Chemistry Study Materials for Class 11
(NCERT Based Notes of Chapter- 07)
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Equilibrium

Reversible and irreversible reactions

A reaction that takes place in only one direction is called an ***irreversible reaction***.

e.g. Reaction between NaOH and HCl $NaOH + HCl \rightarrow NaCl + H_2O$

Reversible reactions are those which take place in both directions. i.e. here reactants combined to form products and the products recombine to form reactants.

E.g. Haber process for the preparation of ammonia

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

The process by which reactants are converted to products is called *forward reaction* and the process by which products recombine to form reactants is called *backward reaction*. After sometimes, the rate of forward reaction becomes equal to the rate of backward reaction and the reaction attains equilibrium. Thus *equilibrium is a state in which the rates of forward and backward reactions are equal*.

Equilibrium is dynamic in nature. i.e. at equilibrium the reaction does not stop. The reactant molecules collide to form products and the product molecules collide to form the reactants and the rates of these reactions are equal.

Physical equilibrium:

Equilibrium involving physical process is called Physical equilibrium E.g. melting of ice, evaporation of water, dissolution of solids or gases in liquids, sublimation

For example when ice and water are kept in a perfectly insulated thermo flask at 273K and 1 atm pressure, there exist equilibrium between ice and water. At this stage, there is no change in the mass of ice and water. i.e. the rate of transfer of molecules from ice to water and the reverse process are equal.

$$H_2O(s) \rightleftharpoons H_2O\ (l)$$

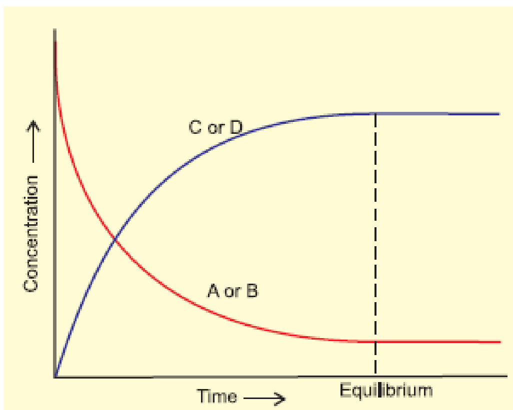
Chemical Equilibrium

Equilibrium associated with chemical reactions is called chemical equilibrium.

At equilibrium, the concentrations of reactants and products are constant.

Consider a hypothetical reaction, $A + B \rightleftharpoons C + D$

As the reaction proceeds, the concentration of the reactants decreases and that of the products increases. After sometimes, the two reactions occur at the same rates and an equilibrium state is reached. This can be illustrated by the following graph.



After the equilibrium is attained, the concentration of the reactants and products become constant.

At equilibrium,

the rate of forward reaction (r_f) = the rate of backward reaction (r_b).

Law of Chemical Equilibrium and Equilibrium Constant

This law was proposed by Goldberg and Waage. It states that *at constant temperature, the product of concentration of the products to that of the reactants, in which each concentration terms is raised to a power which is equal to the stoichiometric coefficients in the balanced chemical equation, has a constant value*. For a general reversible reaction:

$$A + B \rightleftharpoons C + D$$

According to the equilibrium law, $\frac{[C][D]}{[A][B]} = Kc$

Where Kc is called the equilibrium constant

For a general reaction, aA + bB $\rightleftharpoons$ cC + dD,

$$Kc = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

For the reaction; $H_2 + I_2 \rightleftharpoons 2HI$; $\quad Kc = \frac{[HI]^2}{[H_2][I_2]}$

For the reaction; $N_2 + 3H_2 \rightleftharpoons 2NH_3$; $\quad Kc = \frac{[NH_3]^2}{[N_2][H_2]^3}$

Equilibrium constant for the reverse reaction is the inverse of that for the forward reaction. i.e. if the equilibrium constant for the reaction; $H_2 + I_2 \rightleftharpoons 2HI$ is Kc,

Then that for the reverse reaction; $2HI \rightleftharpoons H_2 + I_2$ is 1/Kc.

Equilibrium constant for gaseous reactions

For a reaction involving gases, the concentration terms are replaced by partial pressures.

For example, $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$; $Kp = \frac{P^2_{HI}}{P_{H_2}.P_{I_2}}$

Where Kp is called equilibrium constant in terms of partial pressure, P_{HI}, P_H2 and P_I2 are the partial pressures of HI, H_2 and I_2 respectively.

Relation between Kc and Kp

Consider a general reaction, aA + bB $\rightleftharpoons$ cC + dD

The equilibrium constant in terms of concentration for this reaction is

$$Kc = \frac{[C]^c[D]^d}{[A]^a[B]^b} \quad \ldots\ldots\ldots\ldots (1)$$

And the equilibrium constant in terms of partial pressures is

$$Kp = \frac{P_C^{\,c}.P_D^{\,d}}{P_A^{\,a}.P_B^{\,b}} \quad \ldots\ldots\ldots\ldots (2)$$

From ideal gas equation, PV = nRT, P = nRT/V = CRT (since n/V = C, the concentration)

Therefore, $P_A = C_ART$, $P_B = C_BRT$, $P_C = C_CRT$ and $P_D = C_DRT$

Substitute these values in equation (2), we get;

$$K_p = \frac{[C_C RT]^c . [C_D RT]^d}{[C_A RT]^a [C_B RT]^b}$$

$$\text{Or, } K_p = \frac{C_C^{\,c} . C_D^{\,d}}{C_A^{\,a} . C_B^{\,b}} \quad \frac{(RT)^{c+d}}{(RT)^{a+b}}$$

$$\text{Or, } K_p = \frac{C_C^{\,c} . C_D^{\,d} . [RT]^{(c+d)-(a+b)}}{C_A^{\,a} . C_B^{\,b}}$$

$$\text{Or, } K_p = K_c . (RT)^{\Delta n},$$

$$\text{where } K_c = \frac{C_C^{\,c} . C_D^{\,d}}{C_A^{\,a} . C_B^{\,b}}$$

And Δn is the change in no. of moles of gaseous species. i.e. Δn = no. of moles of gaseous products – no. of moles of gaseous reactants.

Special cases:

i) If $\Delta n = 0$, then $K_p = K_c$

ii) If $\Delta n > 0$, then $K_p > K_c$ and

iii) If $\Delta n < 0$, then $K_p < K_c$

Homogeneous and heterogeneous equilibria

An equilibrium reaction in which all the reactants and products are in the same phase is called homogeneous equilibrium.

e.g. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

Equilibrium reaction in which the reactants and products are in different phases is called heterogeneous equilibrium.

e.g. $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

$H_2O\,(s) \rightleftharpoons H_2O_{(l)}$

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