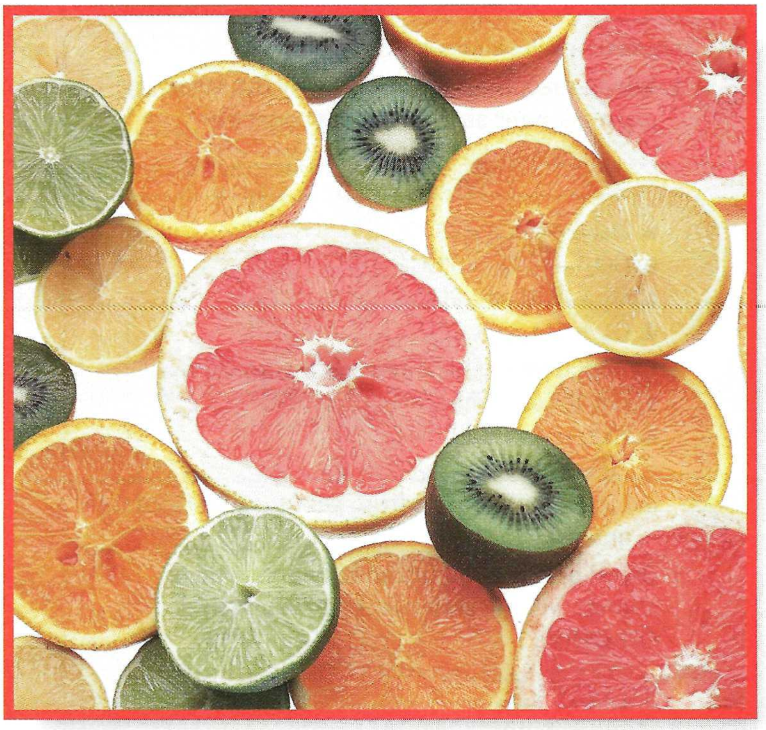
Chapter 20 ACIDS AND BASES



These tart-tasting fruits contain a weak acid called citric acid.

FEATURES

DISCOVER IT! EFFECT OF FOODS ON BAKING SODA

You need baking soda (sodium hydrogen carbonate, $NaHCO_3$), a large paper plate, a knife, paper towels, and a variety of fruits and vegetables (e.g., a celery stalk, a banana, a grape, a tomato, a lemon, an apple, an orange, a grapefruit).

1. Carefully cut the fruits and vegetables into small pieces and place them on the plate. Make sure the pieces are well separated from each other. Be sure to wipe any juice off the knife after cutting each fruit.
2. Sprinkle a pinch of baking soda on each sample.

What do you observe? Is there any relationship between what you observe and which foods you know from experience have a sour taste? After reading about acids and bases, provide an explanation for what you observed.

DESCRIBING ACIDS AND BASES

section 20.1

Some ants can give painful stings when threatened or disturbed. Certain ant species called formicines have poison glands that produce venom containing formic acid. Formicines protect themselves by spraying this venom on their predators. Formic acid can stun or even kill the ants' most common enemies. A formicine attack on a human, however, is much less severe, usually resulting only in blistered skin. **What are some of the properties of acids?**

objectives

- List the properties of acids and bases
- Name an acid or base when given the formula

key terms

- acid
- base

Properties of Acids and Bases

Did you know that acids and bases play a central role in much of the chemistry that affects your daily life? Most manufacturing processes use acids or bases. Your body needs acids and bases to function properly. Vinegar, carbonated drinks, and foods such as citrus fruits contain acids. The electrolyte in a car battery is an acid. Bases are present in many commercial products, including antacid tablets and household cleaning agents. **Figure 20.1** shows some of the many products that contain acids and bases. As you read this chapter, you will learn about the qualitative and quantitative aspects of acids and bases. You will see how these two classes of compounds ionize or dissociate in water. And you will learn how pH is used to describe the concentrations of acidic and basic solutions.

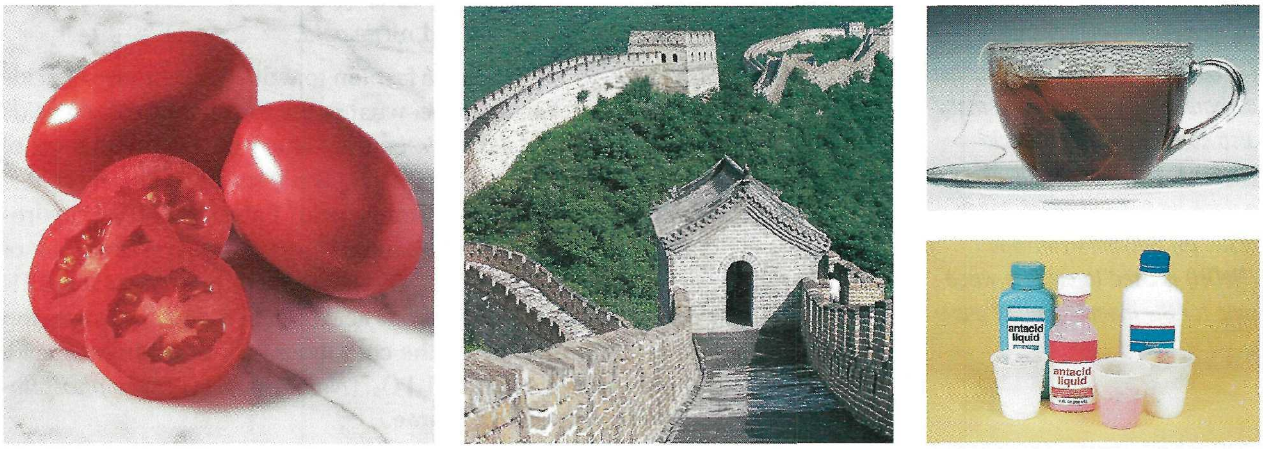
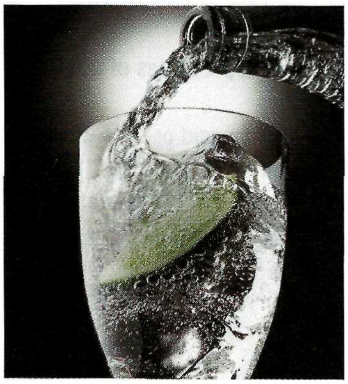


Figure 20.1

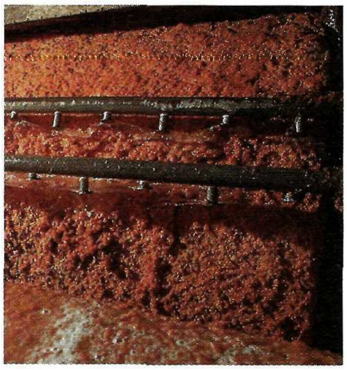
All of these items contain acids or bases, or produce acids or bases upon dissolving in water. Tomatoes contain ascorbic acid; tea contains tannic acid. The base calcium hydroxide is a component of mortar, which was used to build the Great Wall of China. Antacids use a variety of bases to neutralize excess stomach acid.

As you just read and perhaps already knew, many common items contain acids. Acids have several distinctive properties with which you are probably familiar. Acidic compounds give foods a tart or sour taste. For example, vinegar contains ethanoic acid, sometimes called acetic acid. Lemons, which taste sour enough to make your mouth pucker, contain citric acid. What type of acid do you think limes contain?

Aqueous solutions of acids are electrolytes. Recall from Chapter 17 that electrolytes conduct electricity. Some acid solutions are strong electrolytes and others are weak electrolytes. 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. Acids react with compounds containing hydroxide ions to form water and a salt.



(a)



(b)

Figure 20.2

(a) Carbonated soft drinks contain carbonic acid (H_2CO_3), and many also contain phosphoric acid (H_3PO_4). These two acids give a drink its fizz and sharp, tangy taste. (b) Sodium hydroxide (NaOH) is used to prepare wood pulp for the manufacture of paper.

Bases are compounds that react with acids to form water and a salt. Milk of magnesia (a suspension of magnesium hydroxide in water) is a base used to treat the problem of excess stomach acid. Aqueous solutions of bases taste bitter and feel slippery. Like acids, bases will change the color of an acid–base indicator and can be strong or weak electrolytes.

Names and Formulas of Acids and Bases

An **acid** is a compound that produces hydrogen ions when dissolved in water. Therefore, the chemical formulas of acids are of the general form HX, where X is a monatomic or polyatomic anion. When the compound $HCl(g)$ (hydrogen chloride) dissolves in water to form $HCl(aq)$, it is named as an acid. How, then, is an acid named? To illustrate the naming of an acid, consider the following three rules as applied to the acid HX dissolved in water. Notice that the rules focus on the name of the anion, in particular the suffix of the anion name. **Table 20.1** summarizes these rules.

1. When the name of the anion (X) ends in *-ide*, the acid name begins with the prefix *hydro-*. The stem of the anion has the suffix *-ic* and is followed by the word *acid*. Therefore, $HCl(aq)$ (X = chloride) is named *hydro*chlor*ic acid*. $H_2S(aq)$ (X = sulfide) is named *hydro*sulfur*ic acid*.
2. When the anion name ends in *-ite*, the acid name is the stem of the anion with the suffix *-ous*, followed by the word *acid*. Thus $H_2SO_3(aq)$ (X = sulfite) is named sulfur*ous acid*.
3. When the anion name ends in *-ate*, the acid name is the stem of the anion with the suffix *-ic*, followed by the word *acid*. Thus $HNO_3(aq)$ (X = nitrate) is named nitr*ic acid*.

These rules can be used in reverse fashion to write the formulas of acids when given their names. For example, what is the formula of chloric acid? According to Rule 3, chloric acid (*-ic* ending) must be a combination of hydrogen ion (H^+) and chlor*ate* ion (ClO_3^-). The formula of chloric acid is $HClO_3$. What is the formula of hydrobromic acid? Following Rule 1, hydrobromic acid (*hydro-* prefix and *-ic* suffix) must be a combination of hydrogen ion and bromide ion (Br^-). The formula of hydrobromic acid is HBr. What is the formula for phosphorous acid? Using Rule 2, hydrogen ion and phosphite ion (PO_3^{3-}) must be the components of phosphorous acid. The formula of phosphorous acid is H_3PO_3. (*Note:* Do not confuse *phosphorous* with *phosphorus*, the element name.)

Table 20.1

Naming Acids			
Anion ending	**Example**	**Acid name**	**Example**
-ide	Cl^- chlor*ide*	*hydro*-(stem)-*ic acid*	hydrochlor*ic acid*
-ite	SO_3^{2-} sulf*ite*	(stem)-*ous acid*	sulfur*ous acid*
-ate	NO_3^- nitr*ate*	(stem)-*ic acid*	nitr*ic acid*

A **base** is a compound that produces hydroxide ions when dissolved in water. Ionic compounds that are bases are named in the same way as any other ionic compound—the name of the cation followed by the name of the anion. For example, NaOH, a base used in making paper pulp, detergents, and soap, is called sodium hydroxide. What would you call $Ca(OH)_2$? You write the formulas for bases by balancing the ionic charges, just as you do for any ionic compound.

Sample Problem 20-1

Name these compounds as acids.

a. HClO **b.** HCN **c.** H_3PO_4

1. ANALYZE ***Plan a problem-solving strategy.***

The rules for naming acids can be applied because the formulas of the acids are given.

2. SOLVE ***Apply the problem-solving strategy.***

a. Use Rule 2. The anion name (hypochlorite) ends in *-ite,* so add the suffix *-ous* to the anion stem, followed by the word *acid.* The correct name is hypochlorous acid.

b. Use Rule 1. The anion name (cyanide) ends in *-ide,* so this acid name begins with the prefix *hydro-* and ends with the suffix *-ic,* followed by the word *acid.* The correct name is hydrocyanic acid.

c. Use Rule 3. The anion name (phosphate) ends in *-ate.* So add the suffix *-ic* to the anion stem (in this case, modified slightly to *phosphor*), followed by the word *acid.* The correct name is phosphoric acid.

3. EVALUATE ***Do the results make sense?***

The names are consistent with the indicated rules.

Practice Problems

1. Name each acid or base.

a. HF **c.** KOH

b. HNO_3 **d.** H_2SO_4

2. Write formulas for each acid or base.

a. chromic acid

b. iron(II) hydroxide

c. hydriodic acid

d. lithium hydroxide

Chem ASAP!

Problem-Solving 2

Solve Problem 2 with the help of an interactive guided tutorial.

section review 20.1

3. Identify each property as applying to an acid, a base, or both.

a. bitter taste **c.** indicator color change

b. electrolyte **d.** sour taste

4. Write the formula for each acid or base.

a. barium hydroxide **c.** rubidium hydroxide

b. hydrobromic acid **d.** hydroselenic acid

5. Name each acid or base.

a. HF **c.** H_2CO_3

b. $HClO_3$ **d.** $Al(OH)_3$

Chem ASAP! **Assessment 20.1** Check your understanding of the important ideas and concepts in Section 20.1.

section 20.2

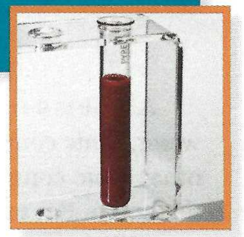
HYDROGEN IONS AND ACIDITY

objectives

- Given the hydrogen-ion or hydroxide-ion concentration, classify a solution as neutral, acidic, or basic
- Convert hydrogen-ion concentrations into values of pH and hydroxide-ion concentrations into values of pOH

key terms

- hydroxide ion (OH^-)
- hydronium ion (H_3O^+)
- self-ionization
- neutral solution
- ion-product constant for water (K_w)
- acidic solution
- basic solution
- alkaline solutions
- pH



A patient is brought to a hospital unconscious and with a fruity odor on his breath. The doctor suspects the patient has fallen into a diabetic coma. To confirm her diagnosis, she orders several tests, including one of the acidity of the patient's blood. The results from this test will be expressed in units of pH, not molar concentration. **How is the pH scale used to indicate the acidity of a solution, and why is this scale used?**

Hydrogen Ions from Water

As you already know, water molecules are highly polar and are in continuous motion, even at room temperature. Occasionally, the collisions between water molecules are energetic enough to transfer a hydrogen ion from one water molecule to another. See **Figure 20.3.** A water molecule that loses a hydrogen ion becomes a negatively charged **hydroxide ion (OH^-).** A water molecule that gains a hydrogen ion becomes a positively charged **hydronium ion (H_3O^+).**

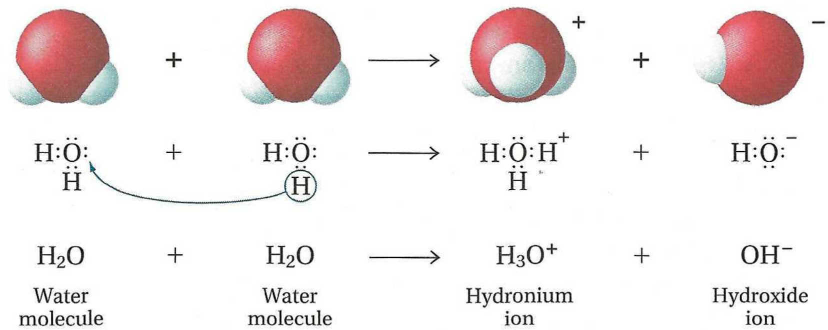


Figure 20.3
What is the name and formula of the particle that results when a water molecule gains a hydrogen ion? How is a hydroxide ion formed from a water molecule?

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

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

In water or aqueous solution, hydrogen ions (H^+) are always joined to water molecules as hydronium ions (H_3O^+). The hydronium ions are themselves solvated to form species such as $H_9O_4^+$. Hydrogen ions in aqueous solution have several names. Some chemists call them protons. Others prefer to call them hydrogen ions, hydronium ions, or solvated 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 ($[H^+]$) and the concentration of hydroxide ions ($[OH^-]$) are each only $1.0 \times 10^{-7}M$. This means that the concentrations of H^+ and OH^- are equal in pure water. Any aqueous solution in which $[H^+]$ and $[OH^-]$ are equal is described as a **neutral solution.**

In any aqueous solution, $[H^+]$ and $[OH^-]$ are interdependent. In other words, when $[H^+]$ increases, $[OH^-]$ decreases. When $[H^+]$ decreases, $[OH^-]$ increases. Le Châtelier's principle, which you learned about in Chapter 19, applies here. If additional ions (either hydrogen ions or hydroxide ions) are added to a solution, the equilibrium shifts. The concentration of the other type of ion decreases. More water molecules are formed in the process.

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

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

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

The product of the concentrations of the hydrogen ions and hydroxide ions in water is called the **ion-product constant for water (K_w).**

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

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. How does hydrochloric acid differ from hydrogen chloride?

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

In such a solution, the hydrogen-ion concentration is greater than the hydroxide-ion concentration. The hydroxide ions are present from the self-ionization of water. An **acidic solution** is one in which $[H^+]$ is greater than $[OH^-]$. Therefore, the $[H^+]$ of an acidic solution is greater than $1.0 \times 10^{-7}M$.

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

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

In such a solution, the hydrogen-ion concentration is less than the hydroxide-ion concentration. The hydrogen ions are present from the self-ionization of water. A **basic solution** is one in which $[H^+]$ is less than $[OH^-]$. Therefore, the $[H^+]$ of a basic solution is less than $1.0 \times 10^{-7}M$. Basic solutions are also known as **alkaline solutions.** Look at **Figure 20.4.** What are the names of the acids and bases shown?

Figure 20.4

(a) One of these two chemical reagents is an acid and the other is a base. Which of these reagents would increase the hydrogen-ion concentration when added to an aqueous solution? Which would increase the hydroxide-ion concentration? (b) Unrefined hydrochloric acid, commonly known as muriatic acid, is used to clean stone buildings and swimming pools. (c) Sodium hydroxide, or lye, is commonly used as a drain cleaner.

(a)

(b)

(c)

Sample Problem 20-2

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?

1. ***ANALYZE*** ***List the knowns and the unknowns.***

   Knowns:
   - $[H^+] = 1.0 \times 10^{-5}M$
   - Ion-product constant for water: $K_w = [H^+] \times [OH^-] = 1 \times 10^{-14}$

   Unknowns:
   - solution = acidic, basic, or neutral?
   - $[OH^-] = ?M$

2. ***CALCULATE*** ***Solve for the unknowns.***

   $[H^+] = 1.0 \times 10^{-5}M$. Because this is greater than $1.0 \times 10^{-7}M$, the solution is acidic. By definition, $K_w = [H^+] \times [OH^-]$.

   Therefore, $[OH^-] = \dfrac{K_w}{[H^+]}$.

   Substituting the known numerical values, $[OH^-]$ is computed as follows.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}M} = 1.0 \times 10^{-9}M$$

3. ***EVALUATE*** ***Do the results make sense?***

   If $[H^+]$ is greater than $1.0 \times 10^{-7}M$, then $[OH^-]$ must be less than $1.0 \times 10^{-7}M$. At $1 \times 10^{-9}M$, $[OH^-]$ is less than $1 \times 10^{-7}M$. Notice that when the concentration of the acid is 1.0×10^{-x} and the concentration of the base is 1.0×10^{-y}, $x + y = 14$. In this special case, you can easily find the value of $[H^+]$ when you know $[OH^-]$, or vice versa. To do so, subtract the exponent of the known, either $[H^+]$ or $[OH^-]$, from 14. For example, if

$$[H^+] = 1.0 \times 10^{-9}M,$$
$$\text{then } [OH^-] = 1.0 \times 10^{-(14-9)}M$$
$$= 1.0 \times 10^{-5}M.$$

Practice Problems

6. If the hydroxide-ion concentration of an aqueous solution is $1.0 \times 10^{-3}M$, what is the $[H^+]$ in the solution? Is the solution acidic, basic, or neutral?

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

The pH Concept

Expressing hydrogen-ion concentration in molarity is cumbersome. A more widely used system for expressing $[H^+]$ is the pH scale, proposed in 1909 by the Danish scientist Søren Sørensen (1868–1939). On the pH scale, which ranges from 0 to 14, neutral solutions have a pH of 7. A pH of 0 is strongly acidic. What is a solution with a pH of 14? Calculating the pH of a solution is straightforward. The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration. The pH may be represented mathematically using the following equation.

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

In a neutral solution, $[H^+] = 1 \times 10^{-7}M$. The pH of a neutral solution is 7.0.

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

To summarize, the pH of pure water or a neutral aqueous solution is 7.0. A solution in which $[H^+]$ is greater than $1 \times 10^{-7}M$ has a pH less than 7.0 and is acidic. A solution with a pH greater than 7 is basic and has a $[H^+]$ of less than $1 \times 10^{-7}M$. See **Figure 20.5** for a visual representation of this information.

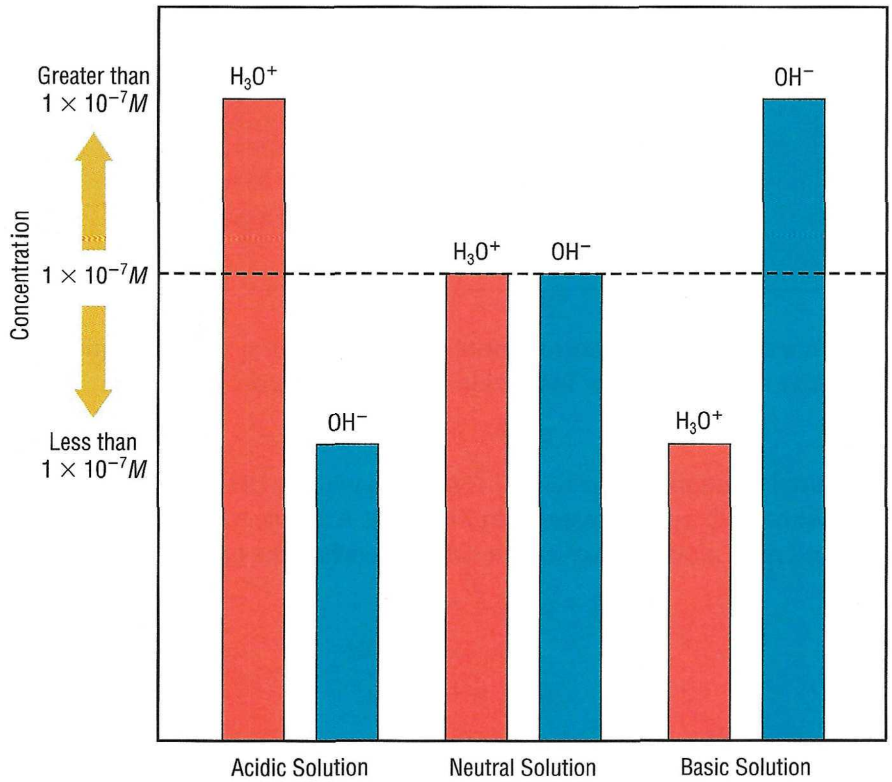


Figure 20.5
In acidic solutions, $[H^+]$ is greater than $[OH^-]$. In basic solutions, $[OH^-]$ is greater than $[H^+]$. Neutral solutions are those in which $[H^+]$ is equal to $[OH^-]$.

- Acidic solution: $pH < 7.0$ $[H^+]$ greater than $1 \times 10^{-7}M$
- Neutral solution: $pH = 7.0$ $[H^+]$ equals $1 \times 10^{-7}M$
- Basic solution: $pH > 7.0$ $[H^+]$ less than $1 \times 10^{-7}M$

The pH values of several common aqueous solutions are listed in **Table 20.2** on the following page. The table also summarizes the relationship among $[H^+]$, $[OH^-]$, and pH. You may notice that pH can sometimes be read from the value of $[H^+]$. If $[H^+]$ is written in scientific notation and has a coefficient of 1, then 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. What is the pH of a solution with $[H^+] = 1 \times 10^{-4}M$?

Table 20.2

Relationship among [H^+], [OH^-], and pH

	[H^+] (mol/L)	[OH^-] (mol/L)	pH	Aqueous system
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; Gastric juice
	1×10^{-2}	1×10^{-12}	2.0	Lemon juice
	1×10^{-3}	1×10^{-11}	3.0	
	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	Sodium hydrogen carbonate, sea water
	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	Washing soda
	1×10^{-13}	1×10^{-1}	13.0	0.1 *M* NaOH
	1×10^{-14}	1×10^{0}	14.0	1 *M* NaOH

In a definition similar to that of pH, the pOH of a solution equals the negative logarithm of the hydroxide-ion concentration.

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

A neutral solution has a pOH of 7. A solution with a pOH less than 7 is basic. A solution with a pOH greater than 7 is acidic. A simple relationship between pH and pOH makes it easy to find either one when the other is known.

$$\text{pH} + \text{pOH} = 14$$
$$\text{pH} = 14 - \text{pOH}$$
$$\text{pOH} = 14 - \text{pH}$$

For pH calculations, you should express the hydrogen-ion concentration in scientific notation. For example, a hydrogen-ion concentration of 0.0010*M*, rewritten as $1.0 \times 10^{-3}M$ in scientific notation, has two significant figures. The pH of this solution is 3.00, with the two numbers to the right of the decimal point representing the two significant figures in the concentration. A solution with a pH of 3.00 is acidic, as shown in **Figure 20.6.** How many significant figures are indicated in a pH of 7.61?

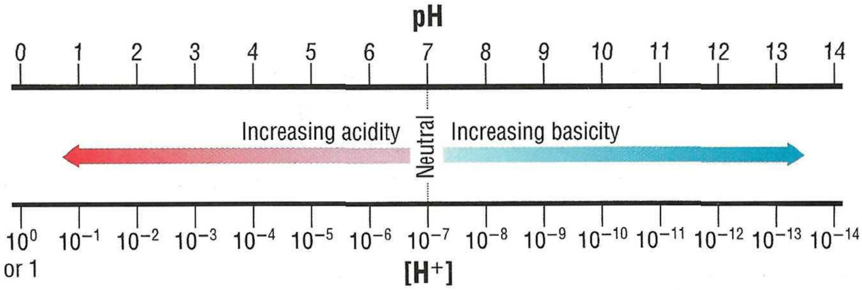


Figure 20.6
The pH scale shows the relationship between pH and the hydrogen-ion concentration. Notice that acids have lower pHs than bases.

CHEMath

USING LOGARITHMS

Logarithms are used throughout mathematics and science. Examples include the decibel scale for loudness, the Richter scale for earthquakes, and the pH scale for acidity. Because the pH scale is a logarithmic scale, it allows a tremendous range of values (from 1 to 10^{-14}) to be expressed as a number between 1 and 14.

The common logarithm of a number is the exponent to which ten must be raised to produce the given number: $\log x = y$ if $x = 10^y$. Thus $\log 0.01 = -2$ because $0.01 = 10^{-2}$, and $\log 10\,000 = 4$ because $10\,000 = 10^4$. The common logarithm is also known as the base-10 logarithm. Logarithms can have other bases, but this textbook only uses common logarithms.

Logarithms of numbers between 1 and 10 can be evaluated directly using **Table B.1** in Appendix B of this textbook. For example, $\log 7.21 = 0.8579$, as shown below.

x	0	1	2
7.1	.8513	.8519	.8525
7.2	.8573	.8579	.8585
7.3	.8633	.8639	.8645

To find the logarithm of a number that is less than 1 or greater than 10, write the number in scientific notation and use the formulas below. See Example 1.

$$\log (a \times b) = \log a + \log b, \text{ and } \log (10^x) = x$$

A logarithm is rounded so that its number of decimal places equals the number of significant figures in the original number.

The antilog of a number x is the number y whose logarithm is x. Because $\log y = x$ means $y = 10^x$, the antilog of x is the same as 10^x. Antilogs of numbers between 0 and 1 can be found directly from **Table B.1.** To find the antilog of any number, write it as a sum of its decimal part (between 0 and 1) and its integer part (less than or equal to the given number). Then use the formula $10^{a+b} = 10^a \times 10^b$. See Example 2.

Many calculators can be used to find logarithms and antilogs. In general, use LOG for a logarithm, or 10^x for an antilog. The exact keystrokes depend on your calculator. Do not confuse the LOG key with the LN key. The LN key gives another kind of logarithm, called the natural logarithm, whose base is the constant $e = 2.718$.

Example 1

Evaluate log 0.0000721.

$$\begin{aligned} \log 0.0000721 &= \log (7.21 \times 10^{-5}) \\ &= \log 7.21 + \log (10^{-5}) \\ &= 0.8579 + (-5) \\ &= -4.1421 \end{aligned}$$

The original number had 3 significant figures, so the result should be rounded to 3 decimal places.

$$\log 0.0000721 = -4.142$$

Example 2

Find the antilog of −8.375.

Write −8.375 as a sum. Note that the decimal part is 0.625, not −0.375 or 0.375.

$$-8.375 = (-8.375 + 9) + (-9) = 0.625 + (-9)$$

In **Table B.1**, the number nearest to 0.625 is 0.6253, and its antilog is 4.22. So,

$$\begin{aligned} \text{antilog}(-8.375) &= 10^{-8.375} = 10^{0.625+(-9)} \\ &= 10^{0.625} \times 10^{-9} \\ &= 4.22 \times 10^{-9} \end{aligned}$$

Practice Problems

Prepare for upcoming problems in this chapter by using **Table B.1** in Appendix B to evaluate the following expressions. Then check your work using a calculator.

A. log 1.68
B. log (3.57×10^4)
C. log (2.18×10^{-8})
D. antilog −8
E. antilog 0.969
F. antilog (−5 + 0.782)
G. log 17 800
H. log 0.0067
I. log 0.0000738
J. antilog 6.281
K. antilog (−3.192)
L. antilog (−5.936)

Sample Problem 20-3

What is the pH of a solution with a hydrogen-ion concentration of $1.0 \times 10^{-10}M$?

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

Knowns:
- $[H^+] = 1.0 \times 10^{-10}M$
- $pH = -\log[H^+]$

Unknown:
- $pH = ?$

2. *CALCULATE* *Solve for the unknown.*

$$\begin{aligned} pH &= -\log[H^+] \\ &= -\log(1.0 \times 10^{-10}) \\ &= -(\log 1.0 + \log 10^{-10}) \\ &= -(0.00 + (-10.00)) \\ &= -(-10.00) \\ &= 10.00 \end{aligned}$$

3. *EVALUATE* *Does the result make sense?*

$[H^+]$ is three orders of magnitude (1000 times) less than $1 \times 10^{-7}M$ (pH 7). Each order-of-magnitude decrease in $[H^+]$ equals an increase of 1 pH unit, so the unknown solution should have a pH of $7 + 3 = 10$. This problem can also be solved by inspection. The coefficient of $[H^+]$ is 1.0; therefore, the pH is the value of the exponent (-10) with the sign changed from minus to plus. However, the answer based on experimental data must be reported as 10.00 to show the two significant figures of the original value of $[H^+]$.

Practice Problems

8. Find the pH of each solution.
   a. $[H^+] = 1.0 \times 10^{-4}M$
   b. $[H^+] = 0.0010M$
   c. $[H^+] = 1.0 \times 10^{-9}M$

9. What are the pH values of the following three solutions, based on their hydrogen-ion concentrations?
   a. $[H^+] = 1.0 \times 10^{-12}M$
   b. $[H^+] = 0.010M$
   c. $[H^+] = 1.0 \times 10^{-4}M$

Chem ASAP!

Problem-Solving 9
Solve Problem 9 with the help of an interactive guided tutorial.

Sample Problem 20-4

The pH of an unknown solution is 6.00. What is its hydrogen-ion concentration?

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

Knowns:
- $pH = 6.00$
- $pH = -\log[H^+]$

Unknown:
- $[H^+] = ?M$

2. *CALCULATE* *Solve for the unknown.*

First, rearrange the equation for the definition of pH to solve for the unknown.

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

Next, substitute the value of pH.

$$-\log[H^+] = 6.00$$

Sample Problem 20-4 (cont.)

Change the signs on both sides of the equation.

$$\log [H^+] = -6.00$$

Finally, determine the number that has a log of -6.00. The antilog of -6.00 is 1.0×10^{-6}. Therefore,

$$[H^+] = 1.0 \times 10^{-6}M.$$

3. *EVALUATE* *Does the result make sense?*

This problem can be solved by inspection. Any integral value of pH, such as 3.00, 5.00, and so forth, can easily be converted to $[H^+]$. The sign of the pH value is changed from plus to minus and is used as the exponent in the numerical expression of $[H^+]$. A decimal point and zeros are attached to the coefficient (1) to obtain the proper number of significant figures. For example, for pH 4.00, $[H^+] = 1.0 \times 10^{-4}M$; for pH 11.00, $[H^+] = 1.0 \times 10^{-11}M$.

Practice Problems

10. Calculate $[H^+]$ for each solution.

a. pH = 5.00 **c.** pH = 6.00
b. pH = 7.00 **d.** pH = 3.00

11. What are the hydrogen-ion concentrations for solutions with the following pH values?

a. 4.00 **b.** 11.00 **c.** 8.00

Calculating pH Values

Most pH values are not whole numbers. For example, **Table 20.2** on page 584 shows that milk of magnesia has a pH of 10.5. Using the definition of pH, this means that $[H^+]$ must equal $1 \times 10^{-10.5}M$. Thus $[H^+]$ must be less than $1 \times 10^{-10}M$ (pH 10.0), but greater than $1 \times 10^{-11}M$ (pH 11.0).

If $[H^+]$ is written in scientific notation but its coefficient is not 1, then you need a table of common logarithms or a hand calculator with a log function key to calculate pH. A four-place table of common logarithms is provided in Appendix B. Sample Problem 20-5 shows how to make such a pH calculation.

Sample Problem 20-5

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

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

Knowns:

- $[OH^-] = 4.0 \times 10^{-11}M$
- $K_w = [OH^-] \times [H^+] = 1 \times 10^{-14}$
- $pH = -\log [H^+]$

Unknown:

- pH = ?

2. *CALCULATE* *Solve for the unknown.*

To calculate pH, first calculate $[H^+]$ by using the definition of K_w.

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

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

$$= 0.25 \times 10^{-3}M$$

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

Sample Problem 20-5 (cont.)

With the value of $[H^+]$ determined, use the definition of pH to solve for the pH.

$$\begin{aligned} pH &= -\log [H^+] \\ &= -\log (2.5 \times 10^{-4}) \\ &= -(\log 2.5 + \log 10^{-4}) \end{aligned}$$

A log table or calculator indicates that $\log 2.5 = 0.40$, and $\log 10^{-4} = -4$. Insert these values to find the pH.

$$\begin{aligned} pH &= -(0.40) - (-4) \\ &= -0.40 + 4 \\ &= 3.60 \end{aligned}$$

3. ***EVALUATE*** ***Does the result make sense?***
A solution in which $[OH^-]$ is less than $1 \times 10^{-7}M$ would be acidic because $[H^+]$ would be greater than $1 \times 10^{-7}M$. $[OH^-]$ is less than $10^{-10}M$ but greater than $10^{-11}M$. Therefore, the solution is somewhere between pOH 10 and pOH 11. Because pOH + pH = 14, the pH should be less than 4 and greater than 3.

Practice Problems

12. Calculate the pH of each solution.
- **a.** $[H^+] = 5.0 \times 10^{-6}M$
- **b.** $[H^+] = 8.3 \times 10^{-10}M$
- **c.** $[H^+] = 2.7 \times 10^{-7}M$

13. Calculate the pH of each solution.
- **a.** $[OH^-] = 4.3 \times 10^{-5}M$
- **b.** $[OH^-] = 2.0 \times 10^{-5}M$
- **c.** $[OH^-] = 4.5 \times 10^{-11}M$

Chem ASAP!

Problem-Solving 13
Solve Problem 13 with the help of an interactive guided tutorial.

You can calculate the hydrogen-ion concentration of a solution if you know the pH. For example, if the solution has a pH of 3.00, then $[H^+] = 1.0 \times 10^{-3}M$. When the pH is not a whole number, you will need log tables or a calculator with a y^x function key to calculate the hydrogen-ion concentration. For example, if the pH is 3.70, the hydrogen-ion concentration must be greater than $1.0 \times 10^{-4}M$ (pH 4.0) and less than $1.0 \times 10^{-3}M$ (pH 3.0). To get an accurate value, use log tables or a calculator.

Sample Problem 20-6

What is $[H^+]$ of a solution if the pH = 3.70?

1. ***ANALYZE*** ***List the knowns and the unknown.***
Knowns:
- pH = 3.70
- $pH = -\log [H^+]$

Unknown:
- $[H^+] = ?M$

2. ***CALCULATE*** ***Solve for the unknown.***
First rearrange the equation $pH = -\log [H^+]$.

$$\log [H^+] = -pH = -3.70$$

A log table cannot be used directly to find a number that has a negative log. To avoid this problem, add and then subtract the whole number that is closest to and larger than the negative log. In this case, the negative log is 3.70, and the whole number is 4.

Sample Problem 20-6 (cont.)

$$\log[H^+] = (-3.70 + 4) - 4$$
$$= 0.30 - 4$$
$$[H^+] = 10^{(0.30-4)}$$
$$= 10^{0.30} \times 10^{-4}$$

From the log table, the number with a log of 0.30 is 2.0; the antilog of 0.30 is thus 2.0. The number with a log of -4 is 10^{-4}. Therefore, $[H^+] = 2.0 \times 10^{-4}M$. A calculator with a y^x function key can be used because $[H^+] = 10^{-pH}$; that is, $[H^+]$ is of the form y^x. Thus change the sign of the given pH, in this case to -3.70. Enter $y = 10$, $x = -3.70$, and press the y^x key in the order required by the calculator. The readout gives 1.995×10^{-4}. Rounded to two significant figures, $[H^+] = 2.0 \times 10^{-4}M$.

3. EVALUATE ***Does the result make sense?***
A solution of pH 3.70 should have a $[H^+]$ between $1 \times 10^{-4}M$ (pH 4) and $1 \times 10^{-3}M$ (pH 3), as found. Because $[H^+]$ is greater than $1 \times 10^{-7}M$, the solution is acidic.

Practice Problems

14. What is the molarity of $[H^+]$ in each solution?
a. pH = 7.30 **c.** pH = 7.05
b. pH = 1.80 **d.** pH = 6.70

15. Calculate the value of $[OH^-]$ in each solution in Practice Problem 14.

Measuring pH

People need to be able to measure the pH of the solutions they use. From maintaining the correct acid–base balance in a swimming pool, to creating soil conditions ideal for plant growth, to making medical diagnoses, pH measurement has valuable applications. For preliminary pH measurements and for small-volume samples, indicators such as the ones shown in **Figure 20.7** are often used. For precise and continuous measurements, a pH meter is preferred.

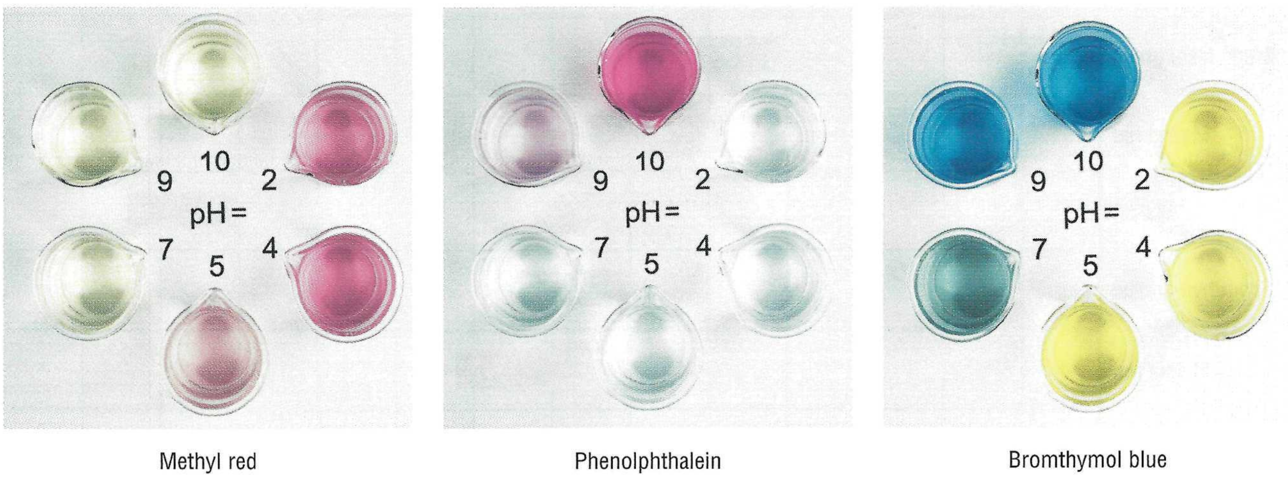


Figure 20.7
Acid–base indicators respond to pH changes over a specific range. Methyl red (left) changes from red to yellow at pH 5–7. Phenolphthalein (center) changes from colorless to pink at pH 7–9. Bromthymol blue (right) changes from yellow to blue at pH 5–7.

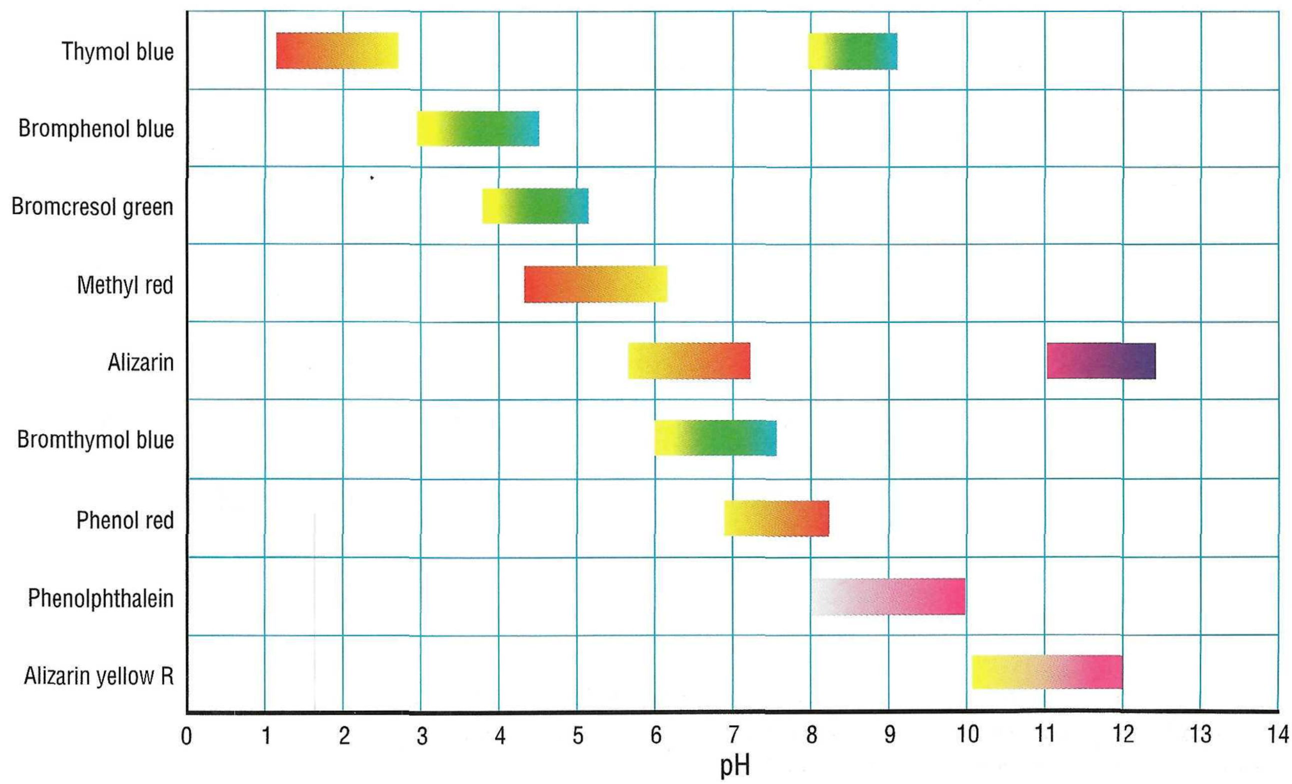
Acid–Base Indicators An indicator (In) is an acid or a base that undergoes dissociation in a known pH range. An indicator is a valuable tool for measuring pH because its acid form and base form have different colors in solution. The following generalized equation represents the dissociation of an indicator (HIn)

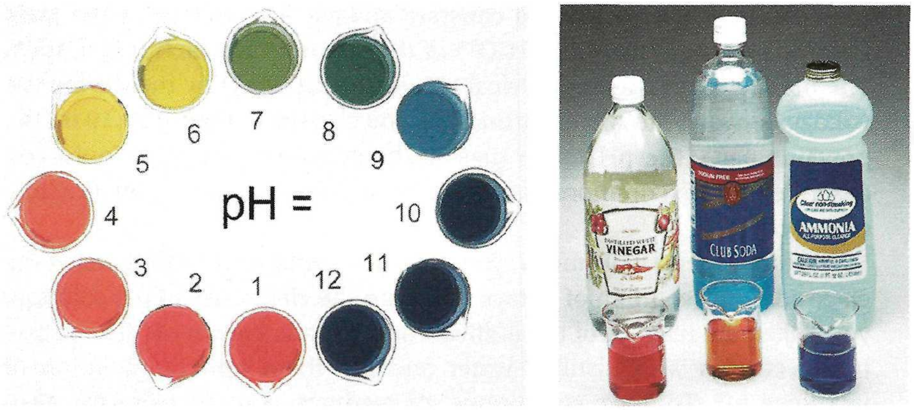
$$\underset{\text{Acid form}}{HIn(aq)} \underset{H^+}{\overset{OH^-}{\rightleftharpoons}} H^+(aq) + \underset{\text{Base form}}{In^-(aq)}$$

The acid form dominates the dissociation equilibrium at low pH (high $[H^+]$), and the base form dominates the equilibrium at high pH (high $[OH^-]$). For each indicator, the change from dominating acid form to dominating base form occurs in a narrow range of approximately two pH units. Within this narrow range, the color of the solution is a mixture of the colors of the acid and the base forms. Knowing the pH range over which this color change occurs can give you 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 observe only the color of the base form. You could eventually zero in on a more precise estimate of the pH of the solution by repeating the experiment with indicators that have different pH ranges for their color changes. Many different indicators are needed to span the entire pH spectrum. **Figure 20.8** shows the pH ranges of some commonly used indicators.

Figure 20.8

Each indicator changes color at a different pH. Which indicator would you choose to show that a reaction solution has changed from pH 3 to pH 4?





(a) (b)

Figure 20.9
You can determine the approximate pH of a substance by testing it with pH indicator and comparing the color with standards. ***(a)*** *Universal indicator solution has been added to solutions of known pH in the range from 1 to 12 to produce a set of reference colors.* ***(b)*** *Universal indicator has been added to samples of some familiar household products: vinegar, soda water, and ammonia solution. Use the colors in* ***(a)*** *to assign pH values to the items in* ***(b)****.*

Although indicators are useful tools, they do have certain characteristics that limit their usefulness. The listed 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 distorted. Dissolved salts in a solution may also affect the indicator's dissociation. Often, using indicator strips can help overcome these problems. An indicator strip is a piece of paper or plastic impregnated with an indicator. The paper is dipped into an unknown solution and compared with a color chart to measure the pH. Some indicator paper is impregnated with multiple indicators that give a palette of colors over a wide pH range. See **Figure 20.9.**

pH Meters A pH meter is used to make rapid, accurate pH measurements. Most chemistry laboratories have a pH meter. If a pH meter is connected to a computer or chart recorder, it can be used to make a continuous recording of pH changes.

Figure 20.10 shows a pH meter. The combination electrode—a glass electrode and a reference electrode—is connected to a millivoltmeter.

(a) (b) (c)

Figure 20.10
(a) *A pH meter is used to measure hydrogen ion concentrations. The instrument gives the user a direct reading of pH. For continuous monitoring, the pH meter may be connected to a strip-chart recorder.* ***(b)*** *The pH of vinegar, a dilute aqueous solution of ethanoic (acetic) acid, is about 3.* ***(c)*** *The pH of milk of magnesia, an aqueous suspension of magnesium hydroxide, is 10.5. In which of the test solutions is* $[H^+] < 1 \times 10^{-7}\,M$*?*

The reference electrode has a constant voltage. The voltage of the glass electrode changes with the [H_3O^+] of the solution in which it is dipped. The pH meter makes an electrical measurement of pH by measuring the voltage between the two electrodes. As you can see in **Figure 20.10** on the previous page, the pH meter gives a direct readout of pH. Would you describe the two common household items shown in this figure as acidic or basic?

Before you use a pH meter, you must first calibrate it by immersing the electrodes in a solution of known pH. With the electrodes in the solution, you adjust the readout of the millivoltmeter to this known pH. Then rinse the electrodes with distilled water and dip them into the solution of unknown pH. To make continuous pH readings, you can leave the electrodes in the solution.

A pH meter is a valuable instrument. In many situations it is easier to use than liquid indicators or indicator strips. Measurements of pH obtained with a pH meter are typically accurate to within 0.01 pH unit of the true pH. The color and cloudiness of the unknown solution do not affect the accuracy of the pH value obtained. Hospitals use pH meters to find small but meaningful changes of pH in blood and other body fluids. Sewage and industrial effluents and soil pH are also easily monitored with a pH meter.

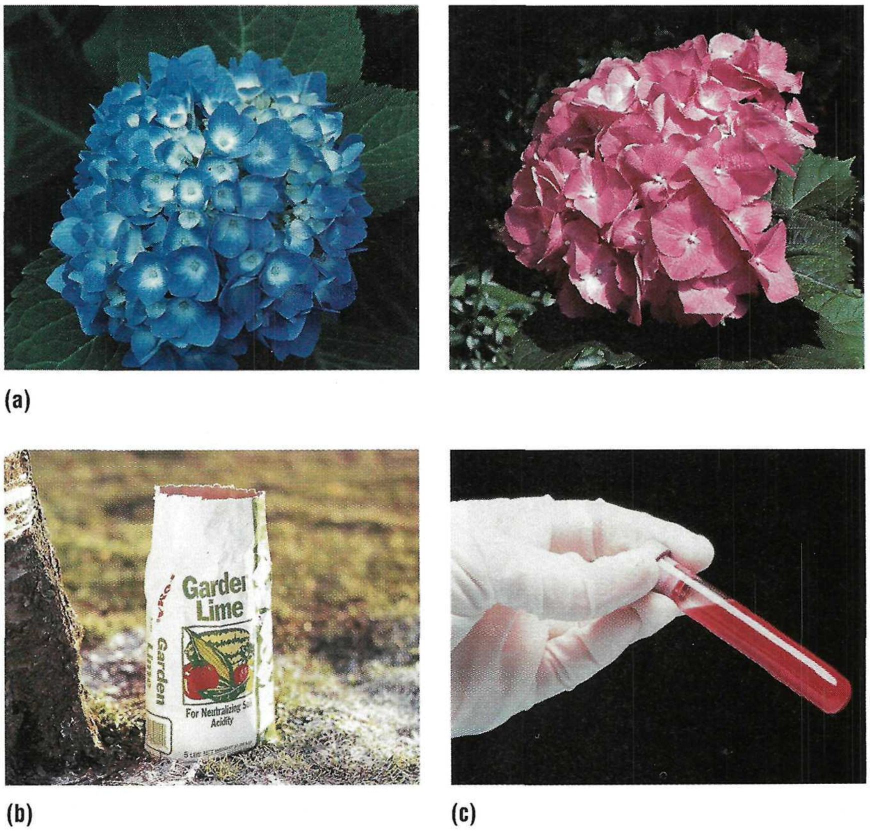
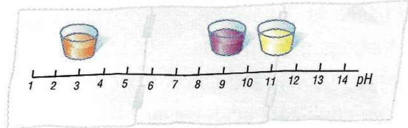


Figure 20.11
(a) Altering soil pH can affect the development of plants. In acidic soils, hydrangeas produce blue flowers. In basic soils, they produce pink flowers. (b) Spreading lime on lawns and gardens neutralizes acidic soils. (c) Blood pH is used to help diagnose illness.

MINI LAB

Indicators from Natural Sources



PURPOSE

To measure the pH of various household materials by using a natural indicator to make an indicator chart.

MATERIALS

- red cabbage leaves
- white vinegar (CH_3COOH)
- household ammonia
- baking soda ($NaHCO_3$)
- 1-cup measure
- 10 clear plastic cups
- 2 jars
- tape
- knife
- pencil
- teaspoon
- ruler
- clean white cloth
- hot water
- 3 sheets of plain white paper
- medicine dropper
- clear plastic wrap
- rubber band
- various household items listed in Step 5

PROCEDURE

1. Put $\frac{1}{2}$ cup of finely chopped red cabbage leaves in a jar and add $\frac{1}{2}$ cup of hot water. Stir and crush the leaves with a spoon. Continue the extraction until the water is distinctly colored.
2. Strain the extract through a piece of cloth into a clean jar. This liquid is your natural indicator.
3. Tape three sheets of paper end to end. Draw a line along the center and label it at 5 cm intervals with the numbers 1 to 14. This is your pH scale.
4. Pour your indicator to about 1 cm depth into each of three plastic cups. To one cup, add several drops of vinegar, to the second add a pinch of baking soda, and to the third add several drops of ammonia. The resulting colors indicate pH values of about 3, 9, and 11, respectively. Place these colored positions on your pH scale.
5. Repeat Step 4 for household items such as table salt, borax, milk, lemon juice, laundry detergent, dish detergent, milk of magnesia, mouthwash, toothpaste, shampoo, and carbonated beverages.

ANALYSIS AND CONCLUSIONS

1. What was the color of the indicator at acidic, neutral, and basic conditions?
2. What chemical changes were responsible for the color changes?
3. Label the materials you tested as acidic, basic, or neutral.
4. Which group contains items used for cleaning or for personal hygiene?

section review 20.2

16. What is true about the relative concentrations of hydrogen ions and hydroxide ions in each kind of solution?

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

17. Determine the pH of each solution.

a. $[H^+] = 1.0 \times 10^{-6}M$ **d.** $[OH^-] = 1.0 \times 10^{-11}M$

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

c. $[OH^-] = 1.0 \times 10^{-2}M$

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

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

Chem ASAP! Assessment 20.2 Check your understanding of the important ideas and concepts in Section 20.2.

Chapter 20 STUDENT STUDY GUIDE

Take It to the NET
For interactive study and review, go to www.phschool.com

KEY TERMS

- acid *p. 578*
- acid dissociation constant (K_a) *p. 601*
- acidic solution *p. 581*
- alkaline solution *p. 581*
- amphoteric *p. 598*
- base *p. 579*
- base dissociation constant (K_b) *p. 603*
- basic solution *p. 581*
- conjugate acid *p. 597*
- conjugate acid–base pair *p. 597*
- conjugate base *p. 597*
- diprotic acid *p. 594*
- hydrogen-ion acceptor *p. 596*
- hydrogen-ion donor *p. 596*
- hydronium ion (H_3O^+) *p. 580*
- hydroxide ion (OH^-) *p. 580*
- ion-product constant for water (K_w) *p. 581*
- Lewis acid *p. 598*
- Lewis base *p. 598*
- monoprotic acid *p. 594*
- neutral solution *p. 580*
- pH *p. 582*
- self-ionization *p. 580*
- strong acid *p. 600*
- strong base *p. 602*
- triprotic acid p. *594*
- weak acid *p. 600*
- weak base *p. 602*

KEY EQUATIONS AND RELATIONSHIPS

- $K_w = [H^+] \times [OH^-] = 1.0 \times 10^{-14} M^2$
- $pH = -\log [H^+]$
- $pOH = -\log [OH^-]$
- $pH + pOH = 14$

CONCEPT SUMMARY

20.1 Describing Acids and Bases

- Acids taste sour, are electrolytes, react with active metals to produce hydrogen, react with bases to form water and salts, and cause indicators to change color.
- Bases taste bitter, are electrolytes, react with acids to form water and salts, and cause indicators to change color.

20.2 Hydrogen Ions and Acidity

- Water molecules dissociate into hydrogen ions (H^+) and hydroxide ions (OH^-).
- On the pH scale, 0 is strongly acidic, 14 is strongly basic, and 7 is neutral. Pure water at 25 °C has a pH of 7.

20.3 Acid–Base Theories

- A Brønsted-Lowry acid is a proton donor, and a Brønsted-Lowry base is a proton acceptor.
- An Arrhenius acid yields hydrogen ions in aqueous solution. An Arrhenius base yields hydroxide ions in aqueous solution.
- A Lewis acid is an electron-pair acceptor, and a Lewis base is an electron-pair donor.
- A conjugate acid–base pair consists of two substances related by the loss or gain of a single hydrogen ion.

20.4 Strengths of Acids and Bases

- The strength of an acid or a base is determined by its degree of ionization in solution. The acid dissociation constant (K_a) is a quantitative measure of acid strength.
- The base dissociation constant (K_b) is a quantitative measure of base strength.

CHAPTER CONCEPT MAP

Use these terms to construct a concept map that organizes the major ideas of this chapter.

Arrhenius	strong base	Brønsted-Lowry	Lewis
acid	strong acid	hydrogen ion	pH
weak acid	base	weak base	theory

Chem ASAP! Concept Map 20
Create your Concept Map using the computer.

Chapter 20 REVIEW

CONCEPT PRACTICE

34. Write formulas for these compounds. *20.1*
   a. nitrous acid
   b. aluminum hydroxide
   c. hydroselenic acid
   d. strontium hydroxide
   e. phosphoric acid
   f. ethanoic acid

35. Write an equation showing the ionization of water. *20.2*

36. What are the concentrations of H^+ and OH^- in pure water at 25 °C? *20.2*

37. How is the pH of a solution calculated? *20.2*

38. Why is the pH of pure water at 25 °C equal to 7.00? *20.2*

39. Calculate the pH for the following solutions and indicate whether each solution is acidic or basic. *20.2*
   a. $[H^+] = 1.0 \times 10^{-2} M$
   b. $[OH^-] = 1.0 \times 10^{-2} M$
   c. $[OH^-] = 1.0 \times 10^{-8} M$
   d. $[H^+] = 1.0 \times 10^{-6} M$

40. What are the hydroxide-ion concentrations for solutions with the following pH values? *20.2*
   a. 4.00
   b. 8.00
   c. 12.00

41. Calculate the pH or $[H^+]$ for each solution. *20.2*
   a. $[H^+] = 2.4 \times 10^{-6} M$
   b. $[H^+] = 9.1 \times 10^{-9} M$
   c. pH = 13.20
   d. pH = 6.70

42. A soft drink has a pH of 3.80. What is the hydrogen-ion concentration in the drink? *20.2*

43. Write the reaction for the dissociation of each compound in water. *20.3*
   a. potassium hydroxide
   b. magnesium hydroxide

44. How did Arrhenius describe acids and bases? *20.3*

45. Classify each as an Arrhenius acid or an Arrhenius base. *20.3*
   a. $Ca(OH)_2$
   b. HNO_3
   c. KOH
   d. C_2H_5COOH
   e. HBr
   f. H_2SO_4

46. Identify each acid in Problem 45 as monoprotic, diprotic, or triprotic. *20.3*

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

1. 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.

3. Which species is the conjugate base of an ammonium ion, NH_4^+?
 a. H_2O **c.** NH_3
 b. OH^- **d.** H_3O^+

4. A solution with a hydrogen ion concentration of $2.3 \times 10^{-8} M$ has a pH between
 a. 2 and 3. **c.** 7 and 8.
 b. 3 and 4. **d.** 8 and 9.

Use the description and data table to answer questions 5–8.

Ethanoic acid (acetic acid) is a weak acid with a K_a of 1.8×10^{-5} at room temperature. A student measured the pH of ethanoic acid solutions at different concentrations. She used the pH values to calculate the percent of ethanoic acid molecules that were ionized in each solution.

Concentration of ethanoic acid (*M*)	Percent of molecules ionized
0.050	1.9
0.040	2.1
0.030	2.5
0.020	3.1
0.010	4.0

5. Write the equilibrium equation for the ionization of ethanoic acid in water.

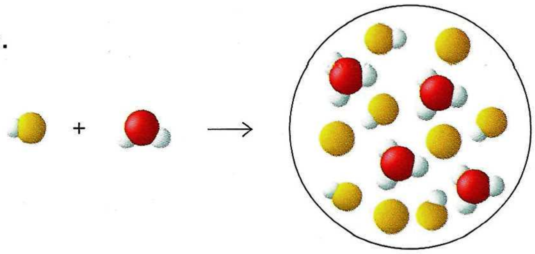
6. Use the data in the table to construct a graph. Make the molarity of ethanoic acid the independent variable.

7. Does the relationship between the variables make sense in terms of Le Châtelier's principle? Explain your answer.

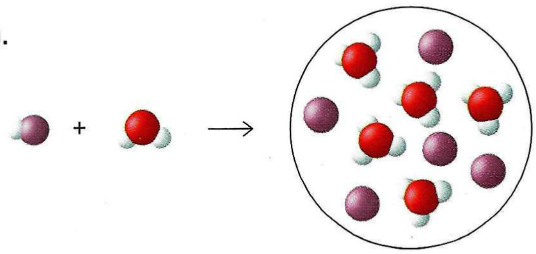
8. Use the graph to estimate the percent of molecules ionized at a concentration of 0.015*M*.

Use the drawings to answer questions 9–11. They show what happens when acids form aqueous solutions. Water molecules have been omitted from the solution windows.

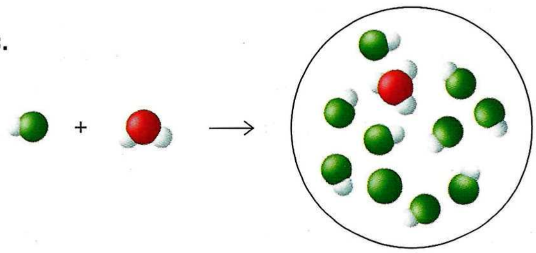
a.



b.



c.



9. Put the acids in order of increasing strength.

10. How many of the acids are strong acids?

11. Which acid is dissociated about 10%?

Write a few sentences in response to each of the statements in questions 12–14.

12. Indicators such as methyl red provide accurate and precise measurements of pH.

13. According to the Arrhenius definition of acids and bases, ammonia qualifies as a base.

14. The strength of an acid or base changes as its concentration changes.