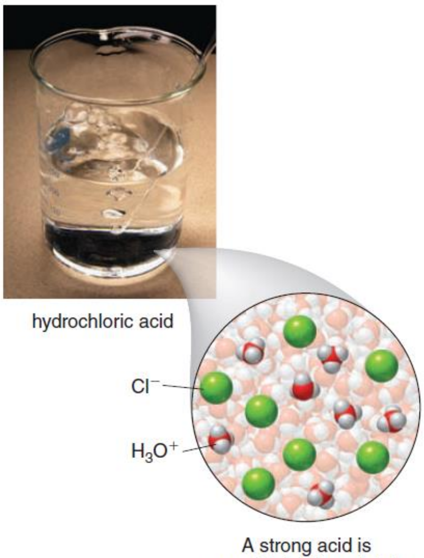
Al-Rasheed University College
Department of Dentistry
1st Stage

MEDICAL CHEMISTRY

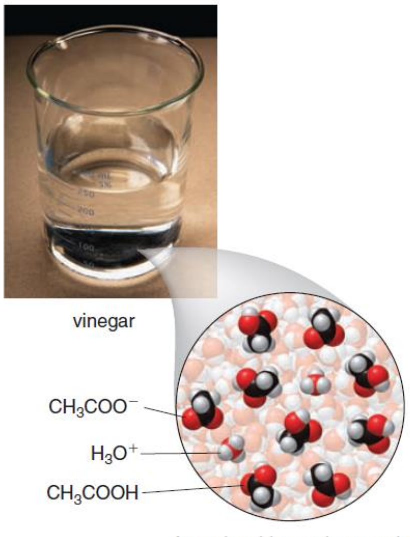
lecture 4
Acids and Bases

Edited By:

Lec. Dr. Haider AbdulKareem Al-Mashhadani



A strong acid is completely dissociated.



A weak acid contains mostly undissociated acid, CH_3COOH.

Introduction

Chemical terms such as *anion* and *cation* may be unfamiliar to most nonscientists, but acid has found a place in everyday language. ***Acid*** comes from the Latin word *acidus*, meaning sour, because when tasting compounds was a routine method of identification, these compounds were found to be sour. ***Acids*** commonly react with bases, and many products, including antacid tablets, glass cleaners, and drain cleaners, all contain bases.

Acids and Base According Arrhenius

- The earliest definition of acids and bases was suggested by Swedish chemist Svante Arrhenius in the late nineteenth century. **According to Arrhenius.**

✓ **An acid** contains a hydrogen atom and dissolves in water to form a hydrogen ion, H^+.

$$HCl \longrightarrow H^+ + Cl^-$$

✓ **A base** contains hydroxide and dissolves in water to form OH^-.

$$NaOH \longrightarrow Na^+ + {}^-OH$$

By this definition, hydrogen chloride (HCl) is an acid because it forms aqueous H^+ and $Cl^–$ when it dissolves in water. Sodium hydroxide (NaOH) is a base because it contains –OH and forms solvated Na^+ and $^–OH$ ions when it dissolves in water.

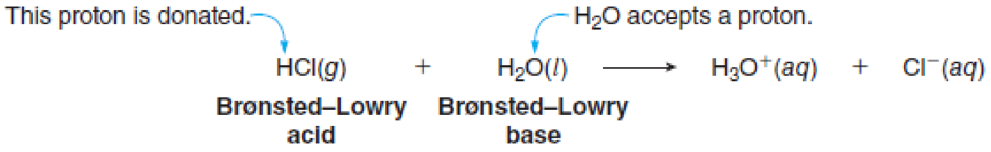
Acids and Base According Brønsted–Lowry

While the Arrhenius definition correctly predicts the behavior of many acids and bases, this definition is limited and sometimes inaccurate. We now know, for example, that the hydrogen ion, H^+, does not exist in water. H^+ is a naked proton with no electrons, and this concentrated positive charge reacts rapidly with a molecule of H_2O to form the hydronium ion, H_3O^+. Although $H^+_{(aq)}$ will sometimes be written in an equation for emphasis, $H_3O^+_{(aq)}$ is actually the reacting species.

Moreover, several compounds contain no hydroxide anions, yet they still exhibit the characteristic properties of a base. Examples include the neutral molecule ammonia (NH_3) and the salt sodium carbonate (Na_2CO_3). As a result, a more general definition of acids and bases, proposed by Johannes Brønsted and

Thomas Lowry in the early twentieth century, is widely used today. In the Brønsted–Lowry definition, acids and bases are classified according to whether they can donate or accept a proton—a positively charged hydrogen ion, H^+.

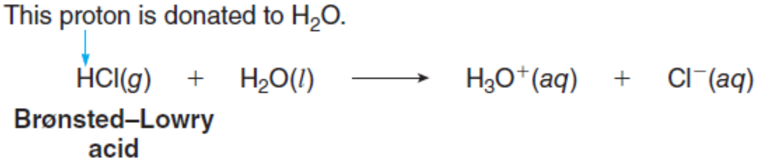
- ✓ A Brønsted–Lowry acid is a proton donor.
- ✓ A Brønsted–Lowry base is a proton acceptor.



- ➢ HCl is a Brønsted–Lowry acid because it donates a proton to the solvent water.
- ➢ H_2O is a Brønsted–Lowry base because it accepts a proton from HCl.

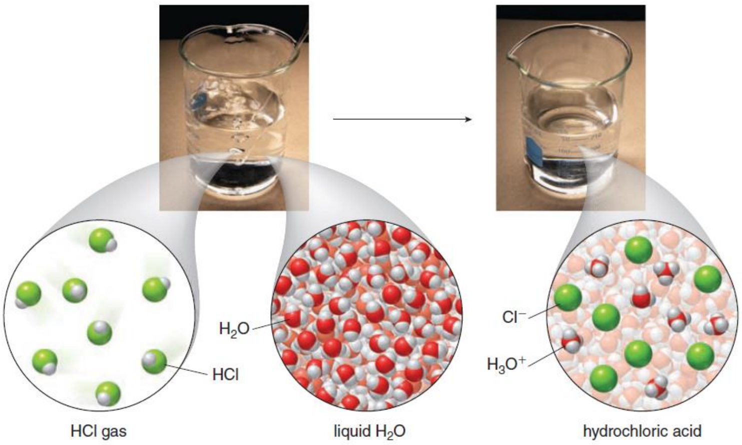
A. *Brønsted–Lowry Acids*

A Brønsted–Lowry acid must contain a hydrogen atom. HCl is a Brønsted–Lowry acid because it donates a proton (H^+) to water when it dissolves, forming the hydronium ion (H_3O^+) and chloride (Cl^-).



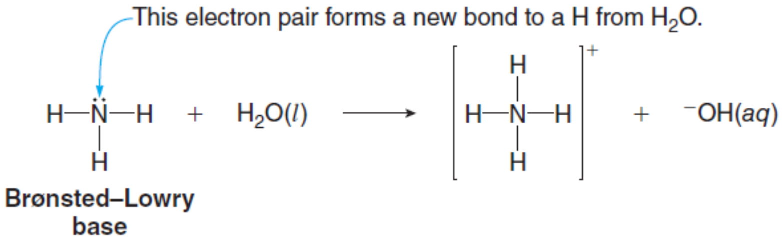
Although hydrogen chloride, HCl, is a covalent molecule and a gas at room temperature, when it dissolves in water it reacts to form two ions, H_3O^+ and Cl^-. An aqueous solution of hydrogen chloride is called hydrochloric acid.

Because a Brønsted–Lowry acid contains a hydrogen atom, a general Brønsted–Lowry acid is often written as HA. A can be a single atom such as Cl or Br. Thus, HCl and HBr are Brønsted–Lowry acids. A can also be a polyatomic ion. Sulfuric acid (H_2SO_4) and nitric acid (HNO_3) are Brønsted–Lowry acids, as well. Carboxylic acids are a group of Brønsted–Lowry acids that contain the atoms COOH arranged so that the carbon atom is doubly bonded to one O atom and singly bonded to another. Acetic acid, CH_3COOH, is a simple carboxylic acid. Although carboxylic acids may contain several hydrogen atoms, the H atom of the OH group is the acidic proton that is donated.

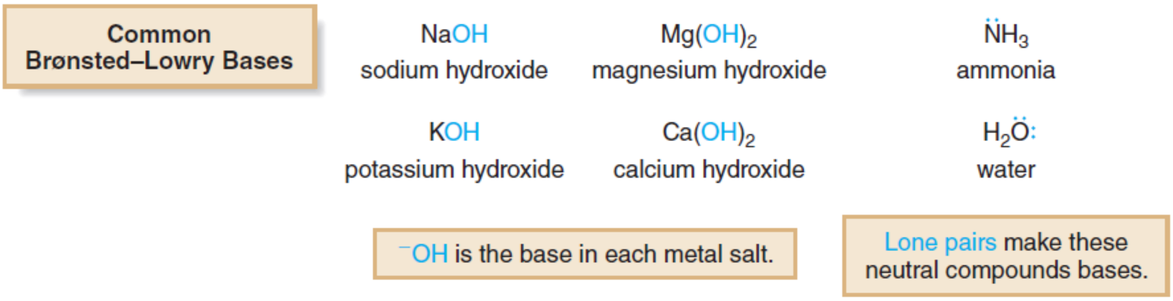


B. Brønsted–Lowry Bases

A Brønsted–Lowry base is a proton acceptor and as such, it must be able to form a bond to a proton. Because a proton has no electrons, a base must contain a lone pair of electrons that can be donated to form a new bond. Thus, ammonia (NH_3) is a Brønsted–Lowry base because it contains a nitrogen atom with a lone pair of electrons. When NH_3 is dissolved in water, its N atom accepts a proton from H_2O, forming an ammonium cation (NH_4^+) and hydroxide (^-OH).



A general Brønsted–Lowry base is often written as B: to emphasize that the base must contain a lone pair of electrons to bond to a proton. A base may be neutral or, more commonly, have a net negative charge. Hydroxide (^-OH) is the most common Brønsted–Lowry base. The source of hydroxide anions can be a variety of metal salts, including NaOH, KOH, $Mg(OH)_2$, and $Ca(OH)_2$. Ammonia (NH_3) and water (H_2O) are both Brønsted–Lowry bases because each contains an atom with a lone pair of electrons.



Acid and Base Strength

Although all Brønsted–Lowry acids contain protons, some acids readily donate protons while others do not. Similarly, some Brønsted–Lowry bases accept a proton much more readily than others. How readily proton transfer occurs is determined by the strength of the acid and base. When a covalent acid dissolves in water, proton transfer forms H_3O^+ and an anion. The splitting apart of a covalent molecule (or an ionic compound) into individual ions is called ***dissociation***. Acids differ in their tendency to donate a proton; that is, acids differ in the extent to which they dissociate in water.

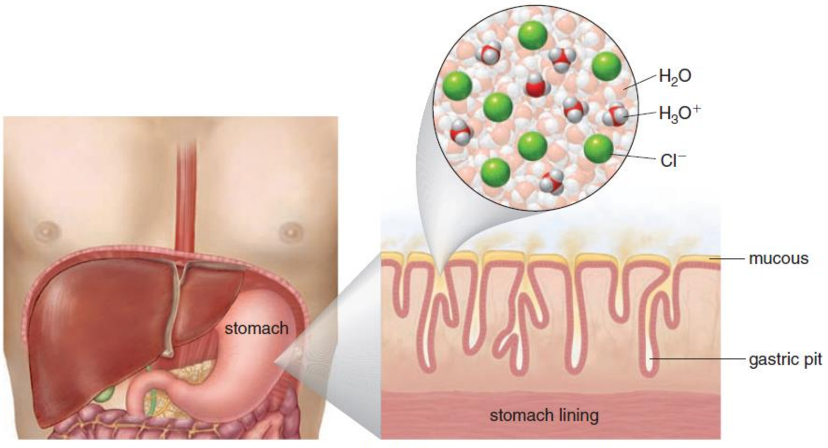
- ✓ ***A strong acid*** readily donates a proton. When a strong acid dissolves in water, essentially 100% of the acid dissociates into ions.
- ✓ ***A weak acid*** less readily donates a proton. When a weak acid dissolves in water, only a small fraction of the acid dissociates into ions.

Common strong acids include **HI, HBr, HCl, H_2SO_4,** and **HNO_3**. When each acid is dissolved in water, 100% of the acid dissociates, forming H_3O^+ and the conjugate base, as shown for HCl and H_2SO_4.

$$\underset{\text{strong acid}}{HCl(g)} + H_2O(l) \longrightarrow H_3O^+(aq) + \underset{\text{conjugate base}}{Cl^-(aq)}$$

$$\underset{\text{strong acid}}{H_2SO_4(l)} + H_2O(l) \longrightarrow H_3O^+(aq) + \underset{\text{conjugate base}}{HSO_4^-(aq)}$$

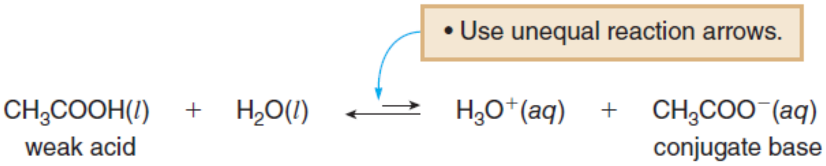
HCl, hydrochloric acid, is secreted by the stomach to digest food, and H_2SO_4, sulfuric acid, is an important industrial starting material in the synthesis of phosphate fertilizers.



The thick mucous layer protects the stomach lining.

Although HCl is a corrosive acid secreted in the stomach, a thick layer of mucous covering the stomach wall protects it from damage by the strong acid. The strong acid HCl is completely dissociated to H_3O^+ and Cl^-.

Acetic acid, CH_3COOH, is a weak acid. When acetic acid dissolves in water, only a small fraction of acetic acid molecules donate a proton to water to form H_3O^+ and the conjugate base, CH_3COO^-. The major species in solution is the undissociated acid, CH_3COOH. Two arrows that are unequal in length ($\overset{\rightarrow}{\longleftarrow}$) are used between the reactants and products to show that both are present in solution. The longer arrow points towards the reactants, since few molecules of acetic acid dissociate. Other weak acids and their conjugate bases are listed in Table 8.1.



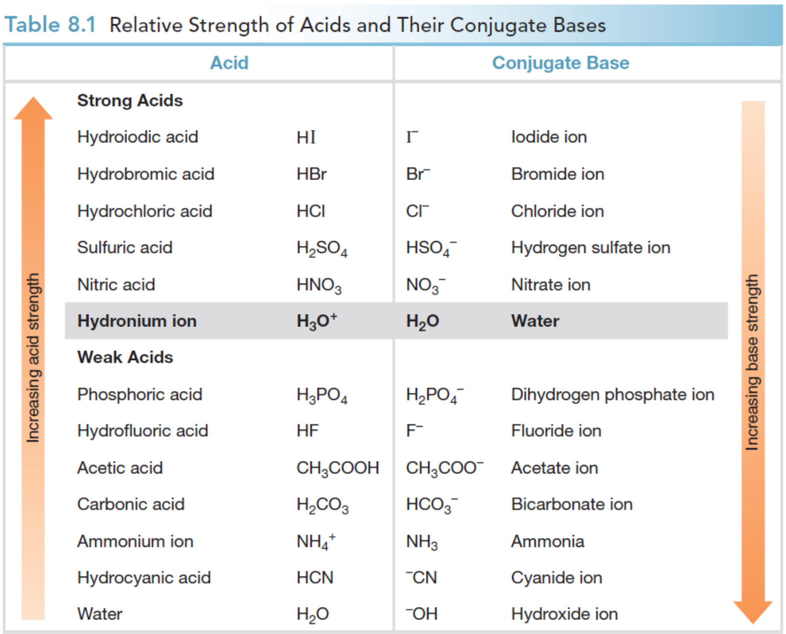
• Use unequal reaction arrows.

$$CH_3COOH(l) + H_2O(l) \overset{\rightarrow}{\longleftarrow} H_3O^+(aq) + CH_3COO^-(aq)$$

weak acid

conjugate base

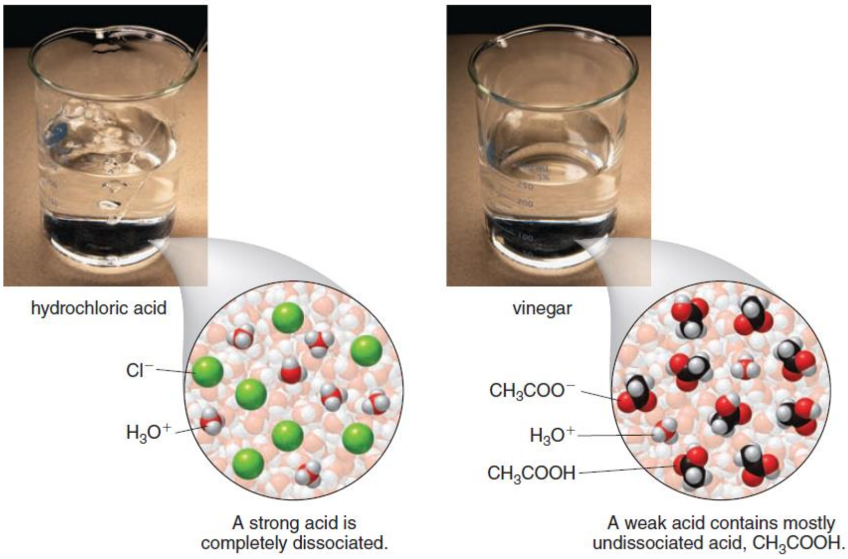
Table 8.1 Relative Strength of Acids and Their Conjugate Bases



Acid		Conjugate Base	
Strong Acids			
Hydroiodic acid	HI	I^-	Iodide ion
Hydrobromic acid	HBr	Br^-	Bromide ion
Hydrochloric acid	HCl	Cl^-	Chloride ion
Sulfuric acid	H_2SO_4	HSO_4^-	Hydrogen sulfate ion
Nitric acid	HNO_3	NO_3^-	Nitrate ion
Hydronium ion	H_3O^+	H_2O	**Water**
Weak Acids			
Phosphoric acid	H_3PO_4	$H_2PO_4^-$	Dihydrogen phosphate ion
Hydrofluoric acid	HF	F^-	Fluoride ion
Acetic acid	CH_3COOH	CH_3COO^-	Acetate ion
Carbonic acid	H_2CO_3	HCO_3^-	Bicarbonate ion
Ammonium ion	NH_4^+	NH_3	Ammonia
Hydrocyanic acid	HCN	^-CN	Cyanide ion
Water	H_2O	^-OH	Hydroxide ion

Figure 8.4

A Strong and Weak Acid Dissolved in Water



- The strong acid HCl completely dissociates into H_3O^+ and Cl^- in water.
- Vinegar contains CH_3COOH dissolved in H_2O. The weak acid CH_3COOH is only slightly dissociated into H_3O^+ and CH_3COO^-, so mostly CH_3COOH is present in solution.

Bases also differ in their ability to accept a proton.

- ☒ ***A strong base*** readily accepts a proton. When a strong base dissolve in water, 100% of the base dissociates into ions.
- ☒ ***A weak base*** less readily accepts a proton. When a weak base dissolve in water, only a small fraction of the base forms ions.

The most common strong base is hydroxide, ^-OH, used as a variety of metal salts, including NaOH and KOH. Solid NaOH dissolves in water to form solvated Na^+ cations and ^-OH anions. In contrast, when NH_3, a weak base, dissolves in water, only a small fraction of NH_3 molecules react to form NH_4^+ and ^-OH. The major species in solution is the undissociated molecule, NH_3.

$$\underset{\text{strong base}}{NaOH(s)} + H_2O(l) \longrightarrow Na^+(aq) + {}^-OH(aq)$$

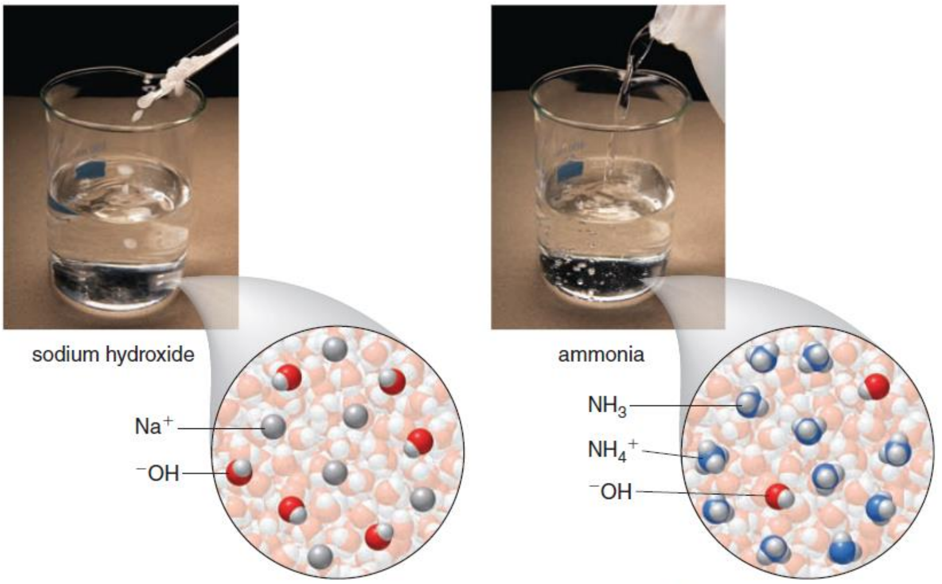
$$\underset{\text{weak base}}{NH_3(g)} + H_2O(l) \rightleftharpoons NH_4^+(aq) + {}^-OH(aq)$$

• Use unequal reaction arrows.

An inverse relationship exists between acid and base strength.

- A strong acid readily donates a proton, forming a weak conjugate base.
- A strong base readily accepts a proton, forming a weak conjugate acid.

? *Why does this inverse relationship exist?* Since a strong acid readily donates a proton, it forms a conjugate base that has little ability to accept a proton. Since a strong base readily accepts a proton, it forms a conjugate acid that tightly holds onto its proton, making it a weak acid.



A strong base is completely dissociated.

A weak base contains mostly undissociated base, NH_3.

- The strong base NaOH completely dissociates into Na^+ and ^-OH in water.
- The weak base NH_3 is only slightly dissociated into NH_4^+ and ^-OH, so mostly NH_3 is present in solution.

Common Acid–Base Reactions

1. *Reaction of Acids with Hydroxide Bases:*

- The reaction of a Brønsted–Lowry acid (HA) with the metal salt of a hydroxide base (MOH) is an example of a neutralization reaction—an acid–base reaction that produces a salt and water as products.

$$\begin{array}{ccccccc} HA_{(aq)} & + & MOH_{(aq)} & \longrightarrow & H\text{—}OH_{(l)} & + & MA_{(aq)} \\ \text{acid} & & \text{base} & & \text{water} & & \text{salt} \end{array}$$

- For example, hydrochloric acid, HCl, reacts with sodium hydroxide, NaOH, to form water and sodium chloride, NaCl.

$$\begin{array}{ccccccc} HCl_{(aq)} & + & NaOH_{(aq)} & \longrightarrow & H\text{—}OH_{(l)} & + & NaCl_{(aq)} \\ \text{acid} & & \text{base} & & \text{water} & & \text{salt} \end{array}$$

- The acid HA donates a proton (H^+) to the ^-OH base to form H_2O.
- The anion A^- from the acid combines with the cation M^+ from the base to form the salt MA.

The important reacting species in this reaction are H^+ from the acid HCl and ^-OH from the base NaOH. To more clearly see the acid–base reaction, we can write an equation that contains only the species that are actually involved in the reaction. Such an equation is called a ***net ionic equation***.

❖ *A net ionic equation contains only the species involved in a reaction.*

To write a net ionic equation for an acid–base reaction, we first write the acid, base, and salt as individual ions in solution. This process is simplified if we use H^+ (not H_3O^+) as the reacting species of the acid, since it is the H^+ ion that is transferred to the base. The reaction of HCl with NaOH using individual ions is then drawn as:

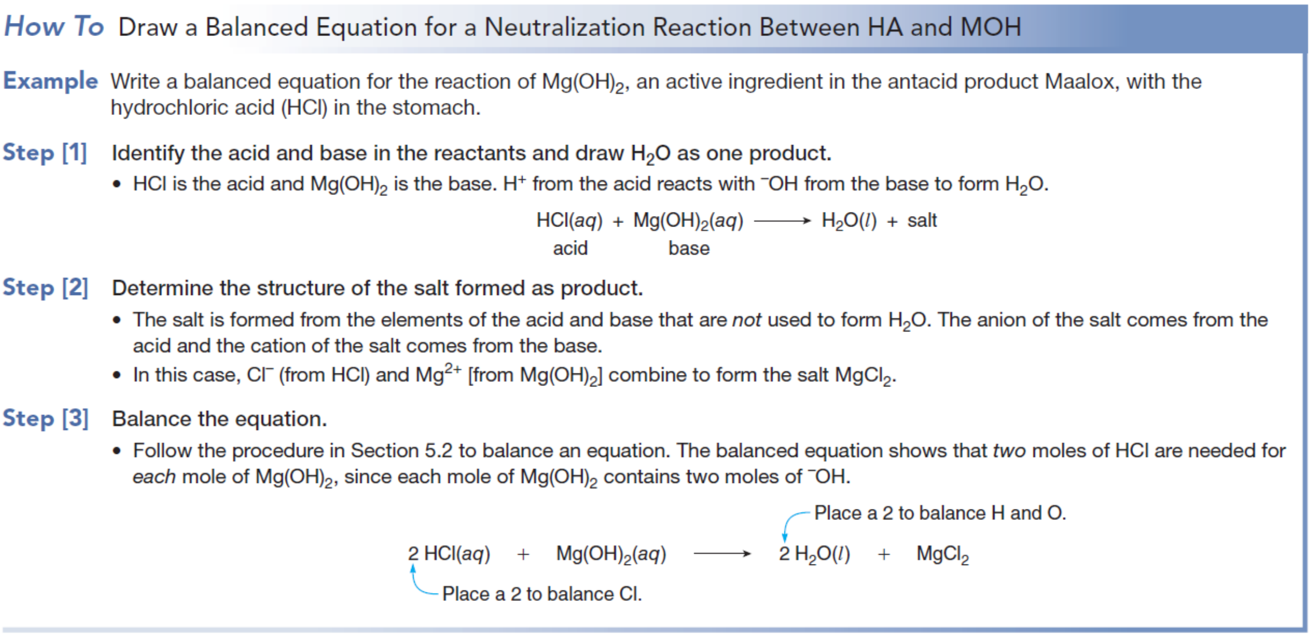
$$H^+(aq) + Cl^-(aq) + Na^+(aq) + {}^-OH(aq) \longrightarrow H{-}OH(l) + Na^+(aq) + Cl^-(aq)$$

Writing the equation in this manner shows that the Na^+ and Cl^- ions are unchanged in the reaction. Ions that appear on both sides of an equation but undergo no change in a reaction are called ***spectator ions***. Removing the spectator ions from the equation gives the net ionic equation.

$$H^+(aq) + \underbrace{\cancel{Cl}^-(aq) + \cancel{Na}^+(aq)}_{\text{Omit the spectator ions.}} + {}^-OH(aq) \longrightarrow H{-}OH(l) + \underbrace{\cancel{Na}^+(aq) + \cancel{Cl}^-(aq)}_{\text{Omit the spectator ions.}}$$

Net ionic equation $H^+(aq) + {}^-OH(aq) \longrightarrow H{-}OH(l)$

- Whenever a strong acid and strong base react, the net ionic equation is always the same H^+ reacts with ^-OH to form H_2O.



How To Draw a Balanced Equation for a Neutralization Reaction Between HA and MOH

Example Write a balanced equation for the reaction of $Mg(OH)_2$, an active ingredient in the antacid product Maalox, with the hydrochloric acid (HCl) in the stomach.

Step [1] Identify the acid and base in the reactants and draw H_2O as one product.

- HCl is the acid and $Mg(OH)_2$ is the base. H^+ from the acid reacts with ^-OH from the base to form H_2O.

$$\underset{\text{acid}}{HCl(aq)} + \underset{\text{base}}{Mg(OH)_2(aq)} \longrightarrow H_2O(l) + \text{salt}$$

Step [2] Determine the structure of the salt formed as product.

- The salt is formed from the elements of the acid and base that are *not* used to form H_2O. The anion of the salt comes from the acid and the cation of the salt comes from the base.
- In this case, Cl^- (from HCl) and Mg^{2+} [from $Mg(OH)_2$] combine to form the salt $MgCl_2$.

Step [3] Balance the equation.

- Follow the procedure in Section 5.2 to balance an equation. The balanced equation shows that *two* moles of HCl are needed for *each* mole of $Mg(OH)_2$, since each mole of $Mg(OH)_2$ contains two moles of ^-OH.

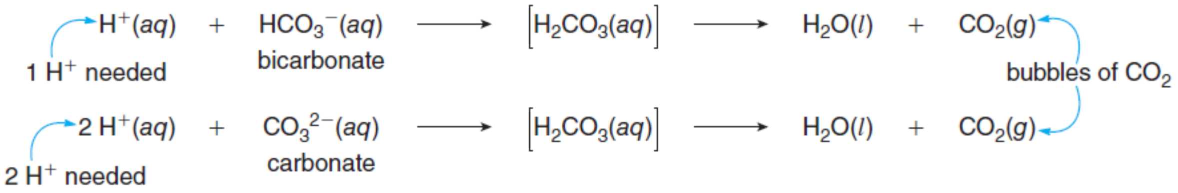
Place a 2 to balance H and O.

$$2\,HCl(aq) + Mg(OH)_2(aq) \longrightarrow 2\,H_2O(l) + MgCl_2$$

Place a 2 to balance Cl.

2. *Reaction of Acids with Bicarbonate and Carbonate*

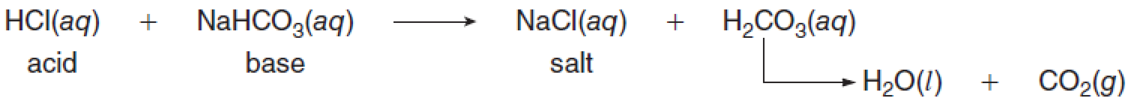
Acids react with the bases bicarbonate (HCO_3^-) and carbonate (CO_3^{2-}). A bicarbonate base reacts with one proton to form carbonic acid, H_2CO_3. A carbonate base reacts with two protons. The carbonic acid formed in these reactions is unstable and decomposes to form CO_2 and H_2O. Thus, when an acid reacts with either base, bubbles of CO_2 gas are given off.



$$H^+(aq) + HCO_3^-(aq) \longrightarrow [H_2CO_3(aq)] \longrightarrow H_2O(l) + CO_2(g)$$

$$2\,H^+(aq) + CO_3^{2-}(aq) \longrightarrow [H_2CO_3(aq)] \longrightarrow H_2O(l) + CO_2(g)$$

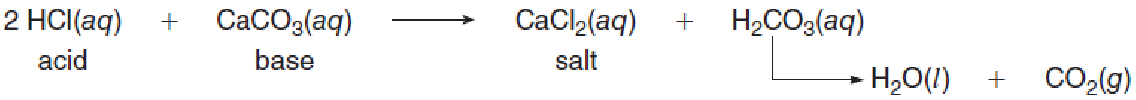
Sodium bicarbonate ($NaHCO_3$), an ingredient in the over-the-counter antacid Alka-Seltzer, is the metal salt of a bicarbonate base that reacts with excess stomach acid, releasing CO_2. Like the neutralization reactions in a salt, NaCl, is formed in which the cation (Na^+) comes from the base and the anion (Cl^-) comes from the acid.



$$\underset{\text{acid}}{HCl(aq)} + \underset{\text{base}}{NaHCO_3(aq)} \longrightarrow \underset{\text{salt}}{NaCl(aq)} + H_2CO_3(aq)$$

$$H_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g)$$

Calcium carbonate ($CaCO_3$), a calcium supplement and antacid in Tums, also reacts with excess stomach acid with release of CO_2. Since each carbonate ion reacts with two protons, the balanced equation shows a 2:1 ratio of HCl to $CaCO_3$.



$$\underset{\text{acid}}{2\,HCl(aq)} + \underset{\text{base}}{CaCO_3(aq)} \longrightarrow \underset{\text{salt}}{CaCl_2(aq)} + H_2CO_3(aq)$$

$$H_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g)$$

The pH Scale

- Knowing the hydronium ion concentration is necessary in many different instances. The blood must have an H_3O^+ concentration in a very narrow range for an individual's good health.
- ✓ Plants thrive in soil that is not too acidic or too basic.
- ✓ The H_3O^+ concentration in a swimming pool must be measured and adjusted to keep the water clean and free from bacteria and algae.

Since values for the hydronium ion concentration are very small, with negative powers of ten, the pH scale is used to more conveniently report $[H_3O^+]$. The pH of a solution is a number generally between 0 and 14, defined in terms of the logarithm (log) of the H_3O^+ concentration.

$$\boxed{\mathrm{pH} = -\log\,[H_3O^+]}$$

A logarithm is an exponent of a power of ten.

The log is the exponent.

$$\log(10^{5}) = 5 \qquad \log(10^{-10}) = -10 \qquad \log(0.001) = \log(10^{-3}) = -3$$

The log is the exponent.

Convert to scientific notation.

In calculating pH, first consider an H_3O^+ concentration that has a coefficient of one when the number is written in scientific notation. For example, the value of $[H_3O^+]$ in apple juice is about 1×10^{-4}, or 10^{-4} written without the coefficient. The pH of this solution is calculated as follows:

$$\mathrm{pH} = -\log\,[H_3O^+] = -\log(10^{-4}) = -(-4) = 4$$

pH of apple juice

Since pH is defined as the negative logarithm of $[H_3O^+]$ and these concentrations have negative exponents (10^{-x}), pH values are positive numbers.

Whether a solution is acidic, neutral, or basic can now be defined in terms of its pH.

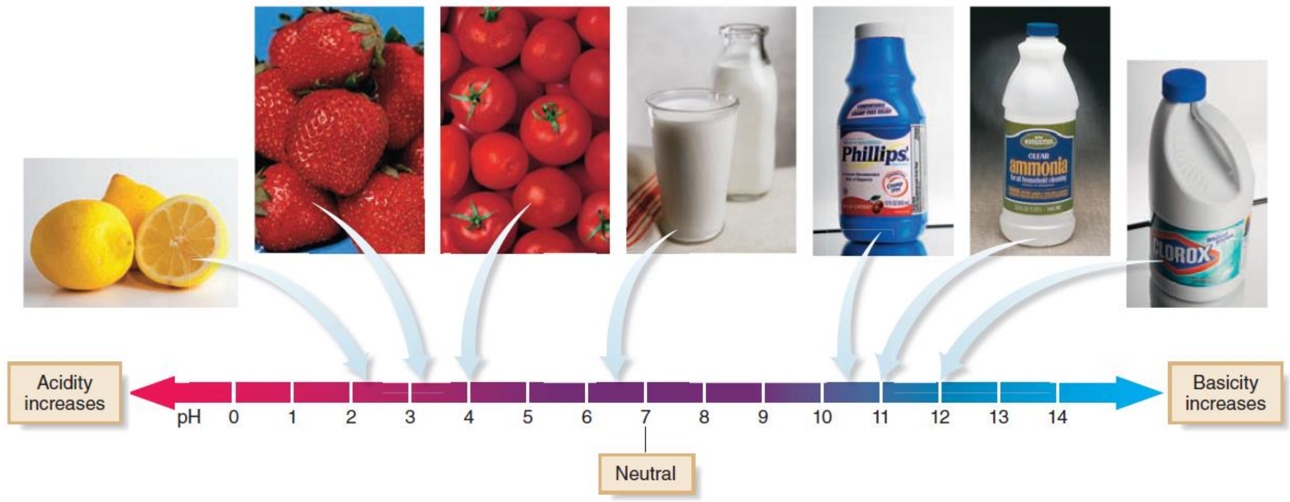
- **Acidic solution:** $\mathbf{pH < 7\ [H_3O^+] > 1 \times 10^{-7}}$
- **Neutral solution:** $\mathbf{pH = 7\ [H_3O^+] = 1 \times 10^{-7}}$
- **Basic solution:** $\mathbf{pH > 7\ [H_3O^+] < 1 \times 10^{-7}}$

? Note the relationship between $[H_3O^+]$ and pH.

The *lower* the pH, the *higher* the concentration of H_3O^+.

The pH of a solution can be measured using a pH meter as shown in Figure 8.6. Approximate pH values are determined using pH paper or indicators that turn different colors depending on the pH of the solution. The pH of various substances is shown in Figure 8.7.

Figure 8.7 The pH of Some Common Substances



The pH of many fruits is less than 7, making them acidic. Many cleaning agents, such as household ammonia and bleach, are basic (pH > 7).

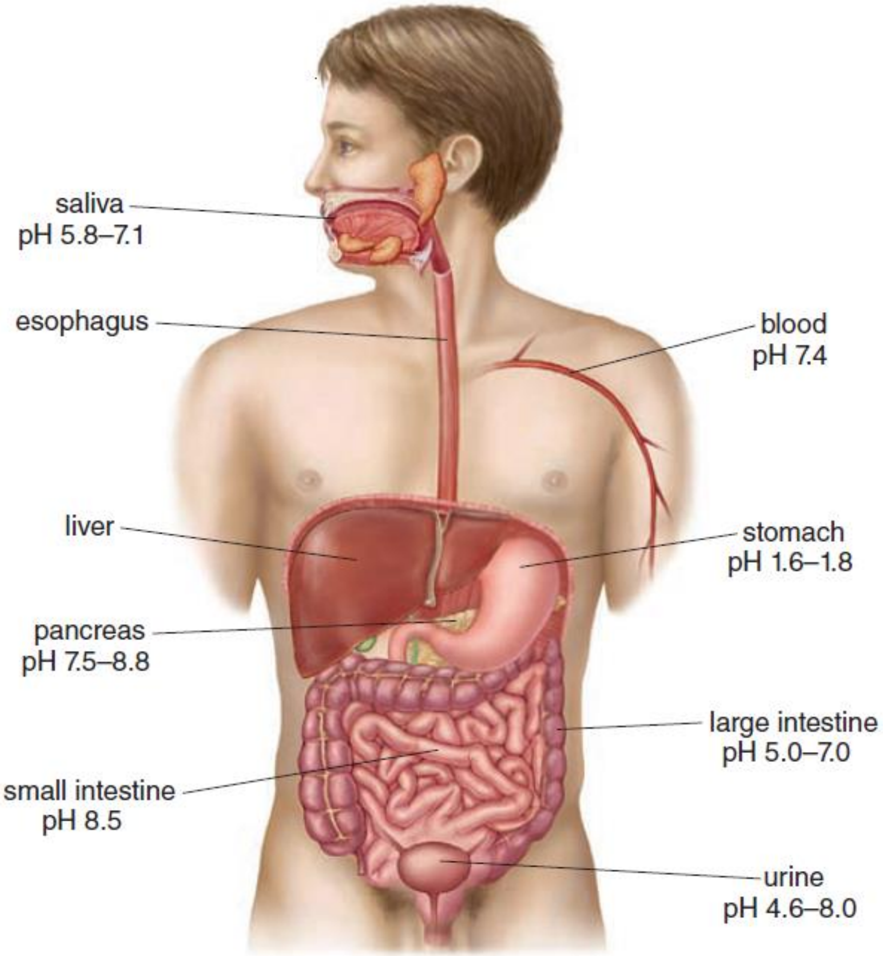
The pH of Body Fluids

The human body contains fluids that vary in pH as shown in Figure 8.8.

While saliva is slightly acidic. The gastric juice in the stomach has the lowest pH found in the body. The strongly acidic environment of the stomach aids in the digestion of food. It also kills many types of bacteria that might be inadvertently consumed along with food and drink.

When food leaves the stomach, it passes to the basic environment of the small intestines. Bases in the small intestines react with acid from the stomach.

The pH of some body fluids must occupy a very narrow range. For example, a healthy individual has a blood pH in the range of 7.35–7.45. Maintaining this pH is accomplished by a complex mechanism described in Section 8. The pH of other fluids can be more variable. Urine has a pH anywhere from 4.6–8.0, depending on an individual's recent diet and exercise.

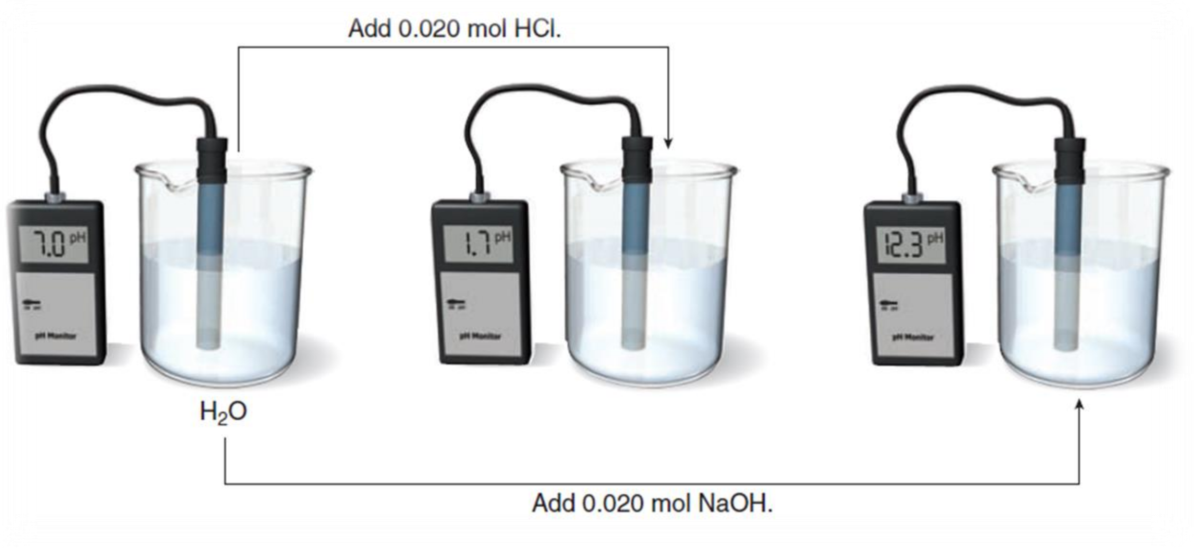


Buffers

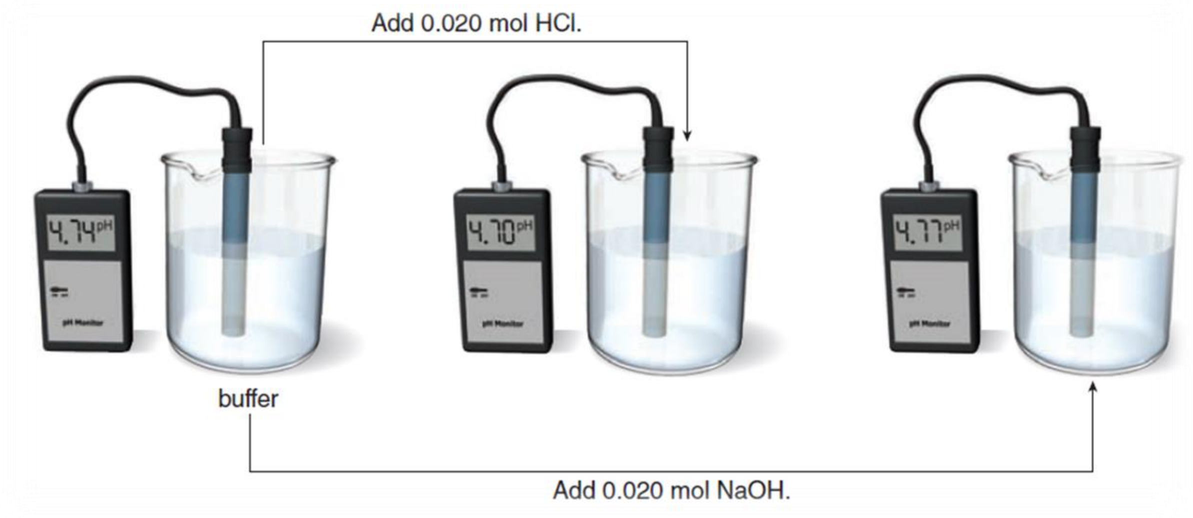
A buffer is a solution whose pH changes very little when acid or base is added. Most buffers are solutions composed of approximately equal amounts of a weak acid and the salt of its conjugate base.

- The weak acid of the buffer reacts with added base, ^{-}OH.
- The conjugate base of the buffer reacts with added acid, H_3O^+.

- The effect of a buffer can be illustrated by comparing the pH change that occurs when a small amount of strong acid or strong base is added to water, with the pH change that occurs when the same amount of strong acid or strong base is added to a buffer.
- When 0.020 mol of HCl is added to 1.0 L of water, the pH changes from 7 to 1.7, and when 0.020 mol of NaOH is added to 1.0 L of water, the pH changes from 7 to 12.3. In this example, addition of a small quantity of a strong acid or strong base to neutral water changes the pH by over 5 pH units.

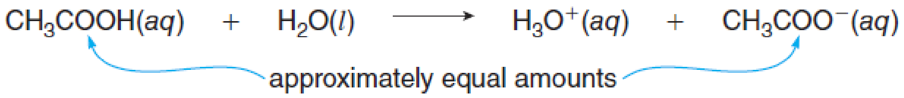


- In contrast, a buffer prepared from 0.50 M acetic acid (CH_3COOH) and 0.50 M sodium acetate ($NaCH_3COO$) has a pH of 4.74. Addition of the same quantity of acid, 0.020 mol HCl, changes the pH to 4.70, and addition of the same quantity of base, 0.020 mol of NaOH, changes the pH to 4.77. In this example, the change of pH in the presence of the buffer is no more than 0.04 pH units!



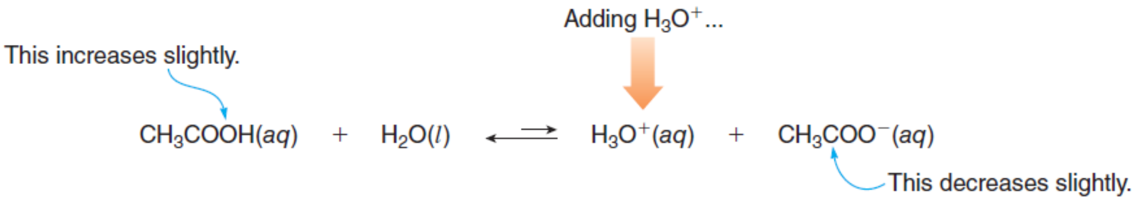
? ***Why is a buffer able to absorb acid or base with very little pH change?***

- Let's use as an example a buffer that contains equal concentrations of acetic acid (CH_3COOH), and the sodium salt of its conjugate base, sodium acetate ($NaCH_3COO$). CH_3COOH is a weak acid, so when it dissolves in water, only a small fraction dissociates to form its conjugate base CH_3COO^-. In the buffer solution, however, the sodium acetate provides an equal amount of the conjugate base.



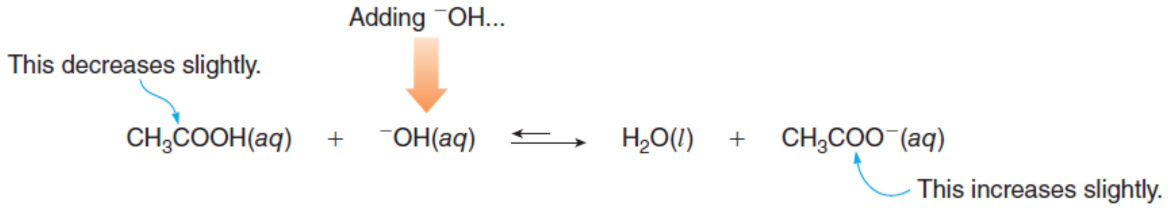
$$CH_3COOH(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + CH_3COO^-(aq)$$

- Suppose a small amount of strong acid is added to the buffer. Added H_3O^+ reacts with CH_3COO^- to form CH_3COOH, so that $[CH_3COO^-]$ decreases slightly and $[CH_3COOH]$ increases slightly, but the $[H_3O^+]$ and therefore the pH change only slightly.



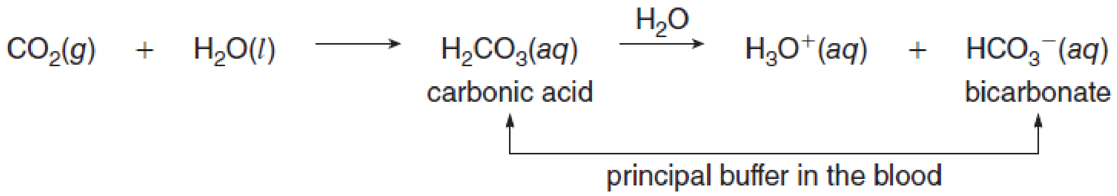
$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

- On the other hand, if a small amount of strong base is added to the buffer, ^-OH reacts with CH_3COOH to form CH_3COO^-, so that $[CH_3COOH]$ decreases slightly and $[CH_3COO^-]$ increases slightly but the $[H_3O^+]$ and therefore the pH change only slightly.

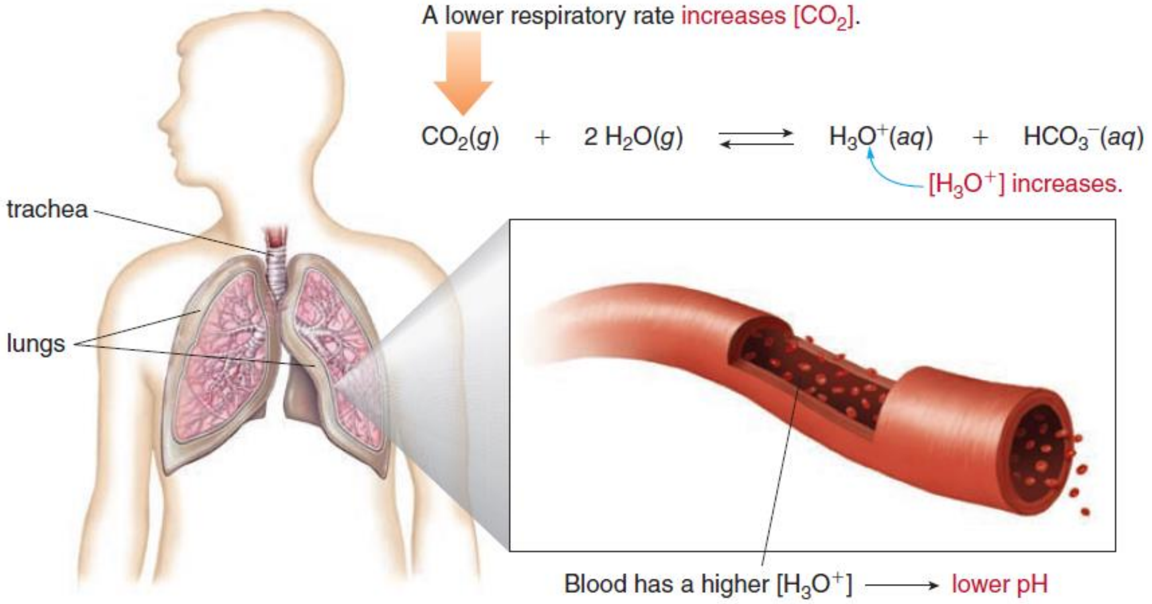


Buffers in the Blood

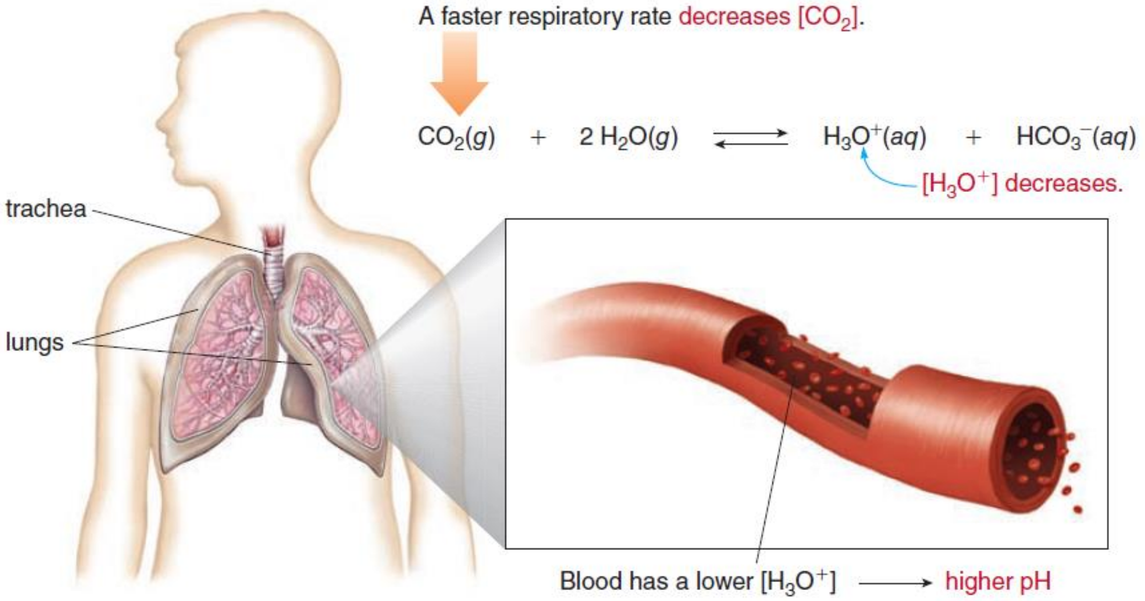
- *The normal blood pH* of a healthy individual is in the range of 7.35 to 7.45. A pH above or below this range is generally indicative of an imbalance in *respiratory* or *metabolic* processes. The body is able to maintain a very stable pH because the blood and other tissues are buffered. The principal buffer in the blood is carbonic acid/bicarbonate (H_2CO_3/HCO_3^-).
- In examining the carbonic acid/bicarbonate buffer system in the blood, two reactions are important. First of all, carbonic acid (H_2CO_3) is formed from CO_2 dissolved in the bloodstream. Second, since carbonic acid is a weak acid, it is also dissociated in water to form its conjugate base, bicarbonate (HCO_3^-). Bicarbonate is also generated in the kidneys.



- CO_2 is constantly produced by metabolic processes in the body and then transported to the lungs to be eliminated. Increasing or decreasing the level of dissolved CO_2 affects the pH of the blood.
- A higher-than-normal CO_2 concentration increases the H_3O^+ concentration and lowers the pH. **Respiratory acidosis** results when the body fails to eliminate adequate amounts of CO_2 through the lungs. This may occur in patients with advanced lung disease or respiratory failure.

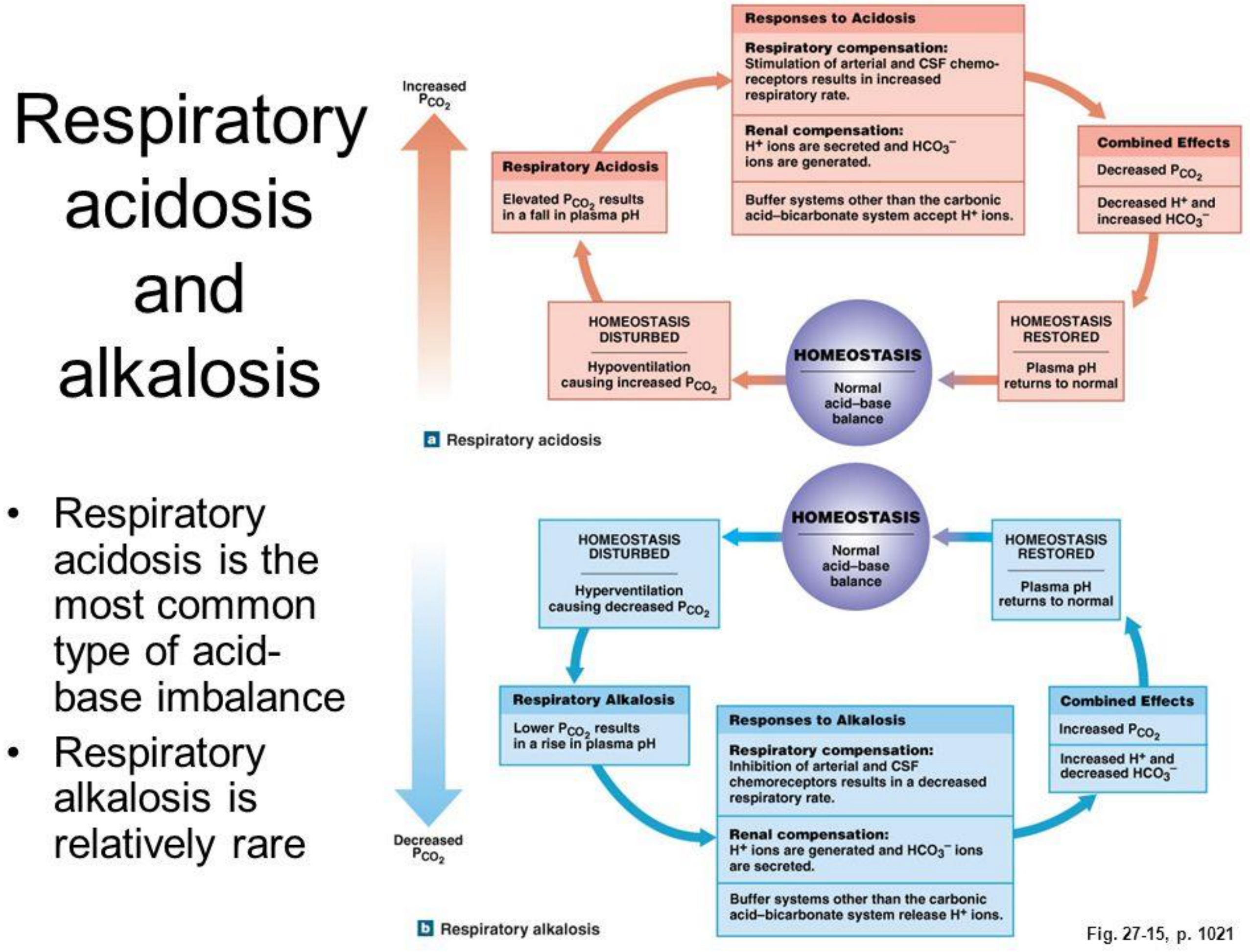


- A lower-than-normal CO_2 concentration decreases the H_3O^+ concentration and raises the pH. **Respiratory alkalosis** is caused by **<u>hyperventilation</u>**, very rapid breathing that occurs when an individual experiences excitement or panic.



- The pH of the blood may also be altered when the metabolic processes of the body are not in balance.
- **Metabolic acidosis** results when excessive amounts of acid are produced and the blood pH falls. This may be observed in patients with severe infections (sepsis). It may also occur in poorly controlled diabetes.

- **Metabolic alkalosis** may occur when recurrent vomiting decreases the amount of acid in the stomach, thus causing a rise in pH.



Respiratory acidosis and alkalosis

- Respiratory acidosis is the most common type of acid-base imbalance
- Respiratory alkalosis is relatively rare