Buffer Solutions

A buffer is an aqueous solution containing a weak acid and its conjugate base or a weak base and its conjugate acid. Buffer solutions are resistant to pH change because of the presence of an Equilibrium between the acid (HA) and its conjugate base (A^-) Or the base (XOH) and its conjugate acid (X^+). The balanced equation for this reaction is:

$$HA \rightleftharpoons H^+ + A^-$$
$$XOH \rightleftharpoons OH^- + X^+$$

Types of buffer solutions.

1. **Weak acid and its conjugate base.**

Example: - acetic acid with sodium acetate

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$
$$CH_3COONa \rightleftharpoons CH3COO^- + Na^+$$

Sodium acetate (conjugate base)

2. **Weak base and its conjugate acid.**

Example: - Ammonia with Ammonium Chloride

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$
$$NH_4CI \rightleftharpoons NH_4^+ + Cl^-$$

Ammonium Chloride (conjugate acid)

Importance of buffer solution

Blood as a Buffer Solution

1. Blood itself tends to be a buffer solution by keeping its pH value constant. Buffer solutions help in the adjustment of the nature of blood. . If the alkaline nature of blood increases, buffer solutions tend to bring down the pH value of blood. The reverse happens if blood becomes acidic. Acidic nature increases the pH value of blood.

Role of Buffers in Human Body

2. Reactions inside the human body take place in the blood plasma. These reactions might fail to happen if the pH changes. For complete reaction to take place, the pH of the blood should remain constant. Biochemical reactions are quite sensitive, These changes also affect the biological activity of a human being.

Importance of Buffer Solutions to Human Body

3. If the pH value of blood remains in either alkaline or acidic form then it could prove harmful to a human being. It may even lead to death. This may prevent the working of some organs also. This serves to be the best example as to why buffer solutions are important for the body. Buffer solutions prevent the body from permanent damage.

Action of Buffers in Blood Plasma

4. When carbon dioxide dissolves in blood, it decreases the pH value, thereby increasing the acidic content of blood. In this case, alkaline buffers come into play. They tend to mix with the plasma of blood and then neutralize its value. The same happens in the plasma when the alkaline value of blood increases.

Changes in Body due to Buffer Solutions

5. Without buffer solutions, our body may undergo a lot of changes. The enzyme action is regulated by blood. Therefore, the change of pH value also affects the enzymes indirectly. Enzyme actions require low energy involvement. Changes in the temperature of body can effect enzyme action to a wider range. This serves to be the primary reason as to why buffer solutions play a major role in the human body.

Determination the PH.

A. Determination the PH for Strong acid and base.

-The pH of strong acid and strong base can be calculated by the relations.

$$\mathbf{PH = -\log [H^+]} \quad \& \quad \mathbf{POH = -\log [OH^-]}$$

-The relation between PH and POH is:-

$$\mathbf{PH + POH = 14}$$

EX: Determine the PH of 0.1M NaOH solution.
Solution:

POH =–log [OH^+] ⇨ POH= –log [0.1] = 1 ⇨ PH + POH = 14 ⇨ PH = 13

H.W: Calculate the PH value of 4.5 x 10^{-3}M solution of KOH.

Answer=11.65

B. Determination the PH for Weak acid.
The pH of Weak acid can be calculated by the relations

	$CH_3COOH \rightleftharpoons$	$CH3COO^- +$	H^+
t=0	M	0	0
t=t	M–X	X	X

(The dissociation constant of weak acid) $Ka = \frac{[H^+]\,[CH3COO^-]}{[CH_3COOH]} \Rightarrow Ka = \frac{X^2}{(M-X)}$

(M–X) = M because the concentration of each product compared with concentration of reactant very small so (X) ignored

$$\mathbf{Ka = \frac{X^2}{M}}$$

Where :-
X = represent $[H^+]$
M= represent the weak acid concentration

C. Determination the PH for Weak Base.
The pH of Weak base can be calculated by the relations

	$NH_4OH \rightleftharpoons$	$NH_4^+ +$	OH^-
t=0	M	0	0
t=t	M–X	X	X

(The dissociation constant of weak base) $Kb = \frac{[NH_4^+]\,[OH^-]}{[NH_4OH]} \Rightarrow Kb = \frac{X^2}{(M-X)}$

(M–X) = M | because the concentration of each product compared with concentration of reactant very small so (X) ignored

$$\mathbf{Kb = \frac{X^2}{M}}$$

Where:-

X =represent $[OH^-]$

M=represent the weak base concentration

-The relation between ka and kb is:-

$$\mathbf{Kw = Ka \times Kb} \Rightarrow \mathbf{1x10^{14} = Ka \times Kb}$$

$$\mathbf{pKa + pKb = 14}$$

EX: Calculate the PH of 0.1m CH_3COOH. the dissociation constant of acetic acid is 1.8×10^{-5}.

Solution:

$$Ka = \frac{X^2}{M} \Rightarrow 1.8 \times 10^{-5} = \frac{X^2}{0.1} \Rightarrow X^2 = 1.8 \times 10^{-6}$$

$$X = 1.3 \times 10^{-3} = [H^+]$$

$$PH = -\log [H^+] \Rightarrow PH = -\log 1.3 \times 10^{-3} \Rightarrow PH = -(\log 1.3 + \log 10^{-3})$$

$$PH = -(0.11-3) \Rightarrow PH = 2.8$$

H.W: Calculate the POH of 0.4M NH_4OH .the dissociation constant of Ammonium Hydroxide is 1.77×10^{-5}.

Answer=2.6

D. Determination the PH for Salts

- The salts are dissociate into ions
- Ions that are part of strong acid or strong base are neutral like Cl^-,NO_3^-,I^-,OH^- ,ClO_4^-,Br^- ,H^+,Na^+,Ba^+,Ca^+,k^+,Li^+,etc.
- For the other ions which belong to the weak acid or weak base (if the ion contain positive charge "cation = acidic" ,and for negative charge " anion = basic"

- If the salt contains acidic and basic ions together ,compare ka vs kb

  ka **=** kb : neutral , ka **>** kb : acidic , ka **<** kb : basic

Examples

1. KCN $\longrightarrow$ K^+ (Neutral) + CN^- (Basic) = Basic salt
2. NH_4NO_3 $\longrightarrow$ NH_4^+ (Acidic) + NO_3^- (Neutral) = Acidic salt
3. NaCl $\longrightarrow$ Na^+ (Neutral) + Cl^- (Neutral) = Neutral salt
4. NH_4CH_3COO $\longrightarrow$ NH_4^+ (Acidic) + CH_3COO^- (Basic) = both weak, the Kb for NH_3 and Ka for CH3COOH= 1.8×10^{-5} (in this case have to know the value of ka and kb to recognize the type of salt).

   $Kw = Ka \times Kb \Rightarrow 10^{-14} = Ka \times 1.8 \times 10^{-5} \Rightarrow Ka\ (NH_4^+) = 5.56 \times 10^{-5}$

   $Kw = Ka \times Kb \Rightarrow 10^{-14} = 1.8 \times 10^{-5} \times Kb \Rightarrow Kb\ (CH_3COO^-) = 5.56 \times 10^{-5}$

   Ka =Kb : (NH_4CH_3COO) is a Neutral salt

5. C_5H_5NHF $\longrightarrow$ $C5H5NH^+$ (Acidic) + F^- (Basic) = both weak , the Kb for $C_5H_5N = 1.8 \times 10^{-9}$ and Ka for HF = 6.8×10^{-4}

   $Kw = Ka \times Kb \Rightarrow 10^{-14} = Ka \times 1.8 \times 10^{-9} \Rightarrow Ka\ (C5H5NH^+) = 5.56 \times 10^{-6}$

   $Kw = Ka \times Kb \Rightarrow 10^{-14} = 6.8 \times 10^{-4} \times Kb \Rightarrow Kb\ (F^-) = 1.47 \times 10^{-11}$

   Ka **>** Kb : (C_5H_5NHF) is an Acidic salt

- **PH for Basic salt for example KF**

  KF $\longrightarrow$ K^+ (Neutral) + F^- (Basic) = Basic salt

  Dissociation KF $\longrightarrow$ $K^+ + F^-$

  Hydrolysis $F^- + H_2O \rightleftharpoons HF + OH^-$

t=0	M	––	0	0
t=t	M–X	––	X	X

$$Kb = \frac{[HF]\ [OH^-]}{[F^-]} \Rightarrow \mathbf{X^2 = Kb \times M}$$

Where :-

X = represent $[OH^-]$

M = represent the Basic salt concentration.

- **PH for Acidic salt for example NH_4NO_3**

$NH_4NO_3 \longrightarrow NH_4^+$ (Acidic) + NO_3^- (Neutral) = Acidic salt

Dissociation $NH_4NO_3 \longrightarrow NH_4^+ + NO_3^-$

Hydrolysis $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

t=0	M	--	0	0
t=t	M–X	--	X	X

$$Ka = \frac{[NH_3]\ [H3O^-]}{[NH_4^+]} \Rightarrow \mathbf{X^2 = Ka \times M}$$

Where :-

X = represent $[H^+]$

M = represent the Acidic salt concentration.

EX: Calculate the PH of a 0.1 M NaCl solution.
Solution:

$NaCl \longrightarrow Na^+$ (Neutral) + Cl^- (Neutral) = Neutral salt $\Rightarrow$ PH=7

EX: Calculate the PH of a 0.1 M KF solution, ka (KF)=7.2×10^{-4}
Solution:

$KF \longrightarrow K^+$ (Neutral) + F^- (Basic) = Basic salt

$Kw = Ka \times Kb \Rightarrow 10^{-14} = 7.2 \times 10^{-4} \times Kb\ (F^-) \Rightarrow Kb\ (F^-) = 1.38 \times 10^{-11}$

$X^2 = Kb \times M \Rightarrow X^2 = 1.38 \times 10^{-11} \times 0.1 \Rightarrow X = 1.17 \times 10^{-6} = [OH^-]$

$POH = -\log [OH^+] \Rightarrow POH = -\log [1.17 \times 10^{-6}] = -(\log 1.17 - 6) = 5.9 \Rightarrow PH = 8.06$

EX: Calculate the PH of a 0.25 M NH_4NO_3 solution, Kb (NH_3)=1.8×10^{-5}

Solution:

$NH_4NO_3 \longrightarrow NH_4^+$ (Acidic) + NO_3^- (Neutral) = Acidic salt

$Kw = Ka \times Kb \Rightarrow 10^{-14} = Ka\ (NH_4^+) \times 1.8 \times 10^{-5} \Rightarrow Kb\ (NH_4^+) = 5.55 \times 10^{-10}$

$X^2 = Ka \times M \Rightarrow X^2 = 5.55 \times 10^{-10} \times 0.25 \Rightarrow X = 1.17 \times 10^{-5} = [H^+]$

$PH = -\log [H^+] \Rightarrow PH = 4.93$

H.W: What is the PH of a 2.5 M NH_4Cl solution, kb (NH_3)=1.8×10^{-5}

Answer=4.43

H.W: hydrogen cyanide, HCN,is a weak acid with Ka =4.9×10^{-10} .what is the PH of 1.5 M potassium cyanide ,KCN?

Answer=11.74

E. Determination the PH for buffer solutions

- solution containing a weak acid and its conjugate base

$$\mathbf{PH = pk_a + \log \frac{[Salt]}{[Acid]}}$$

Pk_a=acid dissociation constant = – log k_a

- solution containing a weak base and its conjugate acid

$$\mathbf{POH = pk_b + \log \frac{[Salt]}{[Base]}}$$

Pk_b=base dissociation constant = – log k_b

EX: A buffer consists of 0.30M acetic acid (k_a 1.4 x 10^{-5}) and 0.30M sodium acetate what is the PH?

Solution:

$$PH = pk_a + \log \frac{[Salt]}{[Acid]}$$

$Pka = -\log ka \Rightarrow Pka = -\log (1.4 \times 10^{-5}) \Rightarrow Pka = -(\log 1.4 + \log 10^{-5})$

$Pka = -(0.14 - 5) \Rightarrow Pka = 4.85$

$$PH = 4.85 + \log \frac{0.3}{0.3}$$

$PH = 4.85$

H.W: A sample of 0.0001M NH3 and 0.2 M NH_4Cl (K_b= 1.8 x 10^{-5}) Calculate the PH?

Answer=5.9