Worksheet 8 Addendum—More Solubility Problems

1. A solution is made with NaI and NaCl such that it is 0.01 M in both I⁻ and Cl⁻. To 1 L of this solution 0.01 moles Ag(NO₃) are added (you can ignore any volume change). The NaI, NaCl, and Ag(NO₃) are completely soluble (as is NaNO₃ but you already knew that). The $K_{sp}$ for AgI is $8.3 \times 10^{-17}$ and for AgCl is $1.8 \times 10^{-10}$.

After the solution has reached equilibrium what are the concentrations of the following?

Will anything precipitate?

Initial concentration of $[Ag^{+}]$ is 0.01 M, $[I^-] = 0.01$ M, $[Cl^-] = 0.01$ M

$$Q_{sp} = [Ag^+][I^-] = (0.01)(0.01) = 10^{-4} \quad \text{AgI could precipitate}$$

$$Q_{sp} = [Ag^+][Cl^-] = (0.01)(0.01) = 10^{-4} \quad \text{AgCl could precipitate}$$

However AgI is much less soluble than AgCl. Assume the AgI precipitates completely to equilibrium

Then you have a saturated solution of AgI

Concentration of $Ag^+$ will be

$$K_{sp} = [Ag^+][I^-] \quad [Ag^+] = \sqrt{K_{sp}} = \sqrt{8.3 \times 10^{-17}} = 9.11 \times 10^{-9}$$

Given this concentration will the AgCl precipitate?

$$Q_{sp} = [Ag^+][Cl^-] = (9.11 \times 10^{-9})(0.01) = 9.11 \times 10^{-11}$$

$$Q_{sp} < K_{sp} \text{ so no AgCl will precipitate}$$

$$[Ag^+] = 9.11 \times 10^{-9} \text{ M}$$

$$[I^-] = 9.11 \times 10^{-9} \text{ M}$$

$$[Cl^-] = 0.01 \text{ M}$$

Are there any solid precipitates? If so how many grams of each.

Only AgI will precipitate. Essentially all the silver will precipitate as AgI. That is 0.01 moles.

$$0.01 \text{ mol}(234.8 \text{ g mol}^{-1}) = 2.35 \text{ g}$$
2. The $K_{sp}$ of PbCl$_2$ is $1.7 \times 10^{-5}$. How many grams of PbCl$_2$ will dissolve in 100 mL of a 0.1 M NaCl solution?

\[
\begin{array}{c|c|c|c|c|c|c}
 & Pb^{2+} & Cl^- \\
I & 0 & .1 \\
C & +x & +2x \\
E & +x & .1+2x \\
\end{array}
\]

\[
K_{sp} = [Pb^{2+}][Cl^-]^2 = (x)(.1 + 2x)^2 \approx (x)(.1)^2
\]

\[
[Pb^{2+}] = K_{sp}/[Cl^-]^2 = (1.7 \times 10^{-5})/(.1)^2 = 1.7 \times 10^{-3}
\]

that will be $(1.7 \times 10^{-3} \text{ M})(.1 \text{ L}) = 1.7 \times 10^{-4} \text{ moles PbCl}_2$

\[
(1.7 \times 10^{-4} \text{ moles})(278.1 \text{ g mol}^{-1}) = 0.047 \text{ g}
\]

3. Will CaF$_2$ be more soluble in acid or base?

F$^-$ is the conjugate base of the weak acid HF. In acid, F$^-$ will form HF allowing more CaF$_2$ to dissolve.

4. Consider the following reactions

AgCN(s) $\rightarrow$ Ag$^+$ (aq) + CN$^-$ (aq)

HCN (aq) $\rightarrow$ H$^+$ (aq) + CN$^-$ (aq)

You a saturated solution of AgCN, what will the effect of each of the following (nothing, more AgCN dissolves, some AgCN precipitates)

What is the concentration of

A. Adding HNO$_3$

Increasing H$^+$ will cause more HCN to form lowering the CN$^-$ concentration. More AgCN will dissolve. (also the Cl$^-$ concentration will increase. If it get high enough AgCl will precipitate causing more AgCN to dissolve)

B. Adding KCN

Adding CN$^-$ will cause some AgCN to precipitate

C. Adding KNO$_3$
Adding K\(^+\) and NO\(_3^-\) will do nothing

5. A blast from the past

\[ AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq) \]
\[ Ag^+(aq) + 2S_2O_3^{2-}(aq) \rightleftharpoons Ag(S_2O_3)_2^{3-}(aq) \]
\[ S_2O_3^{2-}(aq) + H_3O^+(aq) \rightleftharpoons HS_2O_3^-(aq) + H_2O(l) \]

What is the effect of each of these on the solubility of AgBr(s)

1. Adding the soluble salt Kbr

   This will decrease the solubility of the AgBr as the concentration of Br\(^-\) will increase

2. Adding the soluble salt Na\(_2\)S\(_2\)O\(_3\)

   This increase the solubility of the AgBr. The S\(_2\)O\(_3^-\) will react with the silver to form Ag(S\(_2\)O\(_3\))\(_2^{3-}\). This will decrease the Ag\(^+\) concentration leading to more AgBr dissolving.

3. Adding HCl

   Adding HCl will cause the S\(_2\)O\(_3^{2-}\) to form HS\(_2\)O\(_3^-\). This will decrease in S\(_2\)O\(_3^{2-}\). This will caused Ag(S\(_2\)O\(_3\))\(_3^{3-}\) to dissolve forming more Ag\(^+\). This will decrease the solubility of the AgBr

4. Adding solid AgBr

   This will have no effect.