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# Acid-base balance, buffer systems and pH blood

**Lec(4)**

**First stage**

**By**

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# ***Weak and strong acids***

The extent of dissociation decides whether they are strong acids or weak acids. Strong acids dissociate completely in solution, while weak acids ionize incompletely, for example,

$$HCl \longrightarrow H^+ + Cl^- \quad \text{(Complete)}$$

$$H_2CO_3 \rightleftharpoons H^+ + HCO_3^- \quad \text{(Partial)}$$

In a solution of HCl, almost all the molecules dissociate and exist as H+ and Cl– ions. Hence the concentration of H+ is very high and it is a strong acid. But in the case of a weak acid (e.g. acetic acid), it will ionize only partially. So, the number of acid molecules existing in the ionized state is much less, may be only 50%.

# *Constant dissociation*

Since the dissociation of an acid is a freely reversible reaction, at equilibrium the ratio between dissociated and undissociated particle is a constant. The dissociation constant (Ka) of an acid is given by the formula,

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

Where [H+] is the concentration of hydrogen ions, [A–] = the concentration of anions or conjugate base, and [HA] is the concentration of undissociated molecules.

The pH at which the acid is half ionized is called pKa of an acid which is constant at a particular temperature and pressure. Strong acids will have a low pKa and weak acids have a higher pKa

# *The effect of salt upon the dissociation*

The relationship between pH, pKa, concentration of acid and conjugate base (or salt) is expressed by the Henderson-Hasselbalch equation:

$$pH = pKa + \log \frac{\{conjugate\ base\}}{\{acid\}} \text{ or } pH = pKa + \log \frac{\{salt\}}{\{acid\}}$$

When [base] = [acid]; then pH = pKa Therefore, . Thus, when the acid is half ionized, pH and pKa have the same value.

For weak base and its salt:

$$pOH = pKb + \log \frac{\{conjugate\ acid\}}{\{base\}} \text{ or } pOH = pKb + \log \frac{\{salt\}}{\{base\}}$$

Example// calculate the pH of a buffer solution with an acetic acid concentration of 0.1M and a sodium acetate concentration of 0.02M. If you know (Ka=$1.8 \times 10^{-5}$)

answer//

$CH_3COOH$ + $CH_3COONa$ (weak acid and its salt)

pH= pKa + log $\frac{\{conjucate\ base\}}{\{acid\}}$ or pH = pKa + log $\frac{\{salt\}}{\{acid\}}$

pH= -log ($1.8 \times 10^{-5}$) + log $\frac{0.02}{0.1}$

pH= -log1.8-(-5log10)+log0.2

pH= 4.74

# *Cellular mechanisms of regulation of acid-base*

cellular mechanisms of maintaining constancy of hydrogen ions concentration in extra-cellular fluid:

- If pH increases hydrogen ions move from cells to extracellular fluid in exchange of potassium ions that enter the cells and alkalosis is usually accompanied by hypokalemia.
- If pH decreases hydrogen ions enter the cells in exchange of potassium ions that leaves the cells and acidosis may cause hyperkalemia. In such a way electro-neutrality law is maintained by cellular regulation. According to it, the sum of the positive and negative charges (cations and anions) is equal. So, hydrogen to potassium exchange between ECF and ICF should be equal.

# *Buffer solutions*

**Buffers** are solutions which can resist changes in pH when acid or alkali is added

Composition of buffers:

- Buffers are of two types:

a. Mixtures of weak acids with their salt with a strong base

b. Mixtures of weak bases with their salt with a strong acid.

A few examples are given below:

i. $H_2CO_3/NaHCO_3$ (Bicarbonate buffer) (carbonic acid and sodium bicarbonate)

ii. $CH_3COOH/CH_3COONa$ (Acetate buffer) (acetic acid and sodium acetate)

iii. $Na_2HPO_4/NaH_2PO_4$ (Phosphate buffer)

# *types of buffer solutions*

**1- Acidic buffer**

- Consists of weak acid and its salt of strong electrolyte.

e.g. Acetic acid and sodium acetate (CH3COOH/CH3COONa)

- Upon addition of small amount of acid (H+).

Sodium acetate reacts with it giving weakly ionized acetic acid. $H^{+} + CH3COONa \longrightarrow CH3COOH + Na^{+}$

- Upon addition of small amount of a base ($OH^{-}$)

Acetic acid reacts with it and non ionized water is formed.

$OH^{-} + CH3COOH \longrightarrow CH3COO^{-} + H_2O$

# *types of buffer solutions*

**2- Basic buffer**

- Consists of weak base and its salt of strong electrolyte.

e.g. Ammonium hydroxide and ammonium chloride. (NH4OH/NH4Cl)

- Upon addition of small amount of acid (H+).

$$H^{+} + NH_4OH \longrightarrow NH4^{+} + H_2O$$

- Upon addition of a small amount of base ($OH^{-}$)

Acetic acid reacts with it and non ionized water is formed.

$$OH^{-} + NH_4Cl \longrightarrow NH_4OH + Cl^{-}$$

# *Factors affecting pH of a buffer*

- The pH of a buffer solution is determined by two factors:

a. The value of pK: the lower the value of pK, the lower is the pH of the solution.

b. The ratio of salt to acid concentration: Actual concentrations of salt and acid in a buffer solution may be varying widely, with no change in pH, so long as the ratio of the concentrations remains the same.

# ***Buffer capacity '' ß''***

**Buffer capacity**: It is the magnitude of the resistance of a buffer to change in the pH.

- $ß = \Delta ß / \Delta pH$

Where:

- ß is buffer capacity.
- Δß is amount of acid or base added.
- ΔpH is the change in pH.

***The main buffer systems are following:***

1. Bicarbonate buffer: The most important extracellular buffer produced by kidneys, has the largest buffering capacity.

2. Haemoglobin buffer: Main intracellular buffer of the blood.

3. Protein buffer: An extracellular buffer together with bicarbonate buffer, represented by plasma proteins.

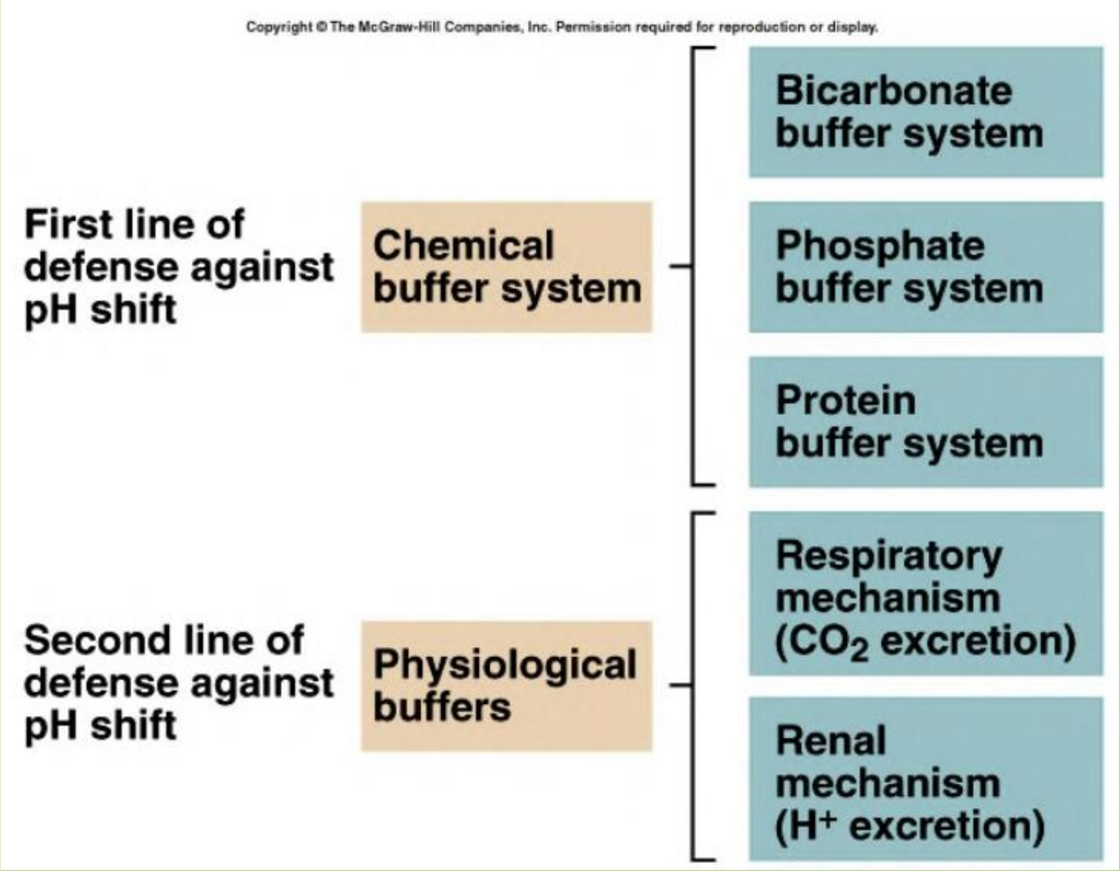
4. Phosphate buffer: It takes part in hydrogen ions excretion in renal tubules, is not of great importance in blood.

Main blood buffer systems:

| Buffer system | Buffering capacity (%) |
|---|---|
| Bicarbonate | 53 |
| Haemoglobin | 35 |
| Protein | 7 |
| Phosphate | 5 |

# Regulation of the blood PH

- There are three primary system that regulate the {H+} conc.
- 1-buffer mechanism
- 2-the respiratory mechanism
- 3-renal mechanism



# Regulation of the blood PH

- Buffers of **extracellular fluid** present in plasma.

1. Bicarbonate buffer .
2. Phosphate buffer
3. Protein buffer .

- Buffers of **intracellular fluid** present in RBCs

1. Bicarbonate buffer
2. Phosphate buffer
3. Hemoglobin buffer .

- The pH of plasma is 7.4 (Range 7.35 - 7.45). In normal life, the variation of plasma pH is very small
- The pH of plasma is maintained within a narrow range. The pH of the interstitial fluid is generally 0.5 units below that of the plasma.
- The role of kidneys in the maintenance of acid-base balance of the body (blood pH) is highly significant.
- The renal mechanism tries to provide a permanent solution to the acid-base disturbances.
- This is in contrast to the temporary buffering system and a short term respiratory mechanism.
- The kidneys regulate the blood pH by maintaining the alkali reserve.