

Solubility Equilibria

"Insoluble" compounds

The Solubility Product Constant

- For a general slightly soluble compound $X_n Y_m$ we have the following equilibrium
 - $X_n Y_m \rightleftharpoons n X^{m+} + m Y^{n-}$
- And the following equilibrium expression $K_c = [X^{m+}]^n [Y^{n-}]^m$
- We call K_c the solubility product constant, K_{sp}

Calculating K_{sp}

- If we know the solubility we can determine the K_{sp} .
- Example:
 - Silver bromide has a solubility of 1.3×10^{-4} g/L. What is the K_{sp} of silver bromide?

Answer

- Find molar solubility first.

$$\frac{1.3 \times 10^{-4} \text{ g AgBr}}{1 \text{ L}} \times \frac{1 \text{ mol AgBr}}{187.77 \text{ g AgBr}} = 6.9 \times 10^{-7} \text{ mol L}^{-1}$$

- For Silver bromide, when we dissolve one mole of the solid we get one mole of Silver ions and one mole of Bromide ions.

Answer

- This means that the molar concentration of each of the ions can be determined from the stoichiometry of the compound.

$$[Ag^+] = [Br^-] = 6.9 \times 10^{-7} \text{ M}$$

- Plugging into the equilibrium expression gives the K_{sp} .

Answer

$$K_{sp} = [Ag^+][Br^-] = (6.9 \times 10^{-7})(6.9 \times 10^{-7}) = 4.8 \times 10^{-13}$$

Finding molar solubility

- Just another type of equilibrium problem.
- Start with equilibrium and K_{sp} .
- Set up ICE table.
- When we solve for x , x will be the molar solubility.

Example

- Find the molar solubility of Lead(II) iodide at 25 C.
- Equilibrium
 - $\text{PbI}_2 \rightleftharpoons \text{Pb}^{2+} + 2 \text{I}^-$ $K_{sp} = 7.9 \times 10^{-9}$
- ICE table

	Pb^{2+}	I^-
I	0	0
C	+x	+2x
E	x	2x

Answer

- Set up calculation

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2 = (x)(2x)^2 = 4x^3 = 7.9 \times 10^{-9}$$

$$x = \sqrt[3]{\frac{7.9 \times 10^{-9}}{4}} = 1.3 \times 10^{-3} \text{ M}$$

- The molar solubility is $1.3 \times 10^{-3} \text{ M}$.

Solubility and the common ion effect

- The presence of one of the ions in a slightly soluble compound in a solution reduces the solubility of the compound according to Le Chatelier's principle.
- If we have a solution of sodium oxalate (very soluble) and add solid calcium oxalate (slightly soluble), the solubility of the calcium oxalate is reduced.

Example

- Calculate the solubility of calcium oxalate in 0.10 M sodium oxalate.
- Equilibrium
 - $\text{CaC}_2\text{O}_4 \rightleftharpoons \text{Ca}^{2+} + \text{C}_2\text{O}_4^{2-}$ $K_{sp} = 2.3 \times 10^{-9}$
- ICE table

	Ca^{2+}	$\text{C}_2\text{O}_4^{2-}$
I	0	0.10
C	+x	+x
E	x	0.10+x

Answer

- Set up calculation

$$K_{sp} = [\text{Ca}^{2+}][\text{C}_2\text{O}_4^{2-}] = (x)(0.10 + x)^2 = 2.3 \times 10^{-9}$$

- This is a slightly soluble salt so x should be a small number ($> 10^{-4}$). We can probably ignore x in $(0.10+x)$. This leaves.

$$K_{sp} = (x)(0.10 + x)^2 \approx (x)(0.10) = 2.3 \times 10^{-9}$$

[Answer]

- Solve for x.

$$x = \frac{2.3 \times 10^{-9}}{0.10} = 2.3 \times 10^{-8}$$

- The molar solubility is now 2.3×10^{-8} .