CHAPTER 10

Acids and Bases

What do cheese, stomach juices, baking soda, oven cleaner, and underarm odour have in common? They are all acidic or basic. Do you know which are acidic and which are basic?

Acids and bases are very important chemicals. They have been used for thousands of years. Vinegar is an acidic solution that is common in many food and cleaning products. It was discovered long ago—before people invented the skill of writing to record its use. Today, acids are also used to manufacture fertilizers, explosives, plastics, motor vehicles, and computer circuit boards.

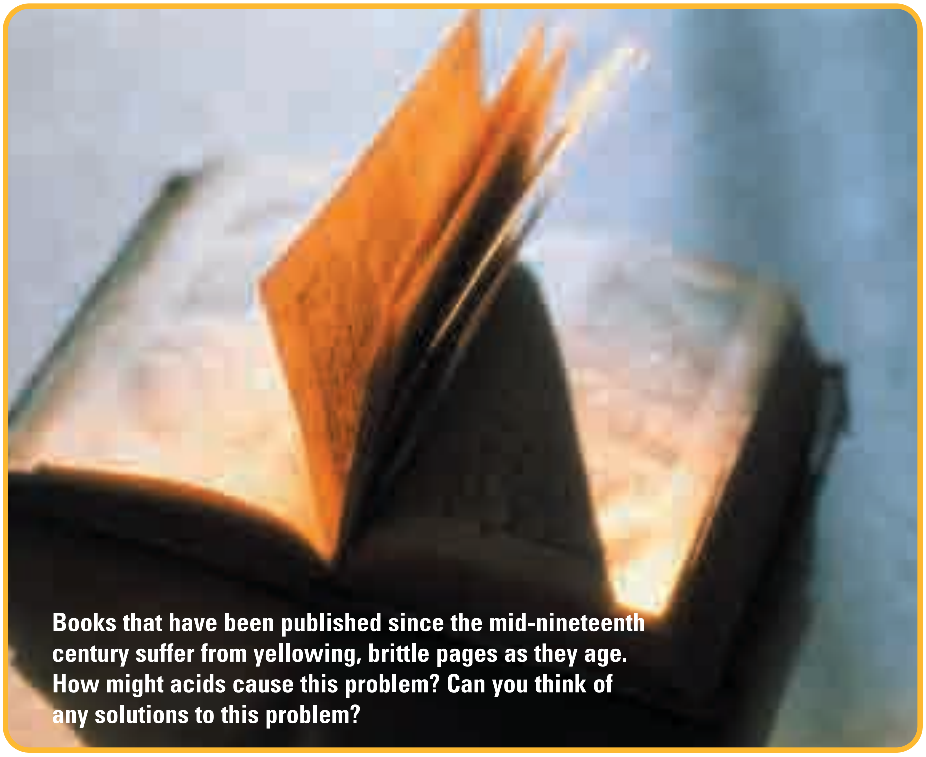
Like acids, bases have numerous uses in the home and in chemical industries. Nearly 5000 years ago, in the Middle East, the Babylonians made soap using the bases in wood ash. Today, one of Canada's most important industries, the pulp and paper industry, uses huge quantities of a base called sodium hydroxide. Sodium hydroxide is also used to manufacture soaps, detergents, dyes, and many other compounds.

In this chapter, you will learn about the properties of acids and bases. You will learn how these properties change when acids and bases react together. As well, you will have a chance to estimate and measure the acidity of aqueous solutions.

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- **describing** and **calculating** the concentration of solutions (Chapter 8, section 8.2)
- **performing** stoichiometry calculations (Chapter 7, section 7.2)
- **naming** and **identifying** polyatomic ions and their formulas (Chapter 3, section 3.4)



Books that have been published since the mid-nineteenth century suffer from yellowing, brittle pages as they age. How might acids cause this problem? Can you think of any solutions to this problem?

10.1 Acid-Base Theories

Section Preview/ Specific Expectations

In this section, you will

- **describe** and **compare** the Arrhenius and Brønsted-Lowry theories of acids and bases
- **identify** conjugate acid-base pairs
- **communicate** your understanding of the following terms: *Arrhenius theory of acids and bases, hydronium ion, Brønsted-Lowry theory of acids and bases, conjugate acid-base pair, conjugate base, conjugate acid*

As you can see in Table 10.1, acids and bases are common products in the home. It is easy to identify some products as acids. Often the word "acid" appears in the list of ingredients. Identifying bases is more difficult. Acids and bases have different properties, however, that enable you to distinguish between them.

Table 10.1 Common Acids and Bases in the Home

Acids	
Product	**Acid(s) contained in the product**
citrus fruits (such as lemons, limes, oranges and tomatoes)	citric acid and ascorbic acid
dairy products (such as cheese, milk, and yogurt)	lactic acid
vinegar	acetic acid
soft drinks	carbonic acid; may also contain phosphoric acid and citric acid
underarm odour	3-methyl-2-hexenoic acid
Bases	
Product	**Base contained in the product**
oven cleaner	sodium hydroxide
baking soda	sodium hydrogen carbonate
washing soda	sodium carbonate
glass cleaner (some brands)	ammonia

Language LINK

The word "acid" comes from the Latin *acidus*, meaning "sour tasting." As you will learn in this chapter, bases are the "base" (the foundation) from which many other compounds form. A base that is soluble in water is called an alkali. The word "alkali" comes from an Arabic word meaning "ashes of a plant." In the ancient Middle East, people rinsed plant ashes with hot water to obtain a basic solution. The basic solution was then reacted with animal fats to make soap.

Properties of Acids and Bases

One way to distinguish acids from bases is to describe their observable properties. For example, acids taste sour, and they change colour when mixed with coloured dyes called indicators. Bases taste bitter and feel slippery. They also change colour when mixed with indicators.

CAUTION You should never taste or touch acids, bases, or any other chemicals. Early chemists used their senses of taste and touch to observe the properties of many chemicals. This dangerous practice often led to serious injury, and sometimes death.

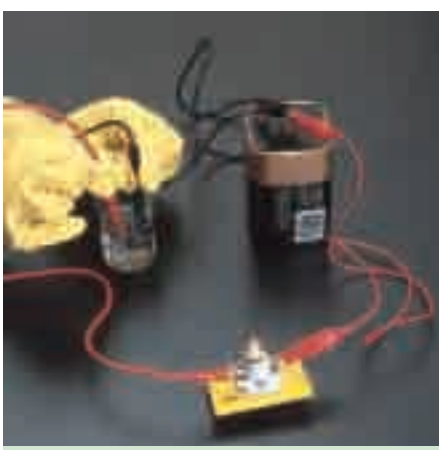
Another property that can be used to distinguish acids from bases is their conductivity in solution. As you can see in Figure 10.1, aqueous solutions of acids and bases conduct electricity. This is evidence that ions are present in acidic and basic solutions. Some of these solutions, such as hydrochloric acid and sodium hydroxide (a base), cause the bulb to glow brightly. Most acidic and basic solutions, however, cause the bulb to glow dimly.

pure water

hydrochloric acid, $HCl_{(aq)}$ (1 mol/L)

acetic acid, $CH_3COOH_{(aq)}$ (1 mol/L)

sodium hydroxide, $NaOH_{(aq)}$ (1 mol/L)



ammonia, $NH_{3(aq)}$ (1 mol/L)

Figure 10.1 **Aqueous solutions of acids and bases can be tested using a conductivity tester. The brightness of the bulb is a clue to the concentration of ions in the solution. Which of these solutions have higher concentrations of ions? Which have lower concentrations?**

mind STRETCH

An oxide is a compound of oxygen with a metal or non-metal. Most metal oxides react with water to form basic solutions. For example, calcium oxide is a metal oxide that is important in the construction industry as an ingredient of cement. Calcium oxide reacts with water to form a basic solution of calcium hydroxide.

$CaO_{(s)} + H_2O_{(\ell)} \rightarrow Ca(OH)_{2(aq)}$

Many municipal water treatment plants use calcium hydroxide to soften very hard water before releasing it for public use. Most non-metal oxides react with water to form acidic solutions. For example, sulfur dioxide gas dissolves in water to form sulfurous acid.

$SO_{2(g)} + H_2O_{(\ell)} \rightarrow H_2SO_{3(aq)}$

The metallic character of the elements in the periodic table, and the acid-base properties of their oxides, show a distinct trend across periods and down groups. Infer what this trend is. In other words, state what you think happens to the acid-base properties of oxides as you go across a period and down a group. Make a quick sketch of the periodic table to illustrate this trend. How would you describe the acid-base properties of the metalloids? (Use your knowledge of the physical properties of the metalloids to help you make your inference.)

Table 10.2 on the next page summarizes the observable properties of acids and bases. These observable properties include their physical characteristics and their chemical behaviour. The Express Lab on page xxx provides you with an opportunity to compare some of these properties. What are acids and bases, however? How does chemical composition determine whether a substance is acidic or basic? You will consider one possible answer to this question starting on page 373.

Table 10.2 Some Observable Properties of Acids and Bases

	Property		
	Taste	**Electrical conductivity in solution**	**Feel of solution**
ACIDS	taste sour	conduct electricity	have no characteristic feel
BASES	taste bitter	conduct electricity	feel slippery
			

	Property		
	Reaction with litmus paper	**Reaction with active metals**	**Reaction with carbonate compounds**
ACIDS	Acids turns blue litmus red	produce hydrogen gas	produce carbon dioxide gas
BASES	Bases turn red litmus blue	do not react	do not react
			

ExpressLab Clean a Penny

Many cleaning products contain an acid or a base. For example, some window cleaners contain vinegar (acetic acid). Other window cleaners contain ammonia (a base). Oven cleaners, however, contain only bases. This activity will help you infer why.

Safety Precautions

Materials

water
vinegar
100 mL graduated cylinder
spoon or scoopula
baking soda
3 small beakers (about 200 mL)
3 tarnished pennies

Procedure

1. Predict which solution(s) will clean the penny best. Give reasons for your prediction.
2. In one beaker, mix 50 mL of vinegar with about 150 mL of water. In a second beaker, mix about 20 mL to 30 mL spoonfuls of baking soda with 150 mL of water. In the third beaker, put only 150 mL of water.
3. Place a tarnished penny in each beaker. Observe what happens for about 15 min.

Analysis

1. Which solution was the best cleaner? How did your observations compare with your prediction?
2. What results would you expect if you tried cleaning a penny in a solution of lemon juice? What if you used a dilute solution of ammonia? **Note:** If you want to test your predictions, ask your teacher for the concentrations of the solutions you should use.
3. The base that is often used in oven cleaners is sodium hydroxide. This base is very corrosive, and it can burn skin easily. A corrosive acid, such as hydrochloric acid, could also remove baked-on grease and grime from ovens. Why are bases a better choice for oven cleaners?

The Arrhenius Theory of Acids and Bases

In Figure 10.1, you saw evidence that ions are present in solutions of acids and bases. When hydrogen chloride dissolves in water, for example, it dissociates (breaks apart) into hydrogen ions and chloride ions.

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

When sodium hydroxide dissolves in water, it dissociates to form sodium ions and hydroxide ions.

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

The dissociations of other acids and bases in water reveal a pattern. This pattern was first noticed in the late nineteenth century by a Swedish chemist named Svanté Arrhenius. (See Figure 10.2.)

$$HBr_{(aq)} \rightarrow \mathbf{H^+}_{(aq)} + Br^-_{(aq)}$$

$$H_2SO_{4(aq)} \rightarrow \mathbf{H^+}_{(aq)} + HSO^-_{4(aq)}$$

$$HClO_{4(aq)} \rightarrow \mathbf{H^+}_{(aq)} + ClO^-_{4(aq)}$$

acids dissociating in water, and their resulting ions

$$LiOH_{(aq)} \rightarrow Li^+_{(aq)} + \mathbf{OH^-}_{(aq)}$$

$$KOH_{(aq)} \rightarrow K^+_{(aq)} + \mathbf{OH^-}_{(aq)}$$

$$Ba(OH)_{2(aq)} \rightarrow Ba^+_{(aq)} + 2\mathbf{OH^-}_{(aq)}$$

bases dissociating in water, and their resulting ions

Figure 10.2 **Svanté Arrhenius (1859–1927).**

Web LINK

www.school.mcgrawhill.ca/resources/

Scientists did not embrace the Arrhenius theory when they first heard about it during the 1880s. Why were they unimpressed with this theory? What was necessary to convince them? Arrhenius is featured on several web sites on the Internet. To link with these web sites, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

In 1887, Arrhenius published a theory to explain the nature of acids and bases. It is called the **Arrhenius theory of acids and bases**.

The Arrhenius Theory of Acids and Bases

- An acid is a substance that dissociates in water to produce one or more hydrogen ions, H^+.
- A base is a substance that dissociates in water to form one or more hydroxide ions, OH^-.

According to the Arrhenius theory, acids increase the concentration of H^+ in aqueous solutions. Thus, an Arrhenius acid must contain hydrogen as the source of H^+. You can see this in the dissociation reactions for acids on the previous page.

Bases, on the other hand, increase the concentration of OH^- in aqueous solutions. An Arrhenius base must contain the hydroxyl group, —OH, as the source of OH^-. You can see this in the dissociation reactions for bases on the previous page.

Limitations of the Arrhenius Theory

The Arrhenius theory is useful if you are interested in the ions that result when an acid or a base dissociates in water. It also helps explain what happens when an acid and a base undergo a neutralization reaction. In such a reaction, an acid combines with a base to form an ionic compound and water. Examine the following reactions:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)}$$

The net ionic equation for this reaction shows the principal ions in the Arrhenius theory.

$$H^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow H_2O_{(\ell)}$$

Since acids and bases produce hydrogen ions and hydroxide ions, water is an inevitable result of acid-base reactions.

Problems arise with the Arrhenius theory, however. One problem involves the ion that is responsible for acidity: H^+. Look again at the equation for the dissociation of hydrochloric acid.

$$HCl_{(aq)} \rightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)}$$

This dissociation occurs in aqueous solution, but chemists often leave out H_2O as a component of the reaction. They simply assume that it is there. What happens if you put H_2O into the equation?

$$HCl_{(aq)} + H_2O_{(\ell)} \rightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)} + H_2O_{(\ell)}$$

Notice that the water is unchanged when the reaction is represented this way. However, you learned earlier that water is a polar molecule. The O atom has a partial negative charge, and the H atoms have partial positive charges. Thus, H_2O must interact in some way with the ions H^+ and Cl^-. In fact, chemists made a discovery in the early twentieth century. They realized that protons do not exist in isolation in aqueous solution. (The hydrogen ion is simply a proton. It is a positively charged nuclear particle.) Instead, protons are always *hydrated*: they are attached to water molecules. A hydrated proton is called a **hydronium ion,** $H_3O^+{}_{(aq)}$. (See Figure 10.3.)

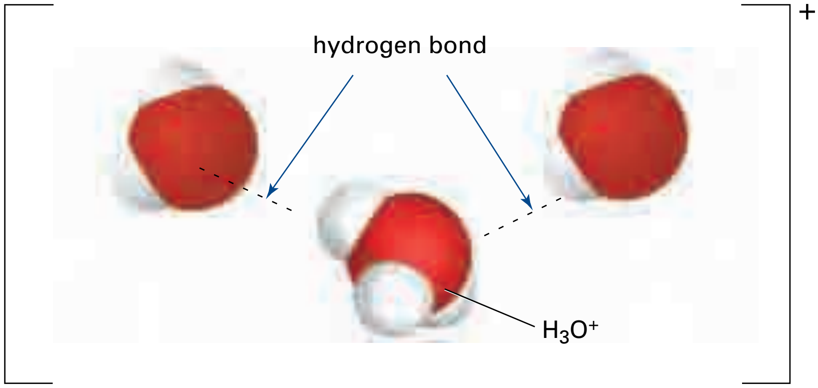


Figure 10.3 For convenience, chemists often use $H^+_{(aq)}$ as a shorthand notation for the hydronium ion, $H_3O^+_{(aq)}$. Hydronium ions do not exist independently. Instead, they form hydrogen bonds with other water molecules. Thus, a more correct formula is $H^+(H_2O)_n$, where *n* is usually 4 or 5.

There is another problem with the Arrhenius theory. Consider the reaction of ammonia, NH_3, with water.

$$NH_{3(g)} + H_2O_{(\ell)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)}$$

Ammonia is one of several substances that produce basic solutions in water. As you can see, ammonia does not contain hydroxide ions. However, it does produce these ions when it reacts with water. Ammonia also undergoes a neutralization reaction with acids. The Arrhenius theory cannot explain the basic properties of ammonia. Nor can it explain the fact that certain other substances, such as salts that contain carbonate ions, also have basic properties.

There is yet another problem with the Arrhenius theory. It is limited to acid-base reactions in a single solvent, water. Many acid-base reactions take place in other solvents, however.

CHECKPOINT

Use the idea of the hydronium ion to complete the following equation:

$$HCl_{(aq)} + H_2O_{(\ell)} \rightarrow$$

The Brønsted-Lowry Theory of Acids and Bases

In 1923, two chemists working independently of each other, proposed a new theory of acids and bases. (See Figure 10.4.) Johannes Brønsted in Copenhagen, Denmark, and Thomas Lowry in London, England, proposed what is called the **Brønsted-Lowry theory of acids and bases.** This theory overcame the problems related to the Arrhenius theory.

The Brønsted-Lowry Theory of Acids and Bases

- An acid is a substance from which a proton (H^+ ion) can be removed.
- A base is a substance that can remove a proton (H^+ ion) from an acid.

Figure 10.4 Johannes Brønsted (1879–1947), left, and Thomas Lowry (1874–1936), right. Brønsted published many more articles about ions in solution than Lowry did. Thus, some chemistry resources refer to the "Brønsted theory of acids and bases."

CHEM FACT

In many chemistry references, Brønsted-Lowry acids are called "proton donors." Brønsted-Lowry bases are called "proton acceptors." Although these terms are common, they create a false impression about the energy that is involved in acid-base reactions. Breaking bonds always requires energy. For example, removing a proton from a hydrochloric acid molecule requires 1.4×10^3 kJ/mol. This is far more energy than the word "donor" implies.

Like an Arrhenius acid, a Brønsted-Lowry acid must contain H in its formula. This means that all Arrhenius acids are also Brønsted-Lowry acids. However, any negative ion (not just OH^-) can be a Brønsted-Lowry base. In addition, water is not the only solvent that can be used.

According to the Brønsted-Lowry theory, there is only one requirement for an acid-base reaction. One substance must provide a proton, and another substance must receive the same proton. In other words, *an acid-base reaction involves the transfer of a proton.*

The idea of proton transfer has major implications for understanding the nature of acids and bases. According to the Brønsted-Lowry theory, any substance can behave as an acid, but only if another substance behaves as a base at the same time. Similarly, any substance can behave as a base, but only if another substance behaves as an acid at the same time.

For example, consider the reaction between hydrochloric acid and water shown in Figure 10.5. In this reaction, hydrochloric acid is an acid because it provides a proton (H^+) to the water. The water molecule receives the proton. Therefore, according to the Brønsted-Lowry theory, water is a base in this reaction. When the water receives the proton, it becomes a hydronium ion (H_3O^+). Notice the hydronium ion on the right side of the equation.

$$\underbrace{HCl_{(aq)}}_{\text{acid}} + \underbrace{H_2O_{(\ell)}}_{\text{base}} \xrightarrow{} \overbrace{H_3O^+_{(aq)}}^{\text{conjugate acid}} + \underbrace{Cl^-_{(aq)}}_{\text{conjugate base}}$$

proton transfer

Figure 10.5 The reaction between hydrochloric acid and water, according to the Brønsted-Lowry theory

Two molecules or ions that are related by the transfer of a proton are called a **conjugate acid-base pair.** (Conjugate means "linked together.") The **conjugate base** of an acid is the particle that remains when a proton is removed from the acid. The **conjugate acid** of a base is the particle that results when the base receives the proton from the acid. In the reaction between hydrochloric acid and water, the hydronium ion is the conjugate acid of the base, water. The chloride ion is the conjugate base of the acid, hydrochloric acid.

According to the Brønsted-Lowry theory, every acid has a conjugate base, and every base has a conjugate acid. The conjugate base and conjugate acid of an acid-base pair are linked by the transfer of a proton. The conjugate base of the acid-base pair has one less hydrogen than the acid. It also has one more negative charge than the acid. The conjugate acid of the acid-base pair has one more hydrogen than the base. It also has one less negative charge than the base.

These ideas about acid-base reactions and conjugate acid-base pairs will become clearer as you study the following Sample Problems and Practice Problems.

Sample Problem

Conjugate Acid-Base Pairs

Problem

Hydrogen bromide is a gas at room temperature. It is soluble in water, forming hydrobromic acid. Identify the conjugate acid-base pairs.

What Is Required?

You need to identify two sets of conjugate acid-base pairs.

What Is Given?

You know that hydrogen bromide forms hydrobromic acid in aqueous solution.

Plan Your Strategy

- Write a balanced chemical equation.
- On the left side of the equation, identify the acid as the molecule that provides the proton. Identify the base as the molecule that accepts the proton.
- On the right side of the equation, identify the particle that has one proton less than the acid on the left side as the conjugate base of the acid. Identify the particle on the right side that has one proton more than the base on the left side as the conjugate acid of the base.

Act on Your Strategy

Hydrogen bromide provides the proton, so it is the Brønsted-Lowry acid in the reaction. Water receives the proton, so it is the Brønsted-Lowry base. The conjugate acid-base pairs are HBr/Br^- and H_2O/H_3O^+.

conjugate acid-base pair

acid (HBr), conjugate base (Br⁻)

$$\underset{\text{acid}}{HBr_{(g)}} + \underset{\text{base}}{H_2O_{(\ell)}} \longrightarrow \underset{\text{conjugate acid}}{H_3O^+_{(aq)}} + \underset{\text{conjugate base}}{Br^-_{(aq)}}$$

conjugate acid-base pair

Check Your Solution

The formulas of the conjugate pairs differ by one proton, H^+, as expected.

Sample Problem

More Conjugate Acid-Base Pairs

Problem

Ammonia is a pungent gas at room temperature. Its main use is in the production of fertilizers and explosives. It is very soluble in water. It forms a basic solution that is used in common products, such as glass cleaners. Identify the conjugate acid-base pairs in the reaction between aqueous ammonia and water.

$$NH_{3(g)} + H_2O_{(\ell)} \rightarrow NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$$

What Is Required?

You need to identify the conjugate acid-base pairs.

What Is Given?

The chemical equation is given.

Plan Your Strategy

- Identify the proton-provider on the left side of the equation as the acid. Identify the proton-remover (or proton-receiver) as the base.
- Identify the conjugate acid and base on the right side of the equation by the difference of a single proton from the acid and base on the left side.

PROBLEM TIP

In the previous Sample Problem, water acted as a base. In this Sample Problem, water acts as an acid. **Remember:** Water can act as a proton-provider (an acid) in some reactions and as a proton-receiver (a base) in others.

Act on Your Strategy

The conjugate acid base pairs are $NH_4{}^+/NH_3$ and H_2O/OH^-.

conjugate acid-base pair

acid · conjugate base

$$NH_{3(g)} + H_2O_{(\ell)} \longrightarrow NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$$

base · conjugate acid

conjugate acid-base pair

Check Your Solution

The formulas of the conjugate pairs differ by a proton, as expected.

Practice Problems

1. Hydrogen cyanide is a poisonous gas at room temperature. When this gas dissolved in water, the following reaction occurs:

$$HCN_{(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+{}_{(aq)} + CN^-{}_{(aq)}$$

Identify the conjugate acid-base pairs.

Continued ...

Continued ...
FROM PAGE 378

2. Sodium acetate is a good electrolyte. In water, the acetate ion reacts as follows:

$$CH_3COO^-_{(aq)} + H_2O_{(\ell)} \rightarrow CH_3COOH_{(aq)} + OH^-_{(aq)}$$

Identify the conjugate acid-base pairs.

3. Write equations to show how the hydrogen sulfide ion, HS^-, can react with water. First show the ion acting as an acid. Then show the ion acting as a base.

Section Wrap-up

The two theories that you have considered in this section attempt to explain the chemical nature of acids and bases. Table 10.3 summarizes the key points of these theories.

In both the Arrhenius theory and the Brønsted-Lowry theory, acids and bases form ions in solution. Many characteristics of acid-base behaviour are linked to the number of ions that form from a particular acid or base. One of these characteristics is strength.

In section 10.2, you will learn why a dilute solution of vinegar is safe to ingest, while the same molar concentration of hydrochloric acid would be extremely poisonous.

Table 10.3 Comparing the Arrhenius Theory and the Brønsted-Lowry Theory

Theory	Arrhenius	Brønstead-Lowry
Acid	any substance that dissociates to form H^+ in aqueous solution	any substance that provides a proton to another substance (or any substance from which a proton may be removed)
Base	any substance that dissociates to form OH^- in aqueous solution	any substance that receives a proton from an acid (or any substance that removes a proton from an acid)
Example	$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$	$HCl_{(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

Section Review

1 I Suppose that you have four unknown solutions, labelled A, B, C, and D. You use a conductivity apparatus to test their conductivity, and obtain the results shown below. Use these results to answer the questions that follow.

Solution	Results of Conductivity Test
A	the bulb glows dimly
B	the bulb glows strongly
C	the bulb does not glow
D	the bulb glows strongly

(a) Which of these solutions has a high concentration of dissolved ions? What is your evidence?

(b) Which of these solutions has a low concentration of dissolved ions? What is your evidence?

(c) Which of the four unknowns are probably aqueous solutions of acids or bases?
(d) Based on these tests alone, can you distinguish the acidic solution(s) from the basic solutions(s)? Why or why not?
(e) Suggest one way that you could distinguish the acidic solution(s) from the basic solution(s).

2 **(a)** K/U Define an acid and a base according to the Arrhenius theory.
(b) Give two examples of observations that the Arrhenius theory can explain.
(c) Give two examples of observations that the Arrhenius theory can *not* explain.

3 **(a)** K/U Define an acid and a base according to the Brønsted-Lowry theory.
(b) What does the Brønsted-Lowry theory have in common with the Arrhenius theory? In what ways is it different?
(c) Which of the two acid-base theories is more comprehensive? (In other words, which explains a broader body of observations?)

4 **(a)** K/U What is the conjugate acid of a base?
(b) What is the conjugate base of an acid?
(c) Use an example to illustrate your answers to parts (a) and (b) above.

5 K/U Write the formula for the conjugate acid of the following:
(a) the hydroxide ion, OH^-
(b) the carbonate ion, CO_3^{2-}

6 K/U Write the formula for the conjugate base of the following:
(a) nitric acid, HNO_3
(b) the hydrogen sulfate ion, HSO_4^-

7 K/U Which of the following compounds is an acid according to the Arrhenius theory?
(a) H_2O
(b) $Ca(OH)_2$
(c) H_3PO_3
(d) HF

8 K/U Which of the following compounds is a base according to the Arrhenius theory?
(a) KOH
(b) $Ba(OH)_2$
(c) HClO
(d) H_3PO_4

9 C Hydrofluoric acid dissociates in water to form fluoride ions.
(a) Write a balanced chemical equation for this reaction.
(b) Identify the conjugate acid-base pairs.
(c) Explain how you know whether or not you have correctly identified the conjugate acid-base pairs.

10 I Identify the conjugate acid-base pairs in the following reactions:
(a) $H_2PO_4^-{}_{(aq)} + CO^{2-}{}_{3(aq)} \rightarrow HPO_4^{2-}{}_{(aq)} + HCO_3^-{}_{(aq)}$
(b) $HCOOH_{(aq)} + CN^-{}_{(aq)} \rightarrow HCOO^-{}_{(aq)} + HCN_{(aq)}$
(c) $H_2PO_4^-{}_{(aq)} + OH^-{}_{(aq)} \rightarrow HPO_4^{2-}{}_{(aq)} + H_2O_{(\ell)}$

Strong and Weak Acids and Bases 10.2

Re-examine Figure 10.1 on page xxx. Look at the photographs of hydrochloric acid and acetic acid. The conductivity tester is testing the same concentrations of both acids. As you can see, the bulb glows brightly in the hydrochloric acid. The bulb glows dimly in the acetic acid. How can these different results be explained?

Section Preview/ Specific Expectations

In this section, you will

- **explain**, in terms of the degree to which they dissociate, the difference between strong and weak acids and bases
- **distinguish** between binary acids and oxoacids
- **define** pH, and experimentally determine the effect on pH of diluting an acidic solution
- **communicate** your understanding of the following terms: *strong acid, weak acid, strong base, weak base, binary acid, oxoacid, pH*

Strong Acids and Weak Acids

You know that ions are present in an aqueous solution of an acid. These ions result from the dissociation of the acid. An acid that dissociates completely into ions in water is called a **strong acid.** For example, hydrochloric acid is a strong acid. *All* the molecules of hydrochloric acid in an aqueous solution dissociate into H^+ and Cl^- ions. The H^+ ions, as you know, bond with surrounding water molecules to form hydronium ions, H_3O^+. (See Figure 10.6.) *The concentration of hydronium ions in a dilute solution of a strong acid is equal to the concentration of the acid.* Thus, a 1.0 mol/L solution of hydrochloric acid contains 1.0 mol/L of hydronium ions. Table 10.4 lists the strong acids.

Table 10.4 Strong Acids

hydrochloric acid, HCl
hydrobromic acid, HBr
hydroiodic acid, HI
nitric acid, HNO_3
sulfuric acid, H_2SO_4
perchloric acid, $HClO_4$

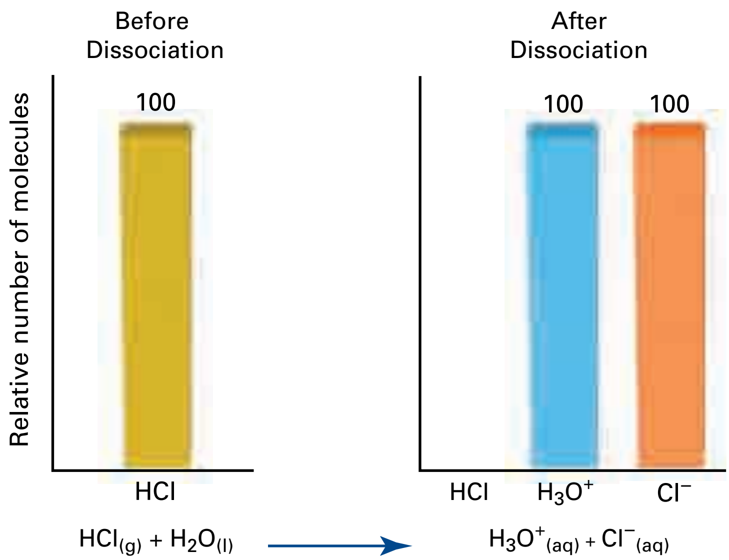


Figure 10.6 When hydrogen chloride molecules enter an aqueous solution, 100% of the hydrogen chloride molecules dissociate. As a result, the solution contains the same percent of H^+ ions (in the form of H_3O^+) and Cl^- ions: 100%.

A **weak acid** is an acid that dissociates very slightly in a water solution. Thus, only a small percentage of the acid molecules break apart into ions. Most of the acid molecules remain intact. For example, acetic acid is a weak acid. On average, only about 1% (one in a hundred) of the acetic acid molecules dissociate at any given moment in a 0.1 mol/L solution. (The number of acid molecules that dissociate depends on the concentration and temperature of the solution.) In fact, *the concentration of hydronium ions in a solution of a weak acid is always less than the concentration of the dissolved acid.* (See Figure 10.7.)

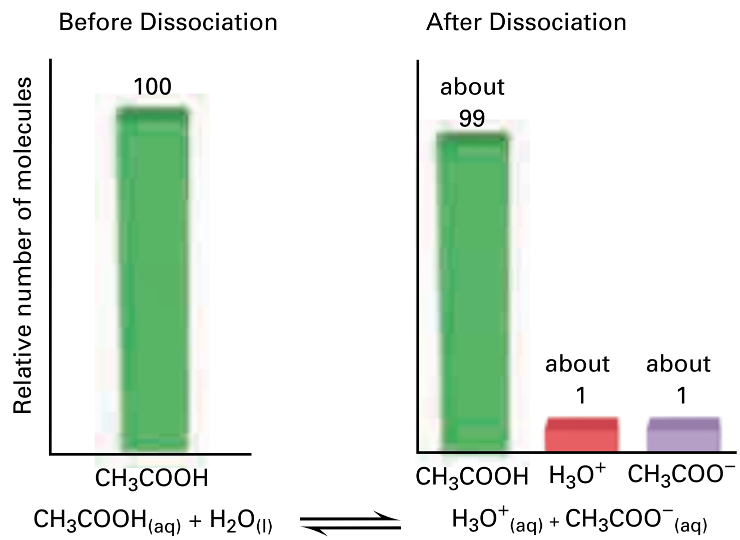


Figure 10.7 **When acetic acid molecules enter an aqueous solution, only about 1% of them dissociate. Thus, the number of acetic acid molecules in solution is far greater than the number of hydronium ions and acetate ions.**

Notice the arrow that is used in the equation in Figure 10.7. It points in both directions, indicating that the reaction is *reversible.* In other words, the products of the reaction also react to produce the original reactants. In this reaction, molecules of acetic acid dissociate just as quickly and as often as the dissociated ions re-associate to produce acetic acid molecules. (Figure 10.8 will help you visualize what happens.)

Most acids are weak acids. Whenever you see a reversible chemical equation involving an acid, you can safely assume that the acid is weak.

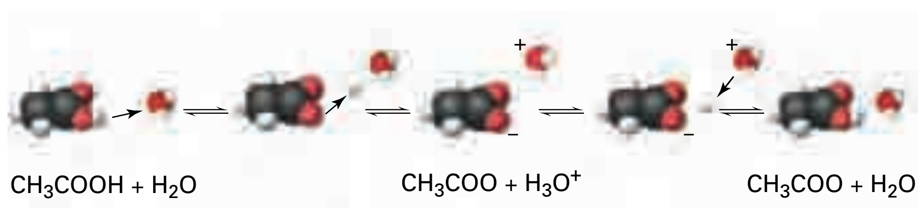


Figure 10.8 **When acetic acid dissolves in water, acetic acid molecules dissociate and re-associate at the same time and at the same rate.**

A few acids contain only a single hydrogen ion that can dissociate. These acids are called *monoprotic acids.* (The prefix mono- means "one." The root -protic refers to "proton.") Hydrochloric acid, hydrobromic acid, and hydroiodic acid are strong monoprotic acids. Hydrofluoric acid, HF, is weak monoprotic acid.

Many acids contain two or more hydrogen ions that can dissociate. For example, sulfuric acid, $H_2SO_{4(aq)}$, has two hydrogen ions that can dissociate. As you know from Table 10.4, sulfuric acid is a strong acid. This is true only for its first dissociation, however.

$$H_2SO_{4(aq)} \rightarrow H^+{}_{(aq)} + HSO_4^-{}_{(aq)}$$

The resulting aqueous hydrogen sulfate ion, $HSO^-{}_4$, is a weak acid. It dissociates to form the sulfate ion in the following reversible reaction:

$$HSO_4^-{}_{(aq)} \rightleftharpoons H^+{}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Thus, acids that contain two hydrogen ions dissociate to form two anions. These acids are sometimes called *diprotic acids.* (The prefix di-, as you know, means "two.") The acid that is formed by the first dissociation is stronger than the acid that is formed by the second dissociation.

Acids that contain three hydrogen ions are called *triprotic acids.* Phosphoric acid, $H_3PO_{4(aq)}$, is a triprotic acid. It gives rise to three anions, as follows:

$$H_3PO_{4(aq)} + H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + H_2PO_4{}^-_{(aq)}$$
$$H_2PO_4{}^-_{(aq)} + H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + HPO_4{}^{2-}_{(aq)}$$
$$HPO_4{}^{2-}_{(aq)} + H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + PO_4{}^{3-}_{(aq)}$$

Here again, the acid that is formed by the first dissociation is stronger than the acid that is formed by the second dissociation. This acid is stronger than the acid that is formed by the third dissociation. Keep in mind, however, that all three of these acids are weak because only a small proportion of them dissociates.

Strong Bases and Weak Bases

Like a strong acid, a **strong base** dissociates completely into ions in water. All oxides and hydroxides of the alkali metals—Group 1 (IA)—are strong bases. The oxides and hydroxides of the alkaline earth metals—Group 2 (IIA)—below beryllium are also strong bases.

Recall that the concentration of hydronium ions in a dilute solution of a strong acid is equal to the concentration of the acid. Similarly, the concentration of hydroxide ions in a dilute solution of a strong base is equal to the concentration of the base. For example, a 1.0 mol/L solution of sodium hydroxide (a strong base) contains 1.0 mol/L of hydroxide ions.

Table 10.5 lists some common strong bases. Barium hydroxide, $Ba(OH)_2$, and strontium hydroxide, $Sr(OH)_2$, are strong bases that are soluble in water. Magnesium oxide, MgO, and magnesium hydroxide, $Mg(OH)_2$, are also strong bases, but they are considered to be insoluble. A small amount of these compounds does dissolve in water, however. Virtually all of this small amount dissociates completely.

Most bases are weak. A **weak base** dissociates very slightly in a water solution. The most common weak base is aqueous ammonia. In a 0.1 mol/L solution, only about 1% of the ammonia molecules react with water to form hydroxide ions. This reversible reaction is represented in Figure 10.9.

Table 10.5 Common Strong Bases

sodium hydroxide, NaOH
potassium hydroxide, KOH
calcium hydroxide, $Ca(OH)_2$
strontium hydroxide, $Sr(OH)_2$
barium hydroxide, $Ba(OH)_2$

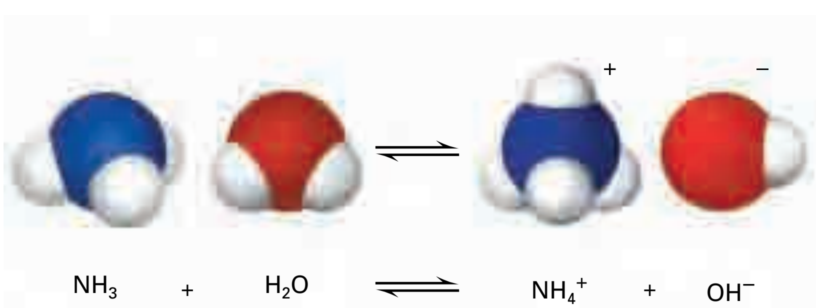


Figure 10.9 **Ammonia does not contain hydroxide ions, so it is not an Arrhenius base. As you can see, however, an ammonia molecule can remove a proton from water, leaving a hydroxide ion behind. Thus, ammonia is a Brønsted-Lowry weak base.**

CHECKPOINT

The terms "strong" and "concentrated" have very different meanings when describing solutions of acids, bases, or salts. Similarly, the terms "weak" and "dilute" have very different meanings. In your own words, summarize the difference between "strong" and "concentrated." Give examples that illustrate this difference. Then summarize the difference between "weak" and "dilute," again giving examples.

mind STRETCH

The conjugate base of a strong acid is always a weak base. Conversely, the conjugate base of a weak acid is always a strong base. Explain this inverse relationship.

Naming Acids and Their Anions

There are two main kinds of acids: binary acids and oxoacids. A **binary acid** is composed of two elements: hydrogen and a non-metal. Two examples of binary acids are hydrofluoric acid and hydrochloric acid. All binary acids have the general formula $HX_{(aq)}$. The H represents one or more hydrogen atoms. The X represents the non-metal. As you can see in Table 10.6, the names of binary acids are made up of the following parts:

- the prefix hydro-
- a root that is formed from the name of the non-metal
- the suffix -ic
- the word "acid" at the end

Table 10.6 Examples of Naming Binary Acids

Binary acid	Prefix	Non-metal root	Suffix
hydrofluoric acid, $HF_{(aq)}$	hydro-	-fluor-	-ic
hydrochloric acid, $HCl_{(aq)}$	hydro-	-chlor-	-ic
hydrosulfuric acid, H_2S	hydro-	-sulfur-	-ic

As you know, anions are formed when binary acids dissociate. The names of these anions end in the suffix -ide. For example, hydroflouric acid forms the anion fluoride, F^-. Hydrochloric acid forms the anion chloride, Cl^-.

An **oxoacid** is an acid formed from a polyatomic ion that contains oxygen, hydrogen, and another element. (Oxoacids are called oxyacids in some chemistry textbooks). In Chapter 3, you learned the names of common polyatomic ions and their valences (oxidation numbers). The names of oxoacids are similar to the names of their polyatomic oxoanions. Only the suffix is different. Study the three rules and examples for naming oxoacids below. Then try the Practice Problems that follow.

1. For anions that end in -ate, the suffix of the acid is -ic. For example, the acid of the chlorate anion ClO_3^-, is chloric acid, $HClO_3$.
2. For anions that end in -ite, the suffix of the acid is -ous. For example, the acid of the chlorite anion, ClO_2^-, is chlorous acid, $HClO_2$.
3. The prefixes hypo- and per- remain as part of the acid name. For example, the acid of the perchlorate anion, ClO_4^-, is perchloric acid, $HClO_4$. The acid of the hypochlorite anion, ClO^-, is hypochlorous acid, HClO.

Practice Problems

4. (a) Write the chemical formula for hydrobromic acid. Then write the name and formula for the anion that it forms.
   (b) Hydrosulfuric acid, H_2S, forms two anions. Name them and write their formulas.
5. Write the chemical formulas for the following acids. Then name and write the formulas for the oxoanions that form from each acid. Refer to Chapter 3, Table 3.5, Names and Valences of Some Common Polyatomic Ions, as necessary.
   (a) nitric acid
   (b) nitrous acid
   (c) hyponitrous acid
   (d) phosphoric acid
   (e) phosphorous acid
   (f) periodic acid

Describing Acid and Base Strength Quantitatively: pH

You are probably familiar with the term "pH" from a variety of sources. Advertisers talk about the "pH balance" of products such as soaps, shampoos, and skin creams. People who own aquariums and swimming pools must monitor the pH of the water. (See Figure 10.10.) Gardeners and farmers use simple tests to determine the pH of the soil. They know that plants and food crops grow best within a narrow range of pH. Similarly, the pH of your blood must remain within narrow limits for you to stay healthy.

Figure 10.10 **Maintaining a safe environment for an aquarium or a swimming pool requires measuring the pH of the water and knowing how to adjust it.**

pH is clearly related to health, and to the proper functioning of products and systems. (Notice that the "p" is always lower case, even at the start of a sentence.) What exactly is pH? How is it measured? To answer these questions, consider a familiar substance: water.

The Power of Hydrogen in Water

As you know, all aqueous solutions contain ions. Even pure water contains a few ions that are produced by the dissociation of water molecules. **Remember:** The double arrow in the equation shows that the reaction is reversible. The ions recombine to form water molecules.

$$H_2O_{(\ell)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

On average, at 25°C, only about two water molecules in a billion are dissociated at any given moment. As you know, it is the ions in solution that conduct electricity. If there are virtually no ions, no electricity is conducted. This is why pure water is such a poor conductor. Chemists have determined that the concentration of hydronium ions in neutral water, at 25˚C, is only 1.0×10^{-7} mol/L. The dissociation of water also produces the same, very small number of hydroxide ions. Therefore, the concentration of hydroxide ions is also 1.0×10^{-7} mol/L.

Chemists sometimes use square brackets around a chemical formula. This shorthand notation means "the concentration of" the chemical inside the brackets. For example, $[H_3O^+]$ is read as "the concentration of hydronium ions." Thus, the concentration of hydronium ions and hydroxide ions in neutral water can be written as

the concentration of hydronium ions

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

the concentration of hydroxide ions

Compared with neutral water, acidic solutions contain a higher concentration of hydronium ions. Basic solutions contain a lower concentration of hydronium ions. Therefore, the dissociation of water provides another way of thinking about acids and bases. *An acid is any compound that increases [H_3O^+] when it is dissolved in water. A base is any compound that increases [OH^-] when it is dissolved in water.* (See Figure 10.11.)

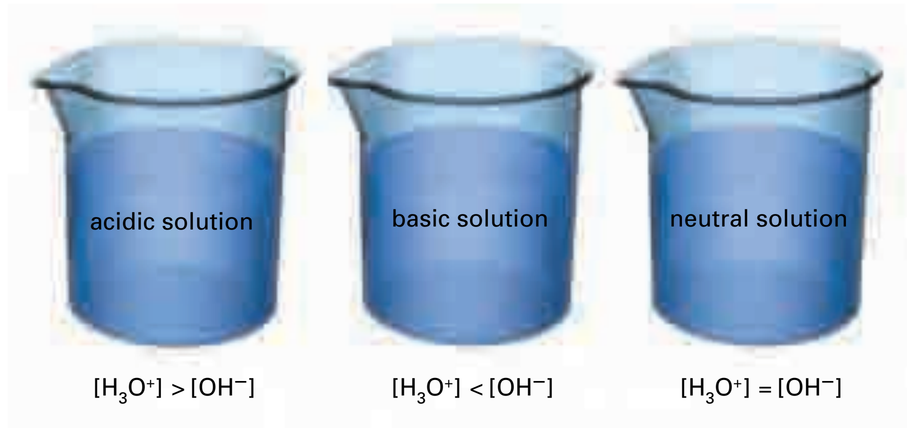


Figure 10.11 **The relationship between the concentrations of hydronium ions and hydroxide ions in a solution determines whether the solution is acidic, basic, or neutral.**

The pH Scale: Measuring by Powers of Ten

The concentration of hydronium ions ranges from about 10 mol/L for a concentrated strong acid to about 10^{-15} mol/L for a concentrated strong base. This wide range of concentrations, and the negative powers of 10, are not very convenient to work with. In 1909, a Danish biochemist, Søren Sørensen, suggested a method for converting concentrations to positive numbers. His method involved using the numerical system of logarithms.

The logarithm of a number is the power to which you must raise 10 to equal that number. For example, the logarithm of 10 is 1, because $10^1 = 10$. The logarithm of 100 is 2, because $10^2 = 100$. (See Appendix E for more information about exponents and logarithms.)

Sørensen defined **pH** as $-\log [H^+]$. Since Sørensen did not know about hydronium ions, his definition of pH is based on Arrhenius' hydrogen ion. Many chemistry references reinterpret the H so that it refers to the Brønsted-Lowry hydronium ion, H_3O^+, instead. This textbook adopts the hydronium ion usage. Thus, the definition for pH becomes $pH = -\log [H_3O^+]$. Recall, though, that chemists use [H^+] as a shorthand notation for [H_3O^+]. As a result, both equations give the same product.

As you can see in Figure 10.12, the "p" in pH stands for the word "power." The power referred to is exponential power: the power of 10. The "H" stands for the concentration of hydrogen ions (or H_3O^+ ions), measured in mol/L.

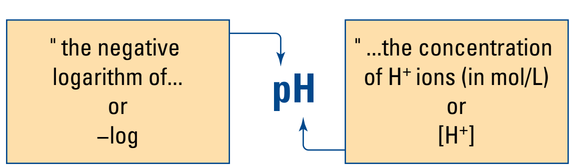


Figure 10.12 **The concept of pH makes working with very small values, such as 0.000 000 000 000 01, much easier.**

The concept of pH allows hydronium (or hydrogen) ion concentrations to be expressed as positive numbers, rather than negative exponents. For example, recall that $[H_3O^+]$ of neutral water at 25°C is 1.0×10^{-7}.

$$\begin{aligned} \therefore \text{pH} &= -\log\,[H_3O^+] \\ &= -\log\,(1.0 \times 10^{-7}) \\ &= -(-7.00) \\ &= 7.00 \end{aligned}$$

$[H_3O^+]$ in acidic solutions is greater than $[H_3O^+]$ in neutral water. For example, if $[H_3O^+]$ in an acid is 1.0×10^{-4} mol/L, this is 1000 times greater than $[H_3O^+]$ in neutral water. Use Table 10.7 to make sure that you understand why. The pH of the acid is 4.00. All acidic solutions have a pH that is less than 7.

Table 10.7 Understanding pH

Range of acidity and basicity	$[H_3O^+]$ (mol/L)	Exponential notation (mol/L)	log	pH (-log $[H_3O^+]$)
strong acid	1	1×10^{0}	0	0
	0.1	1×10^{-1}	−1	1
	0.01	1×10^{-2}	−2	2
	0.001	1×10^{-3}	−3	3
	0.000 1	1×10^{-4}	−4	4
	0.000 01	1×10^{-5}	−5	5
	0.000 001	1×10^{-6}	−6	6
neutral $[H^+] = [OH^-]$ $= 1.0 \times 10^{-7}$	0.000 000 1	1×10^{-7}	−7	7
	0.000 000 01	1×10^{-8}	−8	8
	0.000 000 001	1×10^{-9}	−9	9
	0.000 000 000 1	1×10^{-10}	−10	10
	0.000 000 000 01	1×10^{-11}	−11	11
	0.000 000 000 001	1×10^{-12}	−12	12
	0.000 000 000 000 1	1×10^{-13}	−13	13
strong base	0.000 000 000 000 01	1×10^{-14}	−14	14

$[H_3O^+]$ in basic solutions is less than $[H_3O^+]$ in pure water. For example, if $[H_3O^+]$ in a base is 1.0×10^{-11} mol/L, this is 10 000 times less than $[H_3O^+]$ in neutral water. The pH of the base is 11.00. All basic solutions have a pH that is greater than 7.

The relationship among pH, $[H_3O^+]$, and the strength of acids and bases is summarized in the Concept Connection on the next page. Use the following Sample Problem and Practice Problems to assess your understanding of this relationship. Then, in Investigation 10-A, you will look for a pattern involving the pH of a strong acid, a weak acid, and dilutions of both.

Science LINK

Using logarithms is a convenient way to count a wide range of values by powers of 10. Chemists are not the only scientists who use such logarithms, however. Audiologists (scientists who study human hearing) use logarithms, too. Research the decibel scale to find out how it works. Present your findings in the medium of your choice.

Math LINK

How do you determine the number of significant digits in a pH? You count only the digits to the right of the decimal point. For example, suppose that the concentration of hydronium ions in a sample of orange juice is 2.5×10^{-4} mol/L. This number has two significant digits: the 2 and the 5. The power of 10 only tells us where to place the decimal: 0.000 25. The pH of the sample is $-\log\,(2.5 \times 10^{-4}) = 3.602\,059$. The digit to the left of the decimal (the 3) is derived from the power of 10. Therefore, it is not considered to be a significant digit. Only the two digits to the right of the decimal are significant. Thus, the pH value is rounded off to 3.60.

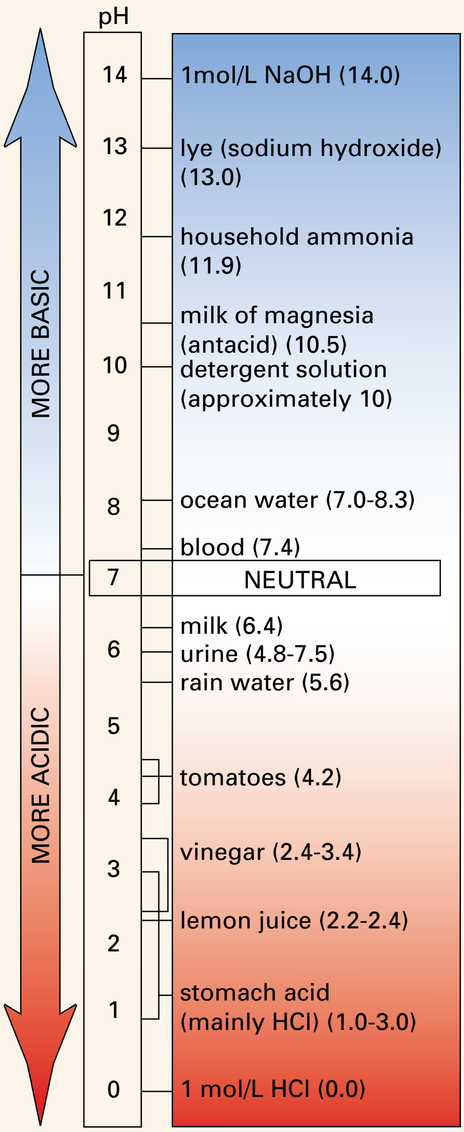
Identify the significant digits in each pH value below.

1. The pH of drain cleaner is 13.1.
2. The pH of milk is 6.4.
3. The pH of vinegar is 2.85.
4. The pH of lemon juice is 2.310.

Concept Organizer

pH, $[H_3O^+]$, and the Strength of Acids and Bases

Type of solution	$[H_3O^+]$ (mol/L)	Concentration of hydronium and hydroxide ions	pH at 25°C
acidic solution	greater than 1×10^{-7}	$[H_3O^+] > [OH^-]$	< 7.00
neutral solution	1×10^{-7}	$[H_3O^+] = [OH^-]$	7.00
basic solution	less than 1×10^{-7}	$[H_3O^+] < [OH^-]$	> 7.00



The pH values of many common solutions fall within a range from 0 to 14, as shown on this pH scale. The table above the pH scale relates the positive pH values to their hydronium ion concentrations and their logarithms.

Sample Problem

Calculating the pH of a Solution

Problem

Calculate the pH of a solution with $[H_3O^+] = 3.8 \times 10^{-3}$ mol/L.

What Is Required?

You need to calculate the pH, given $[H_3O^+]$.

What Is Given?

You know that $[H_3O^+]$ is 3.8×10^{-3} mol/L.

Plan Your Strategy

Use the equation $pH = -\log [H_3O^+]$ to solve for the unknown.

Act on Your Strategy

$$\begin{aligned} pH &= -\log (3.8 \times 10^{-3}) \\ &= 2.42 \end{aligned}$$

Check Your Solution

$[H_3O^+]$ is greater than 1.0×10^{-7} mol/L. Therefore, the pH should be less than 7.00. The solution is acidic, as you would expect.

Practice Problems

6. Calculate the pH of each solution, given the hydronium ion concentration.

(a) $[H_3O^+] = 0.0027$ mol/L
(b) $[H_3O^+] = 7.28 \times 10^{-8}$ mol/L
(c) $[H_3O^+] = 9.7 \times 10^{-5}$ mol/L
(d) $[H_3O^+] = 8.27 \times 10^{-12}$

7. $[H_3O^+]$ in a cola drink is about 5.0×10^{-3} mol/L. Calculate the pH of the drink. State whether the drink is acidic or basic.

8. A glass of orange juice has $[H_3O^+]$ of 2.9×10^{-4} mol/L. Calculate the pH of the juice. State whether the result is acidic or basic.

9. (a) $[H_3O^+]$ in a dilute solution of nitric acid, HNO_3, is 6.3×10^{-3} mol/L. Calculate the pH of the solution.

(b) $[H_3O^+]$ of a solution of sodium hydroxide is 6.59×10^{10} mol/L. Calculate the pH of the solution.

PROBLEM TIP

Appendix E, "Math and Chemistry", explains how you can do these calculations with a calculator.

SKILL FOCUS

- Predicting
- Performing and recording
- Analyzing and interpreting

Investigation 10-A

The Effect of Dilution on the pH of an Acid

In this investigation, you will compare the effects of diluting a strong acid and a weak acid.

In Part 1, you will measure the pH of a strong acid. Then you will perform a series of ten-fold dilutions. That is, each solution will be one-tenth as dilute as the previous solution. You will measure and compare the pH after each dilution.

In Part 2, you will measure the pH of a weak acid with the same initial concentration as the strong acid. Then you will perform a series of ten-fold dilutions with the weak acid. Again, you will measure and compare the pH after each dilution.

Problem

How does the pH of dilutions of a strong acid compare with the pH of dilutions of a weak acid?

Prediction

Predict each pH, and explain your reasoning.

(a) the pH of 0.10 mol/L hydrochloric acid

(b) the pH of the hydrochloric acid after one ten-fold dilution

(c) the pH of the hydrochloric acid after each of six more ten-fold dilutions

(d) the pH of 0.10 mol/L acetic acid, compared with the pH of 0.10 mol/L hydrochloric acid

(e) the pH of the acetic acid after one ten-fold dilution

Safety Precautions

Hydrochloric acid is corrosive. Wash any spills on skin or clothing with plenty of cool water. Inform your teacher immediately.

Materials

100 mL graduated cylinder
100 mL beaker
2 beakers (250 mL)
universal indicator paper and glass rod
pH meter
0.10 mol/L hydrochloric acid (for Part 1)
0.10 mol/L acetic acid (for Part 2)
distilled water

Procedure

Part 1 The pH of Solutions of a Strong Acid

1. Copy the table below into your notebook. Record the pH you predicted for each dilution.
2. Pour about 40 mL of 0.10 mol/L hydrochloric acid into a clean, dry 100 mL beaker. Use the end of a glass rod to transfer a drop of solution to a piece of universal pH paper into the acid. Compare the colour against the colour chart to determine the pH. Record the pH. Then measure and record the pH of the acid using a pH meter. Rinse the electrode with distilled water afterward.

Data Table for Part 1

$[HCl_{(aq)}]$ mol/L	Predicted pH	pH measured with universal indicator	pH measured with pH meter
1×10^{-1}			
1×10^{-2}			
1×10^{-3}			
1×10^{-4}			
1×10^{-5}			
1×10^{-6}			
1×10^{-7}			
1×10^{-8}			

3. Measure 90 mL of distilled water in a 100 mL graduated cylinder. Add 10 mL of the acid from step 2. The resulting 100 mL of solution is one-tenth as concentrated as the acid from step 2. Pour the dilute solution into a clean, dry 250 mL beaker. Use universal pH paper and a pH meter to measure the pH. Record your results.
4. Repeat step 3. Pour the new dilute solution into a second clean, dry beaker. Dispose of the more concentrated acid solution as directed by your teacher. Rinse and dry the beaker so you can use it for the next dilution.
5. Make further dilutions and pH measurements until the hydrochloric acid solution is 1.0×10^{-8} mol/L

Part 2 The pH of Solutions of a Weak Acid

1. Design a table to record your predictions and measurements for 0.10 mol/L and 0.010 mol/L concentrations of acetic acid.
2. Use the same procedure that you used in Part 1 to measure and record the pH of a 0.10 mol/L sample of acetic acid. Then dilute the solution to 0.010 mol/L. Measure the pH again.

Analysis

1. Which do you think gave the more accurate pH: the universal indicator paper or the pH meter? Explain.
2. For the strong acid, compare the pH values you predicted with the measurements you made. How can you explain any differences for the first few dilutions?
3. What was the pH of the solution that had a concentration of 1.0×10^{-8} mol/L? Explain the pH you obtained.
4. Compare the pH of 0.10 mol/L acetic acid with the pH of 0.10 mol/L hydrochloric acid. Why do you think the pH values are different, even though the concentrations of the acids were the same?
5. What effect does a ten-fold dilution of a strong acid (hydrochloric acid) have on the pH of the acid? What effect does the same dilution of a weak acid (acetic acid) have on its pH? Compare the effects for a strong acid and a weak acid. Account for any differences.

Conclusion

6. Use evidence from your investigation to support the conclusion that a weak acid ionizes less than a strong acid of identical concentration.
7. Why is the method for calculating the pH of a strong acid (if it is not too dilute) not appropriate for a weak acid?

Applications

8. Nicotinic acid is a B vitamin. The pH of a 0.050 mol/L solution of this acid is measured to be 3.08. Is it a strong acid or a weak acid? Explain. What would be the pH of a solution of nitric acid having the same concentration?
9. Would you expect to be able to predict the pH of a weak base, given its concentration? Explain. Design an experiment you could perform to check your answer.

Chemistry Bulletin

Science | Technology | Society | Environment

The Chemistry of Oven Cleaning

Oven cleaning is not a job that most people enjoy. Removing baked-on grease from inside an oven requires serious scrubbing. Any chemical oven cleaners that help to make the job easier are usually welcome. Like all chemicals, however, the most effective oven cleaners require attention to safety.

Cleaners that contain strong bases are the most effective for dissolving grease and grime. Bases are effective because they produce soaps when they react with the fatty acids in grease. When a strong base (such as sodium hydroxide, $NaOH$, or potassium hydroxide, KOH) is used on a dirty oven, the fat molecules that make up the grease are split into smaller molecules. Anions from the base then bond with some of these molecules to form soap.

One end of a soap molecule is non-polar (uncharged), so it is soluble in dirt and grease, which are also non-polar. The other end of a soap molecule is polar (charged), so it is soluble in water. Because of its two different properties, soap acts like a "bridge" between the grease and the water. Soap enables grease to dissolve in water and be washed away, thus allowing the cleaner to remove the grease from the oven surface.

Cleaners that contain sodium hydroxide and potassium hydroxide are very effective. They are also caustic and potentially very dangerous. For example, sodium hydroxide, in the concentrations that are used in oven cleaners, can irritate the skin and cause blindness if it gets in the eyes. As well, it is damaging to paints and fabrics.

There are alternatives to sodium hydroxide and other strong base cleaners. One alternative involves using ammonia, NH_3, which is a weak base. If a bowl of dilute ammonia solution is placed in an oven and left for several hours, most of the grease and grime can be wiped off.

Ammonia does not completely ionize in water. Only a small portion dissociates. Although an ammonia solution is less caustic than sodium hydroxide, it can be toxic if inhaled directly. As well, ammonia vapours can cause eye, lung, and skin irritations. At higher concentrations, ammonia can be extremely toxic.

Baking soda is a non-toxic alternative, but it is much less effective. Therefore, it requires even more scrubbing. An abrasive paste can be made by mixing baking soda and water. The basic properties of baking soda also have a small effect on grease and grime if it is applied to the oven and left for several hours.

Making Connections

1. Survey the cleaners in your home or school. Which cleaners contain bases and which contain acids? What cleaning jobs can an acid cleaner perform well? How do most acid cleaners work?
2. Some companies claim to make environmentally sensitive cleaners. Investigate these cleaners. What chemicals do they contain? See if you can infer how they work. You might like to design a controlled experiment to test the effectiveness of several oven cleaners. CAUTION Obtain permission from your teacher before performing such an experiment.

Section Wrap-up

In this section, you considered the relationship among the strength of acids and bases, the concentration of hydronium and hydroxide ions, and pH. Much of the time, you examined acids and bases acting independently of each other. However, acids and bases often interact. In fact, acid-base reactions have many important applications in the home, as well as in the laboratory. In section 10.3, you will investigate acid-base reactions.

Section Review

1. K/U Distinguish, in terms of degree of dissociation, between a strong acid and a weak acid, and a strong base and a weak base.
2. K/U Give one example of the following:
   (a) a weak acid
   (b) a strong acid
   (c) a strong base
   (d) a weak base
3. K/U Formic acid, HCOOH, is responsible for the painful bites of fire ants. Is formic acid strong or weak? Explain.
4. K/U $KMnO_4$ is an intense purple-coloured solid that can be made into a solution to kill bacteria. What is the name of this compound? Give the name and the formula of the acid that forms when $KMnO_4$ combines with water.
5. K/U State the name or the formula for the acid that forms from each of the following anions:
   (a) hydrogen sulfate
   (b) F^-
   (c) HS^-
   (d) bromite
6. K/U Explain the meaning of pH, both in terms of hydrogen ions and hydronium ions.
7. K/U Arrange the following foods in order of increasing acidity: beets, pH = 5.0; camembert cheese, pH = 7.4; egg white, pH = 8.0; sauerkraut, pH = 3.5; yogurt, pH = 4.5.
8. I Calculate the pH of each body fluid, given the concentration of hydronium ions.
   (a) tears, $[H_3O^+] = 4.0 \times 10^{-8}$ mol/L
   (b) stomach acid, $[H_3O^+] = 4.0 \times 10^{-2}$ mol/L
9. I Calculate the pH of the solution that is formed by diluting 50 mL of 0.025 mol/L hydrochloric acid to a final volume of 1.0 L
10. C What is $[H_3O^+]$ in a solution with pH = 0? Why do chemists not usually use pH to describe $[H_3O^+]$ when the pH value would be a negative number?

Unit Issue Prep

As you investigate the contamination of Prince Edward Island's soils with sodium arsenite from pesticides, investigate the link between the pH and solubility. For example, water polluted with sodium arsenite may be treated with lime (calcium oxide), CaO. What is the purpose of this treatment?

10.3 Acid-Base Reactions

Section Preview/ Specific Expectations

In this section, you will

- **perform** calculations involving neutralization reactions
- **determine** the concentration of an acid in solution by conducting a titration
- **communicate** your understanding of the following terms: *neutralization reaction, salt, acid-base indicator, titration, equivalence point, end-point*

Is there a box of baking soda in your refrigerator at home? Baking soda is sodium hydrogen carbonate. (It is also commonly called sodium bicarbonate.) Baking soda removes the odours caused by spoiling foods. The smelly breakdown products of many foods are acids. Baking soda, a base, eliminates the odours by neutralizing the characteristic properties of the acids.

Adding a base to an acid neutralizes the acid's acidic properties. This type of reaction is called a **neutralization reaction.**

There are many different acids and bases. Being able to predict the results of reactions between them is important. Bakers, for example, depend on neutralization reactions to create light, fluffy baked goods. Gardeners and farmers depend on these reactions to modify the characteristics of the soil. Industrial chemists rely on these reactions to produce the raw materials that are used to make a wide variety of chemicals and chemical products. (See Figure 10.13.)

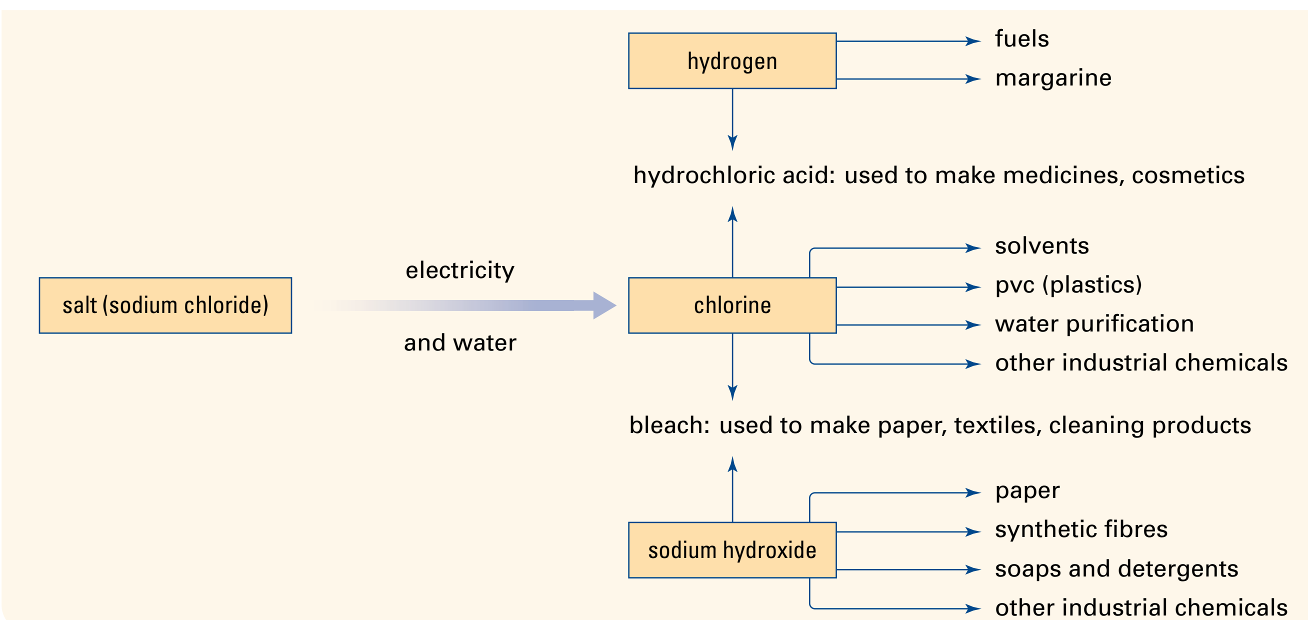


Figure 10.13 **You know sodium chloride as common table salt. As you can see here, however, sodium chloride is anything but common. Sodium chloride is a product of an acid-base reaction between hydrochloric acid and sodium hydroxide.**

Neutralization Reactions

The reaction between an acid and a base produces an ionic compound (a salt) and water.

$$\text{acid} + \text{base} \rightarrow \text{a salt} + \text{water}$$

A **salt** is an ionic compound that is composed of the anion from an acid and a cation from a base. For example, sodium nitrate is a salt that is found in many kitchens. It is often added to processed meat to preserve the colour and to slow the rate of spoiling by inhibiting bacterial growth. Sodium nitrate can be prepared in a laboratory by reacting nitric acid with sodium hydroxide, as shown on the next page.

$$\text{H}\underbrace{\text{NO}_{3(aq)}}_{\text{anion from acid}} + \overbrace{\text{Na}}^{\text{cation from base}}\text{OH}_{(aq)} \rightarrow \text{NaNO}_{3(aq)} + \text{H}_2\text{O}_{(\ell)}$$

The balanced chemical equation for this reaction shows that 1 mol of nitric acid reacts with 1 mol of sodium hydroxide. If equal molar quantities of nitric acid and sodium hydroxide are used, the result is a neutral (pH 7) aqueous solution of sodium nitrate. In fact, *when any strong acid reacts with any strong base in the mole ratio from the balanced chemical equation, a neutral aqueous solution of a salt is formed.* Reactions between acids and bases of different strengths usually do not result in neutral solutions.

For most neutralization reactions, there are no visible signs that a reaction is occurring. How can you determine that a neutralization reaction is taking place? One way is to use an **acid-base indicator.** This is a substance that changes colour in acidic and basic solutions. Most acid-base indicators are weak, monoprotic acids. The undissociated weak acid is one colour. Its conjugate base is a different colour.

$$\overbrace{\text{H (indicator)}_{(aq)}}^{\text{colour 1}} \rightleftharpoons \text{H}^+ + \underbrace{(\text{indicator})^-_{(aq)}}_{\text{colour 2}}$$

In an acidic solution, the indicator does not dissociate very much. It appears as colour 1. In a basic solution, the indicator dissociates much more. It appears as colour 2. Often a single drop of indicator causes a dramatic change in colour. For example, phenolphthalein is an indicator that chemists often use for reactions between a strong acid and a strong base. It is colourless between pH 0 and pH 8. It turns pink between pH 8 and pH 10. (See Figure 10.14.)

Figure 10.14 **A good indicator, such as the phenolphthalein shown here, must give a vivid colour change.**

CHEM FACT

If a small quantity of an acid or a base is spilled in a laboratory, you can use a neutralization reaction to minimize the hazard. To neutralize a basic solution spill, you can add solid sodium hydrogen sulfate or citric acid. For an acidic solution spill, you can use sodium hydrogen carbonate (baking soda). Note that you cannot use a strong acid or base to clean up a spill. This would result in another hazardous spill. As well, the neutralization reaction would generate a lot of heat, and thus produce a very hot solution.

CHECKPOINT

Show that the net ionic equation for the reaction between HNO_3 (a strong acid) and NaOH (a strong base) results in the formation of water.

Calculations Involving Neutralization Reactions

Suppose that a solution of an acid reacts with a solution of a base. You can determine the concentration of one solution if you know the concentration of the other. (This assumes that the volumes of both are accurately measured.) Use the concentration and volume of one solution to determine the amount (in moles) of reactant that it contains. The balanced chemical equation for the reaction describes the mole ratio in which the compounds combine. In the following Sample Problems and Practice Problems, you will see how to do these calculations.

CHEM FACT

An old remedy to relieve the prickly sting of a nettle plant is to rub the area with the leaf of a dock plant. The sting contains an acid. This acid is neutralized by a base that is present in the dock leaf. Bees and ants also have an acidic sting. You can wash the sting with soap, because soap is basic. You can also apply baking soda (a base) to the skin for more effective relief. If you are stung by a wasp, however, you should apply vinegar. The sting of a wasp contains a base.

Sample Problem

Finding Concentration

Problem

13.84 mL of hydrochloric acid, $HCl_{(aq)}$, just neutralizes 25.00 mL of a 0.1000 mol/L solution of sodium hydroxide, $NaOH_{(aq)}$. What is the concentration of the hydrochloric acid?

What Is Required?

You need to find the concentration of the hydrochloric acid.

What Is Given?

Volume of hydrochloric acid, HCl = 13.84 mL
Volume of sodium hydroxide, NaOH = 25.00 mL
Concentration of sodium hydroxide, NaOH = 0.1000 mol/L

Plan Your Strategy

Step 1 Write the balanced chemical equation for the reaction.
Step 2 Calculate the amount (in mol) of sodium hydroxide added, based on the volume and concentration of the sodium hydroxide solution.
Step 3 Determine the amount (in mol) of hydrochloric acid needed to neutralize the sodium hydroxide.
Step 4 Find $[HCl_{(aq)}]$, based on the amount and volume of hydrochloric acid solution needed.

Act on Your Strategy

Step 1 The balanced chemical equation is

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)}$$

Step 2

$$\text{Amount (in mol)} = \text{Concentration (in mol/L)} \times \text{Volume (in L)}$$

$$\text{Amount NaOH (in mol) added} = 0.1000\ \text{mol}/\cancel{\text{L}} \times 0.02500\ \cancel{\text{L}} = 2.500 \times 10^{-3}\ \text{mol}$$

PROBLEM TIPS

1. Make sure that the values you use in your calculations refer to the same reactant. For example, you can use the concentration and volume of sodium hydroxide to find the amount of sodium hydroxide in this problem. You cannot use the concentration of sodium hydroxide and the volume of hydrochloric acid.
2. In the solution, the volumes are converted to litres. If all the volumes are expressed in the same unit, the conversion step is not necessary.
3. Do not drop significant digits, even zeros, during your calculations.

Continued ...

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Step 3 HCl reacts with NaOH in a 1:1 ratio, so there must be 2.500×10^{-3} mol HCl.

Step 4
$$\text{Concentration (in mol/L)} = \frac{\text{Amount (in mol)}}{\text{Volume (in L)}}$$
$$[\text{HCl}_{(aq)}] = \frac{2.500 \times 10^{3}\ \text{mol}}{0.01384\ \text{L}}$$
$$= 0.1806\ \text{mol/L}$$

Therefore, the concentration of hydrochloric acid is 0.1806 mol/L.

Check Your Solution

$[HCl_{(aq)}]$ is greater than $[NaOH_{(aq)}]$. This is reasonable because a smaller volume of hydrochloric acid was required. As well, the balanced equation shows a 1:1 mole ratio between these reactants.

Sample Problem

Finding Volume

Problem

What volume of 0.250 mol/L sulfuric acid, $H_2SO_{4(aq)}$, is needed to react completely with 37.2 mL of 0.650 mol/L potassium hydroxide, $KOH_{(aq)}$?

What Is Required?

You need to find the volume of sulfuric acid.

What Is Given?

Concentration of sulfuric acid, H_2SO_4 = 0.250 mol/L
Concentration of potassium hydroxide, KOH = 0.650 mol/L
Volume of potassium hydroxide, KOH = 37.2 mL.

Plan Your Strategy

Step 1 Write the balanced chemical equation for the reaction.

Step 2 Calculate the amount (in mol) of potassium hydroxide, based on the volume and concentration of the potassium hydroxide solution.

Step 3 Determine the amount (in mol) of sulfuric acid that is needed to neutralize the potassium hydroxide.

Step 4 Find the volume of the sulfuric acid, based on the amount and concentration of sulfuric acid needed.

Act on Your Strategy

Step 1 The balanced chemical equation is
$$H_2SO_{4(aq)} + 2KOH \rightarrow K_2SO_{4(aq)} + 2H_2O_{(\ell)}$$

Step 2
$$\text{Amount (in mol) of KOH} = 0.650\ \text{mol/}\cancel{\text{L}} \times 0.0372\ \cancel{\text{L}}$$
$$= 0.02418\ \text{mol}$$

Continued ...

Continued ...
FROM PAGE 397

Step 3 H_2SO_4 reacts with KOH in a 1:2 mole ratio. The amount of H_2SO_4 needed is

$$\frac{1 \text{ mol } H_2SO_4}{2 \text{ mol KOH}} = \frac{0.01209 \text{ mol } H_2SO_4}{0.2418 \text{ mol KOH}}$$

$$\frac{0.024\ 18 \text{ mol KOH} \times 1 \text{ mol } H_2SO_4}{2 \text{ mol KOH}} = 0.01209 \text{ mol } H_2SO_4$$

Step 4 Amount (in mol) H_2SO_4

$$= 0.01209 \text{ mol}$$

$$= 0.250 \text{ mol/L} \times \text{Volume } H_2SO_{4(aq)} \text{ (in L)}$$

$$\text{Volume } H_2SO_{4(aq)} = \frac{0.01209 \cancel{\text{mol}}}{0.250 \cancel{\text{mol}}/\text{L}}$$

$$= 0.04836 \text{ L}$$

Therefore, the volume of sulfuric acid that is needed is 48.4 mL.

Check Your Solution

The balanced chemical equation shows that half the amount of sulfuric acid will neutralize a given amount of potassium hydroxide. The concentration of sulfuric acid, however, is less than half the concentration of potassium hydroxide. Therefore, the volume of sulfuric acid should be greater than the volume of potassium hydroxide.

Practice Problems

10. 17.85 mL of nitric acid neutralizes 25.00 mL of 0.150 mol/L $NaOH_{(aq)}$. What is the concentration of the nitric acid?

11. What volume of 1.015 mol/L magnesium hydroxide is needed to neutralize 40.0 mL of 1.60 mol/L hydrochloric acid?

12. What volume of 0.150 mol/L hydrochloric acid is needed to neutralize each solution below?

(a) 25.0 mL of 0.135 mol/L sodium hydroxide

(b) 20.0 mL of 0.185 mol/L ammonia solution

(c) 80 mL of 0.0045 mol/L calcium hydroxide

13. What concentration of sodium hydroxide solution is needed for each neutralization reaction?

(a) 37.82 mL of sodium hydroxide neutralizes 15.00 mL of 0.250 mol/L hydrofluoric acid.

(b) 21.56 mL of sodium hydroxide neutralizes 20.00 mL of 0.145 mol/L sulfuric acid.

(c) 14.27 mL of sodium hydroxide neutralizes 25.00 mL of 0.105 mol/L phosphoric acid.

Acid-Base Titration

In the previous Sample Problems and Practice Problems, you were given the concentrations and volumes you needed to solve the problems. What if you did not have some of this information? Chemists often need to know the concentration of an acidic or basic solution. To acquire this information, they use an experimental procedure called a titration. In a **titration**, the concentration of one solution is determined by quantitatively observing its reaction with a solution of known concentration. The solution of known concentration is called a *standard solution*. The aim of a titration is to find the point at which the number of moles of the standard solution is stoichiometrically equal to the original number of moles of the unknown solution. This point is referred to as the **equivalence point.** At the equivalence point, all the moles of hydrogen ions that were present in the original volume of one solution have reacted with an equal number of moles of hydroxide ions from the other solution.

Precise volume measurements are needed when you perform a titration. Chemists use special glass apparatus to collect these measurements. (See Figure 10.15.) As well, an acid-base indicator is needed to monitor changes in pH during the titration.

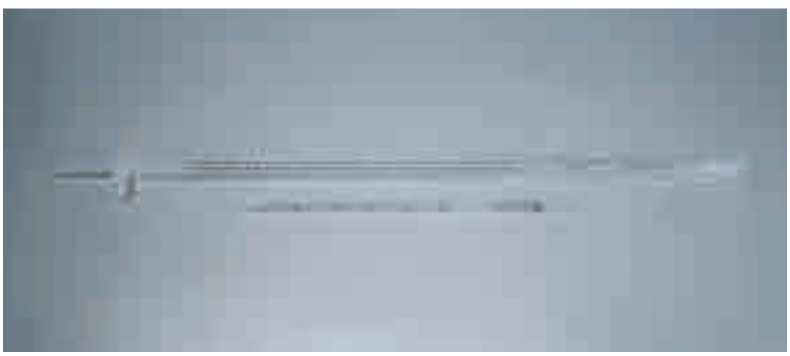

Figure 10.15 **A transfer pipette (bottom) measures a fixed volume of liquid, such as 10.00 mL, 25.00 mL, or 50.0 mL. A burette (top) measures a variable volume of liquid.**

In a titration, a pipette is used to measure a precise volume of standard solution into a flask. The flask sits under a burette that contains the solution of unknown concentration. After adding a few drops of indicator, you take an initial burette reading. Then you start adding the known solution, slowly, to the flask. The **end-point** of the titration occurs when the indicator changes colour. The indicator is chosen so that it matches its equivalance point.

Titration Step by Step

The following pages outline the steps that you need to follow to prepare for a titration. Review these steps carefully. Then observe as your teacher demonstrates them for you. At the end of this section, in Investigation 10-B, you will perform your own titration of a common substance: vinegar.

Web LINK

www.school.mcgrawhill.ca/resources

Sometimes a moving picture is worth a thousand words. To enhance your understanding, your teacher will demonstrate the titration procedure described in this textbook. In addition, some web sites provide downloadable or real-time titration movies to help students visualize the procedure and its techniques. Go to the web site above, then to Science Resources and to Chemistry 11 to see where to go next. Compare the different demonstrations you can find and observe, including your teacher's. Prepare your own set of "Titration Tips" to help you recall important details.

PROBEWARE

If you have access to probeware, try the Chemistry 11 lab, Titrating an Unknown, or a similar lab from a probeware company

TITRATION TIP

Never use your mouth in place of a suction bulb to draw a liquid into a pipette. The liquid could be corrosive or poisonous. As well, you will contaminate the glass stem.

TITRATION TIP

Practice removing the bulb and replacing it with your index finger. You need to be able to perform this action quickly and smoothly.

Rinsing the Pipette

A pipette is used to measure and transfer a precise volume of liquid. You rinse a pipette with the solution whose volume you are measuring. This ensures that any drops that remain inside the pipette will form part of the measured volume.

1. Pour a sample of standard solution into a clean, dry beaker.
2. Place the pipette tip in a beaker of distilled water. Squeeze the suction bulb. Maintain your grip while placing it over the stem of the pipette. (If your suction bulbs have valves, your teacher will show you how to use them.)
3. Relax your grip on the bulb to draw up a small volume of distilled water.
4. Remove the bulb and discard the water by letting it drain out.
5. Rinse the pipette by drawing several millilitres of solution from the beaker into it. Rotate and rock the pipette to coat the inner surface with solution. Discard the rinse. Rinse the pipette twice in this way. It is now ready to fill with standard solution.

Filling the Pipette

6. Place the tip of the pipette below the surface of the solution.
7. Hold the suction bulb loosely on the end of the glass stem. Use the suction bulb to draw liquid up just past the etched volume mark. (See Figure 10.16.)
8. As quickly and smoothly as you can, slide the bulb off and place your index finger over the end of the glass stem.
9. Gently roll your finger slightly away from end of the stem to let solution drain slowly out.
10. When the bottom of the meniscus aligns with the etched mark, as in Figure 10.17, press your finger back over the end of the stem. This will prevent more solution from draining out.
11. Touch the tip of the pipette to the side of the beaker to remove any clinging drop. See Figure 10.18. The measured volume inside the pipette is now ready to transfer to an Erlenmeyer flask or a volumetric flask.

Transferring the Solution

12. Place the tip of the pipette against the inside glass wall of the flask. Let the solution drain slowly, by removing your finger from the stem.
13. After the solution drains, wait several seconds, then touch the tip to the inside wall of the flask to remove any drop on the end. Note: You may notice a small amount of liquid remaining in the tip. The pipette was calibrated to retain this amount. Do not try to remove it.

Figure 10.16 Draw a bit more liquid than you need into the pipette. It is easier to reduce this volume than it is to add more solution to the pipettes.

Figure 10.17 The bottom of the meniscus must align exactly with the etched mark.

Figure 10.18 You can prevent a "stubborn" drop from clinging to the pipette tip by touching the tip to the inside of the glass surface.

Adding the Indicator

14. Add two or three drops of indicator to the flask and its contents. Do not add too much indicator. Using more does not make the colour change easier to see. Also, indicators are usually weak acids. Too much can change the amount of base needed for neutralization. You are now ready to prepare the apparatus for the titration.

Rinsing the Burette

A burette is used to accurately measure the volume of liquid added during a titration experiment. It is a graduated glass tube with a tap at one end.

15. To rinse the burette, close the tap and add about 10 mL of distilled water from a wash bottle.
16. Tip the burette to one side and roll it gently back and forth so that the water comes in contact with all inner surfaces.
17. Hold the burette over a sink. Open the tap, and let the water drain out. While you do this, check that the tap does not leak. Make sure that it turns smoothly and easily.
18. Rinse the burette with 5 mL to 10 mL of the solution that will be measured. Remember to open the tap to rinse the lower portion of the burette. Rinse the burette twice, discarding the liquid each time.

Filling the Burette

19. Assemble a retort stand and burette clamp to hold the burette. Place a funnel in the top of the burette.
20. With the tap closed, add solution until the liquid is above the zero mark. Remove the funnel. Carefully open the tap. Drain the liquid into a beaker until the bottom of the meniscus is at or below the zero mark.
21. Touch the tip of the burette against the beaker to remove any clinging drop. Check that the portion of the burette that is below the tap is filled with liquid and contains no air bubbles.
22. Record the initial burette reading in your notebook.
23. Replace the beaker with the Erlenmeyer flask that you prepared earlier. Place a sheet of white paper under the Erlenmeyer to help you see the indicator colour change that will occur near the end-point.

Reading the Burette

24. A meniscus reader is a small white card with a thick black line on it. Hold the card behind the burette, with the black line just under the meniscus, as in Figure 10.20. Record the volume added from the burette to the nearest 0.05 mL.

Figure 10.20 A meniscus reader helps you read the volume of liquid in the burette more easily

TITRATION TIP

If you are right-handed, the tap should be on your right as you face the burette. Use your left hand to operate the tap. Use your right hand to swirl the liquid in the Erlenmeyer flask. If you are left-handed, reverse this arrangement.

TITRATION TIP

Near the end-point, when you see the indicator change colour as liquid enters the flask from the burette, slow the addition of liquid. The end-point can occur very quickly.

TITRATION TIP

Observe the level of solution in the burette so that your eye is level with the bottom of the meniscus.

Investigation 10-B

SKILL FOCUS
Predicting
Performing and recording
Analyzing and interpreting

The Concentration of Acetic Acid in Vinegar

Vinegar is a dilute solution of acetic acid, CH_3COOH. Only the hydrogen atom that is attached to an oxygen atom is acidic. Thus, acetic acid is monoprotic. As a consumer, you can buy vinegar with different concentrations. For example, the concentration of table vinegar is different from the concentration of the vinegar that is used for pickling foods. To maintain consistency and quality, manufacturers of vinegar need to determine the percent concentration of acetic acid in the vinegar. In this investigation, you will determine the concentration of acetic acid in a sample of vinegar.

Prediction

Which do you predict has the greater concentration of acetic acid: table vinegar or pickling vinegar? Give reasons for your prediction.

Materials

pipette
suction bulb
retort stand
burette
burette clamp
3 beakers (250 mL)
3 Erlenmeyer flasks (250 mL)
labels
meniscus reader
sheet of white paper
funnel
table vinegar
pickling vinegar
sodium hydroxide solution
distilled water
dropper bottle containing phenolphthalein

Safety Precautions

Both vinegar and sodium hydroxide solutions are corrosive. Wash any spills on skin or clothing with plenty of water. Inform your teacher immediately.

Procedure

1. Record the following information in your notebook. Your teacher will tell you the concentration of the sodium hydroxide solution.
   - concentration of $NaOH_{(aq)}$ (in mol/L)
   - type of vinegar solution
   - volume of pipette (in mL)
2. Copy the table below into your notebook, to record your observations.

Burette Readings for the Titration of Acetic Acid

Reading (mL)	Trial 1	Trial 2	Trial 3
final reading			
initial reading			
volume added			

3. Label a clean, dry beaker for each liquid: $NaOH_{(aq)}$, vinegar, and distilled water. Obtain each liquid. Record the type of vinegar you will be testing.
4. Obtain a pipette and a suction bulb. Record the volume of the pipette for trial 1. Rinse it with distilled water, and then with vinegar.
5. Pipette some vinegar into the first Erlenmeyer flask. Record this amount. Add approximately 50 mL of water. Also add two or three drops of phenolphthalein indicator.
6. Set up a retort stand, burette clamp, burette, and funnel. Rinse the burette first with distilled water. Then rinse it with sodium hydroxide solution. Make sure that there are no air bubbles in the burette. Also make sure

that the liquid fills the tube below the glass tap. Remove the funnel before beginning the titration.

7. Place a sheet of white paper under the Erlenmeyer flask. Titrate sodium hydroxide into the Erlenmeyer flask while swirling the contents. The end-point of the titration is reached when a permanent pale pink colour appears. If you are not sure whether you have reached the end-point, take the burette reading. Add one drop of sodium hydroxide, or part of a drop. Observe the colour of the solution. If you go past the end-point, the solution will become quite pink.
8. Repeat the titration twice more. Record your results for each of these trials.
9. When you have finished all three trials, dispose of the chemicals as directed by your teacher. Rinse the pipette and burette with distilled water. Leave the burette tap open.

Analysis

1. Average the two closest burette readings. Average all three readings if they agree within about ±0.2 mL.
2. Write the chemical equation for the reaction of acetic acid with sodium hydroxide.
3. Calculate the concentration of acetic acid in your vinegar sample. Use the average volume and concentration of sodium hydroxide, and the volume of vinegar.
4. Find the molar mass of acetic acid. Then calculate the mass of acid in the volume of vinegar you used.
5. The density of vinegar is 1.01 g/mL. (The density of the more concentrated vinegar solution is greater than the density of the less concentrated solution. You can ignore the difference, however.) Calculate the mass of the vinegar sample. Find the percent by mass of acetic acid in the sample.

Conclusions

6. Compare your results with the results of other students who used the same type of vinegar. Then compare the concentration of acetic acid in table vinegar with the concentration in pickling vinegar. How did your results compare with your prediction?
7. List several possible sources of error in this investigation.

Application

8. Most shampoos are basic. Why do some people rinse their hair with vinegar after washing it?

Section Wrap-up

In this section, as in much of Unit 3, you combined liquid solutes and liquid solvents. You have learned how to describe the concentration of ions that determine the acidic or basic nature of a solution. As well, you performed calculations to determine the concentration of acids and many other substances in solution. In the upcoming unit, you will apply your understanding of stoichiometry and solutions by examining the nature and interactions of substances in the gaseous state.

Section Review

1 K/U Write a generalized word equation to describe what happens during a neutralization reaction.

2 K/U Write a chemical equation for each neutralization reaction.

(a) KOH with HNO_3

(b) HBr with $Ca(OH)_2$

(c) H_3PO_4 with NaOH

(d) $Mg(OH)_2$ (the active ingredient in milk of magnesia, an antacid) with HCl (the acid in your stomach)

3 K/U Distinguish between the equivalence point and the end-point in a titration. Why might they be different? How would this affect the result of a titration?

4 I A 25.0 mL sample of sulfuric acid is completely neutralized by adding 32.8 mL of 0.116 mol/L ammonia solution. Ammonium sulfate, $(NH_4)_2SO_4$, and water are formed. What is the concentration of the sulfuric acid?

5 I The following data were collected during a titration. Calculate the concentration of the sodium hydroxide solution.

Titration Data

Volume of $HCl_{(aq)}$	10.00 mL
Final volume of $NaOH_{(aq)}$	23.08 mL
Initial volume of $NaOH_{(aq)}$	1.06 mL
Concentration of $HCl_{(aq)}$	0.235 mol/L

6 I You should always put the two solutions for a titration experiment in clean, dry beakers. You do not need to dry the Erlenmeyer flask to which you add the solutions, however, if it has been thoroughly rinsed with distilled water. Explain the difference in these procedures.

7 C Suppose that a laboratory technician accidentally spills a dilute solution of a strong acid on her hands, the sleeve of her lab coat, and the laboratory bench. Explain how she would deal with this spill.

CHAPTER 10 Review

Reflecting on Chapter 10

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Compare the properties of acids and bases.
- Distinguish the two theories for explaining the behaviour and composition of acids and bases.
- Identify conjugate acid-base pairs for selected acid-base reactions.
- Describe strong and weak acids and bases on the basis of their dissociation.
- Explain the significance of the concentration of hydronium ion in describing pH.
- Dilute an acid and describe the effect on its pH.
- Describe and conduct a titration procedure.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

Arrhenius theory	weak base
hydronium ion	binary acid
Brønsted-Lowry theory	oxoacid
conjugate acid-base pair	pH
conjugate base	neutralization reaction
conjugate acid	salt
strong acid	acid-base indicator
weak acid	titration
strong base	equivalence point
	end-point

Knowledge/Understanding

1. Use the Arrhenius theory and then the Brønsted-Lowry theory to describe the following concepts. If one of the theories does not apply, state that this is the case.
   (a) composition of an acid and a base
   (b) conductivity of an acidic or basic solution
   (c) interaction between an acid and water
   (d) interaction between a base and water
   (e) conjugate acid-base pairs
   (f) strong and weak acids and bases
   (g) the pH of a solution
2. How does diluting an acidic or basic solution affect the pH of the solution?
3. Codeine is a compound that is extracted from opium. It is used for pain relief. The pH of a 0.020 mol/L solution of codeine is 10.26. Is codeine an acid or a base? Is it strong or weak? Explain how you decided.
4. Sodium hydrogen carbonate, $NaHCO_3$ (commonly called sodium bicarbonate, or bicarbonate of soda), is commonly used in baked goods. It dissolves in water to form an alkaline solution.
   (a) Is the pH of $NaHCO_{3(aq)}$ greater or less than 7.00?
   (b) Write the name and formula of an acid and a base that react together to form this compound. Identify each as strong or weak.
5. Aluminum sulfate, $Al_2(SO_4)_3$, is used to help clarify water. In aqueous solution, it is slightly acidic.
   (a) Is the pH of $Al_2(SO_4)_{3(aq)}$ greater or less than 7.00?
   (b) Write the name and formula of an acid and a base that react together to form aluminum sulfate. Identify each as either strong or weak.
6. Sodium hydrogen carbonate can be used to neutralize an acid. The hydrogen carbonate ion is the conjugate base of which acid?
7. Write the net ionic equation for the reaction of aqueous sodium hydroxide with aqueous nitric acid.
8. In different reactions in aqueous solution, the hydrogen carbonate ion can act as an acid or a base. Write the chemical formula of the conjugate acid and the conjugate base of the hydrogen carbonate ion, $HCO_3{}^-{}_{(aq)}$. Then complete the following equations. State whether the ion is a Brønsted-Lowry acid or a base.
   (a) $HCO_3{}^-{}_{(aq)} + H_3O^+{}_{(aq)} \rightarrow$
   (b) $HCO_3{}^-{}_{(aq)} + OH^-{}_{(aq)} \rightarrow$
9. Which of the following are conjugate acid-base pairs? For those pairs that are not conjugates, write the correct conjugate acid or base for each compound or ion.
   (a) HNO_3/OH^-
   (b) $NH_4{}^+/NH_3$
   (c) $HSO_4{}^-/SO_4{}^{2-}$
   (d) $H_3PO_4/PO_4{}^{3-}$
10. In the laboratory, you have samples of three different acids of equal concentration: a 1.0 mol/L solution of acetic acid, a 1.0 mol/L solution of hydrochloric acid, and a 1.0 mol/L solution of sulfuric acid.
    (a) How would the pH of each acid solution compare? Explain.

(b) If samples of each acid were used in separate titration experiments with 0.50 mol/L sodium hydroxide solution, how would the volume of acid required for neutralization compare? State your reasoning.

11. Write balanced chemical equations for the following reactions:
   (a) calcium oxide with hydrochloric acid
   (b) magnesium with sulfuric acid
   (c) sodium carbonate with nitric acid

12. Domestic bleach is typically a 5% solution of sodium hypochlorite, $NaOCl_{(aq)}$. It is made by bubbling chlorine gas through a solution of sodium hydroxide.
   (a) Write a balanced chemical equation showing the reaction that takes place.
   (b) In aqueous solution, the hypochlorite ion combines with $H^+_{(aq)}$ present in water to form hypochlorous acid. Write the equation for this reaction. Is the hypochlorite ion acting as an acid or a base?

13. In this chapter, you are told that $[H_3O^+]$ in pure water is 1.0×10^{-7} mol/L at 25˚C. Thus, two out of every one billion water molecules have dissociated. Check these data by answering the following questions.
   (a) What is the mass (in g) of 1.0 L of water?
   (b) Calculate the amount (in mol) of water in 1.0 L. This is the concentration of water in mol/L.
   (c) Divide the concentration of hydronium ions by the concentration of water. Your answer should be about 2 ppb.

Inquiry

14. 80.0 mL of 4.00 mol/L, H_2SO_4 are diluted to 400.0 mL by adding water. What is the molar concentration of the sulfuric acid after dilution?

15. In a titration experiment, 25.0 mL of an aqueous solution of sodium hydroxide was required to neutralize 50.0 mL of 0.010 mol/L hydrochloric acid. What is the molar concentration of the sodium hydroxide solution?

16. A burette delivers 20 drops of solution per 1.0 mL. What amount (mol) of $H^+_{(aq)}$ is present in one drop of a 0.20 mol/L HCl solution?

17. How is a 1.0 mol/L solution of hydrochloric acid different from a 1.0 mol/L solution of acetic acid? Suppose that you added a strip of magnesium metal to each acid. Would you observe any differences in the reactions? Explain your answer so that grade 9 students could understand it.

Communication

18. Commercial processors of potatoes remove the skin by using a 10-20% by mass solution of sodium hydroxide. The potatoes are soaked in the solution for a few minutes at 60-70˚C, after which the peel can be sprayed off using fresh water. You work in the laboratory at a large food processor and must analyse a batch of sodium hydroxide solution. You pipette 25.00 mL of $NaOH_{(aq)}$, and find it has a mass of 25.75 g. Then you titrate the basic solution against 1.986 mol/L HCI, and find it requires 30.21 mL of acid to reach an end point.
   (a) Inform your supervisor what the molar concentration of the sodium hydroxide is.
   (b) The mass percent of NaOH present must be a minimum of 10% for the solution to be used. Advise your supervisor whether or not the solution can be used to process more potatoes, and explain your reasoning.

19. Ammonia is an important base, used to make fertilizers, nylon, and nitric acid. The manufacture of ammonia depends on a process discovered by Fritz Haber (1868-1934). After gathering information from print or electronic resources, write an obituary for Haber. Describe his accomplishments and the effect on society of plentiful supplies of ammonia.

Making Connections

20. Limestone, chalk, and marble are all forms of calcium carbonate. Limestone rock can be used to build roads, but it is a very important basic compound used in large quantities by chemical industries. For example, limestone is used directly to make concrete and cement. It is also used in the manufacture of glass and in agriculture. Limestone is often processed to make quicklime, CaO, and hydrated lime (calcium hydroxide), $Ca(OH)_2$.

(a) Research the uses of quicklime and hydrated lime. Investigate one of these uses further.

(b) Design a poster illustrating the use you decided to research. Your poster should be both informative and visually interesting. Include a bibliography showing the resources you found useful.

21. On several occasions during the past few years, you have studied the environmental issue of acid rain. Now that you have further developed your understanding of acids and bases in this chapter, reflect on your earlier understandings.

(a) List two facts about acid rain that you now understand in a more comprehensive way. Explain what is different between your previous and your current understanding in each case.

(b) Identify three questions that your teacher could assign as a research project on acid rain. The emphasis of the research must be on how an understanding of chemistry can contribute clarifying the questions and possible solutions involved in this issue. Develop a rubric that would be used to assess any student who is assigned this research project.

22. Research the use of hypochlorous acid in the management of swimming pools and write a report on your findings. Include a discussion on the importance of controlling pool water.

Answers to Practice Problems and Short Answers to Section Review Questions

Practice Problems: 1. HCN/CN^- and H_2O/H_3O^+ **2.** H_2O/OH^- and CH_3COO^-/CH_3COOH **3.** as an acid: $HS^-_{(aq)} + H_2O_{(\ell)} \leftrightarrow S^{2-}_{(aq)} + H_3O^+_{(aq)}$; as a base: $HS^-_{(aq)} + H_2O_{(\ell)} \leftrightarrow H_2S_{(aq)} + OH^-_{(aq)}$ **4.(a)** HBr; bromate; Br^- **(b)** hydrogen sulfide, HS^-; sulfide, S_2^- **5.(a)** nitric acid, HNO_3; nitrate, NO_3^- **(b)** nitrous acid, HNO_2; nitrite, NO_2^- **(c)** hyponitrous acid, HNO; hyponitrite, NO^- **(d)** phosphoric acid, H_3PO_4; dihydrogen phosphate, $H_2PO_4^-$, hydrogen phosphate, HPO_4^{2-}, phosphate, PO_4^{3-} **(e)** phosphorous acid, H_3PO_3; dihydrogen phosphite, $H_2PO_3^-$, hydrogen phosphite, HPO_3^{2-}, phosphite, PO_3^{3-} **(f)** periodic acid, HIO_4; periodate, IO_4^- **6.(a)** 2.57 **(b)** 7.138 **(c)** 4.01 **(d)** 11.082 **7.** pH = 2.30; acidic **8.** 3.54; acidic **9.(a)** 2.20 **(b)** 9.181; basic **10.** 0.210 mol/L **11.** 31.5 mL **12.(a)** 22.5 mL **(b)** 24.7 mL **(c)** 4.8 mL **13.(a)** 0.0992 mol/L **(b)** 0.269 mol/L **(c)** 0.552 mol/L

Section Review: 10.1: 1.(a) B and D **(b)** A (C, although unintended, would also be correct) **(c)** A, B, D **(d)** no **(e)** litmus test, for example **5.(a)** H_2O **(b)** HCO_3^- **6.(a)** NO_3^- **(b)** SO_4^{2-} **7.** c and d **8.** a and b **9.(a)** $HF_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + F^-_{(aq)}$ **(b)** HF/F^-; H_2O/H_3O^+ **10.(a)** $H_2PO_4^-/HPO_4^{2-}$ and CO_3^{2-}/HCO_3^- **(b)** $HCOOH/HCOO^-$ and CN^-/HCN **(c)** $H_2PO_4^-/HPO_4^{2-}$ and OH^-/H_2O **10.2: 3.** weak **4.** potassium permanganate; permanganic acid, $HMnO_4$ **5.(a)** sulfuric acid, H_2SO_4 **(b)** hydrofluoric acid, HF **(c)** hydrosulfuric acid, H_2S **(d)** bromous acid, HBr_2 **7.** egg white, camembert cheese, beets, yogurt, sauerkraut **8.(a)** 7.4 **(b)** 1.4 **9.** 2.90 **10.** 1.0 mol/L; concentrations greater than this give negative pH values, but this gives no advantage over the actual concentration of H_3O^+ **10.3: 1.** acid + base = salt + water **4.** 0.304 mol/L **5.** 0.107 mol/L