



CHAPTER 7

AN INTRODUCTION TO CHEMICAL REACTIONS

Now that you understand the basic structural differences between different kinds of substances, you are ready to begin learning about the chemical changes that take place as one substance is converted into another. Chemical changes are chemists' primary concern. They want to know what, if anything, happens when one substance encounters another. Do the substances change? How and why? Can the conditions be altered to speed the changes up, slow them down, or perhaps reverse them? Once chemists understand the nature of one chemical change, they begin to explore the possibilities that arise from causing other similar changes.

For example, let's pretend that you just bought an old house *as is*, with the water turned off. On moving day, you twist the hot water tap as far as it will go, and all you get is a slow drip, drip, drip. As if the lack of hot water weren't enough to ruin your day, you also have a toothache because of a cavity that you haven't had time to get filled. As a chemist in training, you want to know what chemical changes have caused your troubles. In this chapter, you will read about the chemical change that causes a solid to form in your hot water pipes, eventually blocking the flow of water through them. In Chapter 6, you found out about a chemical change that will dissolve that solid, and a similar change that dissolves the enamel on your teeth is described in Chapter 8. Chapter 8 will also show you how fluoride in your toothpaste makes a minor chemical change in your mouth that can help fight cavities.

Chemical changes, like the ones mentioned above, are described with chemical equations. This chapter begins with a discussion of how to interpret and write chemical equations.

7.1 Chemical Reactions and Chemical Equations

7.2 Liquid Water and Water Solutions

7.3 Precipitation Reactions

7.4 Chemical Changes and Energy



A chemical reaction causes solids to form in hot water pipes.

Review Skills

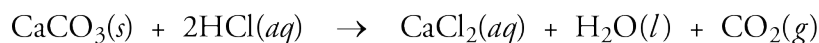
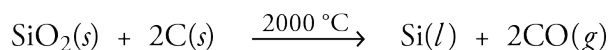
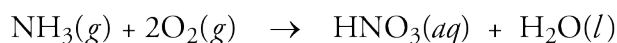
The presentation of information in this chapter assumes that you can already perform the tasks listed below. You can test your readiness to proceed by answering the Review Questions at the end of the chapter. This might also be a good time to read the Chapter Objectives, which precede the Review Questions.

- Write the formulas for the diatomic elements. (Section 3.5)
- Write the definitions of energy, kinetic energy, and potential energy. (Chapter 4 Glossary)
- Describe the relationship between stability, capacity to do work, and potential energy. (Section 4.1)
- Explain why energy must be absorbed to break a chemical bond. (Section 4.1)
- Explain why energy is released when a chemical bond is formed. (Section 4.1)
- Describe the changes that take place during heat transfer between objects at different temperatures. (Section 4.1)
- Predict whether a bond between two atoms of different elements would be a covalent bond or an ionic bond. (Section 5.2)
- Convert between the names and formulas for alcohols, binary covalent compounds, and ionic compounds. (Sections 6.1, 6.2, 6.4, and 6.5)

7.1 Chemical Reactions and Chemical Equations



A **chemical change** or **chemical reaction** is a process in which one or more pure substances are converted into one or more different pure substances. Chemical changes lead to the formation of substances that help grow our food, make our lives more productive, cure our heartburn, and much, much more. For example, nitric acid, HNO_3 , which is used to make fertilizers and explosives, is formed in the chemical reaction of the gases ammonia, NH_3 , and oxygen, O_2 . Silicon dioxide, SiO_2 , reacts with carbon, C , at high temperature to yield silicon, Si —which can be used to make computers—and carbon monoxide, CO . An antacid tablet might contain calcium carbonate, CaCO_3 , which combines with the hydrochloric acid in your stomach to yield calcium chloride, CaCl_2 , water, and carbon dioxide. The chemical equations for these three chemical reactions are below.



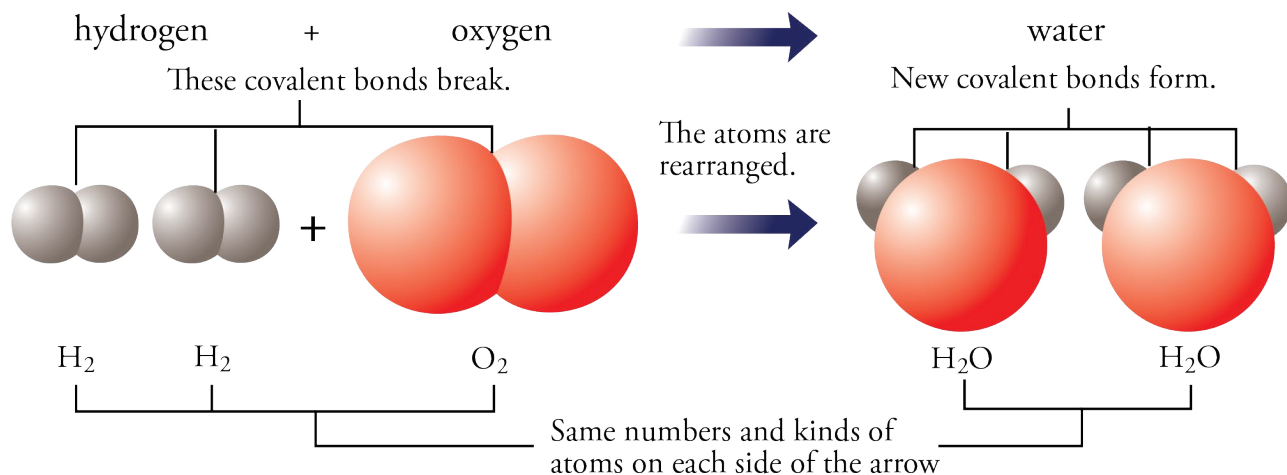
Once you know how to read these chemical equations, they will tell you many details about the reactions that take place.

Interpreting a Chemical Equation

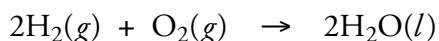
Chemical changes lead to the formation of substances that help grow our food, make our lives more productive, and cure our heartburn.

In chemical reactions, atoms are rearranged and regrouped through the breaking and making of chemical bonds. For example, when hydrogen gas, $\text{H}_2(g)$, is burned in the presence of gaseous oxygen, $\text{O}_2(g)$, a new substance, liquid water, $\text{H}_2\text{O}(l)$, forms. The covalent bonds within the H_2 molecules and O_2 molecules break, and new covalent bonds form between oxygen atoms and hydrogen atoms (Figure 7.1).

Figure 7.1
The Formation of Water from Hydrogen and Oxygen



A chemical equation is a shorthand description of a chemical reaction. The following equation describes the burning of hydrogen gas to form liquid water.

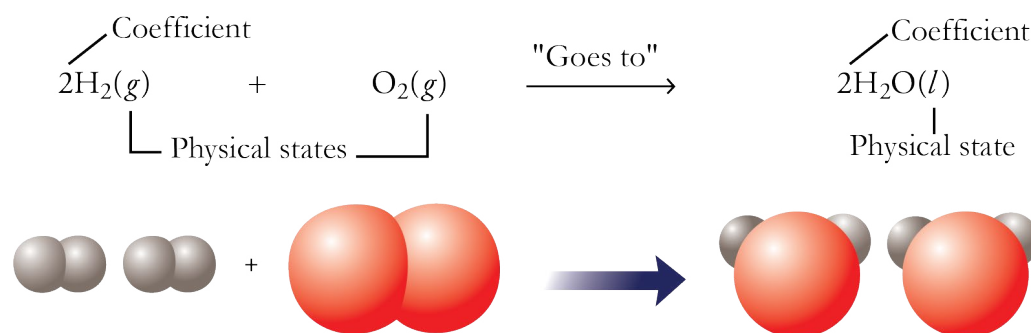


Chemical equations give the following information about chemical reactions.

OBJECTIVE 2

- Chemical equations show the formulas for the substances that take part in the reaction. The formulas on the left side of the arrow represent the **reactants**, the substances that change in the reaction. The formulas on the right side of the arrow represent the **products**, the substances that are formed in the reaction. If there are more than one reactant or more than one product, they are separated by plus signs. The arrow separating the reactants from the products can be read as “goes to” or “yields” or “produces.”
- The physical states of the reactants and products are provided in the equation. A (g) following a formula tells us the substance is a gas. Solids are described with (s) . Liquids are described with (l) . When a substance is dissolved in water, it is described with (aq) for *aqueous*, which means “mixed with water.”
- The relative numbers of particles of each reactant and product are indicated by numbers placed in front of the formulas. These numbers are called **coefficients**. An equation containing correct coefficients is called a balanced equation. For example, the 2's in front of H_2 and H_2O in the equation we saw above are coefficients. If a formula in a balanced equation has no stated coefficient, its coefficient is understood to be 1, as is the case for oxygen in the equation above (Figure 7.2).

OBJECTIVE 3

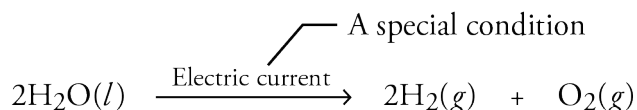


OBJECTIVE 2

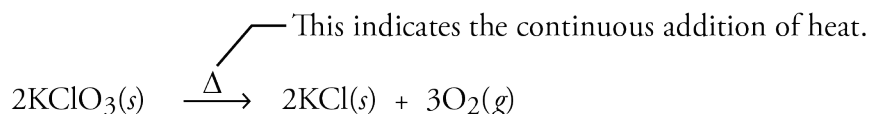
Figure 7.2
The Chemical Equation for the Formation of Water from Hydrogen and Oxygen

- If special conditions are necessary for a reaction to take place, they are often specified above the arrow. Some examples of special conditions are electric current, high temperature, high pressure, and light.

The burning of hydrogen gas must be started with a small flame or a spark, but that is not considered a special condition. There is no need to indicate it above the arrow in the equation for the creation of water from hydrogen and oxygen. However, the conversion of water back to hydrogen and oxygen does require a special condition—specifically, exposure to an electric current:



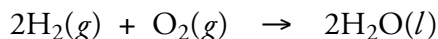
To indicate that a chemical reaction requires the *continuous* addition of heat in order to proceed, we place an upper case Greek delta, Δ , above the arrow in the equation. For example, the conversion of potassium chlorate (a fertilizer and food additive) to potassium chloride and oxygen requires the continuous addition of heat:



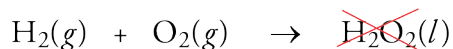
Balancing Chemical Equations

In chemical reactions, atoms are neither created nor destroyed; they merely change partners. Thus the number of atoms of an element in the reaction's products is equal to the number of atoms of that element in the original reactants. The coefficients we often place in front of one or more of the formulas in a chemical equation reflect this fact. They are used whenever necessary to balance the number of atoms of a particular element on either side of the arrow.

For an example, let's return to the reaction of hydrogen gas and oxygen gas to form liquid water. The equation for the reaction between $\text{H}_2(g)$ and $\text{O}_2(g)$ to form $\text{H}_2\text{O}(l)$ shows there are two atoms of oxygen in the diatomic O_2 molecule to the left of the arrow, so there should also be two atoms of oxygen in the product to the right of the arrow. Because each water molecule, H_2O , contains only one oxygen atom, two water molecules must form for each oxygen molecule that reacts. The coefficient 2 in front of the $\text{H}_2\text{O}(l)$ makes this clear. But two water molecules contain four hydrogen atoms, which means that two hydrogen molecules must be present on the reactant side of the equation for the numbers of H atoms to balance (Figure 7.2 on the previous page).



Note that we do not change the subscripts in the formulas, because that would change the identities of the substances. For example, changing the formula on the right of the arrow in the equation above to H_2O_2 would balance the atoms without using coefficients, but the resulting equation would be incorrect.



Water is H_2O , whereas H_2O_2 is hydrogen peroxide, a very different substance from water. (You add water to your hair to clean it; you add hydrogen peroxide to your hair to bleach it.)

The following sample study sheet shows a procedure that you can use to balance chemical equations. It is an approach that chemists often call balancing equations “by inspection.” Examples 7.1 through 7.5, which follow the study sheet, will help to clarify the process.

TIP-OFF You are asked to balance a chemical equation.

GENERAL STEPS

- Consider the first element listed in the first formula in the equation.
If this element is mentioned in two or more formulas on the same side of the arrow, skip it until after the other elements are balanced. (*See Example 7.2.*)
If this element is mentioned in one formula on each side of the arrow, balance it by placing coefficients in front of one or both of these formulas.
- Moving from left to right, repeat the process for each element.
- When you place a number in front of a formula that contains an element you tried to balance previously, recheck that element and put its atoms back in balance. (*See Examples 7.2 and 7.3.*)
- Continue this process until the number of atoms of each element is balanced.

The following strategies can be helpful for balancing certain equations.

STRATEGY 1 Often, an element can be balanced by using the subscript for this element on the left side of the arrow as the coefficient in front of the formula containing this element on the right side of the arrow, and vice versa (using the subscript of this element on the right side of the arrow as the coefficient in front of the formula containing this element on the left side). (*See Example 7.3.*)

STRATEGY 2 It is sometimes easiest, as a temporary measure, to balance the pure nonmetallic elements (H_2 , O_2 , N_2 , F_2 , Cl_2 , Br_2 , I_2 , S_8 , Se_8 , and P_4) with a fractional coefficient ($\frac{1}{2}$, $\frac{3}{2}$, $\frac{5}{2}$, etc.). If you do use a fraction during the balancing process, you can eliminate it later by multiplying each coefficient in the equation by the fraction's denominator (which is usually the number 2). (*See Example 7.4.*)

STRATEGY 3 If polyatomic ions do not change in the reaction, and therefore appear in the same form on both sides of the chemical equation, they can be balanced as though they were single atoms. (*See Example 7.5.*)

STRATEGY 4 If you find an element difficult to balance, leave it for later.

EXAMPLES See Examples 7.1 to 7.5.

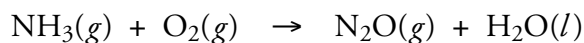
Sample Study Sheet 7.1 Balancing Chemical Equations

OBJECTIVE 4

EXAMPLE 7.1 - Balancing Equations

OBJECTIVE 4

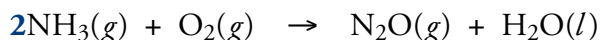
Balance the following equation so that it correctly describes the reaction for the formation of dinitrogen oxide (commonly called nitrous oxide), an anesthetic used in dentistry and surgery.

*Solution*

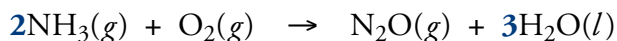
The following table shows that the atoms are not balanced yet.

Element	Left	Right
N	1	2
H	3	2
O	2	2

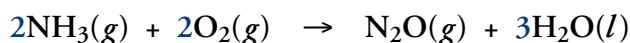
Nitrogen is the first element in the first formula. It is found in one formula on each side of the arrow, so we can try to balance it now. There are two nitrogen atoms on the right side of the equation and only one on the left; we bring them into balance by placing a 2 in front of NH_3 .



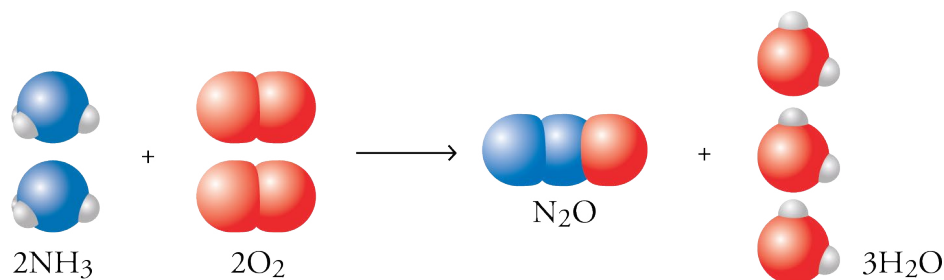
There are now six hydrogen atoms on the left side of the arrow (in the two NH_3 molecules) and only two H's on the right, so we balance the hydrogen atoms by placing a 3 in front of the H_2O . This gives six atoms of hydrogen on each side.



There are now two oxygen atoms on the left and four on the right (in one N_2O and three H_2O 's), so we balance the oxygen atoms by placing a 2 in front of the O_2 .



The following space-filling models show how you might visualize the relative number of particles participating in this reaction. You can see that the atoms regroup but are neither created nor destroyed.

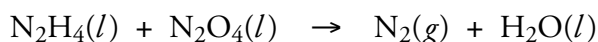


The following table shows that the atoms are now balanced.

Element	Left	Right
N	2	2
H	6	6
O	4	4

EXAMPLE 7.2 - Balancing Equations

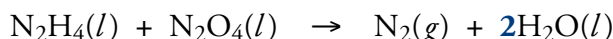
Balance the following equation.



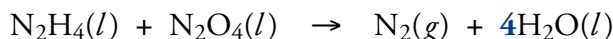
Solution

Nitrogen is the first element in the equation; however, because nitrogen is found in two formulas on the left side of the arrow, we will leave the balancing of the nitrogen atoms until later.

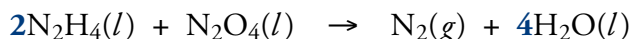
Balance the hydrogen atoms by placing a 2 in front of H_2O .



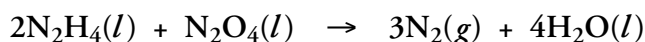
Balance the oxygen atoms by changing the 2 in front of H_2O to a 4.



Because we unbalanced the hydrogen atoms in the process of balancing the oxygen atoms, we need to go back and re-balance the hydrogen atoms by placing a 2 in front of the N_2H_4 .



Finally, we balance the nitrogen atoms by placing a 3 in front of N_2 .



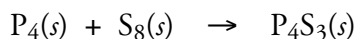
The following table shows that the atoms are now balanced.

Element	Left	Right
N	6	6
H	8	8
O	4	4

OBJECTIVE 4

EXAMPLE 7.3 - Balancing Equations

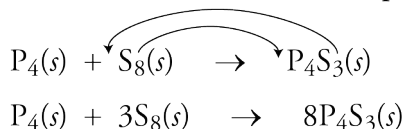
Balance the following equation so that it correctly describes the reaction for the formation of tetraphosphorus trisulfide (used in the manufacture of matches).



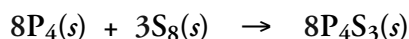
Solution

The phosphorus atoms appear to be balanced at this stage.

We can balance the sulfur atoms by using the subscript for the sulfur on the right (3) as the coefficient for S_8 on the left and using the subscript for the sulfur on the left (8) as the coefficient for the sulfur compound on the right. (Strategy 1)



We restore the balance of the phosphorus atoms by placing an 8 in front of P_4 .



OBJECTIVE 4

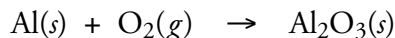


Tetraphosphorus trisulfide is used to make matches.

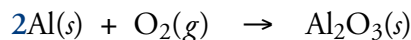
EXAMPLE 7.4 - Balancing Equations

OBJECTIVE 4

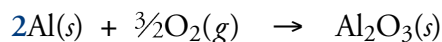
Balance the following equation so that it correctly describes the reaction for the formation of aluminum oxide (used to manufacture glass).

*Solution*

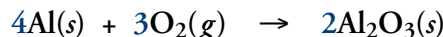
Balance the aluminum atoms by placing a 2 in front of the Al.



There are three oxygen atoms on the right and two on the left. We can bring them into balance by placing $\frac{3}{2}$ in front of the O_2 . Alternatively, we could place a 3 in front of the O_2 and a 2 in front of the Al_2O_3 , but that would un-balance the aluminum atoms. By inserting only one coefficient, in front of the O_2 , to balance the oxygen atoms, the aluminum atoms remain balanced. We arrive at $\frac{3}{2}$ by asking what number times the subscript 2 of the O_2 would give us three atoms of oxygen on the left side: $\frac{3}{2}$ times 2 is 3. (Strategy 2)



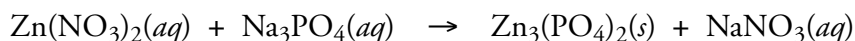
It is a good habit to eliminate the fraction by multiplying all the coefficients by the denominator of the fraction, in this case 2. (Some instructors consider fractional coefficients to be incorrect, so check with your instructor to find out if you will be allowed to leave them in your final answer.)



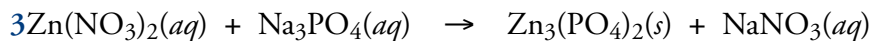
EXAMPLE 7.5 - Balancing Equations

OBJECTIVE 4

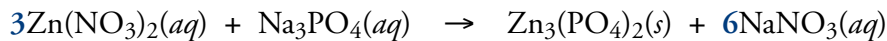
Balance the following equation for the chemical reaction that forms zinc phosphate (used in dental cements and for making galvanized nails).

*Solution*

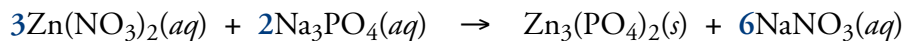
Balance the zinc atoms by placing a 3 in front of $\text{Zn}(\text{NO}_3)_2$.



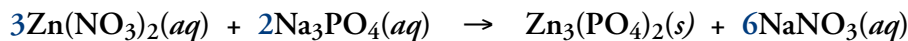
The nitrate ions, NO_3^- , emerge unchanged from the reaction, so we can balance them as though they were single atoms. There are six NO_3^- ions in three $\text{Zn}(\text{NO}_3)_2$. We therefore place a 6 in front of the NaNO_3 to balance the nitrates. (Strategy 3)



Balance the sodium atoms by placing a 2 in front of the Na_3PO_4 .



The phosphate ions, PO_4^{3-} , do not change in the reaction, so we can balance them as though they were single atoms. There are two on each side, so the phosphate ions are balanced. (Strategy 3)



Zinc phosphate is used to galvanize nails.

EXERCISE 7.1 - Balancing Equations

Balance the following chemical equations.

- $$\text{P}_4(s) + \text{Cl}_2(g) \rightarrow \text{PCl}_3(l)$$

Phosphorus trichloride, PCl_3 , is an intermediate for the production of pesticides and gasoline additives.
- $$\text{PbO}(s) + \text{NH}_3(g) \rightarrow \text{Pb}(s) + \text{N}_2(g) + \text{H}_2\text{O}(l)$$

Lead, Pb, is used in storage batteries and as radiation shielding.
- $$\text{P}_4\text{O}_{10}(s) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{PO}_4(aq)$$

Phosphoric acid, H_3PO_4 , is used to make fertilizers and detergents.
- $$\text{Mn}(s) + \text{CrCl}_3(aq) \rightarrow \text{MnCl}_2(aq) + \text{Cr}(s)$$

Manganese(II) chloride, MnCl_2 , is used in pharmaceutical preparations.
- $$\text{C}_2\text{H}_2(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l)$$

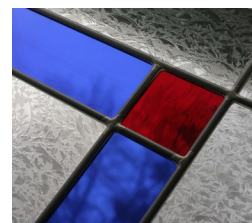
Acetylene, C_2H_2 , is used in welding torches.
- $$\text{Co}(\text{NO}_3)_2(aq) + \text{Na}_3\text{PO}_4(aq) \rightarrow \text{Co}_3(\text{PO}_4)_2(s) + \text{NaNO}_3(aq)$$

Cobalt phosphate, $\text{Co}_3(\text{PO}_4)_2$, is used to color glass and as an additive to animal feed.
- $$\text{CH}_3\text{NH}_2(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{N}_2(g)$$

Methylamine, CH_3NH_2 , is a fuel additive.
- $$\text{FeS}(s) + \text{O}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{Fe}_2\text{O}_3(s) + \text{H}_2\text{SO}_4(aq)$$

Iron(III) oxide, Fe_2O_3 , is a paint pigment.

OBJECTIVE 4



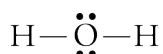
You can find a computer tutorial that will provide more practice balancing equations at the textbook's Web site.

7.2 Liquid Water and Water Solutions

Because many of the chemical changes described in this and other chapters take place in water solutions, it is important that you have a clear image of the nature of liquid water and the process of dissolving substances in water.

Liquid Water

We start with a review of the structure of water molecules. As we discovered in Chapter 5, the Lewis structure of water shows that the oxygen atom has four electron-groups around it: two covalent bonds and two lone pairs.



We predict that the four groups would be distributed in a tetrahedral arrangement to keep their negative charges as far apart as possible. Because the lone pairs are more

repulsive than the bond pairs, the angle between the bond pairs is less than 109.5° . In fact, it is about 105° (Figure 7.3).

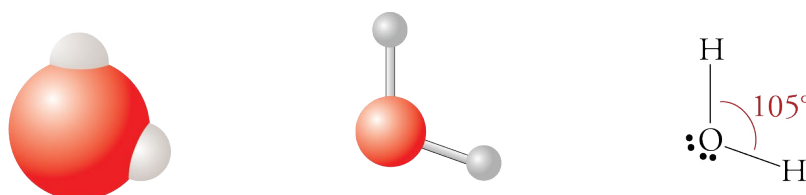


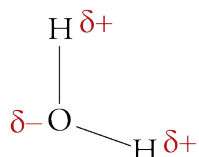
Figure 7.3
Three ways to Describe
a Water Molecule

Space-filling model

Ball-and-stick model

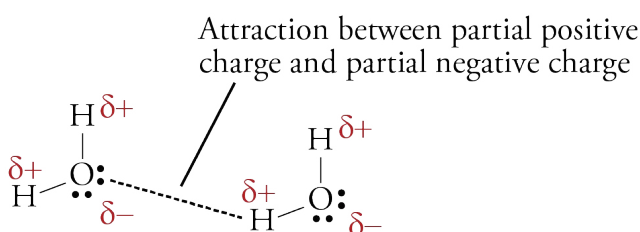
Geometric Sketch

Because oxygen atoms attract electrons much more strongly than do hydrogen atoms, the O-H covalent bond is very polar, leading to a relatively large partial minus charge on the oxygen atom (represented by a δ^-) and a relatively large partial plus charge on the hydrogen atom (represented by a δ^+).



OBJECTIVE 5

The attraction between the region of partial positive charge on one water molecule and the region of partial negative charge on another water molecule tends to hold water molecules close together (Figure 7.4). Remember that opposite charges attract each other and like charges repel each other.



OBJECTIVE 5

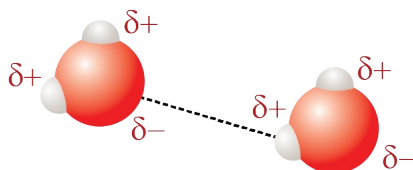
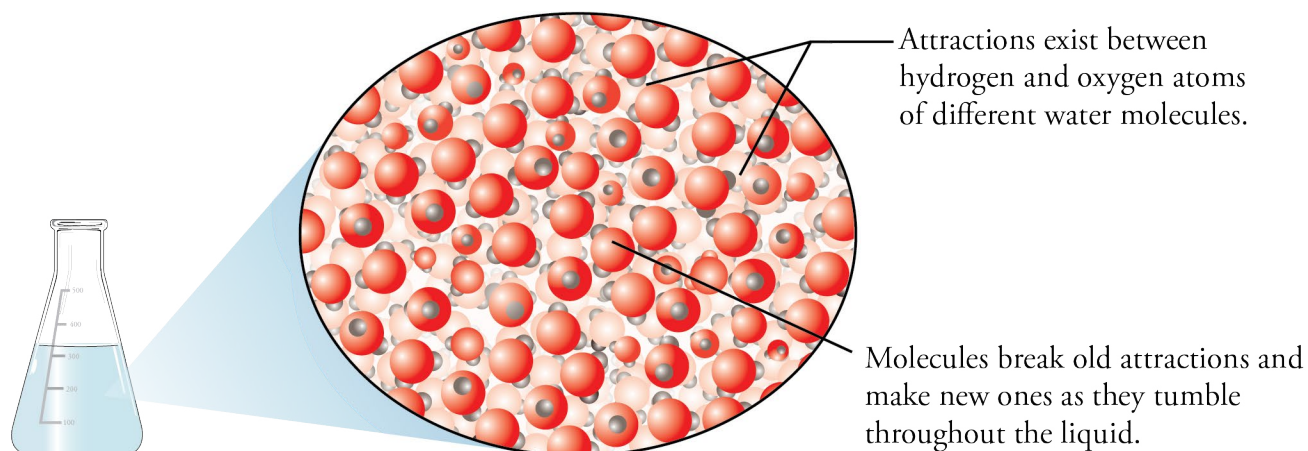


Figure 7.4
Attractions Between
Water Molecules

OBJECTIVE 6

As in other liquids, the attractions between water molecules are strong enough to keep them the same average distance apart, but they are weak enough to allow each molecule to be constantly breaking the attractions that momentarily connect it to some molecules and forming new attractions to other molecules (Figure 7.5). You will find this image of the structure of water useful in developing your understanding of what is happening when salt dissolves in your pasta water and when bubbles form in a soft drink or in a glass of Alka-Seltzer and water.



OBJECTIVE 6

Figure 7.5
Liquid Water

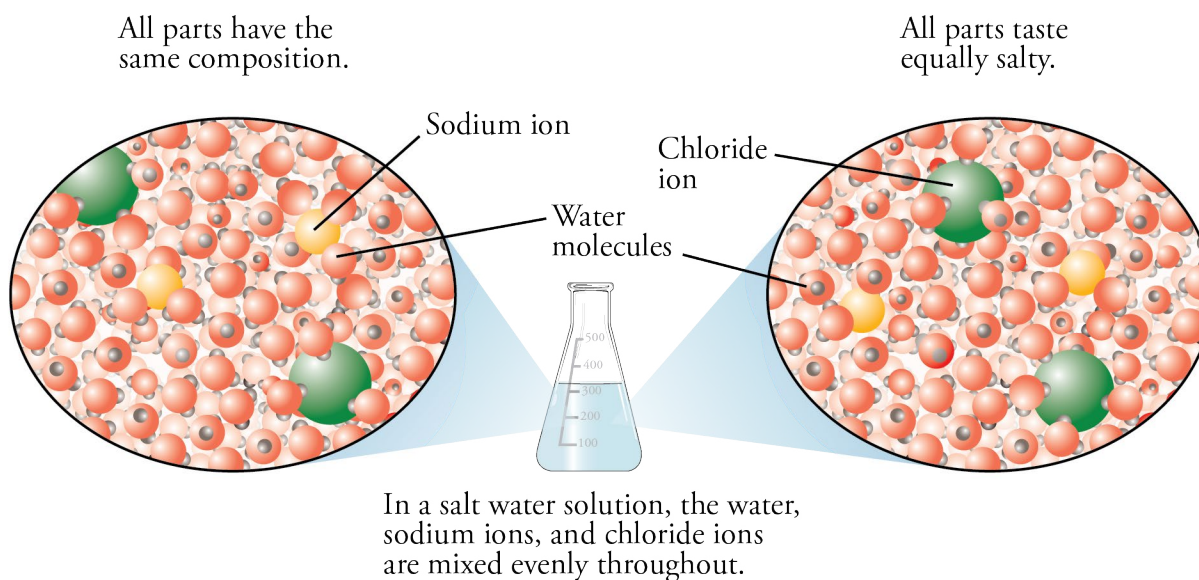
The animation at the textbook's Web site will help you to visualize the structure of water.

Water Solutions of Ionic Compounds

A **solution**, also called a homogeneous mixture, is a mixture whose particles are so evenly distributed that the relative concentrations of the components are the same throughout. When salt dissolves in water, for example, the sodium and chloride ions from the salt spread out evenly throughout the water. Eventually, every part of the solution has the same proportions of water molecules and ions as every other part (Figure 7.6). Ionic compounds are often able to mix with water in this way, in which case we say they are “soluble in water.” Many of the chemical reactions we will study take place in water solutions—that is, solutions in which the substances are dissolved in water. Chemists refer to these as **aqueous solutions**.



An aqueous solution forms when dye is added to water (left). Soon the mixture will be homogeneous (right).

Figure 7.6
Solution (Homogeneous Mixture)

Remember how a model made it easier for us to picture the structures of solids, liquids, and gases in Chapter 3? In order to develop a mental image of the changes that take place on the microscopic level as an ionic compound dissolves in water, chemists have found it helpful to extend that model. We will use table salt or sodium chloride, NaCl, as an example in our description of the process through which an ionic compound dissolves in water.

OBJECTIVE 7

When solid NaCl is first added to water, it settles to the bottom of the container. Like the particles of any solid, the sodium and chloride ions at the solid's surface are constantly moving, although their mutual attractions are holding them more or less in place. For example, if you were riding on a sodium ion at the surface of the solid, you might move out and away from the surface and into the water at one instant, and back toward the surface at the next (pulled by the attractions between your sodium ion and the chloride ions near it). When you move back toward the surface, collisions there might push you away from the surface once again. In this way, all of the ions at the solid's surface, both sodium ions and chloride ions, can be viewed as repeatedly moving out into the water and returning to the solid's surface (Figure 7.7).

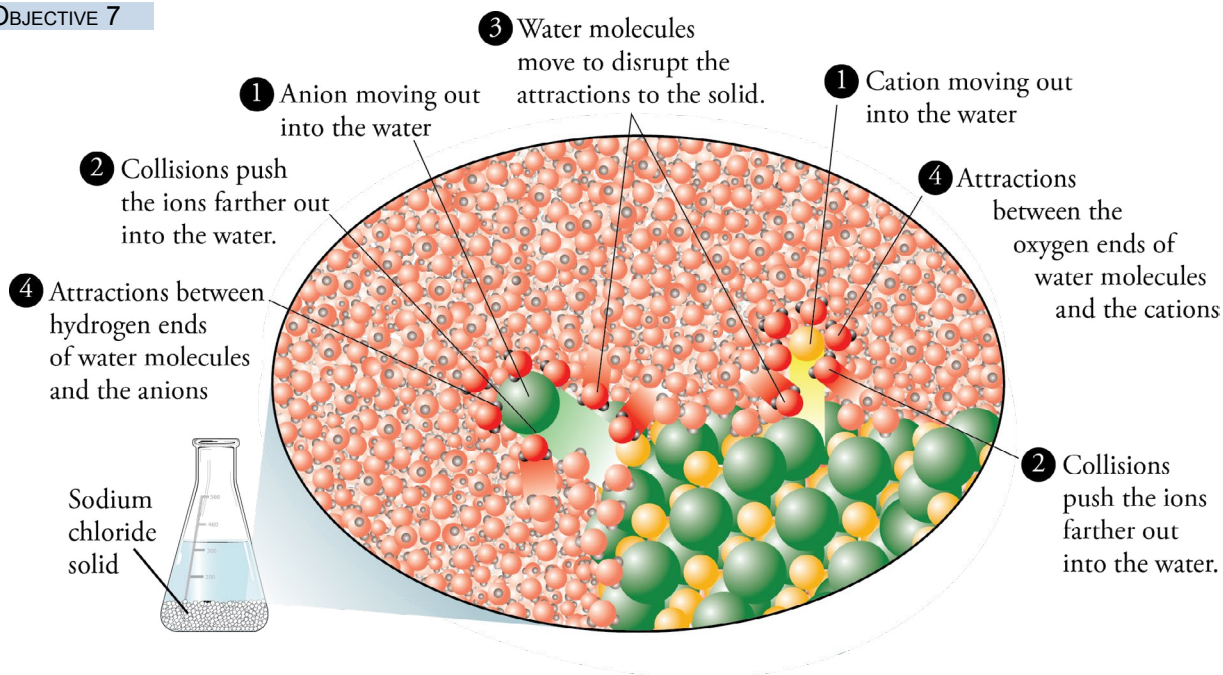
You can find an animation that illustrates the process by which sodium chloride dissolves at the textbook's Web site.

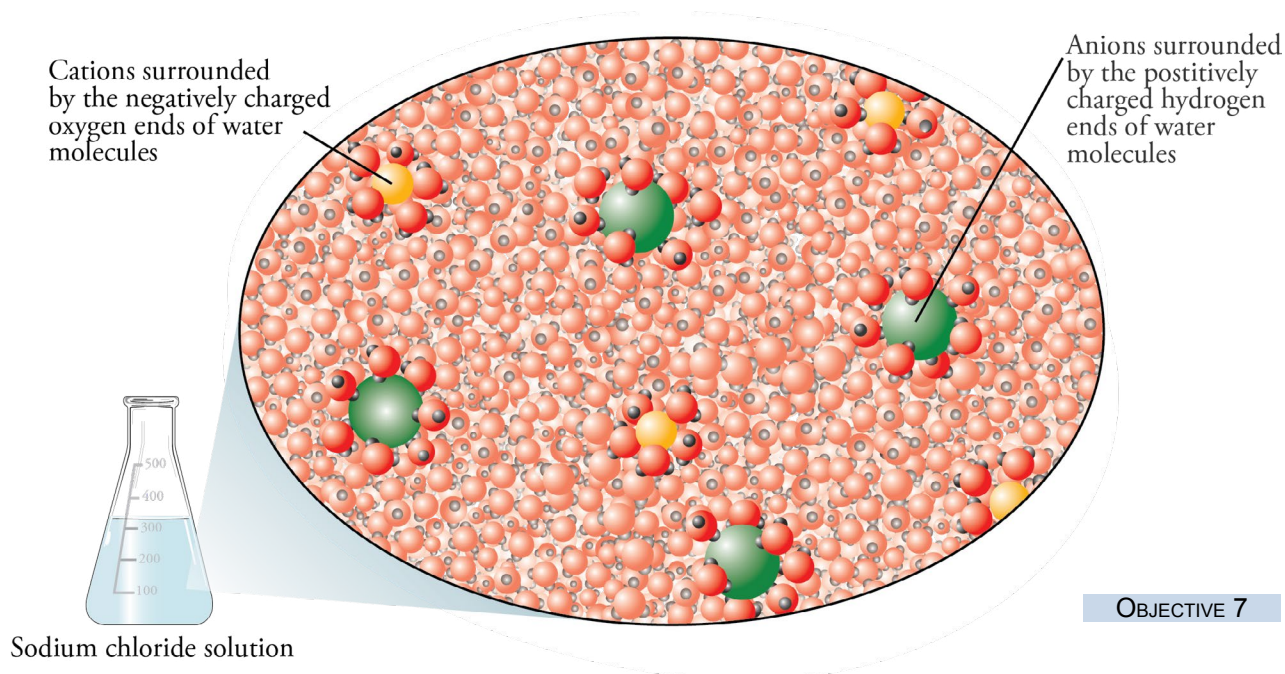
Sometimes when an ion moves out into the water, a water molecule collides with it, pushing the ion out farther into the liquid and helping to break the ionic bond. Other water molecules move into the space between the ion and the solid and shield the ion from the attractions exerted by the ions at the solid's surface (Figure 7.7). The ions that escape the solid are then held in solution by attractions between their own charge and the partial charges of the polar water molecules. The negative oxygen ends of the water molecules surround the cations, and the positive hydrogen ends surround the anions (Figure 7.8).

Figure 7.7

Sodium Chloride Dissolving in Water This shows the mixture immediately after sodium chloride has been added to water. Certain water molecules are highlighted to draw attention to their role in the process.

OBJECTIVE 7



**Figure 7.8**

Aqueous Sodium Chloride This image shows a portion of the solution that forms when sodium chloride dissolves in water. Certain water molecules are highlighted to draw attention to their role in the process.

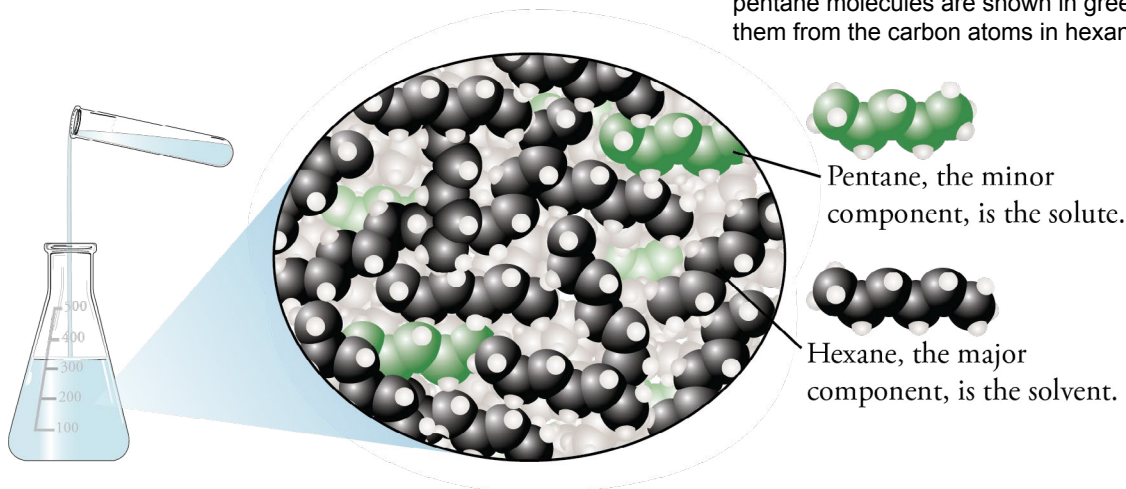
The substances that mix to make a solution, such as our solution of sodium chloride and water, are called the solute and the solvent. In solutions of solids dissolved in liquids, we call the solid the **solute** and the liquid the **solvent**. Therefore, the NaCl in an aqueous sodium chloride solution is the solute, and the water is the solvent. In solutions of gases in liquids, we call the gas the **solute** and the liquid the **solvent**. Therefore, when gaseous hydrogen chloride, $\text{HCl}(g)$, is dissolved in liquid water to form hydrochloric acid, $\text{HCl}(aq)$, the hydrogen chloride is the solute, and water is the solvent. In other solutions, we call the minor component the **solute** and the major component the **solvent**. For example, in a mixture that is 5% liquid pentane, C_5H_{12} , and 95% liquid hexane, C_6H_{14} , the pentane is the solute, and the hexane is the solvent (Figure 7.9).

OBJECTIVE 8

OBJECTIVE 8

Figure 7.9

Liquid-Liquid Solution The carbon atoms in the pentane molecules are shown in green to distinguish them from the carbon atoms in hexane molecules.



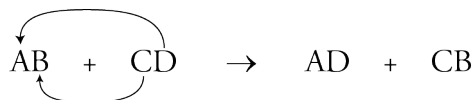
7.3 Precipitation Reactions

The reaction that forms the scale in hot water pipes, eventually leading to major plumbing bills, belongs to a category of reactions called precipitation reactions. So does the reaction of calcium and magnesium ions with soap to create a solid scum in your bathtub and washing machine. Cadmium hydroxide, which is used in rechargeable batteries, is made from the precipitation reaction between water solutions of cadmium acetate and sodium hydroxide. This section describes the changes that occur in precipitation reactions, shows you how to predict when they take place, and shows you how to describe them using chemical equations.

Precipitation Reactions

OBJECTIVE 9

Precipitation reactions, such as the ones we will see in this section, belong to a general class of reactions called **double-displacement reactions**. (Double displacement reactions are also called **double-replacement**, **double-exchange**, or **metathesis reactions**.) Double displacement reactions have the following form, signifying that the elements in two reacting compounds change partners.

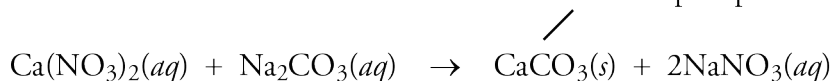


Precipitation reactions take place between ionic compounds in solution. For example, in the precipitation reactions that we will see, A and C represent the cationic (or positively charged) portions of the reactants and products, and B and D represent the anionic (or negatively charged) portions of the reactants and products. The cation of the first reactant (A) combines with the anion of the second reactant (D) to form the product AD, and the cation of the second reactant (C) combines with the anion of the first reactant to form the product CB.

OBJECTIVE 10

Sometimes a double-displacement reaction has one product that is insoluble in water. As that product forms, it emerges, or precipitates, from the solution as a solid. This process is called **precipitation**, such a reaction is called a **precipitation reaction**, and the solid is called the **precipitate**. For example, when water solutions of calcium nitrate and sodium carbonate are mixed, calcium carbonate precipitates from the solution while the other product, sodium nitrate, remains dissolved.

This solid precipitates from the solution. It is a precipitate.



One of the goals of this section is to help you to visualize the process described by this equation. Figures 7.10, 7.11, and 7.12 will help you do this.

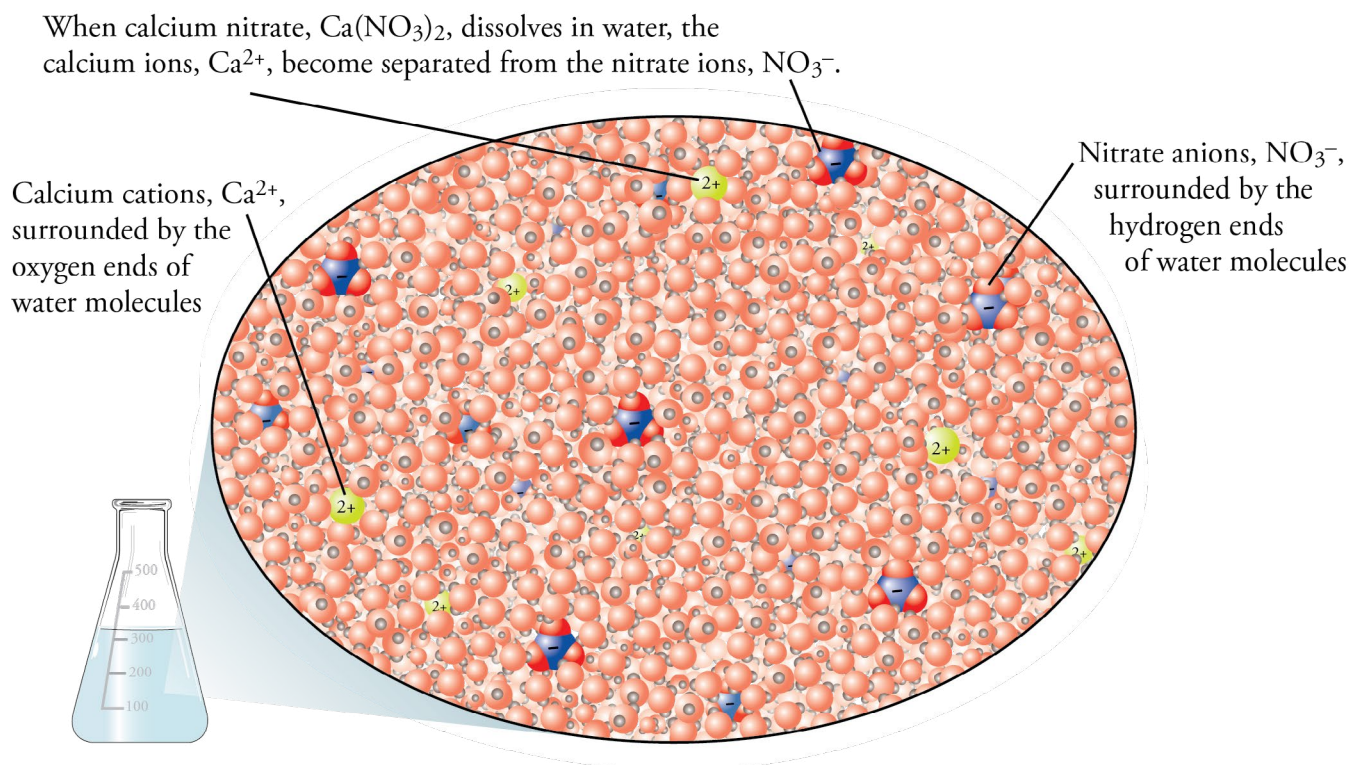
First, let us imagine the particles making up the $\text{Ca}(\text{NO}_3)_2$ solution. Remember that when ionic compounds dissolve, the ions separate and become surrounded by water molecules. When $\text{Ca}(\text{NO}_3)_2$ dissolves in water (Figure 7.10), the Ca^{2+} ions separate from the NO_3^- ions, with the oxygen ends of water molecules surrounding the calcium ions, and the hydrogen ends of water molecules surrounding the nitrate ions.

OBJECTIVE 10

An aqueous solution of sodium carbonate also consists of ions separated and surrounded by water molecules, much like the solution of calcium nitrate. If time were to stop at the instant that the solution of sodium carbonate was added to aqueous calcium nitrate, there would be four different ions in solution surrounded by water molecules: Ca^{2+} , NO_3^- , Na^+ , and CO_3^{2-} . The oxygen ends of the water molecules surround the calcium and sodium ions, and the hydrogen ends of water molecules surround the nitrate and carbonate ions. Figure 7.11 on the next page shows the system at the instant just after solutions of calcium nitrate and sodium carbonate are combined and just before the precipitation reaction takes place. Because a chemical reaction takes place as soon as the $\text{Ca}(\text{NO}_3)_2$ and Na_2CO_3 solutions are combined, the four-ion system shown in this figure lasts for a very short time.

OBJECTIVE 10

Figure 7.10
Aqueous Calcium Nitrate
There are twice as many -1 nitrate ions as +2 calcium ions.



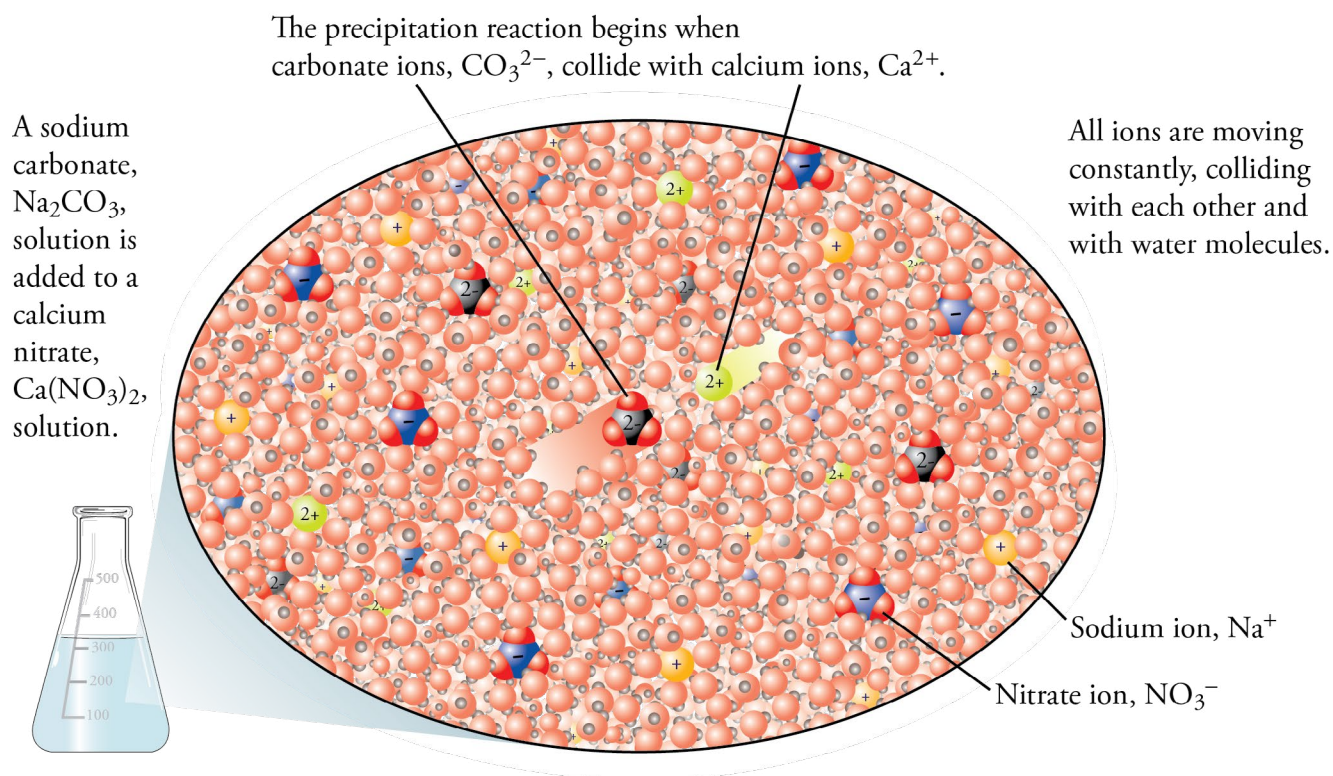


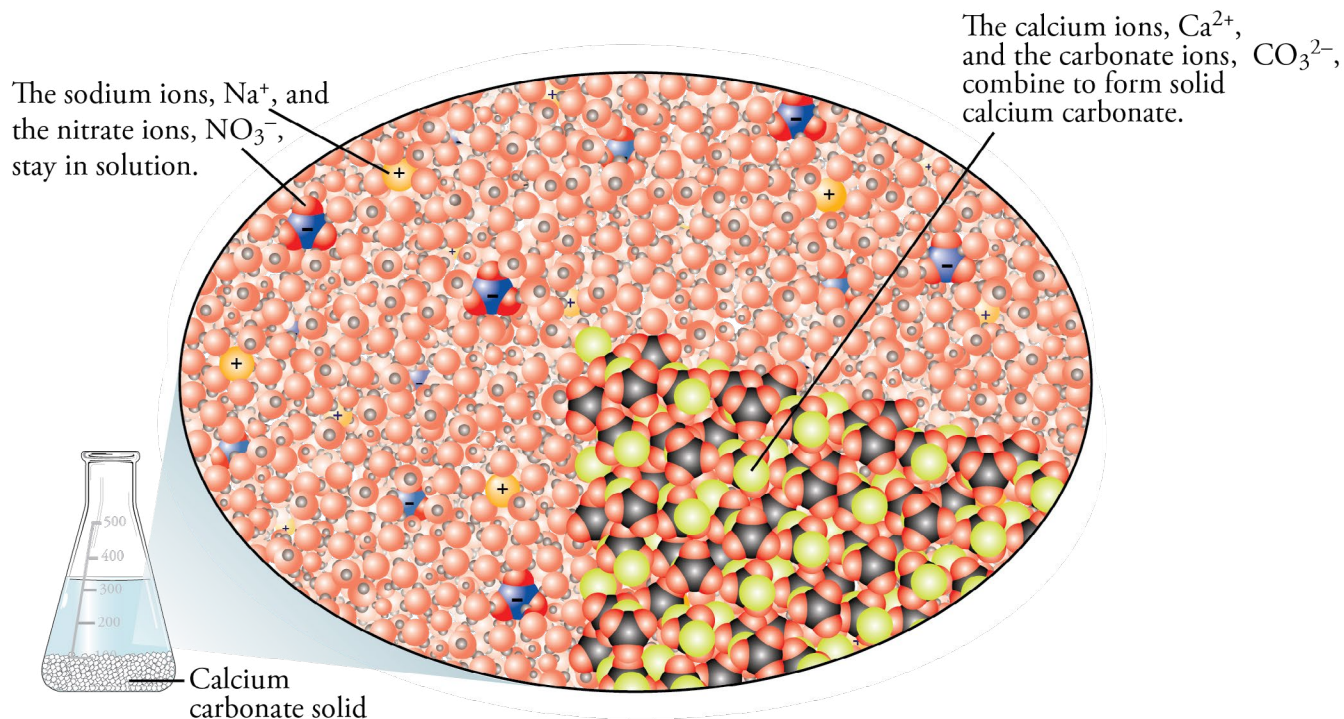
Figure 7.11
Mixture of $\text{Ca}(\text{NO}_3)_2(\text{aq})$ and
 $\text{Na}_2\text{CO}_3(\text{aq})$ at the Instant
They Are Combined

OBJECTIVE 10

OBJECTIVE 10

The ions in solution move in a random way, like any particle in a liquid, so they will constantly collide with other ions. When two cations or two anions collide, they repel each other and move apart. When a calcium ion and a nitrate ion collide, they may stay together for a short time, but the attraction between them is too weak to keep them together when water molecules collide with them and push them apart. The same is true for the collision between sodium ions and carbonate ions. After colliding, they stay together for only an instant before water molecules break them apart again.

When calcium ions and carbonate ions collide, however, they stay together longer because the attraction between them is stronger than the attractions between the other pairs of ions. They might eventually be knocked apart, but while they are together, other calcium ions and carbonate ions can collide with them. When another Ca^{2+} or CO_3^{2-} ion collides with a CaCO_3 pair, a trio forms. Other ions collide with the trio to form clusters of ions that then grow to become small **crystals**—solid particles whose component atoms, ions, or molecules are arranged in an organized, repeating pattern. Many crystals form throughout the system, so the solid CaCO_3 at first appears as a cloudiness in the mixture. The crystals eventually settle to the bottom of the container (Figures 7.11 and 7.12).

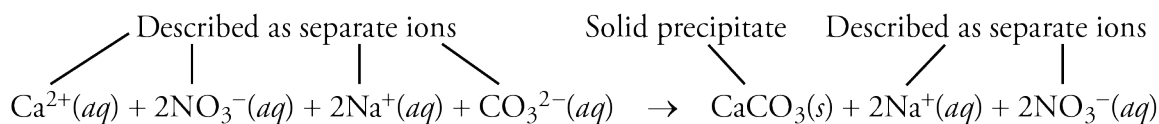


OBJECTIVE 10

Figure 7.12

Product Mixture for the Reaction of $\text{Ca}(\text{NO}_3)_2(aq)$ and $\text{Na}_2\text{CO}_3(aq)$

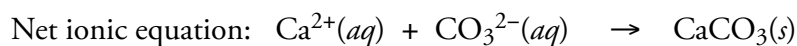
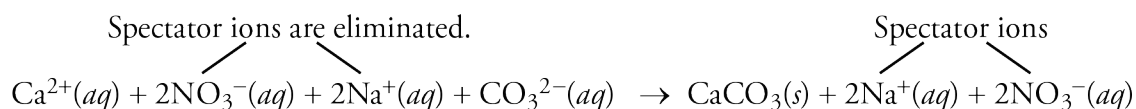
The equation that follows, which is often called a **complete ionic equation**, describes the forms taken by the various substances in solution. The ionic compounds dissolved in the water are described as separate ions, and the insoluble ionic compound is described with a complete formula.



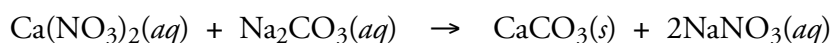
The sodium and nitrate ions remain unchanged in this reaction. They were separate and surrounded by water molecules at the beginning, and they are still separate and surrounded by water molecules at the end. They were important in delivering the calcium and carbonate ions to solution (the solutions were created by dissolving solid calcium nitrate and solid sodium carbonate in water), but they did not actively participate in the reaction. When ions play this role in a reaction, we call them **spectator ions**.

OBJECTIVE 10

Because spectator ions are not involved in the reaction, they are often left out of the chemical equation. The equation written without the spectator ions is called a **net ionic equation**.



We call the equation that shows the complete formulas for all of the reactants and products the **complete equation**, or, sometimes, the **molecular equation**.



You can find an animation that shows this precipitation reaction at the textbook's Web site.

Predicting Water Solubility

In order to predict whether a precipitation reaction will take place when two aqueous ionic compounds are mixed, you need to be able to predict whether the possible products of the double-displacement reaction are soluble or insoluble in water.

When we say that one substance is soluble in another, we mean that they can be mixed to a significant degree. More specifically, chemists describe the **solubility** of a substance as the maximum amount of it that can be dissolved in a given amount of solvent at a particular temperature. This property is often described in terms of the maximum number of grams of solute that will dissolve in 100 milliliters (or 100 grams) of solvent. For example, the water solubility of calcium nitrate is 121.2 g $\text{Ca}(\text{NO}_3)_2$ per 100 mL water at 25 °C. This means that when calcium nitrate is added steadily to 100 mL of water at 25 °C, it will dissolve until 121.2 g $\text{Ca}(\text{NO}_3)_2$ have been added. If more $\text{Ca}(\text{NO}_3)_2$ is added to the solution, it will remain in the solid form.

When we say an ionic solid is insoluble in water, we do not mean that none of the solid dissolves. There are always some ions that can escape from the surface of an ionic solid in water and go into solution. Thus, when we say that calcium carbonate is insoluble in water, what we really mean is that the solubility is very low (0.0014 g CaCO_3 per 100 mL H_2O at 25 °C).

OBJECTIVE 11

Solubility is difficult to predict with confidence. The most reliable way to obtain a substance's solubility is to look it up on a table of physical properties in a reference book. When that is not possible, you can use the following guidelines for predicting whether some substances are soluble or insoluble in water. They are summarized in Table 7.1.

- Ionic compounds with group 1 (or 1A) metallic cations or ammonium cations, NH_4^+ , form soluble compounds no matter what the anion is.

- Ionic compounds with acetate, $C_2H_3O_2^-$, or nitrate, NO_3^- , ions form soluble compounds no matter what the cation is.
- Compounds containing chloride, Cl^- , bromide, Br^- , or iodide, I^- , ions are water-soluble except with silver ions, Ag^+ , and lead(II) ions, Pb^{2+} .
- Compounds containing the sulfate ion, SO_4^{2-} , are water-soluble except with barium ions, Ba^{2+} , and lead(II) ions, Pb^{2+} .
- Compounds containing carbonate, CO_3^{2-} , phosphate, PO_4^{3-} , or hydroxide, OH^- , ions are insoluble in water except with group 1 metallic ions and ammonium ions.

Table 7.1 Water Solubility of Ionic Compounds

OBJECTIVE 11

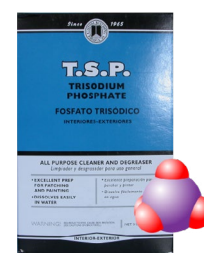
Category	Ions	Except with these ions	Examples
Soluble cations	Group 1 metallic ions and ammonium, NH_4^+	No exceptions	Na_2CO_3 , $LiOH$, and $(NH_4)_2S$ are soluble.
Soluble anions	NO_3^- and $C_2H_3O_2^-$	No exceptions	$Bi(NO_3)_3$, and $Co(C_2H_3O_2)_2$ are soluble.
Usually soluble anions	Cl^- , Br^- , and I^-	Soluble with some exceptions, including with Ag^+ and Pb^{2+}	$CuCl_2$ is water soluble, but $AgCl$ is insoluble.
	SO_4^{2-}	Soluble with some exceptions, including with Ba^{2+} and Pb^{2+}	$FeSO_4$ is water soluble, but $BaSO_4$ is insoluble.
Usually insoluble	CO_3^{2-} , PO_4^{3-} , and OH^-	Insoluble with some exceptions, including with group 1 elements and NH_4^+	$CaCO_3$, $Ca_3(PO_4)_2$, and $Mn(OH)_2$ are insoluble in water, but $(NH_4)_2CO_3$, Li_3PO_4 , and $CsOH$ are soluble.

EXERCISE 7.2 - Predicting Water Solubility

OBJECTIVE 11

Predict whether each of the following is soluble or insoluble in water.

- $Hg(NO_3)_2$ (used to manufacture felt)
- $BaCO_3$ (used to make radiation resistant glass for color TV tubes)
- K_3PO_4 (used to make liquid soaps)
- $PbCl_2$ (used to make other lead salts)
- $Cd(OH)_2$ (in storage batteries)



Sodium phosphate (or trisodium phosphate), Na_3PO_4 , is an all-purpose cleaner.

You can find a computer tutorial that will provide more practice predicting water solubility at the textbook's Web site.

The study sheet on the next page will guide you in predicting whether precipitation reactions take place and help you write chemical equations for precipitation reactions.

Sample Study Sheet 7.2

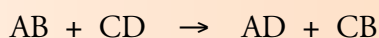
Predicting Precipitation Reactions and Writing Precipitation Equations

OBJECTIVE 12

TIP-OFF You are asked to predict whether a precipitation reaction will take place between two aqueous solutions of ionic compounds, and if the answer is yes, to write the complete equation for the reaction.

GENERAL STEPS

Step 1 Determine the formulas for the possible products using the general double-displacement equation. (Remember to consider ion charges when writing your formulas.)



Step 2 Predict whether either of the possible products is water-insoluble. If either possible product is insoluble, a precipitation reaction takes place, and you may continue with step 3. If neither is insoluble, write “No reaction.”

Step 3 Follow these steps to write the complete equation.

- Write the formulas for the reactants separated by a + sign.
- Separate the formulas for the reactants and products with an arrow.
- Write the formulas for the products separated by a + sign.
- Write the physical state for each formula.
 - The insoluble product will be followed by (s).
 - Water-soluble ionic compounds will be followed by (aq).
- Balance the equation.

EXAMPLES See Examples 7.6-7.8.

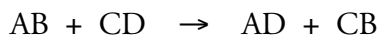
EXAMPLE 7.6 - Predicting Precipitation Reactions

OBJECTIVE 12

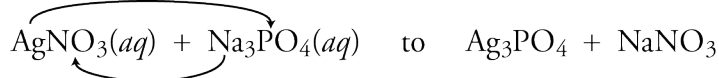
Predict whether a precipitate will form when water solutions of silver nitrate, $\text{AgNO}_3(aq)$, and sodium phosphate, $\text{Na}_3\text{PO}_4(aq)$, are mixed. If there is a precipitation reaction, write the complete equation that describes the reaction.

Solution

Step 1 Determine the possible products using the general double-displacement equation.



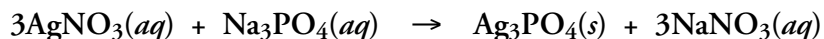
In AgNO_3 , Ag^+ is A, and NO_3^- is B. In Na_3PO_4 , Na^+ is C, and PO_4^{3-} is D. The possible products from the mixture of $\text{AgNO}_3(aq)$ and $\text{Na}_3\text{PO}_4(aq)$ are Ag_3PO_4 and NaNO_3 . (Remember to consider charge when you determine the formulas for the possible products.)



Step 2 Predict whether either of the possible products is water-insoluble.

According to our solubility guidelines, most phosphates are insoluble, and compounds with Ag^+ are not listed as an exception. Therefore, silver phosphate, Ag_3PO_4 , which is used in photographic emulsions, would be insoluble. Because compounds containing Na^+ and NO_3^- are soluble, NaNO_3 is soluble.

Step 3 Write the complete equation. (Don't forget to balance it.)



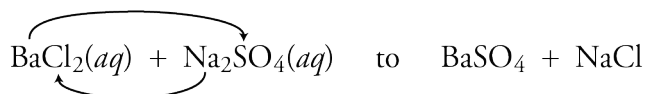
EXAMPLE 7.7 - Predicting Precipitation Reactions

Predict whether a precipitate will form when water solutions of barium chloride, $\text{BaCl}_2(aq)$, and sodium sulfate, $\text{Na}_2\text{SO}_4(aq)$, are mixed. If there is a precipitation reaction, write the complete equation that describes the reaction.

OBJECTIVE 12

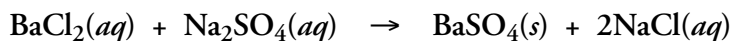
Solution

Step 1 In BaCl_2 , A is Ba^{2+} , and B is Cl^- . In Na_2SO_4 , C is Na^+ , and D is SO_4^{2-} . The possible products from the reaction of $\text{BaCl}_2(aq)$ and $\text{Na}_2\text{SO}_4(aq)$ are BaSO_4 and NaCl .



Step 2 According to our solubility guidelines, most sulfates are soluble, but BaSO_4 is an exception. It is insoluble and would precipitate from the mixture. Because compounds containing Na^+ (and most containing Cl^-) are soluble, NaCl is soluble.

Step 3



This is the reaction used in industry to form barium sulfate, which is used in paint preparations and in x-ray photography.

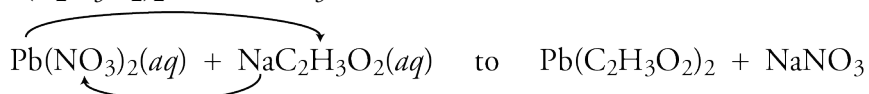
EXAMPLE 7.8 - Predicting Precipitation Reactions

Predict whether a precipitate will form when lead(II) nitrate, $\text{Pb}(\text{NO}_3)_2(aq)$, and sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2(aq)$, are mixed. If there is a precipitation reaction, write the complete equation that describes the reaction.

OBJECTIVE 12

Solution

Step 1 The possible products from the mixture of $\text{Pb}(\text{NO}_3)_2(aq)$ and $\text{NaC}_2\text{H}_3\text{O}_2(aq)$ are $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$ and NaNO_3 .

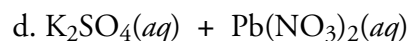
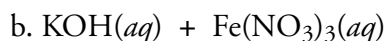
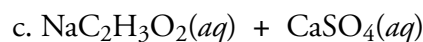
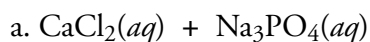


Step 2 According to our solubility guidelines, compounds with nitrates and acetates are soluble, so both $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$ and NaNO_3 are soluble. **There is no precipitation reaction.**

EXERCISE 7.3 - Precipitation Reactions

Predict whether a precipitate will form when each of the following pairs of water solutions is mixed. If there is a precipitation reaction, write the complete equation that describes the reaction.

OBJECTIVE 12



Having Trouble?

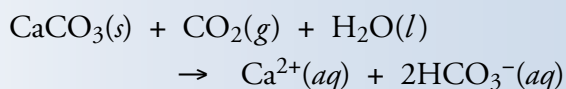
Are you having trouble with the topics in this chapter? People often do. To complete each of the lessons in it successfully, you need to have mastered the skills taught in previous sections. Here is a list of the things you need to know how to do to solve the problems at the end of this chapter. Work through these items in the order presented, and be sure you have mastered each before going on to the next.

- Convert between names and symbols for the common elements. See Table 3.1.
- Identify whether an element is a metal or a nonmetal. See Section 3.3.
- Determine the charges on many of the monatomic ions. See Figure 5.3.
- Convert between the names and formulas for polyatomic ions. See Table 6.3.
- Convert between the names and formulas for ionic compounds. See Section 6.1.
- Balance chemical equations. See Section 7.1.
- Predict the products of double-displacement reactions. See Section 7.3.
- Predict whether ionic compounds are soluble or insoluble in water. See Section 7.3.

SPECIAL TOPIC 7.1 Hard Water and Your Hot Water Pipes

A precipitation reaction that is a slight variation on the one depicted in Figures 7.11 and 7.12 helps explain why a solid scale forms more rapidly in your hot water pipes than in your cold water pipes.

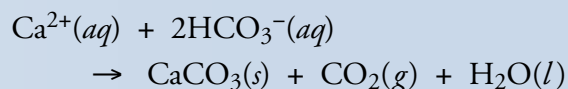
We say water is hard if it contains calcium ions, magnesium ions, and in many cases, iron ions. These ions come from rocks in the ground and dissolve into the water that passes through them. For example, limestone rock is calcium carbonate, $\text{CaCO}_3(s)$, and dolomite rock is a combination of calcium carbonate and magnesium carbonate, written as $\text{CaCO}_3 \cdot \text{MgCO}_3(s)$. Water alone will dissolve very small amounts of these minerals, but carbon dioxide dissolved in water speeds the process.



If the water were removed from the product mixture, calcium hydrogen carbonate,

$\text{Ca}(\text{HCO}_3)_2$, would form, but this compound is much more soluble than calcium carbonate and does not precipitate from our tap water.

When hard water is heated, the reverse of this reaction occurs, and the calcium and hydrogen carbonate ions react to reform solid calcium carbonate.

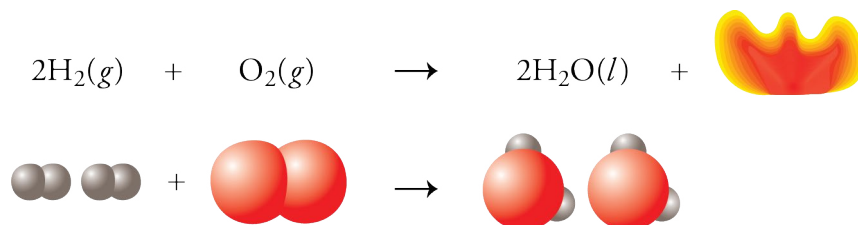


Thus, in your hot water pipes, solid calcium carbonate precipitates from solution and collects as scale on the inside of the pipes. After we have discussed acid-base reactions in Chapter 8, it will be possible to explain how the plumber can remove this obstruction.

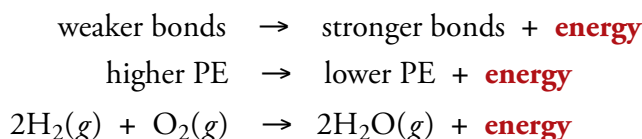


7.4 Chemical Changes and Energy

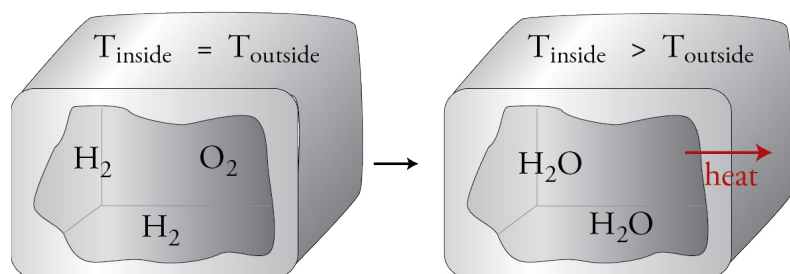
In most chemical reactions, bonds are broken and new ones formed. Section 4.1 showed us that the breaking of bonds requires energy and the formation of bonds releases it. If in a chemical reaction, more energy is released in the formation of new bonds than was necessary to break old bonds, energy is released overall, and the reaction is exergonic. For example, the burning of hydrogen gas is an exergonic process:



On the one hand, energy is required to break the bonds between the hydrogen atoms in the hydrogen molecules, H_2 , and between the oxygen atoms in the oxygen molecules, O_2 . On the other hand, energy is released in the formation of the H–O bonds in water. Because the bonds in the product are more stable, we say that they are stronger than the bonds in the reactants, and the product has lower potential energy than the reactants. Energy is released overall.



To visualize these energy changes at the molecular level, let's picture a container of hydrogen and oxygen that is initially at the same temperature as its surroundings (Figure 7.13). As the reaction proceeds and forms water molecules, some of the potential energy of the system is converted into kinetic energy. (This is similar to the conversion of potential energy to kinetic energy when a coin falls toward the ground.) The higher average kinetic energy of the particles in the product mixture means that the product mixture is at a higher temperature than the initial reactant mixture and therefore at a higher temperature than the surroundings. Thus the higher-temperature products transfer heat to the surroundings. Because the conversion of potential energy



OBJECTIVE 13

Figure 7.13
Exothermic Reaction

Stronger bonds \rightarrow More stable
 \downarrow
 Energy released \leftarrow Lower PE
 \downarrow
 Increases KE_{ave} of product particles
 \downarrow
 Increased $T \rightarrow T_{\text{inside}} > T_{\text{outside}}$
 \downarrow
 Heat transferred to surroundings
 \downarrow
Exothermic

OBJECTIVE 13

OBJECTIVE 13

OBJECTIVE 13

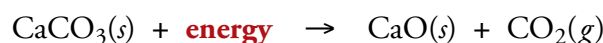
to kinetic energy in the reaction can lead to heat being released from the system, the energy associated with a chemical reaction is often called the *heat of reaction*. A change that leads to heat energy being released from the system to the surroundings is called **exothermic**. (Exergonic means *any form of energy* released, and exothermic means *heat* energy released.)

OBJECTIVE 14

If less energy is released in the formation of the new bonds than is necessary to break the old bonds, energy must be absorbed from the surroundings for the reaction to proceed. The reaction that forms calcium oxide, often called quicklime, from calcium carbonate is an example of this sort of change. (Quicklime has been used as a building material since 1500 BC. It is used today to remove impurities in iron ores. Calcium oxide is also used in air pollution control and water treatment.) In the industrial production of quicklime, this reaction is run at over 2000 °C to provide enough energy to convert the calcium carbonate into calcium oxide and carbon dioxide.



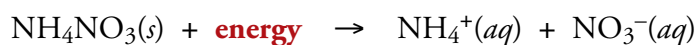
OBJECTIVE 14



OBJECTIVE 14

Collectively, the bonds formed in the products are weaker and, therefore, less stable than those of the reactant, so the products have higher potential energy than the reactants. The high temperature is necessary to provide energy for the change to the greater-potential-energy products. Because energy is absorbed in the reaction, it is endergonic.

Endergonic changes can lead to a transfer of heat *from* the surroundings. Cold packs used to quickly cool a sprained ankle are an example of this kind of change. One kind of cold pack contains a small pouch of ammonium nitrate, NH_4NO_3 , inside a larger pouch of water. When the cold pack is twisted, ammonium nitrate is released into the water, and as it dissolves, the water cools (Figure 7.14).

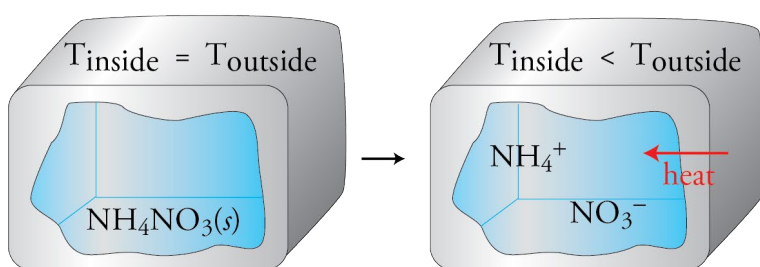


OBJECTIVE 14

The attractions between the particles in the final mixture are less stable and have higher potential energy than the attractions in the separate ammonium nitrate solid and liquid water. The energy necessary to increase the potential energy of the system comes from some of the kinetic energy of the moving particles, so the particles in the final mixture have a lower average kinetic energy and a lower temperature than the

Figure 7.14
Endothermic Reaction

OBJECTIVE 14



Weaker bonds \rightarrow Less stable
 \downarrow
 Energy absorbed \leftarrow Higher PE
 \downarrow
 Decreases KE_{ave} of product particles
 \downarrow
 Decreased T $\rightarrow T_{\text{inside}} < T_{\text{outside}}$
 \downarrow
 Heat transferred to system \rightarrow **Endothermic**

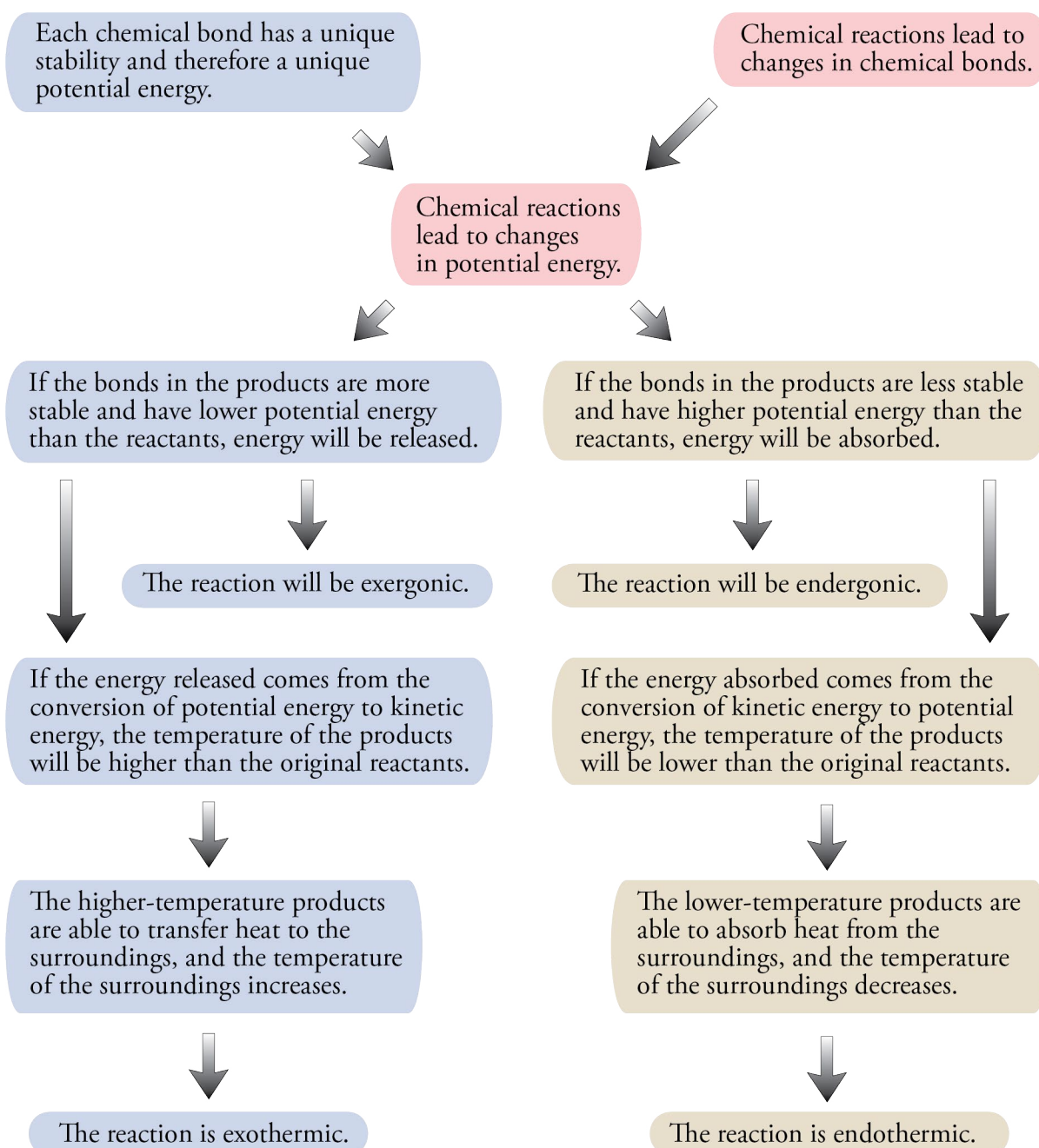
original substances. Heat is transferred from the higher-temperature sprained ankle to the lower-temperature cold pack. A change, such as the one in the cold pack, which leads a system to absorb heat energy from the surroundings, is called an **endothermic** change. (Endergonic means *any form of energy* absorbed, and endothermic means *heat* energy absorbed.)

Figure 7.15 shows the logic sequence that summarizes why chemical reactions either release or absorb energy. It also shows why, in many cases, the change leads to a transfer of heat either to or from the surroundings, with a corresponding increase or decrease in the temperature of the surroundings.

OBJECTIVE 14

OBJECTIVE 15

Figure 7.15
Energy and Chemical
Reactions



Chapter
Glossary

Chemical reaction or chemical change The conversion of one or more pure substances into one or more different pure substances.

Reactants The substances that change in a chemical reaction. Their formulas are on the left side of the arrow in a chemical equation.

Products The substances that form in a chemical reaction. Their formulas are on the right side of the arrow in a chemical equation.

Coefficients The numbers in front of chemical formulas in a balanced chemical equation.

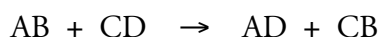
Solution A mixture whose particles are so evenly distributed that the relative concentrations of the components are the same throughout. Solutions can also be called homogeneous mixtures.

Aqueous solution A solution in which water is the solvent.

Solute The gas in a solution of a gas in a liquid. The solid in a solution of a solid in a liquid. The minor component in other solutions.

Solvent The liquid in a solution of a gas in a liquid. The liquid in a solution of a solid in a liquid. The major component in other solutions.

Double-displacement reaction A chemical reaction that has the following form:



Precipitation reaction A reaction in which one of the products is insoluble in water and comes out of solution as a solid.

Precipitate A solid that comes out of solution.

Precipitation The process of forming a solid in a solution.

Crystals Solid particles whose component atoms, ions, or molecules are arranged in an organized, repeating pattern.

Complete ionic equation A chemical equation that describes the actual form for each substance in solution. For example, ionic compounds that are dissolved in water are described as separate ions.

Spectator ions Ions that play a role in delivering other ions into solution to react but that do not actively participate in the reaction themselves.

Complete equation or molecular equation A chemical equation that includes uncharged formulas for all of the reactants and products. The formulas include the spectator ions, if any.

Net ionic equation A chemical equation for which the spectator ions have been eliminated, leaving only the substances actively involved in the reaction.

Solubility The maximum amount of solute that can be dissolved in a given amount of solvent.

Exothermic change Change that leads to heat energy being released from the system to the surroundings.

Endothermic change Change that leads the system to absorb heat energy from the surroundings.

You can test yourself on the glossary terms at the textbook's Web site.

The goal of this chapter is to teach you to do the following.

1. Define all of the terms in the Chapter Glossary.

Section 7.1 Chemical Reactions and Chemical Equations

2. Describe the information given by a chemical equation.
3. Write the symbols used in chemical equations to describe solid, liquid, gas, and aqueous.
4. Balance chemical equations.

Section 7.2 Liquid Water and Water Solutions

5. Describe attractions between H_2O molecules.
6. Describe the structure of liquid water.
7. Describe the process for dissolving an ionic compound in water. Your description should include mention of the nature of the particles in solution and the attractions between the particles in the solution.
8. Given a description of a solution, identify the solute and the solvent.

Section 7.3 Precipitation Reactions

9. Describe double-displacement reactions.
10. Describe precipitation reactions. Your descriptions should include mention of the nature of the particles in the system before and after the reaction, a description of the cause of the reaction, and a description of the attractions between the particles before and after the reaction.
11. Given the formula for an ionic compound, predict whether it is soluble in water or not.
12. Given formulas for two ionic compounds;
 - a. Predict whether a precipitate will form when the water solutions of the two are mixed,
 - b. If there is a reaction, predict the products of the reaction and write their formulas,
 - c. If there is a reaction, write the complete equation that describes the reaction.

Section 7.4 Energy and Chemical Reactions

13. Explain why some chemical reactions release heat to their surroundings.
14. Explain why some chemical reactions absorb heat from their surroundings.
15. Explain why chemical reactions either absorb or release energy.

Chapter Objectives

1. Write the formulas for all of the diatomic elements.
2. Write the definitions of energy, kinetic energy, and potential energy.
3. Describe the relationship between stability, capacity to do work, and potential energy.
4. Explain why energy must be absorbed to break a chemical bond.
5. Explain why energy is released when a chemical bond is formed.
6. Describe the changes that take place during heat transfer between objects at different temperatures.

Review Questions

7. Predict whether atoms of each of the following pairs of elements would be expected to form ionic or covalent bonds.
- | | |
|-------------|-------------|
| a. Mg and F | c. Fe and O |
| b. O and H | d. N and Cl |
8. Write formulas that correspond to the following names.
- | | |
|------------|------------|
| a. ammonia | c. propane |
| b. methane | d. water |
9. Write formulas that correspond to the following names.
- | | |
|------------------------|-----------------------|
| a. nitrogen dioxide | c. dibromine monoxide |
| b. carbon tetrabromide | d. nitrogen monoxide |
10. Write formulas that correspond to the following names.
- | | |
|-----------------------|------------------------------|
| a. lithium fluoride | d. sodium carbonate |
| b. lead(II) hydroxide | e. chromium(III) chloride |
| c. potassium oxide | f. sodium hydrogen phosphate |

Key Ideas

Complete the following statements by writing one of these words or phrases in each blank.

above	liquid
absorbed	major
charges	minor
chemical bonds	negative
coefficients	none
complete formula	organized, repeating
constantly breaking	partial charges
continuous	positive
converted into	precipitate
created	precipitates
δ^-	precipitation
δ^+	released
delta, Δ	same average distance apart
destroyed	same proportions
equal to	separate ions
forming new attractions	shorthand description
from	solute
gas	solvent
heat	strongly
homogeneous mixture	subscripts
left out	very low

11. A chemical change or chemical reaction is a process in which one or more pure substances are _____ one or more different pure substances.
12. In chemical reactions, atoms are rearranged and regrouped through the breaking and making of _____.
13. A chemical equation is a(n) _____ of a chemical reaction.
14. If special conditions are necessary for a reaction to take place, they are often specified _____ the arrow in the reaction's chemical equation.

15. To indicate that a chemical reaction requires the _____ addition of heat in order to proceed, we place an upper-case Greek _____ above the arrow in the reaction's chemical equation.
16. In chemical reactions, atoms are neither _____ nor _____; they merely change partners. Thus the number of atoms of an element in the reaction's products is _____ the number of atoms of that element in the original reactants. The _____ we often place in front of one or more of the formulas in a chemical equation reflect this fact.
17. When balancing chemical equations, we do not change the _____ in the formulas.
18. Because oxygen atoms attract electrons much more _____ than do hydrogen atoms, the O–H covalent bond is very polar, leading to a relatively large partial negative charge on the oxygen atom (represented by a(n) _____) and a relatively large partial positive charge on the hydrogen atom (represented by a(n) _____).
19. As in other liquids, the attractions between water molecules are strong enough to keep them the _____ but weak enough to allow each molecule to be _____ the attractions that momentarily connect it to some molecules and _____ to other molecules.
20. A solution, also called a(n) _____, is a mixture whose particles are so evenly distributed that the relative concentrations of the components are the same throughout.
21. Every part of a water solution of an ionic compound has the _____ of water molecules and ions as every other part.
22. When an ionic compound dissolves in water, the ions that escape the solid are held in solution by attractions between their own _____ and the _____ of the polar water molecules. The _____ oxygen ends of the water molecules surround the cations, and the _____ hydrogen ends surround the anions.
23. In solutions of solids dissolved in liquids, we call the solid the _____ and the liquid the _____.
24. In solutions of gases in liquids, we call the _____ the solute and the _____ the solvent.
25. In solutions of two liquids, we call the _____ component the solute and the _____ component the solvent.
26. Sometimes a double-displacement reaction has one product that is insoluble in water. As that product forms, it emerges, or _____, from the solution as a solid. This process is called _____, and the solid is called a(n) _____.
27. Crystals are solid particles whose component atoms, ions, or molecules are arranged in a(n) _____ pattern.
28. In a complete ionic equation, which describes the forms taken by the various substances in solution, the ionic compounds dissolved in the water are described as _____, and the insoluble ionic compound is described with a(n) _____.
29. Because spectator ions are not involved in the reaction, they are often _____ of the chemical equation.

30. When we say an ionic solid is insoluble in water, we do not mean that _____ of the solid dissolves. Thus, when we say that calcium carbonate is insoluble in water, what we really mean is that the solubility is _____.
31. If in a chemical reaction, more energy is released in the formation of new bonds than was necessary to break old bonds, energy is _____ overall, and the reaction is exergonic.
32. A change that leads to _____ energy being released from the system to the surroundings is called exothermic.
33. If less energy is released in the formation of the new bonds than is necessary to break the old bonds, energy must be _____ from the surroundings for the reaction to proceed.
34. Endergonic changes can lead to a transfer of heat _____ the surroundings.

Chapter Problems

OBJECTIVES 2

OBJECTIVES 3

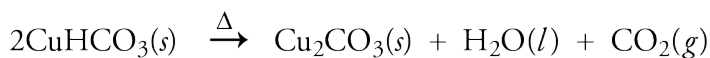
OBJECTIVES 2

OBJECTIVES 3

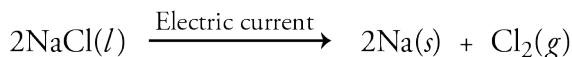
OBJECTIVE 4

Section 7.1 Chemical Reactions and Chemical Equations

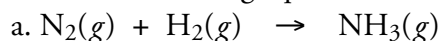
35. Describe the information given in the following chemical equation.



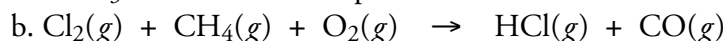
36. Describe the information given in the following chemical equation.



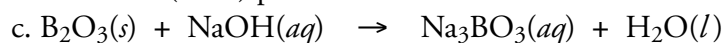
37. Balance the following equations.



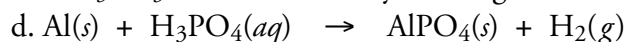
NH_3 is used to make explosives and rocket fuel.



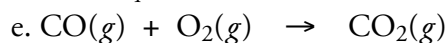
HCl is used to make vinyl chloride, which is then used to make polyvinyl chloride (PVC) plastic.



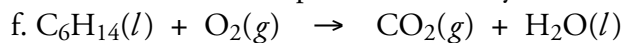
Na_3BO_3 is used as an analytical reagent.



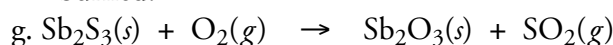
AlPO_4 is used in dental cements.



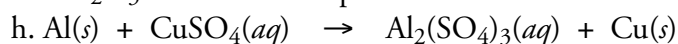
This reaction takes place in the catalytic converter of your car.



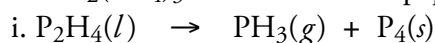
This is one of the chemical reactions that take place when gasoline is burned.



Sb_2O_3 is used to flameproof cloth.

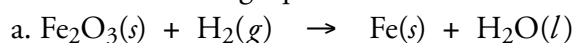


$\text{Al}_2(\text{SO}_4)_3$ has been used in paper production.



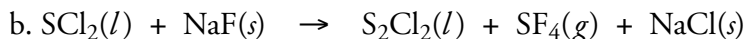
PH_3 is used to make semiconductors.

38. Balance the following equations.

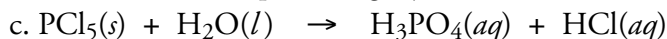


Fe is the primary component in steel.

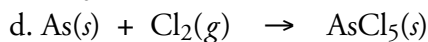
OBJECTIVE 4



S_2Cl_2 is used to purify sugar juices.



H_3PO_4 is used to make fertilizers, soaps, and detergents.



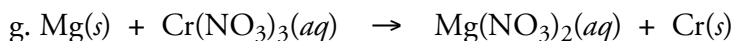
AsCl_5 is an intermediate in the production of arsenic compounds.



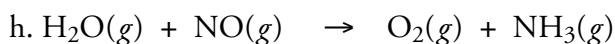
$\text{C}_2\text{H}_5\text{SH}$ is added to natural gas to give it an odor.



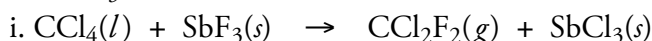
NO_2 is used in rocket fuels.



Cr is used to make stainless steel.

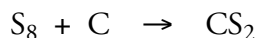
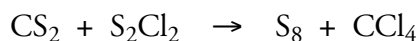
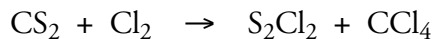


NH_3 is used to make fertilizers.



CCl_2F_2 is a chlorofluorocarbon called CFC-12. Although it has had many uses in the past, its use has diminished greatly due to the damage it can do to our protective ozone layer.

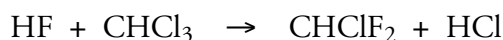
39. Because of its toxicity, carbon tetrachloride is prohibited in products intended for home use, but it is used industrially for a variety of purposes, including the production of chlorofluorocarbons (CFCs). It is made in a three-step process. Balance their equations:



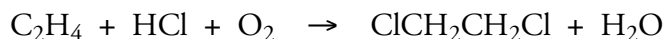
40. Chlorofluorocarbons (CFCs) are compounds that contain carbon, fluorine, and chlorine. Because they destroy ozone that forms a protective shield high in earth's atmosphere, their use is being phased out, but at one time they were widely employed as aerosol propellants and as refrigerants. Balance the following equation that shows how the CFCs dichlorodifluoromethane, CCl_2F_2 , and trichlorofluoromethane, CCl_3F , are produced.



41. Hydrochlorofluorocarbons (HCFCs), which contain hydrogen as well as carbon, fluorine, and chlorine, are less damaging to the ozone layer than chlorofluorocarbons (CFCs). HCFCs are therefore used instead of CFCs for many purposes. Balance the following equation that shows how the HCFC chlorodifluoromethane, CHClF_2 , is made.



42. The primary use of 1,2-dichloroethane, $\text{ClCH}_2\text{CH}_2\text{Cl}$, is to make vinyl chloride, which is then converted into polyvinyl chloride (PVC) for many purposes, including plastic pipes. Balance the following equation, which describes the industrial reaction for producing 1,2-dichloroethane.



OBJECTIVE 4

OBJECTIVE 4

OBJECTIVE 4

OBJECTIVE 4

Section 7.2 Liquid Water and Water Solutions

OBJECTIVE 5

43. What are the particles that form the basic structure of water? Describe the attraction that holds these particles together. Draw a rough sketch that shows the attraction between two water molecules.

OBJECTIVE 6

44. Describe the structure of liquid water.

OBJECTIVE 7

45. Describe the process for dissolving the ionic compound lithium iodide, LiI, in water, including the nature of the particles in solution and the attractions between the particles in the solution.

OBJECTIVE 7

46. Describe the process for dissolving the ionic compound potassium nitrate, KNO_3 , in water, including the nature of the particles in solution and the attractions between the particles in the solution.

OBJECTIVE 7

47. Describe the process for dissolving the ionic compound sodium sulfate, Na_2SO_4 , in water. Include mention of the nature of the particles in solution and the attractions between the particles in the solution.

OBJECTIVE 7

48. Describe the process for dissolving the ionic compound calcium chloride, CaCl_2 , in water. Include mention of the nature of the particles in solution and the attractions between the particles in the solution.

OBJECTIVE 8

49. Solid camphor and liquid ethanol mix to form a solution. Which of these substances is the solute and which is the solvent?

OBJECTIVE 8

50. Gaseous propane and liquid diethyl ether mix to form a solution. Which of these substances is the solute and which is the solvent?

OBJECTIVE 8

51. Consider a solution of 10% liquid acetone and 90% liquid chloroform. Which of these substances is the solute and which is the solvent?

Section 7.3 Precipitation Reactions

OBJECTIVE 10

52. Black and white photographic film has a thin layer of silver bromide deposited on it. Wherever light strikes the film, silver ions are converted to uncharged silver atoms, creating a dark image on the film. Describe the precipitation reaction that takes place between water solutions of silver nitrate, $\text{AgNO}_3(aq)$, and sodium bromide, $\text{NaBr}(aq)$, to form solid silver bromide, $\text{AgBr}(s)$, and aqueous sodium nitrate, $\text{NaNO}_3(aq)$. Include mention of the nature of the particles in the system before and after the reaction, a description of the cause of the reaction, and a description of the attractions between the particles before and after the reaction.



OBJECTIVE 10

53. Magnesium carbonate is used as an anti-caking agent in powders and as an antacid. Describe the precipitation reaction that takes place between water solutions of magnesium nitrate, $\text{Mg}(\text{NO}_3)_2(aq)$, and sodium carbonate, $\text{Na}_2\text{CO}_3(aq)$, to form solid magnesium carbonate, $\text{MgCO}_3(s)$, and aqueous sodium nitrate, $\text{NaNO}_3(aq)$. Include mention of the nature of the particles in the system before and after the reaction, a description of the cause of the reaction, and a description of the attractions between the particles before and after the reaction.

OBJECTIVE 11

54. Predict whether each of the following substances is soluble or insoluble in water.

- Na_2SO_3 (used in water treatment)
- iron(III) acetate (a wood preservative)
- CoCO_3 (a red pigment)
- lead(II) chloride (used in the preparation of lead salts)

55. Predict whether each of the following substances is soluble or insoluble in water.
- MgSO_4 (fireproofing)
 - barium sulfate (used in paints)
 - $\text{Bi}(\text{OH})_3$ (used in plutonium separation)
 - ammonium sulfite (used in medicine and photography)

OBJECTIVE 11

56. Predict whether each of the following substances is soluble or insoluble in water.
- zinc phosphate (used in dental cements)
 - $\text{Mn}(\text{C}_2\text{H}_3\text{O}_2)_2$ (used in cloth dyeing)
 - nickel(II) sulfate (used in nickel plating)
 - AgCl (used in silver plating)

OBJECTIVE 11



OBJECTIVE 11

57. Predict whether each of the following substances is soluble or insoluble in water.
- copper(II) chloride (used in fireworks and as a fungicide)
 - PbSO_4 (a paint pigment)
 - potassium hydroxide (used in soap manufacture)
 - NH_4F (used as an antiseptic in brewing)

58. For each of the following pairs of formulas, predict whether the substances they represent would react in a precipitation reaction. The products formed in the reactions that take place are used in ceramics, cloud seeding, photography, electroplating, and paper coatings. If there is no reaction, write, "No Reaction." If there is a reaction, write the complete equation for the reaction.

OBJECTIVE 12

- $\text{Co}(\text{NO}_3)_2(aq) + \text{Na}_2\text{CO}_3(aq)$
- $\text{KI}(aq) + \text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2(aq)$
- $\text{CuSO}_4(aq) + \text{LiNO}_3(aq)$
- $\text{Ni}(\text{NO}_3)_2(aq) + \text{Na}_3\text{PO}_4(aq)$
- $\text{K}_2\text{SO}_4(aq) + \text{Ba}(\text{NO}_3)_2(aq)$

59. For each of the following pairs of formulas, predict whether the substances they represent would react in a precipitation reaction. If there is no reaction, write, "No Reaction." If there is a reaction, write the complete equation for the reaction.

OBJECTIVE 12

- $\text{NaCl}(aq) + \text{Pb}(\text{NO}_3)_2(aq)$
- $\text{NH}_4\text{Cl}(aq) + \text{CaSO}_3(aq)$
- $\text{NaOH}(aq) + \text{Zn}(\text{NO}_3)_2(aq)$
- $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2(aq) + \text{Na}_2\text{SO}_4(aq)$

60. Phosphate ions find their way into our water system from the fertilizers dissolved in the runoff from agricultural fields and from detergents that we send down our drains. Some of these phosphate ions can be removed by adding aluminum sulfate to the water and precipitating the phosphate ions as aluminum phosphate. Write the net ionic equation for the reaction that forms the aluminum phosphate.

OBJECTIVE 12

61. The taste of drinking water can be improved by removing impurities from our municipal water by adding substances to the water that precipitate a solid (called a flocculent) that drags down impurities as it settles. One way this is done is to dissolve aluminum sulfate and sodium hydroxide in the water to precipitate aluminum hydroxide. Write the complete equation for this reaction.

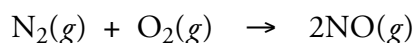
OBJECTIVE 12

OBJECTIVE 12 62. Cadmium hydroxide is used in storage batteries. It is made from the precipitation reaction of cadmium acetate and sodium hydroxide. Write the complete equation for this reaction.

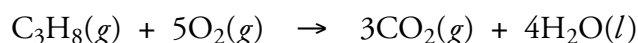
OBJECTIVE 12 63. Chromium(III) phosphate is a paint pigment that is made in a precipitation reaction between water solutions of chromium(III) chloride and sodium phosphate. Write the complete equation for this reaction.

Section 7.4 Chemical Changes and Energy

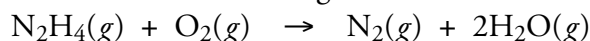
OBJECTIVE 13 64. Consider the following endergonic reaction. In general terms, explain why energy is absorbed in the process of this reaction.



OBJECTIVE 14 65. The combustion of propane is an exergonic reaction. In general terms, explain why energy is released in the process of this reaction.

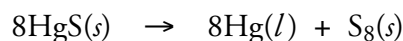


66. Hydrazine, N_2H_4 , is used as rocket fuel. Consider a system in which a sample of hydrazine is burned in a closed container, followed by heat transfer from the container to the surroundings.



- In general terms, explain why energy is released in the reaction.
- Before heat energy is transferred to the surroundings, describe the average internal kinetic energy of the product particles compared to the reactant particles. If the product's average internal kinetic energy is higher than for the reactants, from where did this energy come? If the average internal kinetic energy is lower than for the reactants, to where did this energy go?
- Describe the changes in particle motion as heat is transferred from the products to the surroundings.

OBJECTIVE 14 67. Cinnabar is a natural mercury(II) sulfide, HgS , found near volcanic rocks and hot springs. It is the only important source of mercury, which has many uses, including dental amalgams, thermometers, and mercury vapor lamps. Mercury is formed in the following endothermic reaction when mercury(II) sulfide is heated. Consider a system in which a sample of HgS in a closed container is heated with a Bunsen burner flame.



- Describe the changes in particle motion as heat is transferred from the hot gases of the Bunsen burner flame to the container to the $\text{HgS}(s)$.
- In general terms, explain why energy is absorbed in the reaction.
- Into what form of energy is the heat energy converted for this reaction?

OBJECTIVE 13 68. Even on a hot day, you get cold when you step out of a swimming pool. Suggest a reason why your skin cools as water evaporates from it. Is evaporation an exothermic or endothermic process?

69. Classify each of the following changes as exothermic or endothermic.

- Leaves decaying in a compost heap.
- Dry ice (solid carbon dioxide) changing to carbon dioxide gas.
- Dew forming on a lawn at night.

70. Classify each of the following changes as exothermic or endothermic.
- The nuclear reaction that takes place in a nuclear electrical generating plant.
 - Cooking an egg in boiling water.
 - The breakdown of plastic in the hot sun.
71. Explain why chemical reactions either absorb or evolve energy.

OBJECTIVE 15

Additional Problems

72. Balance the following chemical equations.

OBJECTIVE 4

- $\text{SiCl}_4 + \text{H}_2\text{O} \rightarrow \text{SiO}_2 + \text{HCl}$
- $\text{H}_3\text{BO}_3 \rightarrow \text{B}_2\text{O}_3 + \text{H}_2\text{O}$
- $\text{I}_2 + \text{Cl}_2 \rightarrow \text{ICl}_3$
- $\text{Al}_2\text{O}_3 + \text{C} \rightarrow \text{Al} + \text{CO}_2$

73. Balance the following chemical equations.

OBJECTIVE 4

- $\text{HClO}_4 + \text{Fe}(\text{OH})_2 \rightarrow \text{Fe}(\text{ClO}_4)_2 + \text{H}_2\text{O}$
- $\text{NaClO}_3 \rightarrow \text{NaCl} + \text{O}_2$
- $\text{Sb} + \text{Cl}_2 \rightarrow \text{SbCl}_3$
- $\text{CaCN}_2 + \text{H}_2\text{O} \rightarrow \text{CaCO}_3 + \text{NH}_3$

74. Balance the following chemical equations.

OBJECTIVE 4

- $\text{NH}_3 + \text{Cl}_2 \rightarrow \text{N}_2\text{H}_4 + \text{NH}_4\text{Cl}$
- $\text{Cu} + \text{AgNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{Ag}$
- $\text{Sb}_2\text{S}_3 + \text{HNO}_3 \rightarrow \text{Sb}(\text{NO}_3)_3 + \text{H}_2\text{S}$
- $\text{Al}_2\text{O}_3 + \text{Cl}_2 + \text{C} \rightarrow \text{AlCl}_3 + \text{CO}$

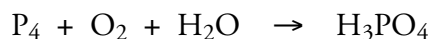
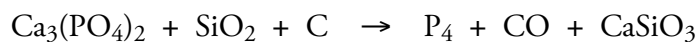
75. Balance the following chemical equations.

OBJECTIVE 4

- $\text{AsH}_3 \rightarrow \text{As} + \text{H}_2$
- $\text{H}_2\text{S} + \text{Cl}_2 \rightarrow \text{S}_8 + \text{HCl}$
- $\text{Co} + \text{O}_2 \rightarrow \text{Co}_2\text{O}_3$
- $\text{Na}_2\text{CO}_3 + \text{C} \rightarrow \text{Na} + \text{CO}$

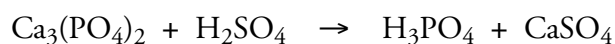
76. Phosphoric acid, H_3PO_4 , is an important chemical used to make fertilizers, detergents, pharmaceuticals, and many other substances. High purity phosphoric acid is made in a two-step process called the furnace process. Balance its two equations:

OBJECTIVE 4



77. For most applications, phosphoric acid is produced by the “wet process”. Balance the following equation that describes the reaction for this process.

OBJECTIVE 4



78. Predict whether each of the following substances is soluble or insoluble in water.

OBJECTIVE 11

- manganese(II) chloride (used as a dietary supplement)
- CdSO_4 (used in pigments)
- copper(II) carbonate (used in fireworks)
- $\text{Co}(\text{OH})_3$ (used as a catalyst)

- OBJECTIVE 11** 79. Predict whether each of the following substances is soluble or insoluble in water.
- copper(II) hydroxide (used as a pigment)
 - BaBr_2 (used to make photographic compounds)
 - silver carbonate (used as a laboratory reagent)
 - $\text{Pb}_3(\text{PO}_4)_2$ (used as a stabilizing agent in plastics)

- OBJECTIVE 12** 80. For each of the following pairs of formulas, predict whether the substances they represent would react to yield a precipitate. (The products formed in the reactions that take place are used to coat steel, as a fire-proofing filler for plastics, in cosmetics, and as a topical antiseptic.) If there is no reaction, write, "No Reaction". If there is a reaction, write the complete equation for the reaction.
- $\text{NaCl}(aq) + \text{Al}(\text{NO}_3)_3(aq)$
 - $\text{Ni}(\text{NO}_3)_2(aq) + \text{NaOH}(aq)$
 - $\text{MnCl}_2(aq) + \text{Na}_3\text{PO}_4(aq)$
 - $\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2(aq) + \text{Na}_2\text{CO}_3(aq)$

- OBJECTIVE 12** 81. For each of the following pairs of formulas, predict whether the substances they represent would react to yield a precipitate. (The products formed in the reactions that take place are used as a catalyst, as a tanning agent, as a pigment, in fertilizers, as a food additive, and on photographic film.) If there is no reaction, write, "No Reaction". If there is a reaction, write the complete equation for the reaction.
- $\text{KOH}(aq) + \text{Cr}(\text{NO}_3)_3(aq)$
 - $\text{Fe}(\text{NO}_3)_3(aq) + \text{K}_3\text{PO}_4(aq)$
 - $\text{NaBr}(aq) + \text{AgNO}_3(aq)$
 - $\text{Mg}(\text{C}_2\text{H}_3\text{O}_2)_2(aq) + \text{NaCl}(aq)$

Before working Chapter Problems 82 through 98, you might want to review the procedures for writing chemical formulas that are described in Chapter 6. Remember that some elements are described with formulas containing subscripts (as in O_2).

- OBJECTIVE 4** 82. Hydrochloric acid is used in the cleaning of metals (called pickling). Hydrogen chloride, used to make hydrochloric acid, is made industrially by combining hydrogen and chlorine. Write a balanced equation, without including states, for this reaction.

- OBJECTIVE 4** 83. Potassium hydroxide has many uses, including the manufacture of soap. It is made by running an electric current through a water solution of potassium chloride. Both potassium chloride and water are reactants, and the products are potassium hydroxide, hydrogen, and chlorine. Write a balanced equation, without including states, for this reaction.

- OBJECTIVE 4** 84. Aluminum sulfate, commonly called alum, is used to coat paper made from wood pulp. (It fills in tiny holes in the paper and thus keeps the ink from running.) Alum is made in the reaction of aluminum oxide with sulfuric acid, H_2SO_4 , which produces aluminum sulfate and water. Write a balanced equation, without including states, for this reaction.

- OBJECTIVE 4** 85. Under the right conditions, methanol reacts with oxygen to yield formaldehyde, CH_2O , and water. Most of the formaldehyde made in this way is used in the production of other substances, including some important plastics. Formaldehyde, itself now a suspected carcinogen, was once used in insulation foams and in plywood adhesives. Write a balanced equation, without including states, for the reaction that produces formaldehyde from methanol.

86. Hydrogen fluoride is used to make chlorofluorocarbons (CFCs) and in uranium processing. Calcium fluoride reacts with sulfuric acid, H_2SO_4 , to form hydrogen fluoride and calcium sulfate. Write a balanced equation, without including states, for this reaction. OBJECTIVE 4
87. Sodium sulfate, which is used to make detergents and glass, is one product of the reaction of sodium chloride, sulfur dioxide, water, and oxygen. The other product is hydrogen chloride. Write a balanced equation, without including states, for this reaction. OBJECTIVE 4
88. Sodium hydroxide, which is often called caustic soda, is used to make paper, soaps, and detergents. For many years, it was made from the reaction of sodium carbonate with calcium hydroxide (also called slaked lime). The products are sodium hydroxide and calcium carbonate. Write a balanced equation, without including states, for this reaction. OBJECTIVE 4
89. In the modern process for making sodium hydroxide, an electric current is run through a sodium chloride solution, forming hydrogen and chlorine along with the sodium hydroxide. Both sodium chloride and water are reactants. Write a balanced equation, without including states, for this reaction. OBJECTIVE 4
90. For over a century, sodium carbonate (often called soda ash) was made industrially by the Solvay process. This process was designed by Ernest Solvay in 1864 and was used in the United States until extensive natural sources of sodium carbonate were found in the 1970s and 1980s. Write a balanced equation, without including states, for each step in the process: OBJECTIVE 4
- Calcium carbonate (from limestone) is heated and decomposes into calcium oxide and carbon dioxide.
 - Calcium oxide reacts with water to form calcium hydroxide.
 - Ammonia reacts with water to form ammonium hydroxide.
 - Ammonium hydroxide reacts with carbon dioxide to form ammonium hydrogen carbonate.
 - Ammonium hydrogen carbonate reacts with sodium chloride to form sodium hydrogen carbonate and ammonium chloride.
 - Sodium hydrogen carbonate is heated and decomposes into sodium carbonate, carbon dioxide, and water.
 - Ammonium chloride reacts with calcium hydroxide to form ammonia, calcium chloride, and water.
91. All of the equations for the Solvay process described in problem 90 can be summarized by a single equation, called a net equation, that describes the overall change for the process. This equation shows calcium carbonate reacting with sodium chloride to form sodium carbonate and calcium chloride. Write a balanced equation, without including states, for this net reaction. OBJECTIVE 4
92. Nitric acid, HNO_3 , which is used to make fertilizers and explosives, is made industrially in the three steps described below. Write a balanced equation, without including states, for each of these steps. OBJECTIVE 4
- Ammonia reacts with oxygen to form nitrogen monoxide and water.
 - Nitrogen monoxide reacts with oxygen to form nitrogen dioxide.
 - Nitrogen dioxide reacts with water to form nitric acid and nitrogen monoxide.

OBJECTIVE 4 93. All of the equations for the production of nitric acid described in problem 92 can be summarized in a single equation, called a net equation, that describes the overall change for the complete process. This equation shows ammonia combining with oxygen to yield nitric acid and water. Write a balanced equation, without including states, for this net reaction.

OBJECTIVE 4 94. Ammonium sulfate, an important component in fertilizers, is made from the reaction of ammonia and sulfuric acid, H_2SO_4 . Write a balanced equation, without including states, for the reaction that summarizes this transformation.

OBJECTIVE 4 95. Hydrogen gas has many practical uses, including the conversion of vegetable oils into margarine. One way the gas is produced by the chemical industry is by reacting propane gas with gaseous water to form carbon dioxide gas and hydrogen gas. Write a balanced equation for this reaction, showing the states of reactants and products.

OBJECTIVE 4 96. Sodium tripolyphosphate, $\text{Na}_5\text{P}_3\text{O}_{10}$, is called a “builder” when added to detergents. It helps to create the conditions in laundry water necessary for the detergents to work most efficiently. When phosphoric acid, H_3PO_4 , is combined with sodium carbonate, three reactions take place in the mixture that lead to the production of sodium tripolyphosphate. Write a balanced equation, without including states, for each of the reactions:

- Phosphoric acid reacts with sodium carbonate to form sodium dihydrogen phosphate, water, and carbon dioxide.
- Phosphoric acid also reacts with sodium carbonate to form sodium hydrogen phosphate, water, and carbon dioxide.
- Sodium dihydrogen phosphate combines with sodium hydrogen phosphate to yield sodium tripolyphosphate and water.

OBJECTIVE 4 97. Pig iron is iron with about 4.3% carbon in it. The carbon lowers the metal’s melting point and makes it easier to shape. To produce pig iron, iron(III) oxide is combined with carbon and oxygen at high temperature. Three changes then take place to form molten iron with carbon dispersed in it. Write a balanced equation, without including states, for each of these changes:

- Carbon combines with oxygen to form carbon monoxide.
- Iron(III) oxide combines with the carbon monoxide to form iron and carbon dioxide.
- Carbon monoxide changes into carbon (in the molten iron) and carbon dioxide.

OBJECTIVE 4 98. The United States chemical industry makes more sulfuric acid than any other chemical. One process uses hydrogen sulfide from “sour” natural gas wells or petroleum refineries as a raw material. Write a balanced equation, without including states, for each of the steps leading from hydrogen sulfide to sulfuric acid:

- Hydrogen sulfide (which could be called dihydrogen monosulfide) combines with oxygen to form sulfur dioxide and water.
- The sulfur dioxide reacts with more hydrogen sulfide to form sulfur and water. (Sulfur is described as both S and S_8 in chemical equations. Use S in this equation).

- c. After impurities are removed, the sulfur is reacted with oxygen to form sulfur dioxide.
- d. Sulfur dioxide reacts with oxygen to yield sulfur trioxide.
- e. Sulfur trioxide and water combine to make sulfuric acid, H_2SO_4 .
99. Assume you are given a water solution that contains either sodium ions or aluminum ions. Describe how you could determine which of these is in solution
100. Assume that you are given a water solution that contains either nitrate ions or phosphate ions. Describe how you could determine which of these is in solution.
101. Write a complete, balanced chemical equation for the reaction between water solutions of iron(III) chloride and silver nitrate. OBJECTIVE 4
102. Write a complete, balanced chemical equation for the reaction between water solutions of sodium phosphate and copper(II) chloride. OBJECTIVE 4
103. When the solid amino acid methionine, $\text{C}_5\text{H}_{11}\text{NSO}_2$, reacts with oxygen gas, the products are carbon dioxide gas, liquid water, sulfur dioxide gas, and nitrogen gas. Write a complete, balanced equation for this reaction. OBJECTIVE 4
104. When the explosive liquid nitroglycerin, $\text{C}_3\text{H}_5\text{N}_3\text{O}_9$, decomposes, it forms carbon dioxide gas, nitrogen gas, water vapor, and oxygen gas. Write a complete, balanced equation for this reaction. OBJECTIVE 4
105. Classify each of the following changes as exothermic or endothermic.
- a. Fuel burning in a camp stove
- b. The melting of ice in a camp stove to provide water on a snow-camping trip
106. Classify each of the following changes as exothermic or endothermic.
- a. Fireworks exploding
- b. Water evaporating on a rain-drenched street
107. Why does your body temperature rise when you exercise?

Discussion Problems

108. The solubility of calcium carbonate is 0.0014 g of CaCO_3 per 100 mL of water at 25 °C, and the solubility of sodium nitrate is 92.1 grams of NaNO_3 per 100 mL of water at 25 °C. We say that calcium carbonate is insoluble in water, and sodium nitrate, NaNO_3 , is soluble.
- a. In Section 6.3, the following statement was made: “*When we say an ionic solid is insoluble in water, we do not mean that none of the solid dissolves. There are always some ions that can escape from the surface of an ionic solid in water and go into solution.*” Discuss the process by which calcium and carbonate ions can escape from the surface of calcium carbonate solid in water and go into solution. You might want to draw a picture to illustrate this process.
- b. If the calcium and carbonate ions are constantly going into solution, why doesn't the calcium carbonate solid all dissolve?
- c. Why do you think sodium nitrate, NaNO_3 , dissolves to a much greater degree than calcium carbonate?
- d. Why is there a limit to the solubility of even the “soluble” sodium nitrate?
109. Only a fraction of the energy released in the combustion of gasoline is converted into kinetic energy of the moving car. Where does the rest go?