Chem12 K_a and K_b : Notes-140

If we have a **weak acid**, only a small percentage of the acid is ionized. Therefore the equilibrium reaction is :

$$HA + H_2O \leftrightarrow H_3O^+ + A^-$$

The equilibrium constant for this reaction is K_a. The equilibrium expression is :

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

Example : Find the $[H_3O^+]$ for a 0.25 M solution of H_2S.

The equilibrium is : $H_2S + H_2O \leftrightarrow H_3O^+ + HS^-$; $K_a = 9.1\times10^{-8}$

We have : $\{[H_3O^+][HS^-]\}/[H_2S] = X^2/0.25 = 9.1\times10^{-8}$. Since $[H_3O^+] = [HS^-]$, we have : $[H_3O^+] = X = 1.5\times10^{-4}$ M.

This type of calculation is valid as long as the acid is weak.

If we have a **weak base**, the equilibrium reaction is :

$$B^- + H_2O \leftrightarrow OH^- + HB$$

The equilibrium constant is K_b and the equilibrium expression is :

$$\{[OH^-][HB]\}/[B^-] = K_b$$

Example : Find the $[OH^-]$ for a 0.45 solution of NH_3. $K_b = 1.8\times10^{-5}$.

$$NH_3 + H_2O \leftrightarrow OH^- + NH_4^+$$

The equilibrium expression is :

$$\{[NH_4^+][OH^-]\}/NH_3 = X^2/0.45 = 1.8\times10^{-5}$$

$$[OH^-] = X = 2.8\times10^{-3} \text{ M.}$$

Important : The equilibrium expression for NH_4^+, the conjugate acid of NH_3 is :

$$\{[H_3O^+][NH_3]\}/[NH_4^+] = K_a$$

Note : $K_b \times K_a = \{[NH_4^+][OH^-]\}/NH_3 \times \{[H_3O^+][NH_3]\}/[NH_4^+] = [OH^-][H_3O^+]$

In general, for conjugate acid-base pairs : $\mathbf{K_a \times K_b = 10^{-14}}$.