

Electrolyte

The substance which conduct electricity in solution or molten state is called electrolyte. Eg: NaOH , H_2SO_4 , $\text{Ca}(\text{OH})_2$ etc.

Types of Electrolyte

1. Strong Electrolyte: Those electrolyte which ionize almost completely and they have high conducting power are called strong electrolyte. Eg: HCl , H_2SO_4 etc

2. Weak Electrolyte: Those electrolyte which ionize partially and have low conducting power are called weak electrolyte.
Eg: HCN , HCOOH , NH_4Cl etc.

Degree of Ionization (α)

The ratio of total number of molecules / reactant ionized to the total number of molecules dissolved, is called degree of ionization or dissociation.

$$\text{i.e. } \alpha = \frac{\text{No. of reactant dissociated}}{\text{Total number of molecules}}$$

Note:

Degree of dissociation (α) for strong electrolytes is nearly equal to 100% / 1 (in term of number)

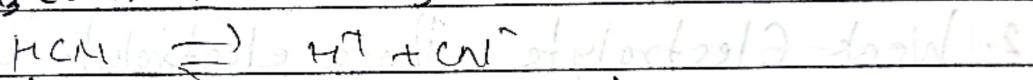
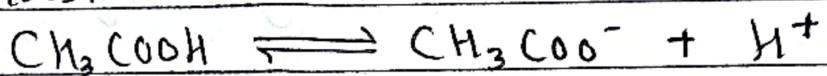
Degree of dissociation (α) for weak electrolytes is around 5% $\rightarrow < 1$ (in terms of number)

Degree of ionization depends upon:

- i. Nature of electrolyte
- ii. Nature of solvent
- iii. Temperature
- iv. Dilution
- v. Presence of other ions in solution.

Ionization of Weak Electrolyte

Let us consider a weak electrolyte, CH_3COOH dissociate as follows:



According to the Law of Mass Action

Rate of backward reaction = rate of forward Rxn

Here,

Rate of forward reaction $\propto [\text{CH}_3\text{COOH}]$

$$\text{i.e. R.O.F} = k_f [\text{CH}_3\text{COOH}]$$

Rate of Backward reaction $\propto [\text{CH}_3\text{COO}^-] [\text{H}^+]$

$$\text{i.e. R.O.B} = k_b [\text{CH}_3\text{COO}^-] [\text{H}^+]$$

At Equilibrium

$$\text{R.O.F} = \text{R.O.B}$$

$$\text{or, } k_f [\text{CH}_3\text{COOH}] = k_b [\text{CH}_3\text{COO}^-] [\text{H}^+]$$

$$\text{or, } \frac{k_f}{k_b} = \frac{[\text{CH}_3\text{COO}^-] [\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

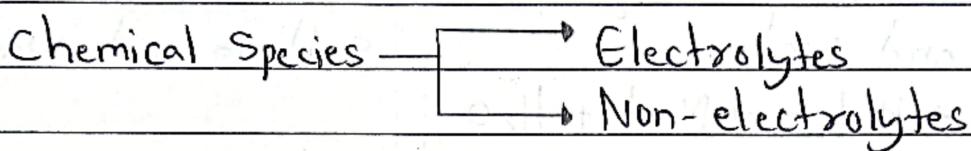
Cases:

1. If $K_{\text{equilibrium}} > 1$, reaction is favourable towards forward direction.
i.e. Product $>$ reactant

2. If $K_{eq} < 1$, reaction is favourable in backward direction, i.e. Product < reactant.

3. If $K_{eq} = 1$ forward and backward reaction is equal, i.e. reactant = Product

Chemical Species



Electrolytes → Strong → Strong Acid = H_2SO_4 , HCl , HNO_3

→ Strong base = $NaOH$, KOH

→ Strong salt = $NaCl$, KCl ,

$CaCO_3$, Na_2CO_3

→ Weak → Weak Acid = HCN , $HCOOH$, H_2S

C_6H_5COOH

→ Weak Base = CH_3-N-CH_3 ,
 $\begin{array}{c} | \\ H \end{array}$

$CH_3-N=CH_3$
 $\begin{array}{c} | \\ CH_3 \end{array}$

→ Weak Salt = NH_4Cl , $(NH_4)_2CO_3$ etc

Non-Electrolytes = Oil, sugar, petrol etc.

Acid and Base Concept : (Classification of Acid and Base)

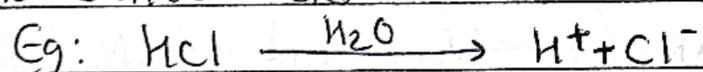
1. Classical Concept

Acid	Base
i. Sour in taste, turns blue litmus into red.	i. Bitter in taste, turns red litmus to blue.
ii. React with base to give salt and water.	iii. React with acid to give salt and water.
$\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$	
iii. React with metal to give $\text{H}_2\uparrow$	

2. Modern Concept of Acid and Base

a. Arrhenius Concept

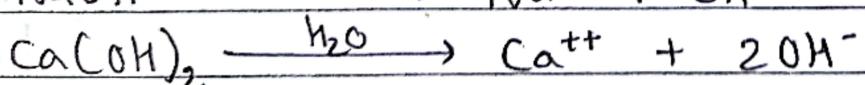
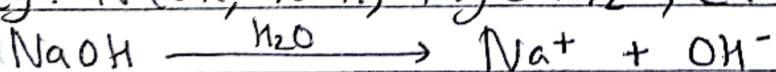
Any chemical species that can produce H^+ in water is called acid.



H_2SO_4 , H_2CO_3 , H_2S , HCN etc.

b. Any species that can produce OH^- in water is base.

Eg: NaOH , KOH , Mg(OH)_2 , Ca(OH)_2 etc.



Merits

- Chemical species that has H and OH can be classified as acid and base.
- 1st to classify acid and base in modern way.

Limitations

- It can't explain acidic and basic nature of species in gaseous phase. Eg: HCl (gas), NH₃ (gas) etc.
- It can't explain acidic nature of electron deficient compounds like AlCl₃, BF₃, BeH₂ etc.
- It can't explain acidic nature of non-metallic oxides like CO₂, NO₂ etc. and basic nature of metallic oxides like Na₂O, CaO, MgO etc.
- This concept only depends on aqueous solution.

b. Bronsted - Lowry Concept (Purely 'H' Concept)

- According to this concept 'H' donor are said to be acid and 'H' acceptors are base despite any medium.

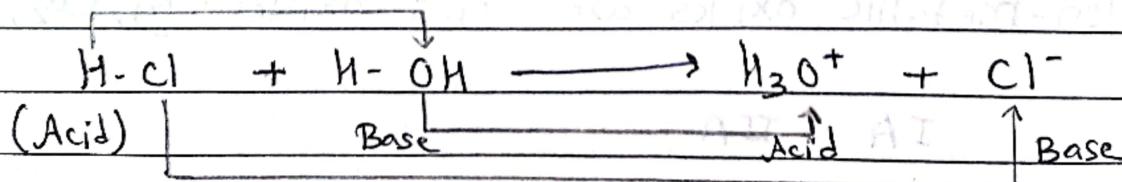
Acid: HCl, H₂SO₄, H₂CO₃, H₂S, HCN etc.

Base: Cl⁻, H₂O⁺, NH₃, SO₄²⁻, CH₃- $\overset{\ddot{\text{N}}}{\underset{\text{H}}{\text{—}}}$ -CH₃ etc.

(2-amine)

(All these species with lone pair and negative charge).

Concept of conjugate Acid and Base Pair



Here,

Acid gives base thus H-Cl and Cl⁻ are conjugate base pair.

Base gives acid thus H_2O and H_3O^+ are conjugate acid-base pair.

Merits of Bronsted-Lowry Concept

- It can also classify chemical species as acid and base but in both gaseous and aqueous stage I phase.
- Negative and neutral species with lone pair are also considered base.

Limitations

- Can't explain acidic nature of oxides like CO_2 , SO_3 and basic property of MgO , CaO etc.
- Fails to explain the acidic behaviour of salt like AlCl_3 , FeCl_3 , etc.

C. Lewis Concept (Purely Electronic Concept)

According to it "All electron donating chemical species are base". Eg: NH_3 , $\text{H}_2\ddot{\text{O}}$, SO_4^{2-} , Cl^- , Br^- , HSO_4^- etc.

According to it "All electron accepting chemical species are acid". Eg: BF_3 , AlCl_3 , BeH_2 etc.

Note:

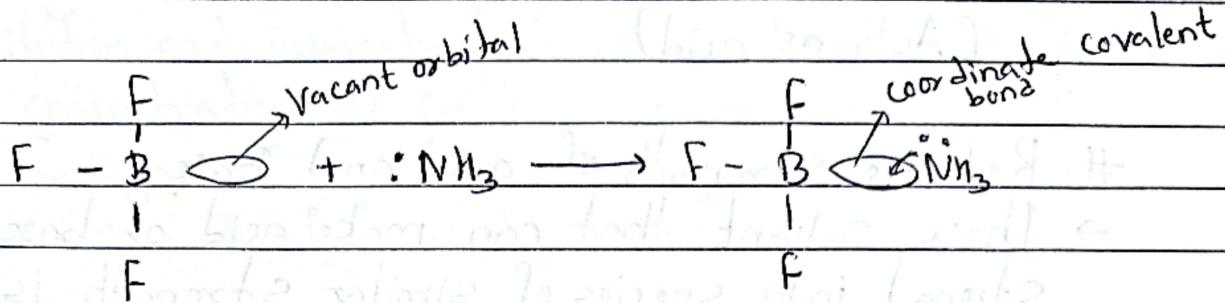
Metallic oxides are basic oxides: Eg Na_2O , MgO , CaO etc.

Non-metallic oxides are acidic oxides: Eg, CO_2 , SO_3 etc.

IA IIA

	H		
Alkali Metal	Li	Be	← Alkaline earth metal
	Na	Mg	
	K	Ca	

According to this concept acid reacts with base to give coordinate covalent bond.



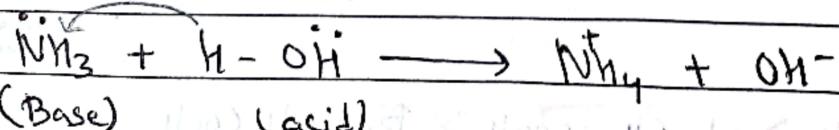
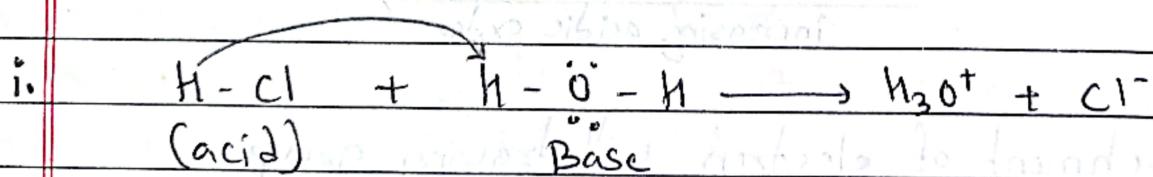
Merits

- It is general concept and can explain about any species as acid and Base.
 - No requirement of a particular medium.
 - Explain metallic oxides as base and non-metallic oxides as acid.
 - Explain electron deficiency deficient compounds like AlCl_3 , BF_3 as acid as they accepts electrons.

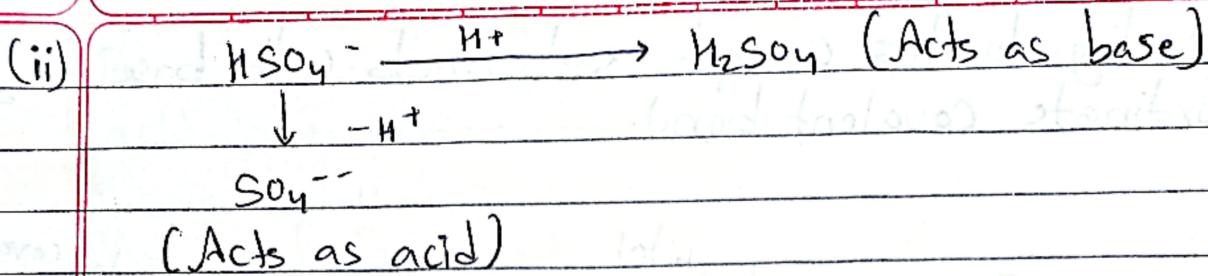
Limitations

- Acid, Base reaction are usually fast but formation of coordinate covalent bond is usually time consuming.

Amphoteric Species (Acts as both acid and Base)



∴ Water in above example behaves as both acid and base so it is an amphoteric in nature.



Relative strength of acid and Base

→ Those solvent that can make acid or base (weak or strong) into species of similar strength is levelling solvent.

Examples:

- i. H_2O if strong acid H_2SO_4 , HCl , HNO_3 and weak acid H_2CO_3 , HCN , H_2S are kept, both types of acid ionize completely.
 \therefore Both seems strong acid in water as Solvent.

- ii. But in glacial acetic acid, strong and weak acid and base are clearly separated. Therefore, glacial acetic acid is called differentiating solvent.

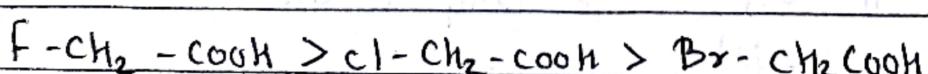
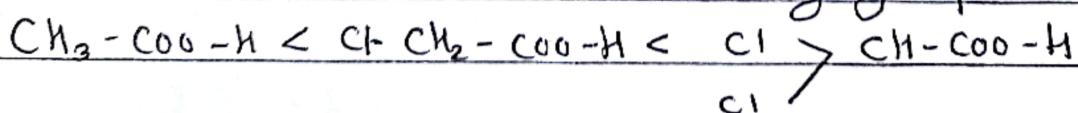
Strength of Acid depends on:

- a. Acidic strength depend on electronegativity

Electronegativity order : $\text{C} < \text{N} < \text{O} < \text{F}$

$\text{CH}_4 < \text{NH}_3 < \text{H}_2\ddot{\text{O}} < \text{HF}$
 Increasing acidic order

- b. Attachment of electron withdrawing group



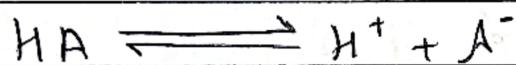
Electro Negativity	F
Cl	
Br	
I	

Ionization of Weak Electrolyte [Ostwald's dilution Law]
 → It states that the degree of dissociation of weak electrolytes is directly proportional to square root of its dilution and inversely proportional to its square root of concentration at constant temperature.

$$\text{i.e. Degree of Dissociation } (\alpha) \propto \frac{1}{\sqrt{c}}$$

$$\text{and, D. of dissociation } (\alpha) \propto \frac{1}{\sqrt{c}}$$

Let us consider a weak electrolyte HA ionized (dissociated) as follows:



Let 'c' be the initial concentration of electrolytes HA and 'α' be the degree of dissociation.



$$\begin{array}{lll} \text{Initial Conc. } c & 0 & 0 \\ \text{final Conc. } & c - c\alpha & c\alpha \end{array}$$

Now,

According to the law of mass action, ionization or dissociation constant is written as;

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

$$\text{or, } K_a = \frac{c\alpha \times c\alpha}{c - c\alpha}$$

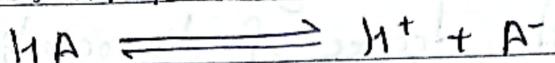
$$\text{or, } K_a = \frac{c\alpha^2}{c - c\alpha} \quad [\alpha \text{ is so negligible for weak acid, so in } 1 - \alpha, \alpha \text{ is ignored}]$$

$$\text{or, } K_a = c\alpha^2$$

$$\text{or, } \alpha = \sqrt{\frac{K_a}{c}}$$

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

When reactant is 1 mole Initially:



Initially: 1 mole 0 mole 0 mole

AT equilibrium:

$$1-\alpha \xrightarrow{V} \alpha \xrightarrow{V} \alpha \xrightarrow{V}$$

Concentration
⇒ mole
volume

Now,

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

$$\text{or, } K_a = \frac{\alpha \times \alpha}{V(1-\alpha)}$$

$$\text{or, } K_a = \frac{\alpha^2}{V^2 \times (1-\alpha)} \quad \text{Note: } V \text{ is constant}$$

$$\text{or, } K_a = \frac{\alpha^2}{V(1-\alpha)} \quad \text{Note: } \alpha \ll 1, \text{ so it is}$$

$$\text{or, } K_a = \frac{\alpha^2}{V} \quad \text{neglected}$$

$$\text{or, } \alpha = \sqrt{K_a \cdot V}$$

$$\therefore \alpha \propto \sqrt{V}$$

Limitation of Ostwald's Dilution Law

→ It is applicable only in case of weak monobasic acid like H_2S , CH_3COOH , HCN and weak monoacidic base like NH_4OH but not valid for dibasic acid and diacidic base.

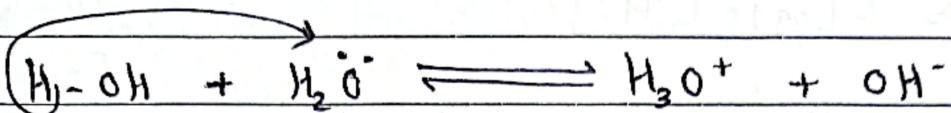
In Expression $K_a = \frac{C\alpha^2}{(1-\alpha)}$, $\alpha=1$ for strong electrolyte

$$\therefore K_a = \frac{C\alpha^2}{1-1}$$

= ∞ (\because This law is not applicable)

Ionization of water

In 1 litre of water at 25°C (standard temperature and pressure)



At Equilibrium,

$$K_w = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}][\text{H}_2\text{O}]}$$

$$\text{or, } K_w [\text{H}_2\text{O}]^2 = [\text{H}_3\text{O}^+][\text{OH}^-]$$

Since,

Very little H_2O is dissociated, experimentally 55.4 mole/litre of water (H_2O) remains undissociated
So, $[\text{H}_2\text{O}]$ can be taken as constant 1.

$$\text{i.e. } K_w \times 1 = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$\text{or, } K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

Experimentally value of $K_w = 10^{-14}$

$$\text{Then, } 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-] \quad \text{--- (1)}$$

Since water is neutral,

$$[\text{H}_3\text{O}^+] = [\text{OH}^-]$$

eq (i) becomes

$$[\text{H}_3\text{O}^+][\text{H}_3\text{O}^+] = 10^{-14}$$

$$\text{or, } [\text{H}_3\text{O}^+]^2 = 10^{-14}$$

$$\text{or, } \text{H}_3\text{O}^+ = (10^{-14})^{1/2}$$

$$\therefore \text{H}_3\text{O}^+ = 10^{-7} \text{ mole/litre}$$

Also,

$$\text{OH}^- = \text{H}_3\text{O}^+ = 10^{-7} \text{ mole/litre}$$

PH : Power of Hydrogen

PH is defined as the negative logarithm of hydrogen ion concentration of base 10.

$$PH = -\log_{10} [H^+]$$

What is the PH of Water?

We know,

$$PH = -\log_{10} [H^+]_{H_2O}$$

$$= -\log_{10} [10^{-7}]$$

$$= -7 \log_{10} (10) = -7 \log_{10} 10$$

$$= 7 \times 1$$

Note: PH of acidic solution $[H^+] > 10^{-7}$ mole/Litre

PH of acidic solution $[H^+] > 10^{-7}$ mole/Litre

$$\text{i.e. } [H^+] = 10^{-6}, 10^{-5}, 10^{-4}, 10^{-3}, 10^{-2}, 10^{-1}$$

PH of basic solution $[H^+] < 10^{-7}$ mole/Litre

$$\text{i.e. } [H^+] = 10^{-8}, 10^{-9}, 10^{-10}, 10^{-11}, 10^{-12}, 10^{-13} \text{ and } 10^{-14}$$

$$\therefore \text{PH} = -\log(10^{-8})$$

$$= 8 \text{ and so on}$$

$\therefore \text{PH}$ of basic solution lies between 8-14.

→ PH scale lies between 0-14.

→ $\text{PH} 7$ in PH scale represent the neutral solution.

P_{Kw} / P_{Ka} / P_{Kb}

$$P_{Kw} = -\log_{10}(K_w)$$

$$P_{Ka} = -\log_{10}(K_a)$$

$$P_{Kb} = -\log_{10}(K_b)$$

Relation between K_a / PK_a and K_b and PK_b .

- If $K_a \uparrow$ $PK_a \downarrow$ then the solution is more acidic. (contain more H^+)
- If $K_b \uparrow$ $PK_b \downarrow$ then the solution is more basic. (contain more OH^-)

$$K_w = [H_3O^+] [OH^-]$$

$$\text{or, } 10^{-14} = [H_3O^+] [OH^-]$$

$$\text{Taking -Log on both side.} \\ -\log_{10}(10^{-14}) = -[\log[H_3O^+] + \log_{10}[OH^-]]$$

$$\text{or, } 14 = -\log_{10}[H_3O^+] - \log_{10}[OH^-]$$

$$\text{or, } 14 = pH + pOH$$

$$\therefore pH + pOH = 14$$

Common ion Effect

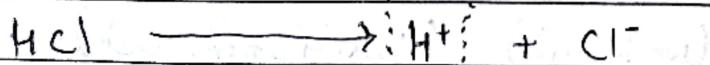
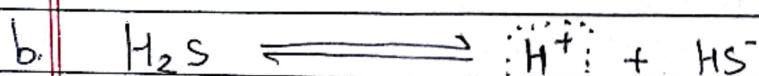
According to it "The degree of dissociation of weak electrolytes is suppressed in the presence of strong electrolyte having one ions common."

Example:



As HCl is strong electrolyte, concentration of H^+ in solution increases.

Thus, the CH_3COO^- and H^+ combines rapidly as per le-chatelier giving the undissociated CH_3COOH .



$\therefore H^+$ and HS^- gives back to H_2S .

pH of Mixture

1. Mixture of Acid and Acid

$$\text{Strength (S)} = \frac{m_1 v_1 + m_2 v_2}{v_1 + v_2}$$

Where, $m_1 v_1$ are conc. of Acid₁ in molarity and volume in litre.

$m_2 v_2$ are conc. of Acid₂ in molarity and volume in litre.

$$\therefore \text{Molarity (M)} = \frac{\text{mole (n)}}{\text{Volume of Solution in Ltr}}$$

$$(\text{F-HO}) \text{ important} \quad n = m \times V$$

2. Mixture of Acid and Base

$$\text{Strength (S)} = \frac{m_1 v_1 - m_2 v_2}{v_1 + v_2}$$

Where, $m_1 v_1$ = mole (n₁) of Acid

$m_2 v_2$ = mole (n₂) of Base

v_1 = Volume of Acid

v_2 = Volume of Base

Now,

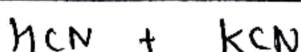
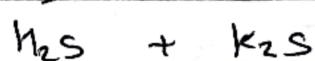
$$\text{pH} = -\log_{10} [S]$$

Buffer Solution

Those Solution which resist the change in pH even if small amount of acid and or base is added on it, is called buffer solution.

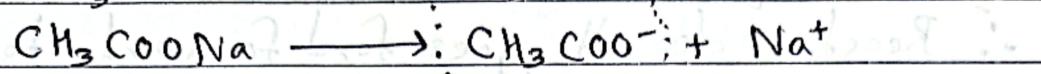
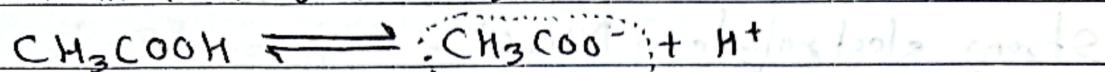
Types (Weak Acid) (WA. + Strong Base)

1. Acidic Buffer : CH₃COOH + CH₃COONa

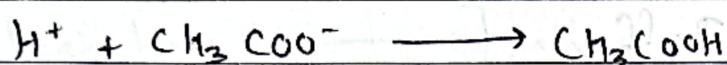


2. Basic Buffer: $\text{NH}_3\text{OH} + \text{NH}_4\text{Cl}$

Mechanism of Acidic Buffer

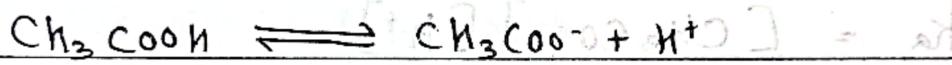


i. If small amount of acid is added i.e. HCl



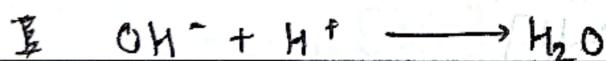
\therefore The loss of CH_3COO^- from CH_3COOH is balanced by CH_3COO^- from CH_3COONa (Strong Electrolyte).

\therefore Effect of common ion $>$ forward reaction of



Thus $[\text{H}^+]$ remain same and hence pH remain same.

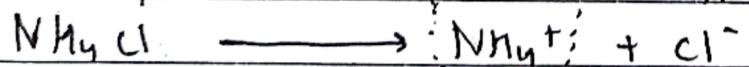
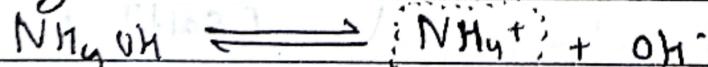
ii. If small amount of base is added



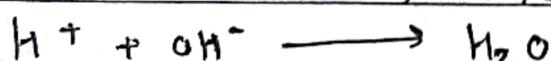
\therefore eff concentration of OH^- decreases thus to fulfill this deficiency reaction goes forward giving more OH^- .

\therefore Effect of forward Rn $>$ Effect of backward Rn.

Working Mechanism of Basic Buffer



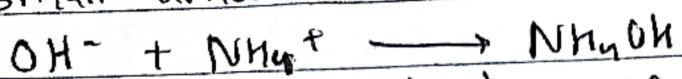
i. If small amount of acid is added.



$\therefore [\text{OH}^-]$ decreases so to balance the $[\text{OH}^-]$ reaction goes forward.

\therefore Effect of forward $>$ Effect of common ion

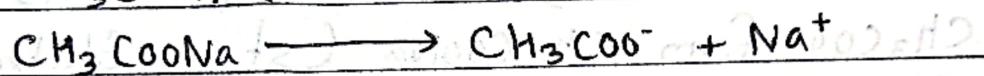
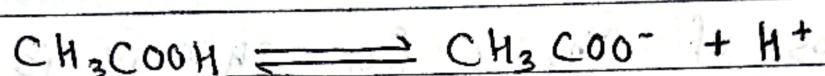
ii. If small amount of base is added



The decrease concentration of NH_4^+ is balanced by strong electrolyte NH_4Cl

∴ Reaction doesn't go forward to give more OH^- .

pH of Acidic Buffer



Now,

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

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

$$\text{or, } [\text{H}^+] = K_a \frac{[\text{Acid}]}{[\text{Salt}]}$$

$$\text{or, } -\log_{10}[\text{H}^+] = -\log_{10} \left(K_a \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]} \right)$$

$$\text{or, } -\log_{10}[\text{H}^+] = -\log_{10} \left(K_a \frac{[\text{Acid}]}{[\text{Salt}]} \right)$$

$$\text{or, } \text{pH} = - \left(\log_{10} K_a + \log_{10} \frac{[\text{Acid}]}{[\text{Salt}]} \right)$$

$$\text{or, } \text{pH} = - \left(\log_{10} K_a + \log_{10} [\text{Acid}] - \log_{10} [\text{Salt}] \right)$$

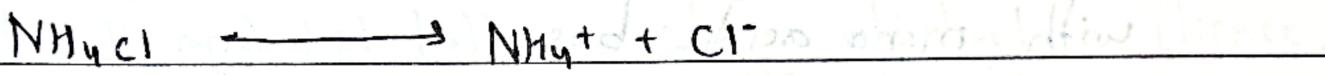
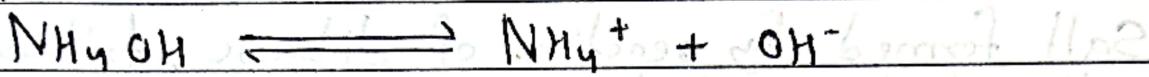
$$\text{or, } \text{pH} = -\log_{10} K_a - \log_{10} [\text{Acid}] + \log_{10} [\text{Salt}]$$

$$\text{or, } \text{pH} = \text{p}K_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]}$$

$$\therefore \text{pH} = \text{p}K_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]}$$

pH of Basic Buffer

Weak base \longrightarrow Salt of weak base and strong acid



Now,

$$K_b = \frac{[\text{NH}_3^+][\text{OH}^-]}{[\text{NH}_3\text{OH}]}$$

$$\text{or, } [\text{OH}^-] = K_b \frac{[\text{NH}_3\text{OH}]}{[\text{NH}_3^+]}$$

$$\text{or, } -\log_{10} [\text{OH}^-] = -\log_{10} \left(\frac{K_b [\text{Base}]}{[\text{Salt}]} \right)$$

$$\text{or, } \text{pOH} = -\log_{10} \left(\frac{K_b [\text{Base}]}{[\text{Salt}]} \right)$$

$$\text{or, } \text{pOH} = -\log_{10} K_b - \log_{10} [\text{Base}] + \log_{10} [\text{Salt}]$$

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

Now,

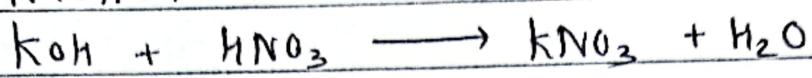
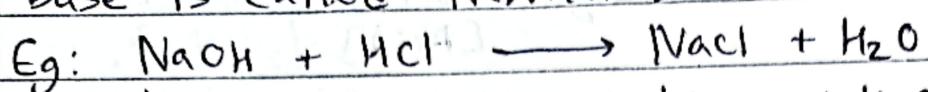
$$\text{pH} + \text{pOH} = 14$$

$$\therefore \text{pH} = 14 - \text{pOH}$$

Types of Salts

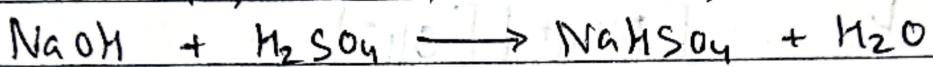
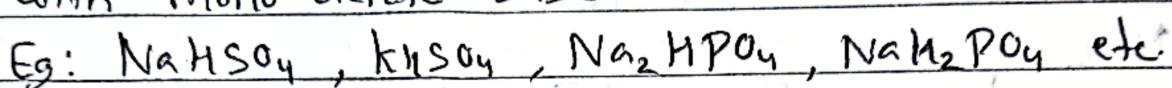
1. Normal Salt

Salts formed by complete neutralization of acid and base is called Normal Salt.



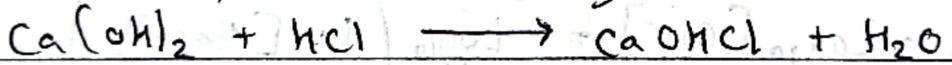
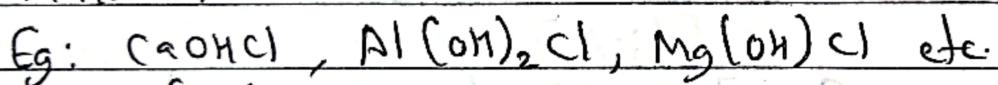
2. Acidic Salt

Salt formed by reaction of dibasic, tribasic acid with mono acidic base.



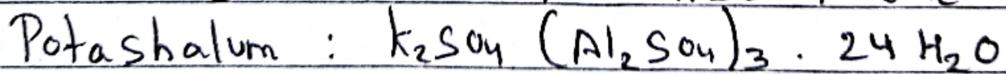
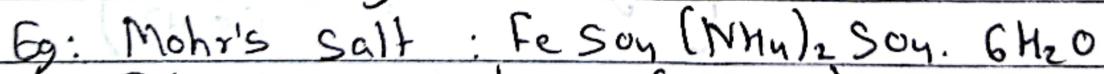
3. Basic Salt

Salt formed by reaction of one diacidic base with monobasic acid.



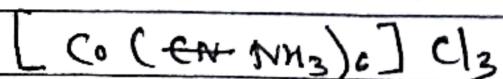
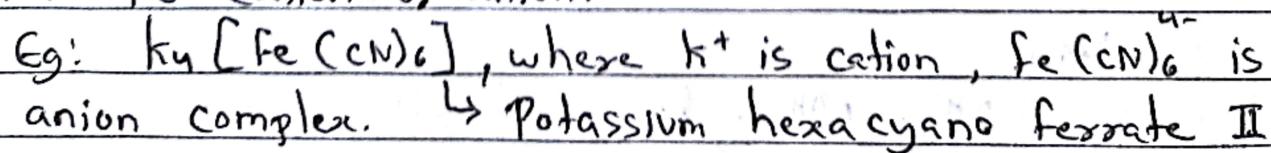
4. Double Salt

Salt formed by combination of two different salt.



5. Complex Salt

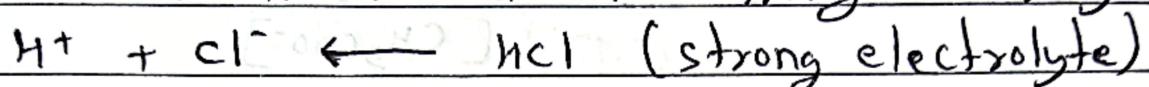
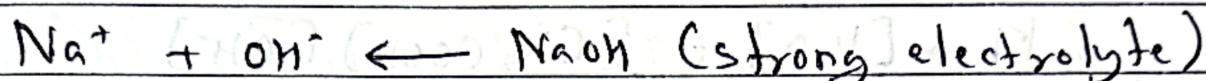
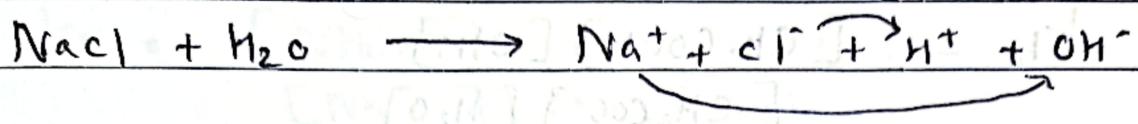
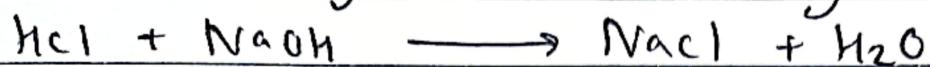
Salt formed by combination of any one of the complex ion. i.e. cation or anion.



Hexaammine cobalt III chloride

Hydrolysis of Salt

1. Salt of Strong Acid and Strong Base



\therefore NaOH and HCl both are strong electrolytes. Hence, dissociate to give $\text{Na}^+ + \text{OH}^-$ respectively.

\therefore No hydrolysis of NaCl occurs.

Hence,

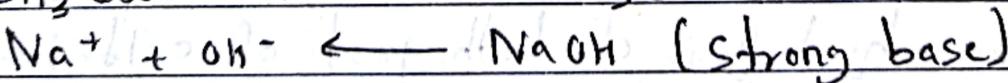
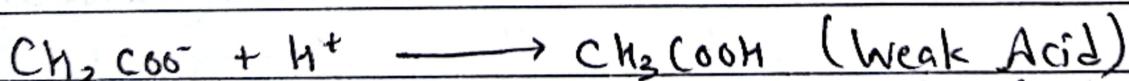
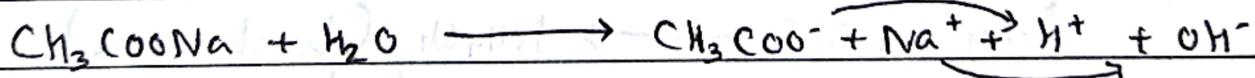
pH of solution of such salt is neutral.

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

$$\text{Or}, \therefore \text{pH} = 7$$



2. Salt of Weak Acid and Strong Base



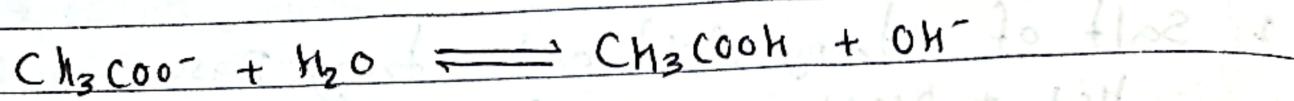
Conclusion:

\rightarrow Hydrolysis of weak part of salt takes place.

\rightarrow Solution of such salt is basic. $[\text{H}^+] < [\text{OH}^-]$

$$\therefore \text{pH} > 7$$

#



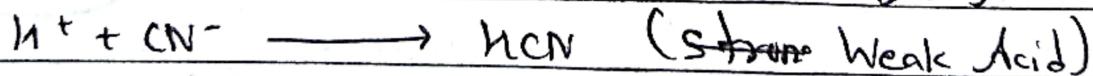
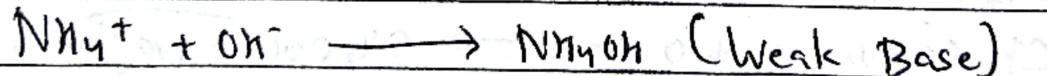
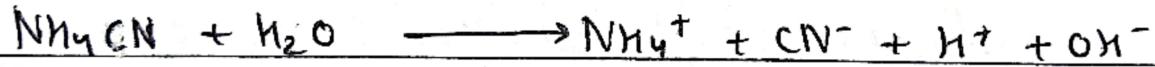
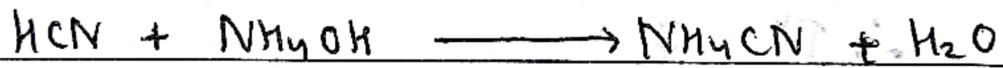
$$K_h = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-][\text{H}_2\text{O}]}$$

$$\text{or, } K_h \times [\text{H}_2\text{O}] = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$

$$\text{or, } K_h = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$

$\therefore K_h$ is the hydrolysis constant of salt obtained from weak acid and strong base.

3. Salt of Weak Acid and Weak Base



→ Here, hydrolysis of both part of salt i.e. NH_4^+ and CN^- are obtained.

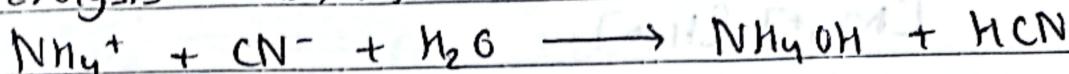
→ Both H^+ and OH^- are consumed.

$$\therefore [\text{H}^+] \neq [\text{OH}^-]$$

∴ Solution is neutral

$$\text{i.e. } \text{pH} = 7$$

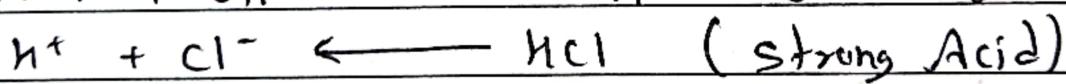
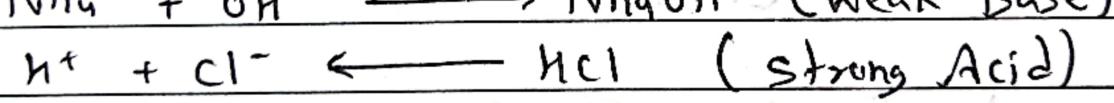
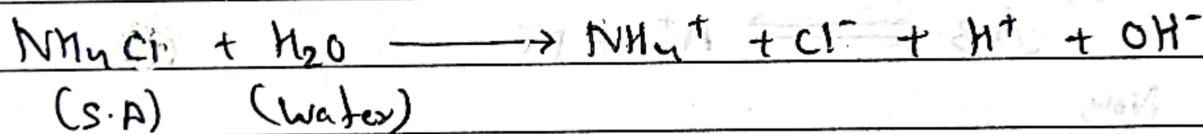
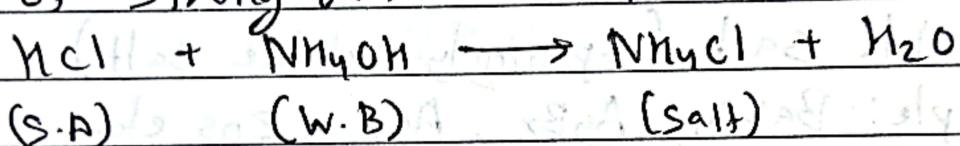
Hydrolysis Constant of Salt of W.A and W.B



$$K_h = \frac{[\text{NH}_4\text{OH}][\text{HCN}]}{[\text{NH}_4^+][\text{CN}^-][\text{H}_2\text{O}]}$$

$$K_h = \frac{[\text{NH}_4\text{OH}][\text{HCN}]}{[\text{NH}_4^+][\text{CN}^-]}$$

4. Salt of Strong Acid and Weak Base



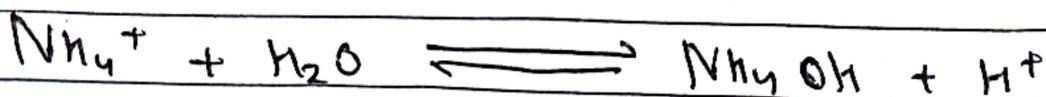
Here hydrolysis of NH_4^+ occurred

$\therefore [\text{H}^+]$ remains in the solution but $[\text{OH}^-]$ is consumed by NH_4^+ .

\therefore Solution is acidic i.e. $[\text{OH}^-] < [\text{H}^+]$

$\therefore \text{pH} < 7$

Hydrolysis constant of salt of strong Acid and weak Base



$$K_h = \frac{[\text{NH}_3\text{OH}][\text{H}^+]}{[\text{NH}_3^+][\text{H}_2\text{O}]}$$

$$\therefore K_h = \frac{[\text{NH}_3\text{OH}][\text{H}^+]}{[\text{NH}_3^+]}$$

Solubility and Solubility product

i. Soluble Salts

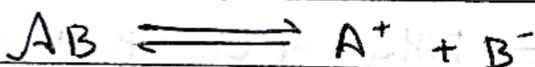
→ All the nitrates are soluble.

→ Salts of Na^+ , K^+ , Cs^+ (Halogen, CO_3 , SO_4 , NO_3)

ii. Insoluble salt (sparingly soluble salt)

Example: BaSO_4 , AgBr , AgI , ZnS etc.

→ AB Type of Salt:



Now,

$$K = \frac{[\text{A}^+][\text{B}^-]}{[\text{AB}]}$$

$$\text{or, } K \times [\text{AB}] = [\text{A}^+][\text{B}^-]$$

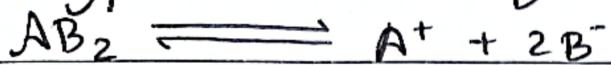
$$\therefore K_{\text{sp}} = [\text{A}^+][\text{B}^-]$$

Solubility

The amount of solute present in 100 gm of solvent in saturated solution of sparingly soluble salts.

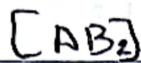
→ It can be written in other ways. gm, mole/litre

AB_2 Type of Salt : Eg - PbCl_2



Now,

$$K = [\text{A}^+] [2\text{B}^-]^2$$



$$\text{or, } K[\text{AB}_2] = [\text{A}^+] [2\text{B}^-]^2$$

When,

Solubility is 's'

$$K_{\text{sp}} = s \times (2s)^2$$

$$K_{\text{sp}} = 4s^3$$

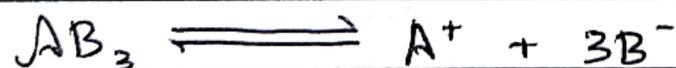
$$\therefore s = \sqrt[3]{K_{\text{sp}} / 4}$$

Hence

Solubility of AB_2 or A_2B type of salt is;

$$s = \sqrt[3]{K_{\text{sp}} / 4}$$

AB_3 type of salt : Eg - AlCl_3



$$K = [\text{A}^+] [3\text{B}^-]^3$$

$$\text{or, } K[\text{AB}_3] = [\text{A}^+] [3\text{B}^-]^3$$

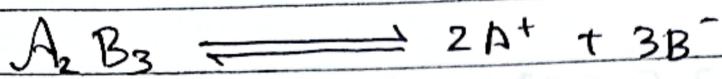
$$\text{or, } K_{\text{sp}} = (s \times (3s)^3)$$

$$\therefore K_{\text{sp}} = 27s^4$$

Now,

$$s = \sqrt[4]{K_{\text{sp}} / 27}$$

A_2B_3 Type of Salt: Eg. AS_2S_3 , $Ca_3(PO_4)_2$



Now,

$$K = \frac{[2A^+]^2 [3B^-]^3}{[A_2B_3]}$$

$$\text{or, } K[A_2B_3] = [2A^+]^2 [3B^-]^3$$

$$\text{or, } K_{sp} = (2s)^2 (3s)^3$$

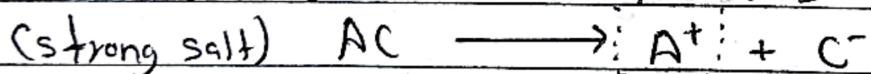
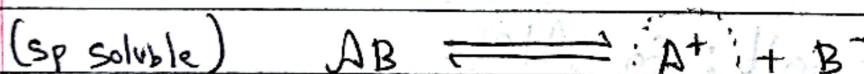
$$\text{or, } K_{sp} = 4s^2 \times 27s^3$$

$$\text{or, } K_{sp} = 108s^5$$

Now,

$$s = \sqrt[5]{\frac{K_{sp}}{108}}$$

Relation between solubility and common ion effect



Common ion effect decreases the solubility of sparingly soluble salt by directing the reaction towards backward direction.

$$\text{Common Ion Effect (CIE)} \propto \frac{1}{\text{Solubility}}$$

Application of Solubility Product in qualitative field

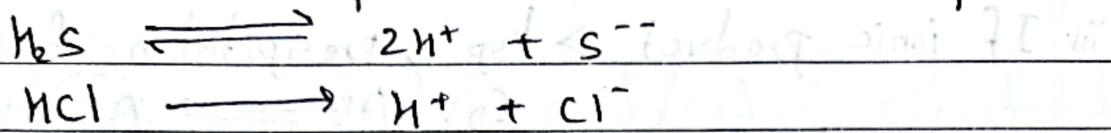
1. Solubility of different types of sparingly soluble salt can be calculated.

$$\text{For } AB \text{ type, } s = \sqrt{K_{sp}}$$

$$\text{For } A_2B \text{ or } AB_2, s = \sqrt[3]{\frac{K_{sp}}{4}}$$

For AB_3 type, $S = \sqrt[4]{\frac{K_{sp}}{27}}$ and so on.

2. Precipitation of Group II metal as their sulphide.



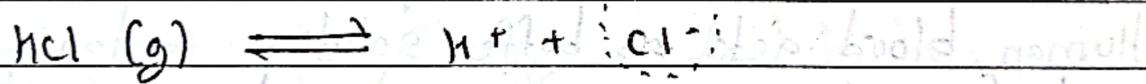
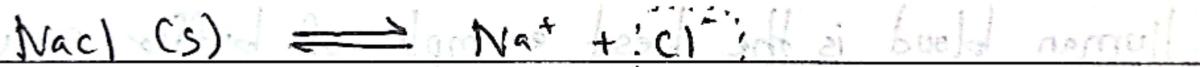
The solubility product of Group III metal is large.

∴ The solubility product of H_2S alone can precipitate both Group II and Group III metal. but when HCl is added, the solubility product of H_2S decreases which is only sufficient to precipitate Group II metal as their Sulphide.

Group II Metals : Mg^{++} , Pb^{++} , Bi^{++} , Cu^{++} , As^{++} , Se^{++} , Sn^{++}

B. In Quantitative field

a. Purification of Common salt

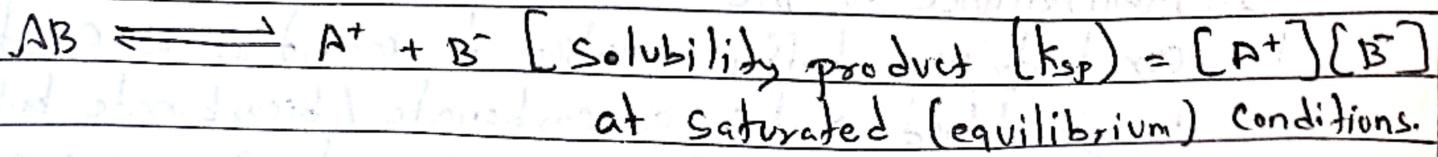
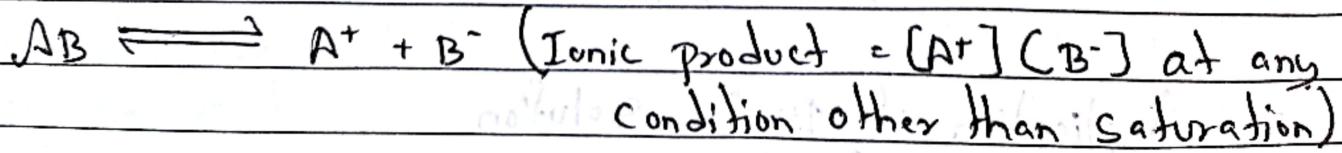


∴ Due to common ion effect the reaction get backward

∴ $NaCl$ get precipitated.

↳ Condition For Precipitation

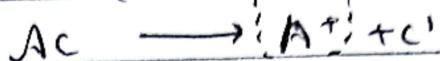
Relation between Solubility and ionic product



Conditions:

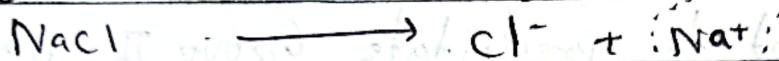
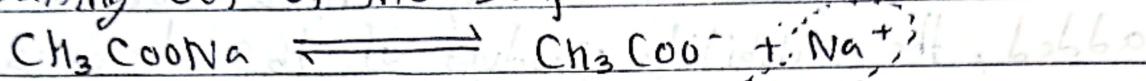
- # Conditions:

 - If ionic product < k_{sp} , more solute can be dissolved.
 - If ionic product = k_{sp} , no more solute can be dissolved
(Solute will be taken as the solution become saturated)
 - If ionic product > k_{sp} , precipitation of solutes occurs.



Application of Common ion effect

Salting out of the Soap

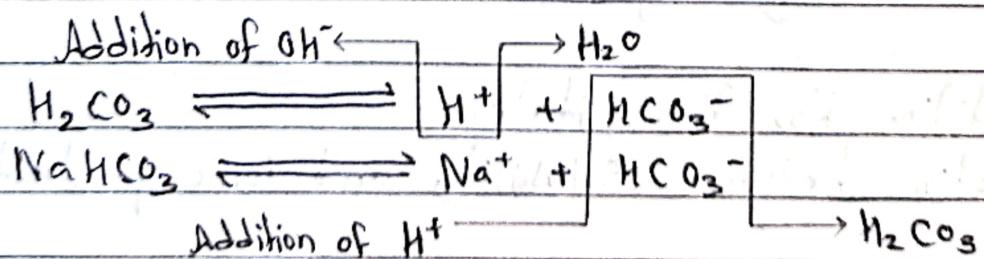


\therefore Due to common Na^+ ion (Conc. of Na^+ increases)
due to which reaction goes backward ~~forward~~ and
form more soap.

Buffer Action of Human Blood

Human blood is the best example of buffer solution.

Human blood acts as buffer solution containing carbonic acid ($H_2O + CO_2 \longrightarrow H_2CO_3$) and sodium bicarbonate ($NaHCO_3$) having $pH = 7.35 \pm 0.5$. The mechanism of buffer action is:



Application of Buffer Solution

1. Maintenance of life

our body uses buffer solution to maintain a constant pH. For eg: blood contain a carbonate / bicarbonate buffer solution that keeps the pH close to 7.35 ± 0.05 .

2. Biochemical Assays

Enzyme activity depends on pH, so that pH during an enzyme assay must stay constant.

3. In Shampoos

Many shampoos use a citric acid/sodium citrate to maintain a slightly acidic "pH balance". This counteracts the basicity of the detergents present in the shampoo.

4. In Laundry Detergents

Many laundry detergent use buffers to prevent their natural ingredients from breaking down.

Titration

Titration is the process of finding the concentration volume of unknown solution by the help of known solution in the presence of or absence of set of indicators.

Indicator	pH range	Acid Medium	Basic Medium
Methyl orange	3.1 - 4.5	Red	Yellow
Methyl Red	4.2 - 6.2	Red	Yellow
Phenol Red	6.4 - 8.2	Yellow	Red
Phenolphthalein	8.2 - 10	Colourless	Pink

Titration of Strong Acid and Strong Base, the pH of the solution lies between (3-11) in such pH value, indicator like methyl orange, methyl red, phenolphthalein etc. can be used.

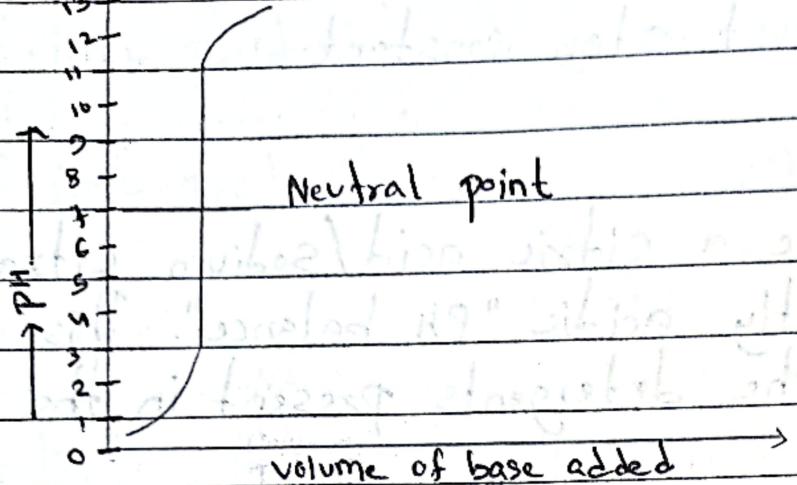


Fig: Titration of S.A and W.B

2. Titration of Strong Acid and Weak Base

When strong acid (HCl , H_2SO_4) is titrated with weak base (NH_4OH) the pH of the solution lies between (3-8) in such pH value, indicator such as methyl orange, methyl red etc. can be used.

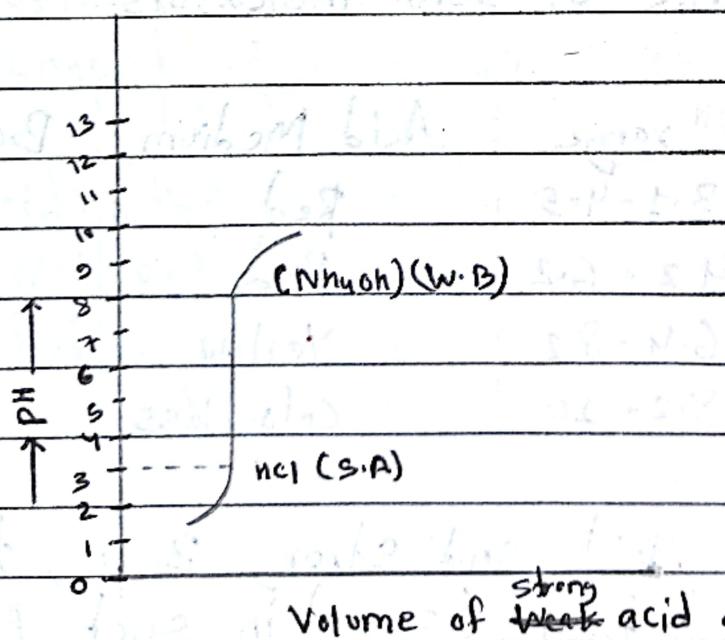


Fig: Titration of Strong Acid Vs St. Weak Base

3. Titration of Weak Acid and Weak Base

When weak acid (HCN , H_2S , HCOOH) is titrated with weak acid base (NH_3OH) the pH of solution lies between (6-8) in such pH. Almost no indicator works between pH (6-8). So end point is obtained by conductometric method.

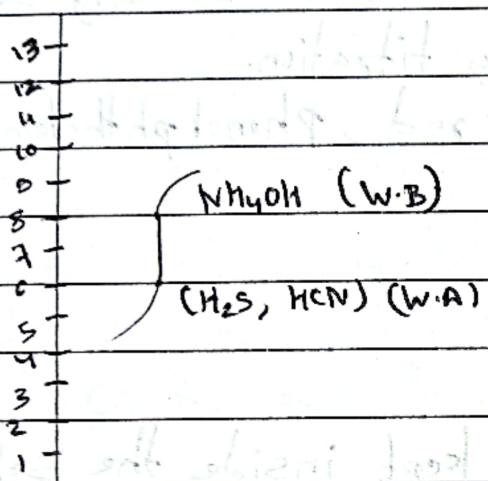
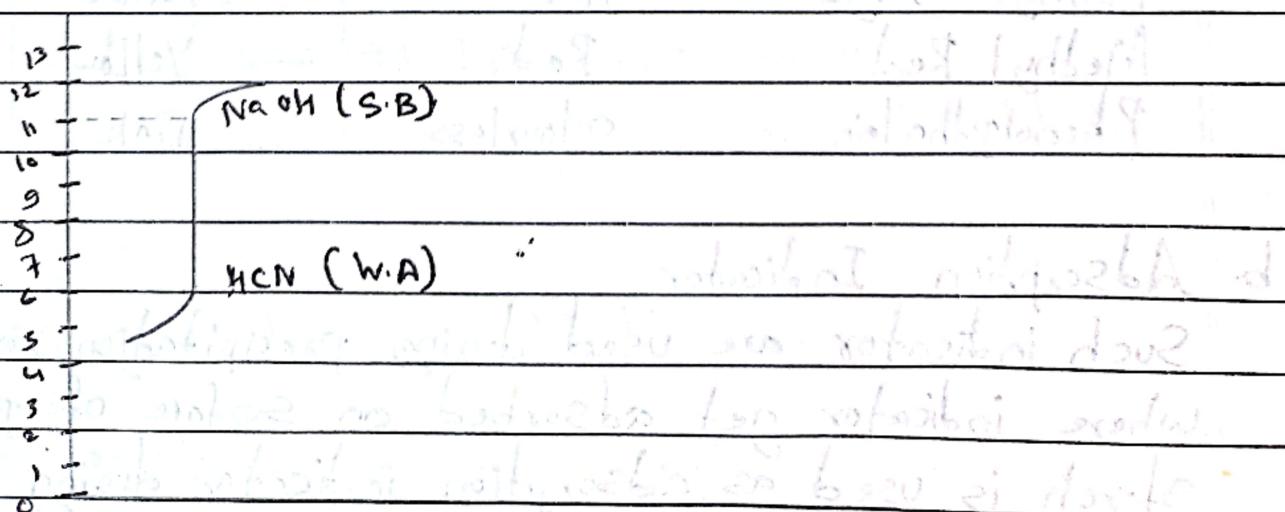


Fig: Titration of WB vs WA

4. Titration of Weak Acid and Strong Base

When weak acid (H_2S , HCOOH , HCN) is titrated with strong base (NaOH), the pH of solution lies between (6-11) in such pH, phenolphthalein is used.



Indicator

Indicator is a chemical substance that indicates the completion of reaction (end point) by changing its colour during titration.

- Indicator may be acidic or basic. So only small amount of is added during titration.
- Eg: Methyl orange, methyl red, phenolphthalein, litmus paper etc.

Types of Indicator

1. Internal Indicator

Those indicator that are kept inside the solution to determine the end point by changing its colour during titration.

a. Acid-Base Indicator

It is used during acid-base titration to detect

Indicator	Acidic medium	Basic medium	pH range
Methyl orange	Red	Yellow	3.1 - 4.4
Methyl Red	Red	Yellow	4.2 - 6.3
Phenolphthalein	colourless	Pink	8.2 - 10

b. Adsorption Indicator

Such indicators are used during precipitation reaction where indicator get adsorbed on surface of ppt.

Starch is used as adsorption indicator during Iodometry or ~~Iodine~~ Iodometry.

c. Self - Indicator

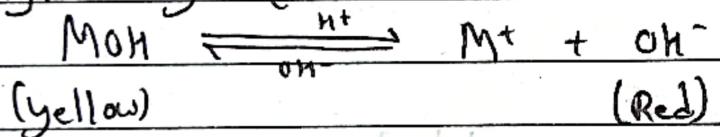
- c. Self - Indicator
 - One of the reaction in titration itself acts as indicator.
 - In titration of oxalic acid with KMnO_4 , KMnO_4 changes its colour to pink at the end of titration. Here KMnO_4 is self-indicator.

2. External Indicators

- Indicator used outside the titrating solution is external indicator.
 - $\text{K}_3[\text{Fe}(\text{CN}_6)]$ is one of the example of external indicator which is kept in reference porcelain plate during titration of $\text{K}_2\text{Cr}_2\text{O}_7$ and ferrous ammonium sulphate.

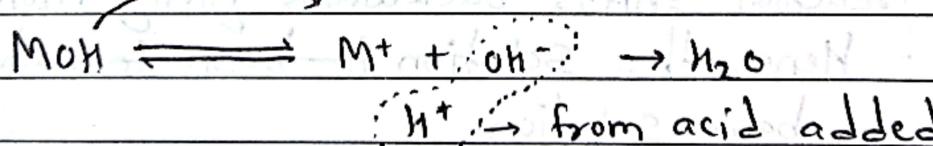
Working Mechanism of Indicator

a. Methyl orange (Mohit)

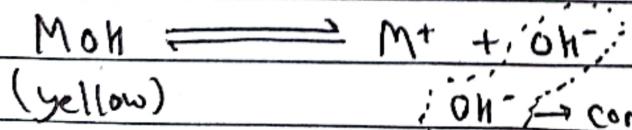


↳ Case I: If acid is added

H^+ ion from added acid consumes OH^- of MnO_4^- . Thus, shifting equilibrium forward giving red colour.



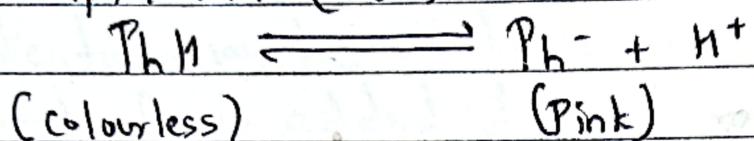
↳ Case II: If base is added



If base is added, the OH^- conc. in solution increases due to common OH^- ion, the reaction shifts backward

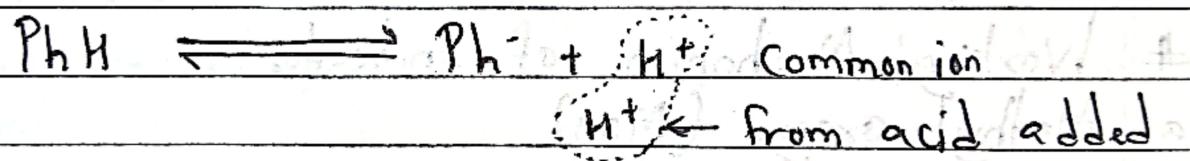
giving more Moh (yellow). In basic solution the methyl orange becomes yellow in colour and the solution becomes yellow.

b. Phenolphthalein (PhH)



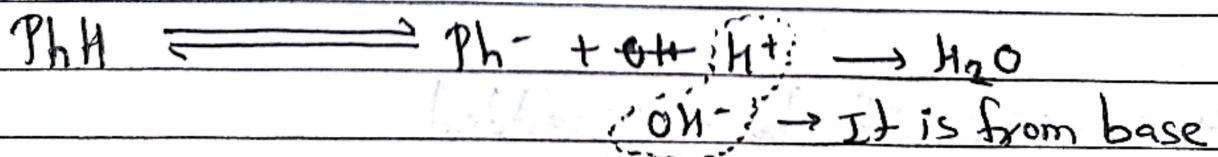
↳ Case I: If acid is added

$[H^+]$ ion in solution increases thus due to common H^+ ion, the reaction shifts backward giving colourless solution.



↳ Case II: If base is added

$[\text{OH}^-]$ from base consumes H^+ from indicator giving H_2O . Therefore, $[\text{H}^+]$ decreases to become equilibrium. The reaction shifts backward giving more Ph^- and H^+ . Hence, the solution is ~~basic in nature~~ becomes pink in basic solution.



Characteristic of Good Indicator

- Indicator can change colour in narrow pH range.
- Indicator can change its colour at the end of the titration.
- The colour of indicator is stable.
- Indicator has different pH range on the basis of acid and base titration. (Strong or weak acid and base).