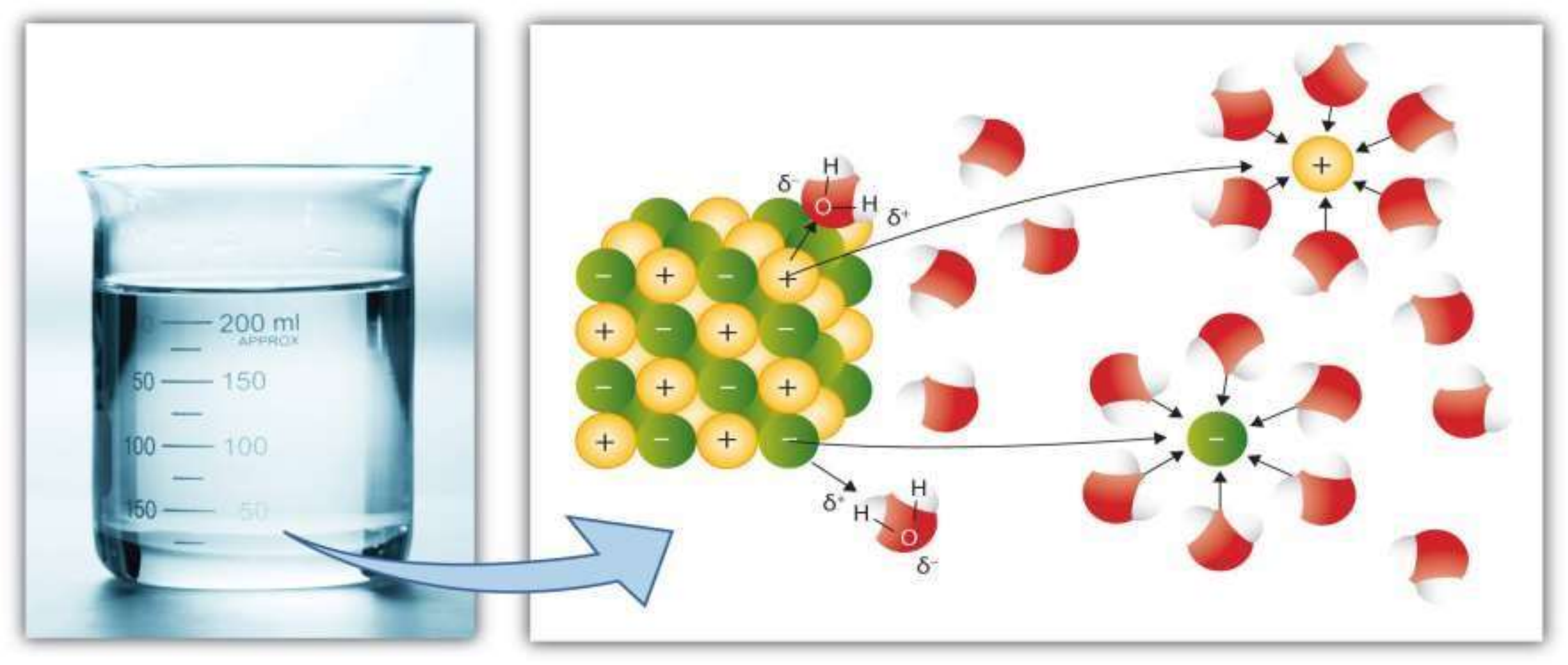
Acid-base equilibrium in the organism

Medical Chemistry

Electrolytes

Electrolytes are substances that dissociate (furnish ions) in water solutions

$$NaCl \rightleftharpoons \mathbf{Na}^{+} + Cl^{-}$$



Strong and weak electrolytes

1. Strong electrolytes are those that dissociate completely or almost completely

Example: HCl, NaOH, KOH, K_2SO_4, NH_4Cl

2. Weak electrolytes are those that only dissociate partially

Example: HCOOH, HCN, NH_4OH, H_2S, H_2O

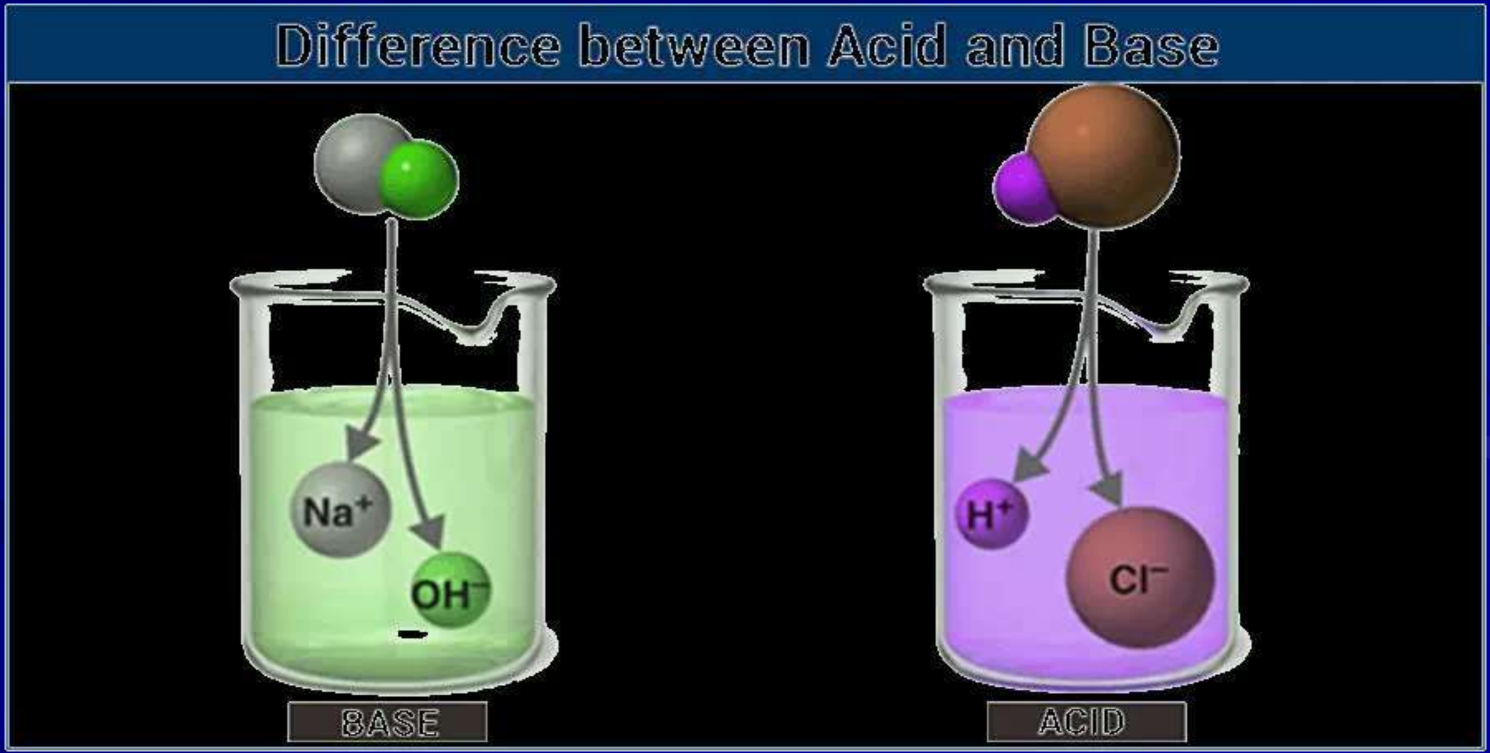
3. Nonelectrolytes do not dissociate

Example: ethanol, acetone, glucose, acetaldehyde

Acids and bases

Acids are compounds that furnish H^+ ions upon dissociation

Bases are compounds that furnish OH^- ions upon dissociation



Strong and weak acids

Strong acids dissociate completely or almost completely

Example: HCl, H_2SO_4, HNO_3, HI, HBr, $HClO_4$

Weak acids dissociate only dissociate partially

Example : acetic acid, citric acid, oxalic acid, lactic acid and other organic acids

Dissociation of acids

An acid dissociates:

$$HA \rightleftharpoons H^+ + A^-$$

The strength of an acid is measured as its dissociation constant, K_a:

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Or, more conveniently as its pK_a:

$$pK_a = -\lg K_a$$

Acidity, pH

Acidity is a measure of molarity of H^+ ions in a solution.

It is represented as pH:

$$pH = -lg[H^+]$$

pH is the negative logarithm of molarity of H^+ ions.

pH

At 25 C, pH range is from 0 to 14

The number of 14 stems from the ionic product of water:

$$K_w = [H^+] \cdot [OH^-] = 10^{-14}\ mol^2/L^2\ (t = 25^o\ C)$$

The ionic product of water stems from the dissociation constant of water:

$$K_{d\,H_2O} = 1.86 \cdot 10^{-16}\ mol/L\ (t = 25^o\ C)$$

pOH

Similarly to pH, pOH is:

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

At 25 C,

$$pOH + pH = 14$$

Neutral, acidic and basic pH

At 25 C, neutral pH is 7,

because in pure water

$$[H^+] = [OH^-] = 10^{-7} \text{ mol/L}$$

Acidic pH is less than 7

Acidic pH is less than 7

Basic pH is more than 7

Basic pH is more than 7

Effect of temperature on pH

With an increase of temperature, pH decreases, because dissociation increases

At 37 C,

$$pOH + pH = 13.6$$

and neutral pH = 6.8

pH of fluids in the organism

Organism fluid or secretion	pH
Blood	7.35 - 7.45
Urine (normal)	5.5 - 7.5
Gastric juice	0.9 - 2.0
Pancreatic juice	7.0 - 8.5
Saliva	6.5 - 7.5
Urine (in various diseases)	4.8 - 8.5

Acid-base metabolism

The organism constantly produces organic acids and carbon dioxide,

that make it acidic

Regulation of acid-base homeostasis is performed by two mechanisms: physiologic and chemical

Physiologic mechanism is the work of organs: kidney and lungs

Chemical mechanism is the action of buffer systems of the organism

Buffer systems

Buffer systems are solutions that resist change in pH upon addition of small amounts of strong acids or strong bases or upon dilution.

Classification of buffer systems

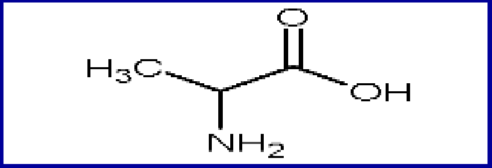
1. **Acid**. They consist of a weak acid and its salt with a strong base.

Example: CH_3COOH / CH_3COONa

2. **Base**. They consist of a weak base and its salt with a strong acid.

Example: NH_4OH / NH_4Cl

3. **Ampholyte**. They are amphoteric electrolytes that have properties of acid and base in same molecule. **Example:** amino acids, proteins



Mechanism of buffering in acid buffer solutions

Dissociation of an acetic acid buffer CH_3COOH/CH_3COONa:

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

$$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$$

Acetic acid is a weak acid, and dissociates weakly. Sodium acetate is strong electrolyte, and dissociates completely.

So the active components of the buffer are CH_3COOH and CH_3COO^-.

Addition of a strong base to an acetic acid buffer solution

$$Na^+ + OH^- + CH_3COOH \rightleftharpoons CH_3COO^- + Na^+ + H_2O$$

or in short:

$$OH^- + CH_3COOH \rightleftharpoons CH_3COO^- + H_2O$$

Free OH^- ions are bound in undissociated molecules of water, and pH of the buffer solution does not change

Addition of a strong acid to an acetic acid buffer solution

$H^+ + Cl^- + CH_3COO^- \rightleftharpoons CH_3COOH + Cl^-$

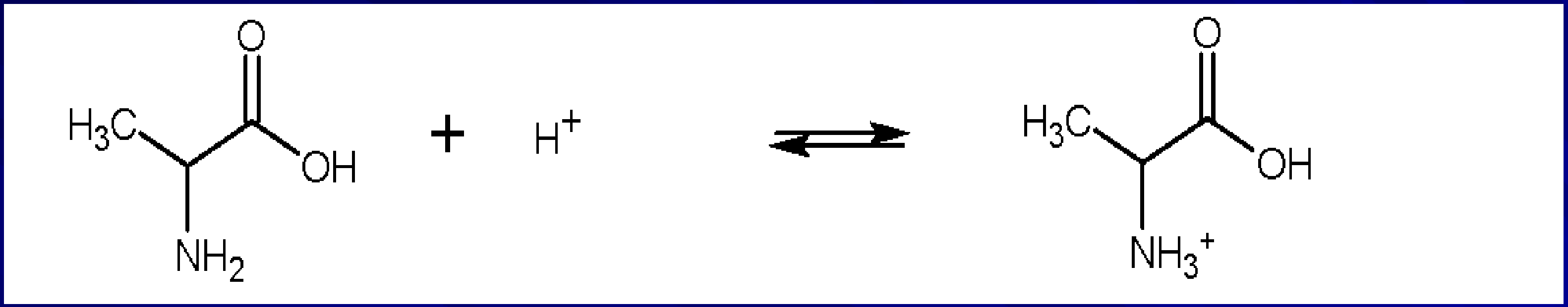
or in short:

$H^+ + CH_3COO^- \rightleftharpoons CH_3COOH$

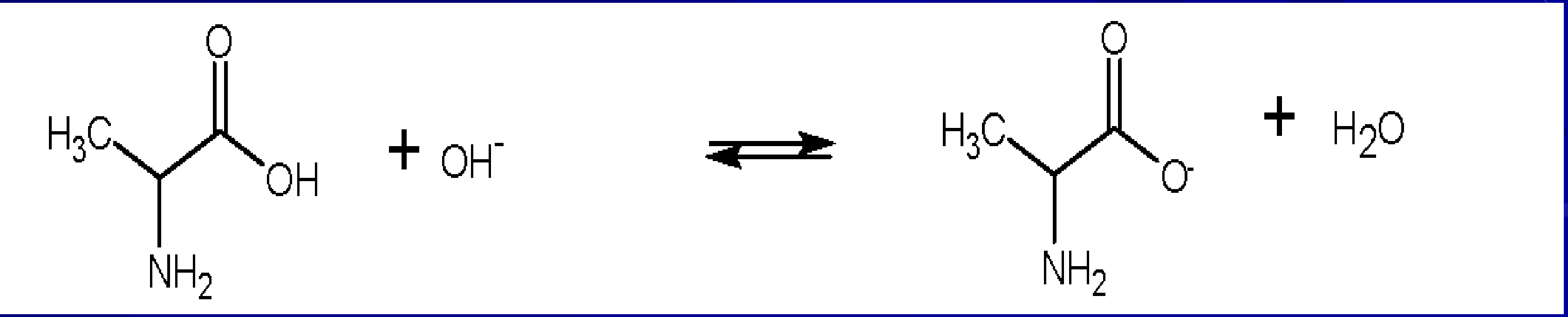
Free H^+ ions are bound in undissociated molecules of weak acid, and pH of the buffer solution does not change

Mechanism of buffering in ampholyte buffer systems

Addition of a strong acid



Addition of a strong base



Buffer capacity

Buffer capacity is the number of equivalents of a strong acid or a strong base that has to be added to 1 liter of a buffer solution to cause a pH change of 1.0 pH unit.

Buffer capacity for acids:

$$B_a = \frac{N(acid) \cdot V(acid)}{V(buffer) \cdot \Delta pH}$$

Buffer capacity for bases:

$$B_a = \frac{N(base) \cdot V(base)}{V(buffer) \cdot \Delta pH}$$

Buffer capacity depends on the ratio of the components and their concentrations.

Henderson-Hasselbach equation

With the equation, we can calculate pH of a buffer solution, or prepare a solution of needed pH.

$$pH = pK_a + \lg \frac{[\text{salt}]}{[\text{acid}]}$$

pH of a buffer solution depends on:

1. the strength of the weak acid (or weak base)
2. ratio of the components
3. temperature

Buffer systems of blood. Hemoglobin and bicarbonate buffer systems

Hemoglobin buffer system is:

deoxygenated hemoglobin HHb / Hb^{-}

oxygenated hemoglobin $HHbO_2 / HbO_2^{-}$

Bicarbonate buffer system is
H_2CO_3/HCO_3^{-}

Interaction of the two buffer systems ensures both buffering and efficient delivery of oxygen to tissues.

Buffer systems of blood.

Mechanism of action of hemoglobin and bicarbonate buffer systems

1. CO_2 is generated in tissues, diffuses to blood plasma, then to red blood cells (RBC).

2. In RBC, the enzyme carbonic anhydrase catalyzes the reaction:

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

Mechanism of action of hemoglobin and bicarbonate buffer systems

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

3. H^+ binds to oxyhemoglobin, makes it release oxygen

4. HCO_3^- diffuses to …, where it is …

5. In the lungs, the process is reversed, and CO_2 is released and exhaled.

Buffer systems of blood. Blood plasma protein buffer system

In proteins, the carboxyl group –COOH and amine group $–NH_2$ of amino acids are capable of buffering.

Protein amino acid side chains that act as buffers are carboxyl groups of glutamate and aspartate and the weakly basic groups of lysine, arginine, and histidine.

Buffer systems of blood.
Phosphate buffer system

Phosphate buffer system $H_2PO_4^-/HPO_4^{2-}$

It has minor role in blood, but it is the principal buffer system of urine

Disorders of acid-base homeostasis (1)

- The normal pH range of blood plasma is 7.35 to 7.45.
- The plasma pH levels either 6.8 or 7.8 are incompatible with life.

Disorders of acid-base homeostasis (2)

- When blood plasma pH goes lower than 7.35 is **acidosis**.
- When pH goes higher than 7.45, it is **alkalosis**.
- Both acidosis and alkalosis can be of two types: **metabolic and respiratory**.

Disorders of acid-base homeostasis (3)

- Blood plasma concentration of bicarbonate HCO_3^- (normal range 22 – 28 mmol/L) is measured.
- Partial pressure of carbon dioxide in arterial blood P_{CO2} (normal range 33 – 44 mm Hg) is measured.

Metabolic acidosis

- Low pH (under 7.35),
  low bicarbonate concentration (less than 22 mmol/L).
- Metabolic acidosis arises because of:

a) increased production of hydrogen ions in metabolism;

b) decreased excretion of hydrogen ions by the kidney;

c) indirectly, with an increased loss of bicarbonate ions.

Respiratory acidosis

It is caused by **hypoventilation.**

The main symptom is an increase of P_{CO2} higher than the normal value.

Metabolic alkalosis

- Loss of hydrogen ions in vomiting or through the kidney.
- Hyperaldosteronism (Conn's syndrome)
- Some diuretic therapies

Respiratory alkalosis

It is caused by hyperventilation.