

10+2 PCM NOTES

BY

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(PDF version handwritten notes of Maths, Physics and Chemistry for 10+2 competitive exams like JEE Main, WBJEE, NEST, IISER Entrance Exam, CUCET, AIPMT, JIPMER, EAMCET etc.)



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With best wishes from Joyoshish Saha

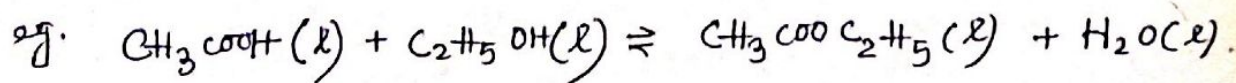
- * In chemical equilibrium rate of forward reaction and rate of backward reaction are same. The process of production of reactants and products never ceases.
- * Equilibrium is possible only in closed system. All measurable properties of system remain constant over time. When equilibrium is attained, it is characterised by a constant value of one of its parameters. Equilibrium can be understood by kinetic and thermodynamical theory. Such as, if 'gibbs function' of product is very less than reactant then forward reaction rate increases.

- * General eqn - reversible $mA + nB \rightleftharpoons xC + yD$.
reaction quotient, $Q = \frac{[C]^x [D]^y}{[A]^m [B]^n}$ [] → molar concentration

If equilibrium achieved then $Q = K \rightarrow$ equilibrium
constant.

- * Static & Dynamic Equilibrium:

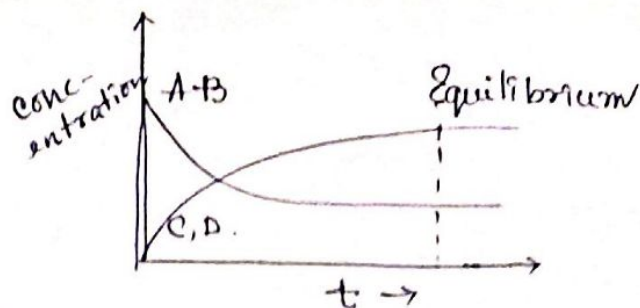
- * Reversible & Irreversible Reactions: If under certain conditions, the forward & backward reaction occur simultaneously, then the reaction is said to be reversible.



Chemical equilibrium is a dynamic equilibrium. Concentrations are constant.



- * In equilibrium when T, P and Conc. are constant.



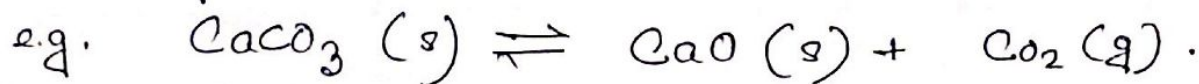
- * Characteristics of Equilibrium: i) Equilibrium is dynamic. ii) It can be achieved from both direction. iii) Can be attained in a closed system. iv) Catalyst can't change the equilibrium state.

- * Equilibrium in Physical Process - Phase Equilibrium: a) liquid - vapour equilibrium, b) Solid - Vapour Equilibrium (Sublimation), c) Solid - Liquid Equilibrium, d) Solute - Solution Equilibrium.

- * Homogeneous Chemical Equilibrium Systems:

- a) Gas-phase equilibrium system
b) Liquid-phase homogeneous system.
c) Solid-phase homogeneous system.

- * Heterogeneous Chemical Equilibrium Systems:



- * Equilibrium Constants:

- i) Concentration Eq. Constant (K_c).
ii) Pressure Eq. Constant (K_p).

[Relation of K_p & $K_c \rightarrow K_p = K_c (RT)^{\Delta n_g}$].

- * Characteristics of Eq. Constant:

- i) The expression of eq. constant depends upon the manner in which the chemical eqn is written.

for reversible reactions,

$$K_{\text{forward}} \times K_{\text{backward}} = 1.$$

ii) The eq. constant tells about the position of equilibrium; how much product can be obtained.

* Le Chatelier Principle: A system in equilibrium, if disturbed by a change in concentration, pressure, temperature, a net change occurs in it in a direction that tends to decrease the disturbing factor.

* Factors affecting equilibrium:

a) Change in Conc., b) Change in Pressure
c) Change in temperature. (for exothermic reaction \rightarrow increase in $T \Rightarrow$ shifting to backward reaction..)

* Application of Le Chatelier Principle:

i) Favourable conditions for synthesis of Ammonia.

ii) Formation of SO_3 . iii) Formation of NO .

* Effect of temperature on Equilibrium —

Vant Hoff Equation: $\frac{d}{dT} (\ln K) = \frac{\Delta H^\circ}{RT^2}$.

$$\frac{d(\ln K)}{d\left(\frac{1}{T}\right)} = -\frac{\Delta H^\circ}{R} \quad \left| \quad \ln\left(\frac{K_2}{K_1}\right) = -\frac{\Delta H^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right) \right.$$

$\Delta H^\circ \rightarrow$ standard heat of reaction at const. pressure.

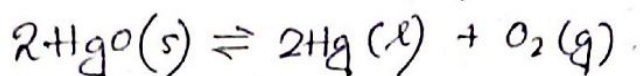
$$2.303 \log\left(\frac{K_2}{K_1}\right) = -\frac{\Delta H^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

* $\Delta G = \Delta G^\circ + RT \ln Q$. [$Q \rightarrow$ reaction quotient]

at equilibrium. $\Delta G = 0$, $\Delta G^\circ = -RT \ln K$.

* Henry's Law: $V \propto p$. ($V \rightarrow$ volume of gas in the solvent).

* Heterogeneous reaction in which three states of matter are present:



* Besides K_p & K_c , a mole fraction equilibrium constant is there (K_x).

$$K_p = K_x \times (P)^{\Delta n}$$

$$K_c = K_x \left(\frac{P}{RT} \right)^{\Delta n}$$

* Relation between K_p , K_c , K_x .

$$K_p = K_c (RT)^{\Delta n} = K_x (P)^{\Delta n}$$

* If a reaction is represented by two or more simple reactions then the equilibrium constant is the product of the eq. constants of primary reactions.

$$K = K_1 \times K_2 \times \dots$$

* Reaction isotherm. $\Delta G = \Delta G^\circ + RT \ln Q$.

* Degree of dissociation: $= -RT \ln K + RT \ln Q$

$$\alpha = \frac{d_t - d_o}{(n-1)d_o}$$

$$\Delta G = RT \ln \frac{Q}{K}$$

$d_t \rightarrow$ theoretical density. } both under same
 $d_o \rightarrow$ experimental " } pressure.
 $n \rightarrow$ one molecule dissociates into n molecules.

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Chemical Equilibrium

3.

c.

If all the d's are taken as vapour densities, then $d_t = \frac{\text{molecular wt.}}{2}$.

* If a chemical reaction is multiplied by a factor, its eq. constant must be raised to a power equal to that factor in order to obtain the eq. constant for the new reaction.

$$K' = (K)^n \quad [n \in \mathbb{I}]$$

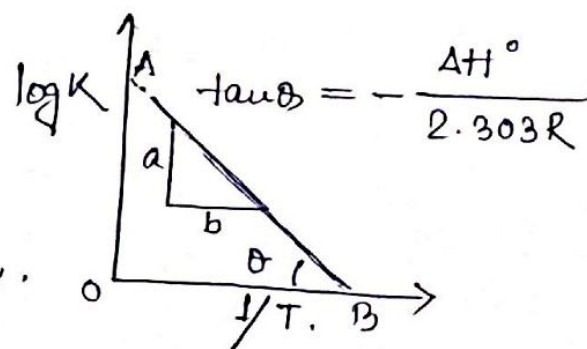
† Eq. constant depends only on Temperature.

$$\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ = -2.303 RT \log K$$

$$\log K = - \frac{\Delta H^\circ}{2.303 R} \cdot \frac{1}{T} + \frac{\Delta S^\circ}{2.303 R}$$

$$AO = \frac{\Delta S^\circ}{2.303 R}$$

ΔH° & ΔS° can be calculated - this way.



$$\Delta H^\circ = -2.303 R \times \tan \theta.$$

$$\Delta S^\circ = 2.303 R \times AO \text{ (intercept at y axis)}.$$

* Volume percentage = mole percent

$$= \frac{\text{no. of moles}}{\text{total moles}} \times 100$$