

1. $pH = -\log_{10} [H^+]$, $pOH = -\log_{10} [OH^-]$.

→ $pH + pOH = 14$. [concentration should be in terms of $\frac{N}{l}$ for calculating pH]

$[H^+] = a \times 10^{-b}$

$pH = \left(b + \log \frac{1}{a}\right) = (b - \log a)$

$\log 10 = 1$

$\log 1 = 0$

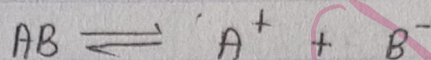
$\log 2 = 0.3010$

$\log 3 = 0.4771$

$\log 5 = 0.7$

$\log 7 = 0.84$

2. Ostwald's Law of Dilution



(t=0) C 0 0

(t=t) c-αx αx αx

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

[if $0 \leq \alpha \leq 5\%$
(1-α=1)]

⇒ $K = C\alpha^2$

→ $\alpha = \sqrt{\frac{K}{C}}$ K depends only on temperature

At constant temp. $\alpha \propto \frac{1}{\sqrt{C}}$, $\alpha \propto \frac{1}{\sqrt{n/V}}$ ⇒ $\alpha = \sqrt{V}$

∴ on infinite dilution ($V \rightarrow \infty$), $\alpha \approx 100\%$

⇒ Weak electrolyte behaves as strong electrolyte

Weak acid (HA)

Weak base (BOH)

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

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

→ $[H^+] = C\alpha = \sqrt{K_a C}$

$[OH^-] = C\alpha = \sqrt{K_b C}$

$pH = -\frac{1}{2} [\log K_a + \log C]$

$pOH = -\frac{1}{2} [\log K_b + \log C]$

$= \frac{1}{2} pK_a - \frac{1}{2} \log C$

$= \frac{1}{2} pK_b - \frac{1}{2} \log C$

Ostwald's law of dilution is applicable only to weak electrolyte

→ Factors affecting α ⇒ (1) temperature (2) dilution (V)

(3) Nature of electrolyte

(4) Dielectric constant (μ) [more μ ⇒ α ↑]

one electrolyte = strong ← (5) Common ion effect (α ↓)
other = weak ← Odd ion effect (α ↑)

3. Water -

ISHITA KANODIA

→ No. of moles of H_2O in 1 L of water = $\frac{1000}{18} = 55.5$ moles
 $[H^+] = 10^{-7} M$
 no. of $H^+ = 10^{-7} N_A$
 $\therefore$ one H^+ ion is obtained from $55.5 \times 10^7 H_2O$ molecules.
 → $\alpha = \frac{10^{-7} N_A}{55.55 N_A} = 18 \times 10^{-8}$ $\% \alpha = 1.8 \times 10^{-7}$

→ $K = \frac{[10^{-7}][10^{-7}]}{55.55} = 1.8 \times 10^{-16}$

→ $K_w = [H^+][OH^-]$ $K_{H_2O} = \frac{[H^+][OH^-]}{[H_2O]}$

$\therefore K_{H_2O} [H_2O] = K_w$

→ $-\log K_w = -\log [H^+] - \log [OH^-]$ [water is neutral]
 $\Rightarrow$ $pK_w = pH + pOH = 14$ $\therefore pH = pOH = \frac{pK_w}{2}$

→ If Temp. $\uparrow$, ionisation $\uparrow \Rightarrow pH \downarrow$, $pOH \downarrow$.
 $\Rightarrow K_w \uparrow$

4. Salts - ① Normal/general salts ⑥ Mixed salts

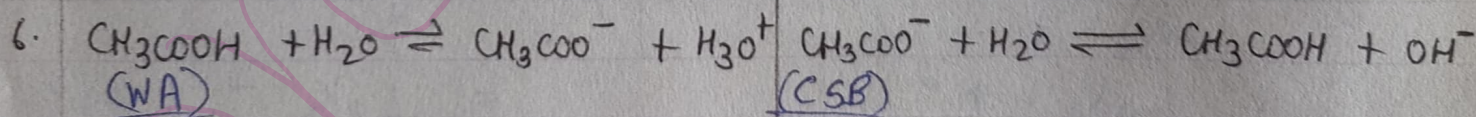
② Acid salts

③ Basic salts

④ Double salts

⑤ Complex salts

5. Strong acid/base $\rightleftharpoons$ weak conjugate base/acid.
 weak acid/base $\rightleftharpoons$ strong conjugate base/acid.



$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$

$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$

$K_a \times K_b = [H^+][OH^-] = K_w$

$\therefore K_a \times K_b = 10^{-14}$

$pK_a + pK_b = pK_w = 14$

ISHITA KANODIA

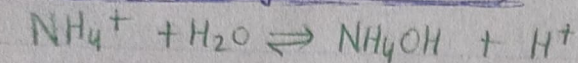
7. Salt Hydrolysis - (reverse of neutralization)

ISHITA KANODIA

not shown by SA-SB salt (pH=7)

a) Cationic Hydrolysis [SA-WB]

step 1 K_h (hydrolysis constant)



$$K_h = \frac{[\text{NH}_4\text{OH}][\text{H}^+]}{[\text{NH}_4^+]}$$

$$K_h = \frac{K_w}{K_b}$$

step 2 Degree of hydrolysis (h)

$$K_h = \frac{Ch^2}{1-h} \approx Ch^2$$

$$h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{K_w}{K_b C}}$$

step 3 pH. $[\text{H}^+] = Ch = \sqrt{\frac{K_w C}{K_b}}$

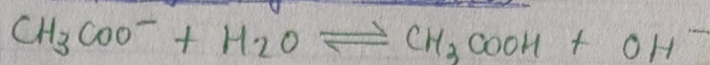
$$\text{pH} = -\frac{1}{2} [\log K_w + \log C - \log K_b]$$

$$\text{pH} = 7 - \frac{1}{2} \text{p}K_b - \frac{1}{2} \log C$$

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

Anionic Hydrolysis [WA-SB]

K_h (hydrolysis constant)



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

$$K_h = \frac{K_w}{K_a}$$

Degree of hydrolysis (h)

$$K_h = \frac{Ch^2}{1-h} \approx Ch^2$$

$$h = \sqrt{\frac{K_w}{K_a C}}$$

$$[\text{OH}^-] = Ch$$

$$\text{pOH} = 7 - \frac{1}{2} \text{p}K_a - \frac{1}{2} \log C$$

$$\text{pH} = 7 + \frac{1}{2} \text{p}K_a + \frac{1}{2} \log C$$

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

b) Hydrolysis of [WA-WB] Salt

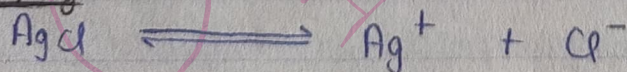
$$K_h = \frac{K_w}{K_a \times K_b}$$

$$K_h = h^2 \therefore h = \sqrt{K_h} = \sqrt{\frac{K_w}{K_a \times K_b}}$$

$$\text{pH} = 7 + \frac{1}{2} \text{p}K_a - \frac{1}{2} \text{p}K_b$$

$$[\text{H}^+] = K_a h$$

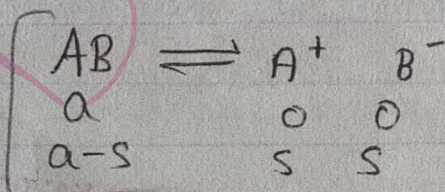
8. Solubility-



$$K = \frac{[\text{Ag}^+][\text{Cl}^-]}{[\text{AgCl}]}$$

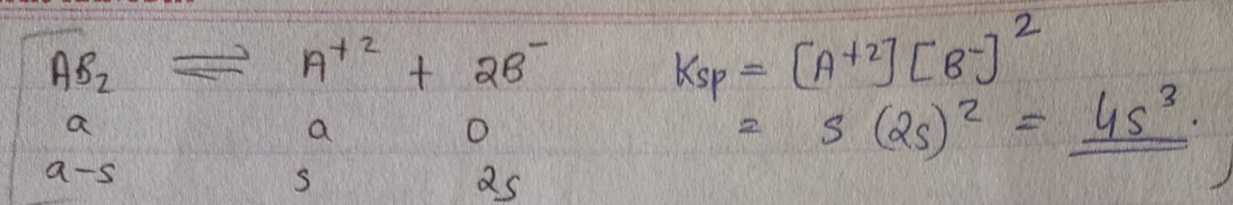
$$K_{sp} = [\text{AgCl}] K = [\text{Ag}^+][\text{Cl}^-]$$

$[\text{AgCl}] \rightarrow$ constant.



$$K_{sp} = [\text{A}^+][\text{B}^-] = s^2$$

ISHITA KANODIA



→ Ionic product [IP] $IP = [Ag^+]_i [Cl^-]_i$
 $K_{sp} = [Ag^+]_{eq} [Cl^-]_{eq} = \text{constant}$

$IP = K_{sp}$ = saturation solution, equilibrium exists

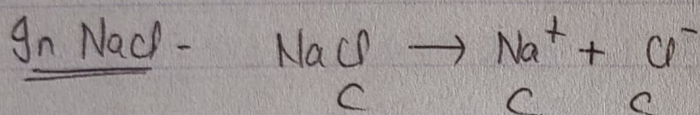
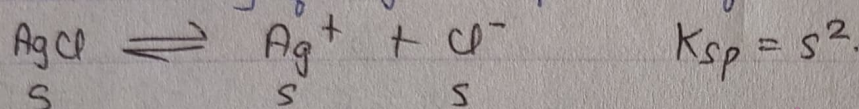
$IP > K_{sp}$ = supersaturated solution, precipitation occurs

$IP < K_{sp}$ = unsaturated solution, precipitation will not occur.

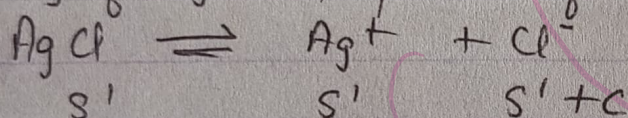
→ Common ion decreases solubility.

Q type

Find solubility of AgCl in presence of C NaCl?



solubility of AgCl in presence of NaCl = s'



$$K_{sp} = s'[s' + C] = s'^2 + s'C \Rightarrow s'C = K_{sp}$$

$$\Rightarrow s' = \frac{K_{sp}}{C}$$

Precipitation occurs in presence of common ion.

9. Isohydric solution - different solutions having the same pH.

$$\begin{array}{c}
 \rightarrow \text{Relative strength (RS)} = \frac{\text{Strength of Acid 1}}{\text{Strength of Acid 2}} = \frac{[H^+]_1}{[H^+]_2} = \frac{C_1 \alpha_1}{C_2 \alpha_2} = \frac{\sqrt{K_{a1} C_1}}{\sqrt{K_{a2} C_2}}
 \end{array}$$

10. Buffer

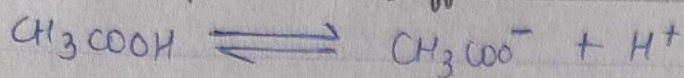
- a solution which resists change in pH on addition of small amount of SA or SB.
- pH does not change appreciably on adding SA/SB.
 - pH of buffer soln. does not depend on volume of solution.
 - ∴ solution can be diluted without change in pH.
 - pH does not change even if it is kept for a long time.

(a) Simple Buffer - salt of WA-WB.

$$pH = 7 + \frac{1}{2} pK_a - \frac{1}{2} pK_b.$$

(b) Mixed Buffer

Acidic Buffer



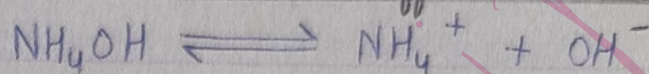
$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$\Rightarrow [H^+] = \frac{K_a [CH_3COOH]}{[CH_3COO^-]}$$

$$\Rightarrow pH = pK_a + \log \frac{[salt]}{[acid]}$$

$$= pK_a + \log \frac{[conjugate base]}{[acid]}$$

Basic Buffer



$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]}$$

$$[OH^-] = \frac{K_b [NH_4OH]}{[NH_4^+]}$$

$$pOH = pK_b + \log \frac{[salt]}{[base]}$$

$$= pK_b + \log \frac{[conjugate acid]}{[base]}$$

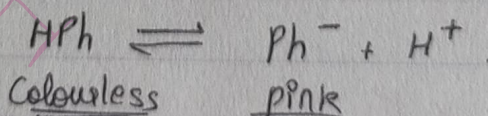
→ Buffer capacity = $\frac{\text{no. of moles of acid or base added per litre}}{\text{change in pH of buffer solution}}$

more the BC $\Rightarrow$ more resistant is the solution to pH change.

II. Indicators - (weak organic acids or bases).

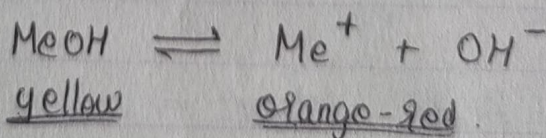
(a) Acidic indicators

Phenolphthalein (HPh)



(b) Basic indicators

Methyl orange (MeOH)



$$\left. \begin{array}{l} pH \text{ of acidic indicator} \\ pOH \text{ of basic indicator} \end{array} \right\} = pK_I + \log \frac{[Ionised form]}{[Unionised form]}$$

ACIDS $\Rightarrow$ SOUR TASTE
 BASES $\Rightarrow$ BITTER TASTE

Indicators	Colour In Acidic Medium	Colour in Basic Medium.
Methyl orange	orange red	yellow
Methyl red	red	yellow
Phenol red	yellow	red
Phenolphthalein	colourless	pink

12. Arrhenius concept

acids $\Rightarrow$ produce free H^+ ions in aq. soln.

bases $\Rightarrow$ produce free OH^- ions in aq. soln.

$\rightarrow$ Brønsted-Lowry Concept

acids $\Rightarrow$ tendency to donate H^+

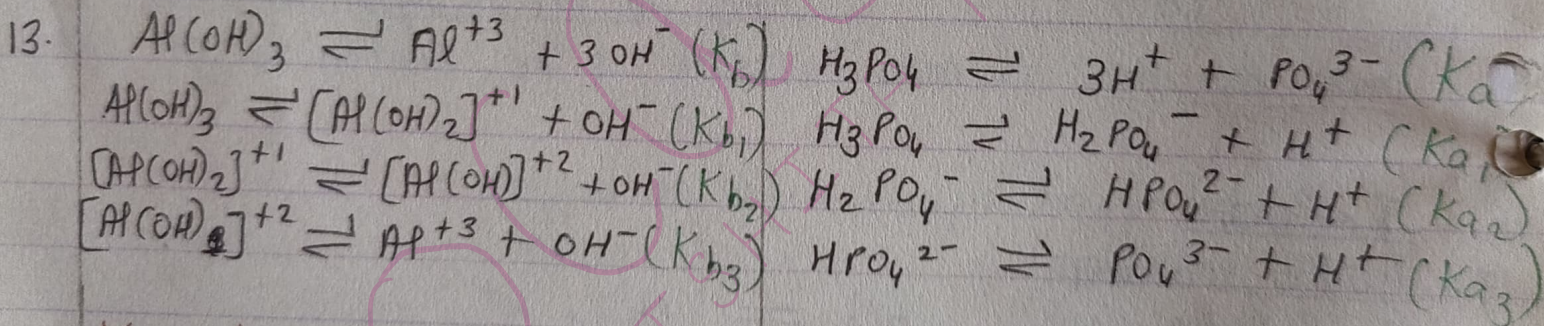
bases $\Rightarrow$ tendency to accept H^+ .

Water = amphoteric / amphiprotic $\Rightarrow$ can act on both acids and bases.

$\rightarrow$ Lewis Concept

acids $\Rightarrow$ tendency to accept lone pair of e^-

bases $\Rightarrow$ tendency to donate lone pair of e^- .



$$K_b = K_{b1} \times K_{b2} \times K_{b3}$$

$$K_{b1} > K_{b2} > K_{b3}$$

$$K_a = K_{a1} \times K_{a2} \times K_{a3}$$

$$K_{a1} > K_{a2} > K_{a3}$$

14. • Protogenic = Acidic solvent $\Rightarrow$ tendency to donate proton
 • Protophilic = Basic solvent $\Rightarrow$ tendency to accept proton
 • Amphiprotic = Amphoteric $\Rightarrow$ tendency to accept or donate proton
 • Aprotic solvents $\Rightarrow$ neither donate nor accept
 • Ampholyte / Zwitter ion $\Rightarrow$ tendency to accept and donate proton
 Eg - Sulphuric acid, glycine.

