{Hg}_2(\text{NO}_3)_2$. Write complete ionic and net ionic equations for the reaction of aqueous $\text{Hg}_2(\text{NO}_3)_2$ with aqueous sodium chloride to form solid Hg_2Cl_2 and aqueous sodium nitrate.

MISSED THIS? Read Section 5.6

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

Acid–Base and Gas-Evolution Reactions

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

MISSED THIS? Read Section 5.7; Watch KCV 5.5, IWE 5.9

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

53. Complete and balance each acid–base equation.

MISSED THIS? Read Section 5.7; Watch IWE 5.9

- a. $\text{H}_2\text{SO}_4(aq) + \text{Ca}(\text{OH})_2(aq) \longrightarrow$
b. $\text{HClO}_4(aq) + \text{KOH}(aq) \longrightarrow$
c. $\text{H}_2\text{SO}_4(aq) + \text{NaOH}(aq) \longrightarrow$

54. Complete and balance each acid–base equation.

- a. $\text{HI}(aq) + \text{LiOH}(aq) \longrightarrow$
b. $\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{Ca}(\text{OH})_2(aq) \longrightarrow$
c. $\text{HCl}(aq) + \text{Ba}(\text{OH})_2(aq) \longrightarrow$

55. Write balanced complete ionic and net ionic equations for each acid–base reaction.

MISSED THIS? Read Section 5.7; Watch KCV 5.5, IWE 5.9

- a. $\text{HBr}(aq) + \text{NaOH}(aq) \longrightarrow$
b. $\text{HF}(aq) + \text{NaOH}(aq) \longrightarrow$
c. $\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{RbOH}(aq) \longrightarrow$

56. Write balanced complete ionic and net ionic equations for each acid–base reaction.

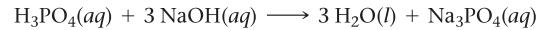
- a. $\text{HI}(aq) + \text{RbOH}(aq) \longrightarrow$
b. $\text{HCHO}_2(aq) + \text{NaOH}(aq) \longrightarrow$
c. $\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{LiOH}(aq) \longrightarrow$

57. A 25.00-mL sample of an unknown HClO_4 solution requires titration with 22.62 mL of 0.2000 M NaOH to reach the equivalence point. What is the concentration of the unknown HClO_4 solution? The neutralization reaction is

MISSED THIS? Read Section 5.7; Watch IWE 5.11



58. A 30.00-mL sample of an unknown H_3PO_4 solution is titrated with a 0.100 M NaOH solution. The equivalence point is reached when 26.38 mL of NaOH solution is added. What is the concentration of the unknown H_3PO_4 solution? The neutralization reaction is



59. Complete and balance each gas-evolution equation.

MISSED THIS? Read Section 5.8; Watch KCV 5.5

- $\text{HBr}(aq) + \text{NiS}(s) \longrightarrow$
- $\text{NH}_4\text{I}(aq) + \text{NaOH}(aq) \longrightarrow$
- $\text{HBr}(aq) + \text{Na}_2\text{S}(aq) \longrightarrow$
- $\text{HClO}_4(aq) + \text{Li}_2\text{CO}_3(aq) \longrightarrow$

60. Complete and balance each gas-evolution equation.

- $\text{HNO}_3(aq) + \text{Na}_2\text{SO}_3(aq) \longrightarrow$
- $\text{HCl}(aq) + \text{KHCO}_3(aq) \longrightarrow$
- $\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{NaHSO}_3(aq) \longrightarrow$
- $(\text{NH}_4)_2\text{SO}_4(aq) + \text{Ca}(\text{OH})_2(aq) \longrightarrow$

Oxidation–Reduction

61. Assign oxidation states to each atom in each element, ion, or compound. **MISSED THIS?** Read Section 5.9; Watch IWE 5.13

- Ag
- Ag^+
- CaF_2
- H_2S
- CO_3^{2-}
- CrO_4^{2-}

62. Assign oxidation states to each atom in each element, ion, or compound.

- Cl_2
- Fe^{3+}
- CuCl_2
- CH_4
- $\text{Cr}_2\text{O}_7^{2-}$
- HSO_4^-

63. What is the oxidation state of Cr in each compound?

MISSED THIS? Read Section 5.9; Watch IWE 5.13

- CrO
- CrO_3
- Cr_2O_3

64. What is the oxidation state of Cl in each ion?

- ClO^-
- ClO_2^-
- ClO_3^-
- ClO_4^-

65. Determine whether each reaction is a redox reaction. For each redox reaction, identify the oxidizing agent and the reducing agent. **MISSED THIS?** Read Section 5.9

- $4 \text{Li}(s) + \text{O}_2(g) \longrightarrow 2 \text{Li}_2\text{O}(s)$
- $\text{Mg}(s) + \text{Fe}^{2+}(aq) \longrightarrow \text{Mg}^{2+}(aq) + \text{Fe}(s)$
- $\text{Pb}(\text{NO}_3)_2(aq) + \text{Na}_2\text{SO}_4(aq) \longrightarrow \text{PbSO}_4(s) + 2 \text{NaNO}_3(aq)$
- $\text{HBr}(aq) + \text{KOH}(aq) \longrightarrow \text{H}_2\text{O}(l) + \text{KBr}(aq)$

66. Determine whether each reaction is a redox reaction. For each redox reaction, identify the oxidizing agent and the reducing agent.

- $\text{Al}(s) + 3 \text{Ag}^+(aq) \longrightarrow \text{Al}^{3+}(aq) + 3 \text{Ag}(s)$
- $\text{SO}_3(g) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{SO}_4(aq)$
- $\text{Ba}(s) + \text{Cl}_2(g) \longrightarrow \text{BaCl}_2(s)$
- $\text{Mg}(s) + \text{Br}_2(l) \longrightarrow \text{MgBr}_2(s)$

67. Determine whether each redox reaction occurs spontaneously in the forward direction. **MISSED THIS?** Read Section 5.9

- $\text{Ni}(s) + \text{Zn}^{2+}(aq) \longrightarrow \text{Ni}^{2+}(aq) + \text{Zn}(s)$
- $\text{Ni}(s) + \text{Pb}^{2+}(aq) \longrightarrow \text{Ni}^{2+}(aq) + \text{Pb}(s)$
- $\text{Al}(s) + 3 \text{Ag}^+(aq) \longrightarrow \text{Al}^{3+}(aq) + 3 \text{Ag}(s)$
- $\text{Pb}(s) + \text{Mn}^{2+}(aq) \longrightarrow \text{Pb}^{2+}(aq) + \text{Mn}(s)$

68. Determine whether each redox reaction occurs spontaneously in the forward direction.

- $\text{Ca}^{2+}(aq) + \text{Zn}(s) \longrightarrow \text{Ca}(s) + \text{Zn}^{2+}(aq)$
- $2 \text{Ag}^+(aq) + \text{Ni}(s) \longrightarrow 2 \text{Ag}(s) + \text{Ni}^{2+}(aq)$
- $\text{Fe}(s) + \text{Mn}^{2+}(aq) \longrightarrow \text{Fe}^{2+}(aq) + \text{Mn}(s)$
- $2 \text{Al}(s) + 3 \text{Pb}^{2+}(aq) \longrightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Pb}(s)$

69. Suppose you wanted to cause Ni^{2+} ions to come out of solution as solid Ni. Which metal could you use to accomplish this?

MISSED THIS? Read Section 5.9

70. Suppose you wanted to cause Pb^{2+} ions to come out of solution as solid Pb. Which metal could you use to accomplish this?

71. Which metal in the activity series reduces Al^{3+} ions but not Na^+ ions? **MISSED THIS?** Read Section 5.9

72. Which metal in the activity series is oxidized with a Ni^{2+} solution but not with a Cr^{3+} solution?

CUMULATIVE PROBLEMS

73. The density of a 20.0% by mass ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$) solution in water is 1.03 g/mL. Find the molarity of the solution.

74. Find the percent by mass of sodium chloride in a 1.35 M NaCl solution. The density of the solution is 1.05 g/mL.

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

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

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

- $\text{HCl}(aq) + \text{Hg}_2(\text{NO}_3)_2(aq) \longrightarrow$
- $\text{KHSO}_3(aq) + \text{HNO}_3(aq) \longrightarrow$
- aqueous ammonium chloride and aqueous lead(II) nitrate
- aqueous ammonium chloride and aqueous calcium hydroxide
- Predict the products and write a balanced molecular equation for each reaction. If no reaction occurs, write “NO REACTION.”
- $\text{H}_2\text{SO}_4(aq) + \text{HNO}_3(aq) \longrightarrow$
- $\text{Cr}(\text{NO}_3)_3(aq) + \text{LiOH}(aq) \longrightarrow$
- aqueous strontium sulfide and aqueous copper(II) sulfate

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

80. An acid solution is 0.100 M in HCl and 0.200 M in H_2SO_4 . What volume of a 0.150 M KOH solution would completely neutralize all the acid in 500.0 mL of this solution?

81. Find the mass of barium metal (in grams) that must react with O_2 to produce enough barium oxide to prepare 1.0 L of a 0.10 M solution of OH^- . (Hint: Barium metal reacts with oxygen to form BaO ; BaO reacts with water to form $\text{Ba}(\text{OH})_2$.)

82. A solution contains Cr^{3+} ions and Mg^{2+} ions. The addition of 1.00 L of 1.51 M NaF solution causes the complete precipitation of these ions as $\text{CrF}_3(s)$ and $\text{MgF}_2(s)$. The total mass of the precipitate is 49.6 g. Find the mass of Cr^{3+} in the original solution.

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

84. We prepare a solution by mixing 0.10 L of 0.12 M sodium chloride with 0.23 L of a 0.18 M MgCl_2 solution. What volume of a 0.20 M silver nitrate solution do we need to precipitate all the Cl^- ion in the solution as AgCl ?

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

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

CHALLENGE PROBLEMS

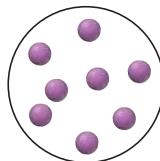
- 87.** A solution contains Ag^+ and Hg^{2+} ions. The addition of 0.100 L of 1.22 M NaI solution is just enough to precipitate all the ions as AgI and HgI_2 . The total mass of the precipitate is 28.1 g. Find the mass of AgI in the precipitate.
- 88.** The water in lakes that have been acidified by acid rain (HNO_3 and H_2SO_4) can be neutralized by a process called liming, in which limestone (CaCO_3) is added to the acidified water. What mass of limestone (in kg) would completely neutralize a 15.2 billion-liter lake that is 1.8×10^{-5} M in H_2SO_4 and 8.7×10^{-6} M in HNO_3 ?
- 89.** Sodium carbonate is often added to laundry detergents to soften hard water and make the detergent more effective. Suppose that a particular detergent mixture is designed to soften hard water that is 3.5×10^{-3} M in Ca^{2+} and 1.1×10^{-3} M in Mg^{2+} and that the average capacity of a washing machine is 19.5 gallons of water. If the detergent requires using 0.65 kg detergent per load of laundry, what percentage (by mass) of the detergent should be sodium

carbonate in order to completely precipitate all of the calcium and magnesium ions in an average load of laundry water?

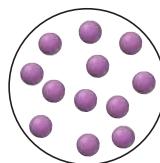
- 90.** Lead poisoning is a serious condition resulting from the ingestion of lead in food, water, or other environmental sources. It affects the central nervous system, leading to a variety of symptoms such as distractibility, lethargy, and loss of motor coordination. Lead poisoning is treated with chelating agents, substances that bind to metal ions, allowing them to be eliminated in the urine. A modern chelating agent used for this purpose is succimer ($\text{C}_4\text{H}_6\text{O}_4\text{S}_2$). Suppose you are trying to determine the appropriate dose for succimer treatment of lead poisoning. What minimum mass of succimer (in mg) is needed to bind all of the lead in a patient's bloodstream? Assume that patient blood lead levels are 45 $\mu\text{g}/\text{dL}$, that total blood volume is 5.0 L, and that 1 mol of succimer binds 1 mol of lead.

CONCEPTUAL PROBLEMS

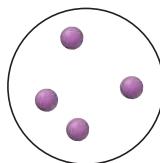
- 91.** The following circle represents 1.0 liter of a solution with a solute concentration of 1 M:



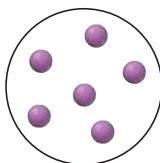
Explain what you would add (the amount of solute or volume of solvent) to the solution to obtain a solution represented by each diagram:



a.



b.

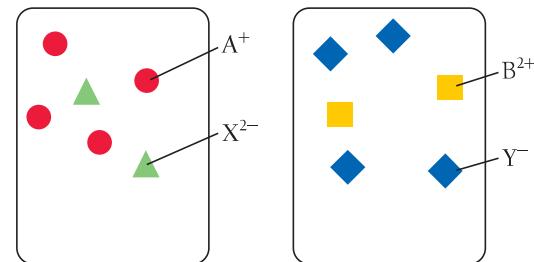


c.

- 92.** Consider the generic ionic compounds with the formulas A_2X and BY_2 and the following solubility rules:

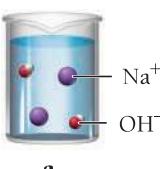
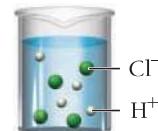
A_2X soluble; BY_2 soluble; AY insoluble; BX soluble.

Assume A^+ ions are circles, B^{2+} ions are squares, X^{2-} ions are triangles, and Y^- ions are diamonds. Solutions of the two compounds (A_2X and BY_2) are represented as follows:



Draw a molecular-level representation showing the result of mixing the two given solutions and write an equation to represent the reaction.

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



a.

b.

c.

d.

QUESTIONS FOR GROUP WORK

Active Classroom Learning

94. Write a detailed set of instructions for making two solutions: (1) 100 mL of 12 M NaOH from solid sodium hydroxide and (2) 1.00 L of 0.1 M NaOH from your first solution. You have in your lab: volumetric flasks marked to contain 100.0 mL and 1.000 L, a graduated cylinder, and a balance.

95. Review the solubility rules. Without referring back to the rules, have each group member list two ionic compounds that are expected to be soluble and two that are expected to be insoluble. Include at least one exception. Check the work of the other members of your group.

96. Define and give an example of each of the following classes of reactions: precipitation, acid-base, gas-evolution, and redox. Each group member can do one, and then present his or her reaction to the group.

97. Using group members to represent atoms, ions, or electrons, act out the reaction $\text{Zn}(s) + \text{Fe}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + \text{Fe}(s)$. Which group member is oxidized? Which is reduced? Which is the oxidizing agent? Which is the reducing agent?



DATA INTERPRETATION AND ANALYSIS

The Flint, Michigan Water Crisis

98. In April of 2014, in an effort to save money, officials in Flint, Michigan, changed their water source from Lake Huron to the Flint River. In subsequent months, residents began complaining about the quality of the water, and General Motors stopped using the water in manufacturing processes because of its corrosiveness. That corrosiveness was causing problems that would soon fuel a national outrage. The water flowed through pipes to taps in homes, and as it flowed through the pipes—many of which contained lead—the corrosive water became contaminated with lead. Routine monitoring of the tap water in select homes did not reveal the magnitude of the problem because samples were collected only after preflushing the tap (allowing the water to run for a time). A Virginia Tech professor and his students began an independent test of the water coming from Flint's taps and got much different results by analyzing the water that initially came from the taps (called 1st draw). Their results—which showed elevated lead levels in the tap water—ultimately forced officials to switch back to the Lake Huron water source.

The table below shows a set of data collected by the Virginia Tech team. The lead levels in water are expressed in units of part per billion (ppb). 1 ppb = 1 g Pb/10⁹ parts solution. Examine the data and answer the questions that follow.

Sample #	Lead Level 1st draw (ppb)	Lead Level 45 sec flush (ppb)	Lead Level 2 min flush (ppb)
1	0.344	0.226	0.145
2	8.133	10.77	2.761
3	1.111	0.11	0.123
4	8.007	7.446	3.384
5	1.951	0.048	0.035
6	7.2	1.4	0.2
7	40.63	9.726	6.132
8	1.1	2.5	0.1
9	10.6	1.038	1.294
10	6.2	4.2	2.3
11	4.358	0.822	0.147
12	24.37	8.796	4.347
13	6.609	5.752	1.433
14	4.062	1.099	1.085
15	29.59	3.258	1.843

Source: FlintWaterStudy.org (2015) "Lead Results from Tap Water Sampling in Flint, MI during the Flint Water Crisis"

- Determine the average value of lead for 1st draw, 45-second flush, and 2-minute flush (round to three significant figures).
- Does the data support the idea that running the tap water before taking a sample made the lead levels in the water appear lower? Why might this occur?
- The EPA requires water providers to monitor drinking water at customer taps. If lead concentrations exceed 15 ppb in 10% or more of the taps sampled, the water provider must notify the customer and take steps to control the corrosiveness of the water. If the water provider in Flint had used 1st-draw samples to monitor lead levels, would they have been required to take action by EPA requirements? If the Flint water provider used 2-min flush samples, would they have had to take action? Which drawing technique do you think more closely mimics the way residents actually use their water?
- Using the highest value of lead from the 1st-draw data set, and assuming a resident drinks 2 L of water per day, calculate the mass of lead that the resident would consume over the course of 1 year. (Assume the water has a density of 1.0 g/mL.)

Cc**ANSWERS TO CONCEPTUAL CONNECTIONS****Molarity**

5.1 (d) $3.0 \text{ L} \times \frac{2.0 \text{ mol}}{\text{L}} = 6.0 \text{ mol}$

Solutions

5.2 (b) The mass of a solution is equal to the mass of the solute plus the mass of the solvent. Although the solute seems to disappear, it really does not, and its mass becomes part of the mass of the solution, in accordance with the law of mass conservation.

Solution Dilution

5.3 (c) Since the volume has doubled, the concentration is halved, so the same volume should contain half as many solute molecules.

Solution Stoichiometry

5.4 (a) A is the limiting reactant. You have equal amounts of both reactants, but the reaction requires twice as much of A as B.

Electrolyte Solutions

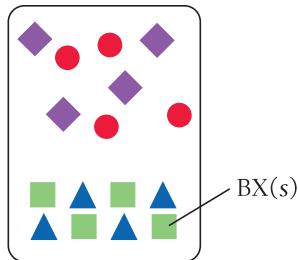
5.5 (a) KBr is an ionic compound and a strong electrolyte.

Solubility of Ionic Compounds

5.6 (c) Compounds containing the NO_3^- ion are soluble with no exceptions.

Precipitation Reactions

5.7 (c)

**Spectator Ions**

5.8 (b) NO_3^- does not participate in the reaction and is therefore a spectator ion.

Acid–Base Titration

5.9 (b) The endpoint occurs when moles of added OH^- equals moles of H^+ in the solution. The solution contains 0.0020 mol H^+ , and 20.0 mL of 0.10 M NaOH only contains 0.0020 mol OH^- .

Oxidation Numbers in Polyatomic Ions

5.10 (a) The charge of a polyatomic ion is the charge associated with the ion *as a whole*. The oxidation states of the individual atoms must sum to the charge of the ion, but they are assigned to *the individual atoms themselves*. Answer **(b)** is incorrect because oxidation state and charge *are not identical*, even though the charge of a *monoatomic* ion is equal to its oxidation state. Answer **(c)** is incorrect because charge *is* a physical property of ions. Conversely, the oxidation states of atoms are *not* real physical properties but an imposed electron bookkeeping scheme.

Oxidation and Reduction

5.11 (d) Since oxidation and reduction must occur together, an increase in the oxidation state of a reactant is always accompanied by a decrease in the oxidation state of another reactant.

Activity Series

5.12 (a) Sodium is highest on the activity series and therefore most easily oxidized.