CHAPTER 15

Acids, Bases, and Salts

Citrus fruits are all acidic substances.

Chapter Outline

Acids are important chemicals. They are used in cooking to produce the surprise of tartness (from lemons) and to release CO_2 bubbles from leavening agents in baking. Vitamin C is an acid that is an essential nutrient in our diet. Our stomachs release acid to aid in digestion. Excess stomach acid can produce heartburn and indigestion. Bacteria in our mouths produce acids that can dissolve tooth enamel to form cavities. In our recreational activities we are concerned about acidity levels in swimming pools and spas. Acids are essential in the manufacture of detergents, plastics, and storage batteries. The acid–base properties of substances are found in all areas of our lives. In this chapter we consider the properties of acids, bases, and salts.

15.1 Acids and Bases

The word *acid* is derived from the Latin *acidus,* meaning "sour" or "tart," and is also related to the Latin word *acetum,* meaning "vinegar." Vinegar has been around since antiquity as a product of the fermentation of wine and apple cider. The sour constituent of vinegar is acetic acid ($HC_2H_3O_2$). Characteristic properties commonly associated with acids include the following:

1. sour taste
2. the ability to change the color of litmus, a vegetable dye, from blue to red
3. the ability to react with
   - metals such as zinc and magnesium to produce hydrogen gas
   - hydroxide bases to produce water and an ionic compound (salt)
   - carbonates to produce carbon dioxide

These properties are due to the hydrogen ions, H^+, released by acids in a water solution.

Classically, a *base* is a substance capable of liberating hydroxide ions, OH^-, in water solution. Hydroxides of the alkali metals (Group 1A) and alkaline earth metals (Group 2A), such as $LiOH$, $NaOH$, KOH, $Ca(OH)_2$, and $Ba(OH)_2$, are the most common inorganic bases. Water solutions of bases are called *alkaline solutions* or *basic solutions.* Some of the characteristic properties commonly associated with bases include the following:

1. bitter or caustic taste
2. a slippery, soapy feeling
3. the ability to change litmus from red to blue
4. the ability to interact with acids

Several theories have been proposed to answer the question "What is an acid and what is a base?" One of the earliest, most significant of these theories was advanced in 1884 by Svante Arrhenius (1859–1927), a Swedish scientist, who stated that "an acid is a hydrogen-containing substance that dissociates to produce hydrogen ions, and a base is a hydroxide-containing substance that dissociates to produce hydroxide ions in aqueous solutions." Arrhenius postulated that the hydrogen ions are produced by the dissociation of acids in water and that the hydroxide ions are produced by the dissociation of bases in water:

$$\underset{\text{acid}}{HA} \xrightarrow{H_2O} H^+(aq) + A^-(aq)$$

$$\underset{\text{base}}{MOH} \xrightarrow{H_2O} M^+(aq) + OH^-(aq)$$

An Arrhenius acid solution contains an excess of H^+ ions.
An Arrhenius base solution contains an excess of OH^- ions.

In 1923, the Brønsted–Lowry proton transfer theory was introduced by J. N. Brønsted (1897–1947), a Danish chemist, and T. M. Lowry (1847–1936), an English chemist. This theory states that an acid is a proton donor and a base is a proton acceptor.

A Brønsted–Lowry acid is a proton (H^+) donor.
A Brønsted–Lowry base is a proton (H^+) acceptor.

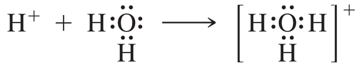
Consider the reaction of hydrogen chloride gas with water to form hydrochloric acid:

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq) \qquad (1)$$

In the course of the reaction, HCl donates, or gives up, a proton to form a Cl^- ion, and H_2O accepts a proton to form the H_3O^+ ion. Thus, HCl is an acid and H_2O is a base, according to the Brønsted–Lowry theory.

hydronium ion

A hydrogen ion (H^+) is nothing more than a bare proton and does not exist by itself in an aqueous solution. In water H^+ combines with a polar water molecule to form a hydrated hydrogen ion (H_3O^+) commonly called a **hydronium ion.** The H^+ is attracted to a polar water molecule, forming a bond with one of the two pairs of unshared electrons:



$$H^+ + H:\overset{\cdot\cdot}{\underset{\overset{\cdot\cdot}{H}}{O}}: \longrightarrow \left[H:\overset{\cdot\cdot}{\underset{\overset{\cdot\cdot}{H}}{O}}:H\right]^+$$

hydronium ion

Note the electron structure of the hydronium ion. For simplicity we often use H^+ instead of H_3O^+ in equations, with the explicit understanding that H^+ is always hydrated in solution.

When a Brønsted–Lowry acid donates a proton, as illustrated in the equation below, it forms the conjugate base of that acid. When a base accepts a proton, it forms the conjugate acid of that base. A conjugate acid and base are produced as products. The formulas of a conjugate acid–base pair differ by one proton (H^+). Consider what happens when HCl(g) is bubbled through water, as shown by this equation:

conjugate acid–base pair

$$\underset{acid}{HCl(g)} + \underset{base}{H_2O(l)} \longrightarrow \underset{base}{Cl^-(aq)} + \underset{acid}{H_3O^+(aq)}$$

conjugate acid–base pair

The conjugate acid–base pairs are $HCl—Cl^-$ and $H_3O^+—H_2O$. The conjugate base of HCl is Cl^-, and the conjugate acid of Cl^- is HCl. The conjugate base of H_3O^+ is H_2O, and the conjugate acid of H_2O is H_3O^+.

Another example of conjugate acid–base pairs can be seen in this equation:

$$\underset{acid}{NH_4^+} + \underset{base}{H_2O} \longrightarrow \underset{acid}{H_3O^+} + \underset{base}{NH_3}$$

Here the conjugate acid–base pairs are $NH_4^+—NH_3$ and $H_3O^+—H_2O$.

Example 15.1 Write the formula for (a) the conjugate base of H_2O and of HNO_3, and (b) the conjugate acid of SO_4^{2-} and of $C_2H_3O_2^-$.

SOLUTION (a) To write the conjugate base of an acid, remove one proton from the acid formula:

$H_2O \xrightarrow{-H^+} OH^-$ (conjugate base)

$HNO_3 \xrightarrow{-H^+} NO_3^-$ (conjugate base)

Note that, by removing an H^+, the conjugate base becomes more negative than the acid by one minus charge.

(b) To write the conjugate acid of a base, add one proton to the formula of the base:

$SO_4^{2-} \xrightarrow{+H^+} HSO_4^-$ (conjugate acid)

$C_2H_3O_2^- \xrightarrow{+H^+} HC_2H_3O_2$ (conjugate acid)

In each case the conjugate acid becomes more positive than the base by a +1 charge due to the addition of H^+.

Remember: The difference between an acid or a base and its conjugate is one proton, H^+.

Practice 15.1

Indicate the conjugate base for these acids:

(a) H_2CO_3 (b) HNO_2 (c) $HC_2H_3O_2$

Practice 15.2

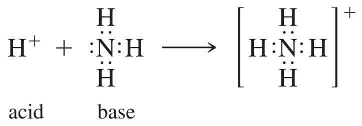
Indicate the conjugate acid for these bases:

(a) HSO_4^- (b) NH_3 (c) OH^-

A more general concept of acids and bases was introduced by Gilbert N. Lewis. The Lewis theory deals with the way in which a substance with an unshared pair of electrons reacts in an acid–base type of reaction. According to this theory a base is any substance that has an unshared pair of electrons (electron pair donor), and an acid is any substance that will attach itself to or accept a pair of electrons.

A Lewis acid is an electron pair acceptor.
A Lewis base is an electron pair donor.

In the reaction

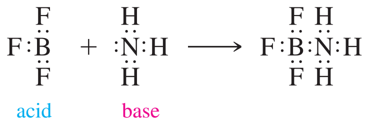


$$H^+ + :\underset{\underset{\displaystyle H}{\cdot\cdot}}{\overset{\overset{\displaystyle H}{\cdot\cdot}}{N}}:H \longrightarrow \left[H:\underset{\underset{\displaystyle H}{\cdot\cdot}}{\overset{\overset{\displaystyle H}{\cdot\cdot}}{N}}:H \right]^+$$

Table 15.1 Summary of Acid–Base Definitions

Theory	Acid	Base
Arrhenius	A hydrogen-containing substance that produces hydrogen ions in aqueous solution	A hydroxide-containing substance that produces hydroxide ions in aqueous solution
Brønsted–Lowry	A proton (H^+) donor	A proton (H^+) acceptor
Lewis	Any species that will bond to an unshared pair of electrons (electron pair acceptor)	Any species that has an unshared pair of electrons (electron pair donor)

The H^+ is a Lewis acid and $:NH_3$ is a Lewis base. According to the Lewis theory, substances other than proton donors (e.g., BF_3) behave as acids:


$$BF_3 + :NH_3 \longrightarrow F_3B:NH_3$$

These three theories, which explain how acid–base reactions occur, are summarized in Table 15.1. We will generally use the theory that best explains the reaction under consideration. Most of our examples will refer to aqueous solutions. Note that in an aqueous acidic solution the H^+ ion concentration is always greater than OH^- ion concentration. And vice versa—in an aqueous basic solution the OH^- ion concentration is always greater than the H^+ ion concentration. When the H^+ and OH^- ion concentrations in a solution are equal, the solution is *neutral*; that is, it is neither acidic nor basic.

15.2 Reactions of Acids

In aqueous solutions, the H^+ or H_3O^+ ions are responsible for the characteristic reactions of acids. The following reactions are in an aqueous medium:

Reaction with Metals Acids react with metals that lie above hydrogen in the activity series of elements to produce hydrogen and an ionic compound (salt) (see Section 17.5):

$$\text{acid} + \text{metal} \longrightarrow \text{hydrogen} + \text{ionic compound}$$

$$2\,HCl(aq) + Ca(s) \longrightarrow H_2(g) + CaCl_2(aq)$$

$$H_2SO_4(aq) + Mg(s) \longrightarrow H_2(g) + MgSO_4(aq)$$

$$6\,HC_2H_3O_2(aq) + 2\,Al(s) \longrightarrow 3\,H_2(g) + 2\,Al(C_2H_3O_2)_3(aq)$$

Acids such as nitric acid (HNO_3) are oxidizing substances (see Chapter 17) and react with metals to produce water instead of hydrogen. For example,

$$3\,Zn(s) + 8\,HNO_3(\text{dilute}) \longrightarrow 3\,Zn(NO_3)_2(aq) + 2\,NO(g) + 4\,H_2O(l)$$

Reaction with Bases The interaction of an acid and a base is called a *neutralization reaction.* In aqueous solutions, the products of this reaction are a salt and water:

acid + base ⟶ salt + water

$HBr(aq) + KOH(aq) \longrightarrow KBr(aq) + H_2O(l)$

$2\,HNO_3(aq) + Ca(OH)_2(aq) \longrightarrow Ca(NO_3)_2(aq) + 2\,H_2O(l)$

$2\,H_3PO_4(aq) + 3\,Ba(OH)_2(aq) \longrightarrow Ba_3(PO_4)_2(s) + 6\,H_2O(l)$

Reaction with Metal Oxides This reaction is closely related to that of an acid with a base. With an aqueous acid solution, the products are a salt and water:

acid + metal oxide ⟶ salt + water

$2\,HCl(aq) + Na_2O(s) \longrightarrow 2\,NaCl(aq) + H_2O(l)$

$H_2SO_4(aq) + MgO(s) \longrightarrow MgSO_4(aq) + H_2O(l)$

$6\,HCl(aq) + Fe_2O_3(s) \longrightarrow 2\,FeCl_3(aq) + 3\,H_2O(l)$

Reaction with Carbonates Many acids react with carbonates to produce carbon dioxide, water, and an ionic compound:

Carbonic acid (H_2CO_3) is not the product because it is unstable and spontaneously decomposes into water and carbon dioxide.

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

acid + carbonate ⟶ salt + water + carbon dioxide

$2\,HCl(aq) + Na_2CO_3(aq) \longrightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$

$H_2SO_4(aq) + MgCO_3(s) \longrightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$

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

15.3 Reactions of Bases

The OH^- ions are responsible for the characteristic reactions of bases. The following reactions are in an aqueous medium:

Reaction with Acids Bases react with acids to produce a salt and water. See reaction of acids with bases in Section 15.2.

amphoteric

Amphoteric Hydroxides Hydroxides of certain metals, such as zinc, aluminum, and chromium, are **amphoteric;** that is, they are capable of reacting as either an acid or a base. When treated with a strong acid, they behave like bases; when reacted with a strong base, they behave like acids:

Strong acids and bases are discussed in Section 15.7.

$Zn(OH)_2(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + 2\,H_2O(l)$

$Zn(OH)_2(s) + 2\,NaOH(aq) \longrightarrow Na_2Zn(OH)_4(aq)$

Reaction of NaOH and KOH with Certain Metals Some amphoteric metals react directly with the strong bases sodium hydroxide and potassium hydroxide to produce hydrogen:

base + metal + water ⟶ salt + hydrogen

$2\,NaOH(aq) + Zn(s) + 2\,H_2O(l) \longrightarrow Na_2Zn(OH)_4(aq) + H_2(g)$

$2\,KOH(aq) + 2\,Al(s) + 6\,H_2O(l) \longrightarrow 2\,KAl(OH)_4(aq) + 3\,H_2(g)$

Practice 15.3

Write the formulas and names of the acid and base from which these salts are formed.

(a) K_3PO_4 (b) $MgBr_2$ (c) LiCl (d) $FeCO_3$

CHEMISTRY IN ACTION • A Cool Fizz

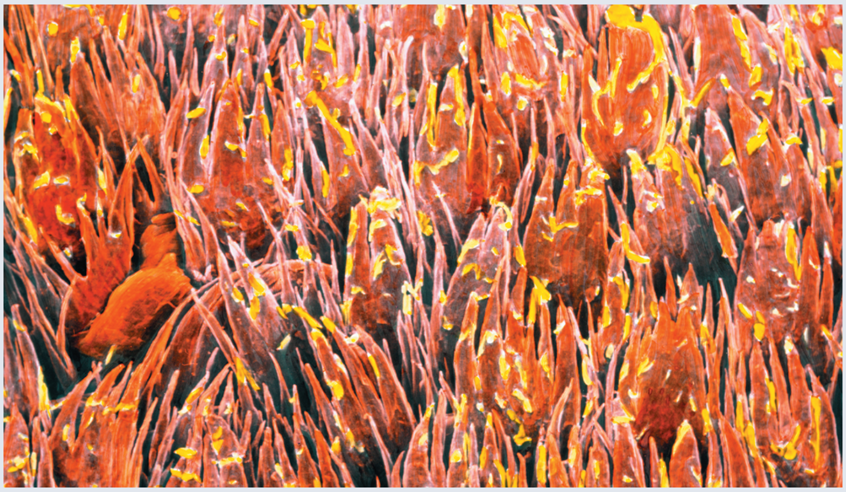
Try this experiment with your friends: Chill some soft drinks well. Next, pour the soft drink into a glass. Stick your tongue into the liquid and time how long you can keep it there.

What causes the tingling we feel in our mouth (or on our tongue)? Many people believe it is the bubbles of carbon dioxide, but scientists have found that is not the answer. The tingling is caused by "chemisthesis." Bruce Bryant at Monell Chemical Senses Center in Philadelphia says that chemisthesis is a chemically induced sensation that does not involve taste or odor receptors. The tongue-tingling response to the soft drink is caused by production of protons (H^+ ions) released when an enzyme (carbonic anhydrase) acts on CO_2. The H^+ ions acidify nerve endings, producing the sensation of tingling.

Carbon dioxide also stimulates other neurons when the drink is cold. This means that at a constant pressure of CO_2, a cold drink will produce more tingling than a room-temperature drink. If the drink is at room temperature, CO_2 will increase the cool feeling. Chilling a soda increases the effect of the protons on the nerve endings. At the same time, the high concentration of CO_2 in a freshly opened soda makes the cold soda feel even colder. Together these chemical effects make it really painful to keep your tongue in the chilled soda.

Chemisthesis is important for survival, as well as for having fun with friends.

Bryant says, "It tells you that something is chemically impinging on your body, that tissue is in imminent danger. Burrowing animals can sense toxic levels of carbon dioxide and even feel a sting when exposed to those levels."



Tongue papillae contain nerve endings that send temperature and tactile information to the brain.

15.4 Salts

Salts are very abundant in nature. Most of the rocks and minerals of Earth's mantle are salts of one kind or another. Huge quantities of dissolved salts also exist in the oceans. Salts can be considered compounds derived from acids and bases. They consist of positive metal or ammonium ions combined with negative nonmetal ions (OH^- and O^{2-} excluded). The positive ion is the base counterpart and the nonmetal ion is the acid counterpart:

$$\overset{\text{base}}{\text{NaOH}} \qquad \overset{\text{acid}}{\text{HCl}} \;\longrightarrow\; \underset{\text{salt}}{\text{NaCl}}$$

Chemists use the terms *ionic compound* and *salt* interchangeably.

Salts are usually crystalline and have high melting and boiling points.

From a single acid such as hydrochloric acid (HCl), we can produce many chloride compounds by replacing the hydrogen with metal ions (e.g., NaCl, KCl, RbCl, $CaCl_2$, $NiCl_2$). Hence the number of known salts greatly exceeds the number of known acids and bases. If the hydrogen atoms of a binary acid are replaced by a nonmetal, the resulting compound has covalent bonding and is therefore not considered to be ionic (e.g., PCl_3, S_2Cl_2, Cl_2O, NCl_3, ICl).

You may want to review Chapter 6 for nomenclature of acids, bases, and salts.

15.5 Electrolytes and Nonelectrolytes

We can show that solutions of certain substances are conductors of electricity with a simple conductivity apparatus, which consists of a pair of electrodes connected to a voltage source through a light bulb and switch (see Figure 15.1). If the medium between the electrodes is a conductor of electricity, the light bulb will glow when the switch is closed. When chemically pure water is placed in the beaker and the switch is closed, the light does not glow, indicating that water is a virtual nonconductor. When we dissolve a small amount of sugar in the water and test the solution, the light still does not glow, showing that a sugar solution is also a nonconductor. But when a small amount of salt, NaCl, is dissolved in water and this solution is tested, the light glows brightly. Thus the salt solution conducts electricity. A fundamental difference exists between the chemical bonding in sugar and that in salt. Sugar is a covalently bonded (molecular) substance; common salt is a substance with ionic bonds.

electrolyte
nonelectrolyte

Substances whose aqueous solutions are conductors of electricity are called **electrolytes.** Substances whose solutions are nonconductors are known as **nonelectrolytes.** The classes of compounds that are electrolytes are acids, bases, and other ionic compounds (salts). Solutions of certain oxides also are conductors because the oxides form an acid or a base when dissolved in water. One major difference between electrolytes and nonelectrolytes is that electrolytes are capable of producing ions in solution, whereas nonelectrolytes do not have this property. Solutions that contain a sufficient number of ions will conduct an electric current. Although pure water is essentially a nonconductor, many city water supplies contain enough dissolved ionic matter to cause the light to glow dimly when the water is tested in a conductivity apparatus. Table 15.2 lists some common electrolytes and nonelectrolytes.

Acids, bases, and salts are electrolytes.

(a) (b) (c)

Figure 15.1
A conductivity apparatus shows the difference in conductivity of solutions. (a) Distilled water does not conduct electricity. (b) Sugar water is a nonelectrolyte. (c) Salt water is a strong electrolyte and conducts electricity.

These strange mineral formations called "tufa" exist at Mono Lake, California. Tufa is formed by water bubbling through sand saturated with NaCl, Na_2CO_3, and Na_2SO_4.

Table 15.2 Representative Electrolytes and Nonelectrolytes

Electrolytes		Nonelectrolytes	
H_2SO_4	$HC_2H_3O_2$	$C_{12}H_{22}O_{11}$ (sugar)	CH_3OH (methyl alcohol)
HCl	NH_3	C_2H_5OH (ethyl alcohol)	$CO(NH_2)_2$ (urea)
HNO_3	K_2SO_4	$C_2H_4(OH)_2$ (ethylene glycol)	O_2
NaOH	$NaNO_3$	$C_3H_5(OH)_3$ (glycerol)	H_2O

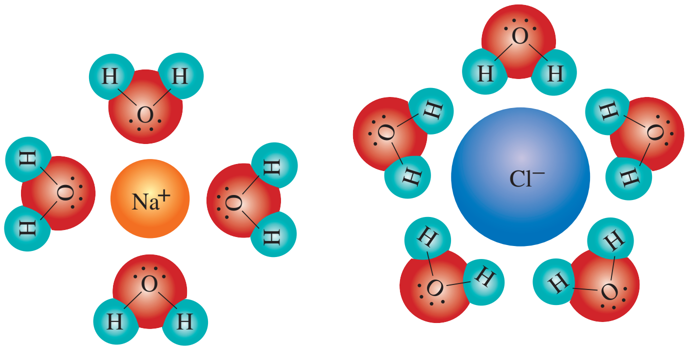
15.6 Dissociation and Ionization of Electrolytes

Arrhenius received the 1903 Nobel Prize in chemistry for his work on electrolytes. He found that a solution conducts electricity because the solute dissociates immediately upon dissolving into electrically charged particles (ions). The movement of these ions toward oppositely charged electrodes causes the solution to be a conductor. According to his theory, solutions that are relatively poor conductors contain electrolytes that are only partly dissociated. Arrhenius also believed that ions exist in solution whether or not an electric current is present. In other words, the electric current does not cause the formation of ions. Remember that positive ions are cations; negative ions are anions.

dissociation

We have seen that sodium chloride crystals consist of sodium and chloride ions held together by ionic bonds. **Dissociation** is the process by which the ions of a salt separate as the salt dissolves. When placed in water, the sodium and chloride ions are attracted by the polar water molecules, which surround each ion as it dissolves. In water, the salt dissociates, forming hydrated sodium

Figure 15.2
Hydrated sodium and chloride ions. When sodium chloride dissolves in water, each Na^+ and Cl^- ion becomes surrounded by water molecules. The negative end of the water dipole is attracted to the Na^+ ion, and the positive end is attracted to the Cl^- ion.



and chloride ions (see Figure 15.2). The sodium and chloride ions in solution are surrounded by a specific number of water molecules and have less attraction for each other than they had in the crystalline state. The equation representing this dissociation is

$$NaCl(s) + (x + y)\,H_2O \longrightarrow Na^+(H_2O)_x + Cl^-(H_2O)_y$$

A simplified dissociation equation in which the water is omitted but understood to be present is

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$$

Remember that sodium chloride exists in an aqueous solution as hydrated ions, not as NaCl units, even though the formula NaCl (or $Na^+ + Cl^-$) is often used in equations.

The chemical reactions of salts in solution are the reactions of their ions. For example, when sodium chloride and silver nitrate react and form a precipitate of silver chloride, only the Ag^+ and Cl^- ions participate in the reaction. The Na^+ and NO_3^- remain as ions in solution:

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

ionization

Ionization is the formation of ions; it occurs as a result of a chemical reaction of certain substances with water. Glacial acetic acid (100% $HC_2H_3O_2$) is a liquid that behaves as a nonelectrolyte when tested by the method described in Section 15.5. But a water solution of acetic acid conducts an electric current (as indicated by the dull-glowing light of the conductivity apparatus). The equation for the reaction with water, which forms hydronium and acetate ions, is

$$\underset{\text{acid}}{HC_2H_3O_2} + \underset{\text{base}}{H_2O} \rightleftharpoons \underset{\text{acid}}{H_3O^+} + \underset{\text{base}}{C_2H_3O_2^-}$$

or, in the simplified equation,

$$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$$

In this ionization reaction, water serves not only as a solvent but also as a base according to the Brønsted–Lowry theory.

Hydrogen chloride is predominantly covalently bonded, but when dissolved in water, it reacts to form hydronium and chloride ions:

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

When a hydrogen chloride solution is tested for conductivity, the light glows brilliantly, indicating many ions in the solution.

Ionization occurs in each of the preceding two reactions with water, producing ions in solution. The necessity for water in the ionization process can be demonstrated by dissolving hydrogen chloride in a nonpolar solvent such as hexane and testing the solution for conductivity. The solution fails to conduct electricity, indicating that no ions are produced.

The terms *dissociation* and *ionization* are often used interchangeably to describe processes taking place in water. But, strictly speaking, the two are different. In the dissociation of a salt, the salt already exists as ions; when it dissolves in water, the ions separate, or dissociate, and increase in mobility. In the ionization process, ions are produced by the reaction of a compound with water.

15.7 Strong and Weak Electrolytes

Electrolytes are classified as strong or weak depending on the degree, or extent, of dissociation or ionization. **Strong electrolytes** are essentially 100% ionized in solution; **weak electrolytes** are much less ionized (based on comparing 0.1 *M* solutions). Most electrolytes are either strong or weak, with a few classified as moderately strong or weak. Most salts are strong electrolytes. Acids and bases that are strong electrolytes (highly ionized) are called *strong acids* and *strong bases.* Acids and bases that are weak electrolytes (slightly ionized) are called *weak acids* and *weak bases.*

strong electrolyte
weak electrolyte

For equivalent concentrations, solutions of strong electrolytes contain many more ions than do solutions of weak electrolytes. As a result, solutions of strong electrolytes are better conductors of electricity. Consider two solutions, 1 *M* HCl and 1 *M* $HC_2H_3O_2$. Hydrochloric acid is almost 100% ionized; acetic acid is about 1% ionized. (See Figure 15.3.) Thus, HCl is a strong acid and $HC_2H_3O_2$ is a weak acid. Hydrochloric acid has about 100 times as many hydronium ions in solution as acetic acid, making the HCl solution much more acidic.

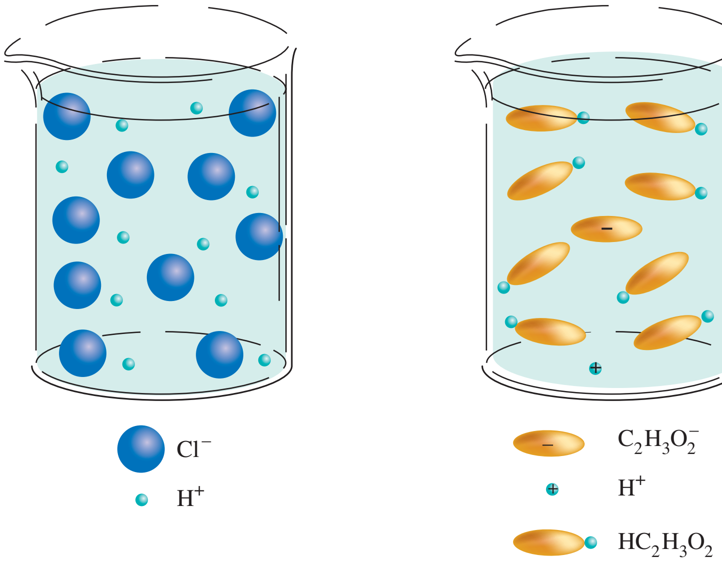
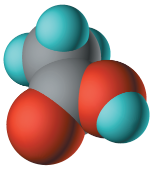
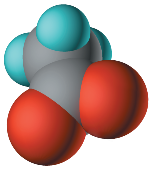


Figure 15.3
HCl solution (left) is 100% ionized, while in $HC_2H_3O_2$ solution (right) almost all of the solute is in molecular form. HCl is a strong acid, while $HC_2H_3O_2$ is a weak acid. Note: The water molecules in the solution are not shown in this figure.



$HC_2H_3O_2$



$C_2H_3O_2^-$

We can distinguish between strong and weak electrolytes experimentally using the apparatus described in Section 15.5. A 1 *M* HCl solution causes the light to glow brilliantly, but a 1 *M* $HC_2H_3O_2$ solution causes only a dim glow. The strong base sodium hydroxide (NaOH) can be distinguished in a similar fashion from the weak base ammonia (NH_3). The ionization of a weak electrolyte in water is represented by an equilibrium equation showing that both the un-ionized and ionized forms are present in solution. In the equilibrium equation of $HC_2H_3O_2$ and its ions, we say that the equilibrium lies "far to the left" because relatively few hydrogen and acetate ions are present in solution:

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

We have previously used a double arrow in an equation to represent reversible processes in the equilibrium between dissolved and undissolved solute in a saturated solution. A double arrow ($\rightleftharpoons$) is also used in the ionization equation of soluble weak electrolytes to indicate that the solution contains a considerable amount of the un-ionized compound in equilibrium with its ions in solution. (See Section 16.1 for a discussion of reversible reactions.) A single arrow is used to indicate that the electrolyte is essentially all in the ionic form in the solution. For example, nitric acid is a strong acid; nitrous acid is a weak acid. Their ionization equations in water may be indicated as

$$HNO_3(aq) \xrightarrow{H_2O} H^+(aq) + NO_3^-(aq)$$

$$HNO_2(aq) \overset{H_2O}{\rightleftharpoons} H^+(aq) + NO_2^-(aq)$$

Practically all soluble salts, acids (such as sulfuric, nitric, and hydrochloric acids), and bases (such as sodium, potassium, calcium, and barium hydroxides) are strong electrolytes. Weak electrolytes include numerous other acids and bases such as acetic acid, nitrous acid, carbonic acid, and ammonia. The terms *strong acid, strong base, weak acid,* and *weak base* refer to whether an acid or base is a strong or weak electrolyte. A brief list of strong and weak electrolytes is given in Table 15.3.

Electrolytes yield two or more ions per formula unit upon dissociation—the actual number being dependent on the compound. Dissociation is complete or nearly complete for nearly all soluble ionic compounds and for certain other strong electrolytes, such as those given in Table 15.3. The following

Table 15.3 Strong and Weak Electrolytes

Strong electrolytes		Weak electrolytes	
Most soluble salts	$HClO_4$	$HC_2H_3O_2$	$H_2C_2O_4$
H_2SO_4	NaOH	H_2CO_3	H_3BO_3
HNO_3	KOH	HNO_2	HClO
HCl	$Ca(OH)_2$	H_2SO_3	NH_3
HBr	$Ba(OH)_2$	H_2S	HF

are dissociation equations for several strong electrolytes. In all cases, the ions are actually hydrated:

$$\mathrm{NaOH} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+}(aq) + \mathrm{OH^-}(aq) \qquad \text{2 ions in solution per formula unit}$$

$$\mathrm{Na_2SO_4} \xrightarrow{\mathrm{H_2O}} 2\,\mathrm{Na^+}(aq) + \mathrm{SO_4^{2-}}(aq) \qquad \text{3 ions in solution per formula unit}$$

$$\mathrm{Fe_2(SO_4)_3} \xrightarrow{\mathrm{H_2O}} 2\,\mathrm{Fe^{3+}}(aq) + 3\,\mathrm{SO_4^{2-}}(aq) \qquad \text{5 ions in solution per formula unit}$$

One mole of NaCl will give 1 mol of Na^+ ions and 1 mol of Cl^- ions in solution, assuming complete dissociation of the salt. One mole of $CaCl_2$ will give 1 mol of Ca^{2+} ions and 2 mol of Cl^- ions in solution:

$$\underset{\text{1 mol}}{\mathrm{NaCl}} \xrightarrow{\mathrm{H_2O}} \underset{\text{1 mol}}{\mathrm{Na^+}(aq)} + \underset{\text{1 mol}}{\mathrm{Cl^-}(aq)}$$

$$\underset{\text{1 mol}}{\mathrm{CaCl_2}} \xrightarrow{\mathrm{H_2O}} \underset{\text{1 mol}}{\mathrm{Ca^{2+}}(aq)} + \underset{\text{2 mol}}{2\,\mathrm{Cl^-}(aq)}$$

Example 15.2

What is the molarity of each ion in a solution of (a) 2.0 *M* NaCl and (b) 0.40 *M* K_2SO_4? Assume complete dissociation.

SOLUTION

(a) According to the dissociation equation,

$$\underset{\text{1 mol}}{\mathrm{NaCl}} \xrightarrow{\mathrm{H_2O}} \underset{\text{1 mol}}{\mathrm{Na^+}(aq)} + \underset{\text{1 mol}}{\mathrm{Cl^-}(aq)}$$

the concentration of Na^+ is equal to that of NaCl: 1 mol NaCl $\longrightarrow$ 1 mol Na^+ and the concentration of Cl^- is also equal to that of NaCl. Therefore the concentrations of the ions in 2.0 *M* NaCl are 2.0 *M* Na^+ and 2.0 *M* Cl^-.

(b) According to the dissociation equation

$$\underset{\text{1 mol}}{\mathrm{K_2SO_4}} \xrightarrow{\mathrm{H_2O}} \underset{\text{2 mol}}{2\,\mathrm{K^+}(aq)} + \underset{\text{1 mol}}{\mathrm{SO_4^{2-}}(aq)}$$

the concentration of K^+ is twice that of K_2SO_4 and the concentration of SO_4^{2-} is equal to that of K_2SO_4. Therefore the concentrations of the ions in 0.40 *M* K_2SO_4 are 0.80 *M* K^+ and 0.40 *M* SO_4^{2-}.

Practice 15.4

What is the molarity of each ion in a solution of (a) 0.050 *M* $MgCl_2$ and (b) 0.070 *M* $AlCl_3$?

Colligative Properties of Electrolyte Solutions

We have learned that when 1 mol of sucrose, a nonelectrolyte, is dissolved in 1000 g of water, the solution freezes at −1.86°C. When 1 mol NaCl is dissolved in 1000 g of water, the freezing point of the solution is not −1.86°C, as might be expected, but is closer to −3.72°C (−1.86 × 2). The reason for the lower freezing point is that 1 mol NaCl in solution produces 2 mol of particles ($2 \times 6.022 \times 10^{23}$ ions) in solution. Thus the freezing point depression produced by 1 mol NaCl is essentially equivalent to that produced by 2 mol of a nonelectrolyte. An electrolyte such as $CaCl_2$, which yields three ions in water, gives a freezing point depression of about three times that of a nonelectrolyte. These freezing point data provide additional evidence that electrolytes dissociate when dissolved in water. The other colligative properties are similarly affected by substances that yield ions in aqueous solutions.

Practice 15.5

What is the boiling point of a 1.5 *m* solution of KCl(*aq*)?

15.8 Ionization of Water

Pure water is a *very* weak electrolyte, but it does ionize slightly. Two equations commonly used to show how water ionizes are

$$\underset{\text{acid}}{H_2O} + \underset{\text{base}}{H_2O} \rightleftharpoons \underset{\text{acid}}{H_3O^+} + \underset{\text{base}}{OH^-}$$

and

$$H_2O \rightleftharpoons H^+ + OH^-$$

The first equation represents the Brønsted–Lowry concept, with water reacting as both an acid and a base, forming a hydronium ion and a hydroxide ion. The second equation is a simplified version, indicating that water ionizes to give a hydrogen and a hydroxide ion. Actually, the proton (H^+) is hydrated and exists as a hydronium ion. In either case equal molar amounts of acid and base are produced so that water is neutral, having neither H^+ nor OH^- ions in excess. The ionization of water at 25°C produces an H^+ ion concentration of 1.0×10^{-7} mol/L and an OH^- ion concentration of 1.0×10^{-7} mol/L. Square brackets, [ ], are used to indicate that the concentration is in moles per liter. Thus $[H^+]$ means the concentration of H^+ is in moles per liter. These concentrations are usually expressed as

$$[H^+] \text{ or } [H_3O^+] = 1.0 \times 10^{-7}\text{ mol/L}$$
$$[OH^-] = 1.0 \times 10^{-7}\text{ mol/L}$$

These figures mean that about two out of every billion water molecules are ionized. This amount of ionization, small as it is, is a significant factor in the behavior of water in many chemical reactions.

15.9 Introduction to pH

The acidity of an aqueous solution depends on the concentration of hydrogen or hydronium ions. The pH scale of acidity provides a simple, convenient, numerical way to state the acidity of a solution. Values on the pH scale are obtained by mathematical conversion of H^+ ion concentrations to pH by the expression

$$\text{pH} = -\log[H^+]$$

pH

where $[H^+] = H^+$ or H_3O^+ ion concentration in moles per liter. The **pH** is defined as the *negative* logarithm of the H^+ or H_3O^+ concentration in moles per liter:

$$pH = -\log[H^+] = -\log(1 \times 10^{-7}) = -(-7) = 7$$

For example, the pH of pure water at 25°C is 7 and is said to be neutral; that is, it is neither acidic nor basic, because the concentrations of H^+ and OH^- are equal. Solutions that contain more H^+ ions than OH^- ions have pH values less than 7, and solutions that contain less H^+ ions than OH^- ions have pH values greater than 7.

$pH < 7.00$ is an acidic solution
$pH = 7.00$ is a neutral solution
$pH > 7.00$ is a basic solution

When $[H^+] = 1 \times 10^{-5}$ mol/L, pH = 5 (acidic)

When $[H^+] = 1 \times 10^{-9}$ mol/L, pH = 9 (basic)

Instead of saying that the hydrogen ion concentration in the solution is 1×10^{-5} mol/L, it's customary to say that the pH of the solution is 5. The smaller the pH value, the more acidic the solution (see Figure 15.4).

The pH scale, along with its interpretation, is given in Table 15.4, and Table 15.5 lists the pH of some common solutions. Note that a change of only one pH unit means a 10-fold increase or decrease in H^+ ion concentration. For example, a solution with a pH of 3.0 is 10 times more acidic than a solution with a pH of 4.0. A simplified method of determining pH from $[H^+]$ follows:

$[H^+] = 1 \times 10^{-5}$ ← pH = this number (5)
pH = 5

when this number
is exactly 1

$[H^+] = 2 \times 10^{-5}$ ← pH is between this number and
next lower number (4 and 5)
pH = 4.7

when this number
is between 1 and 10

Calculating the pH value for H^+ ion concentrations requires the use of logarithms, which are exponents. The **logarithm** (log) of a number is simply the power to which 10 must be raised to give that number. Thus the log of 100 is 2 $(100 = 10^2)$, and the log of 1000 is 3 $(1000 = 10^3)$. The log of 500 is 2.70, but you can't determine this value easily without a scientific calculator.

logarithm

Help on using calculators is found in Appendix II.

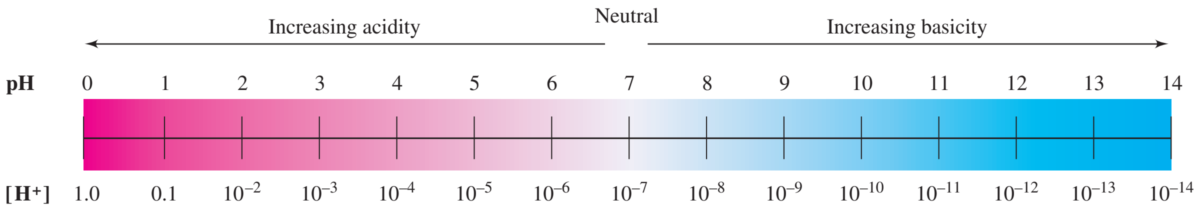


Figure 15.4
The pH scale of acidity and basicity.

Table 15.4 pH Scale for Expressing Acidity

$[H^+]$ (mol/L)	pH	
1×10^{-14}	14	↑
1×10^{-13}	13	
1×10^{-12}	12	Increasing
1×10^{-11}	11	basicity
1×10^{-10}	10	
1×10^{-9}	9	
1×10^{-8}	8	
1×10^{-7}	7	Neutral
1×10^{-6}	6	
1×10^{-5}	5	
1×10^{-4}	4	
1×10^{-3}	3	Increasing
1×10^{-2}	2	acidity
1×10^{-1}	1	
1×10^{0}	0	↓

Table 15.5 The pH of Common Solutions

Solution	pH
Gastric juice	1.0
0.1 *M* HCl	1.0
Lemon juice	2.3
Vinegar	2.8
0.1 *M* $HC_2H_3O_2$	2.9
Orange juice	3.7
Tomato juice	4.1
Coffee, black	5.0
Urine	6.0
Milk	6.6
Pure water (25°C)	7.0
Blood	7.4
Household ammonia	11.0
1 *M* NaOH	14.0

Remember: Change the sign on your calculator since $pH = -\log[H^+]$.

Let's determine the pH of a solution with $[H^+] = 2 \times 10^{-5}$ using a calculator. Enter 2×10^{-5} into your calculator and press the log key. The number $-4.69\ldots$ will be displayed. The pH is then

$$pH = -\log[H^+] = -(-4.69\ldots) = 4.7$$

Next we must determine the correct number of significant figures in the logarithm. The rules for logs are different from those we use in other math operations. The number of decimal places for a log must equal the number of significant figures in the original number. Since 2×10^{-5} has one significant figure, we should round the log to one decimal place $(4.69\ldots) = 4.7$.

Example 15.3 What is the pH of a solution with an $[H^+]$ of (a) 1.0×10^{-11}, (b) 6.0×10^{-4}, and (c) 5.47×10^{-8}?

SOLUTION

(a) $[H^+] = 1.0 \times 10^{-11}$
(2 significant figures)

$$pH = -\log(1.0 \times 10^{-11})$$
$$pH = 11.00$$
(2 decimal places)

(b) $[H^+] = 6.0 \times 10^{-4}$
(2 significant figures)

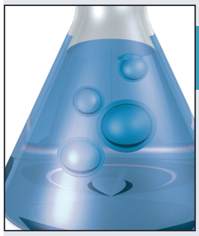
$$\log(6.0 \times 10^{-4}) = -3.22$$
$$pH = -\log[H^+]$$
$$pH = -(-3.22) = 3.22$$
(2 decimal places)

(c) $[H^+] = 5.47 \times 10^{-8}$
(3 significant figures)

$$\log(5.47 \times 10^{-8}) = -7.262$$
$$pH = -\log[H^+]$$
$$pH = -(-7.262) = 7.262$$
(3 decimal places)

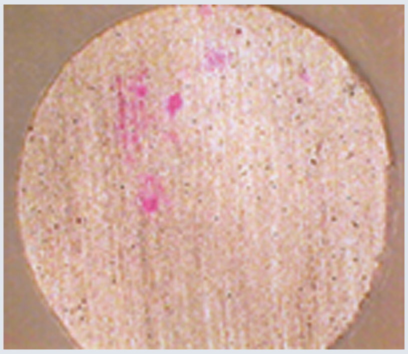
Practice 15.6

What is the pH of a solution with $[H^+]$ of (a) 3.9×10^{-12} *M*, (b) 1.3×10^{-3} *M*, and (c) 3.72×10^{-6} *M*?



CHEMISTRY IN ACTION • In the Pink—A Sign of Corrosion

As our fleet of commercial airplanes ages, scientists are looking for new ways to detect corrosion and stop it *before* it becomes a safety issue. The problem is that planes are subjected to heat, rain, and wind—all factors that can contribute to metal corrosion. How can a maintenance crew detect corrosion on a huge aircraft since many imaging techniques work on only small areas at a time? Imagine a plane that could tell the maintenance crew it needed repair.



Phenolphthalein paint after eight days.

Gerald Frankel and Jim Zhang from the Ohio State University have developed a paint that detects changes in pH. Corrosion of metals (in contact with air and water) results in a chemical reaction that produces hydroxide ions, which increase pH. Frankel and Zhang made a clear acrylic coating that they mixed with phenolphthalein. This acid–base indicator turns bright fuschia when the pH increases above 8.0. The scientists can show the "paint" turns pink at corrosion sites as small as 15 μm deep.

Technicians at Wright-Patterson Air Force Base in Ohio think this could lead to a whole new way to detect corrosion. William Mullins says, "You could walk down the vehicle and see there's a pink spot." This method would be especially good at detecting corrosion concealed around rivets and where metals overlap. The only limitation appears to be that the coating must be clear.

The next time your airline maintenance crew sees pink spots, you may be in for a delay or a change of aircraft!

The measurement and control of pH is extremely important in many fields. Proper soil pH is necessary to grow certain types of plants successfully. The pH of certain foods is too acidic for some diets. Many biological processes are delicately controlled pH systems. The pH of human blood is regulated to very close tolerances through the uptake or release of H^+ by mineral ions, such as HCO_3^-, HPO_4^{2-}, and $H_2PO_4^-$, Changes in the pH of the blood by as little as 0.4 pH unit result in death.

Compounds with colors that change at particular pH values are used as indicators in acid–base reactions. For example, phenolphthalein, an organic compound, is colorless in acid solution and changes to pink at a pH of 8.3. When a solution of sodium hydroxide is added to a hydrochloric acid solution containing phenolphthalein, the change in color (from colorless to pink) indicates that all the acid is neutralized. Commercially available pH test paper contains chemical indicators. The indicator in the paper takes on different colors when wetted with solutions of different pH. Thus the pH of a solution can be estimated by placing a drop on the test paper and comparing the color of the test paper with a color chart calibrated at different pH values. Common applications of pH test indicators are the kits used to measure and adjust the pH of swimming pools, hot tubs, and saltwater aquariums. Electronic pH meters are used for making rapid and precise pH determinations (see Figure 15.5).

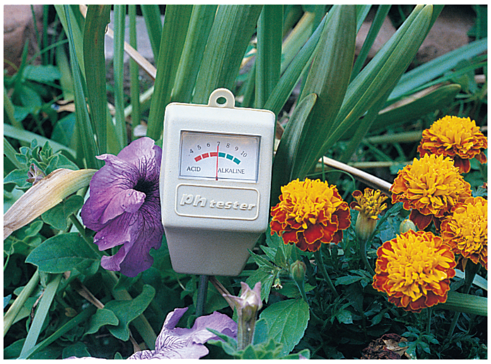


Figure 15.5
The pH of a substance (in this case garden soil) can be measured by using a pH meter.

15.10 Neutralization

neutralization

The reaction of an acid and a base to form a salt and water is known as **neutralization.** We've seen this reaction before, but now with our knowledge about ions and ionization, let's reexamine the process of neutralization.

Consider the reaction that occurs when solutions of sodium hydroxide and hydrochloric acid are mixed. The ions present initially are Na^+ and OH^- from the base and H^+ and Cl^- from the acid. The products, sodium chloride and water, exist as Na^+ and Cl^- ions and H_2O molecules. A chemical equation representing this reaction is

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

This equation, however, does not show that HCl, NaOH, and NaCl exist as ions in solution. The following total ionic equation gives a better representation of the reaction:

$$(\mathrm{H^+} + \mathrm{Cl^-}) + (\mathrm{Na^+} + \mathrm{OH^-}) \longrightarrow \mathrm{Na^+} + \mathrm{Cl^-} + \mathrm{H_2O}(l)$$

spectator ion

This equation shows that the Na^+ and Cl^- ions did not react. These ions are called **spectator ions** because they were present but did not take part in the reaction. The only reaction that occurred was that between the H^+ and OH^- ions. Therefore the equation for the neutralization can be written as this net ionic equation:

$$\underset{\text{acid}}{\mathrm{H^+}(aq)} + \underset{\text{base}}{\mathrm{OH^-}(aq)} \longrightarrow \underset{\text{water}}{\mathrm{H_2O}(l)}$$

titration

This simple net ionic equation represents not only the reaction of sodium hydroxide and hydrochloric acid, but also the reaction of any strong acid with any water-soluble hydroxide base in an aqueous solution. The driving force of a neutralization reaction is the ability of an H^+ ion and an OH^- ion to react and form a molecule of water.

The amount of acid, base, or other species in a sample can be determined by **titration,** which measures the volume of one reagent required to react with a measured mass or volume of another reagent. Consider the titration of an acid with a base. A measured volume of acid of unknown concentration is placed in a flask, and a few drops of an indicator solution are added. Base solution of known concentration is slowly added from a buret to the acid until the indicator changes color. The indicator selected is one that changes color when the stoichiometric quantity (according to the equation) of base has been added to the acid. At this


(a) (b) (c)

Tritation. (a) Before the titration begins, the liqiuid in the flask is colorless and clear. (b) At the endpoint a pole pink color persists in the flash. (c) After the endpoint, the color of the solution turns bright pink, indicating an excess of base.

point, known as the *end point of the titration,* the titration is complete, and the volume of base used to neutralize the acid is read from the buret. The concentration or amount of acid in solution can be calculated from the titration data and the chemical equation for the reaction. Let's look at some examples.

Example 15.4

Suppose that 42.00 mL of 0.150 M NaOH solution are required to neutralize 50.00 mL of hydrochloric acid solution. What is the molarity of the acid solution?

SOLUTION

The equation for the reaction is

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

In this neutralization NaOH and HCl react in a 1 : 1 mole ratio. Therefore the moles of HCl in solution are equal to the moles of NaOH required to react with it. First we calculate the moles of NaOH used, and from this value we determine the moles of HCl:

Data: 42.00 mL of 0.150 M NaOH 50.00 mL HCl
molarity of acid = M (unknown)

Determine the moles of NaOH:

$$M = \mathrm{mol/L} \qquad 42.00\ \mathrm{mL} = 0.04200\ \mathrm{L}$$

$$(0.04200\ \cancel{\mathrm{L}})\left(\frac{0.150\ \mathrm{mol\ NaOH}}{1\ \cancel{\mathrm{L}}}\right) = 0.00630\ \mathrm{mol\ NaOH}$$

Since NaOH and HCl react in a 1 : 1 ratio, 0.00630 mol HCl was present in the 50.00 mL of HCl solution. Therefore, the molarity of the HCl is

$$M = \frac{\mathrm{mol}}{\mathrm{L}} = \frac{0.00630\ \mathrm{mol\ HCl}}{0.05000\ \mathrm{L}} = 0.126\ M\ \mathrm{HCl}$$

Example 15.5

Suppose that 42.00 mL of 0.150 M NaOH solution are required to neutralize 50.00 mL of H_2SO_4 solution. What is the molarity of the acid solution?

SOLUTION

The equation for the reaction is

$$2\ \mathrm{NaOH}(aq) + \mathrm{H_2SO_4}(aq) \longrightarrow \mathrm{Na_2SO_4}(aq) + 2\ \mathrm{H_2O}(l)$$

The same amount of base (0.00630 mol NaOH) is used in this titration as in Example 15.4, but the mole ratio of acid to base in the reaction is 1 : 2. The moles of H_2SO_4 reacted can be calculated:

Data: 42.00 mL of 0.150 M NaOH = 0.00630 mol NaOH

$$(0.00630\ \cancel{\mathrm{mol\ NaOH}})\left(\frac{1\ \mathrm{mol\ H_2SO_4}}{2\ \cancel{\mathrm{mol\ NaOH}}}\right) = 0.00315\ \mathrm{mol\ H_2SO_4}$$

Therefore 0.00315 mol H_2SO_4 was present in 50.00 mL of H_2SO_4 solution. The molarity of the H_2SO_4 is

$$M = \frac{\mathrm{mol}}{\mathrm{L}} = \frac{0.00315\ \mathrm{mol\ H_2SO_4}}{0.05000\ \mathrm{L}} = 0.0630\ M\ \mathrm{H_2SO_4}$$

Example 15.6 A 25.00-mL sample of H_2SO_4 solution required 14.26 mL of 0.2240 M NaOH for complete neutralization. What is the molarity of the sulfuric acid?

SOLUTION The equation for the reaction is

$$2\,\text{NaOH}(aq) + \text{H}_2\text{SO}_4(aq) \longrightarrow \text{Na}_2\text{SO}_4(aq) + 2\,\text{H}_2\text{O}(l)$$

The moles of NaOH needed are

$$\begin{aligned}\text{moles NaOH} &= (\text{V}_{\text{NaOH}})(M_{\text{NaOH}})\\ &= (0.01426\ \text{L})\left(0.2240\,\frac{\text{mol}}{\text{L}}\right)\\ &= 0.003194\ \text{mol NaOH}\end{aligned}$$

Since the mole ratio of acid to base is $\frac{1\ \text{H}_2\text{SO}_4}{2\ \text{NaOH}}$, the moles of acid in the sample are

$$(0.003194\ \text{mol NaOH})\left(\frac{1\ \text{H}_2\text{SO}_4}{2\ \text{NaOH}}\right) = 0.001597\ \text{mol H}_2\text{SO}_4$$

Now to find the molarity of the sample, we divide the moles of acid by its original volume:

$$\frac{0.001597\ \text{mol H}_2\text{SO}_4}{0.02500\ \text{L}} = 0.06388\ M\ \text{H}_2\text{SO}_4$$

Practice 15.7

A 50.0-mL sample of HCl required 24.81 mL of 0.1250 M NaOH for neutralization. What is the molarity of the acid?

15.11 Writing Net Ionic Equations

In Section 15.10, we wrote the reaction of hydrochloric acid and sodium hydroxide in three different equations:

1. $\text{HCl}(aq) + \text{NaOH}(aq) \longrightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$
2. $(\text{H}^+ + \text{Cl}^-) + (\text{Na}^+ + \text{OH}^-) \longrightarrow \text{Na}^+ + \text{Cl}^- + \text{H}_2\text{O}(l)$
3. $\text{H}^+ + \text{OH}^- \longrightarrow \text{H}_2\text{O}$

formula equation
total ionic equation
net ionic equation

In the **formula equation** (1), compounds are written in their molecular, or formula, expressions. In the **total ionic equation** (2), compounds are written to show the form in which they are predominantly present: strong electrolytes as ions in solution and nonelectrolytes, weak electrolytes, precipitates, and gases in their molecular forms. In the **net ionic equation** (3), only those molecules or ions that have reacted are included in the equation; ions or molecules that do not react (the spectators) are omitted.

When balancing equations thus far, we've been concerned only with the atoms of the individual elements. Because ions are electrically charged, ionic equations often end up with a net electrical charge. A balanced equation must have the same net charge on each side, whether that charge is positive, negative, or zero. Therefore, when balancing ionic equations, we must make sure that both the same number of each kind of atom and the same net electrical charge are present on each side.

Here is a list of rules for writing ionic equations:

1. Strong electrolytes in solution are written in their ionic form.
2. Weak electrolytes are written in their molecular form.
3. Nonelectrolytes are written in their molecular form.
4. Insoluble substances, precipitates, and gases are written in their molecular forms.
5. The net ionic equation should include only substances that have undergone a chemical change. Spectator ions are omitted from the net ionic equation.
6. Equations must be balanced, both in atoms and in electrical charge.

Study the following examples. In each one, the formula equation is given. Write the total ionic equation and the net ionic equation for each.

Example 15.7

SOLUTION

$$\mathrm{HNO_3}(aq) + \mathrm{KOH}(aq) \longrightarrow \mathrm{KNO_3}(aq) + \mathrm{H_2O}(l)$$

formula equation

$$(\mathrm{H^+} + \mathrm{NO_3^-}) + (\mathrm{K^+} + \mathrm{OH^-}) \longrightarrow \mathrm{K^+} + \mathrm{NO_3^-} + \mathrm{H_2O}$$

total ionic equation

$$\mathrm{H^+} + \mathrm{OH^-} \longrightarrow \mathrm{H_2O}$$

net ionic equation

The HNO_3, KOH, and KNO_3 are soluble, strong electrolytes. The K^+ and NO_3^- ions are spectator ions, have not changed, and are not included in the net ionic equation. Water is a nonelectrolyte and is written in the molecular form.

Example 15.8

SOLUTION

$$2\,\mathrm{AgNO_3}(aq) + \mathrm{BaCl_2}(aq) \longrightarrow 2\,\mathrm{AgCl}(s) + \mathrm{Ba(NO_3)_2}(aq)$$

formula equation

$$(2\,\mathrm{Ag^+} + 2\,\mathrm{NO_3^-}) + (\mathrm{Ba^{2+}} + 2\,\mathrm{Cl^-}) \longrightarrow 2\,\mathrm{AgCl}(s) + \mathrm{Ba^{2+}} + 2\,\mathrm{NO_3^-}$$

total ionic equation

$$\mathrm{Ag^+} + \mathrm{Cl^-} \longrightarrow \mathrm{AgCl}(s)$$

net ionic equation

Although AgCl is an ionic compound, it is written in the un-ionized form on the right side of the ionic equations because most of the Ag^+ and Cl^- ions are no longer in solution but have formed a precipitate of AgCl. The Ba^{2+} and NO_3^- ions are spectator ions.

Example 15.9

$$Na_2CO_3(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$$

formula equation

SOLUTION

$$(2\,Na^+ + CO_3^{2-}) + (2\,H^+ + SO_4^{2-}) \longrightarrow 2\,Na^+ + SO_4^{2-} + H_2O(l) + CO_2(g)$$

total ionic equation

$$CO_3^{2-} + 2\,H^+ \longrightarrow H_2O(l) + CO_2(g)$$

net ionic equation

Carbon dioxide (CO_2) is a gas and evolves from the solution; Na^+ and SO_4^{2-} are spectator ions.

Example 15.10

$$HC_2H_3O_2(aq) + NaOH(aq) \longrightarrow NaC_2H_3O_2(aq) + H_2O(l)$$

formula equation

SOLUTION

$$HC_2H_3O_2 + (Na^+ + OH^-) \longrightarrow Na^+ + C_2H_3O_2^- + H_2O$$

total ionic equation

$$HC_2H_3O_2 + OH^- \longrightarrow C_2H_3O_2^- + H_2O$$

net ionic equation

Acetic acid, $HC_2H_3O_2$, a weak acid, is written in the molecular form, but sodium acetate, $NaC_2H_3O_2$, a soluble salt, is written in the ionic form. The Na^+ ion is the only spectator ion in this reaction. Both sides of the net ionic equation have a -1 electrical charge.

Example 15.11

$$Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

formula equation

SOLUTION

$$Mg + (2\,H^+ + 2\,Cl^-) \longrightarrow Mg^{2+} + 2\,Cl^- + H_2(g)$$

total ionic equation

$$Mg + 2\,H^+ \longrightarrow Mg^{2+} + H_2(g)$$

net ionic equation

The net electrical charge on both sides of the equation is $+2$.

Example 15.12

$$H_2SO_4(aq) + Ba(OH)_2(aq) \longrightarrow BaSO_4(s) + 2\,H_2O(l)$$

formula equation

SOLUTION

$$(2\,H^+ + SO_4^{2-}) + (Ba^{2+} + 2\,OH^-) \longrightarrow BaSO_4(s) + 2\,H_2O(l)$$

total ionic equation

$$2\,H^+ + SO_4^{2-} + Ba^{2+} + 2\,OH^- \longrightarrow BaSO_4(s) + 2\,H_2O(l)$$

net ionic equation

Barium sulfate ($BaSO_4$) is a highly insoluble salt. If we conduct this reaction using the conductivity apparatus described in Section 15.5, the light glows brightly at first but goes out when the reaction is complete because almost no ions are left in solution. The $BaSO_4$ precipitates out of solution, and water is a nonconductor of electricity.

Practice 15.8

Write the net ionic equation for

$$3\,H_2S(aq) + 2\,Bi(NO_3)_3(aq) \longrightarrow Bi_2S_3(s) + 6\,HNO_3(aq)$$

15.12 Acid Rain

Acid rain is defined as any atmospheric precipitation that is more acidic than usual. The increase in acidity might be from natural or industrial sources. Rain acidity varies throughout the world and across the United States. The pH of rain is generally lower in the eastern United States and higher in the West. Unpolluted rain has a pH of 5.6, and so is slightly acidic. This acidity results from the dissolution of carbon dioxide in the water producing carbonic acid:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

The general process involves the following steps:

1. emission of nitrogen and sulfur oxides into the air
2. transportation of these oxides throughout the atmosphere
3. chemical reactions between the oxides and water, forming sulfuric acid (H_2SO_4) and nitric acid (HNO_3)
4. rain or snow, which carries the acids to the ground

The oxides may also be deposited directly on a dry surface and become acidic when normal rain falls on them.

Acid rain is not a new phenomenon. Rain was probably acidic in the early days of our planet as volcanic eruptions, fires, and decomposition of organic matter released large volumes of nitrogen and sulfur oxides into the atmosphere. Use of fossil fuels, especially since the Industrial Revolution, has made significant changes in the amounts of pollutants being released into the atmosphere. As increasing amounts of fossil fuels have been burned, more and more sulfur and nitrogen oxides have poured into the atmosphere, thus increasing the acidity of rain.

This relief work on St. Bartholomew's Church in New York shows the destructive power of acid rain. The marble is being dissolved slowly over time.

Acid rain affects a variety of factors in our environment. For example, freshwater plants and animals decline significantly when rain is acidic; large numbers of fish and plants die when acidic water from spring thaws enters the lakes. Acidic rainwater leaches aluminum from the soil into lakes, where the aluminum compounds adversely affect the gills of fish. In addition to leaching aluminum from the soil, acid rain also causes other valuable minerals, such as magnesium and calcium, to dissolve and run into lakes and streams. It can also dissolve the waxy protective coat on plant leaves, making them vulnerable to attack by bacteria and fungi.

In our cities, acid rain is responsible for extensive and continuing damage to buildings, monuments, and statues. It reduces the durability of paint and promotes the deterioration of paper, leather, and cloth. In short, we are just beginning to explore the effects of acid rain on human beings and on our food chain.

15.13 Colloids

When we add sugar to a flask of water and shake it, the sugar dissolves and forms a clear homogeneous *solution*. When we do the same experiment with very fine sand and water, the sand particles form a *suspension,* which settles when the shaking stops. When we repeat the experiment again using ordinary cornstarch, we find that the starch does not dissolve in cold water. But if the mixture is heated and stirred, the starch forms a cloudy, opalescent *dispersion*. This dispersion does not appear to be clear and homogeneous like the sugar solution, yet it is not obviously heterogeneous and does not settle like the sand suspension. In short, its properties are intermediate between those of the sugar solution and those of the sand suspension. The starch dispersion is actually a *colloid,* a name derived from the Greek *kolla,* meaning "glue," and was coined by the English scientist Thomas Graham in 1861.

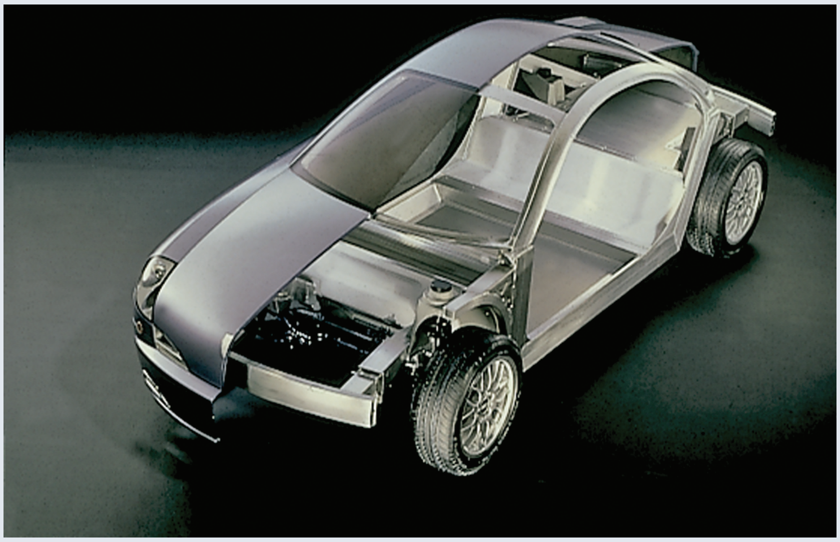
colloid

As it is now used, the word **colloid** means a dispersion in which the dispersed particles are larger than the solute ions or molecules of a true solution and smaller than the particles of a mechanical suspension. The term does not imply a gluelike quality, although most glues are colloidal materials. The size of colloidal particles ranges from a lower limit of about 1 nm (10^{-7} cm) to an upper limit of about 1000 nm (10^{-4} cm). There are eight types of colloids, which are summarized in Table 15.6.

Table 15.6 Types of Colloidal Dispersions

Type	Name	Examples
Gas in liquid	foam	whipped cream, soapsuds
Gas in solid	solid foam	Styrofoam, foam rubber, pumice
Liquid in gas	liquid aerosol	fog, clouds
Liquid in liquid	emulsion	milk, vinegar in oil salad dressing, mayonnaise
Liquid in solid	solid emulsion	cheese, opals, jellies
Solid in gas	solid aerosol	smoke, dust in air
Solid in liquid	sol	india ink, gold sol
Solid in solid	solid sol	tire rubber, certain gems (e.g., rubies)

CHEMISTRY IN ACTION • Foam Cars—Wave of the Future?



Metal foams are a new class of materials that may revolutionize the car industry. Automotive makers know that one key to making cars that are fuel efficient is to decrease weight. Up till now, that weight reduction meant higher cost (in materials like titanium and aluminum) and problems in crash testing (since light vehicles don't absorb energy well).

German automotive supplier Wilhelm Karmann (whose company manufactured the Volkswagen Karmann-Ghia) developed an aluminum foam composite material with some amazing properties. Parts made of this foam composite weigh 30–50% less than an equivalent steel part and are 10 times stiffer. The material is so light it floats in water and although it costs 20–25% more than steel, it could be used for as much as 20% of a compact car. Since the surface of the new material is not smooth, it would likely be used in structural areas of the car (firewalls, roof panels, luggage compartment walls, etc.)

How is aluminum foam made? Two layers of aluminum sheet and a middle powder layer (made of titanium metal hydride and aluminum powder) are rolled together under very high pressure to make a single flat sheet. This sheet metal is then processed in traditional ways to make a variety of 3-D shapes. Then the sheet is placed in a 1148°F oven for two minutes. This quick bake allows the Al metal to melt and mix with H_2 gas (released from the titanium hydride), making foam. The sheet rises just like a cake (increasing five to seven times in thickness). When the foam cools it is a rigidly formed 3-D structure between two aluminum skins.

Not only is the new aluminum foam part much lighter than its steel counterpart; it also performs well in crash tests. One of these days you may climb into a car that is really a foam colloid—with great fuel economy and crash resistance.

The fundamental difference between a colloidal dispersion and a true solution is the size, not the nature, of the particles. The solute particles in a solution are usually single ions or molecules that may be hydrated to varying degrees. Colloidal particles are usually aggregations of ions or molecules. However, the molecules of some polymers, such as proteins, are large enough to be classified as colloidal particles when in solution. To fully appreciate the differences in relative sizes, the volumes (not just the linear dimensions) of colloidal particles and solute particles must be compared. The difference in volumes can be approximated by assuming that the particles are spheres. A large colloidal particle has a diameter of about 500 nm, whereas a fair-sized ion or molecule has a diameter of about 0.5 nm. Thus the diameter of the colloidal particle is about 1000 times that of the solute particle. Because the volumes of spheres are proportional to the cubes of their diameters, we can calculate that the volume of a colloidal particle can be up to a billion ($10^3 \times 10^3 \times 10^3 = 10^9$) times greater than that of a solution particle.

15.14 Properties of Colloids

In 1827, while observing a strongly illuminated aqueous suspension of pollen under a high-powered microscope, Robert Brown (1773–1858) noted that the pollen grains appeared to have a trembling, erratic motion. He later

Brownian movement

determined that this erratic motion is not confined to pollen but is characteristic of colloidal particles in general. This random motion of colloidal particles is called **Brownian movement.** We can readily observe such movement by confining cigarette smoke in a small transparent chamber and illuminating it with a strong beam of light at right angles to the optical axis of the microscope. The smoke particles appear as tiny randomly moving lights because the light is reflected from their surfaces. This motion is due to the continual bombardment of the smoke particles by air molecules. Since Brownian movement can be seen when colloidal particles are dispersed in either a gaseous or a liquid medium, it affords nearly direct visual proof that matter at the molecular level is moving randomly, as postulated by the kinetic-molecular theory.

Tyndall effect

When an intense beam of light is passed through an ordinary solution and viewed at an angle, the beam passing through the solution is hardly visible. A beam of light, however, is clearly visible and sharply outlined when it is passed through a colloidal dispersion. This phenomenon is known as the **Tyndall effect.** The Tyndall effect, like Brownian movement, can be observed in nearly all colloidal dispersions. It occurs because the colloidal particles are large enough to scatter the rays of visible light. The ions or molecules of true solutions are too small to scatter light and therefore do not exhibit a noticeable Tyndall effect.

A beam of lights is visible in a colloid solution (left) but not in a true solution. The size of the particles in the colloid are large enough to scatter the light beam making it visible.

Another important characteristic of colloids is that the particles have relatively huge surface areas. We saw in Section 14.4 that the surface area is increased 10-fold when a 1-cm cube is divided into 1000 cubes with sides of 0.1 cm. When a 1-cm cube is divided into colloidal-size cubes measuring 10^{-6} cm, the combined surface area of all the particles becomes a million times greater than that of the original cube.

Colloidal particles become electrically charged when they adsorb ions on their surfaces. *Adsorption* should not be confused with *absorption.* Adsorption refers to the adhesion of molecules or ions to a surface, whereas absorption refers to the taking in of one material by another material. Adsorption occurs because the atoms or ions at the surface of a particle are not completely surrounded by other atoms or ions as are those in the interior. Consequently, these surface atoms or ions attract and adsorb ions or polar molecules from the dispersion medium onto the surfaces of the colloidal particles. This property is directly related to the large surface area presented by the many tiny particles.

15.15 Applications of Colloidal Properties

Activated charcoal has an enormous surface area, approximately 1 million square centimeters per gram in some samples. Hence, charcoal is very effective in selectively adsorbing the polar molecules of some poisonous gases and is therefore used in gas masks. Charcoal can be used to adsorb impurities from liquids as well as from gases, and large amounts are used to remove substances that have objectionable tastes and odors from water supplies. In sugar refineries, activated charcoal is used to adsorb colored impurities from the raw sugar solutions.

Colloidal particles become electrically charged when they adsorb ions on their surface.

A process widely used for dust and smoke control in many urban and industrial areas was devised by an American, Frederick Cottrell (1877–1948). The Cottrell process takes advantage of the fact that the particulate matter in dust and smoke is electrically charged. Air to be cleaned of dust or smoke is passed between electrode plates charged with a high voltage. Positively charged particles are attracted to, neutralized, and thereby precipitated at the negative electrodes. Negatively charged particles are removed in the same fashion at the positive electrodes. Large Cottrell units are fitted with devices for automatic removal of precipitated material. Small units, designed for removing dust and pollen from air in the home, are now on the market. Unfortunately, Cottrell units remove only particulate matter; they cannot remove gaseous pollutants such as carbon monoxide, sulfur dioxide, and nitrogen oxides.

Thomas Graham found that a parchment membrane would allow the passage of true solutions but would prevent the passage of colloidal dispersions. Dissolved solutes can be removed from colloidal dispersions through the use of such a membrane by a process called **dialysis.** The membrane itself is called a *dialyzing membrane.* Artificial membranes are made from such materials as parchment paper, collodion, or certain kinds of cellophane. Dialysis can be demonstrated by putting a colloidal starch dispersion and some copper(II) sulfate solution in a parchment paper bag and suspending it in running water. In a few hours, the blue color of the copper(II) sulfate has disappeared, and only the starch dispersion remains in the bag.

dialysis

A life-saving application of dialysis has been the development of artificial kidneys. The blood of a patient suffering from partial kidney failure is passed through the artificial kidney machine for several hours, during which time the soluble waste products are removed by dialysis.

Chapter 15 Review

15.1 Acids and Bases

KEY TERM

Hydronium ion

- Characteristic properties of acids include:
  - They taste sour.
  - They change litmus from blue to red.
  - They react with:
    - Metals to form hydrogen gas
    - Hydroxide bases to form water and a salt
    - Carbonates to produce CO_2
- Characteristic properties of bases include:
  - They taste bitter or caustic.
  - They have a slippery feeling.
  - They change litmus from red to blue.
  - They interact with acids to form water and a salt.
- Arrhenius definition of acids and bases:
  - Acids contain excess H^+ ions in aqueous solutions.
  - Bases contain excess OH^- ions in aqueous solutions.
- Brønsted–Lowry definition of acids and bases:
  - Acids are proton donors.
  - Bases are proton acceptors.
- Lewis definition of acids and bases:
  - Acids are electron pair acceptors.
  - Bases are electron pair donors.

15.2 Reactions of Acids

- Acids react with metals above hydrogen in the activity series to form hydrogen gas and a salt.
- Acids react with bases to form water and a salt (neutralization reaction).
- Acids react with metal oxides to form water and salt.
- Acids react with carbonates to form water, a salt, and carbon dioxide.

15.3 Reactions of Bases

KEY TERM

Amphoteric

- Bases react with acids to form water and a salt (neutralization reaction).
- Some amphoteric metals react with NaOH or KOH to form hydrogen and a salt.

15.4 Salts

- Salts can be considered to be derived from the reaction of an acid and a base.
- Salts are crystalline and have high melting and boiling points.

15.5 Electrolytes and Nonelectrolytes

KEY TERMS

Electrolyte
Nonelectrolyte

- A substance whose aqueous solution conducts electricity is called an electrolyte:
  - Aqueous solution contains ions.
- A substance whose aqueous solution does not conduct electricity is called a nonelectrolyte:
  - Aqueous solution contains molecules, not ions.

15.6 Dissociation and Ionization of Electrolytes

KEY TERMS

Dissociation
Ionization

- Dissociation is the process by which the ions of a salt separate as the salt dissolves.
- Ionization is the formation of ions that occurs as a result of a chemical reaction with water.

15.7 Strong and Weak Electrolytes

KEY TERMS

Strong electrolyte
Weak electrolyte

- Strong electrolytes are 100% ionized in solution:
  - Strong acids and strong bases
- Weak electrolytes are much less ionized in solution:
  - Weak acids and bases
- Colligative properties of electrolyte solutions depend on the number of particles produced during the ionization of the electrolyte.

15.8 Ionization of Water

- Water can self-ionize to form H^+ and OH^- ions.
- Concentrations of ions in water at 25°C:
  - $[H^+] = 1.0 \times 10^{-7}$
  - $[OH^-] = 1.0 \times 10^{-7}$

15.9 Introduction to pH

KEY TERMS

pH
Logarithm

- $pH = -\log[H^+]$:
  - $pH < 7$ for acidic solution
  - $pH = 7$ in neutral solution
  - $pH > 7$ for basic solution
- The number of decimal places in a logarithm equals the number of significant figures in the original number.

15.10 Neutralization

KEY TERMS

Neutralization
Spectator ion
Titration

- The reaction of an acid and a base to form water and a salt is called neutralization:
  - The general equation for a neutralization reaction is

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

- The quantitative study of a neutralization reaction is called a titration.

15.11 Writing Net Ionic Equations

KEY TERMS

Formula equation
Total ionic equation
Net ionic equation

- In a formula equation the compounds are written in their molecular, or formula, expressions.
- In the total ionic equation compounds are written as ions if they are strong electrolytes in solution and as molecules if they are precipitates, nonelectrolytes, or weak electrolytes in solution.
- The net ionic equation shows only the molecules or ions that have changed:
  - The spectators (nonreactors) are omitted.

15.12 Acid Rain

- Acid rain is any atmospheric precipitation that is more acidic that usual.
- Acid rain causes significant damage and destruction to our environment.

15.13 Colloids

KEY TERM

Colloid

- A colloid is a dispersion containing particles between 1 nm and 1000 nm.
- Colloid particles are usually aggregates of ions or molecules.
- The difference between a colloid and a true solution is the size of the particles.

15.14 Properties of Colloids

KEY TERMS

Brownian movement
Tyndall effect

- The random motion of particles in a colloid is called Brownion movement.
- Colloid particles cause light from an intense beam to be scattered clearly, showing the path of the light through the colloid.

15.15 Application of Colloidal Properties

KEY TERM

Dialysis

- Activated charcoal can be used to adsorb impurities from liquids and gases.
- Dialysis is based on the idea that a true solution will pass through a parchment membrane but a colloid will not.

Review Questions

All questions with blue numbers have answers in the appendix of the text.

1. Since a hydrogen ion and a proton are identical, what differences exist between the Arrhenius and Brønsted–Lowry definitions of an acid? (Table 15.1)
2. According to Figure 15.1, what type of substance must be in solution for the bulb to light?
3. Which of the following classes of compounds are electrolytes: acids, alcohols, bases, salts? (Table 15.2)
4. What two differences are apparent in the arrangement of water molecules about the hydrated ions as depicted in Figure 15.2?
5. The pH of a solution with a hydrogen ion concentration of 0.003 *M* is between what two whole numbers? (Table 15.4)
6. Which is more acidic, tomato juice or blood? (Table 15.5)
7. Use the three acid–base theories (Arrhenius, Brønsted–Lowry, and Lewis) to define an acid and a base.
8. For each acid–base theory referred to in Question 7, write an equation illustrating the neutralization of an acid with a base.
9. Write the Lewis structure for the (a) bromide ion, (b) hydroxide ion, and (c) cyanide ion. Why are these ions considered to be bases according to the Brønsted–Lowry and Lewis acid–base theories?
10. Into what three classes of compounds do electrolytes generally fall?
11. Name each compound listed in Table 15.3.
12. A solution of HCl in water conducts an electric current, but a solution of HCl in hexane does not. Explain this behavior in terms of ionization and chemical bonding.
13. How do ionic compounds exist in their crystalline structure? What occurs when they are dissolved in water?
14. An aqueous methyl alcohol, CH_3OH, solution does not conduct an electric current, but a solution of sodium hydroxide, NaOH, does. What does this information tell us about the OH group in the alcohol?
15. Why does molten NaCl conduct electricity?
16. Explain the difference between dissociation of ionic compounds and ionization of molecular compounds.
17. Distinguish between strong and weak electrolytes.
18. Explain why ions are hydrated in aqueous solutions.
19. What is the main distinction between water solutions of strong and weak electrolytes?
20. What are the relative concentrations of $H^+(aq)$ and $OH^-(aq)$ in (a) a neutral solution, (b) an acid solution, and (c) a basic solution?
21. Write the net ionic equation for the reaction of a strong acid with a water-soluble hydroxide base in an aqueous solution.
22. The solubility of HCl gas in water, a polar solvent, is much greater than its solubility in hexane, a nonpolar solvent. How can you account for this difference?
23. Pure water, containing equal concentrations of both acid and base ions, is neutral. Why?
24. Indicate the fundamental difference between a colloidal dispersion and a true solution.
25. Explain the process of dialysis, giving a practical application in society.
26. A solution with a pH of 7 is neutral. A solution with a pH less than 7 is acidic. A solution with a pH greater than 7 is basic. What do these statements mean?
27. How does acid rain form?
28. Explain the purpose of a titration.

Paired Exercises

All exercises with blue numbers have answers in the appendix of the text.

1. Identify the conjugate acid–base pairs in the following equations:

   (a) $HC_2H_3O_2 + H_2SO_4 \rightleftharpoons H_2C_2H_3O_2^+ + HSO_4^-$

   (b) The two-step ionization of sulfuric acid,

   $H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^-$

   $HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$

   (c) $HClO_4 + H_2O \longrightarrow H_3O^+ + ClO_4^-$

   (d) $CH_3O^- + H_3O^+ \longrightarrow CH_3OH + H_2O$

2. Identify the conjugate acid–base pairs in the following equations:

   (a) $HCl + NH_3 \longrightarrow NH_4^+ + Cl^-$

   (b) $HCO_3^- + OH^- \rightleftharpoons CO_3^{2-} + H_2O$

   (c) $HCO_3^- + H_3O^+ \rightleftharpoons H_2CO_3 + H_2O$

   (d) $HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$

3. Complete and balance these equations:
(a) $Zn(s) + HCl(aq) \longrightarrow$
(b) $Al(OH)_3(s) + H_2SO_4(aq) \longrightarrow$
(c) $Na_2CO_3(aq) + HC_2H_3O_2(aq) \longrightarrow$
(d) $MgO(s) + HI(aq) \longrightarrow$
(e) $Ca(HCO_3)_2(s) + HBr(aq) \longrightarrow$
(f) $KOH(aq) + H_3PO_4(aq) \longrightarrow$

5. For each of the formula equations in Question 3, write total and net ionic equations.

7. Which of these compounds are electrolytes? Consider each substance to be mixed with water.
(a) HCl
(b) CO_2
(c) $CaCl_2$
(d) $C_{12}H_{22}O_{11}$ (sugar)
(e) C_3H_7OH (rubbing alcohol)
(f) CCl_4 (insoluble)

9. Calculate the molarity of the ions present in these salt solutions. Assume each salt to be 100% dissociated:
(a) 0.015 *M* NaCl
(b) 4.25 *M* $NaKSO_4$
(c) 0.20 *M* $CaCl_2$
(d) 22.0 g KI in 500. mL of solution

11. In Exercise 9, how many grams of each ion would be present in 100. mL of each solution?

13. Calculate the $[H^+]$ for:
(a) a solution with a pH = 8.5
(b) pure water
(c) a solution with a pH = 2.5

15. What is the molar concentration of all ions present in a solution prepared by mixing the following? (Neglect the concentration of H^+ and OH^- from water and assume that volumes of solutions are additive.)
(a) 30.0 mL of 1.0 *M* NaCl and 40.0 mL of 1.0 *M* NaCl
(b) 30.0 mL of 1.0 *M* HCl and 30.0 mL of 1.0 *M* NaOH
(c) 100.0 mL of 0.40 *M* KOH and 100.0 mL of 0.80 *M* HCl

17. Given the data for the following separate titrations, calculate the molarity of the HCl:

	mL HCl	Molarity HCl	mL NaOH	Molarity NaOH
(a)	40.13	*M*	37.70	0.728
(b)	19.00	*M*	33.66	0.306
(c)	27.25	*M*	18.00	0.555

4. Complete and balance these equations:
(a) $Fe_2O_3(s) + HBr(aq) \longrightarrow$
(b) $Al(s) + H_2SO_4(aq) \longrightarrow$
(c) $NaOH(aq) + H_2CO_3(aq) \longrightarrow$
(d) $Ba(OH)_2(s) + HClO_4(aq) \longrightarrow$
(e) $Mg(s) + HClO_4(aq) \longrightarrow$
(f) $K_2O(s) + HI(aq) \longrightarrow$

6. For each of the formula equations in Question 4, write total and net ionic equations.

8. Which of these compounds are electrolytes? Consider each substance to be mixed with water.
(a) $NaHCO_3$ (baking soda)
(b) N_2 (insoluble gas)
(c) $AgNO_3$
(d) HCOOH (formic acid)
(e) RbOH
(f) K_2CrO_4

10. Calculate the molarity of the ions present in these salt solutions. Assume each salt to be 100% dissociated:
(a) 0.75 *M* $ZnBr_2$
(b) 1.65 *M* $Al_2(SO_4)_3$
(c) 900. g $(NH_4)_2SO_4$ in 20.0 L of solution
(d) 0.0120 g $Mg(ClO_3)_2$ in 1.00 mL of solution

12. In Exercise 10, how many grams of each ion would be present in 100. mL of each solution?

14. Calculate the $[H^+]$ for:
(a) a solution with a pH = 2.4
(b) a solution with a pH = 10.0
(c) tap water with a pH = 6.4

16. What is the molar concentration of all ions present in a solution prepared by mixing the following? (Neglect the concentration of H^+ and OH^- from water and assume volumes of solutions are additive.)
(a) 100.0 mL of 2.0 *M* KCl and 100.0 mL of 1.0 *M* $CaCl_2$
(b) 35.0 mL of 0.20 *M* $Ba(OH)_2$ and 35.0 mL of 0.20 *M* H_2SO_4
(c) 1.00 L of 1.0 *M* $AgNO_3$ and 500. mL of 2.0 *M* NaCl

18. Given the data for the following separate titrations, calculate the molarity of the NaOH:

	mL HCl	Molarity HCl	mL NaOH	Molarity NaOH
(a)	37.19	0.126	31.91	*M*
(b)	48.04	0.482	24.02	*M*
(c)	13.13	1.425	39.39	*M*

19. Rewrite the following unbalanced equations, changing them into balanced net ionic equations (assume that all reactions are in water solution):
(a) $K_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow KNO_3(aq) + BaSO_4(s)$
(b) $CaCO_3(s) + HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
(c) $Mg(s) + HC_2H_3O_2(aq) \longrightarrow Mg(C_2H_3O_2)_2(aq) + H_2(g)$

21. For each of the given pairs, determine which solution is more acidic. All are water solutions. Explain your answer.
(a) 1 M HCl or 1 M H_2SO_4?
(b) 1 M HCl or 1 M $HC_2H_3O_2$?

23. What volume (in milliliters) of 0.245 M HCl will neutralize 10.0 g $Al(OH)_3$? The equation is

$$3\,HCl(aq) + Al(OH)_3(s) \longrightarrow AlCl_3(aq) + 3\,H_2O(l)$$

***25.** A 0.200-g sample of impure NaOH requires 18.25 mL of 0.2406 M HCl for neutralization. What is the percent of NaOH in the sample?

***27.** What volume of H_2 gas, measured at 27°C and 700. torr, can be obtained by reacting 5.00 g of zinc metal with 100. mL of 0.350 M HCl? The equation is

$$Zn(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

29. Calculate the pH of solutions having these H^+ ion concentrations:
(a) 0.01 M
(b) 1.0 M
(c) 6.5×10^{-9} M

31. Calculate the pH of
(a) orange juice, 3.7×10^{-4} M H^+
(b) vinegar, 2.8×10^{-3} M H^+

33. Determine whether each of the following is a strong acid, weak acid, strong base, or weak base. Then write an equation describing the process that occurs when the substance is dissolved in water.
(a) NH_3
(b) HCl
(c) KOH
(d) $HC_2H_3O_2$

20. Rewrite the following unbalanced equations, changing them into balanced net ionic equations (assume that all reactions are in water solution):
(a) $H_2S(g) + CdCl_2(aq) \longrightarrow CdS(s) + HCl(aq)$
(b) $Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$
(c) $AlCl_3(aq) + Na_3PO_4(aq) \longrightarrow AlPO_4(s) + NaCl(aq)$

22. For each of the given pairs, determine which solution is more acidic. All are water solutions. Explain your answer.
(a) 1 M HCl or 2 M HCl?
(b) 1 M HNO_3 or 1 M H_2SO_4?

24. What volume (in milliliters) of 0.245 M HCl will neutralize 50.0 mL of 0.100 M $Ca(OH)_2$? The equation is

$$2\,HCl(aq) + Ca(OH)_2(aq) \longrightarrow CaCl_2(aq) + 2\,H_2O(l)$$

***26.** A batch of sodium hydroxide was found to contain sodium chloride as an impurity. To determine the amount of impurity, a 1.00-g sample was analyzed and found to require 49.90 mL of 0.466 M HCl for neutralization. What is the percent of NaCl in the sample?

***28.** What volume of H_2 gas, measured at 27°C and 700. torr, can be obtained by reacting 5.00 g of zinc metal with 200. mL of 0.350 M HCl? The equation is

$$Zn(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

30. Calculate the pH of solutions having these H^+ ion concentrations:
(a) 1×10^{-7} M
(b) 0.50 M
(c) 0.00010 M

32. Calculate the pH of
(a) black coffee, 5.0×10^{-5} M H^+
(b) limewater, 3.4×10^{-11} M H^+

34. Determine whether each of the following is a strong acid, weak acid, strong base, or weak base. Then write an equation describing the process that occurs when the substance is dissolved in water.
(a) $H_2C_2O_4$
(b) $Ba(OH)_2$
(c) $HClO_4$
(d) HBr

Additional Exercises

All exercises with blue numbers have answers in the appendix of the text.

35. Determine whether each of the following describes a substance that is acidic, basic, or neutral.
(a) $[H^+] = 1 \times 10^{-4}$
(b) Phenolphthalein turns pink.
(c) pH = 8
(d) $[H^+] = 1 \times 10^{-9}$
(e) $[OH^-] = 1 \times 10^{-7}$
(f) Blue litmus paper turns red.

36. Draw pictures similar to those found in Figure 15.2 to show what happens to the following compounds when each is dissolved in water.
(a) $CaCl_2$ (b) KF (c) $AlBr_3$

37. What is the concentration of Ca^{2+} ions in a solution of CaI_2 having an I^- ion concentration of 0.520 *M*?

38. If 29.26 mL of 0.430 *M* HCl neutralize 20.40 mL of $Ba(OH)_2$ solution, what is the molarity of the $Ba(OH)_2$ solution? The reaction is

$$Ba(OH)_2(aq) + 2\,HCl(aq) \longrightarrow BaCl_2(aq) + 2\,H_2O(l)$$

39. A 1 *m* solution of acetic acid, $HC_2H_3O_2$, in water freezes at a lower temperature than a 1 *m* solution of ethyl alcohol, C_2H_5OH, in water. Explain.

40. At the same cost per pound, which alcohol, CH_3OH or C_2H_5OH, would be more economical to purchase as an antifreeze for your car? Why?

41. How does a hydronium ion differ from a hydrogen ion?

42. Arrange, in decreasing order of freezing points, 1 *m* aqueous solutions of HCl, $HC_2H_3O_2$, $C_{12}H_{22}O_{11}$ (sucrose), and $CaCl_2$. (List the one with the highest freezing point first.)

43. At 100°C the H^+ concentration in water is about 1×10^{-6} mol/L, about 10 times that of water at 25°C. At which of these temperatures is
(a) the pH of water the greater?
(b) the hydrogen ion (hydronium ion) concentration the higher?
(c) the water neutral?

44. What is the relative difference in H^+ concentration in solutions that differ by one pH unit?

45. What is the mole percent of a 1.00 *m* aqueous solution?

46. A sample of pure sodium carbonate with a mass of 0.452 g was dissolved in water and neutralized with 42.4 mL of hydrochloric acid. Calculate the molarity of the acid:

$$Na_2CO_3(aq) + 2\,HCl(aq) \longrightarrow 2\,NaCl(aq) + CO_2(g) + H_2O(l)$$

47. What volume (mL) of 0.1234 *M* HCl is needed to neutralize 2.00 g $Ca(OH)_2$?

48. How many grams of KOH are required to neutralize 50.00 mL of 0.240 *M* HNO_3?

49. Two drops (0.1 mL) of 1.0 *M* HCl are added to water to make 1.0 L of solution. What is the pH of this solution if the HCl is 100% ionized?

50. What volume of concentrated (18.0 *M*) sulfuric acid must be used to prepare 50.0 L of 5.00 *M* solution?

***51.** If 3.0 g NaOH are added to 500. mL of 0.10 *M* HCl, will the resulting solution be acidic or basic? Show evidence for your answer.

***52.** If 380 mL of 0.35 *M* $Ba(OH)_2$ are added to 500. mL of 0.65 *M* HCl, will the mixture be acidic or basic ? Find the pH of the resulting solution.

***53.** If 50.00 mL of 0.2000 *M* HCl are titrated with 0.2000 *M* NaOH, find the pH of the solution after the following amounts of base have been added:
(a) 0.000 mL (d) 49.00 mL (f) 49.99 mL
(b) 10.00 mL (e) 49.90 mL (g) 50.00 mL
(c) 25.00 mL
Plot your answers on a graph with pH on the *y*-axis and mL NaOH on the *x*-axis.

54. Sulfuric acid reacts with NaOH:
(a) Write a balanced equation for the reaction producing Na_2SO_4.
(b) How many milliliters of 0.10 *M* NaOH are needed to react with 0.0050 mol H_2SO_4?
(c) How many grams of Na_2SO_4 will also form?

***55.** A 10.0-mL sample of HNO_3 was diluted to a volume of 100.00 mL. Then, 25 mL of that diluted solution were needed to neutralize 50.0 mL of 0.60 *M* KOH. What was the concentration of the original nitric acid?

56. The pH of a solution of a strong acid was determined to be 3. If water is then added to dilute this solution, would the pH change? Why or why not? Could enough water ever be added to raise the pH of an acid solution above 7?

Challenge Exercises

All exercises with blue numbers have answers in the appendix of the text.

***57.** An HCl solution has a pH of −0.300. What volume of water must be added to 200 mL of this solution to change the pH to −0.150?

***58.** Lactic acid (found in sour milk) has an empirical formula of $HC_3H_5O_3$. A 1.0-g sample of lactic acid required 17.0 mL of 0.65 *M* NaOH to reach the end point of a titration. What is the molecular formula for lactic acid?

Answers to Practice Exercises

15.1 (a) HCO_3^-, (b) NO_2^-, (c) $C_2H_3O_2^-$

15.2 (a) H_2SO_4, (b) NH_4^+, (c) H_2O

15.3 (a) KOH, potassium hydroxide; H_3PO_4, phosphoric acid
(b) $Mg(OH)_2$, magnesium hydroxide; HBr, hydrobromic acid
(c) LiOH, lithium hydroxide; HCl, hydrochloric acid
(d) $Fe(OH)_2$, iron(II) hydroxide; H_2CO_3, carbonic acid

15.4 (a) 0.050 *M* Mg^{2+}, 0.10 *M* Cl^-,
(b) 0.070 *M* Al^{3+}, 0.21 *M* Cl^-

15.5 approximately 101.54°C

15.6 (a) 11.41, (b) 2.89, (c) 5.429

15.7 0.0620 *M* HCl

15.8 $3\,H_2S(aq) + 2\,Bi^{3+}(aq) \longrightarrow Bi_2S_3(s) + 6\,H^+(aq)$