

# Acids and Bases

## What You'll Learn

- ▶ You will compare acids and bases and understand why their strengths vary.
- ▶ You will define pH and pOH and calculate the pH and pOH of aqueous solutions.
- ▶ You will calculate acid and base concentrations and determine concentrations experimentally.
- ▶ You will explain how buffers resist changes in pH.

## Why It's Important

Acids and bases are present in the soil of Earth, the foods you eat, the products you buy. Amino acids make up the fabric of every organ in your body and are crucial to your existence.



Visit the Chemistry Web site at [chemistrymc.com](http://chemistrymc.com) to find links about acids and bases.

The color of the big-leaf hydrangea can vary from pink to blue depending upon the acidity of the soil in which it is grown.



## DISCOVERY LAB



### Materials

red litmus paper  
blue litmus paper  
microplate  
household products (6-8)  
phenolphthalein

### Investigating What's in Your Cupboards

**Y**ou can learn something about the properties of products in your household by testing them with strips of paper called litmus paper. Can you separate household products into two groups?

#### Safety Precautions



Always wear safety goggles and an apron.

#### Procedure

1. Place three or four drops of each liquid into separate wells of a microplate. Draw a chart to show the position of each liquid.
2. Test each product with litmus paper. Place 2 drops of phenolphthalein in each sample. Record your observations.

#### Analysis

Separate the products into two groups based upon your observations. How do the groups differ? What can you conclude?

## Section

## 19.1

# Acids and Bases: An Introduction

### Objectives

- **Identify** the physical and chemical properties of acids and bases.
- **Classify** solutions as acidic, basic, or neutral.
- **Compare** the Arrhenius and Brønsted-Lowry models of acids and bases.

### Vocabulary

acidic solution  
basic solution  
Arrhenius model  
Brønsted-Lowry model  
conjugate acid  
conjugate base  
conjugate acid-base pair  
amphoteric

When ants sense danger to the ant colony, they emit a substance called formic acid that alerts the entire colony. Acids in rainwater hollow out enormous limestone caverns and destroy valuable buildings and statues. Acids flavor many of the beverages and foods you like, and it's an acid in your stomach that helps digest what you eat. Bases also play a role in your life. The soap you use and the antacid tablet you may take for an upset stomach are bases. Perhaps you have already concluded that the household products you used in the **DISCOVERY LAB** are acids and bases.

### Properties of Acids and Bases

Acids and bases are some of the most important industrial compounds on Earth. In the U.S. alone, industries use 30 to 40 billion kilograms of sulfuric acid each year in the manufacture of products such as plastics, detergents, batteries, and metals.

You are probably already familiar with some of the physical properties of acids and bases. For example, you may know that acidic solutions taste sour. Carbonic and phosphoric acids give many carbonated beverages their sharp taste; citric and ascorbic acids give lemons and grapefruit their mouth-puckering tartness; and acetic acid makes vinegar taste sour. You may also know that basic solutions taste bitter and feel slippery. Just think about the bar of soap that slips from your hand in the shower. **CAUTION:** *You should never attempt to identify an acid or base (or any other substance in the laboratory) by its taste or feel.* The photo on the opposite page shows how the color of a hydrangea depends upon the presence of acids in the soil.

**Figure 19-1**

The blue litmus paper turned pink when it was dipped into the HCl solution, confirming that HCl is an acid. The red litmus paper turned blue after being dipped into the NaOH solution, so NaOH is a base.



Acids can also be identified by their reaction with some metals. Aluminum, magnesium, and zinc react with aqueous solutions of acids to produce hydrogen gas. The reaction between zinc and hydrochloric acid is described by the following equation.



Metal carbonates and hydrogen carbonates also react with aqueous solutions of acids to produce carbon dioxide gas. When vinegar is added to baking soda, a foaming, effervescent reaction occurs between acetic acid ( $\text{HC}_2\text{H}_3\text{O}_2$ ) dissolved in the vinegar solution, and sodium hydrogen carbonate ( $\text{NaHCO}_3$ ). The production of  $\text{CO}_2$  gas accounts for the effervescence.



Geologists identify rocks as limestone ( $\text{CaCO}_3$ ) by using an HCl solution. If a few drops of the acid produce bubbles of carbon dioxide, the rock is limestone.

Practice!

For more practice writing equations for acid reactions, go to

**Supplemental Practice Problems** in Appendix A.

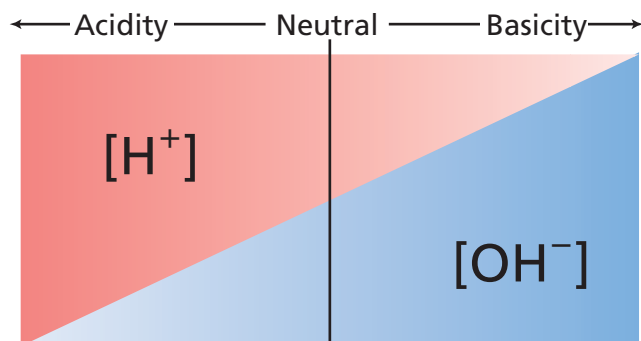
## PRACTICE PROBLEMS

1. Write a balanced formula equation for the reaction that occurs between each of the following pairs of reactants.
  - a. magnesium and nitric acid
  - b. aluminum and sulfuric acid
  - c. calcium carbonate and hydrobromic acid
  - d. potassium hydrogen carbonate and hydrochloric acid

The litmus in litmus paper is one of the dyes commonly used to distinguish solutions of acids and bases, as shown in **Figure 19-1**. Aqueous solutions of acids cause blue litmus paper to turn pink. Aqueous solutions of bases cause red litmus paper to turn blue. With this information you can now identify the two groups of household products you used in the **DISCOVERY LAB**.

Another property of acid and base solutions is their ability to conduct electricity. Pure water is a nonconductor of electricity, but the addition of an acid or base to water causes the resulting solution to become a conductor.

**Ions in solution** Why are some aqueous solutions acidic, others basic, and still others neutral? Neutral solutions are neither acidic nor basic. Scientists have learned that all water (aqueous) solutions contain hydrogen ions ( $\text{H}^+$ )



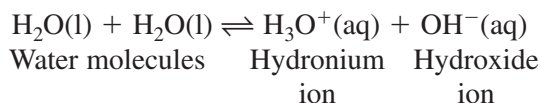
**Figure 19-2**

Note how  $[H^+]$  and  $[OH^-]$  change simultaneously. As  $[H^+]$  decreases to the right,  $[OH^-]$  increases to the right. At what point in the diagram are the two ion concentrations equal?

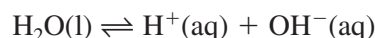
and hydroxide ions ( $OH^-$ ). The relative amounts of the two ions determine whether an aqueous solution is acidic, basic, or neutral.

An **acidic solution** contains more hydrogen ions than hydroxide ions. A **basic solution** contains more hydroxide ions than hydrogen ions. What do you think a neutral solution contains? You are correct if you said a neutral solution contains equal concentrations of hydrogen ions and hydroxide ions. These relationships are illustrated in **Figure 19-2**.

The usual solvent for acids and bases is water. Water produces equal numbers of  $H^+$  ions and  $OH^-$  ions in a process known as self-ionization. In self-ionization, two water molecules react to form a hydronium ion ( $H_3O^+$ ) and a hydroxide ion according to this equilibrium.



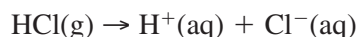
The hydronium ion is a hydrated hydrogen ion, which means that a water molecule is attached to a hydrogen ion by a covalent bond. However, the symbols  $H^+$  and  $H_3O^+$  can be used interchangeably in chemical equations to represent a hydrogen ion in aqueous solution. Thus, a simplified version of the equation for the self-ionization of water is



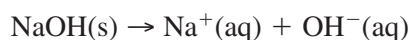
From the equation, you can infer that pure water is neutral because equal numbers of  $H^+$  ions and  $OH^-$  ions are always present.

**The Arrhenius model of acids and bases** If pure water itself is neutral, how does an aqueous solution become acidic or basic? The first person to answer this question was the Swedish chemist Svante Arrhenius, who in 1883 proposed what is now called the Arrhenius model of acids and bases. The **Arrhenius model** states that an acid is a substance that contains hydrogen and ionizes to produce hydrogen ions in aqueous solution. A base is a substance that contains a hydroxide group and dissociates to produce a hydroxide ion in aqueous solution. Some household acids and bases are shown in **Figure 19-3**.

As an example of the Arrhenius model of acids and bases, consider what happens when hydrogen chloride gas dissolves in water.  $HCl$  molecules ionize to form  $H^+$  ions, which make the solution acidic.



When the ionic compound  $NaOH$  dissolves in water, it dissociates to produce  $OH^-$  ions, which make the solution basic.



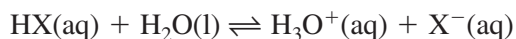
**Figure 19-3**

You probably have some of these products in your home. Examine their labels and the labels of other household products and list the acids and bases you find.

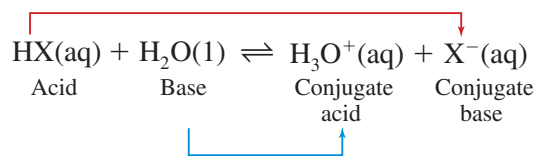
Although the Arrhenius model is useful in explaining many acidic and basic solutions, it has some shortcomings. For example, ammonia (NH<sub>3</sub>) does not contain a hydroxide group, yet ammonia produces hydroxide ions in solution and is a well known base. Clearly, a model that includes all bases is needed.

**The Brønsted-Lowry model** The Danish chemist Johannes Brønsted and the English chemist Thomas Lowry independently proposed a more inclusive model of acids and bases—a model that focuses on the hydrogen ion (H<sup>+</sup>). In the **Brønsted-Lowry model** of acids and bases, an acid is a hydrogen-ion donor and a base is a hydrogen-ion acceptor.

What does it mean to be a hydrogen-ion donor or a hydrogen-ion acceptor? The symbols X and Y may be used to represent nonmetallic elements or negative polyatomic ions. Thus the general formula for an acid can be written as HX or HY. When a molecule of acid, HX, dissolves in water, it donates a H<sup>+</sup> ion to a water molecule. The water molecule acts as a base and accepts the H<sup>+</sup> ion.



On accepting the H<sup>+</sup> ion, the water molecule becomes an acid H<sub>3</sub>O<sup>+</sup>. The hydronium ion (H<sub>3</sub>O<sup>+</sup>) is an acid because it has an extra H<sup>+</sup> ion that it can donate. On donating its H<sup>+</sup> ion, the acid HX becomes a base, X<sup>-</sup>. Why? You are correct if you said X<sup>-</sup> is a base because it has a negative charge and can readily accept a positive hydrogen ion. Thus, an acid-base reaction in the reverse direction can occur. The acid H<sub>3</sub>O<sup>+</sup> can react with the base X<sup>-</sup> to form water and HX and the following equilibrium is established.



The forward reaction is the reaction of an acid and a base. The reverse reaction also is the reaction of an acid and a base. The acid and base that react in the reverse reaction are identified under the equation as a conjugate acid and a conjugate base. A **conjugate acid** is the species produced when a base accepts a hydrogen ion from an acid. The base H<sub>2</sub>O accepts a hydrogen ion from the acid HX and becomes the conjugate acid, H<sub>3</sub>O<sup>+</sup>. A **conjugate base** is the species that results when an acid donates a hydrogen ion to a base. The acid HX donates its hydrogen ion and becomes the conjugate base X<sup>-</sup>. In the reaction above, the hydronium ion (H<sub>3</sub>O<sup>+</sup>) is the conjugate acid of the base H<sub>2</sub>O. The X<sup>-</sup> ion is the conjugate base of the acid HX. Every Brønsted-Lowry interaction involves conjugate acid-base pairs. A **conjugate acid-base pair** consists of two substances related to each other by the donating and accepting of a single hydrogen ion.

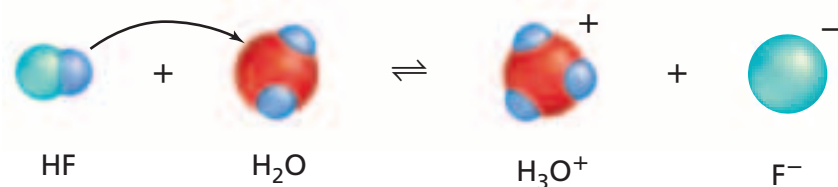
An analogy for conjugate acid-base pairs is shown in **Figure 19-4**. When you have the footbag (Hacky Sack), you are an acid. You pass the footbag (hydrogen ion) to your friend. Now your friend is an acid because she has the footbag (hydrogen ion) to give away and you are a base because you are able to accept the footbag (hydrogen ion). You, with the footbag, and your friend are the acid and base in the forward reaction. In the reverse reaction, your friend with the footbag is the conjugate acid and you are the conjugate base.



**Figure 19-4**

In a game of freestyle footbag, the players stand in a circle. If the footbag is a hydrogen ion, which person is an acid? If the footbag is passed to another person, which person is the acid? Is the other person a conjugate base or a conjugate acid?

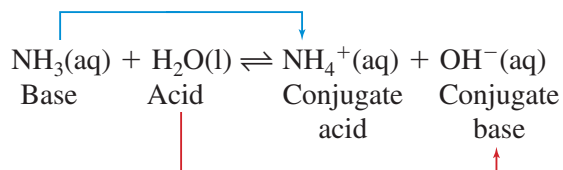
Now, consider the equation for the ionization of hydrogen fluoride in water. According to the Brønsted-Lowry definition, the equation is written this way.



What are the conjugate acid-base pairs? Hydrogen fluoride (HF), the acid in the forward reaction, produces its conjugate base  $\text{F}^-$ , the base in the reverse reaction. Water, the base in the forward reaction, produces its conjugate acid  $\text{H}_3\text{O}^+$ , the acid in the reverse reaction.

Hydrogen fluoride is an acid according to both the Arrhenius and Brønsted-Lowry definitions. All of the acids and bases that fit the Arrhenius definition of acids and bases also fit the Brønsted-Lowry definition. But what about bases such as ammonia, that cannot be considered bases according to the Arrhenius definition because they lack a hydroxide group? Does the Brønsted-Lowry model explain why they are bases?

When ammonia dissolves in water, water is a Brønsted-Lowry acid in the forward reaction. Because the  $\text{NH}_3$  molecule accepts a  $\text{H}^+$  ion to form the ammonium ion ( $\text{NH}_4^+$ ), ammonia is a Brønsted-Lowry base in the forward reaction.



Now look at the reverse reaction. The ammonium ion gives up a  $\text{H}^+$  ion to form the molecule ammonia and thus acts as a Brønsted-Lowry acid. The ammonium ion is the conjugate acid of the base ammonia. The hydroxide ion accepts a  $\text{H}^+$  ion to form a water molecule and is thus a Brønsted-Lowry base. The hydroxide ion is the conjugate base of the acid water.

Recall that when HF dissolves in water, water acts a base; when  $\text{NH}_3$  dissolves in water, water acts as an acid. Depending upon what other substances are in the solution, water can act as either an acid or a base. Water and other substances that can act as both acids and bases are said to be **amphoteric**.

Compare what you have learned about the Arrhenius model and the Brønsted-Lowry model of acids and bases. It should be clear to you that all substances classified as acids and bases by the Arrhenius model are classified as acids and bases by the Brønsted-Lowry model. In addition, some substances *not* classified as bases by the Arrhenius model *are* classified as bases by the Brønsted-Lowry model.

## PRACTICE PROBLEMS

2. Identify the conjugate acid-base pairs in the following reactions.
- $\text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
  - $\text{HBr}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Br}^-(\text{aq})$
  - $\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$
  - $\text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

## Careers Using Chemistry

### Agricultural Technician

*Would you like to work on a large farm, checking soil pH and deciding when to use fertilizers and pesticides? If so, become an agricultural technician.*

Agricultural technicians also work for the government, research groups, and companies that produce fertilizers and other farm chemicals. They know how to modify soil conditions, increase crop yields, help crops resist disease and insects, and store produce. Some agricultural technicians help keep farm animals healthy and productive.



#### Topic: Food Chemistry

To learn more about acids and bases that affect food chemistry, visit the Chemistry Web site at [chemistrymc.com](http://chemistrymc.com)

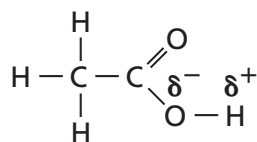
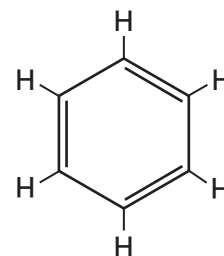
**Activity:** As a class, compile tips for using acids and bases properly in cooking and any chemical equations you can find for the processes that occur in cooking, marinating, and baking.

Practice!

For more practice identifying conjugate acid-base pairs, go to **Supplemental Practice Problems** in Appendix A.

**Figure 19-5**

A partial separation of charge is shown in **a** for the O—H bond and in **b** for the H—F bond. These bonds are polar. In water solution, the  $\text{H}^+$  ionizes, so solutions of HF and ethanoic acid are acidic. The benzene molecule in **c** contains no polar bonds so benzene is not an acid.

**a** Ethanoic acid**b** Hydrogen fluoride**c** Benzene

## Earth Science

### CONNECTION

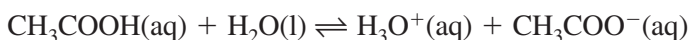
**A**s rain falls, carbon dioxide gas in the air dissolves in the falling water to form a weak acid called carbonic acid. When the acidic rainwater reaches the ground, some water sinks into the pores of the soil to become groundwater. If that groundwater reaches bedrock consisting of limestone, the carbonic acid will slowly dissolve the limestone. Over the course of thousands of years, the limestone is dissolved to the point that huge underground tunnels known as caverns are produced. In many places throughout the cavern, groundwater may drip from the ceiling. As it does so, it deposits some of the dissolved limestone. Thin deposits shaped like icicles on the ceiling are called stalactites. Rounded masses on the floor are called stalagmites. In some cases, stalactites and stalagmites eventually meet in a column or pillar.



## Monoprotic and Polyprotic Acids

You now know that HCl and HF are acids because they can donate a hydrogen ion in an acid-base reaction. From their chemical formulas, you can see that each acid can donate only one hydrogen ion per molecule. An acid that can donate only one hydrogen ion is called a monoprotic acid. Other monoprotic acids are perchloric acid ( $\text{HClO}_4$ ), nitric acid ( $\text{HNO}_3$ ), hydrobromic acid (HBr), and acetic acid ( $\text{CH}_3\text{COOH}$ ). The formula for acetic acid is sometimes written  $\text{HC}_2\text{H}_3\text{O}_2$  and the compound is often called ethanoic acid.

In **Figure 19-5a**, you can see that each ethanoic acid molecule contains four hydrogen atoms. Can ethanoic acid donate more than one hydrogen ion? No; each  $\text{CH}_3\text{COOH}$  molecule contains only one ionizable hydrogen atom.

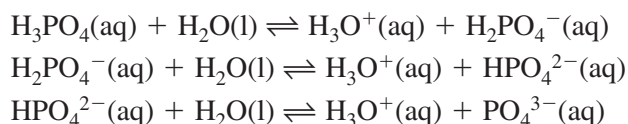


Only one of the four hydrogen atoms in the  $\text{CH}_3\text{COOH}$  molecule can be donated because only those hydrogen atoms bonded to electronegative elements by polar bonds are ionizable.

In an HF molecule, the hydrogen atom is bonded to a fluorine atom, which has the highest electronegativity of all the elements. In **Figure 19-5b**, you can see that the bond linking hydrogen and fluorine is polar; in solution, water facilitates the release of the  $\text{H}^+$  ion. In the  $\text{CH}_3\text{COOH}$  molecule shown in **Figure 19-5a**, three of the four hydrogen atoms are joined to a carbon atom by covalent bonds that are nonpolar because carbon and hydrogen have almost equal electronegativities. Only the hydrogen atom bonded to the electronegative oxygen atom can be released as a  $\text{H}^+$  ion in water. Although a molecule of benzene ( $\text{C}_6\text{H}_6$ ) contains six hydrogen atoms, it is not an acid at all. As **Figure 19-5c** shows, none of the hydrogen atoms are ionizable because they are all joined to carbon atoms by nonpolar bonds.

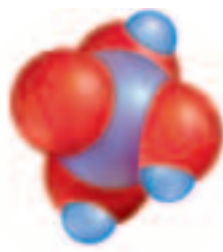
Some acids, however, do donate more than one hydrogen ion. For example, sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and carbonic acid ( $\text{H}_2\text{CO}_3$ ) can donate two hydrogen ions. In each compound, both hydrogen atoms are attached to oxygen atoms by polar bonds. Acids that contain two ionizable hydrogen atoms per molecule are called diprotic acids. In a similar way, phosphoric acid ( $\text{H}_3\text{PO}_4$ ) and boric acid ( $\text{H}_3\text{BO}_3$ ) contain three ionizable hydrogen atoms per molecule. Acids with three hydrogen ions to donate are called triprotic acids. The term polyprotic acid can be used for any acid that has more than one ionizable hydrogen atom. **Figure 19-6** shows models of two polyprotic acids.

All polyprotic acids ionize in steps. The three ionizations of phosphoric acid are described by these equations.





$\text{H}_2\text{SO}_4$   
Sulfuric acid



$\text{H}_3\text{PO}_4$   
Phosphoric acid

**Figure 19-6**

Refer to **Table C-1** in Appendix C for a key to atom color conventions. Then, identify the two hydrogen atoms bonded to oxygen atoms in the sulfuric acid model and the three hydrogen atoms bonded to oxygen atoms in the phosphoric acid model.

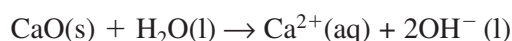
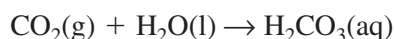
## PRACTICE PROBLEMS

3. Write the steps in the complete ionization of the following polyprotic acids.
- $\text{H}_2\text{Se}$
  - $\text{H}_3\text{AsO}_4$
  - $\text{H}_2\text{SO}_3$



For more practice writing ionization equations for polyprotic acids, go to **Supplemental Practice Problems** in Appendix A.

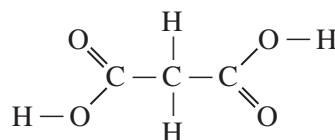
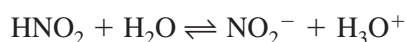
**Anhydrides** Some oxides can become acids or bases by adding the elements contained in water. These compounds are called anhydrides. Oxides of nonmetallic elements, such as carbon, sulfur, or nitrogen, produce an acid in aqueous solution. Oxides of metallic elements usually form basic solutions. For example, carbon dioxide, an oxide of a nonmetal, forms a solution of carbonic acid, whereas, calcium oxide ( $\text{CaO}$ ) forms a basic solution of calcium hydroxide.



Similarly,  $\text{SO}_3$  is the anhydride of  $\text{H}_2\text{SO}_4$  and  $\text{MgO}$  is the anhydride of  $\text{Mg}(\text{OH})_2$ . What acid will form from  $\text{N}_2\text{O}_5$ ?

## Section 19.1 Assessment

- Compare the properties of acidic solutions and basic solutions.
- How do the concentrations of hydrogen ion and hydroxide ion determine whether a solution is acidic, basic, or neutral?
- Based on their formulas, which of the following compounds *may* be Arrhenius acids:  $\text{CH}_4$ ,  $\text{SO}_2$ ,  $\text{H}_2\text{S}$ ,  $\text{Ca}_3(\text{PO}_4)_2$ ? Explain your reasoning.
- Identify the conjugate acid-base pairs in the following equation.
- Thinking Critically** Methylamine ( $\text{CH}_3\text{NH}_2$ ) forms hydroxide ions in aqueous solution. Why is methylamine a Brønsted-Lowry base but not an Arrhenius base?
- Interpreting Scientific Illustrations** In the accompanying structural formula, identify any hydrogen atoms that are likely to be ionizable.



## Objectives

- **Relate** the strength of an acid or base to its degree of ionization.
- **Compare** the strength of a weak acid with the strength of its conjugate base and the strength of a weak base with the strength of its conjugate acid.
- **Explain** the relationship between the strengths of acids and bases and the values of their ionization constants.

## Vocabulary

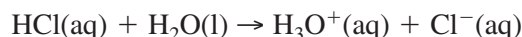
strong acid  
 weak acid  
 acid ionization constant  
 strong base  
 weak base  
 base ionization constant

In the previous section, you learned that one of the properties of acidic and basic solutions is that they conduct electricity. What can electrical conductivity tell you about the hydrogen ions and hydroxide ions in these aqueous solutions?

## Strengths of Acids

To answer this question, suppose you test the electrical conductivities of 0.10M aqueous solutions of hydrochloric acid and acetic acid using a conductivity apparatus similar to the one shown in **Figure 19-7**. When the electrodes are placed in the solutions, both bulbs glow, indicating that both solutions conduct electricity. However, if you compare the brightness of the bulb connected to the hydrochloric acid solution with that of the bulb connected to the acetic acid solution, you can't help but notice a significant difference. The 0.10M HCl solution conducts electricity better than the 0.10M HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> solution. Why is this true if the concentrations of the two acids are both 0.10M?

The answer is that it is ions that carry electricity through the solution and all the HCl molecules that make up the solution are in the form of hydrogen ions and chloride ions. Acids that ionize completely are called **strong acids**. Because strong acids produce the maximum number of ions, they are good conductors of electricity. Other strong acids include perchloric acid (HClO<sub>4</sub>), nitric acid (HNO<sub>3</sub>), hydroiodic acid (HI), sulfuric acid (H<sub>2</sub>SO<sub>4</sub>), and hydrobromic acid (HBr). The ionization of hydrochloric acid in water may be represented by the following equation, which has a single arrow pointing to the right. What does a single arrow to the right mean?



Because hydrochloric acid is virtually 100% ionized, you can consider that the reaction goes to completion and essentially no reaction occurs in the reverse direction.

If the brightly lit bulb of the hydrochloric acid apparatus is due to the large number of ions in solution, then the weakly lit bulb of the acetic acid apparatus must mean that the acetic acid solution has fewer ions. Because the two solutions have the same molar concentrations, you can conclude that acetic acid

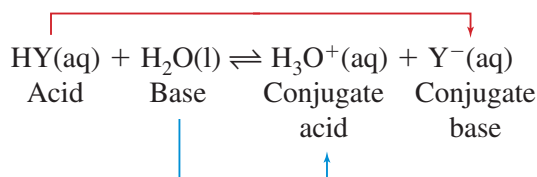
**Figure 19-7**

Both solutions have the same molar concentration, yet HCl produces a brighter light than HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>. Because the amount of electrical current depends upon the number of ions in solution, the HCl solution must contain more ions.





greater attraction for the hydrogen ion? In the case of all strong acids, water is the stronger base. How does the situation differ for the weak acid, HY?



The ionization equilibrium for a weak acid lies far to the left because the conjugate base Y<sup>-</sup> has a greater attraction for the H<sup>+</sup> ion than does the base H<sub>2</sub>O. In the battle of the bases, the conjugate base Y<sup>-</sup> (in the reverse reaction) proves stronger than the base H<sub>2</sub>O (in the forward reaction) and essentially manages to capture the hydrogen ion.

Although the Brønsted-Lowry model helps explain acid strength, the model does not provide a quantitative way to express the strength of an acid or to compare the strengths of various acids. The equilibrium constant expression provides the quantitative measure of acid strength.

## miniLAB

### Acid Strength

**Observing and inferring** The electrical conductivities of solutions of weak acids, such as acetic acid, are related to the degree of ionization of the acid.

**Materials** glacial acetic acid; distilled water; 10-mL graduated cylinder; dropping pipette; 50-mL beaker; 24-well micro plate; conductivity tester with battery; stirring rod

#### Procedure

1. Use a 10-mL graduated cylinder to measure 3 mL of glacial acetic acid. Use a dropping pipette to transfer the 3 mL of glacial acetic acid into well A1 of a 24-well micro plate.
2. Lower the electrodes of a conductivity tester into the glacial acetic acid in well A1. Record your results.
3. Rinse the graduated cylinder with water. Prepare a 6.0M solution of acetic acid by adding 3.4 mL of glacial acetic acid to 6.6 mL of distilled water in the 10-mL graduated cylinder.
4. Empty the 10 mL of diluted acid into a 50-mL beaker. After mixing, transfer 3 mL of the 6.0M acetic acid into well A2. Save the remaining 6.0M acetic acid for procedure step 5. Test and record the conductivity of the solution.
5. Prepare a 1.0M acetic acid solution by adding 1.7 mL of 6.0M acetic acid to 8.3 mL of distilled water in the 10-mL graduated cylinder. Empty

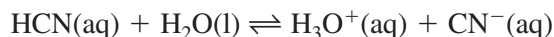
the 10 mL of diluted acid into the rinsed 50-mL beaker. After mixing, transfer 3 mL of the 1.0M acetic acid into well A3. Save the remaining 1.0M acetic acid for procedure step 6. Test and record the conductivity of the solution.

6. Prepare a 0.1M acetic acid solution by adding 1.0 mL of 1.0M acetic acid to 9.0 mL of distilled water in the rinsed 10-mL graduated cylinder. Empty the 10 mL of diluted acid into the rinsed 50-mL beaker. After mixing, transfer 3 mL of the 0.1M acetic acid into well A4. Test and record the conductivity of the solution.

#### Analysis

1. Write the equation for the ionization of acetic acid in water and the equilibrium constant expression. ( $K_{\text{eq}} = 1.8 \times 10^{-5}$ ) What does the size of  $K_{\text{eq}}$  indicate about the degree of ionization of acetic acid?
2. Do the following approximate percents ionization fit your laboratory results: glacial acetic acid, 0.1%; 6.0M acetic acid, 0.2%; 1.0M acetic acid, 0.4%; 0.1M acetic acid, 1.3%? Explain.
3. State a hypothesis that will explain your observations and incorporate your answer to Question 2.
4. Based on your hypothesis, what can you conclude about the need to use large amounts of water for rinsing when acid spills on living tissue?

**Acid ionization constants** As you have learned, a weak acid produces an equilibrium mixture of molecules and ions in aqueous solution. Thus, the equilibrium constant,  $K_{\text{eq}}$ , provides a quantitative measure of the degree of ionization of the acid. Consider hydrocyanic acid (HCN), a deadly poison with applications in the steel industry and in the processing of metal ores. See **Figure 19-9**. The ionization equation and equilibrium constant expression for hydrocyanic acid are



$$K_{\text{eq}} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}][\text{H}_2\text{O}]}$$

The concentration of liquid water in the denominator of the expression is considered to be constant in dilute aqueous solutions, so it can be combined with  $K_{\text{eq}}$  to give a new equilibrium constant,  $K_{\text{a}}$ .

$$K_{\text{eq}} [\text{H}_2\text{O}] = K_{\text{a}} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = 6.2 \times 10^{-10}$$

$K_{\text{a}}$  is called the acid ionization constant. The **acid ionization constant** is the value of the equilibrium constant expression for the ionization of a weak acid. Like all equilibrium constants, the value of  $K_{\text{a}}$  indicates whether reactants or products are favored at equilibrium. For weak acids, the concentrations of the ions (products) in the numerator tend to be small compared to the concentration of un-ionized molecules (reactant) in the denominator. The weakest acids have the smallest  $K_{\text{a}}$  values because their solutions have the lowest concentrations of ions and the highest concentrations of un-ionized acid molecules.  $K_{\text{a}}$  values and ionization equations for several weak acids are listed in **Table 19-2**. Note that for polyprotic acids there is a  $K_{\text{a}}$  value for each ionization and the values decrease for each successive ionization.

**Table 19-2**

Ionization Constants for Weak Acids at 25°C		
Acid	Ionization equation	$K_{\text{a}}$ (298 K)
Hydrosulfuric	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$	$8.9 \times 10^{-8}$
	$\text{HS}^- \rightleftharpoons \text{H}^+ + \text{S}^{2-}$	$1 \times 10^{-19}$
Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$	$6.3 \times 10^{-4}$
Methanoic (Formic)	$\text{HCOOH} \rightleftharpoons \text{H}^+ + \text{HCOO}^-$	$1.8 \times 10^{-4}$
Ethanoic (Acetic)	$\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$	$1.8 \times 10^{-5}$
Carbonic	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$	$4.5 \times 10^{-7}$
	$\text{HCO}_3^- \rightleftharpoons \text{H}^+ + \text{CO}_3^{2-}$	$4.7 \times 10^{-11}$
Hypochlorous	$\text{HClO} \rightleftharpoons \text{H}^+ + \text{ClO}^-$	$4.0 \times 10^{-8}$

## PRACTICE PROBLEMS

**10.** Write ionization equations and acid ionization constant expressions for the following acids.

- $\text{HClO}_2$
- $\text{HNO}_2$
- $\text{HIO}$



**Figure 19-9**

Mine tailings are crushed rock discarded as waste. Hydrocyanic acid and its compounds can be used to extract useable materials from mine tailings.

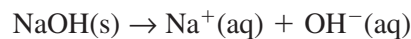


For more practice writing acid ionization constant expressions, go to **Supplemental Practice Problems** in Appendix A.



## Strengths of Bases

What you have learned about acids can be applied to bases except that  $\text{OH}^-$  ions rather than  $\text{H}^+$  ions are involved. For example, the conductivity of a base depends upon the extent to which the base produces hydroxide ions in aqueous solution. **Strong bases** dissociate entirely into metal ions and hydroxide ions. Therefore, metallic hydroxides, such as sodium hydroxide, are strong bases.

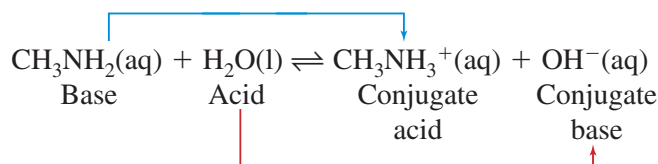


Some metallic hydroxides such as calcium hydroxide have low solubility and thus are poor sources of  $\text{OH}^-$  ions. Note that the solubility product constant,  $K_{\text{sp}}$ , for  $\text{Ca(OH)}_2$  is small, indicating that few hydroxide ions are present in a saturated solution.



Nevertheless, calcium hydroxide and other slightly soluble metallic hydroxides are considered strong bases because all of the compound that dissolves is completely dissociated. The dissociation equations for several strong bases are listed in **Table 19-3**.

In contrast to strong bases, a **weak base** ionizes only partially in dilute aqueous solution to form the conjugate acid of the base and hydroxide ion. The weak base methylamine ( $\text{CH}_3\text{NH}_2$ ) reacts with water to produce an equilibrium mixture of  $\text{CH}_3\text{NH}_2$  molecules,  $\text{CH}_3\text{NH}_3^+$  ions, and  $\text{OH}^-$  ions.



This equilibrium lies far to the left because the base,  $\text{CH}_3\text{NH}_2$ , is weak and the conjugate base,  $\text{OH}^-$  ion, is strong. The hydroxide ion has a much greater attraction for a hydrogen ion than a molecule of methyl amine has.

**Base ionization constants** You won't be surprised to learn that like weak acids, weak bases also form equilibrium mixtures of molecules and ions in aqueous solution. Therefore, the equilibrium constant provides a measure of the extent of the base's ionization. The equilibrium constant for the ionization of methylamine in water is defined by this equilibrium constant expression.

$$K_{\text{b}} = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$$

The constant  $K_{\text{b}}$  is called the base ionization constant. The **base ionization constant** is the value of the equilibrium constant expression for the ionization of a base. The smaller the value of  $K_{\text{b}}$ , the weaker the base.  $K_{\text{b}}$  values and ionization equations for several weak bases are listed in **Table 19-4**.

**Table 19-3**

Common Strong Bases
$\text{NaOH(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$
$\text{KOH(s)} \rightarrow \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$
$\text{RbOH(s)} \rightarrow \text{Rb}^+(\text{aq}) + \text{OH}^-(\text{aq})$
$\text{CsOH(s)} \rightarrow \text{Cs}^+(\text{aq}) + \text{OH}^-(\text{aq})$
$\text{Ca(OH)}_2(\text{s}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$
$\text{Ba(OH)}_2(\text{s}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$

Practice!

For more practice writing base ionization constant expressions, go to **Supplemental Practice Problems** in Appendix A.

## PRACTICE PROBLEMS

- 11.** Write ionization equations and base ionization constant expressions for the following bases.
- hexylamine ( $\text{C}_6\text{H}_{13}\text{NH}_2$ )
  - propylamine ( $\text{C}_3\text{H}_7\text{NH}_2$ )
  - carbonate ion ( $\text{CO}_3^{2-}$ )
  - hydrogen sulfite ion ( $\text{HSO}_3^-$ )

Table 19-4

Ionization Constants of Weak Bases		
Base	Ionization equation	$K_b$ (298 K)
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$5.0 \times 10^{-4}$
Methylamine	$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$4.3 \times 10^{-4}$
Ammonia	$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$	$2.5 \times 10^{-5}$
Aniline	$\text{C}_6\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$4.3 \times 10^{-10}$

Figure 19-10

The photos show a strong acid and a weak acid, a strong base and a weak base. Can you distinguish them? Which of the acids is more concentrated? Which is stronger? Which of the bases is more concentrated? Which is stronger?

**Strong or weak, concentrated or dilute** You have been learning about acids and bases that are often described as weak or strong. When you use the words, weak and strong, do they mean the same as when you describe your cup of tea as being weak or strong? No, there is a difference. When talking about your tea, you could substitute the words *dilute* and *concentrated* for weak and strong. But in talking about acids and bases, the words *weak* and *dilute* have different meanings. Similarly, the words *strong* and *concentrated* are not interchangeable. The terms *dilute* and *concentrated* refer to the number of the acid or base molecules dissolved in a volume of solution. The molarity of the solution is a measure of how dilute or concentrated a solution is. The words *weak* or *strong* refer to the degree to which the acid or base separates into ions. You have already learned that solutions of weak acids and bases contain few ions because few of the molecules are ionized. Solutions of strong acids and bases are completely separated into ions. It's possible to have a dilute solution of a strong acid such as hydrochloric acid, or a concentrated solution of a weak acid such as acetic acid. Which is strong, which is more concentrated, a solution of 0.6M hydrochloric acid or a solution of 6M acetic acid? See Figure 19-10.



## Section 19.2 Assessment

- An acid is highly ionized in aqueous solution. Is the acid strong or weak? Explain your reasoning.
- How is the strength of a weak acid related to the strength of its conjugate base?
- Identify the acid-base pairs in the following.
  - $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCOO}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
  - $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
- $K_b$  for aniline is  $4.3 \times 10^{-10}$ . Explain what this tells you about aniline.
- Thinking Critically** Why is a strong base such as sodium hydroxide generally not considered to have a conjugate acid?
- Predicting** Use Table 19-2 to predict which aqueous solution would have the greater electrical conductivity: 0.1M HClO or 0.1M HF. Explain.



### Objectives

- **Explain** the meaning of pH and pOH.
- **Relate** pH and pOH to the ion product constant for water.
- **Calculate** the pH and pOH of aqueous solutions.

### Vocabulary

ion product constant for water

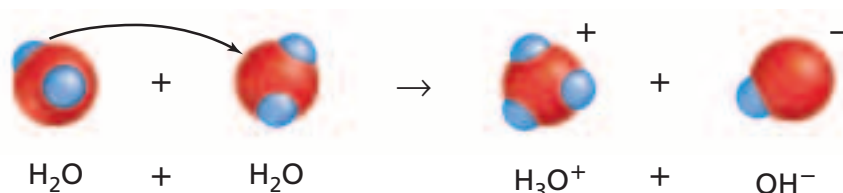
pH

pOH

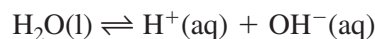
Water not only serves as the solvent in solutions of acids and bases, it also plays a role in the formation of the ions. In aqueous solutions of acids and bases, water sometimes acts as an acid and sometimes as a base. You can think of the self-ionization of water as an example of water assuming the role of an acid and a base in the same reaction.

### Ion Product Constant for Water

Recall from Section 19.1 that pure water contains equal concentrations of  $\text{H}^+$  and  $\text{OH}^-$  ions produced by self-ionization. One molecule of water acts as a Brønsted-Lowry acid and donates a hydrogen ion to a second water molecule. The second molecule of water accepts the hydrogen ion and becomes a hydronium ion. The 1:1 ratio between the products means that equal numbers of hydronium ions and hydroxide ions are formed.



The equation for the equilibrium can be simplified in this way.



The double arrow indicates that this is an equilibrium. Recall that the equilibrium constant expression is written by placing the concentrations of the products in the numerator and the concentrations of the reactants in the denominator. In this example, all terms are to the first power because all the coefficients are 1.

$$K_{\text{eq}} = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

The concentration of pure water is constant so it can be combined with  $K_{\text{eq}}$  by multiplying both sides of the equation by  $[\text{H}_2\text{O}]$ .

$$K_{\text{eq}}[\text{H}_2\text{O}] = K_{\text{w}} = [\text{H}^+][\text{OH}^-]$$

The result is a special equilibrium constant expression that applies only to the self-ionization of water. The constant,  $K_{\text{w}}$ , is called the ion product constant for water. The **ion product constant for water** is the value of the equilibrium constant expression for the self-ionization of water. Experiments show that in pure water at 298 K,  $[\text{H}^+]$  and  $[\text{OH}^-]$  are both equal to  $1.0 \times 10^{-7} M$ . Therefore, at 298 K, the value of  $K_{\text{w}}$  is  $1.0 \times 10^{-14}$ .

$$\begin{aligned} K_{\text{w}} &= [\text{H}^+][\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) \\ K_{\text{w}} &= 1.0 \times 10^{-14} \end{aligned}$$

The product of  $[H^+]$  and  $[OH^-]$  always equals  $1.0 \times 10^{-14}$  at 298 K. This means that if the concentration of  $H^+$  ion increases, the concentration of  $OH^-$  ion must decrease. Similarly, an increase in the concentration of  $OH^-$  ion causes a decrease in the concentration of  $H^+$  ion. You can think about these changes in terms of Le Châtelier's principle, which you learned about in Chapter 18. Adding extra hydrogen ions to the self-ionization of water at equilibrium is a stress on the system. The system reacts in a way to relieve the stress. The added  $H^+$  ions react with  $OH^-$  ions to form more water molecules. Thus, the concentration of  $OH^-$  ion decreases. Example Problem 19-1 shows how you can use  $K_w$  to calculate the concentration of either the hydrogen ion or the hydroxide ion if you know the concentration of the other ion.

## EXAMPLE PROBLEM 19-1

### Using $K_w$ to Calculate $[H^+]$ and $[OH^-]$

At 298 K, the  $H^+$  ion concentration of an aqueous solution is  $1.0 \times 10^{-5}M$ . What is the  $OH^-$  ion concentration in the solution? Is the solution acidic, basic, or neutral?

#### 1. Analyze the Problem

You are given the concentration of  $H^+$  ion and you know that  $K_w$  equals  $1.0 \times 10^{-14}$ . You can use the ion product constant expression to solve for  $[OH^-]$ . Because  $[H^+]$  is greater than  $1.0 \times 10^{-7}$ , you can predict that  $[OH^-]$  will be less than  $1.0 \times 10^{-7}$ .

##### Known

$$[H^+] = 1.0 \times 10^{-5}M$$

$$K_w = 1.0 \times 10^{-14}$$

##### Unknown

$$[OH^-] = ? \text{ mol/L}$$

#### 2. Solve for the Unknown

Write the ion product constant expression.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

Isolate  $[OH^-]$  by dividing both sides of the equation by  $[H^+]$ .

$$[OH^-] = \frac{K_w}{[H^+]}$$

Substitute  $K_w$  and  $[H^+]$  into the expression and solve.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9} \text{ mol/L}$$

Because  $[H^+] > [OH^-]$ , the solution is acidic.

#### 3. Evaluate the Answer

The answer is correctly stated with two significant figures because  $[H^+]$  and  $K_w$  each have two. As predicted, the hydroxide ion concentration,  $[OH^-]$ , is less than  $1.0 \times 10^{-7} \text{ mol/L}$ .



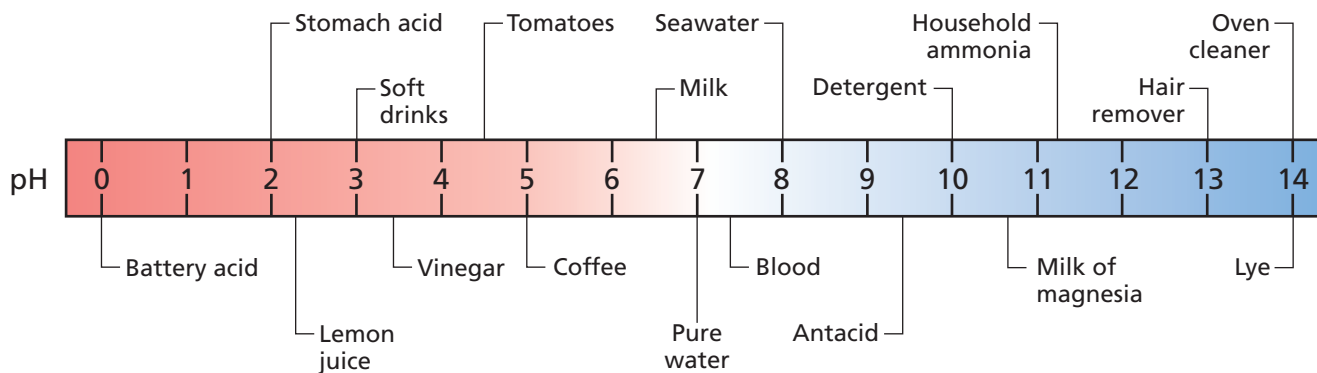
The hydrogen ion and hydroxide ion concentrations of these familiar vegetables are the same as the concentrations calculated in this Example Problem.

## PRACTICE PROBLEMS

18. The concentration of either the  $H^+$  ion or the  $OH^-$  ion is given for three aqueous solutions at 298 K. For each solution, calculate  $[H^+]$  or  $[OH^-]$ . State whether the solution is acidic, basic, or neutral.
- $[H^+] = 1.0 \times 10^{-13}M$
  - $[OH^-] = 1.0 \times 10^{-7}M$
  - $[OH^-] = 1.0 \times 10^{-3}M$



For more practice calculating  $[H^+]$  and  $[OH^-]$  from  $K_w$ , go to **Supplemental Practice Problems** in Appendix A.



**Figure 19-11**

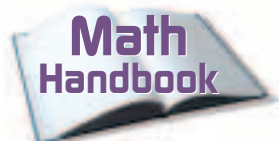
Which has the higher concentration of hydrogen ion, sea water or detergent? How many times higher?

## pH and pOH

As you have probably noted, concentrations of  $\text{H}^+$  ions are often small numbers expressed in exponential notation. Because these numbers are cumbersome, chemists adopted an easier way to express  $\text{H}^+$  ion concentrations using a pH scale based on common logarithms. The **pH** of a solution is the negative logarithm of the hydrogen ion concentration.

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

At 298 K, acidic solutions have pH values below 7. Basic solutions have pH values above 7. Thus, a solution having a pH of 0.0 is strongly acidic; a solution having a pH of 14.0 is strongly basic; and a solution with  $\text{pH} = 7.0$  is neutral. The logarithmic nature of the pH scale means that a change of one pH unit represents a tenfold change in ion concentration. A solution having a pH of 3.0 has ten times the hydrogen ion concentration of a solution with a pH of 4.0. The pH scale and pH values of some common substances are shown in **Figure 19-11**.



Review logarithms in the **Math Handbook** on page 910 of this text.

### EXAMPLE PROBLEM 19-2

#### Calculating pH from $[\text{H}^+]$

What is the pH of a neutral solution at 298 K?

#### 1. Analyze the Problem

You know that in a neutral solution at 298 K,  $[\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$ . You need to find the negative log of  $[\text{H}^+]$ . Because the solution is neutral, you can predict that the pH will be 7.00. You will need a log table or a calculator with a log function.

**Known**

$$[\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$$

**Unknown**

$$\text{pH} = ?$$

#### 2. Solve for the Unknown

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

Substitute  $1.0 \times 10^{-7} \text{ M}$  for  $[\text{H}^+]$  in the equation.

$$\text{pH} = -\log (1.0 \times 10^{-7})$$

$$\text{pH} = -(\log 1.0 + \log 10^{-7})$$

A log table or calculator shows that  $\log 1.0 = 0$  and  $\log 10^{-7} = -7$ . Substitute these numbers in the pH equation.

$$\text{pH} = -[0 + (-7)] = 7.00$$

The pH of the neutral solution at 298 K is 7.00.

### 3. Evaluate the Answer

Values for pH are expressed with as many decimal places as the number of significant figures in the  $H^+$  ion concentration. Thus, the pH is correctly stated with two decimal places. As predicted, the pH value is 7.00.

## PRACTICE PROBLEMS

19. Calculate the pH of solutions having the following ion concentrations at 298 K.
- $[H^+] = 1.0 \times 10^{-2} M$
  - $[H^+] = 3.0 \times 10^{-6} M$
  - $[OH^-] = 8.2 \times 10^{-6} M$



For more practice calculating pH, go to **Supplemental Practice Problems** in Appendix A.

**Using pOH** Sometimes chemists find it convenient to express the basicity, or alkalinity, of a solution on a pOH scale that mirrors the relationship between pH and  $[H^+]$ . The **pOH** of a solution is the negative logarithm of the hydroxide ion concentration.

$$pOH = -\log [OH^-]$$

At 298 K, a solution having a pOH less than 7.0 is basic; a solution having a pOH of 7.0 is neutral; and a solution having a pOH greater than 7.0 is acidic. As with the pH scale, a change of one pOH unit expresses a tenfold change in ion concentration. For example, a solution with a pOH of 2.0 has 100 times the hydroxide ion concentration of a solution with a pOH of 4.0.

A simple relationship between pH and pOH makes it easy to calculate either quantity if the other is known.

$$pH + pOH = 14.00$$

**Figure 19-12** illustrates the relationship between pH and  $[H^+]$  and the relationship between pOH and  $[OH^-]$  at 298 K. Use this diagram as a reference until you become thoroughly familiar with these relationships.

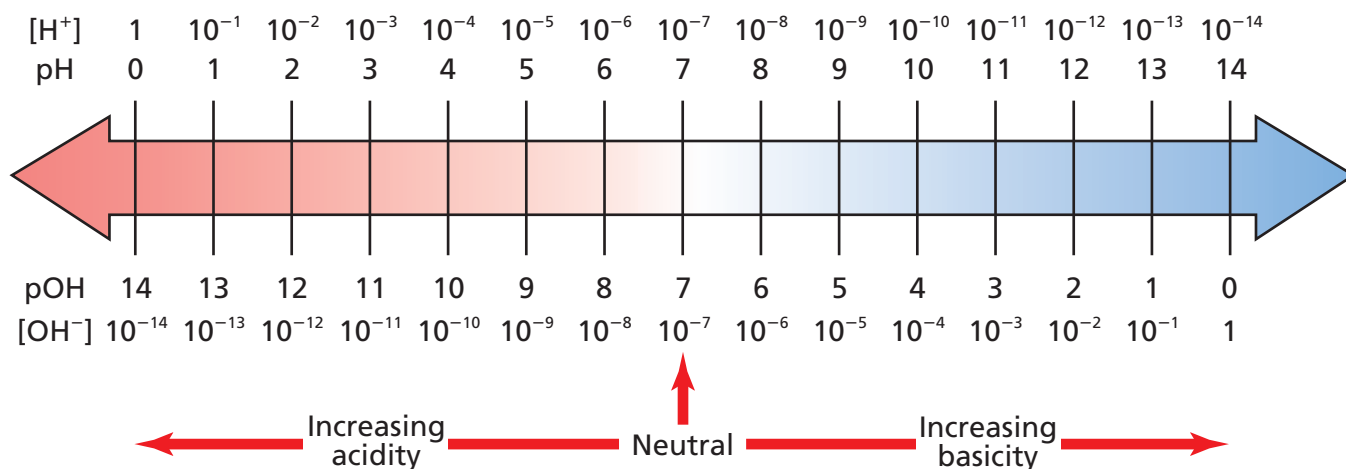


### LAB

See page 961 in Appendix E for **Testing for Ammonia**

**Figure 19-12**

Study this diagram to sharpen your understanding of pH and pOH. Note that at each vertical position, the sum of pH (above the arrow) and pOH (below the arrow) equals 14. Also note that at every position the product of  $[H^+]$  and  $[OH^-]$  equals  $10^{-14}$ .



## EXAMPLE PROBLEM 19-3

### Calculating pOH and pH from $[\text{OH}^-]$

An ordinary household ammonia cleaner is an aqueous solution of ammonia gas with a hydroxide-ion concentration of  $4.0 \times 10^{-3}M$ . Calculate pOH and pH of a typical cleaner at 298 K.

#### 1. Analyze the Problem

You have been given the concentration of hydroxide ion and must calculate pOH and pH. First, you must calculate pOH using the definition of pOH. Then, pH can be calculated using the relationship  $\text{pH} + \text{pOH} = 14.00$ . The negative log of  $10^{-3}$  (the power of 10 of the hydroxide ion concentration) is 3. Therefore, pOH should be close to 3 and pH should be close to 11 or 12.

##### Known

$$[\text{OH}^-] = 4.0 \times 10^{-3}M$$

##### Unknown

$$\text{pOH} = ?$$

$$\text{pH} = ?$$

#### 2. Solve for the Unknown

$$\text{pOH} = -\log [\text{OH}^-]$$

Substitute  $4.0 \times 10^{-3}M$  for  $[\text{OH}^-]$  in the equation.

$$\text{pOH} = -\log (4.0 \times 10^{-3})$$

$$\text{pOH} = -(\log 4.0 + \log 10^{-3})$$

Log tables or your calculator indicate  $\log 4.0 = 0.60$  and  $\log 10^{-3} = -3$ . Substitute these values in the equation.

$$\text{pOH} = -[0.60 + (-3)] = -(0.60 - 3) = 2.40$$

The pOH of the solution is 2.40.

Solve the equation  $\text{pH} + \text{pOH} = 14.00$  for pH by subtracting pOH from both sides of the equation.

$$\text{pH} = 14.00 - \text{pOH}$$

Substitute the value of pOH.

$$\text{pH} = 14.00 - 2.40 = 11.60$$

The pH of the solution is 11.60.

#### 3. Evaluate the Answer

The values of pH and pOH are correctly expressed with two decimal places because the given concentration has two significant figures. As predicted, pOH is close to 3 and pH is close to 12.



The pH paper shows that this common cleaner containing ammonia has a pH greater than 11.

Practice!

For more practice calculating pH and pOH from  $[\text{OH}^-]$ , go to **Supplemental Practice Problems** in Appendix A.

## PRACTICE PROBLEMS

**20.** Calculate the pH and pOH of aqueous solutions having the following ion concentrations.

a.  $[\text{OH}^-] = 1.0 \times 10^{-6}M$

b.  $[\text{OH}^-] = 6.5 \times 10^{-4}M$

c.  $[\text{H}^+] = 3.6 \times 10^{-9}M$

d.  $[\text{H}^+] = 0.025M$

**Calculating ion concentrations from pH** Suppose the pH of a solution is 3.50 and you must determine the concentrations of  $\text{H}^+$  and  $\text{OH}^-$ . The definition of pH relates pH and  $\text{H}^+$  ion concentration and can be solved for  $[\text{H}^+]$ .

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

First you need to multiply both sides of the equation by  $-1$ .

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

To calculate  $[\text{H}^+]$  using this equation you must take the antilog of both sides of the equation.

$$\text{antilog} (-\text{pH}) = [\text{H}^+]$$

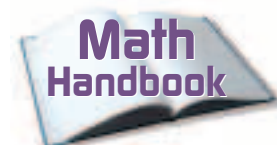
To calculate  $[\text{H}^+]$ , substitute 3.50 for pH in the equation.

$$\text{antilog} (-3.50) = [\text{H}^+]$$

Use a log table or your calculator to determine the antilog of  $-3.50$ . The antilog is  $3.2 \times 10^{-4}$ .

$$[\text{H}^+] = 3.2 \times 10^{-4} \text{ mol/L}$$

You can calculate  $[\text{OH}^-]$  using the relationship  $[\text{OH}^-] = \text{antilog} (-\text{pOH})$ .



Review antilogs in the **Math Handbook** on page 910 of this text.

## EXAMPLE PROBLEM 19-4

### Calculating $[\text{H}^+]$ and $[\text{OH}^-]$ from pH

What are  $[\text{H}^+]$  and  $[\text{OH}^-]$  in a healthy person's blood that has a pH of 7.40? Assume that the temperature is 298 K.

#### 1. Analyze the Problem

You have been given the pH of a solution and must calculate  $[\text{H}^+]$  and  $[\text{OH}^-]$ . You can obtain  $[\text{H}^+]$  using the equation that defines pH. Then, subtract the pH from 14.00 to obtain pOH. The pH is close to 7 but greater than 7, so  $[\text{H}^+]$  should be slightly less than  $10^{-7}$  and  $[\text{OH}^-]$  should be greater than  $10^{-7}$ .

##### Known

$$\text{pH} = 7.40$$

##### Unknown

$$[\text{H}^+] = ? \text{ mol/L}$$

$$[\text{OH}^-] = ? \text{ mol/L}$$

#### 2. Solve for the Unknown

Write the equation that defines pH and solve for  $[\text{H}^+]$ .

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

$$[\text{H}^+] = \text{antilog} (-\text{pH})$$

Substitute the known value of pH.

$$[\text{H}^+] = \text{antilog} (-7.40)$$

Use a log table or your calculator to find the antilog. The antilog of  $-7.40$  is  $4.0 \times 10^{-8}$ .

$$[\text{H}^+] = 4.0 \times 10^{-8} \text{ M}$$

The concentration of hydrogen ion in blood is  $4.0 \times 10^{-8} \text{ M}$ .

To determine  $[\text{OH}^-]$ , calculate pOH using the equation

$$\text{pH} + \text{pOH} = 14.00.$$

Solve for pOH by subtracting pH from both sides of the equation.

$$\text{pOH} = 14.00 - \text{pH}$$

Substitute the known value of pH.

$$\text{pOH} = 14.00 - 7.40 = 6.60$$

Substitute 6.60 for pOH in the equation  $[\text{OH}^-] = \text{antilog} (-\text{pOH})$

$$[\text{OH}^-] = \text{antilog} (-6.60)$$

*Continued on next page*



Blood banks collect blood from healthy people to hold in reserve for persons who need transfusions.

Use a log table or your calculator to find the antilog. The antilog of  $-6.60$  is  $2.5 \times 10^{-7}$ .

$$[\text{OH}^-] = 2.5 \times 10^{-7} M.$$

The concentration of hydroxide ion in blood is  $2.5 \times 10^{-7} M$ .

### 3. Evaluate the Answer

The concentrations of the hydrogen ion and hydroxide ion are correctly stated with two significant figures because the given pH has two decimal places. As predicted,  $[\text{H}^+]$  is less than  $10^{-7}$  and  $[\text{OH}^-]$  is greater than  $10^{-7}$ .

Practice!

For more practice calculating  $[\text{H}^+]$  and  $[\text{OH}^-]$  from pH, go to

**Supplemental Practice Problems** in Appendix A.

## PRACTICE PROBLEMS

21. The pH is given for three solutions. Calculate  $[\text{H}^+]$  and  $[\text{OH}^-]$  in each solution.

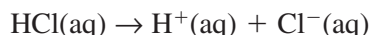
a. pH = 2.37

c. pH = 6.50

b. pH = 11.05

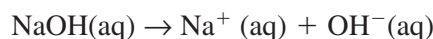
### Calculating the pH of solutions of strong acids and strong bases

Look at the bottles of acid and base solutions in **Figure 19-13**. They are labeled with the number of moles of molecules or formula units that were dissolved in a liter of water ( $M$ ) when the solutions were made. Each of the bottles contains a strong acid or base. Recall from Section 19.2 that strong acids and bases are essentially 100% ionized. That means that this reaction for the ionization of HCl goes to completion.



Every HCl molecule produces one  $\text{H}^+$  ion. The bottle labeled  $0.1M$  HCl contains  $0.1$  mole of  $\text{H}^+$  ions per liter and  $0.1$  mole of  $\text{Cl}^-$  ions per liter. For all strong monoprotic acids, the concentration of the acid is the concentration of  $\text{H}^+$  ion. Thus, you can use the concentration of the acid for calculating pH.

Similarly, the  $0.1M$  solution of the strong base NaOH in **Figure 19-13** is fully ionized.



One formula unit of NaOH produces one  $\text{OH}^-$  ion. Thus, the concentration of the hydroxide ion is the same as the concentration of the solution,  $0.1M$ .

Some strong bases contain two or more hydroxide ions in each formula unit. Calcium hydroxide ( $\text{Ca}(\text{OH})_2$ ) is an example. The concentration of hydroxide ion in a solution of  $\text{Ca}(\text{OH})_2$  is twice the concentration of the ionic compound. Thus, the concentration of  $\text{OH}^-$  in a  $7.5 \times 10^{-4}M$  solution of  $\text{Ca}(\text{OH})_2$  is  $7.5 \times 10^{-4}M \times 2 = 1.5 \times 10^{-3}M$ .



**Figure 19-13**

The label on a bottle of a strong acid or a strong base tells you immediately the concentration of hydrogen ions or hydroxide ions in the solution. That's because, in solution, strong acids and bases exist entirely as ions. What is  $[\text{H}^+]$  in  $0.1M$  HCl? What is  $[\text{OH}^-]$  in  $0.1M$  NaOH?

## PRACTICE PROBLEMS

22. Calculate the pH of the following solutions.

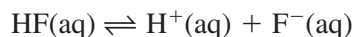
a.  $1.0M$  HI

c.  $1.0M$  KOH

b.  $0.050M$   $\text{HNO}_3$

d.  $2.4 \times 10^{-5}M$   $\text{Mg}(\text{OH})_2$

**Using pH to calculate  $K_a$**  Suppose you measured the pH of a 0.100M solution of the weak acid HF and found it to be 3.20. Would you have enough information to calculate  $K_a$  for HF?



$$K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]}$$

From the pH, you could calculate  $[\text{H}^+]$ . Then, remember that for every mole per liter of  $\text{H}^+$  ion there must be an equal concentration of  $\text{F}^-$  ion. That means that you know two of the variables in the  $K_a$  expression. What about the third,  $[\text{HF}]$ ? The concentration of HF at equilibrium is equal to the initial concentration of the acid (0.100M) minus the moles per liter of HF that dissociated ( $[\text{H}^+]$ ). Example Problem 19-5 illustrates a similar calculation for formic acid.

## EXAMPLE PROBLEM 19-5

### Calculating $K_a$ from pH

The pH of a 0.100M solution of formic acid is 2.38. What is  $K_a$  for HCOOH?

#### 1. Analyze the Problem

You are given the pH of the solution which allows you to calculate the concentration of the hydrogen ion. You know that the concentration of  $\text{HCOO}^-$  equals the concentration of  $\text{H}^+$ . The concentration of un-ionized HCOOH is the difference between the initial concentration of the acid and  $[\text{H}^+]$ .

##### Known

pH = 2.38  
concentration of the solution = 0.100M

##### Unknown

$K_a = ?$

#### 2. Solve for the Unknown

Write the acid ionization constant expression.

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

Use the pH to calculate  $[\text{H}^+]$ .

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

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

Substitute the known value of pH.

$$[\text{H}^+] = \text{antilog}(-2.38)$$

Use a log table or calculator to find the antilog. The antilog of  $-2.38$  is  $4.2 \times 10^{-3}$ .

$$[\text{H}^+] = 4.2 \times 10^{-3}\text{M}$$

$$[\text{HCOO}^-] = [\text{H}^+] = 4.2 \times 10^{-3}\text{M}$$

$[\text{HCOOH}]$  equals the initial concentration minus  $[\text{H}^+]$ .

$$[\text{HCOOH}] = 0.100\text{M} - 4.2 \times 10^{-3}\text{M} = 0.096\text{M}$$

Substitute the known values into the  $K_a$  expression.

$$K_a = \frac{(4.2 \times 10^{-3})(4.2 \times 10^{-3})}{(0.096)} = 1.8 \times 10^{-4}$$

The acid ionization constant for HCOOH is  $1.8 \times 10^{-4}$ .

*Continued on next page*



Natural rubber is an important agricultural export in Southeast Asia. Formic acid is used during the process that converts the milky latex fluid tapped from rubber trees into natural rubber.

### 3. Evaluate the Answer

The calculations are correct and the answer is correctly reported with two significant figures. The  $K_a$  is in the range of reasonably weak acids.

Practice!

For more practice calculating  $K_a$ , go to **Supplemental Practice Problems** in Appendix A.

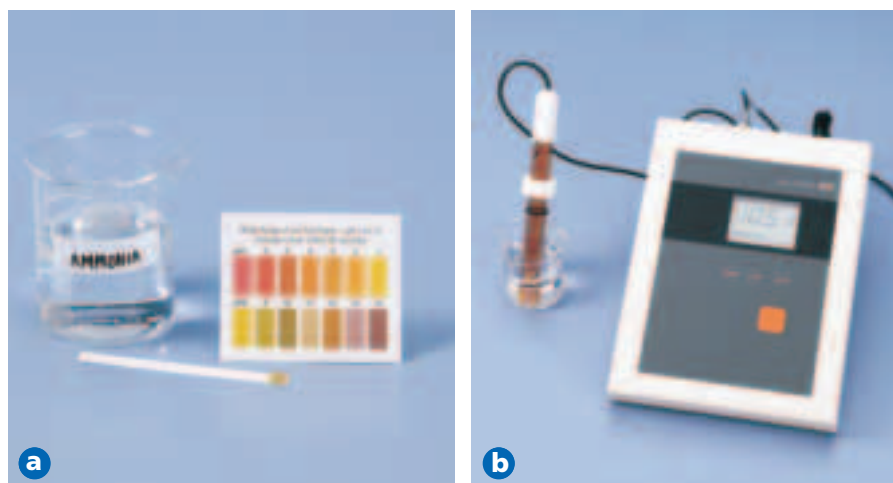
## PRACTICE PROBLEMS

23. Calculate the  $K_a$  for the following acids using the given information.
- 0.220M solution of  $\text{H}_3\text{AsO}_4$ , pH = 1.50
  - 0.0400M solution of  $\text{HClO}_2$ , pH = 1.80

**Measuring pH** Perhaps in an earlier science course, you used indicator paper to measure the pH of a solution. All pH paper is impregnated with one or more substances called indicators that change color depending upon the concentration of hydrogen ion in a solution. When a strip of pH paper is dipped into an acidic or basic solution, the color of the paper changes. To determine the pH, the new color of the paper is compared with standard pH colors on a chart, as shown in **Figure 19-14a**. The pH meter in **Figure 19-14b** provides a more accurate measure of pH. When electrodes are placed in a solution, the meter gives a direct analog or digital readout of pH.

**Figure 19-14**

The approximate pH of a solution can be obtained by wetting a piece of pH paper with the solution and comparing the color of the wet paper with a set of standard colors as shown in **a**. The pH meter in **b** provides a more accurate measurement in the form of a digital display of the pH.



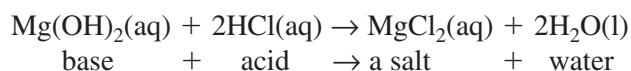
## Section 19.3 Assessment

- What is the relationship between the pH of a solution and the concentration of hydrogen ions in the solution?
- If you know the pOH of a solution, how can you determine its pH?
- How does the ion product constant for water relate to the concentrations of  $\text{H}^+$  and  $\text{OH}^-$  in aqueous solutions?
- Thinking Critically** Why is it logical to assume that the hydrogen ion concentration in an aqueous solution of a strong monoprotic acid equals the molarity of the acid?
- Applying Concepts** Would it be possible to calculate the pH of a weak acid solution if you knew the molarity of the solution and its  $K_a$ ? Explain.

If you were to experience heartburn or indigestion, you might take an antacid to relieve your discomfort. What kind of reaction occurs when magnesium hydroxide, the active ingredient in the common antacid called milk of magnesia, contacts hydrochloric acid produced by the stomach?

## The Reaction Between Acids and Bases

When magnesium hydroxide and hydrochloric acid react, the resulting solution has properties characteristic of neither an acid nor a base. This type of reaction is called a neutralization reaction. A **neutralization reaction** is a reaction in which an acid and a base react in aqueous solution to produce a salt and water. A **salt** is an ionic compound made up of a cation from a base and an anion from an acid. Neutralization is a double-replacement reaction. In the reaction between magnesium hydroxide and hydrochloric acid, magnesium replaces hydrogen in HCl and hydrogen replaces magnesium in  $\text{Mg}(\text{OH})_2$ . The reaction may be described by this balanced formula equation.

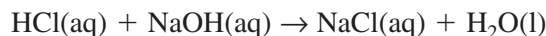


Note that the cation from the base ( $\text{Mg}^{2+}$ ) is combined with the anion from the acid ( $\text{Cl}^-$ ) in the salt  $\text{MgCl}_2$ .

### PRACTICE PROBLEMS

- 29.** Write balanced formula equations for the following acid-base neutralization reactions.
- nitric acid and cesium hydroxide
  - hydrobromic acid and calcium hydroxide
  - sulfuric acid and potassium hydroxide
  - acetic acid and ammonium hydroxide

When considering neutralization reactions, it is important to determine whether all of the reactants and products exist in solution as molecules or formula units. For example, examine the formula equation and complete ionic equation for the reaction between hydrochloric acid and sodium hydroxide.



Because HCl is a strong acid, NaOH a strong base, and NaCl a soluble salt, all three compounds exist as ions in aqueous solution.



The chloride ion and the sodium ion appear on both sides of the equation so they are spectator ions. They can be eliminated to obtain the net ionic equation for the neutralization of a strong acid by a strong base.



## Objectives

- **Write** chemical equations for neutralization reactions.
- **Explain** how neutralization reactions are used in acid-base titrations.
- **Compare** the properties of buffered and unbuffered solutions.

## Vocabulary

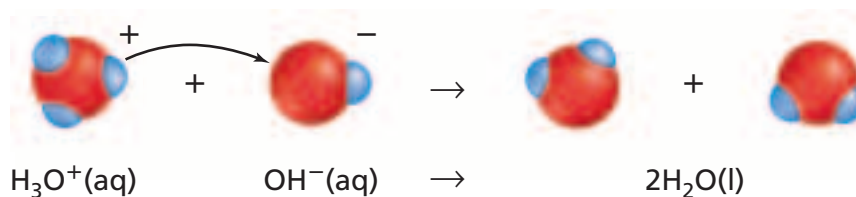
neutralization reaction  
salt  
titration  
equivalence point  
acid-base indicator  
end point  
salt hydrolysis  
buffer  
buffer capacity



For more practice writing neutralization equations, go to **Supplemental Practice Problems** in Appendix A.

**Figure 19-15**

A hydronium ion transfers a hydrogen ion to a hydroxide ion. The loss of the hydrogen ion by  $\text{H}_3\text{O}^+$  results in a water molecule. The gain of a hydrogen ion by  $\text{OH}^-$  also results in a water molecule.

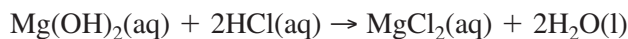


Recall that in an aqueous solution, a  $\text{H}^+$  ion exists as a  $\text{H}_3\text{O}^+$  ion, so the net ionic equation for an acid-base neutralization reaction is

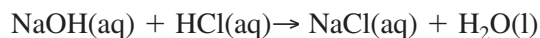


This neutralization reaction is illustrated in **Figure 19-15**. The **How It Works** feature at the end of this chapter shows that this equation does not apply to the neutralization of a strong acid by a weak base.

**Acid-base titration** The stoichiometry of an acid-base neutralization reaction is the same as that of any other reaction that occurs in solution. In the antacid reaction you just read about, one mole of magnesium hydroxide neutralizes two moles of hydrochloric acid.



In the reaction of sodium hydroxide and hydrogen chloride, one mole of sodium hydroxide neutralizes one mole of hydrogen chloride.

**Figure 19-16**

In the titration of an acid by a base, the pH meter measures the pH of the acid solution in the beaker as a solution of base with a known concentration is added from the buret.

Stoichiometry provides the basis for a procedure called titration, which is used to determine the concentrations of acidic and basic solutions. **Titration** is a method for determining the concentration of a solution by reacting a known volume of the solution with a solution of known concentration. If you wished to find the concentration of an acid solution, you would titrate the acid solution with a solution of a base of known concentration. You also could titrate a base of unknown concentration with an acid of known concentration. How is an acid-base titration carried out? **Figure 19-16** illustrates the equipment used for the following titration procedure using a pH meter.

1. A measured volume of an acidic or basic solution of unknown concentration is placed in a beaker. The electrodes of a pH meter are immersed in this solution and the initial pH of the solution is read and recorded.
2. A buret is filled with the titrating solution of known concentration. This solution is called the standard solution.
3. Measured volumes of the standard solution are added slowly and mixed into the solution in the beaker. The pH is read and recorded after each addition. This process continues until the reaction reaches the stoichiometric point, which is the point at which moles of  $\text{H}^+$  ion from the acid equal moles of  $\text{OH}^-$  ion from the base. The stoichiometric point is known as the **equivalence point** of the titration.

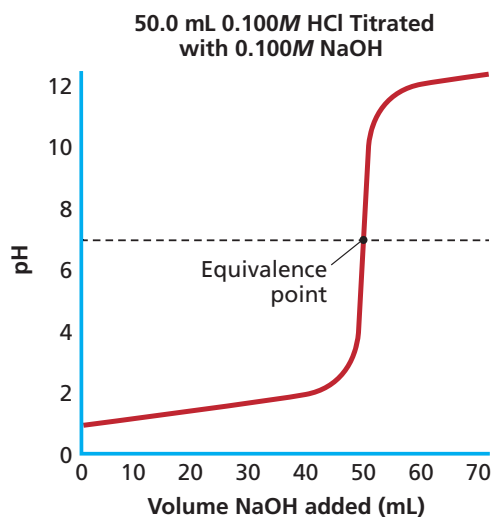
**Figure 19-17** shows how the pH of the solution changes during the titration of 50.0 mL of 0.100M HCl, a strong acid, with 0.100M NaOH, a strong base. The initial pH of the 0.100M HCl is 1.00. As NaOH is added, the acid is neutralized and the solution's pH increases gradually. However, when nearly all of the  $\text{H}^+$  ions from the acid have been used up, the pH increases dramatically with the addition of an exceedingly small volume of NaOH. This

abrupt increase in pH occurs at the equivalence point of the titration. Beyond the equivalence point, the addition of more NaOH again results in a gradual increase in pH.

For convenience, chemists often use a chemical dye rather than a pH meter to detect the equivalence point of an acid-base titration. Chemical dyes whose colors are affected by acidic and basic solutions are called **acid-base indicators**. The chart in **Figure 19-18** shows

the colors of several common acid-base indicators at various positions on the pH scale. The point at which the indicator used in a titration changes color is called the **end point** of the titration. The color change of the indicator selected for an acid-base titration should coincide closely with the equivalence point of the titration because the role of the indicator is to indicate to you, by means of a color change, that just enough of the titrating solution has been added to neutralize the unknown solution. In **Figure 19-18** you can see that bromthymol blue is an excellent choice for an equivalence point near pH 7. Bromthymol blue turns from yellow to blue as the pH of the solution changes from acidic to basic. Thus, the indicator's green transition color can mark an equivalence point near pH 7. What indicator might you choose for a titration that has its equivalence point at pH 5?

You might think that all titrations must have an equivalence point at pH 7 because that's the point at which concentrations of hydrogen ions and hydroxide ions are equal and the solution is neutral. This is not the case, however. Some titrations have equivalence points at  $\text{pH} < 7$  and some have equivalence

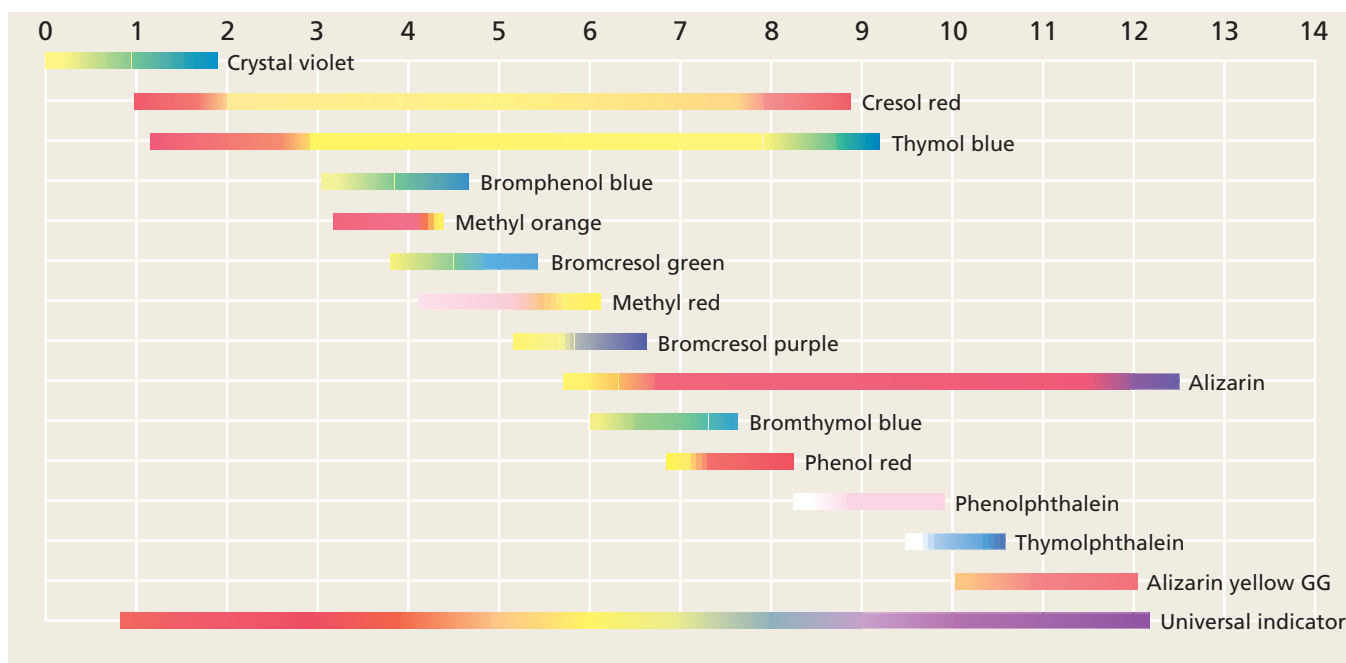


**Figure 19-17**

A steep rise in the pH of the acid solution indicates that all the  $\text{H}^+$  ions from the acid have been neutralized by the  $\text{OH}^-$  ions of the base. The point at which the curve flexes (at its intersection with the dashed line) is the equivalence point of the titration. What is equivalent at this point?

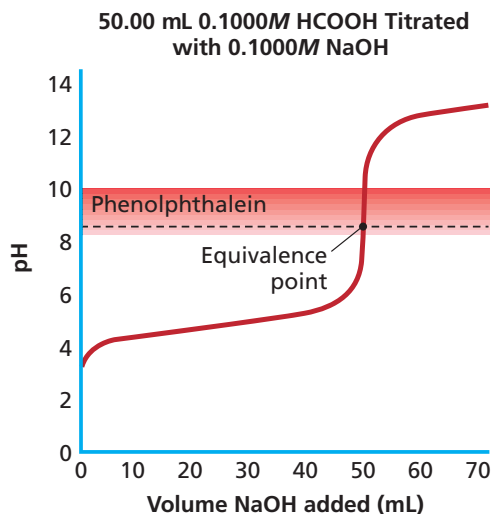
**Figure 19-18**

Choosing the right indicator is important. The indicator must change color at the equivalence point of the titration and the equivalence point is not always at pH 7. Would methyl red be a good choice for the titration graphed in **Figure 19-17**?



**Figure 19-19**

In this titration of a weak acid (HCOOH) by a strong base (NaOH), the equivalence point is not at pH 7. Check **Figure 19-18**. Do you agree that phenolphthalein is the best choice of indicator for this titration? Why?

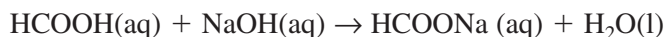


points at pH values  $> 7$ . The pH at the equivalence point of a titration depends upon the relative strengths of the reacting acid and base. **Figure 19-19** shows that the equivalence point for the titration of methanoic acid (a weak acid) with sodium hydroxide (a strong base) lies between pH 8 and pH 9. Therefore, phenolphthalein is a good indicator for this titration. **Figure 19-20** illustrates and describes the titration of a solution of

methanoic acid (HCOOH) of unknown concentration with 0.1000M sodium hydroxide solution.

**Calculating molarity** From the experimental data, the molarity of the unknown HCOOH solution can be calculated by following these steps.

1. Write the balanced formula equation for the acid-base reaction.



2. Calculate the number of moles of NaOH contained in the volume of standard solution added. First, convert milliliters of NaOH to liters by multiplying the volume by a conversion factor that relates milliliters and liters.

$$18.28 \text{ mL NaOH} \times \frac{1 \text{ L NaOH}}{1000 \text{ mL NaOH}} = 0.01828 \text{ L NaOH}$$

Determine the moles of NaOH used by multiplying the volume by a conversion factor that relates moles and liters, the molarity of the solution.

$$0.01828 \text{ L NaOH} \times \frac{0.1000 \text{ mol NaOH}}{1 \text{ L NaOH}} = 1.828 \times 10^{-3} \text{ mol NaOH}$$

**Figure 19-20**

Titration is a precise procedure requiring practice. The white paper under the flask provides a background for viewing the indicator color change.



**a** The buret contains the standard solution (0.1000M NaOH) and the flask contains 25.00 mL HCOOH solution along with a small amount of phenolphthalein indicator.



**b** The standard solution is added slowly to the acid solution in the flask. Although the phenolphthalein indicator turns pink, the color disappears upon mixing until the end point is reached.



**c** The end point of the titration is marked by a permanent but very light pink color. A careful reading of the buret reveals that 18.28 mL 0.1000M NaOH has been added.

3. Use the mole ratio in the balanced equation to calculate the moles of reactant in the unknown solution.

$$1.828 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HCOOH}}{1 \text{ mol NaOH}} = 1.828 \times 10^{-3} \text{ mol HCOOH}$$

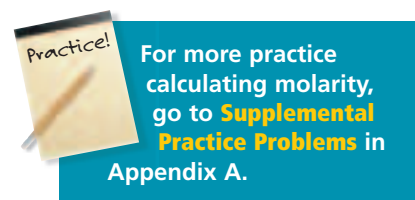
4. Because molarity is defined as moles of solute per liter of solution, calculate the molarity of the unknown solution by dividing the moles of unknown (HCOOH) by the volume of the unknown solution expressed in liters. (25.00 mL = 0.02500 L)

$$M_{\text{HCOOH}} = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{0.02500 \text{ L HCOOH}} = 7.312 \times 10^{-2} M$$

The molarity of the HCOOH solution is  $7.312 \times 10^{-2} M$ , or  $0.07312 M$ . In the **CHEMLAB** at the end of this chapter, you will use titration to standardize a base.

## PRACTICE PROBLEMS

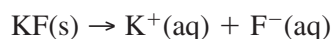
30. What is the molarity of a CsOH solution if 30.0 mL of the solution is neutralized by 26.4 mL 0.250M HBr solution?
31. What is the molarity of a nitric acid solution if 43.33 mL 0.1000M KOH solution is needed to neutralize 20.00 mL of unknown solution?
32. What is the concentration of a household ammonia cleaning solution if 49.90 mL 0.5900M HCl is required to neutralize 25.00 mL solution?



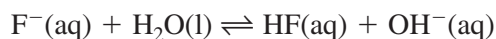
## Salt Hydrolysis

You just learned that a salt is an ionic compound made up of a cation from a base and an anion from an acid. What reaction, if any, occurs when a salt is dissolved in pure water? In **Figure 19-21**, several drops of bromthymol blue indicator solution have been added to 0.10M aqueous solutions of the salts ammonium chloride ( $\text{NH}_4\text{Cl}$ ), sodium nitrate ( $\text{NaNO}_3$ ), and potassium fluoride (KF). Sodium nitrate turns the indicator green, which means that a solution of sodium nitrate is neutral. However, the blue color of the KF solution means that a solution of potassium fluoride is basic, and the yellow color of the ammonium chloride solution indicates that the  $\text{NH}_4\text{Cl}$  solution is acidic. Why are some aqueous salt solutions neutral, some basic, and some acidic?

The answer is that many salts react with water in a process known as salt hydrolysis. In **salt hydrolysis**, the anions of the dissociated salt accept hydrogen ions from water or the cations of the dissociated salt donate hydrogen ions to water. Does a reaction occur when potassium fluoride dissolves in water? Potassium fluoride is the salt of a strong base (KOH) and a weak acid (HF) and dissociates into potassium ions and fluoride ions.



Some fluoride ions react with water molecules to establish this equilibrium.



Note that the fluoride ion acts as a Brønsted-Lowry base and accepts a hydrogen ion from  $\text{H}_2\text{O}$ . Hydrogen fluoride molecules and  $\text{OH}^-$  ions are produced. Although the resulting equilibrium lies far to the left, the potassium fluoride solution is basic because some  $\text{OH}^-$  ions were produced



**Figure 19-21**

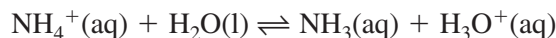
The indicator bromthymol blue provides surprising results when added to three solutions of ionic salts. An ammonium chloride solution is acidic, a sodium nitrate solution is neutral, and a potassium fluoride solution is basic. The explanation has to do with the strengths of the acid and base from which each salt was formed.

making the concentration of  $\text{OH}^-$  ions greater than the concentration of  $\text{H}^+$  ions.

What about ammonium chloride?  $\text{NH}_4\text{Cl}$  is the salt of a weak base ( $\text{NH}_3$ ) and a strong acid ( $\text{HCl}$ ). When dissolved in water, the salt dissociates into ammonium ions and chloride ions.



The ammonium ions then react with water molecules to establish this equilibrium.



The ammonium ion is a Brønsted-Lowry acid in the forward reaction, which produces  $\text{NH}_3$  molecules and hydronium ions. This equilibrium also lies far to the left. However, an ammonium chloride solution is acidic because the solution contains more hydronium ions than hydroxide ions.

The hydrolysis of  $\text{KF}$  produces a basic solution; the hydrolysis of  $\text{NH}_4\text{Cl}$  produces an acidic solution. What about sodium nitrate ( $\text{NaNO}_3$ )? Sodium nitrate is the salt of a strong acid (nitric acid) and a strong base (sodium hydroxide). Little or no salt hydrolysis occurs and a solution of sodium nitrate is neutral.

Practice!

For more practice writing equations for salt hydrolysis reactions, go to **Supplemental Practice Problems** in Appendix A.

## PRACTICE PROBLEMS

- 33.** Write equations for the salt hydrolysis reactions that occur when the following salts are dissolved in water. Classify each solution as acidic, basic, or neutral.
- |                     |                      |
|---------------------|----------------------|
| a. ammonium nitrate | c. potassium sulfate |
| b. rubidium acetate | d. calcium carbonate |

## Buffered Solutions

If you maintain a tropical-fish aquarium similar to the one shown in **Figure 19-22**, you know that the pH of the water must be kept relatively constant if the fish are to survive. Control of pH is important in your body, too. The pH of your blood must be maintained at an average of 7.4. A potentially fatal problem will develop if the pH rises or falls as much as 0.3 pH unit. The gastric juices in your stomach must have a pH between 1.6 and 1.8 to promote digestion of certain foods. How does your body maintain pH values within such narrow limits? It does so by producing buffers.

**Figure 19-22**

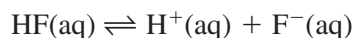
This freshwater aquarium is home to a variety of cichlid. To provide a healthy environment for this fish, it's necessary to know whether it is an African cichlid or a South American cichlid. The African cichlid requires water at a pH of 8.0 to 9.2; the South American cichlid requires a pH of 6.4 to 7.0. What would happen if you put a South American cichlid and some plants into a tank with water at pH of 8.5?



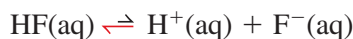
**Buffers** are solutions that resist changes in pH when limited amounts of acid or base are added. For example, adding 0.01 mole of HCl to 1 L of pure water lowers the pH by 5.0 units, from 7.0 to 2.0. Similarly, adding 0.01 mole of NaOH to 1 L of pure water increases the pH from 7.0 to 12.0. But if you add the same amount of either HCl or NaOH to 1 L of a buffered solution, the pH might change by no more than 0.1 unit.

**How do buffers work?** A buffer is a mixture of a weak acid and its conjugate base or a weak base and its conjugate acid. The mixture of ions and molecules in a buffer solution resists changes in pH by reacting with any hydrogen ions or hydroxide ions added to the buffered solution.

Suppose that a buffer solution contains 0.1M concentrations of hydrofluoric acid, and 0.1M fluoride ion (NaF). See **Figure 19-23**. HF is the acid and F<sup>-</sup> is its conjugate base. The following equilibrium would be established.

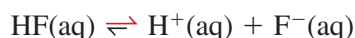


When an acid is added to this buffered solution, the equilibrium shifts to the left. The added H<sup>+</sup> ions react with F<sup>-</sup> ions to form additional HF molecules.



The pH of the solution remains fairly constant because the additional HF molecules do not ionize appreciably.

When a base is added to the hydrofluoric acid/fluoride ion buffer system, the added OH<sup>-</sup> ions react with H<sup>+</sup> ions to form water. This decreases the concentration of hydrogen ions and the equilibrium shifts to the right to replace the H<sup>+</sup> ions.



Although the shift to the right consumes HF molecules and produces additional F<sup>-</sup> ions, the pH remains fairly constant because the H<sup>+</sup> ion concentration has not changed appreciably.

A buffer solution's capacity to resist pH change can be exceeded by the addition of too much acid or base. Excessive acid overwhelms the hydrofluoric acid/fluoride ion buffer system by using up almost all of the F<sup>-</sup> ions. Similarly, too much base overwhelms the same system by using up almost all of the HF molecules. The amount of acid or base a buffer solution can absorb without a significant change in pH is called the **buffer capacity** of the solution. The greater the concentrations of the buffering molecules and ions in the solution, the greater the solution's buffer capacity. Do the **problem-solving LAB** on the next page to find out how buffers work in your blood.

**Choosing a buffer** A buffer can best resist both increases and decreases in pH when the concentrations of the conjugate acid-base pair are equal or nearly equal. Consider the H<sub>2</sub>PO<sub>4</sub><sup>-</sup>/HPO<sub>4</sub><sup>2-</sup> buffer system made by mixing equal molar amounts of NaH<sub>2</sub>PO<sub>4</sub> and Na<sub>2</sub>HPO<sub>4</sub>.



What is the pH of such a buffer solution? The acid ionization constant expression for the equilibrium can provide the answer.

$$K_{\text{a}} = 6.2 \times 10^{-8} = \frac{[\text{H}^{\text{+}}][\text{HPO}_4^{\text{2-}}]}{[\text{H}_2\text{PO}_4^{\text{-}}]}$$



**Figure 19-23**

Hydrofluoric acid is used to etch decorative designs on glass. Beeswax is coated on the glass and the design is drawn in the wax with a metal stylus. When the glass is dipped into hydrofluoric acid solution, the acid etches the glass in the grooves left by the stylus.

Table 19-5

Buffer Systems with Equimolar Components		
Buffer equilibrium	Conjugate acid-base pair in buffered solution	Buffer pH
$\text{HF}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{F}^-(\text{aq})$	$\text{HF}/\text{F}^-$	3.20
$\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$	$\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$	4.76
$\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$	$\text{H}_2\text{CO}_3/\text{HCO}_3^-$	6.35
$\text{H}_2\text{PO}_4^-(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HPO}_4^{2-}(\text{aq})$	$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.21
$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{NH}_4^+/\text{NH}_3$	9.4
$\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{C}_2\text{H}_5\text{NH}_3^+/\text{C}_2\text{H}_5\text{NH}_2$	10.70

Because the solution has been made up with equal molar amounts of  $\text{Na}_2\text{HPO}_4$  and  $\text{NaH}_2\text{PO}_4$ ,  $[\text{HPO}_4^{2-}]$  is equal to  $[\text{H}_2\text{PO}_4^-]$ . Thus, the two terms in the acid ionization expression cancel.

$$6.2 \times 10^{-8} = \frac{[\text{H}^+][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

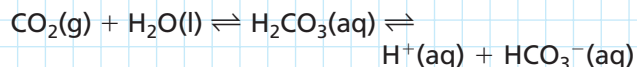
$$6.2 \times 10^{-8} = [\text{H}^+]$$

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

## problem-solving LAB

### How does your blood maintain its pH?

The pH of human blood must be kept within the narrow range of 7.1 to 7.7. Outside this range, proteins in the body lose their structure and ability to function. Fortunately, a number of buffers maintain the necessary acid/base balance. The carbonic acid/hydrogen carbonate buffer is the most important.



As acids and bases are dumped into the bloodstream as a result of normal activity, the blood's buffer systems shift to effectively maintain a healthy pH.

#### Analysis

Depending upon the body's metabolic rate and other factors, the  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$  equilibrium will shift according to Le Châtelier's principle. In addition, the lungs can alter the rate at which  $\text{CO}_2$  is

expelled from the body by breathing and the kidneys can alter the rate of removal of hydrogen carbonate ion.

#### Thinking Critically

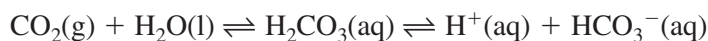
- Calculate the molarity of  $\text{H}^+$  in blood at a normal  $\text{pH} = 7.4$  and at  $\text{pH} = 7.1$ , the lower limit of a healthy pH range. If the blood's pH changes from 7.4 to 7.1, how many times greater is the concentration of hydrogen ion?
- The ratio of  $\text{HCO}_3^-$  to  $\text{CO}_2$  is 20:1. Why is this imbalance favorable for maintaining a healthy pH?
- In the following situations, predict whether the pH of the blood will rise or fall, and which way the  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$  equilibrium will shift.
  - A person with severe stomach flu vomits many times during a 24-hour period.
  - To combat heartburn, a person foolishly takes a large amount of sodium hydrogen carbonate.

$$\text{pH} = -\log (6.2 \times 10^{-8}) = 7.21$$

Thus, when equimolar amounts of each of the components are present in the  $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$  buffer system, the system can maintain a pH close to 7.21. Note that the pH is the negative log of  $K_a$ . **Table 19-5** lists several buffer systems with the pH at which each is effective. The pH values were calculated as the pH for the  $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$  buffer system was calculated above.

**Buffers in Blood** Your blood is a slightly alkaline solution with a pH of approximately 7.4. To be healthy, that pH must be maintained within narrow limits. A condition called acidosis occurs if the pH falls more than 0.3 units below 7.4. An equally serious condition called alkalosis exists if the pH rises 0.3 units. You may have experienced a mild case of acidosis if you have overexerted and developed a cramp in your leg. Cramping results from the formation of lactic acid in muscle tissue.

Your body employs three principle strategies for maintaining proper blood pH. The first is excretion of excess acid or base in the urine. The second is the elimination of  $\text{CO}_2$ , the anhydride of carbonic acid, by breathing. The third involves a number of buffer systems. You have already learned about the  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$  buffer in the **problem-solving LAB**.



The  $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$  buffer system described above is also present in blood. Because the concentration of phosphates in the blood is not as high as the concentration of carbonates, the phosphate buffer plays a lesser role in maintaining blood pH even though it is a more efficient buffer. Other buffer systems, including some associated with the hemoglobin molecule work together with the carbonate and phosphate buffers to maintain pH.

Recall that if acid levels begin to rise, the  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$  equilibrium shifts to the left to consume the added  $\text{H}^+$ . At the same time, the respiratory system is signaled to increase the rate of breathing to eliminate the higher level of  $\text{CO}_2$ . If levels of base begin to rise, the equilibrium shifts to the right as  $\text{H}^+$  neutralizes the base and the respiratory system slows the removal of  $\text{CO}_2$  by depressing the breathing rate. Can you explain why?

Suppose that during a concert a fan becomes overexcited and begins to hyperventilate. Excessive loss of  $\text{CO}_2$  shifts the  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$  equilibrium to the left and increases the pH. The situation can be controlled by helping the person calm down and breathe into a paper bag to recover exhaled  $\text{CO}_2$  as shown in **Figure 19-24**.



**Figure 19-24**

Look at the equation for the carbonate buffer as you think about this situation. As  $\text{CO}_2$  is eliminated through rapid breathing, the equilibrium shifts to the left. The  $[\text{H}^+]$  decreases and pH increases. Re-breathing  $\text{CO}_2$  from a paper bag helps restore the pH balance.

## Section 19.4 Assessment

- 34.** Write the formula equation and the net ionic equation for the neutralization reaction between hydroiodic acid and potassium hydroxide.
- 35.** Explain the difference between the equivalence point and the end point of a titration.
- 36.** Predict the results of two experiments: A small amount of base is added to an unbuffered solution with a pH of 7 and the same amount of base is added to a buffered solution with a pH of 7.
- 37. Thinking Critically** When a salt is dissolved in water, how can you predict whether or not a salt hydrolysis reaction occurs?
- 38. Designing an Experiment** Describe how you would design and carry out a titration in which you use  $0.250M$   $\text{HNO}_3$  to determine the molarity of a cesium hydroxide solution.



# Standardizing a Base Solution by Titration

The procedure called titration can be used to standardize a solution of a base, which means determine its molar concentration. To standardize a base, a solution of the base with unknown molarity is gradually added to a solution containing a known mass of an acid. The procedure enables you to determine when the number of moles of added  $\text{OH}^-$  ions from the base equals the number of moles of  $\text{H}^+$  ion from the acid.

## Problem

How can you determine the molar concentration of a base solution? How do you know when the neutralization reaction has reached the equivalence point?

## Objectives

- **Recognize** the color change of the indicator that shows that the equivalence point has been reached.
- **Measure** the mass of the acid and the volume of the base solution used.
- **Calculate** the molar concentration of the base solution.

## Materials

50-mL buret  
buret clamp  
ring stand  
sodium hydroxide pellets (NaOH)  
potassium hydrogen phthalate ( $\text{KHC}_8\text{H}_4\text{O}_4$ )  
distilled water  
weighing bottle  
spatula

250-mL Erlenmeyer flask  
500-mL Florence flask and rubber stopper  
250-mL beaker  
centigram balance  
wash bottle  
phenolphthalein solution  
dropper

## Safety Precautions



- Always wear safety goggles and a lab apron.

## Pre-Lab

1. What is the equivalence point of a titration?
2. Read the entire CHEMLAB.
3. What is the independent variable? The dependent variable? Constant?
4. When the solid acid dissolves to form ions, how many moles of  $\text{H}^+$  ions are produced for every mole of acid used?
5. What is the formula used to calculate molarity?
6. Prepare a data table that will accommodate multiple titration trials.
7. List safety precautions that must be taken.

Titration Data	
	Trial 1
mass of weighing bottle and acid	
mass of weighing bottle	
mass of solid acid	
moles of acid	
moles of base required	
final reading of base buret	
initial reading of base buret	
volume of base used in mL	
molarity of base	

## Procedure

1. Place approximately 4 g NaOH in a 500-mL Florence flask. Add enough water to dissolve the pellets and bring the volume of the NaOH solution to about 400 mL. **CAUTION:** *The solution will get hot.* Keep the stopper in the flask.
2. Use the weighing bottle to mass by difference about 0.40 g of potassium hydrogen phthalate (molar mass = 204.32 g/mol) into a 250-mL Erlenmeyer flask. Record this data.
3. Using a wash bottle, rinse down the insides of the flask and add enough water to make about 50 mL of solution. Add two drops of phenolphthalein indicator solution.
4. Set up the buret as shown. Rinse the buret with about 10 mL of your base solution. Discard the rinse solution in a discard beaker.



5. Fill the buret with NaOH solution. To remove any air trapped in the tip, allow a small amount of the base to flow from the tip of the buret into the discard beaker. Read the buret to the nearest 0.02 mL and record this initial reading. The meniscus of the solution in the buret should be at eye level when you make a reading.
6. Place a piece of white paper under the buret. Allow the NaOH solution to flow slowly from the buret into the flask containing the acid. Control the flow of the base solution with your left hand, and gently swirl the flask with your right hand.
7. The NaOH solution may be added in a rapid stream of drops until the pink color begins to last

longer as the flask is swirled. At this stage, begin adding the base drop by drop.

8. The equivalence point is reached when one additional drop of base turns the acid in the flask pink. The pink color should persist as the flask is swirled. Record the final volume in the buret.
9. Calculate the molarity of your base using steps 2–5 below.
10. Refill your buret with base. Rinse your Erlenmeyer flask with water. Repeat the titration with additional samples of acid until you get three trials that show close agreement between the calculated values of the molarity.

## Cleanup and Disposal

1. The neutralized solutions can be washed down the sink using plenty of water.

## Analyze and Conclude

1. **Observing and Inferring** Identify the characteristics of this neutralization reaction.
2. **Collecting and Interpreting Data** Complete the data table. Calculate the number of moles of acid used in each trial by dividing the mass of the sample by the molar mass of the acid.
3. **Using Numbers** How many moles of base are required to react with the moles of acid you used?
4. **Using Numbers** Convert the volume of base used from milliliters to liters.
5. **Analyze and Conclude** For each trial, calculate the molar concentration of the base by dividing the moles of base by the volume of base in liters.
6. **Error Analysis** How well did your calculated molarities agree? Explain any irregularities.

## Real-World Chemistry

1. Use what you have learned about titration to design a field investigation to determine whether your area is affected by acid rain. Research the factors that affect the pH of rain, such as location, prevailing winds, and industries. Form a hypothesis about the pH of rain in your area. What equipment will you need to collect samples? To perform the titration? What indicator will you use?

# How It Works

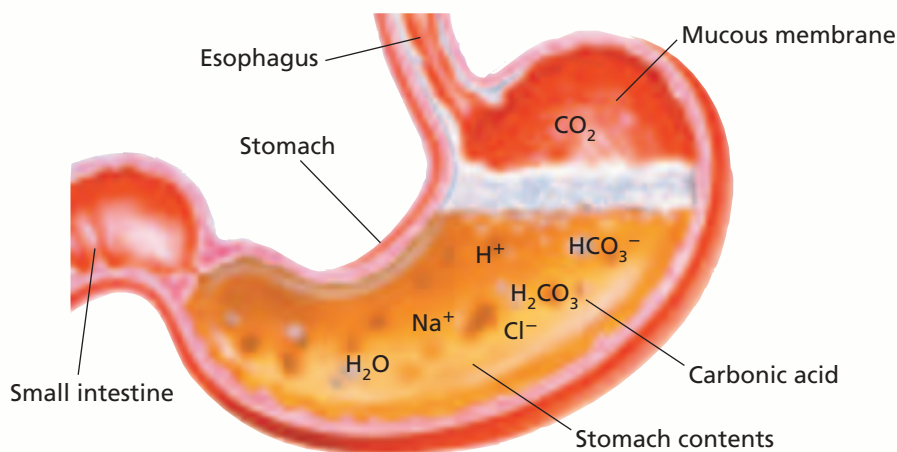
## Antacids

Your stomach is a hollow organ where the food you eat is broken down into a usable form. The stomach contains powerful enzymes and hydrochloric acid, which are responsible for the breakdown process. Once the food is processed, it is released into the small intestine.

The pain and discomfort of indigestion is an indication that normal digestion has been interrupted. Heartburn is an irritation of the esophagus that is caused by stomach acid. Millions of people use antacids to treat indigestion and heartburn. Antacids are bases that neutralize digestive acids.



- 1 Acids and enzymes in the stomach help digest food. The pH in the stomach is about 2.5.
- 2 A basic mucous membrane lines the stomach and protects it from corrosion.



- 4 Carbonic acid decomposes into carbon dioxide gas and water.  
$$\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$
- 3 A sodium hydrogen carbonate (sodium bicarbonate) antacid reacts with hydrochloric acid in the stomach and forms carbonic acid.  
$$\text{NaHCO}_3(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{CO}_3(\text{aq})$$

### Thinking Critically

1. **Applying** Milk of magnesia is a suspension of magnesium hydroxide in water. Write the net ionic equation for the neutralization reaction between magnesium hydroxide and hydrochloric acid.
2. **Comparing and Contrasting** Why is sodium hydrogen carbonate an effective antacid but sodium hydroxide is not?

## Summary

### 19.1 Acids and Bases: An Introduction

- Acidic solutions contain more hydrogen ions than hydroxide ions; neutral solutions contain equal concentrations of hydrogen ions and hydroxide ions; basic solutions contain more hydroxide ions than hydrogen ions.
- In the Arrhenius model, an acid is a substance that contains hydrogen and ionizes in aqueous solution to produce hydrogen ions. An Arrhenius base is a substance that contains a OH group and dissociates in aqueous solution to produce hydroxide ions.
- According to the Brønsted-Lowry model, an acid is a hydrogen ion donor and a base is a hydrogen ion acceptor.
- When a Brønsted-Lowry acid donates a hydrogen ion, a conjugate base is formed; when a Brønsted-Lowry base accepts a hydrogen ion, a conjugate acid is formed.

### 19.2 Strengths of Acids and Bases

- Strong acids and strong bases are completely ionized in dilute aqueous solution. Weak acids and weak bases are partially ionized.
- For weak acids and weak bases, the value of the acid or base ionization constant is a measure of the strength of the acid or base.

### 19.3 What is pH?

- The pH of a solution is the negative log of the hydrogen ion concentration and the pOH is the negative log of the hydroxide ion concentration.
- A neutral solution has a pH of 7.0 and a pOH of 7.0 because the concentrations of hydrogen ion and hydroxide ion are equal. The pH of a solution decreases as the solution becomes more acidic and increases as the solution becomes more basic. The pOH of a solution decreases as the solution becomes more basic and increases as the solution becomes more acidic.
- The ion product constant for water,  $K_w$ , equals the product of the hydrogen ion concentration and the hydroxide ion concentration.

### 19.4 Neutralization

- The general equation for an acid-base neutralization reaction is acid + base  $\rightarrow$  salt + water.
- The net ionic equation for the neutralization of a strong acid by a strong base is  $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ .
- Titration is the process in which an acid-base neutralization reaction is used to determine the concentration of a solution of unknown concentration.
- Buffered solutions contain mixtures of molecules and ions that resist changes in pH.

## Key Equations and Relationships

- $pH = -\log [H^+]$  (p. 610)
- $pOH = -\log [OH^-]$  (p. 611)
- $pH + pOH = 14.00$  (p. 611)
- $K_w = [H^+][OH^-]$  (p. 608)

## Vocabulary

- acid-base indicator (p. 619)
- acid ionization constant (p. 605)
- acidic solution (p. 597)
- amphoteric (p. 599)
- Arrhenius model (p. 597)
- base ionization constant (p. 606)
- basic solution (p. 597)
- Brønsted-Lowry model (p. 598)
- buffer (p. 623)
- buffer capacity (p. 623)
- conjugate acid (p. 598)
- conjugate acid-base pair (p. 598)
- conjugate base (p. 598)
- end point (p. 619)
- equivalence point (p. 618)
- ion product constant for water (p. 608)
- neutralization reaction (p. 617)
- pH (p. 610)
- pOH (p. 611)
- salt (p. 617)
- salt hydrolysis (p. 621)
- strong acid (p. 602)
- strong base (p. 606)
- titration (p. 618)
- weak acid (p. 603)
- weak base (p. 606)

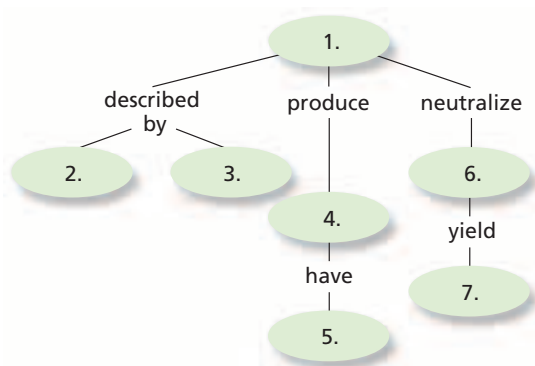




Go to the Chemistry Web site at [chemistrymc.com](http://chemistrymc.com) for additional Chapter 19 Assessment.

## Concept Mapping

39. Use the following words and phrases to complete the concept map: acidic solutions, acids, bases, Arrhenius model,  $\text{pH} < 7$ , a salt plus water, Brønsted-Lowry model.



## Mastering Concepts

40. An aqueous solution tastes bitter and turns litmus blue. Is the solution acidic or basic? (19.1)
41. An acidic solution reacts with magnesium carbonate to produce a gas. What is the formula for the gas? (19.1)
42. In terms of ion concentrations, distinguish between acidic, neutral, and basic solutions. (19.1)
43. Write a balanced chemical equation that represents the self-ionization of water. (19.1)
44. How did Arrhenius describe acids and bases? Why was his description important? (19.1)
45. Table sugar ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ) contains 22 hydrogen atoms per molecule. Does this make table sugar an Arrhenius acid? Explain your answer. (19.1)
46. Classify each of the following compounds as an Arrhenius acid or an Arrhenius base. (19.1)
- |                         |                             |
|-------------------------|-----------------------------|
| a. $\text{H}_2\text{S}$ | c. $\text{Mg}(\text{OH})_2$ |
| b. $\text{RbOH}$        | d. $\text{H}_3\text{PO}_4$  |
47. Explain the difference between a monoprotic acid, a diprotic acid, and a triprotic acid. Give an example of each. (19.1)
48. Why does acid rain dissolve statues made of marble ( $\text{CaCO}_3$ )? Write the formula equation for the reaction between sulfuric acid and calcium carbonate. (19.1)
49. Ammonia contains three hydrogen atoms per molecule. However, an aqueous ammonia solution is basic. Explain using the Brønsted-Lowry model of acids and bases. (19.1)
50. Identify the conjugate acid-base pairs in the equilibrium equation.  
 $\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_2\text{H}_3\text{O}_2^-$  (19.1)
51. Gaseous  $\text{HCl}$  molecules interact with gaseous  $\text{NH}_3$  molecules to form a white smoke made up of solid  $\text{NH}_4\text{Cl}$  particles. Explain whether or not this is an acid-base reaction according to both the Arrhenius model and the Brønsted-Lowry model. (19.1)
52. Explain the difference between a strong acid and a weak acid. (19.2)
53. Why are strong acids and bases also strong electrolytes? (19.2)
54. State whether each of the following acids is strong or weak. (19.2)
- |                    |                      |
|--------------------|----------------------|
| a. acetic acid     | c. hydrofluoric acid |
| b. hydroiodic acid | d. phosphoric acid   |
55. State whether each of the following bases is strong or weak. (19.2)
- |                       |                     |
|-----------------------|---------------------|
| a. rubidium hydroxide | c. ammonia          |
| b. methylamine        | d. sodium hydroxide |
56. How would you compare the strengths of two weak acids (19.2)
- |  |
|--|
| a. experimentally?                                     |
| b. by looking up information in a table or a handbook? |
57. **Figure 19-7** shows the conductivity of two acids. Explain how you could distinguish between solutions of two bases, one containing  $\text{NaOH}$  and the other  $\text{NH}_3$ . (19.2)
58. Explain why the base ionization constant ( $K_b$ ) is a measure of the strength of a base. (19.2)
59. Explain why a weak acid has a strong conjugate base. Give an equation that illustrates your answer. (19.2)
60. Explain why a weak base has a strong conjugate acid. Give an equation that illustrates your answer. (19.2)
61. Explain how you can calculate  $K_a$  for a weak acid if you know the concentration of the acid and its  $\text{pH}$ . (19.2)
62. What is the relationship between the  $\text{pOH}$  and the hydroxide-ion concentration of a solution? (19.3)

- 63.** Solution A has a pH of 2.0. Solution B has a pH of 5.0. Which solution is more acidic? Based on the hydrogen-ion concentrations in the two solutions, how many times more acidic? (19.3)
- 64.** If the concentration of hydrogen ions in an aqueous solution decreases, what must happen to the concentration of hydroxide ions? Why? (19.3)
- 65.** Explain why pure water has a very slight electrical conductivity. (19.3)
- 66.** Why is the pH of pure water 7.0 at 298 K? (19.3)
- 67.** What is a standard solution in an acid-base titration? (19.4)
- 68.** How do you recognize the end point in an acid-base titration? (19.4)
- 69.** Give the name and formula of the acid and the base from which each salt was formed. (19.4)
- |                    |                             |
|--------------------|-----------------------------|
| a. NaCl            | c. $\text{NH}_4\text{NO}_2$ |
| b. $\text{KClO}_3$ | d. CaS                      |
- 70.** How does buffering a solution change the solution's behavior when a base is added? When an acid is added? (19.4)
- 71.** Why are some aqueous salt solutions acidic or basic? (19.4)
- 72.** How does knowing the equivalence point in an acid-base titration help you choose an indicator for the titration? (19.4)
- 73.** An aqueous solution causes bromthymol blue to turn blue and phenolphthalein to turn colorless. What is the approximate pH of the solution? (19.4)
- 74.** In the net ionic equation for the neutralization reaction between nitric acid and magnesium hydroxide, what ions are left out of the equation? (19.4)
- 75.** Describe two ways you might detect the end point of an acid-base titration experimentally. (19.4)
- 76.** What is the approximate pH of the equivalence point in the titration pH curve? (19.4)
- 77.** Define both the equivalence point and the end point of an acid-base titration. Why should you choose an indicator so that the two points are nearly the same pH? (19.4)
- 78.** Explain how you can predict whether an aqueous salt solution is acidic, basic, or neutral by evaluating the strengths of the salt's acid and base parents. (19.4)
- 79.** In a hypochlorous acid/hypochlorite-ion buffer, what chemical species reacts when an acid is added to the solution? (19.4)
- 80.** Arrange the three buffers in order of increasing pH values. In order of increasing buffer capacity.
- |                             |
|-----------------------------|
| a. 1.0M HClO/1.0M NaClO     |
| b. 0.10M HClO/0.10M NaClO   |
| c. 0.010M HClO/0.010M NaClO |

## Mastering Problems

### Equations for Acid and Base Reactions (19.1)

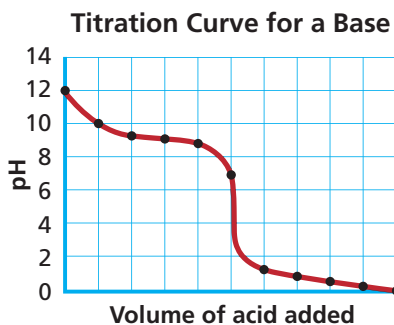
- 81.** Write a balanced formula equation for the reaction between sulfuric acid and calcium metal.
- 82.** Write a balanced formula equation for the reaction between potassium hydrogen carbonate and chlorous acid ( $\text{HClO}_2$ ).
- 83.** Write the balanced chemical equation for the ionization of perchloric acid ( $\text{HClO}_4$ ) in water.
- 84.** Write the balanced chemical equation for the dissociation of solid magnesium hydroxide in water.

### Weak Acids and Bases (19.2)

- 85.** Write the equation for the ionization reaction and the acid ionization constant expression for the  $\text{HS}^-$  ion in water.
- 86.** Write the equation for the ionization reaction and the acid ionization constant expression for the third ionization of phosphoric acid in water.

### $K_w$ , pH, and pOH (19.3)

- 87.** Given the concentration of either hydrogen ion or hydroxide ion, use the ion product constant of water to calculate the concentration of the other ion at 298 K.
- |   |
|---|
| a. $[\text{H}^+] = 1.0 \times 10^{-4}\text{M}$  |
| b. $[\text{OH}^-] = 1.3 \times 10^{-2}\text{M}$ |
- 88.** Calculate the pH at 298 K of solutions having the following ion concentrations.
- |   |
|---|
| a. $[\text{H}^+] = 1.0 \times 10^{-4}\text{M}$  |
| b. $[\text{H}^+] = 5.8 \times 10^{-11}\text{M}$ |



- 89.** Calculate the pOH and pH at 298 K of solutions having the following ion concentrations.
- $[\text{OH}^-] = 1.0 \times 10^{-12}M$
  - $[\text{OH}^-] = 1.3 \times 10^{-2}M$
- 90.** Calculate the pH of each of the following strong acid or strong base solutions at 298 K.
- $2.6 \times 10^{-2}M$  HCl
  - $0.28M$   $\text{HNO}_3$
  - $7.5 \times 10^{-3}M$  NaOH
  - $0.44M$  KOH

### Calculations Using $K_a$ (19.3)

- 91.** A  $8.6 \times 10^{-3}M$  solution of  $\text{H}_3\text{PO}_4$  has a  $\text{pH} = 2.30$ . What is  $K_a$  for  $\text{H}_3\text{PO}_4$ ?
- 92.** What is  $K_a$  for a solution of chloroacetic acid ( $\text{C}_2\text{H}_3\text{ClO}_2$ ) which has a concentration of  $0.112M$  and a pH of 1.92?

### Neutralization Reactions (19.4)

- 93.** Write formula equations for the following acid-base neutralization reactions.
- sulfuric acid + sodium hydroxide
  - methanoic acid + potassium hydroxide
- 94.** Write formula equations and net ionic equations for the hydrolysis of the following salts in water.
- sodium carbonate
  - ammonium bromide
- 95.** In a titration, 33.21 mL  $0.3020M$  rubidium hydroxide solution is required to exactly neutralize 20.00 mL hydrofluoric acid solution. What is the molarity of the hydrofluoric acid solution?
- 96.** A 35.00 mL-sample of NaOH solution is titrated to an endpoint by 14.76 mL  $0.4122M$  HBr solution. What is the molarity of the NaOH solution?

### Mixed Review

Sharpen your problem-solving skills by answering the following.

- 97.** Calculate  $[\text{H}^+]$  and  $[\text{OH}^-]$  in each of the following solutions at 298 K.
- $\text{pH} = 3.00$
  - $\text{pH} = 5.24$
- 98.** Write the equation for the ionization reaction and the base ionization constant expression for ethylamine ( $\text{C}_2\text{H}_5\text{NH}_2$ ) in water.
- 99.** Write net ionic equations for the three ionizations of boric acid ( $\text{H}_3\text{BO}_3$ ) in water. Include  $\text{H}_2\text{O}$  in the three equations. Identify the conjugate acid-base pairs.

- 100.** How many milliliters of  $0.225M$  HCl would be required to titrate 6.00 g KOH?
- 101.** What is the pH of a  $0.200M$  solution of hypobromous acid ( $\text{HBrO}$ ),  $K_a = 2.8 \times 10^{-9}$ .

### Thinking Critically

- 102. Analyzing and Concluding** Is it possible that an acid according to the Arrhenius model is not a Brønsted-Lowry acid? Is it possible that an acid according to the Brønsted-Lowry model is not an Arrhenius acid? Explain and give examples.
- 103. Applying Concepts** Use the ion product constant of water at 298 K to explain why a solution with a pH of 3.0 must have a pOH of 11.0.
- 104. Using Numbers** The ion product constant of water rises with temperature. What is the pH of pure water at 313 K if  $K_w$  at that temperature is  $2.917 \times 10^{-14}$ ?
- 105. Interpreting Scientific Illustrations** Sketch the shape of the approximate pH vs. volume curve that would result from titrating a diprotic acid with a  $0.10M$  NaOH solution.
- 106. Recognizing Cause and Effect** Illustrate how a buffer works using the  $\text{C}_2\text{H}_5\text{NH}_3^+/\text{C}_2\text{H}_5\text{NH}_2$  buffer system. Show with equations how the weak base/conjugate acid system is affected when small amounts of acid and base are added to a solution containing this buffer system.

### Writing in Chemistry

- 107.** Examine the labels of at least two brands of shampoo and record any information regarding pH. Research the pH of skin, hair, and the pH ranges of various shampoos. Write a report summarizing your findings.
- 108.** The twenty amino acids combine to form proteins in living systems. Research the structures and  $K_a$  values for five amino acids. Compare the strengths of these acids with the weak acids in **Table 19-2**.

### Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 109.** What factors determine whether a molecule is polar or nonpolar? (Chapter 9)
- 110.** When 5.00 g of a compound was burned in a calorimeter, the temperature of 2.00 kg of water increased from  $24.5^\circ\text{C}$  to  $40.5^\circ\text{C}$ . How much heat would be released by the combustion of 1.00 mol of the compound (molar mass =  $46.1 \text{ g/mol}$ )? (Chapter 16)

# STANDARDIZED TEST PRACTICE

## CHAPTER 19

Use these questions and the test-taking tip to prepare for your standardized test.

- A carbonated soft drink has a pH of 2.5. What is the concentration of  $H^+$  ions in the soft drink?
  - $3 \times 10^{-12}M$
  - $3 \times 10^{-3}M$
  - $4.0 \times 10^{-1}M$
  - $1.1 \times 10^1M$
- At the equivalence point of a strong acid-strong base titration, what is the approximate pH?
  - 3
  - 5
  - 7
  - 9
- Hydrogen bromide (HBr) is a strong, highly corrosive acid. What is the pOH of a 0.0375M HBr solution?
  - 12.57
  - 12.27
  - 1.73
  - 1.43
- A compound that accepts  $H^+$  ions is
  - an Arrhenius acid.
  - an Arrhenius base.
  - a Brønsted-Lowry acid.
  - a Brønsted-Lowry base.

**Interpreting Tables** Use the table to answer questions 5–7.

- Which of the following acids is the strongest?
  - formic acid
  - cyanoacetic acid
  - lutidinic acid
  - barbituric acid
- What is the acid dissociation constant of propanoic acid?
  - $1.4 \times 10^{-5}$
  - $2.43 \times 10^0$
  - $3.72 \times 10^{-3}$
  - $7.3 \times 10^4$
- What is the pH of a 0.400M solution of cyanoacetic acid?
  - 2.059
  - 1.22
  - 2.45
  - 1.42

- Which of the following is NOT a characteristic of a base?
  - bitter taste
  - ability to conduct electricity
  - reactivity with some metals
  - slippery feel
- Diprotic succinic acid ( $H_2C_4H_4O_4$ ) is an important part of the process that converts glucose to energy in the human body. What is the  $K_a$  expression for the second ionization of succinic acid?
  - $K_a = \frac{[H_3O^+][HC_4H_4O_4^-]}{[H_2C_4H_4O_4]}$
  - $K_a = \frac{[H_3O^+][C_4H_4O_4^{2-}]}{[HC_4H_4O_4^-]}$
  - $K_a = \frac{[H_2C_4H_4O_4]}{[H_3O^+][HC_4H_4O_4^-]}$
  - $K_a = \frac{[H_2C_4H_4O_4]}{[H_3O^+][C_4H_4O_4^{2-}]}$
- A solution of 0.600M HCl is used to titrate 15.00 mL of KOH solution. The endpoint of the titration is reached after the addition of 27.13 mL of HCl. What is the concentration of the KOH solution?
  - 9.000M
  - 1.09M
  - 0.332M
  - 0.0163M

### TEST-TAKING TIP

**Slow Down!** Check to make sure you're answering the question that each problem is posing. Read the questions and every answer choice very carefully. Remember that doing most of the problems and getting them right is always preferable to doing all the problems and getting lots of them wrong.

**Ionization Constants and pH Data for Several Weak Organic Acids**

Acid	Ionization equation	pH of 1.000M solution	$K_a$
Formic	$HCHO_2 \rightleftharpoons H^+ + CHO_2^-$	1.87	$1.78 \times 10^{-4}$
Cyanoacetic	$HC_3H_2NO_2 \rightleftharpoons H^+ + C_3H_2NO_2^-$	?	$3.55 \times 10^{-3}$
Propanoic	$HC_3H_5O_2 \rightleftharpoons H^+ + C_3H_5O_2^-$	2.43	?
Lutidinic	$H_2C_7H_3NO_4 \rightleftharpoons H^+ + C_7H_4NO_4^-$	1.09	$7.08 \times 10^{-3}$
Barbituric	$HC_4H_3N_2O_3 \rightleftharpoons H^+ + C_4H_3N_2O_3^-$	2.01	$9.77 \times 10^{-5}$

