Section 15.10 - Solubility Equilibria

1. **Solubility Rules.** Predict whether each compound is soluble in water or not.
   (a) BaSO₄
   (b) AgCl
   (c) KOH
   (d) (NH₄)₂CO₃
   (e) FeCrO₄

2. Write the reaction for these sparingly soluble salts dissolving in water.
   (a) barium phosphate
   (b) cuprous hydroxide
   (c) manganese (II) sulfide

3. What is the simple equation in your “toolbox” used to calculate solubility?
4. **Relative Solubilities.** The relative values of the Ksp’s can’t always be used to decide which compound is more soluble. For example, based on the Ksp’s below, it appears that AgCl is more soluble. Prove using the simple equation in question 3 that Ag₂CrO₄ is actually more soluble.

\[
\text{Ag}_2\text{CrO}_4 \ (K_{sp} = 2.0 \times 10^{-12})
\]

\[
\text{AgCl} \ (K_{sp} = 1.8 \times 10^{-10})
\]

5. **Environmental Chemistry Application.** Lead pipes were used at one time for delivering drinking water, and in places like Flynt, MI, there are still some lead pipes in the water system. Calculate the maximum possible concentration of lead in water if it comes from lead(II) hydroxide dissolving from the pipes. Note that the EPA limit on lead in drinking water is \(7.2 \times 10^{-8}\) M. \(K_{sp} = 2.8 \times 10^{-16}\)
6. **Common Ion Effect.** Calculate the solubility of PbCl₂ in (a) pure water, and (b) 0.10 M NaCl. 
   \[ K_{sp} = 2.4 \times 10^{-4} \].
7. Predicting if a Precipitate Forms – Comparing Q to Ksp.

   (a) Will a precipitate form when 50.0 mL of $1.2 \times 10^{-3}$ M Pb(NO$_3$)$_2$ is added to 50.0 mL of $2.0 \times 10^{-4}$ M Na$_2$SO$_4$? $K_{sp}$(PbSO$_4$) = $1.8 \times 10^{-8}$

   (b) Will a precipitate of MgF$_2$ form when 300 mL of $1.1 \times 10^{-3}$ M MgCl$_2$ solution is added to 500 mL of $1.2 \times 10^{-3}$ M NaF? $K_{sp}$(MgF$_2$) = $6.9 \times 10^{-9}$
Using pH to Control Solubility in the Qualitative Analysis Lab – Separating Groups II and III.

Qualitative Analysis of Cations

From lecture, we discussed qualitatively how pH and Le Chatelier’s Principle can be used to control the solubility of salts that containing an acidic or basic ion. The example we looked at was CaF$_2$ where the F$^-$ ion is the conjugate base of the weak acid HF:

\[
\text{CaF}_2(s) = \text{Ca}^{2+}(aq) + 2\text{F}^-(aq)
\]

\[
\text{F}^-(aq) + \text{H}_2\text{O}(l) = \text{HF}(aq) + \text{OH}^-(aq)
\]

Using Le Chatelier’s Principle, explain how the solubility of CaF$_2$ can be increased by either increasing or decreasing the pH.
Metal Sulfides and Their Solubility

<table>
<thead>
<tr>
<th>Metal Sulfide</th>
<th>$K_{sp}$</th>
<th>$[S^{2-}]$ Required to Begin Precipitation from 0.0010 M $M^{2+}$ Solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>HgS</td>
<td>$3.0 \times 10^{-53}$</td>
<td></td>
</tr>
<tr>
<td>CuS</td>
<td>$8.7 \times 10^{-36}$</td>
<td></td>
</tr>
<tr>
<td>CdS</td>
<td>$3.6 \times 10^{-29}$</td>
<td></td>
</tr>
<tr>
<td>ZnS</td>
<td>$1.1 \times 10^{-21}$</td>
<td></td>
</tr>
<tr>
<td>CoS</td>
<td>$5.9 \times 10^{-21}$</td>
<td></td>
</tr>
<tr>
<td>MnS</td>
<td>$5.1 \times 10^{-15}$</td>
<td></td>
</tr>
</tbody>
</table>

1. For each metal sulfide, calculate the $[S^{2-}]$ needed to begin precipitation from a 0.0010 M solution of the metal cation. Enter the values into the last column of the table. Show one calculation here.

2. Approximately how many orders of magnitude less is the solubility of CdS compared to ZnS?

3. Using CdS as an example, fill in the pH-dependent equations below. Using Le Chatelier’s Principle, explain how pH can be used to separate it from ZnS.

$$\text{CdS(s)} = \underline{\text{________}} + \underline{\text{________}}$$

$$S^{2-}(aq) + \underline{\text{________}} = \underline{\text{________}} + \underline{\text{________}}$$

Net reaction:
(add together)