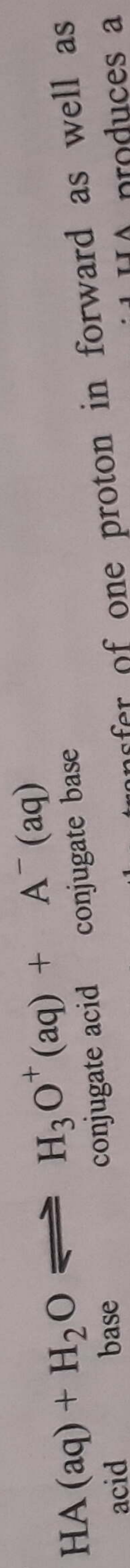
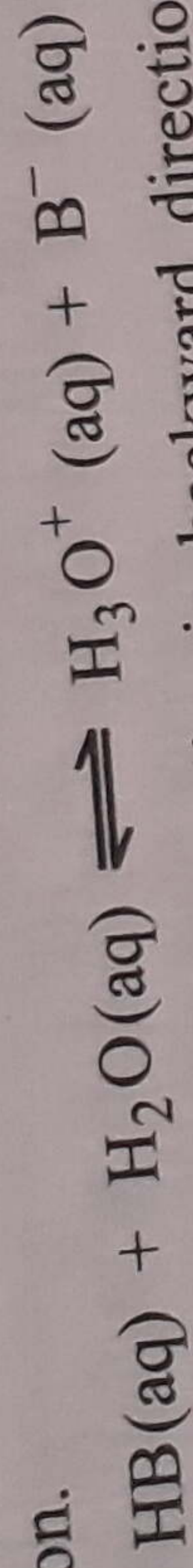


## DISSOCIATION / IONISATION OF ACIDS AND BASES

Arrhenius concept of acids and bases is most useful as most of the ionisations in chemical and biological systems occur in aqueous medium. Acids and bases dissociate in water, to give  $H^+$  and  $OH^-$  ions respectively. Acids as well as bases are of two types. **Strong acids** such as  $HCl$ ,  $HNO_3$ ,  $H_2SO_4$  and  $HClO_4$  are almost completely dissociated in water and produce large number of  $H^+$  ions. Acids like  $HF$ ,  $HCOOH$ ,  $CH_3COOH$ ,  $H_3PO_4$  etc dissociate only to a small extent and produce small number of  $H^+$  ions in solution. These are called **weak acids**. Bases like  $LiOH$ ,  $NaOH$ ,  $Ba(OH)_2$  etc are almost completely dissociated in water and produce large number of  $OH^-$  ions, and are called **strong bases**. Bases like  $NH_4OH$ ,  $Al(OH)_3$  etc dissociate to a small extent producing small number of  $OH^-$  ions in aqueous solution and are called **weak bases**. According to Bronsted-Lowry concept, strong acids are good proton donor and strong bases are good proton acceptor. Let us consider the dissociation of an acid HA in aqueous solution.



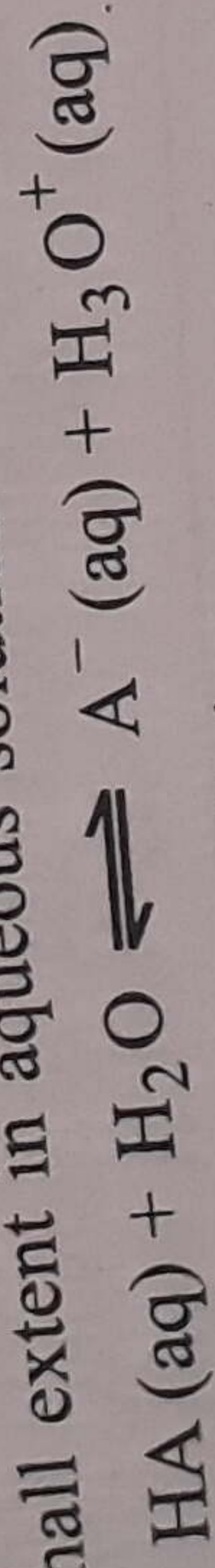
The above equilibrium is dynamic and involves the transfer of one proton in forward as well as reverse direction. If HA is a strong acid, equilibrium lies in forward direction. Strong acid HA produces a weak conjugate base  $A^-$  which will have little tendency to accept proton and shift the equilibrium in backward direction.



If HB is a weak acid, equilibrium lies in backward direction. Weak acid HB will produce a strong base  $B^-$  having strong tendency to accept a proton and shift the equilibrium in backward direction.

### Ionisation Constants of Weak Acids and Weak Bases

Strong acids and strong bases are almost completely dissociated in water. But weak acids and weak bases dissociate to a small extent in aqueous solution. Let us consider the dissociation of a weak acid (HA).



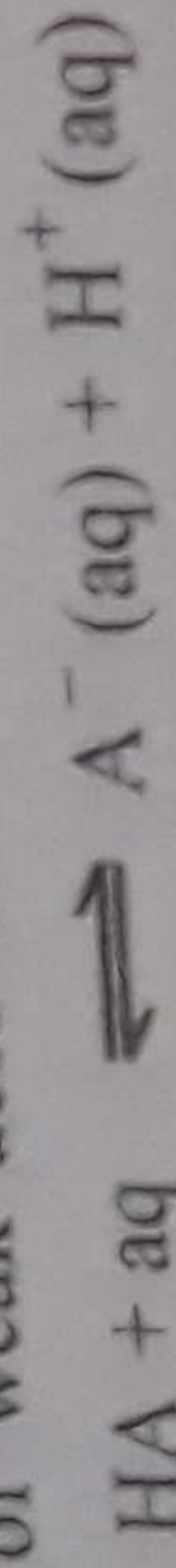
Applying law of chemical equilibrium

$$K = \frac{[A^-][H_3O^+]}{[HA][H_2O]}$$

As  $H_2O$  is solvent and its concentration remains almost constant (55.5 mol/L), let  $K[H_2O] = K_a$

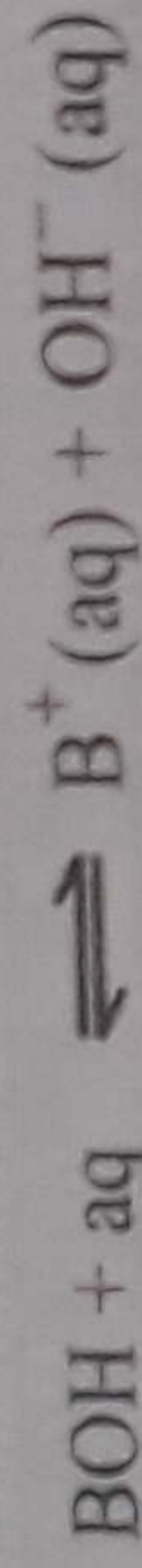
$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

$K_a$  is called **ionisation constant or dissociation constant of weak acid**. The chemical equation for the dissociation of weak acid and  $K_a$  may also be represented as :



$$K_a = \frac{[A^-][H^+]}{[HA]}$$

Similarly, dissociation of a weak base in water may be represented as :



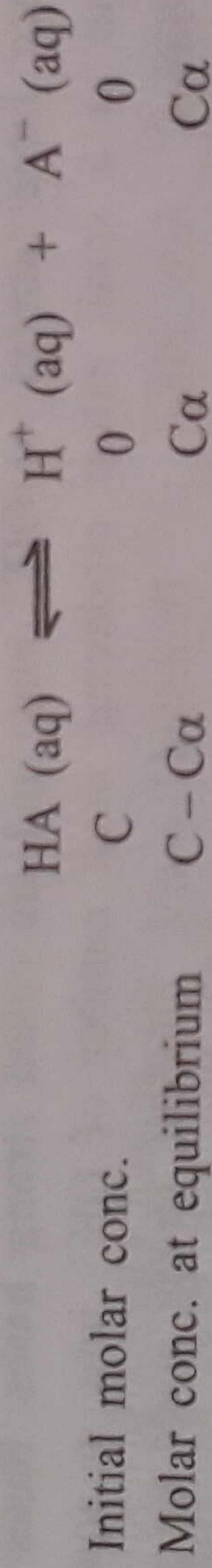
Applying law of chemical equilibrium,

$$K_b = \frac{[B^+][OH^-]}{[BOH]}$$

$K_b$  is called **dissociation constant or ionisation constant of the weak base**.

### Calculation of degree of dissociation ( $\alpha$ ) for the weak acid and weak base

Let the concentration of weak acid (HA) be  $C$  mol/lit and degree of dissociation at this concentration and constant temperature is  $\alpha$ . Moles of acid dissociated will be  $C\alpha$ . Moles of acid left at equilibrium will be  $(C - C\alpha)$ .



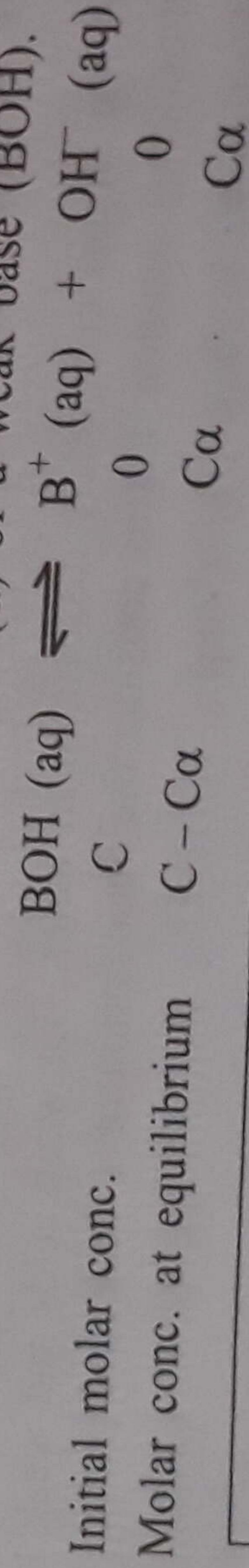
$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \text{or} \quad K_a = \frac{C\alpha \times C\alpha}{C - C\alpha} \quad \text{or} \quad K_a = \frac{C\alpha^2}{1 - \alpha}$$

As degree of dissociation ( $\alpha$ ) is very small,  $(1 - \alpha)$  may be taken equal to one.

$$\therefore K_a = C\alpha^2, \quad \alpha = \sqrt{\frac{K_a}{C}} \quad \text{and} \quad [H^+] = C\alpha \quad \text{or} \quad [H^+] = \sqrt{K_a \times C}$$

$K_a$  is constant at a given temperature.  $K_a$  is a dimensionless quantity with the understanding that the standard state concentration is 1 M.

Let us calculate the degree of dissociation ( $\alpha$ ) of a weak base (BOH).



$$K_b = \frac{[B^+][OH^-]}{[BOH]} \quad \text{or} \quad K_b = \frac{C\alpha \times C\alpha}{C - C\alpha} \quad \text{or} \quad K_b = \frac{C\alpha^2}{1 - \alpha}$$