14 Properties of Acids and Bases

Chapter Preview

14.1 Defining Acids and Bases

14.2 Strong and Weak Acids and Bases

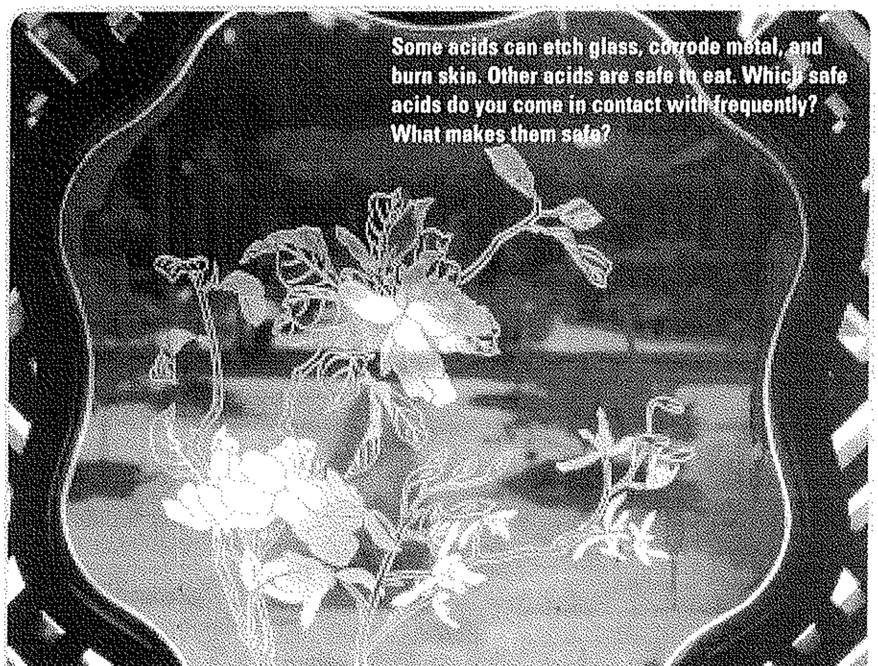
Prerequisite Concepts and Skills

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

- calculating molar concentrations (Chapter 7)
- writing net ionic equations (Chapter 8)
- solving equilibrium problems (Chapter 13)

For many people, the word "acid" evokes the image of a fuming, highly corrosive, dangerous liquid. This image is fairly accurate for concentrated hydrochloric acid, which chemists classify as a strong acid. Most acids, however, are not as corrosive as hydrochloric acid, although they may still be very hazardous. For example, hydrofluoric acid can cause deep, slow-healing tissue burns if it is handled carelessly. It is used by artists and artisans who etch glass. It reacts with the silica in glass to form a compound that dissolves, leaving contoured patterns in the glass surface. Hydrofluoric acid is highly corrosive. Even a 1% solution is considered to be hazardous. Yet chemists classify hydrofluoric acid as a weak acid.

In this chapter, you will examine the properties of acids and bases, starting with those you observe directly. Afterwards, you will learn how chemists developed theories to explain the properties of acids and bases, as well as the properties of the solutions that result when they are mixed together. You will learn how to compare the strengths of acids and bases qualitatively. As well, you will use another equilibrium constant and the pH of an aqueous solution to compare the strengths of acids and bases quantitatively.



Some acids can etch glass, corrode metal, and burn skin. Other acids are safe to eat. Which safe acids do you come in contact with frequently? What makes them safe?

Defining Acids and Bases

14.1

Acids and bases are as much a part of people's lives today as they were thousands of years ago. (See Figure 14.1.) For example, vinegar is an acidic solution that is common in many food and cleaning products. It was discovered long before people invented the skill of writing to record its use. Today, acids are also used in the manufacturing of fertilizers, explosives, medicines, plastics, motor vehicles, and computer circuit boards.

Bases also 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—one of the bases in wood ash. Sodium hydroxide is also used to make soap, detergents, drain cleaner, dyes, medicines, and many other products. Table 14.1 lists some common bases and acids you can find in and around the home. You will work with some of these in Investigation 14-A, as you compare the properties of acids and bases.

Figure 14.1 It is easy to identify some products as acids, because the word "acid" appears in the list of ingredients. Identifying bases is more difficult. What acids and bases do you have in and around your home?

Table 14.1 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

Section Preview/Outcomes

In this section, you will

- **describe** and **compare** theories of acids and bases
- **identify** conjugate acid-base pairs
- **write** chemical equations to show the amphoteric nature of water
- **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, amphoteric*

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.

SKILL FOCUS

Performing and recording

Analyzing and interpreting

Investigation 14-A

Observing Properties of Acids and Bases

In this investigation, you will conduct tests in order to develop your own set of identifying properties and definitions for acids and bases. Two of the tests will involve the use of an indicator—a chemical that changes colour in the presence of acids or bases. One of these indicators, litmus paper, is made from a compound that is extracted from lichens, a plant-like member of the fungi kingdom. Litmus paper is made by dipping paper in a solution made with this compound, litmus.

Question

How can the use of chemical tests enable you to develop a definition for acids and bases?

Safety Precautions

- Hydrochloric acid, sulfuric acid, sodium hydroxide, and ammonia are toxic and/or corrosive. Wash any spills on skin or clothing with plenty of cool water. Inform your teacher immediately.
- If you are using a conductivity tester with two separate electrodes, be extremely careful to keep the electrodes well separated while you perform your tests.
- Phenolphthalein is flammable. Keep it well away from sparks or flames.

Materials

1.0 mol/L solutions of the following:
 hydrochloric acid, $HCl_{(aq)}$, sulfuric acid, $H_2SO_{4(aq)}$, acetic acid, $HC_3COOH_{(aq)}$, sodium hydroxide, $NaOH_{(aq)}$, ammonia, $NH_{3(aq)}$
saturated limewater solution
baking soda (sodium hydrogen carbonate, $NaHCO_{3(s)}$)
phenolphthalein solution
magnesium ribbon, finely cut to 1 mm–2 mm lengths
6 test tubes
25 mL graduated cylinder
glass stirring rod
well plate or glass plate
red litmus paper
blue litmus paper
rubber stopper
wooden splint
evaporating dish or 50 mL beaker
medicine dropper or dropper pipette
scoopula
conductivity apparatus

Procedure

1. Read through the entire Procedure, and design a table to record your observations.
2. Add 10 mL of hydrochloric acid solution to a test tube and place it in the test tube rack. Test the conductivity of the solution.
3. Place a strip of red litmus paper on a glass cover plate, and use a stirring rod to put one drop of the acid on the litmus paper. Do the same with another glass plate and a strip of blue litmus paper. Clean the stirring rod, and record your observations.
4. Read this step completely before proceeding. Add limewater to a test tube until it is half full, and place it in the rack. Obtain a third test tube, and pour half of the hydrochloric acid from the first test tube into it. Still holding this test tube, use the scoopula to gently add about 0.25 mL of baking soda to the acid. Working quickly but carefully, tip the tube to pour the gas into the limewater. Take care not to pour any liquid into the limewater. Place a stopper firmly in the test tube and shake it for several seconds. Then put it in the rack and record your observations.

5. Add six to eight small pieces of magnesium ribbon to the acid in the remaining (second) test tube. Record your observations. Then carefully lower a burning wooden splint into the test tube. Record your observations.
6. Wash the test tubes so they are completely clean. Ensure that no unreacted magnesium ribbon is placed down the sink.
7. Repeat steps 2 to 6 with the sulfuric acid and acetic acid solutions.
8. Add 5 mL of sodium hydroxide solution to a clean test tube. Use the clean stirring rod to put a drop of the sodium hydroxide solution on the strip of red litmus paper. Do the same with a strip of blue litmus paper. Clean the stirring rod, and record your observations.
9. Add six to eight small pieces of magnesium ribbon to the sodium hydroxide solution, and perform the splint test as you did for the acids in step 4. Record your observations.
10. Repeat steps 8 and 9 with the ammonia solution and the limewater.
11. Add 10 mL of sodium hydroxide solution to a clean 50 mL beaker or evaporating dish. With a stirring rod, add two drops of phenolphthalein. Then stir and record your observations. Continue stirring and use a dropper pipette or medicine dropper to add hydrochloric acid solution, slowly, one drop at a time. Stop adding acid when the pink colour you are observing disappears.
12. Test the clear, colourless solution that results from step 10 with blue and red litmus paper. Also test its conductivity. Record your observations.

Analysis

1. Examine the chemical formulas for the acids and bases you tested in this investigation. Identify anything you notice that they have in common.
2. What is the effect of acids and bases on red and blue litmus paper?
3. What reaction, if any, occurs when magnesium is mixed with an acid and with a base? Explain your answer in detail.
4. Describe the reaction that occurs when baking soda is mixed with an acid, and identify the gas that is produced.
5. (a) The reaction that occurs when an acid is mixed with a base is called a neutralization reaction. Write a balanced chemical equation to describe the neutralization reaction that you observed.
   (b) What evidence from the procedure supports your balanced equation?

Conclusion

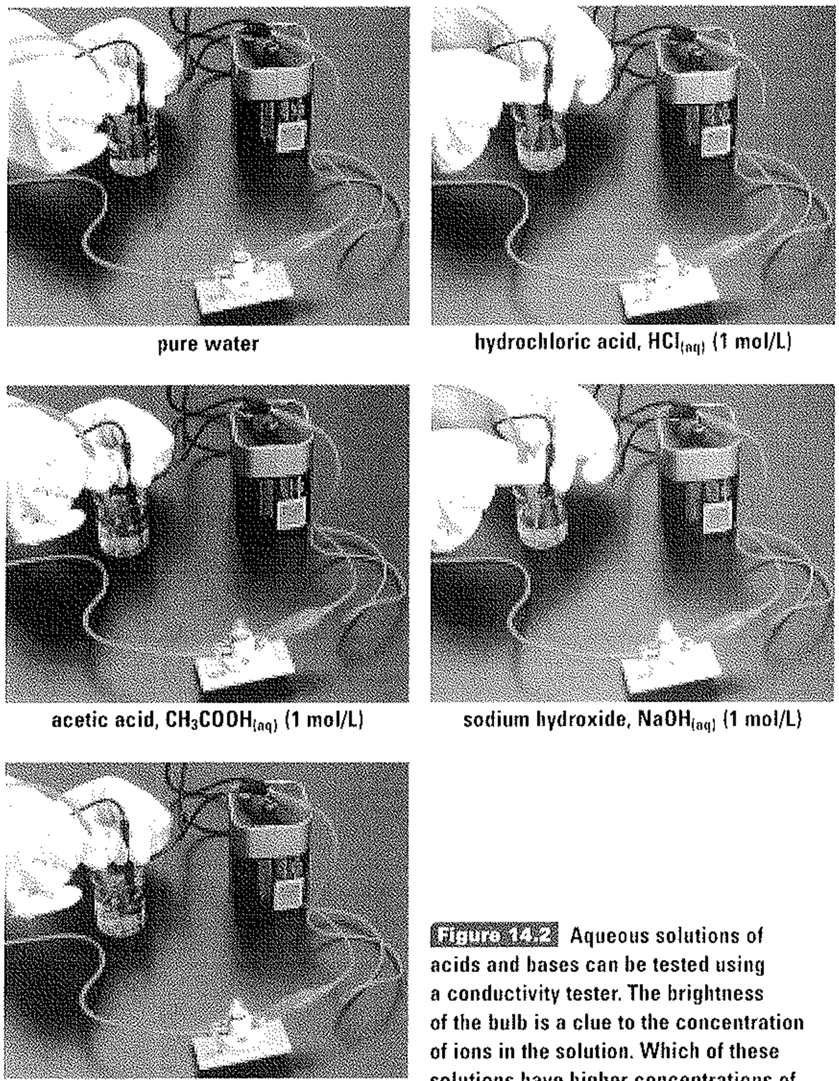
6. (a) Design a chart or concept map to compare the properties of acids and bases, based on your observations.
   (b) Collaborate with the class to write a definition for acids and bases.

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 14.2, 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 14.2 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?

Table 14.2 on the next page summarizes some other observable properties of acids and bases. These properties include their physical characteristics and their chemical behaviour. Taken together, the observable properties of acids and bases constitute their operational definitions. For example, an operational definition of an acid might be, "An acid is a sour-tasting compound that turns blue litmus paper red and reacts with bases to form a salt and water." An operational definition of a base might be, "A base is a bitter-tasting, slippery-feeling compound that turns red litmus paper blue and reacts with acids to form a salt and water." Operational definitions are useful for classifying substances and describing their macroscopic (observable) characteristics and behaviour. However, an operational definition cannot describe or explain the microscopic (unobservable) *causes* of those characteristics and behaviour. A more comprehensive definition of acids and bases requires a theory—something that chemists can use to explain established observations and properties and to make predictions in new situations. You will consider several theories of acids and bases in the remainder of this section.

The Arrhenius Theory of Acids and Bases

In Figure 14.2, 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 14.3.)

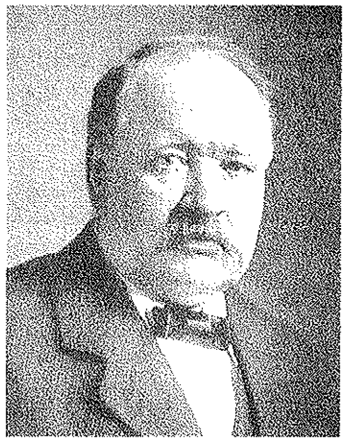


Figure 14.3 **Svanté Arrhenius (1859–1927)—contributed to chemists understanding of electrolytes and reaction rates, as well as acids and bases.**

$$HBr_{(aq)} \rightarrow H^{+}_{(aq)} + Br^{-}_{(aq)}$$
$$H_2SO_{4(aq)} \rightarrow H^{+}_{(aq)} + HSO_4{}^{-}_{(aq)}$$
$$HClO_{4(aq)} \rightarrow H^{+}_{(aq)} + ClO_4{}^{-}_{(aq)}$$

acids dissociating in water, and their resulting ions

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

bases dissociating in water, and their resulting ions

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^-.

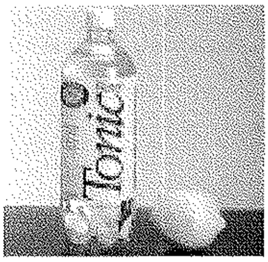
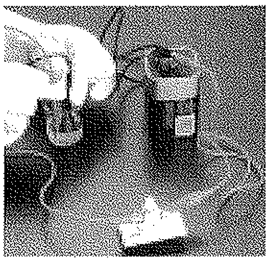
According to the Arrhenius theory, acids increase the concentration of H^+ in aqueous solutions. Thus, *an Arrhenius acid must contain hydrogen as the source of* H^+. You can see this in the preceding dissociation reactions for acids.

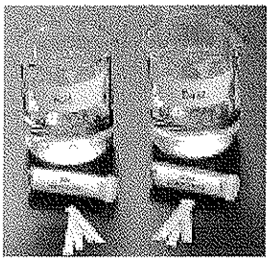
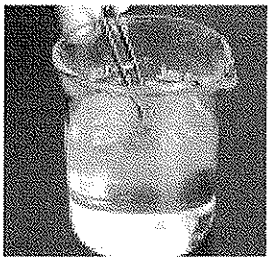
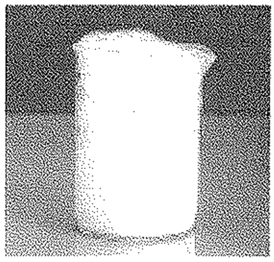
Web LINK

www.mcgrawhill.ca/links/atlchemistry

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 and click on **Web Links**. Write a brief report to answer the above questions.

Table 14.2 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
			

	Property
	Reaction with each other
ACIDS	Acids neutralize basic solutions, forming a salt and water.
BASES	Bases neutralize acidic solutions, forming a salt and water.

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

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 14.4.)

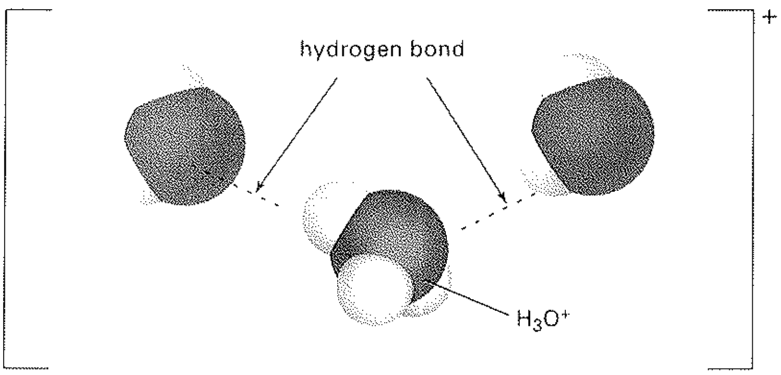


Figure 14.4 For convenience, chemists often use $H^+_{(aq)}$ as a shorthand notation for the hydronium ion, $H_3O^+_{(aq)}$. Hydronium ions do not exist independently. Instead, they form hydrogen bonds with other water molecules. Thus, a more correct formula is $H^+(H_2O)_n$, where n is usually 4 or 5.

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

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

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.

Modernizing the Arrhenius Theory

One of the strengths of the Arrhenius theory is that chemists can modernize it to account for some of its limitations. For example, here again is the dissociation equation for $HCl_{(aq)}$ according to Arrhenius theory.

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

A modernized Arrhenius theory acknowledges the role of water and the production of hydronium ions in this reaction. Thus:

$$HCl_{(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Similarly, the dissociation of sulfuric acid may be reinterpreted in light of a modernized Arrhenius theory.

$$H_2SO_{4(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + HSO_4{}^-_{(aq)}$$

You have read that another limitation of the Arrhenius theory is that it cannot account for the basic properties of aqueous ammonia, because ammonia does not contain hydroxide ions. Thus, the dissociation of ammonia in water would occur by way of a two-step process:

$$NH_{3(g)} + H_2O_{(\ell)} \rightarrow NH_4OH_{(aq)}$$

$$NH_4OH_{(aq)} \rightarrow NH_4{}^+_{(aq)} + OH^-_{(aq)}$$

The compound inserted into this process is ammonium chloride. Its ionic formula, NH_4OH, is written correctly. However, there is *no* experimental evidence to suggest that this hypothesized compound actually exists. Chemists' attempt to "rescue" the Arrhenius theory in this case fails.

There is a further limitation with the Arrhenius theory—one that chemists cannot overcome. The theory assumes that all acid-base reactions occur in a single solvent: water. However, many acid-base reactions may occur in other solvents.

Yet another limitation of the Arrhenius theory is its inability to predict successfully whether a given aqueous compound is acidic or basic. For example, the hydrogen phosphate ion, $HPO_4{}^{2-}$, appears to contain H^+ ions. You would expect that solutions containing this ion would be acidic. And yet, when tested with litmus paper, solutions that contain this ion are found to be *basic*. Similarly, salts that contain carbonate ions also have basic properties when dissolved in water. The Arrhenius theory cannot predict or explain these facts.

Despite its shortcomings, the Arrhenius theory provides chemists—both past and present—with a way of understanding the properties and behaviour of certain compounds when they react with water and with each other. An underlying strength of the Arrhenius theory is that it provides a simple, effective way to understand neutralization reactions. You can see this clearly by examining the ionic and net ionic equations for a typical neutralization reaction.

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

Ionic equation: $H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(\ell)}$

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

The net ionic equation clearly shows that hydrogen ions and hydroxide ions combine in a neutralization reaction to form water.

On the basis of its limitations, however, both the original and the modernized Arrhenius definitions for acids and bases are incomplete. Chemists needed a more comprehensive theory to account for a much broader range of observations and properties.

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 14.5.) 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 limitations of 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.

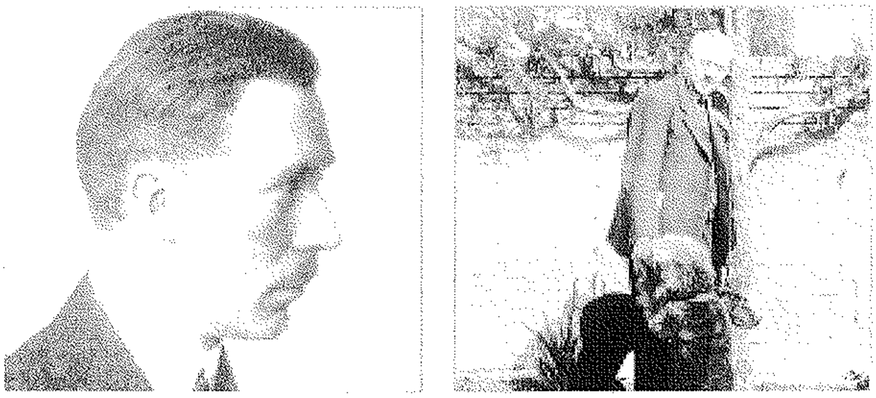


Figure 14.5 Johannes Brønsted (1879–1947), left, and Thomas Lowry (1874–1936), right. Brønsted published many more articles about ions in solution than Lowry did. Thus, some chemistry resources refer to the "Brønsted theory of acids and bases."

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.* Note that the word "proton" refers to the nucleus of a hydrogen atom—an H^+ ion that has been removed from the acid molecule. It does not refer to a proton removed from the nucleus of another atom, such as oxygen or sulfur, that may be present in the acid molecule.

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 14.6. In this reaction, hydrochloric acid is an acid

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.

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.

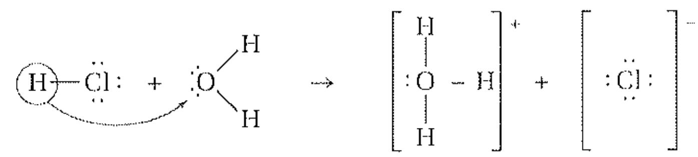


Figure 14.6 **The reaction between hydrochloric acid and water, according to the Brønsted-Lowry theory.**

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. See how this works with another reaction involving acetic acid.

The dissociation of acetic acid in water is represented in Figure 14.7. Acetic acid is a Brønsted-Lowry acid, because it provides a proton (H^+) to the water. In receiving the proton, the water molecule is the base in this reaction, and becomes a hydronium ion. Notice, however, that this reaction is an equilibrium reaction; it proceeds in both directions. When acetic acid reacts with water, only a few ions dissociate. (You will learn about the significance of this fact shortly.) The position of equilibrium lies to the left, and the reverse reaction is favoured. In the reverse reaction, the hydronium ion gives up a proton to the acetate ion. Thus, in the reverse reaction, the hydronium ion is a Brønsted-Lowry acid and the acetate ion is a Brønsted-Lowry base. The acid on the left (CH_3COOH) and the base on the right (CH_3COO^-) differ by one proton, so they are a conjugate acid-base pair. Similarly, H_2O and H_3O^+ are a conjugate acid-base pair, because they, too, differ by one proton. This relationship is shown in Figure 14.8.

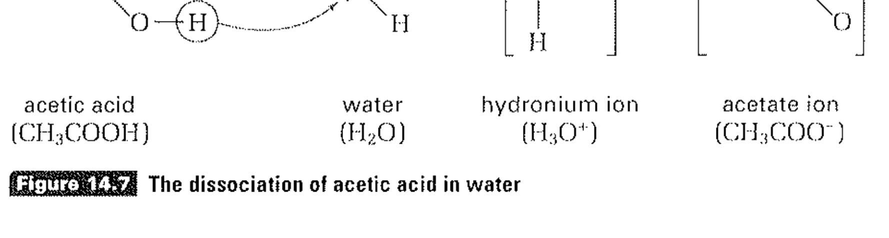


Figure 14.7 **The dissociation of acetic acid in water**

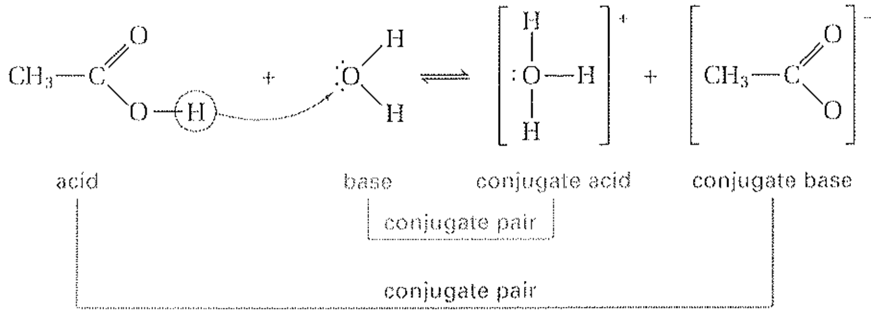


Figure 14.8 Conjugate acid-base pairs in the dissociation of acetic acid in water

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^+.

Continued

Continued...

conjugate acid-base pair

acid — conjugate base

$$HBr_{(g)} + H_2O_{(\ell)} \longrightarrow H_3O^+_{(aq)} + Br^-_{(aq)}$$

base — conjugate acid

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.

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

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.

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

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

base — conjugate acid

conjugate acid-base pair

Check Your Solution

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

PROBLEM TIP

In the previous Sample Problem, water acted as a base. In this Sample Problem, water acts as an acid. Water can act as a proton-provider (an acid) in some reactions and as a proton-receiver (a base) in others. You will learn more about this dual-behaviour of certain substances following the Practice Problems.

Continued

8. Identify the conjugate acid-base pair in each reaction.

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

(b) $CH_3COOH_{(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3COO^-{}_{(aq)} + H_3O^+{}_{(aq)}$

9. For each reaction in question 8, identify the amphoteric chemical species, and identify its role as either an acid or a base.

Section Wrap-up

The theories that you have considered in this section attempt to explain the chemical nature of acids and bases. Table 14.3 summarizes the key points of these theories.

In the Arrhenius theory, the modern 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 the next section, 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 14.3 Comparing the Arrhenius Theory, Modern Arrhenius Theory, and the Brønsted-Lowry Theory.

Theory	Arrhenius	Modern Arrhenius	Brønsted-Lowry
Acid	any substance that dissociates to form H^+ in aqueous solution	any substance that dissociates to form H_3O^+ 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 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	$HCl_{(aq)} \rightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)}$	$HCl_{(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+{}_{(aq)} + Cl^-{}_{(aq)}$	$NH_{3(g)} + H_2O_{(\ell)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$

Section Review

1 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) 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 Sodium bicarbonate, $NaHCO_{3(aq)}$ turns red litmus paper blue and undergoes a neutralization reaction with acids.
(a) Use Arrhenius theory to write a chemical equation showing the dissociation of $NaHCO_{3(aq)}$ in water.
(b) Use a modernized Arrhenius theory to show this same dissociation.
(c) Does the chemical behaviour of $NaHCO_{3(aq)}$ support or refute either forms of the Arrhenius theory? Give reasons to justify your answer.

4 (a) 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?)

5 (a) 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.

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

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

8 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

9 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

10 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.

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

14.2 Strong and Weak Acids and Bases

Section Preview/Outcomes

In this section, you will

- **explain**, in terms of the degree to which they dissociate, the difference between strong and weak acids and bases
- **compare**, qualitatively, the relative strengths of acids and bases
- **predict** whether reactants or products are favoured in an acid-base reaction
- **define** Kw, and relate its value to $[H_3O^+]$ and $[OH^-]$
- **define** and **calculate** values associated with pH, pOH, $[H_3O^+]$, and $[OH^-]$
- **communicate** your understanding of the following terms: *strong acid, weak acid, strong base, weak base, ion product constant for water (K_w), pH, pOH*

You have learned that substances such as water and the hydrogen carbonate ion are amphoteric—that is, they can act either as a Brønsted-Lowry acid or a Brønsted-Lowry base in a particular reaction. For example, the hydrogen carbonate ion behaves as an acid in the first of the two chemical equations below, but as a base in the second.

$$HCO_3^-{}_{(aq)} + HPO_4^{2-}{}_{(aq)} \rightleftharpoons CO_3^{2-}{}_{(aq)} + H_2PO_4^-{}_{(aq)}$$

($HCO_3^-{}_{(aq)}$ behaving as an acid)

$$HCO_3^-{}_{(aq)} + HF_{(aq)} \rightleftharpoons H_2CO_{3(aq)} + F^-{}_{(aq)}$$

($HCO_3^-{}_{(aq)}$ behaving as a base)

How could you predict which role the ion plays in each reaction? The answer to this question involves the strength of the ion in relation to the other substances present in each reaction. Keep this idea in mind as you learn about strong and weak acids and bases. You will return to it later in this section.

Strong and Weak Acids

In terms of acid-base reactions, strength refers to the extent to which a substance dissociates in its solvent. An acid that dissociates completely is termed a **strong acid**. As you can see in Figure 14.9, hydrochloric acid is a strong acid. *All* the molecules of hydrochloric acid in an aqueous solution dissociate into H^+ and Cl^- ions. Table 14.4 lists the strong acids. Note that *the concentration of hydronium ions in a dilute solution of a strong acid is equal to the concentration of the acid.* Therefore, a 1.0 mol/L solution of hydrochloric acid contains 1.0 mol/L of hydronium ions and 1.0 mol/L of chloride ions.

Table 14.4
Common Strong Acids

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

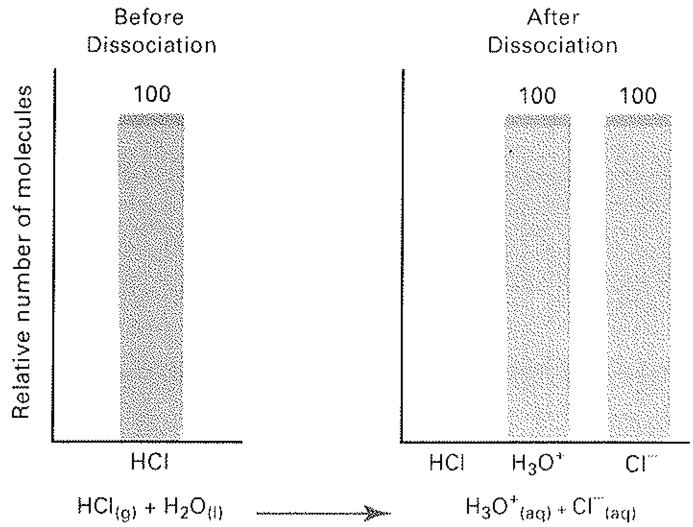
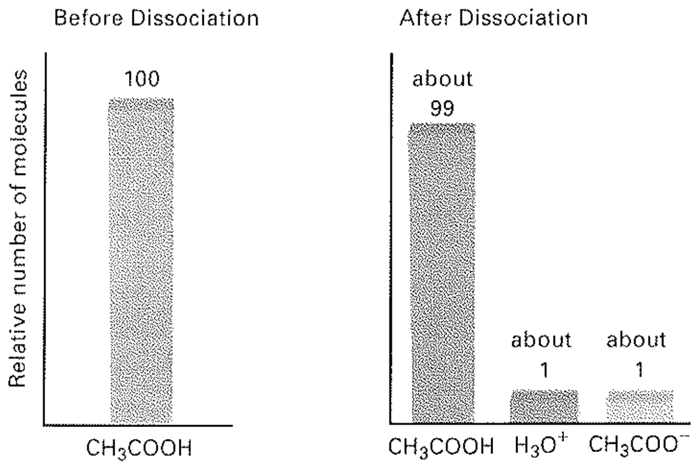


Figure 14.9 When hydrogen chloride molecules enter an aqueous solution, 100% of the hydrogen chloride molecules dissociate. In other words, the percent dissociation of HCl in water is 100%. As a result, the solution contains the same percent of H^+ ions (in the form of H_3O^+) and Cl^- ions: 100%.

A **weak acid** is an acid that dissociates only slightly in a water solution. Thus, only a small percentage of the acid molecules dissociate. Most of the acid molecules remain intact. The majority of acids are weak acids. For example, acetic acid is a weak acid. The percent dissociation of acetic acid molecules is only about 1% in a 0.1 mol/L solution. (The number of acid molecules that dissociate depends on the concentration and temperature of the solution.) Note that the concentration of hydronium ions in a solution of a weak acid is always less than the concentration of the dissolved acid. (See Figure 14.10.)



$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

Figure 14.10 **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.**

CONCEPT CHECK

Notice that the chemical equation for the dissociation of hydrochloric acid—a strong acid—proceeds in a single direction, to the right. The chemical equation for the dissociation of acetic acid—a weak acid—shows an equilibrium reaction. Explain why this makes sense. Look for other instances of single- and double-arrows up to this point in the chapter, and explain what you see.

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. Sulfuric acid is a strong acid, but *only* for its first dissociation.

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

The resulting aqueous hydrogen sulfate ion, HSO^-_4, is a weak acid. It dissociates to form the sulfate ion in the following equilibrium 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 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.

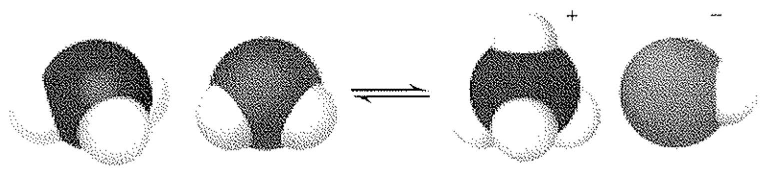
Recall that the concentration of hydronium ions in a dilute solution of a strong acid is equal to the concentration of the acid. Similarly, the concentration of hydroxide ions in a dilute solution of a strong base is equal to the concentration of the base. For example, a 1.0 mol/L solution of sodium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

Table 14.5
Common Strong Bases

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

Table 14.5 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 equilibrium is represented in Figure 14.11.



$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Figure 14.11 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.

Comparing the Relative Strengths of Acids and Bases

Earlier, you saw that the hydrogen carbonate ion may behave as an acid or a base in the following reactions.

$HCO_3^-{}_{(aq)} + HPO_4^{2-}{}_{(aq)} \rightleftharpoons CO_3^{2-}{}_{(aq)} + H_2PO_4^-{}_{(aq)}$

($HCO_3^-{}_{(aq)}$ behaving as an acid)

$HCO_3^-{}_{(aq)} + HF_{(aq)} \rightleftharpoons H_2CO_{3(aq)} + F^-{}_{(aq)}$

($HCO_3^-{}_{(aq)}$ behaving as a base)

You are now in a position to reconsider the question: How could you predict the role that each ion plays?

Over the centuries, chemists have performed countless experiments involving acids and bases. The data from these experiments have enabled chemists to rank acids and bases according to their strengths in relation to

one another, as shown in Figure 14.12. Using these data, you can easily compare the relative strengths of acids and bases. For example, the chart shows that HCO_3^- is above HPO_4^{2-}. Since the acids are listed in order of increasing strength, you know that HCO_3^- is a stronger acid than HPO_4^{2-}. Therefore, you can (if the remainder of the balanced equation were not given to you) correctly predict the products of the reaction. Why? Because you know that the proton from the acid must go to its conjugate base, $H_2PO_4^{2-}$.

ACID STRENGTH

Acid	Base
HCl	Cl^-
H_2SO_4	HSO_4^-
HNO_3	NO_3^-
H_3O^+	H_2O
HSO_4^-	SO_4^{2-}
H_2SO_3	HSO_3^-
H_3PO_4	$H_2PO_4^-$
HF	F^-
CH_3COOH	CH_3COO^-
H_2CO_3	HCO_3^-
H_2S	HS^-
HSO_3^-	SO_3^{2-}
$H_2PO_4^-$	HPO_4^{2-}
NH_4^+	NH_3
HCO_3^-	CO_3^{2-}
HPO_4^{2-}	PO_4^{3-}
H_2O	OH^-
HS^-	S^{2-}
OH^-	O^{2-}

BASE STRENGTH

Figure 14.12 **Relative strengths of selected acids and bases. Notice that the strongest acids appear at the top of the list on the left, and the strongest bases appear at the bottom of the list on the right. Notice also that a stronger acid has a weaker conjugate base, and a stronger base has a weaker conjugate acid.**

Figure 14.12 is useful in another way. You can use it to predict the direction in which an acid-base reaction will proceed. Or, to put it another way, you can use the chart to help you determine whether products or reactants are favoured in an acid-base reaction. The direction of an acid-base reaction usually proceeds from a stronger acid and a stronger base to a weaker acid and a weaker base. If the reaction proceeds to the right (that is, if the stronger acid and stronger base are on the left side of the equation), products are favoured. If the reaction goes to the left (if the stronger acid and stronger base are on the right side of the equation), reactants are favoured. Use the next Sample Problem and Practice Problems to help you understand these ideas better.

Sample Problem

Predicting the Direction of an Acid-Base Reaction

Problem

Predict the direction in which the following reaction will proceed. State whether reactants or products are favoured at equilibrium, and briefly defend your reasoning.

$$SO_4{}^{2-}{}_{(aq)} + CH_3COOH_{(aq)} \rightleftharpoons HSO_4{}^-{}_{(aq)} + CH_2COO^-{}_{(aq)}$$

Solution

Use Figure 14.12 to assess the relative strengths of the acids and bases. You can see that $HSO_4{}^-$ is above CH_3COOH, so $HSO_4{}^-$ is a stronger acid than CH_3COOH. Similarly, CH_2COO^- is lower than $SO_4{}^-$, so CH_2COO^- is a stronger base. The stronger acid and stronger base form the weaker acid and the weaker base, so the direction proceeds in that direction—that is, to the left.

$SO_4{}^{2-}{}_{(aq)}$	+	$CH_3COOH_{(aq)}$	$\rightleftharpoons$	$HSO_4{}^-{}_{(aq)}$	+	$CH_2COO^-{}_{(aq)}$
weaker acid		weaker base		stronger acid		stronger base

the reactants are favoured in this reaction.

Practice Problems

10. Predict the direction for the following equations. State whether reactants or products are favoured, and give reasons to support your decision.

(a) $NH_4{}^+{}_{(aq)} + H_2PO_4{}^-{}_{(aq)} \rightleftharpoons NH_{3(aq)} + H_3PO_{4(aq)}$

(b) $H_2O_{(\ell)} + HS^-{}_{(aq)} \rightleftharpoons OH^-{}_{(aq)} + H_2S_{(aq)}$

(c) $HF_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightleftharpoons F^-{}_{(aq)} + HSO_4{}^-{}_{(aq)}$

11. In which direction will the following reactions proceed? Explain why in each case.

(a) $HPO_4{}^{2-}{}_{(aq)} + NH_4{}^+{}_{(aq)} \rightleftharpoons H_2PO_4{}^- + NH_{3(aq)}$

(b) $H_2SO_{4(aq)} + H_2O_{(\ell)} \rightleftharpoons HSO_4{}^-{}_{(aq)} + OH^-{}_{(aq)}$

(c) $H_2S + NH_{3(aq)} \rightleftharpoons HS^-{}_{(aq)} + NH_4{}^+{}_{(aq)}$

pH: One Method for Describing Acids and Bases Quantitatively

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. 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.

What exactly is pH? How is it measured? To answer these questions, you must look again at the chemical equilibrium of water.

The Ion Product Constant for 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.

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

At 25°C, only about two water molecules in one billion dissociate. This is why pure water is such a poor conductor of electricity. Chemists have determined that the concentration of hydronium ions in pure water at 25°C is 1.0×10^{-7} mol/L. The dissociation of water also produces the same very small number of hydroxide ions: 1.0×10^{-7} mol/L. These concentrations must be the same, because the dissociation of water produces equal number of hydronium and hydroxide ions. That is, the ions are produced in a 1:1 ratio.

Because the dissociation of water is an equilibrium, you can write an equilibrium constant expression for it. However, recall that the concentration of a liquid is itself a constant, so it may be factored out of the expression for K. The resulting constant is called the **ion product constant for water**, and is given the symbol $\boldsymbol{K_w}$.

$$K_w = [H_3O^+][OH^-]$$

The equilibrium value of the concentration ion product, $[H_3O^+][OH^-]$, at 25°C is as follows:

$$\begin{aligned}K_w &= 1.0 \times 10^{-7}\text{ mol/L} \times 1.0 \times 10^{-7}\text{ mol/L}\\ &= 1.0 \times 10^{-14}\end{aligned}$$

The concentration of H_3O^+ in the solution of a strong acid is equal to the concentration of the dissolved acid, unless the solution is very dilute. Consider $[H_3O^+]$ in a solution of 0.1 mol/L hydrochloric acid. All the molecules of HCl dissociate in water, forming a hydronium ion concentration that equals 0.1 mol/L. The increased $[H_3O^+]$ pushes the dissociation reaction between water molecules to the left, in accordance with Le Châtelier's principle. Therefore, the concentration of hydronium ions that results from the dissociation of water is even less than 1×10^{-7} mol/L. This $[H_3O^+]$ is negligible compared with the 0.1 mol/L concentration of the hydrochloric acid. Unless the solution is very dilute (about 1×10^{-7} mol/L), the dissociation of water molecules can be ignored when determining $[H_3O^+]$ of a strong acid.

Similarly, you can determine the concentration of hydroxide ions from the concentration of the dissolved base. If the solution is a strong base, you can ignore the dissociation of water molecules when determining $[OH^-]$, unless the solution is very dilute. When either $[H_3O^+]$ or $[OH^-]$ is known, you can use the ion product constant for water, K_w, to determine the concentration of the other ion. Although the value of K_w for water is 1.0×10^{-14} at 25°C only, you can use this value unless another value is given for a different temperature.

$[H_3O^+]$ and $[OH^-]$ in Aqueous Solutions at 25°C

In an acidic solution, $[H_3O^+]$ is greater than 1.0×10^{-7} mol/L and $[OH^-]$ is less than 1.0×10^{-7} mol/L.

In a neutral solution, both $[H_3O^+]$ and $[OH^-]$ are equal to 1.0×10^{-7} mol/L.

In a basic solution, $[H_3O^+]$ is less than 1.0×10^{-7} mol/L and $[OH^-]$ is greater than 1.0×10^{-7} mol/L.

CONCEPT CHECK

At 24°C or at any temperature other than 25°C, the concentration of hydronium (or hydroxide) ions in water is *not* 1.0×10^{-7} mol/L. Explain in detail why this is the case.

Sample Problem

Determining $[H_3O^+]$ and $[OH^-]$

Problem

Find $[H_3O^+]$ and $[OH^-]$ in each solution.

(a) 2.5 mol/L nitric acid

(b) 0.16 mol/L barium hydroxide

Solution

You know that nitric acid is a strong acid and barium hydroxide is a strong base. Since both dissociate completely in aqueous solutions, you can use their molar concentrations to determine $[H_3O^+]$ or $[OH^-]$. You can find the concentration of the other ion using K_w:

$$K_w = 1.0 \times 10^{-14}$$
$$= [H_3O^+][OH^-]$$

(a) $[HNO_3] = 2.5$ mol/L, so $[H_3O^+] = 2.5$ mol/L

$$[OH^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{2.5}$$
$$= 4.0 \times 10^{-15} \text{ mol/L}$$

(b) $Ba(OH)_2 \xrightarrow{H_2O} Ba^{2+}{}_{(aq)} + 2OH^-{}_{(aq)}$

Each mole of $Ba(OH)_2$ in solution forms two moles of OH^- ions.

$\therefore [OH^-] = 2 \times 0.16 = 0.32$ mol/L

$$[H_3O^+] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{0.32}$$
$$= 3.1 \times 10^{-14} \text{ mol/L}$$

Check Your Solution

For a solution of a strong acid, as in part (a), $[H_3O^+]$ should be greater than 1.0×10^{-14} and $[OH^-]$ should be less than 1.0×10^{-14}. For a solution of strong base, $[OH^-]$ should be greater than, and $[H_3O^+]$ should be less than, 1.0×10^{-14}.

PROBLEM TIP

In the solution for part (b), be sure you understand why $[OH^-] = 0.32$ mol/L, and 0.16 mol/L. If you are unsure, review Chapter 8, page 300. As well, try this problem: What is the molar concentration of $[H_3O^+]$ and $[OH^-]$ in each of the following aqueous solutions of strong acids and strong bases:

- 1.0×10^{-6} mol/L KOH
- 0.100 mol/L $Ba(OH)_2$
- 1.50 mol/L HBr
- 1.0×10^{-2} mol/L HNO_3

Practice Problems

12. Determine $[H_3O^+]$ and $[OH^-]$ in each solution.

(a) 0.45 mol/L hydrochloric acid

(b) 1.1 mol/L sodium hydroxide

13. Determine $[H_3O^+]$ and $[OH^-]$ in each solution.

(a) 0.95 mol/L hydroiodic acid

(b) 0.012 mol/L calcium hydroxide

14. $[OH^-]$ is 5.6×10^{-14} mol/L in a solution of hydrochloric acid. What is the molar concentration of the $HCl_{(aq)}$?

15. $[H_3O^+]$ is 1.7×10^{-14} in a solution of calcium hydroxide. What is the molar concentration of the $Ca(OH)_{2(aq)}$?

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 14.13, 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.

"the negative logarithm of... or −log	pH	"...the concentration of H^+ ions (in mol/L) or $[H^+]$

Figure 14.13 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}.

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

$[H_3O^+]$ 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 14.6 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.

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.

Table 14.6 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 $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$	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.

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.

The relationship among pH, $[H_3O^+]$, and the strength of acids and bases is summarized in Table 14.7. Use the following Sample Problem and Practice Problems to assess your understanding of this relationship.

Table 14.7 The Relation of pH, $[H_3O^+]$, and Acid-Base Strength

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

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

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

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.

PROBLEM TIP

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

Practice Problems

16. 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) $[H_3O^+] = 8.27 \times 10^{-12}$

17. $[H_3O^+]$ 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.

18. 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.

19. (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.

Web LINK

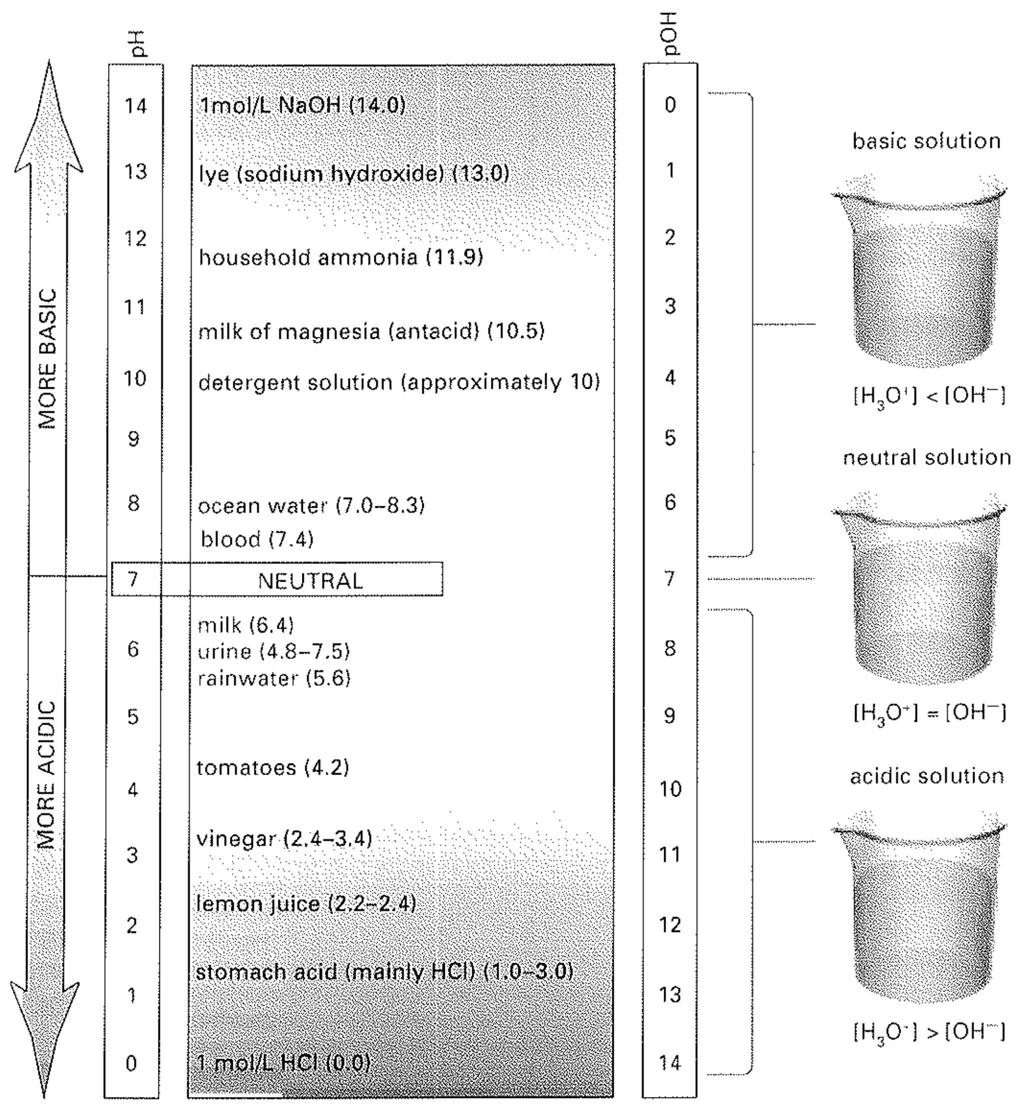
www.mcgrawhill.ca/links/atlchemistry

Many people, for both personal and professional reasons, rely on pH meters to provide quick, reliable pH measurements. Use the Internet to find out how a pH meter works, and what jobs or tasks it is used for. To start your research, go to the web site above and click on **Web Links**. Prepare a brief report, a web page, or a brochure to present your findings.

Just as pH refers to the exponential power of the hydronium ion concentration in a solution, **pOH** refers to the power of hydroxide ion concentration. You can calculate the **pOH** of a solution from the $[OH^-]$. Notice the relationship between pH and pOH shown below, and in the Concept Organizer.

$$pOH = -\log[OH^-]$$
$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ C$$
$$\therefore pH + pOH = 14$$

Concept Organizer pH, pOH, $[H_3O^+]$, $[OH^-]$, and Acid-Base Strength



Math LINK

Prove the relationship pH + pOH = 14 as follows. Record the ion product equation and its value at 25°C. Take the logarithm of both sides. Then reverse the sign of each term. What is your result?

Sample Problem
Calculating pH and pOH

Problem

A liquid shampoo has a hydroxide ion concentration of 6.8×10^{-5} mol/L at 25°C.

(a) Is the shampoo acidic, basic, or neutral?

(b) Calculate the hydronium ion concentration.

(c) What is the pH and the pOH of the shampoo?

Solution

(a) Compare $[OH^-]$ in the shampoo with $[OH^-]$ in neutral water at 25°C.
$[OH^-] = 6.8 \times 10^{-5}$ mol/L, which is greater than 1×10^{-7} mol/L.
Therefore, the shampoo is basic.

(b) Use the equation $[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]}$ to find the hydronium ion concentration.

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-5}}$$
$$= 1.5 \times 10^{-10} \text{ mol/L}$$

(c) Substitute known values into the equations $pH = -\log[H_3O^+]$ and $pOH = -\log[OH^-]$.

$$pH = -\log(1.5 \times 10^{-10})$$
$$= 9.83$$
$$pOH = -\log(6.8 \times 10^{-5})$$
$$= 4.17$$

Check Your Solution

$pH + pOH = 14$

PROBLEM TIP

When you work with logarithms, *the number of significant digits in a number must equal the number of digits after the decimal in the number's logarithm.* Here **1.5** × 10^{-10} has two significant digits. Therefore, the calculated pH, 9.**83**, must have two significant digits after the decimal.

Another Way to Find $[H_3O^+]$ and $[OH^-]$

You can calculate $[H_3O^+]$ or $[OH^-]$ by finding the *antilog* of the pH or pOH.

$$[H_3O^+] = 10^{-pH}$$
$$[OH^-] = 10^{-pOH}$$

If you are using a calculator, you can use it to find the antilog of a number in one of two ways. If the logarithm is entered in the calculator, you can press the two keys INV and LOG in sequence. (Some calculators may have a 10^x button instead.) Alternatively, since $[H_3O^+] = 10^{-pH}$ and $[OH^-] = 10^{-pOH}$, you can enter 10, press the y^x button, enter the negative value of pH (or pOH), and then press =.

Sample Problem

Finding pOH, $[H_3O^+]$, and $[OH^-]$

Problem

If the pH of urine is outside the normal range of values, this can indicate medical problems. Suppose that the pH of a urine sample was measured to be 5.53 at 25°C. Calculate pOH, $[H_3O^+]$, and $[OH^-]$ for the sample.

Solution

You use the known value, pH = 5.53, to calculate the required values.

$$pOH = 14.00 - 5.53$$
$$= 8.47$$

$$[H_3O^+] = 10^{-5.53}$$
$$= 3.0 \times 10^{-6} \text{ mol/L}$$

$$[OH^-] = 10^{-8.47}$$
$$= 3.4 \times 10^{-9} \text{ mol/L}$$

Continued...

Continued...

Check Your Solution

In this problem, the ion product constant is a useful check:

$$[H_3O^+][OH^-] = (3.0 \times 10^{-6}) \times (3.4 \times 10^{-9})$$
$$= 1.0 \times 10^{-14}$$

This value equals the expected value for K_w at 25°C.

Practice Problems

20. $[H_3O^+]$ of a sample of milk is found to be 3.98×10^{-7} mol/L. Is the milk acidic, neutral, or basic? Calculate the pH and $[OH^-]$ of the sample.

21. A sample of household ammonia has a pH of 11.9. What is the pOH and $[OH^-]$ of the sample?

22. Phenol, C_6H_5OH, is used as a disinfectant. An aqueous solution of phenol was found to have a pH of 4.72. Is phenol acidic, neutral, or basic? Calculate $[H_3O^+]$, $[OH^-]$, and pOH of the solution.

23. At normal body temperature, 37°C, the value of K_w for water is 2.5×10^{-14}. Calculate $[H_3O^+]$ and $[OH^-]$ at this temperature. Is pure water at 37°C acidic, neutral, or basic?

24. A sample of baking soda was dissolved in water and the pOH of the solution was found to be 5.81 at 25°C. Is the solution acidic, basic, or neutral? Calculate the pH, $[H_3O^+]$, and $[OH^-]$ of the solution.

25. A chemist dissolved some Aspirin™ in water. The chemist then measured the pH of the solution and found it to be 2.73 at 25°C. What are $[H_3O^+]$ and $[OH^-]$ of the solution?

Dilution Calculations involving Acids and Bases

When chemists go to use an acid in the lab, they commonly use a stock solution of known concentration and dilute it to the concentration they need. A chemist may want to dilute the stock solution either to a specific $[H_3O^+]$, a specific pH, or a specific $[OH^-]$ before use. Thus, calculations involving dilutions of acids and bases are very common in a practical lab setting. You have already studied the basis of these calculations in Unit 3. Where diluting acids and bases is concerned, the main idea is that *the number of moles of acid (or base) remains constant, and thus the number of moles* (n = CV) *before dilution equals the number of moles after dilution.*

This means that acid (or base) dilution problems may be summarized as shown in Figure 14.14. Note that the given or wanted parameters may be any of pH, pOH, $[H^+]$ or $[OH^-]$. This is possible because any given pH or pOH is directly related to the concentration of its respective hydronium or hydroxide ion.

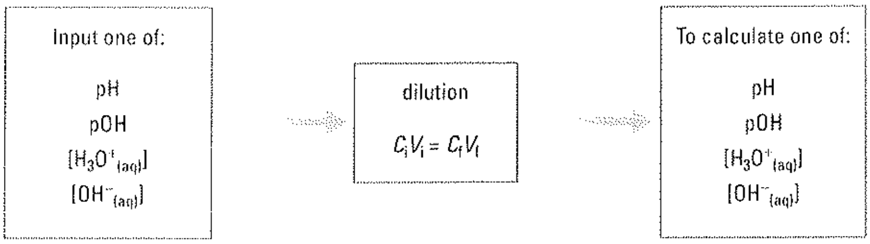


Figure 14.14 **A summary for making dilution calculations involving strengths and concentrations of acids and bases**

Note also that if one of the volumes (V_i or V_f) and all other concentration-related terms are given, the missing volume may be the variable that you must calculate.

After reading these paragraphs, you may feel that the number of possible problems that can be developed are daunting, yet if you remember that it is the number of moles of solute that is constant and that pH/pOH are simply "disguised" versions of $[H^+]/[OH^-]$, then any version of the above problem can be tackled with confidence. Use the Sample Problem and Practice Problems below to develop your calculation skills. You will use these skills in a hands-on setting in Investigation 14-B, which follows.

Sample Problem

Calculating Unknown Quantities Involving Dilution of Acids or Bases

Problem

In preparing a solution, a chemist takes a stock hydrochloric acid solution of pH = 1.50 from the chemical storeroom. If the chemist uses 125 mL of the stock solution and dilutes it to 495 mL, what is the final hydronium ion concentration?

What Is Required?

You need to find the final $[H_3O^+]$ of the solution after it has been diluted (that is, $[H_3O^+]_f$), to determine the final pH (that is pH_f).

What Is Given?

Since this is a dilution problem, you may separate the given quantities as either initial or final quantities in the dilution process.

Initial: $pH_i = 1.50$

$V_i = 125$ mL

Final: $V_f = 495$ mL

Plan Your Strategy

You must first recognize the initial pH as a "disguised form" of the initial hydronium ion concentration, and then calculate the $[H_3O^+]_i$ directly from this given initial pH. Then relate the initial given data to the final given data by remembering that $n = CV$, and that during a dilution, the moles of solute remain constant ($n_i = n_f$).

Continued...

Continued...

PROBLEM TIP

You may prefer to substitute the values directly into the dilution formula rather than rearrange the formula. This procedure will also give the correct result.

PROBLEM TIP

The conversion from mL to L here serves as a reminder that the units for both V_i and V_f are consistent and, thus, cancel. Converting units in litres avoids possible problems with cancellation.

Act on Your Strategy

$[H_3O^+]_i = 10^{-pH_i} = 10^{-1.50} = 0.032$ mol/L H_3O^+

Since $n_i = n_f$,

therefore,

$$[H_3O^+]_iV_i = [H_3O^+]_fV_f$$

Rearranging for $[H_3O^+]_f$:

$$[H_3O^+] = \frac{[H_3O^+]_i}{V_f}$$
$$= \frac{(0.32\ \text{mol/L})(0.125\ \cancel{\text{L}})}{(0.495\ \cancel{\text{L}})}$$
$$= 0.0081\ \text{mol/L}$$

Finally, convert $[H_3O^+]_f$ to the final pH:

$$pH_f = -\log[H_3O^+]_f$$
$$= -\log(0.00799)$$
$$= 2.109$$

Check Your Solution

For dilution problems, you can check the initial and final concentrations by looking at the ratio of the final volume to the initial volume. From the given volumes, the final volume is larger by a factor of almost four (495 mL/125 mL). The final solution also has a concentration that is smaller (that is, diluted) by a factor of almost four (0.0316 mol/L/0.00799 mol/L). Remember also that the pH is a logarithmic scale, so there will not be a direct relationship between the factor of the dilution and the pH; you can only check your final answer against concentration and volume values.

Practice Problems

26. Calculate the pH of a $HNO_{3(aq)}$ solution which is formed by diluting 45 mL of 0.0115 mol/L $HNO_{3(aq)}$ to a final volume of 2.00 L.

27. **(a)** A solution of lithium hydroxide is diluted from 3.25×10^{-3} mol/L to 3.25×10^{-6} mol/L. If the initial volume was 36.0 mL, calculate the final volume of the solution.

(b) Calculate the pOH and pH of the solution.

28. **(a)** Calculate the volume of concentrated hydrochloric acid (12.4 mol/L) required to prepare 950.0 mL of a solution that has a pH of 1.50.

(b) What is the pOH and $[OH^-]$ of the solution?

29. **(a)** Calculate the pOH of a solution that forms when 150.0 mL of 0.0000223 mol/L $Ca(OH)_{2(aq)}$ is diluted to 15.0 L.

(b) Calculate the pH of the final solution.

(c) Calculate the hydronium ion concentration of the final solution.

(d) Is the solution acidic or basic? Briefly explain your choice.

Investigation 14-B

SKILL FOCUS
Performing and recording
Analyzing and interpreting

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

		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}			

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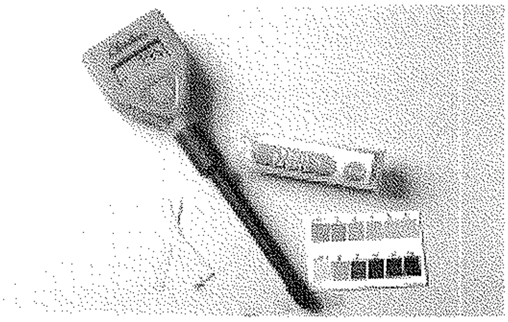

3. Measure 90 mL of distilled water in a 100 mL graduated cylinder. Add 10 mL of the acid from step 2. The resulting 100 mL of solution is one-tenth as concentrated as the acid from step 2. Pour the dilute solution into a clean, dry 250 mL beaker. Use universal pH paper and a pH meter to measure the 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 | Science | 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 their relation to pH and pOH. Earlier in the section, pH was introduced to you as "one method" for quantitatively describing acids and bases. There is another method, which involves equilibrium constants. In the second chapter of this unit, you will learn about these equilibrium constants, and how they apply to acid-base reactions involving weak acids and weak bases. As well, you will look more closely at what happens, both macroscopically and chemically, during neutralization reactions.

Section Review

1. Distinguish, in terms of degree of dissociation, between a strong acid and a weak acid, and a strong base and a weak base.

2. Give one example of the following:
   (a) a weak acid
   (b) a strong acid
   (c) a strong base
   (d) a weak base

3. Explain the meaning of pH, both in terms of hydrogen ions and hydronium ions.

4. 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.

5. 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

6. 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

7. 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?

8. Complete the following table by calculating the missing values and indicating whether each solution is acidic or basic.

$[H_3O^+]$ (mol/L)	pH	$[OH^-]$ (mol/L)	pOH	Acidic or basic?
3.7×10^{-5}	(a)	(b)	(c)	(d)
(e)	10.41	(f)	(g)	(h)
(i)	(j)	7.0×10^{-2}	(k)	(l)
(m)	(n)	(o)	8.9	(p)

14 Review

Reflecting on Chapter 14

- Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:
- Compare the properties and theories of acids and bases.
- Identify conjugate acid-base pairs for selected acid-base reactions, and compare their strengths.
- Distinguish strong and weak acids and bases on the basis of their dissociation in water.
- Outline the relationship among $[H_3O^+]$, pH, $[OH^-]$, and pOH.
- Dilute an acid and describe the effect on its pH.

Reviewing Key Terms

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

Arrhenius theory of acids and bases
hydronium ion
Brønsted-Lowry theory of acids and bases
conjugate acid-base pair
conjugate base
conjugate acid
amphoteric
strong acid
weak acid
strong base
weak base
ion product constant for water (K_w)
pH
pOH

Knowledge/Understanding

1. Use the Arrhenius theory, the modernized Arrhenius theory, and the Brønsted-Lowry theory to describe the following concepts. If one or more of the theories do 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) an aqueous solution of ammonia
   (g) strong and weak acids and bases
   (h) 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. Classify each compound as a strong acid, strong base, weak acid, or weak base.
   (a) phosphoric acid, H_3PO_4 (used in cola beverages and rust-proofing products)
   (b) chromic acid, H_2CrO_4 (used in the production of wood preservatives)
   (c) barium hydroxide, $Ba(OH)_2$, a white, toxic base (can be used to de-acidify paper)
   (d) CH_3NH_2, commonly called methylamine (is responsible for the characteristic smell of fish that are no longer fresh)
6. Write a chemical formula for each acid or base.
   (a) the conjugate base of OH^-
   (b) the conjugate acid of ammonia, NH_3
   (c) the conjugate acid of HCO_3^-
   (d) the conjugate base of HCO_3^-
7. Decide whether each statement is true or false, and explain your reasoning.
   (a) HBr is a stronger acid than HI.
   (b) $HBrO_2$ is a stronger acid than HBrO.
   (c) H_2SO_3 is a stronger acid than HSO_3^-.
8. In each pair of bases, which is the stronger base?
   (a) $HSO_4^-{}_{(aq)}$ or $SO_4^{2-}{}_{(aq)}$
   (b) $S^{2-}{}_{(aq)}$ or $HS^-{}_{(aq)}$
   (c) $HPO_4^{2-}{}_{(aq)}$ or $H_2PO_4^-{}_{(aq)}$
   (d) $HCO_3^-{}_{(aq)}$ or $CO_3^{2-}{}_{(aq)}$
9. Sodium hydrogen carbonate can be used to neutralize an acid. The hydrogen carbonate ion is the conjugate base of which acid?
10. Explain the significance of the following statement, and give an example to illustrate its meaning: Water is amphoteric.

Answers to questions highlighted in red type are provided in Appendix A.

11. 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) $HCO_3^-{}_{(aq)} + H_3O^+{}_{(aq)}$
   (b) $HCO_3^-{}_{(aq)} + OH^-{}_{(aq)}$
12. 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-}
13. How is K_w related to the pH of acids and bases?
14. How are pH and pOH related to each other?

Inquiry

15. 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.
16. 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
   (d) Are products or reactants favoured in each of the above three reactions? Explain how you know.
17. 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?
18. 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.
19. 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?
20. (a) Calculate the pOH of the solution that forms when 375 mL of a 0.000102 mol/L $Ba(OH)_{2(aq)}$ is diluted to 1.50 L.
   (b) Calculate the pH of the final solution.
   (c) Calculate the hydronium ion concentration of the final solution.
21. 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.

Communicate

22. Predict the products of the following aqueous reactions.
   (a) $NO_2^-{}_{(aq)} + H_2O_{(\ell)}$
   (b) $H_2PO_4^-{}_{(aq)} + H_2O_{(\ell)}$
23. Rank the following in order of decreasing acidity: HF, H_2O, OH^-, HCl, HSO_4^-.
24. Rank the following in order of increasing basicity: F^-, Cl^-, HSO_3^-, S^{2-}, NH_3.
25. 40.0 g of sodium hydroxide is dissolved in enough water to make 1.00 L of solution. Afterward, 2.50 mL of the solution is diluted to 250 mL. What is the pH of the final, diluted solution?