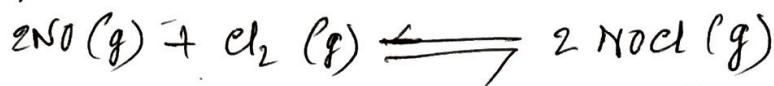


Q1. The value of K_p at $25^\circ C$ for the reaction



is $1.9 \times 10^3 \text{ atm}^{-1}$. Calculate the value of K_c at the same temperature?

Soln. We can write the general expression as

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

$$\text{or, } K_c = \frac{K_p}{(RT)^{\Delta n}}$$

$$\text{Here, } T = 25^\circ C = (25 + 273) K = 298 K$$

$$R = 0.082$$

$$\Delta n = 2 - (2+1) = -1$$

$$K_p = 1.9 \times 10^3$$

Substituting these values in the general expression,

$$K_c = \frac{1.9 \times 10^3}{(0.082 \times 298)^{-1}}$$

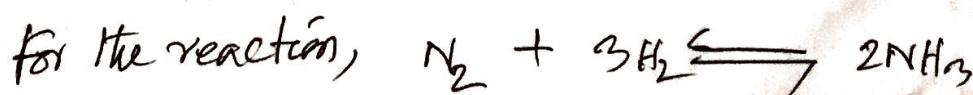
$$= 4.8 \times 10^4. \text{ Ans.}$$



Q2. At $500^\circ C$, the reaction between N_2 and H_2 to form NH_3 has $K_c = 6.0 \times 10^2$. What is the numerical value of K_p for the reaction?

Soln. Here, we will use the general expression

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



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$$\text{we have } \Delta n = \frac{\text{(Sum of quotients of products)}}{\text{(Sum of quotients of reactants)}} - 2$$

$$= 2 - 4 = -2$$

$$\text{So, } K_c = 6.0 \times 10^{-2}$$

$$T = 500^\circ\text{C} = (500 + 273) = 773\text{K}$$

$$R = 0.0821$$

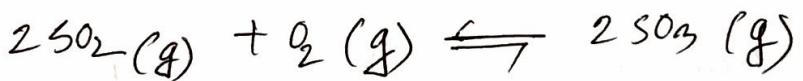
Substituting the value of R , T , K_c and Δn in the general expression, we have

$$K_p = [(6.0 \times 10^{-2})(0.0821) \times (773)]^{-2}$$

$$= 1.5 \times 10^{-5} \text{ Ans.}$$



At equilibrium for the reaction



The concentrations of reactant and products at 727°C were found to be: $\text{SO}_2 = 0.22 \text{ mol L}^{-1}$
 $\text{O}_2 = 0.40 \text{ mol L}^{-1}$ and
 $\text{SO}_3 = 0.33 \text{ mol L}^{-1}$

What is the value of the equilibrium constant, K_c at this temperature?

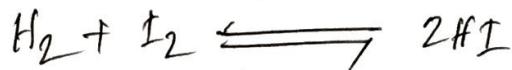
Soln. Write the equilibrium constant expression for the reaction,

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{(0.33)^2}{(0.22)^2 \times (0.40)}$$

$$= 3.7 \text{ mol L}^{-1} \text{ Ans.}$$



Q. At a certain temperature 0.100 mole of H_2 and 0.100 mole of I_2 were placed in a one-litre flask. The purple colour of Iodine vapour was used to monitor the reaction. After a time the equilibrium,



was established and it was found that the concentration of I_2 decreased to 0.020 mole/litre. Calculate the value of K_c for the reaction at the given temperature?

Soln. The equilibrium constant expression for the reaction,

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

The initial moles of H_2 , I_2 and HI present at equilibrium are



moles at equilibrium : $(0.100-x)$ $(0.100-x)$ $2x$

Since the volume of the reaction vessel is one litre, the moles at equilibrium also represent the equilibrium concentrations. Thus—

$$[I_2] = 0.100 - x = 0.020 \quad (\text{given})$$
$$\text{or, } x = 0.100 - 0.020$$
$$= 0.080$$

Knowing the value of x , we can give the equilibrium concentrations of H_2 , I_2 and HI

$$[\text{H}_2] = 0.100 - x = 0.100 - 0.080 = 0.020$$

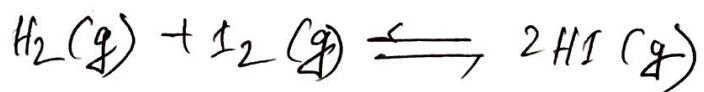
$$[\text{HI}] \leftarrow 2x = 0.160$$

$$[\text{I}_2] = 0.020 \text{ (given)}$$

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2] \cdot [\text{I}_2]} = \frac{(0.160)^2}{(0.02) \times (0.02)} = 64$$



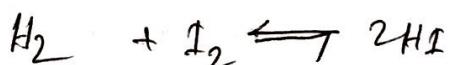
Q. 2. 13.5 ml of HI are produced by the interaction of 8.1 ml of H_2 and 9.3 ml of I_2 vapour at 994°C . Calculate the equilibrium constant at this temperature of the reaction?



Soln The equilibrium constant expression for the reaction is

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2] \cdot [\text{I}_2]}$$

Given reaction,



Equilibrium volumes : $(8.1-x)$ $(9.3-x)$ $2x$

The experimental volume of HI = 13.5 ml (given)

$$2x = 13.5$$

$$\therefore x = 6.75$$

Thus the equilibrium concentrations are :

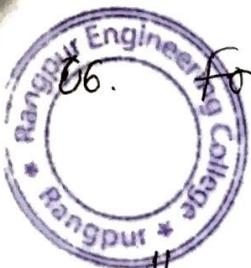
$$[\text{H}_2] = (8.1-x) = 8.1 - 6.75 = 1.35$$

$$[\text{I}_2] = (9.3-x) = (9.3 - 6.75) = 2.55$$

$$[\text{HI}] = 13.5$$

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2] \cdot [\text{I}_2]} = \frac{(13.5)^2}{1.35 \times 2.55}$$

$$= 52.94 \text{ Ans.}$$



Q6. For the reaction



the value of K at 552°C is 0.137. If 5 moles of CO_2 , 5.0 moles of H_2 , 1 mole of CO and 1.0 mole of H_2O are initially present. What are the actual concentrations of CO_2 , H_2 , CO and H_2O at equilibrium?

Soln If x moles of CO_2 has reached at equilibrium x moles of H_2O will be produced

The moles at equilibrium due to the reaction will be



Moles at equilibrium: $5-x$ $5-x$ x x

But 1 mole each of CO and H_2O were present initially. Therefore the actual moles at equilibrium would be

$$\text{CO}_2 = 5-x$$

$$\text{H}_2 = 5-x$$

$$\text{CO} = (1+x)$$

$$\text{H}_2\text{O} = 1+x$$

If V litres be the volume of the reaction vessel, the concentration of the various species are:

$$[\text{CO}_2] = \frac{5-x}{V}$$

$$[\text{H}_2] = \frac{5-x}{V}$$

$$[\text{CO}] = \frac{1+x}{V}$$

$$[\text{H}_2\text{O}] = \frac{1+x}{V}$$

Substituting the values in the equilibrium constant expression, we have —

$$K = \frac{[e_0][H_2O]}{[CO_2][H_2]}$$

$$\text{or, } 0.137 = \frac{[1+x/v][1+x/v]}{[5-x/v][5-x/v]}$$

$$\text{or, } 0.137 = \frac{(1+x)^v/v^v}{(5-x)^v/v^v}$$

$$\text{or, } 0.137 = \frac{(1+x)^v}{(5-x)^v}$$

$$\text{or, } 0.137 = \frac{1+2x+x^v}{25-10x+x^v}$$

$$\text{or, } 0.137 \times (25) - 0.137 \times 10x + 0.137 x^v = 1+2x+x^v$$

$$\text{or, } 0.863x^v + 3.370x - 3.425 = 0$$

Solving for x using the quadratic formula

$$x = 0.62$$

thus the final equilibrium concentration are

$$[e_0] = 1+x = 1+0.62 = 1.62 \text{ mole L}^{-1}$$

$$[H_2O] = 1+x = 1+0.62 = 1.62 \text{ mole L}^{-1}$$

$$[CO_2] = 5-x = 5-0.62 = 4.38 \text{ mole L}^{-1}$$

$$[H_2] = 5-x = 5-0.62 = 4.38 \text{ mole L}^{-1}$$



for the reaction $I_2(g) \rightleftharpoons 2I(g)$ $K = 3.76 \times 10^{-5}$
at $727^\circ C$. Let 1 mole of I_2 be injected into 2 litre
glass box at $727^\circ C$. What will be the concentrations of I_2
and I when the equilibrium is attained?

Soln: The balanced equation for the reaction is

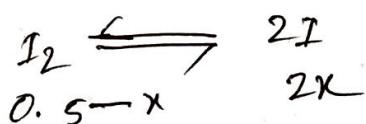


1 mole of I_2 is injected in a 2 litre box. Thus the
initial concentration of I_2 is

$$\frac{1 \text{ mole}}{2 \text{ litre}} = 0.5 \text{ M}$$



Let, x moles of I_2 decompose at equilibrium. Then
the moles of the two species are



Therefore the equilibrium concentrations are

$$[I_2] = 0.5 - x$$

$$[I] = 2x$$

$$\text{and } K = 3.76 \times 10^{-5}$$

And Substituting the values in the equilibrium
constant expression, we have

$$K = \frac{[I]^2}{[I_2]} = 3.76 \times 10^{-5}$$

$$\text{or, } \frac{(2x)^2}{(0.5-x)} = 3.76 \times 10^{-5}$$

∴ Solving for x using the quadratic equation

$$\text{formule, } x = 2.17 \times 10^{-3}$$

Thus the concentration of I_2 and I at equilibrium are —

$$[I] = 2x = 2 \times 2.17 \times 10^{-3} \\ = 4.34 \times 10^{-3} \text{ mole L}^{-1}$$

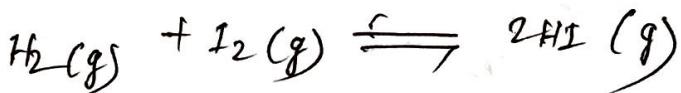
$$[I_2] = 0.5 - x = 0.5 - 2.17 \times 10^{-3} \\ = 0.498 \text{ mole L}^{-1}$$

Ans.



* * ————— Unit * *

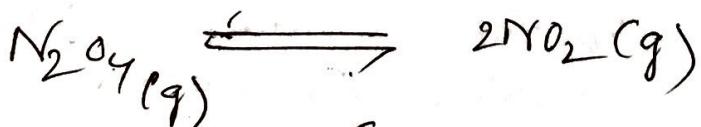
Find out the unit of the following reaction



$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(\text{mole L}^{-1})^2}{\text{mol}^{-1} \text{mol}^{-1}} = \text{No unit}$$

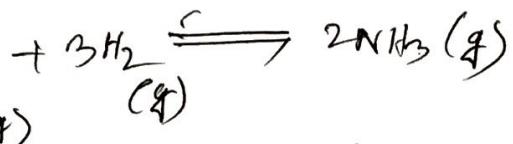
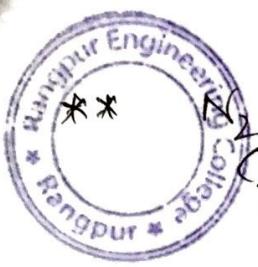
$$K_p = \frac{(P_{HI})^2}{(P_{H_2})(P_{I_2})} = \frac{(\text{atm})^2}{(\text{atm})(\text{atm})} = \text{No units}$$

* * ————— * *



$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(\text{mole L}^{-1})^2}{\text{mol L}^{-1}} = \text{mole}^{-1}$$

$$K_p = \frac{(P_{NO_2})^2}{(P_{N_2O_4})} = \frac{(\text{atm})^2}{(\text{atm})} \\ = \text{atm}$$



$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(mol L^{-1})^2}{(mol L^{-1})(mol L^{-1})^3} = mol^{-1} L^2$$

$$K_p = \frac{(P_{NH_3})^2}{(P_{N_2})(P_{H_2})^3} = \frac{(atm)^2}{(atm)(atm)^3} = atm^{-2} \text{ Ans}$$

* * → *

We know that, $\Delta G^\circ = -RT \ln K$
 or, $\Delta G^\circ = -2.303 RT \log K$

Calculate K for reaction which has ΔG° value -20 kcal at 25°C.

Soln. We know that, $\Delta G^\circ = -2.303 RT \log K$

where $\Delta G^\circ = \text{Standard free Energy}$

$R = \text{gas constant}$

$T = \text{Absolute Temperature}$

$K = \text{equilibrium constant}$

If ΔG° is given in calories, $R = 1.99$

$$T = (25 + 273) K = 298 K$$

$$\therefore -\log K = \frac{-\Delta G^\circ}{(2.303)(1.99)(298)} \\ = \frac{20,000}{1365.75} = 14.2$$

Taking antilogs, $\log K = 0.7 + 14.0$.

$$K = 5 \times 10^4 \text{ Ans.}$$



The Standard free energy change for the reaction,



is $+173.1\text{ kJ}$. Calculate K_p for the reaction at 25°C

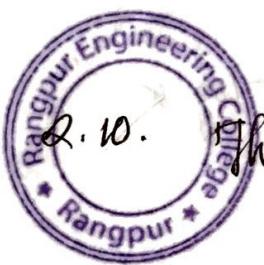
Soln. Here, we will use the expression,

$$\Delta G^\circ = -2.303 RT \log K_p$$

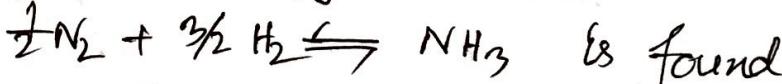
$$\text{or, } \log K_p = -\frac{\Delta G^\circ}{2.303 RT}$$

$$\text{or, } \log K_p = -\frac{1.73 \times 10^5}{(2.303)(8.314)(298)}$$
$$= -30.34$$

Taking anti-log, $K_p = 10^{0.66} \times 10^{-31}$
 $= 9.6 \times 10^{-31}$ Ans.



Q. 10. The value of K_p at 298 K for the reaction



is found to be 826.0 partial pressures being measured in atmospheric units. Calculate ΔG° at 298 K

Soln. We know that, $\Delta G^\circ = -2.303 RT \log K_p$

$$\text{Here, } R = 1.987 \text{ cal mol}^{-1}\text{K}^{-1}$$

$$T = 298\text{ K}$$

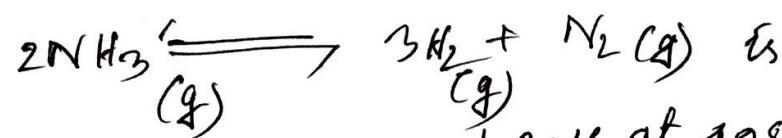
$$K_p = 826 ; \Delta G^\circ = ?$$

Substituting the values we get,

$$\Delta G^\circ = -2.303 \times 1.987 \times 298 \times \log 826$$
$$= 3977.78 \text{ cal}$$



Q. 11. The equilibrium constant K_p for the reaction



1.22×10^{-3} at 298 K and 2.16 at 498 K . Calculate ΔH° for the reaction.

Soln. we know that, $\log \frac{K_p_2}{K_p_1} = -\frac{\Delta H^\circ}{2.303 R} \left(\frac{T_2 - T_1}{T_1 T_2} \right)$

Hence, $K_p_2 = 2.16$; $K_p_1 = 1.22 \times 10^{-3}$
 $T_2 = 498\text{ K}$; $T_1 = 298\text{ K}$
 $R = 1.987 \text{ cal K}^{-1}\text{mol}^{-1}$

Substituting the values we get,

$$\log \frac{2.16}{1.22 \times 10^{-3}} = -\frac{\Delta H^\circ}{2.303 \times 1.987} \left[\frac{498 - 298}{498 \times 298} \right]$$

$$\text{or, } \log \frac{2.16}{1.22 \times 10^{-3}} = -\frac{\Delta H^\circ}{2.303 \times 1.987} \times \frac{200}{498 \times 298}$$

$$\therefore \Delta H^\circ = 11028.9 \text{ cal/s.} \\ = 11.0289 \text{ K cal/s. (Ans)}$$