

SALT OF A WEAK ACID AND A STRONG/WEAK BASE

Salt of a Weak acid and a Strong Base :

The solution of such a salt is basic in nature. The anion of the salt is reactive. It reacts with water to form a **weak acid** and OH^- ions.



Weak acid

Consider, for example, the salt CH_3COONa . It ionises in water completely to give CH_3COO^- and Na^+ ions. CH_3COO^- ions react with water to form a **weak acid**, CH_3COOH and OH^- ions.

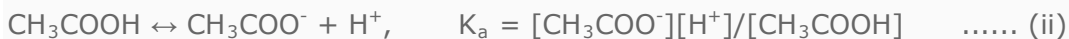


Thus, OH^- ion concentration increases, the solution becomes alkaline.

Applying law of mass action,

$$K_h = [\text{CH}_3\text{COOH}][\text{OH}^-]/[\text{CH}_3\text{CO}^-] = (\text{Cx} \times \text{Cx})/\text{C}(1-x) = (\text{Cx}^2)/(1-x) \dots\dots (i)$$

Other equations present in the solution are:



From eqs. (ii) and (iii),

$$\log [\text{OH}^-] = \log K_w - \log K_a + \log[\text{salt}]/[\text{acid}]$$

$$-\text{pOH} = -\text{p}K_w + \text{p}K_a + \log[\text{salt}]/[\text{acid}]$$

$$\text{p}K_w - \text{pOH} = \text{p}K_a + \log[\text{salt}]/[\text{acid}]$$

$$\text{pH} = \text{p}K_a + \log[\text{salt}]/[\text{acid}]$$

Considering eq. (i) again,

$$K_h = \text{cx}^2/(1-x) \quad \text{or} \quad K_h = \text{Ch}^2/(1-h)$$

When h is very small, $(1-h) \rightarrow 1$

$$\text{or } h^2 = K_h/C$$

$$\text{or } h = \sqrt{K_h/C}$$

$$[\text{OH}^-] = h \times C = \sqrt{CK_h} = \sqrt{C \cdot K_w/K_a}$$

$$[\text{H}^+] = K_w/[\text{OH}^-]$$

$$= K_w/\sqrt{C \cdot K_w/K_a} = \sqrt{(K_a \cdot K_w)/K_c}$$

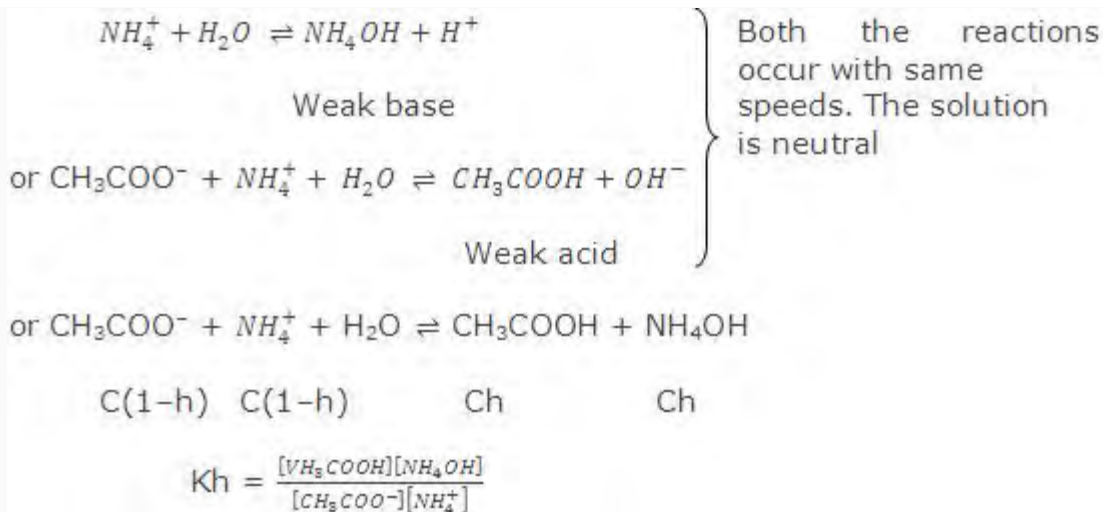
$$-\log [\text{H}^+] = -1/2 \log K_w - 1/2 \log K_a + 1/2 \log C$$

$$\text{pH} = 1/2 \text{p}K_w + 1/2 \text{p}K_a + 1/2 \log C$$

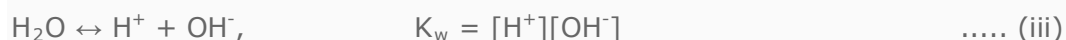
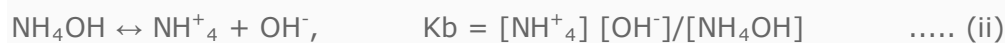
$$= 7 + 1/2 \text{p}K_a + 1/2 \log C.$$

Salt of a weak acid and a weak base:

Maximum **hydrolysis** occurs in the case of such a salt as both the cation and anion are reactive and react with water to produce H^+ and OH^- ions. The solution is generally neutral but it can be either slightly acidic or slightly alkaline if both the reactions take place with slightly different rates. Consider, for example, the salt $\text{CH}_3\text{COONH}_4$. It gives CH_3COO^- and NH_4^+ ions in solution. Both react with water.



Other equilibria which exist in solution are:



From Eqs. (i), (ii) and (iii),

$$K_h = K_w/K_a \cdot K_b = \frac{[CH_3COOH][NH_4OH]}{[CH_3COO^-][NH_4^+]} \dots (iv)$$

Let C be the concentration and h be the degree of **hydrolysis**

$$K_h = h^2/(1-h)^2$$

When h is small, (1-h) → 1.

$$K_h = h^2$$

$$h = \sqrt{K_h} = \sqrt{K_w/K_a \cdot K_b}$$

$$[H^+] K_a \times h$$

$$= K_a \times \sqrt{K_w/K_a \cdot K_b}$$

$$= \sqrt{K_w \cdot K_a/K_b}$$

$$-\log [H^+] = -1/2 \log K_a - 1/2 \log K_w + 1/2 \log K_b$$

$$pH = 1/2 pK_a + 1/2 pK_w - 1/2 pK_b$$

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

When $pK_a = pK_b$, $pH = 7$, i.e., solution will be neutral in nature.

When $pK_a > pK_b$. The solution will be alkaline as the acid will be slightly weaker than base and pH value will be more than 7. In case $pK_a < pK_b$, the solution will be acidic as the acid is relatively stronger than base and pH will be less than 7.

Source : <http://ciseche10.files.wordpress.com/2013/12/ionic-equilibrium.pdf>