NCERT Exercise Questions and Answers

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 rate of evaporation and condensation change initially?

   (c) What happens when equilibrium is restored finally and what will be the final vapour pressure?

Ans:

(a) Initially, the vapour pressure will decrease because the same amount of vapour will be distributed in larger volume.

(b) The rate of evaporation remains constant at constant temperature. But the rate of condensation will decrease initially because there are fewer molecules per unit volume in the vapour phase.

(c) When equilibrium is restored, rate of evaporation becomes equal to the rate of condensation. The final vapour pressure will be the same as the original vapour pressure because vapour pressure depends only on temperature.

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

Ans:

Percentage by volume of gaseous iodine atoms = 40%

$\therefore$ Mole fraction of gaseous iodine atoms = $\frac{40}{100}$

$\therefore$ Partial pressure of gaseous iodine atoms = $\frac{40}{100} \times 10^5\ Pa = 0.4 \times 10^5\ Pa$

Partial pressure of gaseous $I_2 = 0.6 \times 10^5\ Pa$

$$\therefore\ K_p = \frac{P_I^2}{P_{I_2}} = \frac{(0.4 \times 10^5)^2}{0.6 \times 10^5} = \mathbf{2.67 \times 10^4}\ \boldsymbol{Pa}$$

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

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

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

Ans:

(i) $\Delta n\,(g) = (2+1) - 2 = 1$,

$$\therefore\ K_p = K_c RT \qquad \text{or} \qquad K_c = \frac{K_p}{RT} = \frac{1.8 \times 10^{-2}}{0.0821 \times 500} = \mathbf{4.38 \times 10^{-4}}$$

(ii) $\Delta n\,(g) = 1 - 0 = 1$

$$\therefore\ K_c = \frac{K_p}{RT} = \frac{167}{0.0821 \times 1073} = \mathbf{1.895}$$

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

$$NO\,(g) + O_3\,(g) \rightleftharpoons NO_2\,(g) + O_2\,(g)$$

What is K_c' for the reverse reactions?

Ans: K_c' for the reverse reaction $= \dfrac{1}{K_c} = \dfrac{1}{6.3 \times 10^{4}} = 0.159 \times 10^{-14} = \mathbf{1.59 \times 10^{-15}}$

5. Nitric oxide reacts with bromine to form nitrosyl bromide as follows

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

When 0.087 mol of NO and 0.0437 mole 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.

Ans: 2 moles of NO gives 2 moles of $NOBr$

$\therefore$ Number of moles NO reacted to form 0.0518 mol of $NOBr = 0.0518$

Number of moles of Br_2 reacted $= \dfrac{0.0518}{2} = 0.0259$

$\therefore$ Number of moles of NO remaining at equilibrium $= 0.087 - 0.0518 = \mathbf{0.0352\ mol}$

$\therefore$ Number of moles Br_2 at equilibrium $= 0.0437 - 0.0259 = \mathbf{0.0178\ mol}$

6. A sample of $HI\,(g)$ is placed in a 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 equilibrium

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

Ans:

Initial pressure of $HI\,(g)$ = 0.2 atm

Partial pressure of $HI\,(g)$ at equilibrium = 0.04 atm

$\therefore$ Decrease in pressure of $HI\,(g)$ = 0.2 – 0.04 = 0.16 atm

It is because of decomposition to H_2 and I_2.

$\therefore$ Partial pressure of H_2 = Partial pressure of I_2 (g) $= \frac{0.16}{2} = 0.08$ atm

$$\therefore K_p = \frac{P_{H_2} \times P_{I_2}}{P^2_{HI}} = \frac{0.08 \times 0.08}{0.04 \times 0.04} = \mathbf{4.0}$$

7. 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 20L vessel at 500K. K_c for the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ is 1.7×10^2 at 500K. Is the reaction mixture at equilibrium. If not, what is the direction of the net reaction.

Ans:

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

$$\therefore Q_c = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} = \frac{(8.13/20)^2}{\frac{1.57}{20} \times \left(\frac{1.92}{20}\right)^3} = 2.38 \times 10^3$$

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

Since $Q_c > K_c$, the reaction will take place in the backward direction

8. One mole of H_2O and one mole of CO are taken in a 10L vessel and heated to 725K. At Solution

Moles of H_2O reacted = $1 \times \frac{40}{100} = 0.4$

$\therefore$ Number of moles of H_2O at equilibrium = $1 - 0.4 = 0.6$

$$\therefore [H_2O] = \frac{0.6}{10} = 0.06 \text{ mol L}^{-1}$$

Similarly, $[CO] = 0.06$ mol L^{-1}

$[H_2] = \frac{0.4}{10} = 0.04$ mol L^{-1}; $[CO_2] = 0.04$ mol L^{-1}

$$\therefore K_c = \frac{[H_2]\,[CO_2]}{[H_2O]\,[CO]} = \frac{0.04 \times 0.04}{0.06 \times 0.06} = \mathbf{0.44}$$

9. What is the equilibrium concentration of each of the substance in the equilibrium when the initial concentration of ICl was 0.78 M.

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

Ans: Let the equilibrium concentration of I_2 and Cl_2 be x. Then equilibrium concentration of ICl will be $0.78 - 2x$

$$K_c = \frac{[I_2]\,[Cl_2]}{[ICl]^2} \qquad \text{i.e., } 0.14 = \frac{x \times x}{(0.78 - 2x)^2}$$

$$\therefore \frac{x}{0.78 - 2x} = \sqrt{0.14} = 0.37$$

$$\therefore x = (0.78 - 2x)\,0.37 = 0.288 - 0.74\,x$$

$$x + 0.74\,x = 0.288 \qquad \text{i.e., } 1.74\,x = 0.288$$

$$\therefore x = \frac{0.288}{1.74} = 0.165$$

i.e., $[I_2] = [Cl_2] = \mathbf{0.165\ M}$

$[ICl] = 0.78 - 2 \times 0.165 = \mathbf{0.45\ M}$

10. Do 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) $PCl_5\,(g) \rightleftharpoons PCl_3\,(g) + Cl_2\,(g)$

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

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

Ans:

(a) Applying Le Chatelier's principle, as pressure decreases the system will favour the forward direction where there is increase in the number of moles of gaseous products. i.e., Number of moles of reaction products increase.

(b) Decrease of pressure will favour backward direction because of the formation of more gaseous substance. i.e., number of moles of reaction products decrease.

(c) In this reaction, there is no change in the number of moles of gaseous reactants and products. So pressure will not have any effect on the equilibrium. i.e., number of moles of the products will remain the same.

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

Ans: $\underset{\text{(acid)}}{HA} \rightleftharpoons \underset{\text{(anion)}}{A^-} + H^+$

$pH = -\log\,[H^+]$

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

$\therefore [H^+] = \text{Antilog of } \bar{5}.85 = 7.08 \times 10^{-5}$ M

Concentration of A^- = Concentration of H^+ = **7.08×10^{-5} M**

$$K_a = \frac{[A^-][H^+]}{[HA]} = \frac{7.08 \times 10^{-5} \times 7.08 \times 10^{-5}}{0.01} = \mathbf{5.01 \times 10^{-7}}$$

$pKa = -\log K_a = -\log(5.01 \times 10^{-7}) = 7 - 0.699 = 6.301 = \mathbf{6.3}$

12. The degree of ionisation of 0.1 M bromoacetic acid solution is 0.132. Calculate the *pH* of the solution and the *pKa* of bromoacetic acid.

Ans:

$K_a = c\alpha^2 = 0.1 \times 0.132 \times 0.132 = 1.74 \times 10^{-3}$

$\therefore pKa = -\log K_a = -\log(1.74 \times 10^{-3}) = 3 - 0.2405 \approx \mathbf{2.76}$

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

$\therefore pH = -\log(1.32 \times 10^{-2}) = 2 - 0.1205 \approx \mathbf{1.88}$

13. 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*?

Ans: Mass of *KOH* is 1000 mL solution = $0.561 \times 5g$

$$\therefore \text{Concentration of } KOH \text{ solution} = \frac{0.561 \times 5}{56} = 0.05 \text{ M}$$

Since *KOH* completely ionises to give K^{-1} and OH^-, $[K^+] = [OH^-] = [KOH] = \mathbf{0.05\ M}$

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.05} = 2.0 \times 10^{-13} \text{ M}$$

$\therefore pH = -\log(2 \times 10^{-13}) = 13 - 0.3010 = 12.699 \approx \mathbf{12.7}$

14. The solubility of $Sr(OH)_2$ at 298 K is 19.23 g/L of solution. Calculate the concentration of strontium and hydroxyl ions. What is its *pH*? (Atomic mass of *Sr* = 87.6)

Ans: Molar mass of $Sr(OH)_2 = 87.6 + 34 = 121.6 \text{ g mol}^{-1}$

$$\text{Solubility of } Sr(OH)_2 \text{ in mol/L} = \frac{18.23}{121.6} = 0.158 \text{ mol L}^{-1}$$

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

$\therefore [Sr^{2+}] = \mathbf{0.158\ mol\ L^{-1}}$ **(or M)**

$[OH^-] = 2 \times 0.158 = \mathbf{0.316\ mol\ L^{-1}\ (or\ M)}$

$pOH = -\log 0.316 = 0.5$

$\therefore pH = 14 - 0.5 = \mathbf{13.5}$

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

Ans: $K_w = [H^+][OH^-] = 2.7 \times 10^{-14}$

$\therefore [H^+] = \sqrt{2.7 \times 10^{-14}} = 1.643 \times 10^{-7}$

$\therefore pH = -\log(1.643 \times 10^{-7}) = 7 - 0.2156 = 6.7844 \approx 6.78$

16. 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 solution.

Ans: K_{sp} of $Ag_2CrO_4 = 1.1 \times 10^{-12}$

For Ag_2CrO_4, $K_{sp} = 4s^3$

$$\therefore \text{Solubility } (s) \text{ of } Ag_2CrO_4 = \sqrt[3]{\frac{1.1 \times 10^{-12}}{4}} = 0.65 \times 10^{-4} = 6.5 \times 10^{-5}\ \text{M}$$

$$\text{Solubility of } AgBr = \sqrt{K_{sp}} = \sqrt{5 \times 10^{-13}} = \sqrt{50 \times 10^{-14}} = 7.07 \times 10^{-7}\ \text{M}$$

$$\therefore \text{Ratio of concentration of } Ag_2CrO_4 \text{ to } AgBr \text{ solution} = \frac{6.5 \times 10^{-5}}{7.07 \times 10^{-7}} = \mathbf{91.9}$$