

Chapter 10

Solubility

Equilibria

CHAPTER ANALYSIS

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- One of the major chapters
- Relatively straight foreword chapter
- 3 **key** concepts



EXAM

- Always tested
- Consist of both definition & calculations



WEIGHTAGE

- Usually incorporated with other chapters on equilibrium

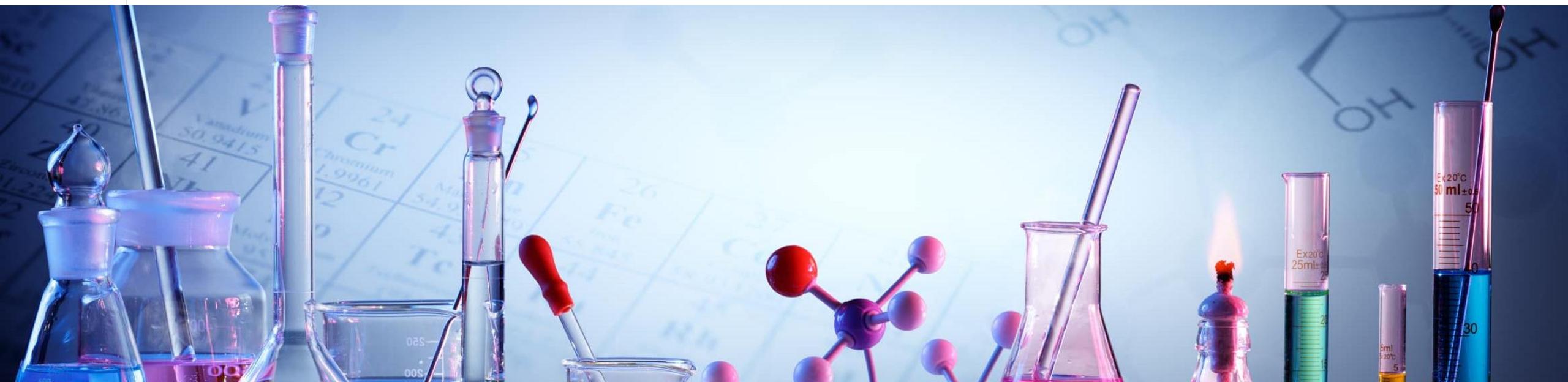
KEY CONCEPT

Solubility & Solubility Product

Ionic Product (IP)

Factors affecting Solubility

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Solubility & Solubility Product

SOLUBILITY

- The solubility of an ionic salt is the **amount of solute** that can **dissolve in a certain quantity** of solvent or solution at a **specified temperature**.
- Units = mol dm⁻³

SATURATED SOLUTION

The solution that contains the **maximum amount of salt** that it can dissolve.

Solubility & Solubility Product

SOLUBILITY PRODUCT

- It is an equilibrium constant.
- It is the **product of the molar concentration of the constituent ions in a saturated solution, raised to the appropriate powers** at a given temperature.



$$K_c = \frac{[A^{y+}]^x (\text{aq})_{\text{eqm}} [B^{x-}]^y (\text{aq})_{\text{eqm}}}{[A_x B_y] (s)_{\text{eqm}}} = [A^{y+}]^x (\text{aq})_{\text{eqm}} [B^{x-}]^y (\text{aq})_{\text{eqm}} = K_{sp}$$

- Units = $(\text{mol dm}^{-3})^{x+y}$

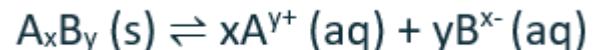
Note: The **solubility of a salt can change** while **solubility product is constant** for a given solute at any specified temperature.

Note: K_{sp} allows a **direct comparison of the solubility** of different salts **only if the mol ratio between the product ions are the same**.

Ionic Product (IP)



It is the **product of the molar concentration of the constituent ions** in the solution, raised to the appropriate powers.



$$IP = [A^{y+}]^x[B^{x-}]^y = K_{sp}$$

Units = (mol dm⁻³)^{x+y}

$IP < K_{sp}$: The solution is not saturated, no precipitation.

$IP = K_{sp}$: The solution is saturated, first trace of precipitation occurs.

$IP > K_{sp}$: The solution is saturated, precipitation occurs.

Note: IP has the same expression as K_{sp} , however for K_{sp} the concentration of the ions are equilibrium concentrations in a saturated solution whereas for IP the concentration of the ions are concentrations at any one time.

Note: When $IP = K_{sp}$, the solution contains the maximum concentration of the product ions.

Factors Affecting Solubility

Factors Affecting Solubility

Temperature

Common Ion

Complex Ion

Factors Affecting Solubility

FACTOR 1: Temperature

Temperature affects K_{sp} and hence the solubility.

FACTOR 2: Common Ion

- Solubility will be **reduced** by the presence of another solute that contains a **common ion**, known as the **common ion effect**.



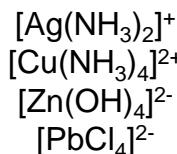
$$K_{sp} = [\text{Ag}^+]_{\text{eqm}}[\text{Cl}^-]_{\text{eqm}}$$

- When a solute of common ion is added (eg: NaCl), $[\text{Cl}^-]$ increases. Based on LCP, this will cause the equilibrium to shift to the left to counteract the increase in $[\text{Cl}^-]$, causing the solid AgCl to be precipitated out. This thus results in a reduction in the solubility of AgCl.

Note: Commonly tested question on common ion effect will be selective precipitation of ions or to calculate the percentage of ions remaining in the solution.

Factors Affecting Solubility

Examples of complex ions formation with cations



FACTOR 3: Complex Ion

A **complex ion** is a charged species which consist of a **central metal ion** linked to surrounding ions or molecules (called ligands) by **dative covalent bonds**.

Solubility will be **increase** by the presence of another solute that result in the formation of a **complex ion**, known as the **complex ion effect**.



$$K_{\text{sp}} = [\text{Ag}^+]_{\text{eqm}}[\text{Cl}^-]_{\text{eqm}}$$

When NH_3 is added into the solution of silver chloride, the complex ion $[\text{Ag}(\text{NH}_3)_2]^+$ is formed.



This reduced the $[\text{Ag}^+]$. Based on LCP, this will cause the equilibrium to shift to the right to counteract the decrease in $[\text{Ag}^+]$, causing more of the solid AgCl to be dissolved. This thus result in an increase in the solubility of AgCl .

Note: The precipitation of the salt will ultimately dependent on the IP vs it's K_{sp} . Sometimes even with the formation of complex ion, IP will still be greater than K_{sp} , causing precipitation to occur.

Practice Questions

Question: The solubility of lead sulfate at 25°C is 1.26×10^{-4} mol dm⁻³. Calculate the K_{sp} of lead sulfate in water at 25°C.

Answer: 1.59×10^{-8} mol² dm⁻⁶

Question: The solubility of calcium fluoride at 25°C is 1.6×10^{-2} g dm⁻³. Calculate the K_{sp} calcium fluoride in water at 25°C.

Answer: 3.44×10^{-11} mol³ dm⁻⁹

Question: Given that the K_{sp} of iron (II) hydroxide is 7.9×10^{-16} mol³ dm⁻⁹ at 25°C, calculate the solubility of iron (II) hydroxide and concentration of Fe²⁺ and OH⁻ at equilibrium.

Answer: 5.82×10^{-6} mol dm⁻³; [Fe²⁺] = 5.82×10^{-6} mol dm⁻³; [OH] = 1.16×10^{-5} mol dm⁻³

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Practice Questions

Question: Determine if a precipitate of lead (II) iodide will be formed when 500 cm³ of 0.1 mol dm⁻³ lead nitrate were added to 500 cm³ of 0.1 mol dm⁻³ of sodium iodide, given that the K_{sp} of lead (II) iodide is 7.0×10^{-9} mol³ dm⁻⁹ at 25°C.

Answer: Since $IP > K_{sp}$, precipitate will be formed.

Question: Determine if a precipitate will be formed when 250 cm³ of 0.02 mol dm⁻³ lead nitrate were added to 150 cm³ of 0.01 mol dm⁻³ of sodium chloride, given that the K_{sp} of lead (II) chloride is 1.6×10^{-5} mol³ dm⁻⁹ at 25°C.

Answer: Since $IP < K_{sp}$, no precipitate will be formed.

Question: Given that the K_{sp} of iron (II) hydroxide in water at 25°C is 7.9×10^{-16} mol³ dm⁻⁹, calculate its solubility in (a) pure water (b) solution containing 0.10 mol dm⁻³ FeSO₄ (c) solution containing 0.10 mol dm⁻³ NaOH.

Answer:

- (a) 5.82×10^{-6} mol dm⁻³
- (b) 4.44×10^{-8} mol dm⁻³
- (c) 7.90×10^{-14} mol dm⁻³

Practice Questions

Question: A solution A is saturated with both calcium hydroxide and calcium sulfate.

The pH of the solution A is 12.3. Calculate the total hydroxide concentration.

- (i) Write an expression for the K_{sp} of calcium hydroxide.
- (ii) Given that K_{sp} of calcium hydroxide is 4.0×10^{-5} , calculate the concentration of calcium ions in solution A.
- (i) Write an expression for the K_{sp} of calcium sulfate
- (ii) Given that K_{sp} of calcium sulfate is 2.5×10^{-5} , calculate the concentration of sulfate ions in solution A.
- (iii) The concentration of sulfate ions in solution A is less than 5.0×10^{-3}

Answer:

- (a) 0.02 mol dm^{-3}
- (b) (i) $[\text{Ca}^{2+}][\text{OH}^-]^2$
- (b) (ii) 0.1 mol dm^{-3}
- (c) (i) $[\text{Ca}^{2+}][\text{SO}_4^{2-}]$
- (c) (ii) $2.5 \times 10^{-4} \text{ mol dm}^{-3}$
- (c) (iii) Presence of common ions Ca^{2+} .

Practice Questions

Question: Given that the K_{sp} of AgCl in water at 25°C is $2.0 \times 10^{-10} \text{ mol}^2 \text{ dm}^{-6}$, calculate its solubility in a solution containing $0.10 \text{ mol dm}^{-3} \text{ Cl}^-$.

Answer: $2.0 \times 10^{-9} \text{ mol dm}^{-3}$

Question: Given that K_{sp} of AgCl is $1.6 \times 10^{-10} \text{ mol}^2 \text{ dm}^{-6}$ and K_{sp} of AgI is $8.0 \times 10^{-17} \text{ mol}^2 \text{ dm}^{-6}$, calculate the K_c for the below equation:

Hint: Linking K_{sp} with K_c .

Answer: $\text{AgCl} (\text{s}) + \text{I}^- (\text{aq}) \rightleftharpoons \text{AgI} (\text{s}) + \text{Cl}^- (\text{aq})$

$$\frac{[\text{Ag}^+][\text{Cl}^-]}{[\text{Ag}^+][\text{I}^-]} = \frac{1.6 \times 10^{-10}}{8.0 \times 10^{-17}} = \frac{[\text{Cl}^-]}{[\text{I}^-]} = K_c = 2 \times 10^6$$

Test yourself!

- (a) show understanding of, and apply, the concept of solubility product, K_{sp}
- (b) calculate K_{sp} from concentrations and *vice versa*
- (c) discuss the effects on the solubility of ionic salts by the following:
 - (i) common ion effect
 - (ii) formation of complex ion, as exemplified by the reactions of halide ions with aqueous silver ions followed by aqueous ammonia

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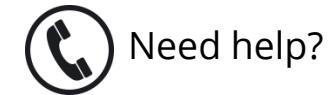
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