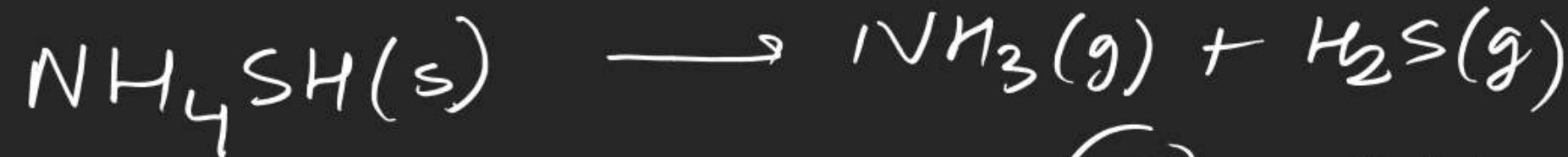


(23)

(28)

$$\frac{5.1 \text{ gm}}{51}$$



$$0.1 \text{ mol}$$

$$\underline{0.1 \times (0.3)} \quad \underline{\underline{0.1 \times 0.3}}$$

$$P = \frac{0.1 \times 0.3 \text{ RT}}{\sqrt{}}$$

(36)

$$K_p = P^2$$

(37)

@ backward

(B) No change

(B)

$$\text{(C)} \quad K_p = \underline{K_c (RT)^{1/2}}$$

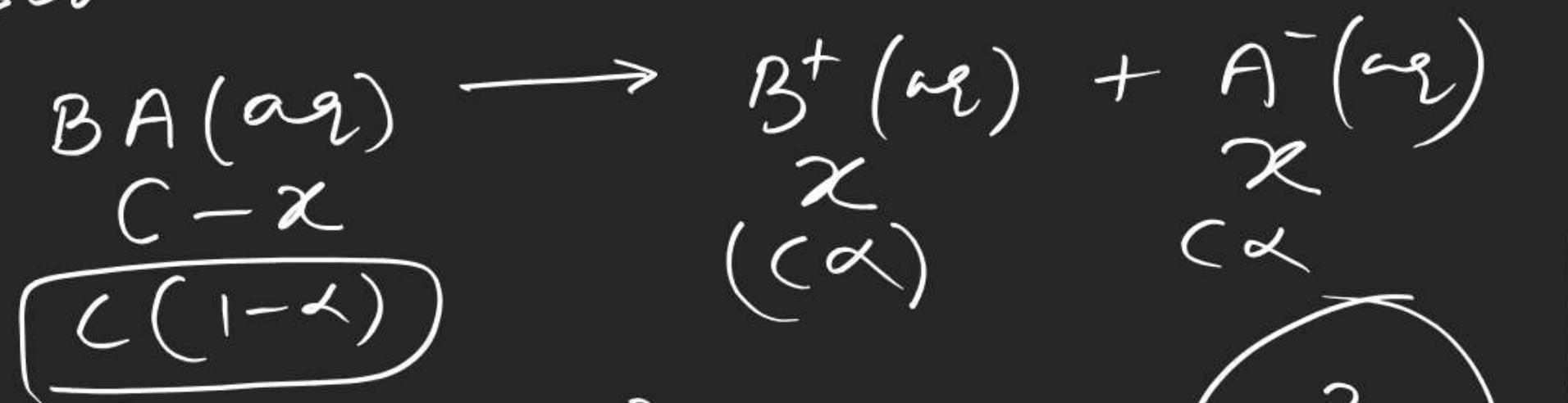
## Ionic Eq/Bm

This chapter mainly deals with

- 1) pH of solution
- 2) Solubility & solubility product
- 3) Indicators

## Arrhenius theory of dissociation:

It states that there exist a dynamic eqm b/w ionised and unionised solute.



$$K_{\text{diss}} = K_C = \frac{[\text{B}^+][\text{A}^-]}{[\text{BA}]} = \frac{\alpha^2}{C - \alpha} = \frac{C\alpha^2}{1 - \alpha}$$

$$= \frac{\alpha \cdot \alpha}{\cancel{C}(1-\alpha)} = \frac{\alpha^2}{1 - \alpha}$$

$\alpha'$  depends on

- 1) Nature of solute
- 2) Nature of solvent :— As dielectric constant  $\gamma_{es} \propto \gamma_{es}$
- 3) Temperature : Since dissociation is an endothermic process therefore as  $T \gamma_{es} \propto \gamma_{es}$

- ④ Concentration

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

$C \downarrow \alpha \uparrow$

Ostwald dilution law :—

As concentration  $\gamma_{es} \propto \gamma_{es}$

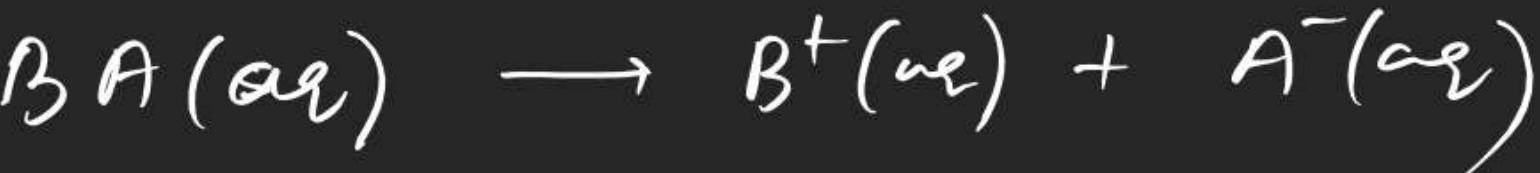
⑤ Common ion effect



$$c - x$$

$$x$$

$$x$$



$$c - y$$

$$\underline{b+y}$$

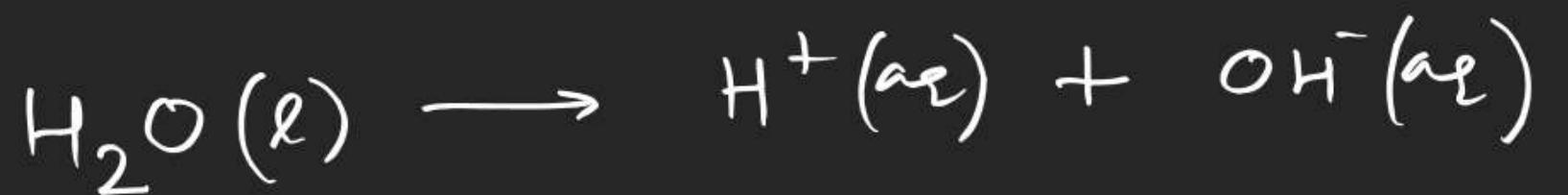
$$y$$

A)  $x = y$

B)  $x > y$

C)  $x < y$

# Dissociation of $H_2O$



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

$$K_{diss}[H_2O] = [H^+][OH^-]$$

$$[H_2O] = \frac{1000}{18}$$

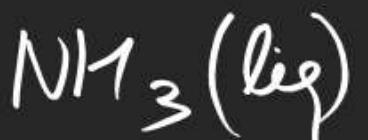
(ionic product of  $H_2O$ )

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

At  $25^\circ C$

$$K_w = 10^{-14} \left(\text{mol/lit}\right)^2$$

As  $T \uparrow$   $K_w \uparrow$  es



$[H^+] > [OH^-]$	Acidic
$[H^+] < [OH^-]$	Basic
$[H^+] = [OH^-]$	Neutral

at  $25^\circ C$ 

$[H^+] = 10^{-7}$

$[OH^-] = 10^{-7}$  Neutral

$[H^+] > 10^{-7}$

$[OH^-] < 10^{-7}$  Acidic

$[H^+] < 10^{-7}$

$[OH^-] > 10^{-7}$  Basic

$[H^+] [OH^-] = 10^{-14}$

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

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

at  $25^\circ C$ 

$[H^+] = 10^{-7}$  pH = 7 Neutral

$[H^+] > 10^{-7}$  pH < 7 Acidic

$[H^+] < 10^{-7}$  pH > 7 Basic

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

$pH + pOH = pK_w$

at  $25^\circ C$ 

$pH + pOH = 14$

Q. At  $80^\circ\text{C}$ ,  $K_w = 4 \times 10^{-12} \text{ M}^2$

find ① pH of pure  $\text{H}_2\text{O}$

② Define a solution as

acidic, basic or neutral

having pH	i	5	Acidic
	ii	6	Basic
	iii	7	Basic

$$\log 2 = 0.3$$

$$[\text{H}^+][\text{OH}^-] = 4 \times 10^{-12}$$

$$[\text{H}^+]^2 = [\text{OH}^-]^2 = 4 \times 10^{-12}$$

$$[\text{H}^+] = 2 \times 10^{-6}$$

$$\text{pH} = 6 - \log 2 = 5.7$$

Chemical eq 16<sup>m</sup>

JEE-Adv

O-II

(29)



$$\frac{1}{K_p}$$

$$P_{NH_3} = \frac{P_1/4 \times (3P_1/4)^3}{P^2}$$

$$\frac{1}{K_p} = \frac{P_1/4 \times (3P_1/4)^3}{P^2}$$

$$P_{NH_3}^2 = K_p \times P^4 \times \frac{27}{2^8}$$

$$P_{NH_3} = K_p^{1/2} \times P^2 \times \frac{3^{3/2}}{2^4}$$

(35)

$$K = \frac{1}{6} = 2$$