$$K = \frac{[P]^c\,[Q]^d}{[M]^a\,[N]^b}$$

Where the letters in brackets represent the molar concentration of dissolved solutes or partial pressures (in atmospheres) if the reacting substances are gases.

**Common Types of Equilibrium – Constant Expressions :**

**1-dissociation of water**

Aqueous solution always contain small amounts of hydronium and hydroxide ion as a consequence of the dissociation reaction

$$2H_2O \longrightarrow H_3O^+ + OH^-$$

The equilibrium constant for this reaction

$$K = \frac{[H_3O^+]\,[OH^-]}{[H_2O]^2}$$

$[H_2O]$ in the equation can be taken as constant

$$K\,[H_2O]^2 = K_W = [H_3O^+]\,[OH^-]$$

Where the new constant $K_W$ is given the special name, the ion – product constant for water ( $\approx 1X\ 10^{-14}$ at 25C°. )

$$pK_W = -\log K_W$$

$$= -\log(1X\ 10^{-14})$$

$$= -(-14) \longrightarrow = 14$$

$$\log K_W = \log[H^+] + \log[OH^-]$$

$$-\log K_W = (-\log[H^+]) + (-\log[OH^-])$$

$$pK_W = pH + pOH$$

$$14 = pH + pOH$$

Example 1 :

What is the pH of a 0.0002 M HCl solution ?

$HCl \longrightarrow H^+ + Cl^-$

0.0002 0.0002 0.0002

$[H^+] = 0.0002\ M = 2 \times 10^{-4}$

$pH = -\log [H^+]$

$= -\log ( 2 \times 10^{-4} )$

$= -(\log 2 + \log 10^{-4})$

$= -\{ 0.3 + ( -4) \}$

$= -( -3.7 )$

$= 3.7$

Example 2 :

What is the pH of a $5 \times 10^{-3}$ M NaOH solution ?

Method 1:

$NaOH \longrightarrow Na^+ + OH^-$

$5 \times 10^{-3}$ $5 \times 10^{-3}$ $5 \times 10^{-3}$

$[OH^-] = 0.005\ M = 5 \times 10^{-3}$

$pOH = -\log [OH^-]$

$= -\log ( 5 \times 10^{-3} )$

$= -(\log 5 + \log 10^{-3})$

$= -\{ 0.7 + (-3) \}$

$= -(-2.3)$

$= 2.3$

$pK_W = pH + pOH$

$14 = pH + pOH$

$14 = pH + 2.3$

$pH = 14 - 2.3$

$pH = 11.7$

Method 1:

$K_W = [H^+] [OH^-]$

$1 X 10^{-14} = [H^+] X 5 X 10^{-3}$

$[H^+] = \frac{1X 10-14}{5 X 10-3}$

$= 0.2 X 10^{-11} \longrightarrow = 2 X 10^{-12}$

$pH = -\log [H^+]$

$= -\log ( 2 \times 10^{-12} )$

$= -(\log 2 + \log 10^{-12})$

$= -\{ 0.3 + (-12) \}$

$= -(-11.7) = 11.7$

2 – **Equilibrium involving slightly ionic solids :**

When an aqueous solution is saturated with a sparingly soluble salt ( e.g. silver chloride ) one or more equilibrium will be established :

$$AgCl_{(S)} \rightleftharpoons AgCl_{(aq)}$$

$$AgCl_{(aq)} \rightleftharpoons Ag^{+} + Cl^{-}$$

$$K = \frac{[AgCl]_{aq}}{[AgCl]_{s}}$$

$$K\,[AgCl]_{s} = K_1 = [AgCl]_{aq}$$

The concentration of AgCl in a saturated solution is constant at any given temperature :

$$K_2 = \frac{[Ag^{+}]\,[Cl^{-}]}{[AgCl]_{aq}}$$

$$K_1 K_2 = K_{SP} = [Ag^{+}]\,[Cl^{-}]$$

The product of these two constant, $K_{SP}$ is called the solubility – product constant.