BC Science Chemistry 12

Chapter – 3 Solubility Equilibrium Answer Key

September 20, 2012

3.1 The Concept of Solubility

Warm Up

1.

	ionic	**molecular**
Potassium phosphate	√	
Xenon hexafluoride		√
Phosphoric acid		√
Sulphur dioxide		√
Ammonium sulphide	√	

2. Ionic compounds will conduct electricity.

Quick Check

1. No. Some salts are weak electrolytes. They do not dissolve to a significant extent.

2. a) molecular

b) ionic

c) ionic

d)molecular

e) ionic

3. Sample A is tap water. It has a low electrical conductivity due to low concentrations of dissolved salts in tap water. Sample B is carbon tetrachloride. It is molecular and does not conduct electricity. Sample C is seawater. It has a significant amount of dissolved ions present so will conduct electricity well.

Quick Check

1. $CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$

2. $(NH_4)_2C_2O_4(aq) \rightarrow 2NH_4^+(aq) + C_2O_4^{2-}(aq)$

3. $Na_3PO_4(aq) \rightarrow 3Na^+(aq) + PO_4^{3-}(aq)$

Quick Check

1. a) $Na_2C_2O_4(aq) \rightarrow 2Na^+(aq) + C_2O_4^{2-}(aq)$

0.45 M 0.90 M 0.45 M total [ions] = 0.90 M + 0.45 M = 1.35 M

b) $AlCl_3(aq) \rightarrow Al^{3+}(aq) + 3Cl^-(aq)$

2.5 M 2.5 M 7.5 M total [ions] = 2.5 M + 7.5 M = 10.0 M

2. $LiOH(aq) \rightarrow Li^+(aq) + OH^-(aq)$

1.0 M 1.0 M 1.0 M total ion concentration = 2.0 M

$CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$

0.8 M 0.8 M 1.6 M total ion concentration = 2.4 M

The $CaCl_2$ will have the greater conductivity because it has a greater ion concentration.

3. $AlBr_3(aq) \rightarrow Al^{3+}(aq) + 3Br^-(aq)$

According to the dissociation equation, 1 mole Al^{3+} : 3 mole Br^-. If $[Br^-]$ = 0.15 M, then $[Al^{3+}]$ = 0.050 M.

Practice Problems: Calculating Ion Concentrations

1. a) $[(NH_4)_3PO_4] = \frac{0.35\ \cancel{g}}{0.650\ L} \times \frac{1\ mol}{149.12\ \cancel{g}} = 3.6 \times 10^{-3}\ M$

$[NH_4^+] = 3(3.6 \times 10^{-3}) = 1.1 \times 10^{-2}\ M$ $[PO_4^{3-}] = 3.6 \times 10^{-3}\ M$

b) $[HCl] = \frac{6.0\ mol}{1\ \cancel{L}} \times \frac{0.015\ \cancel{L}}{0.035\ L} = 2.6\ M$

$[H^+]$ = 2.6 M $[Cl^-]$ = 2.6 M

2. $Al_2(SO_4)_3(aq) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$

0.15 M 0.45 M

mass of $Al_2(SO_4)_3 = 0.250\ L \times \frac{0.15\ \cancel{mol}}{1\ L} \times \frac{342.17\ g}{1\ \cancel{mol}} = 13\ g$

3. $[(NH_4)_2SO_4] = \frac{0.22 \text{ mol}}{1 \text{ L}} \times \frac{0.055 \text{ L}}{0.080 \text{ L}} = 0.15$ M $[NH_4^+] = 0.30$ M, $[SO_4^{2-}] = 0.15$ M

$[(NH_4)_2S] = \frac{0.45 \text{ mol}}{1 \text{ L}} \times \frac{0.025 \text{ L}}{0.080 \text{ L}} = 0.14$ M $[NH_4^+] = 0.28$ M, $[S^{2-}] = 0.14$ M

total $[NH_4^+] = 0.58$ M, $[SO_4^{2-}] = 0.15$ M $[S^{2-}] = 0.14$ M

Practice Problems: Converting Between Units of Solubility

1. $\frac{\text{mol}}{\text{L}}$ MnS $= \frac{4.7 \times 10^{-6} \text{ g}}{0.001 \text{ L}} \times \frac{1 \text{ mol}}{87.01 \text{ g}} = 5.4 \times 10^{-2}$ M

2. Mass $Pb(IO_3)_2 = 0.300 \text{ L} \times \frac{4.5 \times 10^{-5} \text{ mol}}{1 \text{ L}} \times \frac{557.17 \text{ g}}{1 \text{ mol}} = 7.5 \times 10^{-3}$ g

3. $\frac{\text{mol}}{\text{L}}$ $CaCO_3 = \frac{7.1 \times 10^{-4} \text{ g}}{0.100 \text{ L}} \times \frac{1 \text{ mol}}{100.09 \text{ g}} = 7.1 \times 10^{-5}$ M

Quick Check

1. $ZnCO_3(s) \rightleftharpoons Zn^{2+}(aq) + CO_3^{2-}(aq)$

2. $Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$

3. $Fe_2S_3(s) \rightleftharpoons 2Fe^{3+}(aq) + 3S^{2-}(aq)$

3.1 Review

1. a) any soluble salt or strong acid or base: NaCl, HCl, NaOH. Many answers possible.

b) any molecular substance not including weak acids: CCl_4, SO_2, XeF_6

2. Both are acids with the same concentration, but $HClO_4$ is a strong acid that ionizes completely. H_3PO_4 is a weak acid, and remains largely in molecular form. Since there are more ions in the $HClO_4$, it will have a greater electrical conductivity.

3. a) $Mg(ClO_4)_2(aq) \rightarrow 2Mg^{2+}(aq) + 2ClO_4^-(aq)$

b) $CaCr_2O_7(aq) \rightarrow Ca^{2+}(aq) + Cr_2O_7^{2-}(aq)$

c) $Cu(CH_3COO)_2(aq) \rightarrow Cu^{2+}(aq) + 2CH_3COO^-(aq)$

d) $Mn(SCN)_2(aq) \rightarrow Mn^{2+}(aq) + 2SCN^-(aq)$

e) $Al(HC_2O_4)_3(aq) \rightarrow Al^{3+}(aq) + 3HC_2O_4^-(aq)$

f) $Ba(OH)_2 \bullet 8H_2O(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq) + 8\, H_2O(l)$

4. $\frac{\text{mol } BaSO_4}{\text{L}} = \frac{0.0012 \text{ } \cancel{g}}{0.500 \text{ L}} \times \frac{1 \text{ mol}}{233.40 \text{ } \cancel{g}} = 1.0 \times 10^{-5} \text{ M}$

5. mass $CaCO_3 = 0.250 \text{ } \cancel{L} \times \frac{7.1 \times 10^{-5} \text{ } \cancel{mol}}{1 \cancel{L}} \times \frac{100.09 \text{ g}}{1 \text{ } \cancel{mol}} = 1.8 \times 10^{-3} \text{ g}$

6. $[Na_2Cr_2O_7] = \frac{0.50 \text{ } \cancel{g}}{0.150 \text{ L}} \times \frac{1 \text{ mol}}{261.98 \text{ } \cancel{g}} = 0.0126 \text{ M}$

$Na_2Cr_2O_7(aq) \rightarrow 2Na^+(aq) + Cr_2O_7^{2-}(aq)$
0.0126 M   2(0.0126) M   0.0126 M

$[Na^+] = 0.025$ M $[Cr_2O_7^{2-}] = 0.013$ M

7. $[NaCl] = \frac{5.0 \text{ mol}}{1 \text{ } \cancel{L}} \times \frac{60. \text{ } \cancel{L}}{90. \text{ L}} = 3.3 \text{ M}$ $[Na^+] = 3.3$ M $[Cl^-] = 3.3$ M

$[MgCl_2] = \frac{2.4 \text{ mol}}{1 \text{ } \cancel{L}} \times \frac{30. \text{ } \cancel{L}}{90. \text{ L}} = 0.80 \text{ M}$ $[Mg^{2+}] = 0.80$ M $[Cl^-] = 1.6$ M

$[Mg^{2+}] = 0.80$ M, total $[Cl^-] = 4.9$ M, $[Na^+] = 3.3$ M

8. Measure out 1.0 L of water and add solid NaCl until there is undissolved solid remaining. Pour off until 1.0 L solution remains.

9. a$AgBrO_3(s) \rightleftharpoons Ag^+(aq) + BrO_3^-(aq)$

b) $Al_2(CrO_4)_3(s) \rightleftharpoons 2Al^{3+}(aq) + 2CrO_4^{2-}(aq)$

c) $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$

d) $PbSO_4(s) \rightleftharpoons Pb^{2+}(aq) + SO_4^{2-}(aq)$

e) $Cu_3(PO_4)_2(s) \rightleftharpoons 3Cu^{2+}(aq) + 2PO_4^{3-}(aq)$

3.2 Qualitative Analysis – Identifying Unknown Ions

Warm Up

1. $\frac{\text{mol AgCl}}{\text{L}} = \frac{1.4 \times 10^{-3} \text{ } \cancel{g}}{\text{L}} \times \frac{1 \text{ mol}}{143.32 \text{ } \cancel{g}} = 9.8 \times 10^{-6} \text{ M}$

2. $\frac{\text{mol NaCl}}{\text{L}} = \frac{359\ \cancel{g}}{\text{L}} \times \frac{1\ \text{mol}}{58.44\ \cancel{g}} = 6.14\ \text{M}$

3. AgCl has low solubility, and NaCl is soluble.

Practice Problems: Predicting the Solubility of Salts Using the Solubility Table

1.

a) low solubility

b) low solubility

c) soluble

d) soluble

e) soluble

d) soluble

2. (many answers possible)

a) Na_2CO_3

b) $SrSO_4$

c) Ag^+

d) NO_3^-

3. The sodium hydroxide contains sodium ions and hydroxide ions. Sodium ions form soluble compounds, so any precipitate must contain the hydroxide anion. The two compounds possible for the precipitate are $Mg(OH)_2$ and $Sr(OH)_2$. According to the Solubility Table, $Mg(OH)_2$ has low solubility and $Sr(OH)_2$ is soluble. The sample must contain strontium nitrate since no precipitate was observed.

Quick Check

1.

a) $BaSO_4$

b) $Fe(OH)_2$

c) $Zn_3(PO_4)_2$

d) $CrCO_3$

e) $Mn(OH)_2$

2.

a) $Sr(OH)_2(aq)$ + $2AgNO_3(aq)$ → $2AgOH(s)$ + $Sr(NO_3)_2(aq)$

$Sr^{2+}(aq) + 2OH^-(aq) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow 2AgOH(s) + Sr^{2+}(aq) + 2NO_3^-(aq)$

$Ag^+(aq) + OH^-(aq) \rightarrow AgOH(s)$

b)

$MgS(aq)$ + $ZnCl_2(aq)$ → $ZnS(s)$ + $MgCl_2(aq)$

$Mg^{2+}(aq) + S^{2-}(aq) + Zn^{2+}(aq) + 2Cl^-(aq) \rightarrow ZnS(s) + Mg^{2+}(aq) + 2Cl^-(aq)$

$Zn^{2+}(aq) + S^{2-}(aq) \rightarrow ZnS(s)$

c)

$Na_2CO_3(aq)$ + $BaS(aq)$ → $BaCO_3(s)$ + $Na_2S(aq)$

$2Na^+(aq) + CO_3^{2-}(aq) + Ba^{2+}(aq) + S^{2-}(aq) \rightarrow BaCO_3(s) + 2Na^+(aq) + S^{2-}(aq)$

$Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$

Practice Problems: Selective Precipitation

(other answers possible)

1.

First add NaCl. Filter out the $AgCl(s)$

Then add Na_2S. Filter out the $BeS(s)$

Finally add Na_2SO_4. Filter out the $BaSO_4(s)$

2.

First add $Ca(NO_3)_2$. Filter out the $CaSO_4(s)$

Then add $Cu(NO_3)_2$. Filter out the $CuS(s)$

Finally add $AgNO_3$. Filter out the $AgBr(s)$

3.

First add $Sr(NO_3)_2$. Filter out the $Sr_3(PO_4)_2(s)$

Then add $Mg(NO_3)_2$. Filter out the $Mg(OH)_2(s)$

Finally add $AgNO_3$. Filter out the $Ag_2S(s)$

Quick Check

1.

- The color of the solution (for example, solutions containing Cu^{2+} are blue)

- flame test color

- solubility – amount of precipitate formed, the speed at which a precipitate forms, or the ability of a precipitate to re-dissolve on the addition of NH_3 or HNO_3.

2. First add NaCl. A precipitate of both $PbCl_2$ and AgCl may form. Filter out the precipitate, dry it and measure its mass. Put the precipitate into a second beaker. To the original solution, add NaOH. If a precipitate forms, filter out the $Cu(OH)_2(s)$. To the second beaker, add 6 M NH_3. Any AgCl re-dissolves. Filter out the $PbCl_2(s)$. Dry the PbCl2(s) and measure its mass. If the mass had decreased, the presence of AgCl is confirmed.

3. First add HNO_3. The formation of bubbles confirms the presence of CO_3^{2-} ions. Next, add $CaCl_2$. If a precipitate forms, it is $Ca_3(PO_4)_2(s)$. Filter out the precipitate.

Activity: Using Titration to Calculate the Unknown [Cl^-] In A Sample of Seawater

Question: What is the concentration of Cl^- in seawater?

Procedure:

4. Procedures 1-3 were repeated in 2 more trials, and the data below was recorded.

	Trial #1	Trial #2	Trial #3

Initial burette reading (mL)	0.00	14.70	28.60
Final burette reading (mL)	14.70	28.60	42.35
Volume $AgNO_3$ added (mL)	14.70	13.90	13.75
Average volume $AgNO_3$ used (mL)	13.83	X	X

Results and Discussion

1. Write a balanced net ionic equation that represents the reaction between silver ion and chloride ion.

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

2. Calculate the moles of silver nitrate reacted in the titration using the average volume and concentration of the silver nitrate.

$$\text{mol } AgNO_3 = 0.01383\ \cancel{L} \times \frac{0.100\text{ mol}}{1\ \cancel{L}} = 0.00138\text{ mol } AgNO_3$$

3. Using the mole ratio from the balanced net ionic equation, calculate the moles of chloride ion present in the sample.

$$\text{mol } Cl^- = 0.00138\ \cancel{\text{mol } Ag^+} \times \frac{1\text{ mol } Cl^-}{1\ \cancel{\text{mol } Ag^+}} = 0.00138\text{ mol } Cl^-$$

4. Using the volume of the **diluted** sample and the moles of chloride calculated above, calculate the concentration of chloride ion in the diluted sample.

In the diluted sample, the $[Cl^-] = \dfrac{0.00138\text{ mol } Cl^-}{0.0250\text{ L}} = 0.0552\text{ M}$

5. Calculate the concentration of chloride ion in the original sample of seawater.

The original sample was diluted 10x, so in the original sample, the $[Cl^-]$ = 0.552 M.

6. The level of salt in seawater varies across the planet. The average amount of chloride ion in seawater is 21.2 g/L. How does your sample compare?

To convert 0.552 M Cl^- to g/L:

$$\frac{g}{L} = \frac{0.552\ \cancel{mol}}{L} \times \frac{35.45\ g}{1\ \cancel{mol}} = 19.6\ g$$

Our sample has slightly less chloride in it than the average.

7. Silver ions will form a precipitate with both chloride and chromate ions. Use a reference book to look up the solubility of AgCl and Ag_2CrO_4. Calculate the concentration of Ag^+ in a saturated solution of each. Use this information to explain why AgCl precipitates before Ag_2CrO_4.

The solubility of AgCl = 1.3×10^{-5} M

The solubility of Ag_2CrO_4 = 6.5×10^{-5} M

When silver ions are added in the titration, AgCl will precipitate first because it has a lower solubility than Ag_2CrO_4.

Review Questions:

1.

a) soluble

b) low solubility

c) low solubility

d) soluble

e) soluble

f) soluble

2. Yes – there are silver and sulphate ions in solution. A low solubility simply means that very little will dissolve; **not** that it is insoluble.

3. $\frac{\text{mol } AgCH_3COO}{\text{L}} = \frac{11.1\text{ g}}{\text{L}} \times \frac{1\text{ mol}}{166.92\text{ g}} = 0.0665\text{ M}$

$0.0665\text{ M} < 0.1\text{ M}$ so $AgCH_3COO$ would have a low solubility.

4. A formula equation is written with substances as compounds and states shown.

A complete ionic equation is written with soluble ionic compounds and strong acids in dissociated form, and compounds with low solubility undissociated. States are also shown.

A net ionic equation only shows the substances that take part in the reaction. Spectator ions are not shown. States are also shown.

5. A spectator ion is one that does not take part in the reaction. Na^+ ions and NO_3^- ions are often spectators.

6.

a) $(NH_4)_2S(aq) + FeSO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + FeS(s)$

$2NH_4^+(aq) + S^{2-}(aq) + Fe^{2+}(aq) + SO_4^{2-}(aq) \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq) + FeS(s)$

$Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$

b) $H_2SO_3(aq) + CaCl_2(aq) \rightarrow 2HCl(aq) + CaSO_3(s)$

$H_2SO_3(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \rightarrow 2H^+(aq) + 2Cl^-(aq) + CaSO_3(s)$

$H_2SO_3(aq) + Ca^{2+}(aq) + SO_3^{2-}(aq) \rightarrow 2H^+(aq) + CaSO_3(s)$

c) $CuSO_4(aq) + CaS(aq) \rightarrow CuS(s) + CaSO_4(s)$

$Cu^{2+}(aq) + SO_4^{2-}(aq) + Ca^{2+}(aq) + S^{2-}(aq) \rightarrow CuS(s) + CaSO_4(s)$

$Cu^{2+}(aq) + SO_4^{2-}(aq) + Ca^{2+}(aq) + S^{2-}(aq) \rightarrow CuS(s) + CaSO_4(s)$

7. Na^+ and K^+ are alkali ions. Compounds containing alkali ions are soluble in water.

8. First add Na_2SO_4 to precipitate $CaSO_4$: $Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$

Filter out the precipitate

Secondly, add Na_2S to precipitate Cr_2S_3: $2Cr^{3+}(aq) + 3S^{2-}(aq) \rightarrow Cr_2S_3\ (s)$

Filter out the precipitate

Finally add NaOH to precipitate $Mg(OH)_2$: $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

Filter out the precipitate

9. First add $Mg(NO_3)_2$ to precipitate $Mg_3(PO_4)_2$: $3Mg^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Mg_3(PO_4)_2\ (s)$

Filter out the precipitate

Secondly, add $Fe(NO_3)_2$ to precipitate FeS: $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$

Filter out the precipitate

Finally add $AgNO_3$ to precipitate AgCl: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Filter out the precipitate

10. Nitrates are soluble in water. If a precipitate forms when a nitrate is added, the precipitate must contain the cation of the nitrate.

11. AgCl can be dissolved in 6M NH_3, while AgI cannot. This can be used to separate AgCl from AgI or AgBr.

Carbonates will dissolve on the addition of a strong acid such as HNO_3.

12.

a) $2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$

b) $Ag_2CO_3(s) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l) + 2Ag^+(aq)$

$Ag_2CO_3(s) + 4NH_3(aq) \rightarrow 2Ag(NH_3)_2^+(aq) + CO_3^{2-}(aq)$

13. a) Hard water contains Ca^{2+} ions, Fe^{2+} ions and/or Mg^{2+} ions.

Scale is $CaCO_3$ that collects inside water heaters, pipes, kettles and boilers.

b) $CaCO_3(s) + 2HCl(aq) \rightarrow H_2O(l) + CO_2(g) + CaCl_2(aq)$

$CaCO_3(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow H_2O(l) + CO_2(g) + Ca^{2+}(aq) + 2Cl^-(aq)$

$CaCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g) + Ca^{2+}(aq)$

c) Many answers are possible: Na_2CO_3

d) Water softeners remove Ca^{2+} ions and Mg^{2+} ions from hard water. Soap scum can't form without these ions.

3.3 The Solubility Product Constant - K_{sp}

Warm Up

1.

a) the maximum amount of solute that can dissolve in a given volume of solvent at a particular temperature.

b) the answer when numbers are multiplied together

c) a number that does not change

2.

a) $PbCl_2(aq) \rightarrow Pb^{2+}(aq) + 2Cl^-(aq)$

b) $\frac{mol}{L} = \frac{4.4\ \cancel{g}}{L} \times \frac{1\ mol}{278.27\ \cancel{g}} = 0.016\ M$

c) $PbCl_2(aq) \rightarrow Pb^{2+}(aq) + 2Cl^-(aq)$

0.016 M   0.016 M   0.032 M

Quick Check

1. a) $SrCO_3(s) \rightleftharpoons Sr^{2+}(aq) + CO_3^{2-}(aq)$

b) $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$

c) $Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$

2. a) $K_{sp} = [Sr^{2+}][CO_3^{2-}]$

b) $K_{sp} = [Mg^{2+}][OH^-]^2$

c) $K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$

3. The solubility is the maximum *amount* of solute that can be dissolved in a particular volume of solvent. The solubility product constant is the *product* of the ion concentrations raised to the power of the coefficients from the equilibrium.

Practice Problems

1. a)

$CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$

6.1×10^{-5} M 6.1×10^{-5} M 6.1×10^{-5} M

$K_{sp} = [Ca^{2+}][CO_3^{2-}]$

K_{sp} = (6.1 x 10^{-5})(6.1 x 10^{-5})

K_{sp} = 3.7 x 10^{-9}

b)

$Mn(OH)_2(s)$ ⇌ $Mn^{2+}(aq)$ + $2OH^{-}(aq)$

3.6 x 10^{-5} M   3.6 x 10^{-5} M   7.2 x 10^{-5} M

$K_{sp} = [Mn^{2+}][OH^{-}]^2$

K_{sp} = (3.6 x 10^{-5})(7.2 x 10^{-5})^2

K_{sp} = 1.9 x 10^{-13}

c)

Solubility of $BaCrO_4$:

$\frac{\text{mol}}{\text{L}} = \frac{2.8 \times 10^{-3}\ \text{g}}{\text{L}} \times \frac{1\ \text{mol}}{253.33\ \text{g}}$ = 1.1 x 10^{-5} M

$BaCrO_4(s)$ ⇌ $Ba^{2+}(aq)$ + $CrO_4^{2-}(aq)$

1.1 x 10^{-5} M   1.1 x 10^{-5} M   1.1 x 10^{-5} M

$K_{sp} = [Ba^{2+}][CrO_4^{2-}]$

K_{sp} = (1.1 x 10^{-5})(1.1 x 10^{-5})

K_{sp} = 1.2 x 10^{-10}

d)

Solubility of $Ag_2C_2O_4$:

$\frac{\text{mol}}{\text{L}} = \frac{0.033\ \text{g}}{\text{L}} \times \frac{1\ \text{mol}}{303.76\ \text{g}}$ = 1.1 x 10^{-4} M

$Ag_2C_2O_4\ (s)$ ⇌ $2Ag^{+}(aq)$ + $C_2O_4^{2-}(aq)$

1.1 x 10^{-4} M   2.2 x 10^{-4} M   1.1 x 10^{-4} M

$K_{sp} = [Ag^{+}]^2[C_2O_4^{2-}]$

K_{sp} = (2.2 x 10^{-4})^2 (1.1 x 10^{-4})

K_{sp} = 5.2 x 10^{-12}

2.

Solubility of $Mg(OH)_2$:

$\frac{mol}{L} = \frac{5.5 \times 10^{-5} \text{ mol}}{0.500 \text{ L}} = 1.1 \times 10^{-4}$ M

$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$

1.1×10^{-4} M 1.1×10^{-4} M 2.2×10^{-4} M

$K_{sp} = [Mg^{2+}][OH^-]^2$

$K_{sp} = (1.1 \times 10^{-4})(2.2 \times 10^{-4})^2$

$K_{sp} = 5.2 \times 10^{-12}$

3.

Solubility of MgC_2O_4 :

$\frac{mol}{L} = \frac{0.16 \text{ g}}{0.150 \text{ L}} \times \frac{1 \text{ mol}}{112.33 \text{ g}} = 9.5 \times 10^{-3}$ M

$MgC_2O_4 (s) \rightleftharpoons Mg^{2+}(aq) + C_2O_4^{2-}(aq)$

9.5×10^{-3} M 9.5×10^{-3} M 9.5×10^{-3} M

$K_{sp} = [Mg^{2+}][C_2O_4^{2-}]$

$K_{sp} = (9.5 \times 10^{-3}) (9.5 \times 10^{-3})$

$K_{sp} = 9.0 \times 10^{-5}$

Practice Problems:

1. a)

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

s s

$K_{sp} = [Ag^+][Cl^-]$

$1.8 \times 10^{-10} = s^2$

$s = 1.3 \times 10^{-5}$ M

b)

$FeS(s) \rightleftharpoons Fe^{2+}(aq) + S^{2-}(aq)$

$s \qquad s$

$K_{sp} = [Fe^{2+}][S^{2-}]$

$6.0 \times 10^{-19} = s^2$

$s = 7.7 \times 10^{-10}$ M

g = $\underline{7.7 \times 10^{-10}}$ ~~mol~~ x 87.92 g x 1 ~~L~~ = 6.8×10^{-11} g

mL 1 ~~L~~ 1 ~~mol~~ 1000 mL mL

c)

$Pb(IO_3)_2(s) \rightleftharpoons Pb^{2+}(aq) + 2IO_3^-(aq)$

$s \qquad 2s$

$K_{sp} = [Pb^{2+}][IO_3^-]^2$

$3.7 \times 10^{-13} = (s)(2s)^2 = 4s^3$

$s^3 = \frac{3.7 \times 10^{-13}}{4} = 9.25 \times 10^{-14}$

$s = 4.5 \times 10^{-5}$ M

d)

$SrF_2(s) \rightleftharpoons Sr^{2+}(aq) + 2F^-(aq)$

$s \qquad 2s$

$K_{sp} = [Sr^{2+}][F^-]^2$

$4.3 \times 10^{-9} = (s)(2s)^2 = 4s^3$

$s^3 = \frac{4.3 \times 10^{-9}}{4} = 1.1 \times 10^{-9}$

$s = 1.0 \times 10^{-3}$ M

g = $\underline{1.0 \times 10^{-3}}$ ~~mol~~ x 125.62 g = 6.8×10^{-11} g

L 1 L 1 ~~mol~~ L

2.

$Fe(OH)_3(s) \rightleftharpoons Fe^{3+}(aq) + 3OH^-(aq)$

s $3s$

$K_{sp} = [Fe^{3+}][OH^-]^3$

$2.6 \times 10^{-39} = (s)(3s)^3 = 27s^4$

$s^4 = \frac{2.6 \times 10^{-39}}{27} = 9.6 \times 10^{-41}$

$s = 9.9 \times 10^{-11}$ M

$[OH^-] = 3s = 3(9.9 \times 10^{-11}) = 3.0 \times 10^{-10}$ M

3.

$CaC_2O_4(s) \rightleftharpoons Ca^{2+}(aq) + C_2O_4^{2-}(aq)$

s s

$K_{sp} = [Ca^{2+}][C_2O_4^{2-}]$

$2.3 \times 10^{-9} = s^2$

$s = 4.8 \times 10^{-5}$ M

mass (g) $CaC_2O_4 = 0.650\ \cancel{L} \times \frac{4.8 \times 10^{-5}\ \cancel{mol}}{1\ \cancel{L}} \times \frac{128.10\ g}{1\ \cancel{mol}} = 4.0 \times 10^{-3}$ g

Activity: Experimentally Determining the K_{sp} of $CaCO_3$

Question: What is the K_{sp} of $CaCO_3$?

Procedure:

This is only one method. Other answers are possible. This procedure is done at room temperature. If another temperature is used, either heat up or cool down the distilled water to the required temperature.

Reagents: 150 mL distilled water

Calcium carbonate solid

Apparatus:

3 250 mL beakers

1 hot plate

1 scoopula

1 stirring rod

filter paper

funnel

1 100. mL graduated cylinder

1 scale

1 thermometer

Procedure:

1. Prepare a saturated solution of $CaCO_3$ by measuring out 150 mL of distilled water into the first beaker.

2. Continue to add solid $CaCO_3$ and stir until no more $CaCO_3$ dissolves. You should see un-dissolved solid in the solution.

3. Measure and note the temperature of the solution in the data table.

4. Filter the saturated solution of $CaCO_3$ into the second beaker.

5. Using the graduated cylinder, measure 100. mL of the filtered solution.

6. Measure and record the mass of the third beaker in your data table.

7. Pour the 100. mL of filtered solution into the third beaker.

8. Heat the beaker on the hot plate until all water has evaporated. Allow the $CaCO_3$ to cool.

9. Measure and record the mass of the beaker and solid $CaCO_3$.

Data Table:

Temperature of saturated solution:

Mass of beaker:

Mass of beaker and $CaCO_3(s)$:

Results and Discussion:

1. Describe how you would analyze the data and perform calculations to determine the K_{sp}.

Calculate the mass of $CaCO_3$ dissolved in 100. mL of saturated solution by subtracting the Mass of beaker from the Mass of beaker and $CaCO_3(s)$.

Calculate the molar solubility using:

$$s = \frac{\text{mass of } CaCO_3 \text{ (g)}}{0.100 \text{ L}} \times \frac{1 \text{ mol}}{\text{molar mass } CaCO_3 \text{ (g)}}$$

Calculate the solubility product constant by substituting in for s:

$CaCO_3(s) \leftrightarrow Ca^{2+}(aq) + CO_3^{2-}(aq)$

$K_{sp} = [Ca^{2+}][CO_3^{2-}] = s^2$

2. What would be some sources of error?

Not removing all of the solid calcium carbonate when filtering the saturated solution.

Not completely drying the solid calcium carbonate before weighing it.

3. Compare your procedure to that of another group. Was their procedure the same?

Another method may be used such as removing the dissolved calcium ions or carbonate ions by selective precipitation, and then measuring the mass of precipitate formed.

3.3 Review Questions

1.

a) $Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$ $K_{sp} = [Al^{3+}][OH^-]^3$

b) $Cd_3(AsO_4)_2(s) \rightleftharpoons 3Cd^{2+}(aq) + 2AsO_4^{3-}(aq)$ $K_{sp} = [Cd^{2+}]^3[AsO_4^{3-}]^2$

c) $BaMoO_4(s) \rightleftharpoons Ba^{2+}(aq) + MoO_4^{2-}(aq)$ $K_{sp} = [Ba^{2+}][MoO_4^{2-}]$

d $CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$ $K_{sp} = [Ca^{2+}][SO_4^{2-}]$

e) $Pb(IO_3)_2(s) \rightleftharpoons Pb^{2+}(aq) + 2IO_3^-(aq)$ $K_{sp} = [Pb^{2+}][IO_3^-]^2$

f) $Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$ $K_{sp} = [Ag^+]^2[CO_3^{2-}]$

2.

a) $BaSO_3(s) \rightleftharpoons Ba^{2+}(aq) + SO_3^{2-}(aq)$

b) The solubility is the maximum amount of solute dissolved in a given volume of solvent. The solubility product constant is the solubility squared (s^2)

3.

$$ZnCO_3(s) \rightleftharpoons Zn^{2+}(aq) + CO_3^{2-}(aq)$$

1.1×10^{-5} M　　1.1×10^{-5} M

$K_{sp} = [Zn^{2+}][CO_3^{2-}]$

$= (1.1 \times 10^{-5})^2 = 1.2 \times 10^{-10}$

4.

Solubility = $\frac{0.0045 \text{ g}}{0.250 \text{ L}} \times \frac{1 \text{ mol}}{418.58 \text{ g}} = 4.3 \times 10^{-5}$ M

$Ag_3PO_4(s) \rightleftharpoons$	$3Ag^+(aq)$ +	$PO_4^{3-}(aq)$
4.3×10^{-5} M	1.3×10^{-4} M	4.3×10^{-5} M

$K_{sp} = [Ag^+]^3[PO_4^{3-}]$

$= (1.3 \times 10^{-4})^3(4.3 \times 10^{-5} \text{ M})$

$= 9.5 \times 10^{-17}$

5.

$CaSO_4(s) \rightleftharpoons$	$Ca^{2+}(aq)$ +	$SO_4^{2-}(aq)$
	s	s

$K_{sp} = [Ca^{2+}][SO_4^{2-}]$

$9.1 \times 10^{-6} = s^2$

$s = 3.0 \times 10^{-3}$ M

mass gypsum (g) = $0.500 \text{ L} \times \frac{3.0 \times 10^{-3} \text{ mol}}{1 \text{ L}} \times \frac{172.15 \text{ g}}{1 \text{ mol}} = 0.26$ g

6.

Calcite: $CaCO_3(s) \rightleftharpoons$	$Ca^{2+}(aq)$ +	$CO_3^{2-}(aq)$
	s	s

$K_{sp} = [Ca^{2+}][CO_3^{2-}]$

$3.4 \times 10^{-9} = s^2$

$s = 5.8 \times 10^{-5}$ M

g = $\underline{5.8 \times 10^{-5}\ \text{mol}}$ x $\underline{100.09\ \text{g}}$ x $\underline{1\ \text{L}}$ = 5.8×10^{-6} g

mL 1 L 1 mol 1000 mL mL

Argonite: $CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$

s s

$K_{sp} = [Ca^{2+}][CO_3^{2-}]$

$6.0 \times 10^{-9} = s^2$

$s = 7.7 \times 10^{-5}$ M

g = $\underline{7.7 \times 10^{-5}\ \text{mol}}$ x $\underline{100.09\ \text{g}}$ x $\underline{1\ \text{L}}$ = 7.7×10^{-6} g

mL 1 L 1 mol 1000 mL mL7.

Solubility of $Pb_3(AsO_4)_2$:

$\underline{\text{mol}}$ = $\underline{3.0 \times 10^{-5}\ \text{g}}$ x $\underline{1\ \text{mol}}$ = 3.3×10^{-8} M

L L 899.95 g

$Pb_3(AsO_4)_2\ (s) \rightleftharpoons 3Pb^{2+}(aq) + 2AsO_4^{3-}(aq)$

3.3×10^{-8} M 1.0×10^{-7} M 6.7×10^{-8} M

$K_{sp} = [Pb^{2+}]^3[AsO_4^{3-}]^2$

$K_{sp} = (1.0 \times 10^{-7})^3(6.7 \times 10^{-8})^2$

$K_{sp} = 4.4 \times 10^{-36}$

8

$CdCO_3(s) \rightleftharpoons Cd^{2+}(aq) + CO_3^{2-}(aq)$

s s

$K_{sp} = [Cd^{2+}][CO_3^{2-}]$

$1.0 \times 10^{-12} = s^2$

$s = 1.0 \times 10^{-6}$ M

$Cd(OH)_2(s) \rightleftharpoons Cd^{2+}(aq) + 2OH^-(aq)$

s $\quad$ $2\,s$

$K_{sp} = [Cd^{2+}][OH^-]^2$

$7.2 \times 10^{-15} = (s)(2s)^2 = 4s^3$

$s = 1.2 \times 10^{-5}$ M

Disagree. Even though the K_{sp} value for $CdCO_3$ is greater than the K_{sp} for $Cd(OH)_2$, its solubility is less.

9

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

s $\quad$ s

$K_{sp} = [Ag^+][Cl^-]$

$1.8 \times 10^{-10} = s^2$

$s = 1.3 \times 10^{-5}$ M

$Ag_2CrO_4(s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq)$

$2s$ $\quad$ s

$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

$1.1 \times 10^{-12} = (2s)^2(s) = 4s^3$

$s = 6.5 \times 10^{-5}$ M

The solubility of AgCl is less than the solubility of Ag_2CrO_4. All things being equal, the precipitate of AgCl will form first.

10.

$Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$

$2s \qquad s$

$K_{sp} = [Ag^+]^2[CO_3^{2-}]$

$8.5 \times 10^{-12} = (2s)^2(s) = 4s^3$

$s = 1.3 \times 10^{-4}$ M

mass Ag_2CO_3 (g) $= 2.5 \text{ L} \times \frac{1.3 \times 10^{-4} \text{ mol}}{1 \text{ L}} \times \frac{275.75 \text{ g}}{1 \text{ mol}} = 9.0 \times 10^{-2}$ g

3.4 Precipation Formation and the Solubility Product K_{sp}

Warm Up

1. Add solid AgCl until no more solid dissolves. There must be some undissolved solid in the solution. $[Ag^+] = [Cl^-]$.

2. a)

$[AgNO_3] = \frac{0.10 \text{ mol}}{\text{L}} \times \frac{0.025 \text{ L}}{0.040 \text{ L}} = 0.063 \text{ M} = [Ag^+]$

$[NaCl] = \frac{0.085 \text{ mol}}{\text{L}} \times \frac{0.015 \text{ L}}{0.040 \text{ L}} = 0.032 \text{ M} = [Cl^-]$

b) $[Ag^+]$ is not the same as $[Cl^-]$ because they came from different sources. The $[Ag^+]$ came from the $AgNO_3$ whereas the $[Cl^-]$ came from the NaCl.

c) $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Practice Problems: Predicting Whether a Precipitate Will Form

1.

$Pb(NO_3)_2 + 2NaIO_3 \rightarrow Pb(IO_3)_2(s) + 2NaNO_3$

$Pb(IO_3)_2(s) \rightleftharpoons Pb^{2+}(aq) + 2IO_3^-(aq)$

$K_{sp} = [Pb^{2+}][IO_3^-]^2 = 3.7 \times 10^{-13}$

$[Pb(NO_3)_2] = 6.3 \times 10^{-2} \frac{mol}{\cancel{L}} \times \frac{0.0085 \cancel{L}}{1.0085\ L} = 5.3 \times 10^{-4}\ M = [Pb^{2+}]$

$[NaIO_3] = 1.2 \times 10^{-3} \frac{mol}{\cancel{L}} \times \frac{1.0 \cancel{L}}{1.0085\ L} = 1.2 \times 10^{-3}\ M = [IO_3^-]$

$TIP = [Pb^{2+}][IO_3^-]^2$

$= (5.3 \times 10^{-4})(1.2 \times 10^{-3})^2 = 7.6 \times 10^{-10}$

TIP > K_{sp} so a precipitate forms

2.

$NH_4BrO_3 + AgNO_3 \rightarrow AgBrO_3(s) + NH_4NO_3$

$AgBrO_3(s) \rightleftharpoons Ag^+(aq) + BrO_3^-(aq)$

$K_{sp} = [Ag^+][BrO_3^-] = 5.3 \times 10^{-5}$

$[NH_4BrO_3] = 4.5 \times 10^{-3} \frac{mol}{\cancel{L}} \times \frac{0.0015 \cancel{L}}{0.122\ L} = 5.5 \times 10^{-5}\ M = [BrO_3^-]$

$[AgNO_3] = 2.5 \times 10^{-3} \frac{mol}{\cancel{L}} \times \frac{0.1205 \cancel{L}}{0.122\ L} = 2.5 \times 10^{-3}\ M = [Ag^+]$

$TIP = [Ag^+][BrO_3^-]$

$= (2.5 \times 10^{-3})(5.5 \times 10^{-5}) = 1.4 \times 10^{-7}$

TIP < K_{sp} so no precipitate forms

3.

$2NaF + Cd(NO_3)_2 \rightarrow CdF_2(s) + 2NaNO_3$

$CdF_2(s) \rightleftharpoons Cd^{2+}(aq) + 2F^-(aq)$

$K_{sp} = [Cd^{2+}][F^-]^2$

$[NaF] = 0.17 \frac{mol}{\cancel{L}} \times \frac{0.024 \cancel{L}}{0.079 L} = 0.052 M = [F^-]$

$[Cd(NO_3)_2] = 0.22 \frac{mol}{\cancel{L}} \times \frac{0.055 \cancel{L}}{0.079 L} = 0.15 M = [Cd^{2+}]$

TIP = $[Cd^{2+}][F^-]^2 = (0.15)(0.052)^2 = 4.1 \times 10^{-4}$

If no precipitate forms, then TIP < K_{sp} . K_{sp} must be greater than 4.1×10^{-4}.

Practice Problems: Forming a Precipitate In Solution

1.

a)

$SrF_2(s) \rightleftharpoons Sr^{2+}(aq) + 2F^-(aq)$

$K_{sp} = [Sr^{2+}][F^-]^2$

$4.3 \times 10^{-9} = [Sr^{2+}](0.045)^2$

$[Sr^{2+}] = 2.1 \times 10^{-6}$ M

b) 2

$SrCO_3(s) \rightleftharpoons Sr^{2+}(aq) + CO_3^{2-}(aq)$

$K_{sp} = [Sr^{2+}][CO_3^{2-}]$

$5.6 \times 10^{-10} = [Sr^{2+}](2.3 \times 10^{-4})$

$[Sr^{2+}] = 2.4 \times 10^{-6}$ M

c)

$SrSO_4(s) \rightleftharpoons Sr^{2+}(aq) + SO_4^{2-}(aq)$

$K_{sp} = [Sr^{2+}][SO_4^{2-}]$

$3.4 \times 10^{-7} = [Sr^{2+}](0.011)$

$[Sr^{2+}] = 3.1 \times 10^{-5}$ M

2.

$MgCO_3(s) \rightleftharpoons Mg^{2+}(aq) + CO_3^{2-}(aq)$

$K_{sp} = [Mg^{2+}][CO_3^{2-}]$

$6.8 \times 10^{-6} = (3.2 \times 10^{-3})[CO_3^{2-}]$

$[CO_3^{2-}] = 2.1 \times 10^{-3}$ M

$\text{mass } Na_2CO_3 \text{ (g)} = 10.0 \cancel{L} \times \frac{2.1 \times 10^{-3} \text{ mol}}{\cancel{L}} \times \frac{105.99 \text{ g}}{1 \cancel{\text{mol}}} = 2.2 \text{ g}$

3.

$PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq)$

$s \qquad 2s$

$K_{sp} = [Pb^{2+}][I^-]^2$

$8.5 \times 10^{-9} = (s)(2s)^2 = 4s^3$

$s = 1.3 \times 10^{-3}$ M

$[I^-] = 2s = 2.6 \times 10^{-3}$ M

$AgI(s) \rightleftharpoons Ag^+(aq) + I^-(aq)$

$K_{sp} = [Ag^+][I^-]$

$8.5 \times 10^{-17} = [Ag^+](2.6 \times 10^{-3})$

$[Ag^+] = 3.3 \times 10^{-14}$ M

Quick Check:

1.

a) Only changing temperature changes the value of $K_{sp.}$

b) Changing the temperature or the presence of certain ions in the solution can change the solubility of AgCl. For example, if there are Pb^{2+} ions in the solution, the solubility of AgCl would be increased.

2. $Mg(OH)_2$ (*s*) $\rightleftharpoons$ Mg^{2+}(*aq*) + $2OH^-$(*aq*)

To decrease the solubility, we need *more* $Mg(OH)_2$ (*s*), so we need to cause the equilibirum system to shift left. The [Mg^{2+}] or [OH^-] could be increased by adding $Mg(NO_3)_2$ or NaOH to the solution.

3. List 2 substances that would increase the solubility of $Mg(OH)_2$. Use Le Chatelier's Principle and a K_{sp} expression to explain each.

$Mg(OH)_2$ (*s*) $\rightleftharpoons$ Mg^{2+}(*aq*) + $2OH^-$(*aq*)

To increase the solubility, we need more $Mg(OH)_2$ (*s*) to *dissolve*, so we need to cause the equilibirum system to shift right. The [Mg^{2+}] or [OH^-] could be decreased by adding a substance to precipitate one of them out such as $Ca(NO_3)_2$ (to precipitate out the OH^-). Additionally, we could add an acid (such as HCl) to neutralize the OH^- ions.

$K_{sp} = [Mg^{2+}][OH^-]^2$

Since the value of K_{sp} does not change, a decrease in the [Mg^{2+}] (or [OH^-]) causes the equilibrium above to shift right and the concentration of the other ion to increase.

Practice Problems: Calculating Solubility With a Common Ion Present

1

$AgIO_3$(*s*) $\leftrightarrow$	Ag^+(*aq*)	+ IO_3^-(*aq*)
I -	0	0.12
C -	+*x*	+*x*
E -	x	0.12 +*x*

$K_{sp} = [Ag^+][IO_3^-]$

$3.2 \times 10^{-8} = (x)(0.12)$

$x = 2.7 \times 10^{-7}$ M = [Ag^+] = [$AgIO_3$] dissolved

2.

PbI_2(*s*) $\leftrightarrow$	Pb^{2+}(*aq*)	+ $2I^-$(*aq*)
I -	0	0.10

C - $+x$ $+2x$

E - x $0.10 + 2x$

$K_{sp} = [Pb^{2+}][I^-]^2$

$8.5 \times 10^{-9} = (x)(0.10)^2$

$x = 8.5 \times 10^{-7}$ M = $[Pb^{2+}]$ = $[PbI_2]$ dissolved

3.

$BaSO_4(s) \leftrightarrow Ba^{2+}(aq) + SO_4^{2-}(aq)$

I	-	0.050	0
C	-	$+x$	$+x$
E	-	0.050 + x	x

$K_{sp} = [Ba^{2+}][SO_4^{2-}]$

$1.1 \times 10^{-10} = (0.050)(x)$

$x = 2.2 \times 10^{-9}$ M = $[Ba^{2+}]$ = $[BaSO_4]$ dissolved

mass of $BaSO_4$ (g) = $\frac{2.2 \times 10^{-9}\ \cancel{mol}}{L} \times \frac{233.40\ g}{1\ \cancel{mol}} = 5.1 \times 10^{-7}$ g/L

Activity: Experimentally Determining the K_{sp} Of Copper(II) Iodate

Question: What is the approximate value of K_{sp} for copper(II) iodate?

Results and Discussion:

1. Write a balanced formula, complete ionic and net ionic equation for this reaction.

$Cu(NO_3)_2(aq) + 2NaIO_3(aq) \rightarrow 2NaNO_3(aq) + Cu(IO_3)_2(s)$

$Cu^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + 2IO_3^-(aq) \rightarrow + 2Na^+(aq) + 2NO_3^-(aq) + Cu(IO_3)_2(s)$

$Cu^{2+}(aq) + 2IO_3^-(aq) \rightarrow Cu(IO_3)_2(s)$

2. Calculate the $[Cu^{2+}]$ in each of the mixtures.

Mixture 1: $[Cu(NO_3)_2] = \frac{0.010\ mol}{\cancel{L}} \times \frac{0.0100\ \cancel{L}}{0.0200\ L} = 0.0050\ M = [Cu^{2+}]$

Mixture 2: $[Cu(NO_3)_2] = \frac{0.010\ mol}{\cancel{L}} \times \frac{0.0080\ \cancel{L}}{0.0200\ L} = 0.0040\ M = [Cu^{2+}]$

Mixture 3: $[Cu(NO_3)_2] = \frac{0.010\ mol}{\cancel{L}} \times \frac{0.0060\ \cancel{L}}{0.0200\ L} = 0.0030\ M = [Cu^{2+}]$

Mixture 4: $[Cu(NO_3)_2] = \frac{0.010\ mol}{\cancel{L}} \times \frac{0.0040\ \cancel{L}}{0.0200\ L} = 0.0020\ M = [Cu^{2+}]$

Mixture 5: $[Cu(NO_3)_2] = \frac{0.010\ mol}{\cancel{L}} \times \frac{0.0020\ \cancel{L}}{0.0200\ L} = 0.0010\ M = [Cu^{2+}]$

3. Calculate the $[IO_3^-]$ in each of the mixtures.

Mixture 1: $[NaIO_3] = \frac{0.020\ mol}{\cancel{L}} \times \frac{0.0100\ \cancel{L}}{0.0200\ L} = 0.010\ M = [Cu^{2+}]$

Mixture 2: $[NaIO_3] = \frac{0.020\ mol}{\cancel{L}} \times \frac{0.0080\ \cancel{L}}{0.0200\ L} = 0.0080\ M = [Cu^{2+}]$

Mixture 3: $[NaIO_3] = \frac{0.020\ mol}{\cancel{L}} \times \frac{0.0060\ \cancel{L}}{0.0200\ L} = 0.0060\ M = [Cu^{2+}]$

Mixture 4: $[NaIO_3] = \frac{0.020\ mol}{\cancel{L}} \times \frac{0.0040\ \cancel{L}}{0.0200\ L} = 0.0040\ M = [Cu^{2+}]$

Mixture 5: $[NaIO_3] = \frac{0.020\ mol}{\cancel{L}} \times \frac{0.0020\ \cancel{L}}{0.0200\ L} = 0.0020\ M = [Cu^{2+}]$

4. Write the equation for the equilibrium involving the precipitate, and the K_{sp} expression.

$Cu(IO_3)_2(s) \leftrightarrow Cu^{2+}(aq) + 2IO_3^-(aq)$

$K_{sp} = [Cu^{2+}][IO_3^-]^2$

5. Calculate a TIP value for each mixture.

Mixture 1: TIP = $(0.0050)(0.010)^2 = 5.0 \times 10^{-7}$ (ppt)

Mixture 2: TIP = $(0.0040)(0.0080)^2 = 2.6 \times 10^{-7}$ (ppt)

Mixture 3: TIP = $(0.0030)(0.0060)^2 = 1.1 \times 10^{-7}$ (ppt)

Mixture 4: TIP = $(0.0020)(0.0040)^2 = 3.2 \times 10^{-8}$ (no ppt)

Mixture 5: TIP = $(0.0010)(0.0020)^2 = 4.0 \times 10^{-9}$ (no ppt)

6. State the K_{sp} as a range of values from this data.

$3.2 \times 10^{-8} < K_{sp} < 1.1 \times 10^{-7}$

7. Compare your range to the stated K_{sp} value on the K_{sp} table.

$K_{sp} = 6.9 \times 10^{-8}$ which does fit into the range stated in question 6.

Review Questions:

1. a)

$FeS(s) \rightleftharpoons Fe^{2+}(aq) + S^{2-}(aq)$

$K_{sp} = [Fe^{2+}][S^{2}]$

b)

$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$

$K_{sp} = [Mg^{2+}][OH^-]^2$

c)

$Ag_2CrO_4(s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq)$

$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

2. a)

$CaSO_4$

b)

$CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$

$K_{sp} = [Ca^{2+}][SO_4^{2-}]$

c)

Not necessarily. The Ca^{2+} came from the $CaCl_2$ and the SO_4^{2-} came from the H_2SO_4. The concenrations of ions depend on the original concentrations of the $CaCl_2$ and the H_2SO_4.

3.

$CuI_2(aq) \rightarrow Cu^{2+} + 2I^-$

0.015 M 0.015M 0.030 M

$PbI_2(s) \leftarrow > Pb^{2+}(aq) + 2I^-(aq)$

$K_{sp} = [Pb^{2+}][I^-]^2$

$8.5 \times 10^{-9} = (x)(0.030)^2$

$x = 9.4 \times 10^{-6}$ M $= [Pb^{2+}]$

4.

$CaC_2O_4(s) \rightleftharpoons Ca^{2+}(aq) + C_2O_4^{2-}(aq)$

$K_{sp} = [Ca^{2+}][C_2O_4^{2-}]$

$2.3 \times 10^{-9} = (5 \times 10^{-3})[C_2O_4^{2-}]$

$[C_2O_4^{2-}] = 5 \times 10^{-7}$ M

5.

$AgBr(s) \rightleftharpoons Ag^{+}(aq) + Br^-(aq)$

$s \qquad s$

$K_{sp} = [Ag^+][Br^-]$

$5.4 \times 10^{-13} = s^2$

$s = 7.3 \times 10^{-7}$ M $= [Ag^+]$

$Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$

$K_{sp} = [Ag^+]^2 [CO_3^{2-}]$

$8.5 \times 10^{-12} = (7.3 \times 10^{-7})^2 [CO_3^{2-}]$

$[CO_3^{2-}] = 16$ M

6.

$$[AgNO_3] = \frac{0.20 \text{ mol}}{\cancel{\text{L}}} \times \frac{0.0002 \cancel{\text{L}}}{0.1002 \text{ L}} = 4.0 \times 10^{-4} \text{ M} = [Ag^+]$$

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

$K_{sp} = [Ag^+] [Cl^-]$

$1.8 \times 10^{-10} = (4.0 \times 10^{-4}) [Cl^-]$

$[Cl^-] = 4.5 \times 10^{-7}$ M

$$\text{mass NaCl (g)} = 0.100 \cancel{\text{L}} \times \frac{4.5 \times 10^{-7} \cancel{\text{mol}}}{\cancel{\text{L}}} \times \frac{58.44 \text{ g}}{1 \cancel{\text{mol}}} = 2.6 \times 10^{-6} \text{ g}$$

7.

$Sr(NO_3)_2 + ZnSO_4 \rightarrow SrSO_4(s) + Zn(NO_3)_2$

$SrSO_4(s) \rightleftharpoons Sr^{2+}(aq) + SO_4^{2-}(aq)$

$K_{sp} = [Sr^{2+}] [SO_4^{2-}] = 3.4 \times 10^{-7}$

$$[Sr(NO_3)_2] = 0.055 \frac{\text{mol}}{\cancel{\text{L}}} \times \frac{0.0025 \cancel{\text{L}}}{1.5025 \text{ L}} = 9.2 \times 10^{-5} \text{ M} = [Sr^{2+}]$$

$[ZnSO_4] = \frac{0.011 \text{ mol}}{\text{L}} \times \frac{1.5 \text{ L}}{1.5025 \text{ L}} = 1.1 \times 10^{-2} \text{ M} = [SO_4^{2-}]$

TIP = $[Sr^{2+}]\,[SO_4^{2-}]$

$= (9.2 \times 10^{-5})(1.1 \times 10^{-2}) = 1.0 \times 10^{-6}$

TIP > K_{sp} so precipitate forms

8. $Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2(s) + 2NaNO_3$

$PbCl_2\,(s) \rightleftharpoons Pb^{2+}\,(aq) + 2Cl^-(aq)$

$K_{sp} = [Pb^{2+}]\,[Cl^-]^2 = 1.2 \times 10^{-5}$

$[Pb(NO_3)_2] = \frac{0.068 \text{ g}}{2.0 \text{ L}} \times \frac{1 \text{ mol}}{331.39 \text{ g}} = 1.0 \times 10^{-4} \text{ M} = [Pb^{2+}]$

$[NaCl] = 0.080 \text{ M} = [Cl^-]$

TIP = $[Pb^{2+}]\,[Cl^-]^2$

$= (1.0 \times 10^{-4})(0.080) = 8.2 \times 10^{-6}$

TIP < K_{sp} so no precipitate forms

9. a)

$AgCl(s) \rightleftharpoons Ag^+\,(aq) + Cl^-(aq)$ $\quad K_{sp} = [Ag^+]\,[Cl^-] = 1.8 \times 10^{-10}$

$AgI(s) \rightleftharpoons Ag^+\,(aq) + I^-(aq)$ $\quad K_{sp} = [Ag^+]\,[I^-] = 8.5 \times 10^{-17}$

b) See above

c) AgI

d)

$AgI(s) \rightleftharpoons Ag^{+}(aq) + I^{-}(aq)$

$K_{sp} = [Ag^{+}][I^{-}]$

$8.5 \times 10^{-17} = [Ag^{+}](0.020)$

$[Ag^{+}] = 4.3 \times 10^{-15}$ M

e)

$AgCl(s) \rightleftharpoons Ag^{+}(aq) + Cl^{-}(aq)$

$K_{sp} = [Ag^{+}][Cl^{-}]$

$1.8 \times 10^{-10} = [Ag^{+}](0.020)$

$[Ag^{+}] = 9.0 \times 10^{-9}$ M

f)

$4.3 \times 10^{-15}\text{ M} < [Ag^{+}] < 9.0 \times 10^{-9}\text{ M}$

g)

$AgI(s) \rightleftharpoons Ag^{+}(aq) + I^{-}(aq)$

$K_{sp} = [Ag^{+}][I^{-}]$

$8.5 \times 10^{-17} = (9.0 \times 10^{-9})[I^{-}]$

$[I^{-}] = 9.4 \times 10^{-9}$ M

h)

$$\%\ I^{-}\text{ remaining} = \frac{9.4 \times 10^{-9}\text{ M}}{0.020\text{ M}} \times 100\% = 0.00047\ \%$$

so 100.00000 – 0.00047 = 99.999953 % precipitated out.

10.

$MgCO_3(s) \rightleftharpoons Mg^{2+}(aq) + CO_3^{2-}(aq)$

$$[Mg^{2+}] = \frac{12\ \cancel{mg}}{1.0\text{ L}} \times \frac{1\ \cancel{g}}{1000\ \cancel{mg}} \times \frac{1\text{ mol}}{24.31\ \cancel{g}} = 4.9 \times 10^{-4}\text{ M}$$

K_{sp} = $[Mg^{2+}]\,[CO_3^{2-}]$

$6.8 \times 10^{-6} = (4.9 \times 10^{-4})\,[CO_3^{2-}]$

$[CO_3^{2-}] = 1.4 \times 10^{-2}$ M

mass $Na_2CO_3 \cdot 10H_2O$ (g) = 1.0 ~~L~~ x <u>1.4 x 10^{-2} ~~mol~~</u> x <u>286.19 g</u> = 4.0 g

~~L~~ 1 ~~mol~~

11.

$PbI_2(s)$ $\leftarrow$> $Pb^{2+}(aq)$ + $2I^-(aq)$

To decrease the solubility, the equilibrium must be shifted left. By adding $Pb(NO_3)_2$ or KI, the common ions (Pb^{2+} or I^-) cause an increase in their respective concentrations.

12.

$BaSO_4(s)$ $\leftarrow$> $Ba^{2+}(aq)$ + $SO_4^{2-}(aq)$

In Na_2SO_4, the common ion SO_4^{2-} exists, so equilibrium shifts left and solubility is decreased. In water, there are no Ba^{2+} or SO_4^{2-} ions.

13.

$Fe(OH)_3(s)$ $\leftarrow$> $Fe^{3+}(aq)$ + $3OH^-(aq)$

In HCl, the H^+ ions will neutralize the OH^- ions, causing the $[OH^-]$ to decrease. The equilibrium will shift right and the solubility of $Fe(OH)_3$ will increase. It is more soluble in HCl than in water.

Extension

14. a)

$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$

s s

K_{sp} = $[Ba^{2+}]\,[SO_4^{2-}]$

$1.1 \times 10^{-10} = s^2$

$s = 1.0 \times 10^{-5}$ M

mass $BaSO_4$ (g) = 0.200 ~~L~~ x <u>1.0 x 10^{-5} ~~mol~~</u> x <u>233.40 g</u> = 4.7 x 10^{-4} g

L 1 ~~mol~~

b)

$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$

I		0	0.10
C		+x	+x
E		x	0.10 + x

$K_{sp} = [Ba^{2+}][SO_4^{2-}]$

$1.1 \times 10^{-10} = (x)(0.10)$

$x = 1.1 \times 10^{-9}\ M = [Ba^{2+}]$

$$\text{mass } BaSO_4\ (g) = 0.200\ \cancel{L} \times \frac{1.1 \times 10^{-9}\ \cancel{mol}}{\cancel{L}} \times \frac{233.40\ g}{1\ \cancel{mol}} = 5.1 \times 10^{-8}\ g$$