19

Acids, Bases, and Salts

INSIDE:

PearsonChem.com

Artists often use hydrofluoric acid to etch designs on glass.

BIGIDEA

REACTIONS

Essential Questions:

1. *What are the different ways chemists define acids and bases?*
2. *What does the pH of a solution mean?*
3. *How do chemists use acid-base reactions?*

CHEMYSTERY

Paper Trail

The invention of the printing press in 1440 increased the number of books that could be printed and the need for paper on which to print these books. The main ingredient in paper was cotton, which is almost pure cellulose. This type of paper is referred to as rag-based paper. When books printed on high-quality rag-based paper are stored correctly, they can last for hundreds of years.

By 1880, the demand for paper was so great that printers switched from rag-based paper to wood-based paper. Over time, the wood-based paper tended to yellow and crack. Sometimes the paper was so brittle that it crumbled when touched.

Why did a change in the content of paper cause such a dramatic change in the properties of the paper?

▶ Connect to the **BIG**IDEA As you read about acids and bases, think about what could cause paper to crumble.

NATIONAL SCIENCE EDUCATION STANDARDS

A-1, A-2, B-3, G-1, G-2, G-3

19.1 Acid-Base Theories

CHEMISTRY & YOU

Q: *Why are high levels of ammonia harmful to you?* Bracken Cave, near San Antonio, Texas, is home to millions of Mexican free-tailed bats. Nitrogen compounds in bat urine can decompose and release ammonia into the air. Visitors to the cave must wear protective goggles and respirators. They need this protection because of what happens when ammonia reacts with water. Ammonia is an example of a base. In this lesson, you will learn about some of the properties of acids and bases.

Key Questions

How did Arrhenius define an acid and a base?

What distinguishes an acid from a base in the Brønsted-Lowry theory?

How did Lewis define an acid and a base?

Vocabulary

- hydronium ion (H_3O^+)
- conjugate acid
- conjugate base
- conjugate acid-base pair
- amphoteric
- Lewis acid
- Lewis base

Arrhenius Acids and Bases

How did Arrhenius define an acid and a base?

Acids and bases have distinctive properties. Many of the foods you eat, including those shown in Figure 19.1a, contain acids. Acids give foods a tart or sour taste. Lemons, which taste sour enough to make your mouth pucker, contain citric acid. Aqueous solutions of acids are strong or weak electrolytes. Recall that an electrolyte can conduct electricity. The electrolyte in a car battery is an acid. Acids cause certain chemical dyes, called indicators, to change color. Many metals, such as zinc and magnesium, react with aqueous solutions of acids to produce hydrogen gas.

The soap in Figure 19.1b is a familiar material that has the properties of a base. If you have accidentally tasted soap, you know that it has a bitter taste. This bitter taste is a general property of bases, but one that is dangerous to test. The slippery feel of soap is another property of bases. Like acids, bases will cause an indicator to change color. Bases also form aqueous solutions that are strong or weak electrolytes.

Chemists had known the properties of acids and bases for many years. Yet they were not able to propose a theory to explain this behavior. Then, in 1887, the Swedish chemist Svante Arrhenius proposed a new way of defining and thinking about acids and bases. **According to Arrhenius, acids are hydrogen-containing compounds that ionize to yield hydrogen ions (H^+) in aqueous solution. Bases are compounds that ionize to yield hydroxide ions (OH^-) in aqueous solution.**

Figure 19.1 Acids and Bases
Many items contain acids or bases, or produce acids and bases when dissolved in water. **a.** Citrus fruits contain citric acid ($HC_6H_7O_7$). **b.** Many soaps are made using the common base sodium hydroxide (NaOH).

Arrhenius Acids Table 19.1 lists six common acids. They vary in the number of hydrogens they contain that can form hydrogen ions. A hydrogen atom that can form a hydrogen ion is described as *ionizable*. Nitric acid (HNO_3) has one ionizable hydrogen, so nitric acid is classified as a *monoprotic* acid. The prefix *mono-* means "one," and the stem *protic* reflects the fact that a hydrogen ion is a proton. Acids that contain two ionizable hydrogens, such as sulfuric acid (H_2SO_4), are called *diprotic* acids. Acids that contain three ionizable hydrogens, such as phosphoric acid (H_3PO_4), are called *triprotic* acids.

Not all compounds that contain hydrogen are acids. Also, some hydrogens in an acid may not form hydrogen ions. Only a hydrogen that is bonded to a very electronegative element can be released as an ion. Recall that such bonds are highly polar. When a compound that contains such bonds dissolves in water, it releases hydrogen ions. An example is the hydrogen chloride molecule, shown below.

$$\underset{\text{Hydrogen chloride}}{\overset{\delta+}{H}-\overset{\delta-}{Cl}(g)} \xrightarrow{H_2O} \underset{\text{Hydrogen ion}}{H^+(aq)} + \underset{\text{Chloride ion}}{Cl^-(aq)}$$

(hydrochloric acid)

However, in an aqueous solution, hydrogen ions are not present. Instead, the hydrogen ions are joined to water molecules as hydronium ions. A **hydronium ion (H_3O^+)** is the ion that forms when a water molecule gains a hydrogen ion. As seen in Figure 19.2, hydrogen chloride ionizes to form an aqueous solution of hydronium ions and chloride ions.

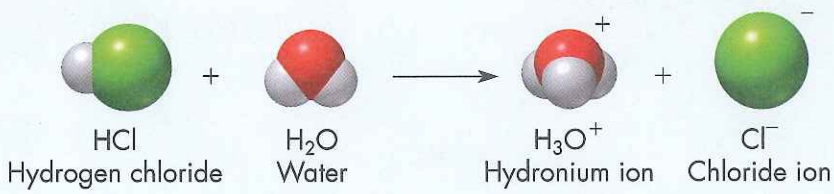
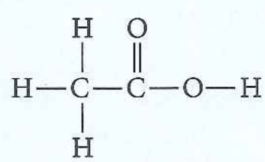


Figure 19.2 Hydrochloric Acid
Hydrochloric acid is actually an aqueous solution of hydrogen chloride. Hydrogen chloride forms hydronium ions, making this compound an acid.
Explain ***Why does hydrogen chloride release a hydrogen ion when dissolved in water?***

See acid dissociation animated online. KINETIC ART

In contrast to hydrogen chloride, methane (CH_4) is an example of a hydrogen-containing compound that is not an acid. The four hydrogen atoms in methane are attached to the central carbon atom by weakly polar C—H bonds. Thus, methane has no ionizable hydrogens and is not an acid. Ethanoic acid (CH_3COOH), which is commonly called acetic acid, is an example of a molecule that contains both hydrogens that do not ionize and a hydrogen that does ionize. Although its molecules contain four hydrogens, ethanoic acid is a monoprotic acid. The structural formula shows why.



Ethanoic acid
(CH_3COOH)

The three hydrogens attached to a carbon atom are in weakly polar bonds. They do not ionize. Only the hydrogen bonded to the highly electronegative oxygen can be ionized. For complex acids, you need to look at the structural formula to recognize which hydrogens can be ionized.

Table 19.1

Some Common Acids

Name	Formula
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulfuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Ethanoic acid	CH_3COOH
Carbonic acid	H_2CO_3

Figure 19.3 Clogged Drains
Sometimes water backs up in a sink because the drain is clogged. A plumber can take apart the pipes to remove a clog, or a drain cleaner containing sodium hydroxide can be used to eat away the clog.

Table 19.2

Some Common Bases		
Name	**Formula**	**Solubility in water**
Sodium hydroxide	$NaOH$	High
Potassium hydroxide	KOH	High
Calcium hydroxide	$Ca(OH)_2$	Very low
Magnesium hydroxide	$Mg(OH)_2$	Very low

Arrhenius Bases Table 19.2 lists four common bases. You may be familiar with the base sodium hydroxide (NaOH), which is also known as lye. Sodium hydroxide is an ionic solid. It dissociates into sodium ions and hydroxide ions in aqueous solution.

$$\underset{\text{Sodium hydroxide}}{NaOH(s)} \xrightarrow{H_2O} \underset{\text{Sodium ion}}{Na^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

Sodium hydroxide is extremely caustic. A caustic substance can burn or eat away materials with which it comes in contact. This property is the reason that sodium hydroxide is a major component of products that are used to clean clogged drains. Figure 19.3 shows a drain cleaner that contains sodium hydroxide.

Potassium hydroxide (KOH) is another ionic solid. It dissociates to produce potassium ions and hydroxide ions in aqueous solution.

$$\underset{\text{Potassium hydroxide}}{KOH(s)} \xrightarrow{H_2O} \underset{\text{Potassium ion}}{K^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

Sodium and potassium are Group 1A elements. Elements in Group 1A, the alkali metals, react violently with water. The products of these reactions are aqueous solutions of a hydroxide and hydrogen gas. The following equation summarizes the reaction of sodium with water.

$$\underset{\text{Sodium metal}}{2Na(s)} + \underset{\text{Water}}{2H_2O(l)} \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Sodium hydroxide and potassium hydroxide are very soluble in water. Thus, making concentrated solutions of these compounds is easy. The solutions would have the typically bitter taste and slippery feel of a base. However, these are not properties that you would want to confirm. The solutions are extremely caustic to the skin. They can cause deep, painful, slow-healing wounds if not immediately washed off.

CHEMISTRY & YOU

Q: *Visitors to Bracken Cave wear protective gear to keep ammonia gas out of their eyes and respiratory tracts. Think about the properties of bases. Why are high levels of ammonia harmful?*

Calcium hydroxide, $Ca(OH)_2$, and magnesium hydroxide, $Mg(OH)_2$, are compounds of Group 2A metals. These compounds are not very soluble in water. Their solutions are always very dilute, even when saturated. A saturated solution of calcium hydroxide has only 0.165 g $Ca(OH)_2$ per 100 g of water. Magnesium hydroxide is even less soluble than calcium hydroxide. A saturated solution has only 0.0009 g $Mg(OH)_2$ per 100 g of water. Figure 19.4 shows a suspension of magnesium hydroxide in water. Some people use this suspension as an antacid and as a mild laxative.

Figure 19.4 Milk of Magnesia
This product is a suspension of magnesium hydroxide in water. Most bases are too caustic to be swallowed. But, the low solubility of magnesium hydroxide makes the suspension safe to consume.

Brønsted-Lowry Acids and Bases

What distinguishes an acid from a base in the Brønsted-Lowry theory?

The Arrhenius definition of acids and bases is not a very broad one. It excludes some substances that have acidic or basic properties, on their own or in solution. For example, sodium carbonate (Na_2CO_3) and ammonia (NH_3) act as bases when they form aqueous solutions. Yet neither of these compounds is a hydroxide-containing compound, so neither compound would be classified as a base by the Arrhenius definition.

In 1923, the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry were working independently. Each chemist proposed the same definition of acids and bases. **According to the Brønsted-Lowry theory, an acid is a hydrogen-ion donor and a base is a hydrogen-ion acceptor.** This theory includes all the acids and bases that Arrhenius defined. It also includes some compounds that Arrhenius did not classify as bases.

You can use the Brønsted-Lowry theory to understand why ammonia is a base. Ammonia gas is very soluble in water. When ammonia dissolves in water, hydrogen ions are transferred from water to ammonia to form ammonium ions and hydroxide ions.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Figure 19.5 illustrates how each water molecule donates a hydrogen ion to ammonia. Ammonia is a Brønsted-Lowry base because it accepts hydrogen ions. Water is a Brønsted-Lowry acid because it donates hydrogen ions.

KINETIC ART *See ammonia in water animated online.*

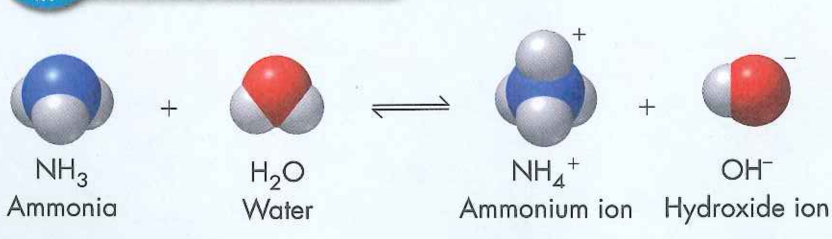


Figure 19.5 Ammonia in Water
When ammonia and water react, water molecules donate hydrogen ions to ammonia molecules. The reaction produces ammonium ions and hydroxide ions.
Explain *Why is ammonia not classified as an Arrhenius base?*

Conjugate Acids and Bases All gases become less soluble in water as the temperature rises. Thus, when the temperature of an aqueous solution of ammonia is increased, ammonia gas is released. This release acts as a stress on the system. In response to this stress, NH_4^+ reacts with OH^- to form more NH_3 and H_2O. In the reverse reaction, ammonium ions donate hydrogen ions to hydroxide ions. Thus, NH_4^+ (the donor) acts as a Brønsted-Lowry acid, and OH^- (the acceptor) acts as a Brønsted-Lowry base. In essence, the reversible reaction of ammonia and water has two acids and two bases.

$$\underset{\text{Base}}{NH_3(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{NH_4^+(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

In the equation, the products of the forward reaction are distinguished from the reactants by the use of the adjective *conjugate*. This term comes from a Latin word meaning "to join together." A **conjugate acid** is the ion or molecule formed when a base gains a hydrogen ion. In the reaction above, NH_4^+ is the conjugate acid of the base NH_3. A **conjugate base** is the ion or molecule that remains after an acid loses a hydrogen ion. In the reaction above, OH^- is the conjugate base of the acid H_2O.

Conjugate acids are always paired with a base, and conjugate bases are always paired with an acid. A **conjugate acid-base pair** consists of two ions or molecules related by the loss or gain of one hydrogen ion. The ammonia molecule and the ammonium ion are a conjugate acid-base pair. The water molecule and the hydroxide ion are also a conjugate acid-base pair.

$$\underset{\text{Base}}{NH_3(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{NH_4^+(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

The dissociation of hydrogen chloride in water provides another example of conjugate acids and bases.

$$\underset{\text{Acid}}{HCl(g)} + \underset{\text{Base}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{H_3O^+(aq)} + \underset{\text{Conjugate base}}{Cl^-(aq)}$$

In this reaction, hydrogen chloride is the hydrogen-ion donor. Thus, it is by definition a Brønsted-Lowry acid. Water is the hydrogen-ion acceptor and a Brønsted-Lowry base. The chloride ion is the conjugate base of the acid HCl. The hydronium ion is the conjugate acid of the base water.

Figure 19.6 shows the reaction that takes place when sulfuric acid dissolves in water. The products of this reaction are hydronium ions and hydrogen sulfate ions. Use the figure to identify the two conjugate acid-base pairs.

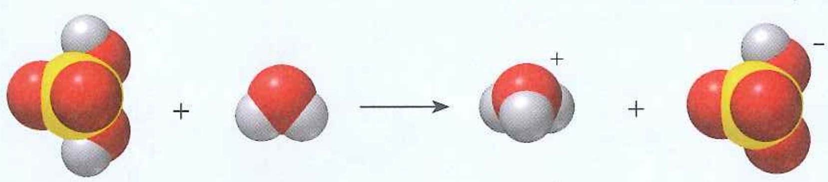


Figure 19.6 Sulfuric Acid When sulfuric acid and water react, they form hydronium ions and hydrogen sulfate ions. **Identify** ***Which product is the conjugate acid, and which is the conjugate base?***

Amphoteric Substances Look at Table 19.3. Note that water appears in both the list of acids and the list of bases. Sometimes water accepts a hydrogen ion. At other times, it donates a hydrogen ion. How water behaves depends on the other reactant. A substance that can act as either an acid or a base is said to be **amphoteric.** Water is amphoteric. In the reaction with hydrochloric acid, water accepts a proton and is therefore a base. In the reaction with ammonia, water donates a proton and is therefore an acid. Look for two other substances in Table 19.3 that are amphoteric.

Table 19.3

Some Conjugate Acid-Base Pairs

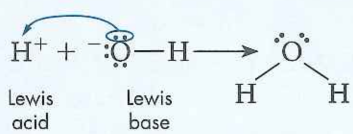
Acid	Base
HCl	Cl^-
H_2SO_4	HSO_4^-
H_3O^+	H_2O
HSO_4^-	SO_4^{2-}
CH_3COOH	CH_3COO^-
H_2CO_3	HCO_3^-
HCO_3^-	CO_3^{2-}
NH_4^+	NH_3
H_2O	OH^-

Lewis Acids and Bases

How did Lewis define an acid and a base?

The work that Gilbert Lewis (1875–1946) did on bonding led to a new concept of acids and bases. **According to Lewis, an acid accepts a pair of electrons and a base donates a pair of electrons during a reaction.** This definition is more general than those offered by Arrhenius or by Brønsted and Lowry. A **Lewis acid** is a substance that can accept a pair of electrons to form a covalent bond. Similarly, a **Lewis base** is a substance that can donate a pair of electrons to form a covalent bond.

The Lewis definitions include all the Brønsted-Lowry acids and bases. Consider the reaction of H^+ and OH^-. The hydrogen ion donates itself to the hydroxide ion. Therefore, H^+ is a Brønsted-Lowry acid and OH^- is a Brønsted-Lowry base. The hydroxide ion can bond to the hydrogen ion because it has an unshared pair of electrons. Thus, OH^- is also a Lewis base, and H^+, which accepts the pair of electrons, is a Lewis acid.



A second example of a reaction between a Lewis acid and a Lewis base is what happens when ammonia dissolves in water. Hydrogen ions from the dissociation of water are the electron-pair acceptor and the Lewis acid. Ammonia is the electron-pair donor and the Lewis base.

Table 19.4 compares the definitions of acids and bases. The Lewis definition is the broadest. It extends to compounds that the Brønsted-Lowry theory does not classify as acids and bases. Sample Problem 19.1 provides some examples of those compounds.

Table 19.4

Acid-Base Definitions

Type	Acid	Base
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

READING SUPPORT

Building Vocabulary: ***Prefixes*** The prefix *amphi-* is from a Greek word meaning "of both kinds." An amphibian is an animal that is capable of living both on land and in the water. ***What does it mean to describe an airplane as* amphibious*?***

Sample Problem 19.1

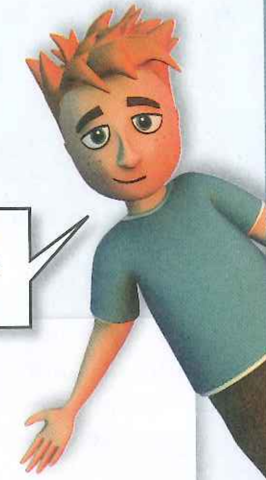
Identifying Lewis Acids and Bases

Identify the Lewis acid and the Lewis base in this reaction between ammonia and boron trifluoride.

$$NH_3 + BF_3 \longrightarrow NH_3BF_3$$

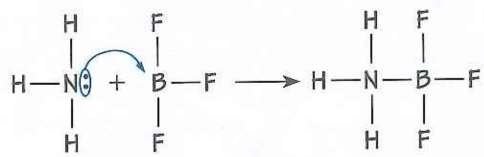
1 Analyze **Identify the relevant concepts.** When a Lewis acid reacts with a Lewis base, the base donates a pair of electrons and the acid accepts the donated pair.

2 Solve **Apply the concepts to this problem.**



Draw electron dot structures to identify which reactant has an unshared pair of electrons.

Identify the reactant with the unshared pair of electrons and the reactant that can accept the pair of electrons.



Ammonia has an unshared pair of electrons to donate. The boron atom can accept the donated electrons.

Classify the reactants based on their behavior.

Lewis bases donate a pair of electrons, so ammonia is the Lewis base. Lewis acids accept a pair of electrons, so boron trifluoride is the Lewis acid.

1. Identify the Lewis acid and Lewis base in each reaction.

a. $H^+ + H_2O \longrightarrow H_3O^+$

b. $AlCl_3 + Cl^- \longrightarrow AlCl_4^-$

2. Predict whether PCl_3 would be a Lewis acid or a Lewis base in typical chemical reactions. Explain your prediction.

19.1 LessonCheck

ONLINE PROBLEMS

3. **Review** What is the Arrhenius definition of an acid and a base?

4. **Describe** How are acids and bases defined by the Brønsted-Lowry theory?

5. **Explain** How did Lewis broaden the definition of acids and bases?

6. **Compare and Contrast** How are the properties of acids and bases similar? How are they different?

7. **Classify** Determine whether the following acids are monoprotic, diprotic, or triprotic:

a. H_2CO_3 **b.** H_3PO_4 **c.** HCl **d.** H_2SO_4

8. **Apply Concepts** Write a chemical equation for the ionization of HNO_3 in water and for the reaction of CO_3^{2-} with water. Identify the hydrogen-ion donor and the hydrogen-ion acceptor in each equation. Then, label each conjugate acid-base pair in the two equations.

BIGIDEA REACTIONS

9. Some household drain cleaners contain both sodium hydroxide and small particles of aluminum or zinc. Research how adding these metals can increase the effectiveness of the product.

19.2 Hydrogen Ions and Acidity

CHEMISTRY & YOU

Q: *What factors do you need to control so a fish has healthy water to live in?* Goldfish can live for 20 years or more in an aquarium if the conditions are right. The water in the aquarium must be cleaned regularly. You must also control the temperature of the water. In this lesson, you will study another factor that affects the ability of fish to survive.

Key Questions

How are [H⁺] and [OH⁻] related in an aqueous solution?

How is pH used to classify a solution as neutral, acidic, or basic?

What are two methods that are used to measure pH?

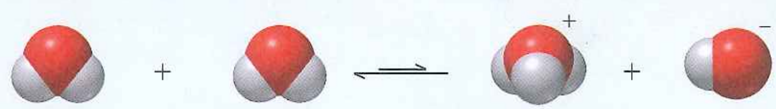
Vocabulary

- self-ionization
- neutral solution
- ion-product constant for water (K_w)
- acidic solution
- basic solution
- pH

Hydrogen Ions From Water

How are $[H^+]$ and $[OH^-]$ related in an aqueous solution?

Water molecules are highly polar and are in constant motion, even at room temperature. On occasion, the collisions between water molecules are energetic enough for a reaction to occur. When this happens, a hydrogen ion is transferred from one water molecule to another, as illustrated below. A water molecule that gains a hydrogen ion becomes a hydronium ion (H_3O^+). A water molecule that loses a hydrogen ion becomes a hydroxide ion (OH^-).



H_2O		H_2O		H_3O^+		OH^-
Water molecule	+	Water molecule	$\rightleftharpoons$	Hydronium ion	+	Hydroxide ion

Self-Ionization of Water The reaction in which water molecules produce ions is called the **self-ionization** of water. This reaction can be written as a simple dissociation.

$$\underset{}{H_2O(l)} \rightleftharpoons \underset{\text{Hydrogen ion}}{H^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

In water or in an aqueous solution, hydrogen ions are always joined to water molecules as hydronium ions. Yet chemists may still refer to these ions as hydrogen ions or even protons. In this textbook, either H^+ or H_3O^+ is used to represent hydrogen ions in aqueous solution.

The self-ionization of water occurs to a very small extent. In pure water at 25°C, the concentration of hydrogen ions is only $1 \times 10^{-7}M$. The concentration of OH^- is also $1 \times 10^{-7}M$ because the numbers of H^+ and OH^- ions are equal in pure water. Any aqueous solution in which $[H^+]$ and $[OH^-]$ are equal is a **neutral solution.**

Figure 19.7 Aged by Acid
Sometimes guitar players want a new guitar to look like it is old or "vintage." The guitarist can remove the shiny new metal parts of the guitar and expose them to hydrochloric acid. The acid will make the metal parts look dull. Both of the guitars in the photo below are new, but the bottom one has been aged with acid.

Ion-Product Constant for Water The ionization of water is a reversible reaction, so Le Châtelier's principle applies. Adding either hydrogen ions or hydroxide ions to an aqueous solution is a stress to the system. In response, the equilibrium will shift toward the formation of water. The concentration of the other ion will decrease. In any aqueous solution, when $[H^+]$ increases, $[OH^-]$ decreases. Likewise, when $[H^+]$ decreases, $[OH^-]$ increases.

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

For aqueous solutions, the product of the hydrogen-ion concentration and the hydroxide-ion concentration equals 1.0×10^{-14}.

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

This equation is true for all dilute aqueous solutions at 25°C. When substances are added to water, the concentrations of H^+ and OH^- may change. However, the product of $[H^+]$ and $[OH^-]$ does not change. The product of the concentrations of the hydrogen ions and hydroxide ions in water is called the **ion-product constant for water (K_w).**

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

Acidic Solutions Not all solutions are neutral. When some substances dissolve in water, they release hydrogen ions. For example, when hydrogen chloride dissolves in water, it forms hydrochloric acid.

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

In hydrochloric acid, the hydrogen-ion concentration is greater than the hydroxide-ion concentration. (The hydroxide ions come from the self-ionization of water.) A solution in which $[H^+]$ is greater than $[OH^-]$ is an **acidic solution.** In acidic solutions, the $[H^+]$ is greater than $1 \times 10^{-7}M$. Figure 19.7 shows a guitar that was artificially aged by using hydrochloric acid.

Basic Solutions When sodium hydroxide dissolves in water, it forms hydroxide ions in solution.

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

In such a solution, the hydrogen-ion concentration is less than the hydroxide-ion concentration. Remember, the hydrogen ions are present from the self-ionization of water. A **basic solution** is one in which $[H^+]$ is less than $[OH^-]$. The $[H^+]$ of a basic solution is less than $1 \times 10^{-7}M$. Basic solutions are also known as alkaline solutions.

Sample Problem 19.2

Using the Ion-Product Constant for Water

If the $[H^+]$ in a solution is $1.0 \times 10^{-5} M$, is the solution acidic, basic, or neutral? What is the $[OH^-]$ of this solution?

KNOWNS

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

$K_w = 1 \times 10^{-14}$

UNKNOWNS

Is the solution acidic, basic, or neutral?

$[OH^-] = ?M$

1 Analyze **List the knowns and the unknowns.** Use the expression for the ion-product constant for water and the known concentration of hydrogen ions to find the concentration of hydroxide ions.

2 Calculate **Solve for the unknowns.**

Use $[H^+]$ to determine whether the solution is acidic, basic, or neutral.

$[H^+]$ is $1.0 \times 10^{-5} M$, which is greater than $1.0 \times 10^{-7} M$. Thus, the solution is acidic.

Rearrange the expression for the ion-product constant to solve for $[OH^-]$.

$$K_w = [H^+] \times [OH^-]$$

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

Substitute the known values of $[H^+]$ and K_w. Then, solve for $[OH^-]$.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}}$$

$$= 1.0 \times 10^{-9} M$$

When you divide numbers written in scientific notation, subtract the exponent in the denominator from the exponent in the numerator.

3 Evaluate **Does the result make sense?** If $[H^+]$ is greater than $1.0 \times 10^{-7} M$, then $[OH^-]$ must be less than $1.0 \times 10^{-7} M$. $1 \times 10^{-9} M$ is less than $1 \times 10^{-7} M$. To check your calculation, multiply the values for $[H^+]$ and $[OH^-]$ to make sure the result equals 1×10^{-14}.

For Problem 11, rearrange the expression for the ion-product constant to solve for $[H^+]$.

10. Classify each solution as acidic, basic, or neutral.

a. $[H^+] = 6.0 \times 10^{-10} M$

b. $[OH^-] = 3.0 \times 10^{-2} M$

c. $[H^+] = 2.0 \times 10^{-7} M$

d. $[OH^-] = 1.0 \times 10^{-7} M$

11. If the hydroxide-ion concentration of an aqueous solution is $1 \times 10^{-3} M$, what is the $[H^+]$ in the solution? Is the solution acidic, basic, or neutral?

The pH Concept

How is pH used to classify a solution as neutral, acidic, or basic?

CHEMISTRY & YOU

Q: *In an aquarium, the pH of water is another factor that affects the ability of fish to survive. Most freshwater fish need a slightly acidic or neutral pH. For a saltwater tank, the ideal pH is slightly basic. What might explain this difference in the ideal pH range?*

Expressing hydrogen-ion concentration in molarity is not practical. A more widely used system for expressing $[H^+]$ is the pH scale, proposed in 1909 by the Danish scientist Søren Sørensen. The pH scale ranges from 0 to 14.

Hydrogen Ions and pH The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration. The pH may be represented mathematically using the following equation:

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

In pure water or a neutral solution, the $[H^+] = 1 \times 10^{-7}M$, and the pH is 7.

$$\begin{aligned}\mathrm{pH} &= -\log(1 \times 10^{-7}) \\ &= -(\log 1 + \log 10^{-7}) \\ &= -(0.0 + (-7.0)) = 7.0\end{aligned}$$

If the $[H^+]$ of a solution is greater than $1 \times 10^{-7}M$, the pH is less than 7.0. If the $[H^+]$ of the solution is less than $1 \times 10^{-7}M$, the pH is greater than 7.0. **A solution with a pH less than 7.0 is acidic. A solution with a pH of 7.0 is neutral. A solution with a pH greater than 7.0 is basic.** Table 19.5 summarizes the relationship among $[H^+]$, $[OH^-]$, and pH. It also indicates the pH values of some common aqueous systems, including milk and blood.

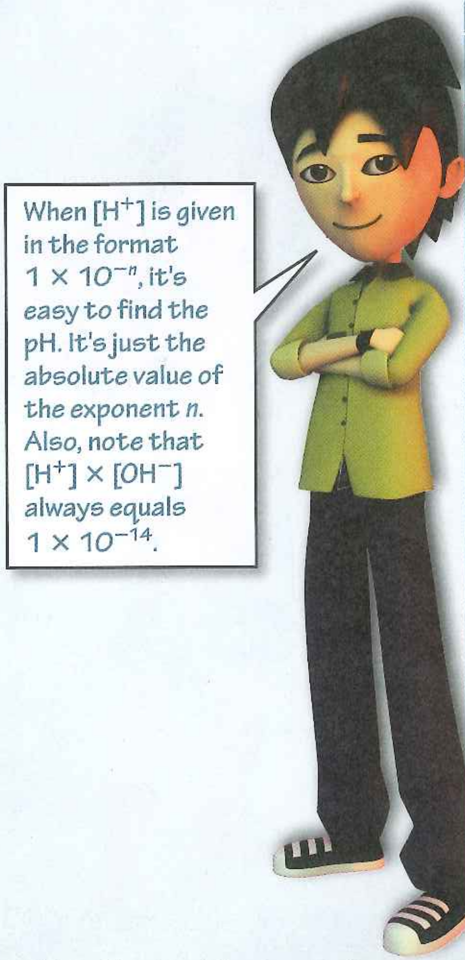


Table 19.5

Relationships Among $[H^+]$, $[OH^-]$, and pH

	$[H^+]$ (mol/L)	$[OH^-]$ (mol/L)	pH	
Increasing acidity ↑	1×10^{0}	1×10^{-14}	0.0	1 *M* HCl
	1×10^{-1}	1×10^{-13}	1.0	0.1 *M* HCl
	1×10^{-2}	1×10^{-12}	2.0	Gastric juice
	1×10^{-3}	1×10^{-11}	3.0	Lemon juice
	1×10^{-4}	1×10^{-10}	4.0	Tomato juice
	1×10^{-5}	1×10^{-9}	5.0	Black coffee
	1×10^{-6}	1×10^{-8}	6.0	Milk
Neutral	1×10^{-7}	1×10^{-7}	7.0	Pure water, Blood
Increasing basicity ↓	1×10^{-8}	1×10^{-6}	8.0	Seawater
	1×10^{-9}	1×10^{-5}	9.0	
	1×10^{-10}	1×10^{-4}	10.0	Milk of magnesia
	1×10^{-11}	1×10^{-3}	11.0	Household ammonia
	1×10^{-12}	1×10^{-2}	12.0	
	1×10^{-13}	1×10^{-1}	13.0	0.1 *M* NaOH
	1×10^{-14}	1×10^{0}	14.0	1 *M* NaOH

Calculating pH From $[H^+]$ Expressing $[H^+]$ in scientific notation can make it easier to calculate pH. For example, you would rewrite $0.0010M$ as $1.0 \times 10^{-3}M$. The coefficient 1.0 has two significant figures. The pH for a solution with this concentration is 3.00. The two numbers to the right of the decimal point represent the two significant figures in the concentration.

It is easy to find the pH for solutions when the coefficient is 1.0. The pH of the solution equals the exponent, with the sign changed from minus to plus. For example, a solution with $[H^+] = 1 \times 10^{-2}M$ has a pH of 2.0. When the coefficient is a number other than 1, you will need to use a calculator with a log function key to calculate pH.

CHEM TUTOR

Sample Problem 19.3

Calculating pH From $[H^+]$

What is the pH of a solution with a hydrogen-ion concentration of $4.2 \times 10^{-10}M$?

1 Analyze List the known and the unknown. To find the pH from the hydrogen-ion concentration, you use the equation $pH = -\log[H^+]$.

KNOWN	UNKNOWN
$[H^+] = 4.2 \times 10^{-10}M$	pH = ?

2 Calculate Solve for the unknown.

Start with the equation for finding pH from $[H^+]$.

$$pH = -\log[H^+]$$

Substitute the known $[H^+]$ and use the log function on your calculator to calculate the pH.

$$\begin{aligned} pH &= -\log(4.2 \times 10^{-10}) \\ &= -(-9.37675) \\ &= 9.37675 \\ &= 9.38 \end{aligned}$$

Round the pH to two decimal places because the hydrogen-ion concentration has two significant figures.

3 Evaluate Does the result make sense? The value of the hydrogen-ion concentration is between $1 \times 10^{-9}M$ and $1 \times 10^{-10}M$. So, the calculated pH should be between 9 and 10, which it is.

12. Find the pH of each solution.
- **a.** $[H^+] = 0.045M$
- **b.** $[H^+] = 8.7 \times 10^{-6}M$
- **c.** $[H^+] = 0.0015M$
- **d.** $[H^+] = 1.2 \times 10^{-3}M$

13. What are the pH values of the following solutions, based on their hydrogen-ion concentrations?
- **a.** $[H^+] = 1.0 \times 10^{-12}M$
- **b.** $[H^+] = 1 \times 10^{-4}M$

Calculating [H^+] From pH You can calculate the hydrogen-ion concentration of a solution if you know the pH. If the pH is an integer, it is easy to find the value of [H^+]. For a pH of 9.0, [H^+] $= 1 \times 10^{-9}M$. For a pH of 4.0, [H^+] is $1 \times 10^{-4}M$.

However, most pH values are not whole numbers. For example, milk of magnesia has a pH of 10.50. The [H^+] must be less than $1 \times 10^{-10}M$ (pH 10.0) but greater than $1 \times 10^{-11}M$ (pH 11.0). The hydrogen-ion concentration is $3.2 \times 10^{-11}M$. When the pH value is not a whole number, you will need a calculator with an antilog (10^x) function to get an accurate value for the hydrogen-ion concentration.

SampleProblem 19.4

Calculating [H^+] From pH

The pH of an unknown solution is 6.35. What is the hydrogen-ion concentration of the solution?

KNOWN
pH = 6.35

UNKNOWN
[H^+] = ?M

1 Analyze **List the known and the unknown.** You will use the antilog function of your calculator to find the concentration.

2 Calculate **Solve for the unknown.**

First, simply swap the sides of the equation for finding pH and substitute the known value.

$$\mathrm{pH} = -\log[H^+]$$
$$-\log[H^+] = \mathrm{pH}$$
$$-\log[H^+] = 6.35$$

Change the signs on both sides of the equation and then solve for the unknown.

$$\log[H^+] = -6.35$$
$$[H^+] = \mathrm{antilog}(-6.35)$$

Use the antilog (10^x) function on your calculator to find [H^+]. Report the answer in scientific notation.

$$[H^+] = 4.5 \times 10^{-7}M$$

On most calculators, use the 2nd or the INV key followed by log to get the antilog.

3 Evaluate **Does the result make sense?** The pH is between 6 and 7. So, the hydrogen ion concentration must be between $1 \times 10^{-6}M$ and $1 \times 10^{-7}M$. The answer is rounded to two significant figures because the pH was measured to two decimal places.

14. Calculate [H^+] for each solution.
a. pH = 5.00
b. pH = 12.83

15. What are the hydrogen-ion concentrations for solutions with the following pH values?
a. 4.00
b. 11.55

Calculating pH From $[OH^-]$ If you know the $[OH^-]$ of a solution, you can find its pH. Recall that the ion-product constant for water defines the relationship between $[H^+]$ and $[OH^-]$. Therefore, you can use the ion-product constant for water to determine $[H^+]$ for a known $[OH^-]$. Then, you use $[H^+]$ to calculate the pH. For practice, try doing Sample Problem 19.5.

Sample Problem 19.5

Calculating pH From $[OH^-]$

What is the pH of a solution if $[OH^-] = 4.0 \times 10^{-11} M$?

KNOWNS	UNKNOWN
$[OH^-] = 4.0 \times 10^{-11} M$	pH = ?
$K_w = 1.0 \times 10^{-14}$	

1 Analyze **List the knowns and the unknown.**
To find $[H^+]$, divide K_w by the known $[OH^-]$. Then, calculate pH as you did in Sample Problem 19.3.

2 Calculate **Solve for the unknown.**

Start with the ion-product constant to find $[H^+]$. Rearrange the equation to solve for $[H^+]$.

$$K_w = [OH^-] \times [H^+]$$

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

Substitute the values for K_w and $[OH^-]$ to find $[H^+]$.

$$[H^+] = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-11}} = 0.25 \times 10^{-3} M$$

$$= 2.5 \times 10^{-4} M$$

Next, use the equation for finding pH. Substitute the value for $[H^+]$ that you just calculated.

$$\text{pH} = -\log [H^+]$$
$$= -\log (2.5 \times 10^{-4})$$

Use a calculator to find the log.

$$= -(-3.60205)$$
$$= 3.60$$

Round the pH to two decimal places because the $[OH^-]$ has two significant figures.

3 Evaluate **Does the result make sense?** A solution in which $[OH^-]$ is less than $1 \times 10^{-7} M$ is acidic because $[H^+]$ is greater than $1 \times 10^{-7} M$. The hydrogen-ion concentration is between $1 \times 10^{-3} M$ and $1 \times 10^{-4} M$. Thus, the pH should be between 3 and 4.

16. Calculate the pH of each solution.
  a. $[OH^-] = 4.3 \times 10^{-5} M$
  b. $[OH^-] = 4.5 \times 10^{-11} M$

17. Calculate the pH of each solution.
  a. $[OH^-] = 5.0 \times 10^{-9} M$
  b. $[OH^-] = 8.3 \times 10^{-4} M$

Measuring pH

What are two methods that are used to measure pH?

In many situations, knowing the pH is useful. A custodian might need to maintain the correct acid-base balance in a swimming pool. A gardener may want to know if a certain plant will thrive in a yard. A doctor might be trying to diagnose a medical condition. **Either acid-base indicators or pH meters can be used to measure pH.**

Acid-Base Indicators An indicator is often used for initial pH measurements and for samples with small volumes. An indicator (HIn) is an acid or a base that dissociates in a known pH range. Indicators work because their acid form and base form have different colors in solution. The following general equation represents the dissociation of an acid-base indicator (HIn).

$$\underset{\text{Acid form}}{\mathrm{HIn}(aq)} \rightleftharpoons \mathrm{H^+}(aq) + \underset{\text{Base form}}{\mathrm{In^-}(aq)}$$

The acid form of the indicator (HIn) is dominant at low pH and high $[H^+]$. The base form (In^-) is dominant at high pH and high $[OH^-]$.

The change from dominating acid form to dominating base form occurs within a narrow range of about two pH units. Within this range, the color of the solution is a mixture of the colors of the acid and the base forms. If you know the pH range over which this color change occurs, you can make a rough estimate of the pH of a solution. At all pH values below this range, you would see only the color of the acid form. At all pH values above this range, you would see only the color of the base form.

For a more precise estimate of the solution's pH, you could repeat the test with indicators that have different pH ranges for their color change. Many indicators are needed to span the entire pH spectrum. Figure 19.8 shows the pH ranges of some common acid-base indicators.

Interpret Graphs

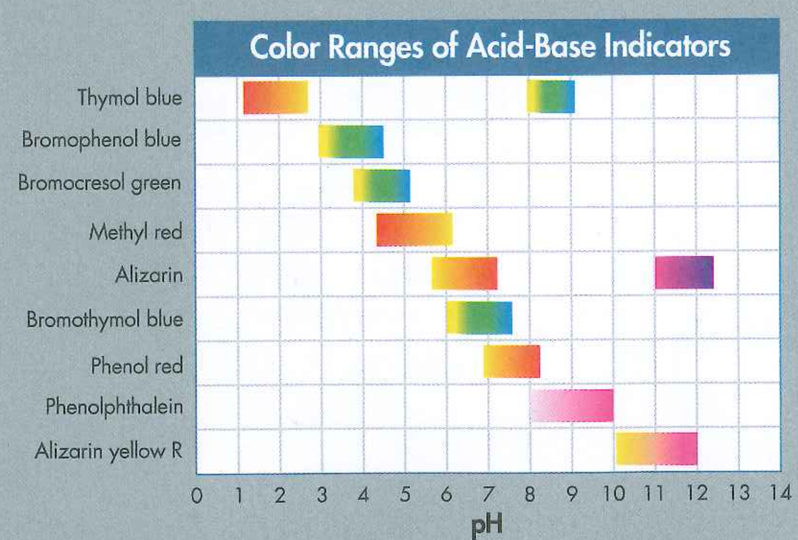


Figure 19.8 Each indicator is useful for a specific range of pH values.
a. Identify At a pH of 12, which indicator would be yellow?
b. Apply Concepts Which indicator could you use to show that the pH of a solution has changed from 3 to 5?
c. Make Generalizations What do you notice about the range over which each indicator changes color?

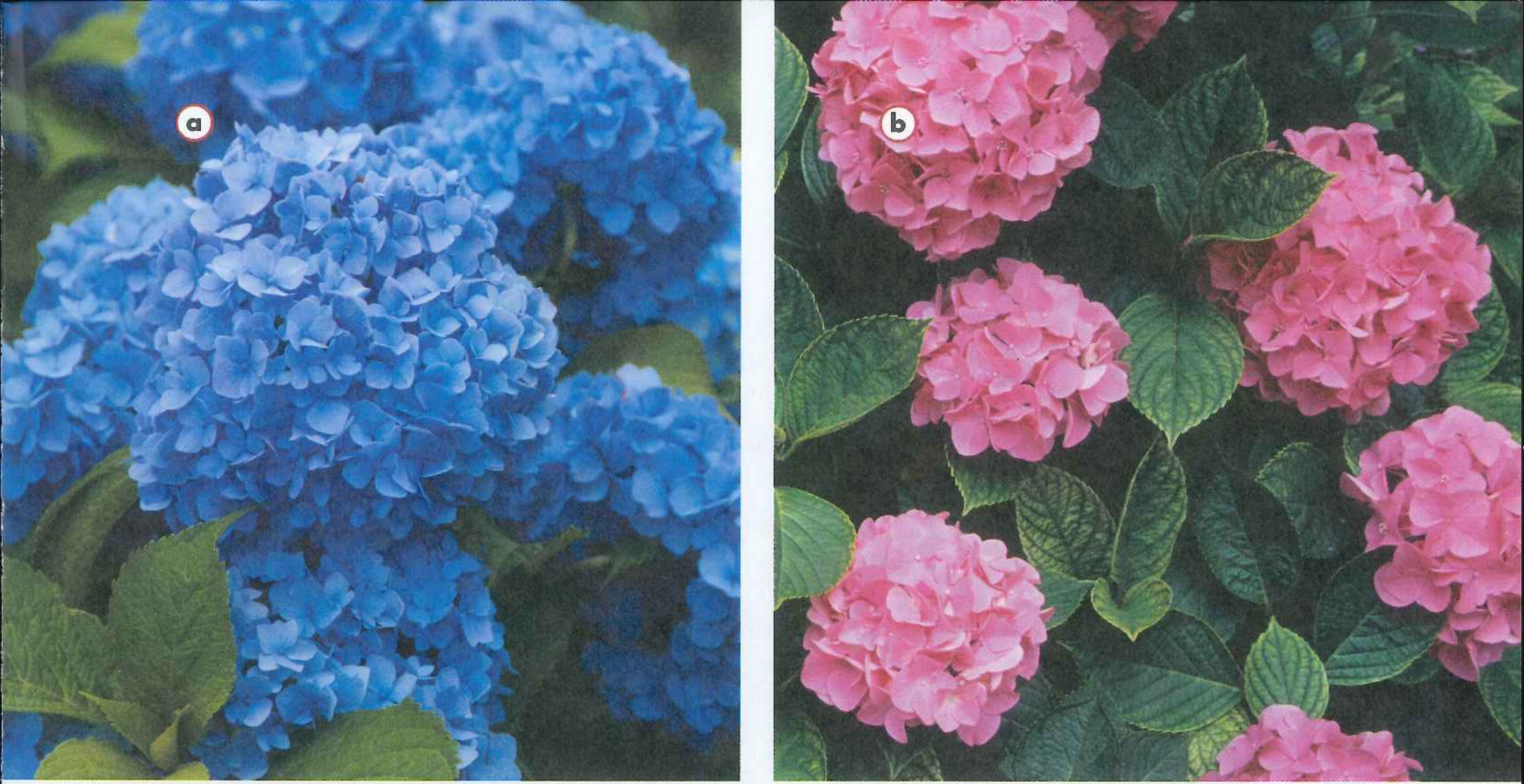


Figure 19.9 Effects of Soil Acidity
Soil pH can affect how plants develop. **a.** In acidic soils, hydrangeas produce blue flowers. **b.** In basic soils, hydrangeas produce pink flowers.

Indicators have certain properties that limit their usefulness. The pH values of indicators are usually given for 25°C. At other temperatures, an indicator may change color at a different pH. If the solution being tested is not colorless, the color of the indicator may be misleading. Dissolved salts in a solution may affect an indicator's dissociation. Using indicator strips can help overcome these problems. An indicator strip is a piece of paper or plastic that has been soaked in an indicator, and then dried. The paper is dipped into an unknown solution. The color that results is compared with a color chart to measure the pH. Some indicator paper has absorbed multiple indicators. The colors that result will cover a wide range of pH values. Before planting the shrub shown in Figure 19.9, you might want to test the pH of your soil.

pH Meters Your chemistry laboratory probably has a pH meter. A pH meter is used to make rapid, continuous measurements of pH. The measurements of pH obtained with a pH meter are typically accurate to within 0.01 pH unit of the true pH. If the pH meter is connected to a computer or chart recorder, the user will have a record of the pH changes.

A pH meter can be easier to use than liquid indicators or indicator strips. As shown in Figure 19.10, the pH reading is visible in a display window on the meter. Hospitals use pH meters to find small but meaningful changes in the pH of blood and other body fluids. Sewage, industrial wastes, and soil pH are also easily monitored with a pH meter. The color and cloudiness of the solution do not affect the accuracy of the pH value obtained.

Figure 19.10 pH Meter
A pH meter provides a quick and accurate way to measure the pH of a system.

Quick Lab

Purpose To measure the pH of household materials using a natural indicator

Materials

- red cabbage leaves
- 1-cup measure
- hot water
- 2 jars
- spoon
- cheesecloth
- 3 sheets of plain white paper
- transparent tape
- metric ruler
- pencil
- 10 small clear plastic cups
- permanent marker
- white vinegar (CH_3COOH)
- baking soda ($NaHCO_3$)
- spatula
- household ammonia
- dropper
- assorted household materials

Indicators From Natural Sources

Procedure

1. Put one-half cup of finely chopped red cabbage leaves in a jar and add one-half cup of hot water. Stir and crush the leaves with a spoon. Continue this process until the water has a distinct color.

2. Strain the mixture through a piece of clean cheesecloth into a clean jar. The liquid that collects in the jar is your indicator.

3. Tape three sheets of paper end to end. Draw a line along the center of the taped sheets. Label the line at 5-cm intervals with the numbers 1 to 14. This labeled line is your pH scale.

4. Use the permanent marker to label three cups vinegar, baking soda, and ammonia. Pour indicator into each cup to a depth of about 1 cm.

5. Add several drops of vinegar to the first cup. Use a spatula to add a pinch of baking soda to the second cup. Add several drops of ammonia to the third cup. The pH values for the solutions of vinegar, baking soda, and household ammonia are about 3, 9, and 11, respectively. Record the colors you observe at the correct locations on your pH scale.

6. Repeat the procedure for household items such as table salt, milk, lemon juice, laundry detergent, milk of magnesia, toothpaste, shampoo, and carbonated beverages.

Analyze and Conclude

1. Observe What color is the indicator in acidic, neutral, and basic solutions?

2. Relate Cause and Effect What caused the color of the indicator to change when a material was added to a cup?

3. Classify Divide the household materials you tested into three groups—acidic, basic, and neutral.

4. Analyze Data Which group contains items used for cleaning? Which group contains items used for personal hygiene?

ONLINE PROBLEMS

19.2 LessonCheck

18. Review How are the concentrations of hydrogen ions and hydroxide ions related in an aqueous solution?

19. Identify What is the range of pH values in the following solutions?

a. basic **b.** acidic **c.** neutral

20. List What methods can you use to measure the pH of a solution?

21. Relate Cause and Effect What happens to the $[H^+]$ as the pH of a solution increases?

22. Calculate Determine the pH of each solution.

a. $[H^+] = 1 \times 10^{-6}M$

b. $[H^+] = 0.00010M$

c. $[OH^-] = 1 \times 10^{-2}M$

d. $[OH^-] = 1 \times 10^{-11}M$

23. Compare In terms of ion concentrations, how do basic solutions differ from acidic solutions?

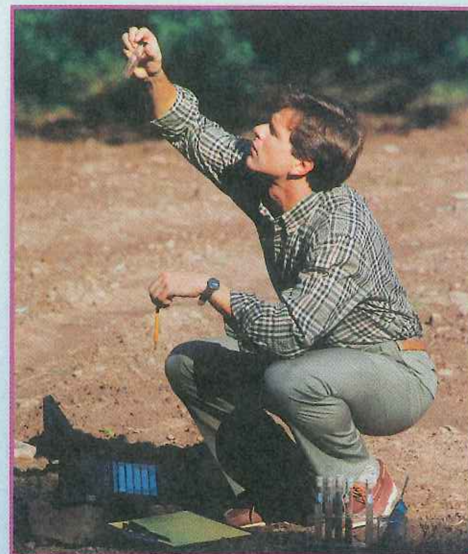
24. Calculate Find the hydroxide-ion concentrations for solutions with the following pH values:

a. 6.00 **b.** 9.00 **c.** 12.00

Agronomist

Do you like the idea of working with plants but find chemistry more interesting than biology? If so, you might want to consider a career in agronomy. Agronomy is a branch of agriculture that deals with the interactions between plants, soils, and the environment. Agronomists use their knowledge of chemistry to help produce healthy crops and increase yields, while preserving the environment.

The opportunities for agronomists extend beyond laboratories and greenhouses. Agronomists also work for business firms, government agencies, conservation groups, philanthropic organizations, and universities. Agronomists can also use their knowledge of water and land management to address such issues as urban area beautification and highway landscaping.



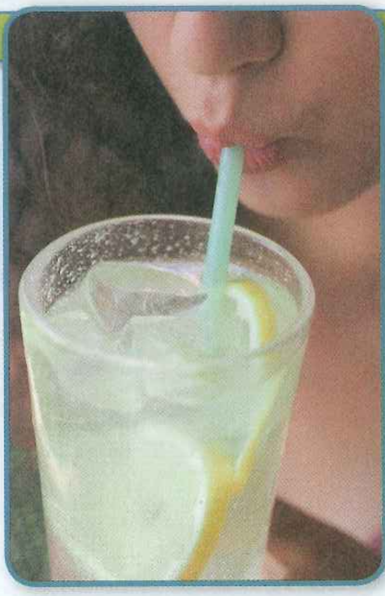
SOIL ACIDITY The pH of the soil is among the most important factors in crop production. Agronomists can help farmers obtain the right soil pH for a specific crop.

FIELD CHEMISTRY Many agronomists have the opportunity to work with local communities. This agronomist is advising a Kenyan farming group on how to improve their crops using science.

Take It Further

1. **Apply Concepts** The ideal soil pH for corn is around 6.0. If $[H^+]$ equals $2.14 \times 10^{-5} M$, is the soil too acidic or too basic for growing corn?
2. **Infer** What are two nonfood items that an agronomist might help produce?

19.3 Strengths of Acids and Bases



CHEMISTRY & YOU

Q: *What makes one acid safer than another?* Lemon juice, which contains citric acid, has a pH of about 2.3. Yet, you consume lemon juice. When you cut a lemon, you usually don't wear gloves or safety goggles. But some acids do require such precautions. This lesson will explain the difference between a "weak" acid such as citric acid and a "strong" acid such as sulfuric acid.

Key Question

How are acids and bases classified as either strong or weak?

Vocabulary

- strong acid
- weak acid
- acid dissociation constant (K_a)
- strong base
- weak base
- base dissociation constant (K_b)

Strong and Weak Acids and Bases

How are acids and bases classified as either strong or weak?

Table 19.6 compares the strengths of some acids and bases. **Acids and bases are classified as strong or weak based on the degree to which they ionize in water.** Hydrochloric acid and sulfuric acid are examples of strong acids. In general, a **strong acid** is completely ionized in aqueous solution.

$$HCl(g) + H_2O(l) \longrightarrow \underbrace{H_3O^+(aq) + Cl^-(aq)}_{100\%}$$

A **weak acid** ionizes only slightly in aqueous solution. The ionization of ethanoic acid (CH_3COOH), a typical weak acid, is not complete.

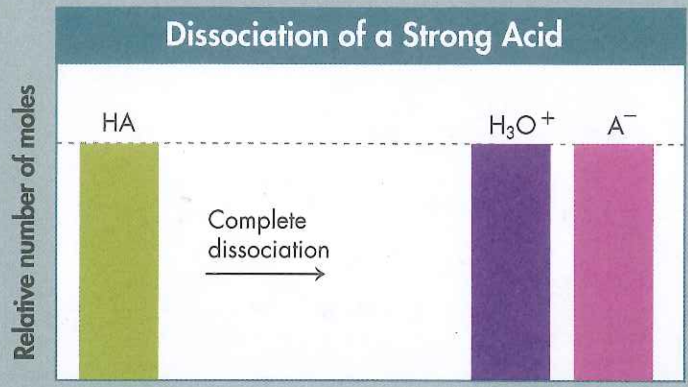
$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons \underbrace{H_3O^+(aq) + CH_3COO^-(aq)}_{<1\%}$$

Table 19.6

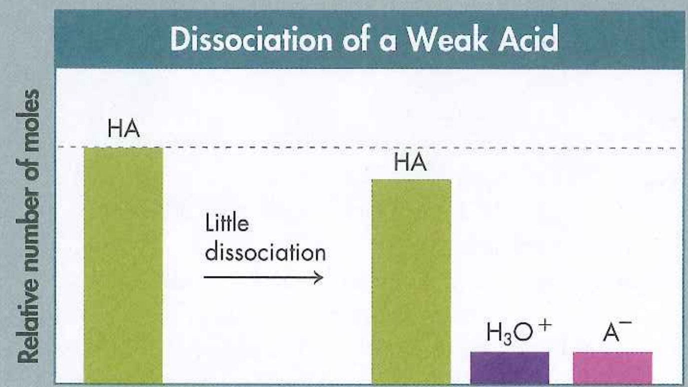
Relative Strengths of Common Acids and Bases

Substance	Formula	Relative strength
Hydrochloric acid	HCl	Strong acids
Nitric acid	HNO_3	
Sulfuric acid	H_2SO_4	
Phosphoric acid	H_3PO_4	↑ Increasing strength of acid
Ethanoic acid	CH_3COOH	
Carbonic acid	H_2CO_3	
Hypochlorous acid	$HClO$	
		Neutral solution
Ammonia	NH_3	↓ Increasing strength of base
Sodium silicate	Na_2SiO_3	
Calcium hydroxide	$Ca(OH)_2$	Strong bases
Sodium hydroxide	$NaOH$	
Potassium hydroxide	KOH	

Interpret Graphs



$$HA(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + A^-(aq)$$



$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

Figure 19.11 Dissociation of an acid (HA) in water yields H_3O^+ and an anion, A^-. The bar graphs compare the extent of dissociation of a strong acid and a weak acid.

a. Explain Why is there only one bar for HA in the graph for the strong acid, but two bars for HA in the graph for the weak acid?

b. Apply Concepts In the graph for the strong acid, why do the bars for H_3O^+ and A^- have the same height as the bar for HA?

c. Infer In the graph for the weak acid, why is the height of the bar for H_3O^+ equal to the distance from the top of the second HA bar to the dotted line?

Hint: The bars represent the relative amounts of the acid and the ions it forms in solution.

Acid Dissociation Constant Figure 19.11 compares the extent of dissociation of strong and weak acids. A strong acid, such as hydrochloric acid, completely dissociates in water. As a result, $[H_3O^+]$ is high in an aqueous solution of a strong acid. By contrast, weak acids remain largely undissociated. For example, in an aqueous solution of ethanoic acid, less than 1 percent of the molecules are ionized at any time.

You can use a balanced equation to write the equilibrium-constant expression for a reaction. The equilibrium-constant expression shown below is for ethanoic acid.

$$K_{eq} = \frac{[H_3O^+] \times [CH_3COO^-]}{[CH_3COOH] \times [H_2O]}$$

For dilute aqueous solutions, the concentration of water is a constant. This constant can be combined with K_{eq} to give an acid dissociation constant. An **acid dissociation constant (K_a)** is the ratio of the concentration of the dissociated form of an acid to the concentration of the undissociated form. The dissociated form includes both the H_3O^+ and the anion.

$$K_{eq} \times [H_2O] = K_a = \frac{[H_3O^+] \times [CH_3COO^-]}{[CH_3COOH]}$$

Table 19.7

Dissociation Constants of Weak Acids		
Acid	**Chemical equation for dissociation**	**K_a (25°C)**
Oxalic acid	$HOOCCOOH(aq) \rightleftharpoons H^+(aq) + HOOCCOO^-(aq)$	5.6×10^{-2}
	$HOOCCOO^-(aq) \rightleftharpoons H^+(aq) + OOCCOO^{2-}(aq)$	5.1×10^{-5}
Phosphoric acid	$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq)$	7.5×10^{-3}
	$H_2PO_4^-(aq) \rightleftharpoons H^+(aq) + HPO_4^{2-}(aq)$	6.2×10^{-8}
	$HPO_4^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4^{3-}(aq)$	4.8×10^{-13}
Methanoic acid	$HCOOH(aq) \rightleftharpoons H^+(aq) + HCOO^-(aq)$	1.8×10^{-4}
Benzoic acid	$C_6H_5COOH(aq) \rightleftharpoons H^+(aq) + C_6H_5COO^-(aq)$	6.3×10^{-5}
Ethanoic acid	$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$	1.8×10^{-5}
Carbonic acid	$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$	4.3×10^{-7}
	$HCO_3^-\ (aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq)$	4.8×10^{-11}

The acid dissociation constant (K_a) reflects the fraction of an acid that is ionized. For this reason, dissociation constants are sometimes called ionization constants. If the degree of dissociation or ionization of the acid in a solution is small, the value of the dissociation constant will be small. Weak acids have small K_a values. If the degree of ionization of an acid is more complete, the value of K_a will be larger. The stronger an acid is, the larger its K_a value will be. For example, nitrous acid (HNO_2) has a K_a of 4.4×10^{-4}, but ethanoic acid (CH_3COOH) has a K_a of 1.8×10^{-5}. This means that nitrous acid is more ionized in solution than ethanoic acid. Therefore, nitrous acid is a stronger acid than ethanoic acid.

Table 19.7 shows the ionization equations and dissociation constants of a few weak acids. Some of the acids have more than one dissociation constant because they have more than one ionizable hydrogen. Oxalic acid, for example, is a diprotic acid. It loses two hydrogens, one at a time. Therefore, it has two dissociation constants. Oxalic acid is found naturally in certain herbs and vegetables, such as those pictured in Figure 19.12.

The acids in Table 19.7 are ranked by the value of the first dissociation constant. Observe what happens to the K_a with each ionization. The K_a decreases from the first ionization to the second. It decreases again from the second ionization to the third.

Figure 19.12 Oxalic Acid
Chives and parsley have relatively high amounts of oxalic acid compared to other fruits and vegetables.

Calculating Dissociation Constants To calculate the acid dissociation constant (K_a) of a weak acid, you need to know the initial molar concentration of the acid and the $[H^+]$ (or alternatively, the pH) of the solution at equilibrium. You can use these data to find the equilibrium concentrations of the acid and the ions. These values are then substituted into the expression for K_a.

In general, you can find the K_a of an acid in water by substituting the equilibrium concentrations of the acid, [HA], the anion from the dissociation of the acid, $[A^-]$, and the hydrogen ion, $[H^+]$, into the equation below.

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Sample Problem 19.6

Calculating a Dissociation Constant

In a 0.1000*M* solution of ethanoic acid, $[H^+] = 1.34 \times 10^{-3}M$. Calculate K_a of this acid. Refer to Table 19.7 for the ionization equation for ethanoic acid.

1 Analyze **List the knowns and the unknown.**

KNOWNS	UNKNOWN
[ethanoic acid] = 0.1000*M*	$K_a = ?$
$[H^+] = 1.34 \times 10^{-3}M$	

2 Calculate **Solve for the unknown.**

Start by determining the equilibrium concentration of the ions.

$$[H^+] = [CH_3COO^-] = 1.34 \times 10^{-3}M$$

Each molecule of CH_3COOH that ionizes gives an H^+ ion and a CH_3COO^- ion.

Then determine the equilibrium concentrations of each component.

$$(0.1000 - 0.00134)M = 0.0987M$$

Concentration	$[CH_3COOH]$	$[H^+]$	$[CH_3COO^-]$
Initial	0.1000	0	0
Change	-1.34×10^{-3}	1.34×10^{-3}	1.34×10^{-3}
Equilibrium	0.0987	1.34×10^{-3}	1.34×10^{-3}

Substitute the equilibrium values into the expression for K_a.

$$K_a = \frac{[H^+] \times [CH_3COO^-]}{[CH_3COOH]} = \frac{(1.34 \times 10^{-3}) \times (1.34 \times 10^{-3})}{0.0987}$$

$$= 1.82 \times 10^{-5}$$

3 Evaluate **Does the result make sense?** The calculated value of K_a is consistent with that of a weak acid.

25. In a 0.1000*M* solution of methanoic acid, $[H^+] = 4.2 \times 10^{-3}M$. Calculate the K_a of this acid.

26. In a 0.2000*M* solution of a monoprotic weak acid, $[H^+] = 9.86 \times 10^{-4}M$. What is the K_a for this acid?

Figure 19.13 Use of Ammonia
Window cleaners often use a solution of ammonia in water to clean glass.
Explain ***Why is this solution relatively safe to use?***

Base Dissociation Constant Just as there are strong acids and weak acids, there are strong bases and weak bases. A **strong base** dissociates completely into metal ions and hydroxide ions in aqueous solution. Some strong bases, such as calcium hydroxide and magnesium hydroxide, are not very soluble in water. The small amounts of these bases that dissolve in water dissociate completely.

A **weak base** reacts with water to form the conjugate acid of the base and hydroxide ions. For a weak base, the amount of dissociation is relatively small. Ammonia is an example of a weak base. One use for an aqueous solution of ammonia is shown in Figure 19.13.

$$\underset{\text{Ammonia}}{NH_3(aq)} + \underset{\text{Water}}{H_2O(l)} \rightleftharpoons \underset{\text{Ammonium ion}}{NH_4^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

When equilibrium is established, only about 1 percent of the ammonia is present as NH_4^+. This ion is the conjugate acid of NH_3. The concentrations of NH_4^+ and OH^- are low and equal. The equilibrium-constant expression for the dissociation of ammonia in water is as follows:

$$K_{eq} = \frac{[NH_4^+] \times [OH^-]}{[NH_3] \times [H_2O]}$$

Recall that the concentration of water is constant in dilute solutions. This constant can be combined with the K_{eq} for ammonia to give a base dissociation constant (K_b) for ammonia.

$$K_{eq} \times [H_2O] = K_b = \frac{[NH_4^+] \times [OH^-]}{[NH_3]}$$

In general, the **base dissociation constant (K_b)** is the ratio of the concentration of the conjugate acid times the concentration of the hydroxide ion to the concentration of the base. The general form of the expression for the base dissociation constant is shown below.

$$K_b = \frac{[\text{conjugate acid}] \times [OH^-]}{[\text{base}]}$$

You can use this equation to calculate the K_b of a weak base. You need to know the initial concentration of the base and the concentration of hydroxide ions at equilibrium. If you know the pH, you can calculate $[H^+]$ and the corresponding $[OH^-]$.

The magnitude of K_b indicates the ability of a weak base to compete with the very strong base OH^- for hydrogen ions. Because bases such as ammonia are weak relative to the hydroxide ion, the K_b for such a base is usually small. The K_b for ammonia is 1.8×10^{-5}. The smaller the value of K_b, the weaker the base.

Table 19.8

Comparing Concentration and Strength of Acids

Acidic solution	Concentration: Quantitative (or Molar)	Concentration: Relative	Strength
Hydrochloric acid	12*M* HCl	Concentrated	Strong
Gastric juice	0.08*M* HCl	Dilute	Strong
Ethanoic acid	17*M* CH_3COOH	Concentrated	Weak
Vinegar	0.2*M* CH_3COOH	Dilute	Weak

CHEMISTRY & YOU

Q: *Despite its relatively low pH, lemon juice is safe to consume because citric acid is a weak acid. Citric acid has three K_a values. What does this information tell you about citric acid?*

Concentration Versus Strength Sometimes people confuse the concepts of concentration and strength. The words *concentrated* and *dilute* indicate how much of an acid or base is dissolved in solution. These terms refer to the number of moles of the acid or base in a given volume. The words *strong* and *weak* refer to the extent of ionization or dissociation of an acid or base.

Table 19.8 shows four possible combinations of concentration and strength for acids. Hydrochloric acid, HCl(*aq*), is a strong acid because it completely dissociates into ions. The gastric juice in your stomach is a dilute solution of HCl. The relatively small number of HCl molecules in a given volume of gastric juice are all dissociated into ions. To summarize, even when concentrated hydrochloric acid is diluted with water, it is still a strong acid. Conversely, ethanoic acid (acetic acid) is a weak acid because it ionizes only slightly in solution. Vinegar is a dilute solution of ethanoic acid. Even at a high concentration, ethanoic acid is still a weak acid.

The same concepts apply to bases. A solution of ammonia can be either dilute or concentrated. However, in any solution of ammonia, the relative amount of ionization will be small. Thus, ammonia is a weak base at any concentration. Likewise, sodium hydroxide is a strong base at any concentration.

ONLINE PROBLEMS

19.3 LessonCheck

27. Review What factor determines whether an acid or base is strong or weak?

28. Compare How do acid dissociation constants vary between strong acids and weak acids?

29. Draw Conclusions Which of the acids in Table 19.6 would you expect to have the lowest dissociation constant?

30. Describe How do you determine the K_a of a weak acid or the K_b of a weak base?

31. Predict Acid HX has a very small K_a. How will the relative amounts of H^+ and HX compare at equilibrium?

32. Calculate A 0.500*M* solution of an acid has a hydronium-ion concentration of $5.77 \times 10^{-6}M$. Calculate the K_a of this acid.

33. Describe Write a chemical equation for the dissociation of each of the following acids and bases in water.

a. nitric acid

b. ethanoic acid

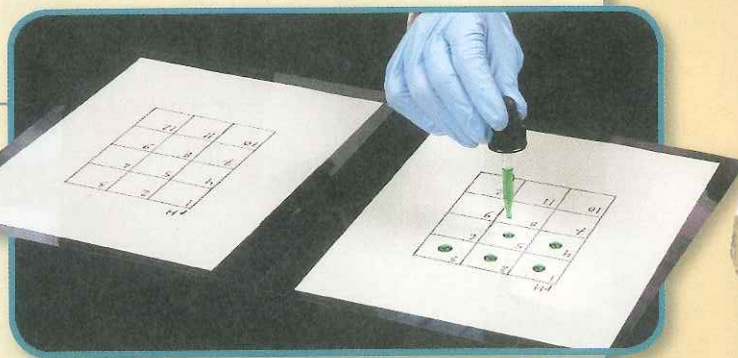
c. ammonia

d. magnesium hydroxide

34. Classify A 15*M* solution of an acid has a K_a of 7.5×10^{-3}. Explain how you would classify this solution in terms of concentration and strength.

Small-Scale Lab

Dissociation Constants of Weak Acids



Purpose

To measure dissociation constants of weak acids

Materials

- paper, pencil, and ruler
- reaction surface
- 12 solutions with different pH values
- bromocresol green
- solutions of other acid-base indicators

Procedure

1. On two separate sheets of paper, draw two grids similar to the one below. Make each square 2 cm on each side.

2. Place a reaction surface over one of the grids and place one drop of bromocresol green in each square.

3. Add one drop of the solution with a pH of 1 to the square labeled 1. Add one drop of the solution with a pH of 2 to the square labeled 2. Continue adding drops in this manner until you have added a drop to each square.

4. Use the second grid as a data table to record your observations for each square.

pH

1	2	3
4	5	6
7	8	9
10	11	12

Analyze and Conclude

1. Observe What colors are the solutions with the lowest pH and the highest pH?

2. Observe At which pH does the bromocresol green change from one color to the other?

3. Infer Acid-base indicators, such as bromocresol green, are usually weak acids. Because bromocresol green has a fairly complex formula, it is convenient to represent its formula as HBCG. HBCG dissociates in water according to the following equation. HBCG and BCG^- are a conjugate acid-base pair.

$$\underset{\text{(yellow)}}{HBCG(aq)} + H_2O(l) \rightleftharpoons \underset{\text{(blue)}}{BCG^-(aq)} + H_3O^+(aq)$$

The K_a expression is

$$K_a = \frac{[BCG^-] \times [H_3O^+]}{[HBCG]}$$

When $[BCG^-] = [HBCG]$, $K_a = [H_3O^+]$.

What color is the conjugate base of HBCG? What color is the conjugate acid of BCG^-?

4. Draw Conclusions At what pH is there an equal amount of the conjugate acid and conjugate base? How can you tell?

5. Calculate What is the K_a for the solution described in Question 4?

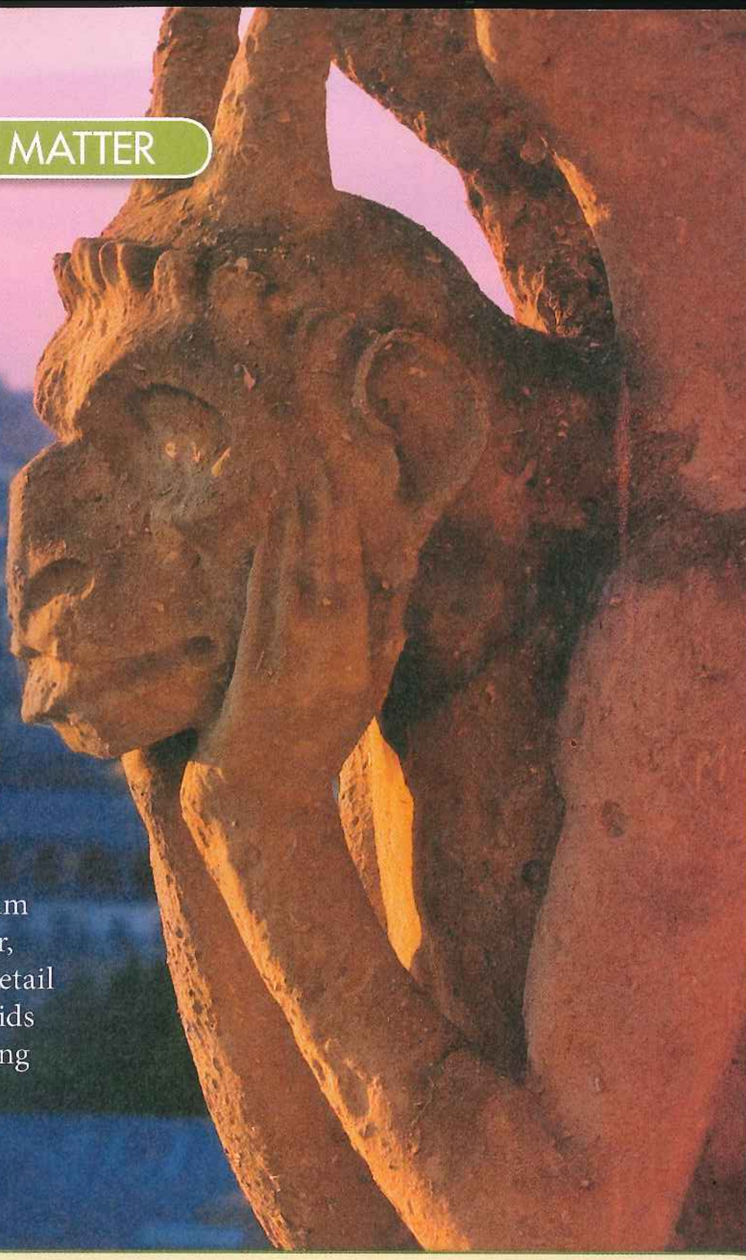
You're the Chemist

1. Design an Experiment Design and carry out an experiment to measure the dissociation constants of some other acid-base indicators. Record the color of each conjugate acid and conjugate base. Calculate the K_a for each acid.

2. Explain How can you measure the dissociation constant of an acid-base indicator? Describe what to do and how to interpret the results.

CHEMISTRY & YOU: EVERYDAY MATTER

Stone Erosion



All around the world, famous stone structures such as the Parthenon in Greece, the Taj Mahal in India, the Mayan carvings in Mexico, and the gargoyles on the facade of the Notre Dame Cathedral are slowly being eroded by acid rain.

Acid rain is rain with a pH of 5.0 or less. It forms when airborne pollutants, namely sulfur dioxide and nitrogen oxides, combine with water vapor in the atmosphere to produce acids.

Many famous buildings, statues, and landmarks are made of marble or limestone. Both of these materials consist of calcium carbonate, $CaCO_3$. The acids in acid rain react with the calcium carbonate in the stones to form calcium ions, water, and carbon dioxide. In this process, the intricate detail of the structure is lost. Explore for yourself how acids affect materials made of calcium carbonate by trying the activity below.

On Your Own

1. For this activity you will need **2 bowls, 2 same-size pieces of chalk made of calcium carbonate, white vinegar, tap water, masking tape, a permanent marker,** and **a paper towel.** Put vinegar in the first bowl. Label the bowl with a piece of masking tape and a marker. Put the same amount of water in the second bowl and label it. *Optional*: If you'd like, you can scratch a design (with a needle, nail, or thumb tack) into each piece of chalk before going to Step 2.

2. Place a piece of chalk in each bowl. Observe what happens to the chalk.

3. After about 5 minutes, remove the chalk pieces from the bowls and place the pieces on a paper towel. Compare the pieces of chalk.

Think About It

1. Compare How did the vinegar affect the chalk compared to the water?

2. Infer What causes the bubbles you see when the chalk is placed in vinegar?

3. Describe Write a balanced equation to explain what happens to the chalk when it is placed in vinegar. *Note*: Let $H^+(aq)$ represent the acid.

4. Draw Conclusions Why does acid rain result in a loss of detail in the gargoyle above? What would be the effect of acid rain on a statue over a long period of time?

19.4 Neutralization Reactions

CHEMISTRY & YOU

Q: *What could cause leaves to turn yellow during the growing season?* You may have noticed yellow leaves like these during a season when the leaves should still be green. This condition is called *chlorosis* because the plant lacks a pigment called chlorophyll. To produce chlorophyll, plants need to absorb nutrients, such as iron, from the soil. Sometimes there is plenty of iron, but it is not taken up by the roots of the plant because the pH of the soil is too high.

Key Questions

What products form when an acid and a base react?

At what point in a titration does neutralization occur?

Vocabulary

- neutralization reaction
- titration
- standard solution
- equivalence point
- end point

Acid-Base Reactions

What products form when an acid and a base react?

Suppose you mix a solution of a strong acid, such as HCl, with a solution of a strong base, such as NaOH. The products are sodium chloride and water.

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

In general, acids and bases react to produce a salt and water. The complete reaction of a strong acid and a strong base produces a neutral solution. Thus, this type of reaction is called a **neutralization reaction.**

When you hear the word *salt,* you may think of the substance that is used to flavor food. Table salt (NaCl) is only one example of a salt. Salts are ionic compounds consisting of an anion from an acid and a cation from a base.

A reaction between an acid and base will go to completion when the solutions contain equal numbers of hydrogen ions and hydroxide ions. The balanced equation provides the correct ratio of acid to base. For hydrochloric acid and sodium hydroxide, the mole ratio is 1:1.

$$\underset{\text{1 mol}}{HCl(aq)} + \underset{\text{1 mol}}{NaOH(aq)} \longrightarrow \underset{\text{1 mol}}{NaCl(aq)} + \underset{\text{1 mol}}{H_2O(l)}$$

For sulfuric acid and sodium hydroxide, the ratio is 1:2. Two moles of the base are required to neutralize one mole of the acid.

$$\underset{\text{1 mol}}{H_2SO_4(aq)} + \underset{\text{2 mol}}{2NaOH(aq)} \longrightarrow \underset{\text{1 mol}}{Na_2SO_4(aq)} + \underset{\text{2 mol}}{2H_2O(l)}$$

Similarly, hydrochloric acid and calcium hydroxide react in a 2:1 ratio.

$$\underset{\text{2 mol}}{2HCl(aq)} + \underset{\text{1 mol}}{Ca(OH)_2(aq)} \longrightarrow \underset{\text{1 mol}}{CaCl_2(aq)} + \underset{\text{2 mol}}{2H_2O(l)}$$

CONCEPTS IN ACTION *Learn more about* **acid-base reactions** *online.*

Sample Problem 19.7

Finding the Moles Needed for Neutralization

The term *neutralization* is used to describe both the reaction and the point at which a neutralization reaction is complete. How many moles of sulfuric acid are required to neutralize 0.50 mol of sodium hydroxide? The equation for the reaction is

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

1 Analyze List the knowns and the unknown. To determine the number of moles of acid, you need to know the number of moles of base and the mole ratio of acid to base.

KNOWNS	UNKNOWN
mol NaOH = 0.50 mol	mol H_2SO_4 = ? mol
1 mol H_2SO_4/2 mol NaOH (from balanced equation)	

2 Calculate Solve for the unknown.

Use the mole ratio of acid to base to determine the number of moles of acid.

$$0.50\ \cancel{\text{mol NaOH}} \times \frac{1\ \text{mol}\ H_2SO_4}{2\ \cancel{\text{mol NaOH}}} = 0.25\ \text{mol}\ H_2SO_4$$

3 Evaluate Does the result make sense? Because the mole ratio of H_2SO_4 to NaOH is 1:2, the number of moles of H_2SO_4 should be half the number of moles of NaOH.

35. How many moles of potassium hydroxide are needed to neutralize 1.56 mol of phosphoric acid?

36. How many moles of sodium hydroxide are required to neutralize 0.20 mol of nitric acid?

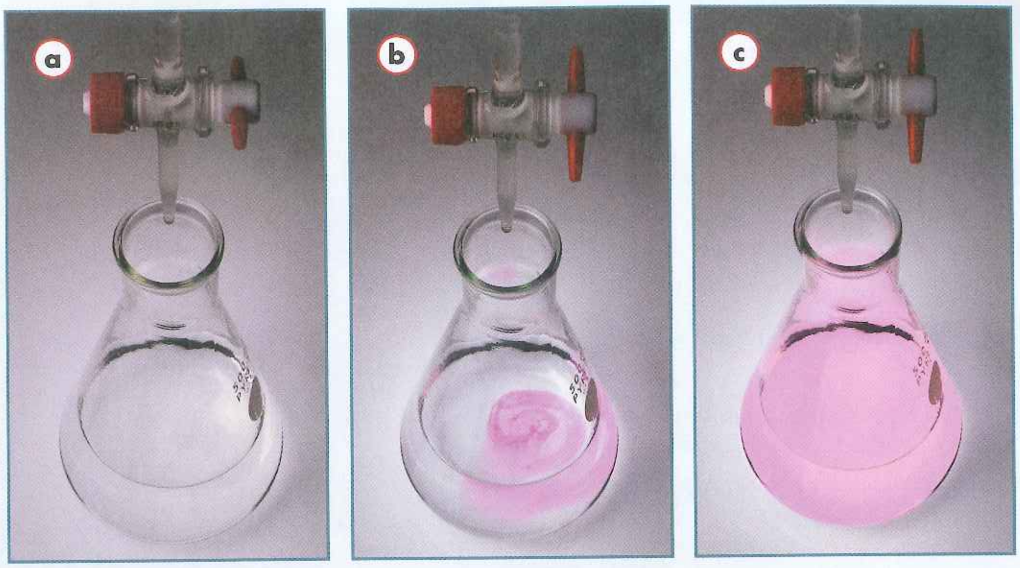
Titration

At what point in a titration does neutralization occur?

You can use a neutralization reaction to determine the concentration of an acid or base. The process of adding a measured amount of a solution of known concentration to a solution of unknown concentration is called a **titration.** The steps in an acid-base titration are as follows.

1. A measured volume of an acid solution of unknown concentration is added to a flask.
2. Several drops of an indicator are added to the solution while the flask is gently swirled.
3. Measured volumes of a base of known concentration are mixed into the acid until the indicator just barely changes color.

Figure 19.14 Titration
The photographs show steps in an acid-base titration. **a.** A flask with a known volume of acid (and some phenolphthalein indicator) is placed beneath a buret that is filled with a base of known concentration. **b.** Base is slowly added from the buret to the acid while the flask is gently swirled. **c.** A change in the color of the solution is the signal that neutralization has occurred.



CHEMISTRY & YOU

Q: *Iron compounds need to dissociate before the iron can be absorbed by plants. However, these compounds become less soluble as the pH rises. For most plants, a pH between 5.0 and 6.5 will provide enough usable iron. How could you change the pH of soil?*

The solution of known concentration is the **standard solution.** The steps in the titration of an acid of unknown concentration with a standard base are shown in Figure 19.14. You can use a similar procedure to find the concentration of a base using a standard acid.

Neutralization occurs when the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions. Two things that are equal in value are said to be *equivalent*. Thus, the point at which neutralization occurs is called the **equivalence point.** The indicator that is chosen for a titration must change color at or near the pH of the equivalence point. The point at which the indicator changes color is the **end point** of the titration.

Figure 19.15 shows how the pH of a solution changes during the titration of a strong acid (HCl) with a strong base (NaOH). The initial acid solution has a low pH (about 1). As NaOH is added, the pH increases because some of the acid reacts with the base. The equivalence point for this reaction occurs at a pH of 7. As the titration nears the equivalence point, the pH rises dramatically because hydrogen ions are being used up. Extending the titration beyond the point of neutralization produces a further increase of pH. If the titration of HCl and NaOH could be stopped right at the equivalence point, the solution in the beaker would consist of only H_2O and NaCl, plus a small amount of indicator.

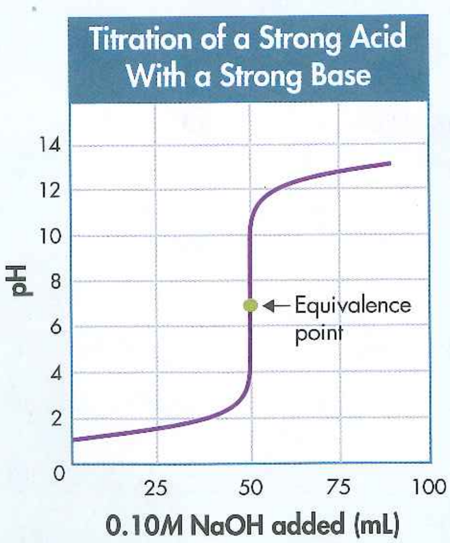


Figure 19.15 Titration Curve
In this titration, 0.10*M* NaOH is slowly added to 50.0 mL of 0.10*M* HCl. The pH of the solution is measured and recorded periodically to construct a titration curve. The equivalence point is located at the midpoint of the vertical part of the curve. Neutralization occurs when 50.0 mL of NaOH have been added to the flask.
Compare *How are [H^+] and [OH^-] related at the equivalence point?*

Sample Problem 19.8

Determining Concentration by Titration

A 25-mL solution of H_2SO_4 is neutralized by 18 mL of 1.0*M* NaOH. What is the concentration of the H_2SO_4 solution? The equation for the reaction is

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

KNOWNS

$[NaOH] = 1.0M$

$V_{NaOH} = 18\ mL = 0.018\ L$

$V_{H_2SO_4} = 25\ mL = 0.025\ L$

UNKNOWN

$[H_2SO_4] = ?M$

1 Analyze **List the knowns and the unknown.** The conversion steps are as follows: L NaOH $\longrightarrow$ mol NaOH $\longrightarrow$ mol H_2SO_4 $\longrightarrow$ *M* H_2SO_4.

2 Calculate **Solve for the unknown.**

Convert volumes to liters because molarity is in moles per liter.

Use the molarity to convert the volume of base to moles of base.

$$0.018\ \cancel{L\ NaOH} \times \frac{1.0\ mol\ NaOH}{1\ \cancel{L\ NaOH}} = 0.018\ mol\ NaOH$$

Use the mole ratio to find the moles of acid.

$$0.018\ \cancel{mol\ NaOH} \times \frac{1\ mol\ H_2SO_4}{2\ \cancel{mol\ NaOH}} = 0.0090\ mol\ H_2SO_4$$

Calculate the molarity by dividing moles of acid by liters of solution.

$$\text{molarity} = \frac{\text{mol of solute}}{\text{L of solution}} = \frac{0.0090\ mol}{0.025\ L} = \boxed{0.36M\ H_2SO_4}$$

3 Evaluate **Does the result make sense?** If the acid had the same molarity as the base (1.0*M*), 50 mL of base would neutralize 25 mL of acid. Because the volume of base is much less than 50 mL, the molarity of the acid must be much less than 1.0*M*.

37. How many milliliters of 0.45*M* HCl will neutralize 25.0 mL of 1.00*M* KOH?

38. What is the molarity of a solution of H_3PO_4 if 15.0 mL is neutralized by 38.5 mL of 0.150*M* NaOH?

19.4 LessonCheck

39. **Review** What are the products of a reaction between an acid and a base?

40. **Explain** Why is the point in the titration when neutralization occurs called the equivalence point?

41. **Calculate** How many moles of HCl are required to neutralize aqueous solutions of these bases?

a. 0.03 mol KOH

b. 2 mol NH_3

c. 0.1 mol $Ca(OH)_2$

42. **Describe** Write complete balanced equations for the following acid-base reactions:

a. $H_2SO_4(aq) + KOH(aq) \longrightarrow$

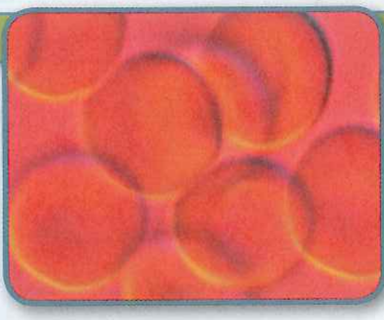
b. $H_3PO_4(aq) + Ca(OH)_2(aq) \longrightarrow$

c. $HNO_3(aq) + Mg(OH)_2(aq) \longrightarrow$

BIGIDEA REACTIONS

43. Review the information on types of chemical reactions in Chapter 11. Which types of reactions are neutralization reactions? Explain your answer.

19.5 Salts in Solution



CHEMISTRY & YOU

Q: *How is the pH of blood controlled in the human body?* Chemical reactions in cells are very sensitive to slight changes in pH. For example, the pH of human blood needs to be kept close to 7.4. A person cannot survive for more than a few minutes if the pH of blood drops below 6.8 or rises above 7.8. This lesson will explain the process that prevents such a life-threatening event.

Key Questions

When is the solution of a salt acidic or basic?

What are the components of a buffer?

Vocabulary

- salt hydrolysis
- buffer
- buffer capacity

Salt Hydrolysis

When is the solution of a salt acidic or basic?

Recall that a salt is one of the products of a neutralization reaction. A salt consists of an anion from an acid and a cation from a base. The solutions of many salts are neutral. Salts that form neutral solutions include sodium chloride and potassium sulfate. Some salts form acidic or basic solutions, as shown in Figure 19.16. The indicator used for Figure 19.16 is called universal indicator because it can be used for a wide range of pH values.

Figure 19.17 shows two titration curves. One curve is for the addition of sodium hydroxide, a strong base, to ethanoic acid, a weak acid. An aqueous solution of sodium ethanoate exists at the equivalence point.

$$\underset{\text{Ethanoic acid}}{CH_3COOH(aq)} + \underset{\text{Sodium hydroxide}}{NaOH(aq)} \longrightarrow \underset{\text{Sodium ethanoate}}{CH_3COONa(aq)} + \underset{\text{Water}}{H_2O(l)}$$

The second titration curve is for the reaction between hydrochloric acid, which is a strong acid, and sodium hydroxide. This second curve should look familiar. It first appeared in Lesson 19.4 in the section on titrations.

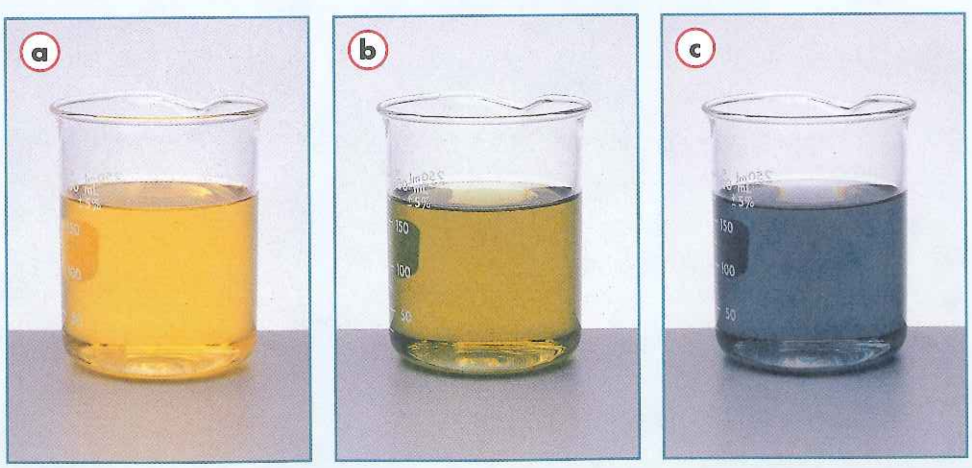
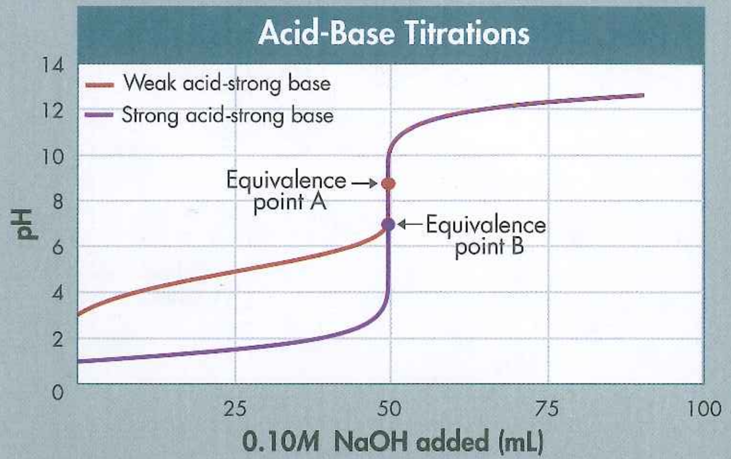


Figure 19.16
The pH of Salt Solutions
Universal indicator was added to 0.10*M* aqueous salt solutions. Based on the indicator color, the solutions can be classified as follows: **a.** Ammonium chloride, $NH_4Cl(aq)$, is acidic (pH of about 5.3). **b.** Sodium chloride, $NaCl(aq)$, is neutral (pH of 7). **c.** Sodium ethanoate, $CH_3COONa(aq)$, is basic (pH of about 8.7).

Interpret Graphs



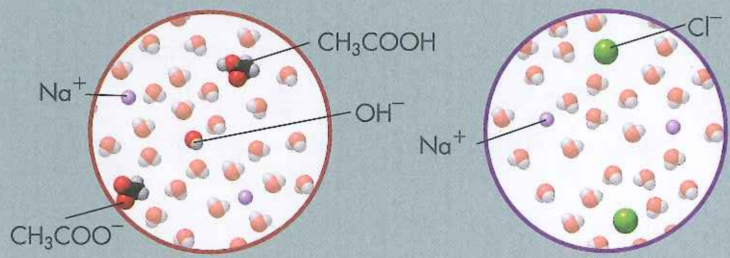


Figure 19.17 The graph compares the titration curves for a weak acid and a strong base with that for a strong acid and a strong base.

a. Read Graphs What is the pH of the equivalence point for each titration?

b. Interpret Diagrams What ions are present in each solution at the equivalence point?

c. Explain Why are the equivalence points of the two titrations different?

Hint: To answer part c, consider which reaction occurs in one solution that does not occur in the other.

The pH at the equivalence point for the weak acid-strong base titration is basic. For a strong acid-strong base titration, the pH at the equivalence point is neutral. This difference in pH exists because hydrolysis occurs with some salts in solution. In **salt hydrolysis,** the cations or anions of a dissociated salt remove hydrogen ions from, or donate hydrogen ions to, water. **Salts that produce acidic solutions have positive ions that release hydrogen ions to water. Salts that produce basic solutions have negative ions that attract hydrogen ions from water.**

Sodium ethanoate (CH_3COONa) is the salt of a weak acid and a strong base. In solution, the salt is completely ionized.

$$\underset{\text{Sodium ethanoate}}{CH_3COONa(aq)} \longrightarrow \underset{\text{Ethanoate ion}}{CH_3COO^-(aq)} + \underset{\text{Sodium ion}}{Na^+(aq)}$$

The ethanoate ion is a Brønsted-Lowry base, which means it is a hydrogen-ion acceptor. It reacts with water to form ethanoic acid and hydroxide ions. At equilibrium, the reactants are favored.

$$\underset{\substack{\text{H}^+\text{ acceptor,}\\ \text{Brønsted-Lowry base}}}{CH_3COO^-(aq)} + \underset{\substack{\text{H}^+\text{ donor,}\\ \text{Brønsted-Lowry acid}}}{H_2O(l)} \rightleftharpoons CH_3COOH(aq) + \underset{\substack{\text{(makes the}\\ \text{solution basic)}}}{OH^-(aq)}$$

This process is called hydrolysis because a hydrogen ion is split off a water molecule. The suffix *-lysis* comes from a Greek word meaning to "separate" or "loosen." In the solution, the hydroxide-ion concentration is greater than the hydrogen-ion concentration. Thus, the solution is basic.

Ammonium chloride (NH_4Cl) is the salt of the strong acid hydrochloric acid (HCl) and the weak base ammonia (NH_3). It is completely ionized in solution.

$$NH_4Cl(aq) \longrightarrow NH_4^+(aq) + Cl^-(aq)$$

The ammonium ion (NH_4^+) is a strong enough acid to donate a hydrogen ion to a water molecule. The products are ammonia molecules and hydronium ions. The reactants are favored at equilibrium, as shown by the relative sizes of the forward and reverse arrows.

$$\underset{\substack{H^+\text{ donor,}\\ \text{Brønsted-Lowry acid}}}{NH_4^+(aq)} + \underset{\substack{H^+\text{ acceptor,}\\ \text{Brønsted-Lowry base}}}{H_2O(l)} \rightleftharpoons NH_3(aq) + \underset{\substack{\text{(makes the}\\ \text{solution acidic)}}}{H_3O^+(aq)}$$

This process is another example of hydrolysis. At equilibrium the $[H_3O^+]$ is greater than the $[OH^-]$. Thus, a solution of ammonium chloride is acidic. To determine if a salt will form an acidic or basic solution, remember the following rules:

Strong acid + **Strong base** ⟶ Neutral solution
Strong acid + Weak base ⟶ **Acidic** solution
Weak acid + **Strong base** ⟶ **Basic** solution

Buffers

What are the components of a buffer?

Suppose you add 10 mL of 0.10*M* sodium hydroxide to 1 L of pure water. The pH will increase about 4 pH units—from 7.0 to about 11.0. This change is a relatively large increase in pH. Now consider a solution containing 0.20*M* each of ethanoic acid and sodium ethanoate. This solution has a pH of 4.76. If you add 10 mL of 0.10*M* sodium hydroxide to 1 L of this solution, the pH increases 0.01 pH units—from 4.76 to 4.77. This is a relatively small change in pH. If 10 mL of acid had been added instead of the base, the amount of change in pH would also have been small.

The solution of ethanoic acid and sodium ethanoate is an example of a buffer. A **buffer** is a solution in which the pH remains fairly constant when small amounts of acid or base are added. **A buffer is a solution of a weak acid and one of its salts or a solution of a weak base and one of its salts.** Figure 19.18 compares what happens when 1.0 mL of 0.01*M* HCl solution is added to an unbuffered solution and to a solution with a buffer.

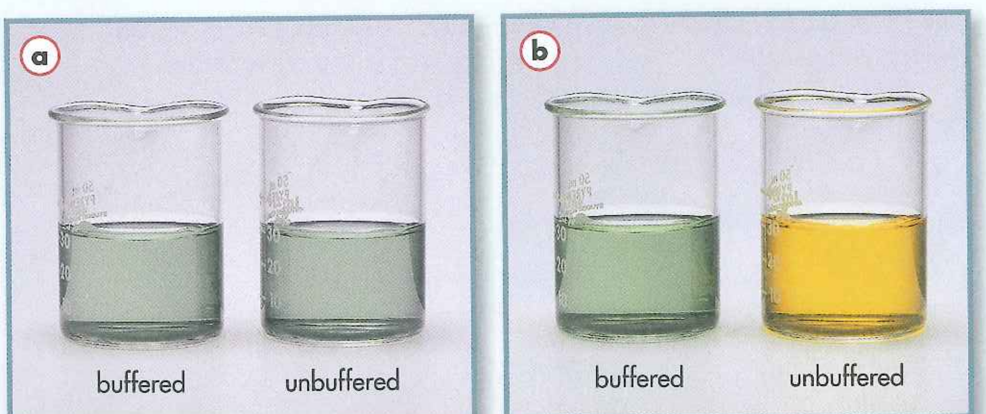


Figure 19.18 Effect of a Buffer In a buffer solution, the pH does not shift dramatically. **a.** The indicator shows that both solutions are basic (pH of about 8). **b.** HCl is added to each solution. The indicator shows no visible pH change in the buffered solution. The color change in the unbuffered solution indicates a change in pH from 8 to about 3. **Predict** *How would the original solutions respond if NaOH were added?*

How Buffers Work A buffer solution is better able to resist drastic changes in pH than is pure water. The reason is fairly simple. A buffer solution contains one component that can react with hydrogen ions (a hydrogen-ion acceptor) and another component that can react with hydroxide ions (a hydrogen-ion donor). These components act as reservoirs of neutralizing power that can be tapped when either hydrogen ions or hydroxide ions are added to the solution.

The ethanoic acid–ethanoate ion buffer can be used to show how a buffer works. When an acid is added to the buffer, the ethanoate ions (CH_3COO^-) act as a hydrogen-ion "sponge." As the ethanoate ions react with the hydrogen ions, they form ethanoic acid. This weak acid does not ionize extensively in water, so the change in pH is very slight.

$$\underset{\text{Ethanoate ion}}{CH_3COO^-(aq)} + \underset{\text{Hydrogen ion}}{H^+(aq)} \rightleftharpoons \underset{\text{Ethanoic acid}}{CH_3COOH(aq)}$$

When hydroxide ions are added to the buffer, the ethanoic acid and the hydroxide ions react to produce water and the ethanoate ion.

$$\underset{\text{Ethanoic acid}}{CH_3COOH(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)} \rightleftharpoons \underset{\text{Ethanoate ion}}{CH_3COO^-(aq)} + \underset{\text{Water}}{H_2O(l)}$$

The ethanoate ion is not a strong enough base to accept hydrogen ions from water to a great extent. Therefore, the reverse reaction is minimal and the change in pH is very slight.

The Capacity of a Buffer Buffer solutions have their limits. As acid is added to an ethanoate buffer, eventually no more ethanoate ions will be present to accept the hydrogen ions. At that point, the buffer can no longer control the pH. The ethanoate buffer also becomes ineffective when too much base is added. In that case, no more ethanoic acid molecules are present to donate hydrogen ions. Adding too much acid or base will exceed the buffer capacity of a solution. The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a significant change in pH occurs.

Table 19.9 lists some common buffer systems. Two of these buffer systems help maintain optimal human blood pH. One is the carbonic acid–hydrogen carbonate buffer system. The other is the dihydrogen phosphate–hydrogen phosphate buffer system.

CHEMISTRY & YOU

Q: *The equilibrium between carbonic acid (H_2CO_3) and hydrogen carbonate ions (HCO_3^-) helps keep the pH of blood within a narrow range (7.35–7.45). If the pH rises, molecules of carbonic acid donate hydrogen ions. What can happen if the pH drops, that is, if the $[H^+]$ increases?*

Table 19.9

Important Buffer Systems

Buffer name	Formulas	Buffer pH*
Ethanoic acid–ethanoate ion	CH_3COOH / CH_3COO^-	4.76
Dihydrogen phosphate ion–hydrogen phosphate ion	$H_2PO_4^- / HPO_4^{2-}$	7.20
Carbonic acid–hydrogen carbonate ion (solution saturated with CO_2)	H_2CO_3 / HCO_3^-	6.46
Ammonium ion–ammonia	NH_4^+ / NH_3	9.25

*Components have concentrations of 0.1*M*.

Sample Problem 19.9

Describing Buffer Systems

Write balanced chemical equations to show how the carbonic acid–hydrogen carbonate buffer can "mop up" added hydroxide ions and hydrogen ions.

1 Analyze **Identify the relevant concepts.** A buffer contains two components: a hydrogen-ion acceptor (which can react with H^+) and a hydrogen-ion donor (which can react with OH^-).

2 Solve **Apply the concepts to this problem.**

Identify the hydrogen-ion acceptor and the hydrogen-ion donor.

H_2CO_3, a weak acid, can release hydrogen ions. HCO_3^- is the conjugate base, which can accept hydrogen ions.

Write the equation for the reaction that occurs when a base is added to the buffer.

When a base is added, the hydroxide ions react with H_2CO_3.

$$H_2CO_3(aq) + (OH^-)(aq) \rightleftharpoons HCO_3^-(aq) + H_2O(l)$$

Write the equation for the reaction that occurs when an acid is added to the buffer.

When an acid is added, the hydrogen ions react with HCO_3^-.

$$HCO_3^-(aq) + (H^+)(aq) \rightleftharpoons H_2CO_3(aq)$$

44. Write equations to show what happens in the following situations:

a. Acid is added to a solution that contains HPO_4^{2-} ions.

b. Base is added to a solution that contains $H_2PO_4^-$ ions.

45. A buffer consists of methanoic acid (HCOOH) and methanoate ion ($HCOO^-$). Write an equation to show what happens when an acid is added to this buffer.

19.5 LessonCheck

46. **Review** What type of salt produces an acidic solution? What type of salt produces a basic solution?

47. **Describe** What types of substances can be combined to make a buffer solution?

48. **Classify** Which of these salts would hydrolyze to produce an acidic aqueous solution, and why?

a. $KC_2H_3O_2$

b. LiCl

c. $NaHCO_3$

d. $(NH_4)_2SO_4$

49. **Identify** Which of the following pairs can form a buffer solution? Explain.

a. NH_3 and HCO_3^-

b. C_6H_5COOH and $C_6H_5COO^-$

50. **Describe** Write a balanced chemical equation to show what happens when an acid is added to an ammonium ion–ammonia buffer. Write an equation to show what happens when a base is added.

51. **Relate Cause and Effect** Use Le Châtelier's principle to explain how a buffer system maintains the pH of a solution.

CHEMISTRY & YOU: GREEN CHEMISTRY

Ocean Buffers

As atmospheric carbon dioxide (CO_2) levels rise due to increased burning of fossil fuels, global warming is not the only potential environmental problem that Earth faces. Additionally, ocean water becomes more acidic, which can disrupt the ocean ecosystem.

The oceans naturally absorb CO_2 from the atmosphere. Some of the absorbed CO_2 is converted to carbonic acid (H_2CO_3), which can lower the pH of ocean water. Fortunately, the oceans have an excellent natural buffer system that helps maintain the optimal pH for supporting ocean life—about 8.2. The ocean's buffer system is largely based on the hydrogen carbonate–carbonate ion buffer system. However, the buffer capacity of the buffers in ocean water is limited, and current human activities are pushing these limits.

Try the following activity at home to gain a better understanding of the effects of CO_2 and sea salt on the pH of ocean water.

On Your Own

1. For this activity, you will need the following materials: **4 drinking glasses, masking tape, a permanent marker, distilled water, carbonated water, sea salt, a measuring cup, ¼ teaspoon measuring spoon,** and **4 pH test strips.** (You can ask your teacher for pH test strips if you do not have them at home.) Use the masking tape and marker to label the glasses 1, 2, 3, and 4.

2. Add ½ cup of distilled water to glasses 1 and 2. Add ½ cup of carbonated water to glasses 3 and 4.

3. Add ¾ teaspoon of sea salt to containers 2 and 4. Stir until the salt is dissolved. (This ratio of sea salt to water is similar to that found in the ocean.)

4. Measure the pH of each solution and record it in a table similar to the one shown to the right.

What Did You Find?

Glass	1	2	3	4
Contents				
pH				

Think About It

1. Compare Determine whether each solution is acidic, basic, or neutral. How do the four solutions differ?

2. Explain How does a solution containing hydrogen carbonate ions (HCO_3^-) and carbonate ions (CO_3^{2-}) act as a buffer? Use chemical equations to support your explanation.

3. Draw Conclusions What does this experiment demonstrate about the effects of dissolved CO_2 (in the carbonated water) and sea salt on the pH of ocean water?

19 Study Guide

BIGIDEA REACTIONS

Chemists define acids and bases according to the ions they yield in aqueous solution. Chemists also define acids and bases based on whether they accept or donate hydrogen ions, and whether they are electron-pair donors or acceptors. The pH of a solution reflects the hydrogen-ion concentration. Chemists use acid-base reactions to determine the concentration of an acid or a base in solution.

19.1 Acid-Base Theories

According to Arrhenius, acids are hydrogen-containing compounds that ionize to yield hydrogen ions in aqueous solution. Bases are compounds that ionize to yield hydroxide ions in aqueous solution.

According to the Brønsted-Lowry theory, an acid is a hydrogen-ion donor and a base is a hydrogen-ion acceptor.

According to Lewis, an acid accepts a pair of electrons and a base donates a pair of electrons.

- hydronium ion (H_3O^+) (647)
- conjugate acid (650)
- conjugate base (650)
- conjugate acid-base pair (650)
- amphoteric (651)
- Lewis acid (651)
- Lewis base (651)

19.2 Hydrogen Ions and Acidity

For aqueous solutions, the product of the hydrogen-ion concentration and the hydroxide-ion concentration equals 1×10^{-14}.

A solution with a pH less than 7.0 is acidic. A solution with a pH of 7 is neutral. A solution with a pH greater than 7.0 is basic.

Either acid-base indicators or pH meters can be used to measure pH.

- self-ionization (653)
- neutral solution (653)
- ion-product constant for water (K_w) (654)
- acidic solution (654)
- basic solution (654)
- pH (656)

Key Equations

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

$$pH = -\log[H^+]$$

19.3 Strengths of Acids and Bases

Acids and bases are classified as strong or weak based on the degree to which they ionize in water.

- strong acid (664)
- weak acid (664)
- acid dissociation constant (K_a) (665)
- strong base (668)
- weak base (668)
- base dissociation constant (K_b) (668)

Key Equation

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

19.4 Neutralization Reactions

In general, acids and bases react to produce a salt and water.

Neutralization occurs when the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

- neutralization reaction (672)
- titration (673)
- standard solution (674)
- equivalence point (674)
- end point (674)

19.5 Salts in Solution

Salts that produce acidic solutions have positive ions that release hydrogen ions to water. Salts that produce basic solutions have negative ions that attract hydrogen ions from water.

A buffer is a solution of a weak acid and one of its salts or a weak base and one of its salts.

- salt hydrolysis (677)
- buffer (678)
- buffer capacity (679)

Math Tune-Up: Acid-Base Problems

Problem	❶ Analyze	❷ Calculate	❸ Evaluate
The pH of an unknown solution is 3.70. What is the hydrogen-ion concentration?	Known: pH = 3.70 Unknown: $[H^+] = ?M$ Use the following equation: $pH = -\log[H^+]$	Rearrange the expression for pH and substitute the known pH value to solve for the unknown: $-\log[H^+] = pH$ $-\log[H^+] = 3.70$ $\log[H^+] = -3.70$ The antilog of −3.70 is 2.0×10^{-4}. Thus, $[H^+] = 2.0 \times 10^{-4}M$.	The pH is between 3 and 4. So, the hydrogen ion concentration must be between $1 \times 10^{-3}M$ and $1 \times 10^{-4}M$. **Note:** To determine the antilog on most calculators, press the 2nd or INV key then the log key.
In a 0.500*M* solution of a weak acid (HA), the $[H^+]$ is 4.02×10^{-3} at equilibrium. Find the K_a for this acid. The acid dissociates as follows: $HA \rightleftharpoons H^+ + A^-$	Knowns: $[HA] = 0.500M$ $[H^+] = 4.02 \times 10^{-3}$ Unknown: $K_a = ?$ Use the general expression for K_a: $K_a = \frac{[H^+] \times [A^-]}{[HA]}$ **Hint:** Review Sample Problem 19.6 for help on finding equilibrium concentrations.	At equilibrium, $[H^+]$ is equal to $[A^-]$: $[H^+] = [A^-] = 4.02 \times 10^{-3}$ Calculate [HA] at equilibrium: $0.500M - 0.00402M = 0.496M$ Substitute the equilibrium concentrations into the equation for K_a and solve: $K_a = \frac{(4.02 \times 10^{-3}) \times (4.02 \times 10^{-3})}{0.496}$ $K_a = 3.26 \times 10^{-5}$	The value of K_a is consistent with that of a weak acid.
How many moles of KOH are needed to neutralize 0.25 mol of H_2SO_4? The equation for the reaction is $2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$	Known: mol H_2SO_4 = 0.25 mol Unknown: mol KOH = ? mol	Use the mole ratio of base to acid (2 mol KOH to 1 mol H_2SO_4) to determine the number of moles of base: $0.25\ \cancel{\text{mol } H_2SO_4} \times \frac{2 \text{ mol KOH}}{1\ \cancel{\text{mol } H_2SO_4}}$ $= 0.50$ mol KOH	The mole ratio of KOH to H_2SO_4 is 2:1. So, the number of moles of KOH should be twice the number of moles of H_2SO_4.

19 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

19.1 Acid-Base Theories

★**52.** How did Arrhenius describe acids and bases?

53. Classify each compound as an Arrhenius acid or an Arrhenius base.

Typo

a. $Ca(OH)_2$ **c.** HNO_3 **e.** HBr
b. ~~C_2H_5COOH~~ CH_3COOH **d.** KOH **f.** H_2SO_4

54. Write an equation for the dissociation of each compound in water.

a. KOH **b.** $Mg(OH)_2$

★**55.** Write balanced equations for the reaction of each metal with water.

a. lithium **b.** barium

56. Identify each reactant in the following equations as a hydrogen-ion donor (acid) or a hydrogen-ion acceptor (base). All the reactions take place in aqueous solution.

a. $HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3^-$
b. $CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$
c. $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
d. $H_2O + CH_3COO^- \rightleftharpoons CH_3COOH + OH^-$

57. Label the conjugate acid-base pairs for each equation in Question 56.

58. What is a Lewis acid? What is a Lewis base?

19.2 Hydrogen Ions and Acidity

59. Write an equation showing the self-ionization of water.

60. What are the concentrations of H^+ and OH^- in pure water at 25°C?

61. How is the pH of a solution calculated?

62. Why is the pH of pure water at 25°C equal to 7.0?

★**63.** Calculate the pH for the following solutions and indicate whether each solution is acidic or basic.

a. $[OH^-] = 1 \times 10^{-2}M$ **b.** $[H^+] = 1 \times 10^{-2}M$

64. What are the hydroxide-ion concentrations for solutions with the following pH values?

a. 4.00 **b.** 8.00 **c.** 12.00

65. Calculate the pH or $[H^+]$ for each solution.

a. $[H^+] = 2.4 \times 10^{-6}M$ **b.** pH = 13.20

19.3 Strengths of Acids and Bases

★**66.** Identify each compound as a strong or weak acid or base.

a. NaOH **b.** NH_3 **c.** H_2SO_4 **d.** HCl

67. Would a strong acid have a large or a small K_a? Explain your answer.

68. Why are $Mg(OH)_2$ and $Ca(OH)_2$ classified as strong bases even though their saturated solutions are only mildly basic?

69. Write the expression for K_a for each acid. Assume only one hydrogen is ionized.

a. HF **b.** H_2CO_3

19.4 Neutralization Reactions

70. Write a general word equation for a neutralization reaction.

71. Identify the products and write balanced equations for each neutralization reaction.

a. $HNO_3(aq) + KOH(aq) \longrightarrow$
b. $HCl(aq) + Ca(OH)_2(aq) \longrightarrow$
c. $H_2SO_4(aq) + NaOH(aq) \longrightarrow$

72. How is it possible to recognize the end point of a titration?

★**73.** What is the molarity of sodium hydroxide if 20.0 mL of the solution is neutralized by each of the following 1.00*M* solutions?

a. 28.0 mL of HCl
b. 17.4 mL of H_3PO_4

19.5 Salts in Solution

74. What kinds of salts hydrolyze water?

★**75.** Write an equation showing why an aqueous solution of sodium hydrogen carbonate is basic.

76. Explain why solutions of salts that hydrolyze water do not have a pH of 7.

77. Predict whether an aqueous solution of each salt will be acidic, basic, or neutral.

a. $NaHCO_3$ **d.** Na_2CO_3
b. NH_4NO_3 **e.** Na_2SO_4
c. KCl **f.** NH_4Cl

78. Explain why a buffered solution cannot absorb an unlimited amount of acid or base.

Understand Concepts

79. Explain how the Lewis theory is a more general classification system than either the Arrhenius descriptions or the Brønsted-Lowry theory.

80. Is it possible to have a concentrated weak acid? Explain.

*81. Write equations showing that the hydrogen phosphate ion (HPO_4^{2-}) is amphoteric.

82. The pH of a 0.5000M HNO_2 solution is 1.83. What is the K_a of this acid?

83. How do the [H^+] and the [OH^-] compare in each type of solution?
    a. neutral solution
    b. basic solution
    c. acidic solution

84. Write the formula and name of the conjugate base of each Brønsted-Lowry acid.
    a. HCO_3^- b. NH_4^+ c. HI d. H_2SO_3

*85. Write the formula and name of the conjugate acid of each Brønsted-Lowry base.
    a. ClO_2^- b. H_2O c. $H_2PO_4^-$ d. NH_3

86. Calculate the [OH^-] or pH of each solution.
    a. pH = 4.60
    b. pH = 9.30
    c. [OH^-] = $1.8 \times 10^{-2}M$
    d. [OH^-] = $7.3 \times 10^{-9}M$

87. Write the three equations for the stepwise ionization of phosphoric acid.

88. Use the Brønsted-Lowry and Lewis definitions of acids and bases to identify each reactant as an acid or a base.
    a. $KOH(aq) + HBr(aq) \longrightarrow KBr(aq) + H_2O(l)$
    b. $HCl(aq) + H_2O(l) \longrightarrow Cl^-(aq) + H_3O^+(aq)$

89. Write the formula for the conjugate base of each of the following acids:
    a. H_2SO_4 b. CH_3COOH c. H_2O

90. Use the phosphate buffer ($H_2PO_4^-/HPO_4^{2-}$) to illustrate how a buffer system works. Use equations to show how the pH of a solution can be kept almost constant when small amounts of acid or base are added.

*91. Write an equation for the reaction of each antacid with hydrochloric acid.
    a. magnesium hydroxide
    b. calcium carbonate
    c. aluminum hydroxide

92. How would the addition of each substance affect the equilibrium between hypochlorous acid and the hypochlorite ion?

$$HOCl(aq) + OH^-(aq) \rightleftharpoons OCl^-(aq) + H_2O(l)$$

    a. HCl b. NaOH

93. The following data were collected from a titration of 50.00 mL of ethanoic acid (CH_3COOH) of unknown concentration with 0.100M NaOH. Plot these data to obtain a titration curve. Place pH on the y-axis.

Volume of NaOH (mL)	pH	Volume of NaOH (mL)	pH
0	3.18	50.00	8.73
10.00	4.15	50.01	8.89
25.00	4.76	51.00	11.00
40.00	5.36	60.00	11.96
49.00	6.45	75.00	12.30
49.99	8.55	100.00	12.52

    a. What is the pH at the end point of this titration?
    b. Use Figure 19.8 to identify one or more acid-base indicators that could be used to determine the end point in this titration.

94. Write an equation to show that an aqueous solution of sodium ethanoate will be basic.

*95. Arrange the following solutions in order of decreasing acidity:
    a. 0.1M NaOH
    b. 0.1M HCl
    c. 0.1M NH_4Cl
    d. 0.1M CH_3COONa

96. Vapors of the strong acid HCl(aq) and the weak base NH_3(aq) combine to form a white salt.

    a. What is the name and formula of the salt?
    b. What kind of solution will this salt form when it dissolves in water?

Think Critically

97. Compare Arrhenius, Brønsted-Lowry, and Lewis all offered explanations for the behavior of acids and bases.

a. Which explanation is easiest for you to understand?
b. How is it possible for all three explanations to be accepted by chemists?

98. Predict The solubility of carbon dioxide in water depends on four different reversible reactions.

$$CO_2(g) \rightleftharpoons CO_2(aq)$$
$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$
$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$
$$HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq)$$

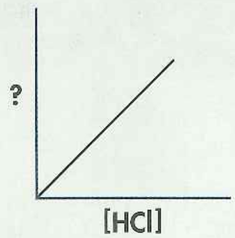
If seawater is slightly alkaline, would you expect the concentration of dissolved CO_2 to be higher or lower than in pure water? Explain your answer.

99. Evaluate Critique the accuracy of each of these statements.

a. Indicators such as methyl red provide accurate and precise measurements of pH.
b. According to the Arrhenius definition of acids and bases, ammonia qualifies as a base.
c. The strength of an acid or base changes as its concentration changes.

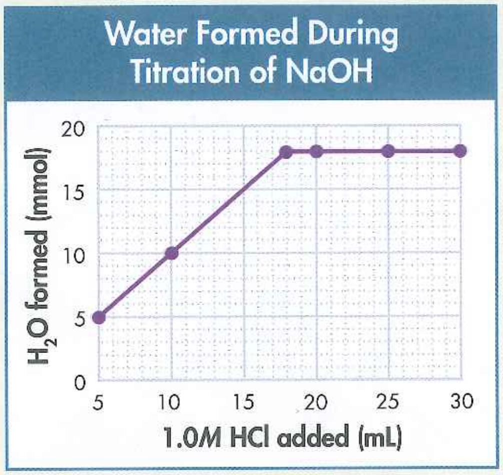
★100. Relate Cause and Effect Use the cyanate buffer $HOCN/OCN^-$ to explain how a buffer system works. Use equations to show how the pH of a solution can be kept almost constant when small amounts of acid or base are added.

★101. Identify Which quantity might correspond to the y-axis on this graph: $[H^+]$, pH, or $[OH^-]$? Explain your answer.



102. Calculate The sugar substitute saccharin ($HNC_7H_4SO_3$) has one acidic hydrogen. A 1.000M aqueous solution of saccharin has a pH of 1.71. Calculate the K_a of saccharin.

103. Interpret Graphs The graph shows the number of millimoles (mmol) of water formed as drops of 1.0M HCl are added to a 25.0-mL sample of NaOH of unknown concentration.



a. Write an equation for the reaction.
b. Estimate the concentration of the NaOH.

104. Calculate Suppose you slowly add 0.1M NaOH to 50.0 mL of 0.1M HCl. What volume of NaOH must you add before neutralization will occur? Explain your reasoning.

105. Predict Will the resulting solutions be neutral, acidic, or basic at the equivalence point for each of the following titrations? Explain.

a. HCl titrated with NaOH
b. NaOH titrated with HCl
c. CH_3COOH titrated with NaOH
d. NH_3 titrated with HCl
e. CH_3COOH titrated with NH_3

★106. Use Models You can use the following expression to find the pH of a solution:

$$pH = -\log[H^+]$$

What expression could you use to find the analogous quantity, the pOH of a solution?

107. Apply Concepts Milk, an aqueous emulsion, has a pH of about 6.7. Calculate the pOH of milk using the equation you derived in Question 106. If the equation you derived is correct, the sum of the values for pH and pOH will equal 14.

108. Apply Concepts Use the expression for K_w to demonstrate the following relationship:

$$pH + pOH = 14$$

Enrichment

109. Calculate What is the pH of a $0.010M$ solution of NaCN ($K_b = 2.1 \times 10^{-5}$)?

***110. Make Generalizations** Show that for any conjugate acid-base pair, $K_a \times K_b = K_w$.

111. Interpret Data The K_w of water varies with temperature, as shown in the table.

Temperature (°C)	K_w	pH
0	1.137×10^{-15}	a. ______
10	2.917×10^{-15}	b. ______
20	6.807×10^{-15}	c. ______
30	1.469×10^{-14}	d. ______
40	2.917×10^{-14}	e. ______
50	5.470×10^{-14}	f. ______

a. Calculate the pH of water for each temperature in the table.
b. Use the data to prepare a graph of pH versus temperature. Use the graph to estimate the pH of water at 5°C.
c. At what temperature is the pH of water approximately 6.85?

112. Calculate What is the molarity of an H_2SO_4 solution if 80.0 mL of the solution reacts with 0.424 g Na_2CO_3?

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

***113. Apply Concepts** The hydrogen carbonate ion–carbonic acid buffer system is an important buffer system in the blood. This system is represented by the following equations:

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

Reactions in cells produce carbon dioxide. Excess carbon dioxide is released through the lungs. How could rapid breathing lead to an abnormally high blood pH (alkalosis)? How could slow breathing lead to an abnormally low blood pH (acidosis)?

114. Calculate Household bleach is a solution of sodium hypochlorite. What is the $[OH^-]$ in an aqueous solution that is 5.0% NaClO by mass? What is the pH of the solution? (The density of the solution is 1.0 g/mL, and $K_a = 3.5 \times 10^{-8}$.)

Write About Science

115. Research The main cause of tooth decay is the weak acid lactic acid (C_2H_5OCOOH). Lactic acid forms when bacteria, such as *Streptococcus mutans,* feed on sugar. In the mouth, sugars are present in the sticky plaque on tooth surfaces. Starting with the information on page R30, research current efforts to fight tooth decay. Write a report summarizing your findings.

116. Connect to the BIGIDEA Hypochlorite salts are used to disinfect swimming pools. On page R30 in the Elements Handbook, read about what happens when chlorine compounds are added to pool water. Use hydrolysis reactions to explain how the pH of the water affects the concentration of hypochlorous acid (HOCl).

CHEMYSTERY

Paper Trail

The wood pulp used to make paper is a suspension of cellulose fibers in water. Wood chips can be ground into a pulp. This is the process used to make newsprint. For higher-quality paper, the pulp is treated chemically to remove parts of the wood other than cellulose.

The paper is often coated with a chemical such as aluminum sulfate to keep it from absorbing too much ink. The chemicals that were used for this purpose often left a residue of acid in the paper. Over time, the acid caused the cellulose fibers to decay.

117. Infer The process of treating the paper is called deacidification. The first step in a popular deacidification method is to immerse the paper in a dilute solution of calcium hydroxide. Write a chemical equation to describe what occurs in this deacidification step.

***118. Connect to the BIGIDEA** What type of reaction is performed in the process of deacidification? Would you expect the pH of the paper to be raised or lowered in the process of deacidification?

Cumulative Review

119. Write the product of each of these combination reactions.

a. $K(s) + O_2(g) \longrightarrow$
b. $Ca(s) + S(s) \longrightarrow$
c. $F_2(g) + Al(s) \longrightarrow$

120. How many grams of oxygen are needed to completely burn 87.4 g of sulfur to form sulfur trioxide?

$$S(s) + O_2(g) \longrightarrow SO_3(g)$$

***121.** Which state of matter is not part of the process of sublimation?

122. State Dalton's law of partial pressures.

123. Which of these laws describes an inverse relationship?

a. Charles's law
b. Boyle's law
c. Gay-Lussac's law

124. Which has the largest particles, a solution, a colloid, or a suspension?

125. Which of these is not an electrolyte?

a. $NaCl(l)$
b. $KNO_2(aq)$
c. $SiO_2(s)$
d. $NaCl(aq)$

***126.** What type of bond is responsible for water's high surface tension?

127. How many grams of potassium chloride are in 45.0 mL of a 5.00% (by mass) solution?

128. How would you prepare 400.0 mL of a 0.680*M* KOH solution?

***129.** How many liters of 8.0*M* HCl are needed to prepare 1.50 L of 2.5*M* HCl?

130. Which of these is an endothermic process? Provide an explanation.

a. burning wax
b. evaporating water
c. melting wax
d. roasting a marshmallow

131. How many joules of heat are required to melt a 55.0-g ice cube at 0°C?

***132.** Make the following conversions:

a. 34.5 cal to joules
b. 250 Cal to kilojoules
c. 0.347 kJ to calories

133. The specific heat capacity of iron is 0.46 J/(g·°C). How many kilojoules of energy are needed to raise the temperature of a 432-g iron bar 14°C?

134. What must be true about the concentration of two ions if precipitation occurs when solutions of the two ions are mixed?

135. Write an equilibrium-constant expression for each equation.

a. $2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$
b. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

136. What is the equilibrium concentration of barium ion in a 1.0-L saturated solution of $BaCO_3$ to which 0.25 mol K_2CO_3 has been added?

***137.** In each pair, which has the higher entropy?

a. $NaCl(s)$ or $NaCl(aq)$
b. $CO_2(s)$ or $CO_2(g)$
c. hot water or cold water

138. How would each change affect the position of equilibrium of this reaction?

$$2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g) + \text{heat}$$

a. increasing the pressure
b. adding a catalyst
c. increasing the concentration of $H_2(g)$
d. cooling the reaction mixture
e. removing water vapor from the container

139. For the reaction $A(g) + B(g) + C(g) \longrightarrow D(g)$, the following data were obtained at a constant temperature. From the data, determine the order of reaction with respect to A, B, and C, and the overall order of reaction.

Initial [A] (mol/L)	Initial [B] (mol/L)	Initial [C] (mol/L)	Initial rate (mol/(L·min))
0.0500	0.0500	0.0100	6.25×10^{-3}
0.1000	0.0500	0.0100	1.25×10^{-2}
0.1000	0.1000	0.0100	5.00×10^{-2}
0.0500	0.0500	0.0200	6.25×10^{-3}

If You Have Trouble With . . .

Question	119	120	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135	136	137	138	139
See Chapter	11	12	13	14	14	15	15	15	16	16	16	17	17	17	17	18	18	18	18	18	18

Standardized Test Prep

Select the choice that best answers each question or completes each statement.

1. If an acid has a measured K_a of 3×10^{-6},
   (A) the acid is a strong acid.
   (B) an aqueous solution of the acid would have a pH < 7.
   (C) the acid is a strong electrolyte.
   (D) All of the above are correct.

2. The pH of a sample of orange juice is 3.5. A sample of tomato juice has a pH of 4.5. Compared to the $[H^+]$ of orange juice, the $[H^+]$ of tomato juice is
   (A) 1.0 times higher. (C) 10 times higher.
   (B) 10 times lower. (D) 1.0 times lower.

Tips for Success

Eliminate Wrong Answers If you don't know which choice is correct, eliminate those you know are wrong. If you can rule out some choices, you'll increase your chances of choosing the correct answer.

3. Which ion or molecule is the conjugate base of the ammonium ion, NH_4^+?
   (A) H_2O (B) OH^- (C) NH_3 (D) H_3O^+

4. How many moles of NaOH are required to neutralize 2.4 mol H_2SO_4?
   (A) 1.2 mol (B) 2.4 mol (C) 3.6 mol (D) 4.8 mol

5. A solution with a hydrogen-ion concentration of $2.3 \times 10^{-8}M$ has a pH between
   (A) 2 and 3. (B) 3 and 4. (C) 7 and 8. (D) 8 and 9.

6. The net ionic equation for the neutralization reaction between solutions of potassium hydroxide and hydrochloric acid is
   (A) $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$
   (B) $KOH(aq) + HCl(aq) \longrightarrow H_2O(l) + KCl(aq)$
   (C) $K^+(aq) + Cl^-(aq) \longrightarrow KCl(aq)$
   (D) $K^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \longrightarrow KCl(aq) + H_2O(l)$

7. Calculate the molarity of an HCl solution if 25.0 mL of the solution is neutralized by 15.5 mL of 0.800M NaOH.
   (A) 0.248M (B) 0.496M (C) 1.29M (D) 0.645M

8. Which combination of compound and ion would not make a useful buffer solution?
   (A) ammonium ion and ammonia
   (B) hydrogen carbonate ion and carbonic acid
   (C) sulfate ion and sulfuric acid
   (D) ethanoate ion and ethanoic acid

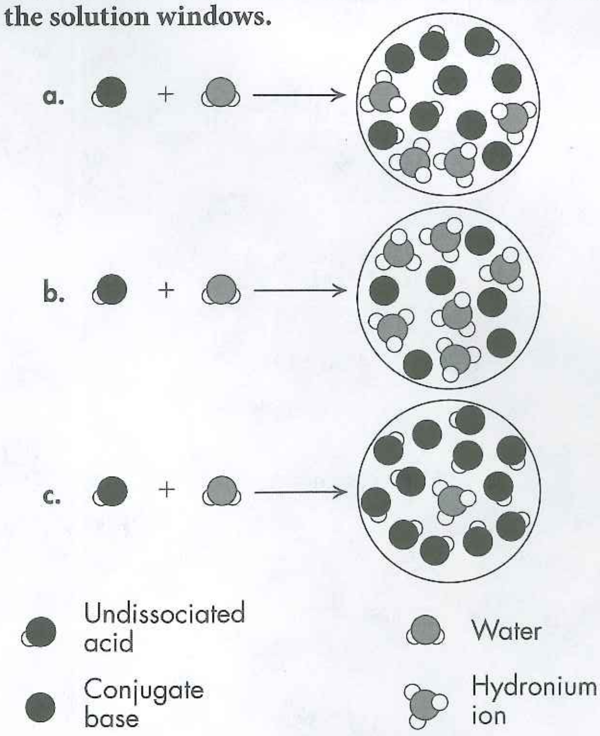
The lettered choices below refer to Questions 9–11. In each formula, P is the cation, and Q is the anion.

(A) PQ (B) P_2Q_3 (C) PQ_3 (D) P_3Q

Which of the choices is the general formula for the salt formed in each of the following neutralization reactions?

9. $H_3PO_4 + NaOH \longrightarrow$
10. $H_2SO_4 + Mg(OH)_2 \longrightarrow$
11. $HNO_3 + Al(OH)_3 \longrightarrow$

Use the drawings below to answer Questions 12 and 13. Water molecules have been omitted from the solution windows.



12. Rank the acids in order of increasing strength.
13. How many of the acids are strong acids?

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10	11	12	13
See Lesson	19.3	19.2	19.1	19.4	19.2	19.4	19.4	19.5	19.4	19.4	19.4	19.3	19.3