

Chapter 10

Acids and Bases

Opening Essay

One of the most concentrated acids in the body is stomach acid, which can be approximated as a 0.05 M hydrochloric acid solution. Special cells in the stomach wall secrete this acid, along with special enzymes, as part of the digestion process. In a laboratory, a 0.05 M solution of hydrochloric acid would dissolve some metals. How does the stomach survive the presence of such a reactive acid?

Actually, the stomach has several mechanisms for withstanding this chemical onslaught. First, the lining of the stomach is coated with a thin layer of mucus that contains some bicarbonate ions (HCO_3^-). These react with the hydrochloric acid to produce water, carbon dioxide, and harmless chloride ions. If any acid penetrates through the mucus, it can attack the surface layer of stomach cells, called the *gastric epithelium*. Cells in the gastric epithelium are being constantly shed, so damaged cells are quickly removed and replaced with healthy cells.

However, if the gastric epithelium is destroyed faster than it can be replaced, the acid may reach the wall of the stomach, resulting in ulcers. If an ulcer grows large enough, it can expose blood vessels in the stomach wall, causing bleeding. In extreme situations, the loss of blood through a severe ulcer can threaten a person's health.

Ulcers can also result from the presence of a certain bacterium—*Helicobacter pylori*—in the stomach. The mechanism for this ulcer formation is not the same as that for ulcers caused by stomach acid and is not completely understood. However, there are two main treatments for ulcers: (1) *antacids* to react chemically with excess hydrochloric acid in the stomach and (2) *antibiotics* to destroy the *H. pylori* bacteria in the stomach.

Many of us are familiar with the group of chemicals called *acids*. But do you know what it takes for a compound to be an acid? Actually, there are several different definitions of *acid* that chemistry uses, and each definition is appropriate under different circumstances. Less familiar—but just as important to chemistry and ultimately to us—is the group of chemicals known as *bases*. Both acids and bases are important enough that we devote an entire chapter to them—their properties and their reactions.

Figure 10.1 "Prevalence of Acids and Bases" illustrates how common acids and bases are in everyday life.

Figure 10.1 Prevalence of Acids and Bases



The products shown in this photograph, all acids or bases, give an idea of how prevalent and important acids and bases are in everyday life.

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10.1 Arrhenius Definition of Acids and Bases

LEARNING OBJECTIVE

1. Recognize a compound as an Arrhenius acid or an Arrhenius base.

One way to define a class of compounds is by describing the various characteristics its members have in common. In the case of the compounds known as acids, the common characteristics include a sour taste, the ability to change the color of the vegetable dye *litmus* to red, and the ability to dissolve certain metals and simultaneously produce hydrogen gas. For the compounds called bases, the common characteristics are a slippery texture, a bitter taste, and the ability to change the color of litmus to blue. Acids and bases also react with each other to form compounds generally known as salts.

Note

Although we include their tastes among the common characteristics of acids and bases, we never advocate tasting an unknown chemical!

Chemists prefer, however, to have definitions for acids and bases in chemical terms. The Swedish chemist Svante Arrhenius developed the first chemical definitions of acids and bases in the late 1800s. Arrhenius defined an **acid** as a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution. Many acids are simple compounds that release a hydrogen cation into solution when they dissolve. Similarly, Arrhenius defined a **base** as a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution. Many bases are ionic compounds that have the hydroxide ion as their anion, which is released when the base dissolves in water.

Many bases and their aqueous solutions are named using the normal rules of ionic compounds that were presented in Chapter 3 "Ionic Bonding and Simple Ionic Compounds", Section 3.4 "Ionic Nomenclature"; that is, they are named as hydroxide compounds. For example, the base sodium hydroxide (NaOH) is both an ionic compound and an aqueous solution. However, aqueous solutions of acids have their own naming rules. The names of *binary acids* (compounds with hydrogen and one

other element in their formula) are based on the root of the name of the other element preceded by the prefix *hydro-* and followed by the suffix *-ic acid*. Thus, an aqueous solution of HCl [designated “HCl(aq)”] is called hydrochloric acid, H₂S(aq) is called hydrosulfuric acid, and so forth. Acids composed of more than two elements (typically hydrogen and oxygen and some other element) have names based on the name of the other element, followed by the suffix *-ic acid* or *-ous acid*, depending on the number of oxygen atoms in the acid’s formula. Other prefixes, like *per-* and *hypo-*, also appear in the names for some acids. Unfortunately, there is no strict rule for the number of oxygen atoms that are associated with the *-ic acid* suffix; the names of these acids are best memorized. Table 10.1 "Formulas and Names for Some Acids and Bases" lists some acids and bases and their names. Note that acids have hydrogen written first, as if it were the cation, while most bases have the negative hydroxide ion, if it appears in the formula, written last.

Note

The name *oxygen* comes from the Latin meaning “acid producer” because its discoverer, Antoine Lavoisier, thought it was the essential element in acids. Lavoisier was wrong, but it is too late to change the name now.

Table 10.1 Formulas and Names for Some Acids and Bases

Formula	Name
Acids	
HCl(aq)	hydrochloric acid
HBr(aq)	hydrobromic acid
HI(aq)	hydriodic acid
H ₂ S(aq)	hydrosulfuric acid
HC ₂ H ₃ O ₂ (aq)	acetic acid
HNO ₃ (aq)	nitric acid
HNO ₂ (aq)	nitrous acid

$\text{H}_2\text{SO}_4(\text{aq})$	sulfuric acid
$\text{H}_2\text{SO}_3(\text{aq})$	sulfurous acid
$\text{HClO}_3(\text{aq})$	chloric acid
$\text{HClO}_4(\text{aq})$	perchloric acid
$\text{HClO}_2(\text{aq})$	chlorous acid
$\text{H}_3\text{PO}_4(\text{aq})$	phosphoric acid
$\text{H}_3\text{PO}_3(\text{aq})$	phosphorous acid
Bases	
$\text{NaOH}(\text{aq})$	sodium hydroxide
$\text{KOH}(\text{aq})$	potassium hydroxide
$\text{Mg}(\text{OH})_2(\text{aq})$	magnesium hydroxide
$\text{Ca}(\text{OH})_2(\text{aq})$	calcium hydroxide
$\text{NH}_3(\text{aq})$	ammonia

EXAMPLE 1

Name each substance.

1. $\text{HF}(\text{aq})$
2. $\text{Sr}(\text{OH})_2(\text{aq})$

Solution

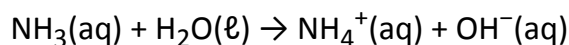
1. This acid has only two elements in its formula, so its name includes the *hydro-* prefix. The stem of the other element's name, fluorine, is *fluor*, and we must also include the *-ic acid* ending. Its name is hydrofluoric acid.
2. This base is named as an ionic compound between the strontium ion and the hydroxide ion: strontium hydroxide.

SKILL-BUILDING EXERCISE

Name each substance.

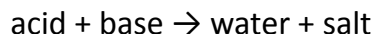
1. $\text{H}_2\text{Se}(\text{aq})$
2. $\text{Ba}(\text{OH})_2(\text{aq})$

Notice that one base listed in Table 10.1 "Formulas and Names for Some Acids and Bases"—ammonia—does not have hydroxide as part of its formula. How does this compound increase the amount of hydroxide ion in aqueous solution? Instead of dissociating into hydroxide ions, ammonia molecules react with water molecules by taking a hydrogen ion from the water molecule to produce an ammonium ion and a hydroxide ion:

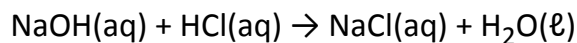


Because this reaction of ammonia with water causes an increase in the concentration of hydroxide ions in solution, ammonia satisfies the Arrhenius definition of a base. Many other nitrogen-containing compounds are bases because they too react with water to produce hydroxide ions in aqueous solution.

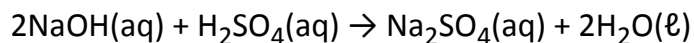
As we noted previously, acids and bases react chemically with each other to form *salts*. A salt is a general chemical term for any ionic compound formed from an acid and a base. In reactions where the acid is a hydrogen ion containing compound and the base is a hydroxide ion containing compound, water is also a product. The general reaction is as follows:



The reaction of acid and base to make water and a salt is called **neutralization**. Like any chemical equation, a neutralization chemical equation must be properly balanced. For example, the neutralization reaction between sodium hydroxide and hydrochloric acid is as follows:



with coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:



Once a neutralization reaction is properly balanced, we can use it to perform stoichiometry calculations, such as the ones we practiced in Chapter 5 "Introduction to Chemical Reactions" and Chapter 6 "Quantities in Chemical Reactions".

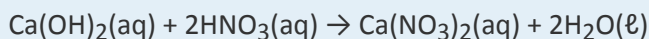
EXAMPLE 2

Nitric acid [$\text{HNO}_3(\text{aq})$] can be neutralized by calcium hydroxide [$\text{Ca}(\text{OH})_2(\text{aq})$].

1. Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.
2. For one reaction, 16.8 g of HNO_3 is present initially. How many grams of $\text{Ca}(\text{OH})_2$ are needed to neutralize that much HNO_3 ?
3. In a second reaction, 805 mL of 0.672 M $\text{Ca}(\text{OH})_2$ is present initially. What volume of 0.432 M HNO_3 solution is necessary to neutralize the $\text{Ca}(\text{OH})_2$ solution?

Solution

1. Because there are two OH^- ions in the formula for $\text{Ca}(\text{OH})_2$, we need two moles of HNO_3 to provide H^+ ions. The balanced chemical equation is as follows:



The salt formed is calcium nitrate.

2. This calculation is much like the calculations we did in Chapter 6 "Quantities in Chemical Reactions". First we convert the mass of HNO_3 to moles using its molar mass of $1.01 + 14.00 + 3(16.00) = 63.01 \text{ g/mol}$; then we use the balanced chemical equation to determine the related number of moles of $\text{Ca}(\text{OH})_2$ needed to neutralize it; and then we convert that number of moles of $\text{Ca}(\text{OH})_2$ to the mass of $\text{Ca}(\text{OH})_2$ using its molar mass of $40.08 + 2(1.01) + 2(16.00) = 74.10 \text{ g/mol}$.

$$16.8 \text{ g } \cancel{\text{HNO}_3} \times \frac{1 \text{ mol } \cancel{\text{HNO}_3}}{63.01 \text{ g } \cancel{\text{HNO}_3}} \times \frac{1 \text{ mol } \cancel{\text{Ca(OH)}_2}}{2 \text{ mol } \cancel{\text{HNO}_3}} \times \frac{74.10 \text{ g } \text{Ca(OH)}_2}{1 \text{ mol } \cancel{\text{Ca(OH)}_2}} = 9.88 \text{ g Ca(OH)}_2 \text{ needed}$$

3. Having concentration information allows us to employ the skills we developed in Chapter 9 "Solutions". First, we use the concentration and volume data to determine the number of moles of Ca(OH)_2 present. Recognizing that $805 \text{ mL} = 0.805 \text{ L}$,

$$0.672 \text{ M Ca(OH)}_2 = \frac{\text{mol Ca(OH)}_2}{0.805 \text{ L soln}}$$

$$(0.672 \text{ M Ca(OH)}_2) \times (0.805 \text{ L soln}) = \text{mol Ca(OH)}_2 = 0.541 \text{ mol Ca(OH)}_2$$

We combine this information with the proper ratio from the balanced chemical equation to determine the number of moles of HNO_3 needed:

$$0.541 \text{ mol } \cancel{\text{Ca(OH)}_2} \times \frac{2 \text{ mol } \text{HNO}_3}{1 \text{ mol } \cancel{\text{Ca(OH)}_2}} = 1.08 \text{ mol HNO}_3$$

Now, using the definition of molarity one more time, we determine the volume of acid solution needed:

$$0.432 \text{ M HNO}_3 = \frac{1.08 \text{ mol HNO}_3}{\text{volume of HNO}_3}$$

$$\text{volume of HNO}_3 = \frac{1.08 \text{ mol HNO}_3}{0.432 \text{ M HNO}_3} = 2.50 \text{ L} = 2.50 \times 10^3 \text{ mL HNO}_3$$

SKILL-BUILDING EXERCISE

Hydrocyanic acid [HCN(aq)] can be neutralized by potassium hydroxide [KOH(aq)].

- Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.
- For one reaction, 37.5 g of HCN is present initially. How many grams of KOH are needed to neutralize that much HCN ?

- In a second reaction, 43.0 mL of 0.0663 M KOH is present initially. What volume of 0.107 M HCN solution is necessary to neutralize the KOH solution?

Note

Hydrocyanic acid (HCN) is one exception to the acid-naming rules that specify using the prefix *hydro-* for binary acids (acids composed of hydrogen and only one other element).

CONCEPT REVIEW EXERCISES

- Give the Arrhenius definitions of an acid and a base.
- What is neutralization?

ANSWERS

- Arrhenius acid: a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution;
Arrhenius base: a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution.
- the reaction of an acid and a base

KEY TAKEAWAY

- An Arrhenius acid increases the H^+ ion concentration in water, while an Arrhenius base increases the OH^- ion concentration in water.

EXERCISES

1. Give two examples of Arrhenius acids.
2. Give two examples of Arrhenius bases.
3. List the general properties of acids.
4. List the general properties of bases.
5. Name each compound.
 - a. HBr(aq)
 - b. $\text{Ca(OH)}_2\text{(aq)}$
 - c. $\text{HNO}_3\text{(aq)}$
 - d. $\text{Fe(OH)}_3\text{(aq)}$
6. Name each compound.
 - a. HI(aq)
 - b. $\text{Cu(OH)}_2\text{(aq)}$
 - c. $\text{H}_3\text{PO}_4\text{(aq)}$
 - d. CsOH(aq)
7. Propose a name for water (H_2O) using the rules for naming acids.
8. Propose a name for hydrogen peroxide (H_2O_2) using the rules for naming acids.
9. Write a balanced chemical equation for the neutralization of $\text{Ba(OH)}_2\text{(aq)}$ with $\text{HNO}_3\text{(aq)}$.
10. Write a balanced chemical equation for the neutralization of $\text{H}_2\text{SO}_4\text{(aq)}$ with $\text{Cr(OH)}_3\text{(aq)}$.
11. How many moles of sodium hydroxide (NaOH) are needed to neutralize 0.844 mol of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$)? (Hint: begin by writing a balanced chemical equation for the process.)
12. How many moles of perchloric acid (HClO_4) are needed to neutralize 0.052 mol of calcium hydroxide [Ca(OH)_2]? (Hint: begin by writing a balanced chemical equation for the process.)
13. Hydrazoic acid (HN_3) can be neutralized by a base.
 - a. Write the balanced chemical equation for the reaction between hydrazoic acid and calcium hydroxide.

- b. How many milliliters of 0.0245 M $\text{Ca}(\text{OH})_2$ are needed to neutralize 0.564 g of HN_3 ?
14. Citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$) has three hydrogen atoms that can form hydrogen ions in solution.
- Write the balanced chemical equation for the reaction between citric acid and sodium hydroxide.
 - If an orange contains 0.0675 g of $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$, how many milliliters of 0.00332 M NaOH solution are needed to neutralize the acid?
15. Magnesium hydroxide [$\text{Mg}(\text{OH})_2$] is an ingredient in some antacids. How many grams of $\text{Mg}(\text{OH})_2$ are needed to neutralize the acid in 158 mL of 0.106 M $\text{HCl}(\text{aq})$? It might help to write the balanced chemical equation first.
16. Aluminum hydroxide [$\text{Al}(\text{OH})_3$] is an ingredient in some antacids. How many grams of $\text{Al}(\text{OH})_3$ are needed to neutralize the acid in 96.5 mL of 0.556 M $\text{H}_2\text{SO}_4(\text{aq})$? It might help to write the balanced chemical equation first.

ANSWERS

1. HCl and HNO_3 (answers will vary)
3. sour taste, react with metals, react with bases, and turn litmus red
- 5.
- hydrobromic acid
 - calcium hydroxide
 - nitric acid
 - iron(III) hydroxide
7. perhaps hydroxic acid
9. $2\text{HNO}_3(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{Ba}(\text{NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O}$
11. 0.844 mol
- 13.
- $2\text{HN}_3(\text{aq}) + \text{Ca}(\text{OH})_2 \rightarrow \text{Ca}(\text{N}_3)_2 + 2\text{H}_2\text{O}$
 - 268 mL

15. 0.488 g

10.2 Brønsted-Lowry Definition of Acids and Bases

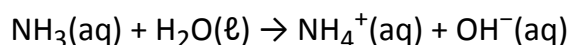
LEARNING OBJECTIVES

1. Recognize a compound as a Brønsted-Lowry acid or a Brønsted-Lowry base.
2. Illustrate the proton transfer process that defines a Brønsted-Lowry acid-base reaction.

Ammonia (NH_3) increases the hydroxide ion concentration in aqueous solution by reacting with water rather than releasing hydroxide ions directly. In fact, the Arrhenius definitions of an acid and a base focus on hydrogen ions and hydroxide ions. Are there more fundamental definitions for acids and bases?

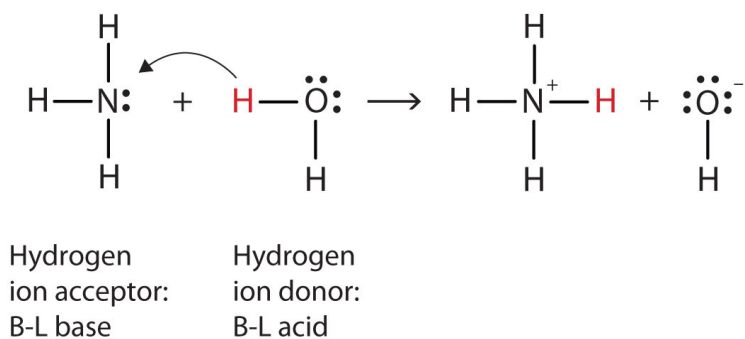
In 1923, the Danish scientist Johannes Brønsted and the English scientist Thomas Lowry independently proposed new definitions for acids and bases. Rather than considering both hydrogen and hydroxide ions, they focused on the hydrogen ion only. A **Brønsted-Lowry acid** is a compound that supplies a hydrogen ion in a reaction. A **Brønsted-Lowry base**, conversely, is a compound that accepts a hydrogen ion in a reaction. Thus, the Brønsted-Lowry definitions of an acid and a base focus on the movement of hydrogen ions in a reaction, rather than on the production of hydrogen ions and hydroxide ions in an aqueous solution.

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:



What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as

follows:



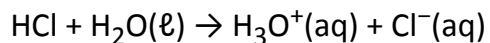
Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+ , which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

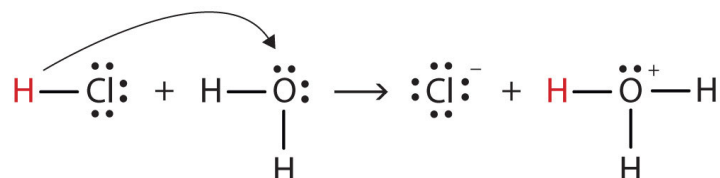
Note

A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like H_5O_2^+ or H_9O_4^+ rather than H_3O^+ . It is simpler, however, to use H_3O^+ .

With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O :



We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H₂O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H₂O is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H₂O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H₂O a base in this circumstance.

Note

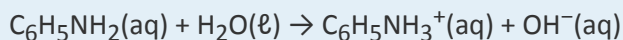
All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. But not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

EXAMPLE 3

Aniline (C₆H₅NH₂) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

Solution

$\text{C}_6\text{H}_5\text{NH}_2$ and H_2O are the reactants. When $\text{C}_6\text{H}_5\text{NH}_2$ accepts a proton from H_2O , it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:



Because $\text{C}_6\text{H}_5\text{NH}_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

SKILL-BUILDING EXERCISE

1. Caffeine ($\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$) is a stimulant found in coffees and teas. When dissolved in water, it can accept a proton from a water molecule. Write the chemical equation for this process and identify the Brønsted-Lowry acid and base.

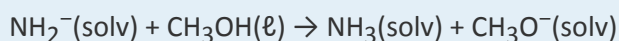
The Brønsted-Lowry definitions of an acid and a base can be applied to chemical reactions that occur in solvents other than water. The following example illustrates.

EXAMPLE 4

Sodium amide (NaNH_2) dissolves in methanol (CH_3OH) and separates into sodium ions and amide ions (NH_2^-). The amide ions react with methanol to make ammonia and the methoxide ion (CH_3O^-). Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

Solution

The equation for the reaction is between NH_2^- and CH_3OH to make NH_3 and CH_3O^- is as follows:



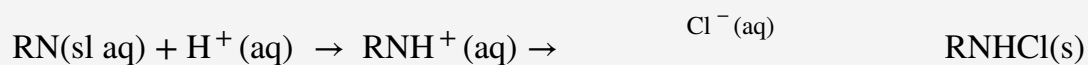
The label (*sol*) indicates that the species are dissolved in some solvent, in contrast to (*aq*), which specifies an aqueous (H_2O) solution. In this reaction, we see that the NH_2^- ion accepts a proton from a CH_3OH molecule to make an NH_3 molecule. Thus, as the proton acceptor, NH_2^- is the Brønsted-Lowry base. As the proton donor, CH_3OH is the Brønsted-Lowry acid.

SKILL-BUILDING EXERCISE

1. Pyridinium chloride ($\text{C}_5\text{H}_5\text{NHCl}$) dissolves in ethanol ($\text{C}_2\text{H}_5\text{OH}$) and separates into pyridinium ions ($\text{C}_5\text{H}_5\text{NH}^+$) and chloride ions. The pyridinium ion can transfer a hydrogen ion to a solvent molecule. Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

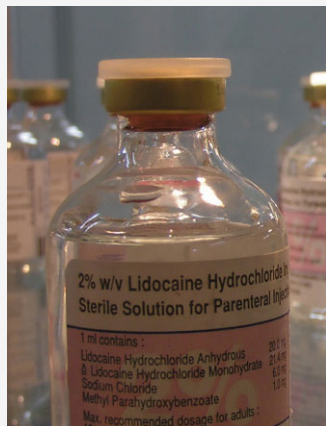
To Your Health: Brønsted-Lowry Acid-Base Reactions in Pharmaceuticals

There are many interesting applications of Brønsted-Lowry acid-base reactions in the pharmaceutical industry. For example, drugs often need to be water soluble for maximum effectiveness. However, many complex organic compounds are not soluble or are only slightly soluble in water. Fortunately, those drugs that contain proton-accepting nitrogen atoms (and there are a lot of them) can be reacted with dilute hydrochloric acid [$\text{HCl}(\text{aq})$]. The nitrogen atoms—acting as Brønsted-Lowry bases—accept the hydrogen ions from the acid to make an ion, which is usually much more soluble in water. The modified drug molecules can then be isolated as chloride salts:



where RN represents some organic compound containing nitrogen. The label (*sl aq*) means “slightly aqueous,” indicating that the compound RN is only slightly soluble. Drugs that are modified in this way are called *hydrochloride salts*. Examples include the powerful painkiller codeine, which is commonly administered as codeine hydrochloride. Acids other than hydrochloric

acid are also used. Hydrobromic acid, for example, gives *hydrobromide salts*. Dextromethorphan, an ingredient in many cough medicines, is dispensed as dextromethorphan hydrobromide. The accompanying figure shows another medication as a hydrochloride salt.



The name of this medicine makes it clear that it exists as the hydrochloride salt.

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CONCEPT REVIEW EXERCISE

1. Give the definitions of a Brønsted-Lowry acid and a Brønsted-Lowry base.

ANSWER

1. A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.

KEY TAKEAWAYS

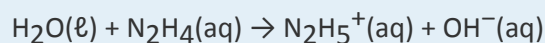
- A Brønsted-Lowry acid is a proton donor, and a Brønsted-Lowry base is a proton acceptor.
- Brønsted-Lowry acid-base reactions are essentially proton transfer reactions.

EXERCISES

1. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.



2. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.



3. Explain why a Brønsted-Lowry acid can be called a proton donor.
4. Explain why a Brønsted-Lowry base can be called a proton acceptor.
5. Write the chemical equation of the reaction of ammonia in water and label the Brønsted-Lowry acid and base.
6. Write the chemical equation of the reaction of methylamine (CH_3NH_2) in water and label the Brønsted-Lowry acid and base.
7. Demonstrate that the dissolution of HNO_3 in water is actually a Brønsted-Lowry acid-base reaction by describing it with a chemical equation and labeling the Brønsted-Lowry acid and base.
8. Identify the Brønsted-Lowry acid and base in the following chemical equation:
$$\text{C}_3\text{H}_7\text{NH}_2(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) \rightarrow \text{C}_3\text{H}_7\text{NH}_3^+(\text{aq}) + \text{H}_2\text{O}(\ell)$$
9. Write the chemical equation for the reaction that occurs when cocaine hydrochloride ($\text{C}_{17}\text{H}_{22}\text{ClNO}_4$) dissolves in water and donates a proton to a water molecule. (When hydrochlorides dissolve in water, they separate into chloride ions and the appropriate cation.)
10. If codeine hydrobromide has the formula $\text{C}_{18}\text{H}_{22}\text{BrNO}_3$, what is the formula of the parent compound codeine?

ANSWERS

1. HCl: Brønsted-Lowry acid; NH₃: Brønsted-Lowry base
3. A Brønsted-Lowry acid gives away an H⁺ ion—nominally, a proton—in an acid-base reaction.
5. NH₃ + H₂O → NH₄⁺ + OH⁻; NH₃: Brønsted-Lowry base; H₂O: Brønsted-Lowry acid
7. HNO₃ + H₂O → H₃O⁺ + NO₃⁻; HNO₃: Brønsted-Lowry acid; H₂O: Brønsted-Lowry base
9. C₁₇H₂₂NO₄⁺ + H₂O → H₃O⁺ + C₁₇H₂₁NO₄

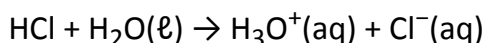
10.3 Water: Both an Acid and a Base

LEARNING OBJECTIVE

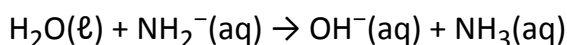
1. Write chemical equations for water acting as an acid and as a base.

Water (H₂O) is an interesting compound in many respects. Here, we will consider its ability to behave as an acid or a base.

In some circumstances, a water molecule will accept a proton and thus act as a Brønsted-Lowry base. We saw an example in the dissolving of HCl in H₂O:



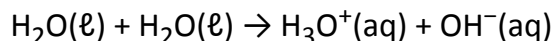
In other circumstances, a water molecule can donate a proton and thus act as a Brønsted-Lowry acid. For example, in the presence of the amide ion (see Example 4 in Section 10.2 "Brønsted-Lowry Definition of Acids and Bases"), a water molecule donates a proton, making ammonia as a product:



In this case, NH₂⁻ is a Brønsted-Lowry base (the proton acceptor).

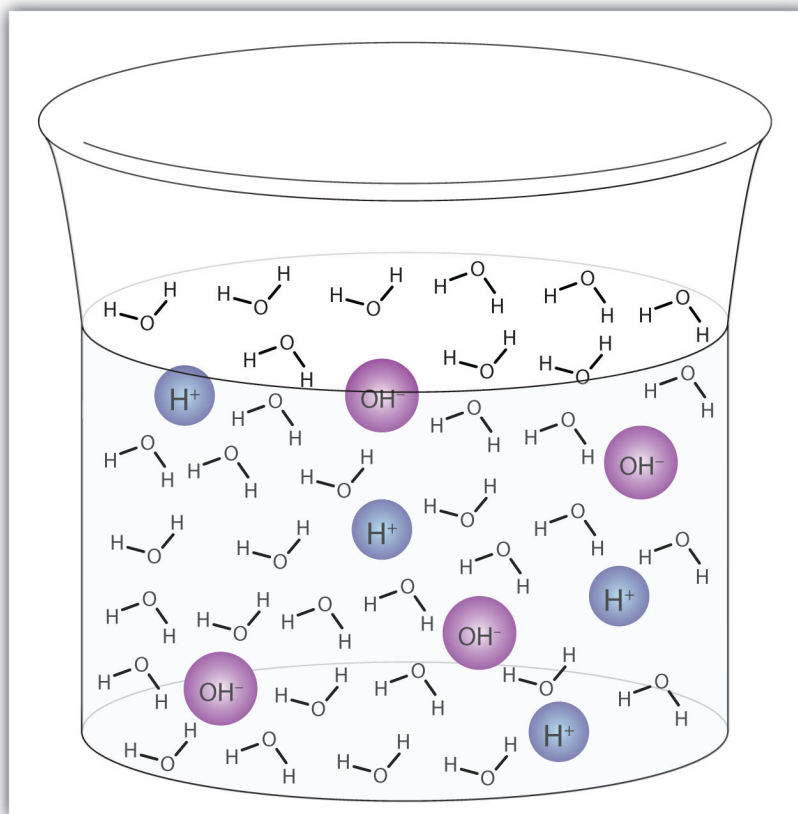
So, depending on the circumstances, H_2O can act as either a Brønsted-Lowry acid or a Brønsted-Lowry base. Water is not the only substance that can react as an acid in some cases or a base in others, but it is certainly the most common example—and the most important one. A substance that can either donate or accept a proton, depending on the circumstances, is called an **amphiprotic** compound.

A water molecule can act as an acid or a base even in a sample of pure water. About 6 in every 100 million ($6 \text{ in } 10^8$) water molecules undergo the following reaction:



This process is called the **autoionization of water** (Figure 10.2 "Autoionization") and occurs in every sample of water, whether it is pure or part of a solution. Autoionization occurs to some extent in any amphiprotic liquid. (For comparison, liquid ammonia undergoes autoionization as well, but only about 1 molecule in a million billion ($1 \text{ in } 10^{15}$) reacts with another ammonia molecule.)

Figure 10.2 *Autoionization*



A small fraction of water molecules—approximately 6 in 100 million—ionize spontaneously into hydronium ions and hydroxide ions. This picture necessarily overrepresents the amount of autoionization that really occurs in pure water.

EXAMPLE 5

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $\text{H}_2\text{O}(\ell) + \text{NO}_2^-(\text{aq}) \rightarrow \text{HNO}_2(\text{aq}) + \text{OH}^-(\text{aq})$
2. $\text{HC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$

Solution

1. In this reaction, the water molecule donates a proton to the NO_2^- ion, making $\text{OH}^-(\text{aq})$. As the proton donor, H_2O acts as a Brønsted-Lowry acid.
2. In this reaction, the water molecule accepts a proton from $\text{HC}_2\text{H}_3\text{O}_2$, becoming $\text{H}_3\text{O}^+(\text{aq})$. As the proton acceptor, H_2O is a Brønsted-Lowry base.

SKILL-BUILDING EXERCISE

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{HCOO}^-(\text{aq})$
2. $\text{H}_2\text{O}(\ell) + \text{PO}_4^{3-}(\text{aq}) \rightarrow \text{OH}^-(\text{aq}) + \text{HPO}_4^{2-}(\text{aq})$

CONCEPT REVIEW EXERCISES

1. Explain how water can act as an acid.
2. Explain how water can act as a base.

ANSWERS

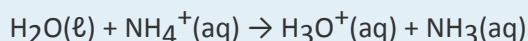
1. Under the right conditions, H₂O can donate a proton, making it a Brønsted-Lowry acid.
2. Under the right conditions, H₂O can accept a proton, making it a Brønsted-Lowry base.

KEY TAKEAWAY

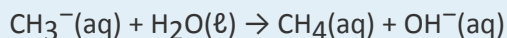
- Water molecules can act as both an acid and a base, depending on the conditions.

EXERCISES

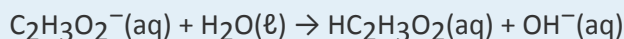
1. Is H₂O(ℓ) acting as an acid or a base?



2. Is H₂O(ℓ) acting as an acid or a base?



3. In the aqueous solutions of some salts, one of the ions from the salt can react with water molecules. In some C₂H₃O₂[−] solutions, the following reaction can occur:



Is H₂O acting as an acid or a base in this reaction?

4. In the aqueous solutions of some salts, one of the ions from the salt can react with water molecules. In some NH₄⁺ solutions, the following reaction can occur:



Is H₂O acting as an acid or a base in this reaction?

5. Aluminum hydroxide [Al(OH)₃] is *amphoteric*; it reacts with both acids and bases. Propose the chemical equations for the reactions of Al(OH)₃ with H⁺ and with OH[−].

6. Based on the information in this section, does ammonia (NH_3) autoionize more or less than water? Write the chemical equation for the autoionization of ammonia.

ANSWERS

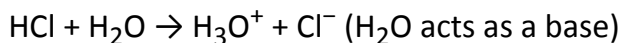
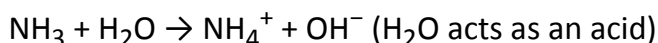
1. base
3. acid
5. $\text{Al}(\text{OH})_3 + \text{H}^+ \rightarrow \text{HAl}(\text{OH})_3^+$; $\text{Al}(\text{OH})_3 + \text{OH}^- \rightarrow \text{Al}(\text{OH})_4^-$

10.4 Autoionization of Water

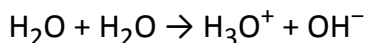
LEARNING OBJECTIVES

1. Describe the autoionization of water.
2. Calculate the concentrations of H_3O^+ and OH^- in solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid or a base:



It may not surprise you to learn, then, that within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:



This occurs only to a very small degree: only about 6 in 10^8 H_2O molecules are participating in this process, which is called the **autoionization of water**. At this level, the concentration of both H_3O^+

(aq) and OH^- (aq) in a sample of pure H_2O is about 1.0×10^{-7} M. We use square brackets [] around a dissolved species to imply the molar concentration of that species. We have

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$$

for *any* sample of pure water because H_2O can act as both an acid and a base. The product of these two concentrations is 1.0×10^{-14} :

$$[\text{H}_3\text{O}^+] \times [\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

In acids, the concentration of H_3O^+ (aq)— $[\text{H}_3\text{O}^+]$ —is greater than 1.0×10^{-7} M, while for bases the concentration of OH^- (aq)— $[\text{OH}^-]$ —is greater than 1.0×10^{-7} M. However, the *product* of the two concentrations— $[\text{H}_3\text{O}^+][\text{OH}^-]$ —is *always* equal to 1.0×10^{-14} , no matter whether the aqueous solution is an acid, a base, or neutral:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted K_w :

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

This relationship is also known as the *ion-product* of water. This means that if you know $[\text{H}_3\text{O}^+]$ for a solution, you can calculate what $[\text{OH}^-]$ has to be for the product to equal 1.0×10^{-14} , or if you know $[\text{OH}^-]$, you can calculate $[\text{H}_3\text{O}^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w .

EXAMPLE 6

What is $[\text{OH}^-]$ of an aqueous solution if $[\text{H}_3\text{O}^+]$ is 1.0×10^{-4} M?

Solution

Using the expression and known value for K_w ,

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} = (1.0 \times 10^{-4})[\text{OH}^-]$$

We solve by dividing both sides of the equation by 1.0×10^{-4} :

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10} \text{ M}$$

It is assumed that the concentration unit is molarity, so $[\text{OH}^-]$ is $1.0 \times 10^{-10} \text{ M}$.

Test Yourself

What is $[\text{H}_3\text{O}^+]$ of an aqueous solution if $[\text{OH}^-]$ is $1.0 \times 10^{-9} \text{ M}$?

Answer

$$1.0 \times 10^{-5} \text{ M}$$

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of H_3O^+ or OH^- ions in the formula unit because $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$ may not be the same as the concentration of the acid or base itself.

EXAMPLE 7

What is $[\text{H}_3\text{O}^+]$ in a 0.0044 M solution of $\text{Ca}(\text{OH})_2$?

Solution

We begin by determining $[\text{OH}^-]$. The concentration of the solute is 0.0044 M, but because $\text{Ca}(\text{OH})_2$ is a strong base, there are two OH^- ions in solution for every formula unit dissolved, so the actual $[\text{OH}^-]$ is two times this, or $2 \times 0.0044 \text{ M} = 0.0088 \text{ M}$. Now we can use the K_w expression:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+](0.0088 \text{ M})$$

Dividing both sides by 0.0088:

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{0.0088} = 1.1 \times 10^{-12} \text{ M}$$

$[\text{H}_3\text{O}^+]$ has decreased significantly in this basic solution.

Test Yourself

What is $[\text{OH}^-]$ in a 0.00032 M solution of H_2SO_4 ? (Hint: assume both H_3O^+ ions ionize.)

Answer

$$1.6 \times 10^{-11} \text{ M}$$

For strong acids and bases, $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ can be determined directly from the concentration of the acid or base itself because these ions are 100% ionized by definition. However, for weak acids and bases, this is not so. The degree, or percentage, of ionization would need to be known before we can determine $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$.

EXAMPLE 8

A 0.0788 M solution of $\text{HC}_2\text{H}_3\text{O}_2$ is 3.0% ionized into H_3O^+ ions and $\text{C}_2\text{H}_3\text{O}_2^-$ ions. What are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ for this solution?

Solution

Because the acid is only 3.0% ionized, we can determine $[\text{H}_3\text{O}^+]$ from the concentration of the acid.

Recall that 3.0% is 0.030 in decimal form:

$$[\text{H}_3\text{O}^+] = 0.030 \times 0.0788 = 0.00236 \text{ M}$$

With this $[\text{H}_3\text{O}^+]$, then $[\text{OH}^-]$ can be calculated as follows:

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{0.00236} = 4.2 \times 10^{-12} \text{ M}$$

This is about 30 times higher than would be expected for a strong acid of the same concentration.

Test Yourself

A 0.0222 M solution of pyridine ($\text{C}_5\text{H}_5\text{N}$) is 0.44% ionized into pyridinium ions ($\text{C}_5\text{H}_5\text{NH}_3^+$) and OH^- ions. What are $[\text{OH}^-]$ and $[\text{H}_3\text{O}^+]$ for this solution?

Answer

$$[\text{OH}^-] = 9.77 \times 10^{-5} \text{ M}; [\text{H}_3\text{O}^+] = 1.02 \times 10^{-10} \text{ M}$$

KEY TAKEAWAY

- In any aqueous solution, the product of $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ equals 1.0×10^{-14} .

EXERCISES

1. Does $[\text{H}_3\text{O}^+]$ remain constant in all aqueous solutions? Why or why not?
2. Does $[\text{OH}^-]$ remain constant in all aqueous solutions? Why or why not?
3. What is the relationship between $[\text{H}_3\text{O}^+]$ and K_w ? Write a mathematical expression that relates them.

4. What is the relationship between $[\text{OH}^-]$ and K_w ? Write a mathematical expression that relates them.
5. Write the chemical equation for the autoionization of water and label the conjugate acid-base pairs.
6. Write the reverse of the reaction for the autoionization of water. It is still an acid-base reaction? If so, label the acid and base.
7. For a given aqueous solution, if $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-3} \text{ M}$, what is $[\text{OH}^-]$?
8. For a given aqueous solution, if $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-9} \text{ M}$, what is $[\text{OH}^-]$?
9. For a given aqueous solution, if $[\text{H}_3\text{O}^+] = 7.92 \times 10^{-5} \text{ M}$, what is $[\text{OH}^-]$?
10. For a given aqueous solution, if $[\text{H}_3\text{O}^+] = 2.07 \times 10^{-11} \text{ M}$, what is $[\text{H}_3\text{O}^+]$?
11. For a given aqueous solution, if $[\text{OH}^-] = 1.0 \times 10^{-5} \text{ M}$, what is $[\text{H}_3\text{O}^+]$?
12. For a given aqueous solution, if $[\text{OH}^-] = 1.0 \times 10^{-12} \text{ M}$, what is $[\text{H}_3\text{O}^+]$?
13. For a given aqueous solution, if $[\text{OH}^-] = 3.77 \times 10^{-4} \text{ M}$, what is $[\text{H}_3\text{O}^+]$?
14. For a given aqueous solution, if $[\text{OH}^-] = 7.11 \times 10^{-10} \text{ M}$, what is $[\text{H}_3\text{O}^+]$?
15. What are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in a 0.344 M solution of HNO_3 ?
16. What are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in a 2.86 M solution of HBr ?
17. What are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in a 0.00338 M solution of KOH ?
18. What are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in a $6.02 \times 10^{-4} \text{ M}$ solution of $\text{Ca}(\text{OH})_2$?
19. If HNO_2 is dissociated only to an extent of 0.445%, what are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in a 0.307 M solution of HNO_2 ?
20. If $(\text{C}_2\text{H}_5)_2\text{NH}$ is dissociated only to an extent of 0.077%, what are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in a 0.0955 M solution of $(\text{C}_2\text{H}_5)_2\text{NH}$?

ANSWERS

1. $[\text{H}_3\text{O}^+]$ varies with the amount of acid or base in a solution.
3. $[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]}$
5. $\text{H}_2\text{O} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{OH}^-$; $\text{H}_2\text{O}/\text{H}_3\text{O}^+$ and $\text{H}_2\text{O}/\text{OH}^-$
7. $1.0 \times 10^{-11} \text{ M}$
9. $1.26 \times 10^{-10} \text{ M}$
11. $1.0 \times 10^{-9} \text{ M}$
13. $2.65 \times 10^{-11} \text{ M}$
15. $[\text{H}_3\text{O}^+] = 0.344 \text{ M}$; $[\text{OH}^-] = 2.91 \times 10^{-14} \text{ M}$
17. $[\text{OH}^-] = 0.00338 \text{ M}$; $[\text{H}_3\text{O}^+] = 2.96 \times 10^{-12} \text{ M}$
19. $[\text{H}_3\text{O}^+] = 0.00137 \text{ M}$; $[\text{OH}^-] = 7.32 \times 10^{-12} \text{ M}$

10.5 The Strengths of Acids and Bases

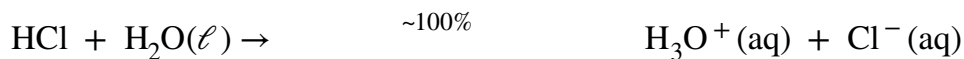
LEARNING OBJECTIVES

1. Describe the difference between strong and weak acids and bases.
2. Describe how a chemical reaction reaches chemical equilibrium.
3. Define the pH scale and use it to describe acids and bases.
4. Determine the pH of acidic and basic solutions.

Acids and bases do not all demonstrate the same degree of chemical activity in solution. Different acids and bases have different strengths.

Strong and Weak Acids

Let us consider the strengths of acids first. A small number of acids ionize completely in aqueous solution. For example, when HCl dissolves in water, every molecule of HCl separates into a hydronium ion and a chloride ion:



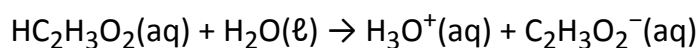
HCl(aq) is one example of a **strong acid**, which is a compound that is essentially 100% ionized in aqueous solution. There are very few strong acids. The important ones are listed in Table 10.2 "Strong Acids and Bases (All in Aqueous Solution)".

Table 10.2 Strong Acids and Bases (All in Aqueous Solution)

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO ₃	Mg(OH) ₂
H ₂ SO ₄	Ca(OH) ₂
HClO ₄	

By analogy, a **strong base** is a compound that is essentially 100% ionized in aqueous solution. As with acids, there are only a few strong bases, which are also listed in Table 10.2 "Strong Acids and Bases (All in Aqueous Solution)".

If an acid is not listed in Table 10.2 "Strong Acids and Bases (All in Aqueous Solution)", it is likely a **weak acid**, which is a compound that is not 100% ionized in aqueous solution. Similarly, a **weak base** is a compound that is not 100% ionized in aqueous solution. For example, acetic acid (HC₂H₃O₂) is a weak acid. The ionization reaction for acetic acid is as follows:



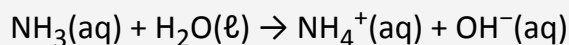
Depending on the concentration of HC₂H₃O₂, the ionization reaction may occur only for 1%–5% of the

acetic acid molecules.

Looking Closer: Household Acids and Bases

Many household products are acids or bases. For example, the owner of a swimming pool may use muriatic acid to clean the pool. Muriatic acid is another name for hydrochloric acid [HCl(aq)]. Vinegar has already been mentioned as a dilute solution of acetic acid [HC₂H₃O₂(aq)]. In a medicine chest, one may find a bottle of vitamin C tablets; the chemical name of vitamin C is ascorbic acid (HC₆H₇O₆).

One of the more familiar household bases is ammonia (NH₃), which is found in numerous cleaning products. As we mentioned previously, ammonia is a base because it increases the hydroxide ion concentration by reacting with water:



Many soaps are also slightly basic because they contain compounds that act as Brønsted-Lowry bases, accepting protons from water and forming excess hydroxide ions. This is one reason that soap solutions are slippery.

Perhaps the most dangerous household chemical is the lye-based drain cleaner. Lye is a common name for sodium hydroxide, although it is also used as a synonym for potassium hydroxide. Lye is an extremely caustic chemical that can react with grease, hair, food particles, and other substances that may build up and form a clog in a pipe. Unfortunately, lye can also attack tissues and other substances in our bodies. Thus, when we use lye-based drain cleaners, we must be very careful not to touch any of the solid drain cleaner or spill the water it was poured into. Safer, nonlye drain cleaners use peroxide compounds to react on the materials in the clog and clear the drain.



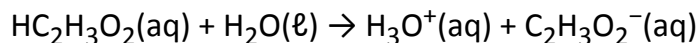
Drain cleaners can be made from a reactive material that is less caustic than a base.

Source: Photo used by permission of Citrasolv, LLC.

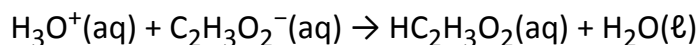
Chemical Equilibrium

The behavior of weak acids and bases illustrates a key concept in chemistry. Does the chemical reaction describing the ionization of a weak acid or base just stop when the acid or base is done ionizing?

Actually, no. Rather, the reverse process—the reformation of the molecular form of the acid or base—occurs, ultimately at the same rate as the ionization process. For example, the ionization of the weak acid $\text{HC}_2\text{H}_3\text{O}_2$ (aq) is as follows:



The reverse process also begins to occur:



Eventually, there is a balance between the two opposing processes, and no additional change occurs.

The chemical reaction is better represented at this point with a double arrow:



The \rightleftharpoons implies that both the forward and reverse reactions are occurring, and their effects cancel each other out. A process at this point is considered to be at **chemical equilibrium (or equilibrium)**. It is important to note that the processes do not stop. They balance out each other so that there is no further net change; that is, chemical equilibrium is a *dynamic equilibrium*.

EXAMPLE 9

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

1. $\text{HNO}_2(\text{aq})$
2. $\text{C}_5\text{H}_5\text{N}(\text{aq})$

Solution

1. $\text{HNO}_2(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{NO}_2^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
2. $\text{C}_5\text{H}_5\text{N}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{C}_5\text{H}_5\text{NH}^+(\text{aq}) + \text{OH}^-(\text{aq})$

SKILL-BUILDING EXERCISE

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

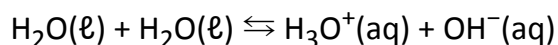
1. $\text{HF}(\text{aq})$

2. AgOH(aq)

Note

Hydrofluoric acid [HF(aq)] is one chemical that reacts directly with glass. (Very few chemicals react with glass.) Hydrofluoric acid is used in glass etching.

Finally, you may realize that the autoionization of water is actually an equilibrium process, so it is more properly written with the double arrow:

**The pH Scale**

One qualitative measure of the strength of an acid or a base solution is the **pH scale**. It is a logarithmic scale that relates the concentration of the hydronium ion in solution. A neutral (neither acidic nor basic) solution, one that has the same concentration of hydronium and hydroxide ions, has a pH of 7. A pH below 7 means that a solution is acidic, with lower values of pH corresponding to increasingly acidic solutions. A pH greater than 7 indicates a basic solution, with higher values of pH corresponding to increasingly basic solutions. Thus, given the pH of several solutions, you can state which ones are acidic, which ones are basic, and which are more acidic or basic than others. Table 10.3 "The pH Values of Some Common Solutions" lists the pH of several common solutions. Notice that some biological fluids are nowhere near neutral.

Table 10.3 The pH Values of Some Common Solutions

Solution	pH
battery acid	0.3
stomach acid	1–2
lemon or lime juice	2.1

vinegar	2.8–3.0
Coca-Cola	3
wine	2.8–3.8
beer	4–5
coffee	5
milk	6
urine	6
pure H ₂ O	7
(human) blood	7.3–7.5
sea water	8
antacid (milk of magnesia)	10.5
NH ₃ (1 M)	11.6
bleach	12.6
NaOH (1 M)	14.0

Weak acids and bases are relatively common. You may notice from Table 10.3 "The pH Values of Some Common Solutions" that many food products are slightly acidic. They are acidic because they contain solutions of weak acids. If the acid components of these foods were strong acids, the food would likely be inedible.

Calculating pH

As we have seen, [H₃O⁺] and [OH⁻] values can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

pH is a logarithmic function of [H₃O⁺]:

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on [H₃O⁺], we can

summarize as follows:

- If $\text{pH} < 7$, then the solution is acidic, and $[\text{H}_3\text{O}^+] > 10^{-7}$
- If $\text{pH} = 7$, then the solution is neutral, and $[\text{H}_3\text{O}^+] = 10^{-7}$
- If $\text{pH} > 7$, then the solution is basic, and $[\text{H}_3\text{O}^+] < 10^{-7}$

pH is a *logarithmic* scale. A solution that has a pH of 1.0 has 10 times the $[\text{H}_3\text{O}^+]$ as a solution with a pH of 2.0, which in turn has 10 times the $[\text{H}_3\text{O}^+]$ as a solution with a pH of 3.0 and so forth.

Using the definition of pH, it is also possible to calculate $[\text{H}_3\text{O}^+]$ from pH and vice versa. The general formula for determining $[\text{H}_3\text{O}^+]$ from pH is as follows:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

You need to determine how to evaluate the above expression on your calculator. Ask your instructor if you have any questions.

EXAMPLE 10

What is the pH of each solution below?

- $[\text{H}_3\text{O}^+] = 1.00 \times 10^{-5} \text{ M}$
- $[\text{H}_3\text{O}^+] = 7.50 \times 10^{-8} \text{ M}$

Solution

- $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.00 \times 10^{-5}) = -(-5.00) = 5.00$
- $\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (7.50 \times 10^{-8}) = -(-7.12) = 7.12$

Test Yourself

What is the pH of a solution that has a hydronium ion concentration of 1.11×10^{-9} M?

Answer

pH = 8.95

EXAMPLE 11

What are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ for an aqueous solution whose pH is 4.88?

Solution

We need to evaluate the expression

$$[\text{H}_3\text{O}^+] = 10^{-4.88}$$

Depending on the calculator you use, the method for solving this problem will vary. In some cases, the “-4.88” is entered and a “10^x” key is pressed; for other calculators, the sequence of keystrokes is reversed. In any case, the correct numerical answer is as follows:

$$[\text{H}_3\text{O}^+] = 1.3 \times 10^{-5} \text{ M}$$

Next, we use the *ion-product* of water relationship, $[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$. From this, $[\text{OH}^-]$ can be determined:

$$[\text{OH}^-] = \frac{1 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.7 \times 10^{-10} \text{ M}$$

Test Yourself

What are $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ for an aqueous solution whose pH is 10.36?

Answer

$[\text{H}_3\text{O}^+] = 4.4 \times 10^{-11} \text{ M}$; $[\text{OH}^-] = 2.3 \times 10^{-4} \text{ M}$

CONCEPT REVIEW EXERCISES

1. Explain the difference between a strong acid or base and a weak acid or base.
2. Explain what is occurring when a chemical reaction reaches equilibrium.
3. Define *pH*.

ANSWERS

1. A strong acid or base is 100% ionized in aqueous solution; a weak acid or base is less than 100% ionized.
2. The overall reaction progress stops because the reverse process balances out the forward process.
3. pH is a measure of the hydronium-ion concentration.

KEY TAKEAWAYS

- Acids and bases can be strong or weak depending on the extent of ionization in solution.
- Most chemical reactions reach equilibrium at which point there is no net change.
- The pH scale is used to succinctly communicate the acidity or basicity of a solution.
- pH is a logarithmic function of $[\text{H}_3\text{O}^+]$.
- $[\text{H}_3\text{O}^+]$ can be calculated directly from pH.

EXERCISES

1. Name a strong acid and a weak acid.
2. Name a strong base and a weak base.
3. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution.
 - a. HF
 - b. $\text{HC}_2\text{H}_3\text{O}_2$
 - c. HCl
 - d. HClO_4
4. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution.
 - a. H_2SO_4
 - b. HSO_4^-
 - c. HPO_4^{2-}
 - d. HNO_3
5. Is each compound a strong base or a weak base? Assume all are in aqueous solution.
 - a. NH_3
 - b. NaOH
 - c. $\text{Mg}(\text{OH})_2$
 - d. $\text{Cu}(\text{OH})_2$
6. Is each compound a strong base or a weak base? Assume all are in aqueous solution.
 - a. KOH
 - b. H_2O
 - c. $\text{Fe}(\text{OH})_2$
 - d. $\text{Fe}(\text{OH})_3$
7. Write the chemical equation for the equilibrium process for each weak acid in Exercise 3.
8. Write the chemical equation for the equilibrium process for each weak acid in Exercise 4.
9. Write the chemical equation for the equilibrium process for each weak base in Exercise 5.

10. Write the chemical equation for the equilibrium process for each weak base in Exercise 6.
11. Which is the stronger acid—HCl(aq) or HF(aq)?
12. Which is the stronger base—KOH(aq) or Ni(OH)₂(aq)?
13. Consider the two acids in Exercise 11. For solutions that have the same concentration, which one would you expect to have a lower pH?
14. Consider the two bases in Exercise 12. For solutions that have the same concentration, which one would you expect to have a higher pH?
15. Consider the list of substances in Table 10.3 "The pH Values of Some Common Solutions". What is the most acidic substance on the list that you have encountered recently?
16. Consider the list of substances in Table 10.3 "The pH Values of Some Common Solutions". What is the most basic substance on the list that you have encountered recently?
17. What is the pH of a solution when [H₃O⁺] is 3.44×10^{-4} M?
18. What is the pH of a solution when [H₃O⁺] is 5.32×10^{-11} M?
19. What is the pH of a solution when [OH⁻] is 3.44×10^{-4} M?
20. What is the pH of a solution when [OH⁻] is 5.32×10^{-11} M?
21. If a solution has a pH of 0.77, what is its [H₃O⁺] and [OH⁻]?
22. If a solution has a pH of 13.09, what is its [H₃O⁺] and [OH⁻]?

ANSWERS

1. strong acid: HCl; weak acid: HC₂H₃O₂ (answers will vary)

3. a. weak
b. weak
c. strong

d. strong

5. a. weak

b. strong

c. strong

d. weak

7. 3a: $\text{HF}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{F}^-(\text{aq})$; 3b: $\text{HC}_2\text{H}_3\text{O}_2(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$

9. 5a: $\text{NH}_3(\text{aq}) + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$; 5d: $\text{Cu}(\text{OH})_2(\text{aq}) \rightleftharpoons \text{Cu}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$

11. $\text{HCl}(\text{aq})$

13. $\text{HCl}(\text{aq})$

15. (answers will vary)

17. 3.46

19. 10.54

21. 0.17 M; 5.88×10^{-14} M

10.6 Buffers

LEARNING OBJECTIVE

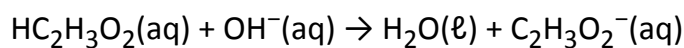
1. Define *buffer* and describe how it reacts with an acid or a base.

As indicated in Section 10.4 "The Strengths of Acids and Bases", weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have a marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1

mL of stomach acid [which we will approximate as 0.05 M HCl(aq)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9—a pH that is not conducive to continued living. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

The mechanism involves a **buffer**, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$, a weak acid) and sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$, a salt derived from that acid). Another example of a buffer is a solution containing ammonia (NH_3 , a weak base) and ammonium chloride (NH_4Cl , a salt derived from that base).

Let us use an acetic acid–sodium acetate buffer to demonstrate how buffers work. If a strong base—a source of OH^- (aq) ions—is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:



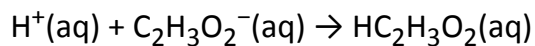
Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

Note

Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

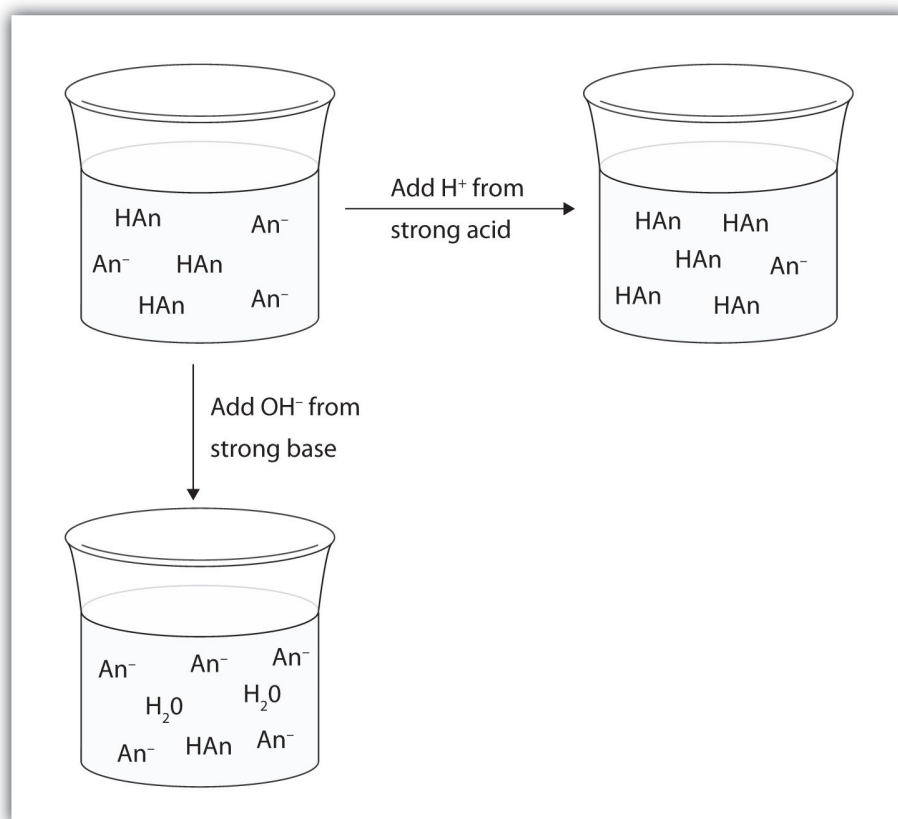
If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come

together to make acetic acid:



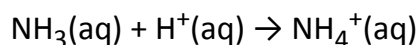
Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure 10.3 "The Action of Buffers" illustrates both actions of a buffer.

Figure 10.3 *The Action of Buffers*

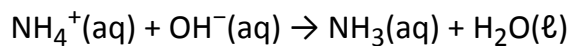


Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing NH₃ and NH₄Cl, ammonia molecules can react with any excess hydrogen ions introduced by strong acids:



while the ammonium ion $[\text{NH}_4^+(\text{aq})]$ can react with any hydroxide ions introduced by strong bases:



EXAMPLE 12

Which solute combinations can make a buffer solution? Assume all are aqueous solutions.

1. HCHO_2 and NaCHO_2
2. HCl and NaCl
3. CH_3NH_2 and $\text{CH}_3\text{NH}_3\text{Cl}$
4. NH_3 and NaOH

Solution

1. Formic acid (HCHO_2) is a weak acid, while NaCHO_2 is the salt made from the anion of the weak acid—the formate ion (CHO_2^-). The combination of these two solutes would make a buffer solution.
2. Hydrochloric acid (HCl) is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
3. Methylamine (CH_3NH_2) is like ammonia with one of its hydrogen atoms substituted with a CH_3 (methyl) group. Because it is not on our list of strong bases, we can assume that it is a weak base. The compound $\text{CH}_3\text{NH}_3\text{Cl}$ is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
4. Ammonia (NH_3) is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

SKILL-BUILDING EXERCISE

Which solute combinations can make a buffer solution? Assume all are aqueous solutions.

1. NaHCO_3 and NaCl

2. H_3PO_4 and NaH_2PO_4
3. NH_3 and $(\text{NH}_4)_3\text{PO}_4$
4. NaOH and NaCl

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain **capacity**. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [H_2CO_3 is another way to write $\text{CO}_2(\text{aq})$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

Career Focus: Blood Bank Technology Specialist

At this point in this text, you should have the idea that the chemistry of blood is fairly complex. Because of this, people who work with blood must be specially trained to work with it properly.

A blood bank technology specialist is trained to perform routine and special tests on blood samples from blood banks or transfusion centers. This specialist measures the pH of blood, types it (according to the blood's ABO+/- type, Rh factors, and other typing schemes), tests it for the presence or absence of various diseases, and uses the blood to determine if a patient has any of several medical problems, such as anemia. A blood bank technology specialist may also interview and prepare donors to give blood and may actually collect the blood donation.

Blood bank technology specialists are well trained. Typically, they require a college degree with at least a year of special training in blood biology and chemistry. In the United States, training must

conform to standards established by the American Association of Blood Banks.

CONCEPT REVIEW EXERCISE

1. Explain how a buffer prevents large changes in pH.

ANSWER

1. A buffer has components that react with both strong acids and strong bases to resist sudden changes in pH.

KEY TAKEAWAY

- A buffer is a solution that resists sudden changes in pH.

EXERCISES

1. Describe a buffer. What two related chemical components are required to make a buffer?
2. Can a buffer be made by combining a strong acid with a strong base? Why or why not?
3. Which solute combinations can make a buffer? Assume all are aqueous solutions.
 - a. HCl and NaCl
 - b. HNO_2 and NaNO_2
 - c. NH_4NO_3 and HNO_3
 - d. NH_4NO_3 and NH_3
4. Which solute combinations can make a buffer? Assume all are aqueous solutions.
 - a. H_3PO_4 and Na_3PO_4
 - b. NaHCO_3 and Na_2CO_3

- c. NaNO_3 and $\text{Ca}(\text{NO}_3)_2$
- d. HN_3 and NH_3
5. For each combination in Exercise 3 that is a buffer, write the chemical equations for the reactions of the buffer components when a strong acid and a strong base is added.
6. For each combination in Exercise 4 that is a buffer, write the chemical equations for the reaction of the buffer components when a strong acid and a strong base is added.
7. The complete phosphate buffer system is based on four substances: H_3PO_4 , H_2PO_4^- , HPO_4^{2-} , and PO_4^{3-} . What different buffer solutions can be made from these substances?
8. Explain why NaBr cannot be a component in either an acidic or a basic buffer.
9. Two solutions are made containing the same concentrations of solutes. One solution is composed of phosphoric acid and sodium phosphate, while the other is composed of hydrocyanic acid and sodium cyanide. Which solution should have the larger capacity as a buffer?
10. Two solutions are made containing the same concentrations of solutes. One solution is composed of ammonia and ammonium nitrate, while the other is composed of sulfuric acid and sodium sulfate. Which solution should have the larger capacity as a buffer?

ANSWERS

1. A buffer resists sudden changes in pH. It has a weak acid or base and a salt of that weak acid or base.
3. a. not a buffer
b. buffer
c. not a buffer
d. buffer
5. 3b: strong acid: $\text{H}^+ + \text{NO}_2^- \rightarrow \text{HNO}_2$; strong base: $\text{OH}^- + \text{HNO}_2 \rightarrow \text{H}_2\text{O} + \text{NO}_2^-$; 3d: strong acid: $\text{H}^+ + \text{NH}_3 \rightarrow \text{NH}_4^+$; strong base: $\text{OH}^- + \text{NH}_4^+ \rightarrow \text{H}_2\text{O} + \text{NH}_3$
7. Buffers can be made by combining H_3PO_4 and H_2PO_4^- , H_2PO_4^- and HPO_4^{2-} , and HPO_4^{2-} and PO_4^{3-} .

9. the phosphoric acid–phosphate buffer

10.7 Acid-Base Titrations

LEARNING OBJECTIVES

1. Describe a titration experiment.
2. Explain what an indicator does.
3. Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because so many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is because they are often used to determine quantitative amounts of one or the other. Performing chemical reactions quantitatively to determine the exact amount of a reagent is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

In a titration, one reagent has a known concentration or amount, while the other reagent has an unknown concentration or amount. Typically, the known reagent (the **titrant**) is added to the unknown quantity and is dissolved in solution. The unknown amount of substance (the **analyte**) may or may not be dissolved in solution (but usually is). The titrant is added to the analyte using a precisely calibrated volumetric delivery tube called a burette (also spelled buret; see Figure 10.4 "Equipment for Titrations"). The burette has markings to determine how much volume of solution has been added to the analyte. When the reaction is complete, it is said to be at the **equivalence point**; the number of moles of titrant can be calculated from the concentration and the volume, and the balanced chemical equation can be used to determine the number of moles (and then concentration or mass) of the unknown reactant.

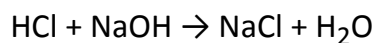
For example, suppose 25.66 mL (or 0.02566 L) of 0.1078 M HCl was used to titrate an unknown sample of NaOH. What mass of NaOH was in the

Figure 10.4

sample? We can calculate the number of moles of HCl reacted:

$$\# \text{ mol HCl} = (0.02566 \text{ L})(0.1078 \text{ M}) = 0.002766 \text{ mol HCl}$$

We also have the balanced chemical reaction between HCl and NaOH:



So we can construct a conversion factor to convert to number of moles of NaOH reacted:

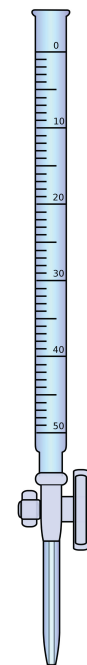
$$0.002766 \text{ mol HCl} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} = 0.002766 \text{ mol NaOH}$$

Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766 \text{ mol NaOH} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 0.1106 \text{ g NaOH}$$

This type of calculation is performed as part of a titration.

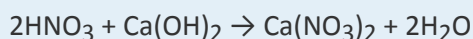
Equipment for Titrations



A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed.

EXAMPLE 13

What mass of $\text{Ca}(\text{OH})_2$ is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO_3 ? The balanced chemical equation is as follows:



Solution

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

$$\# \text{ moles HNO}_3 = (0.04402 \text{ L})(0.0885 \text{ M}) = 0.00390 \text{ mol HNO}_3$$

Using the balanced chemical equation, we can determine the number of moles of Ca(OH)_2 present in the analyte:

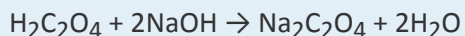
$$0.00390 \text{ mol HNO}_3 \times \frac{1 \text{ mol Ca(OH)}_2}{2 \text{ mol HNO}_3} = 0.00195 \text{ mol Ca(OH)}_2$$

Then we convert this to a mass using the molar mass of Ca(OH)_2 :

$$0.00195 \text{ mol Ca(OH)}_2 \times \frac{74.1 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 0.144 \text{ g Ca(OH)}_2$$

Test Yourself

What mass of $\text{H}_2\text{C}_2\text{O}_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is as follows:



Answer

0.182 g

How does one know if a reaction is at its equivalence point? Usually, the person performing the titration adds a small amount of an **indicator**, a substance that changes color depending on the acidity or basicity of the solution. Because different indicators change colors at different levels of acidity, choosing the correct one is important in performing an accurate titration.

For titrations involving strong acids and strong bases, the pH of the reaction solution itself can be used as an indicator. For that situation, the equivalence point occurs when the pH is equal to 7.

KEY TAKEAWAYS

- A titration is the quantitative reaction of an acid and a base.
- Indicators are used to show that all the analyte has reacted with the titrant.

EXERCISES

1. Define *titration*.
2. What is the difference between the titrant and the analyte?
3. True or false: An acid is always the titrant. Explain your answer.
4. True or false: An analyte is always dissolved before reaction. Explain your answer.
5. If 55.60 mL of 0.2221 M HCl was needed to titrate a sample of NaOH to its equivalence point, what mass of NaOH was present?
6. If 16.33 mL of 0.6664 M KOH was needed to titrate a sample of $\text{HC}_2\text{H}_3\text{O}_2$ to its equivalence point, what mass of $\text{HC}_2\text{H}_3\text{O}_2$ was present?
7. It takes 45.66 mL of 0.1126 M HBr to titrate 25.00 mL of $\text{Ca}(\text{OH})_2$ to its equivalence point. What is the original concentration of the $\text{Ca}(\text{OH})_2$ solution?
8. It takes 9.77 mL of 0.883 M H_2SO_4 to titrate 15.00 mL of KOH to its equivalence point. What is the original concentration of the KOH solution?

ANSWERS

1. a chemical reaction performed in a quantitative fashion

3. False; a base can be a titrant, or the reaction being performed may not even be an acid-base reaction.
5. 0.494 g
7. 0.1028 M

10.8 End-of-Chapter Material

Chapter Summary

To ensure that you understand the material in this chapter, you should review the meanings of the bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

The earliest chemical definition of an acid, the **Arrhenius definition**, says that an **acid** is a compound that increases the amount of hydrogen ion (H^+) in aqueous solution. An **Arrhenius base** is a compound that increases the amount of hydroxide ion (OH^-) in aqueous solution. While most bases are named as ionic hydroxide compounds, aqueous acids have a naming system unique to acids. Acids and bases react together in a characteristic chemical reaction called **neutralization**, in which the products are water and a salt. The principles of stoichiometry, along with the balanced chemical equation for a reaction between an acid and a base, can be used to determine how much of one compound will react with a given amount of the other.

A **Brønsted-Lowry acid** is any substance that donates a proton to another substance. A **Brønsted-Lowry base** is any substance that accepts a proton from another substance. The reaction of ammonia with water to make ammonium ions and hydroxide ions can be used to illustrate Brønsted-Lowry acid and base behavior.

Some compounds can either donate or accept protons, depending on the circumstances. Such compounds are called **amphiprotic**. Water is one example of an amphiprotic compound. One result of water being amphiprotic is that a water molecule can donate a proton to another water

molecule to make a hydronium ion and a hydroxide ion. This process is called the **autoionization of water** and occurs in any sample of water.

Not all acids and bases are equal in chemical strength. A **strong acid** is an acid whose molecules are all dissociated into ions in aqueous solution. Hydrochloric acid is an example of a strong acid. Similarly, a **strong base** is a base whose molecules are dissociated into ions in aqueous solution. Sodium hydroxide is an example of a strong base. Any acid or base whose molecules are not all dissociated into ions in aqueous solution is a **weak acid** or a **weak base**. Solutions of weak acids and weak bases reach a **chemical equilibrium** between the un-ionized form of the compound and the dissociated ions. It is a dynamic equilibrium because acid and base molecules are constantly dissociating into ions and reassociating into neutral molecules.

The **pH** scale is a scale used to express the concentration of hydrogen ions in solution. A neutral solution, neither acidic nor basic, has a pH of 7. Acidic solutions have a pH lower than 7, while basic solutions have a pH higher than 7.

Buffers are solutions that resist dramatic changes in pH when an acid or a base is added to them. They contain a weak acid and a salt of that weak acid, or a weak base and a salt of that weak base. When a buffer is present, any strong acid reacts with the anion of the salt, forming a weak acid and minimizing the presence of hydrogen ions in solution. Any strong base reacts with the weak acid, minimizing the amount of additional hydroxide ions in solution. However, buffers only have limited **capacity**; there is a limit to the amount of strong acid or strong base any given amount of buffer will react with.

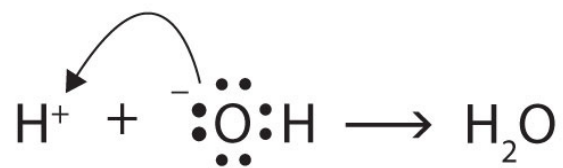
ADDITIONAL EXERCISES

1. The properties of a 1.0 M HCl solution and a 1.0 M HC₂H₃O₂ solution are compared. Measurements show that the hydrochloric acid solution has a higher osmotic pressure than the acetic acid solution. Explain why.
2. Of a 0.50 M HNO₃ solution and a 0.50 M HC₂H₃O₂ solution, which should have the higher boiling point? Explain why.

- The reaction of sulfuric acid $[\text{H}_2\text{SO}_4(\text{aq})]$ with sodium hydroxide $[\text{NaOH}(\text{aq})]$ can be represented by two separate steps, with only one hydrogen ion reacting in each step. Write the chemical equation for each step.
- The reaction of aluminum hydroxide $[\text{Al}(\text{OH})_3(\text{aq})]$ with hydrochloric acid $[\text{HCl}(\text{aq})]$ can be represented by three separate steps, with only one hydroxide ion reacting in each step. Write the chemical equation for each step.
- A friend brings you a small sample of an unknown chemical. Assuming that the chemical is soluble in water, how would you determine if the chemical is an acid or a base?
- A neutral solution has a hydrogen ion concentration of about 1×10^{-7} M. What is the concentration of the hydroxide ion in a neutral solution?
- The Lewis definitions of an acid and a base are based on electron pairs, not protons. A Lewis acid is an electron pair acceptor, while a Lewis base is an electron pair donor. Use Lewis diagrams to show that $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$ is an acid-base reaction in the Lewis sense as well as in the Arrhenius and Brønsted-Lowry senses.
- Given the chemical reaction $\text{NH}_3(\text{g}) + \text{BF}_3(\text{g}) \rightarrow \text{NH}_3\text{—BF}_3(\text{s})$ show that the reaction illustrated by this equation is an acid-base reaction if we use the Lewis definitions of an acid and a base (see Exercise 7). The product contains a bond between the N and B atoms.

ANSWERS

- HCl is a strong acid and yields more ions in solution.
- $\text{H}_2\text{SO}_4 + \text{NaOH} \rightarrow \text{NaHSO}_4 + \text{H}_2\text{O}$; $\text{NaHSO}_4 + \text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$
- One way is to add it to NaHCO_3 ; if it bubbles, it is an acid. Alternatively, add the sample to litmus and look for a characteristic color change (red for acid, blue for base).



7. The O atom is donating an electron pair to the H^+ ion, making the base an electron pair donor and the acid an electron pair acceptor.