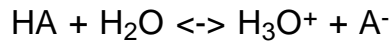


Chem12 K_a and K_b : Notes-140

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



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

$$\frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = K_a$$

Example : Find the $[\text{H}_3\text{O}^+]$ for a 0.25 M solution of H_2S .

The equilibrium is : $\text{H}_2\text{S} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HS}^-$; $K_a = 9.1 \times 10^{-8}$

We have : $\frac{[\text{H}_3\text{O}^+][\text{HS}^-]}{[\text{H}_2\text{S}]} = \frac{X^2}{0.25} = 9.1 \times 10^{-8}$. Since $[\text{H}_3\text{O}^+] = [\text{HS}^-]$, we have : $[\text{H}_3\text{O}^+] = X = 1.5 \times 10^{-4}$ M.

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

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



The equilibrium constant is K_b and the equilibrium expression is :

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

Example : Find the $[\text{OH}^-]$ for a 0.45 solution of NH_3 . $K_b = 1.8 \times 10^{-5}$.



The equilibrium expression is :

$$\frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{X^2}{0.45} = 1.8 \times 10^{-5}$$

$$[\text{OH}^-] = X = 2.8 \times 10^{-3} \text{ M.}$$

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

$$\frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]} = K_a$$

Note : $K_b \times K_a = \frac{[\text{NH}_4^+][\text{OH}^-]}{\text{NH}_3} \times \frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]} = [\text{OH}^-][\text{H}_3\text{O}^+]$

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