Buffers

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* It is a solution which resist pH change upon addition of an acidic or basic component

* It is able to neutralize small amount of added acid or base, thus maintaining pH of solution relatively stable.

* This is important for process or reaction which require a stable pH.

* The pH value of buffer is not changed when keeping for long time.

Buffer Capacity ⇒ It is a quantitative measure of solution's ability to resist the change in pH
So if B.C. is larger then it is a good buffer as it can accomodate more acid/Base without altering the pH significantly.

High Buffer capacity → Effective

lower Buffer capacity → less effective

Buffer system

→ Buffer system can be made of weak acid & it's salt or weak base & it's salt.

Types of Buffer

Acidic Buffer

→ weak acid + salt c̄ strong base

→ ex.

1) $CH_3COOH + CH_3COONa$ — CH_3COONa: salt c̄ strong base (NaoH)

2) $H_2CO_3 + NaHCO_3$

3) $H_3PO_4 + NaH_2PO_4$

H_3PO_4: phosphoric acid; NaH_2PO_4: (sod. di hydrogen phosphate)

4) $HCOOH + HCOONa$

$HCOOH$: (Formic Acid); $HCOONa$: (sod. formate)

Basic Buffer

weak base + salt c̄ strong acid

ex.

1) $NH_4OH + NH_4Cl$ — NH_4Cl: salt c̄ strong acid (Hcl)

2) $NH_3 + NH_4Cl$

Mechanism of Action of Buffer

(1) for Acidic Buffer

Considering the Buffer is "$CH_3COOH + CH_3COONa$"

(i) If a strong acid (HCl) in a small amount is added.

⇒ In Buffer system equillibrium is maintained as follows.

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^- \text{ (Partly ionized)}$$

$$CH_3COONa \rightleftharpoons Na^+ + CH_3COO^- \text{ (fully ionised)}$$

⇒ **When HCl is added.**

$$CH_3COONa \rightleftharpoons Na^+ + CH_3COO^- \qquad H^+ + Cl^-$$

CH_3COO^- + H^+ → CH_3COOH W.A.

⇒ H^+ from HCl quickly react c̄ CH_3COO^- & form unionised CH_3COOH (weak acid).

↓

CH_3COOH is weak acid as compared to HCl. hence pH is slightly affected.

⇒ **When Base is added**

OH^- react c̄ excess of CH_3COOH & form H_2O molecule

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

Summary

Addition of OH^- → H_2O

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

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

Addition of H^+ → CH_3COOH

(2) For Basic Buffer

($NH_4OH + NH_4Cl$)

pH of Buffer is govern by following.

- $NH_4OH \rightleftharpoons NH_4^+ + OH^-$

ionization of Ammonium chloride

$NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$ (complete ionised)

when HCl is added

Addition of H^+ → H_2O

$NH_4OH \rightleftharpoons OH^- + NH_4^-$

$NH_4Cl \rightleftharpoons Cl^- + NH_4^+$

when Base is added ($NaOH$) — OH^- → NH_4OH

Buffer Equation

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(Handerson- Hasselbalch equation)

→ The buffer equation is also k/as "Handerson - Hasselbalch equation".

⇒ With the help of this equation, it is possible to calculate the
→ pH of buffer sol" of k/n con"
→ to make buffer of k/n pH.

⇒ There are two separate equations for Acid & Base buffer.

(1) pH of Acidic Buffer

The dissociation of weak acid HA expressed as.

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

so as per law of Mass action

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

} Ka = equillibrium constant.

so
H^+ con^n ⇒

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

or $[H^+] = K_a \frac{(\text{con}^n \text{ of acid})}{(\ldots)}$

Taking -ve log of both side.

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

$$\Rightarrow -\log[H^+] = -\log Ka - \log \frac{[HA]}{[A^-]} \quad \text{---} (1)$$

where

$$-\log[H^+] = pH, \quad \& \quad -\log Ka = pKa$$

$$\Rightarrow pH = pka - \log \frac{[HA]}{[A^-]} \quad \text{---} (2)$$

(or)

$$\Rightarrow pH = pka + \log \frac{[A^-]}{[HA]} \qquad \frac{[Salt]}{(Acid)}$$

(II) For Base

$$BOH \rightleftharpoons B^+ + OH^-$$

$$\Rightarrow Kb = \frac{[B^+][OH^-]}{[BOH]}$$

$$\Rightarrow [OH^-] = Kb \frac{[BOH]}{[B^+]} \quad \text{---} (1)$$

Taking log of both side

$$\Rightarrow -\log[OH^-] = -\log K_b - \log \frac{[BOH]}{[B^+]} \quad \text{—— II}$$

or

$$\Rightarrow -\log[OH^-] = -\log K_b + \log \frac{[B^+]}{[BOH]} \quad \text{—— (III)}$$

Where

$$-\log[OH^-] = pOH, \quad -\log K_b = pK_b$$

$$\Rightarrow pOH = pK_b + \log \frac{[B^+]}{[BOH]}$$

or

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

$$\Rightarrow pH = 14 - pOH$$

Buffer Capacity (β)

It is defined as the ration of increment of strong acid or base to small change in pH brough about by this addition.

$$\beta = \frac{\Delta A \text{ or } \Delta B}{\Delta pH}$$

ΔA or ΔB = small increment in gm equivalent per litre of strong acid or base.

ΔPH = change in pH.

Application of Buffer

(1) In Biological System :→ Blood pH is maintained at 7.4 by two buffers.

(a) Primary buffer :→ In plasma, compounds like carbonic acid, carbonates, sodium salt of phosphoric acid, these all maintain the pH.

(b) Secondary : In erythrocyte, oxy-hb & Hb acid/alkali, pottasium salt of phosphoric acid.

Preparation of Buffers

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① Preparation of $CH_3COOH + CH_3COONa$ Buffers.

Reagent

① Acetic acid : (0.2M)

→ 1.5 ml gl. acetic acid is made 100 ml c̄ dist. H_2O

② CH_3COONa solution

0.64 gm sodium acetate or 2.72 gm sod. acetate trihydrate is dissolve in 100 ml dist H_2O

Procedure

Take 86.2 ml sod. acetate + 14.8 ml gl. acetic acid

↓

make up vol" 100 ml c̄ dist. H_2O

↓

This give 0.2 M $CH_3COOH + CH_3COONa$ Buffer

↓

Measure the pH c̄ pH meter.

⇒ pH = 4 – 4.6 (c̄ 5N NaOH)

Buffer Isotonic Solution
(iso-osmotic)

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⇒ When two solutions are seperated by semipermiable membrane, & it shows same osmotic pressure.

⇒ ex. Body fluid (blood plasma & lacrymal fluid)

↓

They have an iso-osmotic pressure equall to 0.9 % NaCl sol^n

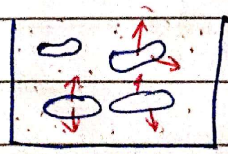
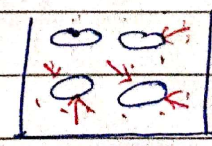
↓

hence 0.9 % NaCl sol^n is isotonic c̄ body fluid.

⇒ Hypertonic :- High osmotic pressure of sol^n

⇒ Hypotonic :- low osmotic pressure of sol^n

Difference B/w Hypertonic & Hypotonic

Hypertonic	Hypotonic
	
H_2O comes out of cell to dilute surrounding	salt solution pass into cell
↓	↓
cell will shrink	cell swell & burst

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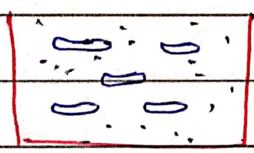
Measurment of Tonicity

(1) Haemolytic Method

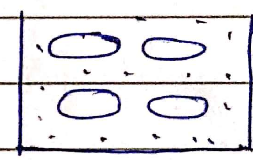
In this method the
↓
R.B.Cs are suspended into the solution to be tested

If cell get shrinkage



↓
solution is Hypertonic

cell get swelled



↓
Sol^n is Hypotonic.

(2) Colligative Method

colligative Property
⇒ Property of solution depends on NO. of solute not Nature of particles.

e.g. ⇒ V.P., Osmotic pressure, B.P., F.P.

* It is based on the principle that "if solutions are isotonic then they have same colligative property.

Method to adjust Tonicity

(1) Freezing point depression method

(2) Molecular Concentration method

(3) Sodium chloride equivallent Method

(4) white vincent Method

(1) freezing point depression Method

following formula is used for calculation of quantity of substance required to make solution isotonic c̄ physiological fluids.

$$\%\ w/v \text{ of adjusting substance} = \frac{0.52 - a}{b}$$

a = freezing point of unadjusted sol" (drug)

b = freezing point of 1% w/v sol" of adjusting substance. (i.e. NaCl)

(2) Molecular Conⁿ Method

$$\text{\% w/v of adjusting substance required} = \frac{0.03\,M}{N}$$

M = Gram molecular weight of substance
N = Number of ion c substance is ionized.

e.g. NaCl

M = wt.
N = Na^+ Cl^-
1 + 1 → 2 ions

(3) Sodium chloride Equivalent Method

⇒ Developed by Mellen & Seltzar
⇒ It is also called "E-Value of drug."

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

L = freezing point of substance

L_{iso} = Specific value of L
M = Mol. wt. of drug.

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⇒ E value represent the "amount of NaCl which will exert toxic effect equall to one gram of that drug.

⇒ The concn of NaCl is isotonic with body fluid → 0.9 gm/100 ml

③ White Vincent Method

This method includes"
"The addition of sufficient quantity of H_2O to drug to prepare isotonic solution.

↓

Make final volume by adding isotonic or buffer solution

$+H_2O$ ↓ iso/Buff

So.

$$V = W \times E \times 111.1$$

V = vol" of isotonic solution that can be prepared by dissolving W gm of drug in H_2O

E = NaCl equivalent of drug.