Ionic Equilibrium

* **Electrolytes** :- A substance in aqueous form which dissociates into ions and can conduct electricity.

It is divided into 2 types :-

i) **Strong electrolyte** :- are electrolytes which completely dissociates into ions in water.
Ex, $NaCl$; HCl, H_2SO_4.

ii) **Weak electrolyte** :- are electrolytes which partially dissociates into ions in water.
Ex, CH_3COOH, NH_4OH, HCN.

* **Ionic Equilibrium** :- Equilibrium involving ions in aqueous solution is called ionic equilibrium.

*
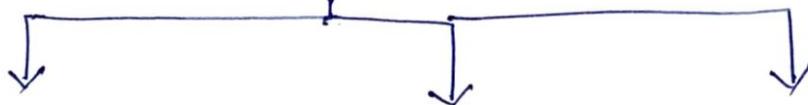
Electrolytes

Acid

(H_2SO_4, HCl, H_2CO_3)

Base

$[NH_4OH, NaOH, Ba(OH)_2]$

Salts

$[NaCl, NaHCO_3, NH_4Cl]$

Ostwald Dilution Law :- (for weak electrolyte)

$$AB \rightleftharpoons A^+ + B^-$$
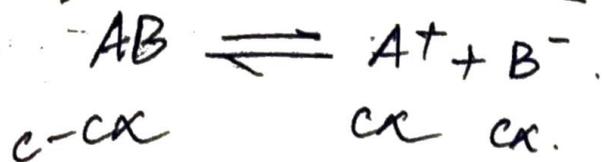
$$c - c\alpha \qquad c\alpha \qquad c\alpha. \qquad \alpha \longrightarrow \text{degree of dissociation}.$$

$$K = \frac{c\alpha^2}{1-\alpha} \ell = c\alpha^2 \quad [\ 1-\alpha \approx 1 \ \text{for weak electrolyte}).$$

$$\therefore \boxed{\alpha = \frac{1}{\sqrt{c}}.}$$

* Concept of Acids and Bases :-

1. Arrhenius Concept :-

* Acids are substances that dissociates in water to give H^+ ions.

* Bases are substances that produce OH^- ions when dissolved in water.

2. Bronsted - Lowry Theory :-

• Acid is a substance that is capable of donating a H^+ ion.

• Bases are substances capable of accepting a H^+ ion.

In short,
Acids are proton donors.
Bases are proton acceptors.

Ex,

$$\text{NH}_3 (aq) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{NH}_4^+ (aq) + \text{OH}^- (aq)$$

adds proton →

loses H^+

conjugate acid

conjugate base
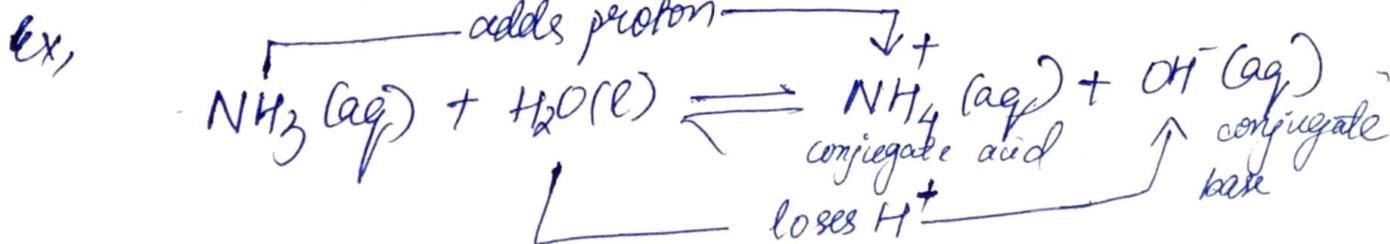
3) Lewis Acid and Bases

Acid — which accepts electron pair.

Base — which donates an electron pair.

- Electron deficient species like $AlCl_3$, Co^{3+}, Mg^{2+} act as Lewis acid.

- Electron rich species like H_2O, NH_3, OH^-, which can donate a pair of electrons act as Lewis base.

~~Ionization of Acids and Bases~~

* Ionization Constant of H_2O.

$$H_2O + H_2O \rightleftharpoons H_3O^+_{(aq)} + OH^-(aq).$$
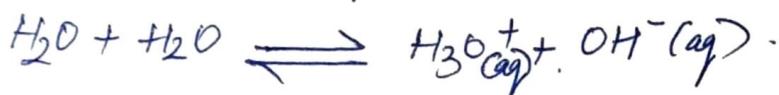
$$K = \frac{[H_3O^+][OH^-]}{[H_2O]}.$$

$$\boxed{K_W = [H_3O^+][OH^-]} \longrightarrow \text{Ionic Product}$$

i) $[H_3O^+] > [OH^-] \longrightarrow$ Acidic.

ii) $[H_3O^+] = [OH^-] \longrightarrow$ Neutral.

iii) $[H_3O^+] < [OH^-] \longrightarrow$ Basic.

* pH scale

$$pH = -\log[H^+].$$ — H^+ concentration in mol/L

For acidic soln, $pH < 7$,

Basic soln, $pH > 7$,

Neutral soln, $pH = 7$.

At 298K, $K_W = [H_3O^+][OH^-] = 10^{-14}$.

So, $\boxed{pK_W = pH + pOH = 14}$

Ionization Constants of Weak Acids and Weak Bases.

$$HX(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + X^-(aq).$$

Initial conc. c $\qquad\qquad\qquad\qquad$ 0 $\qquad$ 0.

$\qquad\qquad$ $c - c\alpha$ $\qquad\qquad\qquad\qquad$ $c\alpha$ $\qquad$ $c\alpha$.

$$Ka = \frac{c^2\alpha^2}{c(1-\alpha)} = \frac{c\alpha^2}{1-\alpha}.$$

Ka = dissociation or ionization constant.

$$MOH(aq) \rightleftharpoons M^+(aq) + OH^-(aq).$$

$$Kb = \frac{[M^+][OH^-]}{[MOH]}.$$

$$Kb = \frac{(c\alpha)^2}{c(1-\alpha)} = \frac{c\alpha^2}{1-\alpha}.$$

$$\boxed{Kw = Ka \times Kb.}$$

$$\therefore \quad pKa + pKb = pKw = 14.$$