EQUILIBRIUM

BUFFER SOLUTIONS

❖ BUFFER SOLUTIONS

A solution that resists change in pH value upon addition of small amount of strong acid or base (less than 1 %) or when solution is diluted is called buffer solution.

The capacity of a solution to resist alteration in its pH value is known as buffer capacity and the mechanism of buffer solution is called buffer action.

Types of Buffer Solutions

(A) Simple buffer solution

(B) Mixed buffer solution

(A) Simple Buffer Solution

A salt of weak acid and weak base in water

e.g. CH_3COONH_4, $HCOONH_4$, AgCN, NH_4CN.

Buffer Action of Simple Buffer Solution

Consider a simple buffer solution of CH_3COONH_4,since it is a salt will dissociated completely.

$$CH_3COONH_4 \longrightarrow CH_3COO^- + NH_4^+$$

If a strong acid such as HCl is added then

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

The H^+ ions from the added acid (HCl) combine with CH_3COO^- ions to form CH_3COOH, which is a weak acid so will not further ionized.

Thus, there is no rise in H^+ ion concentration and the pH remain constant.

$$CH_3COO^- + H^+ \leftrightarrow CH_3COOH \text{ (Weak acid)}$$

If a strong base is added as NaOH

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

$$NH_4^+ + OH^- \leftrightarrow NH_4(OH) \text{ (Weak base)}$$

Thus, change in OH^- ion concentration is resisted by $NH^+{}_4$ ions by forming NH_4OH which is a weak base. So it will not further ionized and pH remains constant.

pH of a simple buffer solution: -

(B) Mixed Buffer Solutions

(a) Acidic Buffer Solution

An acidic buffer solution consists of solution of a weak acid and its salt with strong base. The best-known example is a mixture of solution of acetic acid and its salt with strong base (CH_3COONa). Other example :

$$HCN + KCN, (H_2CO_3 + NaHCO_3) \longrightarrow \text{blood}$$

$$CH_3COOH \leftrightarrow CH_3COO^- + H^+ \text{ (Weakly ionised)}$$

$$CH_3COONa \longrightarrow CH_3COO^- + Na^+ \text{ (Highly ionised)}$$

When a few drops of an acid (HCl) are added to it, the H^+ ions from the added acid (HCl) combine with the CH_3COO^- ions to form CH_3COOH. Thus, there is no rise in H^+ ion concentration and the pH of solution remains constant. On the other hand, when a few drops of base(NaOH) are added, the OH^- of the added base reacts with acetic acid to form unionize water and acetate ions.

$$CH_3COOH + OH^- \leftrightarrow H_2O + CH_3COO^-.$$

Thus there is no increase in OH^- ion concentration and hence the pH of the solution remains constant.

pH of an Acidic Buffer Solution (Henderson Equation)

Consider a buffer mixture (acidic buffer)

$$HA + NaA \quad (CH_3COOH + CH_3COONa)$$

where $A = CH_3COO$, $A^- = CH_3COO^-$

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

$$NaA \longrightarrow Na^{+} + A^{-}$$

Applying law of mass action to dissociation equilibrium of HA

$$K_a = \frac{[H^+][A^-]}{[HA]}; \quad \text{so} \quad [H^+] = \frac{K_a[HA]}{[A^-]}$$

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

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

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$[A^-]$ = Initial concentration of salt as it is mainly comes from salt.

$[HA]$ = Initial concentration of the acid.

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

(it is known as Henderson-Hasselbalch equation.)

(b) Basic Buffer Solution

A basic buffer solution consists of a mixture of a weak base and its salt with strong acid. The best known example is a mixture of NH_4OH and NH_4Cl.

$$NH_4OH \leftrightarrow NH_4^{+} + OH^{-} \qquad \text{(Weakly ionised)}$$

$$NH_4Cl \longrightarrow NH_4^{+} + Cl^{-} \qquad \text{(Highly ionised)}$$

When a few drops of a base (NaOH) are added, the OH^- ions from NaOH combine with NH_4^+ ions to form feebly ionised NH_4OH thus there is no rise in the concentration of OH^-ions and hence the pH value remains constant.

$$NH_4^{+} + OH^{-} \longrightarrow NH_4OH$$

If a few drops of a acid (HCl) are added the H^+ from acid combine with NH_4OH to form H_2O and NH_4^+ ions.

$$NH_4OH + H^{+} \longrightarrow NH_4^{+} + H_2O$$

Thus the addition of acid does not increase the H^+ ion concentration and hence pH remains unchanged.

BUFFER CAPACITY

It is defined as the number of moles of acid (or base) require by one liter of a buffer solution for changing its pH by one unit.

$$\text{Buffer capacity} = \frac{\text{No.of moles of acid or bases added per litre}}{\text{change in pH}}$$

Buffer capacity gives the tendency of buffer to resist change in its pH.

Higher is the buffer capacity, smaller will be the change in pH and more efficient will be the buffer.