

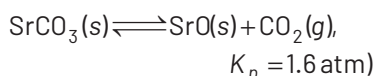
8

Chemical Equilibrium

TOPIC 1

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

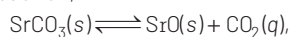
01 A 20 litre container at 400 K contains $\text{CO}_2(\text{g})$ at pressure 0.4 atm and an excess of SrO (neglect the volume of solid SrO). The volume of the container is now decreased by moving the movable piston fitted in the container. The maximum volume of the container, when pressure of CO_2 attains 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,



$K_p = 1.6 \text{ atm} = p_{\text{CO}_2}$ = 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,

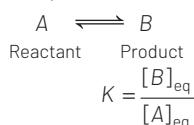
$$T_2 > T_1 \text{ and } H = -ve \\ \Rightarrow K_p > K'_p$$

03 If the value of an equilibrium constant for a particular reaction is 1.6×10^{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.

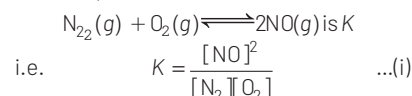
04 If the equilibrium constant for $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$ is K , the equilibrium constant for $\frac{1}{2}\text{N}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightleftharpoons \text{NO}(\text{g})$ will be, [CBSE AIPMT 2015]

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

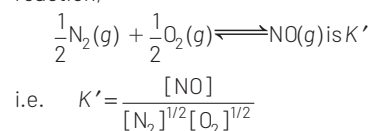
Ans. (a)

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

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

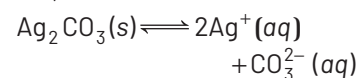
On comparing Eqs. (i) and (ii), we get

$$K = K'^2$$

or

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

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



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

$$(R = 8.314 \text{ JK}^{-1}\text{mol}^{-1})$$

[CBSE AIPMT 2014]

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

Ans. (b)

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

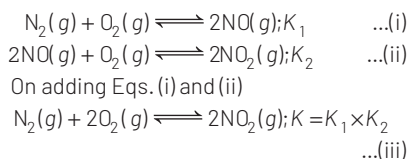
Given, $\Delta G^\circ = +63.3 \text{ kJ}$
 $= 63.3 \times 10^3 \text{ J}$
 Thus, substitute $\Delta G^\circ = 63.3 \times 10^3 \text{ J}$,
 $R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$ and $T = 298 \text{ K}$ [25 + 273 K] from the above equation we get,
 $63.3 \times 10^3 = -2.303 \times 8.314 \times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

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,
 $\text{NO}_2(\text{g}) \rightleftharpoons \frac{1}{2}\text{N}_2(\text{g}) + \text{O}_2(\text{g})$?

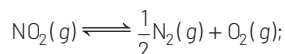
[CBSE AIPMT 2011]

- (a) $1/4 K_1 K_2$ (b) $[1/K_1 K_2]^{1/2}$
 (c) $1/(K_1 K_2)$ (d) $1/(2K_1 K_2)$

Ans. (b)



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



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

$\therefore K = \left[\frac{1}{K_1 K_2} \right]^{1/2}$

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°C are
 1.5×10^{-5} and 4.5×10^{-10} ,

respectively. The equilibrium
 constant for the equilibrium,

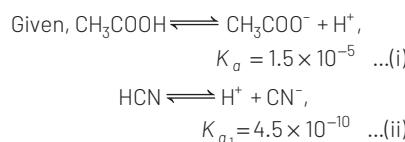


+ CH_3COO^- would be

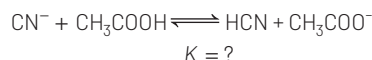
[CBSE AIPMT 2009]

- (a) 3.0×10^5 (b) 3.0×10^{-5}
 (c) 3.0×10^{-4} (d) 3.0×10^4

Ans. (d)



For



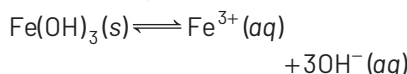
On subtracting Eq. (ii) from Eq. (i), we get



$$K = \frac{K_a}{K_{a1}} = \frac{1.5 \times 10^{-5}}{4.5 \times 10^{-10}}$$

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

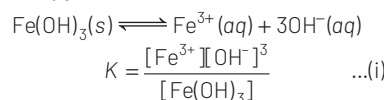
09 If the concentration of OH^- ions in
 the reaction,



is decreased by $1/4$ times, then
 equilibrium concentration of 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 Fe^{3+} is increased x
 times, on decreasing the concentration
 of OH^- by $\frac{1}{4}$ times

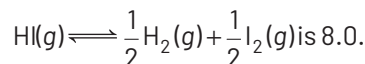
$$K = \frac{[x\text{Fe}^{3+}][\frac{1}{4}\text{OH}^-]^3}{[\text{Fe}(\text{OH})_3]} \dots(ii)$$

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

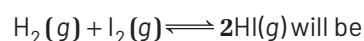
$$\frac{1}{64} \times x = 1$$

$\Rightarrow x = 64$ times

10 The value of equilibrium constant of
 the reaction,



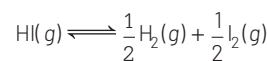
The equilibrium constant of the
 reaction,



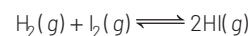
[CBSE AIPMT 2008]

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

Ans. (b)



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



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

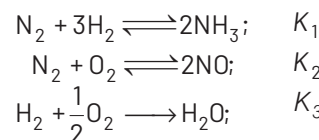
From Eqs. (i) and (ii)

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

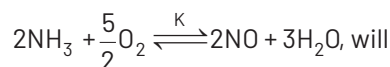
$$K' = \frac{1}{K^2} = \frac{1}{(8)^2} = \frac{1}{64}$$

11 The equilibrium constants of the
 following are

[NEET 2017, CBSE AIPMT 2007]



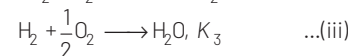
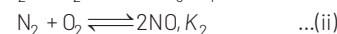
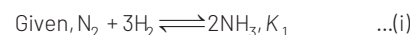
The equilibrium constant (K) of the
 reaction



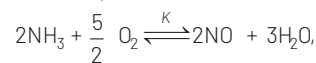
be

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

Ans. (b)

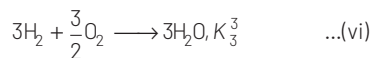
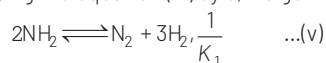


To calculate,

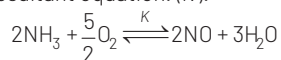


$K = ? \dots(iv)$

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

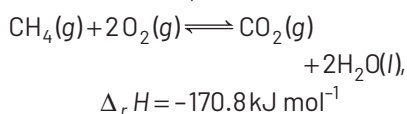


Now, add equation. (ii), (v) and (vi), we get the resultant equation. (iv).



$$\therefore K = \frac{K_2 K_3^3}{K_1}$$

12 For the reaction,

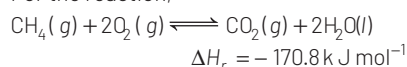


Which of the following statement is not true? **[CBSE AIPMT 2006]**

- (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.

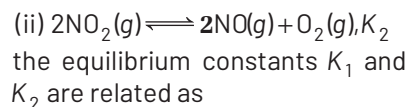
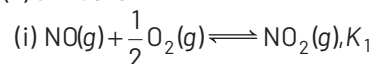
The equilibrium constant $(K_p) = \frac{[\text{CO}_2]}{[\text{CH}_4][\text{O}_2]}$ is not correct expression.

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 CO_2 will increase in same order. Hence, on addition of CH_4 or 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°C



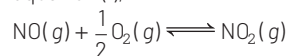
[CBSE AIPMT 2005, 1994]

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

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

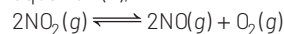
Ans. (c)

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,

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

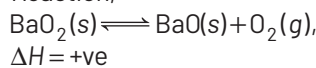
$$= \frac{[\text{NO}][\text{O}_2]^{1/2}}{[\text{NO}_2]}$$

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

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

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

14 Reaction,

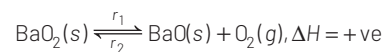


In equilibrium condition, pressure of O_2 depends on

[CBSE AIPMT 2002]

- (a) increased mass of BaO_2
 (b) increased mass of BaO
 (c) increased temperature of equilibrium
 (d) increased mass of BaO_2 and BaO both

Ans. (c)



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

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

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

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_{O_2} = partial pressure of 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 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 BaO_2 would increase and more O_2 is produced.

15 For the equilibrium,



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

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

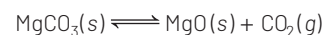
$$\text{(b) } K_p = \frac{[\text{MgO}][\text{CO}_2]}{[\text{MgCO}_3]}$$

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

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

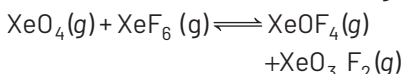
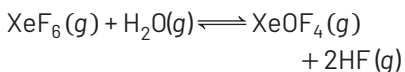
Ans. (a)

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.

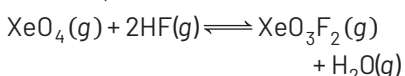


$$\therefore K_p = p_{\text{CO}_2}$$

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



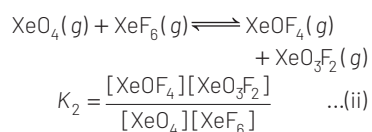
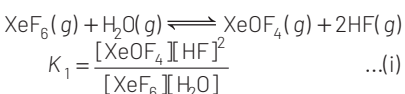
The equilibrium constant of the reaction,



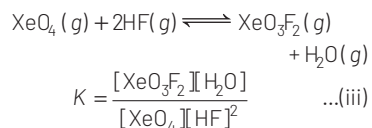
will be [CBSE AIPMT 1998]

- (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,

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

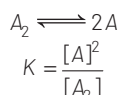
- 17** The equilibrium constants for the reaction, $\text{A}_2 \rightleftharpoons 2\text{A}$ at 500 K and 700 K are 1×10^{-10} and 1×10^{-5} . The

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.

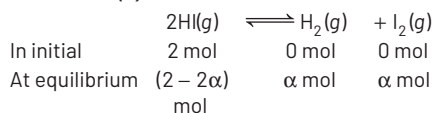
- 18** If α is the fraction of HI dissociated at equilibrium in the reaction, $2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)$ starting with the 2 moles of HI, then the

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)

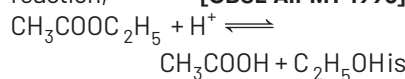


So, at equilibrium total moles

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

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

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



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

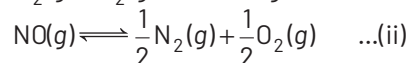
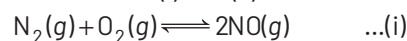
Ans. (d)

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)



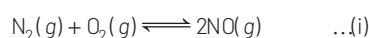
Then,

[CBSE AIPMT 1989]

- (a) $K_1 = \left[\frac{1}{K_2} \right]^2$ (b) $K_1 = K_2^2$
(c) $K_1 = \frac{1}{K_2}$ (d) $K_1 = (K_2)^0$

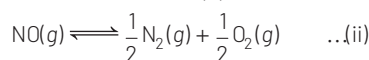
Ans. (a)

Consider reaction (i),



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

Now, consider reaction (ii),



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

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

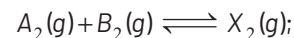
$$\left(\frac{1}{K_2} \right)^2 = \left\{ \frac{[\text{NO}]}{[\text{N}_2]^{1/2}[\text{O}_2]^{1/2}} \right\}^2$$

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

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]



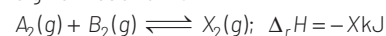
$$\Delta_r H = -X \text{ kJ}$$

- (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)

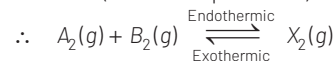
Key Concept The given question is based upon Le-Chatelier's principle. According to this principle, if a stress is applied to a reaction mixture at equilibrium, reaction proceeds in such a direction that relieves the stress.

The given reaction is

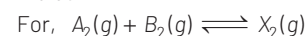


According to Le-Chatelier's principle, with increase in temperature the equilibrium shifts in the direction of endothermic reaction (i.e., heat is absorbed).

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,

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) + \text{heat}$$
the equilibrium shifts in forward direction [CBSE AIPMT 2014]
- by increasing the concentration of $\text{NH}_3(\text{g})$
 - by decreasing the pressure
 - by decreasing the concentrations of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$
 - 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.



- Increasing the concentration of $\text{NH}_3(\text{g})$** On increasing the concentration of $\text{NH}_3(\text{g})$, the equilibrium shifts in the backward direction where concentration of $\text{NH}_3(\text{g})$ decreases.
- Decreasing the pressure** Since, $p \propto n$ (number of moles), therefore, equilibrium shifts in the backward direction where number of moles are increasing.
- Decreasing the concentration of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$** Equilibrium shifts in the backward direction when concentration of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$ decreases.
- 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,

$$3\text{MnO}_4^{2-} + 2\text{H}_2\text{O} \rightleftharpoons 2\text{MnO}_4^- + \text{MnO}_2 + 4\text{OH}^-$$
The reaction can go to completion by removing OH^- ions by adding [NEET 2013]
- HCl
 - KOH
 - CO_2
 - 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.

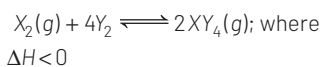
SO_2 reacts with water to give strong acid, so it cannot be used.

- 24** The value of ΔH for the reaction,

$$\text{X}_2(\text{g}) + 4\text{Y}_2(\text{g}) \rightleftharpoons 2\text{XY}_4(\text{g})$$
is less than zero. Formation of $\text{XY}_4(\text{g})$ will be favoured at [CBSE AIPMT 2011]

- low pressure and low temperature
- high temperature and low pressure
- high pressure and low temperature
- 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,

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$$
is given by
- $$Q = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

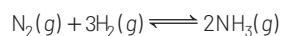
The reaction will proceed towards right side, if [CBSE AIPMT 2003]

- $Q > K_c$
- $Q = 0$
- $Q = K_c$
- $Q < K_c$

where, K_c is the equilibrium constant.

Ans. (d)

For the reaction,



$$Q(\text{Quotient}) = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

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

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]
- one-fourth
 - halved
 - doubled
 - the same

Ans. (d)

Consider a hypothetical change,



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

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]
- temperature to increase
 - temperature to decrease
 - amount of liquid to decrease
 - 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]
- Dissociation constant of a weak acid
 - Entropy change in a reaction
 - Equilibrium constant of a chemical reaction
 - 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.