

pH and Buffer Solution

pH of the solution:- The pH of the solution is the negative logarithm of the concentration of hydrogen ions which it contains.

$$\boxed{\text{pH} = -\log[\text{H}^+]}$$

* The conc. of $[\text{H}^+]$ must be in moles/litre.

- (i) If $\text{pH} = 7$ - solution is neutral.
- (ii) If $\text{pH} < 7$ - " " Acidic.
- (iii) If $\text{pH} > 7$ - " " Basic.

Similarly $\rightarrow \boxed{\text{pOH} = -\log[\text{OH}^-]}$

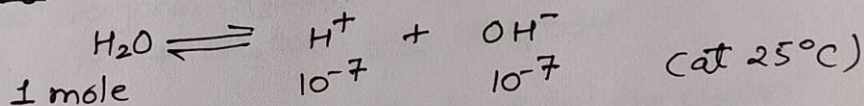
* $[\text{OH}^-]$ - must be in moles/litre.

$$\boxed{\text{pH} + \text{pOH} = 14}$$

$$\boxed{[\text{H}^+][\text{OH}^-] = 10^{-14}}$$

pH Scale:- Given by "Sorensen" - It extends from 0 - 14.

Dissociation of water:- Very less.



Water dissociates very less at 25°C .

* If the conc. of Acid or base is less than 10^{-6} (10^{-7} or 10^{-8} or less) - then only dissociation of H_2O is considered.

* If suppose HCl conc. is 10^{-3}M . then dissociating water is not considered. As from water we get $10^{-7}[\text{H}^+]$ very less as compared to 10^{-3} .

$$\text{As } - 10^{-3} + 10^{-7} \approx 10^{-3}$$

\downarrow
very small as compared to 10^{-3} .

Buffer Solution:-

A buffer solution is one which can resist change in its pH on the addition of an acid or a base.

OR.

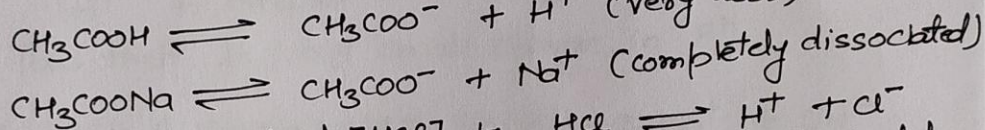
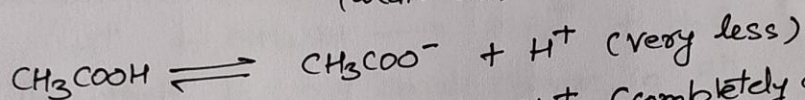
A buffer solution is one which maintains its pH fairly constant even upon the addition of small amount of Acid or base.

e.g. $CH_3COOH + CH_3COONa$ - Acidic buffer
 $NH_4OH + NH_4Cl$ - Basic buffer.

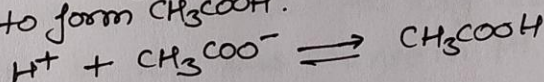
$$\text{Buffer capacity} = \frac{\text{No. of moles of acid or base added to 1 litre of buffer}}{\text{change in } pH}$$

Acidic Buffer & its Mechanism:-

Acidic buffer - $CH_3COOH + CH_3COONa$
(weak acid) (salt)

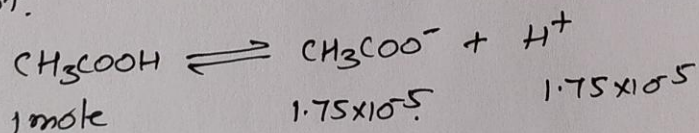


(i) Addition of Strong Acid $[HCl]$:- $HCl \rightleftharpoons H^+ + Cl^-$
The H^+ of the strong acid will be taken up immediately by CH_3COO^- to form CH_3COOH .



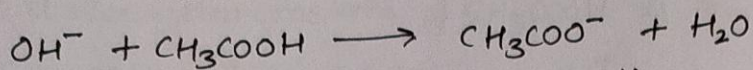
Thus H^+ ions added are neutralized by the CH_3COO^- present in the mixture so there will be very little change in pH of mixture.

* Dissociation constant of CH_3COOH at $25^\circ C$ is 1.75×10^{-5} .
i.e. from 1 mole of CH_3COOH we get only 1.75×10^{-5} moles of H^+ .



(ii) Addition of strong base - $\text{NaOH} \rightleftharpoons \text{Na}^+ + \text{OH}^-$

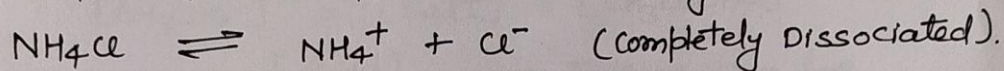
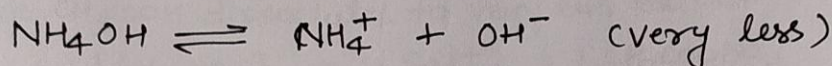
The OH^- ions added are neutralized by the acetic acid present in the mixture.



Thus there is very small change in pH.

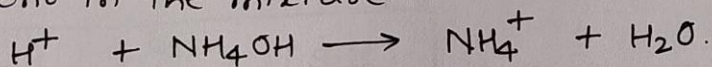
Basic buffer and its Mechanism :-

Basic buffer - NH_4OH + NH_4Cl
(weak base) (Salt)



(i) Addition of strong Acid :- $\text{HCl} \rightleftharpoons \text{H}^+ + \text{Cl}^-$

The H^+ ions added are neutralized by the NH_4OH present in the mixture.

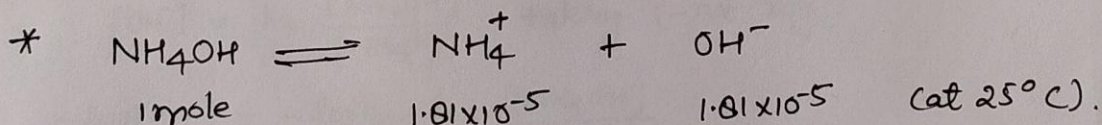


Thus there is very small change in pH.

(ii) Addition of strong base :- $\text{NaOH} \rightleftharpoons \text{Na}^+ + \text{OH}^-$

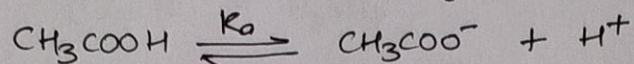
The OH^- ions added are taken up by NH_4^+ ions forming NH_4OH . ~~which~~ which dissociated to very less extent.

So there will be very less change in pH.



Henderson equation :-

- (i) Acrylic Buffer + $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$.
let dissociation constant of CH_3COOH is K_a .



$$\Rightarrow K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

$$\Rightarrow [\text{H}^+] = K_a \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]}$$

Since CH_3COOH dissociates to very less extent.

$$[\text{CH}_3\text{COO}^-] = [\text{Salt}] = [\text{CH}_3\text{COONa}]$$

$$\Rightarrow [\text{H}^+] = K_a \frac{[\text{Acid}]}{[\text{Salt}]} \rightarrow \text{Taking } (-)\text{ve log}$$

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

$$\boxed{\text{pH} = \text{p}^{K_a} + \log \frac{[\text{Salt}]}{[\text{Acid}]}}$$

- (2) Basic buffer + $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$
 $\text{NH}_4\text{OH} \xrightleftharpoons{K_b} \text{NH}_4^+ + \text{OH}^-$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_4\text{OH}]}$$

$$[\text{OH}^-] = K_b \frac{[\text{NH}_4\text{OH}]}{[\text{NH}_4^+]} \text{ taking } (-)\text{ve log.}$$

$$-\log[\text{OH}^-] = -\log K_b - \log \frac{[\text{NH}_4\text{OH}]}{[\text{NH}_4^+]}$$

$$[\text{NH}_4^+] = [\text{Salt}]$$

$$-\log[\text{OH}^-] = -\log K_b - \log \frac{[\text{Base}]}{[\text{Salt}]}$$

$$\boxed{\text{pOH} = \text{p}^{K_b} + \log \frac{[\text{Salt}]}{[\text{Base}]}}$$