

BIG Idea Acids and bases can be defined in terms of hydrogen ions and hydroxide ions or in terms of electron pairs.

18.1 Introduction to Acids and Bases

MAIN Idea Different models help describe the behavior of acids and bases.

18.2 Strengths of Acids and Bases

MAIN Idea In solution, strong acids and bases ionize completely, but weak acids and bases ionize only partially.

18.3 Hydrogen Ions and pH

MAIN Idea pH and pOH are logarithmic scales that express the concentrations of hydrogen ions and hydroxide ions in aqueous solutions.

18.4 Neutralization

MAIN Idea In a neutralization reaction, an acid reacts with a base to produce a salt and water.

ChemFacts

- The optimal pH for aquarium water varies for different aquatic organisms.
- A pH of 8.2 is generally accepted as the average pH of natural seawater, but maintaining that pH in an aquarium does not ensure that all its inhabitants will thrive.
- The South American cichlid, a freshwater fish, requires a pH range of 6.4 to 7.0, whereas the African cichlid thrives in water with a pH between 8.0 and 9.2.



Measure pH

Evaluate results

Start-Up Activities

LAUNCH Lab

What is in your cupboards?

You can learn something about the properties of products in your household by testing them with strips of paper called litmus paper. Can you separate household products into two groups?



Procedure

1. Read and complete the lab safety form.
2. Place three or four drops of several **household products** into separate wells of a **microplate**. Draw a chart to show the position of each liquid.
3. Test each product with **red and blue litmus paper**. Place two drops of **phenolphthalein** in each sample. Record your observations.

WARNING: Phenolphthalein is flammable. Keep away from flames.

Analysis

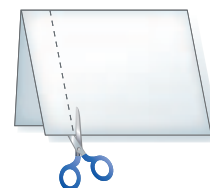
1. **Classify** the products into two groups based on your observations.
2. **Describe** how the groups differ. What can you conclude?

Inquiry Choose one sample that reacted with the phenolphthalein. Can you reverse the reaction? Design an experiment to test your hypothesis.

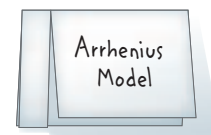
FOLDABLES™ Study Organizer

Acids and Bases Make the following Foldable to compare the main models of acids and bases.

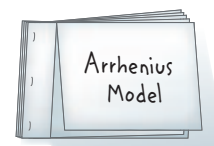
- ▶ **STEP 1** Collect three sheets of paper. Fold the first sheet in half. Measure and draw a line about 3 cm from the left edge. Cut on the line to the fold. Repeat for each sheet of paper.



- ▶ **STEP 2** Label each sheet with the name of a model of acids and bases.



- ▶ **STEP 3** Stack the sheets and staple through all of the narrow flaps.



FOLDABLES Use this Foldable with Section 18.1. As you read the section, make notes about the models of acids and bases, and write generic reactions representative of each model.

Chemistry Online

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- ▶ find the Try at Home Lab, Testing for Ammonia.

Section 18.1

Objectives

- ▶ **Identify** the physical and chemical properties of acids and bases.
- ▶ **Classify** solutions as acidic, basic, or neutral.
- ▶ **Compare** the Arrhenius, Brønsted-Lowry, and Lewis models of acids and bases.

Review Vocabulary

Lewis structure: a model that uses electron-dot structures to show how electrons are arranged in molecules

New Vocabulary

acidic solution
basic solution
Arrhenius model
Brønsted-Lowry model
conjugate acid
conjugate base
conjugate acid-base pair
amphoteric
Lewis model

Introduction to Acids and Bases

MAIN Idea Different models help describe the behavior of acids and bases.

Real-World Reading Link You might not realize it, but acids and bases are two of the most common classifications of substances. You can recognize them by the tart taste of some of your favorite beverages and by the pungent odor of ammonia in some household cleaners.

Properties of Acids and Bases

When ants sense danger to their ant colony, they emit a substance called formic acid that alerts the entire colony. Acids dissolved in rainwater hollow out enormous limestone caverns and destroy valuable buildings and statues over time. Acids flavor many of the beverages and foods you like, and an acid in your stomach helps digest what you eat. Bases also play a role in your life. The soap you use and the antacid tablet you might take for an upset stomach are bases. Many household products, such as those you used in the Launch Lab, are acids and bases.

Physical properties You are probably already familiar with some of the physical properties of acids and bases. For example, you might know that acidic solutions taste sour. Carbonic and phosphoric acids give many carbonated beverages their sharp taste; citric and ascorbic acids give lemons and grapefruit their tartness; and acetic acid makes vinegar taste sour. You might also know that basic solutions taste bitter and feel slippery. Think about how a bar of soap becomes slippery when it gets wet. You should never attempt to identify an acid or a base, or any other substance in the laboratory, by its taste or feel. **Figure 18.1** shows two plants growing in different soils. One grows best in acidic soil, sometimes called “sour” soil. The other thrives in basic or alkaline soil.

■ **Figure 18.1** Rhododendrons flourish in rich, moist soil that is moderately acidic, whereas sempervivum, commonly called hen and chicks, grows best in drier, slightly basic soil.



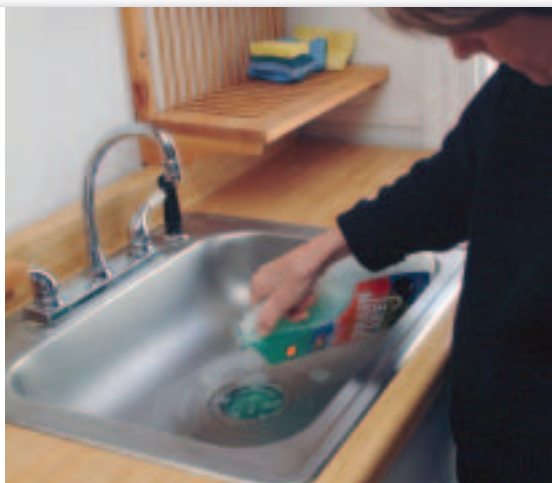
Rhododendron



Sempervivum



Acids turn blue litmus red.



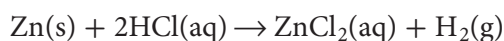
Bases turn red litmus blue.

Electrical conductivity Another physical property of acid and base solutions is the ability to conduct electricity. Pure water is a non-conductor of electricity, but the addition of an acid or base produces ions that cause the resulting solution to become a conductor.

Chemical properties You might have already identified acids and bases by their reaction with litmus paper. Acids can also be identified by their reactions with some metals and metal carbonates.

Reactions with litmus Litmus is one of the dyes commonly used to distinguish solutions of acids and bases, as shown in **Figure 18.2**. Aqueous solutions of acids cause blue litmus paper to turn red. Aqueous solutions of bases cause red litmus paper to turn blue.

Reactions with metals and metal carbonates Magnesium and zinc react with aqueous solutions of acids to produce hydrogen gas. The reaction between zinc and hydrochloric acid is described by the following equation.



Metal carbonates and hydrogen carbonates also react with aqueous solutions of acids to produce carbon dioxide (CO_2) gas. When vinegar is added to baking soda, a foaming reaction occurs between acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) dissolved in the vinegar, and sodium hydrogen carbonate (NaHCO_3). The production of CO_2 gas accounts for the bubbling.



Geologists identify rocks as limestone (primarily CaCO_3) by using a hydrochloric acid solution. If a few drops of the acid produce bubbles of carbon dioxide, the rock contains limestone.

■ **Figure 18.2** The strong acid hydrochloric acid (HCl), also called muriatic acid, is used to clean bricks and concrete. The strong base sodium hydroxide (NaOH) can clear clogged drains.

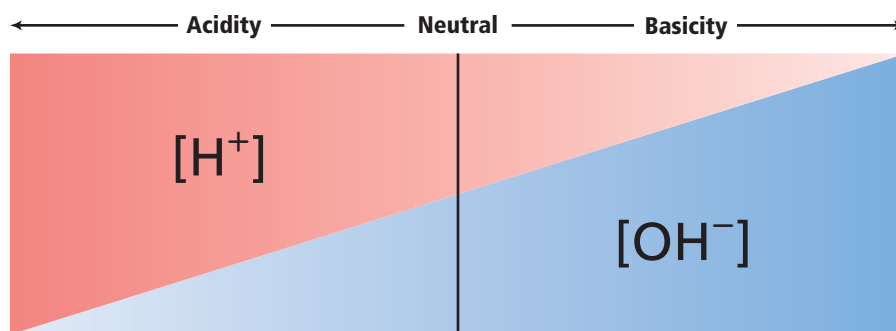
PRACTICE Problems

Extra Practice Page 989 and glencoe.com

- Write balanced equations for the reactions between the following.
 - aluminum and sulfuric acid
 - calcium carbonate and hydrobromic acid
- Challenge** Write the net ionic equation for the reaction in Question 1b.

■ **Figure 18.3** Note how $[H^+]$ and $[OH^-]$ change simultaneously. As $[H^+]$ decreases to the right, $[OH^-]$ increases to the right.

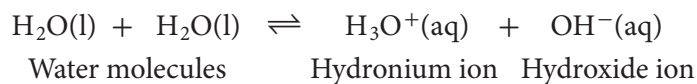
Identify the point in the diagram at which the two ion concentrations are equal.



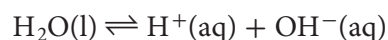
Hydronium and hydroxide ions All water solutions contain hydrogen ions (H^+) and hydroxide ions (OH^-). The relative amounts of the two ions determine whether an aqueous solution is acidic, basic, or neutral. Neutral solutions are neither acidic nor basic.

An **acidic solution** contains more hydrogen ions than hydroxide ions. A **basic solution** contains more hydroxide ions than hydrogen ions. A neutral solution contains equal concentrations of hydrogen ions and hydroxide ions. **Figure 18.3** illustrates these relationships. **Figure 18.4** describes how scientists developed an understanding of acids and bases.

Pure water produces equal numbers of H^+ ions and OH^- ions in a process called self-ionization, in which water molecules react to form a hydronium ion (H_3O^+) and a hydroxide ion.



The hydronium ion is a hydrogen ion which has a water molecule attached to it by a covalent bond. The symbols H^+ and H_3O^+ can be used interchangeably, as this simplified self-ionization equation shows.



■ Figure 18.4 History of Acids and Bases

Current understanding of the structure and behavior of acids and bases is based on the contributions of chemists, biologists, environmental scientists, and inventors over the past 150 years.

1869 Nucleic acids are discovered in cell nuclei. DNA and RNA are examples of nucleic acids.

1909 The development of the pH scale allows scientists to define the acidity of a substance.

1865 The introduction of an antiseptic spray containing carbolic acid marks the beginning of modern antiseptic surgery.

1883 Svante Arrhenius proposes that acids produce hydrogen ions (H^+) and bases produce hydroxide ions (OH^-) when dissolved in water.

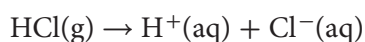
1923 Scientists expand and refine the definition of acids and bases, producing the definitions currently in use.



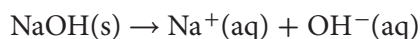
The Arrhenius Model

If pure water itself is neutral, how does an aqueous solution become acidic or basic? The first person to answer this question was Swedish chemist Svante Arrhenius, who in 1883 proposed what is now called the Arrhenius model of acids and bases. The **Arrhenius model** states that an acid is a substance that contains hydrogen and ionizes to produce hydrogen ions in aqueous solution. A base is a substance that contains a hydroxide group and dissociates to produce a hydroxide ion in aqueous solution.

Arrhenius acids and bases As an example of the Arrhenius model of acids and bases, consider what happens when hydrogen chloride gas dissolves in water. HCl molecules ionize to form H^+ ions, which make the solution acidic.



When the ionic compound sodium hydroxide (NaOH) dissolves in water, it dissociates to produce OH^- ions, which make the solution basic.



Although the Arrhenius model is useful in explaining many acidic and basic solutions, it has some shortcomings. For example, ammonia (NH_3) and sodium carbonate (Na_2CO_3) do not contain a hydroxide group, yet both substances produce hydroxide ions in solution and are well-known bases. Sodium carbonate is the compound that causes the alkalinity of Lake Natron, Tanzania, which is shown in **Figure 18.5**. Clearly, a model that includes all bases is needed.



■ **Figure 18.5** Lake Natron in Africa's Great Rift Valley is a naturally basic body of water. Water, laden with dissolved sodium carbonate from surrounding volcanic rocks, drains into the lake but finds no outlet. Evaporation concentrates the mineral leaving a white crust on the surface and strongly alkaline water.

FOLDABLES

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1933–34 Scientists develop portable pH meters.

1980s Silicon-chip pH meters have no glass component. They are now widely used in the food, cosmetic, and pharmaceutical industries.

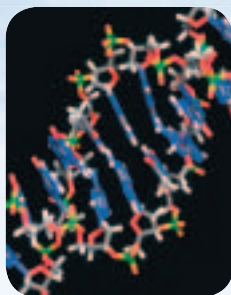
2005 Scientists develop superacids, which are more acidic than 100% sulfuric acid. Applications include producing strong plastics and high-octane gasoline.

1950

1970

1990

2010



1953 James Watson, Francis Crick, and Rosalind Franklin study the nucleic acid DNA, laying the framework for today's biotechnology industry.

1963 Scientists discover acid rain in North America. pH measurements show polluted rain to be 100 times more acidic than unpolluted rain.

Concepts in Motion

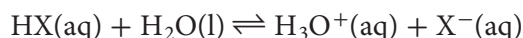
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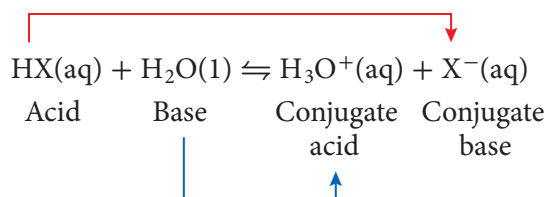
The Brønsted-Lowry Model

Danish chemist Johannes Brønsted and English chemist Thomas Lowry proposed a more inclusive model of acids and bases—a model that focuses on the hydrogen ion (H^+). In the **Brønsted-Lowry model** of acids and bases, an acid is a hydrogen-ion donor. A base is a hydrogen-ion acceptor.

Hydrogen ion donors and acceptors The symbols X and Y represent nonmetallic elements or negative polyatomic ions. Thus, the general formula for an acid can be represented as HX or HY. When a molecule of acid HX dissolves in water, it donates a H^+ ion to a water molecule. The water molecule acts as a base and accepts the H^+ ion.



Upon accepting the H^+ ion, the water molecule becomes an acid, H_3O^+ . The hydronium ion (H_3O^+) is an acid because it has an extra H^+ ion that it can donate. Upon donating its H^+ ion, the acid HX becomes a base, X^- . X^- is a base because it has a negative charge and can readily accept a positive hydrogen ion. Thus, an acid-base reaction in the reverse direction can occur. The acid H_3O^+ can react with the base X^- to form water and HX, establishing the following equilibrium.



■ **Figure 18.6** When a father throws the ball to his son, the father is like a Brønsted-Lowry acid and the son is like a base. After the son catches the ball, he becomes like a conjugate acid.

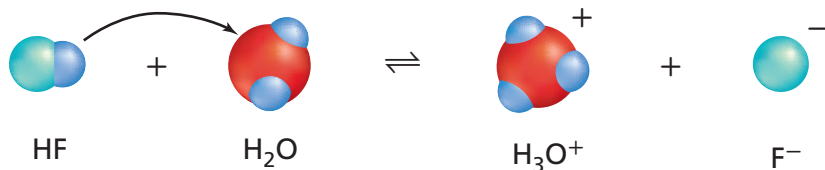


Conjugate acids and bases The forward reaction is the reaction of an acid and a base. The reverse reaction is also the reaction of an acid and a base. The acid and base that react in the reverse reaction are identified under the equation as a conjugate acid and a conjugate base. A **conjugate acid** is the species produced when a base accepts a hydrogen ion. The base H_2O accepts a hydrogen ion from the acid HX and becomes the conjugate acid H_3O^+ . A **conjugate base** is the species that results when an acid donates a hydrogen ion. The acid HX donates its hydrogen ion and becomes the conjugate base X^- . In the reaction shown above, the hydronium ion (H_3O^+) is the conjugate acid of the base H_2O . The X^- ion is the conjugate base of the acid HX. Brønsted-Lowry interactions involve conjugate acid-base pairs. A **conjugate acid-base pair** consists of two substances related to each other by the donating and accepting of a single hydrogen ion.

An analogy for conjugate acid-base pairs is shown in **Figure 18.6**. When the father has the ball in his hand, he is an acid. He throws the ball (a hydrogen ion) to his son. Now his son is the acid because he has the ball (a hydrogen ion) to give away. The father is now a base because he is available to accept the ball (a hydrogen ion). The father is the acid and the son is the base in the forward reaction. In the reverse reaction, the son has the ball and is the conjugate acid while the father is the conjugate base.



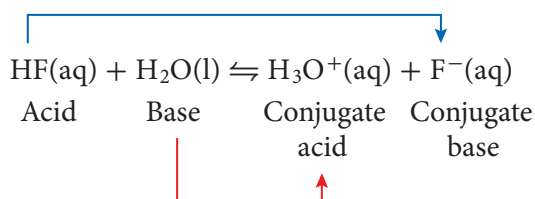
Reading Check Explain how the ion HCO_3^- can be both an acid and a base.



■ **Figure 18.7** Hydrogen fluoride donates a hydrogen ion to a water molecule, so hydrogen fluoride is an acid.

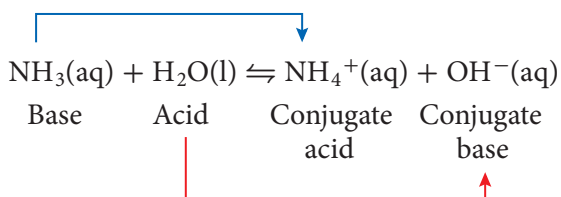
Decide which species is the conjugate base of hydrogen fluoride.

Hydrogen fluoride—a Brønsted-Lowry acid Consider the equation for the ionization of hydrogen fluoride (HF) in water, shown in **Figure 18.7**. What are the conjugate acid-base pairs? Hydrogen fluoride, the acid in the forward reaction, produces its conjugate base F^- , the base in the reverse reaction. Water, the base in the forward reaction, produces its conjugate acid H_3O^+ , the acid in the reverse reaction.



Hydrogen fluoride is used to manufacture a variety of fluorine-containing compounds, such as the nonstick coating on the kitchenware shown in **Figure 18.8**. It is an acid according to both the Arrhenius and Brønsted-Lowry definitions.

Ammonia—a Brønsted-Lowry base All of the acids and bases that fit the Arrhenius definition of acids and bases also fit the Brønsted-Lowry definition. But some other substances that lack a hydroxide group and, therefore, cannot be considered bases according to the Arrhenius definition can be classified as acids according to the Brønsted-Lowry model. One example is ammonia (NH_3). When ammonia dissolves in water, water is a Brønsted-Lowry acid in the forward reaction. Because the NH_3 molecule accepts a H^+ ion to form the ammonium ion (NH_4^+), ammonia is a Brønsted-Lowry base in the forward reaction.



In the reverse reaction, the ammonium ion (NH_4^+) gives up a H^+ ion to form the molecule ammonia and thus acts as a Brønsted-Lowry acid. The ammonium ion is the conjugate acid of the base ammonia. The hydroxide ion accepts a H^+ ion to form a water molecule and is thus a Brønsted-Lowry base. The hydroxide ion is the conjugate base of the acid water.

Water—a Brønsted-Lowry acid and base Recall that when HF dissolves in water, water acts a base; when NH_3 dissolves in water, water acts as an acid. Depending on what other substances are in the solution, water can act as either an acid or a base. Water and other substances that can act as both acids and bases are said to be **amphoteric**.

VOCABULARY

WORD ORIGIN

Conjugate

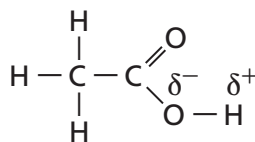
con- prefix; from Latin, meaning *with* or *together*

jugare, verb; from Latin, meaning *to join*

■ **Figure 18.8** To make the smooth, nonstick surface of this kitchenware, hydrogen fluoride is reacted with organic compounds called hydrocarbons to substitute fluorine atoms for hydrogen atoms.



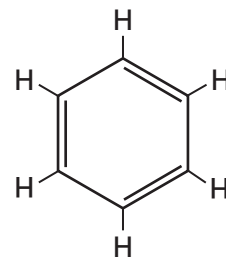
■ **Figure 18.9** Whether a hydrogen is ionizable depends on the polarity of its bond. In acetic acid, oxygen is more electronegative than hydrogen. The bond between oxygen and hydrogen is polar, so the hydrogen atom can ionize in solution. In hydrogen fluoride, fluorine is highly electronegative, so HF is an acid in solution. In benzene, there is little electronegativity difference between the carbon and hydrogen atoms, so benzene is not an acid.



Acetic acid



Hydrogen fluoride



Benzene

PRACTICE Problems

Extra Practice Page 989 and glencoe.com

- Identify the conjugate acid-base pairs in each reaction.
 - $\text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $\text{HBr}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Br}^-(\text{aq})$
 - $\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$
- Challenge** The products of an acid-base reaction are H_3O^+ and SO_4^{2-} . Write a balanced equation for the reaction and identify the conjugate acid-base pairs.

Monoprotic and Polyprotic Acids

From the chemical formulas of HCl and HF, you know that each acid has one hydrogen ion per molecule. An acid that can donate only one hydrogen ion is called a monoprotic acid. Other monoprotic acids are perchloric acid (HClO_4), nitric acid (HNO_3), hydrobromic acid (HBr), and acetic acid (CH_3COOH). Because acetic acid is a monoprotic acid, its formula is often written $\text{HC}_2\text{H}_3\text{O}_2$ to emphasize the fact that only one of the four hydrogen atoms in the molecule is ionizable.

Ionizable hydrogen atoms The difference between acetic acid's ionizable hydrogen atom and the other three hydrogen atoms is that the ionizable atom is bonded to the element oxygen, which is more electronegative than hydrogen. The difference in electronegativity makes the bond between oxygen and hydrogen polar. The structure of acetic acid is shown in **Figure 18.9**, along with structures of the acid HF and the nonacid benzene (C_6H_6). The hydrogen atom in hydrogen fluoride is bonded to the highly electronegative fluorine atom, so the hydrogen-fluorine bond is polar and the fluorine atom is ionizable to a certain extent. However, the hydrogen atoms in benzene are each bonded to a carbon atom. Carbon atoms have about the same electronegativity as hydrogen. These bonds are nonpolar, so benzene is not an acid.

Some acids donate more than one hydrogen ion. For example, sulfuric acid (H_2SO_4) and carbonic acid (H_2CO_3) can donate two hydrogen ions. In each compound, both hydrogen atoms are attached to oxygen atoms by polar bonds. Acids that contain two ionizable hydrogen atoms per molecule are called diprotic acids. Phosphoric acid (H_3PO_4) and boric acid (H_3BO_3) contain three ionizable hydrogen atoms per molecule. Acids with three hydrogen ions to donate are called triprotic acids. The term *polyprotic acid* can be used for any acid that has more than one ionizable hydrogen atom.

Table 18.1

Some Common Acids and Their Conjugate Bases

Acid		Conjugate Base	
Name	Formula	Name	Formula
Hydrochloric acid	HCl	Chloride ion	Cl ⁻
Nitric acid	HNO ₃	Nitrate ion	NO ₃ ⁻
Sulfuric acid	H ₂ SO ₄	Hydrogen sulfate ion	HSO ₄ ⁻
Hydrogen sulfate ion	HSO ₄ ⁻	Sulfate ion	SO ₄ ²⁻
Hydrofluoric acid	HF	Fluoride ion	F ⁻
Hydrocyanic acid	HCN	Cyanide	CN ⁻
Acetic acid	HC ₂ H ₃ O ₂	Acetate ion	C ₂ H ₃ O ₂ ⁻
Phosphoric acid	H ₃ PO ₄	Dihydrogen phosphate ion	H ₂ PO ₄ ⁻
Dihydrogen phosphate ion	H ₂ PO ₄ ⁻	Hydrogen phosphate ion	HPO ₄ ²⁻
Hydrogen phosphate ion	HPO ₄ ²⁻	Phosphate ion	PO ₄ ³⁻
Carbonic acid	H ₂ CO ₃	Hydrogen carbonate ion	HCO ₃ ⁻
Hydrogen carbonate ion	HCO ₃ ⁻	Carbonate ion	CO ₃ ²⁻

All polyprotic acids ionize in steps. The three ionizations of phosphoric acid are described by the following equations.

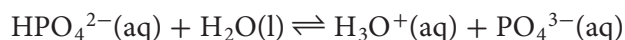
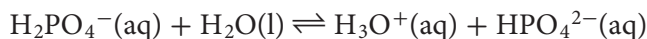
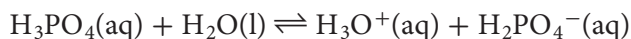


Table 18.1 shows some common monoprotic and polyprotic acids.

The Lewis Model

Notice that all substances classified as acids and bases by the Arrhenius model are classified as acids and bases by the Brønsted-Lowry model. In addition, some substances *not* classified as bases by the Arrhenius model *are* classified as bases by the Brønsted-Lowry model.

Perhaps you will not be surprised, then, you to learn that an even more general model of acids and bases was proposed by American chemist G. N. Lewis (1875–1946). Recall that Lewis developed the electron-pair theory of chemical bonding and introduced Lewis structures to keep track of the electrons in atoms and molecules. He applied his electron-pair theory of chemical bonding to acid-base reactions. Lewis proposed that an acid is an ion or molecule with a vacant atomic orbital that can accept (share) an electron pair. A base is an ion or molecule with a lone electron pair that it can donate (share). According to the **Lewis model**, a Lewis acid is an electron-pair acceptor and a Lewis base is an electron-pair donor. Note that the Lewis model includes all the substances classified as Brønsted-Lowry acids and bases and many more.

FOLDABLES

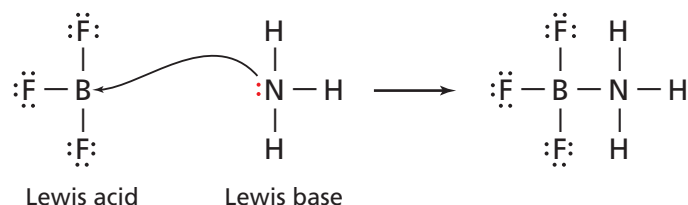
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Electron pair donors and acceptors Consider the reaction between a hydrogen ion (H^+) and a fluoride ion (F^-) to form a hydrogen fluoride (HF) molecule. The role of the electron pair is illustrated through the following Lewis structures.



In this reaction, the H^+ ion is the Lewis acid. Its vacant 1s orbital accepts an electron pair from the F^- ion. The fluoride ion is the Lewis base. It donates a lone electron pair to form the hydrogen-fluorine bond in HF. Note that this reaction also conforms to the Brønsted-Lowry model of acids and bases because H^+ can be considered a hydrogen-ion donor and F^- a hydrogen-ion acceptor.

It might surprise you to learn that the reaction of gaseous boron trifluoride (BF_3) with gaseous ammonia (NH_3) to form BF_3NH_3 is a Lewis acid-base reaction.

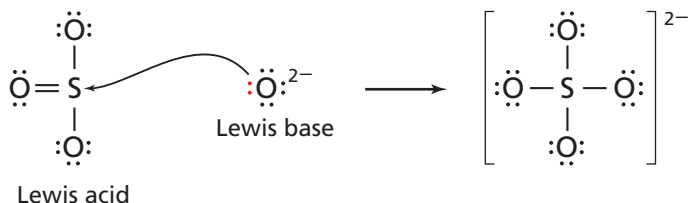


Recall from Chapter 8 that the boron atom in BF_3 has six valence electrons, so a vacant orbital can accept an electron pair from a Lewis base.

Another Lewis acid-base reaction occurs when gaseous sulfur trioxide (SO_3) is brought into contact with solid magnesium oxide (MgO).



The acid-base part of the reaction involves sulfur trioxide (SO_3) and the oxide ion (O^{2-}) of magnesium oxide. The product is the sulfate ion.



Note that the SO_3 molecule, a Lewis acid, accepts an electron pair from the O^{2-} ion, a Lewis base. The Arrhenius, Brønsted-Lowry, and Lewis acid-base models are summarized in **Table 18-2**.

<i>Concepts in Motion</i>		
Table 18.2	Three Models for Acids and Bases	Interactive Table Explore acids and bases at glencoe.com .
Model	Acid Definition	Base Definition
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

VOCABULARY

ACADEMIC VOCABULARY

Conform

to be similar or identical

Their behavior conforms to the expectations of the community.

The reaction of SO_3 and MgO is important because it produces magnesium sulfate, a salt that forms the heptahydrate known as Epsom salt ($\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$). Epsom salt has many uses, including soothing sore muscles and acting as a plant nutrient. The reaction to form magnesium sulfate also has environmental applications. When MgO is injected into the flue gases of coal-fired power plants, such as the one shown in **Figure 18.10**, it reacts with and removes SO_3 . If SO_3 is allowed to enter the atmosphere, it can combine with water in the air to form sulfuric acid, which falls to Earth as acid precipitation.

Connection to Earth Science **Anhydrides** Like the SO_3 molecules you have been reading about, carbon dioxide gas molecules in the air also combine with water molecules in precipitation to form an acid called carbonic acid (H_2CO_3). When the acidic rainwater reaches the ground, some sinks into the soil and reaches limestone bedrock, where it slowly dissolves the limestone. Over the course of thousands of years, the dissolution of limestone creates huge underground caverns. Within a cavern, groundwater might drip from the ceiling and deposit some of the dissolved limestone. Deposits shaped like icicles that form on the ceiling are called stalactites. Rounded masses rising from the floor are called stalagmites.

The formation of caverns occurs because carbon dioxide is an acid anhydride. An acid anhydride is an oxide that can combine with water to form an acid. Other oxides combine with water to form bases. For example, calcium oxide (CaO , lime) forms the base calcium hydroxide $\text{Ca}(\text{OH})_2$. In general, oxides of metallic elements form bases; oxides of nonmetals form acids.



■ **Figure 18.10** Sulfur trioxide, a waste gas from the burning of coal, can be removed from smoke-stack gases by combining it with magnesium oxide in a Lewis acid-base reaction. Note that while there is a good deal of steam coming from the cooling towers, there is little visible smoke from the smokestack.

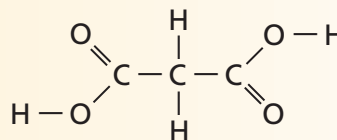
Section 18.1 Assessment

Section Summary

- ▶ The concentrations of hydrogen ions and hydroxide ions determine whether an aqueous solution is acidic, basic, or neutral.
- ▶ An Arrhenius acid must contain an ionizable hydrogen atom. An Arrhenius base must contain an ionizable hydroxide group.
- ▶ A Brønsted-Lowry acid is a hydrogen ion donor. A Brønsted-Lowry base is a hydrogen ion acceptor.
- ▶ A Lewis acid accepts an electron pair. A Lewis base donates an electron pair.

5. **MAIN Idea** **Explain** why many Lewis acids and bases are not classified as Arrhenius or Brønsted-Lowry acids and bases.
6. **Compare** the physical and chemical properties of acids and bases.
7. **Explain** how the concentrations of hydrogen ions and hydroxide ions determine whether a solution is acidic, basic, or neutral.
8. **Explain** why many compounds that contain one or more hydrogen atoms are not classified as Arrhenius acids.
9. **Identify** the conjugate acid-base pairs in the following equation.

$$\text{HNO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{NO}_2^- + \text{H}_3\text{O}^+$$
10. **Write** the Lewis structure for phosphorus trichloride (PCl_3). Is PCl_3 a Lewis acid, a Lewis base, or neither?
11. **Interpret Scientific Illustrations** In the accompanying structural formula, identify any hydrogen atoms that are likely to be ionizable.



Section 18.2

Objectives

- ▶ **Relate** the strength of an acid or base to its degree of ionization.
- ▶ **Compare** the strength of a weak acid with the strength of its conjugate base.
- ▶ **Explain** the relationship between the strengths of acids and bases and the values of their ionization constants.

Review Vocabulary

electrolyte: an ionic compound whose aqueous solution conducts an electric current

New Vocabulary

strong acid
weak acid
acid ionization constant
strong base
weak base
base ionization constant

Strengths of Acids and Bases

MAIN Idea In solution, strong acids and bases ionize completely, but weak acids and bases ionize only partially.

Real-World Reading Link The success of a pass in a football game depends on the passer and the receiver. How ready is the passer to pass the ball? How ready is the receiver to receive the ball? Similarly, in acid and base reactions, the progress of a reaction depends on how readily the acid donates a hydrogen ion and how readily the base accepts a hydrogen ion.

Strengths of Acids

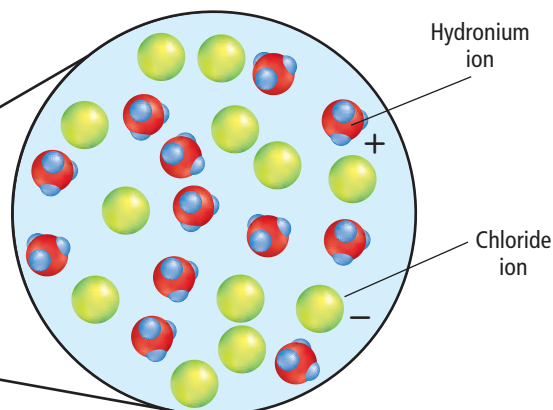
One of the properties of acidic and basic solutions is that they conduct electricity. What can electrical conductivity tell you about the hydrogen ions and hydroxide ions in these aqueous solutions?

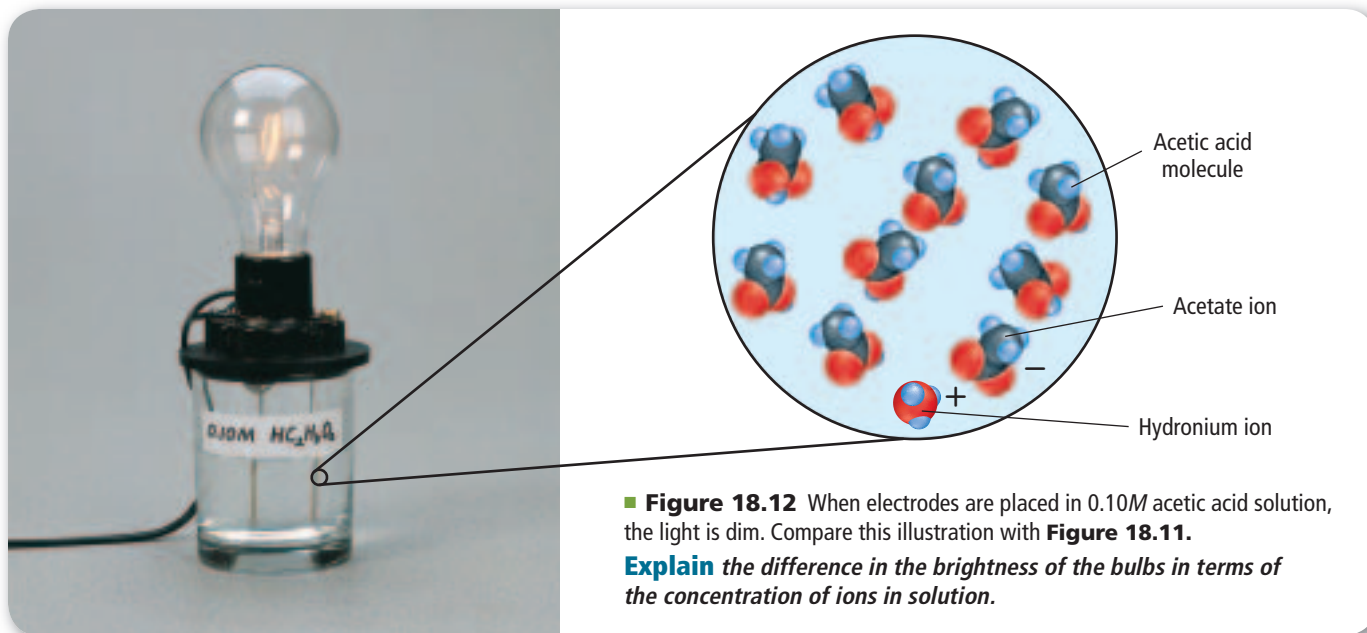
Suppose you test the electrical conductivities of $0.10M$ aqueous solutions of hydrochloric acid and acetic acid. The glow of the bulb in **Figure 18.11** indicates that the solution conducts electricity. However, if you compare the brightness of the bulb connected to the HCl solution in **Figure 18.11** with that of the bulb connected to the $HC_2H_3O_2$ solution in **Figure 18.12**, you should notice a difference. The $0.10M$ HCl solution conducts electricity better than the $0.10M$ $HC_2H_3O_2$ solution. Why is this true if the concentrations of the two acids are both $0.10M$?

Strong acids The answer is that ions carry electric current through the solution and all the HCl molecules contained in the solution are ionized completely into hydronium ions and chloride ions. Acids that ionize completely are called **strong acids**. Because strong acids produce the maximum number of ions, they are good conductors of electricity.



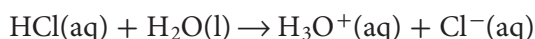
■ **Figure 18.11** The light glows brightly when electrodes are placed in $0.10M$ hydrochloric acid solution because all of the HCl is in the form of hydronium ions and chloride ions.





■ **Figure 18.12** When electrodes are placed in 0.10M acetic acid solution, the light is dim. Compare this illustration with **Figure 18.11**.
Explain the difference in the brightness of the bulbs in terms of the concentration of ions in solution.

The ionization of hydrochloric acid in water can be represented by the following equation, which has a single arrow pointing to the right. Recall that a single arrow means that a reaction goes to completion.



Because strong acids produce the maximum number of ions, their solutions are good conductors of electricity. The names and ionization equations for some strong acids are shown in **Table 18.3**.

Weak acids If the brightly lit bulb of the apparatus containing the HCl solution is due to the large number of ions in solution, shown in **Figure 18.11**, then the weakly lit bulb of the apparatus containing the HC₂H₃O₂ solution, shown in **Figure 18.12**, must mean that the acetic acid solution has fewer ions. Because the two solutions have the same molar concentrations, you can conclude that acetic acid does not ionize completely. An acid that ionizes only partially in dilute aqueous solution is a **weak acid**. Weak acids produce fewer ions and thus cannot conduct electricity as well as strong acids. **Table 18.3** shows ionization equations for some common weak acids.

Concepts In Motion

Interactive Table Explore ionization equations at glencoe.com.

Table 18.3 Ionization Equations

Strong Acids		Weak Acids	
Name	Ionization Equation	Name	Ionization Equations
Hydrochloric	$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$	Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$
Hydroiodic	$\text{HI} \rightarrow \text{H}^+ + \text{I}^-$	Acetic	$\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$
Perchloric	$\text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$	Hydrosulfuric	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$
Nitric	$\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$	Carbonic	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$
Sulfuric	$\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^-$	Hypochlorous	$\text{HClO} \rightleftharpoons \text{H}^+ + \text{ClO}^-$

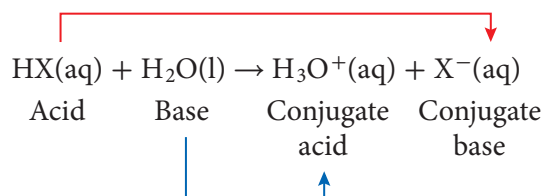
Real-World Chemistry Hydrogen Cyanide



A deadly compound

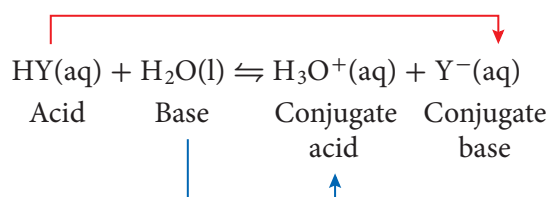
Hydrogen cyanide (HCN) is a poisonous gas found in the exhaust of vehicles, in tobacco and wood smoke, and in smoke from burning nitrogen-containing plastics. Some insects such as millipedes and burnet moths release hydrogen cyanide as a defense mechanism. A solution of hydrogen cyanide in water is called hydrocyanic acid. Fruits that have a pit, such as cherries or peaches, contain cyanohydrins, which convert to hydrocyanic acid in the digestive system if the pits are eaten. However, no hydrocyanic acid is produced in the flesh of these fruits, so the fruit can safely be eaten.

Acid strength and the Brønsted-Lowry model Can the Brønsted-Lowry model explain why HCl ionizes completely but $\text{HC}_2\text{H}_3\text{O}_2$ forms only a few ions? Consider the ionization of any strong acid, HX. Remember that the acid on the reactant side of the equation produces a conjugate base on the product side. Similarly, the base on the reactant side produces a conjugate acid.

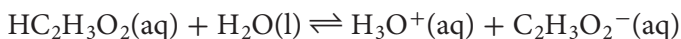


HX represents a strong acid and its conjugate base is weak. That is, HX is nearly 100% ionized because H_2O is a stronger base (in the forward reaction) than is the conjugate base X^- (in the reverse reaction). In other words, the ionization equilibrium lies almost completely to the right because the base H_2O has a much greater attraction for the H^+ ion than does the base X^- . Think of this as the battle of the bases: Which of the two (H_2O or X^-) has a greater attraction for the hydrogen ion? In the case of all strong acids, water is the stronger base. Notice that the equation is shown with a single arrow to the right.

How does the situation differ for any weak acid, HY?



The ionization equilibrium for a weak acid lies far to the left because the conjugate base Y^- has a greater attraction for the H^+ ion than does the base H_2O . In the battle of the bases, the conjugate base Y^- (in the reverse reaction) is stronger than the base H_2O (in the forward reaction) and manages to capture the H^+ ion. In the case of acetic acid, the conjugate base $\text{C}_2\text{H}_3\text{O}_2^-$ (in the reverse reaction) has a stronger attraction for hydrogen ions than does the base H_2O (in the forward reaction).



Notice that the equation is shown with equilibrium arrows.

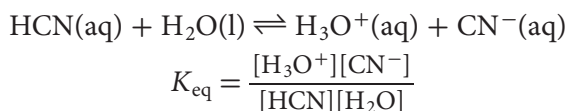


Reading Check Summarize the important difference between strong acids and weak acids in terms of the battle of the bases.

Acid ionization constants Although the Brønsted-Lowry model helps explain acid strength, the model does not provide a quantitative way to express the strength of an acid or to compare the strengths of various acids. The equilibrium constant expression provides the quantitative measure of acid strength.

As you have read, a weak acid produces an equilibrium mixture of molecules and ions in aqueous solution. Thus, the equilibrium constant, K_{eq} , provides a quantitative measure of the degree of ionization of the acid. Consider hydrocyanic acid (HCN), also known as prussic acid which is used in dyeing, engraving, and tempering steel.

The ionization equation and equilibrium constant expression for hydrocyanic acid are as follows.



The concentration of liquid H_2O in the denominator of the expression is considered to be constant in dilute aqueous solutions, so it can be combined with K_{eq} to give a new equilibrium constant, K_{a} .

$$K_{\text{eq}} [\text{H}_2\text{O}] = K_{\text{a}} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = 6.2 \times 10^{-10}$$

K_{a} is called the acid ionization constant. The **acid ionization constant** is the value of the equilibrium constant expression for the ionization of a weak acid. Like all equilibrium constants, the value of K_{a} indicates whether reactants or products are favored at equilibrium. For weak acids, the concentrations of the ions (products) in the numerator tend to be small compared to the concentration of un-ionized molecules (reactant) in the denominator. The weakest acids have the smallest K_{a} values because their solutions have the lowest concentrations of ions and the highest concentrations of un-ionized acid molecules. K_{a} values and ionization equations for several weak acids are listed in **Table 18.4**. Note that polyprotic acids are not necessarily strong acids for any of their ionizations. Each ionization of a polyprotic acid has a K_{a} value, and the values decrease for each successive ionization.

CAREERS IN CHEMISTRY

Nursery Worker The propagation and growth of plants is the primary job of a nursery worker. This involves planting, pruning, transplanting, and selling all kinds of plant material. A nursery worker must know what nutrients are needed for optimum plant growth and what soil conditions, including acidity, foster the strongest growth for each kind of plant. For more information on chemistry careers, visit glencoe.com.

PRACTICE Problems

Extra Practice Page 989 and glencoe.com

12. Write ionization equations and acid ionization constant expressions for each acid.
- a. HClO_2 b. HNO_2 c. HIO
13. Write the first and second ionization equations for H_2SeO_3 .
14. **Challenge** Given the expression $K_{\text{a}} = \frac{[\text{AsO}_4^{3-}][\text{H}_3\text{O}^+]}{[\text{HCN}]}$, write the balanced equation for the corresponding reaction.

Table 18.4

Ionization Constants for Weak Acids

Acid	Ionization Equation	K_{a} (298 K)
Hydrosulfuric, first ionization	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$	8.9×10^{-8}
Hydrosulfuric, second ionization	$\text{HS}^- \rightleftharpoons \text{H}^+ + \text{S}^{2-}$	1×10^{-19}
Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$	6.3×10^{-4}
Hydrocyanic	$\text{HCN} \rightleftharpoons \text{H}^+ + \text{CN}^-$	6.2×10^{-10}
Acetic	$\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$	1.8×10^{-5}
Carbonic, first ionization	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$	4.5×10^{-7}
Carbonic, second ionization	$\text{HCO}_3^- \rightleftharpoons \text{H}^+ + \text{CO}_3^{2-}$	4.7×10^{-11}

Compare Acid Strengths

How can you determine the relative strengths of acid solutions?

Procedure



1. Read and complete the lab safety form.
2. Use a **10-mL graduated cylinder** to measure 3 mL of **glacial acetic acid**. Use a **dropping pipette** to transfer the acid into Well A1 of a **24-well microplate**.
WARNING: Glacial acetic acid is corrosive and toxic by inhalation. Handle with caution.
3. Lower the electrodes of a **conductivity tester** into Well A1. Record your results.
4. Rinse the graduated cylinder and pipette with water. Measure 3 mL of **6.0M acetic acid**, and transfer it to Well A2 of the microplate. Test and record the conductivity of the solution.

5. Repeat Step 4 with **1.0M acetic acid** and **0.10M acetic acid** using wells A3 and A4, respectively.

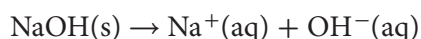
Analysis

1. **Write** the equation for the ionization of acetic acid in water and the equilibrium constant expression ($K_{\text{eq}} = 1.8 \times 10^{-5}$). What does the size of K_{eq} indicate about the degree of ionization?
2. **Explain** whether the following approximate percent ionizations fit your laboratory results: glacial acetic acid, 0.1%; 6.0M acetic acid, 0.2%; 1.0M acetic acid, 0.4%; 0.1M acetic acid, 1.3%.
3. **State** a hypothesis that explains your observations using your answer to Question 2.
4. Utilize your hypothesis to draw a conclusion about the need to use large amounts of water for rinsing when acid spills on living tissue.

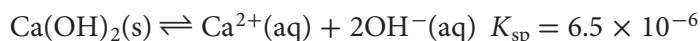
Strengths of Bases

What you have read about acids can be applied to bases, except that OH^- ions, rather than H^+ ions, are involved. For example, the conductivity of a base depends on the extent to which the base produces OH^- ions in aqueous solution.

Strong bases A base that dissociates entirely into metal ions and hydroxide ions is known as a **strong base**. Therefore, metallic hydroxides, such as sodium hydroxide (NaOH), are strong bases.



Some metallic hydroxides, such as calcium hydroxide (Ca(OH)_2) have low solubility and thus are poor sources of OH^- ions. Note that the solubility product constant, K_{sp} , for calcium hydroxide (Ca(OH)_2) is small, indicating that few OH^- ions are present in a saturated solution.



Nevertheless, calcium hydroxide and other slightly soluble metallic hydroxides are considered strong bases because all of the compound that dissolves is completely dissociated. The dissociation equations for several strong bases are listed in **Table 18.5**.

Weak bases In contrast to strong bases, a **weak base** ionizes only partially in dilute aqueous solution. For example, methylamine (CH_3NH_2) reacts with water to produce an equilibrium mixture of CH_3NH_2 molecules, CH_3NH_3^+ ions, and OH^- ions.

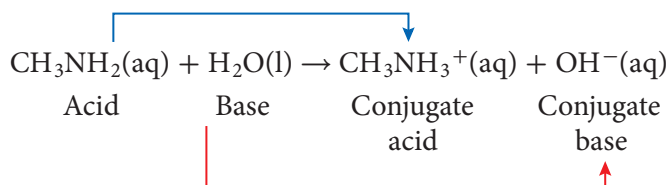


Table 18.5

Dissociation Equations for Strong Bases

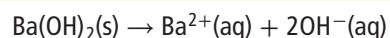
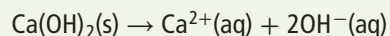
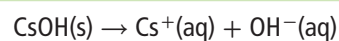
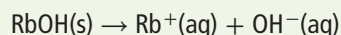
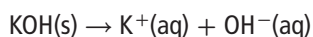
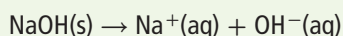


Table 18.6**Ionization Constants of Weak Bases**

Base	Ionization Equation	K_b (298 K)
Ethylamine	$C_2H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_2H_5NH_3^+(aq) + OH^-(aq)$	5.0×10^{-4}
Methylamine	$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$	4.3×10^{-4}
Ammonia	$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$	2.5×10^{-5}
Aniline	$C_6H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_6H_5NH_3^+(aq) + OH^-(aq)$	4.3×10^{-10}

This equilibrium lies far to the left because the base, CH_3NH_2 , is weak and the conjugate base, OH^- ion, is strong. The hydroxide ion has a greater attraction for a hydrogen ion than a molecule of methyl amine has.

Base ionization constants Like weak acids, weak bases also form equilibrium mixtures of molecules and ions in aqueous solution. The equilibrium constant provides a measure of the extent of the base's ionization. The equilibrium constant for the ionization of methylamine in water is defined by the following equilibrium constant expression.

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

The **base ionization constant**, K_b , is the value of the equilibrium constant expression for the ionization of a base. The smaller the value of K_b , the weaker the base. K_b values and ionization equations for several weak bases are listed in **Table 18.6**.

PRACTICE ProblemsExtra Practice Page 989 and glencoe.com

- 15.** Write ionization equations and base ionization constant expressions for the following bases.
- a.** hexylamine ($C_6H_{13}NH_2$) **c.** carbonate ion (CO_3^{2-})
b. propylamine ($C_3H_7NH_2$) **d.** hydrogen sulfite ion (HSO_3^-)
- 16. Challenge** Write an equation for a base equilibrium in which the base in the forward reaction is PO_4^{3-} and the base in the reverse reaction is OH^- .

Section 18.2 Assessment**Section Summary**

- ▶ Strong acids and strong bases are completely ionized in a dilute aqueous solution. Weak acids and weak bases are partially ionized in a dilute aqueous solution.
- ▶ For weak acids and weak bases, the value of the acid or base ionization constant is a measure of the strength of the acid or base.

- **17. MAIN Idea** Describe the contents of dilute aqueous solutions of the strong acid HI and the weak acid HCOOH.
- **18. Relate** the strength of a weak acid to the strength of its conjugate base.
- **19. Identify** the conjugate acid-base pairs in each equation.
 - a.** $HCOOH(aq) + H_2O(l) \rightleftharpoons HCOO^-(aq) + H_3O^+(aq)$
 - b.** $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
- **20. Explain** what the K_b for aniline ($C_6H_5NH_2$) tells you ($K_b = 4.3 \times 10^{-10}$).
- **21. Interpret Data** Use the data in **Table 18.4** to put the seven acids in order according to increasing electrical conductivity.

Section 18.3

Objectives

- **Explain** pH and pOH.
- **Relate** pH and pOH to the ion product constant for water.
- **Calculate** the pH and pOH of aqueous solutions.

Review Vocabulary

Le Châtelier's principle: states that if a stress is applied to a system at equilibrium, the system shifts in the direction that relieves the stress

New Vocabulary

ion product constant for water
pH
pOH

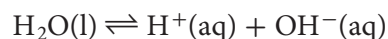
Hydrogen Ions and pH

MAIN Idea pH and pOH are logarithmic scales that express the concentrations of hydrogen ions and hydroxide ions in aqueous solutions.

Real-World Reading Link Think of two children on a seesaw. When one side of a seesaw goes up, the other side goes down. Sometimes, the seesaw is balanced in the middle. The concentrations of hydrogen ions and hydroxide ions in water solutions behave in a similar way.

Ion Product Constant for Water

Recall that pure water contains equal concentrations of H^+ and OH^- ions produced by self-ionization. **Figure 18.13** shows that in self-ionization, equal numbers of hydronium and hydroxide ions are formed. The equation for the equilibrium can be simplified in the following way.



Writing K_w The double arrow indicates that this is an equilibrium. Recall that the equilibrium constant expression is written by placing the concentrations of the products in the numerator and the concentrations of the reactants in the denominator. In this case, all terms are to the first power because all the coefficients in the balanced chemical equation are 1. The concentration of pure water is constant, so $[\text{H}_2\text{O}]$ does not appear in the denominator.

The Ion Product of Water

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

K_w is the ion product constant for water. $[\text{H}^+]$ represents the concentration of the hydrogen ion. $[\text{OH}^-]$ represents the concentration of the hydroxide ion.

In dilute aqueous solutions, the product of the concentrations of the hydrogen ion and the hydroxide ion equals K_w .


The expression for K_w is a special equilibrium constant expression that applies only to water. The constant K_w is called the ion product constant for water. The **ion product constant for water** is the value of the equilibrium constant expression for the self-ionization of water. Experiments show that in pure water at 298 K, $[\text{H}^+]$ and $[\text{OH}^-]$ are both equal to $1.0 \times 10^{-7} \text{M}$. Therefore, at 298 K, the value of K_w is 1.0×10^{-14} .

$$K_w = [\text{H}^+][\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$$
$$K_w = 1.0 \times 10^{-14}$$

■ **Figure 18.13** In the self-ionization of water, one water molecule acts as an acid, and the other acts as a base.



K_w and Le Châtelier's Principle The product of $[H^+]$ and $[OH^-]$ always equals 1.0×10^{-14} at 298 K. This means that if the concentration of H^+ ions increases, the concentration of OH^- ions must decrease. Similarly, an increase in the concentration of OH^- ions causes a decrease in the concentration of H^+ ions. Think about these changes in terms of Le Châtelier's principle, which you read about in Chapter 17. Adding extra hydrogen ions to water at equilibrium is a stress on the system. The system reacts in a way to relieve the stress. The added H^+ ions react with OH^- ions to form more water molecules. Thus, the concentration of OH^- ions decreases. Example Problem 18.1 shows how you can use K_w to calculate the concentration of either H^+ or OH^- if you know the concentration of the other ion.

 **Reading Check** Explain why K_w does not change when the concentration of hydrogen ions increases.

EXAMPLE Problem 18.1

Calculate $[H^+]$ and $[OH^-]$ Using K_w At 298 K, the H^+ ion concentration in a cup of coffee is $1.0 \times 10^{-5}M$. What is the OH^- ion concentration in the coffee? Is the coffee acidic, basic, or neutral?

Math Handbook

Solving Algebraic Equations
pages 954–955

1 Analyze the Problem

You are given the concentration of the H^+ ion, and you know that K_w equals 1.0×10^{-14} . You can use the ion product constant expression to solve for $[OH^-]$. Because $[H^+]$ is greater than 1.0×10^{-7} , you can predict that $[OH^-]$ will be less than 1.0×10^{-7} .

Known

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

$$K_w = 1.0 \times 10^{-14}$$

Unknown

$$[OH^-] = ? \text{ mol/L}$$

2 Solve for the Unknown

Use the ion product constant expression.

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

State the ion product expression.

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

Solve for $[OH^-]$.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9} \text{ mol/L}$$

Substitute $K_w = 1.0 \times 10^{-14}$. Substitute $[H^+] = 1.0 \times 10^{-5}M$ and solve.

Because $[H^+] > [OH^-]$, the coffee is acidic.

3 Evaluate the Answer

The answer is correctly stated with two significant figures because $[H^+]$ and K_w each have two significant figures. As predicted, $[OH^-]$ is less than $1.0 \times 10^{-7} \text{ mol/L}$.

PRACTICE Problems

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22. The concentration of either the H^+ ion or the OH^- ion is given for four aqueous solutions at 298 K. For each solution, calculate $[H^+]$ or $[OH^-]$. State whether the solution is acidic, basic, or neutral.

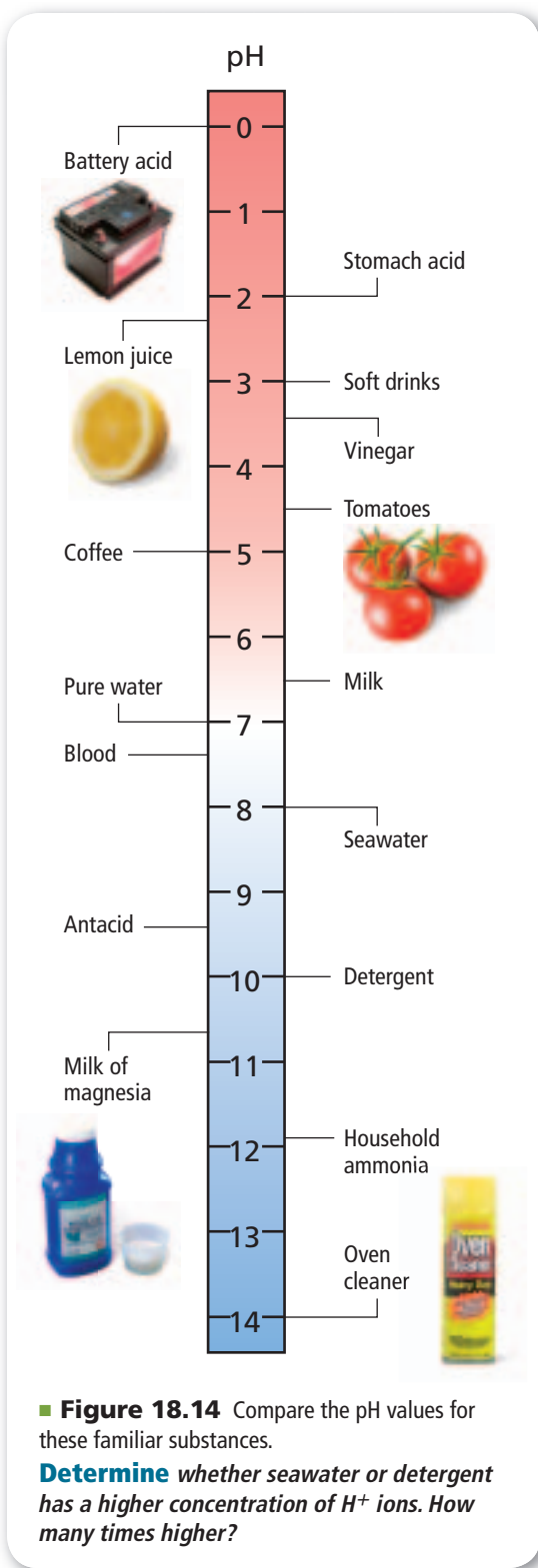
a. $[H^+] = 1.0 \times 10^{-13}M$

c. $[OH^-] = 1.0 \times 10^{-3}M$

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

d. $[H^+] = 4.0 \times 10^{-5}M$

23. Challenge Calculate the number of H^+ ions and the number of OH^- ions in 300 mL of pure water at 298 K.



pH and pOH

Concentrations of H^+ ions are often small numbers expressed in scientific notation. Because these numbers are cumbersome, chemists adopted an easier way to express H^+ ion concentrations.

What is pH? Chemists express the concentration of hydrogen ions using a pH scale based on common logarithms. The **pH** of a solution is the negative logarithm of the hydrogen ion concentration.

pH

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

$[H^+]$ represents the hydrogen ion concentration.

The pH of a solution equals the negative logarithm of the hydrogen ion concentration.

At 298 K, acidic solutions have pH values below 7. Basic solutions have pH values above 7. Thus, a solution with a pH of 0.0 is strongly acidic; a solution with a pH of 14.0 is strongly basic. The logarithmic nature of the pH scale means that a change of one pH unit represents a tenfold change in ion concentration. A solution having a pH of 3.0 has ten times the hydrogen ion concentration of a solution with a pH of 4.0. The pH scale and pH values of some common substances are shown in **Figure 18.14**.

What is pOH? Sometimes it is convenient to express the basicity or alkalinity of a solution on a pOH scale that mirrors the relationship between pH and $[H^+]$. The **pOH** of a solution is the negative logarithm of the hydroxide ion concentration.

pOH

$$pOH = -\log [OH^-]$$

$[OH^-]$ represents the hydroxide ion concentration.

The pOH of a solution equals the negative logarithm of the hydroxide ion concentration.

At 298 K, a solution with a pOH less than 7.0 is basic; a solution with a pOH of 7.0 is neutral; and a solution with a pOH greater than 7.0 is acidic. As with the pH scale, a change of one pOH unit expresses a tenfold change in ion concentration.

A simple relationship between pH and pOH makes it easy to calculate either quantity if the other is known.

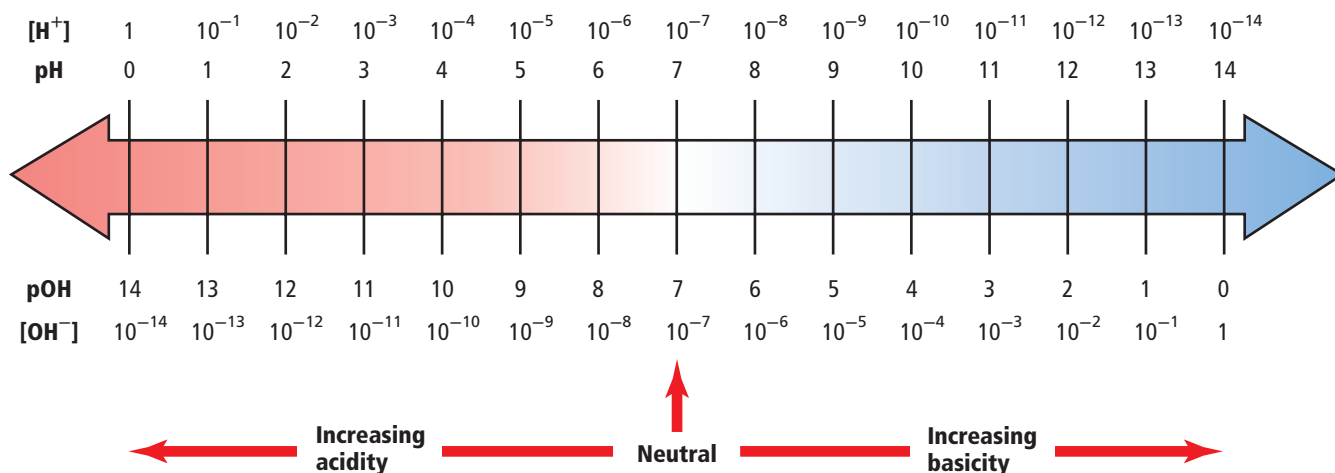
How pH and pOH Are Related

$$pH + pOH = 14.00$$

pH represents $-\log [H^+]$.
pOH represents $-\log [OH^-]$.

The sum of pH and pOH is 14.00.

Figure 18.15 illustrates the relationship between pH and the H^+ concentration and the relationship between pOH and OH^- concentration at 298 K.



■ **Figure 18.15** Study this diagram to sharpen your understanding of pH and pOH. Note that at each vertical position, the sum of pH (above the arrow) and pOH (below the arrow) equals 14. Also note that at every position, the product of [H⁺] and [OH⁻] equals 10⁻¹⁴.

EXAMPLE Problem 18.2

Calculate pH from [H⁺] What is the pH of a neutral solution at 298 K?

Math Handbook

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pages 966–967

1 Analyze the Problem

In a neutral solution at 298 K, [H⁺] = 1.0 × 10⁻⁷ M.
You must find the negative log of [H⁺].

Known

$$[\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$$

Unknown

$$\text{pH} = ?$$

2 Solve for the Unknown

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

State the equation for pH.

$$\text{pH} = -\log (1.0 \times 10^{-7})$$

Substitute [H⁺] = 1.0 × 10⁻⁷ M.

The pH of the neutral solution at 298 K is **7.00**.

3 Evaluate the Answer

Values for pH are expressed with as many decimal places as the number of significant figures in the H⁺ ion concentration. The pH is correctly stated with two decimal places.



Personal Tutor For an online tutorial on logarithms, visit glencoe.com.

PRACTICE Problems

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24. Calculate the pH of solutions having the following ion concentrations at 298 K.
- a. [H⁺] = 1.0 × 10⁻² M b. [H⁺] = 3.0 × 10⁻⁶ M
25. Calculate the pH of aqueous solutions with the following [H⁺] at 298 K.
- a. [H⁺] = 0.0055 M b. [H⁺] = 0.000084 M
26. **Challenge** Calculate the pH of a solution having [OH⁻] = 8.2 × 10⁻⁶ M.

EXAMPLE Problem 18.3

Calculate pOH and pH from $[\text{OH}^-]$ In **Figure 18.16**, a cow is being fed straw and hay that has been treated with ammonia. The addition of ammonia to animal feed promotes protein growth. Another use of ammonia is as a household cleaner, which is an aqueous solution of ammonia gas. A typical cleaner has a hydroxide-ion concentration of $4.0 \times 10^{-3}M$. Calculate the pOH and pH of a cleaner at 298 K.

1 Analyze the Problem

You have been given the concentration of hydroxide ion and must calculate pOH and pH. First, calculate pOH using its definition. Then, calculate pH using the relationship $\text{pH} + \text{pOH} = 14.00$.

Known

$$[\text{OH}^-] = 4.0 \times 10^{-3}M$$

Unknown

$$\text{pOH} = ?$$

$$\text{pH} = ?$$

2 Solve for the Unknown

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

State the equation for pOH.

$$\text{pOH} = -\log (4.0 \times 10^{-3})$$

Substitute $[\text{OH}^-] = 4.0 \times 10^{-3}M$.

The **pOH** of the solution is **2.40**.

Use the relationship between pH and pOH to find the pH.

$$\text{pH} + \text{pOH} = 14.00$$

State the equation that relates pH and pOH.

$$\text{pH} = 14.00 - \text{pOH}$$

Solve for pH.

$$\text{pH} = 14.00 - 2.40 = \mathbf{11.60}$$

Substitute $\text{pOH} = 2.40$.

The **pH** of the solution is **11.60**.

3 Evaluate the Answer

The values of pH and pOH are correctly expressed with two decimal places because the given concentration has two significant figures. Because ammonia is a base, a small pOH value and a large pH value are reasonable.



■ **Figure 18.16** Farmers are able to increase the nutritional value of low-quality vegetable materials such as straw, hay, and other crop residue by immersing the materials in an atmosphere of ammonia gas for three weeks.

PRACTICE Problems

Extra Practice Page 989 and glencoe.com

- Calculate the pH and pOH of aqueous solutions with the following concentrations at 298 K.
 - $[\text{OH}^-] = 1.0 \times 10^{-6}M$
 - $[\text{OH}^-] = 6.5 \times 10^{-4}M$
 - $[\text{H}^+] = 3.6 \times 10^{-9}M$
 - $[\text{H}^+] = 2.5 \times 10^{-2}M$
- Calculate the pH and pOH of aqueous solutions with the following concentration at 298 K.
 - $[\text{OH}^-] = 0.000033M$
 - $[\text{H}^+] = 0.0095M$
- Challenge** Calculate pH and pOH for an aqueous solution containing 1.0×10^{-3} mol of HCl dissolved in 5.0 L of solution.

Calculating ion concentrations from pH Sometimes, you have to calculate the concentration of H^+ ions and OH^- ions from the pH of a solution. Example Problem 18.4 shows how to do this.

EXAMPLE Problem 18.4

Calculate $[\text{H}^+]$ and $[\text{OH}^-]$ from pH What are $[\text{H}^+]$ and $[\text{OH}^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature of the blood is 298 K.

1 Analyze the Problem

You have been given the pH of a solution and must calculate $[\text{H}^+]$ and $[\text{OH}^-]$. You can obtain $[\text{H}^+]$ using the equation that defines pH. Then, subtract the pH from 14.00 to obtain the pOH and use the equation that defines pOH to get $[\text{OH}^-]$.

Known

$$\text{pH} = 7.40$$

Unknown

$$[\text{H}^+] = ? \text{ mol/L}$$

$$[\text{OH}^-] = ? \text{ mol/L}$$

2 Solve for the Unknown

Determine $[\text{H}^+]$.

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

State the equation for pH.

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

Multiply both sides of the equation by -1 .

$$[\text{H}^+] = \text{antilog} (-\text{pH})$$

Take the antilog of each side to solve for $[\text{H}^+]$.

$$[\text{H}^+] = \text{antilog} (-7.40)$$

Substitute $\text{pH} = 7.40$.

$$[\text{H}^+] = 4.0 \times 10^{-8} M$$

A calculator shows that the antilog of -7.40 is 4.0×10^{-8} .

The concentration of H^+ ions in the blood is $4.0 \times 10^{-8} M$.

Determine $[\text{OH}^-]$.

$$\text{pH} + \text{pOH} = 14.00$$

State the equation that relates pH and pOH.

$$\text{pOH} = 14.00 - \text{pH}$$

Solve for pOH.

$$\text{pOH} = 14.00 - 7.40 = 6.60$$

Substitute $\text{pH} = 7.40$.

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

State the equation for pOH.

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

Multiply both sides of the equation by -1 .

$$[\text{OH}^-] = \text{antilog} (-6.60)$$

Take the antilog of each side and substitute $\text{pOH} = 6.60$.

$$[\text{OH}^-] = 2.5 \times 10^{-7} M$$

A calculator shows that the antilog of -6.60 is 2.5×10^{-7} .

The concentration of OH^- ions in the blood is $2.5 \times 10^{-7} M$.

3 Evaluate the Answer

The given pH has two decimal places, so the answers must have two significant figures. A $[\text{H}^+]$ less than 10^{-7} and a $[\text{OH}^-]$ greater than 10^{-7} are reasonable, given the initial pH.

PRACTICE Problems

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30. Calculate $[\text{H}^+]$ and $[\text{OH}^-]$ in each of the following solutions.

a. Milk, $\text{pH} = 6.50$.

c. Milk of magnesia, $\text{pH} = 10.50$

b. Lemon juice, $\text{pH} = 2.37$

d. Household ammonia, $\text{pH} = 11.90$

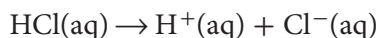
31. Challenge Calculate the $[\text{H}^+]$ and $[\text{OH}^-]$ in a sample of seawater with a $\text{pOH} = 5.60$.



■ **Figure 18.17** The label on a bottle of a strong acid or a strong base tells you the concentration of hydrogen ions or hydroxide ions in the solution. That is because, in solution, strong acids and bases exist entirely as ions.

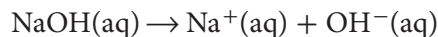
State the $[H^+]$ in the HCl flask and the $[OH^-]$ in the NaOH flask.

Molarity and the pH of strong acids Look at the flasks of acid and base solutions in **Figure 18.17**. The solutions have just been made up and are labeled with their molarity, which is the number of moles of molecules or formula units that were dissolved in 1 L of solution. One flask contains a strong acid (HCl), the other a strong base (NaOH). Recall that strong acids and bases are essentially 100% in the form of ions in solution. That means that the following reaction for the ionization of HCl goes to completion.



Every HCl molecule produces one H^+ ion. The bottle labeled *0.1M HCl* contains 0.1 mol of H^+ ions per liter and 0.1 mol of Cl^- ions per liter. For all strong monoprotic acids, the concentration of the acid is the concentration of H^+ ions. Thus, you can use the molarity of the acid to calculate pH.


Molarity and the pH of strong bases Similarly, the 0.1M solution of the strong base NaOH in **Figure 18.17** is fully ionized.



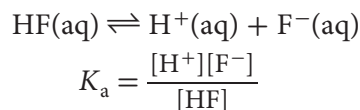
One formula unit of NaOH produces one OH^- ion. Thus, the concentration of the OH^- ions is the same as the molarity of the solution, 0.1M.

Some strong bases, such as calcium hydroxide $Ca(OH)_2$, contain two or more OH^- ions in each formula unit. The concentration of OH^- ion in a solution of $Ca(OH)_2$ is twice the molarity of the ionic compound. For example, the concentration of hydroxide ions in a $7.5 \times 10^{-4}M$ solution of $Ca(OH)_2$ is $7.5 \times 10^{-4}M \times 2 = 1.5 \times 10^{-3}M$.

Although strong acids and strong bases are completely ionized in dilute aqueous solutions, remember that weak acids and weak bases are only partially ionized. Therefore, you must use K_a and K_b values to determine the concentrations of H^+ and OH^- ions in solutions of weak acids and bases.

 **Reading Check** Explain why you cannot obtain the $[H^+]$ directly from the molarity of a weak acid solution.

Calculating K_a from pH Suppose you measured the pH of a 0.100M solution of the weak acid HF and found it to be 3.20. Would you have enough information to calculate K_a for HF?



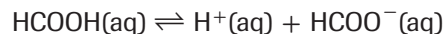
From the pH, you could calculate $[H^+]$. Then, remember that for every mole per liter of H^+ ion there must be an equal concentration of F^- ion. That means you know two of the variables in the K_a expression. What about the third, $[HF]$? The concentration of HF at equilibrium is equal to the initial concentration of the acid (0.100M) minus the moles per liter of HF that dissociated, which is equal to ($[H^+]$).

EXAMPLE Problem 18.5

Calculate K_a from pH Formic acid is used to process latex tapped from rubber trees into natural rubber. The pH of a 0.100M solution of formic acid (HCOOH) is 2.38. What is K_a for HCOOH?

1 Analyze the Problem

You are given the pH of the formic acid solution, which allows you to calculate the concentration of the hydrogen ion.



The balanced chemical equation shows that the concentration of HCOO^- equals the concentration of H^+ . The concentration of un-ionized HCOOH is the difference between the initial concentration of the acid and $[\text{H}^+]$.

Known

$$\text{pH} = 2.38$$

$$\text{concentration of the solution} = 0.100M$$

Unknown

$$K_a = ?$$

2 Solve for the Unknown

Use the pH to calculate $[\text{H}^+]$.

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

$$[\text{H}^+] = \text{antilog} (-\text{pH})$$

$$[\text{H}^+] = \text{antilog} (-2.38)$$

$$[\text{H}^+] = 4.2 \times 10^{-3}M$$

$$[\text{HCOO}^-] = [\text{H}^+] = 4.2 \times 10^{-3}M$$

$[\text{HCOOH}]$ equals the initial concentration minus $[\text{H}^+]$.

$$[\text{HCOOH}] = 0.100M - 4.2 \times 10^{-3}M = 0.096M$$

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

$$K_a = \frac{(4.2 \times 10^{-3})(4.2 \times 10^{-3})}{(0.096)} = 1.8 \times 10^{-4}$$

The acid ionization constant for HCOOH is 1.8×10^{-4} .

Write the equation for pH.

Multiply both sides by -1 and take the antilog of each side.

Substitute $\text{pH} = 2.38$.

A calculator shows that the antilog of -2.38 is 4.2×10^{-3} .

Subtract $[\text{H}^+]$ from the initial $[\text{HCOOH}]$.

State the acid ionization constant expression.

Substitute $[\text{H}^+] = 4.2 \times 10^{-3}M$, $[\text{HCOO}^-] = 4.2 \times 10^{-3}M$, and $[\text{HCOOH}] = 0.096M$.

3 Evaluate the Answer

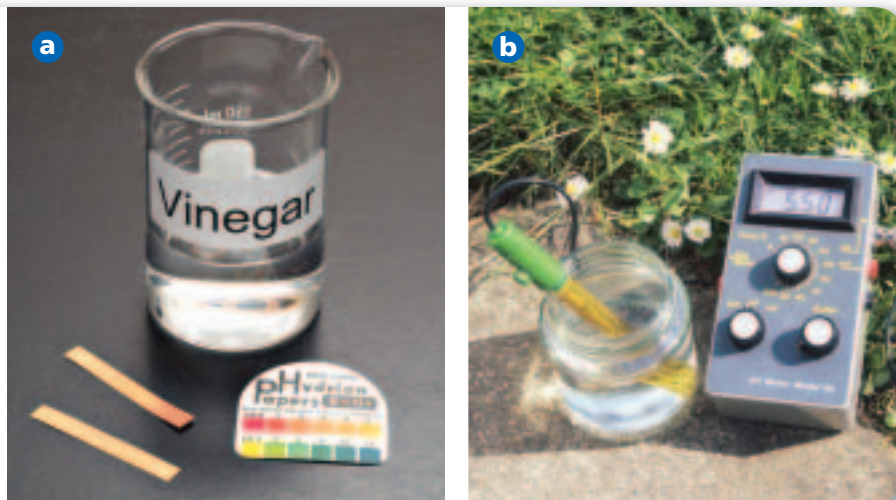
The K_a is reasonable for a weak acid. The answer is correctly reported with two significant figures.

PRACTICE Problems

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32. Calculate the K_a for the following acids using the given information.
- a. 0.220M solution of H_3AsO_4 , pH = 1.50 b. 0.0400M solution of HClO_2 , pH = 1.80
33. Calculate the K_a of the following acids using the given information.
- a. 0.00330M solution of benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$), pOH = 10.70
- b. 0.100M solution of cyanic acid (HCNO), pOH = 11.00
- c. 0.150M solution of butanoic acid ($\text{C}_3\text{H}_7\text{COOH}$), pOH = 11.18
34. **Challenge** Calculate the K_a of a 0.0091M solution of an unknown acid (HX) having a pOH of 11.32. Use **Table 18.4** to identify to acide.

■ **Figure 18.18** The approximate pH of a solution can be obtained by wetting a piece of pH paper with the solution and comparing the color of the wet paper with a set of standard colors as shown in **a**. The portable pH meter in **b**, which is being used to measure the pH of rain water, provides a more accurate measurement in the form of a digital display of the pH.



Measuring pH Perhaps in an earlier science course you used indicator paper to measure the pH of a solution. The litmus paper you used in the Launch Lab is an example of a kind of pH paper. All pH paper is treated with one or more substances called indicators that change color depending on the concentration of hydrogen ions in a solution. Phenolphthalein, which you also used in the Launch Lab, is an example of an indicator. When a strip of pH paper is dipped into an acidic or a basic solution, the color of the paper changes. To determine the pH, the new color of the paper is compared with standard pH colors on a chart, as shown in **Figure 18.18**. The pH meter in **Figure 18.18** provides a more accurate measure of pH. When electrodes are placed in a solution, the meter gives a direct analog or digital readout of pH.

Section 18.3 Assessment

Section Summary

- ▶ The ion product constant for water, K_w , equals the product of the H^+ ion concentration and the OH^- ion concentration.
- ▶ The pH of a solution is the negative log of the hydrogen ion concentration. The pOH is the negative log of the hydroxide ion concentration. pH plus pOH equals 14.
- ▶ A neutral solution has a pH of 7.0 and a pOH of 7.0 because the concentrations of hydrogen ions and hydroxide ions are equal.

- 35. **MAIN Idea Explain** why the pH of an acidic solution is always a smaller number than the pOH of the same solution.
- 36. **Describe** how you can determine the pH of a solution if you know its pOH.
- 37. **Explain** the significance of K_w in aqueous solutions.
- 38. **Explain**, using Le Châtelier's principle, what happens to the $[H^+]$ of a 0.10M solution of acetic acid when a drop of NaOH solution is added.
- 39. **List** the information needed to calculate the K_a of a weak acid.
- 40. **Calculate** The pH of a tomato is approximately 4.50. What are $[H^+]$ and $[OH^-]$ in a tomato?
- 41. **Determine** the pH of a solution that contains 1.0×10^{-9} mol of OH^- ions per liter.
- 42. **Calculate** the pH of the following solutions.

a. 1.0M HI	c. 1.0M KOH
b. 0.050M HNO_3	d. $2.4 \times 10^{-5}M Mg(OH)_2$
- 43. **Interpret Diagrams** Refer to **Figure 18.15** to answer these questions: What happens to the $[H^+]$, $[OH^-]$, pH, and pOH as a neutral solution becomes more acidic? As a neutral solution become more basic?

Section 18.4

Objectives

- Write chemical equations for neutralization reactions.
- Explain how neutralization reactions are used in acid-base titrations.
- Compare the properties of buffered and unbuffered solutions.

Review Vocabulary

stoichiometry: the study of quantitative relationships between the amounts of reactants used and products formed by a chemical reaction; is based on the law of conservation of mass

New Vocabulary

neutralization reaction
salt
titration
titrant
equivalence point
acid-base indicator
end point
salt hydrolysis
buffer
buffer capacity

Neutralization

MAIN Idea In a neutralization reaction, an acid reacts with a base to produce a salt and water.

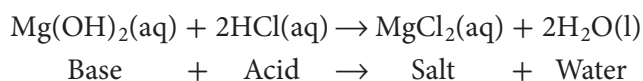
Real-World Reading Link When two teams in a debate present equally convincing arguments, you might find that you are neutral—favoring neither one point of view nor the other. In a similar way, a solution is neutral when the numbers of hydrogen ions and hydroxide ions are equal.

Reactions Between Acids and Bases

If you were to experience heartburn or indigestion, you might take one of the antacids illustrated in **Figure 18.19** to relieve your discomfort. What kind of reaction occurs when magnesium hydroxide ($\text{Mg}(\text{OH})_2$), the active ingredient in milk of magnesia, contacts hydrochloric acid solution (H^+ and Cl^-) produced by the stomach?

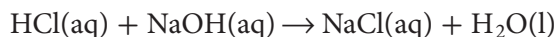
When $\text{Mg}(\text{OH})_2$ and HCl react, a neutralization reaction occurs. A **neutralization reaction** is a reaction in which an acid and a base in an aqueous solution react to produce a salt and water. A **salt** is an ionic compound made up of a cation from a base and an anion from an acid. Neutralization is a double-replacement reaction.

Writing neutralization equations In the reaction between magnesium hydroxide and hydrochloric acid, magnesium replaces hydrogen in HCl and hydrogen replaces magnesium in $\text{Mg}(\text{OH})_2$.



Note that the cation from the base (Mg^{2+}) is combined with the anion from the acid (Cl^-) in the salt MgCl_2 .

When writing neutralization equations, you must know whether all of the reactants and products in the solution exist as molecules or as formula units. For example, examine the formula equation and complete ionic equation for the reaction between hydrochloric acid and sodium hydroxide.



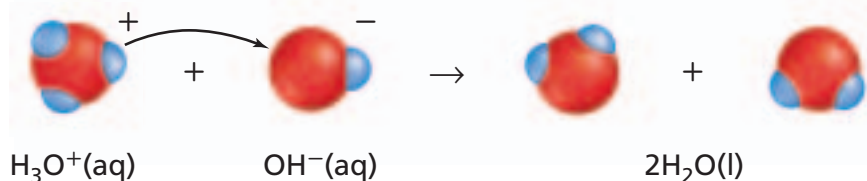
■ **Figure 18.19** A dose of any of these antacids can relieve the symptoms of acid indigestion by reacting with the acidic solution in the stomach and neutralizing it.



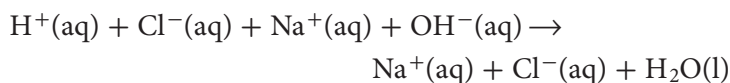
■ **Figure 18.20** A hydronium ion transfers a hydrogen ion to a hydroxide ion. The loss of the hydrogen ion by H_3O^+ results in a water molecule. The gain of a hydrogen ion by OH^- also results in a water molecule.

Concepts in Motion

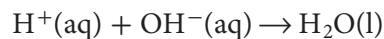
Interactive Figure To see an animation of a neutralization reaction, visit glencoe.com.



Because HCl is a strong acid, NaOH a strong base, and NaCl a soluble salt, all three compounds exist as ions in an aqueous solution.



The chloride ion and the sodium ion appear on both sides of the equation, so they are spectator ions. They can be eliminated to obtain the net ionic equation for the neutralization of a strong acid by a strong base.



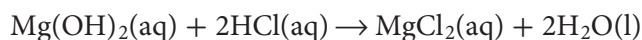
Recall that in an aqueous solution, a H^+ ion exists as a H_3O^+ ion, so the net ionic equation for an acid-base neutralization reaction is



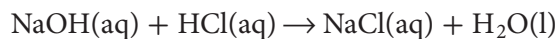
This neutralization reaction is illustrated in **Figure 18.20**.

✓ **Reading Check Demonstrate** that the equation illustrated in **Figure 18.20** represents the neutralization of any strong acid by a strong base by writing the complete ionic equation and the net ionic equation for the neutralization of HNO_3 by KOH .

Acid-base titration The stoichiometry of an acid-base neutralization reaction is the same as that of any other reaction that occurs in solution. In the antacid reaction described above, 1 mol of $\text{Mg}(\text{OH})_2$ neutralizes 2 mol of HCl.



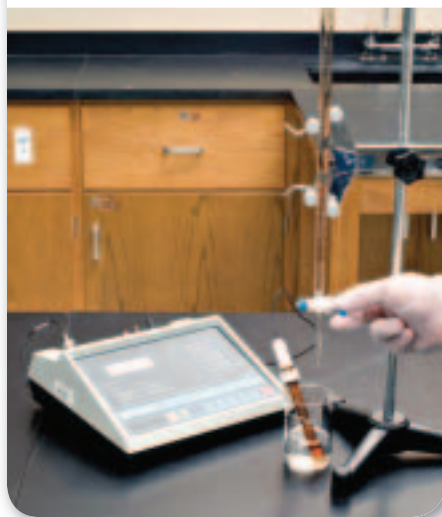
In the reaction of sodium hydroxide and hydrogen chloride, 1 mol of NaOH neutralizes 1 mol of HCl.

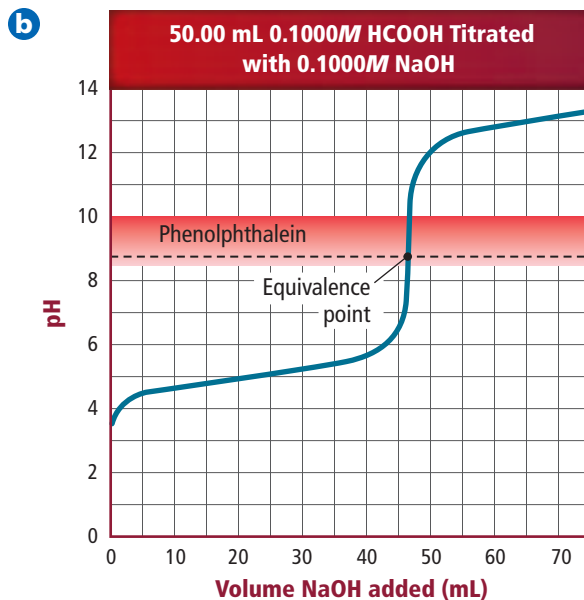
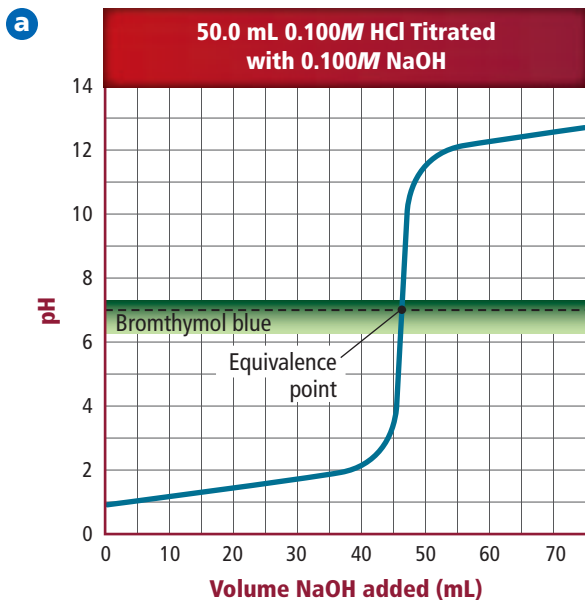


Stoichiometry provides the basis for a procedure called titration, which is used to determine the concentrations of acidic and basic solutions.

Titration is a method for determining the concentration of a solution by reacting a known volume of that solution with a solution of known concentration. If you wish to find the concentration of an acid solution, you would titrate the acid solution with a solution of a base of known concentration. You could also titrate a base of unknown concentration with an acid of known concentration. How is an acid-base titration performed? **Figure 18.21** illustrates one type of setup for the titration procedure outlined on the following page. In this procedure a pH meter is used to monitor the change in the pH as the titration progresses.

■ **Figure 18.21** In the titration of an acid by a base, the pH meter measures the pH of the acid solution in the beaker as a solution of a base with a known concentration is added from the buret.






Titration procedure How is an acid-base titration performed?

1. A measured volume of an acidic or basic solution of unknown concentration is placed in a beaker. The electrodes of a pH meter are immersed in this solution, and the initial pH of the solution is read and recorded.
2. A buret is filled with the titrating solution of known concentration. This is called the standard solution, or **titrant**.
3. Measured volumes of the standard solution are added slowly and mixed into the solution in the beaker. The pH is read and recorded after each addition. This process continues until the reaction reaches the **equivalence point**, which is the point at which moles of H^+ ion from the acid equal moles of OH^- ion from the base.

Figure 18.22a shows how the pH of the solution changes during the titration of 50.0 mL of 0.100M HCl, a strong acid, with 0.100M NaOH, a strong base. The initial pH of the 0.100M HCl is 1.00. As NaOH is added, the acid is neutralized and the solution's pH increases gradually. However, when nearly all of the H^+ ions from the acid have been used up, the pH increases dramatically with the addition of an exceedingly small volume of NaOH. This abrupt increase in pH occurs at the equivalence point of the titration. Beyond the equivalence point, the addition of more NaOH again results in a gradual increase in pH.

You might think that all titrations must have an equivalence point at pH 7 because that is the point at which concentrations of hydrogen ions and hydroxide ions are equal and the solution is neutral. This is not the case, however. Some titrations have equivalence points at pH values less than 7, and some have equivalence points at pH values greater than 7. These differences occur because of reactions between the newly formed salts and water, as you will read later. **Figure 18.22b** shows that the equivalence point for the titration of methanoic acid (a weak acid) with sodium hydroxide (a strong base) lies between pH 8 and pH 9.

 **Graph Check** Identify two ways in which the graphs in **Figure 18.22** are different.

Concepts in Motion

Interactive Figure To see an animation of neutralization reactions, visit glencoe.com.

■ **Figure 18.22** In the titration of a strong acid by a strong base shown in **a**, a steep rise in the pH of the acid solution indicates that all of the H^+ ions from the acid have been neutralized by the OH^- ions of the base. The point at which the curve flexes (at its intersection with the dashed line) is the equivalence point of the titration. Bromthymol blue is an indicator that changes color at this equivalence point. In **b**, a weak acid (HCOOH) is titrated with a strong base (NaOH). The equivalence point is not at a pH of 7. Phenolphthalein is an indicator that changes color at this equivalence point.

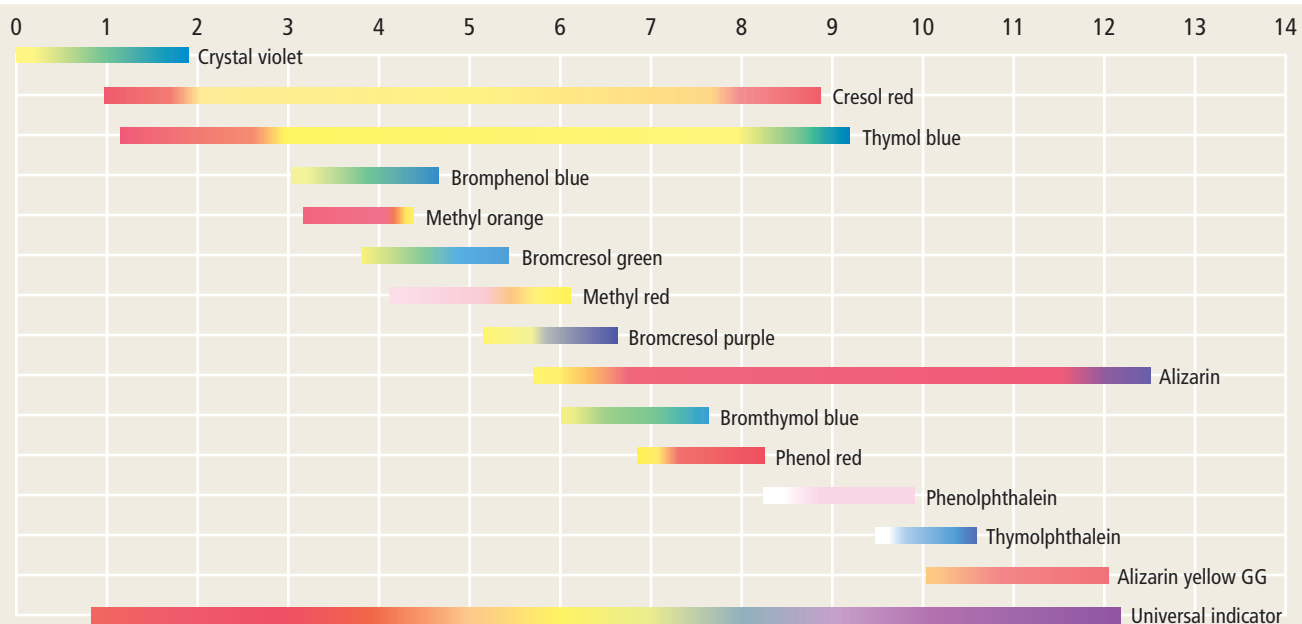
Compare the equivalence points in the two illustrations.

■ **Figure 18.23** The familiar dark color of tea becomes lighter when lemon juice is added. A substance contained in tea is an indicator. Most indicators are large molecules that act as weak acids. Slight differences in bonding patterns when an indicator molecule is ionized or un-ionized account for the color changes.



Acid-base indicators Chemists often use a chemical dye rather than a pH meter to detect the equivalence point of an acid-base titration. Chemical dyes whose colors are affected by acidic and basic solutions are called **acid-base indicators**. Many natural substances act as indicators. If you use lemon juice in your tea, you might have noticed that the brown color of tea gets lighter when lemon juice is added, as shown in **Figure 18.23**. Tea contains compounds called polyphenols that have slightly ionizable hydrogen atoms and therefore are weak acids. Adding acid in the form of lemon juice to a cup of tea depresses the ionization according to Le Châtelier's principle, and the color of the un-ionized polyphenols becomes more apparent. Many of the indicators used by chemists are shown in **Figure 18.24**. As shown in **Figure 18.22**, bromthymol blue is a good choice for a titration of a strong acid with a strong base, and that phenolphthalein changes color at the equivalence point of a titration of a weak acid with a strong base.

■ **Figure 18.24** Choosing the right indicator is important. The indicator must change color at the equivalence point of the titration which is not always at pH 7.

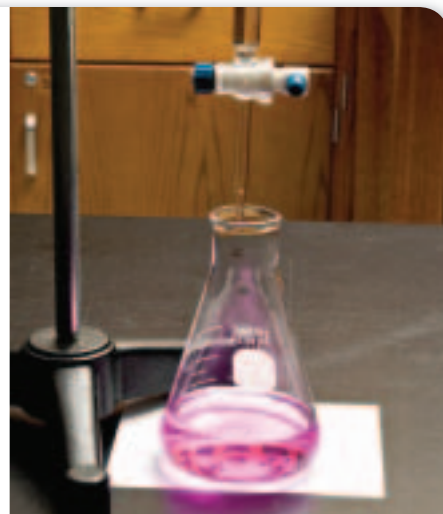




The buret contains the standard solution (0.1000M NaOH), and the flask contains 25.00 mL HCOOH solution along with a small amount of phenolphthalein indicator.



The standard solution is added slowly to the acid solution. The phenolphthalein indicator turns pink, but the color disappears upon mixing, until the end point is reached.



The end point of the titration is marked by a permanent, but very light, pink color. A careful reading of the buret reveals that 18.28 mL 0.1000M NaOH has been added.

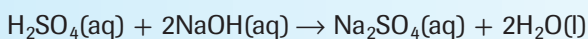
Indicators and titration end point Many indicators used for titration are weak acids. Each has its own particular pH or pH ranges over which it changes color. The point at which the indicator used in a titration changes color is called the **end point** of the titration. It is important to choose an indicator for a titration that will change color at the equivalence point of the titration. Remember that the role of the indicator is to indicate to you, by means of a color change, that just enough of the titrating solution has been added to neutralize the unknown solution. **Figure 18.25** describes the titration of an unknown solution of methanoic acid (HCOOH) with 0.1000M NaOH.

■ **Figure 18.25** Titration is a precise procedure requiring practice. The white paper under the flask provides a background for viewing the indicator color change.

Problem-Solving Strategy

Calculating Molarity

The balanced equation for a titration reaction is the key to calculating the unknown molarity. For example, sulfuric acid is titrated with sodium hydroxide according to this equation.



1. Calculate the moles of NaOH in the standard from the titration data: molarity of the base (M_B) and the volume of the base (V_B).
 $(M_B)(V_B) = (\text{mol/L})(\text{L}) = \text{mol NaOH in standard}$
2. From the equation, you know that the mole ratio of NaOH to H_2SO_4 is 2:1. Two moles of NaOH are required to neutralize 1 mol of H_2SO_4 .
 $\text{mol H}_2\text{SO}_4 \text{ titrated} = \text{mol NaOH in standard} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}$
3. M_A represents the molarity of the acid and V_A represents the volume of the acid in liters. $M_A = \frac{\text{mol H}_2\text{SO}_4 \text{ titrated}}{V_A}$

Apply this strategy as you study Example Problem 18.6.

EXAMPLE Problem 18.6

Math Handbook

Solving Algebraic
Equations
pages 954–955

Molarity from Titration Data A volume of 18.28 mL of a standard solution of 0.1000M NaOH was required to neutralize 25.00 mL of a solution of methanoic acid (HCOOH). What is the molarity of the methanoic acid solution?

1 Analyze the Problem

You are given the molarity and volume of the NaOH solution and the volume of the methanoic acid (HCOOH) solution. The volume of base used is about four-fifths of the volume of the acid, so the molarity of the acid solution should be less than 0.1M.

Known

$$V_A = 25.00 \text{ mL HCOOH}$$

$$V_B = 18.28 \text{ mL NaOH}$$

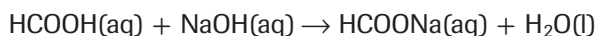
$$M_B = 0.1000M$$

Unknown

$$M_A = ? \text{ mol/L}$$

2 Solve for the Unknown

Write the balanced formula equation for the neutralization reaction.



1 mol NaOH neutralizes 1 mol HCOOH.

$$V_B = 18.28 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.01828 \text{ L}$$

Calculate moles of NaOH.

$$\text{Mol NaOH} = (M_B)(V_B)$$

$$\begin{aligned} \text{Mol NaOH} &= (0.1000 \text{ mol/L})(0.01828 \text{ L}) \\ &= 1.828 \times 10^{-3} \text{ mol NaOH} \end{aligned}$$

Calculate moles of HCOOH.

$$\begin{aligned} 1.828 \times 10^{-3} \text{ mol NaOH} &\times \frac{1 \text{ mol HCOOH}}{1 \text{ mol NaOH}} \\ &= 1.828 \times 10^{-3} \text{ mol HCOOH} \end{aligned}$$

Calculate the molarity of HCOOH.

$$1.828 \times 10^{-3} \text{ mol HCOOH} = (M_A)(V_A)$$

$$M_A = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{V_A}$$

$$V_A = 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.02500 \text{ L HCOOH}$$

$$M_A = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{0.02500 \text{ L HCOOH}} = 7.312 \times 10^{-2} \text{ mol/L}$$

Write the acid to base mole relationship.

Convert volume of base from mL to L.

Apply the relationship between moles of base, molarity of base, and volume of base.

Substitute $M_B = 0.1000M$ and $V_B = 0.01828 \text{ L}$.

Apply the stoichiometric relationship.

Apply the relationship between moles of acids, molarity of acid, and volume of acid.

Solve for M_A .

Convert volume of acid from mL to L.

Substitute $V_A = 0.02500 \text{ L}$.

3 Evaluate the Answer

The answer agrees with the prediction that the molarity of HCOOH is less than 0.1M, and is correctly recorded with four significant figures and the appropriate unit.

PRACTICE Problems

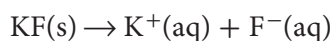
Extra Practice Pages 989–990 and glencoe.com

44. What is the molarity of a nitric acid solution if 43.33 mL of 0.1000M KOH solution is needed to neutralize 20.00 mL of the acid solution?
45. What is the concentration of a household ammonia cleaning solution if 49.90 mL of 0.5900M HCl is required to neutralize 25.00 mL of the solution?
46. **Challenge** How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H₃PO₄?

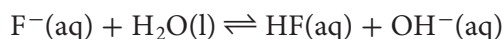
Salt Hydrolysis

In **Figure 18.26**, several drops of bromthymol blue indicator solution have been added to 0.10M aqueous solutions of the salts ammonium chloride (NH_4Cl), sodium nitrate (NaNO_3), and potassium fluoride (KF). Sodium nitrate turns the indicator green, which means that the solution is neutral. The blue color of the KF solution means that the solution is basic, and the yellow color of the ammonium chloride solution indicates that the solution is acidic. Why are some aqueous salt solutions neutral, some basic, and some acidic? Many salts react with water in a process known as salt hydrolysis. In **salt hydrolysis**, the anions of the dissociated salt accept hydrogen ions from water or the cations of the dissociated salt donate hydrogen ions to water.

Salts that produce basic solutions Potassium fluoride is the salt of a strong base (KOH) and a weak acid (HF). It dissociates into potassium ions and fluoride ions.



The K^+ ions do not react with water, but the F^- ion is a weak Brønsted-Lowry base. Some fluoride ions establish this equilibrium with water.

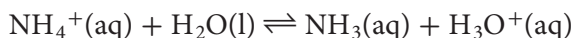


Hydrogen fluoride molecules and OH^- ions are produced. The production of the OH^- ions makes the solution basic.

Salts that produce acidic solutions NH_4Cl is the salt of a weak base (NH_3) and a strong acid (HCl). When dissolved in water, the salt dissociates into ammonium ions and chloride ions.



The Cl^- ions do not react with water, but the NH_4^+ ion is a weak Brønsted-Lowry acid. Ammonium ions react with water molecules to establish this equilibrium.



Ammonia molecules and hydronium ions are produced. The presence of hydronium ions makes the solution acidic.

Salts that produce neutral solutions Sodium nitrate (NaNO_3) is the salt of a strong acid (HNO_3) and a strong base (NaOH). Little or no salt hydrolysis occurs because neither Na^+ nor NO_3^- react with water. Therefore, a solution of sodium nitrate is neutral.



■ **Figure 18.26** The indicator bromthymol blue provides surprising results when added to three solutions of ionic salts. An NH_4Cl solution is acidic, a NaNO_3 solution is neutral, and a KF solution is basic. The explanation has to do with the strengths of the acid and base from which each salt was formed.

PRACTICE Problems

Extra Practice Pages 989–990 and glencoe.com

- 47.** Write equations for the salt hydrolysis reactions occurring when the following salts dissolve in water. Classify each as acidic, basic, or neutral.
- a. ammonium nitrate c. rubidium acetate
 - b. potassium sulfate d. calcium carbonate
- 48. Challenge** Write the equation for the reaction that occurs in a titration of ammonium hydroxide (NH_4OH) with hydrogen bromide (HBr). Will the pH at the equivalence point be greater or less than 7?



■ **Figure 18.27** To provide a healthy environment for these jellies, the pH of the aquarium water at the Monterey Bay Aquarium must be adjusted to stay within the range of 8.1 to 8.4.

Predict what would happen if the pH were allowed to fall to 7.0.

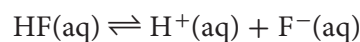
Buffered Solutions

It is important for the jellies shown in **Figure 18.27** that the aquarium water be kept within a narrow pH range. A constant pH is also important in your body. The pH of your blood must be maintained within the range of 7.1 to 7.7. The gastric juices in your stomach must have a pH between 1.6 and 1.8 to promote digestion of certain foods. Your body maintains pH values within such narrow limits by producing buffers.

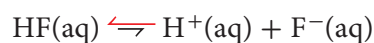
What is a buffer? **Buffers** are solutions that resist changes in pH when limited amounts of acid or base are added. For example, adding 0.01 mol of HCl to 1 L of pure water lowers the pH by 5.0 units, from 7.0 to 2.0. Similarly, adding 0.01 mol of NaOH to 1 L of pure water increases the pH from 7.0 to 12.0. However, if you add the same amount of either HCl or NaOH to 1 L of a buffered solution, the pH might change by no more than 0.1 unit.

How do buffers work? A buffer is a mixture of a weak acid and its conjugate base or a weak base and its conjugate acid. The mixture of ions and molecules in a buffer solution resists changes in pH by reacting with any hydrogen ions or hydroxide ions added to the buffered solution.

Suppose that a buffer solution contains 0.1M concentrations of hydrofluoric acid (HF) and sodium fluoride (NaF). The NaF provides a 0.1M concentration of F^- ions. HF is the acid, and F^- is its conjugate base. The following equilibrium would be established.

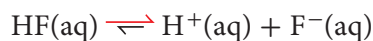


Adding an acid When an acid is added to this buffered solution, as shown in **Figure 18.27b**, the equilibrium shifts to the left. According to Le Châtelier's principle, the added H^+ ions from the acid are a stress on the equilibrium, which is relieved by their reaction with F^- ions to form additional undissociated HF molecules.



Equilibrium is established again with a larger amount of undissociated HF present. However, the pH of the solution has changed little because the shift to the left consumed most of the added H^+ ion.

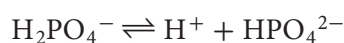
Adding a base When a base is added to the hydrofluoric acid/fluoride ion buffer system, the added OH^- ions react with H^+ ions to form H_2O . This decreases the concentration of H^+ ions, and the equilibrium shifts to the right to replace the H^+ ions.



Although the shift to the right consumes HF molecules and produces additional F^- ions, the pH remains fairly constant because the H^+ ion concentration has not changed appreciably.

A buffer solution's capacity to resist pH change can be exceeded by the addition of too much acid or base. The amount of acid or base a buffer solution can absorb without a significant change in pH is called the **buffer capacity** of the solution. The greater the concentrations of the buffering molecules and ions in the solution, the greater the solution's buffer capacity.

Choosing a buffer A buffer system is most effective when the concentrations of the conjugate acid-base pair are equal or nearly equal. Consider the $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ buffer system made by mixing equal molar amounts of NaH_2PO_4 and NaH_2PO_4 .



What is the pH of such a buffer solution? The acid ionization constant expression for the equilibrium can provide the answer.

$$K_a = 6.2 \times 10^{-8} = \frac{[\text{H}^+][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

Because the solution has been made with equal molar amounts of Na_2HPO_4 and NaH_2PO_4 , $[\text{HPO}_4^{2-}]$ is equal to $[\text{H}_2\text{PO}_4^-]$. Thus, the two terms in the acid ionization expression cancel.

$$6.2 \times 10^{-8} = \frac{[\text{H}^+]\cancel{[\text{HPO}_4^{2-}]}}{\cancel{[\text{H}_2\text{PO}_4^-]}} = [\text{H}^+]$$

$$\text{pH} = -\log [\text{H}^+] = -\log (6.2 \times 10^{-8}) = 7.21$$

Thus, when equimolar amounts of each of the components are present in the $\text{HPO}_4^-/\text{H}_2\text{PO}_4^{2-}$ buffer system, the system can maintain a pH close to 7.21. Note that the pH is the negative log of K_a . **Table 18.7** lists several buffer systems, with the pH at which each is effective.

VOCABULARY

SCIENCE USAGE V. COMMON USAGE

Buffer

Science usage: a solution that resists changes in pH when limited amounts of acid or base are added

The chemist decided to use a buffer consisting of equal molar amounts of formic acid and sodium formate.

Common usage: something that serves as a protective barrier

For the homes along the shore, the high seawall served as a buffer against stormy seas.

Table 18.7

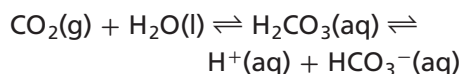
Buffer Systems with Equimolar Components

Buffer Equilibrium	Conjugate Acid-Base Pair in Buffered Solution	Buffer pH
$\text{HF}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{F}^-(\text{aq})$	HF/ F^-	3.20
$\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$	$\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$	4.76
$\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$	$\text{H}_2\text{CO}_3/\text{HCO}_3^-$	6.35
$\text{H}_2\text{PO}_4^-(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HPO}_4^{2-}(\text{aq})$	$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.21
$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{NH}_4^+/\text{NH}_3$	9.4
$\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{C}_2\text{H}_5\text{NH}_3^+/\text{C}_2\text{H}_5\text{NH}_2$	10.70

PROBLEM-SOLVING LAB

Apply Scientific Explanations

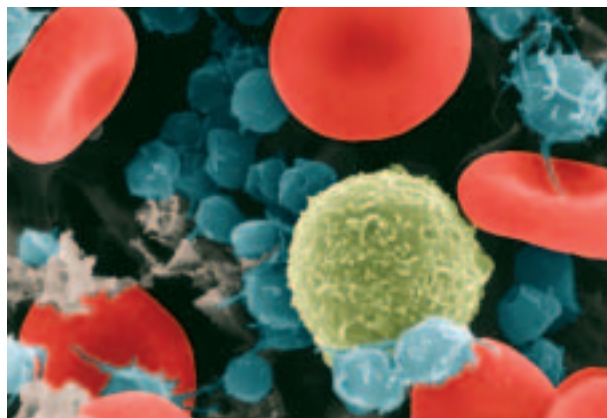
How does your blood maintain its pH? Human blood contains three types of cells. Red blood cells deliver oxygen to every part of the body. White blood cells fight infections, and platelets aid in clotting when bleeding occurs. The critical functions of these cells are impaired if the pH of blood is not maintained within the narrow range of 7.1 to 7.7. Beyond this range, proteins in the body lose their structures and abilities to function. Fortunately, several buffers maintain the necessary acid/base balance. The carbonic acid/hydrogen carbonate ($\text{H}_2\text{CO}_3/\text{HCO}_3^-$) buffer is the most important.



As acids and bases enter the bloodstream as a result of normal activity, the blood's buffer systems shift to effectively maintain a healthful pH.

Analysis

Depending on the body's metabolic rate and other factors, the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium will shift according to Le Châtelier's principle. In addition, the lungs can alter the rate at which CO_2 is expelled from the body by breathing, and the kidneys can alter the rate of removal of HCO_3^- ions.



Think Critically

- Determine** how many times greater the $[\text{H}^+]$ is if the blood's pH changes from pH 7.4 to 7.1.
- Suggest** a reason why a 20:1 ratio of HCO_3^- to CO_2 in the blood is favorable for maintaining a healthy pH.
- Predict** whether, for each situation, the pH of the blood will rise or fall, and which way the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium will shift.
 - A person with a severe stomach virus vomits many times during a 24-h period.
 - To combat heartburn, a person takes too much (NaHCO_3).

Section 18.4 Assessment

Section Summary

- In a neutralization reaction, an acid and a base react to form a salt and water.
- The net ionic equation for the neutralization of a strong acid by a strong base is $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$.
- Titration is the process in which an acid-base neutralization reaction is used to determine the concentration of a solution.
- Buffered solutions contain mixtures of molecules and ions that resist changes in pH.

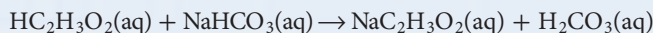
- MAIN Idea** Explain why the net ionic equation for the neutralization reaction of any strong acid with any strong base is always the same.
- Explain** the difference between the equivalence point and the end point of a titration.
- Compare** the results of two experiments: First, a small amount of base is added to an unbuffered solution with a pH of 7. Second, the same amount of base is added to a buffered solution with a pH of 7.
- Calculate** the molarity of a solution of hydrobromic acid (HBr) if 30.35 mL of 0.1000M NaOH is required to titrate 25.00 mL of the acid to the equivalence point.
- Interpret** What substances could be used to make a buffer solution with a pH of 9.4. How should the amounts of the substances be related. Use **Table 18.7**.
- Design an Experiment** Describe how you would design and perform a titration in which you use 0.250M HNO_3 to determine the molarity of a cesium hydroxide solution. Include the formula and net ionic equations.

Everyday Chemistry

Acid-Base Reactions on the Rise

Do you remember how much fun it was to watch a vinegar baking soda volcano erupt? The bubbles of carbon dioxide (CO_2) resulted from a decomposition reaction that quickly followed the acid-base reaction between the vinegar ($\text{HC}_2\text{H}_3\text{O}_2$), an acid, and baking soda (NaHCO_3), a base, as shown below.

Acid-Base Reaction



Decomposition



The release of carbon dioxide as a result of the chemical reaction between an acid and a base, as shown in **Figure 1**, is part of the reason why baked goods rise. An ingredient that causes batter to rise when baked is called a leavening agent. The two main chemical leavening agents are baking soda and baking powder.

Baking soda Sodium hydrogen carbonate, also called sodium bicarbonate, is the chemical name for baking soda. When used in cooking, baking soda reacts with mildly acidic liquids, and carbon dioxide bubbles form. Mildly acidic liquids include vinegar, molasses, honey, citrus juice, buttermilk, and many others.



Figure 1 Carbon dioxide forms bubbles when baking soda, a base, is added to vinegar, an acid.

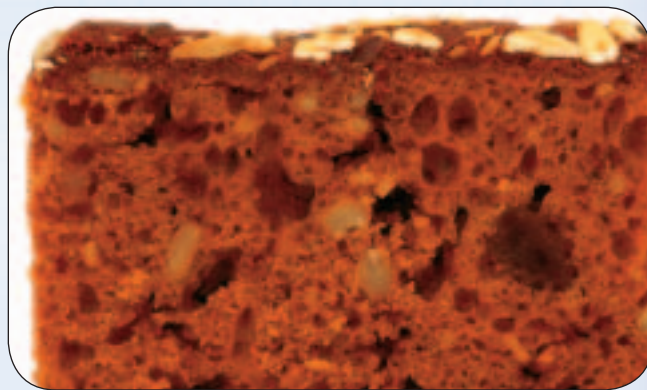


Figure 2 Baking traps the bubbles formed during the reaction between an acid and a base, resulting in a light, airy cake.

Baking soda must be mixed with other dry ingredients and added last to a batter so that the release of carbon dioxide is uniform throughout the batter. This acid-base reaction happens quickly. If baking soda is the only leavening agent in a recipe, the batter must be baked immediately before the bubbles have a chance to escape. Baking causes the bubbles to expand, and the cake rises. As the batter firms, the bubbles are trapped, as shown in **Figure 2**.

Baking powder If a recipe does not include an acidic liquid, baking powder is used. Most baking powder is a mixture of baking soda and two dry acids. One of the acids reacts with the baking soda when it dissolves in the batter, and the other reacts with the baking soda when heated.

Like baking soda, baking powder is mixed with other dry ingredients and added last to a batter. However, batters made with baking powder do not have to be baked immediately.

Sometimes, batters made with mildly acidic liquids include both baking powder and baking soda. Excess acid can disrupt the action of the baking powder. The baking powder provides a reliable source of carbon dioxide, and the baking soda helps to neutralize the acid.

WRITING in Chemistry

Analyze If a recipe calls for flour, salt, sugar, bran cereal, milk, an egg, and shortening or vegetable oil, would you use baking soda or baking powder? Explain. For more information about acids and bases in cooking, visit glencoe.com.

STANDARDIZE A BASE

Background: Titration is a procedure by which the molarity of a base can be determined.

Question: How can you determine the molarity of a solution of a base?

Materials

50-mL buret	spatula
buret clamp	250-mL Erlenmeyer flask
ring stand	500-mL Florence flask
sodium hydroxide pellets (NaOH)	with rubber stopper
potassium hydrogen phthalate (KHC ₈ H ₄ O ₄)	250-mL beaker
distilled water	centigram balance
weighing bottle	wash bottle
	phenolphthalein solution

Safety Precautions

WARNING: Dissolving NaOH in water generates heat. Phenolphthalein is flammable. Keep away from flames.

Procedure

1. Read and complete the lab safety form.
2. Place about 4 g of NaOH in a 500-mL Florence flask. Add enough water to dissolve the pellets and bring the volume of the NaOH solution to about 400 mL. Stopper the flask.
3. Use the weighing bottle to mass by difference about 0.40 g of potassium hydrogen phthalate (KHC₈H₄O₄, molar mass = 204.32 g/mol) into a 250-mL Erlenmeyer flask. Record this mass.
4. Use a wash bottle to rinse the insides of the flask, and add about 50 mL of water. Add two drops of phenolphthalein indicator solution.
5. Rinse the buret with 10 mL of your base solution. Discard the rinse solution in a discard beaker. Attach the buret to the ring stand using the buret clamp.
6. Fill the buret with NaOH solution. The level of the liquid should be at or below the zero mark. To remove any air trapped in the tip of the buret, allow a small amount of the base to flow from the tip into the discard beaker. Read the buret to the nearest 0.02 mL, and record this initial reading.
7. Place a piece of white paper on the base of the ring stand. Swirl the flask while allowing the NaOH solution to flow slowly from the buret into the flask.

Titration Data

	Trial 1
Mass of weighing bottle and acid	
Mass of weighing bottle	
Mass of solid acid	
Moles of acid	
Moles of base required	
Final reading of base buret	
Initial reading of base buret	
Volume of base used in mL	
Molarity of base	

8. When the pink color begins to persist longer as the flask is swirled, add the base drop-by-drop.
9. The end point is reached when one additional drop of base turns the acid pink. The pink color should persist as the flask is swirled. Record the final volume in the buret.
10. Calculate the molarity of your base using Steps 1–4 in the Analyze and Conclude section.
11. Refill the buret. Rinse the flask with water. Repeat the titration until the calculated values of the molarity for three trials show close agreement.
12. **Cleanup and Disposal** Wash the neutralized solutions down the sink with plenty of water.

Analyze and Conclude

1. **Interpret Data** For each titration, calculate the number of moles of acid used by dividing the mass of the sample by the molar mass of the acid.
2. **Infer** How many moles of base are required to react with the moles of acid you used?
3. **Calculate** Convert the volume of base to liters.
4. **Calculate** the molarity of the base by dividing the moles of base by the volume of base in liters.
5. **Error Analysis** Did your calculated molarities agree? Explain any irregularities.

INQUIRY EXTENSION

Design an Experiment Determine the concentration of a vinegar solution without using an indicator.



BIG Idea Acids and bases can be defined in terms of hydrogen ions and hydroxide ions or in terms of electron pairs.

Section 18.1 Introduction to Acids and Bases

MAIN Idea Different models help describe the behavior of acids and bases.

Vocabulary

- acidic solution (p. 636)
- amphoteric (p. 639)
- Arrhenius model (p. 637)
- basic solution (p. 636)
- Brønsted-Lowry model (p. 638)
- conjugate acid (p. 683)
- conjugate acid-base pair (p. 638)
- conjugate base (p. 638)
- Lewis model (p. 641)

Key Concepts

- The concentrations of hydrogen ions and hydroxide ions determine whether an aqueous solution is acidic, basic, or neutral.
- An Arrhenius acid must contain an ionizable hydrogen atom. An Arrhenius base must contain an ionizable hydroxide group.
- A Brønsted-Lowry acid is a hydrogen ion donor. A Brønsted-Lowry base is a hydrogen ion acceptor.
- A Lewis acid accepts an electron pair. A Lewis base donates an electron pair.

Section 18.2 Strengths of Acids and Bases

MAIN Idea In solution, strong acids and bases ionize completely, but weak acids and bases ionize only partially.

Vocabulary

- acid ionization constant (p. 647)
- base ionization constant (p. 649)
- strong acid (p. 644)
- strong base (p. 648)
- weak acid (p. 645)
- weak base (p. 648)

Key Concepts

- Strong acids and strong bases are completely ionized in a dilute aqueous solution. Weak acids and weak bases are partially ionized in a dilute aqueous solution.
- For weak acids and weak bases, the value of the acid or base ionization constant is a measure of the strength of the acid or base.

Section 18.3 Hydrogen Ions and pH

MAIN Idea pH and pOH are logarithmic scales that express the concentrations of hydrogen ions and hydroxide ions in aqueous solutions.

Vocabulary

- ion product constant for water (p. 650)
- pH (p. 652)
- pOH (p. 652)

Key Concepts

- The ion product constant for water, K_w , equals the product of the H^+ ion concentration and the OH^- ion concentration.
$$K_w = [H^+][OH^-]$$
- The pH of a solution is the negative log of the hydrogen ion concentration. The pOH is the negative log of the hydroxide ion concentration. pH plus pOH equals 14.
$$pH = -\log [H^+] \qquad pOH = -\log [OH^-]$$
$$pH + pOH = 14.00$$
- A neutral solution has a pH of 7.0 and a pOH of 7.0 because the concentrations of hydrogen ions and hydroxide ions are equal.

Section 18.4 Neutralization

MAIN Idea In a neutralization reaction, an acid reacts with a base to produce a salt and water.

Vocabulary

- acid-base indicator (p. 662)
- buffer (p. 666)
- buffer capacity (p. 667)
- end point (p. 663)
- equivalence point (p. 661)
- neutralization reaction (p. 659)
- salt (p. 659)
- salt hydrolysis (p. 665)
- titrant (p. 661)
- titration (p. 660)

Key Concepts

- In a neutralization reaction, an acid and a base react to form a salt and water.
- The net ionic equation for the neutralization of a strong acid by a strong base is $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.
- Titration is the process in which an acid-base neutralization reaction is used to determine the concentration of a solution.
- Buffered solutions contain mixtures of molecules and ions that resist changes in pH.

Section 18.3

Mastering Concepts

75. What is the relationship between the pOH and the OH^- ion concentration of a solution?
76. Solution A has a pH of 2.0. Solution B has a pH of 5.0. Which solution is more acidic? Based on the H^+ ion concentrations in the two solutions, how many times more acidic?
77. If the concentration of H^+ ions in an aqueous solution decreases, what must happen to the concentration of OH^- ions? Why?
78. Use Le Châtelier's principle to explain what happens to the equilibrium $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$ when a few drops of HCl are added to pure water.
79. **Common Acids and Bases** Use the data in Table 18.8 to answer the following questions.

Substance	pH
Household ammonia	11.3
Lemon juice	2.3
Antacid	9.4
Blood	7.4
Soft drinks	3.0

- Which substance is the most basic?
- Which substance is closest to neutral?
- Which has a concentration of $\text{H}^+ = 4.0 \times 10^{-10}\text{M}$?
- Which has a pOH of 11.0?
- How many times more basic is antacid than blood?

Mastering Problems

80. What is $[\text{OH}^-]$ in an aqueous solution at 298 K in which $[\text{H}^+] = 5.40 \times 10^{-3}\text{M}$?
81. What are the pH and pOH for the solution described in Question 80?
82. If 5.00 mL of 6.00M HCl is added to 95.00 mL of pure water, the final volume of the solution is 100.00 mL. What is the pH of the solution?
83. Given two solutions, 0.10M HCl and 0.10M HF, which solution has the greater concentration of H^+ ions? Calculate pH values for the two solutions, given that $[\text{H}^+] = 7.9 \times 10^{-3}\text{M}$ in the 0.10M HF.
84. **Metal Cleaner** Chromic acid is used as an industrial cleaner for metals. What is K_a for the second ionization of chromic acid (H_2CrO_4) if a 0.040M solution of sodium hydrogen chromate has a pH of 3.946?

Section 18.4

Mastering Concepts

85. What acid and base must react to produce an aqueous sodium iodide solution?

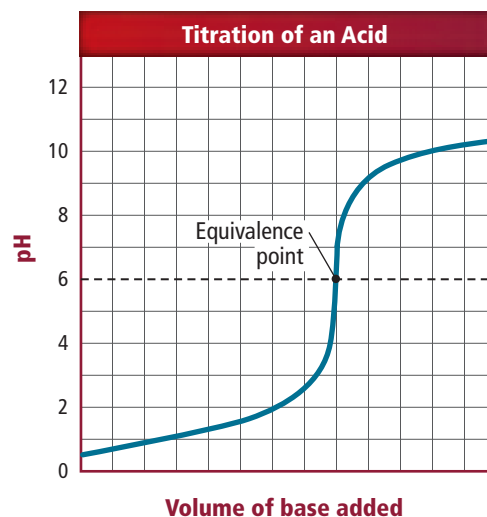


Figure 18.30

86. What acid-base indicators, shown in Figure 18.24, would be suitable for the neutralization reaction whose titration curve is shown in Figure 18.30? Why?
87. When might a pH meter be better than an indicator to determine the end point of an acid-base titration?
88. What happens when an acid is added to a solution containing the HF/F^- buffer system?
89. When methyl red is added to an aqueous solution, a pink color results. When methyl orange is added to the same solution, a yellow color is produced. What is the approximate pH range of the solution? Use Figure 18.24.
90. Give the name and formula of the acid and the base from which each salt was formed.
- NaCl
 - KHCO_3
 - NH_4NO_2
 - CaS

Mastering Problems

91. Write formula equations and net ionic equations for the hydrolysis of each salt in water.
- sodium carbonate
 - ammonium bromide
92. **Air Purifier** Lithium hydroxide is used to purify air by removing carbon dioxide. A 25.00-mL sample of lithium hydroxide solution is titrated to an end point by 15.22 mL of 0.3340M hydrochloric acid solution. What is the molarity of the LiOH solution?
93. In an acid-base titration, 45.78 mL of a sulfuric acid solution is titrated to the end point by 74.30 mL of 0.4388M sodium hydroxide solution. What is the molarity of the H_2SO_4 solution?

Mixed Review

94. Write the equation for the ionization reaction and the base ionization constant expression for ethylamine ($C_2H_5NH_2$) in water.
95. How many milliliters of 0.225M HCl would be required to titrate 6.00 g of KOH?
96. What is the pH of a 0.200M solution of hypobromous acid (HBrO)? $K_a = 2.8 \times 10^{-9}$
97. Which of the following are polyprotic acids? Write successive ionization equations for the polyprotic acids in water.
- | | |
|---------------|---------------|
| a. H_3BO_3 | c. HNO_3 |
| b. CH_3COOH | d. H_2SeO_3 |
98. Write balanced chemical equations for the two successive ionizations of carbonic acid in water. Identify the conjugate-base pair in each of the equations.
99. **Sugar Refining** Strontium hydroxide is used in the refining of beet sugar. Only 4.1 g of strontium hydroxide can be dissolved in 1 L of water at 273 K. Given that its solubility is so low, explain how it is possible that strontium hydroxide is considered a strong base.
100. What are the concentrations of OH^- ions in solutions having pH values of 3.00, 6.00, 9.00, and 12.00 at 298 K? What are the pOH values for the solutions?

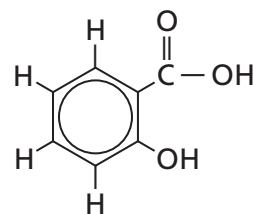


■ Figure 18.31

101. The pH probe in Figure 18.31 is immersed in a 0.200M solution of a monoprotic acid, HA, at 303 K. What is the value of K_a for the acid at 303 K?
102. Write the chemical equation for the reaction that would occur when a base is added to a solution containing the $H_2PO_4^-/HPO_4^{2-}$ buffer system.
103. An aqueous solution buffered by benzoic acid (C_6H_5COOH) and sodium benzoate (C_6H_5COONa) is 0.0500M in both compounds. Given that benzoic acid's K_a equals 6.4×10^{-5} , what is the pH of the solution?

Think Critically

104. **Critique** the following statement: "A substance whose chemical formula contains a hydroxyl group must be considered to be a base."
105. **Analyze and Conclude** Is it possible that an Arrhenius acid is not a Brønsted-Lowry acid? Is it possible that an acid according to the Brønsted-Lowry model is not an Arrhenius acid? Is it possible that a Lewis acid could not be classified as either an Arrhenius or a Brønsted-Lowry acid? Explain and give examples.
106. **Apply Concepts** Use the ion product constant of water at 298 K to explain why a solution with a pH of 3.0 must have a pOH of 11.0.
107. **Identify** the Lewis acids and bases in the following reactions.
- | |
|--|
| a. $H^+ + OH^- \rightleftharpoons H_2O$ |
| b. $Cl^- + BCl_3 \rightleftharpoons BCl_4^-$ |
| c. $SO_3 + H_2O \rightleftharpoons H_2SO_4$ |
108. **Interpret Scientific Illustrations** Sketch the shape of the approximate pH v. volume curve that would result from titrating a diprotic acid with a 0.10M NaOH solution.
109. **Recognize Cause and Effect** Illustrate how a buffer works using the $C_2H_5NH_3^+/C_2H_5NH_2$ buffer system. Show with equations how the weak base/conjugate acid system is affected when small amounts of acid and base are added to a solution containing this buffer system.



■ Figure 18.32

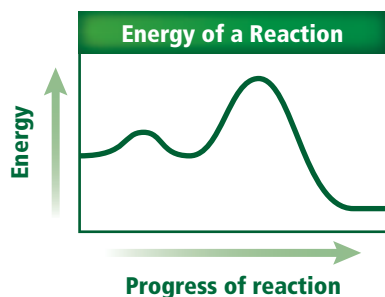
110. **Predict** Salicylic acid, shown in Figure 18.32, is used to manufacture acetylsalicylic acid, commonly known as aspirin. Evaluate the hydrogen atoms in the salicylic acid molecule based on your knowledge about the ionizable hydrogen in the acetic acid molecule, CH_3COOH . Predict which of salicylic acid's hydrogen atoms is likely to be ionizable.
111. **Apply Concepts** Like all equilibrium constants, the value of K_w varies with temperature. K_w equals 2.92×10^{-15} at $10^\circ C$, 1.00×10^{-14} at $25^\circ C$, and 2.92×10^{-14} at $40^\circ C$. In light of this information, calculate and compare the pH values for pure water at these three temperatures. Based on your calculations, is it correct to say that the pH of pure water is always 7.0? Explain.

Challenge Problem

- 112.** You have 20.0 mL of a solution of a weak acid, HX, whose K_a equals 2.14×10^{-6} . The pH of the solution is found to be 3.800. How much distilled water would you have to add to the solution to increase the pH to 4.000?

Cumulative Review

- 113.** What factors determine whether a molecule is polar or nonpolar? (Chapter 8)
- 114.** What property of some liquids accounts for the meniscus that forms at the surface of a solution in a buret? (Chapter 12)
- 115.** Which of the following physical processes are exothermic for water—freezing, boiling, condensing, subliming, evaporating? (Chapter 12)
- 116.** Explain why an air pump gets hot when you pump air into a bicycle tire. (Chapter 13)
- 117.** When 5.00 g of a compound was burned in a calorimeter, the temperature of 2.00 kg of water increased from 24.5°C to 40.5°C. How much heat would be released by the combustion of 1.00 mol of the compound (molar mass = 46.1 g/mol)? (Chapter 15)
- 118.** What is the difference between an exothermic and an endothermic reaction? (Chapter 15)



■ Figure 18.33

- 119.** Figure 18.33 shows how energy changes during the progress of a reaction.
- Is the reaction exothermic or endothermic? (Chapter 15)
 - How many steps are in the reaction mechanism for the reaction? (Chapter 16)
 - Explain how you could use the graph to identify the rate-determining step. (Chapter 16)
- 120.** Hydrogen and fluorine react to form HF according to the following equilibrium equation.
- $$\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightleftharpoons 2\text{HF} \quad \Delta H = -538 \text{ kJ}(\text{g})$$
- Will raising the temperature cause the amount of product to increase? Explain. (Chapter 17)

Additional Assessment

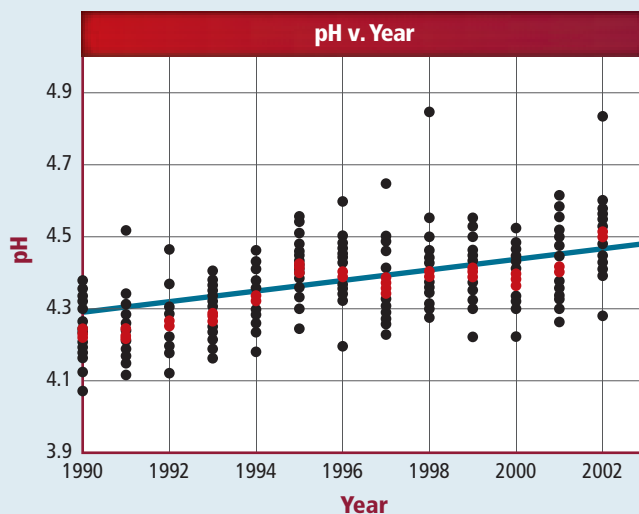
WRITING in Chemistry

- 121. Acid/Base Theories** Imagine that you are the Danish chemist Johannes Brønsted. The year is 1923, and you have formulated a new theory of acids and bases. Write a letter to Swedish chemist Svante Arrhenius in which you discuss the differences between your theory and his and point out the advantages of yours.
- 122. Amino Acids** Twenty amino acids combine to form proteins in living systems. Research the structures and K_a values for five amino acids. Compare the strengths of these acids with the acids in Table 18.4.

DBQ Document-Based Questions

Rainwater Figure 18.34 shows pH measurements made from a number of the monitoring sites in New York state. The pink dot represents the average of the measurement taken at all of the sites at a particular time.

Data obtained from: Acid Deposition in New York 1987–2004. January, 2006. New York State Department of Environmental Conservation.



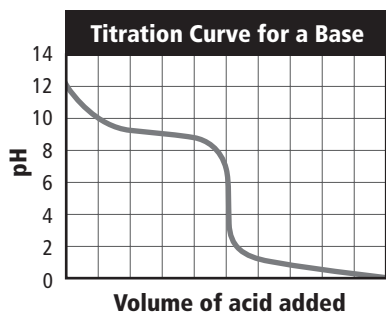
■ Figure 18.34

- 123.** In general, what is the trend in the average pH for the years 1990 to 2003?
- 124.** Calculate the $[\text{H}^+]$ for the lowest and the highest pH measurements recorded on the graph. How many times more acidic is the rainwater having the highest reading than the rainwater with the lowest?
- 125.** What is the pH of the trend line in 2003? How much has the average pH changed between the years 1990 and 2003?

Cumulative Standardized Test Practice

Multiple Choice

Use the graph below to answer Questions 1 and 2.



- What is the pH at the equivalence point of this titration?
 - 10
 - 9
 - 5
 - 1
- Which indicator would be effective for detecting the end point of this titration?
 - methyl orange, with a range of 3.2–4.4
 - phenolphthalein, with a range of 8.2–10
 - bromocresol green, with a range of 3.8–5.4
 - thymol blue, with a range of 8.0–9.6
- Hydrogen bromide (HBr) is a strong, highly corrosive acid. What is the pOH of a 0.0375M HBr solution?
 - 12.574
 - 12.270
 - 1.733
 - 1.433
- Cellular respiration produces about 38 mol of ATP for every mole of glucose consumed:

$$\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} + 38\text{ATP}$$
 If each mole of ATP can release 30.5 kJ of energy, how much energy can be obtained from a candy bar containing 130.0 g of glucose?
 - 27.4 kJ
 - 836 kJ
 - 1159 kJ
 - 3970 kJ

Use the table below to answer Questions 5–7.

Ionization Constants and pH Data for Several Weak Organic Acids		
Acid	pH of 1.000M Solution	K_a
Formic	1.87	1.78×10^{-4}
Cyanoacetic	?	3.55×10^{-3}
Propanoic	2.43	?
Lutidinic	1.09	7.08×10^{-3}
Barbituric	2.01	9.77×10^{-5}

- Which acid is the strongest?
 - formic acid
 - cyanoacetic acid
 - lutidinic acid
 - barbituric acid
- What is the acid dissociation constant of propanoic acid?
 - 1.4×10^{-5}
 - 2.43×10^0
 - 3.72×10^{-3}
 - 7.3×10^4
- What is the pH of a 0.40M solution of cyanoacetic acid?
 - 2.06
 - 1.22
 - 2.45
 - 1.42
- What does a value of K_{eq} greater than 1 mean?
 - More reactants than products exist at equilibrium.
 - More products than reactants exist at equilibrium.
 - The rate of the forward reaction is high at equilibrium.
 - The rate of the reverse reaction is high at equilibrium.
- Magnesium sulfate (MgSO_4) is often added to water-insoluble liquid products of chemical reactions to remove unwanted water. MgSO_4 readily absorbs water to form two different hydrates. One of them is found to contain 13.0% H_2O and 87.0% MgSO_4 . What is the name of this hydrate?
 - magnesium sulfate monohydrate
 - magnesium sulfate dihydrate
 - magnesium sulfate hexahydrate
 - magnesium sulfate heptahydrate

Short Answer

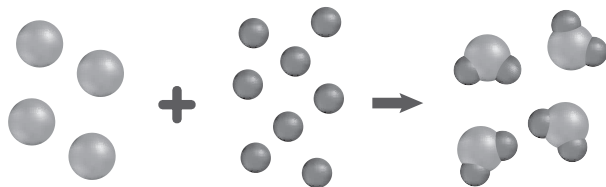
Use the description of an experiment below to answer Questions 10–12.

Two 0.050-mol samples of gas at 20°C are released from the end of a long tube at the same time. One gas is xenon (Xe), and the other is sulfur dioxide (SO₂).

- Explain which gas will have traveled farther after 5 seconds. How can you tell?
- How will increasing the temperature of this experiment affect the rate of effusion of each gas?
- If the pressure on the xenon at the end of the experiment is 0.092 atm, what volume will it occupy?

Extended Response

Use the figure below to answer Question 13.



Atoms of Element A

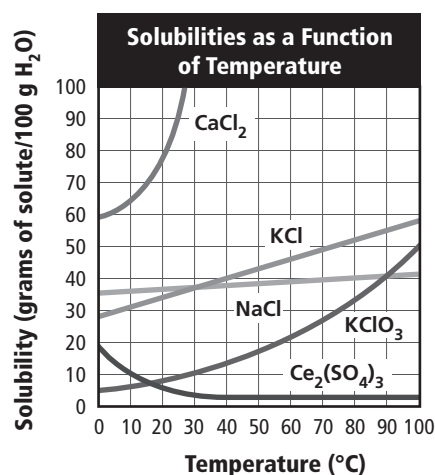
Atoms of Element B

- Explain how the chemical reaction shown in this figure demonstrates the law of conservation of mass.
- Describe lab procedures for preparing a 0.50M aqueous solution of NaOH and a 0.50 m aqueous solution of NaOH.
- Explain how you could express the concentration of the 0.50 m solution in Question 14 as a mole fraction.

SAT Subject Test: Chemistry

- Water has an unusually high boiling point compared to other compounds of similar molar mass because of
 - hydrogen bonding.
 - adhesive forces.
 - covalent bonding.
 - dispersion forces.
 - pi bonds.

Use the graph below to answer Questions 16 and 17.



- Which compound has a solubility of 38 g/100 g H₂O at 50 °C?
 - CaCl₂
 - KCl
 - NaCl
 - KClO₃
 - Ce₂(SO₄)₃
- Which has the greatest increase in solubility as temperature increases?
 - Ce₂(SO₄)₃
 - CaCl₂
 - KClO₃
 - NaCl
 - Ce₂(SO₄)₃

NEED EXTRA HELP?

If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Review Section . . .	18.4	18.4	18.3	11.2	18.2	18.2	18.2	17.3	10.5	12.1	12.1	13.2	11.1	14.2	14.2	12.2	14.3	14.3