Part A: Understanding Precipitation Conceptually

Fill in the Blanks: Fill in the blanks with (<, >, =)

If Q ___ K_{sp} the solution is unsaturated and more of the solid compound can dissolve the solution.

If Q ___ K_{sp} the solution is supersaturated. Any slight disturbance will initiate rapid precipitation.

If Q ___ K_{sp} the solution is saturated and at equilibrium. No additional solid will dissolve in the solution.

Part B: Precipitation

1) What is the minimum concentration of I^- that can cause precipitation of lead iodide from a 0.050 M solution of lead nitrate? The K_{sp} for $PbI_2 = 8.7 \times 10^{-9}$.

2) Will a precipitate of magnesium hydroxide form when 25.0 mL of 0.010 M sodium hydroxide is combined with 75.0 mL of a 0.10 M solution of magnesium chloride? $K_{sp} = 2.1 \times 10^{-13}$.

3) You add aqueous KOH to a solution containing Mg^{2+} (0.059 M) and Ca^{2+} (0.011 M). When the $[OH^-]$ reaches 1.9×10^{-6} M, $Mg(OH)_2$ begins to precipitate out. When you add more OH^-, the $Ca(OH)_2$ precipitates out too. What is the concentration of Mg^{2+} when the Ca^{2+} begins to precipitate out as $Ca(OH)_2$? K_{sp} $Mg(OH)_2 = 2.1 \times 10^{-13}$ *and* K_{sp} $Ca(OH)_2 = 4.68 \times 10^{-6}$

Part B: Qualitative chemical analysis

4) Imagine you work as a chemical engineer for the Sacramento water district and need to test samples of drinking water for possible contamination. Based on tests you have done, you have narrowed down that there is some combination of Ba^{2+}, Ag^{+}, and/or Fe^{2+} ions in the water sample. You perform a series of tests and make the following observations:

Test #1: add 6 M HCl to the sample	Result of test #1: a precipitate forms

Remove the solid from the sample and continue testing the remaining liquid.

Test #2: add NaOH to the sample	Result of test #2: no precipitate forms
Test #3: add $Na_3(PO_4)_2$ to the sample	Result of test #3: a precipitate forms

Based on the test results, which of the possible ions are in the sample? Briefly explain your answer based on the solubility rules shown on the bottom of this page. Draw a flow chart and write net ionic equations for those reactions that form a precipitate.

Table 1: Solubility rules for ionic compounds in water

Soluble salts include:

All Li^{+}, Na^{+}, K^{+}, and NH_4^{+}
All NO_3^{-} and $C_2H_3O_2^{-}$
All Cl^{-}, Br^{-}, and I^{-} [exceptions: Ag^{+}, Pb^{2+}, Hg_2^{2+}]
All SO_4^{2-} [exceptions: Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, or Ag^{+} (slightly soluble)]

Insoluble salts include:

All PO_4^{3-} and CO_3^{2-} [exceptions: Li^{+}, Na^{+}, K^{+}, and NH_4^{+}]
All OH^{-} and S^{2-} [exceptions: Li^{+}, Na^{+}, K^{+}, NH_4^{+}, Ca^{2+}, Sr^{2+}, and Ba^{2+}]

Part C: Complex ion equilibria

5) When an aqueous solution of NaCN is added to AgCl (s), the solid dissolves resulting in the complex ion, $[Ag(CN)_2]^-$ (aq).

a. Write the balanced net ionic equation that shows the formation of $[Ag(CN)_2]^-$ from AgCl(s) and NaCN (aq).

b. What is the value of the equilibrium constant for dissolving AgCl(s) in a NaCN (aq) solution given the Ksp of AgCl(s) = 1.8×10^{-10} and the K_f ($[Ag(CN)_2]^-$) = 1.0×10^{21}.

c. What is the molar solubility of AgCl (s) in a 0.10 M NaCN solution?