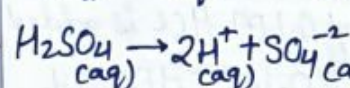
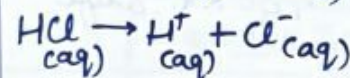
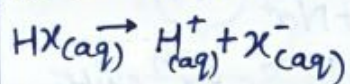


IONIC EQUILIBRIUM

Arrhenius Concept Of Acids and Bases

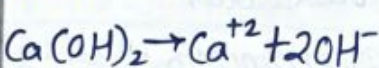
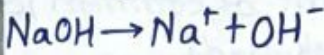
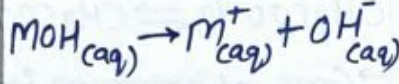
Acids

Substance which give H^+ ion in water.



Bases

Substance which give OH^- ion in water.



Limitation Of Arrhenius Concept

- applicable only to aqueous solution.
- Basicity of NH_3 is not explained (does not possess hydroxyl group).

Bronsted-Lowry Concept / Proton theory

Acids → Which donate H^+ ion / proton donors.

Bases → Which accept H^+ ion / proton acceptors.

Limitations of Bronsted-Lowry

- Doesn't explain about $AlCl_3, BCl_3, \dots$
- Uncertainty in reverse rxn.

Lewis - Acids and Base

Acid → Species which accept electron pair (lone pair)

Base → Species which donate electron pair.

Lewis Acid
e⁻ deficient species
 $AlCl_3, Co^{3+}, Mg^{2+}, H^+, BCl_3, BF_3, SnCl_4, SnCl_2, Ag^+, Cu^{2+}$

Lewis Base
e⁻ rich species.
 $H_2O, NH_3, OH^-, F^-, C_2H_4$

Conjugate - Acid Base Pair

Conjugate Acid → Add H^+

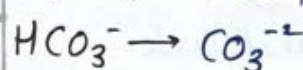
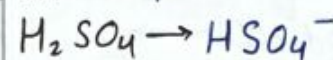
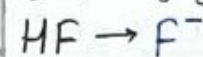
Conjugate Base → Remove H^+

OH^- → Conjugate acid → H_2O

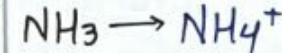
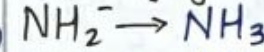
NH_4^+ → Conjugate base → NH_3

If Bronsted acid → strong, then its conjugate base is weak.

Que) Conjugate Base for Bronsted Acids →

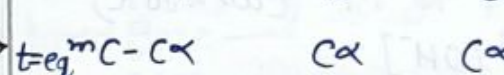
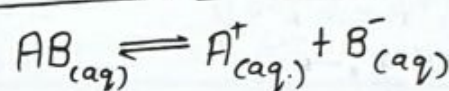


Que) Conjugate Acid for Bronsted Base →



Amphiprotic → H^+ le bhi sakta hai, de bhi sakta hai..

Ionisation Constant



$$K_{eq} = K_i = \frac{[A^+][B^-]}{[AB]} = \frac{C\alpha^2}{C(1-\alpha)} = \boxed{\frac{C\alpha^2}{1-\alpha}}$$

For Weak acid = $K_i = K_a$

For Weak base = $K_i = K_b$

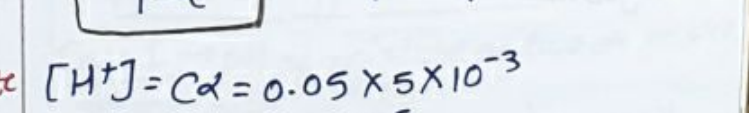
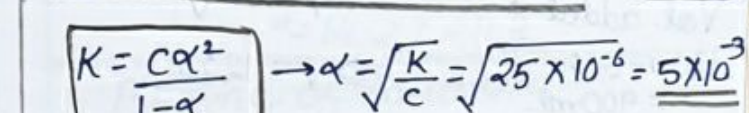
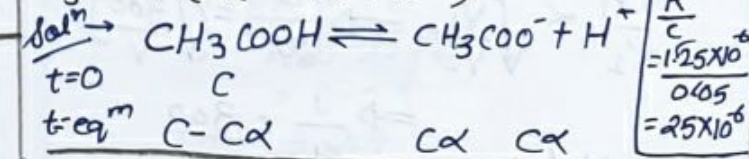
Neglection Theory

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

$$\text{If } \frac{K_i}{C} \leq 6.4 \times 10^{-3} \rightarrow 1-\alpha=1 \Rightarrow K_i = C\alpha^2$$

$$\text{If } \frac{K_i}{C} > 6.4 \times 10^{-3} \rightarrow 1-\alpha \neq 1 \rightarrow \text{Solve quadratic.}$$

Que) Calculate conc. of H^+ ion, in 0.05 M CH_3COOH ($K_a = 1.25 \times 10^{-6}$)



$$\boxed{K = \frac{C\alpha^2}{1-\alpha}} \rightarrow \alpha = \sqrt{\frac{K}{C}} = \sqrt{\frac{1.25 \times 10^{-6}}{0.05}} = \sqrt{25 \times 10^{-6}} = 5 \times 10^{-3}$$

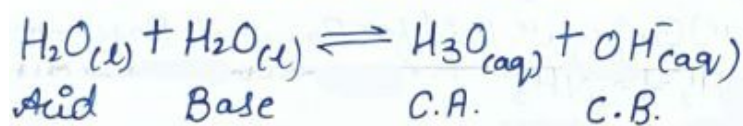
$$[H^+] = C\alpha = 0.05 \times 5 \times 10^{-3} = 25 \times 10^{-5}$$

NEET SLAYER

Quadratic eq.ⁿ = $\frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

Ionisation Constant of Water and its Ionic Product →

→ Water acts as acid and base



$$K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} \quad (K = \text{dissociation constant})$$

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

(K_w = Ionic product of water).

Experimentally: →

$$[\text{H}^+] = [\text{OH}^-] = 10^{-7} \text{ M} \quad (\text{at } 298 \text{ K})$$

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

$$\therefore K_w = 10^{-7} \times 10^{-7} = 10^{-14}$$

↓
Ionisation Constant of H_2O (Temp. dependent)

Ostwald Dilution Law

$$\alpha \propto \sqrt{\text{dilution (Vol)}} \quad \left| \begin{array}{l} \alpha = \text{degree of ionisation} \\ C = \text{concentration} \end{array} \right.$$

$$\alpha \propto \frac{1}{\sqrt{C}}$$

Que) 0.2 M CH_3COOH Solution →

$$V_1 = 300 \text{ ml}, V_2 = ?$$

$$\alpha_1 = \alpha, \alpha_2 = 2\alpha$$

Sol → $\alpha \propto \sqrt{V}$

$$\frac{\alpha_1}{\alpha_2} \propto \sqrt{\frac{V_1}{V_2}} \Rightarrow \left(\frac{\alpha}{2\alpha}\right)^2 = \left(\sqrt{\frac{300}{V}}\right)^2$$

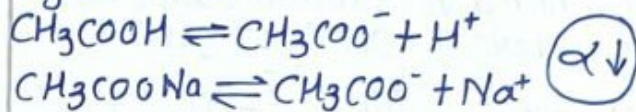
$$\Rightarrow \frac{1}{4} = \frac{300}{V}$$

Vol. added →
1200 - 300
= 900 ml.

$$\Rightarrow V = 1200$$

Common Ion Effect

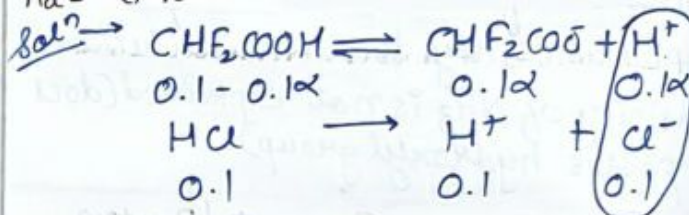
^{is Ka add Ka} ^{added to}
(Strong electrolyte) + Weak electrolyte, having common ion then, degree of ionisation (α) of weak electrolyte ↓
e.g. →



→ Conc. of common ion ↑, So Rx.ⁿ shift backwards.

Que) Conc. of H^+ at eq.^m of 0.1 M HCl is added in an aq. sol.ⁿ containing 0.1 M CHF_2COOH .

$$K_a = 2 \times 10^{-2}$$



$$K_a = \frac{(0.1 + 0.1\alpha)0.1\alpha}{0.1 - 0.1\alpha}$$

$$2 \times 10^{-2} = \frac{0.1\alpha(1 + \alpha)}{(1 - \alpha)}$$

$$0.2 = \frac{\alpha^2 + \alpha}{1 - \alpha}$$

$$\alpha^2 + \alpha = 0.2 - 0.2\alpha$$

$$\alpha^2 + 1.2\alpha - 0.2 = 0$$

$$\alpha = 0.148$$

$$\text{H}^+ = C\alpha = (0.1) \times 0.148$$

pH-Scale

$$P = -\log$$

$$pK_a = -\log K_a$$

$$pK_b = -\log K_b$$

$$pK_w = -\log K_w$$

$$pK_w = -\log 10^{-14}$$

$$pK_w = 14 \rightarrow \text{at } 25^\circ\text{C}$$

Relation b/w p^H and p^{OH} .

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

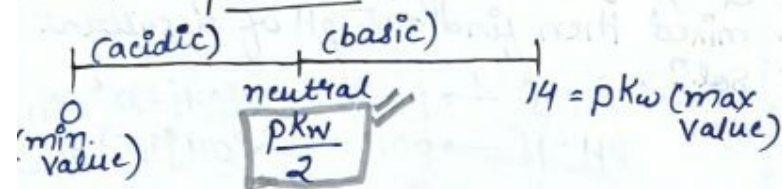
$$\log K_w = -(\log \text{H}^+) - (\log \text{OH}^-)$$

$$pK_w = p^H + p^{OH}$$

$$14 = p^H + p^{OH}$$



pH Scale



Que) $[H^+] = 3.8 \times 10^{-3}$, $pH = ?$

Solⁿ → $[H^+] = 3.8 \times 10^{-3}$ ($\log 3.8 = 0.58$)

$$pH = -\log H^+$$

$$\log = (x \times y^b)$$

$$= b - \log x$$

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

$$= -\log (3.8 \times 10^{-3})$$

$$= 3 - \log 3.8$$

$$= 3 - 0.58$$

$$= 2.42$$

Que) $[HCl] = 1 \times 10^{-8} M$, $pH = ?$

$$-\log(1 \times 10^{-8})$$

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

major mistake

$$\left[\begin{matrix} 10^{-6} \\ 10^{-7} \\ 10^{-8} \end{matrix} \right] [H^+] \text{ consider } H_2O$$

$$[H^+] = 10^{-8} + 10^{-7} = 10^{-7}(10^{-1} + 1)$$

$$= 10^{-7} \left(\frac{11}{10} \right)$$

$$pH = -\log \left(10^{-7} \times \frac{11}{10} \right)$$

$$= 7 - \log \left(\frac{11}{10} \right) \Rightarrow 7 - (\log 11 - \log 10)$$

$$\Rightarrow 7 - (1.04 - 1)$$

$$\Rightarrow 6.96$$

Que) Determine pH at $[H^+] = \frac{a}{b} \times 10^x M$

Solⁿ → $-\log(H^+)$

$$\Rightarrow -\log \left(\frac{a}{b} \times 10^x \right)$$

$$\Rightarrow x - \log \left(\frac{a}{b} \right)$$

$$\Rightarrow x - \log a + \log b$$

Que) Find $[H^+]$ if $pH = 5.7$.

$$\rightarrow 10^{0.3} = 2 / \text{anti log}(0.3) = 2$$

Solⁿ → $[H^+] = 10^{-pH}$

$$= 10^{-5.7} = 10^{-(6+0.3)}$$

$$= 10^{-6} \cdot 10^{0.3}$$

$$= 2 \times 10^{-6}$$

pH Calculation Cases →

- ① Strong Acid / Strong Base.
- ② Mixture of strong acid and Base
- ③ Salt Hydrolysis.
- ④ Buffer Solution.

→ Strong monoprotic acid and base
eg → $HCl, HNO_3, NaOH, KOH$

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

$$pOH = -\log [OH^-]$$

$$\left[\begin{matrix} [H^+] / [OH^-] < 10^{-6} M \\ \text{then, we can't calculate } pH \text{ directly by this formula, we have to consider } H^+ \text{ ion conc. from water also.} \end{matrix} \right]$$

$$\begin{matrix} 10^{-6} + 10^{-7} \\ 10^{-7} + 10^{-7} \\ 10^{-8} + 10^{-7} \end{matrix}$$

Calculate pH of strong diprotic acid, and Base →

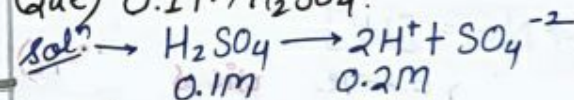
Que) $[H^+]$ concn. of $10^{-8} M HCl$ aq. solⁿ at 298 K ($K_w = 10^{-14}$) is: →

Solⁿ → $10^{-8} + 10^{-7}$

$$\Rightarrow 10^{-7}(10^{-1} + 1) \Rightarrow 10^{-7}(1.1)$$

$$\Rightarrow 1.0525 \times 10^{-7} M$$

Que) $0.1 M H_2SO_4$.



$$pH = -\log(2 \times 10^{-1})$$

$$= 1 - \log 2$$

$$= 1 - 0.3 \Rightarrow 0.7$$

Que) $10^{-3} M Ba(OH)_2$.

Solⁿ → $pOH = 2 \times 10^{-3}$

$$= 3 - \log 2$$

$$= 3 - 0.3$$

$$pOH = 2.7$$

$$pH = 14 - 2.7 = 11.3$$

Mixing of Solution

Case 1 → pH of mixture of two or more strong acid solution.

$$\left[\begin{matrix} \text{Acid 1} \\ N_1 V_1 \end{matrix} \right] + \left[\begin{matrix} \text{Acid 2} \\ N_2 V_2 \end{matrix} \right] = \left[\begin{matrix} \text{Acidic final} \\ N_m V_m \end{matrix} \right] \quad V_m = V_1 + V_2 \dots$$

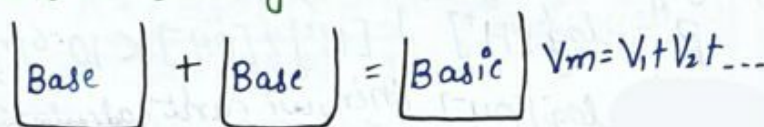
$$N_1 V_1 + N_2 V_2 + \dots = N_m V_m$$

$$N_m = \frac{V_1 N_1 + V_2 N_2 + \dots}{V_m}$$

$$N_m = [H^+]$$

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

Case-2 pH of mixture of two or more strong base solution.



$$N_1 V_1 + N_2 V_2 + \dots = N_m V_m$$

$$N_m = \frac{V_1 N_1 + V_2 N_2 + \dots}{V_m}$$

$$N_m = [OH^-]$$

$$pOH = (14 - pOH = pH)$$

Case-3 → pH of mixture of strong acid and strong base Sol.ⁿ

$$(N_1 V_1 + N_2 V_2 + \dots)_{\text{acid}} - (N_1 V_1 + N_2 V_2 + \dots)_{\text{Base}} = V_m N_m$$

Case 1 → $N_m = +ve$ (acidic solution)
 $\Rightarrow N_m = [H^+]$

Case 2 → $N_m = -ve$ (basic solution)
 $N_m = [OH^-]$

Case 3 → $N_m = 0$ (pH=7)

Que) 100 ml each of 2 solutions having pH=2 and pH=4 are mixed. Find pH of resultant solution.

$$\begin{aligned} \rightarrow H^+ = 10^{-2} &= N_1 \\ \rightarrow H^+ = 10^{-4} &= N_2 \end{aligned} \Rightarrow \frac{10^{-2} \times 100 + 10^{-4} \times 100}{200}$$

$$\Rightarrow \frac{1 + 0.01}{200}$$

$$H^+ \Rightarrow \frac{1.01}{200}$$

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

Que) equal vol. of pH=8 and pH=11 are mixed then find out pH of resultant Sol.ⁿ

$$\begin{aligned} \text{Sol.}^n \rightarrow pH=8 &\rightarrow pOH=6 \rightarrow [OH^-] = 10^{-6} \\ \text{Sol.}^n \rightarrow pH=11 &\rightarrow pOH=3 \rightarrow [OH^-] = 10^{-3} \end{aligned}$$

$$\frac{10^{-6} \times 1 + 10^{-3} \times 1}{2} = \frac{10^{-3}(10^{-3} + 1)}{2}$$

$$= 10^{-3} \times \frac{1}{2} = 0.5 \times 10^{-3}$$

$$\rightarrow 3(-\log 0.5) \rightarrow pOH \rightarrow pH = 14 - pOH$$

Hydrolysis of Salts and pH of their Solution

Anion → from acid
 $ClO_4^-, SO_4^{2-}, NO_3^-$
 Br^-, Cl^-

Cation → from base
 $Na^+, K^+, Ca^{2+}, Ba^{2+}$
 Basicity increase

Salt →

- Salt of weak acid and strong base
- Salt of strong acid and weak base
- Salt of weak acid and weak base
- Salt of strong acid and strong base

	SA and SB Salt	SA and WB Salt	WA and SB Salt	WA and WB Salt
Hydrolysis	—	Cationic	Anionic	Cationic and anionic
Nature of Solution	Neutral solution	Acidic solution	Basic solution	Almost neutral sol
$K_h = \frac{K_w}{K_{\text{Weak}}}$ ↓ Hydrolysis constant	7	$K_h = \frac{K_w}{K_b}$	$K_h = \frac{K_w}{K_a}$	$K_h = \frac{K_w}{K_a K_b}$
h, degree of hydrolysis	—	$h = \sqrt{\frac{K_h}{C}}$ $h = \sqrt{\frac{K_w}{K_b \cdot C}}$	$h = \sqrt{\frac{K_h}{C}}$ $h = \sqrt{\frac{K_w}{K_a \cdot C}}$	$h = \sqrt{K_h}$ $h = \sqrt{\frac{K_w}{K_a K_b}}$
pH	7	$pH = 7 - \frac{1}{2} pK_b - \frac{1}{2} \log C$	$pH = 7 + \frac{1}{2} pK_a + \frac{1}{2} \log C$	$pH = 7 + \frac{1}{2} pK_a - \frac{1}{2} pK_b$

Buffer Solution

Solution which resist change in pH on dilution / with addition of small amount of acids or alkali.

Buffer Solution

Simple Buffer Sol.ⁿ
Salt of W.A. and W.B.

Mixed Buffer Sol.ⁿ

Acidic Buffer Sol.ⁿ / Basic Buffer Sol.ⁿ

Mixed Buffer Solution

Acidic Buffer Sol.ⁿ

Sol.ⁿ of weak acid and salt of its conjugate ion.

e.g. $(\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa})$
 $\text{HCN} + \text{KCN}$

$\text{H}_2\text{CO}_3 + \text{NaHCO}_3$: Blood

$$\text{pH} = \text{pK}_a + \log \left(\frac{[\text{conjugate Base}]}{[\text{Acid}]}\right)$$

$$\text{pH} = \text{pK}_a + \log \left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$$

Basic Buffer Sol.ⁿ

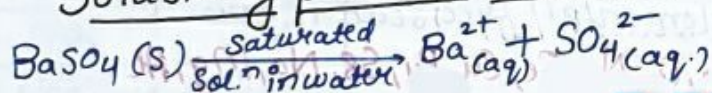
Sol.ⁿ of weak base and salt of its conjugate ion.

e.g. $(\text{NH}_4\text{OH} + \text{NH}_4\text{Cl})$

$$\text{pOH} = \text{pK}_b + \log \left(\frac{[\text{conjugate acid}]}{[\text{Base}]}\right)$$

$$\text{pOH} = \text{pK}_b + \log \left(\frac{[\text{Salt}]}{[\text{Base}]}\right)$$

Solubility product Constant

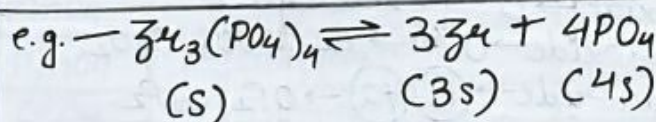


$$K = \frac{[\text{Ba}^{2+}][\text{SO}_4^{2-}]}{[\text{BaSO}_4]}$$

$$K[\text{BaSO}_4] = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$

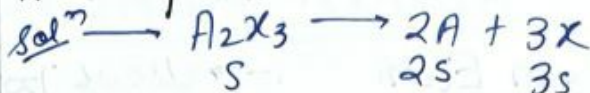
Solubility product constant,
Simply Solubility product



$$\begin{aligned} K_{sp} &= [\text{Zr}]^3 [\text{PO}_4]^4 \\ &= (3s)^3 (4s)^4 \\ &= 6912(s^7) \end{aligned}$$

Que) Calculate the solubility of Al_2X_3 in pure water assuming that neither kind of ion reacts with water the solubility product of

$$\text{Al}_2\text{X}_3, K_{sp} = 1.1 \times 10^{-23}$$

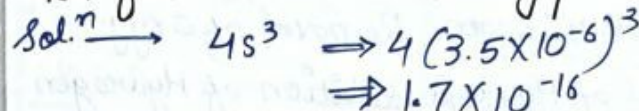


$$K_{sp} = (2s)^2 (3s)^3$$

$$1.1 \times 10^{-23} = 108 s^5$$

$$s = \frac{(1.1 \times 10^{-23})^{1/5}}{108}$$

Que) Solubility of M_2S Salt is 3.5×10^{-6} then find out its solubility product?



Condition for precipitation Ionic product (Q_{sp}) \rightarrow

Q_{sp} : Contains conc of ions at any time.

K_{sp} : Contains only eq.^m concentration.

$Q_{sp} < K_{sp}$: Unsaturated sol.ⁿ and precipitation will not occur.

$Q_{sp} = K_{sp}$: Saturated sol.ⁿ and solubility eq.^m exist.

$Q_{sp} > K_{sp}$: Super saturated and hence, precipitation will occur.

NEET SLAYER

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