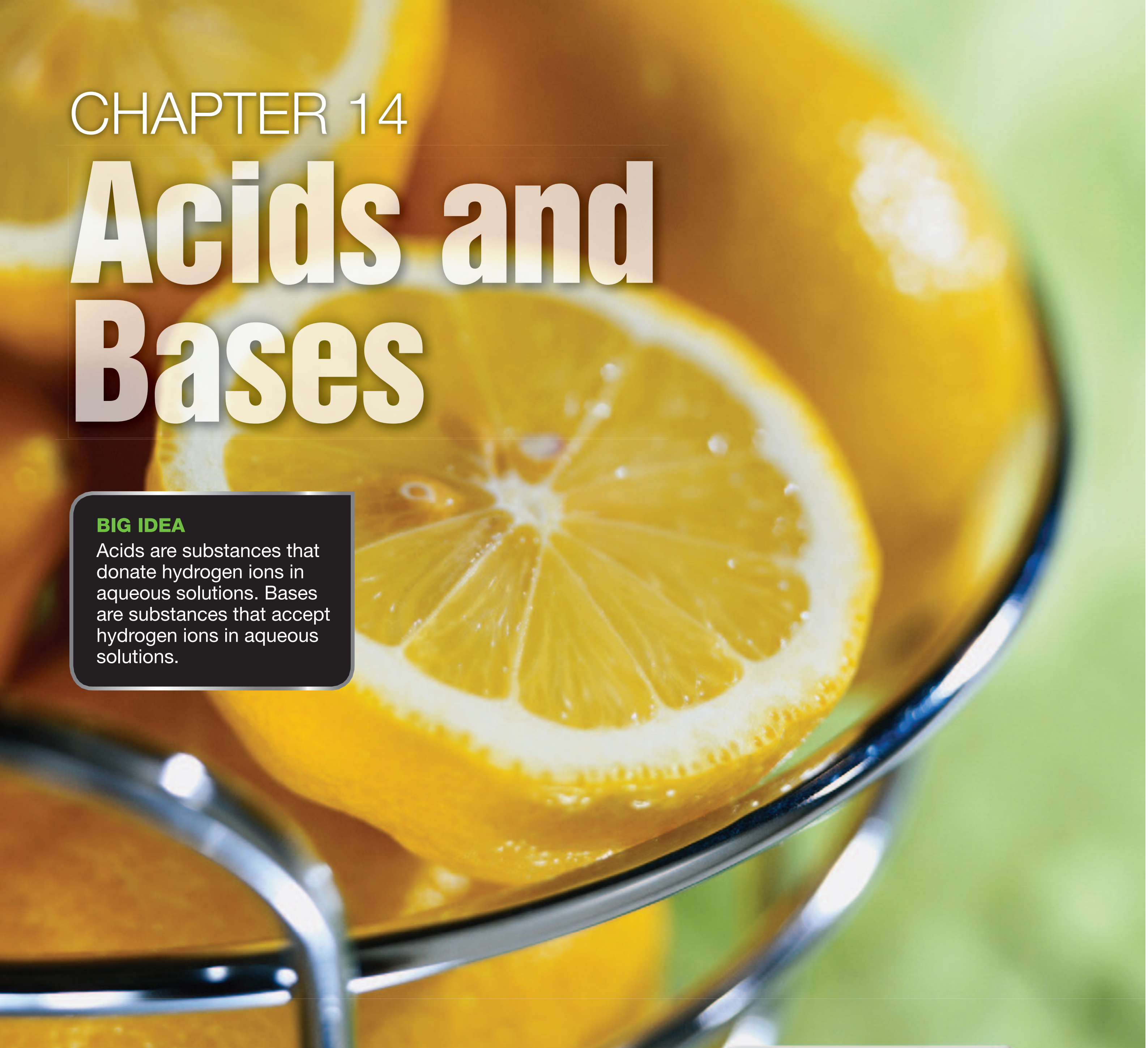

CHAPTER 14

Acids and Bases

BIG IDEA
Acids are substances that donate hydrogen ions in aqueous solutions. Bases are substances that accept hydrogen ions in aqueous solutions.

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SECTION 1
Properties of Acids and Bases

SECTION 2
Acid-Base Theories

SECTION 3
Acid-Base Reactions

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Acids and Bases

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SECTION 1

Properties of Acids and Bases

Key Terms

binary acid
oxyacid
Arrhenius acid
Arrhenius base
strong acid
weak acid

Main Ideas

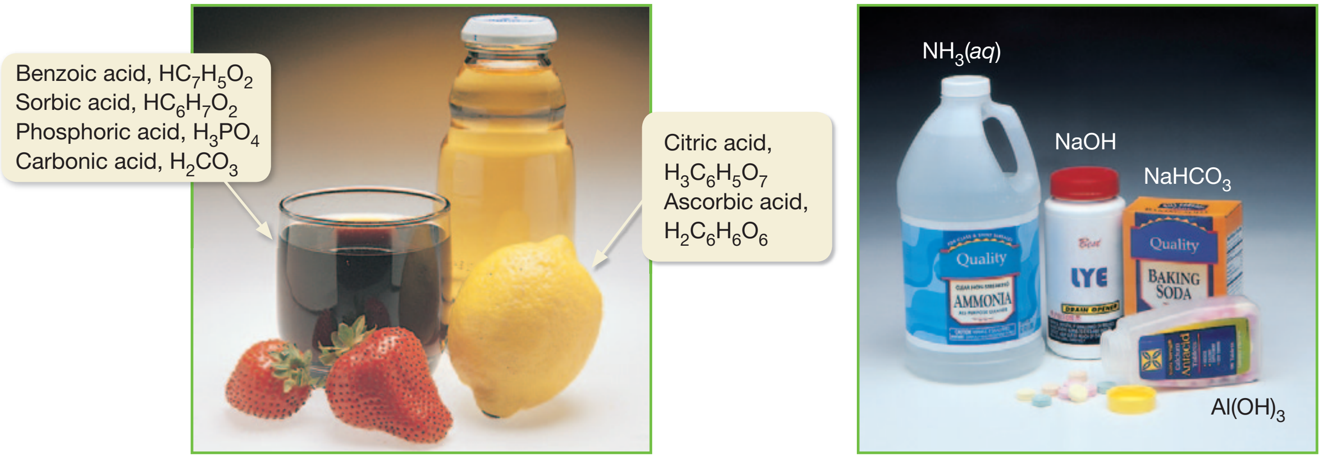
- Acids are identified by their properties.
- Some acids are useful in industry.
- The properties of bases differ from those of acids.
- Arrhenius acids and bases produce ions in solution.

How many foods can you think of that are sour? Chances are that almost all the foods you thought of, like those in **Figure 1.1a,** 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 1.1b.** 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 cleaners. Milk of magnesia is a suspension in water of magnesium hydroxide, $Mg(OH)_2$, 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, $Al(OH)_3$, and sodium hydrogen carbonate, $NaHCO_3$, are also bases commonly found in antacids.

FIGURE 1.1

Common Acids and Bases



(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. Antacids contain bases such as aluminum hydroxide.

FIGURE 1.2

Acid Indicator A strip of pH paper dipped into vinegar turns red, showing that vinegar is an acid.

MAIN IDEA

Acids are identified by their properties.

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 1.2.**
3. *Some acids react with active metals and 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-displacement reactions with certain acids. Hydrogen gas is formed as a product, as shown by the reaction of barium with sulfuric acid.

$$Ba(s) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + H_2(g)$$

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. *Acids conduct electric current*. Some acids completely separate into ions in water and are strong electrolytes. Other acids are weak 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. Names for some binary acids are given in **Figure 1.3.**

FIGURE 1.3

NAMES OF BINARY ACIDS

Formula	Acid name	Molecule name
HF	hydrofluoric acid	hydrogen fluoride
HCl	hydrochloric acid	hydrogen chloride
HBr	hydrobromic acid	hydrogen bromide
HI	hydriodic acid	hydrogen iodide
H_2S	hydrosulfuric acid	hydrogen sulfide

In pure form, each compound listed in the table is a gas. Aqueous solutions of these compounds are known by the acid names. Specific rules for naming binary compounds are listed below.

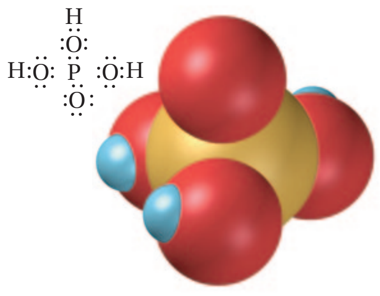
Binary Acid Nomenclature

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*.

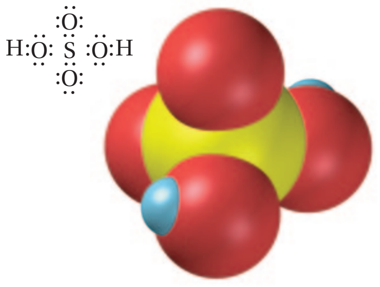
An oxyacid is an acid that is a compound of hydrogen, oxygen, and a third element, usually a nonmetal. Nitric acid, HNO_3, is an oxyacid. The structures of two other oxyacids are shown in **Figure 1.4.** Oxyacids are one class of ternary acids, which are acids that contain three different elements. Usually, the elements in an oxyacid formula are written as one or more hydrogen atoms followed by a polyatomic anion. But as you can see from the structures, the H atoms are bonded to O atoms. The names of oxyacids follow a pattern, and the names of their anions are based on the names of the acids. Some common oxyacids and their anions are given in **Figure 1.5.** Many of these names should be familiar to you.

FIGURE 1.4

Oxyacid Structure Although the chemical formula shows hydrogen atoms and a polyatomic ion, notice that the hydrogens are bonded to the oxygen in an oxyacid.



(a) Structure of H_3PO_4



(b) Structure of H_2SO_4

FIGURE 1.5

NAMES OF COMMON OXYACIDS AND OXYANIONS

Formula	Acid name	Anion
CH_3COOH	acetic acid	CH_3COO^-, acetate
H_2CO_3	carbonic acid	CO_3^{2-}, carbonate
$HClO$	hypochlorous acid	ClO^-, hypochlorite
$HClO_2$	chlorous acid	ClO_2^-, chlorite
$HClO_3$	chloric acid	ClO_3^-, chlorate
$HClO_4$	perchloric acid	ClO_4^-, perchlorate
HIO_3	iodic acid	IO_3^-, iodate
HNO_2	nitrous acid	NO_2^-, nitrite
HNO_3	nitric acid	NO_3^-, nitrate
H_3PO_3	phosphorous acid	PO_3^{3-}, phosphite
H_3PO_4	phosphoric acid	PO_4^{3-}, phosphate
H_2SO_3	sulfurous acid	SO_3^{2-}, sulfite
H_2SO_4	sulfuric acid	SO_4^{2-}, sulfate

MAIN IDEA

Some acids are useful in industry.

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.

Sulfuric Acid

Sulfuric acid is the most commonly produced industrial chemical in the world. More than 37 million metric 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 dehydrating (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. Dissolving the acid in water makes the acid more stable. Solutions of nitric acid are widely used in industry. Nitric acid also stains proteins yellow. The feather in **Figure 1.6** 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 pharmaceuticals. Initially, nitric acid solutions are colorless; however, upon standing, they gradually become yellow because of slight decomposition to brown nitrogen dioxide gas.

FIGURE 1.6

Nitric Acid and Proteins

Concentrated nitric acid stains a feather yellow.

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.

Concentrated solutions of hydrochloric acid, commonly referred to as *muriatic acid*, can be found in hardware stores. It is used to maintain the correct acidity in swimming pools and to clean masonry.

Acetic Acid

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

MAIN IDEA

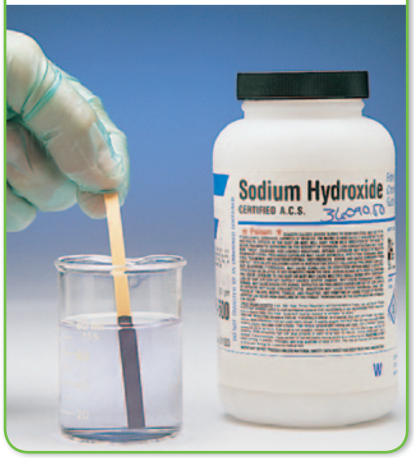
The properties of bases differ from those of acids.

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. Taste should NEVER be used to test a substance to see if it is a base, because many bases are caustic; they attack the skin, causing severe burns.
2. *Bases change the color of acid-base indicators*. As **Figure 1.7** shows, an acid-base indicator changes to a different color in a basic solution than in an acidic solution.
3. *Dilute aqueous solutions of bases feel slippery*. You encounter this property of aqueous bases whenever you wash with soap.

FIGURE 1.7

Base Indicator pH paper turns blue in the presence of this solution of sodium hydroxide.



4. *Bases react with acids to produce salts and water*. It could also be said that "neutralization" of the base occurs when these two substances react to produce a salt and water.
5. *Bases conduct electric current*. Like acids, bases form ions in aqueous solutions and are thus electrolytes.

Unlike acids, there are no special rules for naming basic compounds. The names of bases follow the conventions for chemical compounds outlined in the chapter "Chemical Formulas and Chemical Compounds."

QuickLAB 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 then pour off and save the cabbage juice. This solution is an acid-base indicator.
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.

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?

MATERIALS

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

SAFETY

Wear safety goggles, gloves, and an apron.

Red cabbage can be used to make an acid-base indicator by extracting its anthocyanin pigment.

MAIN IDEA

Arrhenius acids and bases produce ions in solution.

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 aqueous acids are electrolytes.

Because acid molecules are sufficiently polar, water molecules attract one or more of their hydrogen ions. Negatively charged anions are left behind. As explained in a previous chapter, the hydrogen ion in aqueous solution is typically represented as H_3O^+, the hydronium ion. The ionization of an HNO_3 molecule is shown by the following equation. **Figure 1.8** also shows how the hydronium ion forms when nitric acid reacts with water.

$$HNO_3(l) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$$

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

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

FIGURE 1.8

Arrhenius Acids 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.

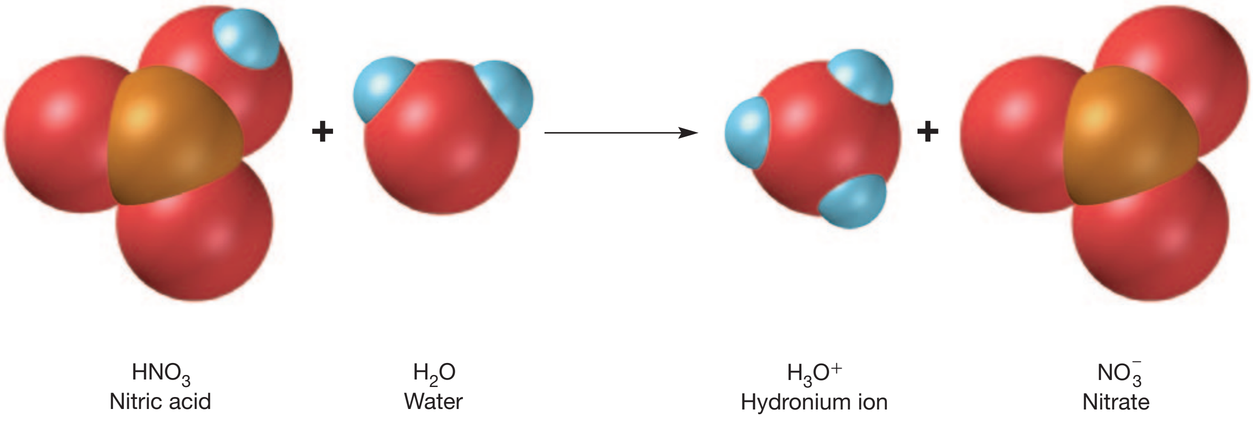


FIGURE 1.9

COMMON AQUEOUS ACIDS

Strong acids	Weak acids
$HI + H_2O \longrightarrow H_3O^+ + I^-$	$HSO_4^- + H_2O \rightleftarrows H_3O^+ + SO_4^{2-}$
$HClO_4 + H_2O \longrightarrow H_3O^+ + ClO_4^-$	$H_3PO_4 + H_2O \rightleftarrows H_3O^+ + H_2PO_4^-$
$HBr + H_2O \longrightarrow H_3O^+ + Br^-$	$HF + H_2O \rightleftarrows H_3O^+ + F^-$
$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$	$CH_3COOH + H_2O \rightleftarrows H_3O^+ + CH_3COO^-$
$H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^-$	$H_2CO_3 + H_2O \rightleftarrows H_3O^+ + HCO_3^-$
$HClO_3 + H_2O \longrightarrow H_3O^+ + ClO_3^-$	$H_2S + H_2O \rightleftarrows H_3O^+ + HS^-$
	$HCN + H_2O \rightleftarrows H_3O^+ + CN^-$
	$HCO_3^- + H_2O \rightleftarrows H_3O^+ + CO_3^{2-}$

Strength of Acids

A strong acid is one that ionizes completely in aqueous solution. A strong acid is a strong electrolyte. Perchloric acid, $HClO_4$, hydrochloric acid, HCl, and nitric acid, HNO_3, are examples of strong acids. In water, 100% of the acid molecules are ionized. 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.

An acid that releases few hydrogen ions in aqueous solution is a weak acid. The aqueous solution of a weak acid contains hydronium ions, anions, and dissolved acid molecules. Hydrocyanic acid is an example of a weak acid. In aqueous solution, both the ionization of HCN and the reverse reaction occur simultaneously. In a 1 M solution of HCN, there will be only two H^+ ions and two CN^- ions out of 100,000 molecules. The other 99,998 molecules remain as HCN.

$$HCN(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + CN^-(aq)$$

Common aqueous acids are listed in **Figure 1.9.** Each strong acid ionizes completely in aqueous solution to give up one hydrogen ion per molecule. Notice that the number of hydrogen atoms in the formula does not indicate acid strength. Molecules with multiple hydrogen atoms may not readily give them up. 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.

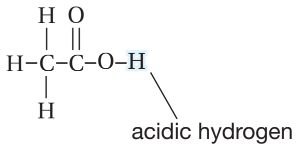
Organic acids, which contain the acidic carboxyl group —COOH, are generally weak acids. For example, acetic acid, CH_3COOH, ionizes slightly in water to give hydronium ions and acetate ions, CH_3COO^-.

$$CH_3COOH(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + CH_3COO^-(aq)$$

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 1.10.**

FIGURE 1.10

Acetic Acid Acetic acid contains four hydrogen atoms, but only one of them is “acidic” and forms the hydronium ion in solution.



$$H-\overset{\overset{H}{|}}{\underset{\underset{H}{|}}{C}}-\overset{\overset{O}{\|}}{C}-O-H$$

acidic hydrogen

Aqueous Solutions of Bases

Most bases are ionic compounds containing metal cations and the hydroxide anion, OH^-. Because these bases are ionic, they dissociate when dissolved in water. When a base completely dissociates in water to yield aqueous OH^- ions, the solution is considered a strong base. Sodium hydroxide, NaOH, is a common laboratory base. It is water-soluble and dissociates as shown by the equation below.

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

As you will remember from learning about the periodic table, Group 1 elements are the alkali metals. This group gets its name from the fact that the hydroxides of Li, Na, K, Rb, and Cs all form alkaline (basic) solutions.

Not all bases are ionic compounds. A base commonly used in household cleaners is ammonia, 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.

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

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, KOH, is a strong base because it completely dissociates into its ions in dilute aqueous solutions.

$$KOH(s) \xrightarrow{H_2O} K^+(aq) + OH^-(aq)$$

Strong bases are strong electrolytes, just as strong acids are strong electrolytes. **Figure 1.11** lists some strong bases.

✓ **CHECK FOR UNDERSTANDING**

Differentiate What is the difference between the strength and the concentration of an acid or base?

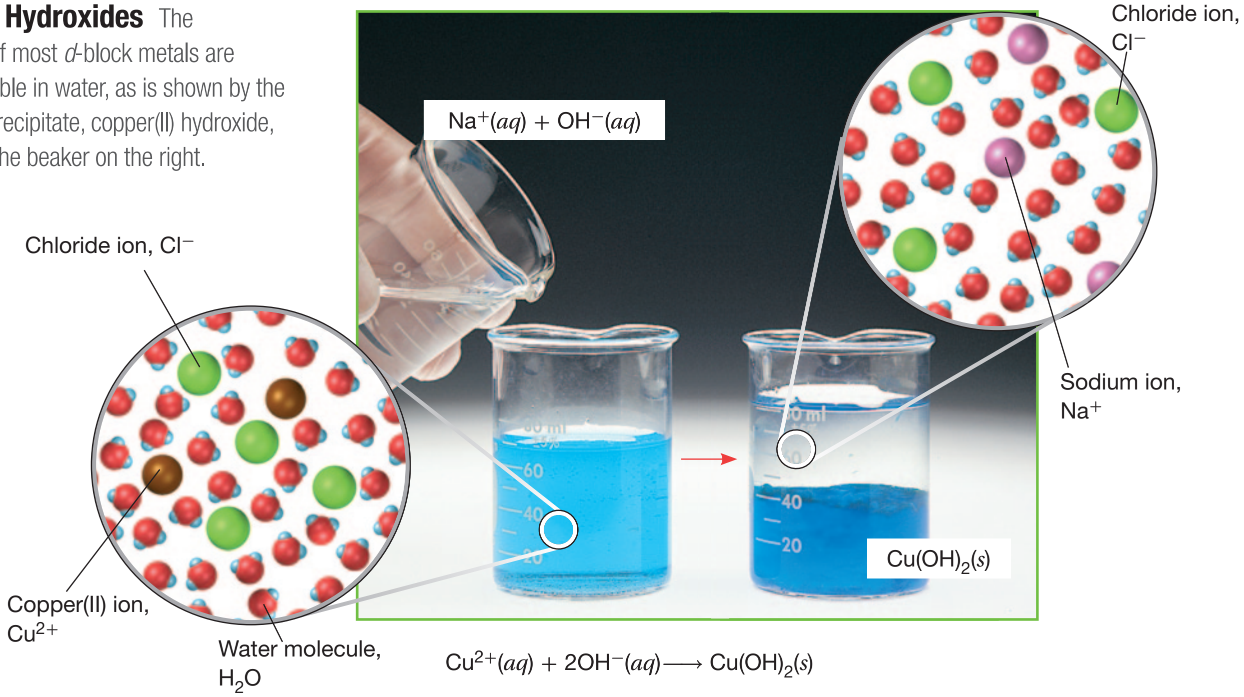
FIGURE 1.11

COMMON AQUEOUS BASES

Strong bases	**Weak bases**
$Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^-$	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
$Sr(OH)_2 \longrightarrow Sr^{2+} + 2OH^-$	$C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + OH^-$
$Ba(OH)_2 \longrightarrow Ba^{2+} + 2OH^-$	
$NaOH \longrightarrow Na^+ + OH^-$	
$KOH \longrightarrow K^+ + OH^-$	
$RbOH \longrightarrow Rb^+ + OH^-$	
$CsOH \longrightarrow Cs^+ + OH^-$	

FIGURE 1.12

Insoluble Hydroxides The hydroxides of most *d*-block metals are nearly insoluble in water, as is shown by the gelatinous precipitate, copper(II) hydroxide, $Cu(OH)_2$, in the beaker on the right.



$$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$$

Bases that are not very soluble do not produce a large number of hydroxide ions when added to water. Some metal hydroxides, such as $Cu(OH)_2$, are not very soluble in water, as seen in **Figure 1.12.** They cannot produce concentrated alkaline solutions. The alkalinity of aqueous solutions depends on the concentration of 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 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, codeine, $C_{18}H_{21}NO_3$, a pain reliever and common cough suppressant found in prescription cough medicine, is a weak base.

SECTION 1 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. a. What are five general properties of aqueous acids?

  b. Name some common substances that have one or more of these properties.

2. Name the following acids.

  a. HBrO **b.** $HBrO_3$

3. Write the chemical formulas for the following common bases.

  a. lithium hydroxide
  b. sodium hydroxide

4. a. What are five general properties of aqueous bases?

  b. Name some common substances that have one or more of these properties.

5. a. Why are strong acids also strong electrolytes?

  b. Is every strong electrolyte also a strong acid?

Critical Thinking

6. RELATING IDEAS A classmate states, "All compounds containing H atoms are acids, and all compounds containing OH groups are bases." Do you agree? Give examples.

Acid Water—A Hidden Menace

Many people are unaware of the pH of the tap water in their home until they are confronted with such phenomena as a blue ring materializing around a porcelain sink drain, a water heater suddenly giving out, or tropical fish that keep dying. Each of these events could be due to acidic water. Acidic water can also cause the amount of lead in the water to rise.

The possibility of lead poisoning from home water supplies is a concern. Many older homes still have lead pipes in their plumbing, though most modern homes use copper piping. Copper pipe joints, however, are often sealed with lead-containing solder. Highly acidic water can leach out both the lead from the solder joints and copper from the pipes themselves, which turns the sink drain blue. In addition, people who are in the habit of filling their tea kettles and coffee pots in the morning without letting the tap run awhile first could be adding copper and lead ions to their tea or coffee.

Lead poisoning is of particular concern in young children. The absorption rate of lead in the intestinal tract of a child is much higher than in that of an adult, and lead poisoning can permanently impair a child's rapidly growing nervous system. The good news is that lead poisoning and other effects of acidic water in the home can be easily prevented by following these tips:

1. Monitor the pH of your water on a regular basis, especially if you have well water. This can easily be done with pH test kits (see photograph) that are sold in hardware or pet stores—many tropical fish are intolerant of water with a pH that is either too high (basic) or too low (acidic). The pH of municipal water supplies is already regulated, but regularly checking your water's pH yourself is a good idea.
2. In the morning, let your water tap run for about half a minute before you fill your kettle or drink the water. If the water is acidic, the first flush of water will have the highest concentration of lead and copper ions.
3. Install an alkali-injection pump, a low-cost, low-maintenance solution that can save your plumbing and lessen the risk of lead poisoning from your own water supply. The pump injects a small amount of an alkali (usually potassium carbonate or sodium carbonate) into your water-pressure tank each time your well's pump starts. This effectively neutralizes the acidity of your water.

Question

Suppose that you have tested the drinking water in your home and found it to be too acidic. Design a solution to correct this problem. In your solution, explain the steps you will take to make the water less acidic and what evidence you will collect to show that your solution has worked.

The pH of your home's water supply can be easily monitored using a test kit, such as the one shown here.

SECTION 2

Main Ideas

- Brønsted-Lowry acids and bases donate or accept protons.
- A Lewis acid or base accepts or donates a pair of electrons.

Acid-Base Theories

Key Terms

Brønsted-Lowry acid
Brønsted-Lowry base
Brønsted-Lowry acid-base reaction
monoprotic acid
polyprotic acid
diprotic acid
triprotic acid
Lewis acid
Lewis base
Lewis acid-base reaction

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.

MAIN IDEA

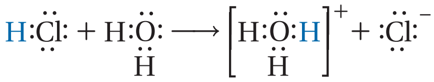
Brønsted-Lowry acids and bases donate or accept protons.

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 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.

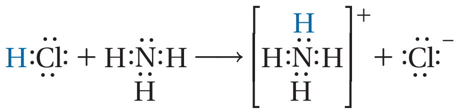
Hydrogen chloride acts as a Brønsted-Lowry acid when it reacts with ammonia. The HCl transfers protons to NH_3 much as it does in water.

$$HCl + NH_3 \longrightarrow NH_4^+ + Cl^-$$

A proton is transferred from the hydrogen chloride molecule, HCl, to the ammonia molecule, NH_3. The ammonium ion, NH_4^+, is formed. Electron-dot formulas show the similarity of this reaction to the reaction of HCl with water.



$$\mathrm{H{:}\overset{..}{\underset{..}{Cl}}{:} + H{:}\underset{H}{\overset{..}{O}}{:} \longrightarrow \left[H{:}\underset{H}{\overset{..}{O}}{:}H\right]^+ + {:}\overset{..}{\underset{..}{Cl}}{:}^-}$$



$$\mathrm{H{:}\overset{..}{\underset{..}{Cl}}{:} + H{:}\underset{H}{\overset{..}{N}}{:}H \longrightarrow \left[\begin{matrix} H \\ H{:}\overset{..}{\underset{..}{N}}{:}H \\ H \end{matrix}\right]^+ + {:}\overset{..}{\underset{..}{Cl}}{:}^-}$$

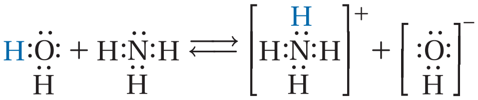
In both reactions, hydrogen chloride is a Brønsted-Lowry acid.

FIGURE 2.1

Brønsted-Lowry Acid-Base Reactions 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 that is dispersed in air.

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.

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



$$\mathrm{H}:\overset{\cdot\cdot}{\underset{\underset{\mathrm{H}}{\cdot\cdot}}{\mathrm{O}}}: + \mathrm{H}:\overset{\cdot\cdot}{\underset{\underset{\mathrm{H}}{\cdot\cdot}}{\mathrm{N}}}:\mathrm{H} \rightleftharpoons \left[\mathrm{H}:\overset{\overset{\mathrm{H}}{\cdot\cdot}}{\underset{\underset{\mathrm{H}}{\cdot\cdot}}{\mathrm{N}}}:\mathrm{H}\right]^+ + \left[:\overset{\cdot\cdot}{\underset{\underset{\mathrm{H}}{\cdot\cdot}}{\mathrm{O}}}:\right]^-$$

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. Instead, it is the OH^- ion produced in solution that 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 2.1** shows the reaction between the Brønsted-Lowry acid HCl and the Brønsted-Lowry base 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, $HClO_4$, hydrochloric acid, HCl, and nitric acid, HNO_3, are all monoprotic. The following equation shows how a molecule of the monoprotic acid HCl donates a proton to a water molecule. The HCl ionizes to form H_3O^+ ions and Cl^- ions. The Cl^- has no hydrogens to lose, so HCl has only one ionization step.

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

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A polyprotic acid is an acid that can donate more than one proton per molecule. Sulfuric acid, H_2SO_4, and phosphoric acid, H_3PO_4, are examples of polyprotic acids. The ionization of a polyprotic acid occurs in stages. The acid loses its hydrogen ions 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, HSO_4^-.

$$H_2SO_4(l) + H_2O(l) \longrightarrow H_3O^+(aq) + HSO_4^-(aq)$$

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

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

All stages of ionization of a polyprotic acid occur in the same solution. Sulfuric acid solutions therefore contain H_3O^+, HSO_4^-, and SO_4^{2-} ions. Note that in sulfuric acid solutions, there are many more hydrogen sulfate and hydronium ions than there are sulfate 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 2.2.**

FIGURE 2.2

Monoprotic and Diprotic Acids Hydrochloric acid, HCl, is a strong monoprotic acid. A dilute HCl solution contains hydronium ions and chloride ions. Sulfuric acid, H_2SO_4, is a strong diprotic acid. A dilute H_2SO_4 solution contains hydrogen sulfate ions from the first ionization, sulfate ions from the second ionization, and hydronium ions from both ionizations.

$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$

$H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^-$
$HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$

CRITICAL THINKING

Explain Using the Brønsted-Lowry definition, is it possible to have an acid without an accompanying base?

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.

$$H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + H_2PO_4^-(aq)$$

$$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$$

$$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + PO_4^{3-}(aq)$$

A solution of phosphoric acid contains H_3O^+, H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}, and 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.

MAIN IDEA

A Lewis acid or base accepts or donates a pair of electrons.

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.**

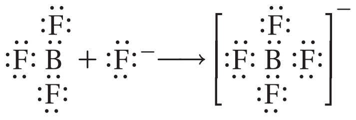
The Lewis definition is the broadest of the three acid definitions you have read about so far. It applies to any species that can accept an electron pair to form a covalent bond with another species. A bare proton (hydrogen ion) is a Lewis acid in reactions in which it forms a covalent bond, as shown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

$$H^+(aq) + :NH_3(aq) \longrightarrow [H\text{–}NH_3]^+(aq) \text{ or } [NH_4]^+(aq)$$

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.

$$Ag^+(aq) + 2:NH_3(aq) \longrightarrow [H_3N\text{–}Ag\text{–}NH_3]^+(aq) \text{ or } [Ag(NH_3)_2]^+$$

A 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.



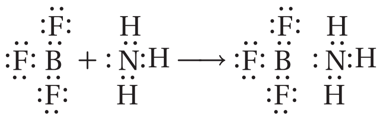
$$BF_3(aq) + F^-(aq) \longrightarrow BF_4^-(aq)$$

FIGURE 2.3

ACID-BASE DEFINITIONS

Type	Acid	Base
Arrhenius	H^+ or H_3O^+ producer	OH^- producer
Brønsted-Lowry	proton (H^+) donor	proton (H^+) 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^- donates an electron pair to boron trifluoride. F^- acts as a Lewis base.

$$BF_3(aq) + :\ddot{\underset{..}{F}}:^-(aq) \longrightarrow BF_4^-(aq)$$

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^- 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 **Figure 2.3.**

SECTION 2 FORMATIVE ASSESSMENT

Reviewing Main Ideas

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

$$H_2CO_3 + H_2O \rightleftharpoons HCO_3^- + H_3O^+$$

2. For the reaction below, label each reactant as an electron-pair acceptor or electron-pair donor and as a Lewis acid or a Lewis base.

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

Critical Thinking

3. ANALYZING INFORMATION For the following three reactions, identify the reactants that are Arrhenius bases, Brønsted-Lowry bases, and/or Lewis bases. State which type(s) of bases each reactant is. Explain your answers.

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

b. $HF(aq) + H_2O(l) \longrightarrow F^-(aq) + H_3O^+(aq)$

c. $H^+(aq) + NH_3(aq) \longrightarrow NH_4^+(aq)$

Acid-Base Reactions

SECTION 3

Main Ideas

- Brønsted-Lowry reactions involve conjugate acid-base pairs.
- Some substances act as either acids or bases.
- Neutralization reactions produce water and a salt.

Key Terms

conjugate base
conjugate acid
amphoteric
neutralization
salt

In the previous sections, 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.

MAIN IDEA

Brønsted-Lowry reactions involve conjugate acid-base pairs.

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 hydrofluoric acid.

$$\underset{\text{acid}}{HF(aq)} + H_2O(l) \rightleftharpoons \underset{\text{conjugate base}}{F^-(aq)} + H_3O^+(aq)$$

In this reaction, the water molecule is a Brønsted-Lowry base. It accepts a proton from HF to form H_3O^+, 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.**

$$HF(aq) + \underset{\text{base}}{H_2O(l)} \rightleftharpoons F^-(aq) + \underset{\text{conjugate acid}}{H_3O^+(aq)}$$

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.

$$\underset{\text{acid}_1}{HF(aq)} + \underset{\text{base}_2}{H_2O(l)} \rightleftharpoons \underset{\text{base}_1}{F^-(aq)} + \underset{\text{acid}_2}{H_3O^+(aq)}$$

The subscripts designate the two conjugate acid-base pairs: (1) HF and F^- and (2) H_3O^+ and H_2O. In every conjugate acid-base pair, the acid has one more proton than its conjugate base.

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 acid is a strong acid. It gives up protons readily. Therefore, the Cl^- ion has little tendency to attract and retain a proton. Consequently, the Cl^- ion is an extremely weak base.

$$\underset{\text{stronger acid}}{HCl(g)} \quad + \quad \underset{\text{stronger base}}{H_2O(l)} \quad \longrightarrow \quad \underset{\text{weaker acid}}{H_3O^+(aq)} \quad + \quad \underset{\text{weaker base}}{Cl^-(aq)}$$

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.*

Using Strength to Predict Reactions

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, $HClO_4$, and water.

$$\underset{\text{stronger acid}}{HClO_4(aq)} \quad + \quad \underset{\text{stronger base}}{H_2O(l)} \quad \longrightarrow \quad \underset{\text{weaker acid}}{H_3O^+(aq)} \quad + \quad \underset{\text{weaker base}}{ClO_4^-(aq)}$$

The hydronium ion is too weak an acid to compete successfully with perchloric acid in donating a proton; $HClO_4$ is the stronger acid. In this reaction, the perchlorate ion, ClO_4^-, and H_2O are both bases. Because $HClO_4$ is a very strong acid, ClO_4^- is an extremely weak base. Therefore, H_2O competes more strongly than 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.

$$\underset{\text{weaker acid}}{CH_3COOH(aq)} \quad + \quad \underset{\text{weaker base}}{H_2O(l)} \quad \longleftarrow \quad \underset{\text{stronger acid}}{H_3O^+(aq)} \quad + \quad \underset{\text{stronger base}}{CH_3COO^-(aq)}$$

The H_3O^+ ion concentration in this solution is much lower than it was in the $HClO_4$ solution because acetic acid is a weak acid. The CH_3COOH molecule does not compete successfully with the H_3O^+ ion in donating protons to a base. The acetate ion, CH_3COO^-, is a stronger base than H_2O. Therefore, the H_2O molecule does not compete successfully with the CH_3COO^- ion in accepting a proton. The H_3O^+ ion is the stronger acid, and the CH_3COO^- ion is the stronger base. Thus, the reverse 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 an acid-base reaction to form products completely, the reactants must be much stronger as acids and bases than the products.

CROSS-DISCIPLINARY CONNECTION

"It's a Bitter Pill" S.T.E.M.

Have you ever wondered about the origin of the saying, "It's a bitter pill to swallow"? This saying is used to describe a situation that is difficult to accept. Many medications contain bases, and a bitter taste is a property of bases. So, many medications actually have a bitter taste. If you look at the chemical formulas of the components of medications, you will see that they often contain nitrogen. One such component is caffeine, which acts as a stimulant on the central nervous and respiratory systems. Its molecular formula is $C_8H_{10}O_2N_4$. Like ammonia, caffeine has basic properties because it has a nitrogen that can accept a proton.

The table in **Figure 3.1** shows that a very strong acid, such as $HClO_4$, has a very weak conjugate base, ClO_4^-. The strongest base listed in the table, the hydride ion, H^-, has the weakest conjugate acid, H_2. In aqueous solutions, all of the strong acids are 100% ionized, forming hydronium ions along with their anion. The acids below hydronium ion in **Figure 3.1** do not ionize 100% in water. Water will react as an acid if a very strong base, such as hydride ion, is present.

FIGURE 3.1

RELATIVE STRENGTHS OF ACIDS AND BASES

Conjugate acid	Formula	Conjugate base	Formula
hydriodic acid*	HI	iodide ion	I^-
perchloric acid*	$HClO_4$	perchlorate ion	ClO_4^-
hydrobromic acid*	HBr	bromide ion	Br^-
hydrochloric acid*	HCl	chloride ion	Cl^-
sulfuric acid*	H_2SO_4	hydrogen sulfate ion	HSO_4^-
chloric acid*	$HClO_3$	chlorate ion	ClO_3^-
nitric acid*	HNO_3	nitrate ion	NO_3^-
hydronium ion	H_3O^+	water	H_2O
chlorous acid	$HClO_2$	chlorite ion	ClO_2^-
hydrogen sulfate ion	HSO_4^-	sulfate ion	SO_4^{2-}
phosphoric acid	H_3PO_4	dihydrogen phosphate ion	$H_2PO_4^-$
hydrofluoric acid	HF	fluoride ion	F^-
acetic acid	CH_3COOH	acetate ion	CH_3COO^-
carbonic acid	H_2CO_3	hydrogen carbonate ion	HCO_3^-
hydrosulfuric acid	H_2S	hydrosulfide ion	HS^-
dihydrogen phosphate ion	$H_2PO_4^-$	monohydrogen phosphate ion	HPO_4^{2-}
hypochlorous acid	$HClO$	hypochlorite ion	ClO^-
ammonium ion	NH_4^+	ammonia	NH_3
hydrogen carbonate ion	HCO_3^-	carbonate ion	CO_3^{2-}
monohydrogen phosphate ion	HPO_4^{2-}	phosphate ion	PO_4^{3-}
water	H_2O	hydroxide ion	OH^-
ammonia	NH_3	amide ion†	NH_2^-
hydrogen	H_2	hydride ion†	H^-

↑ Increasing acid strength

↓ Increasing base strength

* Strong acids
† Strong bases

FIGURE 3.2

Water as an Acid Calcium hydride, CaH_2, reacts vigorously with water to produce hydrogen gas. Water acts as an acid in this reaction because the hydride ion is a very strong base.

$CaH_2(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + 2H_2(g)$

CHECK FOR UNDERSTANDING

Identify Identify the conjugate acid and base in the reaction of calcium hydride with water.

MAIN IDEA

Some substances act as either acids or bases.

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.

$$\underset{\text{acid}_1}{H_2SO_4(aq)} + \underset{\text{base}_2}{H_2O(l)} \longrightarrow \underset{\text{acid}_2}{H_3O^+(aq)} + \underset{\text{base}_1}{HSO_4^-(aq)}$$

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

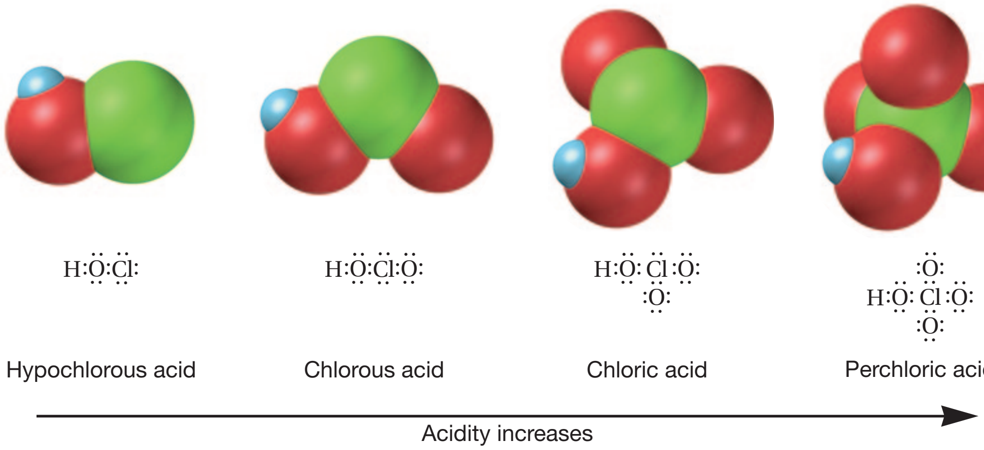
$$\underset{\text{base}_1}{NH_3(g)} + \underset{\text{acid}_2}{H_2O(l)} \rightleftharpoons \underset{\text{acid}_1}{NH_4^+(aq)} + \underset{\text{base}_2}{OH^-(aq)}$$

Likewise, water is acting as an acid in the reaction shown in **Figure 3.2.**

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 it is reacting—if it's a stronger acid than the substance, the substance will act as a base. If, however, it's a stronger base than the substance, the substance will act as an acid.

FIGURE 3.3

O—H Bonds and Acid Strength 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 increasing O—H bond polarity can be seen in the increasing acid strength from hypochlorous acid to perchloric 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 3.3 (on the previous page) 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 pulled 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. 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.

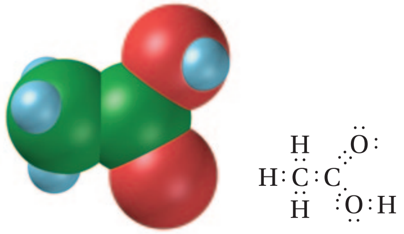
basic	*amphoteric*	*acidic*
$Cr(OH)_2$ chromium(II) hydroxide	$Cr(OH)_3$ chromium(III) hydroxide	H_2CrO_4 chromic acid

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

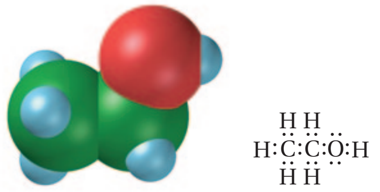
Consider also the compounds shown in **Figure 3.4.** In acetic acid, but not in ethanol, a second oxygen atom is bonded to the carbon atom connected to the –OH group. That explains why acetic acid is acidic but ethanol is not, even though the same elements form each compound.

FIGURE 3.4

Polarity and Acids



(a) CH_3COOH
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) C_2H_5OH
Ethanol is essentially neutral. It has no second oxygen atom, so the O—H bond is less polar than in acetic acid, and it is a much weaker acid.

MAIN IDEA

Neutralization reactions produce water and a salt.

Neutralization reactions commonly occur in nature. One example is when an antacid tablet such as milk of magnesia, $Mg(OH)_2$, neutralizes stomach acid, HCl. The products are $MgCl_2$ and H_2O. Some neutralization reactions produce gases. Sodium bicarbonate, $NaHCO_3$, and tartaric acid, $H_2C_4H_4O_6$, are two components in baking powder. When water is added, the two compounds produce carbon dioxide. The escaping carbon dioxide causes foods, such as biscuits, to rise.

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Neutralization Reactions

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Strong Acid-Strong Base Neutralization

An acid-base reaction occurs in aqueous solution between hydrochloric acid, a strong acid that completely ionizes to produce H_3O^+, and sodium hydroxide, a strong base that completely dissociates to produce OH^-. The formula equation for this reaction is written as follows.

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

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

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

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

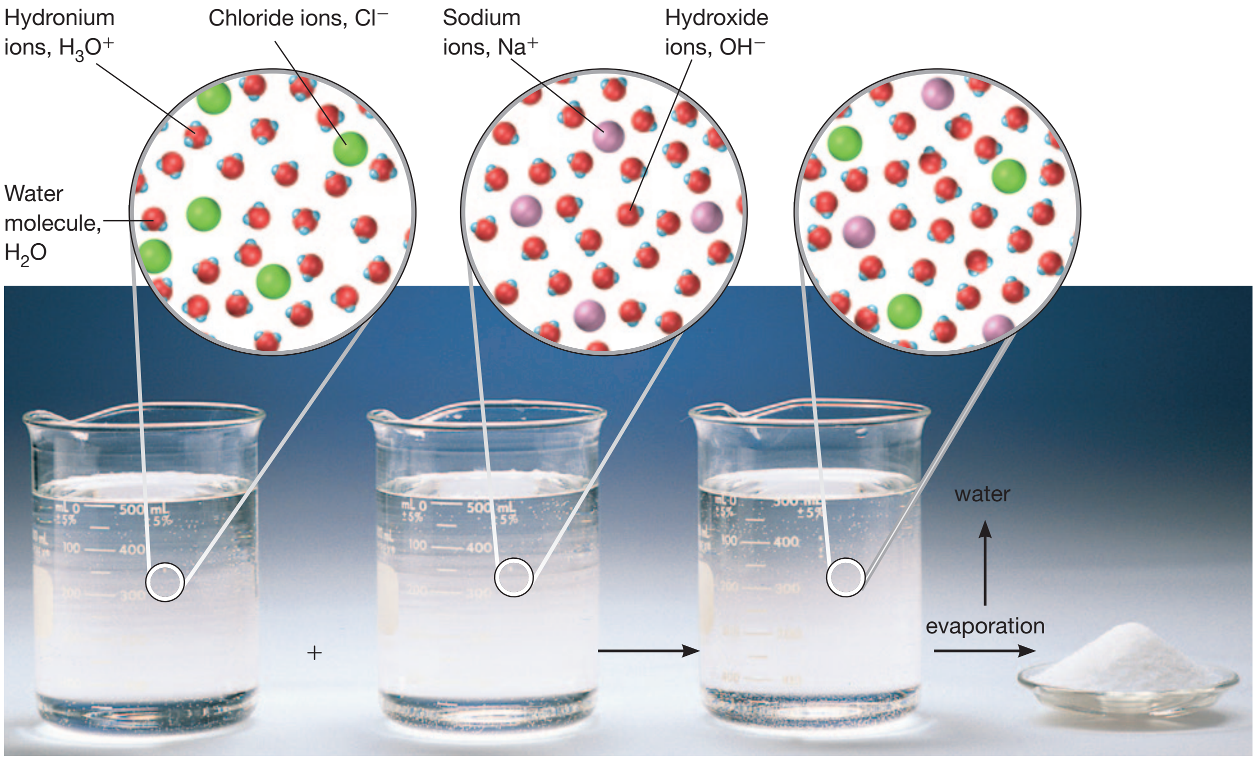
$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

If the two solutions are mixed, as in **Figure 3.5,** a reaction occurs between the aqueous H_3O^+ and OH^- ions. Notice that sodium chloride, NaCl, and water are produced. The full ionic equation is shown below.

$$H_3O^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + 2H_2O(l)$$

FIGURE 3.5

Neutralization When aqueous hydrochloric acid, HCl, reacts with aqueous sodium hydroxide, NaOH, the reaction produces aqueous sodium chloride, NaCl. Ions that are present in each solution are represented by the models.



Because they appear on both sides of the overall ionic equation, Na^+ and 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.

$$H_3O^+(aq) + OH^-(aq) \longrightarrow 2H_2O(l)$$

There are equal numbers of H_3O^+ and 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.**

A neutralization reaction also produces a salt. **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 NO, NO_2, CO_2, SO_2, and 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. For example, sulfur from the burning of oil and coal forms sulfur dioxide, SO_2. The SO_2 is then converted to SO_3, sulfur trioxide, which reacts with water in the atmosphere to produce sulfuric acid, as shown below.

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

Rainwater is normally slightly acidic due to carbon dioxide in the atmosphere, but sometimes rain is very acidic and is called *acid rain*. **Figure 3.6** shows a forest that was damaged by severe acid rain. Acid rain can erode statues and affect ecosystems, such as water environments and forests. In the 1970s, scientists found that acid rain was causing the fish populations in some lakes and streams to decline. When fish are completely eliminated from lakes and streams because of acid rain, the biodiversity of the ecosystem decreases. Because of amendments to the Clean Air Act in 1990, a limit was set on the amount of SO_2 that power plants are permitted to emit. This limit has decreased but not eliminated acid rain in the United States.

FIGURE 3.6

Acid Precipitation Acid precipitation causes extensive environmental damage.

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SECTION 3 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Complete and balance the equations for the following acid-base reactions:

- **a.** $H_2CO_3 + Sr(OH)_2 \longrightarrow$
- **b.** $HClO_4 + NaOH \longrightarrow$
- **c.** $HBr + Ba(OH)_2 \longrightarrow$
- **d.** $NaHCO_3 + H_2SO_4 \longrightarrow$

2. Consider the equation for acetic acid plus water.

$$CH_3COOH + H_2O \longrightarrow CH_3COO^- + H_3O^+$$

- **a.** Refer to **Figure 3.1** to compare the strengths of the two acids in the equation. Do the same for the two bases.
- **b.** Determine which direction—forward or reverse—is favored in the reaction.

Critical Thinking

3. INFERRING RELATIONSHIPS Explain how the presence of several oxygen atoms in a compound containing an $-OH$ group can make the compound acidic.

Math Tutor

Writing Equations for Ionic Reactions

Many chemical reactions that occur in water solutions are reactions involving ions. Soluble ionic compounds dissociate into ions when they dissolve, and some molecular compounds, including acids, ionize when they dissolve. An ionic equation represents the species actually present more accurately than an equation that uses full formulas.

Problem-Solving TIPS

- All soluble ionic compounds are dissociated into ions. Therefore, soluble ionic compounds are shown as the separated ions in the full ionic equation. Strong acids and bases are also shown as the separated ions in the full ionic equation because they are 100% ionized.
- Ions that do not take part in the reaction are called *spectator ions*. In other words, spectator ions stay in solution and will be labeled "(*aq*)" on both sides of the equation. Eliminating spectator ions reduces the "clutter" of the full ionic equation and produces a net ionic equation that shows only the species that actually react.

Sample Problem

Write the net ionic equation for the reaction of aqueous ammonium sulfate and aqueous barium nitrate to produce a precipitate of barium sulfate. The balanced formula equation is $(NH_4)_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow 2NH_4NO_3(aq) + BaSO_4(s)$

Rewrite the equation in full ionic form; because ammonium sulfate and barium nitrate are soluble, they are written as separated ions:

$$2NH_4^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \longrightarrow 2NH_4^+(aq) + 2NO_3^-(aq) + BaSO_4(s)$$

Eliminating spectator ions, NH_4^+ and NO_3^-, yields the net ionic equation:

$$SO_4^{2-}(aq) + Ba^{2+}(aq) \longrightarrow BaSO_4(s)$$

Write full and net ionic equations for the reaction that occurs when hydrochloric acid solution is combined with silver nitrate solution.

Hydrochloric acid is a strong acid, so it is completely ionized in solution. Silver nitrate is a soluble ionic compound, so its ions are separated in solution. Although most chlorides are soluble, silver chloride is not, so silver chloride will precipitate. The balanced formula equation is

$$HCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + HNO_3(aq)$$

The full ionic equation is

$$H_3O^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow H_3O^+(aq) + NO_3^-(aq) + AgCl(s)$$

Eliminate spectator ions to obtain the net ionic equation:

$$Cl^-(aq) + Ag^+(aq) \longrightarrow AgCl(s)$$

Practice

Answers in Appendix E

1. Aqueous copper(II) sulfate reacts with aqueous sodium sulfide to produce a black precipitate of copper(II) sulfide. Write the formula equation, the full ionic equation, and the net ionic equation for this reaction.
2. Write full and net ionic equations for the reaction that occurs when a solution of cadmium chloride, $CdCl_2$, is mixed with a solution of sodium carbonate, Na_2CO_3. Cadmium carbonate is insoluble.

CHAPTER 14 Summary

BIG IDEA Acids are substances that donate hydrogen ions in aqueous solutions. Bases are substances that accept hydrogen ions in aqueous solutions.

SECTION 1 Properties of Acids and Bases

- Acids have a sour taste and react with active metals. Acids change the colors of acid-base indicators, react with bases to produce salts and water, and conduct electricity in aqueous solutions.
- 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 in aqueous solution.
- 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 the acid or base ionizes or dissociates in aqueous solutions.

KEY TERMS

binary acid
oxyacid
Arrhenius acid
Arrhenius base
strong acid
weak acid

SECTION 2 Acid-Base Theories

- 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.
- 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. Polyprotic acids include both diprotic and triprotic acids.

KEY TERMS

Brønsted-Lowry acid
Brønsted-Lowry base
Brønsted-Lowry acid-base reaction
monoprotic acid
polyprotic acid
diprotic acid
triprotic acid
Lewis acid
Lewis base
Lewis acid-base reaction

SECTION 3 Acid-Base Reactions

- 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 the weaker acid and weaker base.
- The acidic or basic behavior of a molecule containing —OH groups depends on the electronegativity of other atoms in the molecule and on the number of oxygen atoms bonded to the atom that is connected to the —OH group.
- A neutralization reaction produces water and an ionic compound called a salt.
- Acid rain can create severe ecological problems.

KEY TERMS

conjugate base
conjugate acid
amphoteric
neutralization
salt

CHAPTER 14 Review

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SECTION 1

Properties of Acids and Bases

REVIEWING MAIN IDEAS

1. Compare the general properties of acids with the general properties of bases.

2. a. Distinguish between binary acids and oxyacids in terms of their component elements and the systems used in naming them.
b. Give three examples of each type of acid.

3. Identify and describe the characteristic properties of five common acids used in industry. Give some examples of the typical uses of each.

4. Although HCl(*aq*) exhibits properties of an Arrhenius acid, pure HCl gas and HCl dissolved in a nonpolar solvent exhibit none of the properties of an Arrhenius acid. Explain why.

5. a. What distinguishes strong acids from weak acids?
b. Give two examples each of strong acids and weak acids.

6. H_3PO_4, which contains three hydrogen atoms per molecule, is a weak acid, whereas HCl, which contains only one hydrogen atom per molecule, is a strong acid. Explain why.

7. a. What compounds are strong Arrhenius bases?
b. Give an example of an aqueous solution of a strong base and one of a weak base.

PRACTICE PROBLEMS

8. Name each of the following binary acids:
a. HCl **b.** H_2S

9. Name each of the following oxyacids:
a. HNO_3 **c.** $HClO_3$
b. H_2SO_3 **d.** HNO_2

10. Write formulas for the following binary acids and common bases:
a. hydrofluoric acid **c.** sodium bicarbonate
b. hydriodic acid **d.** aluminum hydroxide

11. Write formulas for the following oxyacids:
a. perbromic acid
b. chlorous acid
c. phosphoric acid
d. hypochlorous acid

SECTION 2

Acid-Base Theories

REVIEWING MAIN IDEAS

12. Distinguish between a monoprotic, a diprotic, and a triprotic acid. Give an example of each.

13. Which of the three acid definitions is the broadest? Explain.

PRACTICE PROBLEMS

14. a. Write the balanced equations that describe the two-step ionization of sulfuric acid in a dilute aqueous solution.
b. How do the degrees of ionization in the two steps compare?

15. Dilute HCl(*aq*) and KOH(*aq*) are mixed in chemically equivalent quantities. Write the following:
a. formula equation for the reaction
b. full ionic equation
c. net ionic equation

16. Repeat item 15, but mix $H_3PO_4(aq)$ and NaOH(*aq*).

17. Write the formula equation and net ionic equation for each of the following reactions:
a. $Zn(s) + HCl(aq) \longrightarrow$
b. $Al(s) + H_2SO_4(aq) \longrightarrow$

18. Write the formula equation and net ionic equation for the reaction between Ca(*s*) and HCl(*aq*).

SECTION 3

Acid-Base Reactions

REVIEWING MAIN IDEAS

19. Define and give an equation to illustrate each of the following substances:
a. a conjugate base
b. a conjugate acid

20. **a.** What is the relationship between the strength of an acid and the strength of its conjugate base?
  b. What is the relationship between the strength of a base and the strength of its conjugate acid?

21. **a.** What trend is there in the favored direction of proton-transfer reactions?
  b. What determines the extent to which a proton-transfer reaction occurs?

22. **a.** What is meant by the term *amphoteric?*
  b. Give an example of a substance or ion that has amphoteric characteristics.

23. For each reaction listed, identify the proton donor or acid and the proton acceptor or base. Label each conjugate acid-base pair.
  a. $CH_3COOH(aq) + H_2O(l) \rightleftarrows H_3O{+}(aq) + CH_3COO^-(aq)$
  b. $HCO_3^-(aq) + H_2O(l) \rightleftarrows H_2CO_3(aq) + OH^-(aq)$
  c. $HNO_3 + SO_4^{2-} \longrightarrow HSO_4^- + NO_3^-$

24. Using the information given in **Figure 3.1,** determine the following relative to HF, H_2S, HNO_3, and CH_3COOH:
  a. strongest acid
  b. weakest acid
  c. strongest conjugate base among the four conjugate bases produced by the acids listed
  d. weakest conjugate base among the four conjugate bases produced by the acids listed

25. 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.

PRACTICE PROBLEMS

26. Complete the following neutralization reactions. Balance each reaction, and then write the overall ionic and net ionic equation for each.
  a. $HCl(aq) + NaOH(aq) \longrightarrow$
  b. $HNO_3(aq) + KOH(aq) \longrightarrow$
  c. $Ca(OH)_2(aq) + HNO_3(aq) \longrightarrow$
  d. $Mg(OH)_2(aq) + HCl(aq) \longrightarrow$

27. Write the formula equation, the overall ionic equation, and the net ionic equation for the neutralization reaction involving aqueous solutions of H_3PO_4 and $Mg(OH)_2$. 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 water and the non-metallic oxide to form an acid.
  a. $CO_2(g) + H_2O(l) \longrightarrow$
  b. $SO_3(g) + H_2O(l) \longrightarrow$
  c. $N_2O_5(g) + H_2O(l) \longrightarrow$

29. Write the formula equation, the overall ionic equation, and the net ionic equation for a neutralization reaction that would form each of the following salts.
  a. $RbClO_4$ **c.** $CaCl_2$
  b. $BaSO_4$ **d.** K_2SO_4

30. Zinc reacts with 100.0 mL of 6.00 M cold, aqueous sulfuric acid through single replacement.
  a. How many grams of zinc sulfate can be produced?
  b. How many liters of hydrogen gas could be released at STP?

31. A 211 g sample of barium carbonate, $BaCO_3$, reacts with a solution of nitric acid to give barium nitrate, carbon dioxide, and water. If the acid is present in excess, what mass and volume of dry carbon dioxide gas at STP will be produced?

32. A seashell that is composed largely of calcium carbonate reacts with a solution of HCl. As a result, 1500 mL of dry CO_2 gas at STP is produced. The other products are $CaCl_2$ and H_2O.
  a. How many grams of $CaCO_3$ are consumed in the reaction?
  b. What volume of 2.00 M 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 it normally is. 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, 3.50×10^8 kg SO_2 may be produced. If this amount of SO_2 were converted to H_2SO_4 according to the two-step process given below, how many kilograms of H_2SO_4 would be produced from such an eruption?

$$SO_2 + \frac{1}{2}O_2 \longrightarrow SO_3$$
$$SO_3 + H_2O \longrightarrow H_2SO_4$$

Mixed Review

REVIEWING MAIN IDEAS

34. Suppose that dilute $HNO_3(aq)$ and $LiOH(aq)$ are mixed in chemically equivalent quantities. Write the following for the resulting 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, H_3PO_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.** $HClO_4$
- **e.** H_3PO_4
- **f.** hydrobromic acid
- **g.** HClO
- **h.** H_2CO_3
- **i.** sulfuric acid

CRITICAL THINKING

38. Analyzing Conclusions In the 18th century, Antoine Lavoisier experimented with oxides, such as CO_2 and SO_2. He observed that they formed acidic solutions. His observations led him to infer that to exhibit acidic behavior, a substance must contain oxygen. However, today that inference is known to be incorrect. Provide evidence to refute Lavoisier's conclusion.

USING THE HANDBOOK

39. Group 16 of the *Elements Handbook* (Appendix A) contains a section covering the acid-base chemistry of oxides. Review this material, and answer the following questions:
- **a.** What types of compounds form acidic oxides?
- **b.** What is an acid anhydride?
- **c.** What are three examples of compounds that are classified as acid anhydrides?
- **d.** What types of compounds form basic oxides? Why are they basic oxides?

40. **a.** Look at Table 7A in the *Elements Handbook* (Appendix A). What periodic trends regarding the acid-base character of oxides do you notice?
- **b.** How is the nature of the product affected by the concentrations of the reactants?

RESEARCH AND 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 solutions are buffered and what kinds of materials are used as buffers. Write a brief report on your findings.

43. Research how to determine whether the soil around your house is acidic or basic using pH paper obtained from your teacher. Write a brief description of what you should do. Then follow the directions, and test the soil. Find one type of plant that would grow well in the type of soil around your home and one that would not grow well.

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 the most widely used antacids. These ingredients act to drive the neutralization reactions. Examine the labels of several common antacids, and identify the active ingredients.

45. Design an experiment that compares three brands of antacids in terms of the speed of symptom relief and amount of acid neutralized.

Standards-Based Assessment

Record your answers on a separate piece of paper.

MULTIPLE CHOICE

1 Which is the correct name for the base $CuOH_2$?

A copper hydroxide
B copper(II) hydroxide
C copper(I) hydroxide
D copper hydroxide(I)

2 Which of the following formula-name pairs below is written correctly?

A $Na(OH)_1$ = sodium hydroxide
B NH_4OH = ammonium hydroxide
C $Sr(OH)_2$ = strontium(II) hydroxide
D $Cu(OH)_2$ = copper(I) hydroxide

3 The following table shows some common bases.

Hydroxides
sodium hydroxide
potassium hydroxide
ammonium hydroxide
lithium hydroxide

How is the hydroxide part of a chemical formula for the bases in the table expressed?

A OH
B NH_3
C O
D H

4 The following table shows some oxides.

Oxides
lithium oxide
calcium oxide
barium oxide
ferrous oxide

How is the oxide part of a chemical formula for the bases in the table expressed?

A OH
B NH_3
C O
D H

5 There are two classes of acids: organic and inorganic. The following table shows both organic and inorganic acids.

Acids
citric acid
hydrochloric acid
hydrofluoric acid
nitric acid

Which one of the acids in the table is an organic acid?

A citric acid
B hydrochloric acid
C hydrofluoric acid
D nitric acid

6 Which of the following properties is a common property of acids?

A pH values < 7
B insoluble in water
C does not conduct electricity
D does not react with metals

7 Which of the following statements is true for the reaction below?

$$HF(aq) + HPO_4^{2-}(aq) \leftrightarrows F^-(aq) + H_2PO_4^-(aq)$$

A HF is the base.
B HPO_4^{2-} is the acid.
C F^- is the conjugate base.
D $H_2PO_4^-$ is the conjugate base.

GRIDDED RESPONSE

8 How many hydrogen atoms in CH_3COOH can be ionized?

Test Tip

Double check (with a calculator, if permitted) all mathematical computations involved in answering a question.