7 Acids and Bases

Please Pass the Protons

▼

Bases have a bitter taste
And potent cleaning power;
But acids are in many foods
Like lemons, they taste sour.

There are two classes of chemical compounds that are related to each other and are quite important in our daily lives. They are *acids* and *bases*. Some familiar acids are vinegar (acetic acid), vitamin C (ascorbic acid), and battery acid (sulfuric acid). Some familiar bases are lye (sodium hydroxide), baking soda (sodium bicarbonate), and ammonia.

Air and water pollution often involve acids and bases. Acid rain, for example, is a serious environmental problem, and alkaline (basic) water is sometimes undrinkable.

Did you know that the four tastes are related to acid–base chemistry? Acids taste *sour*, bases taste *bitter*, and the compounds formed when acids react with bases (salts) taste *salty*. The sweet taste is more complicated. To taste sweet a compound must have both an acidic-type group and a basic-type group, plus just the right geometry to fit the sweet-taste receptor.

In this chapter we discuss some of the chemistry of acids and bases. You use them every day, and you will probably be hearing and reading about them as long as you live. We hope that what you learn here will help you gain a better understanding of these important classes of compounds.

SECTION 7.1

Acids and Bases: Experimental Definitions

How can you tell if a compound is an acid or a base? Let us begin by listing a few of their properties.

Acids are compounds that

1. cause litmus indicator dye to turn red.
2. taste sour.

◄ The chemical substances we call acids and bases are all around us, even in the food we eat. [*Gerald Zanetti/The Stock Market.*]

Figure 7.1 Some common acids (left) and bases (right). Acids and bases are components of many familiar consumer products. [*Robert Mathena/Fundamental Photographs.*]

3. dissolve active metals (such as zinc or iron) producing hydrogen gas.
4. react with bases to form water and ionic compounds called *salts*.

Acids and bases are chemical opposites, so their properties are quite different. **Bases** are compounds that

1. cause litmus indicator dye to turn blue.
2. taste bitter.
3. feel slippery on the skin.
4. react with acids to form water and salts.

Foods that are acidic can be identified by their sour taste. Vinegar and lemon juice are good examples. Vinegar is a solution of acetic acid (around 5%) in water. Lemons, limes, and other citrus fruits contain citric acid. Lactic acid gives yogurt its tart taste, and phosphoric acid is often added to carbonated drinks to impart tartness. The bitter taste of tonic water, on the other hand, is due to the presence of quinine, which is a base.

Although all acids have a sour taste and all bases are bitter, a taste test is hardly the best general purpose test for determining whether a substance is an acid or a base. Some acids and bases are highly poisonous, and nearly all of them are quite corrosive unless they have been greatly diluted. Some common acids and bases are pictured in Figure 7.1.

The most common way to identify a substance as an acid or a base is the litmus test. If you dip a strip of neutral (violet-colored) litmus paper into an

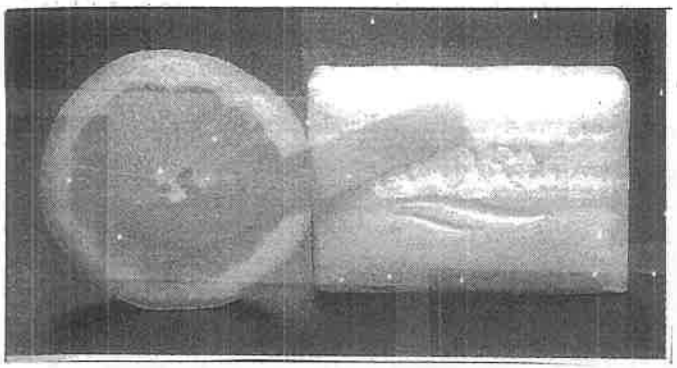


Figure 7.2 Some properties of acids and bases. The paper strip is impregnated with methyl red, an acid–base indicator. Lemon juice is acidic, as demonstrated by the red color of the indicator. The soap is basic, as shown by the fact that the red dye is changed to yellow. *[Carey B. Van Loon.]*

unknown solution and it turns pink, the solution is acidic. If it turns blue, the solution is basic. If the strip does not turn pink or blue, the solution is neither acidic or basic.

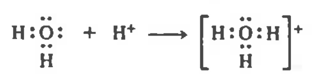
Litmus is only one of several hundred known compounds that are acid–base **indicators** (Figure 7.2). Many natural food colors, such as those in grape juice, red cabbage, and blueberries, are acid–base indicators. So are the colors in most flower petals.

SECTION 7.2

Acids Explained: Hydronium Ions

We know that acids have certain characteristic properties, but *why* do they have these properties? Experimental evidence indicates that what acids have in common are *hydrogen ions* (H^+). (Since these are hydrogen atoms from which the electrons have been removed, they are also called *protons*.)

Table 7.1 lists some common acids. Notice that each formula contains one or more hydrogen atoms. In water solution the hydrogen ions attach themselves to molecules of water, thus forming **hydronium ions** (H_3O^+).



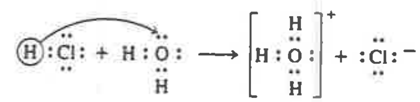
Water — Hydronium ion

Table 7.1 Some Familiar Acids

Name	Formula	Classification
Sulfuric acid	H_2SO_4	Strong
Nitric acid	HNO_3	Strong
Hydrochloric acid	HCl	Strong
Phosphoric acid	H_3PO_4	Moderate
Hydrogen sulfate ion	HSO_4^-	Moderate
Lactic acid	$CH_3CHOHCOOH$	Weak
Acetic acid	CH_3COOH	Weak
Carbonic acid	H_2CO_3	Weak
Boric acid	H_3BO_3	Very Weak
Hydrocyanic acid	HCN	Very Weak

Acids are *proton* (H^+) *donors.*

The hydrogen ions, or protons, have been donated by the acid molecules to the water molecules, so the acids are acting as **proton donors**. In the case of HCl, hydrogen chloride gas dissolves in water to form hydrochloric acid.



Acid (proton donor) — Hydrochloric acid

Notice that the HCl molecule donates a proton to the water molecule producing a hydronium ion. Other acids react in a similar way, donating hydrogen ions to water to produce hydronium ions. Even when the solvent is something other than water, the acid acts as a proton donor, transferring H^+ ions to the solvent molecules.

Nonmetal Oxides: Acid Anhydrides

Many acids are made by the reaction of nonmetal oxides with water. For example, sulfur trioxide reacts with water to form sulfuric acid.

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

Similarly, carbon dioxide reacts with water to form carbonic acid.

$$CO_2 + H_2O \longrightarrow H_2CO_3$$

This is a general reaction of nonmetal oxides.

$$\text{Nonmetal oxide} + H_2O \longrightarrow \text{Acid}$$

Nonmetal oxides are called *acid anhydrides.* **Anhydride** means "without water." These reactions explain why rainwater is acidic, particularly that which forms in air polluted with sulfur oxides (Chapter 12).

▶ EXAMPLE 7.1

Give the formula for the acid formed when sulfur dioxide reacts with water.

SOLUTION
Simply write the equation for the reaction.

$$SO_2 + H_2O \longrightarrow H_2SO_3$$

Exercise 7.1
Give the formula for the acid formed when dinitrogen pentoxide (N_2O_5) reacts with water. [*Hint: Two* molecules of acid are formed.]

SECTION 7.3

Bases Explained: Hydroxide Ions

Experimental evidence indicates that the properties of bases in water are due to OH^-, the **hydroxide ion**. Table 7.2 lists some common bases. Most of these are ionic compounds containing positive metal ions, such as Na^+ or Ca^{2+}, and negative hydroxide ions (OH^-). When the compounds dissolve in water, they all provide OH^- ions, and thus they are all bases.

Ammonia seems out of place in Table 7.2, because it contains no hydroxide ions. How can it be a base? The only way to get hydroxide ions from ammonia in water is for the ammonia molecule to accept a proton from water, leaving a hydroxide ion.

Table 7.2 Common Bases

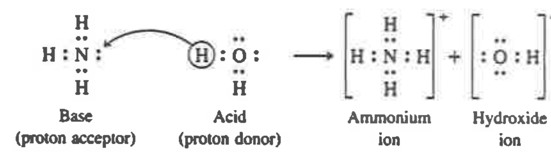
Name	Formula	Classification
Sodium hydroxide	$NaOH$	Strong
Potassium hydroxide	KOH	Strong
Lithium hydroxide	$LiOH$	Strong
Calcium hydroxide	$Ca(OH)_2$	(see text)
Magnesium hydroxide	$Mg(OH)_2$	(see text)
Ammonia	NH_3	Weak

Figure 7.3 An acid is a proton donor. A base is a proton acceptor.

$$NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$$

The NH_4^+ is the ammonium ion.

How can ammonia accept a proton from water? Recall that the nitrogen atom has an unshared pair of electrons. This pair can be used to attach a proton. Removal of a proton from water leaves a hydroxide ion, which is negatively charged because it still has the electrons that bound the "lost" hydrogen.



In water the properties of bases are those of hydroxide ions, just as the properties of acids are those of hydronium ions. When acids react with bases, then, it is really the reaction of hydronium ions with hydroxide ions. In fact, the reaction is simply a proton transfer from H_3O^+ to OH^-.

$$\underset{\text{Proton donor}}{H_3O^+} + \underset{\text{Proton acceptor}}{OH^-} \longrightarrow 2\,H_2O$$

Bases are *proton* (H^+) *acceptors.*

We can now generalize the definition of bases. Just as an acid is a proton donor, a *base* is a **proton acceptor** (Figure 7.3). This definition includes not only hydroxide ions but also neutral molecules such as ammonia. It also includes other negative ions such as carbonate (CO_3^{2-}) or bicarbonate (HCO_3^-).

$$\underset{\text{Proton donor}}{H_3O^+} + \underset{\text{Proton acceptor}}{CO_3^{2-}} \longrightarrow H_2O + HCO_3^-$$

$$\underset{\text{Proton donor}}{H_3O^+} + \underset{\text{Proton acceptor}}{HCO_3^-} \longrightarrow H_2O + H_2CO_3$$

The idea of an acid as a proton donor and a base as a proton acceptor greatly expands our concept of acids and bases.

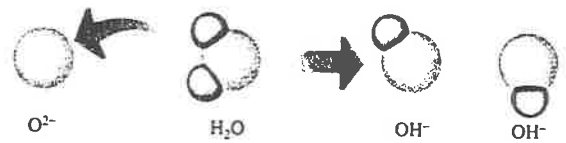


Figure 7.4 Metal oxides are basic because the oxide ion reacts with water to form two hydroxide ions.

Metal Oxides: Basic Anhydrides

Just as acids can be made from nonmetal oxides, many common bases can be made from metal oxides. For example, calcium oxide (lime) reacts with water to form calcium hydroxide (slaked lime).

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Another example is the reaction of lithium oxide with water to form lithium hydroxide.

$$Li_2O + H_2O \longrightarrow 2\,LiOH$$

In general, metal oxides react with water to form bases (Figure 7.4). These metal oxides are called *basic anhydrides.*

$$\text{Metal oxide} + H_2O \longrightarrow \text{Base}$$

▶ EXAMPLE 7.2

What base is formed by the addition of water to barium oxide (BaO)?

SOLUTION
Simply write the equation for the reaction.

$$BaO + H_2O \longrightarrow Ba(OH)_2$$

Exercise 7.2
What base is formed by the addition of water to potassium oxide (K_2O)? [*Hint:* Two moles of base are formed for each mole of potassium oxide.]

SECTION 7.4

Strong and Weak Acids and Bases

When gaseous hydrogen chloride (HCl) reacts with water, it reacts completely to give hydronium ions and chloride ions. Essentially no HCl molecules remain.

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

A strong acid (or base) is one that exists almost completely as ions in solution. The word *strong* does not refer to the amount of acid (or base) in the solution. A solution that contains a relatively large amount of acid or base in a given volume of solution is called a *concentrated* solution. (The acid or base may be strong or weak.) A solution with only a little solute in that same volume of solution is said to be *dilute.*

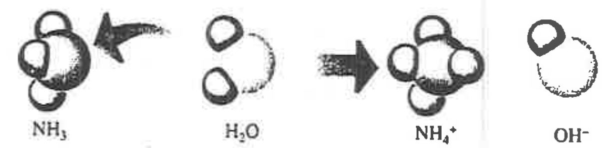


Figure 7.5 Ammonia is a base because it accepts a proton from water. A solution of ammonia in water contains ammonium ions and hydroxide ions. Only a small fraction of the ammonia molecules react, however; most remain unchanged. Ammonia is therefore a weak base.

The poisonous gas hydrogen cyanide (HCN) also reacts with water to produce hydronium ions and cyanide ions.

$$HCN + H_2O \longrightarrow H_3O^+ + CN^-$$

But in the latter case the reaction takes place only to a slight extent. Less than 1% of the HCN molecules react to produce hydronium ions. Most of them remain intact as HCN molecules. Acids like HCl that react completely with water are called **strong acids**. Those that react only slightly with water are **weak acids**. There are not many strong acids. The first three acids listed in Table 7.1 (sulfuric, nitric, and hydrochloric) are the only common ones. Most acids are weak.

Bases are also classified as strong or weak, depending on whether they react completely or only slightly with water to form hydroxide ions. Perhaps the most familiar **strong base** is sodium hydroxide (NaOH). It exists as sodium ions and hydroxide ions even in the solid state. Other strong bases include potassium hydroxide (KOH) and the hydroxides of all the other Group IA metals. Group IIA hydroxides are also strong bases [except for $Be(OH)_2$], but $Ca(OH)_2$ is only slightly soluble in water and $Mg(OH)_2$ is nearly insoluble, so these are not very practical strong bases. The most familiar **weak base** is ammonia (NH_3). It reacts with water to a slight extent to produce ammonium ions (NH_4^+) and hydroxide ions (Figure 7.5).

$$NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$$

Neutralization

When an acid reacts with a base, the products are water and a **salt**. If a solution containing hydronium ions (an acid) is mixed with another solution containing exactly the same amount of hydroxide ions (a base), the resulting solution no longer affects litmus, and it no longer tastes sour or bitter (it tastes salty). It is no longer either acidic or basic; it is neutral. The reaction of an acid with a base is called **neutralization** (Figure 7.6). In water it is simply the reaction of hydronium ions with hydroxide ions to form water molecules.

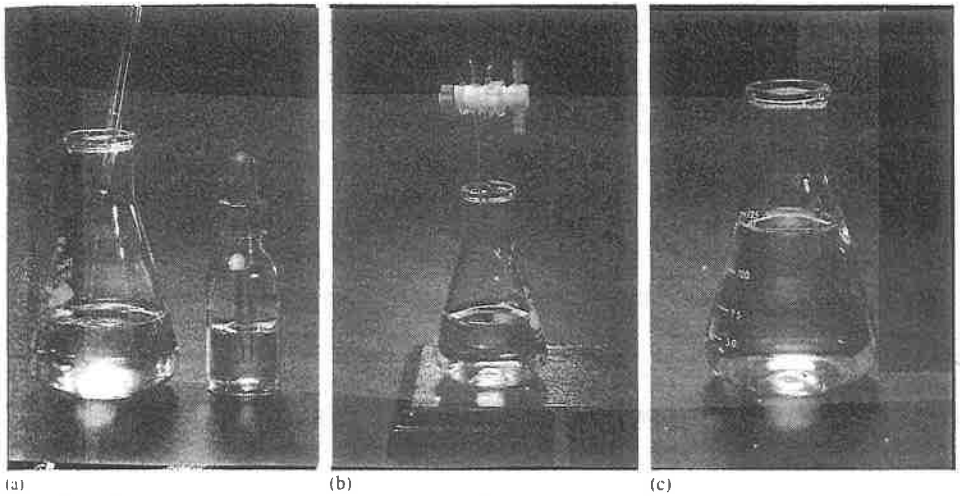


Figure 7.6 The amount of acid (or base) in a solution is determined by careful neutralization. In the experiment shown here, a 5.00-mL sample of vinegar, a small amount of water, and a few drops of phenolphthalein (an acid–base indicator) are added to a flask (a). A solution of 0.1000 M NaOH is added slowly from a buret (a device for precise measurement of volumes of solutions) (b). As long as the acid is in excess, the solution is colorless. When the acid has been neutralized and a tiny excess of base is present, the phenolphthalein indicator turns pink (c). [*Carey B. Van Loon.*]

$$H_3O^+ + OH^- \longrightarrow 2\,H_2O$$

If sodium hydroxide is being neutralized by hydrochloric acid, the products are water and sodium chloride (ordinary table salt).

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

EXAMPLE 7.3

Write the equation for the neutralization reaction between potassium hydroxide and nitric acid.

SOLUTION

$$KOH + HNO_3 \longrightarrow KNO_3 + H_2O$$

Exercise 7.3

Write the equation for the neutralization reaction between calcium hydroxide and hydrochloric acid.

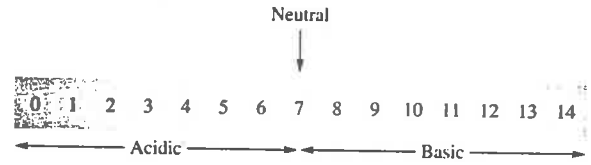


Figure 7.7 The pH scale. A change in pH of one unit means a tenfold change in the hydronium ion concentration.

SECTION 7.5

The pH Scale

Because they are so very tiny, atoms, molecules, and ions are usually counted in *moles* (see Section 6.8). In solutions the concentrations of ions are measured in moles per liter. A solution of 1 molar hydrochloric acid (1 M HCl), for example, contains 1 mole of H_3O^+ ions per liter of solution. In other words, 1 L of 1 M HCl contains 6×10^{23} hydronium ions; 0.5 L of 0.001 M HCl would contain 3×10^{20} hydronium ions.

In describing the acidity of a particular solution one might say: "The hydronium ion concentration of this solution is 1×10^{-3} mole per liter." But rarely will you hear anyone make such a statement. Instead the solution will simply be identified as "pH 3." When people talk about degree of acidity or basicity, they usually do it in terms of pH.

The pH scale is a convenient acidity scale proposed in 1909 by the Danish biochemist S. P. L. Sorensen. It mainly extends from 0 to 14. The neutral point on the scale is 7, with values below 7 becoming increasingly acidic and those above 7 increasingly basic. Thus pH 6 is only very slightly acidic, whereas pH 12 is strongly basic (Figure 7.7).

The numbers on the scale were not chosen arbitrarily. They were directly related to the concentration of hydronium ions. We might expect that pure water would be completely in the form of H_2O molecules; but it turns out that one out of every 500 million molecules is split into ions. That causes the concentration of hydronium ions in pure water to be 0.0000001 mole per liter, or 1×10^{-7} M. Can you see why 7 is the pH of pure water? It is simply the power of 10 for the molar concentration of H_3O^+, with the negative sign removed. (The H in pH stands for *hydrogen* and the p for *power*.) We can define **pH** as the negative logarithm of the molar concentration of hydronium ion (Table 7.3).

Table 7.3 Relationship Between pH and Concentration of Hydronium Ions

Concentration of H_3O^+ (mol/L)	pH
1×10^{0}	0
1×10^{-1}	1
1×10^{-2}	2
1×10^{-3}	3
1×10^{-4}	4
1×10^{-5}	5
1×10^{-6}	6
1×10^{-7}	7
1×10^{-8}	8
1×10^{-9}	9
1×10^{-10}	10
1×10^{-11}	11
1×10^{-12}	12
1×10^{-13}	13
1×10^{-14}	14

Notice that although pH is an acidity scale, the pH goes down when acidity goes up. Not only is the relationship an inverse one, but it is also logarithmic. A decrease of one pH unit represents a tenfold increase in acidity, and when pH goes down by two units, acidity increases by a factor of 100. The relationship may seem strange at first, but once you understand the pH scale, you appreciate its convenience.

Table 7.3 summarizes the relationship between hydronium ion concentration and pH. A pH of 4 means a hydronium ion concentration of 1 ×

Table 7.4 The Approximate pH of Some Common Solutions

Solution	pH
Hydrochloric acid (4%)	0
Gastric juice	1.6–1.8
Lemon juice	2.1
Vinegar (4%)	2.5
Soft drinks	2.0–4.0
Rainwater (thunderstorm)	3.5–4.2
Milk	6.3–6.6
Urine	5.5–7.0
Rainwater (unpolluted)	5.6
Saliva	6.2–7.4
Pure water	7.0
Blood	7.4
Fresh egg white	7.6–8.0
Bile	7.8–8.6
Milk of magnesia	10.5
Washing soda	12.0
Sodium hydroxide (4%)	13.0

10^{-4} mol/L, or 0.0001 M. If the concentration of hydronium ion is 0.01 M, or 1×10^{-2} M, then the pH is 2. The pH values for various common solutions are listed in Table 7.4.

▶ EXAMPLE 7.4

What is the pH of a solution that has a hydronium ion concentration of 1.0×10^{-5} M?

SOLUTION

The exponent is −5; the pH is therefore 5.

Exercise 7.4

What is the pH of a solution that has a hydronium ion concentration of 1.0×10^{-11} M?

▶ EXAMPLE 7.5

What is the hydronium ion concentration in a solution that has a pH of 4?

SOLUTION

The pH value is the negative exponent of 10, so the hydronium ion concentration is 1.0×10^{-4} M.

The pH of a solution is defined as the negative logarithm of the hydronium ion concentration,

$$pH = -\log [H_3O^+]$$

where the brackets indicate molar concentration. Perhaps the relationship is easier to see when the equation is written:

$$[H_3O^+] = 10^{-pH}$$

Normally pH is used to describe the acidity of a solution (rather than pOH), whether the solution is acidic or basic.

Drugs such as Zantac and Tagamet are used to treat people with ulcers. These drugs inhibit the release of hydrochloric acid by the stomach.

All antacids are basic compounds. They act by neutralizing hydronium ions in stomach acid.

Exercise 7.5

What is the hydronium ion concentration of a solution that has a pH of 2?

In the case of basic solutions, pOH is related to $[OH^-]$ just as pH is related to $[H_3O^+]$. In 0.01 M NaOH solution the $[OH^-] = 0.01 = 1 \times 10^{-2}$, and therefore the pOH = 2. To convert pOH to pH, simply subtract the pOH from 14.

$$pH + pOH = 14$$

$$pH = 14 - pOH$$

If the pOH = 2, then the pH = 14 − 2 = 12.

▶ EXAMPLE 7.6

What is the pH of a KOH solution that has a hydroxide ion concentration of 1×10^{-4}?

SOLUTION

The exponent is −4; therefore, the pOH is 4. Since pH = 14 − pOH, the pH = 14 − 4 = 10.

Exercise 7.6

What is the pH of a solution of ammonia that has a hydroxide ion concentration of 1×10^{-3}?

SECTION 7.6

Antacids: A Basic Remedy

The stomach secretes hydrochloric acid (HCl) as an aid in the digestion of food. Sometimes overindulgence or emotional stress leads to a condition of **hyperacidity** (too much acid). There are thousands of brands of **antacids** (Figure 7.8) sold in the United States to treat this condition. Despite the many brand names, there are only a few different antacid ingredients, primarily sodium bicarbonate, calcium carbonate, aluminum hydroxide, and magnesium hydroxide (Table 7.5.)

Sodium bicarbonate ($NaHCO_3$), commonly called baking soda, is probably safe and effective for most people, but overuse can make the blood too alkaline, a condition called **alkalosis**.

Calcium carbonate ($CaCO_3$) is safe in small amounts, but regular use can cause constipation. It also appears that calcium carbonate can actually cause *increased* acid secretion after a few hours.

pH

For coffee it's 5; for tomatoes it's 4;
While household ammonia's 11 or more.
It's 7 for water, if in a pure state,
But rainwater's 6, and seawater is 8.
It's basic at 10, quite acidic at 2,
And well above 7 when litmus turns blue.
Some find it a puzzlement. Doubtless their fog
Has something to do with that negative log!

Table 7.5 Some Common Antacids

Commercial Product	Antacid Ingredient(s)
Alka-Seltzer	$NaHCO_3$, citric acid, aspirin
Amphojel	$Al(OH)_3$
Baking soda	$NaHCO_3$
DiGel	$CaCO_3$
Maalox	$Al(OH)_3$, $Mg(OH)_2$
Milk of magnesia	$Mg(OH)_2$
Mylanta	$Al(OH)_3$, $Mg(OH)_2$
Rolaids	$AlNa(OH)_2CO_3$
Rolaids Sodium-Free*	$CaCO_3$, $Mg(OH)_2$
Tums	$CaCO_3$

*Sodium-containing antacids are not recommended for people with hypertension (high blood pressure).

Figure 7.8 A great variety of antacids are available to consumers. All antacids are basic compounds. [*Robert Mathena/Fundamental Photographs.*]

You can make your own aspirin-free "Alka-Seltzer." Simply place half a teaspoon of baking soda in a glass of orange juice. (What is the acid and what is the base in this reaction?)

Claims of "fast action" are almost meaningless. All acid–base reactions are almost instantaneous. Some tablets may dissolve a little slower than others. You can speed their action by chewing them.

Aluminum hydroxide [$Al(OH)_3$], like calcium carbonate, can cause constipation. There is also some concern that antacids containing aluminum ions can deplete the body of essential phosphate ions.

A suspension of magnesium hydroxide [$Mg(OH)_2$] in water is sold as "milk of magnesia." Magnesium carbonate ($MgCO_3$) is also used as an antacid. In small doses magnesium compounds act as antacids, but in large doses they act as laxatives.

Although antacids are generally safe for occasional use, they can interact with other medications; and anyone who has severe or repeated attacks of indigestion should consult a physician. Self-medication can sometimes be dangerous.

SECTION 7.7

Acids, Bases, and Human Health

Concentrated strong acids and bases are corrosive poisons that can cause serious chemical burns. Once the chemical agents are removed, the injuries are similar to burns from heat, and they are often treated the same way.

Sulfuric acid (H_2SO_4) is by far the leading chemical product in the United States. About 40 billion kg is produced each year, most of it for making other industrial chemicals, especially fertilizers. Around the home we use sulfuric acid in automobile batteries and in some drain cleaners. Besides being a strong acid, sulfuric acid is also a powerful dehydrating agent that can react with water in the cells. Acid rain mainly results from sulfuric acid mists produced by the burning of high sulfur coal.

Hydrochloric acid (also called muriatic acid) is used in industry to remove rust from metal, in construction to remove excess mortar from bricks, and in the home to remove lime deposits from toilet bowls. Concentrated solutions (about 38% HCl) cause severe burns, but dilute solutions can be used safely in the home if handled carefully. The gastric juice in your stomach is a solution containing around 0.5% hydrochloric acid.

Lime (CaO) is the cheapest and most widely used commercial base. It is made by heating limestone ($CaCO_3$) to drive off CO_2. It is the fifth most produced industrial chemical in the United States with an annual production of about 17 billion kg. It is used to make mortar and cement and also to "sweeten" acidic soil (Figure 7.9).

Sodium hydroxide (commonly known as lye) is the strong base most often used in the home. It is used as an oven cleaner in products such as Easy Off, and it is used to open clogged drains with products such as Drano.

Both acids and bases, even in dilute solutions, break down the protein molecules in living cells. Generally, the fragments are not able to carry out the functions of the original proteins. In cases of severe exposure, the fragmentation continues until the tissue has been completely destroyed. And, within

Figure 7.9 Treating the soil with lime makes it "sweeter" (less acid). [*Michael P. Gadomski/Photo Researchers, Inc.*]

living cells, proteins function properly only at an optimum pH. If the pH changes much in either direction, the proteins can't carry out their usual functions.

When they are misused, acids and bases can be damaging to human health. But acids and bases affect human health in more subtle—and ultimately more important—ways. A delicate balance must be maintained between acids and bases in the blood and body fluids. If the acidity of the blood changes very much, the blood loses its capacity to carry oxygen. Fortunately, the body has a complex but efficient mechanism for maintaining proper acid–base balance. (Consult reference 3 for an explanation of this mechanism.)

▸ Summary

1. *Acids* are *proton donors.* They taste sour, turn litmus red, react with active metals to form hydrogen, and react with bases to form salts and water.
2. *Bases* are *proton acceptors.* They taste bitter, turn litmus blue, feel slippery to the skin, and react with acids to form salts and water.
3. When an acid dissolves in water, water molecules pick up hydrogen ions (H^+), or protons, from the acid molecules to produce *hydronium ions* (H_3O^+).
4. When a base dissolves in water, *hydroxide ions* (OH^-) are formed.
5. Nonmetal oxides (acid *anhydrides*) react with water to form acids; metal oxides (basic *anhydrides*) react with water to form bases.

6. A *strong acid* or *strong base* when dissolved in water is almost completely in the form of ions.
7. A *weak acid* or *weak base* reacts only slightly with water to produce ions.
9. Sulfuric, hydrochloric, and nitric acids are strong. Acetic and most other acids are weak.
9. Group IA hydroxides are strong bases, and so are most of the Group IIA hydroxides. Ammonia is a weak base.
10. When an acid reacts with a base, the products are a *salt* and water, and the process is called *neutralization.*
11. The *pH* scale is an acidity scale. A pH of 7 is neutral; pH values lower than 7 are increasingly acidic; and pH values greater than 7 are increasingly basic.
12. *Antacids* are products used to neutralize excess stomach acid; but taking too much antacid can produce a condition of *alkalosis.*

▸ Key Terms

acid 7.1	hydronium ion 7.2	pH 7.5	strong base 7.4
alkalosis 7.6	hydroxide ion 7.3	proton acceptor 7.3	weak acid 7.4
anhydride 7.2	hyperacidity 7.6	proton donor 7.2	weak base 7.4
antacid 7.6	indicator 7.1	salt 7.4	
base 7.1	neutralization 7.4	strong acid 7.4	

▸ Review Questions

1. Define and illustrate the following terms.
   a. acid b. base c. salt
2. List four general properties of acidic solutions.
3. List four general properties of basic solutions.
4. What ion is responsible for the properties of acidic solutions (in water)?
5. What ion is responsible for the properties of basic solutions (in water)?
6. Give the formulas and the names of two strong acids and two weak acids.
7. Give the formulas and the names for two strong bases and one weak base.
8. Strong acids and weak acids both have properties characteristic of hydronium ions. How do strong and weak acids differ?
9. Give a general definition for an acid. Write an equation that illustrates your definition.
10. What is meant by the proton as used in acid–base chemistry? How does it differ from the proton of nuclear chemistry (Chapter 4)?
11. What is an acid anhydride?
12. What is a basic anhydride?
13. How does one brand of household ammonia differ from another?
14. What is meant by the neutralization of an acid or base?
15. Describe the taste and effect on litmus of a solution that has been neutralized.
16. Magnesium hydroxide is completely ionic, even in the solid state; yet it can be taken internally as an antacid. Explain why it does not cause injury as sodium hydroxide would.
17. Why do people take antacids?
18. Name some of the active ingredients in antacids.
19. What is alkalosis?
20. Indicate whether each of the following pH values represents an acidic, basic, or a neutral solution.
    a. 4 b. 7 c. 3.5 d. 9
21. Lime juice has a pH of about 2. Is lime juice acidic or basic?