DCAM classes
Dynamic Classes for Academic Mastery

Chemistry

(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.

a) What is the initial effect of the change on vapour pressure?

b) How do rates of evaporation and condensation change initially?

c) 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: $[SO_2]$= 0.60 M, $[O_2]$ = 0.82 M and $[SO_3]$ = 1.90 M ?

$$2SO_2(g)+O_2(g) \longleftrightarrow 2SO_3(g)$$

Answer 7.2:

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

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$
$$= \frac{(1.90)^2 M^2}{(0.60)^2(0.821)M^3}$$
$$= 12.239\, M^{-1} (\text{approximately})$$

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

Question 7.3:

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

$$I_2\ (g) \longleftrightarrow 2I\ (g)$$

Calculate K_p for the equilibrium.

Answer 7.3:

Partial pressure of I atoms,

$$p_I = \frac{40}{100} \times p_{total}$$

$$= \frac{40}{100} \times 10^5$$
$$= 4 \times 10^4 \text{ Pa}$$

Partial pressure of I_2 molecules,

$$p_{I_2} = \frac{60}{100} \times p_{total}$$

$$= \frac{60}{100} \times 10^5$$
$$= 6 \times 10^4 \text{ Pa}$$

Now, for the given reaction,

$$K_p = \frac{(pI)^2}{p_{I_2}}$$

$$= \frac{(4 \times 10^4)^2 \text{ Pa}^2}{6 \times 10^4 \text{ Pa}}$$
$$= 2.67 \times 10^4 \text{ Pa}$$

Question 7.4:

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

(i) $2NOCl(g) \longleftrightarrow 2NO(g) + Cl_2(g)$

(ii) $2Cu(NO_3)_2(s) \longleftrightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$

(iii) $CH_3COOC_2H_5(aq) + H_2O(l) \longleftrightarrow CH_3COOH(aq) + C_2H_5OH(aq)$

(iv) $Fe^{3+}(aq) + 3OH^-(aq) \longleftrightarrow Fe(OH)_3(s)$

(v) $I_2(s) + 5F_2 \longleftrightarrow 2IF_5$

Answer 7.4:

(i) $$K_c = \frac{\left[NO_{(g)}\right]^2\left[Cl_{2(g)}\right]}{\left[NOCl_{(g)}\right]^2}$$

(ii) $$K_c = \frac{\left[CuO_{(s)}\right]^2\left[NO_{2(g)}\right]^4\left[O_{2(g)}\right]}{\left[Cu(NO_3)_{2(s)}\right]^2}$$

$$= \left[NO_{2(g)}\right]^4\left[O_{2(g)}\right]$$

(iii) $$K_c = \frac{\left[CH_3COOH_{(aq)}\right]\left[C_2H_5OH_{(aq)}\right]}{\left[CH_3COOC_2H_{5(aq)}\right]\left[H_2O_{(l)}\right]} = \frac{\left[CH_3COOH_{(aq)}\right]\left[C_2H_5OH_{(aq)}\right]}{\left[CH_3COOC_2H_{5(aq)}\right]}$$

(iv) $$K_c = \frac{\left[Fe(OH)_{3(s)}\right]}{\left[Fe^{3+}{}_{(aq)}\right]\left[OH^-_{(aq)}\right]^3}$$

$$= \frac{1}{\left[Fe^{3+}{}_{(aq)}\right]\left[OH^-_{(aq)}\right]^3}$$

(v) $$K_c = \frac{[IF_5]^2}{\left[I_{2(s)}\right][F_2]^5}$$

$$= \frac{[IF_5]^2}{[F_2]^5}$$

Question 7.5:

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

(i) $2NOCl(g) \longleftrightarrow 2NO(g) + Cl_2(g); \quad K_p = 1.8 \times 10^{-2}$ at 500 K

(ii) $CaCO_3(s) \longleftrightarrow CaO(s) + CO_2(g); \quad K_p = 167$ at 1073 K

Answer 7.5:

The relation between K_p and K_c is given as:

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

(a) Here,

$\Delta n = 3 - 2 = 1$

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

$T = 500$ K

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

Now,

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

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

$$\Rightarrow K_c = \frac{1.8 \times 10^{-2}}{0.0831 \times 500}$$

$$= 4.33 \times 10^{-4}\ (\text{approximately})$$

(b) Here,

$\Delta n = 2 - 1 = 1$

$R = 0.0831$ barLmol^{-1}K^{-1} T

= 1073 K

$K_p = 167$

Now,

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

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

$$\Rightarrow K_c = \frac{167}{0.0831 \times 1073}$$

$$= 1.87\ (\text{approximately})$$

Question 7.6:

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

$$NO(g) + O_3(g) \longleftrightarrow NO_2(g) + O_2(g)$$

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 6.3×10^{14}.

Then, K_c for the reverse reaction will be,

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

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),

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

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.

Question 7.8:

Reaction between N_2 and O_2 takes place as follows:

$$2N_2(g) \quad + \quad O_2(g) \quad \longleftrightarrow \quad 2N_2O(g)$$

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 × 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:

	$2N_{2(g)}$	+	$O_{2(g)}$	$\longleftrightarrow$	$2N_2O_{(g)}$
Initial conc.	0.482 mol		0.933 mol		0
At equilibrium	$(0.482-x)$ mol		$(1.933-x)$ mol		x mol

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 × 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,

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

Now,

$$K_C=\frac{\left[N_2O_{(g)}\right]^2}{\left[N_{2(g)}\right]^2\left[O_{2(g)}\right]}$$

$$\Rightarrow 2.0\times10^{-37}=\frac{\left(\frac{x}{10}\right)^2}{(0.0482)^2(0.0933)}$$

$$\Rightarrow \frac{x^2}{100}=2.0\times10^{-37}\times(0.0482)^2\times(0.0933)$$

$$\Rightarrow x^2=43.35\times10^{-40}$$

$$\Rightarrow x=6.6\times10^{-20}$$

$$[N_2O]=\frac{x}{10}=\frac{6.6\times10^{-20}}{10}$$

$$=6.6\times10^{-21}$$

Question 7.9:

Nitric oxide reacts with Br_2 and gives nitrosyl bromide as per reaction given below:

$$2NO(g) + Br_2(g) \longleftrightarrow 2NOBr(g)$$

When 0.087 mol of NO and 0.0437 mol of Br_2 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_2.

Answer 7.9:

The given reaction is:

$$\underset{2\text{mol}}{2NO_{(g)}} + \underset{1\text{mol}}{Br_{2(g)}} \longleftrightarrow \underset{2\text{mol}}{2NOBr_{(g)}}$$

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:

[NO] = 0.087 mol [Br_2] = 0.0437 mol

Therefore, the amount of NO present at equilibrium is:

[NO] = 0.087 – 0.0518 = 0.0352 mol

And, the amount of Br present at equilibrium is:

[Br_2] = 0.0437 – 0.0259 = 0.0178 mol

Question 7.10:

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

$$2SO_{2(g)} + O_{2(g)} \longleftrightarrow 2SO_{3(g)}$$

What is K_c at this temperature?

Answer 7.10:

For the given reaction,

Δn = 2 – 3 = – 1

T = 450 K

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

K_p = 2.0 × 10^{10} bar $^{-1}$

We know that,

$$K_P = K_C \left(RT\right)\Delta n$$

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

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

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

$$= 74.79 \times 10^{10}\ \text{L mol}^{-1}$$

$$= 7.48 \times 10^{11}\ \text{L mol}^{-1}$$

$$= 7.48 \times 10^{11}\ \text{M}^{-1}$$

Question 7.11:

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

$$2\text{HI}(g) \longleftrightarrow \text{H}_2(g) + \text{I}_2(g)$$

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:

	$2HI_{(g)}$	$\longleftrightarrow$	$H_{2(g)}$	+	$I_{2(g)}$
Initial conc.	0.2 atm		0		0
At equilibrium	0.04 atm		$\frac{0.16}{2}$ $= 0.08\,\text{atm}$		$\frac{2.15}{2}$ $= 0.08\,\text{atm}$

Therefore,

$$K_p = \frac{p_{H_2} \times p_{I_2}}{p_{HI}^2}$$
$$= \frac{0.08 \times 0.08}{(0.04)^2}$$
$$= \frac{0.0064}{0.0016}$$
$$= 4.0$$

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

$$N_2(g) + 3H_2(g) \longleftrightarrow 2NH_3(g) \text{ is } 1.7 \times 10^2$$

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

Answer 7.12:

The given reaction is:

$$N_{2(g)} + 3H_{2(g)} \longleftrightarrow 2NH_{3(g)}$$

The given concentration of various species is

$$[N_2] = \frac{1.57}{20} \text{ mol L}^{-1} \qquad [H_2] = \frac{1.92}{20} \text{ mol L}^{-1}$$
$$[NH_3] = \frac{8.13}{20} \text{ mol L}^{-1}$$

Now, reaction quotient Q_c is:

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

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,

$$K_C = \frac{[NH_3]^4[O_2]^5}{[NO]^4[H_2O]^6}$$

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:

$$4NO_{(g)} + 6H_2O_{(g)} \longleftrightarrow 4NH_{3(g)} + 5O_{2(g)}$$

Question 7.14:

One mole of H_2O and one mole of CO are taken in 10 L vessel and heated to 725 K. At equilibrium 40% of water (by mass) reacts with CO according to the equation,

$$H_2O(g) + CO(g) \longleftrightarrow H_2(g) + CO_2(g)$$

Calculate the equilibrium constant for the reaction.

Answer 7.14:

The given reaction is:

	$H_2O_{(g)}$	+ $CO_{(g)}$	$\longleftrightarrow$	$H_{2(g)}$	+ $CO_{2(g)}$
Initial conc.	$\frac{1}{10}$ M	$\frac{1}{10}$ M		0	0
At equilibrium	$\frac{1-0.4}{10}$ M $= 0.06$ M	$\frac{1-0.4}{10}$ M $= 0.06$ M		$\frac{0.4}{10}$ M $= 0.04$ M	$\frac{0.4}{10}$ M $= 0.04$ M

Therefore, the equilibrium constant for the reaction,

$$K_C = \frac{[H_2][CO_2]}{[H_2O][CO]}$$
$$= \frac{0.04 \times 0.04}{0.06 \times 0.06}$$
$$= 0.444 \text{ (approximately)}$$

Question 7.15:

At 700 K, equilibrium constant for the reaction

$$H_{2(g)} + I_{2(g)} \longleftrightarrow 2HI_{(g)}$$

is 54.8. If 0.5 molL^{-1} of $HI_{(g)}$ is present at equilibrium at 700 K, what are the concentration of $H_{2(g)}$ and $I_{2(g)}$ assuming that we initially started with $HI_{(g)}$ and allowed it to reach equilibrium at 700 K?

Answer 7.15:

It is given that equilibrium constant K_c for the reaction

$H_{2(g)} + I_{2(g)} \longleftrightarrow 2HI_{(g)}$ is 54.8.

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

$2HI_{(g)} \longleftrightarrow H_{2(g)} + I_{2(g)}$

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

Let the concentrations of hydrogen and iodine at equilibrium be x molL^{-1}

$$[H_2] = [I_2] = x\,\text{mol L}^{-1}$$

$$\text{Therefore, } \frac{[H_2][I_2]}{[HI]^2} = K'_c$$

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

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

$$\Rightarrow x = 0.06754$$

$$x = 0.068\,\text{molL}^{-1}\,(\text{approximately})$$

Hence, at equilibrium, $[H_2] = [I_2] = 0.068\ \text{mol L}^{-1}$.

Question 7.16:

What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78 M?

$$2ICl_{(g)} \longleftrightarrow I_{2(g)} + Cl_{2(g)}; \quad K_c = 0.14$$

Answer 7.16:

The given reaction is:

	$2ICl_{(g)}$	$\longleftrightarrow$	$I_{2(g)}$	$Cl_{2(g)}$
Initial conc.	0.78 M		0	0
At equilibrium	$(0.78-2x)$ M		xM	xM

Now, we can write, $\frac{[I_2][Cl_2]}{[ICl]^2} = K_C$

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

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

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

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

$$\Rightarrow 1.748x = 0.292$$

$$\Rightarrow x = 0.167$$

Hence, at equilibrium,

$$[H_2] = [I_2] = 0.167 \text{ M}$$

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

$$= 0.446 \text{M}$$

Question 7.17:

K_p = 0.04 atm at 899 K for the equilibrium shown below. What is the equilibrium concentration of C_2H_6 when it is placed in a flask at 4.0 atm pressure and allowed to come to equilibrium?

$$C_2H_6(g) \longleftrightarrow C_2H_4(g) + H_2(g)$$

Answer 7.17:

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

	$C_2H_{6(g)}$ $\longleftrightarrow$	$C_2H_{4(g)}$ +	$H_{2(g)}$
Initial conc.	4.0atm	0	0
At equilibrium	$4.0-p$	p	p

We can write,

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

$$\Rightarrow \frac{p \times p}{40-p} = 0.04$$

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

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

$$\text{Now, } p = \frac{-0.04 \pm \sqrt{(0.04)^2 - 4 \times 1 \times (-0.16)}}{2 \times 1}$$

$$= \frac{-0.04 \pm 0.80}{2}$$

$$= \frac{0.76}{2} \qquad \text{(Taking positive value)}$$

$$= 0.38$$

Hence, at equilibrium,

$$[C_2H_6] - 4 - p = 4 - 0.38$$

$$= 3.62 \text{ atm}$$

Question 7.18:

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

$$CH_3COOH\ (l) + C_2H_5OH\ (l) \longleftrightarrow CH_3COOC_2H_5\ (l) + H_2O\ (l)$$

(i) Write the concentration ratio (reaction quotient), *Qc*, for this reaction (note: water is not in excess and is not a solvent in this reaction)

(ii) 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.

(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, $Q_C = \dfrac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]}$

(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:

	$CH_3COOH_{(l)}$ +	$C_2H_5OH_{(l)} \longleftrightarrow$	$CH_3COOC_2H_{5(l)}$ +	$H_2O_{(l)}$
Initial conc.	$\frac{1}{V}$ M	$\frac{0.18}{V}$ M	0	0
At equilibrium	$\frac{1-0.171}{V} = \frac{0.829}{V}$ M	$\frac{0.18-0.171}{V} = \frac{0.009}{V}$ M	$\frac{0.171}{V}$ M	$\frac{0.171}{V}$ M

Therefore, equilibrium constant for the given reaction is:

$$K_C = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]}$$

$$= \frac{\frac{0.171}{V} \times \frac{0.171}{V}}{\frac{0.829}{V} \times \frac{0.009}{V}} = 3.919$$

$$= 3.92 \left(\text{approximately}\right)$$

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

	$CH_3COOH_{(l)}$ +	$C_2H_5OH_{(l)} \longleftrightarrow$	$CH_3COOC_2H_{5(l)}$ +	$H_2O_{(l)}$
Initial conc.	$\frac{1.0}{V}$ M	$\frac{0.5}{V}$ M	0	0
After some time	$\frac{10-0.214}{V} = \frac{0.786}{V}$ M	$\frac{0.5-0.214}{V} = \frac{0.286}{V}$ M	$\frac{0.214}{V}$ M	$\frac{0.214}{V}$ M

Therefore, the reaction quotient is,

$$Q_c = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]}$$

$$= \frac{\frac{0.214}{V} \times \frac{0.214}{V}}{\frac{0.786}{V} \times \frac{0.286}{V}}$$

$$= 0.2037$$

$$= 0204 \text{ (approximately)}$$

Since $Q_C < K_C$, equilibrium has not been reached.

Question 7.19:

A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was attained, concentration of PCl_5 was found to be 0.5×10^{-1} mol L^{-1}. If value of K_c is 8.3×10^{-3}, what are the concentrations of PCl_3 and Cl_2 at equilibrium?

$$PCl_5 \text{ (g)} \longleftrightarrow PCl_3 \text{ (g)} + Cl_2\text{(g)}$$

Answer 7.19:

Let the concentrations of both PCl_3 and Cl_2 at equilibrium be x molL^{-1}. The given reaction is:

	$PCl_{5(g)}$	$\longleftrightarrow$	$PCl_{3(g)}$	+	$Cl_{2(g)}$
At equilibrium	0.5×10^{-1} mol L^{-1}		x mol L^{-1}		x mol L^{-1}

It is given that the value of equilibrium constant, K_C is 8.3×10^{-3}.

Now we can write the expression for equilibrium as:

$$\frac{[PCl_3][Cl_2]}{[PCl_5]} = K_C$$

$$\Rightarrow \frac{x \times x}{0.5 \times 10^{-1}} = 8.3 \times 10^{-3}$$

$$\Rightarrow x^2 = 4.15 \times 10^{-4}$$

$$\Rightarrow x = 2.04 \times 10^{-2}$$

$$= 0.0204$$

$$= 0.02 \text{(approximately)}$$

Therefore, at equilibrium,

$$[PCl_3] = [Cl_2] = 0.02 \text{ mol } L^{-1}.$$

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_2.

FeO (s) + CO (g) $\longleftrightarrow$ Fe (s) + CO_2 (g); K_p= 0.265 at 1050 K.

What are the equilibrium partial pressures of CO and CO_2 at 1050 K if the initial partial pressures are: p_{CO} = 1.4 atm and p_{CO_2}= 0.80 atm?

Answer 7.20:

For the given reaction,

$FeO_{(g)}$	+	$CO_{(g)}$	$\longleftrightarrow$	$Fe_{(s)}$	+	$CO_{2(g)}$
Initialy,		1.4 atm				0.80 atm

$$Q_P = \frac{p_{CO_2}}{p_{CO}} = \frac{0.80}{1.4} = 0.571$$

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_2 will decrease.

Now, let the increase in pressure of CO = decrease in pressure of CO_2 be *p.* Then, we can write,

$$K_P = \frac{p_{CO_2}}{p_{CO}}$$
$$\Rightarrow 0.265 = \frac{0.80 - p}{1.4 + p}$$
$$\Rightarrow 0.371 + 0.265\,p = 0.80 - p$$
$$\Rightarrow 1.265\,p = 0.429$$
$$\Rightarrow p = 0.339 \text{ atm}$$

Therefore, equilibrium partial of $CO_2, p_{CO_2} = 0.80 - 0.339 = 0.461$ atm.

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

Question 7.21:

Equilibrium constant, K_c for the reaction

$N_2(g) + 3H_2(g) \longleftrightarrow 2NH_3(g)$ at 500 K is 0.061.

At a particular time, the analysis shows that composition of the reaction mixture is 3.0 mol L^{-1} N_2, 2.0 mol L^{-1} H_2 and 0.5 mol L^{-1} NH_3. Is the reaction at equilibrium? If not in which direction does the reaction tend to proceed to reach equilibrium?

Answer 7.21:

The given reaction is:

$$\begin{array}{lccc} & N_{2(g)} + & 3H_{2(g)} \longleftrightarrow & 2NH_{3(g)} \\ \text{At a particular time:} & 3.0\,\text{mol L}^{-1} & 2.0\,\text{mol L}^{-1} & 0.5\,\text{mol L}^{-1} \end{array}$$

Now, we know that,

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$
$$= \frac{(0.5)^2}{(3.0)(2.0)^3}$$
$$= 0.0104$$

It is given that $K_c = 0.061$.

Since $Q_c \neq K_c$, the reaction is not at equilibrium.

Since $Q_c < K_c$, the reaction will proceed in the forward direction to reach equilibrium.

Question 7.22:

Bromine monochloride, BrCl decomposes into bromine and chlorine and reaches the equilibrium:

$$2BrCl\,(g) \longleftrightarrow Br_2\,(g) + Cl_2\,(g)$$

for which K_c= 32 at 500 K.

If initially pure BrCl is present at a concentration of 3.3×10^{-3} molL^{-1}, what is its molar concentration in the mixture at equilibrium?

Answer 7.22:

Let the amount of bromine and chlorine formed at equilibrium be *x*. The given reaction is:

	$2BrCl_{(g)}$	$\longleftrightarrow$	$Br_{2(g)}$	+	$Cl_{2(g)}$
Initial conc.	3.3×10^{-3}		0		0
At equilibrium	$3.3\times10^{-3}-2x$		x		x

Now, we can write,

$$\frac{[Br_2][Cl_2]}{[BrCl]^2}=K_C$$

$$\Rightarrow\frac{x\times x}{\left(3.3\times10^{-3}-2x\right)^2}=32$$

$$\Rightarrow\frac{x}{3.3\times10^{-3}-2x}=5.66$$

$$\Rightarrow x=18.678\times10^{-3}-11.32x$$

$$\Rightarrow 12.32x=18.678\times10^{-3}$$

$$\Rightarrow x=1.5\times10^{-3}$$

Therefore, at equilibrium,

$$\begin{aligned}[BrCl]&=3.3\times10^{-3}-\left(2\times1.5\times10^{-3}\right)\\&=3.3\times10^{-3}-3.0\times10^{-3}\\&=0.3\times10^{-3}\\&=3.0\times10^{-4}\text{ molL}^{-1}\end{aligned}$$

Question 7.23:

At 1127 K and 1 atm pressure, a gaseous mixture of CO and CO_2 in equilibrium with solid carbon has 90.55% CO by mass

$$C\ (s)+CO_2\ (g)\longleftrightarrow 2CO\ (g)$$

Calculate K_c for this reaction at the above temperature.

Answer 7.23:

Let the total mass of the gaseous mixture be 100 g.

Mass of CO = 90.55 g

And, mass of CO_2 = (100 – 90.55) = 9.45 g

Now, number of moles of CO, $n_{CO} = \frac{90.55}{28} = 3.234 \text{ mol}$

Number of moles of CO_2, $n_{CO_2} = \frac{9.45}{44} = 0.215 \text{ mol}$

Partial pressure of CO,

$$\begin{aligned} p_{CO} &= \frac{n_{CO}}{n_{CO} + n_{CO_2}} \times p_{total} \\ &= \frac{3.234}{3.234 + 0.215} \times 1 \\ &= 0.938 \text{ atm} \end{aligned}$$

Partial pressure of CO_2,

$$\begin{aligned} p_{CO_2} &= \frac{n_{CO_2}}{n_{CO} + n_{CO_2}} \times p_{total} \\ &= \frac{0.215}{3.234 + 0.215} \times 1 \\ &= 0.062 \text{ atm} \end{aligned}$$

$$\begin{aligned} \text{Therefore, } K_p &= \frac{[CO]^2}{[CO_2]} \\ &= \frac{(0.938)^2}{0.062} \\ &= 14.19 \end{aligned}$$

For the given reaction,

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

We know that,

$$\begin{aligned} & K_p = K_c (RT)^{\Delta n} \\ & \Rightarrow 14.19 = K_c (0.082 \times 1127)^1 \\ & \Rightarrow K_c = \frac{14.19}{0.082 \times 1127} \\ & = 0.154 (\text{approximately}) \end{aligned}$$

Question 7.24:

Calculate

a) $\Delta G°$ and

b) the equilibrium constant for the formation of NO_2 from NO and O_2 at 298 K

$$NO\ (g) + \tfrac{1}{2} O_2\ (g) \longleftrightarrow NO_2\ (g)$$

Where:

$\Delta_f G°\ (NO_2) = 52.0$ kJ/mol

$\Delta_f G°\ (NO) = 87.0$ kJ/mol

$\Delta_f G°\ (O_2) = 0$ kJ/mol

Answer 7.24:

(a) For the given reaction,

$\Delta G° = \Delta G°(\text{ Products}) - \Delta G°(\text{ Reactants})$

$\Delta G° = 52.0 - \{87.0 + 0\}$

$= -\ 35.0 \text{ kJ mol}^{-1}$

(b) We know that,

$\Delta G° = \text{RT} \log K_c$

$\Delta G° = 2.303 \text{ RT} \log K_c$

$$K_c = \frac{-35.0 \times 10^{-3}}{-2.303 \times 8.314 \times 298}$$
$$= 6.134$$
$$\therefore K_c = \text{antilog}\ (6.134)$$
$$= 1.36 \times 10^6$$

Hence, the equilibrium constant for the given reaction K_c is 1.36×10^6

Question 7.25:

Does the number of moles of reaction products increase, decrease or remain same when each of the following equilibria is subjected to a decrease in pressure by increasing the volume?

(a) $PCl_5\ (g) \longleftrightarrow PCl_3\ (g) + Cl_2\ (g)$

(b) $CaO\,(s) + CO_2\,(g) \longleftrightarrow CaCO_3\,(s)$

(c) $3Fe\,(s) + 4H_2O\,(g) \longleftrightarrow Fe_3O_4\,(s) + 4H_2\,(g)$

Answer 7.25:

(a) The number of moles of reaction products will increase. According to Le Chatelier's principle, if pressure is decreased, then the equilibrium shifts in the direction in which the number of moles of gases is more. In the given reaction, the number of moles of gaseous products is more than that of gaseous reactants. Thus, the reaction will proceed in the forward direction. As a result, the number of moles of reaction products will increase.

(b) The number of moles of reaction products will decrease.

(c) The number of moles of reaction products remains the same.

Question 7.26:

Which of the following reactions will get affected by increasing the pressure?

Also, mention whether change will cause the reaction to go into forward or backward direction.

(i) $COCl_2\,(g) \longleftrightarrow CO\,(g) + Cl_2\,(g)$

(ii) $CH_4\,(g) + 2S_2\,(g) \longleftrightarrow CS_2\,(g) + 2H_2S\,(g)$

(iii) $CO_2\,(g) + C\,(s) \longleftrightarrow 2CO\,(g)$

(iv) $2H_2\,(g) + CO\,(g) \longleftrightarrow CH_3OH\,(g)$

(v) $CaCO_3\,(s) \longleftrightarrow CaO\,(s) + CO_2\,(g)$

(vi) $4\,NH_3\,(g) + 5O_2\,(g) \longleftrightarrow 4NO\,(g) + 6H_2O(g)$

Answer7.26:

The reactions given in (i), (iii), (iv), (v), and (vi) will get affected by increasing the pressure.

The reaction given in (iv) will proceed in the forward direction because the number of moles of gaseous reactants is more than that of gaseous products.

The reactions given in (i), (iii), (v), and (vi) will shift in the backward direction because the number of moles of gaseous reactants is less than that of gaseous products.

Question 7.27:

The equilibrium constant for the following reaction is 1.6 $\times 10^5$ at 1024 K.

$$H_2(g) + Br_2(g) \longleftrightarrow 2HBr(g)$$

Find the equilibrium pressure of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024 K.

Answer 7.27:

Given,

K_p for the reaction i.e., $H_{2(g)} + Br_{2(g)} \longleftrightarrow 2HBr_{(g)}$ is 1.6×10^5.

Therefore, for the reaction $2HBr_{(g)} \longleftrightarrow H_{2(g)} + Br_{2(g)}$, the equilibrium constant will be,

$$K'_p = \frac{1}{K_p}$$
$$= \frac{1}{1.6\times10^5}$$
$$= 6.25\times10^{-6}$$

Now, let p be the pressure of both H_2 and Br_2 at equilibrium.

	$2HBr_{(g)}$	$\longleftrightarrow$	$H_{2(g)}$	+	$Br_{2(g)}$
Initial conc.	10		0		0
At equilibrium	$10-2p$		p		p

Now, we can write,

$$\frac{p_{HBr}\times p_{_2}}{p_{HBr}^2} = K'_p$$

$$\frac{p\times p}{(10-2p)^2} = 6.25\times10^{-6}$$

$$\frac{p}{10-2p} = 2.5\times10^{-3}$$

$$p = 2.5\times10^{-2} - \left(5.0\times10^{-3}\right)p$$

$$p + \left(5.0\times10^{-3}\right)p = 2.5\times10^{-2}$$

$$\left(1005\times10^{-3}\right)p = 2.5\times10^{-2}$$

$$p = 2.49\times10^{-2}\text{ bar} = 2.5\times10^{-2}\text{ bar}\left(\text{approximately}\right)$$

Therefore, at equilibrium,

$$[H_2] = [Br_2] = 2.49 \times 10^{-2} \text{ bar}$$
$$[HBr] = 10 - 2 \times (2.49 \times 10^{-2}) \text{ bar}$$
$$= 9.95 \text{ bar} = 10 \text{ bar (approximately)}$$

Question 7.28:

Dihydrogen gas is obtained from natural gas by partial oxidation with steam as per following endothermic reaction:

$$CH_4\ (g) + H_2O\ (g) \longleftrightarrow CO\ (g) + 3H_2\ (g)$$

(a) Write as expression for K_p for the above reaction.

(b) How will the values of K_p and composition of equilibrium mixture be affected by
(i) Increasing the pressure

(ii) Increasing the temperature

(iii) Using a catalyst?

Answer 7.28:

(a) For the given reaction,

$$K_p = \frac{p_{CO} \times p_{H_2}^3}{p_{CH_4} \times p_{H_2O}}$$

(b)

(i) According to Le Chatelier's principle, the equilibrium will shift in the backward direction.

(ii) According to Le Chatelier's principle, as the reaction is endothermic, the equilibrium will shift in the forward direction.

(iii) The equilibrium of the reaction is not affected by the presence of a catalyst. A catalyst only increases the rate of a reaction. Thus, equilibrium will be attained quickly.

Question 7.29:

Describe the effect of:

a) Addition of H_2

b) Addition of CH_3OH

c) Removal of CO

d) Removal of CH_3OH on the equilibrium of the reaction:

$$2H_2(g) + CO(g) \longleftrightarrow CH_3OH(g)$$

Answer 7.29:

(a) According to Le Chatelier's principle, on addition of H_2, the equilibrium of the given reaction will shift in the forward direction.

(b) On addition of CH_3OH, the equilibrium will shift in the backward direction.

(c) On removing CO, the equilibrium will shift in the backward direction.

(d) On removing CH_3OH, the equilibrium will shift in the forward direction.

Question 7.30:

At 473 K, equilibrium constant K_c for decomposition of phosphorus pentachloride, PCl_5 is 8.3×10^{-3}. If decomposition is depicted as,

$$PCl_5(g) \longleftrightarrow PCl_3(g) + Cl_2(g) \qquad \Delta_r H° = 124.0 \text{ kJmol}^{-1}$$

a) Write an expression for K_c for the reaction.

b) What is the value of K_c for the reverse reaction at the same temperature?

c) What would be the effect on K_c if

(i) more PCl_5 is added

(ii) pressure is increased?

(iii) The temperature is increased?

Answer 7.30:

(a) $$K_c \frac{\left[PCl_{3(g)}\right]\left[Cl_{2(g)}\right]}{\left[PCl_{5(g)}\right]}$$

(b) Value of K_c for the reverse reaction at the same temperature is:

$$K'_c = \frac{1}{K_c}$$
$$= \frac{1}{8.3 \times 10^{-3}} = 1.2048 \times 10^2$$
$$= 120 - 48$$

(c) (i) K_c would remain the same because in this case, the temperature remains the same.

(ii) K_c is constant at constant temperature. Thus, in this case, K_c would not change.

(iii) In an endothermic reaction, the value of K_c increases with an increase in temperature. Since the given reaction in an endothermic reaction, the value of K_c will increase if the temperature is increased.

Question 7.31:

Dihydrogen gas used in Haber's process is produced by reacting methane from natural gas with high temperature steam. The first stage of two stage reaction involves the formation of CO and H_2. In second stage, CO formed in first stage is reacted with more steam in water gas shift reaction,

$$CO\ (g) + H_2O\ (g) \longleftrightarrow CO_2\ (g) + H_2\ (g)$$

If a reaction vessel at 400°C is charged with an equimolar mixture of CO and steam such that $p_{CO} = p_{H_2O} =$ 4.0 bar, what will be the partial pressure of H_2 at equilibrium? K_p= 10.1 at 400°C

Answer 7.31:

Let the partial pressure of both carbon dioxide and hydrogen gas be p. The given reaction is:

	$CO_{(g)}$	+	$H_2O_{(g)}$	$\longleftrightarrow$	$CO_{2(g)}$ +	$H_{2(g)}$
Initial conc.	4.0 bar		4.0 bar		0	0
At equilibrium	$4.0 - p$		$4.0 - p$		p	p

It is given that $K_p = 10.1$.

Now,

$$\frac{p_{CO_2} \times p_{H_2}}{p_{CO} \times p_{H_2O}} = K_p$$

$$\Rightarrow \frac{p \times p}{(4.0 - p)(4.0 - p)} = 10.1$$

$$\Rightarrow \frac{p}{4.0 - p} = 3.178$$

$$\Rightarrow p = 12.712 - 3.178p$$

$$\Rightarrow 4.178p = 12.712$$

$$\Rightarrow p = 3.04$$

Hence, at equilibrium, the partial pressure of H_2 will be 3.04 bar.

Question 7.32:

Predict which of the following reaction will have appreciable concentration of reactants and products:

a) $Cl_2\ (g) \longleftrightarrow 2\ Cl\ (g)\ ; \quad K_c = 5\ \times 10^{-39}$

b) $Cl_2\ (g) + 2\ NO\ (g) \longleftrightarrow 2\ NOCl\ (g)\ ; \quad K_c = 3.7 \times 10^{8}$

c) $Cl_2\ (g) + 2\ NO_2\ (g) \longleftrightarrow 2\ NO_2Cl\ (g)\ ; \quad K_c = 1.8$

Answer 7.32:

If the value of K_c lies between 10^{-3} and 10^{3}, a reaction has appreciable concentration of reactants and products. Thus, the reaction given in (c) will have appreciable concentration of reactants and products.

Question 7.33:

The value of K_c for the reaction $3O_2$ (g) $\longleftrightarrow$ $2O_3$ (g) is 2.0 $\times 10^{-50}$ at 25°C. If the equilibrium concentration of O_2 in air at 25°C is 1.6 $\times 10^{-2}$, what is the concentration of O_3?

Answer 7.33:

The given reaction is:

$$3O_{2(g)} \longleftrightarrow 2O_{3(g)}$$

$$\text{Then, } K_c = \frac{\left[O_{3(g)}\right]^2}{\left[O_{2(g)}\right]^3}$$

It is given that $K_C = 2.0 \times 10^{-50}$ and $\left[O_{2(g)}\right] = 1.6 \times 10^{-2}$.

Then, we have,

$$2.0 \times 10^{-50} = \frac{\left[O_{3(g)}\right]^2}{\left[1.6 \times 10^{-2}\right]^3}$$

$$\Rightarrow \left[O_{3(g)}\right]^2 = 2.0 \times 10^{-50} \times \left(1.6 \times 10^{-2}\right)^3$$

$$\Rightarrow \left[O_{3(g)}\right]^2 = 8.192 \times 10^{-56}$$

$$\Rightarrow \left[O_{3(g)}\right] = 2.86 \times 10^{-28}\ \text{M}$$

Hence, the concentration of O_3 is 2.86×10^{-28} M.

Question 7.34:

The reaction, $CO(g) + 3H_2(g) \longleftrightarrow CH_4(g) + H_2O(g)$ is at equilibrium at 1300 K in a 1L flask. It also contain 0.30 mol of CO, 0.10 mol of H_2 and 0.02 mol of H_2O and an unknown amount of CH_4 in the flask. Determine the concentration of CH_4 in the mixture.

The equilibrium constant, K_c for the reaction at the given temperature is 3.90.

Answer 7.34:

Let the concentration of methane at equilibrium be *x*.

$$\begin{array}{lccccccc} & CO_{(g)} & + & 3H_{2(g)} & \longleftrightarrow & CH_{4(g)} & + & H_2O_{(g)} \\ \text{At equilibrium} & \frac{0.3}{1} = 0.3\text{M} & & \frac{0.1}{1} = 0.1\text{M} & & x & & \frac{0.02}{1} = 0.02\text{M} \end{array}$$

It is given that $K_c = 3.90$.

Therefore,

$$\frac{\left[CH_{4(g)}\right]\left[H_2O_{(g)}\right]}{\left[CO_{(g)}\right]\left[H_{2(g)}\right]^3} = K_c$$

$$\Rightarrow \frac{x \times 0.02}{0.3 \times (0.1)^3} = 3.90$$

$$\Rightarrow x = \frac{3.90 \times 0.3 \times (0.1)^3}{0.02}$$

$$= \frac{0.00117}{0.02}$$

$$= 0.0585\text{M}$$

$$= 5.85 \times 10^{-2}\text{ M}$$

Hence, the concentration of CH_4 at equilibrium is 5.85×10^{-2} M.

Question 7.35:

What is meant by the conjugate acid-base pair? Find the conjugate acid/base for the following species:

$$HNO_2, CN^-, HClO_4, F^-, OH^-, CO_3^{2-}, \text{and } S^-$$

Answer 7.35:

A conjugate acid-base pair is a pair that differs only by one proton.
The conjugate acid-base for the given species is mentioned in the table below.

Species	Conjugate acid-base
HNO_2	NO_2^- (base)
CN^-	HCN (acid)
$HClO_4$	ClO_4^- (base)
F^-	HF (acid)
OH^-	H_2O (acid) /O^{2-} (base)
CO_3^{2-}	HCO_3^- (acid)
S_{2-}	HS^- (acid)

Question 7.36:

Which of the followings are Lewis acids? H_2O, BF_3, H^+, and NH_4^+

Answer 7.36:

Lewis acids are those acids which can accept a pair of electrons. For example, BF_3, H^+, and NH_4^+ are Lewis acids.

Question 7.37:

What will be the conjugate bases for the Brönsted acids: HF, H_2SO_4 and HCO_3?

Answer 7.37:

The table below lists the conjugate bases for the given Bronsted acids.

Bronsted acid	Conjugate base
HF	F^-
H_2SO_4	HSO_4^-
HCO_3^-	CO_3^{2-}

Question 7.38:

Write the conjugate acids for the following Brönsted bases: NH_2^-, NH_3 and $HCOO^-$.

Answer 7.38:

The table below lists the conjugate acids for the given Bronsted bases.

Bronsted base	Conjugate acid
NH_2^-	NH_3
NH_3	NH_4^+
$HCOO^-$	$HCOOH$

Question 7.39:

The species: H_2O, HCO_3^-, HSO_4^-, and NH_3 can act both as Brönsted acids and bases. For each case give the corresponding conjugate acid and base.

Answer 7.39:

The table below lists the conjugate acids and conjugate bases for the given species.

Species	Conjugate acid	Conjugate base
H_2O	H_3O^+	OH^-
HCO_3^-	H_2CO_3	CO_3^{2-}
HSO_4^-	H_2SO_4	SO_4^{2-}
NH_3	NH_4^+	NH_2^-

Question 7.40:

Classify the following species into Lewis acids and Lewis bases and show how these act as Lewis acid/base:

(a) OH^-

(b) F^-

(c) H^+

(d) BCl_3.

Answer 7.40:

(a) OH^- is a Lewis base since it can donate its lone pair of electrons.

(b) F^- is a Lewis base since it can donate a pair of electrons.

(c) H^+ is a Lewis acid since it can accept a pair of electrons.

(d) BCl_3 is a Lewis acid since it can accept a pair of electrons.

Question 7.41:

The concentration of hydrogen ion in a sample of soft drink is 3.8 × 10^{-3} M. what is its pH?

Answer 7.41:

Given,

$$\left[H^+\right] = 3.8 \times 10^{-3}\ M$$

$\therefore$ pH value of soft drink

$$= -\log\left[H^+\right]$$

$$= -\log\left(3.8 \times 10^{-3}\right)$$

$$= -\log 3.8 - \log 10^{-3}$$

$$= -\log 3.8 + 3$$

$$= -0.58 + 3$$

$$= 2.42$$

Question 7.42:

The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.

Answer 7.42:

Given, pH

= 3.76

It is known that,

$$\begin{aligned} &pH = -\log\left[H^+\right] \\ &\Rightarrow \log\left[H^+\right] = -pH \\ &\Rightarrow \left[H^+\right] = \text{antilog}(-pH) \\ &\qquad\qquad = \text{antilog}(-3.76) \\ &\qquad\qquad = 1.74 \times 10^{-4}\ M \end{aligned}$$

Hence, the concentration of hydrogen ion in the given sample of vinegar is 1.74 × 10^{-4} M.

Question 7.43:

The ionization constant of HF, HCOOH and HCN at 298K are 6.8×10^{-4}, 1.8×10^{-4} and 4.8×10^{-9} respectively. Calculate the ionization constants of the corresponding conjugate base.

Answer 7.43:

It is known that,

$$K_b = \frac{K_w}{K_a}$$

Given,

K_a of HF = 6.8×10^{-4}

Hence, K_b of its conjugate base F^-

$$= \frac{K_w}{K_a}$$

$$= \frac{10^{-14}}{6.8 \times 10^{-4}}$$

$$= 1.5 \times 10^{-11}$$

Given,

K_a of HCOOH = 1.8×10^{-4}

Hence, K_b of its conjugate base $HCOO^-$

$$= \frac{K_w}{K_a}$$

$$= \frac{10^{-14}}{1.8 \times 10^{-4}}$$

$$= 5.6 \times 10^{-11}$$

Given,

K_a of HCN = 4.8×10^{-9}

Hence, K_b of its conjugate base CN^-

$$= \frac{K_w}{K_a}$$

$$= \frac{10^{-14}}{4.8 \times 10^{-9}}$$

$$= 2.08 \times 10^{-6}$$

Question 7.44:

The ionization constant of phenol is 1.0×10^{-10}. What is the concentration of phenolate ion in 0.05 M solution of phenol? What will be its degree of ionization if the solution is also 0.01M in sodium phenolate?

Answer 7.44:

Ionization of phenol:

$$\begin{array}{lccc} & C_6H_5OH + H_2O \longleftrightarrow & C_6H_5O^- & + H_3O^+ \\ \text{Initial conc.} & 0.05 & 0 & 0 \\ \text{At equilibrium} & 0.05 - x & x & x \end{array}$$

$$K_a = \frac{[C_6H_5O^-][H_3O^+]}{[C_6H_5OH]}$$

$$K_a = \frac{x \times x}{0.05 - x}$$

As the value of the ionization constant is very less, x will be very small. Thus, we can ignore x in the denominator.

$$\therefore x = \sqrt{1 \times 10^{-10} \times 0.05}$$

$$= \sqrt{5 \times 10^{-12}}$$

$$= 2.2 \times 10^{-6}\,M = [H_3O^+]$$

Since $[H_3O^+] = [C_6H_5O^-]$,

$[C_6H_5O^-] = 2.2 \times 10^{-6}\,M$.

Now, let α be the degree of ionization of phenol in the presence of 0.01 M C_6H_5ONa.

$$\begin{array}{lcc} & C_6H_5ONa \longrightarrow & C_6H_5O^- + Na^+ \\ \text{Conc.} & & 0.01 \end{array}$$

Also,

$$\begin{array}{lccccc} & C_6H_5OH & + \; H_2O \longleftrightarrow & C_6H_5O^- & + & H_3O^+ \\ \text{Conc.} & 0.05 - 0.05\alpha & & 0.05\alpha & & 0.05\alpha \end{array}$$

$$[C_6H_5OH] = 0.05 - 0.05\alpha \; ; \; 0.05 \text{ M}$$

$$[C_6H_5O^-] = 0.01 + 0.05\alpha \; ; \; 0.01\text{M}$$

$$[H_3O^+] = 0.05\alpha$$

$$K_a = \frac{[C_6H_5O^-][H_3O^+]}{[C_6H_5OH]}$$

$$K_a = \frac{(0.01)(0.05\alpha)}{0.05}$$

$$1.0\times10^{-10} = .01\alpha$$

$$\alpha = 1\times10^{-8}$$

Question 7.45:

The first ionization constant of H_2S is 9.1×10^{-8}. Calculate the concentration of HS^- ion in its 0.1 M solution. How will this concentration be affected if the solution is 0.1 M in HCl also? If the second dissociation constant of H_2S is 1.2×10^{-13}, calculate the concentration of S^{2-} under both conditions.

Answer 7.45:

(i) To calculate the concentration of HS^- ion:

Case I (in the absence of HCl):

Let the concentration of HS^- be x M.

$$\begin{array}{llll} & H_2S \longleftrightarrow & H^+ + & HS^- \\ C_i & 0.1 & 0 & 0 \\ C_f & 0.1-x & x & x \end{array}$$

$$\text{Then, } K_{a_1} = \frac{[H^+][HS^-]}{[H_2S]}$$

$$9.1\times10^{-8} = \frac{(x)(x)}{0.1-x}$$

$$(9.1\times10^{-8})(0.1-x) = x^2$$

Taking $0.1-x\,M \;;\; 0.1M$, we have $(9.1\times10^{-8})(0.1) = x^2$.

$$9.1\times10^{-9} = x^2$$

$$x = \sqrt{9.1\times10^{-9}}$$

$$= 9.54\times10^{-5}\,M$$

$$\Rightarrow [HS^-] = 9.54\times10^{-5}\,M$$

Case II (in the presence of HCl):

In the presence of 0.1 M of HCl, let $[HS^-]$ be y M

$$\begin{array}{llll} \text{Then,} & H_2S \longleftrightarrow & HS^- + & H^+ \\ C_i & 0.1 & 0 & 0 \\ C_f & 0.1-y & y & y \end{array}$$

$$\text{Now, } K_{a_1} = \frac{[HS^-][H^+]}{[H_2S]}$$

$$K_{a_1} = \frac{[y](0.1+y)}{(0.1-y)}$$

$$9.1\times10^{-8} = \frac{y\times0.1}{0.1} \qquad (\because 0.1-y \;;\; 0.1M)$$

$$(\text{and } 0.1+y \;;\; 0.1M)$$

$$9.1\times10^{-8} = y$$

$$\Rightarrow [HS^-] = 9.1\times10^{-8}$$

To calculate the concentration of $\left[S^{2-}\right]$

Case I (in the absence of 0.1 M HCl):

$$HS^{-} \longleftrightarrow H^{+} + S^{2-}$$

$\left[HS^{-}\right] = 9.54 \times 10^{-5}\,M$ (From first ionization, case I)

Let $\left[S^{2-}\right]$ be X.

Also, $\left[H^{+}\right] = 9.54 \times 10^{-5}\,M$ (From first ionization, case I)

$$K_{a_2} = \frac{\left[H^{+}\right]\left[S^{2-}\right]}{\left[HS^{-}\right]}$$

$$K_{a_2} = \frac{\left(9.54 \times 10^{-5}\right)(X)}{9.54 \times 10^{-5}}$$

$$1.2 \times 10^{-13} = X = \left[S^{2-}\right]$$

Case II (in the presence of 0.1 M HCl):

Again, let the concentration of HS^{-} be X' M.

$\left[HS^{-}\right] = 9.1 \times 10^{-8}\,M$ (From first ionization, case II)

$\left[H^{+}\right] = 0.1\,M$ (From HCl, case II)

$$\left[S^{2-}\right] = X'$$

$$\text{Then, } K_{a_2} = \frac{\left[H^+\right]\left[S^{2-}\right]}{\left[HS^-\right]}$$

$$1.2\times10^{-13} = \frac{(0.1)(X')}{9.1\times10^{-8}}$$

$$10.92\times10^{-21} = 0.1X'$$

$$\frac{10.92\times10^{-21}}{0.1} = X'$$

$$X' = \frac{1.092\times10^{-20}}{0.1}$$

$$=1.092\times10^{-19}\text{ M}$$

$$\Rightarrow K_{a_1} = 1.74\times10^{-5}$$

Question 7.46:

The ionization constant of acetic acid is 1.74×10^{-5}. Calculate the degree of dissociation of acetic acid in its 0.05 M solution. Calculate the concentration of acetate ion in the solution and its pH.

Answer 7.46:

Method 1

1) $CH_3COOH \longleftrightarrow CH_3COO^- + H^+ \quad K_a = 1.74\times10^{-5}$

2) $H_2O + H_2O \longleftrightarrow H_3O^+ + OH^- \quad K_w = 1.0\times10^{-14}$

Since $Ka >> K_{w,}$:

$$\begin{array}{lllll} & CH_3COOH + H_2O & \longleftrightarrow & CH_3COO^- & + H_3O^+ \\ C_i = & 0.05 & & 0 & 0 \\ & 0.05-.05\alpha & & 0.05\alpha & 0.05\alpha \end{array}$$

$$K_a = \frac{(.05\alpha)(.05\alpha)}{(.05-0.05\alpha)}$$

$$= \frac{(.05\alpha)(0.05\alpha)}{.05(1-\alpha)}$$

$$= \frac{.05\alpha^2}{1-\alpha}$$

$$1.74\times10^{-5}=\frac{0.05\alpha^2}{1-\alpha}$$

$$1.74\times10^{-5}-1.74\times10^{-5}\alpha=0.05\alpha^2$$

$$0.05\alpha^2+1.74\times10^{-5}\alpha-1.74\times10^{-5}$$

$$D=b^2-4ac$$

$$=\left(1.74\times10^{-5}\right)^2-4(.05)\left(1.74\times10^{-5}\right)$$

$$=3.02\times10^{-25}+.348\times10^{-5}$$

$$\alpha=\sqrt{\frac{K_a}{c}}$$

$$\alpha=\sqrt{\frac{1.74\times10^{-5}}{.05}}$$

$$=\sqrt{\frac{34.8\times10^{-5}\times10}{10}}$$

$$=\sqrt{3.48\times10^{-6}}$$

$$=CH_3COOH\longleftrightarrow CH_3COO^-+H^+$$

$$\underline{\alpha 1.86\times10^{-3}}$$

$$\left[CH_3COO^-\right]=0.05\times1.86\times10^{-3}$$

$$=\frac{0.93\times10^{-3}}{1000}$$

$$=\underline{.000093}$$

Method 2

Degree of dissociation,

$$\alpha=\sqrt{\frac{K_a}{c}}$$

c = 0.05 M

K_a = 1.74 × 10^{-5}

$$\text{Then, } \alpha = \sqrt{\frac{1.74\times10^{-5}}{.05}}$$

$$\alpha = \sqrt{34.8\times10^{-5}}$$

$$\alpha = \sqrt{3.48}\times10^{-4}$$

$$\alpha = 1.8610^{-2}$$

$$CH_3COOH \longleftrightarrow CH_3COO^- + H^+$$

Thus, concentration of CH_3COO^- = c.α

$$= .05\times1.86\times10^{-2}$$

$$= .093\times10^{-2}$$

$$= .00093\,M$$

$$\text{Since } \left[oAc^-\right] = \left[H^+\right],$$

$$\left[H^+\right] = .00093 = .093\times10^{-2}.$$

$$pH = -\log\left[H^+\right]$$

$$= -\log\left(.093\times10^{-2}\right)$$

$$\therefore pH = 3.03$$

Hence, the concentration of acetate ion in the solution is 0.00093 M and its Ph is 3.03.

Question 7.47:

It has been found that the pH of a 0.01M solution of an organic acid is 4.15. Calculate the concentration of the anion, the ionization constant of the acid and its pK_a.

Answer 7.47:

Let the organic acid be HA.

$$\Rightarrow HA \longleftrightarrow H^+ + A^-$$

Concentration of HA = 0.01 M pH

= 4.15

$$-\log\left[H^+\right] = 4.15$$

$$\left[H^+\right] = 7.08 \times 10^{-5}$$

Now,

$$K_a = \frac{\left[H^+\right]\left[A^-\right]}{[HA]}$$

$$\left[H^+\right] = \left[A^-\right] = 7.08 \times 10^{-5}$$

$$[HA] = 0.01$$

Then,

$$K_a = \frac{\left(7.08 \times 10^{-5}\right)\left(7.08 \times 10^{-5}\right)}{0.01}$$

$$K_a = 5.01 \times 10^{-7}$$

$$pK_a = -\log K_a$$

$$= -\log\left(5.01 \times 10^{-7}\right)$$

$$pK_a = 6.3001$$

Question 7.48:

Assuming complete dissociation, calculate the pH of the following solutions:

(i) 0.003 M HCl

(ii) 0.005 M NaOH

(iii) 0.002 M HBr

(iv) 0.002 M KOH

Answer 7.48:

(i) 0.003MHCl:

$$H_2O + HCl \longleftrightarrow H_3O^+ + Cl^-$$

Since HCl is completely ionized,

$$\left[H_3O^+\right] = [HCl].$$

$$\Rightarrow \left[H_3O^+\right] = 0.003$$

Now,

$$pH = -\log[H_3O^+] = -\log(.003)$$
$$= 2.52$$

Hence, the pH of the solution is 2.52**.**

(ii) 0.005MNaOH:

$$NaOH_{(aq)} \longleftrightarrow Na^+_{(aq)} + HO^-_{(aq)}$$
$$[HO^-] = [NaOH]$$
$$\Rightarrow [HO^-] = .005$$
$$pOH = -\log[HO^-] = -\log(.005)$$
$$pOH = 2.30$$
$$\therefore pH = 14 - 2.30$$
$$= 11.70$$

Hence, the pH of the solution is 11.70.

(iii) 0.002 HBr:

$$HBr + H_2O \longleftrightarrow H_3O^+ + Br^-$$
$$[H_3O^+] = [HBr]$$
$$\Rightarrow [H_3O^+] = .002$$
$$\therefore pH = -\log[H_3O^+]$$
$$= -\log(0.002)$$
$$= 2.69$$

Hence, the pH of the solution is 2.69.

(iv) 0.002 M KOH:

$$KOH_{(aq)} \longleftrightarrow K^+_{(aq)} + OH^-_{(aq)}$$
$$[OH^-] = [KOH]$$
$$\Rightarrow [OH^-] = .002$$
$$\text{Now, } pOH = -\log[OH^-]$$
$$= 2.69$$
$$\therefore pH = 14 - 2.69$$
$$= 11.31$$

Hence, the pH of the solution is 11.31.

Question 7.49:

Calculate the pH of the following solutions:

a) 2 g of TlOH dissolved in water to give 2 litre of solution.

b) 0.3 g of $Ca(OH)_2$ dissolved in water to give 500 mL of solution.

c) 0.3 g of NaOH dissolved in water to give 200 mL of solution.

d) 1mL of 13.6 M HCl is diluted with water to give 1 litre of solution.

Answer 7.49:

(a) For 2g of TlOH dissolved in water to give 2 L of solution:

$$\left[\mathrm{TlOH}_{(aq)}\right] = \frac{2}{2}\,\mathrm{g/L}$$

$$= \frac{2}{2} \times \frac{1}{221}\,\mathrm{M}$$

$$= \frac{1}{221}\,\mathrm{M}$$

$$\mathrm{TlOH}_{(aq)} \longrightarrow \mathrm{Tl}^{+}_{(aq)} + \mathrm{OH}^{-}_{(aq)}$$

$$\left[\mathrm{OH}^{-}_{(aq)}\right] = \left[\mathrm{TlOH}_{(aq)}\right] = \frac{1}{221}\,\mathrm{M}$$

$$K_w = \left[\mathrm{H}^+\right]\left[\mathrm{OH}^-\right]$$

$$10^{-14} = \left[\mathrm{H}^+\right]\left(\frac{1}{221}\right)$$

$$221 \times 10^{-14} = \left[\mathrm{H}^+\right]$$

$$\Rightarrow \mathrm{pH} = -\log\left[\mathrm{H}^+\right] = -\log\left(221 \times 10^{-14}\right)$$

$$= -\log\left(2.21 \times 10^{-12}\right)$$

$$= 11.65$$

(b) For 0.3 g of $Ca(OH)_2$ dissolved in water to give 500 mL of solution:

$$Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^-$$

$$[Ca(OH)_2] = 0.3 \times \frac{1000}{500} = 0.6M$$

$$[OH^-_{aq}] = 2 \times [Ca(OH)_{2aq}] = 2 \times 0.6$$
$$= 1.2M$$

$$[H^+] = \frac{K_w}{[OH^-_{aq}]}$$
$$= \frac{10-14}{1.2}M$$
$$= 0.833 \times 10^{-14}$$
$$pH = -\log(0.833 \times 10^{-14})$$
$$= -\log(8.33 \times 10^{-13})$$
$$= (-0.902 + 13)$$
$$= 12.098$$

(c) For 0.3 g of NaOH dissolved in water to give 200 mL of solution:

$$NaOH \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$$[NaOH] = 0.3 \times \frac{1000}{200} = 1.5M$$

$$[OH^-_{aq}] = 1.5M$$

$$\text{Then, } [H^+] = \frac{10^{-14}}{1.5}$$
$$= 6.66 \times 10^{-13}$$
$$pH = -\log(6.66 \times 10^{-13})$$
$$= 12.18$$

(d) For 1mL of 13.6 M HCl diluted with water to give 1 L of solution:

13.6 × 1 mL = M_2 × 1000 mL

(Before dilution) (After dilution)

$13.6 \times 10^{-3} = M_2 \times 1L$ M_2

$= 1.36 \times 10^{-2}$ $[H^+] = 1.36 \times 10^{-2}$ pH = $-\log(1.36 \times 10^{-2})$

= (− 0.1335 + 2) = 1.866 … 1.87

Question 7.50:

The degree of ionization of a 0.1M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the pK_a of bromoacetic acid.

Answer 7.50:

Degree of ionization, α = 0.132

Concentration, c = 0.1 M

Thus, the concentration of $H_3O^+ = c.\alpha$

= 0.1 × 0.132

= 0.0132

$$\begin{aligned} pH &= -\log\left[H^+\right] \\ &= -\log(0.0132) \\ &= 1.879 : 1.88 \end{aligned}$$

Now,

$$\begin{aligned} K_a &= C\alpha^2 \\ &= 0.1\times(0.132)^2 \end{aligned}$$

$$K_a = .0017$$

$$pK_a = 2.75$$

Question 7.51:

The degree of ionization of a 0.1M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the pK_a of bromoacetic acid.

Answer 7.51:

Degree of ionization, α = 0.132

Concentration, c = 0.1 M

Thus, the concentration of $H_3O^+ = c.\alpha$

= 0.1 × 0.132

= 0.0132

$$\begin{aligned} pH &= -\log\left[H^+\right] \\ &= -\log(0.0132) \\ &= 1.879 : 1.88 \end{aligned}$$

Now,

$$K_a = C\alpha^2$$
$$= 0.1\times(0.132)^2$$
$$K_a = .0017$$
$$pK_a = 2.75$$

Question 7.52:

What is the pH of 0.001 M aniline solution? The ionization constant of aniline can be taken from Table 7.7. Calculate the degree of ionization of aniline in the solution. Also calculate the ionization constant of the conjugate acid of aniline.

Answer 7.52:

K_b = 4.27 × 10^{-10}

c = 0.001M pH

=?

α =?

$$k_b = c\alpha^2$$
$$4.27\times10^{-10} = 0.001\times\alpha^2$$
$$4270\times10^{-10} = \alpha^2$$
$$65.34\times10^{-5} = \alpha = 6.53\times10^{-4}$$
$$\text{Then, } [\text{anion}] = c\alpha = .001\times65.34\times10^{-5}$$
$$= .065\times10^{-5}$$
$$\text{pOH} = -\log(.065\times10^{-5})$$
$$= 6.187$$
$$\text{pH} = 7.813$$

Now,

$$K_a \times K_b = K_w$$
$$\therefore 4.27\times10^{-10}\times K_a = K_w$$
$$K_a = \frac{10^{-14}}{4.27\times10^{-10}}$$
$$= 2.34\times10^{-5}$$

Thus, the ionization constant of the conjugate acid of aniline is 2.34 × 10^{-5}.

Question 7.53:

Calculate the degree of ionization of 0.05M acetic acid if its pK_a value is 4.74.

How is the degree of dissociation affected when its solution also contains (a) 0.01 M (b) 0.1 M in HCl?

Answer 7.53:

$$c = 0.05 \text{ M}$$

$$pK_a = 4.74$$

$$pK_a = -\log(K_a)$$

$$K_a = 1.82 \times 10^{-5}$$

$$K_a = c\alpha^2 \qquad \alpha = \sqrt{\frac{K_a}{c}}$$

$$\alpha = \sqrt{\frac{1.82 \times 10^{-5}}{5 \times 10^{-2}}} = 1.908 \times 10^{-2}$$

When HCl is added to the solution, the concentration of H^+ ions will increase. Therefore, the equilibrium will shift in the backward direction i.e., dissociation of acetic acid will decrease.

Case I: When 0.01 M HCl is taken.

Let x be the amount of acetic acid dissociated after the addition of HCl.

	CH_3COOH	$\longleftrightarrow$	H^+	+	CH_3COO^-
Initial conc.	0.05M		0		0
After dissociation	$0.05 - x$		$0.01 + x$		x

As the dissociation of a very small amount of acetic acid will take place, the values i.e., $0.05 - x$ and $0.01 + x$ can be taken as 0.05 and 0.01 respectively.

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$\therefore K_a = \frac{(0.01)x}{0.05}$$

$$x = \frac{1.82\times10^{-5}\times0.05}{0.01}$$

$$x = 1.82\times10^{-3}\times0.05\,M$$

Now,

$$\alpha = \frac{\text{Amount of acid dissociated}}{\text{Amount of acid taken}}$$

$$= \frac{1.82\times10^{-3}\times0.05}{0.05}$$

$$= 1.82\times10^{-3}$$

Case II: When 0.1 M HCl is taken.

Let the amount of acetic acid dissociated in this case be *X*. As we have done in the first case, the concentrations of various species involved in the reaction are:

$$[CH_3COOH] = 0.05 - X \,;\, 0.05M$$

$$[CH_3COO^-] = X$$

$$[H^+] = 0.1 + X \,;\, 0.1M$$

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$\therefore K_a = \frac{(0.1)X}{0.05}$$

$$x = \frac{1.82\times10^{-5}\times0.05}{0.1}$$

$$x = 1.82\times10^{-4}\times0.05\,M$$

Now,

$$\alpha = \frac{\text{Amount of acid dissociated}}{\text{Amount of acid taken}}$$

$$= \frac{1.82\times10^{-4}\times0.05}{0.05}$$

$$= 1.82\times10^{-4}$$

Question 7.54:

The ionization constant of dimethylamine is 5.4×10^{-4}. Calculate its degree of ionization in its 0.02 M solution. What percentage of dimethylamine is ionized if the solution is also 0.1 M in NaOH?

Answer 7.54:

$$K_b = 5.4 \times 10^{-4}$$
$$c = 0.02 \text{ M}$$
$$\text{Then, } \alpha = \sqrt{\frac{K_b}{c}}$$
$$= \sqrt{\frac{5.4 \times 10^{-4}}{0.02}}$$
$$= 0.1643$$

Now, if 0.1 M of NaOH is added to the solution, then NaOH (being a strong base) undergoes complete ionization.

$$\underset{}{NaOH_{(aq)}} \longleftrightarrow \underset{0.1\text{ M}}{Na^{+}_{(aq)}} + \underset{0.1\text{ M}}{OH^{-}_{(aq)}}$$

And,

$$\underset{\substack{(0.02-x)\\ ;\ 0.02\text{M}}}{(CH_3)_2NH} \quad + \quad H_2O \longleftrightarrow \underset{x}{(CH_3)_2NH_2^{+}} \quad + \quad \underset{\substack{x\\ ;\ 0.1\text{M}}}{\bar{O}H}$$

$$\text{Then, } \left[(CH_3)_2NH_2^{+}\right] = x$$
$$\left[OH^{-}\right] = x + 0.1 \ ; \ 0.1$$
$$\Rightarrow K_b = \frac{\left[(CH_3)_2NH_2^{+}\right]\left[OH^{-}\right]}{\left[(CH_3)_2NH\right]}$$
$$5.4 \times 10^{-4} = \frac{x \times 0.1}{0.02}$$
$$x = 0.0054$$

It means that in the presence of 0.1 M NaOH, 0.54% of dimethylamine will get dissociated.

Question 7.55:

Calculate the hydrogen ion concentration in the following biological fluids whose pH are given below:

(a) Human muscle-fluid, 6.83

(b) Human stomach fluid, 1.2

(c) Human blood, 7.38

(d) Human saliva, 6.4.

Answer 7.55:

(a) Human muscle fluid 6.83:

pH = 6.83 pH = – log $[H^+]$

∴6.83 = – log $[H^+]$

$[H^+]$ =1.48 × 10^{-7} M

(b) Human stomach fluid, 1.2:

pH =1.2

1.2 = – log $[H^+]$

∴$[H^+]$ = 0.063

(c) Human blood, 7.38:

pH = 7.38 = – log $[H^+]$

∴ $[H^+]$ = 4.17 × 10^{-8} M

(d) Human saliva, 6.4:

pH = 6.4

6.4 = – log $[H^+]$

$[H^+]$ = 3.98 × 10^{-7}

Question 7.56:

The pH of milk, black coffee, tomato juice, lemon juice and egg white are 6.8, 5.0, 4.2, 2.2 and 7.8 respectively. Calculate corresponding hydrogen ion concentration in each.

Answer 7.56:

The hydrogen ion concentration in the given substances can be calculated by using the given relation: pH = –log $[H^+]$

(i) pH of milk = 6.8

Since, pH = –log $[H^+]$

6.8 = –log $[H^+]$ log

$[H^+]$ = –6.8

$[H^+]$ = anitlog(–6.8)

= 1.5×19^{-7} M

(ii) pH of black coffee = 5.0

Since, pH = –log $[H^+]$

5.0 = –log $[H^+]$ log

$[H^+]$ = –5.0

$[H^+]$ = anitlog(–5.0)

= 10^{-5} M

(iii) pH of tomato juice = 4.2

Since, pH = –log $[H^+]$
4.2 = –log $[H^+]$ log

$[H^+]$ = –4.2

$[H^+]$ = anitlog(–4.2)

= 6.31×10^{-5} M

(iv) pH of lemon juice = 2.2

Since, pH = –log $[H^+]$

2.2 = –log $[H^+]$ log

$[H^+]$ = –2.2

$[H^+]$ = anitlog(−2.2)

= 6.31×10^{-3} M

(v) pH of egg white = 7.8

Since, pH = −log $[H^+]$

7.8 = −log $[H^+]$ log

$[H^+]$ = −7.8

$[H^+]$ = anitlog(−7.8)

=1.58×10^{-8} M

Question 7.57:

If 0.561 g of KOH is dissolved in water to give 200 mL of solution at 298 K. Calculate the concentrations of potassium, hydrogen and hydroxyl ions. What is its pH?

Answer 7.57:

$$\left[\mathrm{KOH}_{aq}\right]=\frac{0.561}{\frac{1}{5}}\mathrm{g}/L$$

$$=2.805\,\mathrm{g}/L$$

$$=2.805\times\frac{1}{56.11}\ \mathrm{M}$$

$$=.05\ \mathrm{M}$$

$$\mathrm{KOH}_{(aq)}\rightarrow\mathrm{K^+}_{(aq)}+\mathrm{OH^-}_{(aq)}$$

$$\left[\mathrm{OH^-}\right]=.05\ \mathrm{M}=\left[\mathrm{K^+}\right]$$

$$\left[\mathrm{H^+}\right]\left[\mathrm{H^-}\right]=K_w$$

$$\left[\mathrm{H^+}\right]\frac{K_w}{\left[\mathrm{OH^-}\right]}$$

$$=\frac{10^{-14}}{0.05}=2\times10^{-13}\ \mathrm{M}$$

$$\therefore\ \mathrm{pH}=12.70$$

Question 7.58:

The solubility of $Sr(OH)_2$ at 298 K is 19.23 g/L of solution. Calculate the concentrations of strontium and hydroxyl ions and the pH of the solution.

Answer 7.58:

Solubility of $Sr(OH)_2$ = 19.23 g/L

Then, concentration of $Sr(OH)_2$

$$= \frac{19.23}{121.63} \text{ M}$$
$$= 0.1581 \text{ M}$$
$$Sr(OH)_{2(aq)} \longrightarrow Sr^{2+}_{(aq)} + 2(OH^-)_{(aq)}$$
$$\therefore [Sr^{2+}] = 0.1581 \text{ M}$$
$$[OH^-] = 2 \times 0.1581 \text{ M} = 0.3126 \text{ M}$$

Now,

$$K_w = [OH^-][H^+]$$
$$\frac{10^{-14}}{0.3126} = [H^+]$$
$$\Rightarrow [H^+] = 3.2 \times 10^{-14}$$
$$\therefore \text{pH} = 13.495 \; ; \; 13.50$$

Question 7.59:

The ionization constant of propanoic acid is 1.32×10^{-5}. Calculate the degree of ionization of the acid in its 0.05M solution and also its pH. What will be its degree of ionization if the solution is 0.01M in HCl also?

Answer 7.59:

Let the degree of ionization of propanoic acid be α.

Then, representing propionic acid as HA, we have:

$$\begin{array}{lllll} HA & + \quad H_2O & \longleftrightarrow & H_3O^+ & + \quad A^- \\ (.05-0.0\alpha)\approx .05 & & & .05\alpha & .05\alpha \end{array}$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$= \frac{(.05\alpha)(.05\alpha)}{0.05} = .05\alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{.05}} = 1.63\times10^{-2}$$

$$\text{Then,}[H_3O^+] = .05\alpha = .05\times1.63\times10^{-2} = K_b.15\times10^{-4}\ \text{M}$$

$$\therefore\ pH = 3.09$$

In the presence of 0.1M of HCl, let α´ be the degree of ionization.

$$\text{Then,}[H_3O^+] = 0.01$$

$$[A^-] = 005\alpha'$$

$$[HA] = .05$$

$$K_a = \frac{0.01\times.05\alpha'}{.05}$$

$$1.32\times10^{-5} = .01\times\alpha'$$

$$\alpha' = 1.32\times10^{-3}$$

Question 7.60:

The pH of 0.1M solution of cyanic acid (HCNO) is 2.34. Calculate the ionization constant of the acid and its degree of ionization in the solution.

Answer 7.60:

c = 0.1 M

pH = 2.34

$$-\log\left[H^+\right] = pH$$

$$-\log\left[H^+\right] = 2.34$$

$$\left[H^+\right] = 4.5\times10^{-3}$$

Also,

$$\left[H^+\right] = c\alpha$$

$$4.5\times10^{-3} = 0.1\times\alpha$$

$$\frac{4.5\times10^{-3}}{0.1} = \alpha$$

$$\alpha = 45\times10^{-3} = .045$$

Then,

$$\begin{aligned} K_a &= c\alpha^2 \\ &= 0.1\times\left(45\times10^{-3}\right)^2 \\ &= 202.5\times10^{-6} \\ &= 2.02\times10^{-4} \end{aligned}$$

Question 7.61:

The ionization constant of nitrous acid is 4.5×10^{-4}. Calculate the pH of 0.04 M sodium nitrite solution and also its degree of hydrolysis.

Answer 7.61:

$NaNO_2$ is the salt of a strong base (NaOH) and a weak acid (HNO_2).

$$NO_2^- + H_2O \longleftrightarrow HNO_2 + OH^-$$

$$K_h = \frac{[HNO_2][OH^-]}{[NO_2^-]}$$

$$\Rightarrow \frac{K_w}{K_a} = \frac{10^{-14}}{4.5\times10^{-4}} = .22\times10^{-10}$$

Now, If x moles of the salt undergo hydrolysis, then the concentration of various species present in the solution will be:

$$[NO_2^-] = .04 - x ;\ 0.04$$

$$[HNO_2] = x$$

$$[OH^-] = x$$

$$K_h = \frac{x^2}{0.04} = 0.22\times10^{-10}$$

$$x^2 = .0088\times10^{-10}$$

$$x = .093\times10^{-5}$$

$$\therefore [OH^-] = 0.093\times10^{-5}\ \text{M}$$

$$[H_3O^+] = \frac{10^{-14}}{.093\times10^{-5}} = 10.75\times10^{-9}\ \text{M}$$

$$\Rightarrow \text{pH} = -\log(10.75\times10^{-9})$$

$$= 7.96$$

Therefore, degree of hydrolysis

$$= \frac{x}{0.04} = \frac{.093\times10^{-5}}{.04}$$

$= 2.325 \times 10^{-5}$

Question 7.62:

A 0.02 M solution of pyridinium hydrochloride has pH = 3.44. Calculate the ionization constant of pyridine

Answer 7.62:

pH = 3.44

We know that,

pH = − log [H⁺]

$$\therefore \left[H^+\right] = 3.63 \times 10^{-4}$$

$$\text{Then, } K_h = \frac{\left(3.63 \times 10^{-4}\right)^2}{0.02} \qquad (\because \text{concentration} = 0.02 \text{ M})$$

$$\Rightarrow K_h = 6.6 \times 10^{-6}$$

$$\text{Now, } K_h = \frac{K_w}{K_a}$$

$$\Rightarrow K_a = \frac{K_w}{K_h} = \frac{10^{-14}}{6.6 \times 10^{-6}}$$

$$= 1.51 \times 10^{-9}$$

Question 7.63:

Predict if the solutions of the following salts are neutral, acidic or basic:

NaCl, KBr, NaCN, NH_4NO_3, $NaNO_2$ and KF

Answer 7.63:

(i) NaCl:

$$\underset{}{NaCl} + H_2O \longleftrightarrow \underset{\text{Strong base}}{NaOH} + \underset{\text{Strong acid}}{HCl}$$

Therefore, it is a neutral solution.

(ii) KBr:

$$KBr + H_2O \longleftrightarrow \underset{\text{Strong base}}{KOH} + \underset{\text{Strong acid}}{HBr}$$

Therefore, it is a neutral solution.

(iii) NaCN:

$$\underset{}{NaCN} + H_2O \longleftrightarrow \underset{\text{Weak acid}}{HCN} + \underset{\text{Strong base}}{NaOH}$$

Therefore, it is a basic solution.

(iv) NH_4NO_3

$$NH_4NO_3 + H_2O \longleftrightarrow \underset{\text{Weak base}}{NH_4OH} + \underset{\text{Strong acid}}{HNO_3}$$

Therefore, it is an acidic solution.

(v) $NaNO_2$

$$NaNO_2 + H_2O \longleftrightarrow \underset{\text{Strong base}}{NaOH} + \underset{\text{Weak acid}}{HNO_2}$$

Therefore, it is a basic solution.

(vi) KF

$$KF + H_2O \longleftrightarrow \underset{\text{Strong base}}{KOH} + \underset{\text{Weak acid}}{HF}$$

Therefore, it is a basic solution.

Question 7.64:

The ionization constant of chloroacetic acid is 1.35×10^{-3}. What will be the pH of 0.1M acid and its 0.1M sodium salt solution?

Answer 7.64:

It is given that K_a for $ClCH_2COOH$ is 1.35×10^{-3}.

$$\Rightarrow K_a = c\alpha^2$$

$$\therefore \alpha = \sqrt{\frac{K_a}{c}}$$

$$= \sqrt{\frac{1.35\times10^{-3}}{0.1}} \qquad (\because \text{ concentration of acid } = 0.1\text{m})$$

$$\alpha = \sqrt{1.35\times10^{-2}}$$

$$= 0.116$$

$$\therefore \left[H^+\right] = c\alpha = 0.1\times0.116$$

$$= .0116$$

$$\Rightarrow \text{pH} = -\log\left[H^+\right] = 1.94$$

$ClCH_2COONa$ is the salt of a weak acid i.e., $ClCH_2COOH$ and a strong base i.e., NaOH.

$$ClCH_2COO^- + H_2O \longleftrightarrow ClCH_2COOH + OH^-$$

$$K_h = \frac{\left[ClCH_2COOH\right]\left[OH^-\right]}{\left[ClCH_2COO^-\right]}$$

$$K_h = \frac{K_w}{K_a}$$

$$K_h = \frac{10^{-14}}{1.35\times10^{-3}}$$

$$= 0.740\times10^{-11}$$

$$\text{Also, } K_h = \frac{x^2}{0.1} \qquad \left(\text{where } x \text{ is the concentration of } OH^- \text{and } ClCH_2COOH\right)$$

$$0.740\times10^{-11} = \frac{x^2}{0.1}$$

$$0.074\times10^{-11} = x^2$$

$$\Rightarrow x^2 = 0.74\times10^{-12}$$

$$x = 0.86\times10^{-6}$$

$$\left[OH^-\right] = 0.86\times10^{-6}$$

$$\therefore \left[H^+\right] = \frac{K_w}{0.86\times10^{-6}}$$

$$= \frac{10^{-14}}{0.86\times10^{-6}}$$

$$\left[H^+\right] = 1.162\times10^{-8}$$

$$\text{pH} = -\log\left[H^+\right]$$

$$= 7.94$$

Question 7.65:

Ionic product of water at 310 K is 2.7 × 10^{-14}. What is the pH of neutral water at this temperature?

Answer 7.65:

Ionic product,

$$K_w = [H^+][OH^-]$$
$$\text{Let}[H^+] = x.$$
$$\text{Since}[H^+] = [OH^-], K_w = x^2.$$
$$\Rightarrow K_w \text{ at } 310\text{K is } 2.7\times10^{-14}.$$
$$\therefore 2.7\times10^{-14} = x^2$$
$$\Rightarrow x = 1.64\times10^{-7}$$
$$\Rightarrow [H^+] = 1.64\times10^{-7}$$
$$\Rightarrow pH = -\log[H^+]$$
$$= -\log[1.64\times10^{-7}]$$
$$= 6.78$$

Hence, the pH of neutral water is 6.78.

Question 7.66:

Calculate the pH of the resultant mixtures:

a) 10 mL of 0.2M $Ca(OH)_2$ + 25 mL of 0.1M HCl

b) 10 mL of 0.01M H_2SO_4 + 10 mL of 0.01M $Ca(OH)_2$

c) 10 mL of 0.1M H_2SO_4 + 10 mL of 0.1M KOH

Answer 7.66:

(a) $\text{Moles of } H_3O^+ = \frac{25\times0.1}{1000} = .0025\,\text{mol}$

$$\text{Moles of } OH^- = \frac{10 \times 0.2 \times 2}{1000} = .0040\,\text{mol}$$

Thus, excess of OH^- = .0015 mol

$$\left[OH^-\right] = \frac{.0015}{35 \times 10^{-3}}\,\text{mol/L} = .0428$$
$$pOH = -\log[OH]$$
$$= 1.36$$
$$pH = 14 - 1.36$$
$$= 12.63 \qquad \text{(not matched)}$$

(b) $$\text{Moles of } H_3O^+ = \frac{2 \times 10 \times 0.01}{1000} = .0002\,\text{mol}$$

$$\text{Moles of } OH^- = \frac{2 \times 10 \times .01}{1000} = .0002\,\text{mol}$$

Since there is neither an excess of H_3O^+ or OH^-

(c) $$\text{Moles of } H_3O^+ = \frac{2 \times 10 \times 0.1}{1000} = .002\,\text{mol}$$

$$\text{Moles of } OH^- = \frac{10 \times 0.1}{1000} = 0.001\,\text{mol}$$

Excess of H_3O^+ = .001 mol

$$\text{Thus, } \left[H_3O^+\right] = \frac{.001}{20 \times 10^{-3}} = \frac{10^{-3}}{20 \times 10^{-3}} = .05$$
$$\therefore\ pH = -\log(0.05)$$

= 1.30

The solution is neutral. Hence, pH = 7.

Question 7.67:

Determine the solubilities of silver chromate, barium chromate, ferric hydroxide, lead chloride and mercurous iodide at 298K from their solubility product constants given in Table 7.9 (page 221). Determine also the molarities of individual ions.

Answer 7.67:

(1) Silver chromate:

$$Ag_2CrO_4 \longrightarrow 2Ag^+ + CrO_4^{2-}$$

Then,

$$K_{sp} = \left[Ag^+\right]^2 \left[CrO_4^{2-}\right]$$

Let the solubility of Ag_2CrO_4 be s.

$$\Rightarrow \left[Ag^+\right] 2s \text{ and} \left[CrO_4^{2-}\right] = s$$

Then,

$$K_{sp} = (2s)^2 . s = 4s^3$$
$$\Rightarrow 1.1 \times 10^{-12} = 4s^3$$
$$.275 \times 10^{-12} = s^3$$
$$s = 0.65 \times 10^{-4} \text{ M}$$

Molarity of Ag^+ = 2s = 2 × 0.65 × 10^{-4} = 1.30 × 10^{-4} M

Molarity of CrO_4^{2-} = *s* = 0.65 × 10^{-4} M

(2) Barium chromate:

$$BaCrO_4 \longrightarrow Ba^{2+} + CrO_4^{2-}$$
$$\text{Then, } K_{sp} = \left[Ba^{2+}\right]\left[CrO_4^{2-}\right]$$

Let the solubility of $BaCrO_4$. be s.

So, $\left[Ba^{2+}\right]$ = s and $\left[CrO_4^{2-}\right]$= s

$$\Rightarrow K_{SP} = s^2$$
$$\Rightarrow 1.2 \times 10^{-10} = s^2$$
$$\Rightarrow s = 1.09 \times 10^{-5} \text{ M}$$

Molarity of Ba^{2+} = Molarity of $CrO_4^{2-} = s = 1.09\times10^{-5}$ M

(3) Ferric hydroxide:

$$Fe(OH)_3 \longrightarrow Fe^{2+} + 3OH^-$$

$$K_{sp} = [Fe^{2+}][OH^-]^3$$

Let s be the solubility of $Fe(OH)_3$.

$$\text{Thus,} [Fe^{3+}] = s \text{ and } [OH^-] = 3s \quad \Rightarrow K_{sp} = s.(3s)^3$$

$$= s.27s^3$$

$$K_{sp} = 27s^4$$

$$1.0\times10^{-38} = 27s^4$$

$$.037\times10^{-38} = s^4$$

$$.00037\times10^{-36} = s^4 \quad \Rightarrow 1.39\times10^{-10}\text{ M} = S$$

Molarity of $Fe^{3+} = s = 1.39\times10^{-10}$ M

Molarity of $OH^- = 3s = 4.17\times10^{-10}$ M

(4) Lead chloride:

$$PbCl_2 \longrightarrow Pb^{2+} + 2Cl^-$$

$$K_{SP} = [Pb^{2+}][Cl^-]^2$$

Let K_{SP} be the solubility of $PbCl_2$.

$$[PB^{2+}] = s \text{ and } [Cl^-] = 2s$$

$$\text{Thus, } K_{sp} = s.(2s)^2$$

$$= 4s^3$$

$$\Rightarrow 1.6\times10^{-5} = 4s^3$$

$$\Rightarrow 0.4\times10^{-5} = s^3$$

$$4\times10^{-6} = s^3 \Rightarrow 1.58\times10^{-2}\text{M} = S.1$$

Molarity of $PB^{2+} = s = 1.58\times10^{-2}$M

Molarity of chloride = $2s = 3.16\times10^{-2}$M

(5) Mercurous iodide:

$$Hg_2I_2 \longrightarrow Hg^{2+} + 2I^-$$
$$K_{sp} = \left[Hg_2^{2+}\right]^2\left[I^-\right]^2$$

Let s be the solubility of Hg_2I_2.

$$\Rightarrow \left[Hg_2^{2+}\right] = s \text{ and} \left[I^-\right] = 2s$$

$$\text{Thus, } K_{sp} = s(2s)^2 \Rightarrow K_{sp} = 4s^3$$
$$4.5 \times 10^{-29} = 4s^3$$
$$1.125 \times 10^{-29} = s^3$$
$$\Rightarrow s = 2.24 \times 10^{-10} \text{ M}$$

Molarity of $Hg_2^{2+} = s = 2.24 \times 10^{-10}$ M

Molarity of $I^- = 2s = 4.48 \times 10^{-10}$ M

Question 7.68:

The solubility product constant of Ag_2CrO_4 and AgBr are 1.1×10^{-12} and 5.0×10^{-13} respectively. Calculate the ratio of the molarities of their saturated solutions.

Answer 7.68:

Let *s* be the solubility of Ag_2CrO_4.

$$\text{Then, } Ag_2CrO_4 \longleftrightarrow Ag^{2+} + 2CrO_4^-$$
$$K_{sp} = (2s)^2 . s = 4s^3$$
$$1.1 \times 10^{-12} = 4s^3$$
$$s = 6.5 \times 10^{-5} \text{ M}$$

Let *s*´ be the solubility of AgBr.

$$AgBr_{(s)} \longleftrightarrow Ag^+ + Br^-$$
$$K_{sp} = s'^2 = 5.0 \times 10^{-13}$$
$$\therefore s' = 7.07 \times 10^{-7} \text{ M}$$

Therefore, the ratio of the molarities of their saturated solution is

$$\frac{s}{s'} = \frac{6.5 \times 10^{-5}\text{M}}{7.07 \times 10^{-7}\text{M}} = 91.9.$$

Question 7.69:

Equal volumes of 0.002 M solutions of sodium iodate and cupric chlorate are mixed together. Will it lead to precipitation of copper iodate?

(For cupric iodate K_{sp} = 7.4 × 10^{-8}).

Answer 7.69:

When equal volumes of sodium iodate and cupric chlorate solutions are mixed together, then the molar concentrations of both solutions are reduced to half i.e., 0.001 M. Then,

$$\underset{0.001M}{Na10_3} \rightarrow Na^+ + \underset{0.001M}{10_3^-}$$

$$\underset{0.001M}{Cu(ClO_3)_2} \rightarrow Cu^{2+} + \underset{0.001M}{2ClO_3^-}$$

Now, the solubility equilibrium for copper iodate can be written as:

$$Cu(10_3)_2 \rightarrow Cu^{2+}_{(aq)} + 210^-_{3(aq)}$$

Ionic product of copper iodate:

$$= [Cu^{2+}][10_3^-]^2$$
$$= (0.001)(0.001)^2$$
$$= 1 \times 10^{-9}$$

Since the ionic product (1 × 10^{-9}) is less than K_{sp} (7.4 × 10^{-8}), precipitation will not occur.

Question 7.70:

The ionization constant of benzoic acid is 6.46 × 10^{-5} and K_{sp} for silver benzoate is 2.5 × 10^{-13}. How many times is silver benzoate more soluble in a buffer of pH 3.19 compared to its solubility in pure water?

Answer 7.70:

Since pH = 3.19,

$$\left[H_3O^+\right] = 6.46\times10^{-4}\ \text{M}$$

$$C_6H_5COOH + H_2O \longleftrightarrow C_6H_5COO^- + H_3O$$

$$K_a \frac{\left[C_6H_5COO^-\right]\left[H_3O^+\right]}{\left[C_6H_5COOH\right]}$$

$$\frac{\left[C_6H_5COOH\right]}{\left[C_6H_5COO^-\right]} = \frac{\left[H_3O^+\right]}{K_a} = \frac{6.46\times10^{-4}}{6.46\times10^{-5}} = 10$$

Let the solubility of C_6H_5COOAg be *x* mol/L.

Then,

$$\left[Ag^+\right] = x$$

$$\left[C_6H_5COOH\right] + \left[C_6H_5COO^-\right] = x$$

$$10\left[C_6H_5COO^-\right] + \left[C_6H_5COO^-\right] = x$$

$$\left[C_6H_5COO^-\right] = \frac{x}{11}$$

$$K_{sp}\left[Ag^+\right]\left[C_6H_5COO^-\right]$$

$$2.5\times10^{-13} = x\left(\frac{x}{11}\right)$$

$$x = 1.66\times10^{-6}\ \text{mol/L}$$

Thus, the solubility of silver benzoate in a pH 3.19 solution is 1.66×10^{-6} mol/L.

Now, let the solubility of C_6H_5COOAg be *x′* mol/L.

$$\text{Then, } \left[Ag^+\right] = x'\text{M and } \left[CH_3COO^-\right] = x'\text{M.}$$

$$K_{sp} = \left[Ag^+\right]\left[CH_3COO^-\right]$$

$$K_{sp} = (x')^2$$

$$x' = \sqrt{K_{sp}} = \sqrt{2.5\times10^{-13}} = 5\times10^{-7}\ \text{mol/L}$$

$$\therefore \frac{x}{x'} = \frac{1.66\times10^{-6}}{5\times10^{-7}} = 3.32$$

Hence, C_6H_5COOAg is approximately 3.317 times more soluble in a low pH solution.

Question 7.71:

What is the maximum concentration of equimolar solutions of ferrous sulphate and sodium sulphide so that when mixed in equal volumes, there is no precipitation of iron sulphide? (For iron sulphide, K_{sp} = 6.3×10^{-18}).

Answer 7.71:

Let the maximum concentration of each solution be *x* mol/L. After mixing, the volume of the concentrations of each solution will be reduced to half i.e., x/2 .

$$\therefore [FeSO_4] = [Na_2S] = \frac{x}{2} M$$

$$\text{Then,} [Fe^{2+}] = [FeSO_4] = \frac{x}{2} M$$

$$\text{Also, } [S^{2-}] = [Na_2S] = \frac{x}{2} M$$

$$FeS_{(s)} \longleftrightarrow Fe^{2+}{}_{(aq)} + S^{2-}{}_{(aq)}$$

$$K_{sp} = [Fe^{2+}][S^{2-}]$$

$$6.3 \times 10^{-18} = \left(\frac{x}{2}\right)\left(\frac{x}{2}\right)$$

$$\frac{x^2}{4} = 6.3 \times 10^{-18}$$

$$\Rightarrow x = 5.02 \times 10^{-9}$$

If the concentrations of both solutions are equal to or less than 5.02×10^{-9} M, then there will be no precipitation of iron sulphide.

Question 7.72:

What is the minimum volume of water required to dissolve 1g of calcium sulphate at 298 K? (For calcium sulphate, K_{sp} is 9.1×10^{-6}).

Answer 7.72:

$$CaSO_{4(s)} \longleftrightarrow Ca^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

$$K_{sp} = [Ca^{2+}][SO_4^{2-}]$$

Let the solubility of $CaSO_4$ be *s*.

$$\text{Then, } K_{sp} = s^2$$
$$9.1 \times 10^{-6} = s^2$$
$$s = 3.02 \times 10^{-3} \text{ mol/L}$$

Molecular mass of $CaSO_4$ = 136 g/mol

Solubility of $CaSO_4$ in gram/L

= $3.02 \times 10^{-3} \times 136$

= 0.41 g/L

This means that we need 1L of water to dissolve 0.41g of $CaSO_4$

Therefore, to dissolve 1g of $CaSO_4$ we require $= \frac{1}{0.41} \text{L} = 2.44 \text{L}$ of water.

Question 7.73:

The concentration of sulphide ion in 0.1M HCl solution saturated with hydrogen sulphide is 1.0×10^{-19} M. If 10 mL of this is added to 5 mL of 0.04 M solution of the following: $FeSO_4$, $MnCl_2$, $ZnCl_2$ and $CdCl_2$. in which of these solutions precipitation will take place?

$$\text{Given } K_{sp} \text{ for FeS} = 6.3 \times 10^{-18}, \text{ MnS} = 2.5 \times 10^{-13}, \text{ ZnS} = 1.6 \times 10^{-24},$$
$$\text{CdS} = 8.0 \times 10^{-27}$$

Answer 7.73:

For precipitation to take place, it is required that the calculated ionic product exceeds the K_{sp} value.

Before mixing:

$$\left[S^{2-}\right] = 1.0 \times 10^{-19} \text{ M} \quad \left[M^{2+}\right] = 0.04\text{M}$$
$$\text{volume} = 10 \text{ mL} \qquad \text{volume} = 5 \text{ mL}$$

After mixing:

$$[S^{2-}] = ? \qquad [M^{2+}] = ?$$

$$\text{volume} = (10+5) = 15 \text{ mL} \qquad \text{volume} = 15 \text{ mL}$$

$$[S^{2-}] = \frac{1.0\times10^{-19}\times10}{15} = 6.67\times10^{-20} \text{ M}$$

$$[M^{2+}] = \frac{0.04\times5}{15} = 1.33\times10^{-2} \text{ M}$$

$$\begin{aligned}\text{Ionic product} &= [M^{2+}][S^{2-}] \\ &= (1.33\times10^{-2})(6.67\times10^{-20}) \\ &= 8.87\times10^{-22}\end{aligned}$$

This ionic product exceeds the K_{sp} of Zns and CdS. Therefore, precipitation will occur in $CdCl_2$ and $ZnCl_2$ solutions.