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.

$A^- + H_2O; \leftrightarrow HA + OH^-$

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.

$CH_3COO^- + H_2O \leftrightarrow CH_3COOH + OH^-$

C(1-x) Cx Cx

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

Applying law of mass action,

$K_h = [CH_3COOH][OH^-]/[CH_3CO^-] = (Cx\times Cx)/C(1-x) = (Cx^2)/(1-x)$) ...... (i)

Other equations present in the solution are:

$CH_3COOH \leftrightarrow CH_3COO^- + H^+$, $K_a = [CH_3COO^-][H^+]/[CH_3COOH]$ ...... (ii)

$H_2O \leftrightarrow H^+ + OH^-$, $K_w = [H^+][OH^-]$ .... (iii)

From eqs. (ii) and (iii),

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

$-pOH = -pK_w + pK_a + \log[\text{salt}]/[\text{acid}]$

$pK_w - pOH = pK_a + \log[\text{salt}]/[\text{acid}]$

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

Considering eq. (i) again,

$K_h = cx^2/(1-x)$ or $K_h = Ch^2/(1-h)$

When h is very small, (1-h) → 1

or $h^2 = K_h/C$

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

$[OH^-] = h \times C = \sqrt{(CK_h)} = \sqrt{(C*K_w/K_a)}$

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

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

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

$pH = 1/2pK_w + 1/2pK_a + 1/2\log C$

$= 7 + 1/2pK_a + 1/2\log C$.

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 OFT 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 CH_3COONH_4. It gives CH_3COO^- and ions in solution. Both react with water.

$$NH_4^+ + H_2O \rightleftharpoons NH_4OH + H^+$$

Weak base

or $CH_3COO^- + NH_4^+ + H_2O \rightleftharpoons CH_3COOH + OH^-$

Weak acid

Both the reactions occur with same speeds. The solution is neutral

or $CH_3COO^- + NH_4^+ + H_2O \rightleftharpoons CH_3COOH + NH_4OH$

C(1–h)   C(1–h)   Ch   Ch

$$Kh = \frac{[VH_3COOH][NH_4OH]}{[CH_3COO^-][NH_4^+]}$$

Other equilibria which exist in solution are:

$CH_3COOH \leftrightarrow CH_3COO^- + H^+$, $K_a = [CH^3COO^-][H^+]/[CH_3COOH]$ ..... (i)

$NH_4OH \leftrightarrow NH^+{}_4 + OH^-$, $Kb = [NH^+{}_4]\,[OH^-]/[NH_4OH]$ ..... (ii)

$H_2O \leftrightarrow H^+ + OH^-$, $K_w = [H^+][OH^-]$ ..... (iii)

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

$K_h = Kw/K_a.K_b = [CH_3COOH][NH_4OH]/[CH_3COO^-][NH^+{}_4]$ .... (iv)

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

$K_h = h^2/(1\text{-}h)^2$

When h is small, $(1\text{-}h) \rightarrow 1$.

$K_h = h^2$

$h = \sqrt{Kh} = \sqrt{K_w/K_{a*}K_b}$

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

$= K_a \times \sqrt{K_w/K_{a*}K_b}$

$= \sqrt{K_{w\,*}\,K_a/K_b}$

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

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

$= 7 + 1/2pK_a - 1/2pK_b$

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