

Solubility Product

1. Do precipitation reactions reach a measurable equilibrium?

Yes they do.

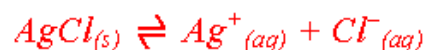
2. What is a solubility product?

The solubility product, K_{sp} , is a measure of this equilibrium.

3. What is the common ion effect?

When a salt (with an ion in common with a precipitate) is added to a solution of the solid precipitate and a decrease in the solubility of the precipitate is observed.

Consider the following equilibrium:



The addition of NaCl to a solution that contains an AgCl precipitate demonstrates the effects of the common ion. Using Le Châtelier's Principle, we can see that the addition of Cl^- (by way of NaCl) would cause equilibrium to shift away from the ions and towards the solid precipitate. Shifting left would mean that AgCl would dissociate less or be less soluble.

4. Calculate the solubility for the following solutions

You handle these problems as you have for every equilibrium calculation thus far – an ICE chart. In this case you are looking at the equilibrium reaction for the solids dissociation into its component ions.

When the question asks for the “solubility” of the salt, it is asking you to solve for the x value in the ICE chart.

a. Ag_2CO_3 . ($K_{\text{sp}} = 8.1 \times 10^{-12}$)

	$\text{Ag}_2\text{CO}_3 (s)$	\rightleftharpoons	$2\text{Ag}^+ (aq)$	+	$\text{CO}_3^{2-} (aq)$
<i>I</i>	—		0		0
<i>C</i>	—		+2x		+x
<i>E</i>	—		2x		x

Remember that because solids do not affect equilibrium, we do not include any values for them in the ICE chart calculation. They are irrelevant most of the time.

Now all we have to do is plug our equilibrium values into the K_{sp} expression for this equilibrium reaction.

$$K_{\text{sp}} = [\text{Ag}^+]^2[\text{CO}_3^{2-}] = (2x)^2(x) = 8.1 \times 10^{-12}$$

$$4x^3 = 8.1 \times 10^{-12}$$

$$x = \boxed{1.27 \times 10^{-4} \text{ M}}$$

b. Ag_2CO_3 in 0.10M AgNO_3 .

The variation in this question, as compared to the previous one, is that we have to take the common ion effect into consideration. Both compounds contain Ag^+ . Remember that this is supposed to limit the solubility of the precipitate in solution. The only difference between is that there will now be an initial concentration of Ag^+ in the ICE chart calculation.

	$\text{Ag}_2\text{CO}_3 (s)$	\rightleftharpoons	$2\text{Ag}^+ (aq)$	+	$\text{CO}_3^{2-} (aq)$
<i>I</i>	—		0.10		0
<i>C</i>	—		+2x		+x
<i>E</i>	—		0.10 + 2x		x

small K, small x approx.

$$K_{sp} = [Ag^+]^2[CO_3^{2-}] = (0.10+2x)^2(x) = (0.10)^2(x) = 8.1 \times 10^{-12}$$

$$x = \boxed{8.1 \times 10^{-10} M}$$

c. Ag_2CO_3 in 0.20M Na_2CO_3

	$Ag_2CO_3 (s)$	\rightleftharpoons	$2Ag^+ (aq)$	+	$CO_3^{2-} (aq)$
<i>I</i>	—		0		0.20
<i>C</i>	—		+2x		+x
<i>E</i>	—		0.10 + 2x		0.20 + x

small K, small x approx.

$$K_{sp} = [Ag^+]^2[CO_3^{2-}] = (2x)^2(0.20 + x) = (2x)^2(0.20) = 8.1 \times 10^{-12}$$

$$x = \boxed{3.18 \times 10^{-6} M}$$