Chemist’s Daily
Chapter 16 - Ksp

1. Write out the Ksp and equilibrium expression for:

\[
PbCO_3 \quad PbCO_3 \leftrightarrow Pb^{2+} + CO_3^{2-}
Ksp = [Pb^{2+}][CO_3^{2-}]
\]

\[
Ag_2S \quad Ag_2S \leftrightarrow 2 Ag^+ + S^{2-}
Ksp = [Ag^+]^2[S^{2-}]
\]

\[
Ca(OH)_2 \quad Ca(OH)_2 \leftrightarrow Ca^{2+} + 2 OH^-
Ksp = [Ca^{2+}][OH^-]^2
\]

2. Calculate the molar solubility of copper(I) iodide.

Looking in your appendix, Ksp (CuI) = 1 x 10^{-12}

\[
CuI \leftrightarrow Cu^+ + I^-
Ksp = [Cu^+][I^-]
1\times10^{-12} = (x)(x) = x^2
x = 1\times10^{-6}
So the molar solubility is 1 x 10^{-6}
\]

3. The molar solubility of silver sulfate is 1.5 x 10^{-2} M. What is the Ksp for this salt?

\[
Ag_2SO_4 \leftrightarrow 2 Ag^+ + SO_4^{2-}
\]

1.5 x 10^{-2} mols of silver sulfate can dissolve per liter. This leads to:

\[
[Ag^+] = 2\times1.5\times10^{-2} = 3.0\times10^{-2} M
[SO_4^{2-}] = 1.5\times10^{-2} M
Ksp = [Ag^+]^2[SO_4^{2-}] = (3.0\times10^{-2} M)^2(1.5\times10^{-2} M) = 1.4\times10^{-5}
\]
1. Exactly 200 mL of a 0.0040 M BaCl\(_2\) solution are added to exactly 600 mL of 0.0080 M K\(_2\)SO\(_4\). Will a precipitate form?

Solubility rules tell us the only possible precipitate is barium sulfate, so we need only concern ourselves with the initial concentrations of Ba\(^{2+}\) and SO\(_4^{2-}\). We are diluting both solutions, so we must calculate the new initial concentrations:

\[
200 \text{ mL} (0.0040 \text{ M Ba}^{2+}) = 800 \text{ mL} (? \text{ M Ba}^{2+}) \\
[Ba^{2+}] = 0.0010 \text{ M}
\]

\[
600 \text{ mL} (0.0080 \text{ M SO}_4^{2-}) = 800 \text{ mL} (? \text{ M SO}_4^{2-}) \\
[SO_4^{2-}] = 0.0060 \text{ M}
\]

The ion product is 
\[
[Ba^{2+}] [SO_4^{2-}] = (0.0010) (0.0060) = 6.0 \times 10^{-6}
\]
Comparing this to the Ksp for barium sulfate, 1.1 \times 10^{-10}, we find it to be larger than the Ksp so a precipitate will form.

2. What is the solubility of silver chloride in a 6.5 \times 10^{-3} M silver nitrate solution? Express your answer in g/L.

\[
\text{AgCl} \rightleftharpoons \text{Ag}^+ + \text{Cl}^-
\]

Initial: \(6.5 \times 10^{-3} \text{ M}\)
End: \(6.5 \times 10^{-3} \text{ M}\) \(x\ \text{M}\)

\[
\text{Ksp} = [\text{Ag}^+] [\text{Cl}^-] \\
1.8 \times 10^{-10} = (6.5 \times 10^{-3}) (x) \\
x = 2.8 \times 10^{-8}
\]

solubility = \((2.8 \times 10^{-8} \text{ mol} \text{ L}^{-1}) (143.323 \text{ g mol}^{-1}) = 4.0 \times 10^{-6} \text{ g/L}\)