

# Chemistry

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(Chapter – 7) (Equilibrium)  
(Class – XI)

## Question 7.1:

A liquid is in equilibrium with its vapour in a sealed container at a fixed temperature. The volume of the container is suddenly increased.

- What is the initial effect of the change on vapour pressure?
- How do rates of evaporation and condensation change initially?
- What happens when equilibrium is restored finally and what will be the final vapour pressure?

## Answer 7.1:

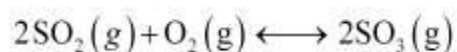
**(a)** If the volume of the container is suddenly increased, then the vapour pressure would decrease initially. This is because the amount of vapour remains the same, but the volume increases suddenly. As a result, the same amount of vapour is distributed in a larger volume.

**(b)** Since the temperature is constant, the rate of evaporation also remains constant. When the volume of the container is increased, the density of the vapour phase decreases. As a result, the rate of collisions of the vapour particles also decreases. Hence, the rate of condensation decreases initially.

**(c)** When equilibrium is restored finally, the rate of evaporation becomes equal to the rate of condensation. In this case, only the volume changes while the temperature remains constant. The vapour pressure depends on temperature and not on volume. Hence, the final vapour pressure will be equal to the original vapour pressure of the system.

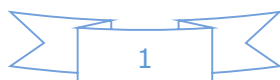
## Question 7.2:

What is  $K_c$  for the following equilibrium when the equilibrium concentration of each substance is:  $[\text{SO}_2] = 0.60 \text{ M}$ ,  $[\text{O}_2] = 0.82 \text{ M}$  and  $[\text{SO}_3] = 1.90 \text{ M}$  ?



## Answer 7.2:

The equilibrium constant ( $K_c$ ) for the give reaction is:



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$$\begin{aligned}K_c &= \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} \\&= \frac{(1.90)^2 \text{ M}^2}{(0.60)^2 (0.821) \text{ M}^3} \\&= 12.239 \text{ M}^{-1} \text{ (approximately)}\end{aligned}$$

Hence,  $K$  for the equilibrium is  $12.239 \text{ M}^{-1}$ .

### Question 7.3:

At a certain temperature and total pressure of  $10^5 \text{ Pa}$ , iodine vapour contains 40% by volume of I atoms



Calculate  $K_p$  for the equilibrium.

### Answer 7.3:

Partial pressure of I atoms,

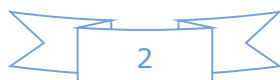
$$\begin{aligned}p_i &= \frac{40}{100} \times p_{\text{total}} \\&= \frac{40}{100} \times 10^5 \\&= 4 \times 10^4 \text{ Pa}\end{aligned}$$

Partial pressure of  $\text{I}_2$  molecules,

$$\begin{aligned}p_{\text{I}_2} &= \frac{60}{100} \times p_{\text{total}} \\&= \frac{60}{100} \times 10^5 \\&= 6 \times 10^4 \text{ Pa}\end{aligned}$$

Now, for the given reaction,

$$\begin{aligned}K_p &= \frac{(p_I)^2}{p_{\text{I}_2}} \\&= \frac{(4 \times 10^4)^2 \text{ Pa}^2}{6 \times 10^4 \text{ Pa}} \\&= 2.67 \times 10^4 \text{ Pa}\end{aligned}$$



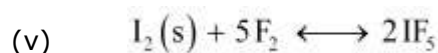
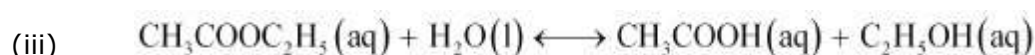
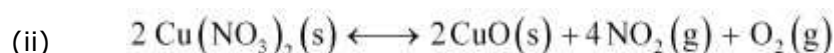
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## Question 7.4:

Write the expression for the equilibrium constant,  $K_c$  for each of the following reactions:



## Answer 7.4:

(i) 
$$K_c = \frac{[\text{NO}(\text{g})]^2 [\text{Cl}_2(\text{g})]}{[\text{NOCl}(\text{g})]^2}$$

(ii) 
$$K_c = \frac{[\text{CuO}(\text{s})]^2 [\text{NO}_2(\text{g})]^4 [\text{O}_2(\text{g})]}{[\text{Cu}(\text{NO}_3)_2(\text{s})]^2}$$

$$= [\text{NO}_2(\text{g})]^4 [\text{O}_2(\text{g})]$$

(iii) 
$$K_c = \frac{[\text{CH}_3\text{COOH}(\text{aq})][\text{C}_2\text{H}_5\text{OH}(\text{aq})]}{[\text{CH}_3\text{COOC}_2\text{H}_5(\text{aq})][\text{H}_2\text{O}(\text{l})]} = \frac{[\text{CH}_3\text{COOH}(\text{aq})][\text{C}_2\text{H}_5\text{OH}(\text{aq})]}{[\text{CH}_3\text{COOC}_2\text{H}_5(\text{aq})]}$$

(iv) 
$$K_c = \frac{[\text{Fe}(\text{OH})_3(\text{s})]}{[\text{Fe}^{3+}(\text{aq})][\text{OH}^-(\text{aq})]^3}$$
$$= \frac{1}{[\text{Fe}^{3+}(\text{aq})][\text{OH}^-(\text{aq})]^3}$$

(v) 
$$K_c = \frac{[\text{IF}_5]^2}{[\text{I}_2(\text{s})][\text{F}_2]^5}$$
$$= \frac{[\text{IF}_5]^2}{[\text{F}_2]^5}$$

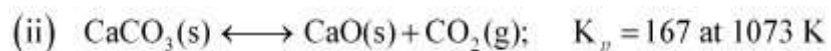
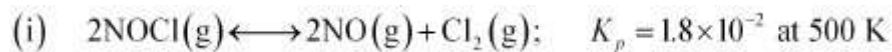
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## Question 7.5:

Find out the value of  $K_c$  for each of the following equilibria from the value of  $K_p$ :



## Answer 7.5:

The relation between  $K_p$  and  $K_c$  is given as:

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

(a) Here,

$$\Delta n = 3 - 2 = 1$$

$$R = 0.0831 \text{ barLmol}^{-1}\text{K}^{-1}$$

$$T = 500 \text{ K}$$

$$K_p = 1.8 \times 10^{-2}$$

Now,

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

$$\Rightarrow 1.8 \times 10^{-2} = K_c (0.0831 \times 500)^1$$

$$\Rightarrow K_c = \frac{1.8 \times 10^{-2}}{0.0831 \times 500}$$
$$= 4.33 \times 10^{-4} \text{ (approximately)}$$

(b) Here,

$$\Delta n = 2 - 1 = 1$$

$$R = 0.0831 \text{ barLmol}^{-1}\text{K}^{-1}$$

$$T = 1073 \text{ K}$$

$$K_p = 167$$

Now,

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

$$\Rightarrow 167 = K_c (0.0831 \times 1073)^{\Delta n}$$

$$\Rightarrow K_c = \frac{167}{0.0831 \times 1073}$$
$$= 1.87 \text{ (approximately)}$$



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## Question 7.6:

For the following equilibrium,  $K_c = 6.3 \times 10^{14}$  at 1000 K



Both the forward and reverse reactions in the equilibrium are elementary bimolecular reactions. What is  $K_c$ , for the reverse reaction?

## Answer 7.6:

It is given that  $K_c$  for the forward reaction is  $6.3 \times 10^{14}$ .

Then,  $K_c$  for the reverse reaction will be,

$$\begin{aligned} K'_c &= \frac{1}{K_c} \\ &= \frac{1}{6.3 \times 10^{14}} \\ &= 1.59 \times 10^{-15} \end{aligned}$$

## Question 7.7:

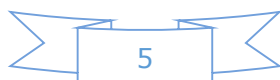
Explain why pure liquids and solids can be ignored while writing the equilibrium constant expression?

## Answer 7.7:

For a pure substance (both solids and liquids),

$$\begin{aligned} [\text{Pure substance}] &= \frac{\text{Number of moles}}{\text{Volume}} \\ &= \frac{\text{Mass/molecular mass}}{\text{Volume}} \\ &= \frac{\text{Mass}}{\text{Volume} \times \text{Molecular mass}} \\ &= \frac{\text{Density}}{\text{Molecular mass}} \end{aligned}$$

Now, the molecular mass and density (at a particular temperature) of a pure substance is always fixed and is accounted for in the equilibrium constant. Therefore, the values of pure substances are not mentioned in the equilibrium constant expression.



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## Question 7.8:

Reaction between  $N_2$  and  $O_2$  takes place as follows:

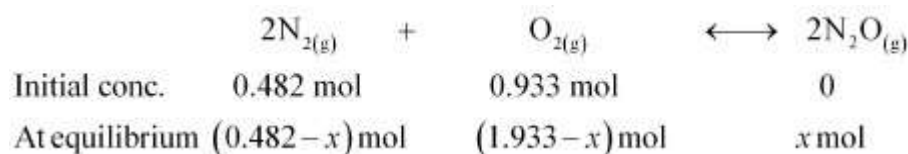


If a mixture of 0.482 mol of  $N_2$  and 0.933 mol of  $O_2$  is placed in a 10 L reaction vessel and allowed to form  $N_2O$  at a temperature for which  $K_c = 2.0 \times 10^{-37}$ , determine the composition of equilibrium mixture.

## Answer 7.8:

Let the concentration of  $N_2O$  at equilibrium be  $x$ .

The given reaction is:



Therefore, at equilibrium, in the 10 L vessel:

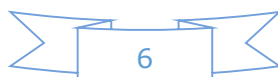
$$[N_2] = \frac{0.482 - x}{10}, [O_2] = \frac{0.933 - x/2}{10}, [N_2O] = \frac{x}{10}$$

The value of equilibrium constant i.e.  $K_c = 2.0 \times 10^{-37}$  is very small. Therefore, the amount of  $N_2$  and  $O_2$  reacted is also very small. Thus,  $x$  can be neglected from the expressions of molar concentrations of  $N_2$  and  $O_2$ . Then,

$$[N_2] = \frac{0.482}{10} = 0.0482 \text{ mol L}^{-1} \text{ and } [O_2] = \frac{0.933}{10} = 0.0933 \text{ mol L}^{-1}$$

Now,

$$\begin{aligned} K_c &= \frac{[N_2O_{(g)}]^2}{[N_{2(g)}]^2 [O_{2(g)}]} \\ \Rightarrow 2.0 \times 10^{-37} &= \frac{\left(\frac{x}{10}\right)^2}{(0.0482)^2 (0.0933)} \\ \Rightarrow \frac{x^2}{100} &= 2.0 \times 10^{-37} \times (0.0482)^2 \times (0.0933) \\ \Rightarrow x^2 &= 43.35 \times 10^{-40} \\ \Rightarrow x &= 6.6 \times 10^{-20} \\ [N_2O] &= \frac{x}{10} = \frac{6.6 \times 10^{-20}}{10} \\ &= 6.6 \times 10^{-21} \end{aligned}$$

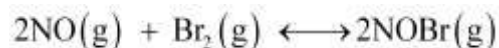


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## Question 7.9:

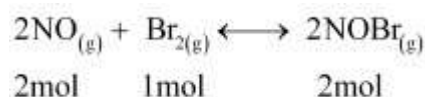
Nitric oxide reacts with Br<sub>2</sub> and gives nitrosyl bromide as per reaction given below:



When 0.087 mol of NO and 0.0437 mol of Br<sub>2</sub> are mixed in a closed container at constant temperature, 0.0518 mol of NOBr is obtained at equilibrium. Calculate equilibrium amount of NO and Br<sub>2</sub>.

## Answer 7.9:

The given reaction is:



Now, 2 mol of NOBr are formed from 2 mol of NO. Therefore, 0.0518 mol of NOBr are formed from 0.0518 mol of NO.

Again, 2 mol of NOBr are formed from 1 mol of Br.

Therefore, 0.0518 mol of NOBr are formed from  $\frac{0.0518}{2}$  mol of Br, or 0.0259 mol of NO.

The amount of NO and Br present initially is as follows:

$$[\text{NO}] = 0.087 \text{ mol} \quad [\text{Br}_2] = 0.0437 \text{ mol}$$

Therefore, the amount of NO present at equilibrium is:

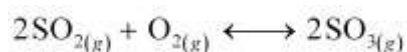
$$[\text{NO}] = 0.087 - 0.0518 = 0.0352 \text{ mol}$$

And, the amount of Br present at equilibrium is:

$$[\text{Br}_2] = 0.0437 - 0.0259 = 0.0178 \text{ mol}$$

## Question 7.10:

At 450 K,  $K_p = 2.0 \times 10^{10}$ /bar for the given reaction at equilibrium.



What is  $K_c$  at this temperature?

## Answer 7.10:

For the given reaction,

$$\Delta n = 2 - 3 = -1$$

$$T = 450 \text{ K}$$



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$$R = 0.0831 \text{ bar L bar K}^{-1} \text{ mol}^{-1}$$

$$K_p = 2.0 \times 10^{10} \text{ bar}^{-1}$$

We know that,

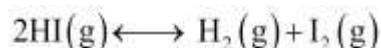
$$K_p = K_c (RT)^{\Delta n}$$

$$\Rightarrow 2.0 \times 10^{10} \text{ bar}^{-1} = K_c (0.0831 \text{ L bar K}^{-1} \text{ mol}^{-1} \times 450 \text{ K})^{-1}$$

$$\begin{aligned} \Rightarrow K_c &= \frac{2.0 \times 10^{10} \text{ bar}^{-1}}{(0.0831 \text{ L bar K}^{-1} \text{ mol}^{-1} \times 450 \text{ K})^{-1}} \\ &= (2.0 \times 10^{10} \text{ bar}^{-1}) (0.0831 \text{ L bar K}^{-1} \text{ mol}^{-1} \times 450 \text{ K}) \\ &= 74.79 \times 10^{10} \text{ L mol}^{-1} \\ &= 7.48 \times 10^{11} \text{ L mol}^{-1} \\ &= 7.48 \times 10^{11} \text{ M}^{-1} \end{aligned}$$

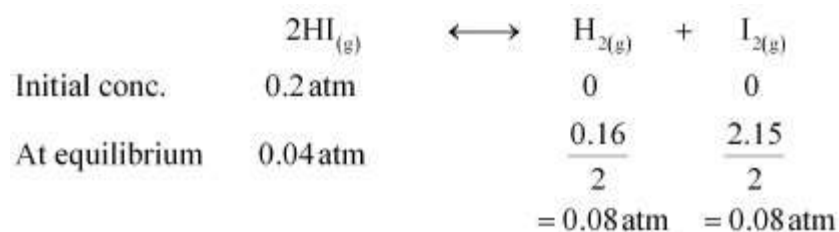
## Question 7.11:

A sample of  $\text{HI}_{(g)}$  is placed in flask at a pressure of 0.2 atm. At equilibrium the partial pressure of  $\text{HI}_{(g)}$  is 0.04 atm. What is  $K_p$  for the given equilibrium?

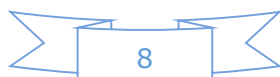


## Answer 7.11:

The initial concentration of HI is 0.2 atm. At equilibrium, it has a partial pressure of 0.04 atm. Therefore, a decrease in the pressure of HI is  $0.2 - 0.04 = 0.16$ . The given reaction is:



Therefore,





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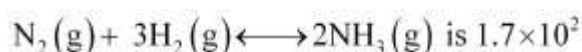
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$$\begin{aligned}K_p &= \frac{P_{H_2} \times P_{I_2}}{P_{HI}^2} \\&= \frac{0.08 \times 0.08}{(0.04)^2} \\&= \frac{0.0064}{0.0016} \\&= 4.0\end{aligned}$$

Hence, the value of  $K_p$  for the given equilibrium is 4.0.

## Question 7.12:

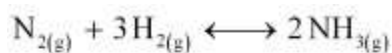
A mixture of 1.57 mol of  $N_2$ , 1.92 mol of  $H_2$  and 8.13 mol of  $NH_3$  is introduced into a 20 L reaction vessel at 500 K. At this temperature, the equilibrium constant,  $K_c$  for the reaction



Is the reaction mixture at equilibrium? If not, what is the direction of the net reaction?

## Answer 7.12:

The given reaction is:



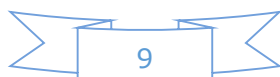
The given concentration of various species is

$$[N_2] = \frac{1.57}{20} \text{ mol L}^{-1} \quad [H_2] = \frac{1.92}{20} \text{ mol L}^{-1}$$

$$[NH_3] = \frac{8.13}{20} \text{ mol L}^{-1}$$

Now, reaction quotient  $Q_c$  is:

$$\begin{aligned}Q_c &= \frac{[NH_3]^2}{[N_2][H_2]^3} \\&= \frac{\left(\frac{8.13}{20}\right)^2}{\left(\frac{1.57}{20}\right)\left(\frac{1.92}{20}\right)^3} \\&= 2.4 \times 10^3\end{aligned}$$



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Since,  $Q_c \neq K_c$ , the reaction mixture is not at equilibrium.  
Again,  $Q_c > K_c$ . Hence, the reaction will proceed in the reverse direction.

## Question 7.13:

The equilibrium constant expression for a gas reaction is,

$$K_c = \frac{[\text{NH}_3]^4 [\text{O}_2]^5}{[\text{NO}]^4 [\text{H}_2\text{O}]^6}$$

Write the balanced chemical equation corresponding to this expression.

## Answer 7.13:

The balanced chemical equation corresponding to the given expression can be written as:



## Question 7.14:

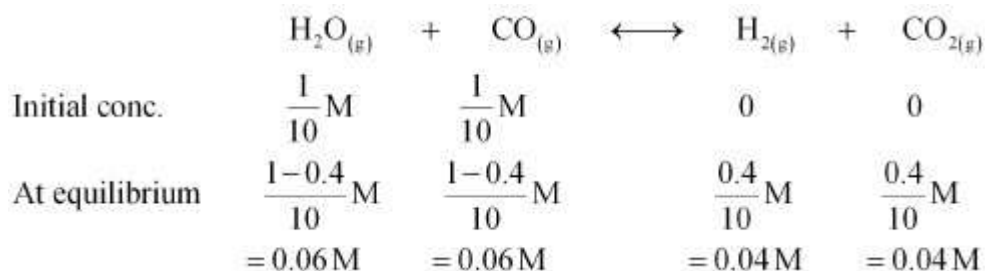
One mole of  $\text{H}_2\text{O}$  and one mole of  $\text{CO}$  are taken in 10 L vessel and heated to 725 K. At equilibrium 40% of water (by mass) reacts with  $\text{CO}$  according to the equation,



Calculate the equilibrium constant for the reaction.

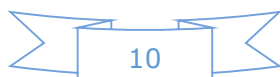
## Answer 7.14:

The given reaction is:



Therefore, the equilibrium constant for the reaction,

$$\begin{aligned}K_c &= \frac{[\text{H}_2][\text{CO}_2]}{[\text{H}_2\text{O}][\text{CO}]} \\&= \frac{0.04 \times 0.04}{0.06 \times 0.06} \\&= 0.444 \text{ (approximately)}\end{aligned}$$

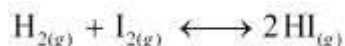


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## Question 7.15:

At 700 K, equilibrium constant for the reaction



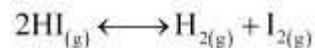
is 54.8. If  $0.5 \text{ mol L}^{-1}$  of  $\text{HI}_{(g)}$  is present at equilibrium at 700 K, what are the concentration of  $\text{H}_{2(g)}$  and  $\text{I}_{2(g)}$  assuming that we initially started with  $\text{HI}_{(g)}$  and allowed it to reach equilibrium at 700 K?

## Answer 7.15:

It is given that equilibrium constant  $K_c$  for the reaction



Therefore, at equilibrium, the equilibrium constant  $K'_c$  for the reaction



$[\text{HI}] = 0.5 \text{ mol L}^{-1}$  will be  $1/54.8$ .

Let the concentrations of hydrogen and iodine at equilibrium be  $x \text{ mol L}^{-1}$

$$[\text{H}_2] = [\text{I}_2] = x \text{ mol L}^{-1}$$

$$\text{Therefore, } \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = K'_c$$

$$\Rightarrow \frac{x \times x}{(0.5)^2} = \frac{1}{54.8}$$

$$\Rightarrow x^2 = \frac{0.25}{54.8}$$

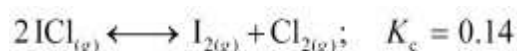
$$\Rightarrow x = 0.06754$$

$$x = 0.068 \text{ mol L}^{-1} \text{ (approximately)}$$

Hence, at equilibrium,  $[\text{H}_2] = [\text{I}_2] = 0.068 \text{ mol L}^{-1}$ .

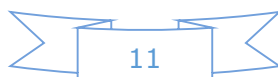
## Question 7.16:

What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of  $\text{ICl}$  was  $0.78 \text{ M}$ ?



## Answer 7.16:

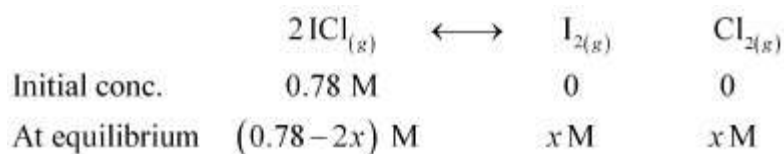
The given reaction is:



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Now, we can write,  $\frac{[\text{I}_2][\text{Cl}_2]}{[\text{ICl}]^2} = K_c$

$$\Rightarrow \frac{x \times x}{(0.78 - 2x)^2} = 0.14$$

$$\Rightarrow \frac{x^2}{(0.78 - 2x)^2} = 0.14$$

$$\Rightarrow \frac{x}{0.78 - 2x} = 0.374$$

$$\Rightarrow x = 0.292 - 0.748x$$

$$\Rightarrow 1.748x = 0.292$$

$$\Rightarrow x = 0.167$$

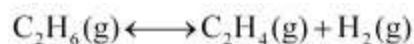
Hence, at equilibrium,

$$[\text{H}_2] = [\text{I}_2] = 0.167 \text{ M}$$

$$[\text{HI}] = (0.78 - 2 \times 0.167) \text{ M}$$
$$= 0.446 \text{ M}$$

## Question 7.17:

$K_p = 0.04$  atm at 899 K for the equilibrium shown below. What is the equilibrium concentration of  $\text{C}_2\text{H}_6$  when it is placed in a flask at 4.0 atm pressure and allowed to come to equilibrium?



## Answer 7.17:

Let  $p$  be the pressure exerted by ethene and hydrogen gas (each) at equilibrium. Now, according to the reaction,

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	$C_2H_6(g)$	$\longleftrightarrow$	$C_2H_4(g)$	+	$H_2(g)$
Initial conc.	4.0atm		0		0
At equilibrium	$4.0 - p$		$p$		$p$

We can write,

$$\frac{P_{C_2H_4} \times P_{H_2}}{P_{C_2H_6}} = K_p$$

$$\Rightarrow \frac{p \times p}{4.0 - p} = 0.04$$

$$\Rightarrow p^2 = 0.16 - 0.04p$$

$$\Rightarrow p^2 + 0.04p - 0.16 = 0$$

$$\begin{aligned} \text{Now, } p &= \frac{-0.04 \pm \sqrt{(0.04)^2 - 4 \times 1 \times (-0.16)}}{2 \times 1} \\ &= \frac{-0.04 \pm 0.80}{2} \\ &= \frac{0.76}{2} \quad (\text{Taking positive value}) \\ &= 0.38 \end{aligned}$$

Hence, at equilibrium,

$$\begin{aligned} [C_2H_6] - 4 - p &= 4 - 0.38 \\ &= 3.62 \text{ atm} \end{aligned}$$

## Question 7.18:

Ethyl acetate is formed by the reaction between ethanol and acetic acid and the equilibrium is represented as:



- Write the concentration ratio (reaction quotient),  $Q_c$ , for this reaction (note: water is not in excess and is not a solvent in this reaction)
- At 293 K, if one starts with 1.00 mol of acetic acid and 0.18 mol of ethanol, there is 0.171 mol of ethyl acetate in the final equilibrium mixture. Calculate the equilibrium constant.

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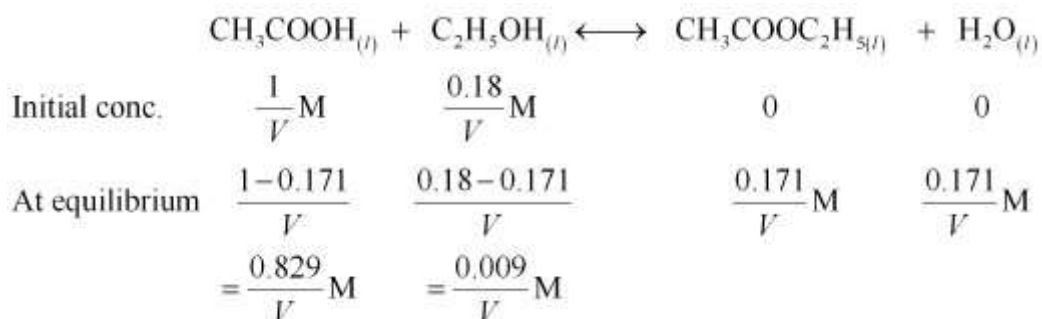
(iii) Starting with 0.5 mol of ethanol and 1.0 mol of acetic acid and maintaining it at 293 K, 0.214 mol of ethyl acetate is found after sometime. Has equilibrium been reached?

**Answer 7.18:**

(i) Reaction quotient,  $Q_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]}$

(ii) Let the volume of the reaction mixture be  $V$ . Also, here we will consider that water is a solvent and is present in excess.

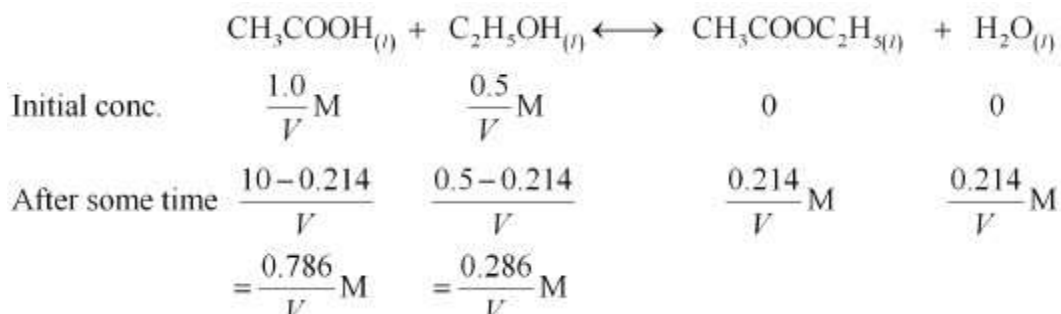
The given reaction is:



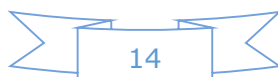
Therefore, equilibrium constant for the given reaction is:

$$\begin{aligned} K_c &= \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]} \\ &= \frac{\frac{0.171}{V} \times \frac{0.171}{V}}{\frac{0.829}{V} \times \frac{0.009}{V}} = 3.919 \\ &= 3.92 \text{ (approximately)} \end{aligned}$$

(iii) Let the volume of the reaction mixture be  $V$ .



Therefore, the reaction quotient is,



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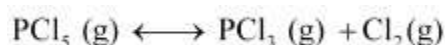
(Class – XI)

$$\begin{aligned} Q_c &= \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]} \\ &= \frac{\frac{0.214}{V} \times \frac{0.214}{V}}{\frac{0.786}{V} \times \frac{0.286}{V}} \\ &= 0.2037 \\ &= 0.204 \text{ (approximately)} \end{aligned}$$

Since  $Q_c < K_c$ , equilibrium has not been reached.

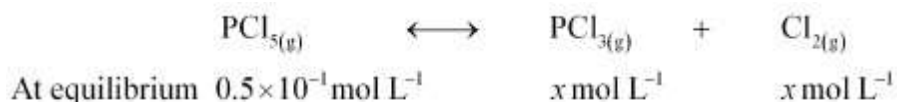
### Question 7.19:

A sample of pure  $\text{PCl}_5$  was introduced into an evacuated vessel at 473 K. After equilibrium was attained, concentration of  $\text{PCl}_5$  was found to be  $0.5 \times 10^{-1} \text{ mol L}^{-1}$ . If value of  $K_c$  is  $8.3 \times 10^{-3}$ , what are the concentrations of  $\text{PCl}_3$  and  $\text{Cl}_2$  at equilibrium?



### Answer 7.19:

Let the concentrations of both  $\text{PCl}_3$  and  $\text{Cl}_2$  at equilibrium be  $x \text{ mol L}^{-1}$ . The given reaction is:



It is given that the value of equilibrium constant,  $K_c$  is  $8.3 \times 10^{-3}$ .

Now we can write the expression for equilibrium as:

$$\begin{aligned} \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} &= K_c \\ \Rightarrow \frac{x \times x}{0.5 \times 10^{-1}} &= 8.3 \times 10^{-3} \\ \Rightarrow x^2 &= 4.15 \times 10^{-4} \\ \Rightarrow x &= 2.04 \times 10^{-2} \\ &= 0.0204 \\ &= 0.02 \text{ (approximately)} \end{aligned}$$

Therefore, at equilibrium,

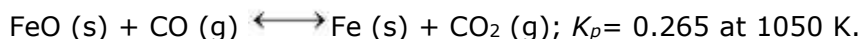
$$[\text{PCl}_3] = [\text{Cl}_2] = 0.02 \text{ mol L}^{-1}$$

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## Question 7.20:

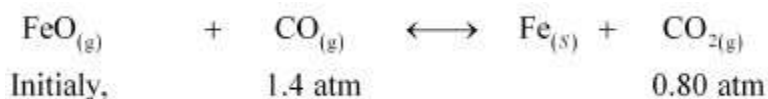
One of the reactions that takes place in producing steel from iron ore is the reduction of iron (II) oxide by carbon monoxide to give iron metal and CO<sub>2</sub>.



What are the equilibrium partial pressures of CO and CO<sub>2</sub> at 1050 K if the initial partial pressures are:  $p_{\text{CO}} = 1.4 \text{ atm}$  and  $p_{\text{CO}_2} = 0.80 \text{ atm}$ ?

## Answer 7.20:

For the given reaction,



$$\begin{aligned} Q_p &= \frac{P_{\text{CO}_2}}{P_{\text{CO}}} \\ &= \frac{0.80}{1.4} \\ &= 0.571 \end{aligned}$$

It is given that  $K_p = 0.265$ .

Since  $Q_p > K_p$ , the reaction will proceed in the backward direction.

Therefore, we can say that the pressure of CO will increase while the pressure of CO<sub>2</sub> will decrease.

Now, let the increase in pressure of CO = decrease in pressure of CO<sub>2</sub> be  $p$ . Then, we can write,

$$\begin{aligned} K_p &= \frac{P_{\text{CO}_2}}{P_{\text{CO}}} \\ \Rightarrow 0.265 &= \frac{0.80 - p}{1.4 + p} \\ \Rightarrow 0.371 + 0.265 p &= 0.80 - p \\ \Rightarrow 1.265 p &= 0.429 \\ \Rightarrow p &= 0.339 \text{ atm} \end{aligned}$$

Therefore, equilibrium partial of CO<sub>2</sub>,  $p_{\text{CO}_2} = 0.80 - 0.339 = 0.461 \text{ atm}$ .

And, equilibrium partial pressure of CO,  $p_{\text{CO}} = 1.4 + 0.339 = 1.739 \text{ atm}$ .