8.7: Acid–Base Reactions

An important class of reactions that occurs in aqueous solution is the acid–base reaction. In an **acid–base reaction** (also called a **neutralization reaction**), an acid reacts with a base and the two neutralize each other, producing water (or in some cases a weak electrolyte). In this section, we discuss how to define acids and bases, how to name acids, and how to write equations for the reactions between acids and bases.

> We discuss acids and bases in more detail in Chapter 16.

Properties of Acids and Bases

As we saw in Section 8.4, we can define *acids* as molecular compounds that release hydrogen ions (H^+) when dissolved in water. For example, HCl(*aq*) is an acid because it produces H^+ ions in solution:

$$HCl\,(aq) \rightarrow H^+\,(aq) + Cl^-\,(aq)$$

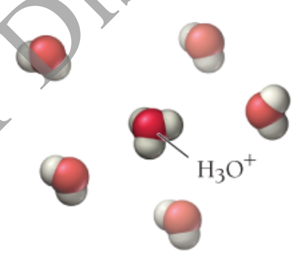
This definition of an acid is the **Arrhenius definition**.

An H^+ ion is a proton. In solution, bare protons normally associate with water molecules to form **hydronium ions** (Figure 8.15):

$$H^+\,(aq) + H_2O\,(l) \rightarrow H_3O^+\,(aq)$$

Figure 8.15 The Hydronium Ion

Protons normally associate with water molecules in solution to form H_3O^+ ions.



Chemists use $H^+\,(aq)$ and $H_3O^+\,(aq)$ interchangeably to mean the same thing—a hydronium ion. The chemical equation for the ionization of HCl and other acids is often written to show the association of the proton with a water molecule to form the hydronium ion:

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

As we discussed in Section 9.4, some acids are weak acids—they do not completely ionize in solution. We represent the ionization of a weak acid with opposing half arrows.

$$HC_2H_3O_2\,(aq) \rightleftharpoons H^+\,(aq) + C_2H_3O_2{}^-\,(aq)$$

Some acids—called **polyprotic acids**—contain more than one ionizable proton and release them sequentially. For example, sulfuric acid, H_2SO_4, is a **diprotic acid**. It is strong in its first ionizable proton, but weak in its second:

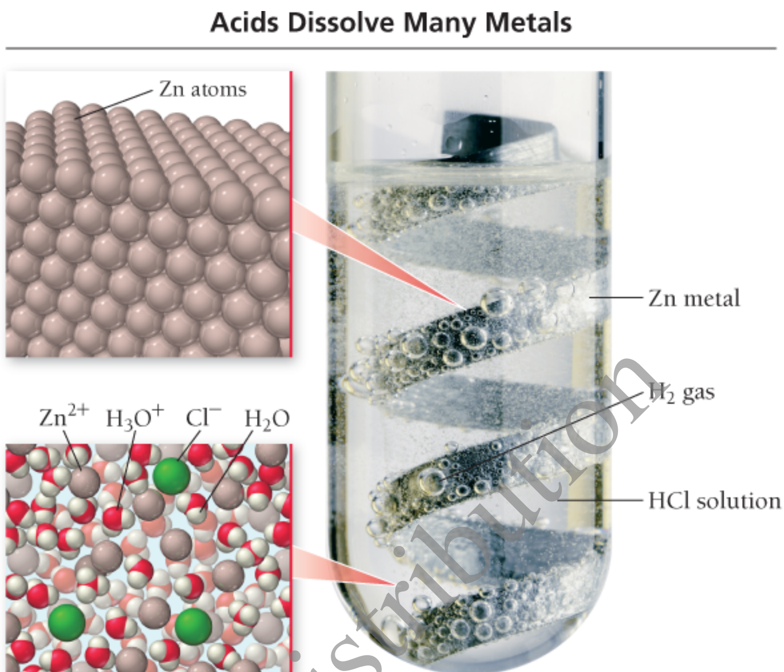
$$H_2SO_4\,(aq) \rightarrow H^+\,(aq) + HSO_4{}^-\,(aq)$$
$$HSO_4{}^-\,(aq) \rightleftharpoons H^+\,(aq) + SO_4{}^{2-}\,(aq)$$

$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$

Acids are characterized by their sour taste and their ability to dissolve many metals. For example, hydrochloric acid is present in stomach fluids; its sour taste becomes painfully obvious during vomiting. Hydrochloric acid dissolves some metals. For example, if you put a strip of zinc into a test tube of hydrochloric acid, it slowly dissolves as the $H^+(aq)$ ions convert the zinc metal into $Zn^{2+}(aq)$ cations (Figure 8.16). Acids are present in foods such as lemons and limes and are used in household products such as toilet bowl cleaners and Lime-Away®.

Figure 8.16 Hydrochloric Acid Dissolving Zinc Metal

The zinc atoms are ionized to zinc ions, which dissolve in the water. The HCl forms H_2 gas, which is responsible for the bubbles you can see in the test tube.



Many fruits are acidic and have the characteristically sour taste of acids.

According to the Arrhenius definition, **bases** produce OH^- in solution. Sodium hydroxide (NaOH) is a base because it produces OH^- ions in solution:

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

In analogy to diprotic acids, some bases, such as $Sr(OH)_2$, produce two moles of OH^- per mole of the base:

$$Sr(OH)_2(aq) \rightarrow Sr^{2+}(aq) + 2\ OH^-(aq)$$

Table 8.2 lists common acids and bases. Acids and bases are present in many everyday substances. We have already mentioned that foods such as citrus fruits and vinegar contain acids. Soap, baking soda, and milk of magnesia all contain bases.

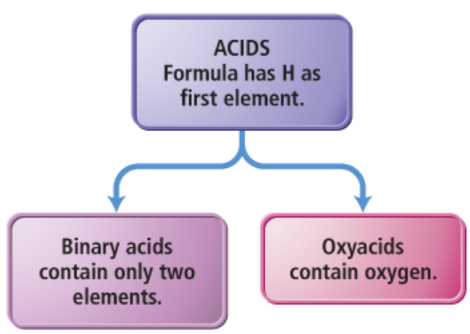
Table 8.2 **Some Common Acids and Bases**

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	$NaOH$
Hydrobromic acid	HBr	Lithium hydroxide	$LiOH$
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Acetic acid	$HC_2H_3O_2$ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

*Ammonia does not contain OH, but it produces OH in a reaction with water that occurs only to a small extent: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

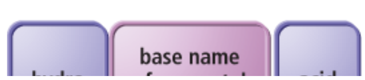
Many common household products are bases.

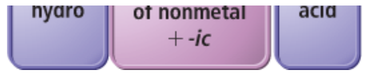
We can categorize acids into two types: binary acids and oxyacids.



Naming Binary Acids

Binary acids are composed of hydrogen and a nonmetal. Names for binary acids have the form:





For example, HCl(*aq*) is hydro*chlor*ic acid and HBr(*aq*) is hydro*brom*ic acid.

$$HCl\,(aq)\ \text{hydrochloric acid}\quad HBr\,(aq)\ \text{hydrobromic acid}$$

Example 8.9 Naming Binary Acids

Name HI(*aq*).

SOLUTION

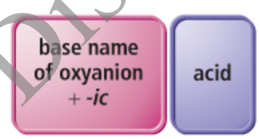
The base name of I is *iod* so HI(*aq*) is hydroiodic acid.

$$HI\,(aq)\quad \text{hydroiodic acid}$$

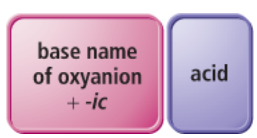
FOR PRACTICE 8.9 Name HF(*aq*).

Naming Oxyacids

Oxyacids contain hydrogen and an oxyanion (an anion containing a nonmetal and oxygen). The common oxyanions are listed in the table of polyatomic ions in Chapter 4 (Table 4.4). For example, $HNO_3\,(aq)$ contains the nitrate $(NO_3{}^-)$ ion, $H_2SO_3\,(aq)$ contains the sulfite $(SO_3{}^{2-})$ ion, and $H_2SO_4\,(aq)$ contains the sulfate $(SO_4{}^{2-})$ ion. Oxyacids are a combination of one or more H^+ ions with an oxyanion (see Table 4.4). The number of H^+ ions depends on the charge of the oxyanion; the formula is always charge-neutral. The names of oxyacids depend on the ending of the oxyanion and take the following forms:



oxyanions ending with *-ate*



oxyanions ending with *-ite*

So $HNO_3\,(aq)$ is nitric acid (oxyanion is nitrate), and $H_2SO_3\,(aq)$ is sulfurous acid (oxyanion is sulfite).

$$HNO_3\,(aq)\quad \text{nitric acid}\quad H_2SO_3\,(aq)\quad \text{sulfurous acid}$$

Example 8.10 Naming Oxyacids

Name $HC_2H_3O_2\,(aq)$.

SOLUTION The oxyanion is acetate, which ends in *-ate;* therefore, the name of the acid is *acetic acid.*

$$HC_2H_3O_2\,(aq)\quad \text{acetic acid}$$

FOR PRACTICE 8.10 Name $HNO_2\,(aq)$

FOR MORE PRACTICE 8.10 Write the formula for perchloric acid.

Acid–Base Reactions

Our stomachs contain hydrochloric acid ($HCl(aq)$), which acts in the digestion of food. Certain foods or stress, however, can increase the stomach's acidity to uncomfortable levels, causing acid stomach or heartburn. Antacids are over-the-counter medicines that work by reacting with and neutralizing stomach acid. Antacids employ different bases as neutralizing agents. Milk of magnesia, for example, contains $Mg(OH)_2$ and Mylanta® contains $Al(OH)_3$. All antacids, regardless of the base they employ, neutralize stomach acid and relieve heartburn through *acid–base reactions*.

When an acid and a base mix, the $H^+\,(aq)$ from the acid—whether it is weak or strong—combines with the $OH^-\,(aq)$ from the base to form $H_2O\,(l)$ (Figure 8.17). Consider the reaction between hydrochloric acid and sodium hydroxide:

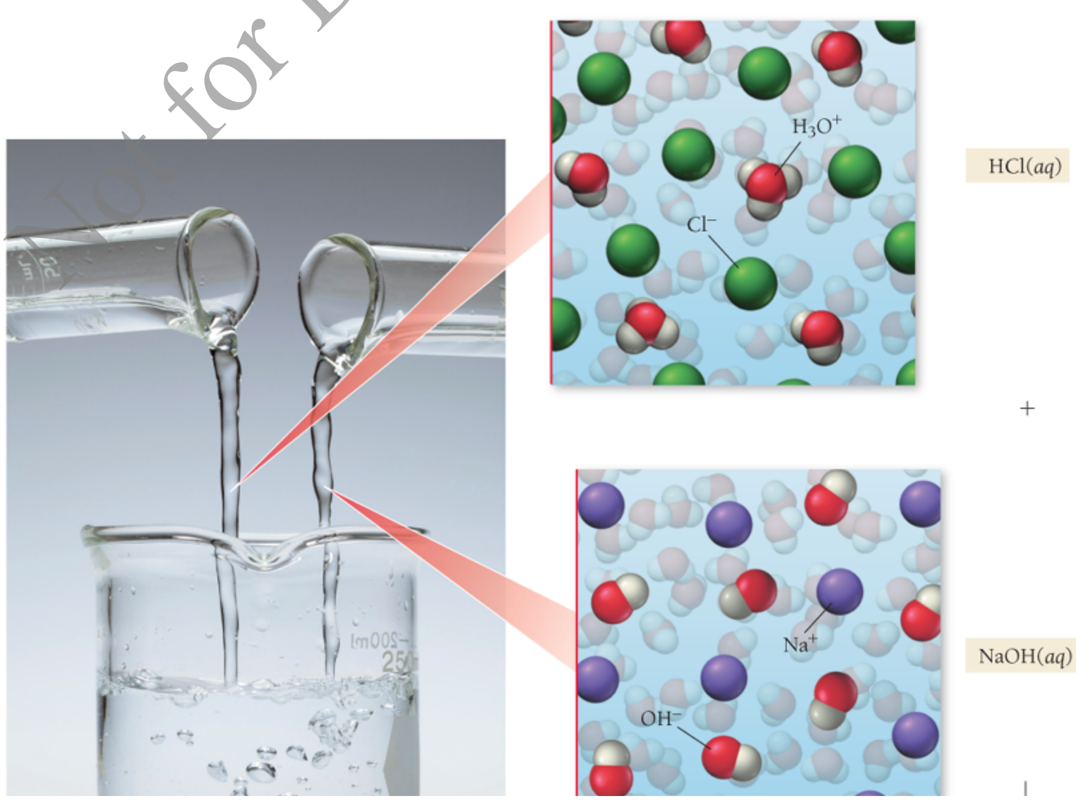
$$\underset{\text{Acid}}{HCl(aq)} + \underset{\text{Base}}{NaOH(aq)} \longrightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{Salt}}{NaCl(aq)}$$

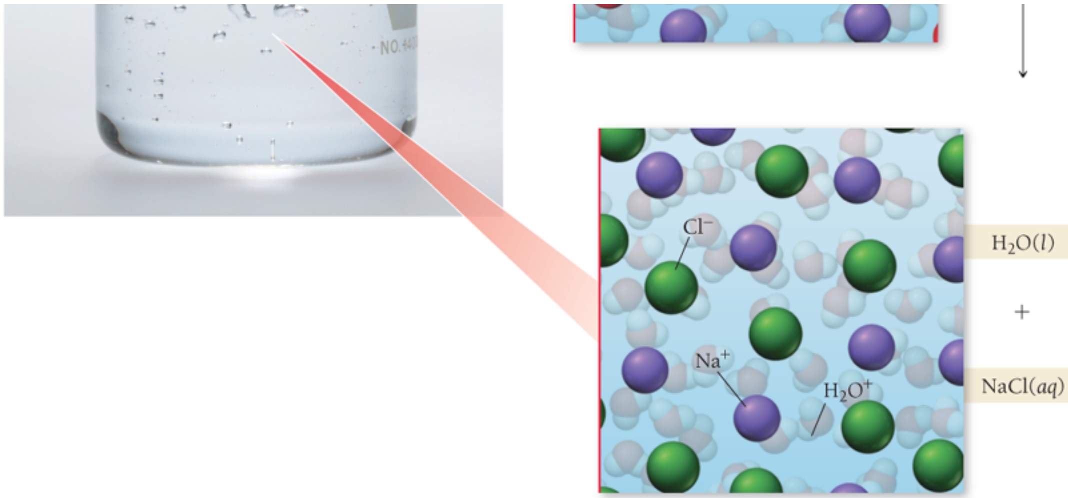
Figure 8.17 Acid–Base Reaction

Acid–Base Reaction

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

The reaction between hydrochloric acid and sodium hydroxide forms water and a salt, sodium chloride, which remains dissolved in the solution.





Acid–base reactions generally form water and an ionic compound—called a **salt**—that usually remains dissolved in the solution. The net ionic equation for acid–base reactions involving a strong acid is:

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

> The word *salt* in this sense applies to any ionic compound and is therefore more general than the common usage, which refers only to table salt (NaCl).

However, if the acid is a weak acid, the net ionic equation is slightly different. For example, consider the acid–base equation between hydrofluoric acid and sodium hydroxide:

$$\underset{\text{acid}}{HF\,(aq)} + \underset{\text{base}}{NaOH\,(aq)} \rightarrow \underset{\text{water}}{H_2O\,(l)} + \underset{\text{salt}}{NaF\,(aq)}$$

The complete ionic equation and net ionic equation for this reaction are:

$$\begin{array}{rcll} HF\,(aq) + Na^+\,(aq) + OH^-\,(aq) & \rightarrow & H_2O\,(l) + Na^+\,(aq) + F^-\,(aq) & \text{(Complete ionic equation)} \\ HF\,(aq) + OH^-\,(aq) & \rightarrow & H_2O\,(l) + F^-\,(aq) & \text{(Net ionic equation)} \end{array}$$

Notice that since HF is a weak acid, we do not show it as ionized in the ionic equations.

Another example of an acid–base reaction is the reaction between sulfuric acid and potassium hydroxide:

$$\underset{\text{acid}}{H_2SO_4\,(aq)} + \underset{\text{base}}{2\,KOH\,(aq)} \rightarrow \underset{\text{water}}{2\,H_2O\,(l)} + \underset{\text{salt}}{K_2SO_4\,(aq)}$$

Again, notice the pattern of acid and base reacting to form water and a salt.

$$\textbf{Acid} + \textbf{Base} \rightarrow \textbf{Water} + \textbf{Salt} \quad \text{(acid-base reactions)}$$

When writing equations for acid–base reactions, write the formula of the salt using the procedure for writing formulas of ionic compounds demonstrated in Section 4.6.

Example 8.11 Writing Equations for Acid–Base Reactions Involving a Strong Acid

Write a molecular and net ionic equation for the reaction between aqueous HI and aqueous $Ba(OH)_2$.

SOLUTION First identify these substances as an acid and a base. Begin by writing the unbalanced equation in which the acid and the base combine to form water and a salt.

$$\underset{\text{acid}}{HI\,(aq)} + \underset{\text{base}}{Ba(OH)_2\,(aq)} \rightarrow \underset{\text{water}}{H_2O\,(l)} + \underset{\text{salt}}{BaI_2\,(aq)}$$

Balance the equation; this is the molecular equation.

Molecular equation:

$$2\ \mathrm{HI}\,(aq) + \mathrm{Ba(OH)_2}\,(aq) \rightarrow 2\ \mathrm{H_2O}\,(l) + \mathrm{BaI_2}\,(aq)$$

Write the net ionic equation by removing the spectator ions.

Net ionic equation:

$$2\ \mathrm{H^+}\,(aq) + 2\ \mathrm{OH^-}\,(aq) \rightarrow 2\ \mathrm{H_2O}\,(l) \text{ or simply}$$
$$\mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{H_2O}\,(l)$$

FOR PRACTICE 8.11 Write a molecular and a net ionic equation for the reaction that occurs between aqueous HBr and aqueous LiOH.

Example 8.12 Writing Equations for Acid–Base Reactions Involving a Weak Acid

Write a molecular equation, complete ionic equation, and net ionic equation for the reaction between aqueous acetic acid ($\mathrm{HC_2H_3O_2}$) and aqueous potassium hydroxide (KOH).

SOLUTION Begin by writing the molecular equation in which the acid and the base combine to form water and a salt. (The equation is already balanced.)

Molecular equation:

$$\mathrm{HC_2H_3O_2}\,(aq) + \mathrm{KOH}\,(aq) \rightarrow \mathrm{H_2O}\,(l) + \mathrm{KC_2H_3O_2}\,(aq)$$

Write the complete ionic equation by separating aqueous ionic compounds into their constituent ions. Do not separate $\mathrm{HC_2H_3O_2}\,(aq)$ because it is a weak acid (and a weak electrolyte).

Complete ionic equation:

$$\mathrm{HC_2H_3O_2}\,(aq) + \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{H_2O}\,(l) + \mathrm{K^+}\,(aq) + \mathrm{C_2H_3O_2}^-\,(aq)$$

Write the net ionic equation by eliminating the spectator ions.

Net ionic equation:

$$\mathrm{HC_2H_3O_2}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{H_2O}\,(l) + \mathrm{C_2H_3O_2}^-\,(aq)$$

FOR PRACTICE 8.12 Write the net ionic equation for the reaction between $\mathrm{HCHO_2}$ (a weak acid) and NaOH.

Acid–Base Titrations

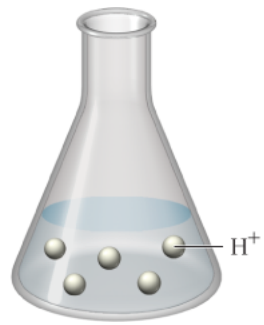
We can apply the principles of acid–base neutralization and stoichiometry to a common laboratory procedure called a *titration*. In a **titration**, a substance in a solution of known concentration is reacted with another substance in a solution of unknown concentration. For example, consider the following acid–base reaction:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{H_2O}\,(l) + \mathrm{NaCl}\,(aq)$$

The net ionic equation for this reaction eliminates the spectator ions:

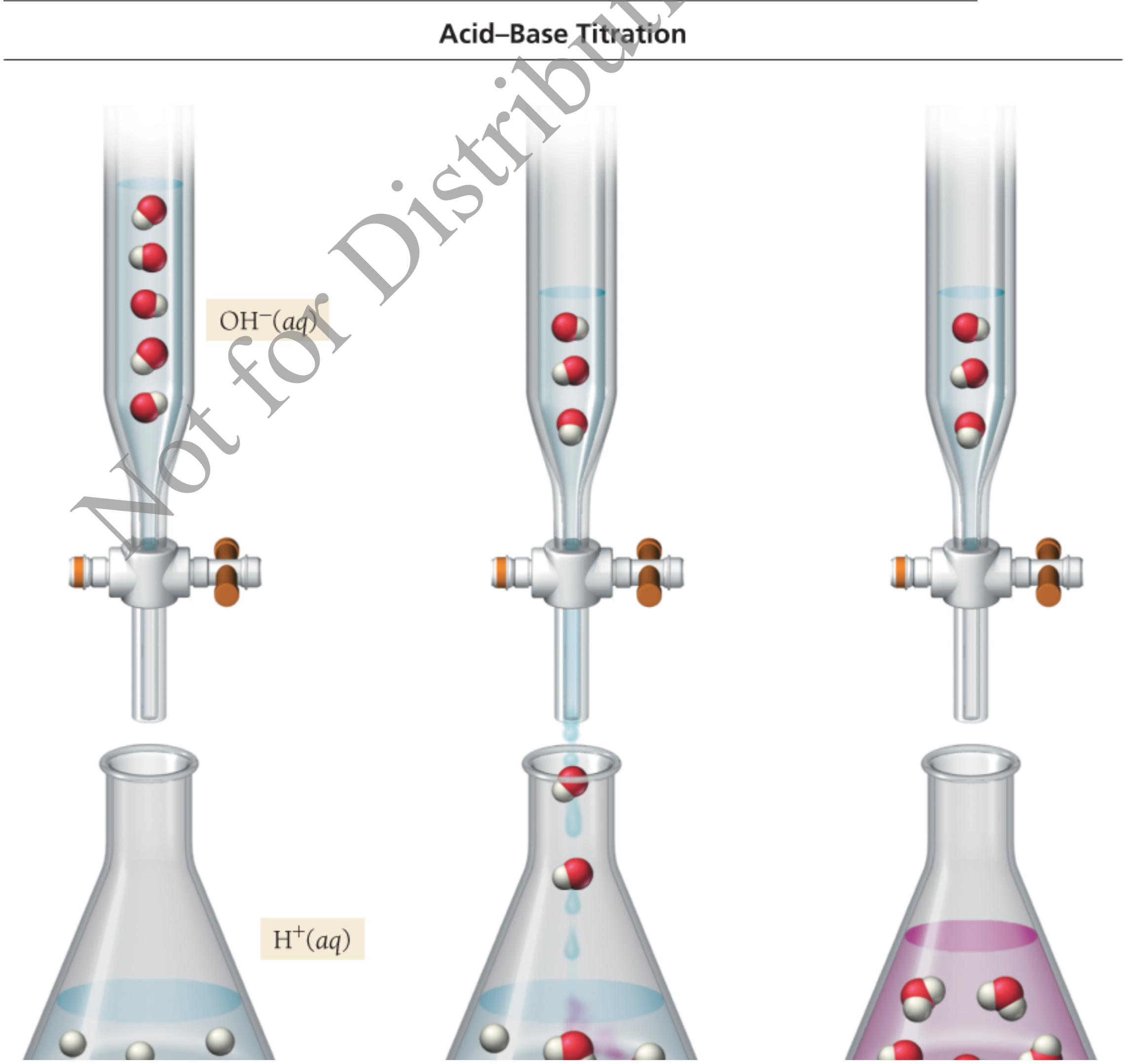
$$H^+\,(aq) + OH^-\,(aq) \rightarrow H_2O\,(l)$$

Suppose we have an HCl solution represented by the following molecular diagram (we have omitted the Cl^- ions and the H_2O molecules not involved in the reaction from this representation for clarity):



In titrating this sample, we slowly add a solution of known OH^- concentration, as shown in the molecular diagrams in Figure 8.18. As the OH^- is added, it reacts with and neutralizes the H^+, forming water. At the **equivalence point**—the point in the titration when the number of moles of OH^- equals the number of moles of H^+ in solution—the titration is complete. The equivalence point is typically signaled by an **indicator**, a dye whose color depends on the acidity or basicity of the solution (Figure 8.19).

Figure 8.18 Acid–Base Titration



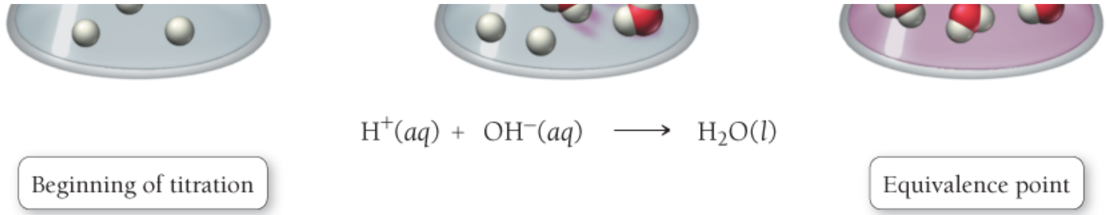


Figure 8.19 **Titration**

In this titration, NaOH is added to a dilute HCl solution. When the NaOH and HCl reach stoichiometric proportions (the equivalence point), the phenolphthalein indicator changes color to pink.

Indicator in Titration

We cover acid–base titrations and indicators in more detail in Chapter 16. In most laboratory titrations, the concentration of one of the reactant solutions is unknown, and the concentration of the other is precisely known. By carefully measuring the volume of each solution required to reach the equivalence point, we can determine the concentration of the unknown solution, as demonstrated in Example 8.13.

Example 8.13 Acid–Base Titration

The titration of a 10.00-mL sample of an HCl solution of unknown concentration requires 12.54 mL of 0.100 M NaOH solution to reach the equivalence point. What is the concentration of the unknown HCl solution in M?

SORT You are given the volume and concentration of NaOH solution required to titrate a given volume of HCl solution. You are asked to find the concentration of the HCl solution.

GIVEN: 12.54 mL of NaOH solution; 0.100 M NaOH solution; 10.00 mL of HCl solution

FIND: concentration of HCl solution

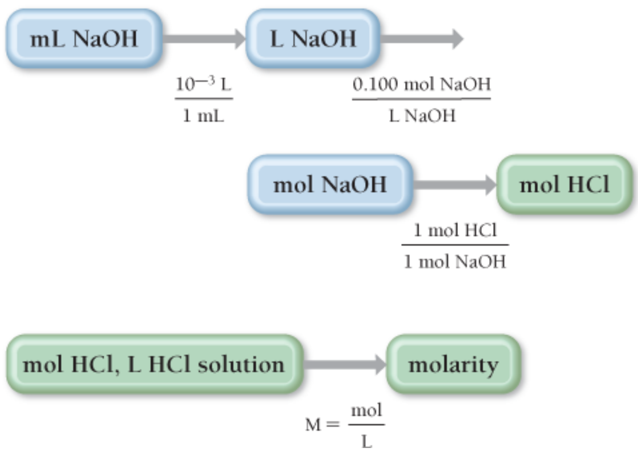
STRATEGIZE Since this problem involves an acid–base neutralization reaction between HCl and NaOH, start by writing the balanced equation.

The first part of the conceptual plan has the form: volume A → moles A → moles B. The concentration of the NaOH solution is a conversion factor between moles and volume of NaOH. The balanced equation provides the relationship between number of moles of NaOH and number of moles of HCl.

In the second part of the conceptual plan, use the number of moles of HCl (from the first part) and the volume of HCl solution (given) to calculate the molarity of the HCl solution.

$HCl\,(aq) + NaOH\,(aq) \rightarrow H_2O\,(l) + NaCl\,(aq)$

CONCEPTUAL PLAN



RELATIONSHIPS USED

$10^{-3}\ L = 1\ mL$

$M\ (NaOH) = \dfrac{0.100\ mol\ NaOH}{L\ NaOH}$

1 mol HCl: 1 mol NaOH

$$\text{molarity (M)} = \frac{\text{moles of solute (mol)}}{\text{volume of solution (L)}}$$

SOLVE First determine the number of moles of HCl in the unknown solution.

Next divide the number of moles of HCl by the volume of the HCl solution in L. 10.0 mL is equivalent to 0.010 L.

SOLUTION

$$12.54\ \cancel{mL}\ \cancel{NaOH} \times \frac{10^{-3}\ \cancel{L}}{1\ \cancel{mL}} \times \frac{0.100\ \cancel{mL}\ \cancel{NaOH}}{\cancel{L}\ \cancel{NaOH}} \times \frac{1\ mol\ HCl}{1\ \cancel{mol}\ \cancel{NaOH}} = 1.25 \times 10^{-3}\ mol\ HCl$$

$$\text{molarity} = \frac{1.25 \times 10^{-3}\ mol\ HCl}{0.01000\ L} = 0.125\ M\ HCl$$

CHECK The units of the answer (M HCl) are correct. The magnitude of the answer (0.125 M) is reasonable because it is similar to the molarity of the NaOH solution, as expected from the reaction stoichiometry (1 mol HCl reacts with 1 mol NaOH) and the similar volumes of NaOH and HCl.

FOR PRACTICE 8.13 The titration of a 20.0-mL sample of H_2SO_4 solution of unknown concentration requires 22.87 mL of a 0.158 M KOH solution to reach the equivalence point. What is the concentration of the unknown H_2SO_4 solution?

FOR MORE PRACTICE 8.13 What volume (in mL) of 0.200 M NaOH do you need to titrate 35.00 mL of 0.140 M HBr to the equivalence point?