

Lecture Presentation

Chapter 9

Acids, Bases and Buffers in the Body

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Outline

- 9.1 Acids and Bases—Definitions
- 9.2 Strong Acids and Bases
- 9.3 Chemical Equilibrium
- 9.4 Weak Acids and Bases
- 9.5 pH and the pH Scale
- 9.6 pK_a
- 9.7 Amino Acids—Common Biological Weak Acids
- 9.8 Buffers and Blood—The Bicarbonate Buffer System

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9.1 Acids and Bases—Definitions

Acids

- Swedish chemist Svante Arrhenius described acids as substances that dissociate, producing hydrogen ions (H^+) when dissolved in water.
- The presence of hydrogen ions gives acids their sour taste and allows acids to corrode some metals.
- In the early twentieth century Johannes Brønsted and Thomas Lowry, working independently, expanded the definition of an acid: an acid is a compound that *donates* a proton.

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9.1 Acids and Bases—Definitions

- A proton is a subatomic particle and H^+ is an ion.
- Hydrogen atoms contain one proton and one electron.
- A hydrogen ion—a hydrogen atom that has lost its electron—and a proton are one and the same.
- In an aqueous solution, the partial negative charge on the oxygen atom in water is strongly attracted to the positive charge of a proton.
- The proton and the oxygen atom in water form a covalent bond, creating a **hydronium ion, H_3O^+** .

$$\begin{array}{c} \text{H}-\ddot{\text{O}}: \\ | \\ \text{H} \end{array} + \text{H}^+ \longrightarrow \begin{array}{c} \text{H}-\ddot{\text{O}}-\text{H} \\ | \\ \text{H} \end{array}$$

Water Proton Hydronium ion

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9.1 Acids and Bases—Definitions

Bases

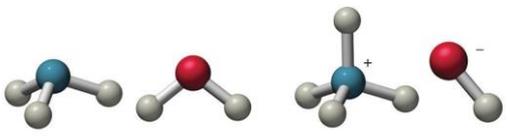
- According to Arrhenius, **bases** are ionic compounds that, when dissolved in water, dissociate to form a metal ion and a hydroxide ion (OH^-).
- Most Arrhenius bases are formed from Group 1A and 2A metals, such as NaOH, KOH, LiOH, and $Ca(OH)_2$.
- Hydroxide bases are characterized by a bitter taste and a slippery feel.

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9.1 Acids and Bases—Definitions

Bases

- A Brønsted–Lowry base is a compound that *accepts* a proton.



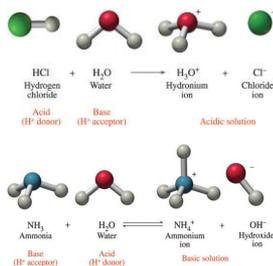
$$\begin{array}{c} \text{NH}_3 \\ \text{Ammonia} \end{array} + \begin{array}{c} \text{H}_2\text{O} \\ \text{Water} \end{array} \rightleftharpoons \begin{array}{c} \text{NH}_4^+ \\ \text{Ammonium ion} \end{array} + \begin{array}{c} \text{OH}^- \\ \text{Hydroxide ion} \end{array}$$

Base (H⁺ acceptor) Acid (H⁺ donor) Basic solution

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9.1 Acids and Bases—Definitions

- Water can act as an acid or a base by donating or accepting a proton.



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9.2 Strong Acids and Bases

- Strong acids break up completely into ions when placed in water, forming hydronium ions and anions.

TABLE 9.1 Six Common Strong Acids

Acid Name	Formula
Perchloric acid	HClO ₄
Sulfuric acid*	H ₂ SO ₄
Hydroiodic acid	HI
Hydrobromic acid	HBr
Hydrochloric acid	HCl
Nitric acid	HNO ₃

*Only the first proton is 100% dissociated. The product, HSO₄⁻, is a weak acid.

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9.2 Strong Acids and Bases

- Acids that only partially dissociate (~5%) are **weak acids**.
- One example is acetic acid (CH₃COOH), which is the main component of vinegar.



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9.2 Strong Acids and Bases

- Strong bases, like NaOH (also known as lye), are used in household products such as oven cleaners and drain openers.
- Arrhenius bases such as LiOH, KOH, NaOH, and Ca(OH)₂ are **strong bases** that dissociate completely (100%) in water to give an aqueous solution of a metal ion and a hydroxide ion.
- Bases that only partially dissociate (~5%) are **weak bases**.
- Many common weak bases contain ammonia (NH₃).

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9.2 Strong Acids and Bases

Neutralization

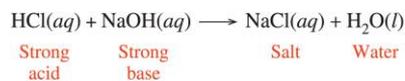
- What happens when a strong acid and strong base are mixed?
- Because both completely dissociate, the water contains metal cations and nonmetal anions as well as hydronium and hydroxide ions.
- The hydronium and hydroxide ions combine to form water molecules, producing a lot of heat.
- The metal cations and the anions remain in solution. If the water were removed, an ionic compound would remain.

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9.2 Strong Acids and Bases

Neutralization

- The reaction of a strong acid and strong base always produces water and a **salt**.
- This reaction is called **neutralization**.



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9.2 Strong Acids and Bases

Completing Neutralization Reactions

- **Step 1: Form the products.**
 - The products will always be (a) a salt and (b) H₂O.
 - The salt produced must be a neutral ionic compound.
- **Step 2: Balance the chemical equation.**
 - This is done by adding coefficients in front of the product or reactant compounds where appropriate.
 - The same number of atoms must appear in both the reactants and products.

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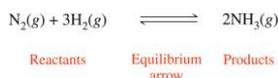
9.2 Strong Acids and Bases

Antacids

- Antacids are used to neutralize excess stomach acid (HCl).
- Some antacids are mixtures of aluminum hydroxide and magnesium hydroxide.
 - Aluminum hydroxide produces constipation and binds phosphate in the intestinal tract, which may cause weakness and loss of appetite.
 - Magnesium hydroxide has a laxative effect.
- When carbonates are used to neutralize acid, the reaction produces a salt, water, and carbon dioxide gas.
- When calcium carbonate is used, about 10% of the calcium is absorbed into the bloodstream.
- Calcium carbonate is not recommended for people who have peptic ulcers or a tendency to form kidney stones.
- Sodium bicarbonate can affect the acidity level of the blood and elevate sodium levels in the body fluids. It is also not recommended in the treatment of peptic ulcers.

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9.3 Chemical Equilibrium



- The generation of ammonia is a reversible reaction.
- Once ammonia is formed, the reaction will reverse, re-forming nitrogen and hydrogen.
- Eventually, the rate of the formation of ammonia and the rate of reformation of nitrogen and hydrogen gases become equal.
- This balance of the rates of the reactions is **chemical equilibrium**.
- An equilibrium arrow is used to indicate that both the forward and reverse reactions take place simultaneously.
- Because the rates of the reactions are equal, there is no net change in amounts.

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9.3 Chemical Equilibrium

- If we measure the concentrations of ammonia, nitrogen, and hydrogen present, the fraction of products to reactants would be a constant.
- This is the **equilibrium constant**, K , and it is a characteristic of equilibrium reactions at a given temperature.
- The equilibrium constant, K , is equal to the molar concentration of the products divided by that of the reactants.

$$K = \frac{[\text{products}]}{[\text{reactants}]}$$

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9.3 Chemical Equilibrium

- If there is more than one reactant or product, the concentrations are multiplied together.
- The superscripts in the expression come from the coefficients (number of moles of each) found in the *balanced* chemical equation.
- For an equilibrium reaction of the form

$$aA + bB \rightleftharpoons cC + dD$$
 the general equilibrium expression is given as

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

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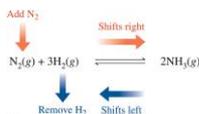
9.3 Chemical Equilibrium

- Only substances whose concentrations change appear in an equilibrium expression.
- Solids (s) and pure liquids (l) have constant concentrations so do not appear in the equilibrium expression.
- If K is equal to 1, the ratio of products:reactants is 1:1 or $[\text{products}] = [\text{reactants}]$.
- A value of K greater than (>) 1 indicates that $[\text{products}] > [\text{reactants}]$.
- A value of K less than (<) 1 indicates that $[\text{products}] < [\text{reactants}]$.

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9.3 Chemical Equilibrium

- According to **Le Châtelier's principle**, applying stress to the equilibrium will cause the rate of the forward or reverse reaction to change to offset the stress and regain equilibrium.

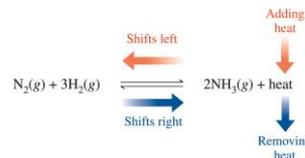


- If one side of the reaction gains a substance, the reaction shifts to the other side to regain its equilibrium.
- If one side of the reaction loses a substance, the reaction will shift toward that side in order to regain its equilibrium.

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9.3 Chemical Equilibrium

- If heat is a product of the reaction, if the temperature of the reaction is raised (heat added), the rate of the reverse reaction increases to offset the stress of adding heat. This causes the equilibrium to shift to the left.
- If the reaction were cooled down (heat removed), the rate of the forward reaction would increase to replenish the heat produced, shifting the equilibrium to the right.



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9.3 Chemical Equilibrium

TABLE 9.4 Effects of Changes on Equilibrium

Factor	Change (stress)	Reaction Shifts toward
Concentration	Add reactant	Right
	Remove reactant	Left
	Add product	Left
	Remove product	Right
Temperature	Raise temperature of endothermic reaction	Right
	Lower temperature of endothermic reaction	Left
	Raise temperature of exothermic reaction	Left
	Lower temperature of exothermic reaction	Right

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9.4 Weak Acids and Bases

- All the principles of equilibrium apply to weak acids and bases because weak acids and bases only partially dissociate, establishing an equilibrium in aqueous solution.

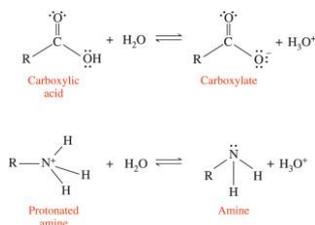


$$K = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

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9.4 Weak Acids and Bases

- All weak acids dissociate by donating a proton to form a hydronium ion. Each weak acid has an acid dissociation constant, or K_a .
- The larger the K_a , the stronger the acid.



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9.4 Weak Acids and Bases

TABLE 9.5 K_a Values for Substances Acting as Weak Acids (25 °C)

Name	Formula	K_a
Hydrogen sulfate ion	HSO_4^-	1.0×10^{-2}
Phosphoric acid	H_3PO_4	7.5×10^{-3}
Hydrofluoric acid	HF	6.5×10^{-4}
Nitrous acid	HNO_2	4.5×10^{-4}
Formic acid	HCOOH	1.8×10^{-4}
Acetic acid	CH_3COOH	1.75×10^{-5}
Carbonic acid	H_2CO_3	4.5×10^{-7}
Water	H_2O	1.0×10^{-7}
Dihydrogen phosphate ion	H_2PO_4^-	6.6×10^{-8}
Ammonium ion	NH_4^+	6.3×10^{-10}
Hydrocyanic acid	HCN	6.2×10^{-10}
Bicarbonate ion	HCO_3^-	4.8×10^{-11}
Hydrogen phosphate ion	HPO_4^{2-}	1.0×10^{-12}



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9.4 Weak Acids and Bases

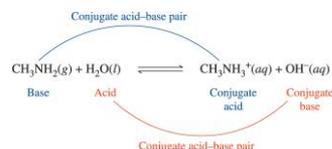
- According to the Brønsted–Lowry theory, the reaction between an acid and base involves a proton transfer.
- If a weak acid is mixed with water, water will act as a base.
- Consider the dissociation of acetic acid, CH_3COOH :



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9.4 Weak Acids and Bases

- The acid CH_3COOH donates a proton to a molecule of water, forming a hydronium ion, H_3O^+ .
- What remains of the acid after the donation, CH_3COO^- , an acetate anion, is the **conjugate base** of CH_3COOH .
- In the reverse reaction, the CH_3COO^- acts as a base and accepts the proton from the hydronium ion.
- Molecules or ions related by the loss or gain of one H^+ are **conjugate acid–base pairs**.



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9.4 Weak Acids and Bases

Rules for writing products of a weak acid–base equation

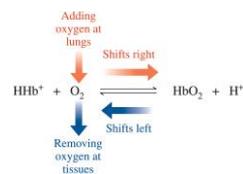
- The conjugates will always appear on the product side of a chemical equilibrium.
- When a proton is donated, the conjugate base formed will have one more negative charge than the acid from which it was formed.
- When a proton is accepted, the conjugate acid formed will have one more positive charge than the base from which it was formed.
- The total charge of the reactants equals the total charge of the products.

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9.4 Weak Acids and Bases

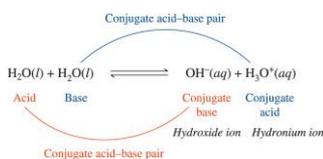
Weak Acids, Oxygen Transport, and Le Châtelier's Principle

- The binding and release of oxygen in the body are controlled by weak acid–base equilibria.
- The protein hemoglobin (Hb) carries oxygen from the lungs to the tissues via the bloodstream.
- Both protons (H^+) and oxygen (O_2) bind to hemoglobin, but with opposite affinity.
- If the concentrations of acid or oxygen change, the equilibrium will shift.



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9.5 pH and the pH Scale



- This reaction, which is always present in water, is the **autoionization of water**.
- The equilibrium constant expression for water, K_w (keeping in mind that pure liquid water will not appear), is

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

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9.5 pH and the pH Scale

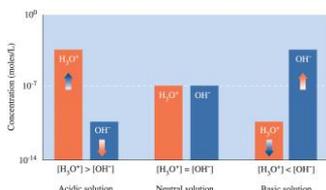
- In pure water, both the hydroxide and hydronium ion are being formed equally, so in pure water $[\text{H}_3\text{O}^+] = [\text{OH}^-]$.
- This is referred to as the autoionization of water.
- At 25°C , both these values are $1 \times 10^{-7} \text{ M}$, and the solution is neutral.

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

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9.5 pH and the pH Scale

- If an acid is added to water, there is an increase in $[H_3O^+]$ and a decrease in $[OH^-]$, which makes the solution acidic.
- If base is added, $[OH^-]$ increases and $[H_3O^+]$ decreases, making a basic solution.
- Most aqueous solutions are *not* neutral and have unequal concentrations of H_3O^+ and OH^- .



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9.5 pH and the pH Scale

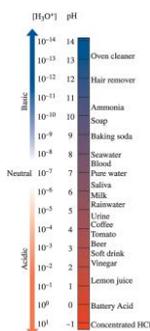
- The amount of H_3O^+ in an aqueous solution defines the acidity of the solution.
- Concentrations of H_3O^+ usually range from about 1 M to 1×10^{-14} M (really close to zero).
- Because of this extremely wide range, it is more useful to compare $[H_3O^+]$ by the "log" function because it gives a set of numbers that usually falls between 0 and 14.
- This set of numbers describes the **pH** scale. pH can be determined from the $[H_3O^+]$ as

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

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9.5 pH and the pH Scale

- Living things prefer a constant pH.
- Normal blood pH is strictly regulated between 7.35 and 7.45.
- The pH of a solution is commonly measured either electronically by using an instrument called a pH meter or visually using pH paper.



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9.5 pH and the pH Scale

Calculating pH

- Given a 0.050 M HCl solution, it is straightforward to calculate pH.
- Because strong acids fully ionize in solution, $[HCl] = [H_3O^+]$.
- pH can be directly calculated for a strong acid:

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

$$pH = -\log(0.050)$$

$$pH = 1.3$$

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9.5 pH and the pH Scale

Calculating $[H_3O^+]$

- If the pH of a solution is 3.00, the corresponding $[H_3O^+]$ can be found by multiplying by -1 and finding the inverse log.
- Inverse log is 10^x , so this can also be solved directly.

$$[H_3O^+] = \text{INV log}(-3.00) = 0.00100$$

or

$$[H_3O^+] = 10^{-3.00} = 0.00100$$

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9.5 pH and the pH Scale

Calculating pH, pOH, H^+ , OH^- :

$$pH + pOH = 14 \rightarrow pH = 14 - pOH, \quad pOH = 14 - pH$$

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

$$pOH = -\log[OH^-]$$

$$[H^+][OH^-] = 1 \times 10^{-14} \rightarrow [H^+] = \frac{1 \times 10^{-14}}{[OH^-]}, \quad [OH^-] = \frac{1 \times 10^{-14}}{[H^+]}$$

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

$$[OH^-] = 10^{-pOH}$$

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Condition	H ⁺	pH	OH ⁻	pOH
Neutral	1x10 ⁻⁷	7	1x10 ⁻⁷	7
Acidic	> 1x10 ⁻⁷	< 7	< 1x10 ⁻⁷	> 7
Basic	< 1x10 ⁻⁷	> 7	> 1x10 ⁻⁷	< 7

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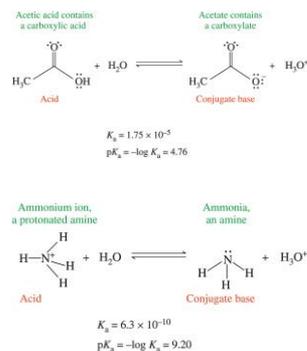
9.5 pH and the pH Scale

- The amount of H₃O⁺ in an aqueous solution determines acidity.
 - Water ionizes slightly, producing hydronium (H₃O⁺) and hydroxide (OH⁻) ions. An excess of H₃O⁺ in a solution makes a solution acidic. An excess of OH⁻ makes a solution basic. When the concentrations of hydronium and hydroxide ions are equal, a solution is neutral.
 - The pH scale is a measure of acidity with values typically falling between 0-14. Neutral solutions have a pH value of 7, acidic solutions have values less than 7, and basic solutions have pH values higher than 7. The pH is mathematically related to the concentration of H₃O⁺ by the following equation: **pH = -log[H₃O⁺]**.
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9.6 pK_a

- Acid strength can be compared by comparing K_a values.
 - Values closer to 1 indicate a stronger acid.
 - As seen with the pH scale, it is easier to compare whole numbers than those in scientific notation.
 - To make this comparison for acid strength, we can use pK_a values, which range from 0 to about 60.
 - The smaller the pK_a, the stronger the acid.
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9.6 pK_a



9.6 pK_a

TABLE 9.7 pK_a and K_a Values for Substances Acting as Weak Acids (25 °C)

Name	Formula	pK _a	K _a
Hydrogen sulfate ion	HSO ₄ ⁻	2.00	1.0 × 10 ⁻²
Phosphoric acid	H ₃ PO ₄	2.12	7.5 × 10 ⁻³
Hydrofluoric acid	HF	3.19	6.5 × 10 ⁻⁴
Nitrous acid	HNO ₂	3.35	4.5 × 10 ⁻⁴
Formic acid	HCOOH	3.74	1.8 × 10 ⁻⁴
Acetic acid	CH ₃ COOH	4.76	1.75 × 10 ⁻⁵
Carbonic acid	H ₂ CO ₃	6.35	4.5 × 10 ⁻⁷
Water	H ₂ O	7.00	1.0 × 10 ⁻⁷
Dihydrogen phosphate ion	H ₂ PO ₄ ⁻	7.18	6.6 × 10 ⁻⁸
Ammonium ion	NH ₄ ⁺	9.20	6.3 × 10 ⁻¹⁰
Hydrocyanic acid	HCN	9.21	6.2 × 10 ⁻¹⁰
Bicarbonate ion	HCO ₃ ⁻	10.32	4.8 × 10 ⁻¹¹
Hydrogen phosphate ion	HPO ₄ ²⁻	12.00	1.0 × 10 ⁻¹²

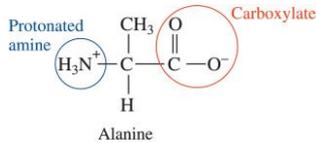
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9.6 pK_a

- The pH value changes with [H₃O⁺], yet the pK_a value is constant for a specific weak acid at a certain temperature
 - The smaller the pK_a value for a weak acid, the stronger the acid.
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9.7 Amino Acids—Common Biological Weak Acids

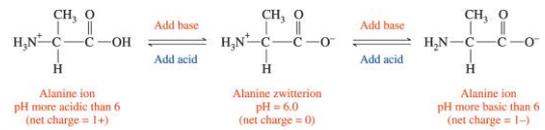


- This molecule is alanine, an **amino acid**.
- The amino acid shown is in the form that predominates at a physiological pH of 7.4.
- This ionic form containing no net charge (+ and – cancel each other out) is called a **zwitterion**.

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9.7 Amino Acids—Common Biological Weak Acids

- Each molecule has a unique pH value at which only the zwitterion is present.
- This is the **isoelectric point (pI)**.
- At the pI, the negative charge on the carboxylate is balanced by the positive charge on the ammonium ion, and the net charge is zero.



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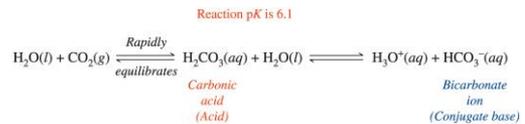
9.7 Amino Acids—Common Biological Weak Acids

- An amino acid is a biological molecule that contains the acid-base functional groups amine and carboxylic acid.
- The charge of functional group changes as the pH changes. At acidic pH, the acid forms of these functional groups is the main form present (protonated amine and carboxylic acid, net charge = +1). At basic pH, the conjugate base form predominates (amine and carboxylate, net charge = -1). The isoelectric point (pI) exists when the solution contains no net charge.

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9.8 Buffers and Blood—The Bicarbonate Buffer System

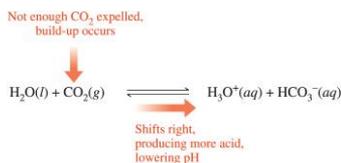
- A solution that contains a weak acid and its conjugate base or a weak base and its conjugate acid is a **buffer**.
- A buffer solution will resist a change in its pH if small amounts of acid or base are added.
- Blood is buffered by the bicarbonate buffer system.
- Dissolved CO₂ produced during cellular respiration travels through the bloodstream to the lungs.
- This dissolved CO₂ is rapidly equilibrated through carbonic acid into bicarbonate ions.



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9.8 Buffers and Blood—The Bicarbonate Buffer System

- The ability of an organism to regulate its internal environment by adjusting factors such as pH, temperature, and solute concentration is called **homeostasis**.
- A person who **hypoventilates** causes CO₂ gas to build up in the bloodstream.
- A buildup of CO₂ ultimately produces more H₃O⁺ in the bicarbonate equilibrium, making the blood more acidic.
- This condition is **respiratory acidosis**.
- It can be treated with intravenous bicarbonate.



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9.8 Buffers and Blood—The Bicarbonate Buffer System

- A person who **hyperventilates** exhales too much CO₂ from the lungs.
- This has the effect of drawing H₃O⁺ from the bloodstream, making the blood more basic.
- This condition is **respiratory alkalosis**.
- This can easily be reversed by having the person breathe into a paper bag, which keeps some of the previously exhaled CO₂ available for the next breath.



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9.8 Buffers and Blood—The Bicarbonate Buffer System

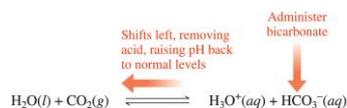
TABLE 9.8 Acidosis and Alkalosis: Symptoms, Causes, and Treatments

Respiratory Acidosis: $\text{CO}_2 \uparrow$ pH ↓		Metabolic Acidosis: $\text{H}_3\text{O}^+ \uparrow$ pH ↓	
Symptoms:	Failure to ventilate, suppression of breathing, disorientation, weakness, coma	Symptoms:	Increased ventilation, fatigue, confusion
Causes:	Lung disease blocking gas diffusion (e.g., emphysema, pneumonia, bronchitis, and asthma); depression of respiratory center by drugs, cardiopulmonary arrest, stroke, polio/myelitis, or nervous system disorders	Causes:	Renal disease, including hepatitis and cirrhosis; increased acid production in diabetes mellitus, hyperthyroidism, alcoholism, and starvation; loss of alkali in diarrhea; acid retention in renal failure
Treatment:	Correction of disorder, infusion of bicarbonate	Treatment:	Sodium bicarbonate given orally, dialysis for renal failure, insulin treatment for diabetic ketosis
Respiratory Alkalosis: $\text{CO}_2 \downarrow$ pH ↑		Metabolic Alkalosis: $\text{H}_3\text{O}^+ \downarrow$ pH ↑	
Symptoms:	Increased rate and depth of breathing, numbness, light-headedness, tetany	Symptoms:	Depressed breathing, apathy, confusion
Causes:	Hyperventilation because of anxiety, hysteria, fever, exercise; reaction to drugs such as salicylate, quinine, and antihistamines; conditions causing hypoxia (e.g., pneumonia, pulmonary edema, and heart disease)	Causes:	Vomiting, diseases of the adrenal glands, ingestion of excess alkali
Treatment:	Elimination of anxiety-producing state, rebreathing into a paper bag	Treatment:	Infusion of saline solution, treatment of underlying disease, administer NH_4Cl

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9.8 Buffers and Blood—The Bicarbonate Buffer System

- Chemical reactions that occur in our bodies can change the pH of blood by producing too much or too little H_3O^+ .
- Diabetics often use fatty acids as a carbon source for energy production.
- A by-product of fatty acid chemical breakdown is acid production. This causes **metabolic acidosis**.
- This can be reversed with intravenous bicarbonate.



- Metabolic alkalosis** can occur with excessive vomiting.
- To lower the pH, ammonium chloride can be given.

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9.8 Buffers and Blood—The Bicarbonate Buffer System

- Buffer solutions consist of approximately equal amounts of a weak acid and its conjugate base.
- Buffers resist changes in pH when acid or base is added to a solution.
- The bicarbonate buffer is an important buffer system in the blood.
- Blood pH is maintained in a narrow range of 7.35-7.45. If the blood pH drops below this range, a condition called acidosis occurs. If the blood p becomes elevated, a condition called alkalosis exists.

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