

Chapter 4

Page 219, Quick Check

- No. Some salts are weak electrolytes. They do not dissolve to a significant extent.
- a) molecular
b) ionic
c) ionic
d) molecular
e) ionic
- 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.

Page 220, Quick Check

- $\text{CaCl}_2(aq) \rightarrow \text{Ca}^{2+}(aq) + 2\text{Cl}^-(aq)$
- $(\text{NH}_4)_2\text{C}_2\text{O}_4(aq) \rightarrow 2\text{NH}_4^+(aq) + \text{C}_2\text{O}_4^{2-}(aq)$
- $\text{Na}_3\text{PO}_4(aq) \rightarrow 3\text{Na}^+(aq) + \text{PO}_4^{3-}(aq)$

Page 221, Quick Check

- a) $\text{Na}_2\text{C}_2\text{O}_4(aq) \rightarrow 2\text{Na}^+(aq) + \text{C}_2\text{O}_4^{2-}(aq)$
0.45 M 0.90 M 0.45 M
total [ions] = 0.90 M + 0.45 M = 1.35 M
b) $\text{AlCl}_3(aq) \rightarrow \text{Al}^{3+}(aq) + 3\text{Cl}^-(aq)$
2.5 M 2.5 M 7.5 M
total [ions] = 2.5 M + 7.5 M = 10.0 M
- $\text{LiOH}(aq) \rightarrow \text{Li}^+(aq) + \text{OH}^-(aq)$
1.0 M 1.0 M 1.0 M
total ion concentration = 2.0 M
 $\text{CaCl}_2(aq) \rightarrow \text{Ca}^{2+}(aq) + 2\text{Cl}^-(aq)$
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.
- $\text{AlBr}_3(aq) \rightarrow \text{Al}^{3+}(aq) + 3\text{Br}^-(aq)$
According to the dissociation equation, 1 mole Al^{3+} : 3 mole Br^- . If $[\text{Br}^-] = 0.15 \text{ M}$, then $[\text{Al}^{3+}] = 0.050 \text{ M}$.

Page 222, Practice Problems 4.1.1

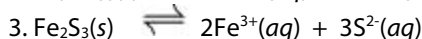
- a) $[\text{NH}_4^+] = 3(3.6 \times 10^{-3}) = 1.1 \times 10^{-2} \text{ M}$
 $[\text{PO}_4^{3-}] = 3.6 \times 10^{-3} \text{ M}$
b) $[\text{H}^+] = 2.6 \text{ M}$ $[\text{Cl}^-] = 2.6 \text{ M}$
- mass of $\text{Al}_2(\text{SO}_4)_3 = 13 \text{ g}$
- total $[\text{NH}_4^+] = 0.58 \text{ M}$, $[\text{SO}_4^{2-}] = 0.15 \text{ M}$ $[\text{S}^{2-}] = 0.14 \text{ M}$

Page 223, Practice Problems 4.1.2

- $5.4 \times 10^{-2} \text{ M}$
- Mass $\text{Pb}(\text{IO}_3)_2 = 7.5 \times 10^{-3} \text{ g}$
- $\text{mol CaCO}_3 = 7.1 \times 10^{-5} \text{ M}$

Page 224, Quick Check

- $\text{ZnCO}_3(s) \rightleftharpoons \text{Zn}^{2+}(aq) + \text{CO}_3^{2-}(aq)$



Page 226, 4.1 Review

- 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
- 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.
- a) $\text{Mg}(\text{ClO}_4)_2(aq) \rightarrow 2\text{Mg}^{2+}(aq) + 2\text{ClO}_4^-(aq)$
b) $\text{CaCr}_2\text{O}_7(aq) \rightarrow \text{Ca}^{2+}(aq) + \text{Cr}_2\text{O}_7^{2-}(aq)$
c) $\text{Cu}(\text{CH}_3\text{COO})_2(aq) \rightarrow \text{Cu}^{2+}(aq) + 2\text{CH}_3\text{COO}^-(aq)$
d) $\text{Mn}(\text{SCN})_2(aq) \rightarrow \text{Mn}^{2+}(aq) + 2\text{SCN}^-(aq)$
e) $\text{Al}(\text{HC}_2\text{O}_4)_3(aq) \rightarrow \text{Al}^{3+}(aq) + 3\text{HC}_2\text{O}_4^-(aq)$
f) $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}(aq) \rightarrow \text{Ba}^{2+}(aq) + 2\text{OH}^-(aq) + 8 \text{H}_2\text{O}(l)$
- $\text{mol BaSO}_4 = 1.0 \times 10^{-5} \text{ M}$
- mass $\text{CaCO}_3 = 1.8 \times 10^{-3} \text{ g}$
- $[\text{Na}^+] = 0.025 \text{ M}$ $[\text{Cr}_2\text{O}_7^{2-}] = 0.013 \text{ M}$
- $[\text{Mg}^{2+}] = 0.80 \text{ M}$, total $[\text{Cl}^-] = 4.9 \text{ M}$, $[\text{Na}^+] = 3.3 \text{ M}$
- 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.
- a) $\text{AgBrO}_3(s) \rightleftharpoons \text{Ag}^+(aq) + \text{BrO}_3^-(aq)$
b) $\text{Al}_2(\text{CrO}_4)_3(s) \rightleftharpoons 2\text{Al}^{3+}(aq) + 2\text{CrO}_4^{2-}(aq)$
c) $\text{Mg}(\text{OH})_2(s) \rightleftharpoons \text{Mg}^{2+}(aq) + 2\text{OH}^-(aq)$
d) $\text{PbSO}_4(s) \rightleftharpoons \text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq)$
e) $\text{Cu}_3(\text{PO}_4)_2(s) \rightleftharpoons 3\text{Cu}^{2+}(aq) + 2\text{PO}_4^{3-}(aq)$

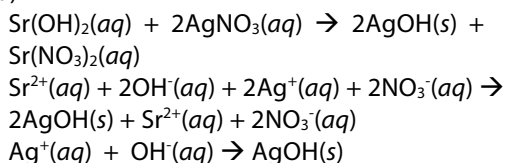
Page 230, Practice Problems 4.2.1

- a) low solubility
b) low solubility
c) soluble
d) soluble
e) soluble
d) soluble
- (many answers possible)
a) Na_2CO_3
b) SrSO_4
c) Ag^+
d) NO_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.

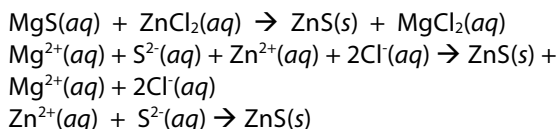
Page 231, Quick Check

1. a) BaSO₄
- b) Fe(OH)₂
- c) Zn₃(PO₄)₂
- d) CrCO₃
- e) Mn(OH)₂

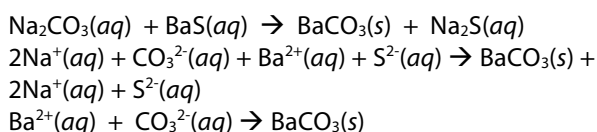
2. a)



b)



c)



Page 233, Practice Problems 4.2.2

(other answers possible)

1. First add NaCl. Filter out the AgCl(s)
Then add Na₂S. Filter out the BeS(s)
Finally add Na₂SO₄. Filter out the BaSO₄(s)
2. First add Ca(NO₃)₂. Filter out the CaSO₄(s)
Then add Cu(NO₃)₂. Filter out the CuS(s)
Finally add AgNO₃. Filter out the AgBr(s)
3. First add Sr(NO₃)₂. Filter out the Sr₃(PO₄)₂(s)
Then add Mg(NO₃)₂. Filter out the Mg(OH)₂(s)
Finally add AgNO₃. Filter out the Ag₂S(s)

Page 234, Quick Check

1. The color of the solution (for example, solutions containing Cu²⁺ 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 NH₃ or HNO₃.
2. First add NaCl. A precipitate of both PbCl₂ 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 NH₃. Any AgCl re-dissolves. Filter out the PbCl₂(s). Dry the PbCl₂(s) and measure its mass. If the mass had decreased, the presence of AgCl is confirmed.
3. First add HNO₃. The formation of bubbles confirms the presence of CO₃²⁻ ions. Next, add CaCl₂. If a precipitate forms, it is Ca₃(PO₄)₂(s). Filter out the precipitate.

Page 238, 4.2 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. mol AgCH₃COO = 0.0665 M

0.0665 M < 0.1 M so AgCH₃COO would have a low solubility.

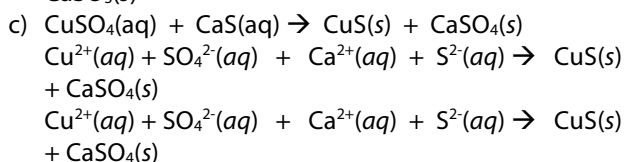
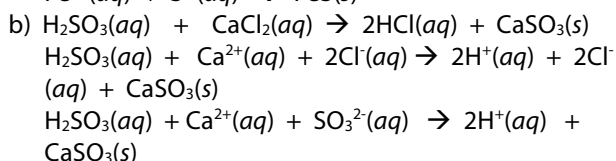
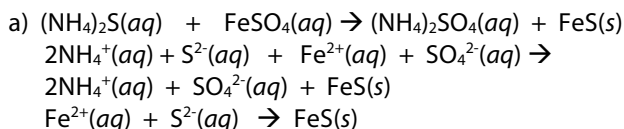
4. A formula equation is written with substances as compounds and states shown.

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 NO₃⁻ ions are often spectators.

6.



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

8. First add Na₂SO₄ to precipitate CaSO₄: Ca²⁺(aq) + SO₄²⁻(aq) → CaSO₄(s)

Filter out the precipitate

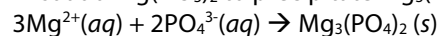
Secondly, add Na₂S to precipitate Cr₂S₃: 2Cr³⁺(aq) + 3S²⁻(aq) → Cr₂S₃(s)

Filter out the precipitate

Finally add NaOH to precipitate Mg(OH)₂: Mg²⁺(aq) + 2OH⁻(aq) → Mg(OH)₂(s)

Filter out the precipitate

9. First add Mg(NO₃)₂ to precipitate Mg₃(PO₄)₂:



Filter out the precipitate

Secondly, add Fe(NO₃)₂ to precipitate FeS: Fe²⁺(aq) + S²⁻(aq) → FeS(s)

Filter out the precipitate

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

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) $2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{Ag}_2\text{CO}_3(s)$

b) $\text{Ag}_2\text{CO}_3(s) + 2\text{H}^+(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + 2\text{Ag}^+(aq)$

$\text{Ag}_2\text{CO}_3(s) + 4\text{NH}_3(aq) \rightarrow 2\text{Ag}(\text{NH}_3)_2^+(aq) + \text{CO}_3^{2-}(aq)$

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) $\text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + \text{CaCl}_2(aq)$

$\text{CaCO}_3(s) + 2\text{H}^+(aq) + 2\text{Cl}^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + \text{Ca}^{2+}(aq) + 2\text{Cl}^-(aq)$

$\text{CaCO}_3(s) + 2\text{H}^+(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + \text{Ca}^{2+}(aq)$

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.

Page 242, Quick Check

1. a) $\text{SrCO}_3(s) \rightleftharpoons \text{Sr}^{2+}(aq) + \text{CO}_3^{2-}(aq)$

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

c) $\text{Ca}_3(\text{PO}_4)_2(s) \rightleftharpoons 3\text{Ca}^{2+}(aq) + 2\text{PO}_4^{3-}(aq)$

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.

Page 244, Practice Problems 4.3.1

1. a) $K_{sp} = 3.7 \times 10^{-9}$

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

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

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

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

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

Page 247, Practice Problems 4.3.2

1. a) $1.3 \times 10^{-5}\text{ M}$

b) $6.8 \times 10^{-11}\text{ g / mL}$

c) $4.5 \times 10^{-5}\text{ M}$

d) $6.8 \times 10^{-11}\text{ g}$

2. $3.0 \times 10^{-10}\text{ M}$

3. $4.0 \times 10^{-3}\text{ g}$

Page 249, 4.3 Review Questions

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

b) $\text{Cd}_3(\text{AsO}_4)_2(s) \rightleftharpoons 3\text{Cd}^{2+}(aq) + 2\text{AsO}_4^{3-}(aq)$ $K_{sp} = [\text{Cd}^{2+}]^3[\text{AsO}_4^{3-}]^2$

c) $\text{BaMoO}_4(s) \rightleftharpoons \text{Ba}^{2+}(aq) + \text{MoO}_4^{2-}(aq)$ $K_{sp} = [\text{Ba}^{2+}][\text{MoO}_4^{2-}]$

d) $\text{CaSO}_4(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq)$ $K_{sp} = [\text{Ca}^{2+}][\text{SO}_4^{2-}]$

e) $\text{Pb}(\text{IO}_3)_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{IO}_3^-(aq)$ $K_{sp} = [\text{Pb}^{2+}][\text{IO}_3^-]^2$

f) $\text{Ag}_2\text{CO}_3(s) \rightleftharpoons 2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq)$ $K_{sp} = [\text{Ag}^+]^2[\text{CO}_3^{2-}]$

2.

a) $\text{BaSO}_3(s) \rightleftharpoons \text{Ba}^{2+}(aq) + \text{SO}_3^{2-}(aq)$

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

4. $K_{sp} = 9.5 \times 10^{-17}$

5. mass gypsum (g) = 0.26 g

6. $7.7 \times 10^{-6}\text{ g}$

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

8. Disagree. Even though the K_{sp} value for CdCO_3 ($1.0 \times 10^{-6}\text{ M}$) is greater than the K_{sp} for $\text{Cd}(\text{OH})_2$ ($1.2 \times 10^{-5}\text{ M}$) its solubility is less.

9. The solubility of AgCl ($1.3 \times 10^{-5}\text{ M}$) is less than the solubility of Ag_2CrO_4 ($6.5 \times 10^{-5}\text{ M}$). All things being equal, the precipitate of AgCl will form first.

10. $9.0 \times 10^{-2}\text{ g}$

Page 252, Practice Problems 4.4.1

1. $\text{TIP} = 7.6 \times 10^{-10}$

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

2. $\text{TIP} = 1.4 \times 10^{-7}$

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

3. $\text{TIP} = 4.1 \times 10^{-4}$

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

Page 254, Practice Problems 4.4.2

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

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

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

2. 2.2 g

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

Page 256, 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* $\text{Mg}(\text{OH})_2(s)$, so we need to cause the equilibrium system to shift left. The $[\text{Mg}^{2+}]$ or $[\text{OH}^-]$ could be increased by adding $\text{Mg}(\text{NO}_3)_2$ or NaOH to the solution.
- To increase the solubility, we need more $\text{Mg}(\text{OH})_2(s)$ 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 $\text{Ca}(\text{NO}_3)_2$ (to precipitate out the OH^-). Additionally, we could add an acid (such as HCl) to neutralize the OH^- ions.
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.

Page 257 Practice Problems 4.4.3

- $x = 2.7 \times 10^{-7} \text{ M} = [\text{Ag}^+] = [\text{AgIO}_3]$ dissolved
- $x = 8.5 \times 10^{-7} \text{ M} = [\text{Pb}^{2+}] = [\text{PbI}_2]$ dissolved
- $5.1 \times 10^{-7} \text{ g/L}$

Page 260, 4.4 Review Questions

- $\text{FeS}(s) \rightleftharpoons \text{Fe}^{2+}(aq) + \text{S}^{2-}(aq)$
 $K_{sp} = [\text{Fe}^{2+}][\text{S}^{2-}]$
 - $\text{Mg}(\text{OH})_2(s) \rightleftharpoons \text{Mg}^{2+}(aq) + 2\text{OH}^-(aq)$
 $K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2$
 - $\text{Ag}_2\text{CrO}_4(s) \rightleftharpoons 2\text{Ag}^+(aq) + \text{CrO}_4^{2-}(aq)$
 $K_{sp} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}]$
- CaSO_4
 - $\text{CaSO}_4(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq)$
 $K_{sp} = [\text{Ca}^{2+}][\text{SO}_4^{2-}]$
- Not necessarily. The Ca^{2+} came from the CaCl_2 and the SO_4^{2-} came from the H_2SO_4 . The concentrations of ions depend on the original concentrations of the CaCl_2 and the H_2SO_4 .
- $[\text{Pb}^{2+}] = 9.4 \times 10^{-6} \text{ M}$
- $[\text{C}_2\text{O}_4^{2-}] = 5 \times 10^{-7} \text{ M}$
- $[\text{CO}_3^{2-}] = 16 \text{ M}$
- $2.6 \times 10^{-6} \text{ g}$
- $\text{TIP} = [\text{Sr}^{2+}][\text{SO}_4^{2-}] = (9.2 \times 10^{-5})(1.1 \times 10^{-2}) = 1.0 \times 10^{-6}$
 $\text{TIP} > K_{sp}$ so precipitate forms
- $\text{TIP} = [\text{Pb}^{2+}][\text{Cl}^-]^2 = (1.0 \times 10^{-4})(0.080)^2 = 6.6 \times 10^{-7}$
 $\text{TIP} < K_{sp}$ so no precipitate forms
- $K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$
 $K_{sp} = [\text{Ag}^+][\text{I}^-] = 8.5 \times 10^{-17}$
 - See above
 - AgI
 - $[\text{Ag}^+] = 4.3 \times 10^{-15} \text{ M}$
 - $[\text{Ag}^+] = 9.0 \times 10^{-9} \text{ M}$
 - $4.3 \times 10^{-15} \text{ M} < [\text{Ag}^+] < 9.0 \times 10^{-9} \text{ M}$
 - $[\text{I}^-] = 9.4 \times 10^{-9} \text{ M}$
 - $\% \text{ I}^- \text{ remaining} = \frac{9.4 \times 10^{-9} \text{ M}}{0.020 \text{ M}} \times 100\% = 0.00047\%$
so $100.00000 - 0.00047 = 99.99953\%$ precipitated out.

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

Extension

- $4.7 \times 10^{-4} \text{ g}$
 - $5.1 \times 10^{-8} \text{ g}$