

10+2 PCM NOTES

BY

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(PDF version handwritten notes of Maths, Physics and Chemistry for 10+2 competitive exams like JEE Main, WBJEE, NEST, IISER Entrance Exam, CUCET, AIPMT, JIPMER, EAMCET etc.)

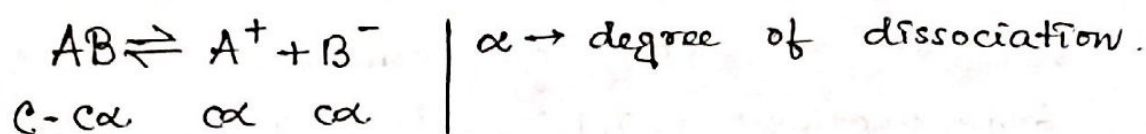


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With best wishes from Joyoshish Saha

- * Strong & Weak Electrolytes: On the basis of to what extent an electrolyte gets ionised. Ionised almost completely, then strong. Degree of ionisation (α) = (no. of moles dissociated) / (Total no. of moles taken). α & electrical conductivity (σ) of strong electrolytes have high values. Strong - NaOH, HCl, H_2SO_4 , HNO_3 , NaCl, $BaCl_2$. Weak - CH_3COOH , HCN, NH_4OH , H_3BO_3 .

- * Ostwald Dilution Law: (for weak electrolyte).



$$K = \frac{C\alpha^2}{1-\alpha} = C\alpha^2 \quad (1-\alpha \approx 1 \text{ for weak electrolyte})$$

$\alpha \propto \frac{1}{\sqrt{C}}$ if V is the volume of solution containing 1 mole of solute, then

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

[$1-\alpha \approx 1$ only if $\alpha < 5\%$]

- * Concept of Acids and Bases:

1. Arrhenius Concept: acid - hydrogen containing compound that gives free hydrogen ions when dissolved in water.

base - hydroxyl group containing compound that gives hydroxyl ions when dissolved in water.

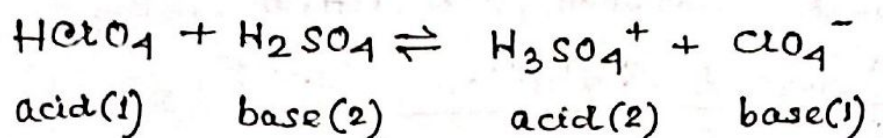
• limitations: i) limited to aqueous medium; ii) H^+ combines with H_2O to form

H_3O^+ , Hydronium ion; iii) fails to explain acidic character of CO_2 , SO_2 , SO_3 etc and basic character of NH_3 , CaO , MgO etc. (These limitations)

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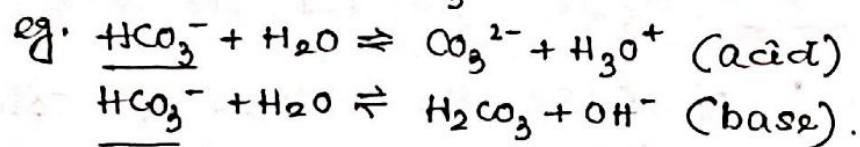
can be overcome in the presence of water acting as solvent - $\text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3$ (acid),
 $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4\text{OH}$ (base).

2. Bronsted-Lowry's Protonic Concept: acid-substance that can give a proton & base-acceptant of proton. [A substance that can act both as an acid and base in different reactions is called amphiprotic] • an acid-base pair which differ by a proton is called conjugate acid-base pair. $\text{Con. acid} \rightleftharpoons \text{Con. base} + \text{H}^+$



If acid is strong, its con. base is weak & vice-versa.

Amphiprotics - H_2O , HCO_3^- , HSO_4^- , H_2PO_4^- , HPO_4^{2-} , H_2PO_3^- , HSO_3^- .



3. Lewis Concept: acid-substance that can accept a pair of electrons (electrophile), base-substance that can furnish a pair of electrons to form a co-ordinate bond (nucleophile). [All Lewis bases are Bronsted bases, but all Lewis acids are not Bronsted acids]. All simple cations or molecules whose central atom has incomplete octet, the molecules in which central atom can expand its octet, molecules having multiple bonds between atoms of different electronegativity can act as Lewis acid. All simple atom, molecules having one or more lone pairs of electrons can act as Lewis bases.

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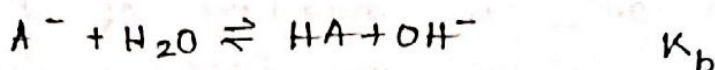
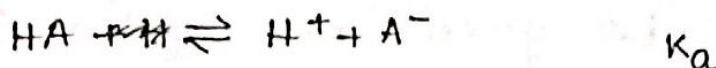
* Ionic Product of Water: $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$

$$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2} \Rightarrow K_w = K \times [H_2O]^2 = [H_3O^+][OH^-].$$

at 298 K, $K_w = 10^{-14} \text{ (mol/L)}^2$.

for pure water $[H_3O^+] = [OH^-] = \sqrt{K_w} = 10^{-7} \text{ mol/L}$.

• for conjugate acid-base pair HA & A^- in aqueous solution



Higher the pK_a value, lower is acid strength, and higher the basic strength.

$$\begin{aligned} \Rightarrow \log K_w &= \log K_a + \log K_b \\ \Rightarrow pK_w &= pK_a + pK_b \\ &= 14 \text{ at } 298 \text{ K} \end{aligned}$$

$$[p \rightarrow -\log()]$$

* pH scale: [Søren Sørensen (1909)]

$$pH = -\log[H^+] = \log\left[\frac{1}{[H^+]}\right] \quad \begin{array}{l} H^+ \text{ concentration in} \\ \text{mole/L.} \end{array}$$

at 298 K, for neutral solution $pH = 7$.

acidic, $pH < 7$; basic, $pH > 7$.

$pH + pOH = 14$

• In calculation of pH of strong acid or strong base, the acid/base is the only significant source of H_3O^+ or OH^- .

* Total H_3O^+ or OH^- in the mixture of Strong acids or bases: $[H_3O^+] = \frac{\sum N_1 V_1}{\sum V}$, $[OH^-] = \frac{\sum N_2 V_2}{\sum V}$

the resultant solⁿ of strong acid (N_1, V_1) & strong base (N_2, V_2) is - [$N \rightarrow$ normality
 $V \rightarrow$ volume]

Joyoshish Saha i) neutral, if $N_1 V_1 = N_2 V_2$.

ii) basic, if $N_2 V_2 > N_1 V_1$, $[OH^-]_{\text{resultant}} = \frac{N_2 V_2 - N_1 V_1}{V_1 + V_2}$

iii) acidic, if $N_1 V_1 > N_2 V_2$, $[H_3O^+]_{\text{resultant}} = \frac{N_1 V_1 - N_2 V_2}{V_1 + V_2}$

• pH of weak monobasic acid: $HA + H_2O \rightleftharpoons A^- + H_3O^+$

$$K_a = \frac{c\alpha^2}{1-\alpha} = c\alpha^2 [1-\alpha \approx 1] \quad c-\alpha \quad c\alpha \quad c\alpha$$

$$\alpha = \sqrt{\frac{K_a}{c}} \quad [H_3O^+] = c\alpha = \sqrt{K_a c} \quad pH = -\log \sqrt{K_a c}$$

$$\boxed{pH = \frac{1}{2} [pK_a - \log c]}$$

pH of weak monoacidic base: $\boxed{pOH = \frac{1}{2} [pK_b - \log c]}$

• total $[H_3O^+]$ in mixture of two weak acids can be given as $[H_3O^+] = (K_{a1}c_1 + K_{a2}c_2)^{1/2}$.

• total $[OH^-]$ in mixture of two weak bases can be given as $[OH^-] = (K_{b1}c_1 + K_{b2}c_2)^{1/2}$.

• total $[H_3O^+]$ in a mixture of weak acid & strong acid - $\frac{c_2 + \sqrt{c_2^2 + 4K_a c_1}}{2}$, $c_1 \rightarrow$ weak acid (K_a), $c_2 \rightarrow$ strong acid.

* Relative Strength of Weak acids & bases:

Strength determined by dissociation rate.

$$\frac{\text{strength of acid (I)}}{\text{strength of acid (II)}} = \sqrt{\frac{K_{a1}}{K_{a2}}} \quad \left. \begin{array}{l} \text{higher } pK_a, \\ \text{lower acid} \\ \text{strength.} \end{array} \right\}$$

* Common Ion Effect. degree of dissociation of a weak electrolyte is suppressed by the addition of another strong electrolyte containing common ion. This is common ion effect.

$AB \rightleftharpoons A^+ + B^-$, $K = \frac{[A^+][B^-]}{[AB]}$, when another electrolyte containing A^+ , B^- is added, equilibrium will shift to the left side.

* Salt hydrolysis: Salts in water react with H_2O to form acid-base. It is called hydrolysis.

1. Salt of weak acid | being hydrolysed the solution and strong base | will be basic.

$$K_h = \frac{K_w}{K_a}, \alpha = \sqrt{\frac{K_w}{K_a c}}$$

$$[H^+] = \sqrt{\frac{K_w K_a}{c}}, [OH^-] = \sqrt{\frac{K_w c}{K_a}}$$

$$pH = 7 + \frac{1}{2} (pK_a + \log c).$$

2. Salt of Weak base | the solution will be and strong acid | acidic.

$$K_h = \frac{K_w}{K_b}, \alpha = \sqrt{\frac{K_w}{K_b c}}$$

$$[H^+] = \sqrt{\frac{K_w c}{K_b}}, pH = 7 - \frac{1}{2} (pK_b + \log c).$$

3. Salt of Weak acid | $K_h = \frac{K_w}{K_a K_b}, h = \sqrt{\frac{K_w}{K_a \times K_b}}$ and weak base

$h \rightarrow$ dissociation constant

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

* Buffer solutions: Resists a change in its pH when such a change is caused by the addition of small amount of acid / base.

1. Buffer of a weak acid and its salt with a strong base: $pH = pK_a + \log \frac{[con. base]}{[Acid]}$

(Henderson-Hasselbalch eqn)

2. Buffer of a weak base & its salt with a strong acid: $pOH = pK_b + \log \frac{[con. acid]}{[base]}$

3. Salt Buffer: $[H^+] = \sqrt{\frac{K_w K_a}{K_b}}$

* Buffer Capacity: The no. of moles of acid / base that has to be added to one litre of buffer solution to change its pH by one unit.

$$\text{Buffer capacity} = \frac{\text{No. of moles of acid/base to be added/L}}{\Delta \text{pH}}$$

P.e. maximum of $\frac{[\text{salt}]}{[\text{acid}]} = 1$, $\text{pH} = \text{pK}_a$ (for acidic buffer)

$$\frac{[\text{salt}]}{[\text{base}]} = 1, \text{pOH} = \text{pK}_b, \text{for basic buffer}$$

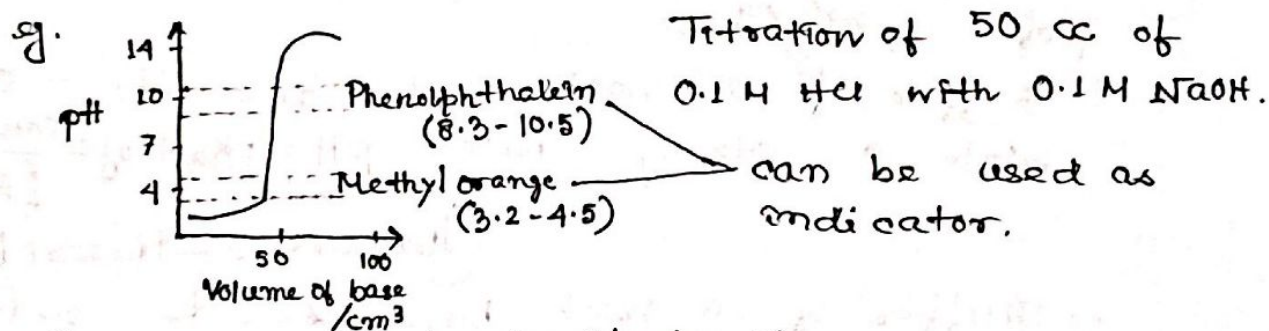
• An acidic buffer will be effective with the pH range $(\text{pK}_a - 1)$ to $(\text{pK}_a + 1)$.

• An basic buffer will be effective with the pH range $\text{pK}_w - (\text{pK}_b \pm 1)$.

* Acid-Base indicators : that changes colour according to hydrogen ion concentration of the solution.

$\text{pH} = \text{pK}$ (K is indicator dissociation constant) at the end point (when it changes colour).

A proper acid-base indicator is the one whose pH range falls on the vertical portion of the titration curve.



• Commonly used indicators:

Indicator	pH range	colour change	neutral colour pH
Methyl orange	3.2-4.5	pink to yellow	3.7
Methyl red	4.4-6.5	red to yellow	5.1
Litmus	5.5-7.5	red to blue	7.0
Phenol red	6.8-8.4	yellow to red	7.8
Phenolphthalein	8.3-10.5	colourless to pink	9.6

c.

* Solubility and Solubility Product: Amount of salt that makes the solution saturated per litre of solⁿ is called solubility of salt. $S, \text{ solubility} = \sqrt{K_{sp}}$

For a solute, $A_x B_y(s) \rightleftharpoons x A^{y+} + y B^{x-}$ $K_{sp} = [A^+][B^-]$

$$K_{sp} = [A^{y+}]^x [B^{x-}]^y = (xS)^x (yS)^y = x^x \cdot y^y \cdot S^{(x+y)}$$

* Calculation of Solubilities:

a) AgCl (salt of strong acid, strong base) in water:

At saturation point, $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
S moles S moles

$$K_{sp} = S^2 \Rightarrow \boxed{S = \sqrt{K_{sp}}}$$

AgCl in a solⁿ, having 0.1 M in AgNO₃

$AgCl \rightleftharpoons Ag^+ + Cl^-$; K_{sp} , solubility product of AgCl 10^{-10} .

In water, $S = 10^{-5}$. In the presence of AgNO₃ S will decrease (common ion effect). So,

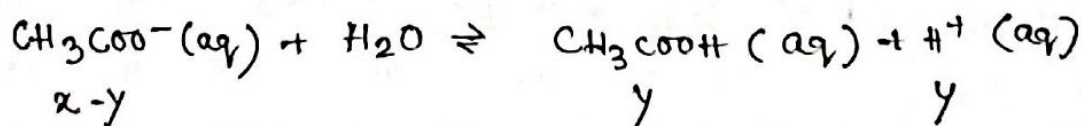
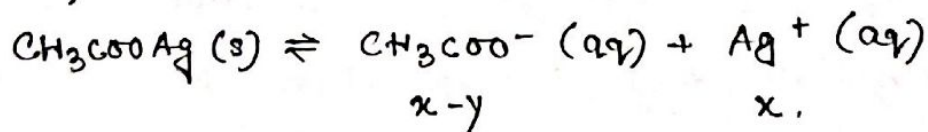
$[Ag^+]$ from AgCl $< 10^{-5}$ M. Say, the solubility here, S'

so, $[Ag^+]$ from AgNO₃ $\rightarrow 0.1$ M, $[Cl^-] = S'$ M

$$K_{sp} = 0.1 \times S' \Rightarrow S' = \frac{K_{sp}}{0.1} = 10^{-9} \text{ M.}$$

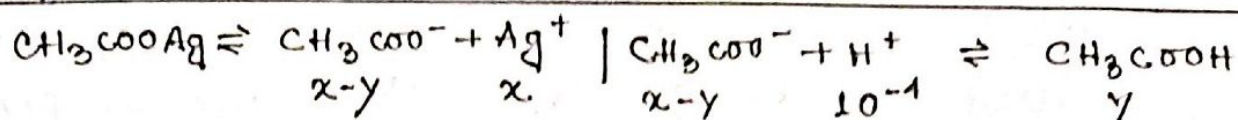
b. CH₃COOAg (salt of weak acid, strong base) in water:

say, solubility is x M.



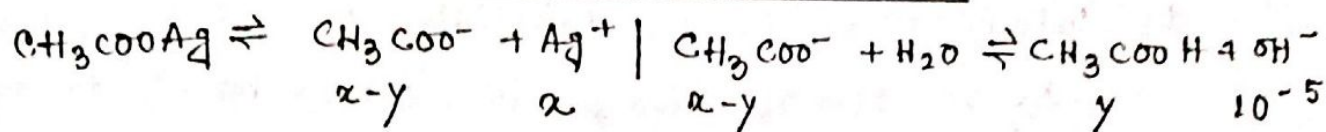
$$K_{sp} = x(x-y), \quad \frac{K_w}{K_a} = \frac{y^2}{x-y}$$

CH_3COOAg (in an acid buffer of $\text{pH} = 4$, assuming buffer does not have any common ion) :



$$K_{sp} = (x-y)x, \quad \frac{K_{sp}}{K_a} = \frac{y}{10^{-4}(x-y)}$$

CH_3COOAg (in a buffer of $\text{pH} = 9$) :

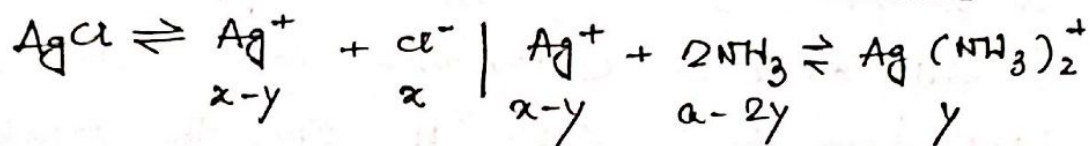


$$K_{sp} = x(x-y), \quad \frac{K_w}{K_a} = \frac{y \times 10^{-5}}{x-y}$$

for OH^- ions, suppression of degree of hydrolysis

$$x^2 = K_{sp} (\text{CH}_3\text{COOAg}) \quad [\text{neglecting } y]$$

c. AgCl (in aqueous solⁿ containing NH_3) :



$$K_{sp} = (x-y)x \quad (\text{NH}_3 \text{ initially 'a' M}).$$

$$\text{formation constant of } \text{Ag}(\text{NH}_3)_2^+, \quad K_f = \frac{y}{(x-y)(a-2y)}$$

* Precipitation of Salts : Ionic Product (IP) -

product of concentrations of ions present in solution. (reaction quotient)

g) a) $\text{IP} > K_{sp}$, precipitation occurs till $\text{IP} = K_{sp}$

b) $\text{IP} < K_{sp}$, a precipitate will not be formed & the solⁿ will be unsaturated.

c) $\text{IP} = K_{sp}$, a precipitate will not form as the solⁿ is saturated in that salt

c.

Ionic Equilibrium.

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* Types of Lewis Acids: i) molecules having central atom with incomplete octet e.g. BF_3 , AlCl_3 ; ii) simple cations, e.g. Ag^+ , Cu^{2+} , Fe^{3+} ; iii) molecules having central atom with empty d-orbitals e.g. SnCl_4 , SiF_4 , PF_5 etc; iv) molecules containing multiple bonds between different atoms e.g. $\text{O}=\text{C}=\text{O}$.

* pH of boiling water is 6.5625. It does not mean that boiling water is not neutral. It is due to greater dissociation of H_2O .

* pH can be zero in 1 N HCl solⁿ or it can be negative for more concentrated solⁿ like 2N, 3N, 10N etc.

* The buffer system present in blood is $\text{H}_2\text{CO}_3 + \text{NaHCO}_3$.