

More Stuff about Aqueous Equilibria: CH 17 Part 5

3/22/19 ①

Electrolytes are ionic compounds

- A strong electrolyte dissociates completely into ions when dissolved in water
- A weak electrolyte only dissociates partially when dissolved in water.

Solubility of Ionic compounds

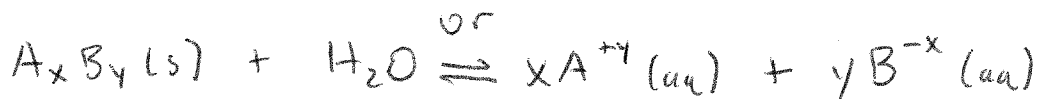
- Not all ionic compounds dissolve significantly in water
- A list of solubility rules is used to decide what combinations of ions will dissolve.

Solubility Equilibria

- For "insoluble salts" how much do they dissolve?
- There is an equilibrium constant called solubility-product constant, K_{sp} .
- Assume that MA is an ionic solid dissolving in water:



$$K_{sp} = [M^+][A^-]$$



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

Solubility Product Examples



$$K_{sp} = [Ba^{2+}][SO_4^{2-}]$$



$$K_{sp} = [Ba^{2+}]^3 [PO_4^{3-}]^2$$

Solubility vs Solubility Product

(2)

- K_{sp} is not the same as solubility, but it can be used to calculate solubility.
- Solubility is the quantity of a substance that dissolves to form a saturated solution.

Calculating Solubility from K_{sp}

~~1~~ K_{sp} for CaF_2 is 3.9×10^{-11} at 25°C . What is molar solubility?

① Write rxn: $\text{CaF}_2(\text{s}) \rightleftharpoons \text{Ca}^{2+}(\text{aq}) + 2\text{F}^{-}(\text{aq})$

② Write K_{sp} : $K_{sp} = [\text{Ca}^{2+}][\text{F}^{-}]^2 = 3.9 \times 10^{-11}$

③ ICE table:

	$\text{CaF}_2(\text{s})$	$[\text{Ca}^{2+}](\text{M})$	$[\text{F}^{-}](\text{M})$
I	—	0	0
C	—	+x	+2x
E	—	x	2x

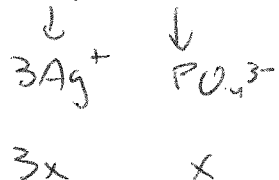
④ Substitute the equilibrium concentration values from the table into the solubility-product equation:

$$K_{sp} = [\text{Ca}^{2+}][\text{F}^{-}]^2 = 3.9 \times 10^{-11}$$

$$= (x)(2x)^2 = 4x^3$$

$$x = 2.1 \times 10^{-4} \text{ M}$$

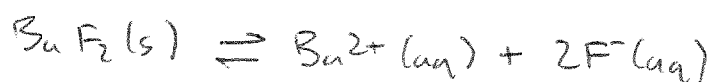
Solubility of Ag_3PO_4 ?



$$K_{sp} = 1.8 \times 10^{-8} = [\text{Ag}^{+}]^3 [\text{PO}_4^{3-}]$$
$$(3x)^3 (x)$$

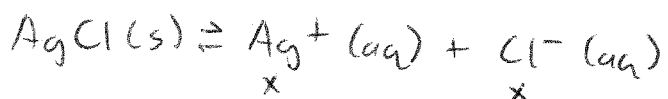
$$1.8 \times 10^{-8} = 27x^4$$

BaF_2 has a solubility of $1.82 \times 10^{-2} \text{ M}$

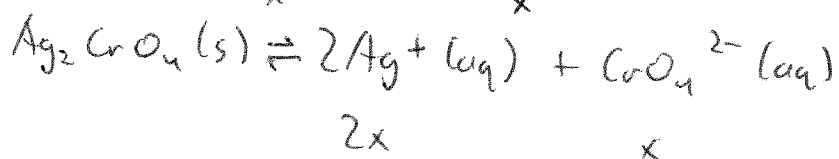


$$\begin{aligned} K_{sp} &= [\text{Ba}^{2+}][\text{F}^-]^2 \\ &= (1.82 \times 10^{-2})(2 \times 1.82 \times 10^{-2})^2 \\ &= 2.4 \times 10^{-5} \end{aligned}$$

Question: Which is more soluble, AgCl or Ag_2CrO_4 ?



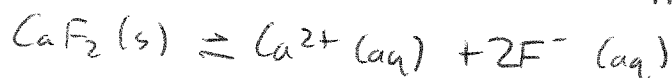
$$\begin{aligned} K_{sp} &= 1.6 \times 10^{-10} \\ x^2 &= 1.6 \times 10^{-10} \quad x = 1.3 \times 10^{-5} \\ K_{sp} &= 9.0 \times 10^{-10} \end{aligned}$$



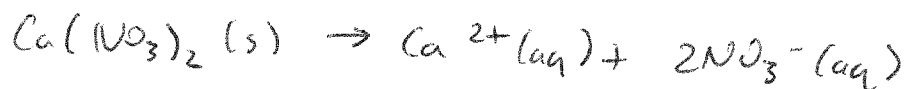
$$\begin{aligned} (2x)^2 x &= 4x^3 = 9.0 \times 10^{-10} \\ x &= 1.3 \times 10^{-4} \text{ M} \end{aligned}$$

Ag_2CrO_4 because it produces more ions

Assume we have a solution of calcium fluoride happy at equilibrium



What happens if we add some solid calcium nitrate?



Factors Affecting Solubility

- The Common-Ion Effect

If one of the ions in a soln. equilibrium is added to the soln., the solubility of the salt will decrease

For example, calcium fluoride is less soluble if either Ca^{2+} ions or F^- are added

What is the molar solubility of CaF_2 in $0.010 \text{ M } \text{Ca}(\text{NO}_3)_2$? (4)



$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = 3.9 \times 10^{-11}$$

	$\text{CaF}_2(s)$	$[\text{Ca}^{2+}](\text{M})$	$[\text{F}^-](\text{M})$
I	—	0.010	0
C	—	+x	+2x
E	—	0.010 + x	2x

$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = 3.9 \times 10^{-11}$$
$$= (0.010 + x)(2x)^2$$

Assume that $x \ll 0.010$, so that $0.010 + x = 0.010$

$$3.9 \times 10^{-11} = (0.010)(2x)^2$$

$$x = 3.1 \times 10^{-5} \text{ M}$$

Solubility of AgCl in $0.55 \text{ M } \text{NaCl}$?

$$[\text{Cl}^-] = 0.55 \text{ M} + x \quad \leftarrow \text{Cl coming from AgCl}$$
$$\approx 0.55 \text{ M}$$

$$[\text{Ag}^+] = x$$

$$K_{sp} = 1.6 \times 10^{-10} = (0.55 + x)x$$

$$x = 2.9 \times 10^{-10} \text{ M}$$