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## **6 ACIDS, BASES, AND SALTS**

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### **6.1 THE IMPORTANCE OF ACIDS, BASES, AND SALTS**

Almost all inorganic compounds and many organic compounds can be classified as acids, bases, or salts. Some of these types of compounds were mentioned in earlier chapters and are discussed in greater detail in this chapter.

Acids, bases, and salts are vitally involved with life processes, agriculture, industry, and the environment. The most widely produced chemical is an acid, sulfuric acid. The second-ranking chemical, lime, is a base. Another base, ammonia, ranks fourth in annual chemical production. Among salts, sodium chloride is widely produced as an industrial chemical, potassium chloride is a source of essential potassium fertilizer, and sodium carbonate is used in huge quantities for glass and paper manufacture, and for water treatment.

The salt content and the acid–base balance of blood must stay within very narrow limits to keep a person healthy, or even alive. Soil with too much acid or excessive base will not support good crop growth. Too much salt in irrigation water may prevent crops from growing. This is a major agricultural problem in arid regions of the world such as the mid-East and California's Imperial Valley. The high salt content of irrigation water discharged to the Rio Grande River has been a source of dispute between the U.S. and Mexico that has been resolved to a degree by installation of a large desalination (salt removal) plant by the U.S.

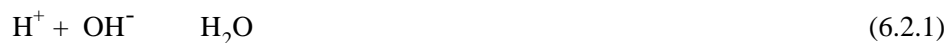
From the above discussion it is seen that acids, bases and salts are important to human health and welfare. This chapter discusses their preparation, properties, and naming.

### **6.2 THE NATURE OF ACIDS, BASES, AND SALTS**

#### **Hydrogen Ion and Hydroxide Ion**

Recall that an ion is an atom or group of atoms having an electrical charge. In discussing acids and bases 2 very important ions are involved. One of these is the

**hydrogen ion,  $H^+$** . It is always produced by acids. The other is the **hydroxide ion,  $OH^-$** . It is always produced by bases. These two ions react together,



to produce water. This is called a **neutralization reaction**. It is one of the most important of all chemical reactions.

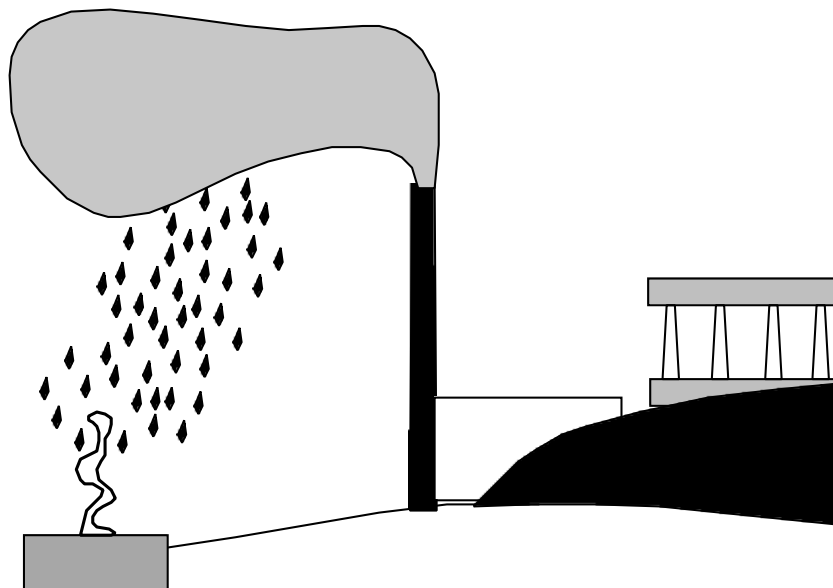
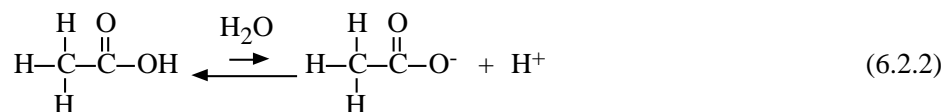


Figure 6.1 Acid rain resulting from the introduction of sulfuric, nitric, and hydrochloric acids into the atmosphere by the burning of fossil fuels damages buildings, statues, crops, and electrical equipment in some areas of the world, including parts of the northeastern U.S.

## Acids

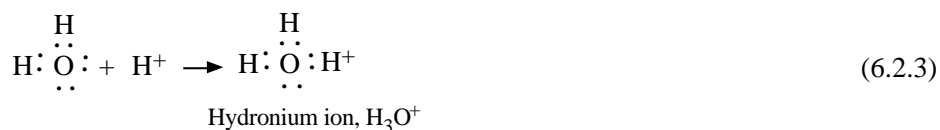
An **acid** is a substance that produces hydrogen ions. For example, HCl in water is entirely in the form of  $H^+$  ions and  $Cl^-$  ions. These 2 ions in water form hydrochloric acid. Acetic acid, which is present in vinegar, also produces hydrogen ions in water:



Acetic acid demonstrates two important characteristics of acids. First, many acids contain H that is not released by the acid molecule to form  $H^+$ . Of the 4 hydrogens in  $\text{CH}_3\text{CO}_2\text{H}$ , only the one bonded to oxygen is ionizable to form  $H^+$ . The second important point about acetic acid has to do with how much of it is ionized to form  $H^+$  and acetate ion,  $\text{CH}_3\text{CO}_2^{2-}$ . Most of the acetic acid remains as molecules of  $\text{CH}_3\text{CO}_2\text{H}$  in solution. In a 1 molar solution of acetic acid (containing 1 mol of

acetic acid per liter of solution) only about 0.5% of the acid is ionized to produce an acetate ion and a hydrogen ion. Of a thousand molecules of acetic acid, 995 remain as unionized  $\text{CH}_3\text{CO}_2\text{H}$ . Therefore, acetic acid is said to be a weak acid. This term will be discussed later in the chapter.

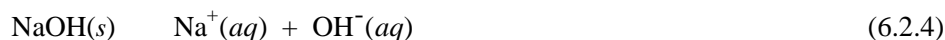
A hydrogen ion in water is strongly attracted to water molecules. Hydrogen ions react with water,



to form  $\text{H}_3\text{O}^+$  or clusters with even more water molecules such as  $\text{H}_5\text{O}_2^+$  or  $\text{H}_7\text{O}_3^+$ . The hydrogen ion in water is frequently shown as  $\text{H}_3\text{O}^+$ . In this book, however, it is simply indicated as  $\text{H}^+$ .

## Bases

A **base** is a substance that produces hydroxide ion and/or accepts  $\text{H}^+$ . Many bases consist of metal ions and hydroxide ions. For example, solid sodium hydroxide dissolves in water,



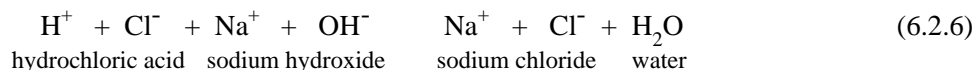
to yield a solution containing  $\text{OH}^-$  ions. When ammonia gas is bubbled into water, a few of the  $\text{NH}_3$  molecules remove hydrogen ion from water and produce ammonium ion,  $\text{NH}_4^+$ , and hydroxide ion as shown by the following reaction:



Only about 0.5% of the ammonia in a 1M solution goes to  $\text{NH}_4^+$  and  $\text{OH}^-$ . Therefore, as discussed later in the chapter,  $\text{NH}_3$  is called a **weak base**.

## Salts

Whenever an acid and a base are brought together, water is always a product. But a negative ion from the acid and a positive ion from the base are always left over as shown in the following reaction:



Sodium chloride dissolved in water is a solution of a **salt**. A salt is made up of a positively charged ion called a *cation* and a negatively charged ion called an *anion*. If the water were evaporated, the solid salt made up of cations and anions would remain as crystals. A salt is a chemical compound made up of a cation (other than  $\text{H}^+$ ) and an anion (other than  $\text{OH}^-$ ).

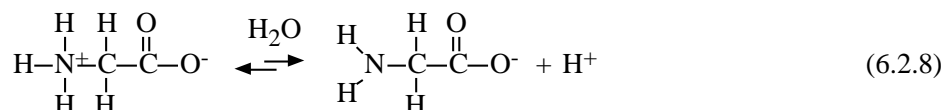
## Amphoteric Substances

Some substances, called **amphoteric substances**, can act both as an acid and a base. The simplest example is water. Water can split apart to form a hydrogen ion and a hydroxide ion.

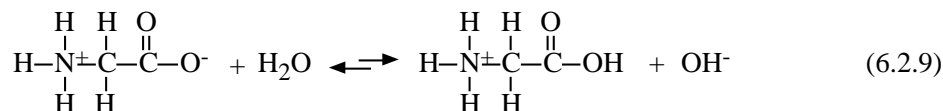


Since it produces a hydrogen ion, water is an acid. However, the fact that it produces a hydroxide ion also makes it a base. This reaction occurs only to a very small extent. In pure water only one out of 10 million molecules of water is in the form of  $\text{H}^+$  and  $\text{OH}^-$ . Except for this very low concentration of these two ions that can exist together,  $\text{H}^+$  and  $\text{OH}^-$  react strongly with each other to form water.

Another important substance that can be either an acid or base is glycine. Glycine is one of the amino acids that is an essential component of the body's protein. It can give off a hydrogen ion



or it can react with water to release a hydroxide ion from the water:

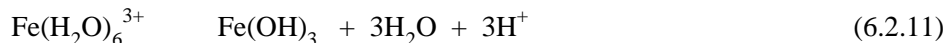


## Metal Ions as Acids

Some metal ions are acids. As an example, consider iron(III) ion,  $\text{Fe}^{3+}$ . This ion used to be commonly called ferric ion. When iron(III) chloride,  $\text{FeCl}_3$ , is dissolved in water,



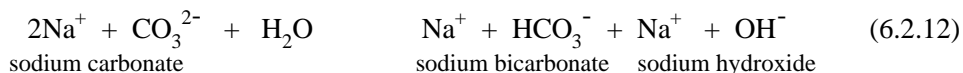
it produces chloride ions and triply charged iron(III) ions. Each iron(III) ion is bonded to 6 water molecules. The iron(III) ion surrounded by water is called a **hydrated ion**. This hydrated iron(III) ion can lose hydrogen ions and form a slimy brown precipitate of iron(III) hydroxide,  $\text{Fe}(\text{OH})_3$ :



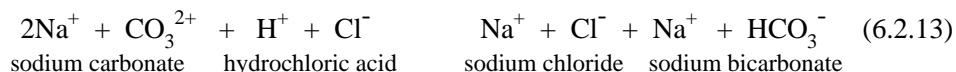
It is this reaction that is partly responsible for the acid in iron-rich acid mine water. It is also used to purify drinking water. The gelatinous  $\text{Fe}(\text{OH})_3$  settles out, carrying the impurities to the bottom of the container, and the water clears up.

## Salts that Act as Bases

Some salts that do not contain hydroxide ion produce this ion in solution. The most widely used of these is sodium carbonate,  $\text{Na}_2\text{CO}_3$ , which is commonly known as soda ash. Millions of pounds of soda ash are produced each year for the removal of hardness from boiler water, for the treatment of waste acid, and for many other industrial processes. Sodium carbonate reacts in water



to produce hydroxide ion. If  $\text{H}^+$ , such as from hydrochloric acid, is already present in the water, sodium carbonate reacts with it as follows:



## Salts that Act as Acids

Some salts act as acids. Salts that act as acids react with hydroxide ions. Ammonium chloride,  $\text{NH}_4\text{Cl}$ , is such a salt. This salt is also called “sal ammoniac.” As a “flux” added to solder used to solder copper plumbing or automobile radiators, ammonium chloride dissolves coatings of corrosion on the metal surfaces so that the solder can stick. In the presence of a base,  $\text{NH}_4\text{Cl}$  reacts with the hydroxide ion



to produce ammonia gas and water.

## 6.3 CONDUCTANCE OF ELECTRICITY BY ACIDS, BASES, AND SALTS IN SOLUTION

When acids, bases, or salts are dissolved in water, charged ions are formed. When HCl gas is dissolved in water,



all of it goes to  $\text{H}^+$  and  $\text{Cl}^-$  ions. Acetic acid in water also forms a few ions,



but most of it stays as  $\text{CH}_3\text{CO}_2\text{H}$ . Sodium hydroxide in water is all in the form of  $\text{Na}^+$  and  $\text{OH}^-$  ions. The salt,  $\text{NaCl}$ , is all present as  $\text{Na}^+$  and  $\text{Cl}^-$  ions in water.

One of the most important properties of ions is that they conduct electricity in water. Water containing ions from an acid, base, or salt will conduct electricity much like a metal wire. Consider what would happen if very pure distilled water were

made part of an electrical circuit as shown in Figure 6.2. The light bulb would not glow at all. This is because pure water does not conduct electricity. However, if a solution of salt water, such as oil well brine, is substituted for the distilled water, the bulb will glow brightly, as shown in Figure 6.2. Salty water conducts electricity because of the ions that it contains. Even tap water has some ions dissolved in it, which is why one may experience a painful, even fatal, electric shock by touching an electrical fixture while bathing.

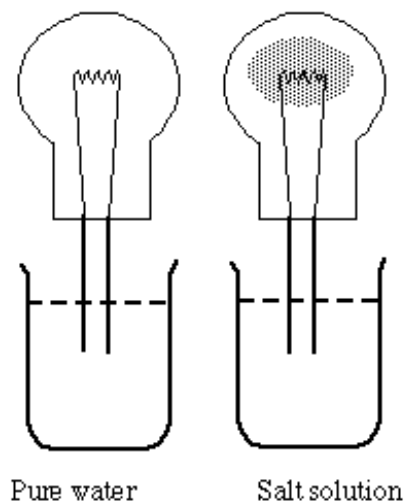


Figure 6.2 Pure water does not conduct electricity, whereas water containing dissolved salt conducts electricity very well.

## Electrolytes

Materials that conduct electricity in water are called **electrolytes**. These materials form ions in water. The charged ions allow the electrical current to flow through the water. Materials, such as sugar, that do not form ions in water are called nonelectrolytes. Solutions of nonelectrolytes in water do not conduct electricity. A solution of brine conducts electricity very well because it contains dissolved NaCl. All of the NaCl in the water is in the form of  $\text{Na}^+$  and  $\text{Cl}^-$ . The NaCl is completely ionized, and it is a strong electrolyte. An ammonia water solution (used for washing windows) does not conduct electricity very well. That is because only a small fraction of the  $\text{NH}_3$  molecules react,



to form the ions that let electricity pass through the water. Ammonia is a weak electrolyte. (Recall that it is also a weak base.) Nitric acid,  $\text{HNO}_3$  is a strong electrolyte because it is completely ionized to  $\text{H}^+$  and  $\text{NO}_3^-$  ions. Acetic acid is a weak electrolyte, as well as a weak acid. The base, sodium hydroxide, is a strong electrolyte. All salts are strong electrolytes because they are always completely ionized in water. Acids and bases can be weak or strong electrolytes.

In the laboratory, the strength of an electrolyte can be measured by how well it conducts electricity in solution, as shown in Figure 6.3. The ability of a solution to conduct electrical current is called its **conductivity**.

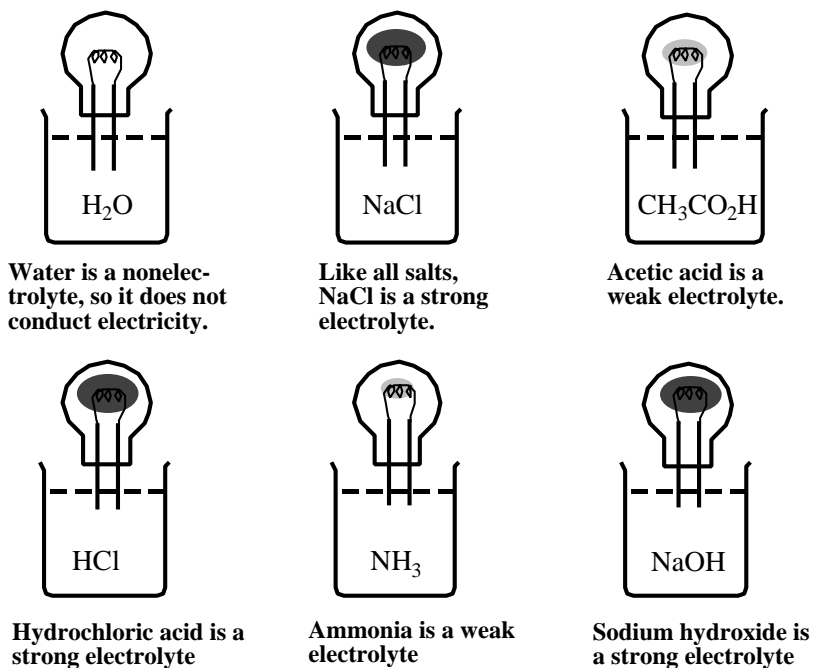


Figure 6.3 The electrical conductivity of a solution can be determined by placing the solution in an electrical circuit and observing how well electricity is conducted by the solution. **Strong electrolytes** conduct electricity well; **weak electrolytes** conduct it poorly. This principle is used in water analysis to determine the total salt concentrations in water.

When electricity is passed through solutions of acids, bases, or salts, chemical reactions occur. One such reaction is the breakdown of water to hydrogen and oxygen. Electricity passing through a solution is widely used to separate and purify various substances.

#### 6.4. DISSOCIATION OF ACIDS AND BASES IN WATER

It has already been seen that acids and bases come apart in water to form ions. When acetic acid splits up in water,



it forms hydrogen ions and acetate ions. The process of forming ions is called **ionization**. Another term is commonly employed. When the acetic acid molecule comes apart, it is said to **dissociate**. The process is called **dissociation**.

There is a great difference in how much various acids and bases dissociate. Some, like HCl or NaOH, are completely dissociated in water. Because of this, hydrochloric acid is called a **strong acid**. Sodium hydroxide is a **strong base**. Some acids such as acetic acid are only partly dissociated in water. They are called weak



acids. Ammonia,  $\text{NH}_3$ , reacts only a little bit in water to form an ammonium ion ( $\text{NH}_4^+$ ) and a hydroxide ion ( $\text{OH}^-$ ). It is a weak base.

The extent of dissociation is a very important property of an acid or base. The 3% or so acetic acid solution used to make up oil and vinegar salad dressing lends a pleasant taste to the lettuce and tomatoes. There is not much of the  $\text{H}^+$  ion in the acetic acid. If 3% HCl had been used instead, nobody could eat the salad. All of the H in HCl is in the form of  $\text{H}^+$ , and a 3% solution of hydrochloric acid is very sour indeed. Similarly, a several percent solution of  $\text{NH}_3$  in water makes a good window-washing agent, helping to dissolve grease and grime on the window surface. If a similar concentration of sodium hydroxide were used to clean windows, they would soon become permanently fogged because the  $\text{OH}^-$  in the strong base eventually reacts with glass and etches it. However, sodium hydroxide solutions are used to clean ovens, where a very strong base is required to break down the charred, baked-on grease.

Table 6.1 shows some acids and the degree to which they are dissociated. It allows comparison of the strengths of these acids.

**Table 6.1 Dissociation of Acids**

Acid formula	Acid name	Common uses	Percent Dissociated in 1 M solution	Strength
$\text{H}_2\text{SO}_4$	Sulfuric	Industrial chemical	100	Strong
$\text{HNO}_3$	Nitric	Industrial chemical	100	Strong
$\text{H}_3\text{PO}_4$	Phosphoric	Fertilizer, food additive	8	Moderately weak
$\text{H}_3\text{C}_6\text{H}_5\text{O}_7$	Citric	Fruit drinks	3	Weak
$\text{CH}_3\text{CO}_2\text{H}$	Acetic	Foods, industry	0.4	Weak
HClO	Hypochlorous	Disinfectant	0.02	Weak
HCN	Hydrocyanic	Very poisonous industrial chemical, electroplating waste	0.002	Very weak
$\text{H}_3\text{BO}_4$	Boric acid	Antiseptic, ceramics	0.002	Very weak

The percentage of acid molecules that are dissociated depends upon the concentration of the acid. The lower the concentration, the higher the percentage of dissociated molecules. This may be understood by looking again at the reaction,



for the dissociation of acetic acid. At high concentrations, there will be more crowding together of  $\text{H}^+$  and  $\text{CH}_3\text{CO}_2^-$  ions. This forces them back together to form  $\text{CH}_3\text{CO}_2\text{H}$  again. At low concentrations, there are fewer  $\text{H}^+$  and  $\text{CH}_3\text{CO}_2^-$  ions. They are more free to roam around the solution alone, and there is less pressure for them

to form  $\text{CH}_3\text{CO}_2\text{H}$ . It is somewhat like the seating which occurs on a bus. If there are few passengers, they will spread out and not sit next to each other, that is, they will be dissociated. If there are many passengers, they will, of course, have to occupy adjacent seats.

An idea of the effect of concentration upon the dissociation of a weak acid can be obtained from the percentage of acid molecules that have dissociated to ions at several different concentrations. This is shown for acetic acid in [Table 6.2](#).

**Table 6.2 Percent Dissociation of Acetic Acid at Various Concentrations**

Total acetic acid concentration	Percent dissociated to $\text{H}^+$ and $\text{CH}_3\text{CO}_2^-$
1 mol/liter	0.4
0.1 mol/liter	1.3
0.01 mol/liter	4.1
$1 \times 10^{-3}$ mol/liter	12
$1 \times 10^{-4}$ mol/liter	34
$1 \times 10^{-5}$ mol/liter	71
$1 \times 10^{-6}$ mol/liter	95
$1 \times 10^{-7}$ mol/liter	99

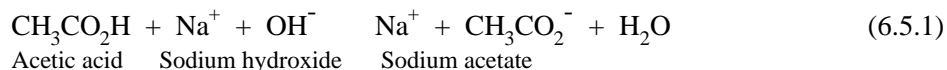
[Table 6.2](#) shows that, in a 1 M solution, less than 1% of acetic acid is dissociated. In a one-thousandth M (0.001 M) solution, 12 out of 100 molecules of acetic acid are in the form of  $\text{H}^+$  and acetate ions. In a one-millionth M (0.000001 M) solution only 5 out of 100 acetic acid molecules are present as  $\text{CH}_3\text{CO}_2\text{H}$ .

It is important to know the difference between the strength of an acid or base in solution and the concentration of the solution. A strong acid is one that is all in the form of  $\text{H}^+$  ions and anions. It may be very concentrated or very dilute. A weak acid does not give off much  $\text{H}^+$  to water solution. It may also range in concentration from a very dilute solution to a very concentrated one. Similar arguments apply to bases.

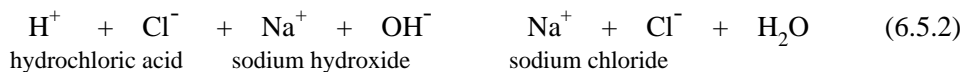
## 6.5 THE HYDROGEN ION CONCENTRATION AND BUFFERS

It is important to make the distinction between the concentration of  $\text{H}^+$  and the concentration of an acid. To show this difference, compare 1 M solutions of acetic acid and hydrochloric acid. The concentration of  $\text{H}^+$  in a 1 M solution of  $\text{CH}_3\text{CO}_2\text{H}$  is only 0.0042 mole/liter. The concentration of  $\text{H}^+$  in a 1 M solution of HCl is 1 mole/liter. A liter of a 1 M solution of HCl contains 240 as many  $\text{H}^+$  ions as a liter of a 1 M solution of acetic acid.

Consider, however, the amount of NaOH that will react with 1.00 liter of 1.00 M acetic acid. The reaction is



Exactly 1.00 mole of NaOH reacts with the 1.00 mole of acetic acid contained in a liter of a 1.00 M solution of this acid. Exactly the same amount of NaOH reacts with the HCl in 1.00 liter of 1.00 M HCl.



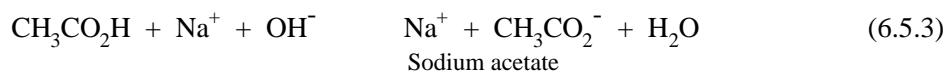
Therefore, even though acetic acid is a weaker acid than hydrochloric acid, equal volumes of each, with the same molar concentration, will react with the same number of moles of base.

In many systems the concentration of  $\text{H}^+$  is very important. For a person to remain healthy the  $\text{H}^+$  concentration in blood must stay within a very narrow range. If the  $\text{H}^+$  concentration is too high in a boiler system, the pipes may become corroded through in a short time. If the  $\text{H}^+$  concentration becomes too high or too low in a lake, plant and animal life cannot thrive in it.

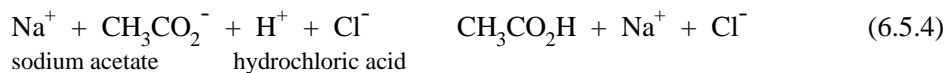
## Buffers

Fortunately, there are mixtures of chemicals that keep the  $\text{H}^+$  concentration of a solution relatively constant. Reasonable quantities of acid or base added to such solutions do not cause large changes in  $\text{H}^+$  concentration. Solutions that resist changes in  $\text{H}^+$  concentration are called **buffers**.

To understand how a buffer works, consider a typical buffer system. A solution containing both acetic acid and sodium acetate is a good buffer. The acetic acid in the solution is present as undissociated  $\text{CH}_3\text{CO}_2\text{H}$ . The  $\text{H}^+$ , which is in solution, is there because a very small amount of the  $\text{CH}_3\text{CO}_2\text{H}$  has dissociated to  $\text{H}^+$  and  $\text{CH}_3\text{CO}_2^-$  ions. The sodium acetate is present as  $\text{Na}^+$  ion and  $\text{CH}_3\text{CO}_2^-$  ion. If some base, such as NaOH, is added, some of the acetic acid reacts.



This reaction changes some of the acetic acid to sodium acetate, but it does not change the hydrogen ion concentration much. If a small amount of hydrochloric acid is added to the buffer mixture of acetic acid and sodium acetate, some of the sodium acetate is changed to acetic acid.



The acetate ion acts like a sponge for  $\text{H}^+$  and prevents the concentration of the added hydrogen ion from becoming too high.

Buffers can also be made from a mixture of a weak base and a salt of the base. A mixture of  $\text{NH}_3$  and  $\text{NH}_4\text{Cl}$  is such a buffer. Mixtures of two salts can be buffers. A mixture of  $\text{NaH}_2\text{PO}_4$  and  $\text{Na}_2\text{HPO}_4$  is a buffer made from salts. It is one of the very common phosphate buffers, such as those that occur in body fluids.

## 6.6 pH AND THE RELATIONSHIP BETWEEN HYDROGEN ION AND HYDROXIDE ION CONCENTRATIONS

Because of the fact that water itself produces both hydrogen ion and hydroxide ion



there is always some  $\text{H}^+$  and some  $\text{OH}^-$  in any solution. Of course, in an acid solution, the concentration of  $\text{OH}^-$  must be very low. In a solution of base the concentration of  $\text{OH}^-$  is very high and that of  $\text{H}^+$  is very low. There is a definite relationship between the concentration of  $\text{H}^+$  and the concentration of  $\text{OH}^-$ . It varies a little with temperature. At  $25^\circ\text{C}$  (about room temperature) the following relationship applies:

$$[\text{H}^+][\text{OH}^-] = 1.00 \times 10^{-14} = K_w \quad (\text{at } 25^\circ\text{C}) \quad (6.6.2)$$

If the value of either  $[\text{H}^+]$  or  $[\text{OH}^-]$  is known, the value of the other can be calculated by substituting into the  $K_w$  expression. For example, in a solution of 0.100 M HCl in which  $[\text{H}^+] = 0.100 \text{ M}$ ,

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1.00 \times 10^{-14}}{0.100} = 1.00 \times 10^{-13} \text{ M} \quad (6.6.3)$$

**Acids**, such as HCl and  $\text{H}_2\text{SO}_4$ , produce  $\text{H}^+$  ion, whereas bases, such as sodium hydroxide and calcium hydroxide ( $\text{NaOH}$  and  $\text{Ca}(\text{OH})_2$ , respectively), produce hydroxide ion,  $\text{OH}^-$ . Molar concentrations of hydrogen ion,  $[\text{H}^+]$ , range over many orders of magnitude and are conveniently expressed by pH defined as

$$\text{pH} = -\log[\text{H}^+] \quad (6.6.4)$$

In absolutely pure water the value of  $[\text{H}^+]$  is exactly  $1 \times 10^{-7}$  mole/L; therefore, the pH of pure water is 7.00, and the solution is **neutral** (neither acidic nor basic). **Acidic** solutions have pH values of less than 7 and **basic** solutions have pH values of greater than 7. [Table 6.3](#) gives some example hydrogen ion concentrations and the corresponding pH values.

As seen in [Table 6.3](#), when the  $\text{H}^+$  ion concentration is 1 times 10 to a power (the superscript number, such as -2, -7, etc.) the pH is simply the negative value of that power. Thus, when  $[\text{H}^+]$  is  $1 \times 10^{-3}$ , the pH is 3; when  $[\text{H}^+]$  is  $1 \times 10^{-4}$ , the pH is 4. That is because the log of  $1 \times 10^{-3}$  is -3 and that of  $1 \times 10^{-4}$  is -4. Therefore, the negative logs are 3 and 4, respectively, because the sign is reversed. What about the pH of a solution with a hydrogen ion concentration between  $1 \times 10^{-4}$  and  $1 \times 10^{-3}$ , such as  $3.16 \times 10^{-4}$ ? Since  $[\text{H}^+]$  is between  $1 \times 10^{-4}$  and  $1 \times 10^{-3}$  M, the pH is obviously going to be between 3 and 4. The pH is calculated very easily on an electronic calculator by entering  $3.16 \times 10^{-4}$  on the keyboard and pressing the “log” button. The log of the number is -3.50, and the pH is 3.50.

**Table 6.3 Values of [H<sup>+</sup>] and Corresponding pH Values**

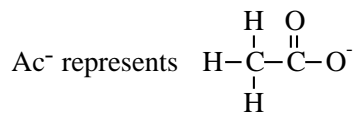
[H <sup>+</sup> ], mol/L	pH
1.00	0.00
0.100	1.00
1.00 × 10 <sup>-3</sup>	3.00
2.25 × 10 <sup>-6</sup> (10 <sup>-5.65</sup> )	5.65
1.00 × 10 <sup>-7</sup>	7.00
1.00 × 10 <sup>-9</sup>	9.00
5.17 × 10 <sup>-9</sup> (10 <sup>-8.29</sup> )	8.29
1.00 × 10 <sup>-13</sup>	13.00
1.00 × 10 <sup>-14</sup>	14.00
1.00 × 10 <sup>-2</sup>	2.00

### Acid-Base Equilibria

Many of the phenomena in aquatic chemistry and geochemistry involve **solution equilibrium**. In a general sense, solution equilibrium deals with the extent to which **reversible** acid-base, solubilization (precipitation), complexation, or oxidation-reduction reactions proceed in a forward or backward direction. This is expressed for a generalized equilibrium reaction



There are several major kinds of equilibria in aqueous solution. The one under consideration here is acid-base equilibrium as exemplified by the ionization of acetic acid, HAc,



for which the acid dissociation constant is

$$\frac{[\text{H}^+][\text{Ac}^-]}{[\text{HAc}]} = K = 1.75 \times 10^{-5} \quad (\text{at } 25^\circ\text{C}) \quad (6.6.7)$$

As an example of an acid-base equilibrium problem, consider water in equilibrium with atmospheric carbon dioxide. The value of [CO<sub>2</sub>(aq)] in water at 25°C in equilibrium with air that is 350 parts per million CO<sub>2</sub> (close to the concentration of this gas in the atmosphere) is 1.146 × 10<sup>-5</sup> moles/liter (M). The carbon dioxide dissociates partially in water to produce equal concentrations of H<sup>+</sup> and HCO<sub>3</sub><sup>-</sup>:



so that:

$$[\text{H}^+] = [\text{HCO}_3^-] \quad (6.6.9)$$

The concentrations of  $\text{H}^+$  and  $\text{HCO}_3^-$  are calculated from  $K_{a1}$ :

$$K_{a1} = \frac{[\text{H}^+][\text{HCO}_3^-]}{[\text{CO}_2]} = \frac{[\text{H}^+]^2}{1.146 \times 10^{-5}} = 4.45 \times 10^{-7} \quad (6.6.10)$$

$$[\text{H}^+] = \frac{K_{a1}[\text{CO}_2]}{[\text{HCO}_3^-]} = \frac{[\text{H}^+]^2}{1.146 \times 10^{-5}} = 4.45 \times 10^{-7} \quad (6.6.11)$$

Since  $[\text{H}^+] = [\text{HCO}_3^-]$ , this relationship simplifies to

$$[\text{H}^+] = [\text{HCO}_3^-] = (1.146 \times 10^{-5} \times 4.45 \times 10^{-7})^{1/2} = 2.25 \times 10^{-6} \quad (6.6.12)$$

$$\text{pH} = 5.65$$

This calculation explains why pure water that has equilibrated with the unpolluted atmosphere is slightly acidic, with a pH somewhat less than 7.

## 6.7 PREPARATION OF ACIDS

Acids can be prepared in several ways. In discussing their preparation, it is important to keep in mind that acids usually contain nonmetals. All acids either contain ionizable hydrogen or produce it when dissolved in water. Furthermore, the hydrogen has to be ionizable; it must have the ability to form  $\text{H}^+$  ion. Finally, more often than not, acids contain oxygen.

A simple way to make an acid is to react hydrogen with a nonmetal that forms a compound with hydrogen that will form  $\text{H}^+$  ion in water. Hydrochloric acid can be made by reacting hydrogen and chlorine



and adding the hydrogen chloride product to water. Other acids that consist of hydrogen combined with a nonmetal are HF, HBr, HI, and  $\text{H}_2\text{S}$ . Hydrocyanic acid, HCN, is an "honorary member" of this family of acids, even though it contains three elements.

Sometimes a nonmetal reacts directly with water to produce acids. The best example of this is the reaction of chlorine with water



to produce hydrochloric acid and hypochlorous acid.

Many very important acids are produced when nonmetal oxides react with water. One of the best examples is the reaction of sulfur trioxide with water

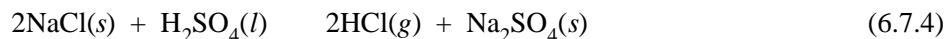


to produce sulfuric acid. Other examples are shown in [Table 6.4](#).

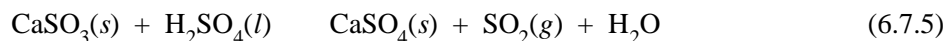
**Table 6.4 Important Acids Produced when Nonmetal Oxides React with Water**

Oxide reacted	Acid formula	Acid Name	Use and Significance of Acid
$\text{SO}_3$	$\text{H}_2\text{SO}_4$	Sulfuric	Major industrial chemical, constituent of acid rain
$\text{SO}_2$	$\text{H}_2\text{SO}_3$	Sulfurous	Paper making, scrubbed from stack gas containing $\text{SO}_2$
$\text{N}_2\text{O}_5$	$\text{HNO}_3$	Nitric	Synthesis of chemicals, constituent of acid rain
$\text{N}_2\text{O}_3$	$\text{HNO}_2$	Nitrous	Unstable, toxic to ingest, few uses
$\text{P}_4\text{O}_{10}$	$\text{H}_3\text{PO}_4$	Phosphoric	Fertilizer, chemical synthesis

Volatile acids—those that evaporate easily—can be made from salts and nonvolatile acids. The most common nonvolatile acid so used is sulfuric acid,  $\text{H}_2\text{SO}_4$ . When solid  $\text{NaCl}$  is heated in contact with concentrated sulfuric acid,

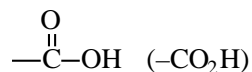


$\text{HCl}$  gas is given off. This gas can be collected in water to make hydrochloric acid. Similarly when calcium sulfite is heated with sulfuric acid,



sulfur dioxide is given off as a gas. It can be collected in water to produce sulfurous acid,  $\text{H}_2\text{SO}_3$ .

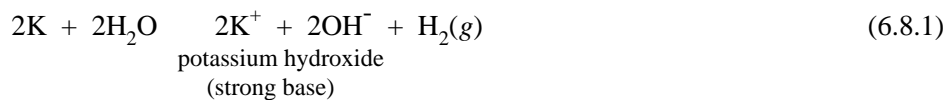
Organic acids, such as acetic acid,  $\text{CH}_3\text{CO}_2\text{H}$ , have the group



attached to a hydrocarbon group. These **carboxylic acids** are discussed further in Chapter 10.

## 6.8 PREPARATION OF BASES

Bases can be prepared in several ways. Many bases contain metals and some metals react directly with water to produce a solution of base. Lithium, sodium, and potassium react very vigorously with water to produce their hydroxides:



Many metal oxides form bases when they are dissolved in water. When waste liquor (a concentrated solution of salts and materials extracted from wood) from the sulfite paper-making process is burned to produce energy and reclaim magnesium hydroxide, the magnesium in the ash is recovered as MgO. This is added to water



to produce the magnesium hydroxide used with other chemicals to break down the wood and produce paper fibers. Other important bases and the metal oxides from which they are prepared are given in [Table 6.5](#).

**Table 6.5 Important Bases Produced when Metal Oxides React with Water**

Oxide reacted	Base formula	Base Name	Use and Significance of Base
Li <sub>2</sub> O	LiOH	Lithium hydroxide	Constituent of some lubricating greases
Na <sub>2</sub> O	NaOH	Sodium hydroxide	Soap making, many industrial uses, removal of H <sub>2</sub> S from petroleum
K <sub>2</sub> O	KOH	Potassium hydroxide	Alkaline battery manufacture
MgO	Mg(OH) <sub>2</sub>	Magnesium hydroxide	Paper making, medicinal uses
CaO	Ca(OH) <sub>2</sub>	Calcium hydroxide	Water purification, soil treatment to neutralize excessive acidity

Many important bases cannot be isolated as the hydroxides but produce OH<sup>-</sup> ion in water. A very good example is ammonia, NH<sub>3</sub>. Ammonium hydroxide, NH<sub>4</sub>OH, cannot be obtained in a pure form. Even when ammonia is dissolved in water, very little NH<sub>4</sub>OH is present in the solution. However, ammonia does react with water,



to give an ammonium ion and a hydroxide ion. Since only a small percentage of the ammonia molecules react this way, ammonia is a weak base.

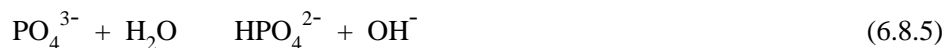
Many salts that do not themselves contain hydroxide ion act as bases by reacting with water to produce OH<sup>-</sup>. Sodium carbonate, Na<sub>2</sub>CO<sub>3</sub>, is the most widely used of these salts. When sodium carbonate is placed in water, the carbonate ion reacts with water





to form a hydroxide ion and a bicarbonate ion ( $\text{HCO}_3^-$ ). Commercial grade sodium carbonate, soda ash, is used very widely for neutralizing acid in water treatment and other applications. It is used in phosphate-free detergents. It is a much easier base to handle and use than sodium hydroxide. Whereas sodium hydroxide rapidly absorbs enough water from the atmosphere to dissolve itself to make little puddles of highly concentrated NaOH solution that are very harmful to the skin, sodium carbonate does not absorb water nearly so readily. It is not as dangerous to the skin.

Trisodium phosphate,  $\text{Na}_3\text{PO}_4$ , is an even stronger base than sodium carbonate. The phosphate ion reacts with water



to yield a high concentration of hydroxide ions. This kind of reaction with water is called a **hydrolysis reaction**.

Many organic compounds are bases. Most of these contain nitrogen. One of these is trimethylamine,  $(\text{CH}_3)_3\text{N}$ . This compound is one of several that give dead fish their foul smell. It reacts with water



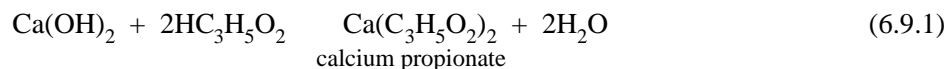
to produce hydroxide ion. Like most organic bases it is a weak base.

## 6.9 PREPARATION OF SALTS

Many salts are important industrial chemicals. Others are used in food preparation or medicine. A huge quantity of  $\text{Na}_2\text{CO}_3$  is used each year, largely to treat water and to neutralize acid. Over 1.5 million tons of  $\text{Na}_2\text{SO}_4$  are used in applications such as inert filler in powdered detergents. Approximately 30,000 tons of sodium thiosulfate,  $\text{Na}_2\text{S}_2\text{O}_3$ , are used each year in developing photographic film and in other applications. Canadian mines produce more than 10 million tons of KCl each year for use as fertilizer. Lithium carbonate,  $\text{Li}_2\text{CO}_3$ , is used as a medicine to treat some kinds of manic-depressive illness. Many other examples of the importance of salts could be given.

Whenever possible, salts are obtained by simply mining them. Many kinds of salts can be obtained by evaporating water from a few salt-rich inland sea waters or from brines pumped from beneath the ground. However, most salts cannot be obtained so directly and must be made by chemical processes. Some of these processes will be discussed.

One way of making salts already discussed in this chapter is to react an acid and a base to produce a salt and water. Calcium propionate, which is used to preserve bread is made by reacting calcium hydroxide and propionic acid,  $\text{HC}_3\text{H}_5\text{O}_2$ :



Almost any salt can be made by the reaction of the appropriate acid and base.

In some cases, a metal and a nonmetal will react directly to make a salt. If a strip of magnesium burns (explodes would be a better description) in an atmosphere of chlorine gas,

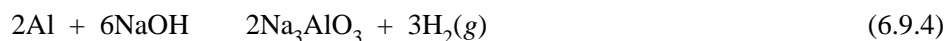


magnesium chloride salt is produced.

Metals react with acids to produce a salt and hydrogen gas. Calcium placed in sulfuric acid will yield calcium sulfate.



Some metals react with strong bases to produce salts. Aluminum metal reacts with sodium hydroxide to yield sodium aluminate,  $\text{Na}_3\text{AlO}_3$ .



In cases where a metal forms an insoluble hydroxide, addition of a base to a salt of that metal can result in the formation of a new salt. If potassium hydroxide is added to a solution of magnesium sulfate



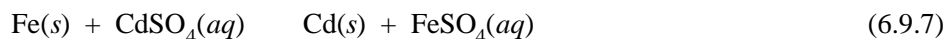
the insoluble magnesium hydroxide precipitates out of the solution, leaving potassium sulfate salt in solution.

If the anion in a salt can form a volatile acid, a new salt can be formed by adding a nonvolatile acid, heating to drive off the volatile product, and collecting the volatile acid in water. If nonvolatile sulfuric acid is heated with NaCl,



HCl gas is given off and sodium sulfate remains behind.

Some metals will displace other metals from a salt. Advantage is taken of this for the removal of toxic heavy metals from water solutions of the metals' salts by reaction with a more active metal, a process called **cementation**. For example, metallic iron can be reacted with wastewater containing dissolved toxic cadmium sulfate,

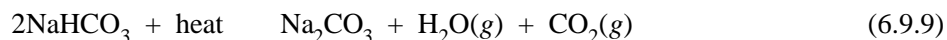


to isolate solid cadmium metal and leave solid cadmium metal and a new salt, iron(II) sulfate.

Finally, there are many special commercial processes for making specific salts. One such example is the widely used Solvay Process for making sodium bicarbonate and sodium carbonate. In this process, a sodium chloride solution is saturated with ammonia gas, then saturated with carbon dioxide and finally cooled. The reaction that occurs is



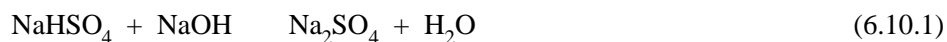
and sodium bicarbonate (baking soda) precipitates from the cooled solution. When the sodium bicarbonate is heated, it is converted to sodium carbonate:



## 6.10 ACID SALTS AND BASIC SALTS

### Acid Salts

Some compounds are crosses between acids and salts. Other salts are really crosses between bases and salts. The acid salts contain hydrogen ion. This hydrogen ion can react with bases. One example of this is sodium hydrogen sulfate,  $\text{NaHSO}_4$ , which reacts with sodium hydroxide,



to give sodium sulfate and water. Some other examples of acid salts are shown in [Table 6.6](#).

**Table 6.6 Some Important Salts**

Acid salt formula	Acid salt name	Typical use
$\text{NaHCO}_3$	Sodium hydrogen carbonate	Food preparation (baking soda)
$\text{NaH}_2\text{PO}_4$	Sodium dihydrogen phosphate	To prepare buffers
$\text{Na}_2\text{HPO}_4$	Disodium hydrogen phosphate	To prepare buffers
$\text{KH}_4\text{C}_4\text{H}_4\text{O}_6$	Potassium hydrogen tartrate	Dry acid in baking powder <sup>1</sup>

<sup>1</sup> Cream of tartar baking powder consists of a mixture of potassium hydrogen tartrate and sodium hydrogen carbonate. When this mixture contacts water in a batch of dough, the reaction



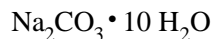
occurs, and the small bubbles of carbon dioxide it produces cause the dough to rise.

### Basic Salts

Some salts contain hydroxide ions. These are known as **basic salts**. One of the most important of these is calcium hydroxyapatite,  $\text{Ca}_5\text{OH}(\text{PO}_4)_3$ . Commonly known as *hydroxyapatite*, this salt occurs in the mineral rock phosphate, which is the source of essential phosphate fertilizer. The heavy metals, in particular, have a tendency to form basic salts. Many rock-forming minerals are basic salts.

## 6.11 WATER OF HYDRATION

Water is frequently bound to other chemical compounds. Water bound to a salt in a definite proportion is called **water of hydration**. An important example is sodium carbonate decahydrate:



Sodium carbonate decahydrate

This salt is used in detergents, as a household cleaner, and for water softening. When dissolved in water, it yields a basic solution by reacting with the water to produce hydroxide ion,  $\text{OH}^-$ . Such a solution tends to dissolve grease, so it can be used as a household cleaner. Sodium carbonate decahydrate is used in detergents, because they work best in basic solutions. The carbonate ion from this salt reacts with calcium ion (which causes water hardness),



This removes hardness from the water by producing solid calcium carbonate,  $\text{CaCO}_3$ . Because of its ability to remove water hardness, sodium carbonate decahydrate is a good water softener, another reason that it is used in detergents. Sodium carbonate without water,  $\text{Na}_2\text{CO}_3$ , is said to be **anhydrous**. It should not be used in consumer products such as powdered detergents because of its strong attraction for water. If anhydrous  $\text{Na}_2\text{CO}_3$  were present in a detergent that was accidentally ingested, it would draw water from the tissue in the mouth and throat, greatly increasing the harm done.

## 6.12 NAMES OF ACIDS, BASES, AND SALTS

### Acids

**Acids** consisting of hydrogen and one other element are named for the element with a *hydro-* prefix and an ending of *-ic*. These acids include HF, hydrofluoric; HCl, hydrochloric; HBr, hydrobromic; and  $\text{H}_2\text{S}$ , hydrosulfuric acid. Another acid named in this manner is HCN, hydrocyanic acid.

The names of many common acids end with the suffix *-ic*, as is the case with acetic acid, nitric acid ( $\text{HNO}_3$ ), and sulfuric acid. In some cases where the anion of the acid contains oxygen, there are related acids with different numbers of oxygen atoms in the anion. When this occurs, the acid with one less oxygen than the “*-ic*” acid has a name ending with *-ous*. For example,  $\text{H}_2\text{SO}_4$  is sulfuric acid and  $\text{H}_2\text{SO}_3$  is sulfurous acid. Similarly,  $\text{HNO}_3$  is nitric acid and  $\text{HNO}_2$  is nitrous acid. One more oxygen in the anion than the “*-ic*” acid is indicated by a *per-* prefix and an *-ic* suffix on the acid name. One less oxygen in the anion than in the “*-ous*” acid gives the acid a name with a *hypo-* prefix and an *-ous* suffix. These rules are shown for the oxyacids (acids containing oxygen in the anion) of chlorine in [Table 6.7](#).

**Table 6.7 Names of the Oxyacids of Chlorine**

Acid formula	Acid name <sup>1</sup>	Anion name <sup>1,2</sup>
HClO <sub>4</sub>	<i>Perchloric acid</i>	<i>Perchlorate</i>
HClO <sub>3</sub>	<i>Chloric acid</i>	<i>Chlorate</i>
HClO <sub>2</sub>	<i>Chlorous acid</i>	<i>Chlorite</i>
HClO	<i>Hypochlorous acid</i>	<i>Hypochlorite</i>

<sup>1</sup> Italicized letters are used with the names only to emphasize the prefixes and suffixes.

<sup>2</sup> Names of anions, such as ClO<sub>4</sub><sup>-</sup> formed by removal of H<sup>+</sup> ion from the acid.

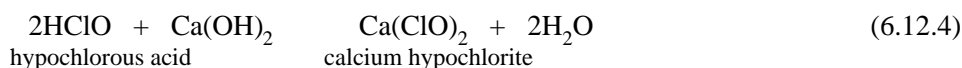
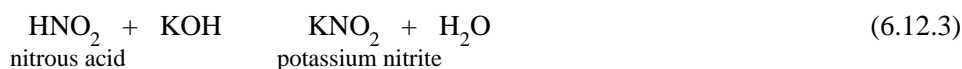
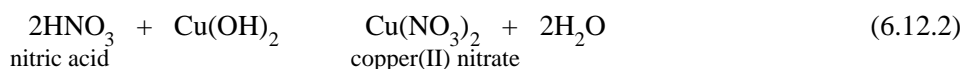
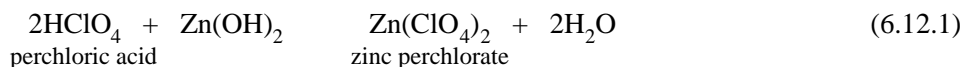
## Bases

Bases that contain hydroxide ion are named very simply by the rules of nomenclature for ionic compounds. The name consists of the name of the metal followed by hydroxide. As examples, LiOH is lithium hydroxide, KOH is potassium hydroxide, and Mg(OH)<sub>2</sub> is magnesium hydroxide.

## Salts

Salts are named according to the name of the cation followed by the name of the anion. The names of the more important ions are listed in Table 6.8. One important cation, ammonium ion, NH<sub>4</sub><sup>+</sup>, and a number of anions are polyatomic ions, meaning that they consist of 2 or more atoms per ion. Table 6.8 includes some polyatomic ions. Note that most polyatomic ions contain oxygen as one of the elements.

As illustrated in Table 6.8, the names of anions are based upon the names of the acids from which they are formed by removal of H<sup>+</sup> ions. A “*hydro -ic*” acid yields an “*-ide*” anion and, therefore, an “*-ide*” salt. For example, hydrochloric acid reacts with a base to give a chloride salt. An “*-ic*” acid yields an “*-ate*” salt; for example, calcium sulfate is the salt that results from a reaction of sulfuric acid with calcium hydroxide. The anion contained in an “*-ous*” acid is designated by the suffix “*-ite*” in a salt; sulfurous acid, H<sub>2</sub>SO<sub>3</sub> reacts with NaOH to give the salt Na<sub>2</sub>SO<sub>3</sub> sodium sulfite. A “*per -ic*” acid, such as perchloric acid, reacts with a base, such as NaOH, to give a “*per -ate*” salt, for example, sodium perchlorate, NaClO<sub>4</sub>. A “*hypo -ous*” acid, such as hypochlorous acid, reacts with a base, KOH, for example, to give a “*hypo -ite*” salt, such as potassium hypochlorite, KClO. Some additional examples are illustrated by the following reactions:



**Table 6.8 Some Important Ions**

+1 charge	+2 charge	+3 charge	-1 charge	-2 charge	-3 charge
H <sup>+</sup> , hydrogen	Mg <sup>2+</sup> , magnesium	Al <sup>3+</sup> , aluminum	H <sup>-</sup> , hydride	O <sup>2-</sup> , oxide	N <sup>3-</sup> , nitride
Li <sup>+</sup> , lithium	Ca <sup>2+</sup> , calcium	Fe <sup>3+</sup> , iron(III) <sup>2</sup>	F <sup>-</sup> , fluoride	S <sup>2-</sup> , sulfide	PO <sub>4</sub> <sup>3-</sup> , phosphate
Na <sup>+</sup> , sodium	Ba <sup>2+</sup> , barium	Cr <sup>3+</sup> , chrom- ium(III)	Cl <sup>-</sup> , chloride	SO <sub>4</sub> <sup>2-</sup> , sulfate	
K <sup>+</sup> , potassium	Fe <sup>2+</sup> , iron (II)		Br <sup>-</sup> , bromide	SO <sub>3</sub> <sup>2-</sup> , sulfite	
NH <sub>4</sub> <sup>+</sup> , ammonium	Zn <sup>2+</sup> , zinc		I <sup>-</sup> , iodide	CO <sub>3</sub> <sup>2-</sup> , carbonate	
Cu <sup>+</sup> , copper (I) <sup>1</sup>	Cu <sup>2+</sup> , copper(II) <sup>2</sup>		C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup> , acetate	CrO <sub>4</sub> <sup>2-</sup> , chromate	
Ag <sup>+</sup> , silver	Cr <sup>2+</sup> , chromium(II)		OH <sup>-</sup> , hydroxide	Cr <sub>2</sub> O <sub>7</sub> <sup>2-</sup> , dichromate	
	Pb <sup>2+</sup> , lead(II) <sup>1</sup>		CN <sup>-</sup> , cyanide	O <sub>2</sub> <sup>2-</sup> , peroxide	
	Hg <sup>2+</sup> , mercury(II) <sup>2</sup>		NO <sub>3</sub> <sup>-</sup> , nitrate	HPO <sub>4</sub> <sup>2-</sup> , monohydrogen phosphate	
	Sn <sup>2+</sup> , tin(II) <sup>2</sup>		H <sub>2</sub> PO <sub>4</sub> <sup>-</sup> , dihydrogen phosphate		
			HCO <sub>3</sub> <sup>-</sup> , hydrogen carbonate <sup>3</sup>		
			HSO <sub>4</sub> <sup>-</sup> , hydrogen sulfate <sup>3</sup>		
			MnO <sub>4</sub> <sup>-</sup> , permanganate		

<sup>1</sup> These metals can also exist as ions with a higher charge and may also be designated by their Latin names with an *-ous* ending as follows: Cu<sup>+</sup>, cuprous; Fe<sup>2+</sup>, ferrous; Cr<sup>2+</sup>, chromous; Pb<sup>2+</sup>, plumbous; Mn<sup>2+</sup>, manganous; Sn<sup>2+</sup>, stannous.

<sup>2</sup> These metals can also exist as ions with a lower charge and may also be designated by their Latin names with an *-ic* ending as follows: Cu<sup>2+</sup>, cupric; Hg<sup>2+</sup>, mercuric; Fe<sup>3+</sup>, ferric; Cr<sup>3+</sup>, chromic.

<sup>3</sup> The ions HCO<sub>3</sub><sup>-</sup> and HSO<sub>4</sub><sup>-</sup> are known as bicarbonate and bisulfate, respectively.

Using the information given in Table 6.8, it is possible to figure out the formulas and give the names of a very large number of ionic compounds. To do that, simply observe the following steps:

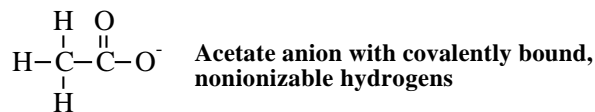
1. Choose the cation and the anion of the compound. The name of the compound is simply the name of the cation followed by the name of the anion. For example, when  $\text{Fe}^{3+}$  is the cation and  $\text{SO}_4^{2-}$  is the anion; the name of the ionic compound is iron(III) sulfate.
2. Choose subscripts to place after the cation and anion in the chemical formula of the compound such that multiplying the subscript of the cation times the charge of the cation gives a number equal in magnitude and opposite in sign from that of the product of the anion's subscript times the anion's charge. In the example of iron(III) sulfate, a subscript of 2 for  $\text{Fe}^{3+}$  gives  $2 \times (3+) = 6+$ , and a subscript of 3 for  $\text{SO}_4^{2-}$  gives  $3 \times (2-) = 6-$ , thereby meeting the condition for a neutral compound.
3. Write the compound formula. If the subscript after any polyatomic ion is greater than 1, put the formula of the ion in parentheses to show that the subscript applies to all the atoms in the ion. Omit the charges on the ions because they make the compound formula too cluttered. In the example under consideration, the formula of iron(III) sulfate is  $\text{Fe}_2(\text{SO}_4)_3$ .

Exercise: Match each cation in the left column below with each anion in the right column and give the formulas and names of each of the resulting ionic compounds.



Answers: a-1, NaBr, sodium bromide; a-2,  $\text{Na}_2\text{CO}_3$ , sodium carbonate; a-3,  $\text{Na}_3\text{PO}_4$ , sodium phosphate; b-1,  $\text{CaBr}_2$ , calcium bromide; b-2,  $\text{CaCO}_3$ , calcium carbonate; b-3,  $\text{Ca}_3(\text{PO}_4)_2$ , calcium phosphate; c-1,  $\text{AlBr}_3$ , aluminum bromide; c-2,  $\text{Al}_2(\text{CO}_3)_3$ , aluminum carbonate; c-3,  $\text{AlPO}_4$ , aluminum phosphate.

The names and formulas of ionic compounds that contain hydrogen in the name and formula of the anion are handled just like any other ionic compound. Therefore,  $\text{NaHCO}_3$  is sodium hydrogen carbonate,  $\text{Ca}(\text{H}_2\text{PO}_4)_2$  is calcium dihydrogen phosphate, and  $\text{K}_2\text{HPO}_4$  is sodium monohydrogen phosphate. The acetate ion,  $\text{C}_2\text{H}_3\text{O}_2^-$  also contains hydrogen, but, as noted previously, its hydrogen is covalently bonded to a C atom, as shown by the structure



and cannot form  $\text{H}^+$  ions, whereas the anions listed with hydrogen in their names can produce  $\text{H}^+$  ion when dissolved in water.

## CHAPTER SUMMARY

The chapter summary below is presented in a programmed format to review the main points covered in this chapter. It is used most effectively by filling in the blanks, referring back to the chapter as necessary. The correct answers are given at the end of the summary.

<sup>1</sup> \_\_\_\_\_ ion is produced by acids and <sup>2</sup> \_\_\_\_\_ by bases. A neutralization reaction is <sup>3</sup> \_\_\_\_\_. Hydrogen ion,  $H^+$ , in water is bonded to <sup>4</sup> \_\_\_\_\_ and is often represented as <sup>5</sup> \_\_\_\_\_. A **base** is a substance that accepts  $H^+$  and produces hydroxide ion. Although  $NH_3$  does not contain hydroxide ions, it undergoes the reaction <sup>6</sup> \_\_\_\_\_ to produce  $OH^-$  in water. The two products produced whenever an acid and a base react together are <sup>7</sup> \_\_\_\_\_. A salt is made up of <sup>8</sup> \_\_\_\_\_  
<sup>9</sup> \_\_\_\_\_. An amphoteric substance is one that  
<sup>9</sup> \_\_\_\_\_. In water, a metal ion is bonded to <sup>10</sup> \_\_\_\_\_ in a form known as a <sup>11</sup> \_\_\_\_\_. In terms of acid–base behavior, some metal ions act as <sup>12</sup> \_\_\_\_\_. In water solution sodium carbonate acts as a <sup>13</sup> \_\_\_\_\_ and undergoes the reaction <sup>14</sup> \_\_\_\_\_. Salts that act as acids react with <sup>15</sup> \_\_\_\_\_  
\_\_\_\_\_. Pure water conducts electricity <sup>16</sup> \_\_\_\_\_, a solution of acetic acid conducts electricity <sup>17</sup> \_\_\_\_\_, and a solution of  $HCl$  conducts electricity <sup>18</sup> \_\_\_\_\_. These differences are due to differences in concentrations of <sup>19</sup> \_\_\_\_\_ in the water. Materials that conduct electricity in water are called <sup>20</sup> \_\_\_\_\_. Materials that do not form ions in water are called <sup>21</sup> \_\_\_\_\_. The reaction



may be classified as <sup>22</sup> \_\_\_\_\_ or <sup>23</sup> \_\_\_\_\_ and when the acetic acid molecule comes apart, it is said to <sup>24</sup> \_\_\_\_\_. A base that is completely dissociated in water is called a <sup>25</sup> \_\_\_\_\_ and an acid that is only slightly dissociated is called a <sup>26</sup> \_\_\_\_\_. At high concentrations the percentage of dissociation of a weak acid is <sup>27</sup> \_\_\_\_\_ than at lower concentrations. Buffers are <sup>28</sup> \_\_\_\_\_. A buffer can be made from a mixture of a weak base and <sup>29</sup> \_\_\_\_\_. The reaction that results in the production of very low concentrations of ions in even pure water is <sup>30</sup> \_\_\_\_\_ and the relationship between the concentrations of these ions in water is <sup>31</sup> \_\_\_\_\_. In absolutely pure water the value of  $[H^+]$  is exactly <sup>32</sup> \_\_\_\_\_, the pH is <sup>33</sup> \_\_\_\_\_, so that the solution is said to be <sup>34</sup> \_\_\_\_\_. Acidic solutions have pH values of <sup>35</sup> \_\_\_\_\_ and basic solutions have pH values of <sup>36</sup> \_\_\_\_\_. In a general sense solution equilibrium deals with the extent to which reversible acid-base, solubilization (precipitation), complexation, or oxidation-reduction reactions <sup>37</sup> \_\_\_\_\_  
\_\_\_\_\_. As an example of acid-base equilibrium



the reaction for the ionization of acetic acid, HAc, is <sup>38</sup> \_\_\_\_\_  
for which the acid dissociation constant is <sup>39</sup> \_\_\_\_\_. Some  
ways to prepare acids are <sup>40</sup> \_\_\_\_\_

Some ways to prepare bases are <sup>41</sup> \_\_\_\_\_

The reaction of an ion with water such as



is an example of a <sup>42</sup> \_\_\_\_\_. The most obvious way  
to prepare a salt is by <sup>43</sup> \_\_\_\_\_. Active  
metals react with acids to produce <sup>44</sup> \_\_\_\_\_. Other  
than reacting with acids, some metals react with <sup>45</sup> \_\_\_\_\_. If the anion  
in a salt can form a volatile acid, a new salt can be formed by <sup>46</sup> \_\_\_\_\_  
\_\_\_\_\_. Some metals will displace other metals from a salt. If magnesium, a  
highly reactive metal is added to a solution of copper sulfate, the reaction that occurs  
is <sup>47</sup> \_\_\_\_\_. NaHSO<sub>4</sub>, which has an ionizable  
hydrogen, is an example of <sup>48</sup> \_\_\_\_\_, whereas Ca<sub>5</sub>OH(PO<sub>4</sub>)<sub>3</sub> is an  
example of <sup>49</sup> \_\_\_\_\_. The water in CuSO<sub>4</sub>•5H<sub>2</sub>O is called <sup>50</sup> \_\_\_\_\_  
\_\_\_\_\_. The names of HClO<sub>4</sub>, HClO<sub>3</sub>, HClO<sub>2</sub>, and HClO are,  
respectively, <sup>51</sup> \_\_\_\_\_.

The names of NaClO<sub>4</sub>, NaClO<sub>3</sub>, NaClO<sub>2</sub>, and NaClO are, respectively, <sup>52</sup> \_\_\_\_\_

The name of a base containing a metal consists of <sup>53</sup> \_\_\_\_\_  
\_\_\_\_\_. As examples, LiOH is lithium hydroxide, KOH is  
potassium hydroxide, and Mg(OH)<sub>2</sub> is magnesium hydroxide. The name of a salt is  
<sup>54</sup> \_\_\_\_\_. The names  
of the ions Ca<sup>2+</sup>, Fe<sup>3+</sup>, H<sup>-</sup>, SO<sub>3</sub><sup>2-</sup>, and C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup> are, respectively, <sup>55</sup> \_\_\_\_\_

In writing the formulas of ionic compounds, choose subscripts to place after the  
cation and anion in the chemical formula of the compound such that multiplying the  
subscript of the cation times the charge of the cation gives a number <sup>56</sup> \_\_\_\_\_  
\_\_\_\_\_ and opposite in sign from that of the product of <sup>57</sup> \_\_\_\_\_

## Answers to Chapter Summary

1.  $\text{H}^+$
2.  $\text{OH}^-$
3.  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
4. water molecules
5. hydronium ion,  $\text{H}_3\text{O}^+$
6.  $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$
7. water and a salt
8. a cation (other than  $\text{H}^+$ ) and an anion (other than  $\text{OH}^-$ )
9. can act as either an acid or a base
10. water molecules
11. hydrated ion
12. acids
13. base
14.  $2\text{Na}^+ + \text{CO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{Na}^+ + \text{HCO}_3^- + \text{Na}^+ + \text{OH}^-$
15. hydroxide ions
16. not at all
17. poorly
18. very well
19. ions
20. electrolytes
21. nonelectrolytes
22. ionization
23. dissociation
24. dissociate
25. strong base
26. weak acid
27. lower
28. solutions that resist changes in  $\text{H}^+$  concentration
29. a salt of the base
30.  $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$
31.  $[\text{H}^+][\text{OH}^-] = 1.00 \times 10^{-14} = K_w$
32.  $1 \times 10^{-7}$  mole/L
33. 7.00
34. neutral
35. less than 7
36. greater than 7
37. proceed in a forward or backward direction
38.  $\text{HAc} \rightleftharpoons \text{H}^+ + \text{Ac}^-$
39.  $\frac{[\text{H}^+][\text{Ac}^-]}{[\text{HAc}]} = K = 1.75 \times 10^{-5}$
40. reaction of hydrogen with a nonmetal, reaction of a nonmetal directly with water, reaction of a nonmetal oxide with water, production of volatile acids by reaction of salts of the acids with nonvolatile acids
41. reaction of active metals directly with water, reaction of metal oxides with

water, reaction of a basic compound that does not itself contain hydroxide with water

42. hydrolysis
43. reaction of an acid with a base
44. a salt and hydrogen gas
45. strong bases
46. adding a nonvolatile acid
47.  $\text{Mg}(s) + \text{CuSO}_4(aq) \rightarrow \text{Cu}(s) + \text{MgSO}_4(aq)$
48. an acid salt
49. a basic salt
50. water of hydration
51. perchloric acid, chloric acid, chlorous acid, and hypochlorous acid
52. sodium perchlorate, sodium chlorate, sodium chlorite, and sodium hypochlorite
53. the name of the metal followed by hydroxide
54. the name of the cation followed by the name of the anion
55. calcium, iron(III), hydride, sulfite, and acetate
56. equal in magnitude
57. the anion's subscript times the anion's charge

## QUESTIONS AND PROBLEMS

1. Give the neutralization reaction for each of the acids in the left column reacting with each of the bases in the right column, below:

(1) Hydrocyanic acid	(a) Ammonia
(2) Acetic acid	(b) Sodium hydroxide
(3) Phosphoric acid	(c) Calcium hydroxide
2. Methylamine is an organic amine, formula  $\text{H}_3\text{C-NH}_2$ , that acts as a weak base in water. By analogy with ammonia, suggest how it might act as a base.
3. In a 1 molar solution of acetic acid (containing 1 mol of acetic acid per liter of solution) only about 0.5% of the acid is ionized to produce an acetate ion and a hydrogen ion. Calculate the number of moles of  $\text{H}^+$  in a liter of such a solution.
4. What does  $\text{H}_3\text{O}^+$  represent in water?
5. When exactly 1 mole of  $\text{NaOH}$  reacts with exactly 1 mole of  $\text{H}_2\text{SO}_4$ , the product is an acid salt. Show the production of the acid salt with a chemical reaction.
6. An amphoteric substance can be viewed as one that may either accept or produce an ion of  $\text{H}^+$ . Using that definition, explain how  $\text{H}_2\text{O}$  is amphoteric.
7. Explain how  $\text{Fe}^{3+}$  ion dissolved in water can be viewed as an acidic hydrated ion.
8. Cyanide ion,  $\text{CN}^-$ , has a strong attraction for  $\text{H}^+$ . Show how this explains why  $\text{NaCN}$  acts as a base.
9. Separate solutions containing 1 mole per liter of  $\text{NH}_3$  and 1 mole per liter of acetic acid conduct electricity poorly, whereas when such solutions are mixed, the resulting solution conducts well. Explain.

10. What characteristic of solutions of electrolytes enables them to conduct electricity well?
11. A solution containing 6 moles of  $\text{NH}_3$  dissolved in a liter of solution would be relatively highly concentrated. Explain why it would not be correct, however, to describe such a solution as a “strong base” solution.

12. The dissociation of acetic acid can be represented by



Explain why this reaction can be characterized as an ionization of acetic acid. Explain on the basis of the “crowding” concept why the percentage of acetic acid molecules dissociated is less in relatively concentrated solutions of the acid.

13. A solution containing 0.1 mole of  $\text{HCl}$  per liter of solution has a low pH of 1, whereas a solution containing 0.1 mole of acetic acid per liter of solution has a significantly higher pH. Explain.
14. Explain why a solution containing both  $\text{NH}_3$  and  $\text{NH}_4\text{Cl}$  acts as a buffer. In so doing, consider reactions of  $\text{NH}_3$ ,  $\text{NH}_4^+$  ion,  $\text{H}^+$  ion,  $\text{OH}^-$  ion, and  $\text{H}_2\text{O}$ .
15.  $\text{NaH}_2\text{PO}_4$  and  $\text{Na}_2\text{HPO}_4$  dissolved in water produce  $\text{H}_2\text{PO}_4^-$  and  $\text{HPO}_4^{2-}$  ions, respectively. Show by reactions of these ions with  $\text{H}^+$  and  $\text{OH}^-$  ions why a solution consisting of a mixture of both  $\text{NaH}_2\text{PO}_4$  and  $\text{Na}_2\text{HPO}_4$  dissolved in water acts as a buffer.
16. What is the expression and value for  $K_w$ ? What is the reaction upon which this expression is based?
17. On the basis of pH, distinguish among acidic, basic, and neutral solutions.
18. Give the pH values corresponding to each of the following values of  $[\text{H}^+]$ :  
 (a)  $1.00 \times 10^{-4}$  mol/L, (b)  $1.00 \times 10^{-8}$  mol/L, (c)  $5.63 \times 10^{-9}$  mol/L, (d)  $3.67 \times 10^{-6}$  mol/L.
19. Why does solution equilibrium deal only with reversible reactions?
20. Write an equilibrium constant expression for the reaction
- $$\text{CO}_3^{2-} + \text{H}_2\text{O} \rightleftharpoons \text{HCO}_3^- + \text{OH}^-$$
21. Calculate  $[\text{H}^+]$  in a solution of carbon dioxide in which  $[\text{CO}_2(aq)]$  is  $3.25 \times 10^{-4}$  moles/liter.
22.  $\text{Cl}_2$  and  $\text{F}_2$  are both halogens. Suggest acids that might be formed from the reaction of  $\text{F}_2$  with  $\text{H}_2$  and with  $\text{H}_2\text{O}$ .
23. Suggest the acid that might be formed by reacting  $\text{S}$  with  $\text{H}_2$ .
24. Suggest the acid or base that might be formed by the reaction of each of the following oxides with water: (a)  $\text{N}_2\text{O}_3$ , (b)  $\text{CO}_2$ , (c)  $\text{SO}_3$ , (d)  $\text{Na}_2\text{O}$ , (e)  $\text{CaO}$ , (f)  $\text{Cl}_2\text{O}$ .
25. Knowing that  $\text{H}_2\text{SO}_4$  is a non-volatile acid, suggest the acid that might be formed by the reaction of  $\text{H}_2\text{SO}_4$  with water.

26. Acetic acid is a carboxylic acid. Formic acid is the lowest carboxylic acid, and it contains only 1 C atom per molecule. What is its formula?
27. A base can be prepared by the reaction of calcium metal with hot water. Give the reaction and the name of the base product.
28. A base can be prepared by the reaction of sodium oxide with water. Give the reaction and the name of the base product.
29. Give the reactions by which the following act as bases in water: (a)  $\text{NH}_3$ , (b)  $\text{Na}_2\text{CO}_3$ , (c)  $\text{Na}_3\text{PO}_4$ , and dimethylamine,  $(\text{CH}_3)_2\text{NH}$ .
30. Choosing from the reagents  $\text{H}_2\text{SO}_4$ ,  $\text{HCl}$ ,  $\text{Mg}(\text{OH})_2$ , and  $\text{LiOH}$  give reactions that illustrate “the most straightforward” means of preparing salts.
31. Choosing from the reagents  $\text{NaOH}$ ,  $\text{HCl}$ ,  $\text{CaO}$ ,  $\text{Mg}$ ,  $\text{F}_2$  and  $\text{Al}$  give reactions that illustrate the preparation of salts by (a) reaction of a metal and a nonmetal that will react directly to make a salt, (b) reaction of a metal with acid, (c) reaction of a metal with strong base, (d) reaction of a salt with a nonvolatile acid, (e) cementation.
32. Describe what is meant by an acid salt.
33. Describe what is meant by a basic salt.
34. Explain how sodium carbonate decahydrate illustrates water of hydration. Why is it less hazardous to skin than is anhydrous sodium carbonate? Illustrate with a chemical reaction why it is also a basic salt.
35. Give the names of each of the following acids: (a)  $\text{HBr}$ , (b)  $\text{HCN}$ , (c)  $\text{HClO}$ , (d)  $\text{HClO}_2$ , (e)  $\text{HClO}_3$ , (f)  $\text{HClO}_4$ , (g)  $\text{HNO}_2$ .
36. Give the names of (a)  $\text{LiOH}$ , (b)  $\text{Ca}(\text{OH})_2$ , and (c)  $\text{Al}(\text{OH})_3$ .
37. Give the names of (a)  $\text{MgSO}_3$ , (b)  $\text{NaClO}$ , (c)  $\text{Ca}(\text{ClO}_4)_2$ , (d)  $\text{KNO}_3$ , (d)  $\text{Ca}(\text{NO}_2)_2$
38. Match each cation in the left column below with each anion in the right column and give the formulas and names of each of the resulting ionic compounds.

(a) $\text{Li}^+$	(1) $\text{CN}^-$
(b) $\text{Ca}^{2+}$	(2) $\text{SO}_3^{2-}$
(c) $\text{Fe}^{3+}$	(3) $\text{NO}_3^-$

39. What color is litmus in (A) acid and (B) base?
40. Give the formulas of each of the following:

Magnesium acetate  
 Calcium monohydrogen phosphate  
 Aluminum sulfate  
 Calcium hypochlorite

41. Using Lewis (electron-dot) structures, show the reaction between hydronium ion and hydroxide ion.
42. A common error in speaking the chemical language is to confuse acidic (pronounced uh-sid-ik) with acetic (pronounced uh-seat-ik). What is the correct meaning of each of these terms? What is the difference between ammonia and ammonium?
43. The following is a list that contains the names of three cations and three anions: hypochlorite, hydrogen, sodium, sulfate, calcium, nitrate. List the three cations. List the three anions. Give the formulas of nine compounds that can be made by various combinations of these.
44. Write a chemical reaction in which  $\text{NaHCO}_3$  acts as an acid. Write another in which it acts as a base, remembering that if  $\text{H}_2\text{CO}_3$  is produced in solution it largely goes to carbon dioxide gas and water.
45. Write the Lewis structure of the hydronium ion,  $\text{H}_3\text{O}^+$ .
46. Explain by chemical reactions how a mixture of  $\text{NaHCO}_3$  and  $\text{Na}_2\text{CO}_3$  in water would act as a buffer.
47. A solid known to be either  $\text{NaCl}$  or  $\text{Na}_2\text{SO}_4$  was moistened with concentrated  $\text{H}_2\text{SO}_4$  and heated, giving off a gas that turned moist blue litmus paper red. What was the solid?
48. Formulas of some chemical compounds are given in the lefthand column. Match the formula of each compound with its correct name in the righthand column.

$\text{CaO}$	Potassium sulfide
$\text{SiO}_2$	Dinitrogen pentoxide
$\text{K}_2\text{S}$	Nitrogen dioxide
$\text{AlCl}_3$	Silicon dioxide
$\text{NO}_2$	Potassium bromide
$\text{N}_2\text{O}_5$	Sodium iodide
$\text{NaI}$	Calcium oxide
$\text{KBr}$	Magnesium fluoride
$\text{MgF}_2$	Calcium fluoride
$\text{CaF}_2$	Aluminum chloride

49. Match the names in the right column with the formulas in the left column.

$\text{Na}_2\text{CO}_3$	Calcium sulfate
$\text{CaSO}_3$	Potassium perchlorate
$\text{Al}(\text{OH})_3$	Sodium carbonate
$\text{CaSO}_4$	Calcium phosphate
$\text{NaNO}_2$	Aluminum hydroxide
$\text{Ca}_3(\text{PO}_4)_2$	Calcium sulfite
$\text{NaNO}_3$	Calcium hypochlorite
$\text{Ca}(\text{ClO})_2$	Sodium nitrate
$\text{KClO}_4$	Sodium nitrite

50. Match the names in the right column with the formulas in the left column.

$K_2HPO_4$	Sodium hydrogen sulfate
$KHCO_3$	Sodium hydrogen oxalate
$NaHSO_4$	Dipotassium hydrogen phosphate
$KH_2PO_4$	Sodium hydrogen phthalate
$NaHC_8H_4O_4$	Potassium hydrogen carbonate
$NaHC_2O_4$	Potassium dihydrogen phosphate

51. Fill in each of the following blanks with the number corresponding to the meaning of each of the prefixes. The first one is done for you as an example.

tetra 4 mono \_\_\_\_\_ octa \_\_\_\_\_ deca \_\_\_\_\_ di \_\_\_\_\_

penta \_\_\_\_\_ hepta \_\_\_\_\_ tri \_\_\_\_\_ nona \_\_\_\_\_ sexa \_\_\_\_\_

52. Give the correct name to each of the following compounds. The first one is done for you as an example.

$N_2O_5$  Dinitrogen pentoxide  $N_2O_4$  \_\_\_\_\_

$NO_2$  \_\_\_\_\_  $N_2O_3$  \_\_\_\_\_

$NO$  \_\_\_\_\_  $N_2O$  \_\_\_\_\_

53. Iron in a compound can also be designated as ferrous or ferric. Similarly copper (Cu) may be called cuprous or cupric. Tin (Sn) may be called stannous or stannic. Name each of the following compounds with two acceptable names. The first one is done for you as an example.

$FeCl_2$  ferrous chloride or iron (II) chloride

$FeCl_3$  \_\_\_\_\_ or \_\_\_\_\_

$CuCl$  \_\_\_\_\_ or \_\_\_\_\_

$CuCl_2$  \_\_\_\_\_ or \_\_\_\_\_

$SnCl_2$  \_\_\_\_\_ or \_\_\_\_\_

$SnCl_4$  \_\_\_\_\_ or \_\_\_\_\_

54. Dry  $CaSO_4$  absorbs enough water to yield a product with a specific number or waters of hydration. Exactly 136 g. of  $CaSO_4$  exposed to humid air gained enough water to weigh exactly 172 g. What is the formula of the product with the waters of hydration?

55. In each of the following chemical reactions fill in the formula of the missing compound. The rest of the chemical equation is balanced.

$2 N_2 + 3 O_2$  \_\_\_\_\_ (nitrogen trioxide)

$KOH + SO_2$  \_\_\_\_\_ (potassium hydrogen sulfite)

$KOH + H_3PO_4$  \_\_\_\_\_ (potassium dihydrogen phosphate)

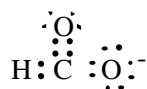
56. Give one or two examples of compounds in which each of the following prefixes or suffixes is used in the compound name.

-ic acetic acid hydrochloric acid

-ide \_\_\_\_\_

-ous \_\_\_\_\_  
 hypo- \_\_\_\_\_  
 per- \_\_\_\_\_  
 -ite \_\_\_\_\_  
 -ate \_\_\_\_\_

57. In this chapter, it was mentioned that a certain group was characteristic of organic acids. The Lewis structure of acetic acid was also given and it was shown how many of its Hs can form  $H^+$  ion. The Lewis formula of the formate ion produced by the ionization of formic acid is,



From this information, explain how many ionizable hydrogens formic acid has, and why it has that number.

58. How do you explain the fact that a 0.01 M solution of HCl contains more  $H^+$  ions than does as 0.1 M (ten-fold higher concentration) of acetic acid?
59. How is it explained that a solution of  $FeCl_3$  is acidic?
60. Explain how a mixture of  $NH_3$  and  $NH_4Cl$  dissolved in solution can act as a buffer.
61. Hydrochloric acid, HCl, is often found in rainwater in ocean coastal areas in which the atmosphere is polluted by sulfuric acid. How might this occur?
62. Classify each of the following as strong electrolytes, weak electrolytes, or nonelectrolytes: (A) Solution of NaCl, (B) vinegar, (C) pure water, (D) solution of sugar, (E) solution of HCl, (F) solution of ammonia, (G) solution of sodium hydroxide.
63. Match the value of  $[H^+]$  in the left column below with the pH in the right column.
- |                            |        |
|----------------------------|--------|
| A. $1.00 \times 10^{-8} M$ | 1.8.37 |
| B. $1.00 \times 10^{-9} M$ | 2.8.00 |
| C. $4.28 \times 10^{-9} M$ | 3.7.70 |
| D. $2.00 \times 10^{-8} M$ | 4.9.00 |
64. Exactly 1.00 mole of an acid was dissolved in 1.00 liter of water solution. The pH of the resulting solution was 2.00. Referring to [Tables 6.1](#) and [6.3](#), classify the acid in regard to its strength.