

# *Acids and Bases*

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*Acids and bases change the color of compounds called indicators.*

# Properties of Acids and Bases

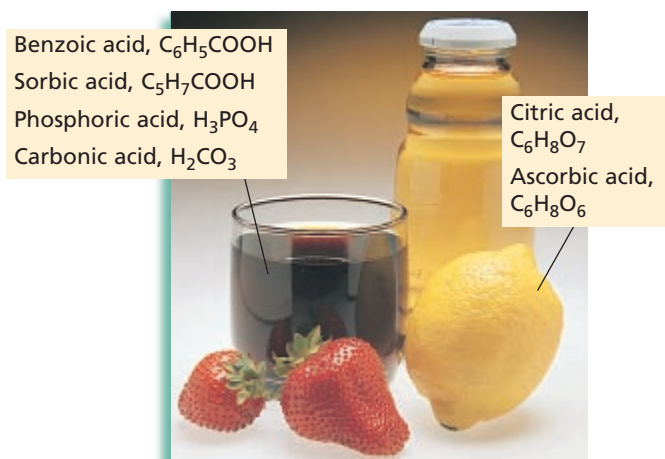
## SECTION 15-1

### OBJECTIVES

- List five general properties of aqueous acids and bases.
- Name common binary acids and oxyacids, given their chemical formulas.
- List five acids commonly used in industry and the laboratory, and give two properties of each.
- Define *acid* and *base* according to Arrhenius's theory of ionization.
- Explain the differences between strong and weak acids and bases.

**H**ow many foods can you think of that are sour? Chances are that almost all the foods you thought of, like those in Figure 15-1(a), owe their sour taste to an acid. Sour milk contains *lactic acid*. Vinegar, which can be produced by fermenting juices, contains *acetic acid*. *Phosphoric acid* gives a tart flavor to many carbonated beverages. Most fruits contain some kind of acid. Lemons, oranges, grapefruits, and other citrus fruits contain *citric acid*. Apples contain *malic acid*, and grape juice contains *tartaric acid*.

Many substances known as bases are commonly found in household products, such as those in Figure 15-1(b). Household ammonia is an ammonia-water solution that is useful for all types of general cleaning. Sodium hydroxide, NaOH, known by the common name *lye*, is present in some commercial drain and oven cleaners. Milk of magnesia is a suspension in water of magnesium hydroxide,  $\text{Mg}(\text{OH})_2$ , which is not very water-soluble. It is used as an antacid to relieve discomfort caused by excess hydrochloric acid in the stomach. Aluminum hydroxide,  $\text{Al}(\text{OH})_3$ , and sodium hydrogen carbonate,  $\text{NaHCO}_3$ , are also bases commonly found in antacids.



(a)



(b)

**FIGURE 15-1** (a) Fruits and fruit juices contain acids such as citric acid and ascorbic acid. Carbonated beverages contain benzoic acid, phosphoric acid, and carbonic acid. (b) Many household cleaners contain bases such as ammonia and sodium hydroxide.

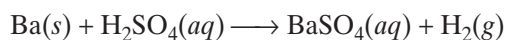


**FIGURE 15-2** A strip of pH paper dipped into vinegar turns red, showing that vinegar is an acid.

## Acids

Acids were first recognized as a distinct class of compounds because of the common properties of their aqueous solutions. These properties are listed below.

1. *Aqueous solutions of acids have a sour taste.* Taste, however, should NEVER be used as a test to evaluate any chemical substance. Many acids, especially in concentrated solutions, are corrosive; that is, they destroy body tissue and clothing. Many are also poisons.
2. *Acids change the color of acid-base indicators.* When pH paper is used as an indicator, the paper turns certain colors in acidic solution. This reaction is demonstrated in Figure 15-2.
3. *Some acids react with active metals to release hydrogen gas,  $H_2$ .* Recall that metals can be ordered in terms of an activity series. Metals above hydrogen in the series undergo single-replacement reactions with certain acids. Hydrogen gas is formed as a product, as shown by the reaction of barium with sulfuric acid.



4. *Acids react with bases to produce salts and water.* When chemically equivalent amounts of acids and bases react, the three properties just described disappear because the acid is “neutralized.” The reaction products are water and an ionic compound called a *salt*.
5. *Some acids conduct electric current.* Acids that form many ions in aqueous solution are electrolytes.

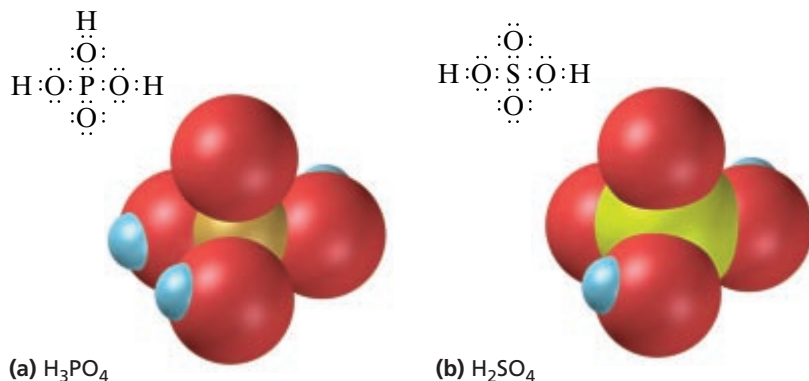
## Acid Nomenclature

A **binary acid** is an acid that contains only two different elements: hydrogen and one of the more electronegative elements. Many common inorganic acids are binary acids. The hydrogen halides—HF, HCl, HBr, and HI—are all binary acids.

The procedure used to name binary acids is illustrated by the examples given in Table 15-1. In pure form, each acid listed in the table is a gas. Aqueous solutions of these compounds are known by the acid names. From the table you can see that naming binary compounds can be summarized as follows.

**TABLE 15-1** Names of Binary Acids

Formula	Acid name
HF	hydrofluoric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydriodic acid
H <sub>2</sub> S	hydrosulfuric acid



**FIGURE 15-3** Structures of (a) phosphoric acid and (b) sulfuric acid

### Binary Acid Nomenclature

1. The name of a binary acid begins with the prefix *hydro-*.
2. The root of the name of the second element follows this prefix.
3. The name then ends with the suffix *-ic*.

An **oxyacid** is an acid that is a compound of hydrogen, oxygen, and a third element, usually a nonmetal. Nitric acid,  $\text{HNO}_3$ , is an oxyacid. The structures of two other oxyacids are shown in Figure 15-3. Oxyacids are one class of ternary acids, which are acids that contain three different elements. Usually, the elements in an oxyacid formula are written as one or more hydrogen atoms followed by a polyatomic anion. The name of an oxyacid is based on this anion. Some common oxyacids and their anions are given in Table 15-2.

**TABLE 15-2** Names of Common Oxyacids and Oxyanions

Formula	Acid name	Anion
$\text{CH}_3\text{COOH}$	acetic acid	$\text{CH}_3\text{COO}^-$ , acetate
$\text{H}_2\text{CO}_3$	carbonic acid	$\text{CO}_3^{2-}$ , carbonate
$\text{HIO}_3$	iodic acid	$\text{IO}_3^-$ , iodate
$\text{HClO}$	hypochlorous acid	$\text{ClO}^-$ , hypochlorite
$\text{HClO}_2$	chlorous acid	$\text{ClO}_2^-$ , chlorite
$\text{HClO}_3$	chloric acid	$\text{ClO}_3^-$ , chlorate
$\text{HClO}_4$	perchloric acid	$\text{ClO}_4^-$ , perchlorate
$\text{HNO}_2$	nitrous acid	$\text{NO}_2^-$ , nitrite
$\text{HNO}_3$	nitric acid	$\text{NO}_3^-$ , nitrate
$\text{H}_3\text{PO}_3$	phosphorous acid	$\text{PO}_3^{3-}$ , phosphite
$\text{H}_3\text{PO}_4$	phosphoric acid	$\text{PO}_4^{3-}$ , phosphate
$\text{H}_2\text{SO}_3$	sulfurous acid	$\text{SO}_3^{2-}$ , sulfite
$\text{H}_2\text{SO}_4$	sulfuric acid	$\text{SO}_4^{2-}$ , sulfate



## Some Common Industrial Acids

The properties of acids make them important chemicals both in the laboratory and in industry. Sulfuric acid, nitric acid, phosphoric acid, hydrochloric acid, and acetic acid are all common industrial acids.

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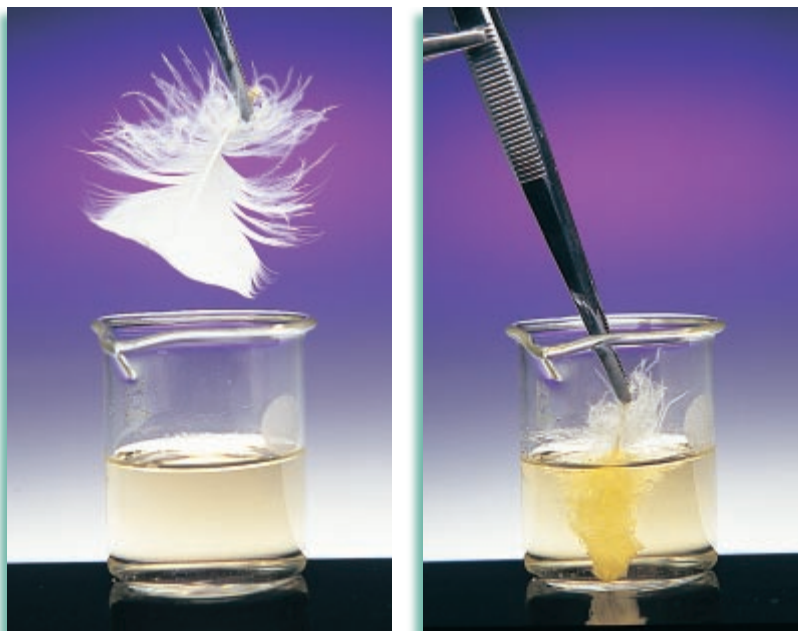
### Sulfuric Acid

Sulfuric acid is the most commonly produced industrial chemical in the world. More than 47 million tons of it are made each year in the United States alone. Sulfuric acid is used in large quantities in petroleum refining and metallurgy as well as in the manufacture of fertilizer. It is also essential to a vast number of industrial processes, including the production of metals, paper, paint, dyes, detergents, and many chemical raw materials. Sulfuric acid is the acid used in automobile batteries.

Because it attracts water, concentrated sulfuric acid is an effective dehydration (water-removing) agent. It can be used to remove water from gases with which it does not react. Sugar and certain other organic compounds are also dehydrated by sulfuric acid. Skin contains organic compounds that are attacked by concentrated sulfuric acid, which can cause serious burns.

### Nitric Acid

Pure nitric acid is a volatile, unstable liquid rarely used in industry or laboratories. Dissolving the acid in water provides stability. Nitric acid stains proteins yellow. The feather in Figure 15-4 was stained by nitric acid. The acid has a suffocating odor, stains skin, and can cause serious burns. It is used in making explosives, many of which are nitrogen-containing compounds. It is also used to make rubber, plastics, dyes, and



**FIGURE 15-4** Concentrated nitric acid stains a feather yellow.

pharmaceuticals. Initially, nitric acid solutions are colorless; however, upon standing, they gradually become yellow because of slight decomposition to brown nitrogen dioxide gas.

## Phosphoric Acid

Phosphorus, along with nitrogen and potassium, is an essential element for plants and animals. The bulk of phosphoric acid produced each year is used directly for manufacturing fertilizers and animal feed. Dilute phosphoric acid has a pleasant but sour taste and is not toxic. It is used as a flavoring agent in beverages and as a cleaning agent for dairy equipment. Phosphoric acid is also important in the manufacture of detergents and ceramics.

## Hydrochloric Acid

The stomach produces HCl to aid in digestion. Industrially, hydrochloric acid is important for “pickling” iron and steel. Pickling is the immersion of metals in acid solutions to remove surface impurities. This acid is also used in industry as a general cleaning agent, in food processing, in the activation of oil wells, in the recovery of magnesium from sea water, and in the production of other chemicals.

A dilute solution of hydrochloric acid, commonly referred to as muriatic acid, may be found in hardware stores. It is used to maintain the correct acidity in swimming pools and for general cleaning of masonry.

## Acetic Acid

Concentrated acetic acid is a clear, colorless, pungent-smelling liquid known as glacial acetic acid. This name derives from the fact that pure acetic acid has a freezing point of only 17°C. It can form crystals in a cold room. The fermentation of certain plants produces vinegars containing acetic acid. White vinegar contains 4–8% acetic acid.


Acetic acid is important industrially in synthesizing chemicals used in the manufacture of plastics. It is a raw material in the production of food supplements—for example, lysine, an essential amino acid. Acetic acid is also used as a fungicide.

## Bases

How do bases differ from acids? You can answer this question by comparing the following properties of bases with those of acids.

1. *Aqueous solutions of bases taste bitter.* You may have noticed this fact if you have ever gotten soap, a basic substance, in your mouth. As with acids, taste should NEVER be used to test a substance to see if it is a base. Many bases are caustic; they attack the skin and tissues, causing severe burns.
2. *Bases change the color of acid-base indicators.* As Figure 15-5 shows, an indicator will be a different color in a basic solution than it would be in an acidic solution.

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**FIGURE 15-5** pH paper turns blue in the presence of this solution of sodium hydroxide.



Wear Safety Goggles and an Apron.



## Materials

- dishwashing liquid, dishwasher detergent, laundry detergent, laundry stain remover, fabric softener, and bleach
- fresh red cabbage
- hot plate
- beaker, 500 mL or larger
- beakers, 50 mL
- mayonnaise, baking powder, baking soda, white vinegar, cider vinegar, lemon juice, soft drinks, mineral water, and milk
- spatula
- tap water
- tongs

# Household Acids and Bases

## Question

Which of the household substances are acids, and which are bases?

## Procedure

Record all your results in a data table.

1. To make an acid-base indicator, extract juice from red cabbage. First, cut up some red cabbage and place it in a large beaker. Add enough water so that the beaker is half full. Then bring the mixture to a boil. Let it cool, and pour off the cabbage juice. Save the solution.
2. Assemble foods, beverages, and cleaning products to be tested.
3. If the substance being tested is a liquid, pour about 5 mL into a small beaker. If it is a solid, place a small amount into a beaker, and moisten it with about 5 mL of water.
4. Add a drop or two of the red cabbage juice to the solution being tested, and note the color. The solution will turn red if it is acidic and green if it is basic.



Red cabbage can be made into an acid-base indicator.

## Discussion

1. Are the cleaning products acids, bases, or neither?
  2. What are acid/base characteristics of foods and beverages?
  3. Did you find consumer warning labels on basic or acidic products?
- 
3. *Dilute aqueous solutions of bases feel slippery.* You encounter this property of aqueous bases whenever you wash with soap.
  4. *Bases react with acids to produce salts and water.* The properties of an acid disappear with the addition of an equivalent amount of a base. It could also be said that “neutralization” of the base occurs when these two substances react to produce salts and water.
  5. *Bases conduct electric current.* Like acids, bases form ions in aqueous solutions and are thus electrolytes.

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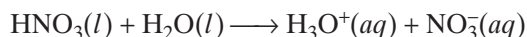
## Arrhenius Acids and Bases

Svante Arrhenius, a Swedish chemist who lived from 1859 to 1927, understood that aqueous solutions of acids and bases conducted electric current. Arrhenius therefore theorized that acids and bases must produce ions in solution. An **Arrhenius acid** is a chemical compound that increases the concentration of hydrogen ions,  $H^+$ , in aqueous solution. In other words, an acid will ionize in solution, increasing the number of hydrogen ions present. An **Arrhenius base** is a substance that increases the concentration of hydroxide ions,  $OH^-$ , in aqueous solution. Some bases are ionic hydroxides. These bases dissociate in solution to release hydroxide ions into the solution. Other bases are substances that react with water to remove a hydrogen ion, leaving hydroxide ions in the solution.

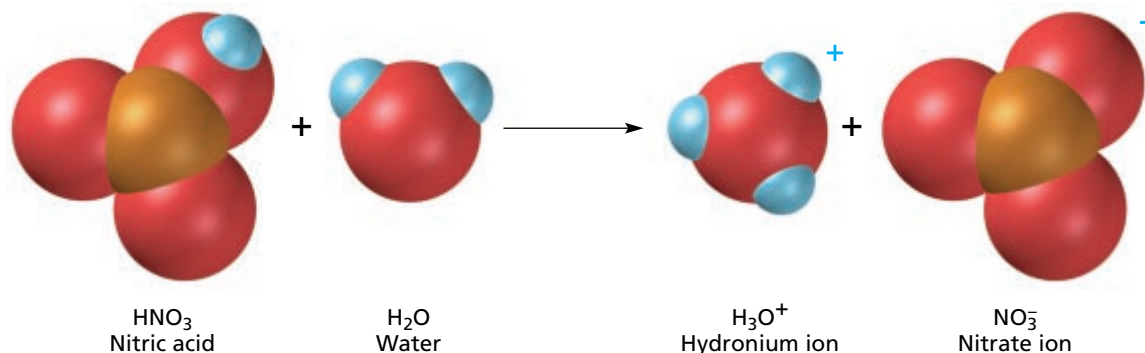
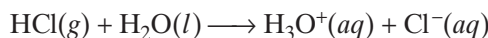
### Aqueous Solutions of Acids

The acids described by Arrhenius are molecular compounds with ionizable hydrogen atoms. Their water solutions are known as *aqueous acids*. All pure aqueous acids are electrolytes.

Acid molecules are sufficiently polar so that one or more hydrogen ions are attracted by water molecules. Negatively charged anions are left behind. As explained in Chapter 14, the hydrogen ion in aqueous solution is best represented as  $H_3O^+$ , the hydronium ion. The ionization of an  $HNO_3$  molecule is shown by the following equation. Figure 15-6 shows how the hydrogen atoms combine with water to form a hydronium ion when nitric acid is diluted.



Similarly, ionization of a hydrogen chloride molecule in hydrochloric acid can be represented in the following way.



**FIGURE 15-6** Arrhenius's observations form the basis of a definition of acids. Arrhenius acids, such as the nitric acid shown here, produce hydronium ions in aqueous solution.



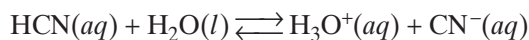
**TABLE 15-3 Common Aqueous Acids**

Strong acids	Weak acids
$\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-$	$\text{HSO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{SO}_4^{2-}$
$\text{HClO}_4 + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+ + \text{ClO}_4^-$	$\text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{H}_2\text{PO}_4^-$
$\text{HCl} + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$	$\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-$
$\text{HNO}_3 + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+ + \text{NO}_3^-$	$\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^-$
$\text{HBr} + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+ + \text{Br}^-$	$\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HCO}_3^-$
$\text{HI} + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+ + \text{I}^-$	$\text{H}_2\text{S} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HS}^-$
	$\text{HCN} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CN}^-$
	$\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CO}_3^{2-}$

## Strength of Acids

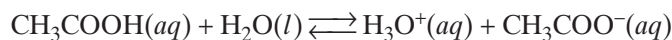
A **strong acid** is one that ionizes completely in aqueous solution. A strong acid is a strong electrolyte. Perchloric acid,  $\text{HClO}_4$ , hydrochloric acid,  $\text{HCl}$ , and nitric acid,  $\text{HNO}_3$ , are examples of strong acids. The strength of an acid depends on the polarity of the bond between hydrogen and the element to which it is bonded and the ease with which that bond can be broken. Acid strength increases with increasing polarity and decreasing bond energy.

Acids that are weak electrolytes are known as **weak acids**. The aqueous solution of a weak acid contains hydronium ions, anions, and dissolved acid molecules. Hydrocyanic acid is an example of a weak electrolyte. In aqueous solution, both the ionization of  $\text{HCN}$  and the reverse reaction occur simultaneously. Although hydronium and cyanide ions are present in solution, the reverse reaction is favored. Most of the solution is composed of hydrogen cyanide and water.



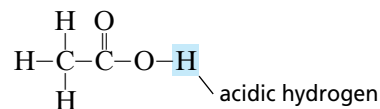
Common aqueous acids are listed in Table 15-3. Each strong acid is assumed to ionize completely to give up one hydrogen ion. Notice that the number of hydrogen atoms in the formula does not affect acid strength. Molecules with multiple hydrogen atoms may not readily give up each hydrogen. The fact that phosphoric acid has three hydrogen atoms per molecule does not mean that it is a strong acid. None of these ionize completely in solution, so phosphoric acid is weak.

*Organic acids*, which contain the acidic carboxyl group  $-\text{COOH}$ , are generally weak acids. For example, acetic acid,  $\text{CH}_3\text{COOH}$ , ionizes slightly in water to give hydronium ions and acetate ions,  $\text{CH}_3\text{COO}^-$ .



A molecule of acetic acid contains four hydrogen atoms. However, only one of the hydrogen atoms is ionizable. The hydrogen atom in the

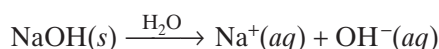
carboxyl group in acetic acid is the one that is “acidic” and forms the hydronium ion. This acidic hydrogen can be seen in the structural diagram in Figure 15-7.



**FIGURE 15-7** Acetic acid contains four hydrogen atoms, but only one of them is “acidic” or forms the hydronium ion in solution.

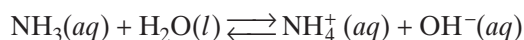
## Aqueous Solutions of Bases

Most bases are ionic compounds containing metal cations and the hydroxide anion,  $\text{OH}^-$ . Because these bases are ionic, they dissociate to some extent when placed in solution. *When a base completely dissociates in water to yield aqueous  $\text{OH}^-$  ions, the solution is referred to as alkaline.* Sodium hydroxide,  $\text{NaOH}$ , is a common laboratory base. It is water-soluble and dissociates as shown by the equation below.



You will remember from Chapter 5 that sodium is one of the alkali metals. This group gets its name from the fact that the hydroxides of Li, Na, K, Rb, and Cs all form alkaline solutions.

Not all bases are ionic compounds. A base commonly used in household cleaners is ammonia,  $\text{NH}_3$ , which is molecular. Ammonia is a base because it produces hydroxide ions when it reacts with water molecules, as shown in the equation below.



## Strength of Bases

As with acids, the strength of a base also depends on the extent to which the base dissociates, or adds hydroxide ions to the solution. For example, potassium hydroxide,  $\text{KOH}$ , is a strong base because it completely dissociates into its ions in dilute aqueous solutions.

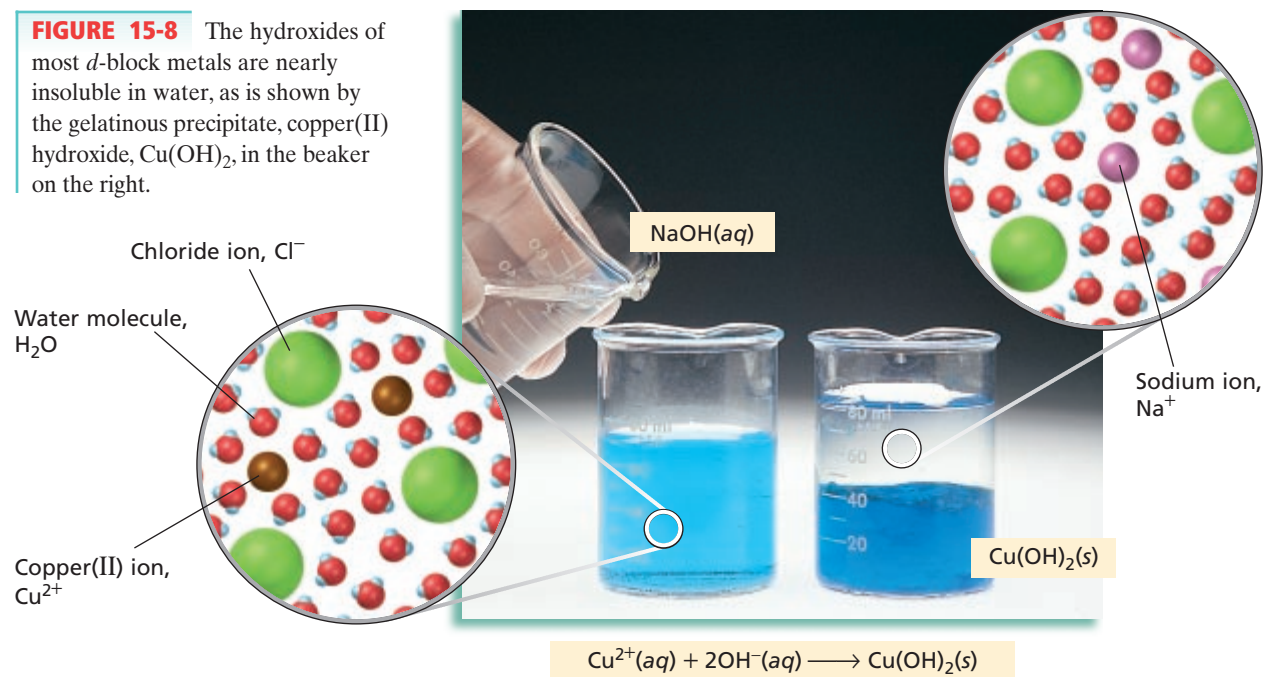


*Strong bases* are strong electrolytes, just as strong acids are strong electrolytes. Table 15-4 lists some strong bases.

**TABLE 15-4 Common Aqueous Bases**

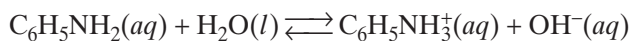
Strong bases	Weak bases
$\text{Ca}(\text{OH})_2 \longrightarrow \text{Ca}^{2+} + 2\text{OH}^-$	$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$
$\text{Sr}(\text{OH})_2 \longrightarrow \text{Sr}^{2+} + 2\text{OH}^-$	$\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+ + \text{OH}^-$
$\text{Ba}(\text{OH})_2 \longrightarrow \text{Ba}^{2+} + 2\text{OH}^-$	
$\text{NaOH} \longrightarrow \text{Na}^+ + \text{OH}^-$	
$\text{KOH} \longrightarrow \text{K}^+ + \text{OH}^-$	
$\text{RbOH} \longrightarrow \text{Rb}^+ + \text{OH}^-$	
$\text{CsOH} \longrightarrow \text{Cs}^+ + \text{OH}^-$	

**FIGURE 15-8** The hydroxides of most *d*-block metals are nearly insoluble in water, as is shown by the gelatinous precipitate, copper(II) hydroxide,  $\text{Cu}(\text{OH})_2$ , in the beaker on the right.



Bases that are not very soluble do not produce a large number of hydroxide ions when added to water. Some metal hydroxides, such as  $\text{Cu}(\text{OH})_2$ , are not very soluble in water, as seen in Figure 15-8. They cannot produce strongly alkaline solutions. The alkalinity of aqueous solutions depends on the concentration of  $\text{OH}^-$  ions in solution. It is unrelated to the number of hydroxide ions in the undissolved compound.

Now consider ammonia, which is highly soluble but is a weak electrolyte. The concentration of  $\text{OH}^-$  ions in an ammonia solution is relatively low. Ammonia is therefore a *weak base*. Many organic compounds that contain nitrogen atoms are also weak bases. For example, aniline, a substance used to make dyes, is a weak base.



## SECTION REVIEW

- What are five general properties of aqueous acids?
  - Name some common substances that have one or more of these properties.
- Name the following acids: a.  $\text{HBrO}$  b.  $\text{HBrO}_3$ .
- What are five general properties of aqueous bases?
  - Name some common substances that have one or more of these properties.
- Why are strong acids also strong electrolytes?
  - Is every strong electrolyte also a strong acid?



## Logic in the Laboratory

From "Acid and Water: A Socratic Dialogue," by David Todd, in *The Journal of Chemical Education*.

**Tutor:** . . . tell me, how does one set about diluting an acid with water?

**Student:** The rule is: pour the acid into the water.

**Tutor:** Why so? . . .

**Student:** I believe much heat is given off if you do it the wrong way, and the mixture can boil up in your face.

**Tutor:** Indeed, that is so . . . But is there not heat also developed if you pour the acid into the water?

**Student:** (thoughtfully) I suppose so. But then I can only assume since there is a rule, that the heat developed is a lot less if you do it that way.

**Tutor:** Let us reason together. Have you ever heard of Hess's Law?

**Student:** . . . Doesn't it have something to do with A going to B and the heat change involved?

**Tutor:** Very good . . . Now suppose I start with 100 g of water in one container at 25°C, and 100 g of concentrated sulfuric acid at 25°C in another container. We can call these two items your A . . . Now let us assume that the two have been mixed—regardless of the mode of mixing—do we not obtain a diluted acid that is 100 g of acid and 100 g of water, and we can call this B?

**Student:** (cautiously) Well—almost. But . . . concentrated sulfuric acid is 96% by weight  $\text{H}_2\text{SO}_4$

and 4% water. This means that B is 96 g of pure sulfuric acid and 104 g of water.

**Tutor:** . . . Oh, excellent! You have had good teachers . . . But the main point is that B has the same composition regardless of the route by which it is obtained. Agreed?

**Student:** That most assuredly must be so.

**Tutor:** Now let us return to Hess's Law. It states that the heat change involved in going from A to B is the same regardless of the path taken.

**Student:** . . . Yes—I remember it now. You mean this law says that the same heat is evolved (or absorbed) if I pour the acid into the water, or vice versa?

**Tutor:** Yes, indeed.

**Student:** (now bewildered) You mean that the rule is nonsense, and therefore useless?

**Tutor:** Oh no, not at all! . . .

**Student:** . . . Hmm. Then in that case there must be some other factor involved.

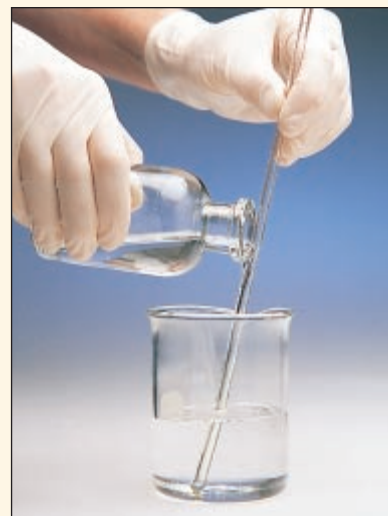
**Tutor:** Indeed there is . . . May I drop a hint? . . . Which is the more dense—concentrated sulfuric acid or water?

**Student:** The acid . . .

**Tutor:** Good. Now if I put water on the acid, does it float or sink?

**Student:** Of course it will stay on the top.

**Tutor:** Right. And it will begin to mix . . .



*Always dilute by pouring acid into water.*

**Student:** I get it—it reacts on the surface, generates a lot of heat, and some of the diluted acid can boil up in my face.

**Tutor:** Exactly. And if I pour the acid into . . .

**Student:** (interrupting) Yes, yes—of course. The acid falls down through the water generating the heat in the entire body of the liquid—not just on the surface. So it won't form steam and boil up in my face . . . now I see the reason for the rule.

### Reading for Meaning

In your own words, sum up the reason behind the rule for diluting acid with water.

### Read Further

Hydrofluoric acid is a fairly weak acid. It does not burn the skin the way sulfuric acid and other strong acids do. Find an explanation of why this acid can greatly damage body tissue when it comes into contact with the skin.



## OBJECTIVES

- Define and recognize *Brønsted-Lowry acids* and *bases*.
- Define a *Lewis acid* and a *Lewis base*.
- Name compounds that are acids under the Lewis definition but are not acids under the Brønsted-Lowry definition.



**Module 8:** Strong and Weakly Ionized Species, pH, and Titrations

# Acid-Base Theories

For most uses, scientists found the Arrhenius definition of acids and bases to be adequate. However, as scientists further investigated acid-base behavior, they found that some substances acted as acids or bases when they were not in a water solution. Because the Arrhenius definition requires that the substances be aqueous, the definitions of acids and bases had to be revised.

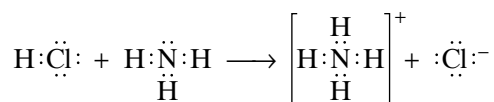
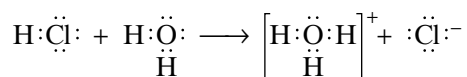
## Brønsted-Lowry Acids and Bases

In 1923, the Danish chemist J. N. Brønsted and the English chemist T. M. Lowry independently expanded the Arrhenius acid definition. A **Brønsted-Lowry acid** is a molecule or ion that is a proton donor. Because  $\text{H}^+$  is a proton, all acids as defined by Arrhenius donate protons to water and are Brønsted-Lowry acids as well. Substances other than molecules, such as certain ions, can also donate protons. Such substances are not Arrhenius acids but are included in the category of Brønsted-Lowry acids.

Hydrogen chloride acts as a Brønsted-Lowry acid when it is dissolved in ammonia. It transfers protons to the solvent much as it does in water.

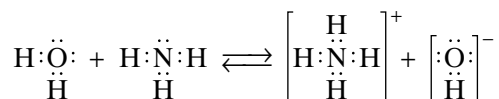
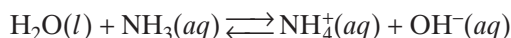


A proton is transferred from the hydrogen chloride molecule, HCl, to the ammonia molecule,  $\text{NH}_3$ . The ammonium ion,  $\text{NH}_4^+$ , is formed. Electron-dot formulas show the similarity of this reaction to the reaction of HCl with water.



In both reactions, hydrogen chloride is a Brønsted-Lowry acid.

Water can also act as a Brønsted-Lowry acid. Consider, for example, the following reaction, in which the water molecule donates a proton to the ammonia molecule.





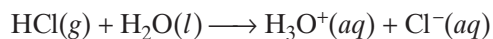
**FIGURE 15-9** Hydrogen chloride gas escapes from a hydrochloric acid solution and combines with ammonia gas that has escaped from an aqueous ammonia solution. The resulting cloud is solid ammonium chloride.

A **Brønsted-Lowry base** is a molecule or ion that is a proton acceptor. In the reaction between hydrochloric acid and ammonia, ammonia accepts a proton from the hydrochloric acid. It is a Brønsted-Lowry base. The Arrhenius hydroxide bases, such as NaOH, are not, strictly speaking, Brønsted-Lowry bases. That is because as compounds they are not proton acceptors. The  $\text{OH}^-$  ion produced in solution is the Brønsted-Lowry base. It is the species that can accept a proton.

In a **Brønsted-Lowry acid-base reaction**, protons are transferred from one reactant (the acid) to another (the base). Figure 15-9 shows the reaction between the Brønsted-Lowry acid HCl and the Brønsted-Lowry base  $\text{NH}_3$ .

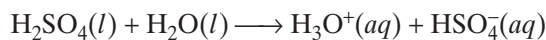
## Monoprotic and Polyprotic Acids

An acid that can donate only one proton (hydrogen ion) per molecule is known as a **monoprotic acid**. Perchloric acid,  $\text{HClO}_4$ , hydrochloric acid, HCl, and nitric acid,  $\text{HNO}_3$ , are all monoprotic. The following equation shows how a molecule of the monoprotic acid HCl donates a proton to a water molecule. HCl is completely ionized; it has no more hydrogen atoms to lose.

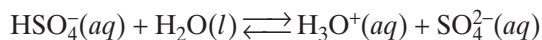


A **polyprotic acid** is an acid that can donate more than one proton per molecule. Sulfuric acid,  $\text{H}_2\text{SO}_4$ , and phosphoric acid,  $\text{H}_3\text{PO}_4$ , are examples of polyprotic acids. The ionization of a polyprotic acid occurs in stages. The acid loses its hydrogen atoms one at a time. Sulfuric acid

ionizes in two stages. In its first ionization, sulfuric acid is a strong acid. It is completely converted to hydrogen sulfate ions,  $\text{HSO}_4^-$ .



The hydrogen sulfate ion is itself a weak acid. It establishes the following equilibrium in solution.

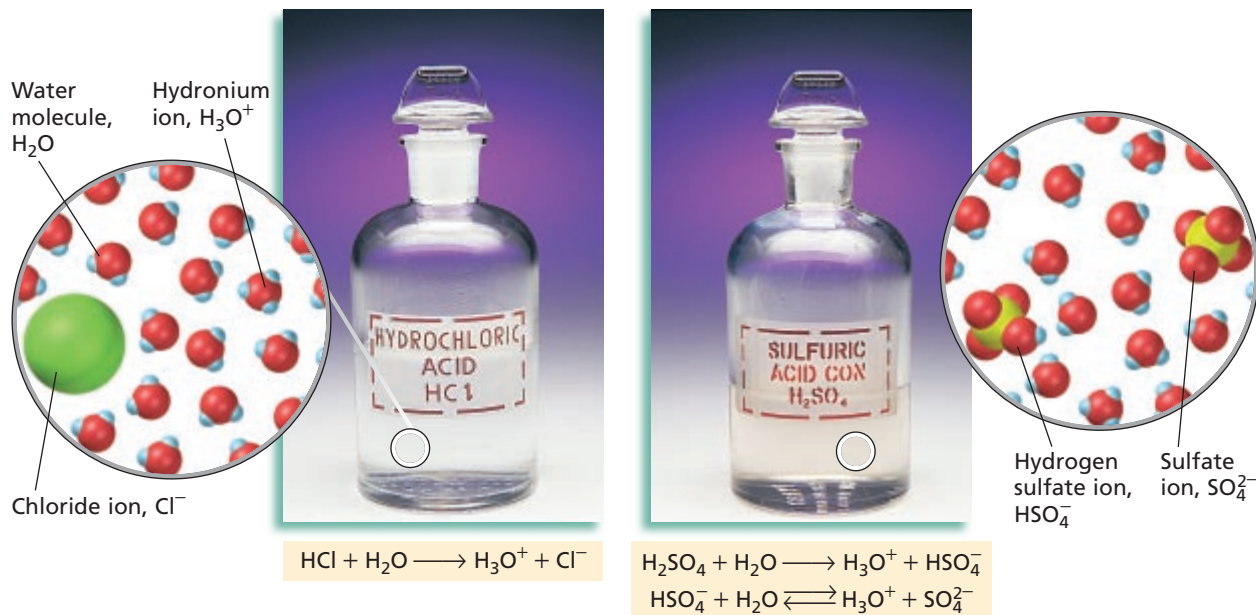
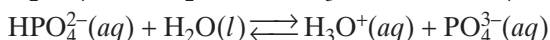
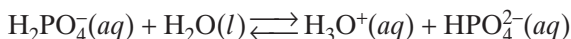
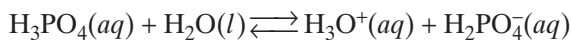


All stages of ionization of a polyprotic acid occur in the same solution. Sulfuric acid solutions therefore contain  $\text{H}_3\text{O}^+$ ,  $\text{HSO}_4^-$ , and  $\text{SO}_4^{2-}$  ions.

Sulfuric acid is the type of polyprotic acid that *can donate two protons per molecule*, and it is therefore known as a **diprotic acid**. Ionizations of a monoprotic acid and a diprotic acid are shown in Figure 15-10.

Phosphoric acid is the type of polyprotic acid known as a **triprotic acid**—an acid able to donate three protons per molecule. The equations for these reactions are shown below.

**FIGURE 15-10** Hydrochloric acid,  $\text{HCl}$ , is a strong monoprotic acid. A dilute  $\text{HCl}$  solution contains hydronium ions and chloride ions. Sulfuric acid,  $\text{H}_2\text{SO}_4$ , is a strong diprotic acid. A dilute  $\text{H}_2\text{SO}_4$  solution contains hydrogen sulfate ions from the first ionization, sulfate ions from the second ionization, and hydronium ions from both ionizations.



A solution of phosphoric acid contains  $\text{H}_3\text{O}^+$ ,  $\text{H}_3\text{PO}_4$ ,  $\text{H}_2\text{PO}_4^-$ ,  $\text{HPO}_4^{2-}$ , and  $\text{PO}_4^{3-}$ . As with most polyprotic acids, the concentration of ions formed in the first ionization is the greatest. There are lesser concentrations of the respective ions from each succeeding ionization. Phosphoric acid is a weak acid in each step of its ionization.

## Lewis Acids and Bases

The Arrhenius and Brønsted-Lowry definitions describe most acids and bases. Both definitions assume that the acid contains or produces hydrogen ions. A third acid classification, based on bonding and structure, includes, as acids, substances that do not contain hydrogen at all. This definition of acids was introduced in 1923 by G. N. Lewis, the American chemist whose name was given to electron-dot structures. Lewis's definition emphasizes the role of electron pairs in acid-base reactions. A **Lewis acid** is an atom, ion, or molecule that accepts an electron pair to form a covalent bond.

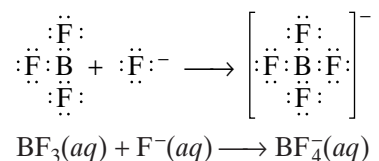
The Lewis definition is the broadest of the three acid definitions you have read about so far. It applies to any species that can accept an electron pair to form a covalent bond with another species. A bare proton (hydrogen ion) is a Lewis acid in reactions in which it forms a covalent bond, as shown below.



The formula for a Lewis acid need not include hydrogen. Even a silver ion can be a Lewis acid, accepting electron pairs from ammonia to form covalent bonds.



Any compound in which the central atom has three valence electrons and forms three covalent bonds can react as a Lewis acid. It does so by accepting a pair of electrons to form a fourth covalent bond, completing an electron octet. Boron trifluoride, for example, is an excellent Lewis acid. It forms a fourth covalent bond with many molecules and ions. Its reaction with a fluoride ion is shown below.

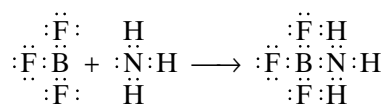




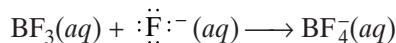
**TABLE 15-5 Acid-Base Systems**

Type	Acid	Base
Arrhenius	H <sup>+</sup> or H <sub>3</sub> O <sup>+</sup> producer	OH <sup>-</sup> producer
Brønsted-Lowry	proton (H <sup>+</sup> ) donor	proton (H <sup>+</sup> ) acceptor
Lewis	electron-pair acceptor	electron-pair donor

The Lewis definition of acids can apply to species in any phase. For example, boron trifluoride is a Lewis acid in the gas-phase combination with ammonia.



A **Lewis base** is an atom, ion, or molecule that donates an electron pair to form a covalent bond. An anion is a Lewis base in a reaction in which it forms a covalent bond by donating an electron pair. In the example of boron trifluoride reacting with the fluoride anion, F<sup>-</sup> donates an electron pair to boron trifluoride. F<sup>-</sup> acts as a Lewis base.



A **Lewis acid-base reaction** is the formation of one or more covalent bonds between an electron-pair donor and an electron-pair acceptor.

Note that although the three acid-base definitions differ, many compounds may be categorized as acids or bases according to all three descriptions. For example, ammonia is an Arrhenius base because OH<sup>-</sup> ions are created when ammonia is in solution, it is a Brønsted-Lowry base because it accepts a proton in an acid-base reaction, and it is a Lewis base in all reactions in which it donates its lone pair to form a covalent bond. A comparison of the three acid-base definitions is given in Table 15-5.

## SECTION REVIEW

- Label each reactant and each product in the reaction below as a proton donor or a proton acceptor and as acidic or basic.  

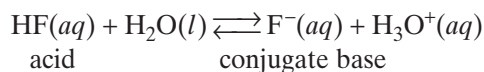
$$\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{HCO}_3^- + \text{H}_3\text{O}^+$$
- Consider the following three reactions. Identify the Arrhenius bases, Brønsted-Lowry bases, and Lewis bases in these reactions. Explain your answers.
  - $\text{NaOH}(s) \longrightarrow \text{Na}^+(aq) + \text{OH}^-(aq)$
  - $\text{HF}(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{F}^-(aq) + \text{H}_3\text{O}^+(aq)$
  - $\text{H}^+(aq) + \text{NH}_3(aq) \longrightarrow \text{NH}_4^+(aq)$

# Acid-Base Reactions

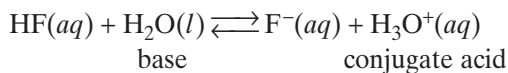
**I**n the previous section, you learned about three acid-base theories: Arrhenius, Brønsted-Lowry, and Lewis. The Brønsted-Lowry theory is especially useful for describing acid-base reactions that take place in aqueous solutions. This section will use the Brønsted-Lowry description to explore reactions between acids and bases.

## Conjugate Acids and Bases

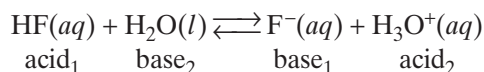
The Brønsted-Lowry definitions of acids and bases provide a basis for studying proton-transfer reactions. Suppose that a Brønsted-Lowry acid gives up a proton; the remaining ion or molecule can re-accept that proton and can act as a base. Such a base is known as a conjugate base. Thus, *the species that remains after a Brønsted-Lowry acid has given up a proton is the **conjugate base** of that acid.* For example, the fluoride ion is the conjugate base of hydrogen fluoride.



In this reaction, the water molecule is a Brønsted-Lowry base. It accepts a proton to form  $\text{H}_3\text{O}^+$ , which is an acid. The hydronium ion is the conjugate acid of water. *The species that is formed when a Brønsted-Lowry base gains a proton is the **conjugate acid** of that base.*



In general, Brønsted-Lowry acid-base reactions are equilibrium systems meaning that both the forward and reverse reactions occur. They involve two acid-base pairs, known as conjugate acid-base pairs.



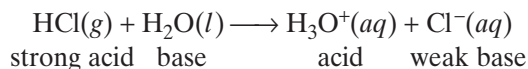
The subscripts designate the two conjugate acid-base pairs: (1) HF and  $\text{F}^-$  and (2)  $\text{H}_3\text{O}^+$  and  $\text{H}_2\text{O}$ . In every conjugate acid-base pair, the acid has one more proton than its conjugate base.

## OBJECTIVES

- Describe a conjugate acid, a conjugate base, and an amphoteric compound.
- Explain the process of neutralization.
- Explain how acid rain damages marble structures.

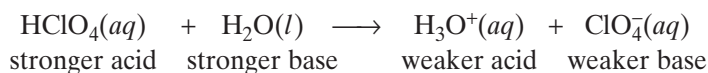
## Strength of Conjugate Acids and Bases

The extent of the reaction between a Brønsted-Lowry acid and base depends on the relative strengths of the acids and bases involved. Consider the following example. Hydrochloric is a strong acid. It gives up protons readily. It follows that the  $\text{Cl}^-$  ion has little tendency to attract and retain a proton. Consequently, the  $\text{Cl}^-$  ion is an extremely weak base.



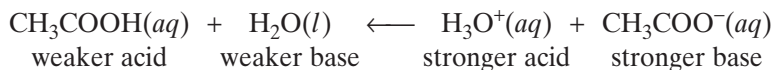
This observation leads to an important conclusion: *the stronger an acid is, the weaker its conjugate base; the stronger a base is, the weaker its conjugate acid.*

This concept allows strengths of different acids and bases to be compared to predict the outcome of a reaction. As an example, consider the reaction of perchloric acid,  $\text{HClO}_4$ , and water.



The hydronium ion is too weak an acid to compete successfully with perchloric acid in donating a proton;  $\text{HClO}_4$  is the stronger acid. In this reaction, the perchlorate ion,  $\text{ClO}_4^-$ , and  $\text{H}_2\text{O}$  are both bases. Because  $\text{HClO}_4$  is a very strong acid,  $\text{ClO}_4^-$  is an extremely weak base. Therefore,  $\text{H}_2\text{O}$  competes more strongly than  $\text{ClO}_4^-$  to acquire a proton. The reaction proceeds such that the stronger acid reacts with the stronger base to produce the weaker acid and base.

Now consider a comparable reaction between water and acetic acid.



The  $\text{H}_3\text{O}^+$  ion concentration in this solution is much lower than it was in the  $\text{HClO}_4$  solution because acetic acid is a weak acid. The  $\text{CH}_3\text{COOH}$  molecule does not compete successfully with the  $\text{H}_3\text{O}^+$  ion in donating protons to a base. The acetate ion,  $\text{CH}_3\text{COO}^-$ , is a stronger base than  $\text{H}_2\text{O}$ . Therefore, the  $\text{H}_2\text{O}$  molecule does not compete successfully with the  $\text{CH}_3\text{COO}^-$  ion in accepting a proton. The  $\text{H}_3\text{O}^+$  ion is the stronger acid, and the  $\text{CH}_3\text{COO}^-$  ion is the stronger base. Thus, the reaction to the left is more favorable.

Note that in the reactions for both perchloric acid and acetic acid, the favored direction is toward the weaker acid and the weaker base. This observation leads to a second important general conclusion: *proton-transfer reactions favor the production of the weaker acid and the weaker base.* For a reaction to approach completion, the reactants must be much stronger as an acid and as a base than the products.

By comparing many different acids and bases, a table of relative strengths, such as Table 15-6, can be assembled. Note that a very strong acid, such as  $\text{HClO}_4$ , has a very weak conjugate base,  $\text{ClO}_4^-$ . The strongest base listed in the table, the hydride ion,  $\text{H}^-$ , has the weakest

**TABLE 15-6 Relative Strengths of Acids and Bases**

	Conjugate acid	Formula	Conjugate base	Formula	
Strong acids	chloric acid	$\text{HClO}_3$	chlorate ion	$\text{ClO}_3^-$	Very weak bases
	hydrobromic acid	$\text{HBr}$	bromide ion	$\text{Br}^-$	
	hydrochloric acid	$\text{HCl}$	chloride ion	$\text{Cl}^-$	
	hydriodic acid	$\text{HI}$	iodide ion	$\text{I}^-$	
	nitric acid	$\text{HNO}_3$	nitrate ion	$\text{NO}_3^-$	
	perchloric acid	$\text{HClO}_4$	perchlorate ion	$\text{ClO}_4^-$	
	sulfuric acid	$\text{H}_2\text{SO}_4$	hydrogen sulfate ion	$\text{HSO}_4^-$	
Increasing acid strength	hydronium ion	$\text{H}_3\text{O}^+$	water	$\text{H}_2\text{O}$	Increasing base strength
	chlorous acid	$\text{HClO}_2$	chlorite ion	$\text{ClO}_2^-$	
	hydrogen sulfate ion	$\text{HSO}_4^-$	sulfate ion	$\text{SO}_4^{2-}$	
	phosphoric acid	$\text{H}_3\text{PO}_4$	dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^-$	
	hydrofluoric acid	$\text{HF}$	fluoride ion	$\text{F}^-$	
	acetic acid	$\text{CH}_3\text{COOH}$	acetate ion	$\text{CH}_3\text{COO}^-$	
	carbonic acid	$\text{H}_2\text{CO}_3$	hydrogen carbonate ion	$\text{HCO}_3^-$	
	hydrosulfuric acid	$\text{H}_2\text{S}$	hydrosulfide ion	$\text{HS}^-$	
	dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^-$	hydrogen phosphate ion	$\text{HPO}_4^{2-}$	
	hypochlorous acid	$\text{HClO}$	hypochlorite ion	$\text{ClO}^-$	
	ammonium ion	$\text{NH}_4^+$	ammonia	$\text{NH}_3$	
	hydrogen carbonate ion	$\text{HCO}_3^-$	carbonate ion	$\text{CO}_3^{2-}$	
	hydrogen phosphate ion	$\text{HPO}_4^{2-}$	phosphate ion	$\text{PO}_4^{3-}$	
	water	$\text{H}_2\text{O}$	hydroxide ion	$\text{OH}^-$	
	ammonia	$\text{NH}_3$	amide ion	$\text{NH}_2^-$	
	hydrogen	$\text{H}_2$	hydride ion	$\text{H}^-$	

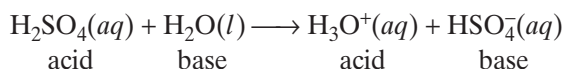
conjugate acid,  $\text{H}_2$ . A violent proton-transfer reaction could result from bringing together a very strong acid and a very strong base in certain proportions because the reaction has almost no tendency to go in the reverse direction. Such a reaction would give off a great deal of heat and would be dangerous. In fact, even the reaction between hydride ions and water—a much weaker acid than perchloric acid—is quite vigorous. The reaction is illustrated in Figure 15-11.

## Amphoteric Compounds

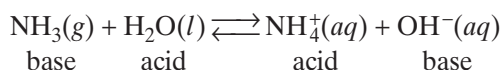
You have probably noticed that water can be either an acid or a base. *Any species that can react as either an acid or a base is described as amphoteric.* For example, consider the first ionization of sulfuric acid, in which water acts as a base.



**FIGURE 15-11** Because  $\text{H}^-$  is an extremely strong base, calcium hydride,  $\text{CaH}_2$ , reacts vigorously with water to produce hydrogen gas. The hydride ion accepts a proton from water, which acts as an acid in this reaction.

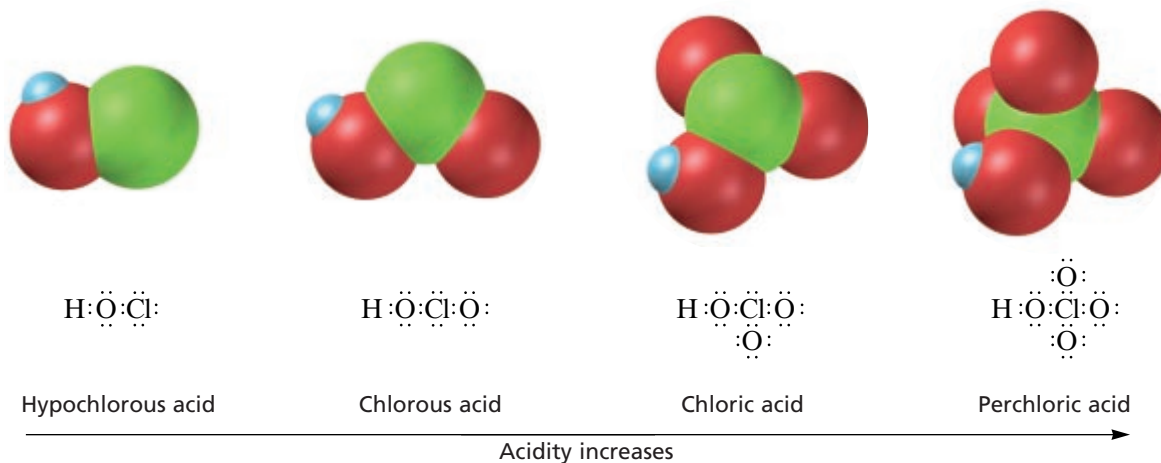


However, water acts as an acid in the following reaction.



**FIGURE 15-12** Each oxyacid of chlorine contains one chlorine atom and one hydrogen atom. They differ in the number of oxygen atoms they contain. The effect of the changing O–H bond polarity can be seen in the increasing acid strength from hypochlorous acid to perchloric acid.

Thus, water can act as either an acid or a base and is amphoteric. Such a substance acts as either an acid or a base, depending on the strength of the acid or base with which they are reacting. For example, if water reacts with a compound that is a stronger acid than itself, water acts as a base. If water reacts with a weaker acid, water acts as an acid.



## –OH in a Molecule

Molecular compounds containing –OH groups can be acidic or amphoteric. The covalently bonded –OH group in an acid is referred to as a *hydroxyl group*. For the compound to be acidic, a water molecule must be able to attract a hydrogen atom from a hydroxyl group. This occurs more easily when the O–H bond is very polar. Any feature of a molecule that increases the polarity of the O–H bond increases the acidity of a molecular compound. The small, more-electronegative atoms of nonmetals at the upper right in the periodic table form compounds with acidic hydroxyl groups. All oxyacids are molecular electrolytes that contain one or more of these O–H bonds. Such compounds include chloric and perchloric acids.

Figure 15-12 shows the electron-dot formulas of the four oxyacids of chlorine. Notice that all of the oxygen atoms are bonded to the chlorine atom. Each hydrogen atom is bonded to an oxygen atom. Aqueous solutions of these molecules are acids because the O–H bonds are broken as the hydrogen is attracted away by water molecules.

The behavior of a compound is affected by the number of oxygen atoms bonded to the atom connected to the –OH group. The larger the number of such oxygen atoms is, the more acidic the compound is likely to be. The electronegative oxygen atoms draw electron density away from the O–H bond and make it more polar. For example, chromium forms three different compounds containing –OH groups, as shown below.

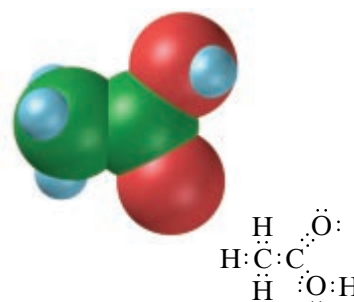
<i>basic</i>	<i>amphoteric</i>	<i>acidic</i>
$\text{Cr}(\text{OH})_2$	$\text{Cr}(\text{OH})_3$	$\text{H}_2\text{CrO}_4$
chromium(II) hydroxide	chromium(III) hydroxide	chromic acid

Notice that as the number of oxygen atoms increases, so does the acidity of the compound.

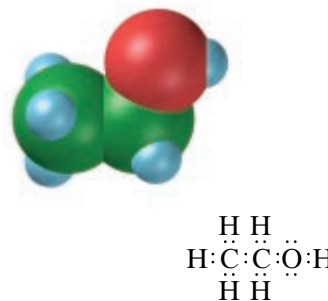
Consider also the compounds shown in Figure 15-13. In acetic acid, but not in ethanol, a second oxygen atom is bonded to the carbon atom connected to the –OH group. That explains why acetic acid is acidic but ethanol is not, even though the same elements form each compound.

## Neutralization Reactions

There are many common examples of acidic compounds reacting with basic compounds, each neutralizing the other. Sodium bicarbonate,  $\text{NaHCO}_3$ , and tartaric acid,  $\text{C}_4\text{H}_6\text{O}_6$ , are two components in baking powder. When allowed to react in solution, the two compounds produce carbon dioxide. The escaping carbon dioxide causes foods, such as biscuits, to rise. An antacid soothes an overly acidic stomach by neutralizing the stomach acid.



(a)  $\text{CH}_3\text{COOH}$   
Acetic acid

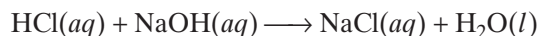


(b)  $\text{C}_2\text{H}_5\text{OH}$   
Ethanol

**FIGURE 15-13** (a) Acetic acid is acidic. The second oxygen atom on the carbon draws electron density away from the –OH group, making the O–H bond more polar. (b) Ethanol is essentially neutral. It has no second oxygen atom, so ethanol is less polar than acetic acid, and is a much weaker acid.

## Strong Acid-Strong Base Neutralization

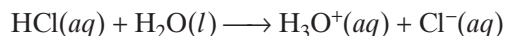
An acid-base reaction occurs in aqueous solution between hydrochloric acid, a strong acid that completely dissociates to produce  $\text{H}_3\text{O}^+$ , and sodium hydroxide, a strong base that completely dissociates to produce  $\text{OH}^-$ . The formula equation for this reaction is written as follows.



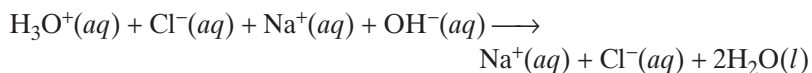
In an aqueous solution containing 1 mol of sodium hydroxide,  $\text{NaOH}$  dissociates as represented by the following equation.



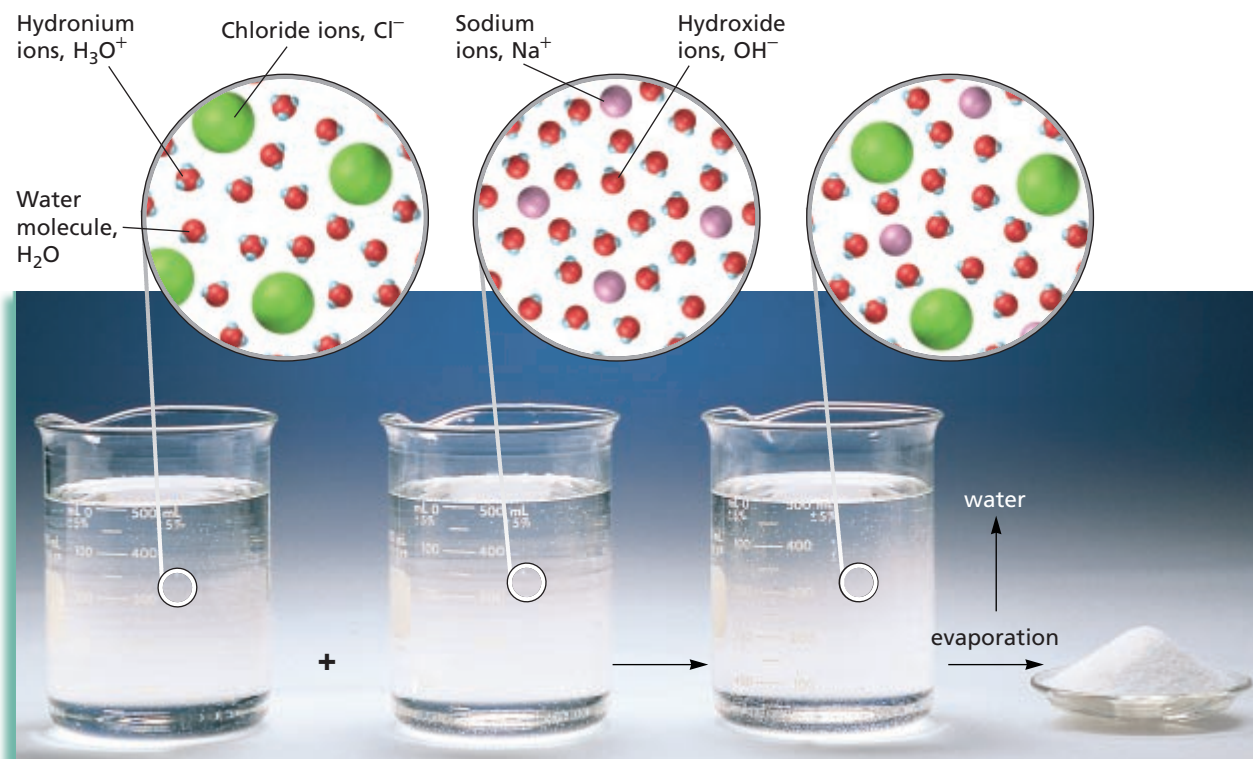
A solution containing 1 mol of hydrochloric acid ionizes as represented by the following equation.



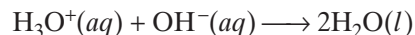
If the two solutions are mixed, as in Figure 15-14, a reaction occurs between the aqueous ions. Notice that sodium chloride,  $\text{NaCl}$ , and water are produced. The overall ionic equation is shown below.



**FIGURE 15-14** When aqueous hydrochloric acid,  $\text{HCl}$ , reacts with aqueous sodium hydroxide,  $\text{NaOH}$ , the reaction produces aqueous sodium chloride,  $\text{NaCl}$ . Ions that are present in each solution are represented by the models.



Because they appear on both sides of the overall ionic equation,  $\text{Na}^+$  and  $\text{Cl}^-$  are spectator ions. The only participants in the reaction are the hydronium ion and the hydroxide ion, as shown in the following net ionic equation.

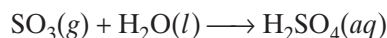


There are equal numbers of  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$  ions in this reaction, and they are fully converted to water. In aqueous solutions, **neutralization** is the reaction of hydronium ions and hydroxide ions to form water molecules.

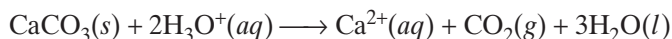
Notice that water is not the only product of a neutralization. A salt is also produced. A **salt** is an ionic compound composed of a cation from a base and an anion from an acid.

## Acid Rain

Many industrial processes produce gases such as  $\text{NO}$ ,  $\text{NO}_2$ ,  $\text{CO}_2$ ,  $\text{SO}_2$ , and  $\text{SO}_3$ . These compounds can dissolve in atmospheric water to produce acidic solutions that fall to the ground in the form of rain or snow. The following reaction shows how sulfur trioxide,  $\text{SO}_3$ , dissolves in water to produce sulfuric acid.



Marble found in many buildings and statues is composed of calcium carbonate,  $\text{CaCO}_3$ . When acid snow or rain falls on these structures, the following reaction takes place.



The products of the reaction of an acid with any carbonate are a salt, water, and carbon dioxide. An example of the effects of acid rain can be seen in Figure 15-15; the marble in this statue has been eroded by the acidic precipitation. Even though the products themselves are harmless, the reaction has serious consequences.



**FIGURE 15-15** Acid precipitation causes extensive damage to buildings and other structures.

## SECTION REVIEW

- Complete and balance the equations for the following acid-base reactions:
  - $\text{H}_2\text{CO}_3 + \text{Sr}(\text{OH})_2 \longrightarrow$
  - $\text{HClO}_4 + \text{NaOH} \longrightarrow$
  - $\text{HBr} + \text{Ba}(\text{OH})_2 \longrightarrow$
  - $\text{NaHCO}_3 + \text{H}_2\text{SO}_4 \longrightarrow$
- Consider the equation for acetic acid plus water.
$$\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \longrightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_3\text{O}^+$$
  - Refer to Table 15-6 to compare the strengths of the two acids in the equation. Do the same for the two bases.
  - Determine which direction—forward or reverse—is favored in the reaction.
- Explain how the presence of several oxygen atoms in a compound containing an  $-\text{OH}$  group can make the compound acidic.



# CHAPTER 15 REVIEW

## CHAPTER SUMMARY

- 15-1**
- Acids have a sour taste and react with active metals. They change the colors of acid-base indicators, and react with bases to produce salts and water. Some conduct electricity.
  - Bases have a bitter taste, feel slippery to the skin in dilute aqueous solutions, change the colors of acid-base indicators, react with acids to produce salts and water, and conduct electricity.

### Vocabulary

alkaline (461)

Arrhenius acid (459)

Arrhenius base (459)

binary acid (454)

- An Arrhenius acid contains hydrogen and ionizes in aqueous solution to form hydrogen ions. An Arrhenius base produces hydroxide ions in aqueous solution.
- The strength of an Arrhenius acid or base is determined by the extent to which it ionizes or dissociates in aqueous solutions.

oxyacid (455)

strong acid (460)

weak acids (460)

- 15-2**
- A Brønsted-Lowry acid is a proton donor. A Brønsted-Lowry base is a proton acceptor.
  - A Lewis acid is an electron-pair acceptor. A Lewis base is an electron-pair donor.

### Vocabulary

amphoteric (470)

Brønsted-Lowry acid (464)

Brønsted-Lowry acid-base reaction (465)

Brønsted-Lowry base (465)

diprotic acid (466)

Lewis acid (467)

- Acids are described as monoprotic, diprotic, or triprotic, depending on whether they can donate one, two, or three protons per molecule, respectively, in aqueous solutions.

Lewis acid-base reaction (468)

Lewis base (468)

monoprotic acid (465)

polyprotic acid (465)

triprotic acid (466)

- 15-3**
- In every Brønsted-Lowry acid-base reaction, there are two conjugate acid-base pairs.
  - A strong acid has a weak conjugate base; a strong base has a weak conjugate acid.
  - Proton-transfer reactions favor the production of weaker acids and bases.
  - The acidic or basic behavior of a molecule containing  $\text{—OH}$  groups may depend on the

### Vocabulary

amphoteric (471)

conjugate acid (469)

conjugate base (469)

electronegativity of other atoms in the molecule and the number of oxygen atoms bonded to the atom connected to the  $\text{—OH}$  group.

- A neutralization reaction produces water and an ionic compound called a salt.
- Acid rain neutralizes the calcium carbonate in marble structures, causing them to deteriorate.

neutralization (475)

salt (475)

## REVIEWING CONCEPTS

- Compare and contrast the general properties of acids and bases. (15-1)
- Distinguish between binary acids and oxyacids in terms of their component elements and the systems used in naming them.
  - Give three examples of each. (15-1)
- Identify and describe the characteristic properties of five common acids used in industry. Give some examples of the typical uses of each. (15-1)
- Although  $\text{HCl}(aq)$  exhibits Arrhenius acidic properties, pure  $\text{HCl}$  gas and  $\text{HCl}$  dissolved in a nonpolar solvent exhibit no acidic properties in the Arrhenius sense. Explain why. (15-1)
- What distinguishes strong acids from weak acids?
  - Give two examples of each. (15-1)
- $\text{H}_3\text{PO}_4$ , which contains three hydrogen atoms per molecule, is a weak acid, whereas  $\text{HCl}$ , which contains only one hydrogen atom per molecule, is a strong acid. Explain why. (15-1)
- What determines the strength of an Arrhenius base?
  - Give one example each of solutions that are strongly and weakly basic. (15-1)
- Distinguish among a monoprotic, a diprotic, and a triprotic acid. Give an example of each. (15-2)
- Which of the three acid definitions is the broadest? Explain. (15-2)
- Define and give an equation to illustrate each of the following:
  - a conjugate base
  - a conjugate acid (15-3)
- What is the relationship between the strength of an acid and that of its conjugate base?
  - What is the relationship between the strength of a base and its conjugate acid? (15-3)
- What trend is there in the favored direction of proton-transfer reactions?
  - What determines the extent to which a proton-transfer reaction occurs? (15-3)
- What is meant by the term *amphoteric*?
  - Give an example of a substance or ion with amphoteric characteristics. (15-3)
- For each reaction listed, identify the proton donor or acid and the proton acceptor or base. Label each conjugate acid-base pair.
  - $\text{CH}_3\text{COOH}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CH}_3\text{COO}^-(aq)$
  - $\text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(aq) + \text{OH}^-(aq)$
  - $\text{HNO}_3 + \text{SO}_4^{2-} \longrightarrow \text{HSO}_4^- + \text{NO}_3^-$  (15-3)
- Based on the information given in Table 15-6, determine the following relative to  $\text{HF}$ ,  $\text{H}_2\text{S}$ ,  $\text{HNO}_3$ , and  $\text{CH}_3\text{COOH}$ :
  - strongest acid
  - weakest acid
  - strongest conjugate base among the four produced by the acids listed
  - weakest conjugate base among the four produced (15-3)
- Explain why the conjugate base of a strong acid is a weak base and the conjugate acid of a strong base is a weak acid. (15-3)

## PROBLEMS

## Acid Nomenclature

- Name each of the following binary acids:
  - $\text{HCl}$
  - $\text{H}_2\text{S}$
- Name each of the following oxyacids:
  - $\text{HNO}_3$
  - $\text{H}_2\text{SO}_3$
  - $\text{HClO}_3$
  - $\text{HNO}_2$
- Write formulas for each of the following binary acids:
  - hydrofluoric acid
  - hydriodic acid
- Write formulas for each of the following oxyacids:
  - perbromic acid
  - chlorous acid
  - phosphoric acid
  - hypochlorous acid

**Acid-Base Theory**

21. a. Write the balanced equations that describe the two-stage ionization of sulfuric acid in a dilute aqueous solution.  
b. How do the degrees of ionization in the two steps compare?
22. Dilute  $\text{HCl}(aq)$  and  $\text{KOH}(aq)$  are mixed in chemically equivalent quantities. Write the following:  
a. formula equation for the reaction  
b. overall ionic equation  
c. net ionic equation
23. Repeat item 22 with  $\text{H}_3\text{PO}_4(aq)$  and  $\text{NaOH}(aq)$ .
24. Write the formula equation and net ionic equation for each of the following reactions:  
a.  $\text{Zn}(s) + \text{HCl}(aq) \longrightarrow$   
b.  $\text{Al}(s) + \text{H}_2\text{SO}_4(aq) \longrightarrow$
25. Write the formula equation and net ionic equation for the reaction between  $\text{Ca}(s)$  and  $\text{HCl}(aq)$ .

**Neutralization Reactions**

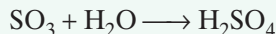
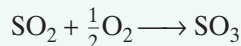
26. Complete the following neutralization reactions. Balance each reaction, and then write the overall ionic and net ionic equation for each.  
a.  $\text{HCl}(aq) + \text{NaOH}(aq) \longrightarrow$   
b.  $\text{HNO}_3(aq) + \text{KOH}(aq) \longrightarrow$   
c.  $\text{Ca}(\text{OH})_2(aq) + \text{HNO}_3(aq) \longrightarrow$   
d.  $\text{Mg}(\text{OH})_2(aq) + \text{HCl}(aq) \longrightarrow$
27. Write the formula equation, the overall ionic equation, and the net ionic equation for the neutralization reaction involving aqueous solutions of  $\text{H}_3\text{PO}_4$  and  $\text{Mg}(\text{OH})_2$ . Assume that the solutions are sufficiently dilute so that no precipitates form.
28. Write the balanced chemical equation for each of the following reactions between an acid and a carbonate:  
a.  $\text{BaCO}_3(s) + \text{HCl}(aq) \longrightarrow$   
b.  $\text{MgCO}_3(s) + \text{HNO}_3(aq) \longrightarrow$   
c.  $\text{Na}_2\text{CO}_3(s) + \text{H}_2\text{SO}_4(aq) \longrightarrow$   
d.  $\text{CaCO}_3(s) + \text{H}_3\text{PO}_4(aq) \longrightarrow$
29. Write the formula equation, the overall ionic equation, and the net ionic equation for a neu-

tralization reaction that would form each of the following salts.

- a.  $\text{RbClO}_4$                       c.  $\text{CaCl}_2$   
b.  $\text{BaSO}_4$                       d.  $\text{K}_2\text{SO}_4$

**Stoichiometry**

30. Zinc reacts with 100. mL of 6.00 M cold, aqueous sulfuric acid through single replacement.  
a. How many grams of zinc sulfate are produced?  
b. How many liters of hydrogen gas would be released at STP?
31. A 211 g sample of barium carbonate,  $\text{BaCO}_3$ , is placed in a solution of nitric acid. Assuming that the acid is present in excess, what mass and volume of dry carbon dioxide gas at STP will be produced?
32. A seashell, composed largely of calcium carbonate, is placed in a solution of  $\text{HCl}$ . As a result, 1500 mL of dry  $\text{CO}_2$  gas at STP is produced. The other products are  $\text{CaCl}_2$  and  $\text{H}_2\text{O}$ .  
a. Based on this information, how many grams of  $\text{CaCO}_3$  are consumed in the reaction?  
b. What volume of 2.00 M  $\text{HCl}$  solution is used in this reaction?
33. *Acid precipitation* is the term generally used to describe rain or snow that is more acidic than normal. One cause of acid precipitation is the formation of sulfuric and nitric acids from various sulfur and nitrogen oxides produced in volcanic eruptions, forest fires, and thunderstorms. In a typical volcanic eruption, for example,  $3.50 \times 10^8$  kg of  $\text{SO}_2$  may be produced. If this amount of  $\text{SO}_2$  were converted to  $\text{H}_2\text{SO}_4$  according to the two-step process given below, how many kilograms of  $\text{H}_2\text{SO}_4$  would be produced from such an eruption?

**MIXED REVIEW**

34. Suppose that dilute  $\text{HNO}_3(aq)$  and  $\text{LiOH}(aq)$  are mixed in chemically equivalent quantities. Write the following for the reaction:  
a. formula equation  
b. overall ionic equation  
c. net ionic equation

35. Write the balanced chemical equation for the reaction between hydrochloric acid and magnesium metal.
36. Write equations for the three-step ionization of phosphoric acid,  $\text{H}_3\text{PO}_4$ . Compare the degree of ionization for the three steps.
37. Name or give the molecular formula for each of the following acids:
  - a. HF
  - b. acetic acid
  - c. phosphorous acid
  - d.  $\text{HClO}_4$
  - e.  $\text{H}_3\text{PO}_4$
  - f. hydrobromic acid
  - g.  $\text{HClO}$
  - h.  $\text{H}_2\text{CO}_3$
  - i. sulfuric acid

### CRITICAL THINKING

38. **Analyzing Conclusions** In the eighteenth century, Antoine Lavoisier experimented with oxides such as  $\text{CO}_2$  and  $\text{SO}_2$ . He observed that they formed acidic solutions. His observations led him to infer that for a substance to exhibit acidic behavior, it must contain oxygen. However, today that is known to be incorrect. Provide evidence to refute Lavoisier's conclusion.



### HANDBOOK SEARCH

39. Group 16 of the *Elements Handbook* contains a section covering the acid-base chemistry of oxides. Review this material and answer the following:
  - a. What type of compounds form acidic oxides?
  - b. What is an acidic anhydride?
  - c. List three examples of compounds that are classified as acidic anhydrides.
  - d. What type of compounds form basic oxides? Why are they basic oxides?
40. a. Look at Table 7A in the *Elements Handbook*. What periodic trends do you notice regarding acid-base character of oxides?
  - b. How is the nature of the product affected by the concentration of  $\text{NaOH}$  in a reaction with  $\text{CO}_2$ ?

### RESEARCH & WRITING

41. Explain how sulfuric acid production serves as a measure of a country's economy. Write a report on your findings.
42. **Performance** Conduct library research to find out about the buffering of solutions. Include information on why buffering is typically carried out and on the kinds of materials used. Write a brief report on your findings.
43. Obtain some pH paper from your teacher. Determine whether the soil around your house is acidic or basic. Find one type of plant that would grow well in that type of soil and one that would not.

### ALTERNATIVE ASSESSMENT

44. Antacids are designed to neutralize excess hydrochloric acid secreted by the stomach during digestion. Carbonates, bicarbonates, and hydroxides are the active ingredients in bringing about the neutralization reactions in the most widely used antacids. Examine the labels of several common antacids, and identify the active ingredients.
45. **Performance** Design an experiment that compares three brands of antacids in terms of reaction speed and amount of acid neutralized.