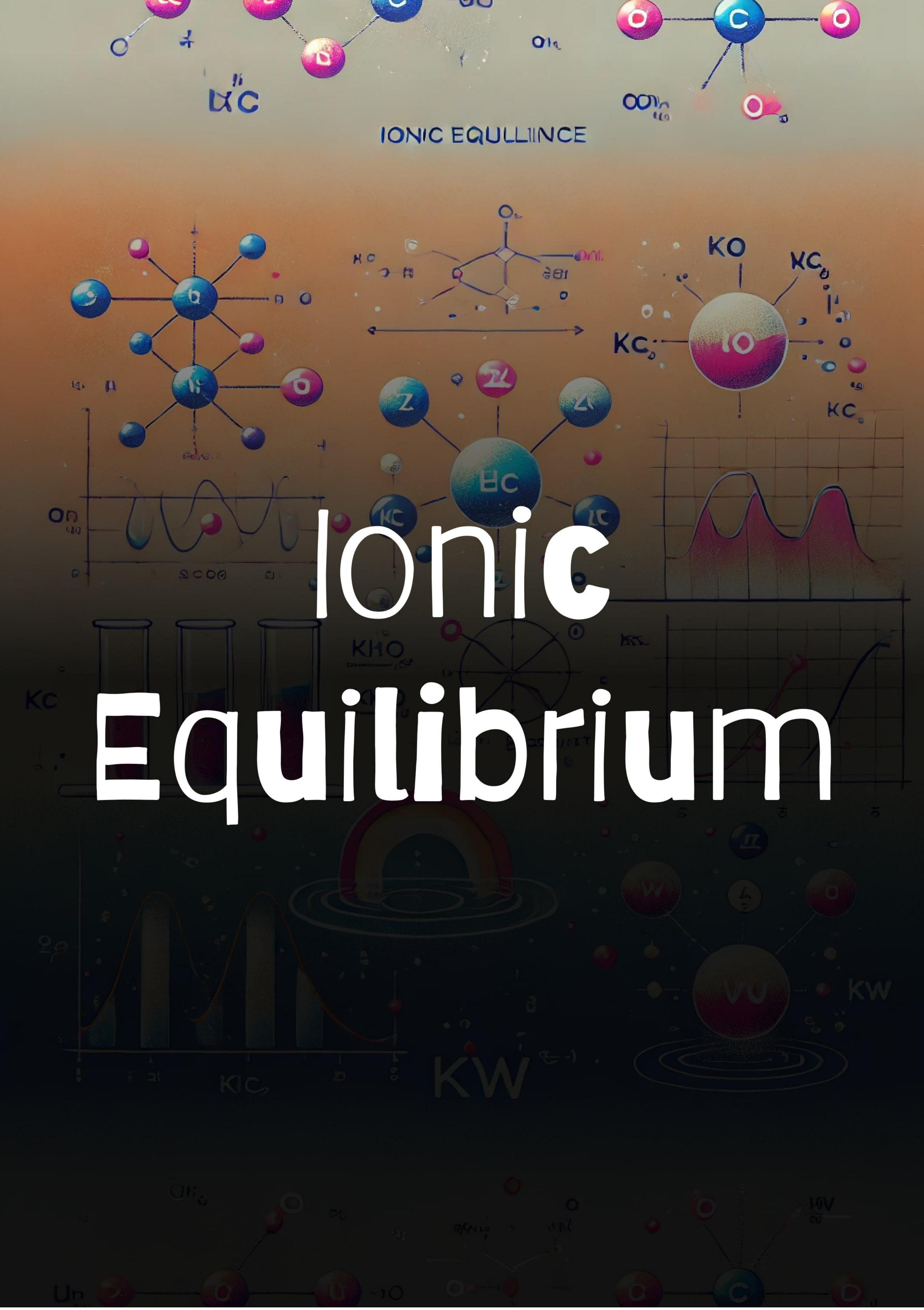

Ionic Equilibrium

Get More Learning Materials Here : CLICK HERE www.studentbro.in

Ionic Equilibrium

① A/B Theories
② pH scale
③ Buffer solution
④ salt hydrolysis
⑤ solubility Product

- A/B Theories :

① Arrhenius A/B Theory :

Acid – It produces H^+ when dissolved in H_2O

Base – It produces OH^- when dissolved in H_2O

② B/L – A/B Theory :

Acid – H^+ donor

Base – H^+ acceptor

conjugate acid – add of H^+

conjugate base – Removal of H^+

- conjugated A/B pair:

A/B pair which differ by only one H^+

③ Lewis A/B Theory :

Acid : – $e^\ominus$ pair acceptor

To accept $e^\ominus$ pair [L·P] = Vacant orbital required.

Base :– $e^\ominus$ pair donor

To donate $e^\ominus$ pair [L·P] = L·P on central atom of molecule/ion required

- Lewis Acid :

$$\left.\begin{matrix} ①\ BF_3 \\ ②\ AlCl_3 \end{matrix}\right\}$$

All Group–13 elements.

B, Al, Ga, In, Tl

Get More Learning Materials Here : CLICK HERE www.studentbro.in

③ Al^{+3} } All most all cations.

④ Fe^{+2}

⑤ SF_6 } sp^3d, sp^3d^2, sp^3d^3

⑥ PCl_5 } 4V.O 3V.O 2V.O

⑦ CO_2 — multiple bonds within dissimilar atoms.

⑧ O — Group-16 elements.

α ⑨ NH_4^+ / H_3O^+ / PH_4^+

⑩ Cl_2O_7 — multiple bonds within dissimilar atoms.

- Lewis Base : G-15 / G-16 / G-17

1LP 2LP 3LP

① NH_3

② $C_6H_5NH_2$

③ H—O—H

④ R—O—H

⑤ R—SH

⑥ SO_4^{-2} } All most all anions.

⑦ N^{-3}

⑧ CN^-

⑨ $CH_2 = CH_2$

⑩ $CH \equiv CH$

- Ionic product of water:

① $Kw = [H^+][OH^-]$

② At room temperature / 25°C → $[H^+] = [OH^-] = 10^{-7}\ M = 10^{-7}\ mol/L$

③ $Kw = 10^{-14}\ M^2$ at 25°C

Get More Learning Materials Here : CLICK HERE www.studentbro.in

④ $[H^+] = \frac{K_w}{[OH^-]} = \frac{1\times10^{-14}}{[OH^-]}$ $[OH^-] = \frac{K_w}{[H^+]} = \frac{1\times10^{-14}}{[H^+]}$

- Ostwald dilution law for weak electrolyte [WA/WB]:

① strong electrolytes [S·A / S·B / salt] undergo 100% dissociation [$\alpha = 100\%$]

② weak electrolytes [WA/WB]
Degree of dissociation = very very less

α = degree of dissociation or ionisation
Ka = dissociation constant of WA.
Kb = dissociation constant of WB.

W·A	W·B
① $[H^+] = C\cdot\alpha$	① $[OH^-] = C\cdot\alpha$
② $K_a = C\alpha^2$ $\alpha = \sqrt{\frac{K_a}{C}}$	② $K_b = C\alpha^2$ $\alpha = \sqrt{\frac{K_b}{C}}$
③ $[H^+] = \sqrt{K_a\cdot C}$	③ $[OH^-] = \sqrt{K_b\cdot C}$

- p^H scale:

① $pH = -\log[H^+]$
② $pOH = -\log[OH^-]$
③ $pK_a = -\log[K_a]$
④ $pK_b = -\log[K_b]$

a) $pH + pOH = 14$
$[H^+]\cdot[OH^-] = 10^{-14}$

b) $pK_a + pK_b = 14$
$K_a\cdot K_b = 10^{-14}$

⑤ $[H^+] = \text{Antilog}[-pH]$

Get More Learning Materials Here : CLICK HERE www.studentbro.in

c) For Acids $pH < 7$
For Base $pH > 7$
For neutral $pH = 7$

$\log 1 = 0$
$2 = 0.30$
$3 = 0.47$
$4 = 0.60$
$5 = 0.69$
$6 = 0.77$
$7 = 0.84$
$8 = 0.90$
$9 = 0.95$
$10 = 1$

Q. 2×10^{-4} M H_2SO_4 soln pH value is

→ $H_2SO_4 \longrightarrow 2H^+ + SO_4^{-2}$
2×10^{-4} $\quad 2\times2\times10^{-4}$ $\quad 2\times10^{-4}$ M
$\quad 4\times10^{-4}$

$pH = 4 - \log 4$
$= 4 - 0.60$
$= 3.40.$

Q. 0.005 M NaOH solⁿ pH value is

→ $NaOH \longrightarrow Na^+ + OH^-$
5×10^{-3} M $\quad 5\times10^{-3}$ M

$pOH = 3 - \log 5$
$= 3 - 0.69$
$= 2.31$

$pH = 14 - pOH$
$pH = 14 - 2.31$
$pH = 11.69.$

Q. 0.01 M CH_3COOH undergo 0.01 % dissociation pH = ?

→ $C = 0.01$ M $\quad \alpha = 0.01\%$
$= 10^{-2}$ $\quad = 10^{-4}$

$[H^+] = C\alpha$
$= 10^{-2} \times 10^{-4}$
$= 10^{-6}$

$pH = \underline{\underline{6}}$

Get More Learning Materials Here : CLICK HERE www.studentbro.in

Q. 0.56 g of KOH dissolved in 200 ml of H_2O pH = ?

→ 0.56 g KOH ——— $[OH^-] = N_{Base}$.

$$E.wt = \frac{M.wt}{x\text{-factor}} = \frac{56}{1}$$

$$N = \frac{wt}{E.wt} \times \frac{1000}{Vml}$$

$$= \frac{0.56}{56} \times \frac{1000}{200}$$

$$N_{Base} = [OH^-] = 0.05$$
$$= 5 \times 10^{-2}$$

$$pOH = 2 - \log 5$$
$$= 2 - 0.69$$
$$= 1.31.$$

$$pH = 14 - 1.31$$
$$= 12.69.$$

Q. pH of 0.1 M HCNO is 2.34 Then calculate Ka and α values ?

→ pH = 2.34

$[H^+] = \text{Antilog}\ [-2.34]$

$-3 + 0.64 \rightarrow 6.4$

6.4 → ≤ 3 (−1)
6.4 → 4,5·9 (−1·5)
6.4 → ≥ 6 (−2)

$[H^+] = 4.4 \times 10^{-3}$

① $[H^+] = C\alpha$

$4.4 \times 10^{-3} = 10^{-1} \times \alpha$

$\alpha = 4.4 \times 10^{-2}$

$\alpha = 4.4\%$

② $[H^+] = \sqrt{Ka \cdot C}$

$4.4 \times 10^{-3} = \sqrt{Ka \cdot 10^{-1}}$

$4.4 \times 4.4 \times 10^{-6} = Ka \cdot 10^{-1}$

$Ka = 4.4 \times 4.4 \times 10^{-5}$ //

Q. $Ka\ (HCN) = 1 \times 10^{-5}$ pH = ?

conc. of HCN = 0.05 M

Get More Learning Materials Here : CLICK HERE www.studentbro.in

→ $Ka = 1 \times 10^{-5}$

$c = 0.05$

$[H^+] = \sqrt{Ka \cdot c}$

$= \sqrt{10^{-5} \times 5 \times 10^{-2}}$

$= \sqrt{5 \times 10^{-7}}$

$= \sqrt{50 \times 10^{-8}}$

$= 7 \times 10^{-4}$ (app)

$pH = 4 - \log 7$

$= 4 - 0.84 = 3.16$

Q: At room temperature conc. of Hydronium ion $= 1 \times 10^{-10}$ Then pOH of solution.

→ $[H_3O]^+ = [OH]^- = ?$

$[H^+] = 10^{-10}$

$Kw = [H^+][OH^-]$

$1 \times 10^{-14} = 10^{-10}[OH^-]$

$[OH^-] = 10^{-4}$

$pOH = 4$ //

- If equal volume of two acids mixed

$pH =$ Lower $pH + 0.30$

- If equal volume of two bases mixed

$pH =$ Higher $pH - 0.30$.

Q. 4g of NaOH dissolved in 1000 ml of H_2O conc of H^+ ions is

→ $[OH^-] = N_{Base} = \frac{Wt}{Ewt} \times \frac{1000}{Vml}$

$[OH^-] = \frac{4}{40} \times \frac{1000}{1000}$

$[OH^-] = 10^{-1}$

$Kw = [H^+][OH^-]$

$1 \times 10^{-14} = 10^{-1}[H^+]$

$[H^+] = 1 \times 10^{-13}$

Get More Learning Materials Here : CLICK HERE www.studentbro.in

Page No.
Date

Q. 100 ml of 0.2 M HCl ; 100 ml of 0.4 M HNO_3 are mixed pH = ?

→ If two or more only acids or only bases are mixed

$pH = -\log[H^+]$ $\curvearrowright$ $N_{RS} = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}$

$$[H^+] = \frac{0.2 \times 100 + 0.4 \times 100}{100 + 100}$$

$$= \frac{60}{200} = 0.3 = 3 \times 10^{-1}$$

$pH = 1 - \log 3$

$= 1 - 0.47$

Normality = Molarity × x-fact

Q. 200 ml of 0.2M HCl and 400 ml of 0.1 M NaCl pH ?
are mixed.

→ When A and B Both are mixed

① $N_aV_a > N_bV_b$ [Acidic] → $[H^+]_{RS}$ → $N_{RS} = \frac{N_aV_a - N_bV_b}{V_a + V_b}$

② $N_bV_b > N_aV_a$ [Basic] → $[OH^-]_{RS}$ → $N_{RS} = \frac{N_bV_b - N_aV_a}{V_a + V_b}$

③ $N_aV_a = N_bV_b$ [Neutral] → pH = 7

$N_aV_a = 0.2 \times 200 = 40$

$N_bV_b = 0.1 \times 400 = 40$

pH = 7

Q. 1 ml of 10^{-2} M is diluted upto 100 ml pH_{RS} = ?

→ $HCl \rightarrow H^+ + Cl^-$

$10^{-2}M \quad 10^{-2}M \quad 10^{-2}M$

$[H^+] = 10^{-2}$

pH = 2

Get More Learning Materials Here : CLICK HERE www.studentbro.in

- pH = 2 (Acid)

10 times diluted pH inc. by 1 unit (3)
100 times diluted pH inc. by 2 unit (4)
1000 times diluted pH inc. by 3 unit (5)

- pH = 13 (Basic)

10 times diluted pH dec by 1 unit (12)
100 times diluted pH dec by 2 units (11)
1000 times diluted pH dec by 3 units (10)

- Solubility product concept:

Solubility : The maxi. amount [moles or g] salt that can be dissolved in 1L of solvent (H_2O).

- Solubility depends upon:

① Nature of salt.
② Nature of solvent (H_2O)
③ Temp.
④ Common - Ion effect.

- soluble salts $\propto$ spiringly soluble salt ✓

NaCl	AgCl	$PbSO_4$
NaOH	$AgNO_3$	
$CaCl_2$	$BaSO_4$	
$Fe_2(SO_4)_3$	CaF_2	

- Solubility $\propto \dfrac{1}{\text{Temp}}$
- Solubility $\propto \dfrac{1}{\text{C·I effect}}$

- Ksp : product of molar conc of ions which is soluble in H_2O.

Ksp = $[Ag^+][Cl^-]$

Get More Learning Materials Here : CLICK HERE www.studentbro.in

I. Relation Between ksp and s:

$A_xB_y \longrightarrow x^x y^y s^{x+y}$

- Solubility = Molarity $= n \times \frac{1}{V_L}$ = mole/L

Q. The solubility of CaF_2 in 0.1 NaF solution. If $K_{sp} = 5.3 \times 10^{-11}$ (CaF_2)

→ CaF_2 $\quad K_{sp} = [Ca^{+2}][F^-]^2 \quad CaF_2 \rightleftharpoons Ca^{+2} + 2F^-$

$\downarrow \quad 5 \times 10^{-11} = s \times (10^{-1})^2 \quad S \quad 0.1M$

0.1 M NaF $\quad S = 5 \times 10^{-9}$ //

Q. Find out solubility of $Ni(OH)_2$ In 0.1 M NaOH If $K_{sp}(Ni(OH)_2 = 2 \times 10^{-15}$

→ $Ni(OH)_2 \quad K_{sp} = [Ni^{+2}][OH]^2 \quad Ni(OH)_2 \rightleftharpoons Ni^{+2} + 2OH^{\ominus}$

$\downarrow \quad 2 \times 10^{-15} = s \times [10^{-1}]^2 \quad S \quad 0.1M$

0.1 M NaOH $\quad S = 2 \times 10^{-13}$ //

Q pH of salt soln of $Ca(OH)_2 = 9$ The ksp of $Ca(OH)_2$ is

→ $Ca(OH)_2 \longrightarrow$ Base pH = 9 pOH = 5 $[OH^-] = 10^{-5}$

$Ca(OH)_2 \rightleftharpoons Ca^{+2} + 2OH^-$ $\quad 2S = 10^{-5}$

$S \quad 2S \quad S = 10^{-5} \times \frac{1}{2}$

$K_{sp} = [Ca^{+2}][OH^-]^2$

$= \frac{1}{2} \times 10^{-5} \times 10^{-5} \times 10^{-5}$

$= 0.25 \times 10^{-15}$

Get More Learning Materials Here : CLICK HERE www.studentbro.in

Page No.
Date

Q. Conc. of Ag^+ ions in Ag_2CrO_4 = 2.2×10^{-4} mol/L
Then ksp of Ag_2CrO_4 = ?

→

$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^-$ $2s = 2.2 \times 10^{-4}$
2s s $s = 1.1 \times 10^{-4}$

$Ksp = [Ag^+]^2 [CrO_4^-]$
$Ksp = (2.2 \times 10^{-4})^2 (1.1 \times 10^{-4})$ //

- Concept :-

① unsaturated solⁿ ⟶ KIP < KSP
② saturated solⁿ ⟶ KIP = KSP
③ super saturated Solⁿ ⟶ KIP > Ksp (ppt)

Ksp – sat. solⁿ.
KIP – not a sat. solⁿ

Q. Eq. vol of 0.002 M NaI and $Cu(ClO_3)_2$ are mixed. If Ksp of $Cu(ClO_3)_2$ = 7.4×10^{-8}
In this CuI_2 salt get ppt or not ?

→ $CuI_2 \rightleftharpoons Cu^{+2} + 2I^-$
0.001 M   2 × 0.001 M

$KIP = [Cu^{+2}][I^-]^2$
$= (10^{-3})(2 \times 10^{-3})^2$
$KIP = 4 \times 10^{-9}$ KIP < Ksp → unsaturated solⁿ
$Ksp = 7.4 \times 10^{-8}$ No precipitation.

- Buffer solution - The solution can't change its pH upon addition of small amount of A/B.

① Acidic B.S

WA + salt of its conjugate Base

$CH_3COOH + CH_3COO^{\ominus}Na^+$

$HCN + K^+CN^-$

$H_2CO_3 + Na_2CO_3$

$HF + NaF$

$H_3PO_4 + Na_3PO_4$

S.A
- $HClO_4$
- HI
- HBr
- H_2SO_4
- HCl
- HNO_3
- $HClO_3$

② Basic B.S

WB + salt of its conj. Acid

$NH_4OH + NH_4^+Cl^-$

$NH_4OH + NH_4NO_3$

$NH_4OH + (NH_4)_2SO_4$

$NH_4OH + (NH_4)_3PO_4$

$NH_4OH + NH_4F$

S.B

IA	IIA
NaOH	$Mg(OH)_2$
KOH	$Ca(OH)_2$
RbOH	$Sr(OH)_2$
CsOH	$Ba(OH)_2$

- ABS ⟶ $pH = pKa + \log \frac{(salt)}{[Acid]}$

BBS ⟶ $pOH = pKb + \log \frac{(salt)}{[Base]}$

① N/M
② NV/MV
③ n
④ No. of eq.

Q. A vessel contain 0.8 moles NH_4Cl and 0.2 moles of NH_4OH If $Kb = 2 \times 10^{-5}$ pH of solution is

→

$Kb = 2 \times 10^{-5}$

$pKb = 5 - \log 2$

$= 5 - 0.30$

$= 4.70$

BBS ⟶ $pOH = pKb + \log \frac{[salt]}{[Base]}$

$pOH = 4.70 + \log \frac{0.8}{0.2}$

$pOH = 5.30.$

$pH + 5.30 = 14$

$pH = 14 - 5.30$ //

Get More Learning Materials Here : CLICK HERE www.studentbro.in

Q. Ka of HCN = 5×10^{-10} To main pH = 9 The volume of 5M KCN solution required to add 10 ml of 2M HCN.

→ ABS = $pH = pKa + \log \frac{(salt)}{(Acid)}$

$Ka = 5 \times 10^{-10}$

$pKa = 10 - \log 5$

$= 10 - 0.69$

$= 9.31$

$9 = 9.31 + \log \frac{M_S V_S}{M_A V_A}$

$-0.30 = \log \frac{5 \cdot V_S}{2 \times 10}$

$\log 2 = \log \frac{20}{5 \times V_S}$ //

Q. The dissociation constant of WB = 10^{-4} The ratio of salt and WB to main pH = 9.

→ pH = 9 pOH = 5 $Kb = 10^{-4}$

$pKb = 4$

BBS → $pOH = pKb + \log \frac{(salt)}{(Base)}$

$5 = 4 + \log \frac{(S)}{(B)}$

$1 = \frac{\log (S)}{\log (B)}$ $\log 10 = \frac{\log (S)}{\log (B)}$ //

Q. In an acidic B.S (pH = 4.4) and $\frac{(salt)}{(Acid)} = 2:1$ Value of Ka ?

→ $pH = pKa + \log \frac{(salt)}{(Acid)}$

$pKa = 4.10$

$4.4 = pKa + \log 2$

$-5 + 0.90$ ↘

$Ka = 7 \times 10^{-5}$

$pKa = 4.4 - 0.30$

$pKa = 4.10$

$Ka = 7 \times 10^{-5}$

9.0 ←
≤4 (−1)
4 to 6 (−1.5)
≥6 (−2)

Get More Learning Materials Here : CLICK HERE www.studentbro.in

- Buffer capacity :- The no. of moles of A or B added to 1L of B.S to change the pH value.

$$B.C = \frac{\text{no. of moles of A/B added to 1L B.S}}{\text{change in pH value.}}$$

- Salt hydrolysis ⟶ salt + $H_2O \longrightarrow A + B$

① SA + SB salt — Neutral ph = 7 Both H^+ or OH^- ions

② WA + SB salt — Basic ph > 7 $OH^\ominus$ ions anionic

③ SA + WB salt — Acidic ph < 7 H^+ ions cationic

④ WB + WA salt — Neutral ph = 0 No H^+ or $OH^\ominus$ ions

① $\underset{}{CH_3COONa} + H_2O \longrightarrow \underset{WA}{CH_3COOH} + \underset{SB}{NaOH}$

$CH_3COO^\ominus Na^\oplus + H_2O \longrightarrow CH_3COOH + Na^+ + OH^\ominus$

$\longrightarrow CH_3COO^\ominus + H_2O \longrightarrow CH_3COOH + OH^\ominus$

1) CH_3COONa is a salt of WA and SB
2) Nature of salt ⟶ $OH^\ominus$ (Basic) $\boxed{pH > 7}$
3) It undergo anionic hydrolysis
4) This salt gives OH^- ion.

Q. pH of NH_4CN [$Ka = 9 \times 10^{-10}$ $Kb = 2 \times 10^{-5}$]

→ WA + WB ⟶ $pH = 7 + \frac{1}{2}(PKa - PKb)$

$$pH = 7 + \frac{1}{2}(9.05 - 4.70)$$

Get More Learning Materials Here : CLICK HERE www.studentbro.in

Q. The hydrolysis constant of 0.5 M NaCN [Ka of HCN = 2×10^{-6}]

→ ① $Kh = \frac{Kw}{Ka} = \frac{1\times10^{-14}}{2\times10^{-6}} = 0.5\times10^{-8} = 5\times10^{-9}$

② $h = ?$

$Kh = ch^2$

$5\times10^{9} = 5\times10^{-1}\,h^2$

$h^2 = 10^{-8}$

$h = 10^{-2}\,\%$.

③ $pH = 7 + \frac{1}{2}(pKa + \log c)$

$pH = 7 + \frac{1}{2}(5.70 + 0.31)$ //

●	Kh value	h value	[H+] or [OH−]	pH value
SA + SB	No hydrolysis	—	Neutral	pH = 7
SA + WB	cationic $Kh = \frac{Kw}{Kb}$	$h = \sqrt{\frac{Kw}{Kb\cdot c}}$	$[H^+] = ch$ $[H^+] = \sqrt{\frac{Kw\cdot c}{Kb}}$	Acidic nature pH < 7 $pH = 7 - \frac{1}{2}(pKa + \log c)$
WA + SB	Basic Anionic $Kh = \frac{Kw}{Ka}$	$h = \sqrt{\frac{Kw}{Ka\cdot c}}$	$[OH^-] = ch$ $[OH^-] = \sqrt{\frac{Kw\cdot c}{Ka}}$	Basic nature ph > 7 $pH = 7 + \frac{1}{2}(pKb + \log c)$
WA + WB	Neutral $Kh = \frac{Kw}{Ka\cdot Kb}$	$h = \sqrt{\frac{Kw}{Ka\cdot Kb}}$	$[H^+] = Ka\cdot h$ $[H^+] = \sqrt{\frac{Kw\cdot Ka}{Kb}}$	$pH = 7 + \frac{1}{2}(pKa - pKb)$

h = degree of hydrolysis

Kh = Hydrolysis constant.

Get More Learning Materials Here : CLICK HERE www.studentbro.in