

## 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

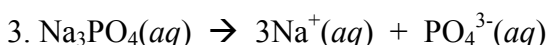
c) ionic

d) molecular

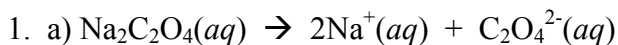
e) ionic

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.

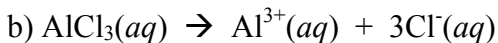
##### Quick Check



### Quick Check



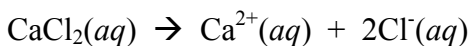
0.45 M                      0.90 M              0.45 M              total [ions] = 0.90 M + 0.45 M = 1.35 M



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

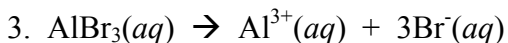


1.0 M              1.0 M              1.0 M              total ion concentration = 2.0 M



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

The  $\text{CaCl}_2$  will have the greater conductivity because it has a greater ion concentration.



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

### Practice Problems: Calculating Ion Concentrations

1. a)  $[(\text{NH}_4)_3\text{PO}_4] = \frac{0.35 \text{ g}}{149.12 \text{ g}} \times \frac{1 \text{ mol}}{1 \text{ mol}} = 3.6 \times 10^{-3} \text{ M}$

$[\text{NH}_4^+] = 3(3.6 \times 10^{-3}) = 1.1 \times 10^{-2} \text{ M}$        $[\text{PO}_4^{3-}] = 3.6 \times 10^{-3} \text{ M}$

b)  $[\text{HCl}] = \frac{6.0 \text{ mol}}{0.015 \text{ L}} = 2.6 \text{ M}$

$[\text{H}^+] = 2.6 \text{ M}$        $[\text{Cl}^-] = 2.6 \text{ M}$



0.15 M                                      0.45 M  
mass of  $\text{Al}_2(\text{SO}_4)_3 = 0.250 \text{ L} \times \frac{0.15 \text{ mol}}{1 \text{ L}} \times \frac{342.17 \text{ g}}{1 \text{ mol}} = 13 \text{ g}$

$$3. \text{[(NH}_4\text{)}_2\text{SO}_4] = \frac{0.22 \text{ mol}}{1 \text{ L}} \times \frac{0.055 \text{ L}}{0.080 \text{ L}} = 0.15 \text{ M} \quad [\text{NH}_4^+] = 0.30 \text{ M}, [\text{SO}_4^{2-}] = 0.15 \text{ M}$$

$$[\text{(NH}_4\text{)}_2\text{S}] = \frac{0.45 \text{ mol}}{1 \text{ L}} \times \frac{0.025 \text{ L}}{0.080 \text{ L}} = 0.14 \text{ M} \quad [\text{NH}_4^+] = 0.28 \text{ M}, [\text{S}^{2-}] = 0.14 \text{ M}$$

$$\text{total } [\text{NH}_4^+] = 0.58 \text{ M}, [\text{SO}_4^{2-}] = 0.15 \text{ M} \quad [\text{S}^{2-}] = 0.14 \text{ M}$$

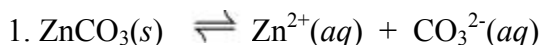
### Practice Problems: Converting Between Units of Solubility

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

$$2. \text{Mass Pb(IO}_3\text{)}_2 = 0.300 \text{ L} \times \frac{4.5 \times 10^{-5} \text{ mol}}{1 \text{ L}} \times \frac{557.17 \text{ g}}{1 \text{ mol}} = 7.5 \times 10^{-3} \text{ g}$$

$$3. \frac{\text{mol}}{\text{L}} \text{ CaCO}_3 = \frac{7.1 \times 10^{-4} \text{ g}}{0.100 \text{ L}} \times \frac{1 \text{ mol}}{100.09 \text{ g}} = 7.1 \times 10^{-5} \text{ M}$$

### Quick Check

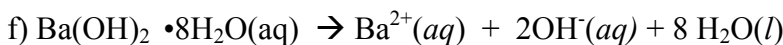
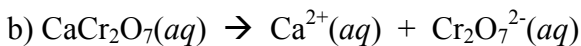
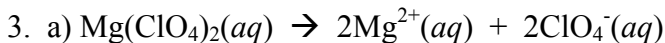


### 3.1 Review

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<sub>4</sub>, SO<sub>2</sub>, XeF<sub>6</sub>

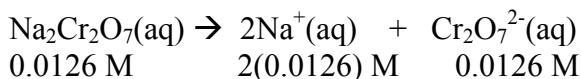
2. Both are acids with the same concentration, but HClO<sub>4</sub> is a strong acid that ionizes completely. H<sub>3</sub>PO<sub>4</sub> is a weak acid, and remains largely in molecular form. Since there are more ions in the HClO<sub>4</sub>, it will have a greater electrical conductivity.



$$4. \frac{\text{mol BaSO}_4}{\text{L}} = \frac{0.0012 \text{ g}}{0.500 \text{ L}} \times \frac{1 \text{ mol}}{233.40 \text{ g}} = 1.0 \times 10^{-5} \text{ M}$$

$$5. \text{mass CaCO}_3 = 0.250 \text{ L} \times \frac{7.1 \times 10^{-5} \text{ mol}}{1 \text{ L}} \times \frac{100.09 \text{ g}}{1 \text{ mol}} = 1.8 \times 10^{-3} \text{ g}$$

$$6. [\text{Na}_2\text{Cr}_2\text{O}_7] = \frac{0.50 \text{ g}}{0.150 \text{ L}} \times \frac{1 \text{ mol}}{261.98 \text{ g}} = 0.0126 \text{ M}$$



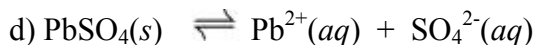
$$[\text{Na}^+] = 0.025 \text{ M} \quad [\text{Cr}_2\text{O}_7^{2-}] = 0.013 \text{ M}$$

$$7. [\text{NaCl}] = \frac{5.0 \text{ mol}}{1 \text{ L}} \times \frac{90. \text{ L}}{30. \text{ L}} = 3.3 \text{ M} \quad [\text{Na}^+] = 3.3 \text{ M} \quad [\text{Cl}^-] = 3.3 \text{ M}$$

$$[\text{MgCl}_2] = \frac{2.4 \text{ mol}}{1 \text{ L}} \times \frac{90. \text{ L}}{30. \text{ L}} = 0.80 \text{ M} \quad [\text{Mg}^{2+}] = 0.80 \text{ M} \quad [\text{Cl}^-] = 1.6 \text{ M}$$

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

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.



## 3.2 Qualitative Analysis – Identifying Unknown Ions

### Warm Up

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

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

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)  $\text{Na}_2\text{CO}_3$

b)  $\text{SrSO}_4$

c)  $\text{Ag}^+$

d)  $\text{NO}_3^-$

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  $\text{Mg}(\text{OH})_2$  and  $\text{Sr}(\text{OH})_2$ . According to the Solubility Table,  $\text{Mg}(\text{OH})_2$  has low solubility and  $\text{Sr}(\text{OH})_2$  is soluble. The sample must contain strontium nitrate since no precipitate was observed.

### Quick Check

1.

a)  $\text{BaSO}_4$

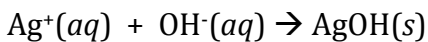
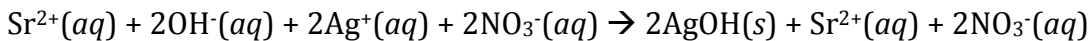
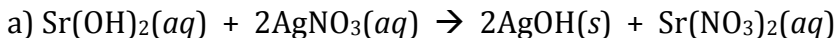
b)  $\text{Fe}(\text{OH})_2$

c)  $\text{Zn}_3(\text{PO}_4)_2$

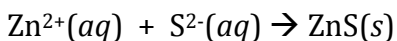
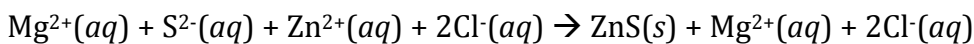
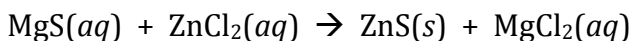
d)  $\text{CrCO}_3$

e)  $\text{Mn(OH)}_2$

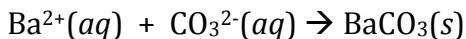
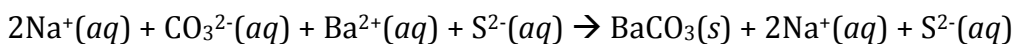
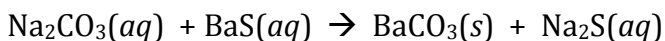
2.



b)



c)



### Practice Problems: Selective Precipitation

(other answers possible)

1.

First add  $\text{NaCl}$ . Filter out the  $\text{AgCl}(\text{s})$

Then add  $\text{Na}_2\text{S}$ . Filter out the  $\text{BeS}(\text{s})$

Finally add  $\text{Na}_2\text{SO}_4$ . Filter out the  $\text{BaSO}_4(\text{s})$

2.

First add  $\text{Ca(NO}_3)_2$ . Filter out the  $\text{CaSO}_4(\text{s})$

Then add  $\text{Cu(NO}_3)_2$ . Filter out the  $\text{CuS}(\text{s})$

Finally add  $\text{AgNO}_3$ . Filter out the  $\text{AgBr}(\text{s})$

3.

First add  $\text{Sr}(\text{NO}_3)_2$ . Filter out the  $\text{Sr}_3(\text{PO}_4)_2(s)$

Then add  $\text{Mg}(\text{NO}_3)_2$ . Filter out the  $\text{Mg}(\text{OH})_2(s)$

Finally add  $\text{AgNO}_3$ . Filter out the  $\text{Ag}_2\text{S}(s)$

### Quick Check

1.

- The color of the solution (for example, solutions containing  $\text{Cu}^{2+}$  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  $\text{NH}_3$  or  $\text{HNO}_3$ .

2. First add  $\text{NaCl}$ . A precipitate of both  $\text{PbCl}_2$  and  $\text{AgCl}$  may form. Filter out the precipitate, dry it and measure its mass. Put the precipitate into a second beaker. To the original solution, add  $\text{NaOH}$ . If a precipitate forms, filter out the  $\text{Cu}(\text{OH})_2(s)$ . To the second beaker, add 6 M  $\text{NH}_3$ . Any  $\text{AgCl}$  re-dissolves. Filter out the  $\text{PbCl}_2(s)$ . Dry the  $\text{PbCl}_2(s)$  and measure its mass. If the mass had decreased, the presence of  $\text{AgCl}$  is confirmed.

3. First add  $\text{HNO}_3$ . The formation of bubbles confirms the presence of  $\text{CO}_3^{2-}$  ions. Next, add  $\text{CaCl}_2$ . If a precipitate forms, it is  $\text{Ca}_3(\text{PO}_4)_2(s)$ . Filter out the precipitate.

#### Activity: Using Titration to Calculate the Unknown $[\text{Cl}^-]$ In A Sample of Seawater

**Question:** What is the concentration of  $\text{Cl}^-$  in seawater?

**Procedure:**

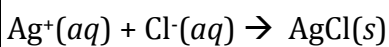
4. Procedures 1-3 were repeated in 2 more trials, and the data below was recorded.

	Trial #1	Trial #2	Trial #3
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<b>Initial burette reading (mL)</b>	0.00	14.70	28.60
<b>Final burette reading (mL)</b>	14.70	28.60	42.35
<b>Volume AgNO<sub>3</sub> added (mL)</b>	14.70	13.90	13.75
<b>Average volume AgNO<sub>3</sub> used (mL)</b>	13.83	X	X

### Results and Discussion

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



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

$$\text{mol AgNO}_3 = 0.01383 \text{ L} \times \frac{0.100 \text{ mol}}{1 \text{ L}} = 0.00138 \text{ mol AgNO}_3$$

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

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

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.

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

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  $[\text{Cl}^-] = 0.552 \text{ M}$ .

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  $\text{Cl}^-$  to g/L:

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

$$\text{L} \quad \text{L} \quad 1 \text{ mol}$$

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  $\text{AgCl}$  and  $\text{Ag}_2\text{CrO}_4$ . Calculate the concentration of  $\text{Ag}^+$  in a saturated solution of each. Use this information to explain why  $\text{AgCl}$  precipitates before  $\text{Ag}_2\text{CrO}_4$ .

The solubility of  $\text{AgCl} = 1.3 \times 10^{-5} \text{ M}$

The solubility of  $\text{Ag}_2\text{CrO}_4 = 6.5 \times 10^{-5} \text{ M}$

When silver ions are added in the titration,  $\text{AgCl}$  will precipitate first because it has a lower solubility than  $\text{Ag}_2\text{CrO}_4$ .

### 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.

$$3. \text{ mol AgCH}_3\text{COO} = \frac{11.1 \text{ g}}{\text{L}} \times \frac{1 \text{ mol}}{166.92 \text{ g}} = 0.0665 \text{ M}$$

0.0665 M < 0.1 M so AgCH<sub>3</sub>COO would have a low solubility.

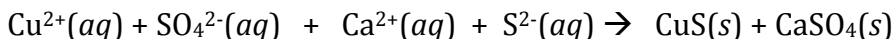
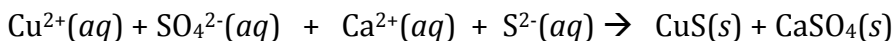
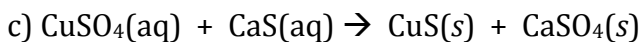
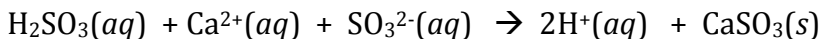
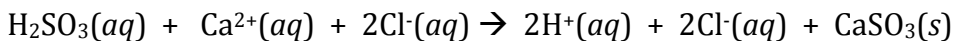
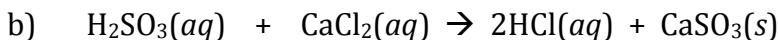
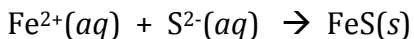
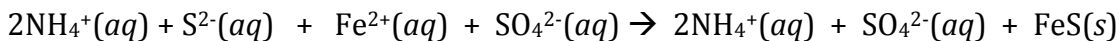
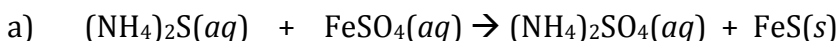
4. A formula equation is written with substances as compounds and states shown.

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<sup>+</sup> ions and NO<sub>3</sub><sup>-</sup> ions are often spectators.

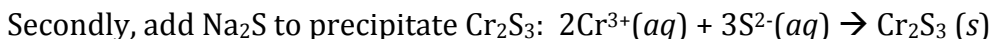
6.



7. Na<sup>+</sup> and K<sup>+</sup> are alkali ions. Compounds containing alkali ions are soluble in water.



Filter out the precipitate



Filter out the precipitate

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

Filter out the precipitate

9. First add  $\text{Mg(NO}_3)_2$  to precipitate  $\text{Mg}_3(\text{PO}_4)_2$ :  $3\text{Mg}^{2+}(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq}) \rightarrow \text{Mg}_3(\text{PO}_4)_2(\text{s})$

Filter out the precipitate

Secondly, add  $\text{Fe(NO}_3)_2$  to precipitate  $\text{FeS}$ :  $\text{Fe}^{2+}(\text{aq}) + \text{S}^{2-}(\text{aq}) \rightarrow \text{FeS}(\text{s})$

Filter out the precipitate

Finally add  $\text{AgNO}_3$  to precipitate  $\text{AgCl}$ :  $\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(\text{s})$

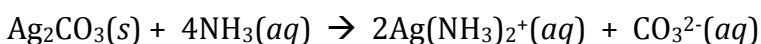
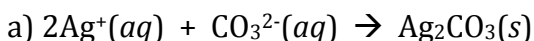
Filter out the precipitate

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.  $\text{AgCl}$  can be dissolved in 6M  $\text{NH}_3$ , while  $\text{AgI}$  cannot. This can be used to separate  $\text{AgCl}$  from  $\text{AgI}$  or  $\text{AgBr}$ .

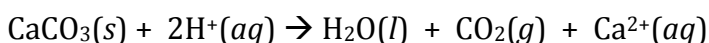
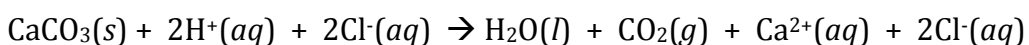
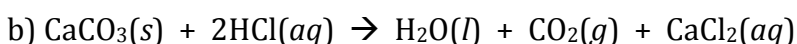
Carbonates will dissolve on the addition of a strong acid such as  $\text{HNO}_3$ .

12.



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

Scale is  $\text{CaCO}_3$  that collects inside water heaters, pipes, kettles and boilers.



c) Many answers are possible:  $\text{Na}_2\text{CO}_3$

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

### 3.3 The Solubility Product Constant - $K_{sp}$

## Warm Up

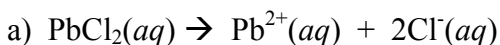
1.

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

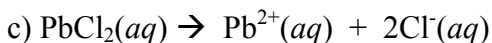
b) the answer when numbers are multiplied together

c) a number that does not change

2.

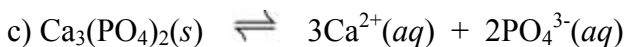
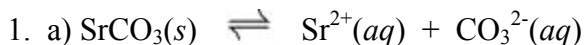


b)  $\frac{\text{mol}}{\text{L}} = \frac{4.4 \text{ g}}{\text{L}} \times \frac{1 \text{ mol}}{278.27 \text{ g}} = 0.016 \text{ M}$



0.016 M      0.016 M      0.032 M

## Quick Check



2. a)  $K_{sp} = [\text{Sr}^{2+}][\text{CO}_3^{2-}]$

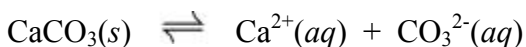
b)  $K_{sp} = [\text{Mg}^{2+}][\text{OH}^{-}]^2$

c)  $K_{sp} = [\text{Ca}^{2+}]^3[\text{PO}_4^{3-}]^2$

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)



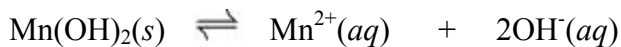
$6.1 \times 10^{-5} \text{ M}$        $6.1 \times 10^{-5} \text{ M}$        $6.1 \times 10^{-5} \text{ M}$

$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

$$K_{sp} = (6.1 \times 10^{-5})(6.1 \times 10^{-5})$$

$$K_{sp} = 3.7 \times 10^{-9}$$

b)



$$3.6 \times 10^{-5} \text{ M} \quad 3.6 \times 10^{-5} \text{ M} \quad 7.2 \times 10^{-5} \text{ M}$$

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

$$K_{sp} = (3.6 \times 10^{-5})(7.2 \times 10^{-5})^2$$

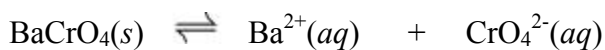
$$K_{sp} = 1.9 \times 10^{-13}$$

c)

Solubility of  $\text{BaCrO}_4$  :

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

$$\text{L} \quad \text{L} \quad 253.33 \text{ g}$$



$$1.1 \times 10^{-5} \text{ M} \quad 1.1 \times 10^{-5} \text{ M} \quad 1.1 \times 10^{-5} \text{ M}$$

$$K_{sp} = [\text{Ba}^{2+}][\text{CrO}_4^{2-}]$$

$$K_{sp} = (1.1 \times 10^{-5})(1.1 \times 10^{-5})$$

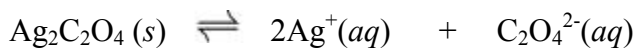
$$K_{sp} = 1.2 \times 10^{-10}$$

d)

Solubility of  $\text{Ag}_2\text{C}_2\text{O}_4$  :

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

$$\text{L} \quad \text{L} \quad 303.76 \text{ g}$$



$$1.1 \times 10^{-4} \text{ M} \quad 2.2 \times 10^{-4} \text{ M} \quad 1.1 \times 10^{-4} \text{ M}$$

$$K_{sp} = [\text{Ag}^{+}]^2[\text{C}_2\text{O}_4^{2-}]$$

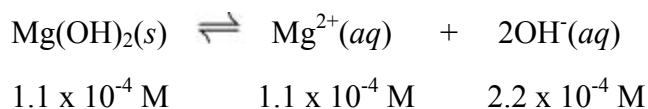
$$K_{sp} = (2.2 \times 10^{-4})^2(1.1 \times 10^{-4})$$

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

2.

Solubility of  $\text{Mg}(\text{OH})_2$  :

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



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

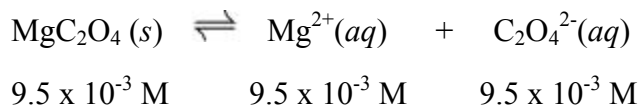
$$K_{sp} = (1.1 \times 10^{-4})(2.2 \times 10^{-4})^2$$

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

3.

Solubility of  $\text{MgC}_2\text{O}_4$  :

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



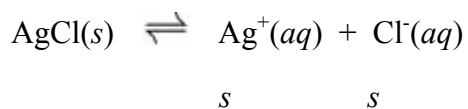
$$K_{sp} = [\text{Mg}^{2+}][\text{C}_2\text{O}_4^{2-}]$$

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

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

### Practice Problems:

1. a)

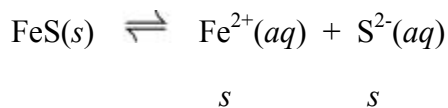


$$K_{sp} = [\text{Ag}^{+}][\text{Cl}^{-}]$$

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

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

b)



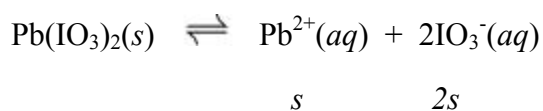
$$K_{sp} = [\text{Fe}^{2+}][\text{S}^{2-}]$$

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

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

$$g = \frac{7.7 \times 10^{-10} \text{ mol}}{1 \text{ L}} \times \frac{87.92 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 6.8 \times 10^{-11} \text{ g/mL}$$

c)



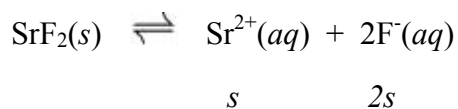
$$K_{sp} = [\text{Pb}^{2+}][\text{IO}_3^-]^2$$

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

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

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

d)



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

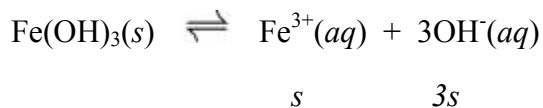
$$4.3 \times 10^{-9} = (s)(2s)^2 = 4s^3$$

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

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

$$g = \frac{1.0 \times 10^{-3} \text{ mol}}{1 \text{ L}} \times \frac{125.62 \text{ g}}{1 \text{ mol}} = 6.8 \times 10^{-11} \text{ g/L}$$

2.



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

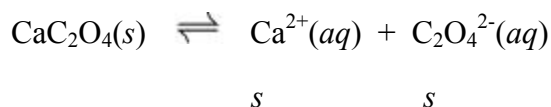
$$2.6 \times 10^{-39} = (s)(3s)^3 = 27s^4$$

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

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

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

3.



$$K_{sp} = [\text{Ca}^{2+}][\text{C}_2\text{O}_4^{2-}]$$

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

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

$$\text{mass (g) CaC}_2\text{O}_4 = 0.650 \text{ L} \times \frac{4.8 \times 10^{-5} \text{ mol}}{1 \text{ L}} \times \frac{128.10 \text{ g}}{1 \text{ mol}} = 4.0 \times 10^{-3} \text{ g}$$

**Activity: Experimentally Determining the  $K_{sp}$  of  $\text{CaCO}_3$**

**Question:** What is the  $K_{sp}$  of  $\text{CaCO}_3$ ?

**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  $\text{CaCO}_3$  by measuring out 150 mL of distilled water into the first beaker.
2. Continue to add solid  $\text{CaCO}_3$  and stir until no more  $\text{CaCO}_3$  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  $\text{CaCO}_3$  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  $\text{CaCO}_3$  to cool.
9. Measure and record the mass of the beaker and solid  $\text{CaCO}_3$ .

### Data Table:

Temperature of saturated solution:

Mass of beaker:

Mass of beaker and  $\text{CaCO}_3(s)$ :

### Results and Discussion:

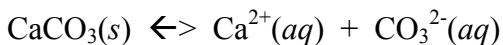
1. Describe how you would analyze the data and perform calculations to determine the  $K_{sp}$ .

Calculate the mass of  $\text{CaCO}_3$  dissolved in 100. mL of saturated solution by subtracting the Mass of beaker from the Mass of beaker and  $\text{CaCO}_3(s)$ .

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  $s$ :



$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}] = s^2$$

2. What would be some sources of error?

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

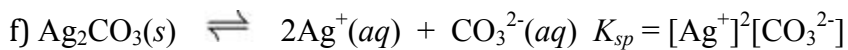
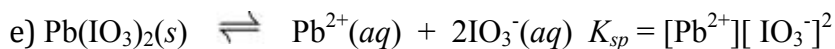
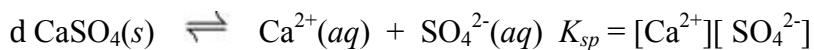
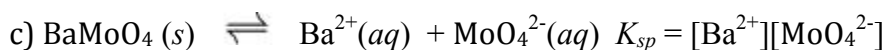
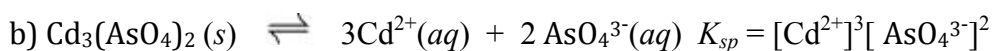
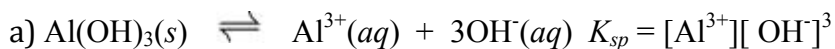
Not completely drying the solid calcium carbonate before weighing it.

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.

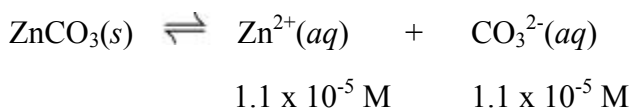


2.



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

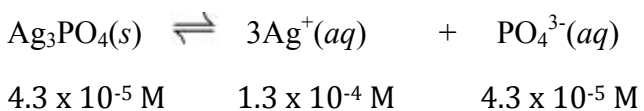
3.



$$\begin{aligned} K_{sp} &= [\text{Zn}^{2+}][\text{CO}_3^{2-}] \\ &= (1.1 \times 10^{-5})^2 = 1.2 \times 10^{-10} \end{aligned}$$

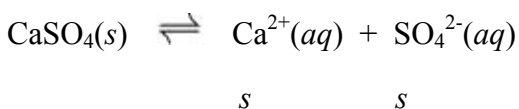
4.

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



$$K_{sp} = [\text{Ag}^+]^3[\text{PO}_4^{3-}]$$
$$= (1.3 \times 10^{-4})^3(4.3 \times 10^{-5} \text{ M})$$
$$= 9.5 \times 10^{-17}$$

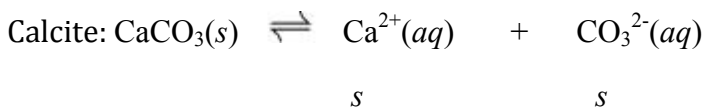
5.



$$K_{sp} = [\text{Ca}^{2+}][\text{SO}_4^{2-}]$$
$$9.1 \times 10^{-6} = s^2$$
$$s = 3.0 \times 10^{-3} \text{ M}$$

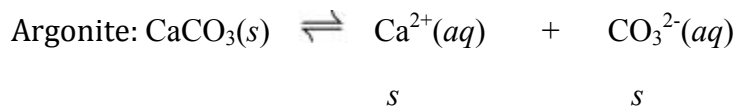
$$\text{mass gypsum (g)} = 0.500 \text{ L} \times \frac{3.0 \times 10^{-3} \text{ mol}}{1 \text{ L}} \times \frac{172.15 \text{ g}}{1 \text{ mol}} = 0.26 \text{ g}$$

6.



$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$
$$3.4 \times 10^{-9} = s^2$$
$$s = 5.8 \times 10^{-5} \text{ M}$$

$$\frac{\text{g}}{\text{mL}} = \frac{5.8 \times 10^{-5} \text{ mol}}{1 \text{ L}} \times \frac{100.09 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 5.8 \times 10^{-6} \frac{\text{g}}{\text{mL}}$$



$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

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

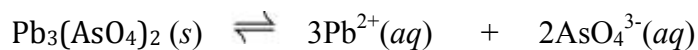
$$s = 7.7 \times 10^{-5} \text{ M}$$

$$\frac{\text{g}}{\text{mL}} = \frac{7.7 \times 10^{-5} \text{ mol}}{1 \text{ L}} \times \frac{100.09 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 7.7 \times 10^{-6} \frac{\text{g}}{\text{mL}}$$

Solubility of  $\text{Pb}_3(\text{AsO}_4)_2$  :

$$\frac{\text{mol}}{\text{L}} = \frac{3.0 \times 10^{-5} \text{ g}}{\text{L}} \times \frac{1 \text{ mol}}{899.95 \text{ g}} = 3.3 \times 10^{-8} \text{ M}$$

$$\text{L} \qquad \text{L} \qquad 899.95 \text{ g}$$



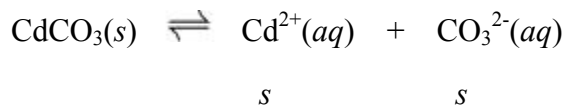
$$3.3 \times 10^{-8} \text{ M} \qquad 1.0 \times 10^{-7} \text{ M} \qquad 6.7 \times 10^{-8} \text{ M}$$

$$K_{sp} = [\text{Pb}^{2+}]^3[\text{AsO}_4^{3-}]^2$$

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

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

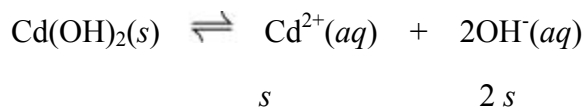
8



$$K_{sp} = [\text{Cd}^{2+}][\text{CO}_3^{2-}]$$

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

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



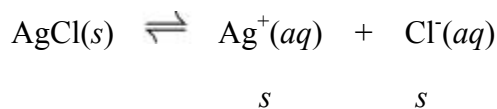
$$K_{sp} = [\text{Cd}^{2+}][\text{OH}^-]^2$$

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

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

Disagree. Even though the  $K_{sp}$  value for  $\text{CdCO}_3$  is greater than the  $K_{sp}$  for  $\text{Cd(OH)}_2$ , its solubility is less.

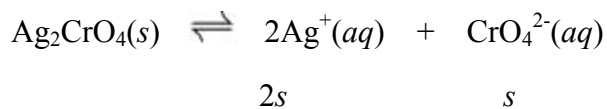
9



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

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

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



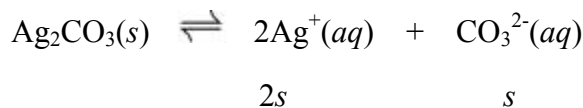
$$K_{sp} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}]$$

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

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

The solubility of  $\text{AgCl}$  is less than the solubility of  $\text{Ag}_2\text{CrO}_4$ . All things being equal, the precipitate of  $\text{AgCl}$  will form first.

10.



$$K_{sp} = [\text{Ag}^+]^2[\text{CO}_3^{2-}]$$

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

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

$$\text{mass Ag}_2\text{CO}_3 (g) = 2.5 \text{ L} \times \frac{1.3 \times 10^{-4} \text{ mol}}{1 \text{ L}} \times \frac{275.75 \text{ g}}{1 \text{ mol}} = 9.0 \times 10^{-2} \text{ g}$$

### 3.4 Precipitation Formation and the Solubility Product $K_{sp}$

#### Warm Up

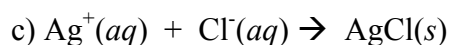
1. Add solid AgCl until no more solid dissolves. There must be some undissolved solid in the solution.  $[\text{Ag}^+] = [\text{Cl}^-]$ .

2. a)

$$[\text{AgNO}_3] = \frac{0.10 \text{ mol}}{1 \text{ L}} \times \frac{0.025 \text{ L}}{0.040 \text{ L}} = 0.063 \text{ M} = [\text{Ag}^+]$$

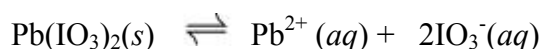
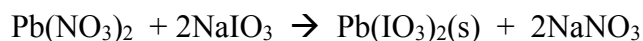
$$[\text{NaCl}] = \frac{0.085 \text{ mol}}{1 \text{ L}} \times \frac{0.015 \text{ L}}{0.040 \text{ L}} = 0.032 \text{ M} = [\text{Cl}^-]$$

b)  $[\text{Ag}^+]$  is not the same as  $[\text{Cl}^-]$  because they came from different sources. The  $[\text{Ag}^+]$  came from the  $\text{AgNO}_3$  whereas the  $[\text{Cl}^-]$  came from the  $\text{NaCl}$ .



#### Practice Problems: Predicting Whether a Precipitate Will Form

1.



$$K_{sp} = [\text{Pb}^{2+}] [\text{IO}_3^-]^2 = 3.7 \times 10^{-13}$$

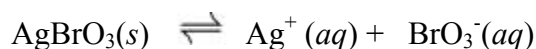
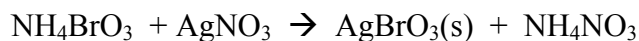
$$[\text{Pb}(\text{NO}_3)_2] = 6.3 \times 10^{-2} \frac{\text{mol}}{\text{L}} \times \frac{0.0085 \text{ L}}{1.0085 \text{ L}} = 5.3 \times 10^{-4} \text{ M} = [\text{Pb}^{2+}]$$

$$[\text{NaIO}_3] = 1.2 \times 10^{-3} \frac{\text{mol}}{\text{L}} \times \frac{1.0 \text{ L}}{1.0085 \text{ L}} = 1.2 \times 10^{-3} \text{ M} = [\text{IO}_3^-]$$

$$\begin{aligned} \text{TIP} &= [\text{Pb}^{2+}] [\text{IO}_3^-]^2 \\ &= (5.3 \times 10^{-4})(1.2 \times 10^{-3})^2 = 7.6 \times 10^{-10} \end{aligned}$$

TIP >  $K_{sp}$  so a precipitate forms

2.



$$K_{sp} = [\text{Ag}^+] [\text{BrO}_3^-] = 5.3 \times 10^{-5}$$

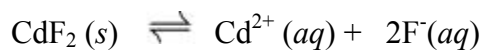
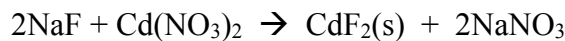
$$[\text{NH}_4\text{BrO}_3] = 4.5 \times 10^{-3} \frac{\text{mol}}{\text{L}} \times \frac{0.0015 \text{ L}}{0.122 \text{ L}} = 5.5 \times 10^{-5} \text{ M} = [\text{BrO}_3^-]$$

$$[\text{AgNO}_3] = 2.5 \times 10^{-3} \frac{\text{mol}}{\text{L}} \times \frac{0.1205 \text{ L}}{0.122 \text{ L}} = 2.5 \times 10^{-3} \text{ M} = [\text{Ag}^+]$$

$$\begin{aligned} \text{TIP} &= [\text{Ag}^+] [\text{BrO}_3^-] \\ &= (2.5 \times 10^{-3})(5.5 \times 10^{-5}) = 1.4 \times 10^{-7} \end{aligned}$$

TIP <  $K_{sp}$  so no precipitate forms

3.



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



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

$$[\text{Cd}(\text{NO}_3)_2] = 0.22 \frac{\text{mol}}{\text{L}} \times \frac{0.055 \text{ L}}{0.079 \text{ L}} = 0.15 \text{ M} = [\text{Cd}^{2+}]$$

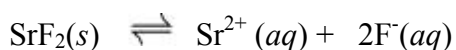
$$\text{TIP} = [\text{Cd}^{2+}] [\text{F}^-]^2 = (0.15)(0.052)^2 = 4.1 \times 10^{-4}$$

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

### Practice Problems: Forming a Precipitate In Solution

1.

a)

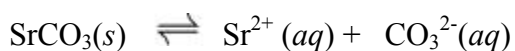


$$K_{sp} = [\text{Sr}^{2+}] [\text{F}^-]^2$$

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

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

b) 2

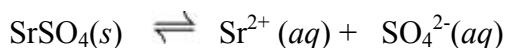


$$K_{sp} = [\text{Sr}^{2+}] [\text{CO}_3^{2-}]$$

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

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

c)

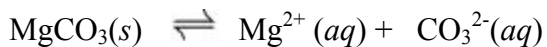


$$K_{sp} = [\text{Sr}^{2+}] [\text{SO}_4^{2-}]$$

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

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

2.



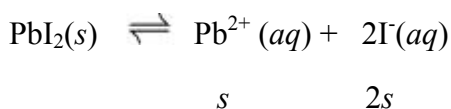
$$K_{sp} = [\text{Mg}^{2+}][\text{CO}_3^{2-}]$$

$$6.8 \times 10^{-6} = (3.2 \times 10^{-3})[\text{CO}_3^{2-}]$$

$$[\text{CO}_3^{2-}] = 2.1 \times 10^{-3} \text{ M}$$

$$\text{mass Na}_2\text{CO}_3 \text{ (g)} = 10.0 \cancel{\text{ L}} \times \frac{2.1 \times 10^{-3} \text{ mol}}{\cancel{\text{L}}} \times \frac{105.99 \text{ g}}{1 \cancel{\text{ mol}}} = 2.2 \text{ g}$$

3.

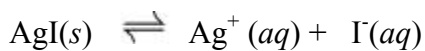


$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$$

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

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

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



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

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

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

**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.



To decrease the solubility, we need *more*  $\text{Mg(OH)}_2(s)$ , so we need to cause the equilibrium system to shift left. The  $[\text{Mg}^{2+}]$  or  $[\text{OH}^{-}]$  could be increased by adding  $\text{Mg(NO}_3)_2$  or NaOH to the solution.

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



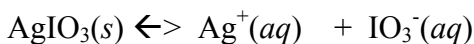
To increase the solubility, we need more  $\text{Mg(OH)}_2(s)$  to *dissolve*, so we need to cause the equilibrium system to shift right. The  $[\text{Mg}^{2+}]$  or  $[\text{OH}^{-}]$  could be decreased by adding a substance to precipitate one of them out such as  $\text{Ca(NO}_3)_2$  (to precipitate out the  $\text{OH}^{-}$ ). Additionally, we could add an acid (such as HCl) to neutralize the  $\text{OH}^{-}$  ions.

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

Since the value of  $K_{sp}$  does not change, a decrease in the  $[\text{Mg}^{2+}]$  (or  $[\text{OH}^{-}]$ ) 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



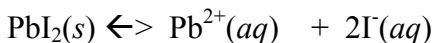
I	-	0	0.12
C	-	+x	+x
E	-	x	0.12 + x

$$K_{sp} = [\text{Ag}^{+}][\text{IO}_3^{-}]$$

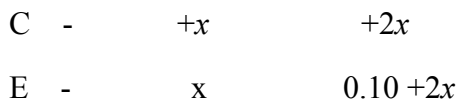
$$3.2 \times 10^{-8} = (x)(0.12)$$

$$x = 2.7 \times 10^{-7} \text{ M} = [\text{Ag}^{+}] = [\text{AgIO}_3] \text{ dissolved}$$

2.



I	-	0	0.10
---	---	---	------

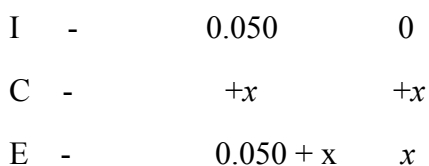
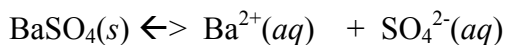


$$K_{sp} = [\text{Pb}^{2+}] [\text{I}^-]^2$$

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

$$x = 8.5 \times 10^{-7} \text{ M} = [\text{Pb}^{2+}] = [\text{PbI}_2] \text{ dissolved}$$

3.



$$K_{sp} = [\text{Ba}^{2+}] [\text{SO}_4^{2-}]$$

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

$$x = 2.2 \times 10^{-9} \text{ M} = [\text{Ba}^{2+}] = [\text{BaSO}_4] \text{ dissolved}$$

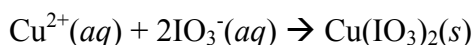
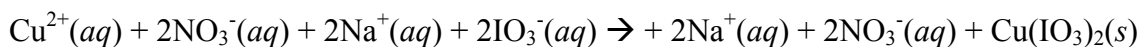
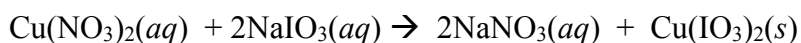
$$\text{mass of BaSO}_4 \text{ (g)} = \frac{2.2 \times 10^{-9} \text{ mol}}{\text{L}} \times \frac{233.40 \text{ g}}{1 \text{ mol}} = 5.1 \times 10^{-7} \text{ g/L}$$

### 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.



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

$$\text{Mixture 1: } [\text{Cu}(\text{NO}_3)_2] = \frac{0.010 \text{ mol} \times 0.0100 \text{ L}}{0.0200 \text{ L}} = 0.0050 \text{ M} = [\text{Cu}^{2+}]$$

$$\text{Mixture 2: } [\text{Cu}(\text{NO}_3)_2] = \frac{0.010 \text{ mol} \times 0.0080 \text{ L}}{0.0200 \text{ L}} = 0.0040 \text{ M} = [\text{Cu}^{2+}]$$

$$\text{Mixture 3: } [\text{Cu}(\text{NO}_3)_2] = \frac{0.010 \text{ mol} \times 0.0060 \text{ L}}{0.0200 \text{ L}} = 0.0030 \text{ M} = [\text{Cu}^{2+}]$$

$$\text{Mixture 4: } [\text{Cu}(\text{NO}_3)_2] = \frac{0.010 \text{ mol} \times 0.0040 \text{ L}}{0.0200 \text{ L}} = 0.0020 \text{ M} = [\text{Cu}^{2+}]$$

$$\text{Mixture 5: } [\text{Cu}(\text{NO}_3)_2] = \frac{0.010 \text{ mol} \times 0.0020 \text{ L}}{0.0200 \text{ L}} = 0.0010 \text{ M} = [\text{Cu}^{2+}]$$

3. Calculate the  $[\text{IO}_3^-]$  in each of the mixtures.

$$\text{Mixture 1: } [\text{NaIO}_3] = \frac{0.020 \text{ mol} \times 0.0100 \text{ L}}{0.0200 \text{ L}} = 0.010 \text{ M} = [\text{Cu}^{2+}]$$

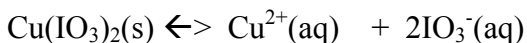
$$\text{Mixture 2: } [\text{NaIO}_3] = \frac{0.020 \text{ mol} \times 0.0080 \text{ L}}{0.0200 \text{ L}} = 0.0080 \text{ M} = [\text{Cu}^{2+}]$$

$$\text{Mixture 3: } [\text{NaIO}_3] = \frac{0.020 \text{ mol} \times 0.0060 \text{ L}}{0.0200 \text{ L}} = 0.0060 \text{ M} = [\text{Cu}^{2+}]$$

$$\text{Mixture 4: } [\text{NaIO}_3] = \frac{0.020 \text{ mol} \times 0.0040 \text{ L}}{0.0200 \text{ L}} = 0.0040 \text{ M} = [\text{Cu}^{2+}]$$

$$\text{Mixture 5: } [\text{NaIO}_3] = \frac{0.020 \text{ mol} \times 0.0020 \text{ L}}{0.0200 \text{ L}} = 0.0020 \text{ M} = [\text{Cu}^{2+}]$$

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



$$K_{sp} = [\text{Cu}^{2+}] [\text{IO}_3^-]^2$$

5. Calculate a TIP value for each mixture.

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

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

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

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

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

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

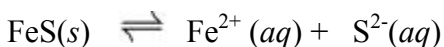
$$3.2 \times 10^{-8} < K_{sp} < 1.1 \times 10^{-7}$$

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)



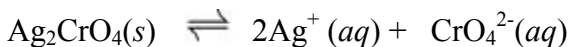
$$K_{sp} = [\text{Fe}^{2+}] [\text{S}^{2-}]$$

b)



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

c)

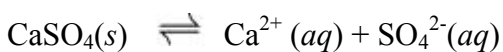


$$K_{sp} = [\text{Ag}^+]^2 [\text{CrO}_4^{2-}]$$

2. a)

CaSO<sub>4</sub>

b)



$$K_{sp} = [\text{Ca}^{2+}][\text{SO}_4^{2-}]$$

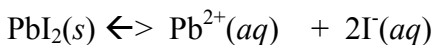
c)

Not necessarily. The Ca<sup>2+</sup> came from the CaCl<sub>2</sub> and the SO<sub>4</sub><sup>2-</sup> came from the H<sub>2</sub>SO<sub>4</sub>. The concentrations of ions depend on the original concentrations of the CaCl<sub>2</sub> and the H<sub>2</sub>SO<sub>4</sub>.

3.



$$0.015 \text{ M} \quad 0.015 \text{ M} \quad 0.030 \text{ M}$$

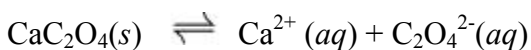


$$K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2$$

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

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

4.

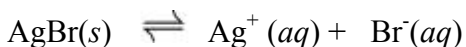


$$K_{sp} = [\text{Ca}^{2+}][\text{C}_2\text{O}_4^{2-}]$$

$$2.3 \times 10^{-9} = (5 \times 10^{-3})[\text{C}_2\text{O}_4^{2-}]$$

$$[\text{C}_2\text{O}_4^{2-}] = 5 \times 10^{-7} \text{ M}$$

5.

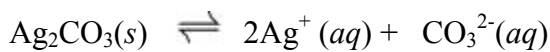


$s$                        $s$

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

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

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



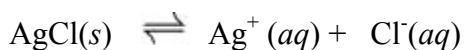
$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]$$

$$8.5 \times 10^{-12} = (7.3 \times 10^{-7})^2 [\text{CO}_3^{2-}]$$

$$[\text{CO}_3^{2-}] = 16 \text{ M}$$

6.

$$[\text{AgNO}_3] = \frac{0.20 \text{ mol}}{\text{L}} \times \frac{0.0002 \text{ L}}{0.1002 \text{ L}} = 4.0 \times 10^{-4} \text{ M} = [\text{Ag}^+]$$



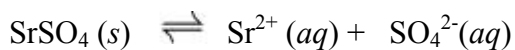
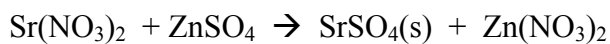
$$K_{sp} = [\text{Ag}^+] [\text{Cl}^-]$$

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

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

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

7.



$$K_{sp} = [\text{Sr}^{2+}] [\text{SO}_4^{2-}] = 3.4 \times 10^{-7}$$

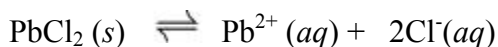
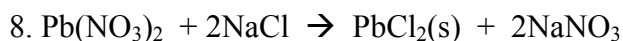
$$[\text{Sr}(\text{NO}_3)_2] = 0.055 \frac{\text{mol}}{\text{L}} \times \frac{0.0025 \text{ L}}{1.5025 \text{ L}} = 9.2 \times 10^{-5} \text{ M} = [\text{Sr}^{2+}]$$



$$[\text{ZnSO}_4] = 0.011 \frac{\text{mol}}{\text{L}} \times \frac{1.5 \text{ L}}{1.5025 \text{ L}} = 1.1 \times 10^{-2} \text{ M} = [\text{SO}_4^{2-}]$$

$$\begin{aligned} \text{TIP} &= [\text{Sr}^{2+}] [\text{SO}_4^{2-}] \\ &= (9.2 \times 10^{-5})(1.1 \times 10^{-2}) = 1.0 \times 10^{-6} \end{aligned}$$

TIP >  $K_{sp}$  so precipitate forms



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

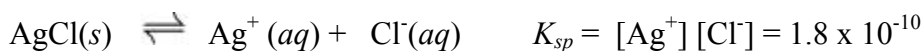
$$[\text{Pb}(\text{NO}_3)_2] = \frac{0.068 \text{ g}}{2.0 \text{ L}} \times \frac{1 \text{ mol}}{331.39 \text{ g}} = 1.0 \times 10^{-4} \text{ M} = [\text{Pb}^{2+}]$$

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

$$\begin{aligned} \text{TIP} &= [\text{Pb}^{2+}] [\text{Cl}^{-}]^2 \\ &= (1.0 \times 10^{-4})(0.080) = 8.2 \times 10^{-6} \end{aligned}$$

TIP <  $K_{sp}$  so no precipitate forms

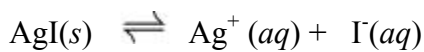
9. a)



b) See above

c) AgI

d)



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

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

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

e)



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

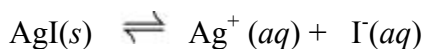
$$1.8 \times 10^{-10} = [\text{Ag}^+](0.020)$$

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

f)

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

g)



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

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

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

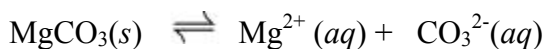
h)

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

$$0.020 \text{ M}$$

so  $100.00000 - 0.00047 = 99.99953\%$  precipitated out.

10.



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

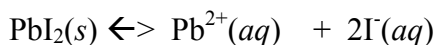
$$K_{sp} = [\text{Mg}^{2+}] [\text{CO}_3^{2-}]$$

$$6.8 \times 10^{-6} = (4.9 \times 10^{-4}) [\text{CO}_3^{2-}]$$

$$[\text{CO}_3^{2-}] = 1.4 \times 10^{-2} \text{ M}$$

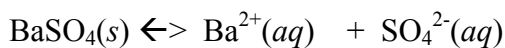
$$\text{mass Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O (g)} = 1.0 \text{ L} \times \frac{1.4 \times 10^{-2} \text{ mol}}{\text{L}} \times \frac{286.19 \text{ g}}{1 \text{ mol}} = 4.0 \text{ g}$$

11.



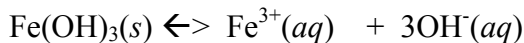
To decrease the solubility, the equilibrium must be shifted left. By adding  $\text{Pb}(\text{NO}_3)_2$  or  $\text{KI}$ , the common ions ( $\text{Pb}^{2+}$  or  $\text{I}^{-}$ ) cause an increase in their respective concentrations.

12.



In  $\text{Na}_2\text{SO}_4$ , the common ion  $\text{SO}_4^{2-}$  exists, so equilibrium shifts left and solubility is decreased. In water, there are no  $\text{Ba}^{2+}$  or  $\text{SO}_4^{2-}$  ions.

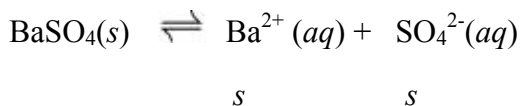
13.



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

### Extension

14. a)



$$K_{sp} = [\text{Ba}^{2+}] [\text{SO}_4^{2-}]$$

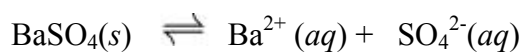
$$1.1 \times 10^{-10} = s^2$$

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

$$\text{mass BaSO}_4 \text{ (g)} = 0.200 \text{ L} \times \frac{1.0 \times 10^{-5} \text{ mol}}{\text{L}} \times \frac{233.40 \text{ g}}{1 \text{ mol}} = 4.7 \times 10^{-4} \text{ g}$$

L 1 mol

b)



I 0 0.10

C +x +x

E x 0.10 + x

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$

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

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

$$\text{mass BaSO}_4(\text{g}) = 0.200 \text{ L} \times \frac{1.1 \times 10^{-9} \text{ mol}}{\text{L}} \times \frac{233.40 \text{ g}}{1 \text{ mol}} = 5.1 \times 10^{-8} \text{ g}$$

L 1 mol