Republic of Iraq

Ministry of Higher Education for Sciences

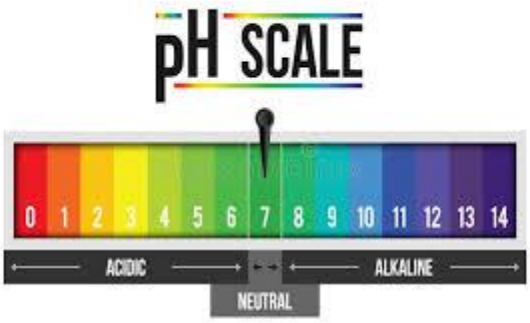
Health and Medical Colleges

Department of Anesthesia Techniques

CHEMISTRY

Title of the lecture:

The PH concept, Acid-base balance, chemical equilibrium, common ion effect.



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Introduction to pH

pH is a scale that expresses the acidity or alkalinity of a solution.

It is determined using the equation:

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

It depends on the concentration of hydrogen ions in the solution.

The pH Scale

The pH scale ranges from 0 to 14

0 - 6.9: The solution is acidic (e.g., lemon juice or vinegar).

٧: The solution is neutral (e.g., pure water).

7.1 - 14: The solution is basic (e.g., soap or ammonia solution).

Examples of pH for Common Substances

1. Lemon juice: (acidic).
2. Vinegar: (acidic).
3. Pure water: (neutral).
4. Ammonia solution: (basic).

Importance of pH in Daily Life

human body

1. Human blood has a pH of around.
2. Even small changes in blood pH can lead to serious health issues.

agriculture

1. Soil quality depends on pH.
2. Acidic or basic soils affect plant growth.

industry

1. Water treatment depends on controlling pH to make it potable.
2. Pharmaceutical production relies on pH control to ensure compound efficacy.

Relationship Between pH and Hydroxide Ions

In aqueous solutions, the relationship between $[H^+]$ and $[OH^-]$ is expressed by the water ionization constant: $[H^+].[OH^-] = 1\times10^{-14}$

If the solution is acidic, the concentration of $[H^+]$ is higher than $[OH^-]$.

If the solution is basic, the concentration of $[OH^-]$ is higher than $[H^+]$.

Methods for Measuring pH

1

Litmus Paper

- Changes color to indicate whether the solution is acidic or basic.

2

pH Meter:

- Provides an accurate measurement of pH.

Definitions of Acids and Bases

Arrhenius Theory

- **Acid:** A substance that increases the concentration of hydrogen ions in a solution.
- **Base:** A substance that increases the concentration of hydroxide ions in a solution.

Bronsted-Lowry Theory

- **Acid:** A substance that donates hydrogen ions.
- **Base:** A substance that accepts hydrogen ions.

Lewis Theory

- **Acid:** A substance that accepts a pair of electrons.
- **Base:** A substance that donates a pair of electrons.

Acid-Base Balance in Solutions:

When an acid or a base dissolves in water, a dynamic equilibrium forms between reactants and products. This balance is represented by the following reactions:

1. Acid Ionization

- **Strong acids ionize completely.**

Such\ HCl $\longrightarrow$ H+ + Cl

- **Weak acids ionize partially.**

Such\ $CH_3COOH \longrightarrow CH_3COO^- + H^+$

2. Base Ionization:

- **Strong bases ionize completely.**

Such\ $NaOH \longrightarrow Na^+ + OH^-$

- **Weak bases ionize partially.**

Such\ $NH_3 + H_2O \longrightarrow [NH4]^+ + [OH^-]$

Common Ion Effect

The common ion effect is a chemical phenomenon that occurs when an ion, identical to one already present in a chemical equilibrium, is added to a solution. This causes the equilibrium to shift according to Le Chatelier's Principle.

Result of the Effect

The dissociation or solubility of the original substance in the solution decreases.

Explanation:

When the concentration of one ion in the equilibrium increases, the reaction shifts in the opposite direction to reduce the impact of this increase.

How Does the Common Ion Effect Occur?

Example 1: Reducing Solubility of Sparingly Soluble Salts

- **When silver chloride is dissolved in water:**
- **If sodium chloride is added to the solution, dissociates and releases ions.**
- **The increase in concentration [Cl^-] shifts the equilibrium to the left, reducing the solubility of [AgCl].**

Example 2: Reducing Ionization of Weak Acids

- **When acetic acid is dissolved:**
- **If sodium acetate is added, the increase in[$CH3COO^-$] reduces the ionization of acetic acid, making the solution less acidic.**

Applications of the Common Ion Effect

1. In Analytical Chemistry

- **Salt Precipitation:**
- **The common ion effect is used to facilitate the precipitation of salts.**
- **Example: Precipitation of barium sulfate by adding a source of sulfate ions .**

2. In Industry

- **The effect is used to reduce the solubility of substances when separation or purification of specific compounds is required.**

3. In Biological Systems:

- **Regulating acid-base balance in blood and tissues, where ions such as bicarbonate help minimize pH changes.**

4. In Pharmaceuticals:

- **Enhancing the stability of certain drugs by controlling their dissociation using the common ion effect.**

Buffers and Buffer Systems of Physiological Importance in Living Systems

- Buffers are chemical systems composed of a weak acid and its conjugate base (or a weak base and its conjugate acid).
- They resist significant changes in pH upon the addition of small amounts of acid or base.

Primary purpose

Maintain pH stability in biological environments.

Mechanism of Action

• When an acid (H^+) is added:

- The conjugate base in the buffer reacts with the added ions, reducing the acid's effect.

• When a base (OH^-) is added:

- The weak acid in the buffer reacts with the added base, reducing its impact.

Importance of Buffers in Living Systems

Maintaining Blood pH

- Blood pH is tightly regulated between 7.35 and 7.45.
- Deviations from this range can cause serious health issues, such as acidosis or alkalosis.

Stabilizing Biochemical Processes

- Many enzymes function optimally at a specific pH.
- Large pH changes can hinder enzymatic activity or denature proteins

Gas Transport

- The transport of oxygen and carbon dioxide in the blood depends on acid-base balance.

a. Bicarbonate Buffer System:

• Components:

• Carbonic acid (H_2CO_3)

• Bicarbonate ion (HCO_3^-)

• Significance: • This system maintains blood pH. • When increases, it combines with to form, which is exhaled as through the lungs.

c. Protein Buffer System:

• Proteins contain functional groups such as:

• Carboxyl groups ($COOH^-$): Act as acids.

• Amino groups ($-NH_3$): Act as bases.

• Example: Hemoglobin buffers blood pH by binding to excess.

Physiological Buffer Systems

b. Phosphate Buffer System:

• Components:

• Dihydrogen phosphate (H_2PO_4)

• Hydrogen phosphate (HPO_4^-)

• Significance:

• Plays a crucial role in intracellular pH regulation (pH ~7.2).

d. Ammonia Buffer System:

• Components:

• Ammonia (NH3)

• Ammonium ion (NH_4^+) •

Significance: • Helps eliminate excess acid in the kidneys.

Applications of Buffers in Medicine

Treating Acidosis or Alkalosis

Bicarbonate or phosphate buffers are used to adjust pH levels.

Preserving Medications

Buffers stabilize the pH of pharmaceutical formulations to maintain their efficacy.