Buffer Solutions

They are solutions containing both weak acids and their salts or weak bases and their salts. They have the capacity of resisting changes of pH when either acids or alkalis are added to them for example suppose 1ml of 0.1 M HCL is added to 99 ml of pure water pH=7. The $[H^+]$ of the mixture will be about 0.0001M and the pH=3. Suppose howefer,1ml of 0.1M HCl is added to 99ml of buffer solution containing 0.1 M acetic acid and 0.1M sodium acetate. The pH of the buffer solution to begin with 4.73. When the HCl is added, the pH becomes about 4.72 .

A similar condition exists when alkali is added to the buffer mixture 1ml of 0.1 M NaOH added to 99ml of pure gives a pH about 11. Where as if it is added to 99 ml of acetate buffer solution (pH=4.73), the pH becomes 4.74

The chemical mechanism according to which buffer function may be illustrated by what happens when NaOH and HCl are added to an acetate buffer solution are as follows.

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+ \quad \text{weakly dissociated}$$

$$CH_3COONa \longrightarrow CH_3COO^- + Na^+ \quad \text{completely dissociated}$$

Upon addition of NaOH

$$CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$$

CH_3COONa

The addition of the alkalis decreases the CH_3COOH in the buffer and increases the CH_3COONa. The pH of the solution increases in proportion to the change in ratio of salt to acid in the buffer solution.

Upon addition of HCl

CH_3COOH

$$CH_3COONa + HCl \longrightarrow CH_3COOH + NaCl$$

In this case the HCl reacts decrease CH_3COONa in the buffer . the pH of the solution falls in proportion to the change in ratio of salt to acid in the solution.

The Henderson – Hasselbalach equation for buffer solution :

In the equilibrium reactions for the dissociation of HA and BA in a buffer solution , we have

$$HA \rightleftharpoons H^+ + A^- \quad \textbf{weakly dissociated}$$

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

$$BA \longrightarrow B^+ + A^- \quad \textbf{completely dissociated}$$

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

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

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

$$[H^+] = K_a \frac{[A^-]}{[HA]} \quad \text{or} \quad [H^+] = K_a \frac{[HA]}{[BA]} \quad \text{or} \quad [H^+] = K_a \frac{[Acid]}{[Salt]}$$

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

$$= -\left(\log K_a + \log \frac{[HA]}{[BA]}\right)$$

$$= -\log K_a - \log \frac{[HA]}{[BA]}$$

$$\text{and } -\log K_a = pK_a \quad , \quad -\log \frac{[HA]}{[BA]} = \log \frac{[BA]}{[HA]}$$

$$pH = pK_a + \log \frac{[BA]}{[HA]} \quad \text{or} \quad pH = pK_a + \log \frac{[salt]}{[acid]}$$

A similar equation that would apply to a basic buffer is :

$$pOH = pK_b + \log \frac{[salt]}{[base]}$$

Capacity :

It is the larger the concentration of the components of the buffer, the more effective it is at resisting changes in pH.

$$pH = pK_a + \log \frac{[\mathbf{salt}]}{[\mathbf{acid}]}$$

Buffered solutions are very frequently used in (quantitative) analytical chemistry for example.

1- Ensure the proper progress of a reaction by removing H_3O^+ or OH^- reaction product as the save formed.

2- Control the level of pH during precipitation process.

3- Maintain a required level of pH in order to guarantee the proper reaction of an indicator substance.

4 – Are encountered at certain stages of titration processes.