CH7 – EQUILIBRIUM

Question 1. A liquid is in equilibrium with its vapours in a sealed container at a fixed temperature. The volume of the container is suddenly increased, (i) What is the initial effect of the change on the vapour pressure? (ii) How do the rates of evaporation and condensation change initially? (iii) What happens when equilibrium is restored finally and what will be the final vapour pressure?

Answer: (i) On increasing the volume of the container, the vapour pressure will initially decrease because the same amount of vapours are now distributed over a larger space.

(ii) On increasing the volume of the container, the rate of evaporation will increase initially because now more space is available. Since the amount of the vapours per unit volume decrease on increasing the volume, therefore, the rate of condensation will decrease initially.

(iii) Finally, equilibrium will be restored when the rates of the forward and backward processes become equal. However, the vapour pressure will remain unchanged because it depends upon the temperature and not upon the volume of the container.

Question 2. What is K_c for the following reaction in state of equilibrium?

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

Given: $[SO_2] = 0.6\ M$; $[O_2] = 0.82\ M$; *and* $[SO_3] = 1.90\ M$

Answer:

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

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(1.9\ M) \times (1.9\ M)}{(0.6\ M) \times (0.6\ M) \times (0.82\ M)}$$

$$= 12.229\ M^{-1} = \mathbf{12.229\ L\ mol^{-1}}$$

Question 3.

At a certain temperature and total pressure of 10^5 *Pa, iodine vapours contain 40% by volume of iodine atoms in the equilibrium* $I_2(g) \rightleftharpoons 2I(g)$. *Calculate* K_p *for the equilibrium.*

Answer:

According to available data:

Total pressure of equilibrium mixture = 10^5 Pa

$$\text{Partial pressure of iodine atoms } (I) = \frac{40}{100} \times (10^5 \text{ Pa}) = 0.4 \times 10^5 \text{ Pa}$$

$$\text{Partial pressure of iodine molecules } (I_2) = \frac{60}{100} \times (10^5 \text{ Pa}) = 0.6 \times 10^5 \text{ Pa}$$

$$\underset{(0.6 \times 10^5 \text{ Pa})}{I_2(g)} \rightleftharpoons \underset{(0.4 \times 10^5 \text{ Pa})}{2I(g)}$$

$$K_p = \frac{p_{I^2}}{p_I} = \frac{(0.4 \times 10^5 \text{Pa})^2}{(0.6 \times 10^5 \text{Pa})} = \mathbf{2.67 \times 10^4 \text{ Pa}}$$

Question 4. Write the expression for the equilibrium constant for each of the following reactions

(i) $2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$
(ii) $2Cu(NO_3)_2(s) \rightleftharpoons 2CuO(s) + 4NO_2(g) + O_2(g)$
(iii) $CH_3COOC_2H_5(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + C_2H_5OH(aq)$
(iv) $Fe^{3+}(aq) + 3OH^-(aq) \rightleftharpoons Fe(OH)_3(s)$
(v) $I_2(s) + 5F_2(g) \rightleftharpoons 2IF_5(l)$

Answer:

(i) $$K_c = \frac{[NO(g)]^2[Cl_2(g)]}{[NOCl(g)]^2}$$

(ii) $$K_c = \frac{[CuO(g)]^2[NO_2(g)]^4[O_2(g)]}{[Cu(NO_3)_2(s)]^2} = [NO_2(g)]^4[O_2(g)]$$

(iii) $$K_c = \frac{[CH_3COOH(aq)][C_2H_5OH(aq)]}{[CH_3COOC_2H_5(aq)][H_2O(l)]}$$

$$= \frac{[CH_3COOH(aq)][C_2H_5OH(aq)]}{[CH_3COOC_2H_5(aq)]}$$

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

(v) $$K_c = \frac{[IF_5(l)]^2}{[I_2(s)][F_2(g)]^5} = \frac{[IF_5(l)]^2}{[F_2(g)]^5}$$

Question 5. Find the value of K_c for each of the following equilibria from the value of K

(a) $2NOCl\ (g) \rightleftharpoons 2NO\ (g) + Cl_2\ (g)\ ;\ K_p = 1.8 \times 10^{-2}$ atm at 500 K
(b) $CaCO_3\ (s) \rightleftharpoons CaO\ (s) + CO_2\ (g)\ ;\ K_p = 167$ atm at 1073 K.

Answer:

K_p and K_c are related to each other as $K_p = K_c\ (RT)^{\Delta ng}$
The value of K_c can be calculated as follows:

(a) $2NOCl\ (g) \rightleftharpoons 2NO\ (g) + Cl_2$

$$K_p = 1.8 \times 10^{-2}\ \text{atm},$$
$$\Delta^{ng} = 3 - 2 = 1\ ;\ R = 0.0821\ \text{litre atm}\ K^{-1}\ \text{mol}^{-1}\ ;\ T = 500\ \text{K}$$

$$\therefore \quad K_c = \frac{K_p}{(RT)^{\Delta ng}} = \frac{(1.8 \times 10^{-2}\ \text{atm})}{(0.0821\ \text{L atm K}^{-1}\ \text{mol}^{-1} \times 500\ \text{K})^1}$$
$$= \mathbf{4.4 \times 10^{-4}\ mol\ L^{-1}}$$

(b) $CaCO_3\ (s) \rightleftharpoons CaO\ (s) + CO_2\ (g)$

$$K_p = 167\ \text{atm}, \quad \Delta^{ng} = 1$$
$$R = 0.0821\ \text{liter atm}\ K^{-1}\ \text{mol}^{-1}\ ;\ T = 1073\ K$$
$$K_c = \frac{K_p}{(RT)^{\Delta ng}} = \frac{(167\ \text{atm})}{(0.0821\ \text{L atm K}^{-1}\ \text{mol}^{-1} \times 1073\ \text{K})^1}$$
$$= \mathbf{1.9\ mol\ L^{-1}}$$

Question 6. For the following equilibrium, K =6.3 x 10^{14} at 1000 K. $NO(g)+O_3$ —–>$NO_2(g)$ + $O_2(g)$ Both the forward and reverse reactions in the equilibrium are elementary bimolecular reactions. What is K_c for the reverse reaction?

Answer:

For the reverse reaction $K_c = \dfrac{1}{K_c} = \dfrac{1}{6.3 \times 10^{14}} = \mathbf{1.59 \times 10^{-15}}$.

Question 7. Explain why pure liquids and solids can be ignored while writing the value of equilibrium constants.

Answer: This is because molar concentration of a pine solid or liquid is independent of the amount present.

$$\text{Molar concentration} = \frac{\text{No. of moles}}{\text{volume}} \times \frac{\text{Mass}}{\text{volume}} \times \text{Density}$$

Since density of pure liquid or solid is fixed and molar mass is also fixed. Therefore molar concentration are constant.

Question 8. Reaction between nitrogen and oxygen takes place as follows:

$$2N_2(g) + O_2(g) \rightleftharpoons 2N_2O(g)$$

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

Answer: Let x moles of $N_2(g)$ take part in the reaction. According to the equation, x/2 moles of O_2 (g) will react to form x moles of $N_2O(g)$. The molar concentration per litre of different species before the reaction and at the equilibrium point is:

	$2N_2(g)$	+ $O_2(g)$	$\rightleftharpoons 2N_2O(g)$
Initial mole/litre:	$\frac{0.482}{10}$	$\frac{0.933}{10}$	Zero
Mole/litre at eqm. point:	$\frac{0.482-x}{10}$	$\frac{0.933-\frac{x}{2}}{10}$	$\frac{x}{10}$

The value of equilibrium constant (2.0×10^{-37}) is extremely small. This means that only small amounts of reactants have reacted. Therefore, is extremely small and can be omitted as far as the reactants are concerned.

Applying Law of Chemical Equilibrium $K_c = \frac{[N_2O(g)]^2}{[N_2(g)]^2[O_2(g)]}$

$$2.0 \times 10^{-37} = \frac{\left(\frac{x}{10}\right)^2}{\left(\frac{0.482}{10}\right)^2 \times \left(\frac{0.933}{10}\right)} = \frac{0.01\,x^2}{2.1676 \times 10^{-4}}$$

$$x^2 = 43.352 \times 10^{-40} \quad \text{or} \quad x = 6.6 \times 10^{-20}$$

As x is extremely small, it can be neglected.

Thus, in the equilibrium mixture

Molar conc. of N_2 = $\mathbf{0.0482\ mol\ L^{-1}}$

Molar conc. of O_2 = $\mathbf{0.0933\ mol\ L^{-1}}$

Molar conc. of N_2O = $0.1 \times x = 0.1 \times 6.6 \times 10^{-20}\ mol\ L^{-1}$
= $\mathbf{6.6 \times 10^{-21}\ mol\ L^{-1}}$

Question 9. Nitric oxide reacts with bromine and gives nitrosyl bromide as per reaction given below:

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

When 0.087 mole of NO and 0.0437 mole of Br_2 are mixed in a closed container at constant temperature, 0.0518 mole of NOB_r is obtained at equilibrium. Determine the compositions of the equilibrium mixture.

Answer: The balanced chemical equation for the reaction is:

According to the equation, 2 moles of NO (g) react with 1 mole of Br_2 (g) to form 2 moles of NOB_r (g). The composition of the equilibrium mixture can be calculated as follows:

No. of moles of NOB_r (g) formed at equilibrium = 0.0518 mol (given)

No. of moles of NO (g) taking part in reaction = 0.0518 mol

No. of moles of NO (g) left at equilibrium = 0.087 – 0.0518 = 0.0352 mol

No. of moles of Br_2 (g) taking part in reaction = 1/2 x 0.0518 = 0.0259 mol

No. of moles of Br_2 (g) left at equilibrium = 0.0437 – 0.0259 = 0.0178 mol

The initial molar concentration and equilibrium molar concentration of different species may be represented as:

$2NO (g) + Br_2(g)$ ————————> $2NOB_r(g)$

Initial moles 0.087 0.0437 0

Moles at eqm. point: 0.0352 0.0178 0.0518

Question 10.

At 450 K, $K_p = 2.0 \times 10^{10}$ bar^{-1} for the equilibrium reaction:

$$2SO_2\,(g) + O_2\,(g) \rightleftharpoons 2SO_3\,(g)$$

What is K_c at this temperature?

Answer:

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

K_p = 2.0×10^{10} bar^{-1} ; R = 0.083 L bar K^{-1} mol^{-1} ; T = 450 K; $\Delta^{ng} = 2 - 3 = -1$

K_c = $(2.0 \times 10^{10}\ \text{bar}^{-1}) \times [(0.083\ \text{L bar K}^{-1}\ \text{mol}^{-1}) \times (450\ \text{K})]^{-(-1)}$

= 7.47×10^{11} mol^{-1} L = $\mathbf{7.47 \times 10^{11}\ M^{-1}}$

Question 11. A sample of HI (g) is placed in a flask at a pressure of 0.2 atm. At equilibrium partial pressure of HI (g) is 0.04 atm. What is K_p for the given equilibrium?

$$2HI\,(g) \rightleftharpoons H_2(g) + I_2\,(g)$$

Answer:

pHI = 0.04 atm, pH_2 = 0.08 atm; pI_2 = 0.08 atm

$$K_p = \frac{pH_2 \times pI_2}{p^2_{HI}} = \frac{(0.08\text{ atm}) \times (0.08\text{ atm})}{(0.04\text{ atm}) \times (0.04\text{ atm})} = \mathbf{4.0}$$

Question 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

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is 1.7×10^{-2}.

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

Answer:

The reaction is: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$$\text{Concentration quotient } (Q_c) = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(8.13/20\text{ mol L}^{-1})^2}{(1.57/20\text{ mol L}^{-1}) \times (1.92/20\text{ mol L}^{-1})^3}$$

$$= \mathbf{2.38 \times 10^3}$$

The equilibrium constant (K_c) for the reaction = 1.7×10^{-2}

As $Q_c \neq K_c$; this means that the reaction is not in a state of equilibrium.

Question 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: Balanced chemical equation for the reaction is 4

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

Question 14. If l mole of H20 and 1 mole of CO are taken in a 10 litre vessel and heated to 725 K, at equilibrium point 40 percent of water (by mass) reacts with carbon monoxide according to equation.

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

Calculate the equilibrium constant for the reaction.

Answer: Number of moles of water originally present = 1 mol

Percentage of water reacted =40%

Number of moles of water reacted = 1 x 40/100 = 0.4 mol

Number of moles of water left = (1 – 0.4) = 0.6 mole According to the equation, 0.4 mole of water will react with 0.4 mole of carbon monoxide to

form 0.4 mole of hydrogen and 0.4 mole of carbon dioxide.
Thus, the molar cone, per litre of the reactants and products before the reaction and at the equilibrium point are as follows:

	$H_2O(g)$ +	$CO(g)$ $\rightleftharpoons$	$H_2(g)$ +	$CO_2(g)$
Initial moles/litre	$\frac{1}{10}$	$\frac{1}{10}$	0	0
Mole/litre at the equilibrium point	$\frac{1-0.4}{10}=\frac{0.6}{10}$	$\frac{1-0.4}{10}=\frac{0.6}{10}$	$\frac{0.4}{10}$	$\frac{0.4}{10}$

Applying law of chemical equilibrium,

$$\text{Equilibrium constant } (K_c) = \frac{[H_2(g)][CO_2(g)]}{[H_2O(g)][CO(g)]} = \frac{\left(\frac{0.4}{10}\text{mol L}^{-1}\right)\times\left(\frac{0.4}{10}\text{mol L}^{-1}\right)}{\left(\frac{0.6}{10}\text{mol L}^{-1}\right)\times\left(\frac{0.6}{10}\text{mol L}^{-1}\right)}$$

$$= \frac{0.16}{0.36} = \mathbf{0.44}$$

Question 15. What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78 M?

$2ICl(g) \rightleftharpoons I_2(g) + Cl_2(g)\,;\ K_c = 0.14$

Answer:

Suppose at equilibrium, the molar concentration of both I_2 (g) and Cl_2 (g) is x mol L^{-1}.

	$2ICl(g)$ $\rightleftharpoons$	$I_2(g)$ +	$Cl_2(g)$
Initial molar conc.	0.78 M	0	0
Eqn. molar conc.	$(0.78-2x)$M	x	x

$$K_c = \frac{[I_2(g)][Cl_2(g)]}{[ICl(g)^2]} = \frac{(x)\times(x)}{(0.78-2x)^2}$$

$$\frac{x}{(0.78-2x)} = (0.14)^{1/2} = 0.374 \quad \text{or} \quad x = 0.374\,(0.78-2x)$$

$$x = 0.292 - 0.748x \quad \text{or} \quad 1.748x = 0.292\,;\ x = \frac{0.292}{1.748} = 0.167$$

$$[ICl] = (0.78 - 2\times 0.167) = (0.78 - 0.334) = \mathbf{0.446\ M}$$

$$[I_2] = 0.167\text{ M}\,;\ [Cl_2] = \mathbf{0.167\ M}$$

Question 16. K =0.04 atm at 898 K for the equilibrium shown below. What is the equilibrium concentration ok C_2H_6 when it is placed in a flask at 4 atm pressure,and allowed to come to equilibrium.

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

Answer:

The equilibrium in the reaction is:

	$C_2H_6(g)$	$\rightleftharpoons$	$C_2H_4(g)$	+	$H_2(g)$
Initial pressure:	4 atm		0		0
Eqn. pressure:	$(4-p)$ atm		p atm		p atm

$$K_p = \frac{p_{C_2H_4} \times p_{H_2}}{p_{C_2H_6}} \quad \text{or} \quad 0.04 = \frac{p^2}{(4-p)}$$

$$p^2 = 0.04\,(4-p) \quad \text{or} \quad p^2 + 0.04\,p - 0.16 = 0$$

$$p = \frac{(-0.04) \pm \sqrt{0.0016 - 4\,(-0.16)}}{2}$$

$$= \frac{(-0.04) \pm 0.8}{2} = \frac{0.76}{2} = 0.38$$

Equilibrium pressure or concentration of C_2H_6 = (4 – 0.38) = **3.62 atm.**

Question 17. The ester, ethyl acetate is formed by the reaction of ethanol and acetic acid and the equilibrium is represented as:

$CH_3COOH(l) + C2H5OH(l)$———> $CH_3COOC_2H_5(l) + H_2O(l)$

(i) Write the concentration ratio (concentration quotient) Q for this reaction. Note that water is not in excess and is not a solvent in this reaction.

(ii) At 293 K, if one starts with 1.000 mol of acetic acid and 0.180 mol of ethanol, there is 0.171 mol of ethyl acetate in the final equilibrium mixture. Calculate the equilibrium constant.

(iii) Starting mth 0.50 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 some time. Has equilibrium been reached?

Answer:

(*i*) The concentration ratio (Concentration quotient) Q_c for the reaction is:

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

(*ii*)

	$CH_3COOH(l)$ +	$C_2H_5OH(l)$ $\rightleftharpoons$	$CH_3COOC_2H_5(l)$ +	$H_2O(l)$
Initial molar conc.	1.0 mol	0.18 mol	0	0
Molar conc. at equilibrium point	(1 – 0.171) = 0.829 mol	(0.18 – 0.171) = 0.009 mol	0.171 mol	0.171 mol

Applying Law of Chemical equilibrium,

$$K_c = \frac{[CH_3COOC_2H_5](l)[H_2O(l)]}{[CH_3COOH(l)][C_2H_5OH(l)]}$$

$$= \frac{(0.171\text{ mol}) \times (0.171\text{ mol})}{(0.829\text{ mol})\,(0.009\text{ mol})} = \mathbf{3.92}$$

(*iii*)

	$CH_3COOH(l)$ +	$C_2H_5OH(l)$ $\rightleftharpoons$	$CH_3COOC_2H_5(l)$ +	$H_2O(l)$
Initial molar conc.	1.0 mol	0.5 mol	0.214 mol	0.214 mol
Molar conc. at equilibrium	1.0 – 0.214 = 0.786 mol	0.5 – 0.214 = 0.286 mol		

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

$$= \frac{(0.214\text{ mol}) \times (0.214\text{ mol})}{(0.786\text{ mol})\,(0.286\text{ mol})} = \mathbf{0.204}$$

Since Q_c is less than K_c this means that the equilibrium has not been reached. The reactants are still taking part in the reaction to form the products.

Question 18. A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was reached, the concentration of PCl_5 was found to be 0.5×10^{-1} mol L^{-1}. If K_c is 8.3×10^{-3} what are the concentrations of PCl_3 and Cl_2 at equilibrium?

Answer: Let the initial molar concentration of PCl_5 per litre = x mol

Molar concentration of PCl_5 at equilibrium = 0.05 mol

∴ Moles of PCl_5 decomposed = (x – 0.05) mol

Moles of PCl_3 formed = (x – 0.05) mol

Moles of Cl_2 formed = (x – 0.05) mol

The molar conc./litre of reactants and products before the reaction and at the equilibrium point are:

	PCl_5	$\rightleftharpoons$	PCl_3	+	Cl_2
Initial moles/litre	x		0		0
Moles/litre at eqm. point	0.05		$(x - 0.05)$		$(x - 0.05)$

Equilibrium constant (K_c) = 8.3×10^{-3} = 0.0083

Applying Law of chemical equilibrium,

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}; \quad 0.0083 = \frac{(x - 0.05) \times (x - 0.05)}{0.05}$$

$$(x - 0.05)^2 = 0.0083 \times 0.05 = 4.15 \times 10^{-4}$$

$$(x - 0.05) = (4.15 \times 10^{-4})^{1/2} = 2.037 \times 10^{-2} = 0.02 \text{ moles}$$

$$x = 0.05 + 0.02 = 0.07 \text{ mol}$$

The molar concentration per litre of PCl_3 at eqm. = 0.07 – 0.05 = **0.02 mol**

The molar concentration per litre of Cl_2 at eqm. = 0.07 – 0.05 = **0.02 mol.**

Question 19. 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 $C0_2$

$FeO(s) + CO(g)$ ———>$Fe(s) + C0_2(g)$; $K_p = 0.265$ atm at 1050 K

What are the equilibrium partial pressures of CO and $C0_2$ at 1050 K if the initial pressures are: $P_{CO} = 1.4$ atm and $P_{CO2} = 0.80$ atm?

Answer:

	$FeO(s) + CO(g)$	$\rightleftharpoons$	$Fe(s) + CO_2(g)$
Initial pressure:	1.4 atm		0.8 atm

$$Q_p = \frac{p_{CO_2}}{p_{CO}} = \frac{(0.8 \text{ atm})}{(1.4 \text{ atm})} = 0.571$$

Since $Q_p > K_p (0.265)$, this means that the reaction will move in the backward direction to attain the equilibrium. Therefore, partial pressure of $C0_2$ will decrease while that of CO will increase so that the equilibrium may be attained again. Let p atm be the decrease in the partial pressure of $C0_2$. Therefore, the partial pressure of CO will increase by the same magnitude i.e., p atm.

$$p_{CO_2} = (0.8 - p) \text{ atm}; \; p_{CO}(g) = (1.4 + p) \text{ atm}$$

At equilibrium, $$K_p = \frac{p_{CO_2}}{p_{CO}} = \frac{(0.8 - p)\text{ atm}}{(1.4 + p)\text{ atm}} = \frac{(0.8 - p)}{(1.4 + p)}$$

or $$0.265 = \frac{(0.8-p)}{(1.4+p)}$$

$$0.371 + 0.265\,p = 0.8 - p \quad \text{or} \quad 1.265\,p = 0.8 - 0.371 = 0.429$$

$$p = 0.429/1.265 = 0.339 \text{ atm}$$

$$(p_{CO})_{eq} = (1.4 + 0.339) = \mathbf{1.739\ atm}$$

$$(p_{CO_2})_{eq} = (0.8 - 0.339) = \mathbf{1.461\ atm}$$

Question 20.

Equilibrium constant K_c for the reaction, N_2 (g) + $3H_2$ (g) $\rightleftharpoons$ $2NH_3$ (g) at 500 K is 0.061. At particular time, the anaylsis shows that the composition of the reaction mixture is: 3.0 mol L^{-1} of N_2 ; 2.0 mol L^{-1} of H_2; 0.50 mol L^{-1} of NH_3. Is the reaction at equilibrium? If not, in which direction does the reaction tend to proceed to reach the equilibrium?

Answer:

The given reaction is:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

According to available data.

$$N_2 = [3.0]\,;\, H_2 = [2.0]\,;\, NH_3 = [0.50]$$

$$Q_c = \frac{[NH_3(g)]^2}{[N_2(g)]\,[H_2(g)]^3} = \frac{[0.50]^2}{[3.0]\,[2.0]^3} = \frac{0.25}{24} = \mathbf{0.0104.}$$

Since the value of Q_c is less than that of K_c (0.061), the reaction is not in a state of equilibrium. It will proceed in the **forward direction** till Q_c becomes the same as K_c.

Question 21. Bromine monochloride (B_rCl) decomposes into bromine and chlorine and reaches the equilibrium:

$$2BrCl(g) \rightleftharpoons Br_2(g) + Cl_2(g)$$

The value of Kc is 32 at 500 K. If initially pure B_rCl is present at a concentration of 3.3 x10^{-3}mol L^{-1}what is its molar concentration in the mixture at equilibrium?

Answer: Let x moles of B_rCl decompose in order to attain the equilibrium. The initial molar concentration and the molar concentration at equilibrium point of different species may be represented as follows:

	$2BrCl(g) \rightleftharpoons$	$Br_2(g)$ +	$Cl_2(g)$
Initial moles/litre	0.0033	0	0
Moles/litre at eqm. point	$0.0033 - x$	$x/2$	$x/2$

Applying Law of chemical equilibrium, $K_c = \frac{[Br_2][Cl_2]}{[BrCl]^2}$ or $32 = \frac{(x/2) \times (x/2)}{(0.0033 - x)^2}$

On taking the square root, $5.656 = \frac{x/2}{(0.0033 - x)}$

$$\frac{x}{(0.0033 - x)} = 11.31 \quad \text{or} \quad 12.31x = 0.037 \quad ; \quad x = \frac{0.037}{12.31} = 0.003$$

$\therefore$ Molar concentration of BrCl at equilibrium point = 0.0033 − 0.003
= 0.0003 mol L^{-1} = **3×10^{-4} mol L^{-1}**

Question 22. At 1127 K and 1 atmosphere pressure, a gaseous mixture of CO and CO_2 in equilibrium with solid carbon has 90.55% CO by mass.

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

Calculate K_c for the reaction at the above temperature.

Answer:

Step I: Calculation of K_p for the reaction
Let the total mass of the gaseous mixture = 100 g
Mass of CO in the mixture = 90.55 g
Mass of CO_2 in the mixture = (100 – 90.55) = 9.45 g

$$\text{No. of moles of CO} = \frac{90.55\text{g}}{(28g\ \text{mol}^{-1})} = 3.234\ \text{mol}$$

$$\text{No. of moles of } CO_2 = \frac{9.45}{(44g\ \text{mol}^{-1})} = 0.215\ \text{mol}$$

$$p_{CO}\ \text{in the mixture} = \frac{(3.234\ \text{mol})}{(3.234 + 0.215)} \times 1\ \text{atm} = \frac{(3.234\ \text{mol})}{(3.449\ \text{mol})} \times 1\ \text{atm} = 0.938\ \text{atm}$$

$$p_{CO_2}\ \text{in the mixture} = \frac{(0.215\ \text{mol})}{(3.449\ \text{mol})} \times 1\ \text{atm} = 0.062\ \text{atm}$$

$$\begin{array}{llcl} & C(s) + CO_2(g) & \rightleftharpoons & 2CO(g) \\ \text{Eqm. pressure} & \quad 0.062\ \text{atm} & & 0.938\ \text{atm} \end{array}$$

$$K_p = \frac{p^2CO}{pCO_2} = \frac{(0.938\ \text{atm})^2}{(0.062\ \text{atm})} = 14.19\ \text{atm}$$

Step II. Calculation of K_c for the reaction.

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

$K_p = 14.19$ atm, $R = 0.0821\ \text{L atm K}^{-1}\ \text{mol}^{-1}$, $T = 1127$ K ; $\Delta^{ng} = 2 - 1 = 1$

$$K_c = \frac{(14.19\ \text{atm})}{(0.0821\ \text{L atm K}^{-1}\ \text{mol}^{-1}) \times (1127\ \text{K})^1} = \mathbf{6.46}$$

Question 23. Calculate (a) ΔG^{-} and (b) the equilibrium constant for the formation of $N0_2$ from NO and 0_2 at 298 K

$NO(g) + 1/2\ O_2(g) \rightleftharpoons NO_2(g)$ *where*

$\Delta_f G^{\ominus}\ (NO_2) = 52.0$ *kJ/mol,* $\Delta_f G^{\ominus}\ (NO) = 87.0$ *kJ/mol,* $\Delta_f G^{\ominus}\ (O_2) = 0$ *kJ/mol,*

Answer:

Step I. Calculation of $\Delta G^{\ominus}$

$$\Delta G^{\ominus} = \Delta_f G^{\ominus}\ (NO_2) - [\Delta_f G^{\ominus}(1/2O_2)]$$
$$= 52.0 - (87 + 0) = -35\ \text{kJ mol}^{-1}$$

Step II. Calculation of K_c

$$\Delta G^{\ominus} = -2.303\ RT \log K_c$$

$$\log K_c = -\frac{\Delta G^{\ominus}}{2.303\ RT} = -\frac{(-35 \times 10^3\ \text{J mol}^{-1})^3}{2.303 \times (8.314\ \text{kJ mol}^{-1}) \times (298\ \text{K})} = 6.134$$

$$K_c = \text{Antilog } 6.314 = \mathbf{1.36 \times 10^6.}$$

Question 24. 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 bp increasing the volume?

$$\begin{array}{ll}
(i) & PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) \\
(ii) & CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s) \\
(iii) & 3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)
\end{array}$$

Answer: (i) Pressure will increase in the forward reaction and number of moles of products will increase.

(ii) Pressure will increase in backward reaction and number of moles of products will decrease.

(iii) The change in pressure will have no effect on the equilibrium constant and there will be no change in the number of moles.

Question 25. Which of the following reactions will get affected by increase in pressure ? Also mention whether the change will cause the reaction to go to the right or left direction.

$$\begin{array}{ll}
(i) & CH_4(g) + 2S_2(g) \rightleftharpoons CS_2(g) + 2H_2S(g) \\
(ii) & CO_2(g) + C(s) \rightleftharpoons 2CO(g) \\
(iii) & 4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) \\
(iv) & C_2H_4(g) + H_2(g) \rightleftharpoons C_2H_6(g) \\
(v) & COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g) \\
(vi) & CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)
\end{array}$$

Answer: Only those reactions will be affected by increasing the pressure in which the number of moles of the gaseous reactants and products are different ($n_p \neq n_r$) (gaseous). With the exception of the reaction (1); all the reamaining five reactions will get affected by increasing the pressure. In general,

- The reaction will go to the left if $n_p > n_r$.
- The reaction will go to the right if $n_r > n_p$.

  Keeping this in mind,

(i) Increase in pressure will not affect equilibrium because $n_p = n_r = 3$.

(ii) Increase in pressure will favour backward reaction because n_p (2) > n_r (1)

(iii) Increase in pressure will favour backward reaction because n_p (10) >

n_r (9)

(iv) Increase in pressure will favour forward reaction because n_p (1) < n_r (2)

(v) Increase in pressure will favour backward reaction because n_p (2) > n_r(1)

(vi) Increase in pressure will favour backward reaction because n_p (1) > n_r (0).

Question 26. The equilibrium constant for the following reaction is 1.6 x 10^5at 1024 K.

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

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

Ans.

Step I. *Calculation of* K_p.

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

$$K_p = K_c(RT)^{\Delta n} = K_c\,(RT)^0 \qquad (\because \Delta_n = 2 - 2 = \text{zero})$$

$$K_p = K_c = 1.6 \times 10^5.$$

Step II. *Calculation of partial pressure of gases*

	$2HBr(g)$	$\rightleftharpoons$	$H_2(g)$	+	$Br_2(g)$
Initial pressure	10 bar		zero		zero
Eqm. pressure	(10 – P) bar		P/2 bar		P/2 bar

$$K_p' = \frac{pH_2 \times pBr_2}{p^2HBr} \text{ or } \frac{1}{1.6\times10^5} = \frac{P/2\times P/2}{(10-P)^2} = \frac{P^2}{4(10-P)^2}$$

On taking square root; $\dfrac{P^2}{4(10-P)^2} = \left(\dfrac{1}{1.6\times10^5}\right)^{1/2}$ or $\dfrac{2(10-P)}{P} = (1.6\times10^5)^{1/2}$

$$= 4 \times 10^2$$

$$20 - 2P = 4 \times 10^2\ P \quad \text{or} \quad P(4 \times 10^2 + 2) = 20$$

or $$P = \frac{20}{(400+2)} = \frac{20}{402} = 0.050 \text{ bar}$$

pH_2 = 0.025 bar; pBr_2 = **0.025 bar:** pHBr = 10 – 0.05 = 9.95 bar ≈ 10.0 bar.

Question 27. Hydrogen gas is obtained from the natural gas by partial oxidation with steam as per the following endothermic reaction:

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

Write the expression for Kp for the above reaction

How will the value of K_p and composition of equilibrium mixture be affected by:

(i) increasing the pressure, (ii) increasing the temperature, (iii) using a catalyst?

Answer: The expression for K_p for the reaction is:

$$K_p = \frac{(pCO)\times(pH_2)^3}{(pCH_4)\times(pH_2O)}$$

(i) By increasing the pressure, the number of moles per unit volume will increase. In order to decrease the same, the equilibrium gets shifted to the left or in the backward direction. As a result, more of reactants will be formed and the value of K_p will decrease.

(ii) If the temperature is increased, according to Le Chatelier's principle, the forward reaction will be favoured as it is endothermic. Therefore, the equilibrium gets shifted to the right and the value of K_p will increase.

(iii) The addition of catalyst will not change the equilibrium since it influences both the forward and the backward reactions to the same extent. But it will be attained more quickly.

Question 28. What is the effect of:

(i) addition of H_2 (ii) addition of CH_3OH

(iii) removal of CO (iv) removal of CH_3OH

On the equilibrium $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

Answer: (i) Equilibrium will be shifted in the forward direction.

(ii) Equilibrium will be shifted in the backward direction.

(iii) Equilibrium will be shifted in the backward direction.

(iv) Equilibrium will be shifted in the forward direction.

Question 29. At 473 K, the equilibrium constant K_c for the decomposition of phosphorus pentachloride (PCl_5) is 8.3 x 10^{-3} . if decomposition proceeds as:

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)\ ;\ \Delta H = +124.0\ kJ\ mol^{-1},$$

(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 of PCl_5 is added (ii) Temperature is increased.

Answer:

(a) The expression for $K_c = \frac{[PCl_3(g)][Cl_2(g)]}{[PCl_5(g)]}$

(b) For reverse reaction $(K_c') = \frac{PCl_5(g)}{[PCl_3(g)][Cl_2(g)]} = \frac{1}{8.3 \times 10^{-3}} = 120.48$

(c) (i) By adding more of PCl_5, value of K_c will remain constant because there is no change in temperature.

(ii) By increasing the temperature, the forward reaction will.be favoured since it is endothermic in nature. Therefore, the value of equilibrium constant will increase.

Question 30. 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.

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

If a reaction vessel at 400°C is charged with an equimolar mixture of CO and steam so that $P_{CO} = PH_2O = 4.0$ bar, what will be the partial pressure of H_2 at equilibrium? $K_p = 0.1$ at 400°C.

Answer:

Let the partial pressure of hydrogen (H_2) at equilibrium point = p bar

	$CO(g)$ +	$H_2O(g)$ $\rightleftharpoons$	$CO_2(g)$ +	$H_2(g)$
Initial pressure:	4.0 bar	4.0 bar	0	0
Eqm. pressure:	$(4-p)$bar	$(4-p)$ bar	p bar	p bar

$$K_p = \frac{p_{CO_2} \times p_{H_2}}{p_{CO} \times p_{H_2O}} \quad \text{or} \quad 0.1 = \frac{(p \text{ bar}) \times (p \text{ bar})}{(4-p) \text{ bar} \times (4-p) \text{ bar}}$$

$$\frac{p^2}{(4-p)^2} = 0.1 \quad \text{or} \quad \frac{p}{(4-p)} = (0.1)^{1/2} = 0.316$$

$$p = 0.316\,(4-p) \quad \text{or} \quad p = 1.264 - 0.316\,p$$

$$1.316\,p = 1.264 \quad \text{or} \quad p = \frac{1.264}{1.316} = \mathbf{0.96\ bar}$$

Question 31. Predict which of the following will have appreciable concentration of reactants and products:

(a) $Cl_2(g) \rightleftharpoons 2Cl(g)$; $K_c = 5 \times 10^{-39}$
(b) $Cl_2(g) + 2NO\,(g) \rightleftharpoons 2NOCl\,(g)$; $K_c = 3.7 \times 10^8$
(c) $Cl_2(g) + 2NO_2(g) \rightleftharpoons 2NO_2Cl\,(g)$; $K_c = 1.8$.

Answer: Following conclusions can be drawn from the values of K_c .

(a) Since the value of K_c is very small, this means that the molar concentration of the products is very small as compared to that of the reactants.

(b) Since the value of K_c is quite large, this means that the molar concentration of the products is very large as compared to that of the reactants.

(c) Since the value of K_c is 1.8, this means that both the products and reactants have appreciable concentration.

Question 32.The value of K_c for the reaction 30_2(g) —>20_3(g) is 2.0 x 10^{-50} at 25°C. If equilibrium concentration of 02 in air at 25°C is 1.6 x 10^{-2}, what is the concentration of O_3?

Answer:

$$3O_2(g) \rightleftharpoons 2O_3(g)$$

$$K_c = \frac{[O_3]^2}{[O_2]^3} \quad \text{or} \quad (2.0 \times 10^{-50}) = \frac{[O_3]^2}{(1.6 \times 10^{-2})^3}$$

$$\text{or} \quad [O_3]^2 = (2.0 \times 10^{-50}) \times (1.6 \times 10^{-2})^3$$

$$[O_3]^2 = 8.192 \times 10^{-56} \quad \text{or} \quad [O_3] = (8.192 \times 10^{-56})^{1/2} = \mathbf{2.86 \times 10^{-28}\ M.}$$

Question 33.

The reaction $CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$ 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:

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

According to available data

$$K_c = \frac{[CH_4] \times [H_2O]}{[CO] \times [H_2]^3} \quad \text{or} \quad 3.90 = \frac{[CH_4] \times [0.02]}{[0.30] \times [0.1]^3}$$

$$[CH_4] = \frac{(3.9) \times (0.30) \times (0.001)}{(0.02)} = \mathbf{5.85 \times 10^{-2}\ M}$$

Question 34. What is meant by conjugate acid-base pair? Find the conjugate acid/base for the following species: HNO_2, CH^-, $HClO_4$, OH^-, CO_3^{2-}, S^{2-}

Answer: An acid-base pair which differs by a proton only (HA———> A^- + H^+) is known as conjugate acid-base pair.

Conjugate acid:HCN,H_20,HCO_3^-,HS^-.

Conjugate base:NO_2^-,ClO_4^-,O_2^-

Question 35. Which of the following are Lewis Acids?

H_2O,BF_3, H^+ and NH^{4+},

Answer: BF_3, H^+ ions are Lewis acids.

Question 36. What will be the conjugate bases for the Bronsted acids?HF, H_2S0_4 and H_2C0_3?

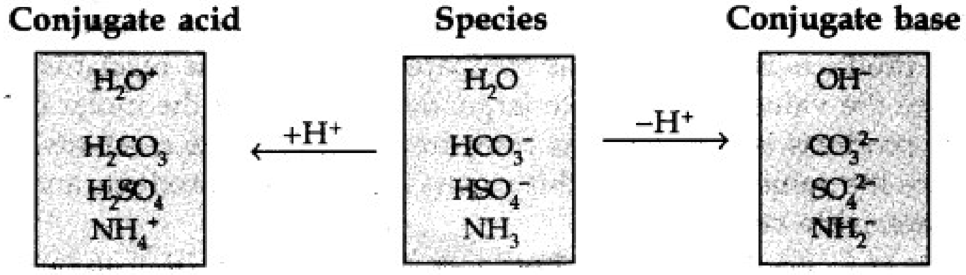
Answer: Conjugate bases: F^-, HSO^-_4 , HCO^-_3.

Question 37. Write the conjugate acids for the following Bronsted bases: NH_2, NH_3 and $HCOO^-$

Answer: NH_3, NH_4^+ and HCOOH

Question 38.The species H_20, HCO_3^-, HSO_4^- and NH_3 can act both as Bronsted acid and base. For each case, give the corresponding conjugate acid and base.

Answer:



Question 39. Classify the following species into Lewis acids and Lewis bases and show how these can act as Lewis acid/Lewis base?

(a) OH^- ions (b) F^- (c) H^+ (d) BCl_3

Answer: (a) OH^- ions can demate an electron pair and act as Lewis base.

(b) F^- ions can donate an electron pair and act'as Lewis base.

(c) H^+ ions can accept an electron pair and act as Lewis acid.

(d) BCl_3 can accept an electron pair since Boron atom is electron deficient. It is a Lewis acid.

Question 40. The concentration of hydrogen ions in a sample of soft drink is 3.8×10^{-3} M. What is the pH value?

Answer: pH = – log [H^+] = – log (3.8×10^{-3}) = – log 3.8 + 3 = 3 – 0.5798 = 2.4202 = 2.42

Question 41. The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.

Answer: pH = – log [H^+] or log [H^+] = – pH = – 3.76 = 4.24

.-. [H^+] = Antilog 4.24 = 1.738×10^{-4} = 1.74 x 10-4 M

Question 42. The ionization constant of HF, HCOOH and HCN at 298 K are is 6.8×10^{-4} , 1.8×10^{-4} and 4.8×10^{-9} respectively, Calculate the ionization constant of the corresponding conjugate base.

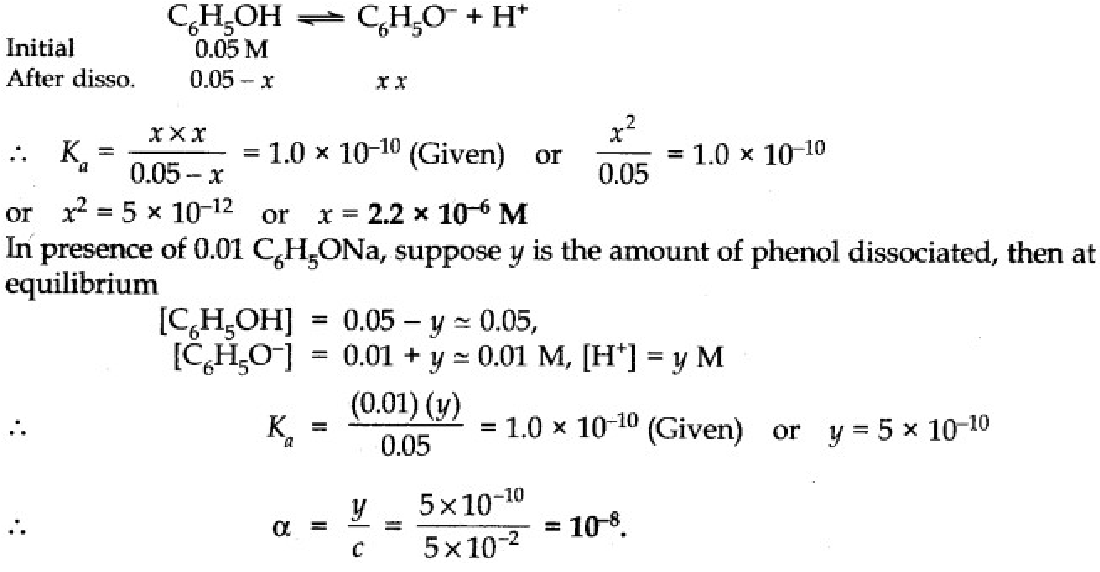
Answer: For F^-, $K_b = K_w/K_a = 10^{-14}/(6.8 \times 10^{-4}) = 1.47 \times 10^{-11} = 1.5 \times 10^{-11}$.

For HCOO-, $K_b = 10^{-14}/(1.8 \times 10^{-4}) = 5.6 \times 10^{-11}$

For CN^-, $K_b = 10^{-14}/(4.8 \times 10^{-9}) = 2.08 \times 10^{-6}$

Question 43. 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 0.01 M in sodium phenolate?

Answer:



$$C_6H_5OH \rightleftharpoons C_6H_5O^- + H^+$$

Initial 0.05 M

After disso. $0.05 - x$ x x

$$\therefore \quad K_a = \frac{x \times x}{0.05 - x} = 1.0 \times 10^{-10} \text{ (Given)} \quad \text{or} \quad \frac{x^2}{0.05} = 1.0 \times 10^{-10}$$

or $x^2 = 5 \times 10^{-12}$ or $x = \mathbf{2.2 \times 10^{-6}\ M}$

In presence of 0.01 C_6H_5ONa, suppose y is the amount of phenol dissociated, then at equilibrium

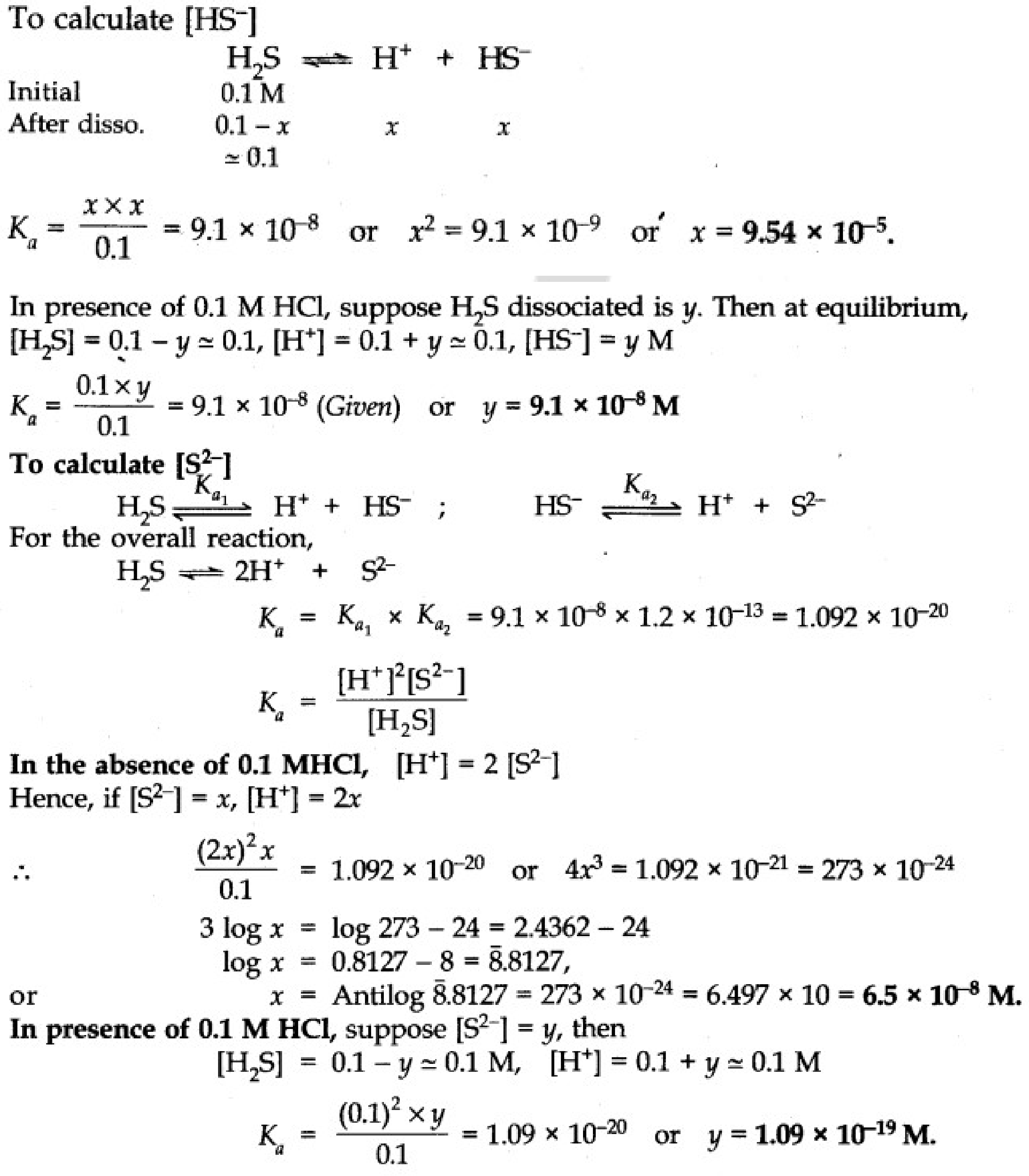
$$[C_6H_5OH] = 0.05 - y \simeq 0.05,$$
$$[C_6H_5O^-] = 0.01 + y \simeq 0.01\ M,\ [H^+] = y\ M$$

$$\therefore \quad K_a = \frac{(0.01)(y)}{0.05} = 1.0 \times 10^{-10} \text{ (Given)} \quad \text{or} \quad y = 5 \times 10^{-10}$$

$$\therefore \quad \alpha = \frac{y}{c} = \frac{5 \times 10^{-10}}{5 \times 10^{-2}} = \mathbf{10^{-8}}.$$

Question 44. The- first ionization constant of H_2S is 9.1×10^{-8}. Calculate the concentration of HS^- ions in its 0.1 M solution and 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:



To calculate $[HS^-]$

$$\begin{array}{lccc} & H_2S & \rightleftharpoons H^+ & + \; HS^- \\ \text{Initial} & 0.1\text{ M} & & \\ \text{After disso.} & 0.1 - x & x & x \\ & \simeq 0.1 & & \end{array}$$

$$K_a = \frac{x \times x}{0.1} = 9.1 \times 10^{-8} \quad \text{or} \quad x^2 = 9.1 \times 10^{-9} \quad \text{or} \quad x = \mathbf{9.54 \times 10^{-5}}.$$

In presence of 0.1 M HCl, suppose H_2S dissociated is y. Then at equilibrium,
$[H_2S] = 0.1 - y \simeq 0.1$, $[H^+] = 0.1 + y \simeq 0.1$, $[HS^-] = y$ M

$$K_a = \frac{0.1 \times y}{0.1} = 9.1 \times 10^{-8} \text{ (Given)} \quad \text{or} \quad y = \mathbf{9.1 \times 10^{-8}\ M}$$

To calculate $[S^{2-}]$

$$H_2S \xrightleftharpoons{K_{a_1}} H^+ + HS^- \quad ; \qquad HS^- \xrightleftharpoons{K_{a_2}} H^+ + S^{2-}$$

For the overall reaction,

$$H_2S \rightleftharpoons 2H^+ + S^{2-}$$

$$K_a = K_{a_1} \times K_{a_2} = 9.1 \times 10^{-8} \times 1.2 \times 10^{-13} = 1.092 \times 10^{-20}$$

$$K_a = \frac{[H^+]^2[S^{2-}]}{[H_2S]}$$

In the absence of 0.1 MHCl, $[H^+] = 2\,[S^{2-}]$
Hence, if $[S^{2-}] = x$, $[H^+] = 2x$

$$\therefore \quad \frac{(2x)^2 x}{0.1} = 1.092 \times 10^{-20} \quad \text{or} \quad 4x^3 = 1.092 \times 10^{-21} = 273 \times 10^{-24}$$

$$3 \log x = \log 273 - 24 = 2.4362 - 24$$

$$\log x = 0.8127 - 8 = \bar{8}.8127,$$

$$\text{or} \quad x = \text{Antilog } \bar{8}.8127 = 273 \times 10^{-24} = 6.497 \times 10 = \mathbf{6.5 \times 10^{-8}\ M.}$$

In presence of 0.1 M HCl, suppose $[S^{2-}] = y$, then

$$[H_2S] = 0.1 - y \simeq 0.1\text{ M}, \quad [H^+] = 0.1 + y \simeq 0.1\text{ M}$$

$$K_a = \frac{(0.1)^2 \times y}{0.1} = 1.09 \times 10^{-20} \quad \text{or} \quad y = \mathbf{1.09 \times 10^{-19}\ M.}$$

Question 45. 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 ions in the solution and its pH.

Answer:

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

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} = \frac{[H^+]^2}{[CH_3COOH]}$$

or $$[H^+] = \sqrt{K_a\,[CH_3COOH]} = \sqrt{(1.74\times10^{-5})(5\times10^{-2})} = 9.33\times10^{-4}\text{ M}$$

$$[CH_3COO^-] = [H^+] = \mathbf{9.33\times10^{-4}\ M}$$

$$pH = -\log(9.33\times10^{-4}) = 4 - 0.9699 = 4 - 0.97 = \mathbf{3.03}$$

Question 46. It has been found that the pH of a 0.01 M 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:

$$HA \rightleftharpoons H^+$$

$$pH = -\log[H^+] \quad \text{or} \quad \log[H^+] = -4.15 = \bar{5}.85$$

$$\therefore \quad [H^+] = 7.08\times10^{-5}\text{ M} = 7.08\times10^{-5}\text{ M}$$

$$[A^-] = [H^+] = 7.08\times10^{-5}\text{ M}$$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(7.08\times10^{-5})(7.08\times10^{-5})}{10^{-2}} = \mathbf{5.0\times10^{-7}}$$

$$pK_a = -\log K_a = -\log(5.0\times10^{-7}) = 7 - 0.699 = \mathbf{6.301}$$

Question 47. Assuming complete dissociation, calculate the pH of the following solutions:

(a) 0.003 M HCl (b) 0.005 M NaOH (c) 0.002 M HBr (d) 0.002 M KOH

Answer:

(*a*) $HCl + aq \rightarrow H^+ + Cl^-$, $\therefore [H^+] = [HCl] = 3\times10^{-3}$ M, $pH = -\log(3\times10^{-3}) = 2.52$

(*b*) $NaOH + aq \rightarrow Na^+ + OH^-$

$$\therefore \quad [OH^-] = 5\times10^{-3}\text{ M},\ [H^+] = 10^{-14}/(5\times10^{-3}) = 2\times10^{-12}\text{ M}$$

$$pH = -\log(2\times10^{-12}) = \mathbf{11.70}$$

(*c*) $HBr + aq \rightarrow H^+ + Br^-$, $\therefore$ $[H^+] = 2\times10^{-3}$ M, $pH = -\log(2\times10^{-3}) = 2.70$

(*d*) $KOH + aq \rightarrow K^+ + OH^-$,

$$\therefore \quad [OH^+] = 2\times10^{-3}\text{ M},\ [H^+] = 10^{-14}/(2\times10^{-3}) = 5\times10^{-12}$$

$$pH = -\log(5\times10^{-12}) = \mathbf{11.30}$$

Question 48. Calculate the pH of the following solutions:

(a) 2g ofTlOH dissolved in water to give 2 litre of the solution

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

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

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

Answer:

$$(a)\ \text{Molar conc. of TlOH} = \frac{2g}{(204+16+1)\,g\,\text{mol}^{-1}} \times \frac{1}{2\,\text{L}} = 4.52 \times 10^{-3}\ \text{M}$$

$$[\text{OH}^-] = [\text{TlOH}] = 4.52 \times 10^{-3}\ \text{M}$$

$$[\text{H}^+] = 10^{-14}/(4.52 \times 10^{-3}) = 2.21 \times 10^{-12}\ \text{M}$$

$$\therefore \quad \text{pH} = -\log(2.21 \times 10^{-12}) = 12 - (0.3424) = 11.66$$

$$(b)\ \text{Molar conc. of Ca(OH)}_2 = \frac{0.3\,g}{(40+34)g\,\text{mol}^{-1}} \times \frac{1}{0.5\,\text{L}} = 8.11 \times 10^{-3}\ \text{M}$$

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

$$[\text{OH}^-] = 2[\text{Ca(OH)}_2] = 2 \times (8.11 \times 10^{-3})\ \text{M} = 16.22 \times 10^{-3}\ \text{M}$$

$$\text{pOH} = -\log(16.22 \times 10^{-3}) = 3 - 1.2101 = 1.79$$

$$\text{pH} = 14 - 1.79 = \mathbf{12.21}$$

$$(c)\ \text{Molar conc. of NaOH} = \frac{0.3\,g}{40\,g\,\text{mol}^{-1}} \times \frac{1}{0.2\,\text{L}} = 3.75 \times 10^{-2}\ \text{M}$$

$$[\text{OH}^-] = 3.75 \times 10^{-2}\ \text{M}$$

$$\text{pOH} = -\log(3.75 \times 10^{-2}) = 2 - 0.0574 = 1.43$$

$$\text{pH} = 14 - 1.43 = \mathbf{12.57}$$

$$(d)\ M_1V_1 = M_2V_2 \therefore 13.6\ \text{M} \times 1\text{m L} = M_2 \times 1000\ \text{mL} \quad \therefore M_2 = 1.36 \times 10^{-2}\ \text{M}$$

$$[\text{H}^+] = [\text{HCl}] = 1.36 \times 10^{-2}\ \text{M},\ \text{pH} = -\log(1.36 \times 10^{-2}) = 2 - 0.1335 \simeq \mathbf{1.87}$$

Question 49. The degree of ionization of a 0.1 M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the PK_a of bromoacetic acid.

Answer:

$$\text{CH}_2\text{(Br) COOH} \rightleftharpoons \text{CH}_2\text{(Br)COO}^- + \text{H}^+$$

	$\text{CH}_2\text{(Br) COOH}$	$\text{CH}_2\text{(Br)COO}^-$	H^+
Initial conc.	C	0	0
Conc. at eqm.	C – Cα	C α	C α

$$K_a = \frac{C\alpha \,.\, C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{1-\alpha} \simeq C\alpha^2 = 0.1 \times (0.132)^2 = 1.74 \times 10^{-3}$$

$$\text{p}K_a = -\log(1.74 \times 10^{-3}) = 3 - 0.2405 = \mathbf{2.76}$$

$$[\text{H}^+] = C\,\alpha = 0.1 \times 0.132 = 1.32 \times 10^{-2}\ \text{M}$$

$$\text{pH} = -\log(1.32 \times 10^{-2}) = 2 - 0.1206 = \mathbf{1.88}$$

Question 50. The pH of0.005 M codeine ($C_{18}H_{21}N0_3$) solution is 9.95. Calculate the ionization constant and PK_b.

Answer:

$$\text{Cod} + \text{H}_2\text{O} \rightleftharpoons \text{Cod H}^+ + \text{OH}^-$$

$$\text{pH} = 9.95 \quad \therefore \text{pOH} = 14 - 9.95 = 4.05,\quad \textit{i.e.,}\ -\log[\text{OH}^-] = 4.05$$

$$\text{or} \quad \log[\text{OH}^-] = -4.05 = \bar{5}.95 \quad \text{or} \quad [\text{OH}^-] = 8.913 \times 10^{-5}\ \text{M}$$

$$K_b = \frac{[\text{CodH}^+][\text{OH}^-]}{[\text{Cod}]} = \frac{[\text{OH}^-]^2}{[\text{Cod}]} = \frac{(8.91 \times 10^{-5})^2}{5 \times 10^{-3}} = \mathbf{1.588 \times 10^{-6}}$$

$$\text{p}K_b = -\log(1.588 \times 10^{-6}) = 6 - 0.1987 = \mathbf{5.8}$$

MORE QUESTIONS SOLVED

NCERT Solutions for Class 11 Chemistry Chapter 7 Very Short Answer Type Questions

Question 1. What is meant by equilibrium?

Answer: Equilibrium is a state at which rate of forwarding reaction is equal to the rate of backward reaction.

Question 2. State the law of mass action?

Answer: It states that the rate at which a substance reacts is directly proportional to its molar concentration.

Question 3. What is meant by reaction quotient?

Answer: It is defined as the ratio of the product of molar concentration of products to the product of molar concentration of reactants at any stage of the reaction.

$$\text{Cod} + H_2O \rightleftharpoons \text{Cod } H^+ + OH^-$$

$$pH = 9.95 \quad \therefore \; pOH = 14 - 9.95 = 4.05, \quad \textit{i.e.}, -\log[OH^-] = 4.05$$

$$\text{or} \quad \log[OH^-] = -4.05 = \bar{5}.95 \quad \text{or} \quad [OH^-] = 8.913 \times 10^{-5}\,\text{M}$$

$$K_b = \frac{[\text{CodH}^+][OH^-]}{[\text{Cod}]} = \frac{[OH^-]^2}{[\text{Cod}]} = \frac{(8.91 \times 10^{-5})^2}{5 \times 10^{-3}} = \mathbf{1.588 \times 10^{-6}}$$

$$pK_b = -\log(1.588 \times 10^{-6}) = 6 - 0.1987 = \mathbf{5.8}$$

Question 4. Define ionic equilibrium.

Answer: The equilibrium between ions and unionised molecules is called ionic equilibrium.

Question 5. What is meant by ionic product of water (k_w)?

Answer: It is the product of concentration of [H_3O^+] and [OH^-] at a specific temperature.

$Kw = [H_3O^+]\,[OH^-]$

= 1.0 x 10-14 at 298 K

Question 6. Define solubility product.

Answer: It is product of molar concentration of ion raised to the power of number of ions produced per compound in saturated solution.

Question 7. How does common ion affect the solubility of electrolyte?

Answer: Solubility of electrolyte decreases due to common ion effect.

Question 8. Write conjugate add and conjugate base of H_2O?

Answer: Conjugate acid is H_3O^+ and conjugate base is OH^-.

Question 9. Give two characteristics of a buffer solution.

Answer:

- Its pH does not change on the addition of small amount of acid or base.
- Its pH does not change on dilution or standing.

Question 10. How does a catalyst affect the equilibrium constant?

Answer: The equilibrium constant is not affected by a catalyst.

Question 11. State Ostwald's dilution law.

Answer: Ostwald's dilution law states that the degree of dissociation of weak electrolyte is inversely proportional to square root of its concentration.

$$\alpha = \sqrt{\frac{K_a}{C}}, \alpha = \sqrt{\frac{K_b}{C}}$$

Where, Kfl and Kfc are acid dissociation and base dissociation constants.

Question 12. What is basic buffer?

Answer: Basic buffer is the buffer whose pH is more than 7. It is a mixture containing weak base and its salt with a strong acid, e.g., $NH_4Cl + NH_4OH$

Question 13. $S0_3^{2-}$is Bronsted base or acid and why?

Answer: $S0_3^{2-}$is Bronsted base because it can accept H^+.

Question 14.

$$H_2 (g) + I_2(g) \rightleftharpoons 2HI (g)$$

What is the relationship between K_p and K_c ?

Answer: $K_p = K_c$

because $\Delta n = 0$.

Question 15. Define common ion effect.

Answer: The supervision in concentration of one of the ions by adding other ion as common ion is called common ion effect.

NCERT Solutions for Class 11 Chemistry Chapter 7 Short Answer Type Questions

Question 1. The following concentration were obtained for the formation of NH_3 from N_2 and H_2 at equilibrium at 500 K.$[N_2(g)] = 1.5 \times 10^{-2}$ M $[H_2 (g)] = 3.0 \times 10^{-2}$ M $[NH_3] = 1.2 \times 10^{-2}$ M. Calculate equilibrium constant.

Answer:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3\ (g)$$

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

$$= \frac{[1.2\times10^{-2}M]^2}{[1.5\times10^{-2}M][3.0\times10^{-2}M]^3}$$

$$= \mathbf{3.55 \times 10^{-2}\ M}$$

Question 2. Write the equilibrium constant (K_c) expression for the following reactions.

(i) $Cu^{2+}(aq) + 2\ Ag\ (s) \rightleftharpoons Cu(s) + 2Ag^+\ (aq)$

(ii) $4HCl\ (g) + O_2(g) \rightleftharpoons 2Cl_2\ (g) + 2H_2O(g)$

Answer:

(i) $$K_c = \frac{[Ag^+(aq)]^2}{[Cu^{2+}(aq)]}$$

(ii) $$K_c = \frac{[Cl_2(g)]^2[H_2O(g)]^2}{[HCl(g)]^4[O_2(g)]}$$

Question 3. Given the equilibrium N_2O_4 (g) ——->$2NO_2$ (g) K=0.15 atm at 298 K

(a) What is K_p using pressure in torr?

(b) What is K_c using units of moles per litre.

Answer:

(a) $$K_p = \frac{(760\ \text{torr})\times(0.15\ \text{atm})}{(1\ \text{atm})}$$

$$= \mathbf{1.14 \times 10^2\ torr}$$

(b)

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

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

$$= \frac{(0.15 \text{ atm})}{\left(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}\right)^{2-1}}$$

$$= \mathbf{6.13 \times 10^{-3} \text{ mol L}^{-1}}$$

Question 4. In the reaction A + B———> C + D, what will happen to the equilibrium if concentration of A is increased?

(b) The equilibrium constant for a reaction is 2×10^{-23} at 25°C and 2×10^{-2} at 50°C. Is the reaction endothermic or exothermic?

(c) Mention at least three ways by which the concentration of $S0_3$ can be increased in the following reaction in a state of equilibrium.

Answer: (a) The reaction will shift in the forward direction.

(b) Endothermic

(c) (i) lowering the temperature (ii) increasing pressure.

(iii) increasing concentration of oxygen.

Question 5. (i) Define Le Chatelier’s principle.

(ii) Following reactions occur in a Blast furnace.

$Fe_20_3(s) + 3CO(g) \text{ ————>} 2Fe(s) + 3CO_2(g)$

use Le chatelier’s principle to predict the direction of reaction when equilibrium mixture is disturbed by

(a) adding Fe_20_3 (b) removing CO_2 .

(c) removing CO.

Answer: (a) When a system under equilibrium is subjected to a change in temperature, pressure or concentration, then the equilibrium shifts in such a direction so as to undo the effect of the change.

(ii) (a) On adding $Fe_20_3(s)$, the equilibrium will remain unaffected.

(b) By removing CO_2 (g), the equilibrium will be shifted in the forward direction.

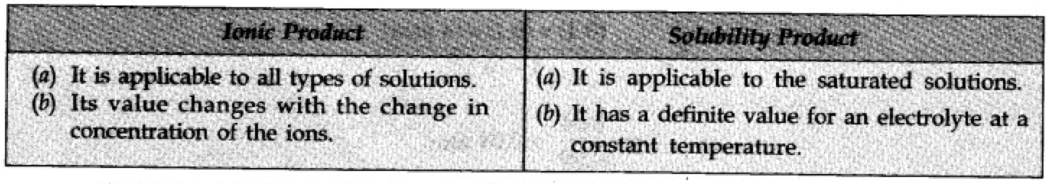
(c) By removing CO(g), the equilibrium will be shifted in the backward direction.

NCERT Solutions for Class 11 Chemistry Chapter 7 Long Answer Type Questions

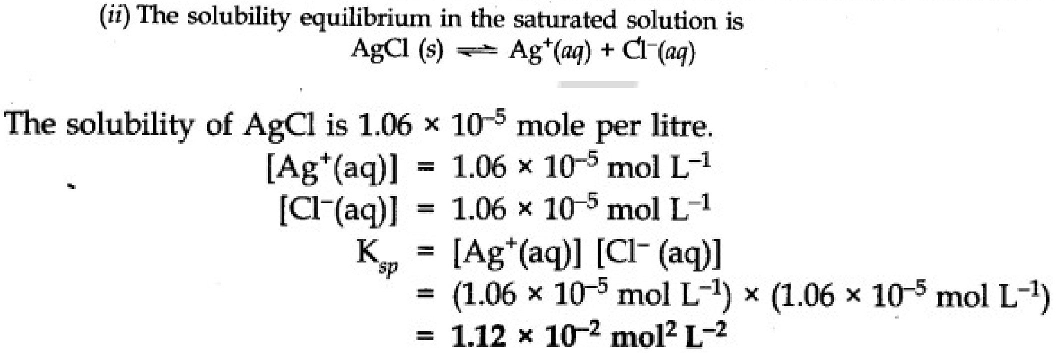
Question 1. (i) Point out the differences between ionic product and solubility product.

(ii) The solubility of AgCl in water at 298 K is 1.06×10^{-5} mole per litre. Calculate its solubility product at this temperature.

Answer: (i)



Ionic Product	*Solubility Product*
(*a*) It is applicable to all types of solutions. (*b*) Its value changes with the change in concentration of the ions.	(*a*) It is applicable to the saturated solutions. (*b*) It has a definite value for an electrolyte at a constant temperature.



(*ii*) The solubility equilibrium in the saturated solution is

$$AgCl\,(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$$

The solubility of AgCl is 1.06×10^{-5} mole per litre.

$$[Ag^+(aq)] = 1.06 \times 10^{-5}\ mol\ L^{-1}$$
$$[Cl^-(aq)] = 1.06 \times 10^{-5}\ mol\ L^{-1}$$
$$K_{sp} = [Ag^+(aq)]\,[Cl^-(aq)]$$
$$= (1.06 \times 10^{-5}\ mol\ L^{-1}) \times (1.06 \times 10^{-5}\ mol\ L^{-1})$$
$$= \mathbf{1.12 \times 10^{-2}\ mol^2\ L^{-2}}$$

Question 2. At certain temperature and under a pressure of 4 atm, PCl_5 is 10% dissociated. Calculate the pressure at which PCl_5 will be 20% dissociated at temperature remaining constant.

Answer: Calculation of K_p

$$\begin{array}{ccccc} PCl_5(g) & \rightleftharpoons & PCl_3(g) & + & Cl_2(g) \\ 1 & & 0 & & 0 \\ (1-\alpha) & & \alpha & & \alpha \end{array}$$

Total no. of moles in the equilibrium mixture $= 1 - \alpha + \alpha + \alpha$
$= (1 + \alpha)$ mol.

Let the total pressure of equilibrium mixture = p atm

Partial pressure of PCl_5

$$p_{PCl_5} = \frac{1-\alpha}{1+\alpha} \times p \text{ atm}$$

Partial pressure of $PCl_3 = \dfrac{\alpha}{1+\alpha} \times p$ atm

Partial pressure of Cl_2

$$p_{Cl_2} = \frac{\alpha}{(1+\alpha)} \times p \text{ atm}$$

$$K_p = \frac{p_{PCl_3} \times p_{Cl_2}}{p_{PCl_5}}$$

$$= \frac{\left(\dfrac{\alpha}{1+\alpha} p \text{ atm}\right) \times \left(\dfrac{\alpha}{1+\alpha} p \text{ atm}\right)}{\dfrac{1-\alpha}{1+\alpha} p \text{ atm}} = \frac{\alpha^2 p}{1-\alpha^2} \text{ atm}$$

$$P = 4 \text{ atm and } \alpha = 10\% = \frac{10}{100} = 0.1$$

$$K_p = \frac{(0.1) \times (0.1) \times (4 \text{ atm})}{1-(0.1)^2}$$

$$= \frac{0.04}{0.99} = 0.04 \text{ atm.}$$

Calculation of P under new condition

$$\alpha = 0.2, K_p = 0.04 \text{ atm}$$

$$K_p = \frac{\alpha^2 p}{1-\alpha^2} \text{ or } p = \frac{K_p(1-\alpha^2)}{\alpha^2}$$

$$= \frac{(0.04 \text{ atm})[(1-(0.2)^2]}{(0.2)^2} = \frac{0.04 \text{ atm} \times 0.96}{0.04}$$

$$= \mathbf{0.96 \text{ atm.}}$$

NCERT Solutions for Class 11 Chemistry Chapter 7 Multiple Choice Questions

Question 1. The equilibrium expression, $K_c = [C0_2]$ represents the reaction.

(a) $C(s) + O_2(g) \rightleftharpoons CO_2(g)$ (b) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

(c) $CO(g) + \frac{1}{2} O_2(g) \rightleftharpoons CO_2(g)$ (d) $CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s)$

Question 2. Hydrogen molecule (H_2) can be dissociated into hydrogen atoms (H). Which one of the following changes will not increase the number of atoms present at equilibrium?

(a) adding H atoms (b) increasing the temperature

(c) increasing the total pressure (d) increasing the volume of the container

Question 3. What is the expression for K_{eq}? for the reaction

$$2N_2O\,(g) + O_2\,(g) \rightleftharpoons 4NO\,(g)?$$

(a) $\frac{[N_2][O_2]}{[NO]}$ (b) $\frac{[NO]^4}{[N_2O]^2}$ (c) $\frac{[NO]^4}{[N_2O]^2\,[O_2]}$ (d) $\frac{[N_2O]^2\,[O_2]}{[NO]^4}$

Question 4. A catalyst will increase the rate of a chemical reaction by

(a) shifting the equilibrium to the right (b) shifting the equilibrium to the left (c) lowering the activation energy (d) increasing the activation energy

Question 5. What is the correct expression for the representation of the solubility product constant of $Ag_2\,Cr0_4$?

(a) $[Ag^+]^2\,[CrO_4^{2-}]$ (b) $[2Ag^+]\,[CrO_4^{2-}]$ (c) $[Ag^+]\,[CrO_4^{2-}]$ (d) $[2Ag^+]^2\,[CrO_4^{2-}]$

Question 6. In a closed system

$$A(S) \rightleftharpoons 4\,B(g) + 3\,C(g)$$

If partial pressure of C is doubled, then partial pressure of B will be

(a) $2\sqrt{2}$ times the original value (b) $\frac{1}{2}$ times the original value

(c) 2 times of the original value (d) $\frac{1}{2\sqrt{2}}$ times of the original value

Question 7. $H_2 + S$ ——->H_2S + energy.

In this reversible reaction, select the factor which will shift the equilibrium to the right.

(a) adding heat (b) adding H_2S

(c) blocking hydrogen gas reaction (d) removing hydrogen sulphide gas

Question 8. What effect does a catalyst have on the equilibrium position of a reaction?

(a) a catalyst favours the formation of products
(b) a catalyst favours the formation of reactants
(c) a catalyst does not change the equilibrium position of a reaction
(d) a catalyst may favour reactants or product formation, depending upon the directiofi in which the reaction is written.

Question 9. A chemist dissolves an excess of $BaSO_4$ in pure water at 25°C if its $K_{sp}= 1 \times 10^{-10}$ what is the concentration of barium in the water?
(a) 10^{-4} M (b) 10^{-5} M (C) 10^{-15} M (d) 10^{-6} M

Question 10. If in a mixture where Q = k is combined, then what happens?
(a) the reaction shift towards products
(b) the reaction shift towards reactants
(c) nothing appears to happen, but forward and reverse are continuing at the same rate
(d) nothing happens

Answer: 1.(b) 2.(c) 3.(c) 4.(c) 5.(a)
6.(d) 7.(a) 8.(c) 9.(c) 10.(c)

NCERT Solutions for Class 11 Chemistry Chapter 7 HOTS Questions

Question 1. For the equilibrium 2 NOCl(g)——-> 2NO(g) + Cl_2(g) the value of the equilibrium constant Kc is 3.75×10^{-6} at 1069 K. Calculate the K_p for the reaction at this temperature?

Answer: We know that $K_p = K_c(RT)^{\Delta n}$
For the above reaction, Δn = (2 + 1) – 2 = 1 K_p = 3.75 x 10’6 (0.0831 x 1069)
$K_p = 0.033$.

Question 2. The values of Ksp of two sparingly soluble salts $Ni(OH)_2$ and AgCN are 2.0×10^{-15} and 6×10^{-17} respectively. Which salt is more soluble? Explain.

Answer:

$AgCN \rightleftharpoons Ag^+ + CN^-$

$Ksp = [Ag^+][CN^-] = 6 \times 10^{-17}$

$Ni(OH)_2 \rightleftharpoons Ni^{2+} + 2OH^-$

$Ksp = [Ni^{2+}][OH^-]^2 = 2 \times 10^{-15}$

Let $[Ag^+] = S_1$, then $[CN^-] = S_1$

Let $[Ni^{2+}] = S_2$, then $[OH^-] = 2S_2$

$S_1^2 = 6 \times 10^{-17}, S_1 = 7.8 \times 10^{-9}$

$(S_2)(2S_2)^2 = 2 \times 10^{-15}, S_2 = 0.58 \times 10^{-4}$

$Ni(OH)_2$ is more soluble than AgCN.

Question 3. The value of Kc for the reaction 2A——>B + C is 2×10^{-3}. At a given time, the composition of reaction mixture is [A] = [B] = [C] = 3×10^{-4} M. In which direction the reaction will proceed?

Answer: For the reaction the reaction quotient Q_c is given by Q_c = [B][C]/[A]2 as [A] = [B] = [C] = 3 x 10-4 M $Q_c = (3 \times 10^{-4})(3 \times 10^{-3})/(3 \times 10) = 1$

as $Q_c > K_c$, so, the reaction will proceed in the reverse direction.

Question 4. PCl_5, PCl_3 and Cl_2 are at equilibrium at 500 K and having concentration 1.59M PCl_5 1.59M Cl_2 and 1.41M PCl_5. Calculate K_c for the reaction PCl_5———>PC_{13}+ Cl_2

Answer: The equilibrium constant Kc for the above reaction can be written as:

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

$$= \frac{(1.59)^2}{1.41} = 1.79$$

Question 5. Dihydrogen gas is obtained from natural gas by partial oxidation with steam as per following endothermic reaction:

$CH_4(g) + H_2O(g)$ ——> $CO(g) + 3\ H_2(g)$

(a) Write an 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:

$$(a)\ K_p = \frac{[p_{CO}][p_{H_2}]^3}{[p_{CH_4}][p_{H_2O}]}$$

(b) (i) value of K_p will not change, equilibrium will shift in backward direction.

(ii) value of K_p will increase and reaction will proceed in forward direction.

(iii) no effect.

* *