ACID BASES BUFFERS

CHAPTER - 1ST | **UNIT - 2ND**

• SYLLABUS :

Introduction, theories of Acid and Bases, Importance of acids and bases.
Buffers - Introduction, Buffer capacity, Buffer Equation (Handerson-Hasselbalch Equation), characteristics of buffer solution, Types of Buffer solution, Buffers in pharmaceutical systems, Buffered Isotonic Solutions, Pharmaceutical Importance of Buffers.

ACIDS AND BASES

These are popular chemicals which interact with each other in the formation of salt and water. $HCl + NaOH \longrightarrow NaCl$.

• Acids :

These are those substances which can be identified by their sour taste.
An acid is a molecules which can donate an H^+ ion (proton ions).
Acids are known to turn blue litmus red.

$$\boxed{\text{Blue litmus paper}} \xrightarrow{\text{Acids}} \boxed{\text{Red litmus paper}}$$

eg. $HCl \longrightarrow H^+ + Cl^-$, H_2SO_4, CH_3COOH
Vinegar, lemon Juice, Vitamin C etc..

• Bases :

These are those substances which can be

identified by their bitter taste, slippery texture.

A base is a molecules which can accept a hydrogen ion (H^+) or release OH^- ions.

Bases are known to turn red litmus blue.

Red litmus paper $\xrightarrow{\text{Bases}}$ Blue litmus Paper

eg. $NaOH \longrightarrow Na^+ + OH^-$, Baking Soda, Soaps, Ammonia (NH_3).

THEORIES OF ACID AND BASES

These are those theories which are basically used to define and explain the concept of Acids and Bases.

It is of three types :-

1. Arrhenius theory
2. Bronsted - Lowry theory
3. lewis acid-base theory

1. Arrhenius theory :

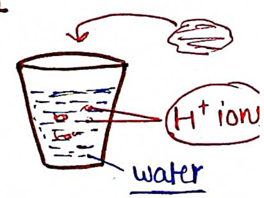
It is most commonly used concept for acids and bases, and it is developed by Savante Arrhenius in 1884.

According to this theory

a. Any substance which produces/gives H^+ ions when dissolved in aqueous solution (water) are known as Acids.

Example → HCl, H_2SO_4

$HCl \xrightarrow[\text{medium}]{\text{aqueous}} H^+ + Cl^-$



b. Any substance which gives OH^- ions when dissolved in aqueous solution (water) are known as Bases.

examples → NaOH, $Mg(OH)_2$, $Ca(OH)_2$ etc..

$NaOH \xrightarrow[\text{medium}]{\text{aqueous}} Na^+ + OH^-$

- It is the first scientific theory and the simplest theory used in case of aqueous solution

• Limitations :

- This theory explain the acidic and basic solution only in aqueous medium and fails to explain in non-aqueous medium.
- There are many acids or bases which does not contain H^+/OH^- ions but they are acidic / Basic in nature. eg. NO_2^-, NH_3 (Bases) SO_2, BF_3 (Acids).

② Bronsted - Lowry theory :

J.N. Bronsted and J.M. Lowry in 1923 gave a broader concept of acids and bases.

According to this theory :

- These are those substances which are capable to donate the 'proton' (H^+ ion) are known as **Acids** - Proton Donor.
- These are those who accepts the proton (H^+). are known as **Base** - Proton Acceptor.

eg.

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

$NH_3 + H^+ \longrightarrow NH_4^+ \longrightarrow$ Base

⊙ Conjugate Acid - Base Pairs :

In Acid-Base reaction, when acid react with base it produces a new base by donating its proton (H^+) and that base is known as conjugate base.

$$\boxed{\underset{\text{Acid}}{HCl} + \underset{\text{Base}}{NH_3} \xrightarrow{H^+} \underset{\text{Conjugate acid}}{NH_4^+} + \underset{\text{Conjugate Base}}{Cl^-}}$$

The substance which accept protons are known as conjugate acid.

⊙ Limitation :

It fails to explain about those molecules which does not contain H^+/OH^- ions.

eg. SO_2, BF_3, N_2O_4 etc..

③ Lewis theory of Acids and Bases:

This theory was given by G.N. Lewis in early 1930s.

According to this theory

- Those substances which accepts the lone pair of electron are known as $\boxed{\text{acids}}$ i.e. Electron-pair acceptor. (Lewis acid).

  eg. Cu^{2+}, BF_3, Fe^{3+} etc..

- Those substances which have lone pair and donate the lone pair of electrons are known as $\boxed{\text{Base}}$ or Lewis Base i.e. Electron pair donor.

  eg. $\ddot{N}H_3$, $:\ddot{O}H^-$, C_2H_4 (ethylene) etc...

- A Lewis acid accepts an electron pair from a Lewis base, forming a coordinate covalent bond which form Lewis adduct complex.

$$\underset{\text{Lewis acid}}{BF_3} + \underset{\text{Lewis base}}{\ddot{N}H_3} \longrightarrow \underset{\text{Complex}}{BF_3 \leftarrow NH_3}$$

- **Limitation:**

  It fails to explain the acid-base reaction that do not involve the formation of a coordinate covalent bond.

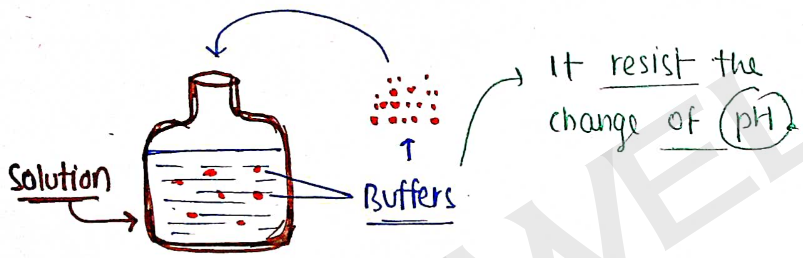
Importance of Acids and Bases:

- Used in the preperation of suitable salts.
- Various acids and bases are used during acid-base titrations.
- Various acids and bases are found in human body and used to maintain pH of GIT, urine, blood eg. HCl for digestion.
- Used in daily households eg. Vinegar, Citric acid (Lemon) and Sulphuric acid
- Soaps etc..

BUFFERS OR BUFFER SOLUTIONS

These are those solutions which resist the change of pH of solution, when a small amount of either an acid (H^+ ions) or a base (OH^- ions) is added.

eg. Ammonium acetate solution



- **Characteristics of buffer Solution:**

- It has reserve acidity or alkalinity, thus a definite pH.
- It does not change pH, even if stored for long duration
- there is no alteration in its pH on dilution.

TYPES:

It is of two types:

① Acidic Buffers:

Those solutions which are added in the acidic solutions.

It is made up by mixing a weak acid and its salts with a strong base

eg. CH_3COOH (Acetic acid) + CH_3COONa (sodium acetate)

② Basic Buffers:

Those buffers which are added in the basic solutions.

It is made up by mixing a weak base and its salts with a strong acid.

eg. NH_4OH (Ammonium hydroxide) + NH_4Cl (Ammonium chloride)

• Buffer Capacity:

The amount of acid or base that must be added to the buffer to produce a unit change of pH.
It helps to know the effectiveness of a buffer on a quantitative basic.

$$\boxed{\beta = \frac{\Delta B}{\Delta pH}}$$

where, β = Buffer Capacity, ΔB = Amount of acid / base
ΔpH = pH changes

BUFFER EQUATION

Also known as Handerson - Hasselbalch Equation.

It is used to calculate the pH of a buffer solution and the changes in pH which occurs during addition of an acid or base.

• for Acidic buffers [weak acid and its salts]

The pH of acidic buffer can be calculated from the dissociation constant (Ka) of the weak acid and its salt.

• let's take HA (weak acid) and BA (salt).

$$HA \rightleftharpoons H^+ + A^- \quad \text{(weak acid)}$$
$$BA \longrightarrow B^+ + A^- \quad \text{(its salt)}$$

Common ion effect

• By applying law of mass action,

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

On rearrange,

$$[H^+] = \frac{K_a\,[HA]}{[A^-]}, \quad \text{where, } HA \rightarrow \text{Acid}, \; A^- \rightarrow \text{Salt}$$

$$[H^+] = K_a \frac{[\text{Acid}]}{[\text{Salt}]}$$

[Acid] → concentration of acid
[Salt] → concentration of its salt

- Taking (-log) on both sides,

$$-\log[H^+] = -\log\left[K_a \frac{(Acid)}{(Salt)}\right]$$

$$-\log[H^+] = -\log K_a - \log\frac{[Acid]}{[Salt]}$$

Now, $\because$ $pH = -\log[H^+]$, $pK_a = -\log K_a$

So,

$$pH = pK_a - \log\frac{[Acid]}{[Salt]}$$

On Rearrange,

$$\boxed{pH = pK_a + \log\frac{[Salt]}{[Acid]}}$$

Also known as Henderson - Hasselbatch Equation.

⊙ for Basic Buffer (weak base and its salts)

$$\boxed{pOH = pK_b + \log\frac{[Salt]}{[Base]}}$$

BUFFERS IN PHARMACEUTICAL SYSTEMS

The buffers play an important role in pharmaceutical preparation to ensure pH condition for the medicines.

- to preserve pH of pharmaceuticals.
- to maintain solubility of pharmaceuticals by adjusted its pH.
- to maintain stability of pharmaceuticals.

eg. Phosphate buffer, Borate buffer.

- **preparation:**

- By mixing a weak acid and its salt with a strong base.

eg. $CH_3COOH + CH_3COONa$

Boric acid + Borax.

BUFFER ISOTONIC SOLUTIONS

These are those solution which have same osmotic pressure as that of body fluids.

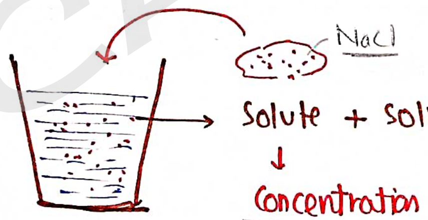
These solution are meant for application of body.

eg. Blood → 0.9% w/v NaCl solution.

• Tonicity:

It is defined as, it is the amount/concentration of solute in any solution.

eg. 0.9% w/v NaCl solution i.e. 9 gm NaCl solution is dissolved in 1000 ml of water.



Solute + solvent
↓
Concentration of solute.

• It is of three types :-

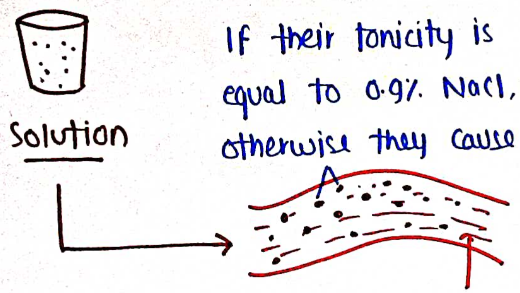
i) Isotonic solution → A buffer solution that have same osmotic pressure as body fluids i.e. 0.9% w/v NaCl solution.
= 0.9% NaCl

ii) Hypotonic solution → A buffer solution that have less solute concentration (osmotic pressure) than 0.9% NaCl solution.
< 0.9% NaCl

iii) Hypertonic solution → A buffer solution that have high solute concentration (osmotic pressure) than 0.9% NaCl solution.
> 0.9% NaCl

◉ The concept of isotonicity is used during preparation of medicament / I.V. fluids for body.

- Because, if the tonicity of these medicament is different than blood / body fluids, then they cause irritation or maybe serious damage.



If their tonicity is equal to 0.9% NaCl, then ok otherwise they cause irritation → RBC → Swell / Shrink

Blood tonicity is equal to 0.9% w/v NaCl solution.

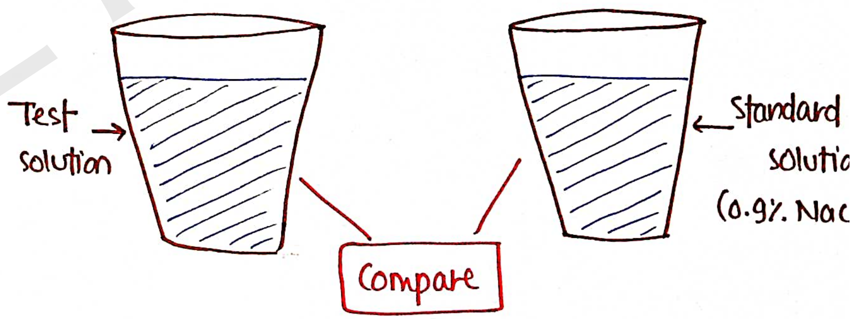
- So, Isotonic solution are those solution which contain 0.9% w/v NaCl solution.

• Method to determine Isotonicity :

i) Cryoscopic method (Colligative methods)
ii) Hemolytic Method

① **Cryoscopic Method** → This method is depends upon colligative properties of solution such as freezing point, Boiling point, Vapour pressure and temperature difference

- Take two solution, one standard isotonic solution (0.9% NaCl) and other is test solution (whose tonicity, we have to determine).
- Now compare their colligative properties with standard solution and determine the tonicity of solution.

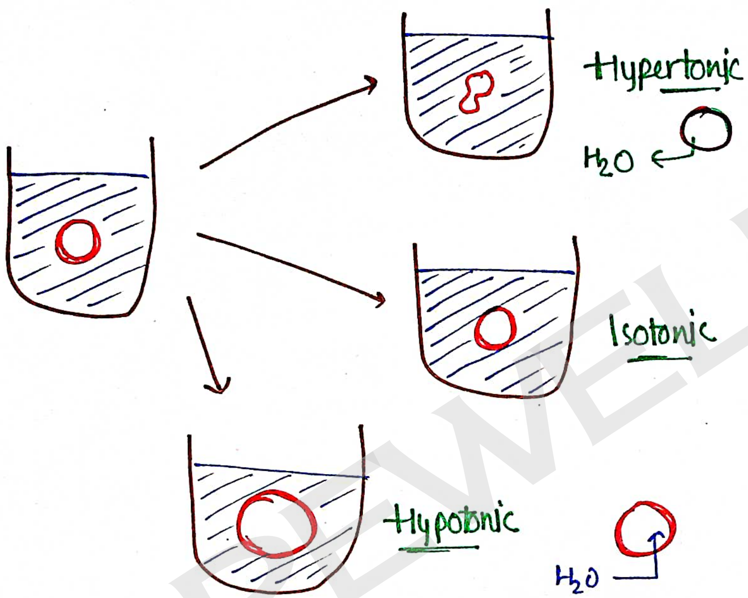


- freezing point
- Boiling point
- Vapour pressure

eg. f.p. of std (standard) 0.9% NaCl = −0.52°C

Same → Isotonic Solution
> 0.9% → Hypertonic solution
< 0.9% → Hypotonic solution

(ii) Hemolytic Method → The effect of various solution of the drug was observed on the appearance of RBC (Red blood cells) suspended in solution.



- According to osmosis, solvent particle move from low concentration to high concentration.

- concentration of solution > concⁿ of RBC (0.9%)

  so, solvent (H_2O) move from low to High, or RBC to solution, this cause cell shrinkage Hypertonic solution.

- Concentration of solution = concⁿ of RBC (0.9%)

  so, RBC cells remain constant (Isotonic Solⁿ).

- Concentration of solution < concⁿ of RBC (0.9%)

  so, Solvent (H_2O) move/diffuse from solution to RBC cells, this cause cell swelling Hypotonic solution.

- **Method of adjusting tonicity:**

  we can make solution isotonic by using two methods

  (i) Class 1^{st} — Cryoscopic & sodium chlorid E.

  (ii) Class 11^{nd} — White Vincent & Sprowls method

(I) Class IST :

(a) Cryoscopic method (freezing point depression method) :-

This method is used for hypotonic solution, i.e. concentration of solution is less than 0.9% w/v NaCl.

Sodium chloride (NaCl) is added to solution to make it isotonic.

$$\boxed{W\% = \frac{0.52 - a}{b}}$$

where,

- w = amount of adjusting substance
- a = freezing point of 1% solution of unadjusted solution.
- b = freezing point of 1% solution of adjusting solution.

(b) Sodium Chloride Equivalent (E) :-

This is used for hypotonic solution, In this add sodium chloride in solution to make it isotonic

$$\boxed{E = \frac{17 \times L_{iso}}{M}}$$

where,

- E = Sodium Chloride Equivalent / Amount of NaCl required.
- Liso = Liso constant value
- M = Molecular weight of drug solution.

(II) Class IIND :

This method is used for hypertonic solution i.e. concentration of solution is more than 0.9% w/v NaCl.

Water is added to solution to make it isotonic.

ⓐ White - Vincet method :-

$$\boxed{V = W \times E \times 111.1}$$

where,

V = volume of water added in solution to make it isotonic

W = weight of drug in gram

E = Equivalent weight of drugs (sodium chloride equivalent).

ⓑ Sprowls method :-

It is the simplification of white and vincet method, Here weight of drug (W) is set to constant value of 0.3

$$\boxed{V = 0.3 \times E \times 111.1}$$

$$\downarrow$$

$$\boxed{V = 33.33 \times E}$$

- **Pharmaceutical Importance of buffers**

- Solubility
- Color
- Stability
- Patient comfort
- Optimum pH condition
- Study and research purpose