

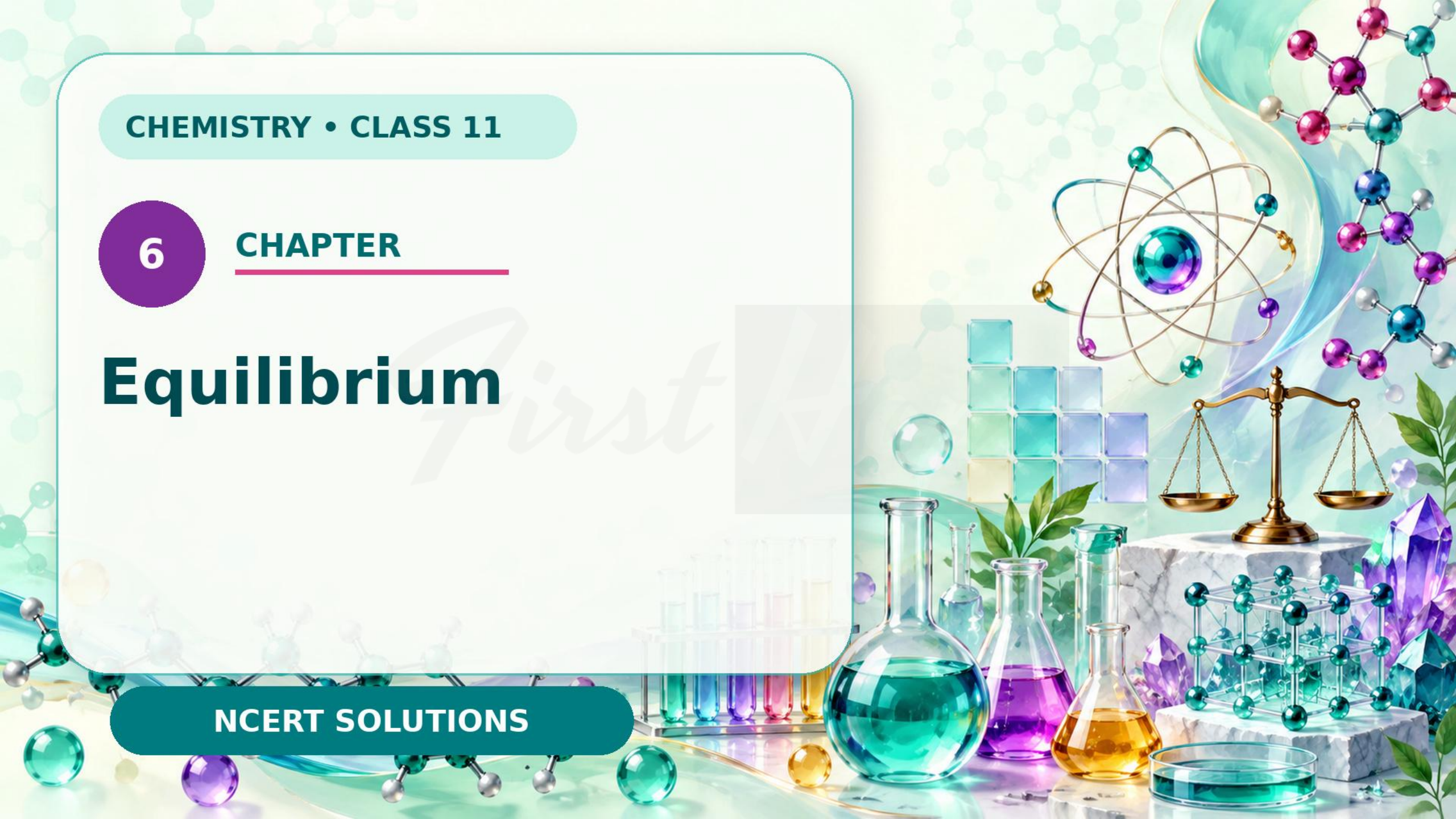
CHEMISTRY • CLASS 11

6

CHAPTER

Equilibrium

NCERT SOLUTIONS



EXERCISE SOLUTIONS: 6.1

6.1 A liquid is in equilibrium with its vapour in a sealed container at fixed temperature. The container volume is suddenly increased. (a) What happens initially to vapour pressure? (b) How do evaporation and condensation rates change initially? (c) What happens when equilibrium is restored, and what is the final vapour pressure?

Answer:

(a) Immediately after expansion →

The same amount of vapour first occupies a larger volume, so its pressure decreases.

$$P_{\text{initial after expansion}} < P_{\text{saturated}}$$

(b) Initial change in rates →

Condensation slows because fewer vapour molecules strike the liquid surface. Evaporation is initially unchanged if temperature and exposed liquid area stay the same. Net evaporation therefore occurs.

$$\text{Rate}(\text{evaporation}) > \text{Rate}(\text{condensation})$$

(c) After equilibrium is restored

More liquid evaporates until the two rates become equal again. At the same temperature, the original saturated vapour pressure returns, provided some liquid remains.

$$P_{\text{final}} = \text{original } P_{\text{saturated}}$$

If all the liquid evaporates before saturation is reached, the final vapour pressure is lower.

EXERCISE SOLUTIONS: 6.2 – 6.3

6.2 Find K_c for $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$, with equilibrium concentrations $[\text{SO}_2] = 0.60 \text{ M}$, $[\text{O}_2] = 0.82 \text{ M}$ and $[\text{SO}_3] = 1.90 \text{ M}$.

Answer:

Use balanced-equation coefficients as powers. The final rounded value has two significant figures.

$$\begin{aligned} K_c &= \frac{[\text{SO}_3]^2}{[[\text{SO}_2]^2 [\text{O}_2]} \\ &= \frac{(1.90)^2}{[(0.60)^2 (0.82)]} \\ &= 12.23 \text{ L mol}^{-1} \approx 12 \text{ L mol}^{-1} \end{aligned}$$

6.3 At total pressure 10^5 Pa , iodine vapour contains 40% I atoms by volume. Find K_p for $\text{I}_2(\text{g}) \rightleftharpoons 2\text{I}(\text{g})$.

Answer:

For an ideal gas mixture, volume fraction equals mole fraction.

$$p(\text{I}) = 0.40(10^5) = 4.0 \times 10^4 \text{ Pa}$$

$$p(\text{I}_2) = 0.60(10^5) = 6.0 \times 10^4 \text{ Pa}$$

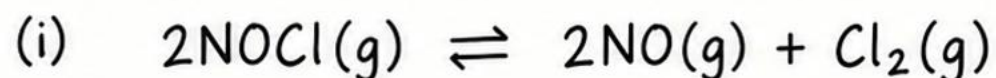
$$\begin{aligned} K_p &= \frac{p(\text{I})^2}{p(\text{I}_2)} \\ &= \frac{(4.0 \times 10^4)^2}{(6.0 \times 10^4)} \\ &= 2.67 \times 10^4 \text{ Pa} \end{aligned}$$

Units follow the textbook concentration and pressure forms of the equilibrium constants.

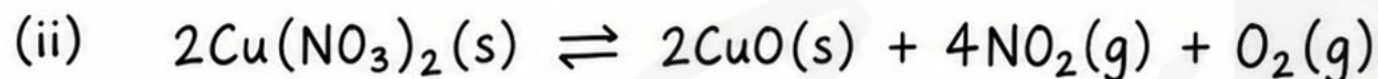
EXERCISE SOLUTIONS: 6.4

6.4 Write the concentration equilibrium-constant expression for each reaction.

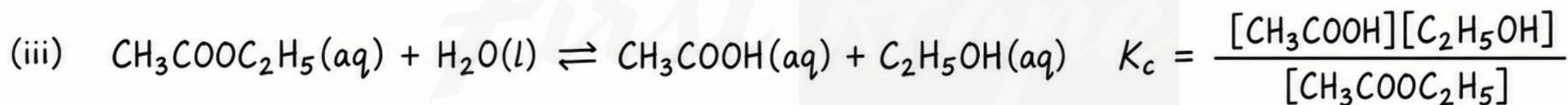
Answer:



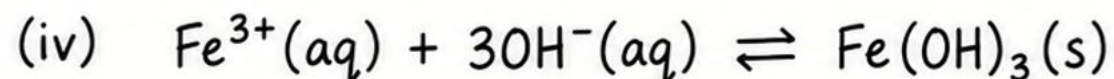
$$K_c = \frac{[\text{NO}]^2 [\text{Cl}_2]}{[\text{NOCl}]^2}$$



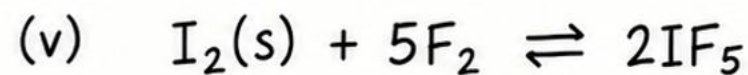
$$K_c = [\text{NO}_2]^4 [\text{O}_2]$$



$$K_c = \frac{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]}{[\text{CH}_3\text{COOC}_2\text{H}_5]}$$



$$K_c = \frac{1}{[\text{Fe}^{3+}][\text{OH}^-]^3}$$

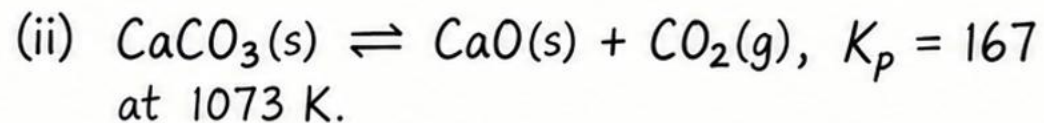
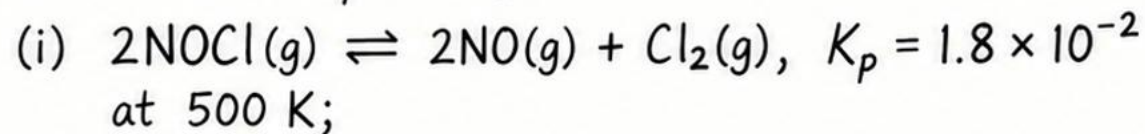


$$K_c = \frac{[\text{IF}_5]^2}{[\text{F}_2]^5}$$

Pure solids and pure liquid water are omitted. In the last reaction, the expression assumes fluorine and iodine pentafluoride are gaseous; their phases were not specified.

EXERCISE SOLUTIONS: 6.5–6.6

6.5 Convert K_p to K_c :



Answer: Use the textbook pressure convention in atmospheres. Both reactions gain one mole of gas.

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

$$\Delta n_g = 1$$

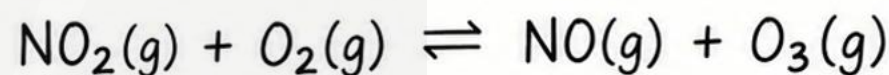
$$R = 0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1}$$

$$\begin{aligned} \text{(i)} \quad K_c &= \frac{0.018}{0.0821 \times 500} \\ &= 4.38 \times 10^{-4} \text{ mol L}^{-1} \end{aligned}$$

$$\begin{aligned} \text{(ii)} \quad K_c &= \frac{167}{0.0821 \times 1073} \\ &= 1.90 \text{ mol L}^{-1} \end{aligned}$$

6.6 For $\text{NO}(g) + \text{O}_3(g) \rightleftharpoons \text{NO}_2(g) + \text{O}_2(g)$, $K_c = 6.3 \times 10^{14}$ at 1000 K. Both directions are elementary bimolecular reactions. Find K_c for the reverse reaction.

Answer:



$$K_c(\text{reverse}) = \frac{1}{K_c(\text{forward})}$$

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

Reversing a reaction inverts its equilibrium constant. The stated reaction order is not needed for this calculation.

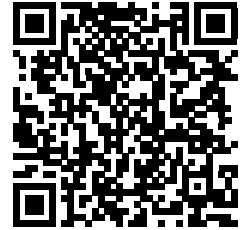


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