

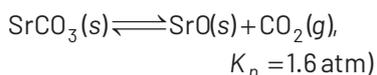
# 8

## Chemical Equilibrium

### TOPIC 1

#### The Chemical Equilibrium, Basic Law of Mass Action and Equilibrium Constant

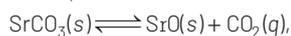
**01** A 20 litre container at 400 K contains  $\text{CO}_2(\text{g})$  at pressure 0.4 atm and an excess of  $\text{SrO}$  (neglect the volume of solid  $\text{SrO}$ ). The volume of the container is now decreased by moving the movable piston fitted in the container. The maximum volume of the container, when pressure of  $\text{CO}_2$  attains its maximum value, will be [NEET 2017] (Given that :



(a) 5 L (b) 10 L (c) 4 L (d) 2 L

**Ans. (a)**

For the reaction,



$K_p = 1.6 \text{ atm} = p_{\text{CO}_2} = \text{maximum pressure of CO}_2$

Given,  $p_1 = 0.4 \text{ atm}$ ,  $V_1 = 20 \text{ L}$ ,  $T_1 = 400 \text{ K}$

$p_2 = 1.6 \text{ atm}$ ,  $V_2 = ?$ ,  $T_2 = 400 \text{ K}$

At constant temperature,  $p_1 V_1 = p_2 V_2$

$$0.4 \times 20 = 1.6 \times V_2$$

$$V_2 = \frac{0.4 \times 20}{1.6} = 5 \text{ L}$$

**02** For a given exothermic reaction,  $K_p$  and  $K'_p$  are the equilibrium constants at temperatures  $T_1$  and  $T_2$ , respectively. Assuming that heat of reaction is constant in temperature range between  $T_1$  and  $T_2$ , it is readily observed that [CBSE AIPMT 2014]

- (a)  $K_p > K'_p$  (b)  $K_p < K'_p$   
 (c)  $K_p = K'_p$  (d)  $K_p = \frac{1}{K'_p}$

**Ans. (a)**

The equilibrium constant at two different temperatures for a thermodynamic process is given by

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

Here,  $K_1$  and  $K_2$  are replaced by  $K_p$  and  $K'_p$ .

$$\text{Therefore, } \log \frac{K'_p}{K_p} = \frac{\Delta H^\circ}{2.303R} \left[ \frac{1}{T_1} - \frac{1}{T_2} \right]$$

For exothermic reaction,

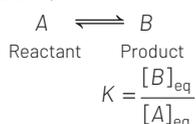
$$T_2 > T_1 \text{ and } H = -ve \\ \Rightarrow K_p > K'_p$$

**03** If the value of an equilibrium constant for a particular reaction is  $1.6 \times 10^{12}$ , then at equilibrium the system will contain

- (a) all reactants [CBSE AIPMT 2015]  
 (b) mostly reactants  
 (c) mostly products  
 (d) similar amounts of reactants and products

**Ans. (c)**

For a reaction,



$$1.6 \times 10^{12} = \frac{[B]_{\text{eq}}}{[A]_{\text{eq}}}$$

$$\therefore [B]_{\text{eq}} \gg [A]_{\text{eq}}$$

So, mostly the product will be present in the equilibrium mixture.

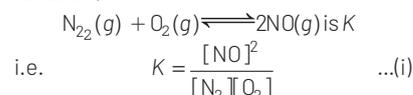
**04** If the equilibrium constant for  $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$  is  $K$ , the equilibrium constant for  $\frac{1}{2}\text{N}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightleftharpoons \text{NO}(\text{g})$  will be, [CBSE AIPMT 2015]

- (a)  $K^{1/2}$  (b)  $\frac{1}{2}K$  (c)  $K$  (d)  $K^2$

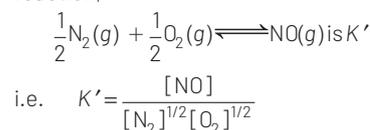
**Ans. (a)**

**Plan** As we can see the reaction for which we have to find out equilibrium constant is different only in stoichiometric coefficient as compared to the given reaction. Hence, we can find equilibrium constant for the required reaction with the help of mentioned equilibrium constant in the problem.

Given, equilibrium constant for the reaction,



Let equilibrium constant for the reaction,



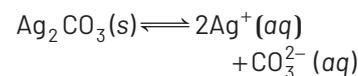
On squaring both sides

$$K'^2 = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} \quad \dots(\text{ii})$$

On comparing Eqs. (i) and (ii), we get

$$K = K'^2 \\ \text{or } K' = \sqrt{K}$$

**05** Using the Gibbs energy change,  $\Delta G^\circ = +63.3 \text{ kJ}$  for the following reaction,



the  $K_{\text{sp}}$  of  $\text{Ag}_2\text{CO}_3(\text{s})$  in water at  $25^\circ\text{C}$  is

( $R = 8.314 \text{ JK}^{-1}\text{mol}^{-1}$ ) [CBSE AIPMT 2014]

- (a)  $3.2 \times 10^{-26}$  (b)  $8.0 \times 10^{-12}$   
 (c)  $2.9 \times 10^{-3}$  (d)  $7.9 \times 10^{-2}$

**Ans. (b)**

$\Delta G^\circ$  is related to  $K_{\text{sp}}$  by the equation,  $\Delta G^\circ = -2.303RT \log K_{\text{sp}}$

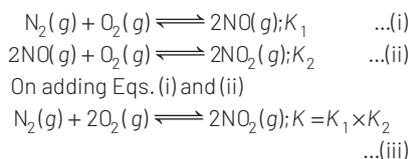
Given,  $\Delta G^\circ = +63.3 \text{ kJ}$   
 $= 63.3 \times 10^3 \text{ J}$   
 Thus, substitute  $\Delta G^\circ = 63.3 \times 10^3 \text{ J}$ ,  
 $R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$  and  $T = 298 \text{ K}$  [25 + 273 K] from the above equation we get,  
 $63.3 \times 10^3 = -2.303 \times 8.314 \times 298 \log K_{sp}$   
 $\therefore \log K_{sp} = -11.09$   
 $\Rightarrow K_{sp} = \text{antilog}(-11.09)$   
 $K_{sp} = 8.0 \times 10^{-12}$

**06** For the reaction,  
 $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$ , the  
 equilibrium constant is  $K_1$ . The  
 equilibrium constant is  
 $K_2$  for the reaction,  
 $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ . What is  
 K for the reaction,  
 $\text{NO}_2(\text{g}) \rightleftharpoons \frac{1}{2}\text{N}_2(\text{g}) + \text{O}_2(\text{g})$ ?

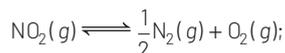
[CBSE AIPMT 2011]

- (a)  $1/(4K_1K_2)$  (b)  $[1/K_1K_2]^{1/2}$   
 (c)  $1/(K_1K_2)$  (d)  $1/(2K_1K_2)$

**Ans. (b)**



On dividing (iii) by  $\frac{1}{2}$  and on reversing we  
 get,



So,  $K = \frac{(\text{N}_2)^{1/2}(\text{O}_2)}{(\text{NO}_2)}$

$\therefore K = \left[ \frac{1}{K_1K_2} \right]^{1/2}$

**07** In which of the following  
 equilibrium  $K_c$  and  $K_p$  are not  
 equal? [CBSE AIPMT 2010]

- (a)  $2\text{NO}(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + \text{O}_2(\text{g})$   
 (b)  $\text{SO}_2(\text{g}) + \text{NO}_2(\text{g}) \rightleftharpoons \text{SO}_3(\text{g}) + \text{NO}(\text{g})$   
 (c)  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$   
 (d)  $2\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{CO}_2(\text{g})$

**Ans. (d)**

**Key Idea** The reaction for which the  
 number of moles of gaseous products ( $n_p$ )  
 is not equal to the number of moles of  
 gaseous reactants ( $n_R$ ), has different  
 value of  $K_c$  and  $K_p$ .

From the equation,  $K_p = K_c \times (RT)^{\Delta n_g}$

where,  $[\Delta n_g \text{ gaseous} = n_p - n_R]$

- (a)  $n_p = n_R = 2$ , thus,  $K_p = K_c$   
 (b)  $n_p = n_R = 2$ , thus,  $K_p = K_c$   
 (c)  $n_p = n_R = 2$ , thus,  $K_p = K_c$   
 (d)  $n_p = 2$ ,  $n_R = 1$ , thus,  $K_p \neq K_c$

**08** The dissociation constants for  
 acetic acid and HCN at  $25^\circ\text{C}$  are  
 $1.5 \times 10^{-5}$  and  $4.5 \times 10^{-10}$ ,

respectively. The equilibrium  
 constant for the equilibrium,

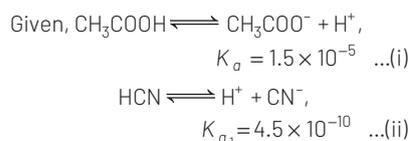


+  $\text{CH}_3\text{COO}^-$  would be

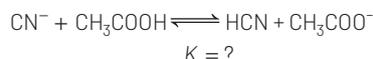
[CBSE AIPMT 2009]

- (a)  $3.0 \times 10^5$  (b)  $3.0 \times 10^{-5}$   
 (c)  $3.0 \times 10^{-4}$  (d)  $3.0 \times 10^4$

**Ans. (d)**



For



On subtracting Eq. (ii) from Eq. (i), we get



$$K = \frac{K_a}{K_{a_1}} = \frac{1.5 \times 10^{-5}}{4.5 \times 10^{-10}}$$

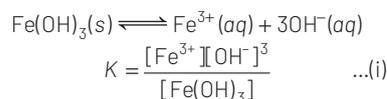
$$= \frac{10^5}{3} = 3.33 \times 10^4 \approx 3 \times 10^4$$

**09** If the concentration of  $\text{OH}^-$  ions in  
 the reaction,  
 $\text{Fe}(\text{OH})_3(\text{s}) \rightleftharpoons \text{Fe}^{3+}(\text{aq})$   
 $+ 3\text{OH}^-(\text{aq})$

is decreased by  $1/4$  times, then  
 equilibrium concentration of  $\text{Fe}^{3+}$   
 will increase by [CBSE AIPMT 2008]

- (a) 8 times (b) 16 times  
 (c) 64 times (d) 4 times

**Ans. (c)**



To maintain equilibrium constant, let the  
 concentration of  $\text{Fe}^{3+}$  is increased  $x$   
 times, on decreasing the concentration  
 of  $\text{OH}^-$  by  $\frac{1}{4}$  times

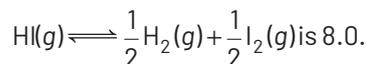
$$K = \frac{[x\text{Fe}^{3+}][\frac{1}{4}\text{OH}^-]^3}{[\text{Fe}(\text{OH})_3]} \quad \dots(\text{ii})$$

By dividing eq. (ii) by (i) we get

$$\frac{1}{64} \times x = 1$$

$\Rightarrow x = 64$  times

**10** The value of equilibrium constant of  
 the reaction,



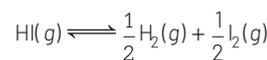
The equilibrium constant of the  
 reaction,



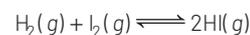
[CBSE AIPMT 2008]

- (a)  $\frac{1}{16}$  (b)  $\frac{1}{64}$  (c) 16 (d)  $\frac{1}{8}$

**Ans. (b)**



$$K = \frac{[\text{I}_2]^{1/2}[\text{H}_2]^{1/2}}{[\text{HI}]} \quad \dots(\text{i})$$



$$K' = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \quad \dots(\text{ii})$$

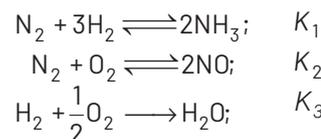
From Eqs. (i) and (ii)

$$K \times \sqrt{K'} = 1$$

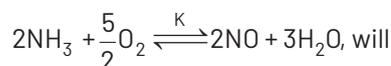
$$K' = \frac{1}{K^2} = \frac{1}{(8)^2} = \frac{1}{64}$$

**11** The equilibrium constants of the  
 following are

[NEET 2017, CBSE AIPMT 2007]



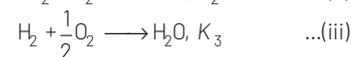
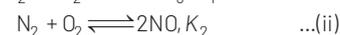
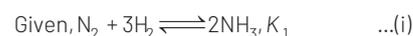
The equilibrium constant (K) of the  
 reaction



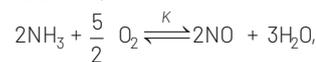
be

- (a)  $K_1K_3^3/K_2$  (b)  $K_2K_3^3/K_1$   
 (c)  $K_2K_3/K_1$  (d)  $K_2^3K_3/K_1$

**Ans. (b)**

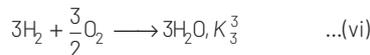
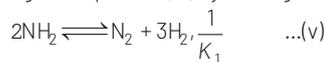


To calculate,

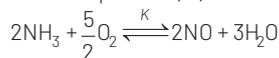


$$K = ? \quad \dots(\text{iv})$$

On reversing the equation (i) and multiplying the equation (iii) by 3, we get

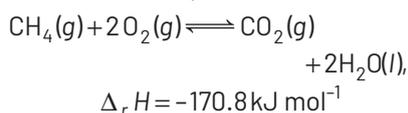


Now, add equation. (ii), (v) and (vi), we get the resultant equation. (iv).



$$\therefore K = \frac{K_2 K_1^3}{K_1}$$

**12** For the reaction,

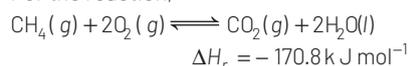


Which of the following statement is not true? **[CBSE AIPMT 2006]**

- (a) At equilibrium, the concentrations of  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{l})$  are not equal  
 (b) The equilibrium constant for the reaction is given by  $K_p = \frac{[\text{CO}_2]}{[\text{CH}_4][\text{O}_2]}$   
 (c) Addition of  $\text{CH}_4(\text{g})$  or  $\text{O}_2(\text{g})$  at equilibrium will cause a shift to the right  
 (d) The reaction is exothermic

**Ans. (b)**

For the reaction,



This equilibrium is an example of heterogeneous chemical equilibrium. Hence, for it

$$K_c = \frac{[\text{CO}_2]}{[\text{CH}_4][\text{O}_2]^2} \quad \dots(\text{i})$$

(equilibrium constant on the basis of concentration)

$$\text{and } K_p = \frac{p_{\text{CO}_2}}{p_{\text{CH}_4} \times p_{\text{O}_2}^2} \quad \dots(\text{ii})$$

(equilibrium constant according to partial pressure)

Thus, in this concentration of  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{l})$  are not equal at equilibrium.

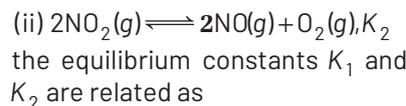
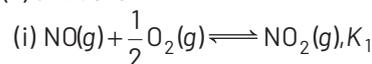
The equilibrium constant  $(K_p) = \frac{[\text{CO}_2]}{[\text{CH}_4][\text{O}_2]}$  is not correct expression.

On adding  $\text{CH}_4(\text{g})$  or  $\text{O}_2(\text{g})$  at equilibrium,  $K_c$  will be decreased according to expression (i) but  $K_c$  remains constant at constant temperature for a reaction, so for maintaining the constant value of  $K_c$ ,

the concentration of  $\text{CO}_2$  will increased in same order. Hence, on addition of  $\text{CH}_4$  or  $\text{O}_2$  equilibrium will cause to the right.

Combustion reaction is an example of exothermic reaction.

**13** In the two gaseous reactions (i) and (ii) at  $25^\circ\text{C}$



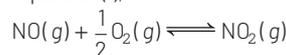
**[CBSE AIPMT 2005, 1994]**

$$\text{(a) } K_2 = \frac{1}{K_1} \quad \text{(b) } K_2 = K_1^{1/2}$$

$$\text{(c) } K_2 = \frac{1}{K_1^2} \quad \text{(d) } K_2 = K_1^2$$

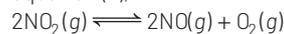
**Ans. (c)**

For equation (i),



$$K_1 = \frac{[\text{NO}_2]}{[\text{NO}][\text{O}_2]^{1/2}} \quad \dots(\text{i})$$

For equation (ii),



$$K_2 = \frac{[\text{NO}]^2[\text{O}_2]}{[\text{NO}_2]^2} \quad \dots(\text{ii})$$

Now, on reversing equation (i), we get,

$$\frac{1}{K_1} = \frac{1}{\frac{[\text{NO}_2]}{[\text{NO}][\text{O}_2]^{1/2}}}$$

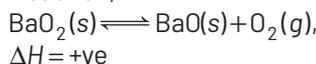
$$= \frac{[\text{NO}][\text{O}_2]^{1/2}}{[\text{NO}_2]}$$

$$\left(\frac{1}{K_1}\right)^2 = \left\{\frac{[\text{NO}][\text{O}_2]^{1/2}}{[\text{NO}_2]}\right\}^2$$

$$= \frac{[\text{NO}]^2[\text{O}_2]}{[\text{NO}_2]^2} = K_2$$

$$\frac{1}{K_1^2} = K_2$$

**14** Reaction,

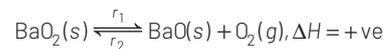


In equilibrium condition, pressure of  $\text{O}_2$  depends on

**[CBSE AIPMT 2002]**

- (a) increased mass of  $\text{BaO}_2$   
 (b) increased mass of  $\text{BaO}$   
 (c) increased temperature of equilibrium  
 (d) increased mass of  $\text{BaO}_2$  and  $\text{BaO}$  both

**Ans. (c)**



According to law of mass action, the rate of forward reaction =  $r_1$

$$r_1 \propto [\text{BaO}_2]$$

$$\text{or } r_1 = k_1 [\text{BaO}_2]$$

$\text{BaO}_2$  is solid substance in pure state concentration = 1m

$$\text{then, } r_1 = k_1$$

Similarly the rate of backward reaction =  $r_2$

$$r_2 \propto [\text{BaO}][\text{O}_2]$$

$$\text{or } r_2 = k_2 [\text{BaO}][\text{O}_2]$$

$\therefore$  Concentration of solid  $[\text{BaO}] = 1$   $[\text{O}_2(\text{g})]$

$$\therefore r_2 = k_2 [\text{O}_2]$$

At equilibrium,

$$r_1 = r_2$$

$$K_1 = K_2 [\text{O}_2]$$

$$\text{or } K_1 = K_2 \cdot p_{\text{O}_2}$$

where,  $p_{\text{O}_2}$  = partial pressure of  $\text{O}_2$

$$\text{or } \frac{K_1}{K_2} = p_{\text{O}_2} \quad (\text{equilibrium constant})$$

$$\therefore \frac{K_1}{K_2} = K$$

$$\text{or } K = p_{\text{O}_2}$$

So, from the above it is clear that pressure of  $\text{O}_2$  does not depend upon the concentration of reactants. The given equation is an endothermic reaction. If the temperature of such reaction is increased, then dissociation of  $\text{BaO}_2$  would increase and more  $\text{O}_2$  is produced.

**15** For the equilibrium,



which of the following expressions is correct? **[CBSE AIPMT 2000]**

$$\text{(a) } K_p = p_{\text{CO}_2}$$

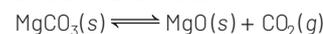
$$\text{(b) } K_p = \frac{[\text{MgO}][\text{CO}_2]}{[\text{MgCO}_3]}$$

$$\text{(c) } K_p = \frac{p_{\text{MgO}} \cdot p_{\text{CO}_2}}{p_{\text{MgCO}_3}}$$

$$\text{(d) } K_p = \frac{p_{\text{MgO}} + p_{\text{CO}_2}}{p_{\text{MgCO}_3}}$$

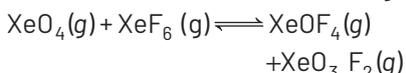
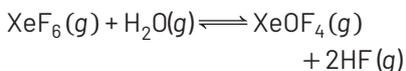
**Ans. (a)**

In heterogeneous system,  $K_c$  and  $K_p$  are not depend upon the concentration or pressure of solid substance. Hence, at equilibrium their concentration or pressure are assumed as one.

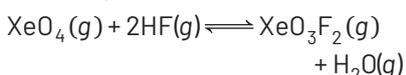


$$\therefore K_p = p_{\text{CO}_2}$$

- 16** If  $K_1$  and  $K_2$  are the respective equilibrium constants for the two reactions,



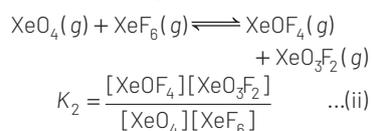
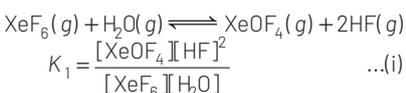
The equilibrium constant of the reaction,



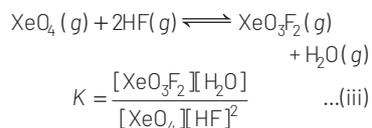
will be [CBSE AIPMT 1998]

- (a)  $K_1/(K_2)^2$  (b)  $K_1 \cdot K_2$   
(c)  $K_1/K_2$  (d)  $K_2/K_1$

**Ans. (d)**



For the reaction,



By dividing eq. (ii) by (i) we get,

$$K = \frac{K_2}{K_1}$$

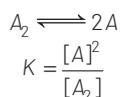
- 17** The equilibrium constants for the reaction,  $\text{A}_2 \rightleftharpoons 2\text{A}$  at 500 K and 700 K are  $1 \times 10^{-10}$  and  $1 \times 10^{-5}$ . The

given reaction is [CBSE AIPMT 1996]

- (a) exothermic (b) slow  
(c) endothermic (d) fast

**Ans. (b)**

For the reaction,



The value of equilibrium constant is very less and hence, the product concentration is also very less. So, the reaction is slow.

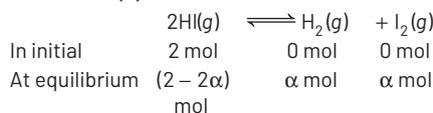
- 18** If  $\alpha$  is the fraction of HI dissociated at equilibrium in the reaction,  $2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)$  starting with the 2 moles of HI, then the

total number of moles of reactants and products at equilibrium are

[CBSE AIPMT 1996]

- (a)  $2 + 2\alpha$  (b) 2  
(c)  $1 + \alpha$  (d)  $2 - \alpha$

**Ans. (b)**



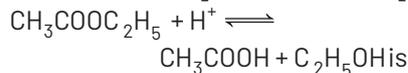
So, at equilibrium total moles

$$= 2 - 2\alpha + \alpha + \alpha$$

$$= 2 - 2\alpha + 2\alpha = 2$$

- 19** The rate constants for forward and backward reaction of hydrolysis of ester are  $1.1 \times 10^{-2}$  and  $1.5 \times 10^{-3}$  per minute. Equilibrium constant for the reaction,

[CBSE AIPMT 1995]



- (a) 4.33 (b) 5.33  
(c) 6.33 (d) 7.33

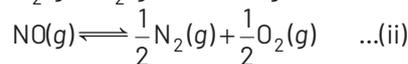
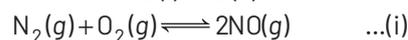
**Ans. (d)**

Equilibrium constant,  $K = \frac{k_f}{k_b}$

$$= \frac{\text{Forward rate constant}}{\text{Backward rate constant}}$$

$$K = \frac{1.1 \times 10^{-2}}{1.5 \times 10^{-3}} = \frac{1.1 \times 10}{1.5} = \frac{11}{1.5} = 7.33$$

- 20**  $K_1$  and  $K_2$  are equilibrium constant for reactions (i) and (ii)



Then,

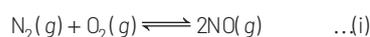
[CBSE AIPMT 1989]

$$(a) K_1 = \left[ \frac{1}{K_2} \right]^2 \quad (b) K_1 = K_2^2$$

$$(c) K_1 = \frac{1}{K_2} \quad (d) K_1 = (K_2)^0$$

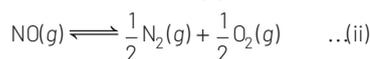
**Ans. (a)**

Consider reaction (i),



$$K_1 = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$$

Now, consider reaction (ii),



$$K_2 = \frac{[\text{N}_2]^{1/2}[\text{O}_2]^{1/2}}{[\text{NO}]}$$

$$\frac{1}{K_2} = \frac{1}{\frac{[\text{N}_2]^{1/2}[\text{O}_2]^{1/2}}{[\text{NO}]}} = \frac{[\text{NO}]}{[\text{N}_2]^{1/2}[\text{O}_2]^{1/2}}$$

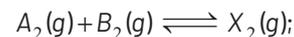
$$\left( \frac{1}{K_2} \right)^2 = \left\{ \frac{[\text{NO}]}{[\text{N}_2]^{1/2}[\text{O}_2]^{1/2}} \right\}^2$$

$$= \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} = K_1$$

## TOPIC 2

### Factors Affecting Equilibrium and Le-Chatelier's Principles

- 21** Which one of the following conditions will favour maximum formation of the product in the reaction, [NEET 2018]



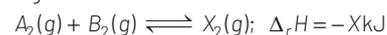
$$\Delta_r H = -X \text{ kJ}$$

- (a) High temperature and high pressure  
(b) Low temperature and low pressure  
(c) Low temperature and high pressure  
(d) High temperature and low pressure

**Ans. (c)**

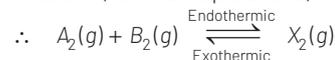
**Key Concept** The given question is based upon Le-Chatelier's principle. According to this principle, if a stress is applied to a reaction mixture at equilibrium, reaction proceeds in such a direction that relieves the stress.

The given reaction is



According to Le-Chatelier's principle, with increase in temperature the equilibrium shifts in the direction of endothermic reaction (i.e., heat is absorbed).

Alternatively, the decrease in temperature shifts the equilibrium towards the direction of exothermic reaction (i.e. heat is produced).



Similarly, an increase in pressure will shift the equilibrium to that direction which leads to decrease in total number of gaseous moles. Whereas, a decrease in the pressure will shift the equilibrium to that direction which leads to an increase in total number of gaseous moles.



$$\Delta n_g = 1 - 2 = -1$$

Thus, low temperature and high pressure will favour maximum formation of the product in the given reaction.

**22** For the reversible reaction,  
 $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{heat}$   
 the equilibrium shifts in forward direction [CBSE AIPMT 2014]

- (a) by increasing the concentration of  $NH_3(g)$   
 (b) by decreasing the pressure  
 (c) by decreasing the concentrations of  $N_2(g)$  and  $H_2(g)$   
 (d) by increasing pressure and decreasing temperature

**Ans. (d)**

Any change in the concentration, pressure and temperature of the reaction results in change in the direction of equilibrium. This change in the direction of equilibrium is governed by Le-Chatelier's principle. According to this equilibrium shifts in the opposite direction to undo the change.



- (a) **Increasing the concentration of  $NH_3(g)$**  On increasing the concentration of  $NH_3(g)$ , the equilibrium shifts in the backward direction where concentration of  $NH_3(g)$  decreases.  
 (b) **Decreasing the pressure** Since,  $p \propto n$  (number of moles), therefore, equilibrium shifts in the backward direction where number of moles are increasing.  
 (c) **Decreasing the concentration of  $N_2(g)$  and  $H_2(g)$**  Equilibrium shifts in the backward direction when concentration of  $N_2(g)$  and  $H_2(g)$  decreases.  
 (d) **Increasing pressure and decreasing temperature** On increasing pressure, equilibrium shifts in the forward direction where number of moles decreases. It is an example of exothermic reaction therefore decreasing temperature favours the forward direction.

**23**  $KMnO_4$  can be prepared from  $K_2MnO_4$  as per reaction,  
 $3MnO_4^{2-} + 2H_2O \rightleftharpoons 2MnO_4^- + MnO_2 + 4OH^-$   
 The reaction can go to completion by removing  $OH^-$  ions by adding [NEET 2013]

- (a) HCl (b) KOH  
 (c)  $CO_2$  (d)  $SO_2$

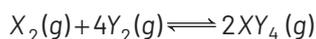
**Ans. (c)**

Since,  $OH^-$  are generated from weak acid ( $H_2O$ ), and a weak acid (like  $CO_2$ ) should be used to remove it. Because if we add strong acid like (HCl) it reverse the reaction. KOH increases the concentration of  $OH^-$ , thus again shifts the reaction in backward side.

$CO_2$  combines with  $OH^-$  to give carbonate which is easily removed.

$SO_2$  reacts with water to give strong acid, so it cannot be used.

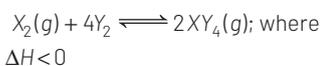
**24** The value of  $\Delta H$  for the reaction,



is less than zero. Formation of  $XY_4(g)$  will be favoured at [CBSE AIPMT 2011]

- (a) low pressure and low temperature  
 (b) high temperature and low pressure  
 (c) high pressure and low temperature  
 (d) high temperature and high pressure

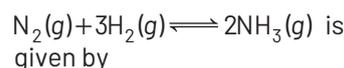
**Ans. (c)**



and  $\Delta n < 0$  [ $\Delta n = n_p - n_R$ ]

$\therefore$  The forward reaction is favoured at high pressure and low temperature. (According to Le-Chatelier's principle)

**25** The reaction quotient ( $Q$ ) for the reaction,



$$Q = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

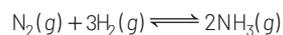
The reaction will proceed towards right side, if [CBSE AIPMT 2003]

- (a)  $Q > K_c$  (b)  $Q = 0$   
 (c)  $Q = K_c$  (d)  $Q < K_c$

where,  $K_c$  is the equilibrium constant.

**Ans. (a)**

For the reaction,



$$Q(\text{Quotient}) = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$\Delta n_g = 2 - 4 = -2$$

At equilibrium  $Q$  is equal to  $K_c$  but for the progress of reaction towards right side,  $Q > K_c$

**26** For a reversible reaction, if the concentrations of the reactants are doubled, the equilibrium constant will be [CBSE AIPMT 2000]

- (a) one-fourth (b) halved  
 (c) doubled (d) the same

**Ans. (d)**

Consider a hypothetical change,



For this reaction,  $K_{eq} = \frac{[C][D]}{[A][B]}$

For the above reaction if concentration of reactants are doubled then the rate of forward reaction increases for a short time but after sometime equilibrium will be established. So, concentration has no effect on equilibrium constant. It remains unchanged after increasing the concentration of reactants.

**27** According to Le-Chatelier's principle, adding heat to a solid  $\rightleftharpoons$  liquid equilibrium will cause the [CBSE AIPMT 1993]

- (a) temperature to increase  
 (b) temperature to decrease  
 (c) amount of liquid to decrease  
 (d) amount of solid to decrease

**Ans. (d)**

When we add heat to the equilibrium between solid and liquid, then the equilibrium shifts towards liquid and hence, the amount of solid decrease and amount of liquid increase.

**28** Which one of the following information can be obtained on the basis of Le-Chatelier's principle? [CBSE AIPMT 1992]

- (a) Dissociation constant of a weak acid  
 (b) Entropy change in a reaction  
 (c) Equilibrium constant of a chemical reaction  
 (d) Shift in equilibrium position on changing value of a constant

**Ans. (d)**

Le-Chatelier's and Braun French chemists made certain generalisations to explain the effect of changes in concentrations, temperature or pressure on the state of system in equilibrium. When a system is subjected to a change in one of these factors, the equilibrium gets disturbed and the system re-adjusts itself until it return to equilibrium.