

Solubility Equilibrium & Ksp Notes

Solubility Refresher:

Solute: substance that is dissolved into something else. The minor component of a solution.

Solvent: substance that makes of the major component of a solution. (water generally)

Saturated Solution: solution that has the maximum amount of solute dissolved in it.

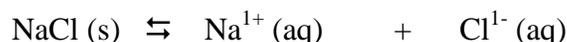
Solubility:

- Refers to the quantity that will dissolve to form a saturated solution.
- Typical units: grams of solute per Liter of solution
- **Molar solubility** is the maximum molarity of a solute.

Molar solubility = MAXIMUM MOLARITY

Factors Affecting Solubility:

1. “**Common Ion Effect**” (LeChatelier’s Principle)



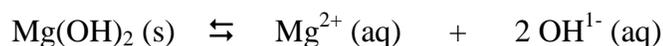
Addition of either $\text{Na}^{1+} \text{ (aq)}$ or $\text{Cl}^{1-} \text{ (aq)}$ shifts equilibrium to the left, reducing solubility. This can be done with aqueous HCl because it dissociates to form H^{1+} and Cl^{1-} ions. Can also be done with aqueous Na_2SO_4 since it exists as separate Na^{1+} and SO_4^{2-} ions.

BIG IDEA: Adding a “common ion” decreases the solubility of a salt.

In which aqueous system is PbI_2 least soluble?

- (a) 0.5 M HI
- (b) H_2O
- (c) 0.2 M HI
- (d) 1.0 M HNO_3
- (e) 0.8 M KI

2. Solubility and pH



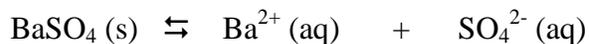
Addition of H^{1+} to this will shift this equilibrium to the right.

The H^{1+} will react with the OH^{1-} to produce to water.

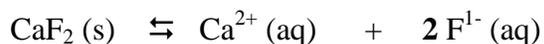
Thus, the equilibrium will shift to the right to produce more OH^{1-} , increasing the solubility.

Mg(OH)_2 is essentially insoluble in cold water 0.0009 grams per 100 mL

Consider the following process:



$$K =$$



$$K =$$

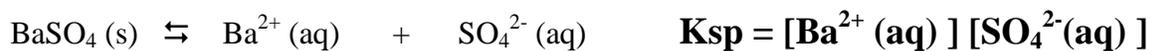
Solubility Product: (K_{sp})

Equals the product of the concentration of the ions involved in the equilibrium raised to the power of their coefficients in the equilibrium expression.

K_{sp} is the equilibrium constant for the equilibrium that exists between a solid ionic solute and its ions in a **saturated** aqueous solution.

The smaller the K_{sp} value, the lower the solubility.

Precipitation of Ions:



Q = initial reaction conditions

$$\mathbf{Q = [\text{Ba}^{2+}(\text{aq})][\text{SO}_4^{2-}(\text{aq})]}$$

If $Q > K_{sp}$ “solvent cannot dissolve that many ions” so precipitate forms until $Q = K_{sp}$

If $Q = K_{sp}$ “equilibrium achieved”....we have a perfectly saturated solution

If $Q < K_{sp}$ “still able to dissolve more ions” so solid dissolves until $Q = K_{sp}$

Types of Ksp Problems

I. Dissolving the solid ionic compound into pure water. Solving for a missing variable in a mathematical expression.

The Ksp for CaF₂ is 3.9×10^{-11} . What is the molar solubility (max molarity) of a CaF₂ solution?

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

A saturated solution of sodium chloride has a molarity of 5.65 M. What is the Ksp for NaCl?

II. Dissolving the solid ionic compound into a “Common-Ion” Solution

Picture this as trying to dissolve a compound into a pre-existing solution containing one of the ions.

What will be the molar solubility of AgCl solution that already contains 0.75 M NaCl? $K_{sp} = 1.8 \times 10^{-10}$.

What is the molar solubility of CaF₂ in a 0.50 M KF solution? The Ksp of CaF₂ is 3.9×10^{-11} .

III. Will a precipitate form when two substances are mixed? Compare Q to K_{sp}.

Throw all of the ions into a beaker of pure water. Are the concentrations of the ions high enough to “find each other” to make a precipitate? Must calculate the molarities of each ion and then plug the ion molarities into the K_{sp} math expression.

A person mixes 100.0 mL of 0.0016 M CaCl₂ and 50.0 mL of 0.0081 M K₂SO₄.

If the K_{sp} for CaSO₄ is 2.4×10^{-5} , will a CaSO₄ precipitate be observed?

Compare this value (Q) to the K_{sp} textbook value. If $Q > K_{sp}$, a ppt will form. If $Q < K_{sp}$ then no precipitate forms. If $Q = K_{sp}$, the solution is fully saturated. “On the brink of a precipitate!”

IV. Selective Precipitation: Which sparingly soluble solid will form a precipitate first?

A beaker contains 250 mL of a mixture that is labeled as 0.10 M Sr(NO₃)₂ and 0.00050 M Pb(NO₃)₂. A person slowly adds solid Na₂CO₃ to the beaker until a precipitate starts to form.

Substance	K _{sp} Value
SrCO ₃	9.3×10^{-10}
PbCO ₃	7.4×10^{-14}

Given these conditions, what substance (SrCO₃ or PbCO₃) starts to precipitate first? What is the concentration of the [CO₃²⁻] when the first precipitate starts to form?