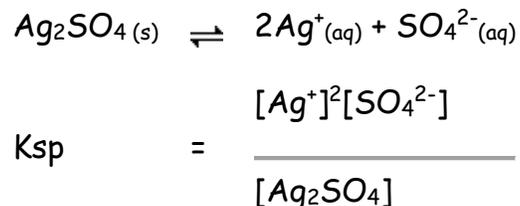


4.0 The Solubility Product Constant

Since saturated solutions are equilibrium systems, we can apply the equilibrium expression as follows:



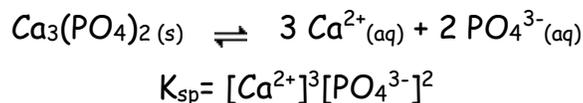
where our equilibrium constant is now called K_{sp} , where "sp" stands for "solubility product".

Don't forget that in equilibrium expressions solids and liquids are NOT included.

Therefore: $K_{sp} = [\text{Ag}^+]^2[\text{SO}_4^{2-}]$

Example: Write the expression for the solubility product constant, K_{sp} , for $\text{Ca}_3(\text{PO}_4)_2$.

Solution: Write the dissolving reaction (remember that polyatomic ions remain together as a unit and do not break apart into separate elements). Write the K_{sp} expression.



4.1 The Meaning of K_{sp}

The K_{sp} is a measure of the solubility of an ionic salt.

The larger the value of the K_{sp} , the greater is the solubility of the salt.

- FeS has a K_{sp} of 4×10^{-19} which means that in a saturated solution there would be few Fe^{2+} or S^{2-} ions.
- PbCl_2 has a K_{sp} of 1.8×10^{-4} which means a saturated solution of PbCl_2 would have many Pb^{2+} and Cl^- ions.

Don't forget that you can only calculate a K_{sp} if the solution is saturated. Only saturated salt solutions are in equilibrium.

4.2 Calculations Involving K_{sp}

(a) ... when concentration of a saturated solution is known

Example: The concentration of lead ions in a saturated solution of PbI_2 at $25^\circ C$ is $1.30 \times 10^{-3} M$. What is its K_{sp} ?



$$K_{sp} = [Pb^{2+}][I^-]^2$$

Now use the equation ratio to find the concentration of the ions.

$$[PbI_2] = 1.30 \times 10^{-3} M$$

$$[Pb^{2+}] = 1.30 \times 10^{-3} M$$

$$[I^-] = 2 \times 1.30 \times 10^{-3} = 2.60 \times 10^{-3} M$$

Substitute values into the K_{sp} expression and solve for the unknown:

$$K_{sp} = [Pb^{2+}][I^-]^2$$

$$= (1.30 \times 10^{-3})(2.60 \times 10^{-3})^2$$

$$= (1.30 \times 10^{-3})(6.76 \times 10^{-6})$$

$$K_{sp} = 8.79 \times 10^{-9} M$$

(b) ... when the amount of solute required to saturate a solution is given

Example: Calculate the K_{sp} for $CaCl_2$ if $2.00 \times 10^2 g$ of $CaCl_2$ is required to saturate 100.0 mL of solution.



$$K_{sp} = [Ca^{2+}][Cl^-]^2$$

First calculate molarity.

$$\text{Molarity} = \frac{200 \text{ g} \times \frac{1 \text{ mole}}{111.1 \text{ g}}}{0.100 \text{ L}} = 18.001 \text{ M}$$

Now use the equation ratio to find the concentration of the ions.

$$[\text{CaCl}_2] = 18.0 \text{ M}$$

$$[\text{Ca}^{2+}] = 18.0 \text{ M}$$

$$[\text{Cl}^-] = 36.0 \text{ M}$$

Substitute values into the K_{sp} expression and solve for the unknown:

$$K_{sp} = [18.0][36.0]^2$$

$$K_{sp} = 2.33 \times 10^4 \text{ M}$$

(c) ... when K_{sp} is known.

Example: K_{sp} for MgCO_3 at 25°C is 2.0×10^{-8} . What are the ion concentrations in a saturated solution at this temperature?



$$K_{sp} = [\text{Mg}^{2+}][\text{CO}_3^{2-}]$$

We will let our unknown ion concentrations equal x . The balanced equation tells us that both Mg^{2+} and CO_3^{2-} will have the same concentration!

Substitute values into the equation and solve for the unknown:

$$K_{sp} = [\text{Mg}^{2+}][\text{CO}_3^{2-}]$$

$$2.0 \times 10^{-8} = (x)(x) = x^2$$

$$x^2 = 2.0 \times 10^{-8}$$

$$x = \sqrt{2.0 \times 10^{-8}}$$

$$= 1.4 \times 10^{-4} \text{ M}$$

$$x = [\text{Mg}^{2+}] = 1.4 \times 10^{-4} \text{ M}$$

$$x = [\text{CO}_3^{2-}] = 1.4 \times 10^{-4} \text{ M}$$