IONIC EQUILIBRIUM

Electrolytes and Non Electrolytes

Electrolytes - Which dissociate into ions

Strong Elec. ($\alpha=1$)	• All strong acids and salts (HCl, HNO$_3$) etc. • All Alkali metal hydroxides and Ba(OH)$_2$
Weak Elec. ($0< \alpha <1$)	• HNO$_2$, H$_2$SO$_3$, HClO, All oxy acids of P, All organic acids, HCN, H$_3$BO$_3$, H$_2$CO$_3$ etc. • All Alkaline hydroxides except Ba(OH)$_2$, All organic bases, D-series hydroxides.

Theories of acids and bases

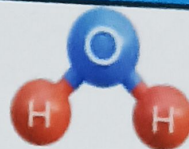| Theory	Arrhenius	Bronsted	Lewis
ACIDS	Donates H$^+$	Donates H$^+$	Lone Pair Acceptor
BASES	Donates OH$^-$	Accepts H$^+$	Lone Pair Donor

NOTE

• Strong Acid $\rightarrow$ H$^+$ + Weak Conjugate base **and**
• Strong Base + H$^+$ $\rightarrow$ Weak Conjugate base **VICE VERSA**

• Amphiprotic acids are H$^+$ donor and acceptor, H

27

Ionic product of water

At 25°C, $K_w = [H^+][OH^-] = 10^{-14}$
For Pure water, $[H^+] = [OH^-] = 10^{-7}$

K_w increases with increase in temperature.

pH Calculation

$pH = -\log[H^+]$	$pOH = -\log[OH^-]$

At 25°C, $pH + pOH = pK_w$

Case I : Strong acids/bases with $[H^+] \geq 10^{-6}$

- Directly take the log of concentration of H^+

Case II : Strong acids/bases with $[H^+] < 10^{-6}$

$pH = -\log(H^+_{acid} + H^+_{water})$ *Here,* $H^+_{water} = 10^{-7}$

Case III : Mixture of Strong acids/bases

Two Strong Acids	Two Strong Bases
$[H^+] = \dfrac{N_1V_1 + N_2V_2}{V_1 + V_2}$	$[OH^-] = \dfrac{N_1V_1 + N_2V_2}{V_1 + V_2}$

Case IV : Mixture of Strong acids/bases

H_3PO_4 will have three dissociation constants : Ka_1, Ka_2 & Ka_3. The calculation of pH is dependent upon,	$pH = \dfrac{1}{2}[pKa_1 - \log c]$ not dependent upon the other constants

28

4. Add Equation with Eq. constant K_2 = $K.K_2$
5. Subtract Equation with Eq. constant K_2 = K/K_2

24

Case V : Case of Weak mono acids and bases

For a weak acid, $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{c\alpha \times c\alpha}{c(1-\alpha)} = \frac{c\alpha^2}{1-\alpha}$$

More the value of K_a
Stronger is the acid

$$1-\alpha \approx \quad \alpha = \sqrt{\frac{K_a}{c}}$$

α = Degree of Dissociation
c = Concentration of HA

$$pH = -\log_{10}\sqrt{K_a c} = \frac{1}{2}(pK_a - \log_{10}c)$$

Salt Hydrolysis

Hydrolysis Constant, K_H	pH of solution
Weak Acid - Strong Base (pH > 7)	
$K_h = \dfrac{ch^2}{1-h} = \dfrac{K_w}{K_b}$	$pH = \dfrac{1}{2}[pK_w - (pK_b + \log c)]$
Weak Base - Strong Acid (pH < 7)	
$K_h = \dfrac{ch^2}{1-h} = \dfrac{K_w}{K_a}$	$pH = \dfrac{1}{2}[pK_w + (pK_a + \log c)]$
Weak Base - Weak Acid (pH depends on K_a and K_b)	
$K_h = \dfrac{h^2}{(1-h)^2} = \dfrac{K_w}{K_a \cdot K_b}$	$pH = \dfrac{1}{2}[pK_w + (pK_a - pK_b)]$

29

Buffer Solution

Acidic Buffer	Basic Buffer
WA + Salt of WA&SB	WB + Salt of WB&SA
• H_2CO_3 + $NaHCO_3$ • CH_3COOH + CH_3COONa	• NH_4OH + NH_4Cl • CH_3NH_2 + $CHNH_3{}^+Cl^-$

Henderson Equation

$$pH = pK_a + log_{10} \frac{[Salt]}{[Acid]} \text{ or } \frac{[Conjugated\ Base]}{[Acid]}$$

Solubility Product (K_{sp})

For a salt A_xB_y $$s = \left[\frac{(K_{sp})}{x^x \cdot y^y}\right]^{\frac{1}{x+y}}$$	In case of Common ion, the Solubility decreases as reaction moves backwards. E.g. AgCl in Cl^- $$s = \frac{K_{sp}{}^{AgCl}}{[Cl^-]}$$

• In case of complex formation, solubility of salt is maximum.

Ionic Product of salt

For a salt A_xB_y

$$K_{IP} = [A^{y+}]^x [B^{x-}]^y$$

$K_{IP} > K_{sp}$ Ppt occurs

$K_{IP} < K_{sp}$ Unsaturated soln

$K_{IP} = K_{sp}$ Saturated s30