

~~23~~

23

5.1 gm

51



0.1 mol



$$\underline{0.1 \times 0.3}$$

$$\underline{0.1 \times 0.3}$$

$$P = \frac{0.1 \times 0.3 RT}{V}$$

36

$$K_c = K_p (RT)^{-1/2}$$

$$K_p = P^2$$

37

a) backward

b) No change

$$\textcircled{c} \underline{K_p = K_c (RT)^{1/2}}$$

B

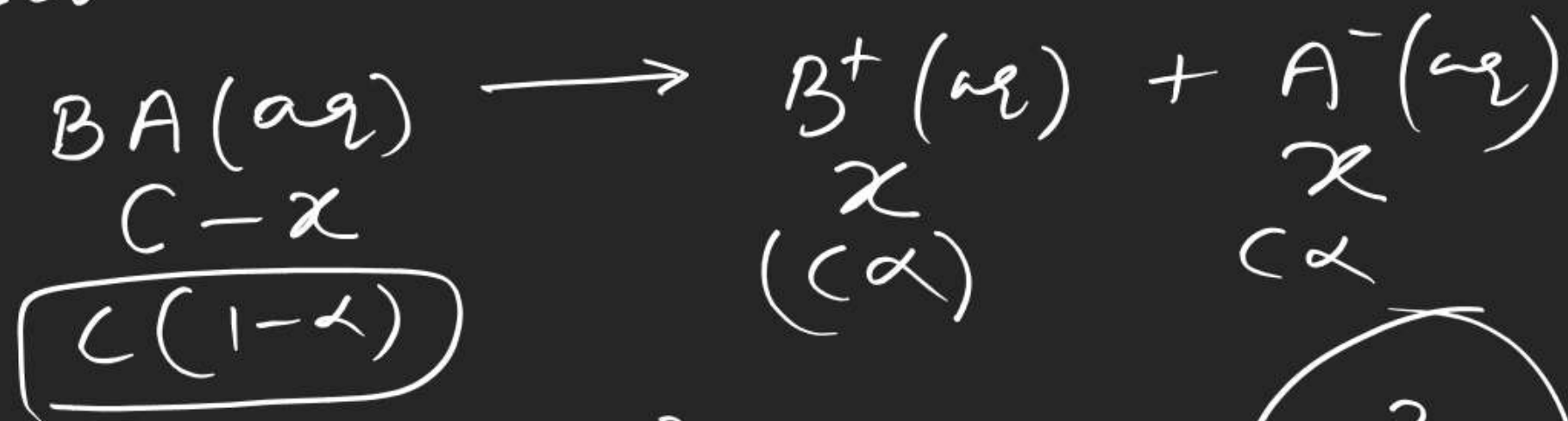
Ionic Eq/bm

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 eq^lbm betⁿ ionised and unionised solute.



$$\begin{aligned}
 \underline{K_{\text{diss}}} &= K_c = \frac{[\text{B}^+][\text{A}^-]}{[\text{BA}]} = \frac{\alpha^2}{C - \alpha} = \frac{C\alpha^2}{1 - \alpha} \\
 &= \frac{\cancel{C}\alpha \cdot C\alpha}{C(1 - \alpha)} = \frac{C\alpha^2}{1 - \alpha}
 \end{aligned}$$

α depends on

1) Nature of solute

2) Nature of solvent : — As dielectric constant \uparrow $\alpha \propto \uparrow$

3) Temperature : Since dissociation is an endothermic process therefore as $T \uparrow$ $\alpha \propto \uparrow$

④ Concentration

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

$$C \downarrow \alpha \uparrow$$

Ostwald dilution law : \rightarrow

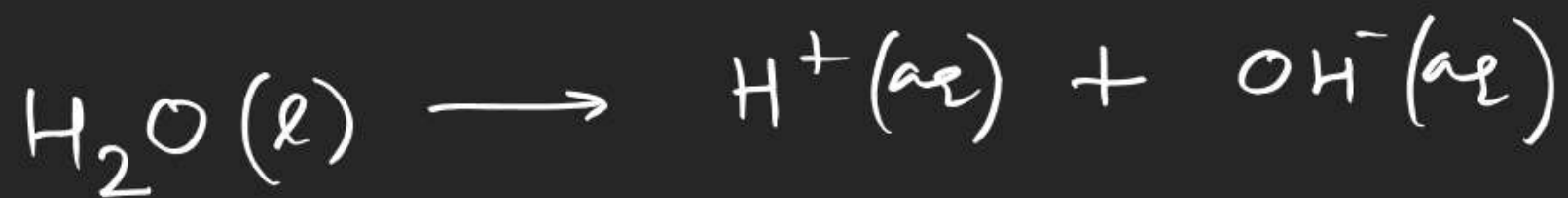
As concentration \downarrow $\alpha \propto \uparrow$

⑤ Common ion effect



- (A) $x = y$
 ✓ (B) $x > y$
 (C) $x < y$

Dissociation of H_2O



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

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

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

(ionic product
of H_2O)

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

At $25^\circ C$

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

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

$NH_3(liq)$



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

at 25°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°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°C

$$pH + pOH = 14$$

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

find (1) pH of pure H_2O

(2) Define a solution as
acidic, basic or neutral

having pH	(i)	5	Acidic
	(ii)	6	Basic
	(iii)	7	<u>Basic</u>

$$\log 2 = 0.3$$

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

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

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

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

Chemical eq 16^m

JEE-Adv

O-II

(29)

 $\frac{1}{K_p}$ 

$$\frac{1}{K_p} = \frac{P/4 \times \left(3P/4\right)^3}{P_{\text{NH}_3}^2}$$

$$P_{\text{NH}_3}^2 = K_p \times P^4 \times \frac{27}{2^8}$$

$$P_{\text{NH}_3} = K_p^{1/2} \times P^2 \times \frac{3^{3/2}}{2^4}$$

(35)

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