

Section 7.6: Solubility Equilibria and the Solubility Product Constant

Tutorial 1 Practice, page 462

1. Solution: $[\text{MgCO}_3(\text{s})]$ is constant, so it is not included in the equation. Therefore,
 $K_{\text{sp}} = [\text{Mg}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})]$

Statement: The solubility product constant equation for the solubility equilibrium is
 $K_{\text{sp}} = [\text{Mg}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})]$

2. Given: iron(III) hydroxide dissolves in water

Required: K_{sp} equation

Solution:

Step 1. Write a balanced equation for the dissolution of iron(III) hydroxide in water.



Step 2. $[\text{Fe}(\text{OH})_3(\text{s})]$ is constant, so it is not included in the equation. Therefore,

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

Statement: The solubility product constant equation for iron(III) hydroxide is

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

3. Given: At 25 °C: $[\text{Ca}^{2+}(\text{aq})] = 4.53 \times 10^{-7} \text{ mol/L}$; $[\text{PO}_4^{3-}(\text{aq})] = 3.02 \times 10^{-7} \text{ mol/L}$

Required: K_{sp} of calcium phosphate at 25 °C

Solution:

Step 1. Write a balanced equation for the solution equilibrium.



Step 2. Write the solubility product constant equation for the equilibrium, omitting any solids.

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

Step 3. Substitute the given ion concentrations into the K_{sp} equation and solve.

$$\begin{aligned} K_{\text{sp}} &= [\text{Ca}^{2+}(\text{aq})]^3[\text{PO}_4^{3-}(\text{aq})]^2 \\ &= (4.53 \times 10^{-7})^3(3.02 \times 10^{-7})^2 \end{aligned}$$

$$K_{\text{sp}} = 8.48 \times 10^{-33}$$

Statement: The K_{sp} of calcium phosphate in a saturated solution at 25 °C is 8.48×10^{-33} .

Tutorial 2 Practice, page 464

1. Given: solubility of $\text{MgF}_2(\text{s}) = 1.72 \times 10^{-3} \text{ g/100 mL}$ at 25 °C

Required: K_{sp} of $\text{MgF}_2(\text{s})$ at 25 °C

Solution:

Step 1. Write the balanced equation.



Step 2. Write the solubility product constant equation.

$$K_{\text{sp}} = [\text{Mg}^{2+}(\text{aq})][\text{F}^-(\text{aq})]^2$$

Step 3. Determine the equilibrium concentrations in mol/L of the dissolved ions. Use an ICE table. Since the concentration of solid MgF_2 remains constant, it is not included.

	$\text{MgF}_2(\text{s})$	\rightleftharpoons	$\text{Mg}^{2+}(\text{aq})$	+	$2\text{F}^{-}(\text{aq})$
I	—		0.00		0.00
C	—		+ x		+ $2x$
E	—		x		$2x$

Step 4. Calculate all concentrations in mol/L. To convert given mass solubility of solid magnesium fluoride to molar solubility, first calculate the molar mass of magnesium fluoride. Then, use the molar mass to convert mass solubility to molar solubility.

$$\text{molar mass of MgF}_2 = \left(24.31 \frac{\text{g}}{\text{mol}}\right) + (2) \left(19.00 \frac{\text{g}}{\text{mol}}\right)$$

$$\text{molar mass of MgF}_2 = 62.31 \text{ g/mol}$$

$$\text{molar solubility of MgF}_2 = \frac{1.72 \times 10^{-3} \text{ g}}{100 \text{ mL}} \times \frac{1 \text{ mol}}{62.31 \text{ g}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$\text{molar solubility of MgF}_2 = 2.76 \times 10^{-4} \text{ mol/L}$$

Step 5. Write the K_{sp} equation using the expressions in the ICE table, and then substitute the molar solubility of solid magnesium fluoride for x .

$$K_{\text{sp}} = [\text{Mg}^{2+}(\text{aq})][\text{F}^{-}(\text{aq})]^2$$

$$= (x)(2x)^2$$

$$= (2.76 \times 10^{-4})(5.52 \times 10^{-4})^2$$

$$K_{\text{sp}} = 8.41 \times 10^{-11}$$

Statement: The solubility product constant of solid magnesium fluoride is 8.41×10^{-11} .

2. Given: K_{sp} of AgI(s) at 25 °C = 8.5×10^{-17}

Required: molar solubility of AgI(s) at 25 °C

Solution:

Step 1. Write the balanced chemical equation.



Step 2. Write the solubility product constant equation.

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

$$[\text{Ag}^{+}(\text{aq})] = [\text{I}^{-}(\text{aq})]$$

$$K_{\text{sp}} = [\text{Ag}^{+}(\text{aq})]^2$$

Step 3. Determine the equilibrium concentrations of the dissolved ions.

From the balanced equation, the molar ratio is 1:1:1, so the concentrations of these ions are equal at any point in time. Since the solubility of solid silver iodide is 8.5×10^{-17} mol/L, then a maximum of 8.5×10^{-17} mol of solid silver iodide can dissolve in water to produce 1.0 L of solution at 25 °C. This will form 8.5×10^{-17} mol of silver ions and 8.5×10^{-17} mol of iodide ions. Therefore, the equilibrium concentrations of these ions are also equal to 8.5×10^{-17} mol/L.

Step 4. Substitute the equilibrium concentrations into the K_{sp} equation and solve for K_{sp} .

$$K_{sp} = [\text{Ag}^+(\text{aq})]^2$$

$$8.5 \times 10^{-7} = [\text{Ag}^+(\text{aq})]^2$$

$$[\text{Ag}^+(\text{aq})] = \sqrt{8.5 \times 10^{-7}}$$

$$[\text{Ag}^+(\text{aq})] = 9.2 \times 10^{-9} \text{ mol/L}$$

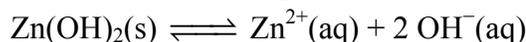
Statement: The molar solubility of silver iodide at 25 °C is 9.2×10^{-9} mol/L.

3. Given: K_{sp} of $\text{Zn}(\text{OH})_2(\text{s}) = 7.7 \times 10^{-17}$ at 25 °C

Required: molar solubility of $\text{Zn}(\text{OH})_2(\text{s})$ at 25 °C

Solution:

Step 1. Write the balanced equation for the solution equilibrium.



Step 2. Write the solubility product constant equation.

$$K_{sp} = [\text{Zn}^{2+}(\text{aq})][\text{OH}^-(\text{aq})]^2$$

Step 3. Determine the equilibrium concentrations in mol/L of the dissolved ions. Use an ICE table. Since the concentration of solid $\text{Zn}(\text{OH})_2$ remains constant, it is not included.

	$\text{Zn}(\text{OH})_2(\text{s})$	\rightleftharpoons	$\text{Zn}^{2+}(\text{aq})$	+	$2 \text{OH}^-(\text{aq})$
I	—		0.00		0.00
C	—		+x		+2x
E	—		x		2x

Step 4. Substitute the equilibrium concentration expressions from the ICE table into the K_{sp} equation and solve for x.

$$K_{sp} = [\text{Zn}^{2+}(\text{aq})][\text{OH}^-(\text{aq})]^2$$

$$K_{sp} = (x)(2x)^2$$

$$7.7 \times 10^{-17} = 4x^3$$

$$x = \sqrt[3]{1.925 \times 10^{-17}}$$

$$x = 2.7 \times 10^{-6}$$

$[\text{Zn}^{2+}(\text{aq})]$ = molar solubility of $\text{Zn}(\text{OH})_2(\text{s})$

molar solubility of $\text{Zn}(\text{OH})_2(\text{s}) = 2.7 \times 10^{-6}$ mol/L

Statement: The molar solubility of zinc hydroxide at 25 °C is 2.7×10^{-6} mol/L.

4. Given: K_{sp} of $\text{HgCl} = 1.5 \times 10^{-18}$; K_{sp} of $\text{CuCl} = 1.7 \times 10^{-7}$

Required: compare molar solubilities of HgCl(s) and CuCl(s)

Solution:

Step 1. Determine the molar solubility of mercury(I) chloride:

The balanced chemical equation is



The solubility product constant equation is

$$K_{sp} = [\text{Hg}^+(\text{aq})][\text{Cl}^-(\text{aq})]$$

Determine the equilibrium concentrations in mol/L of the dissolved ions. Use an ICE table. Since the concentration of solid mercury(I) chloride remains constant, it is not included.

	$\text{HgCl(s)} \rightleftharpoons \text{Hg}^+(\text{aq}) + \text{Cl}^-(\text{aq})$	
I	—	0.00 0.00
C	—	+x +x
E	—	x x

Substitute the equilibrium concentration expressions from the ICE table into the K_{sp} equation and solve for x .

$$K_{sp} = [\text{Hg}^+(\text{aq})][\text{Cl}^-(\text{aq})]^2$$

$$= (x)(x)$$

$$K_{sp} = x^2$$

$$x = \sqrt{K_{sp}}$$

$$= \sqrt{1.5 \times 10^{-18}}$$

$$x = 1.2 \times 10^{-9}$$

molar solubility of $\text{HgCl} = 1.2 \times 10^{-9}$ mol/L

Step 2. Determine the molar solubility of copper(I) chloride. Following the same procedure as for mercury(I) chloride in step 1,

$$K_{sp} = [\text{Cu}^+(\text{aq})][\text{Cl}^-(\text{aq})]^2$$

$$K_{sp} = x^2$$

$$x = \sqrt{K_{sp}}$$

$$= \sqrt{1.7 \times 10^{-7}}$$

$$x = 4.1 \times 10^{-4}$$

molar solubility of $\text{CuCl} = 4.1 \times 10^{-4}$ mol/L

Step 3. Compare the molar solubilities of HgCl(s) and CuCl(s) :

$$1.2 \times 10^{-9} \text{ mol/L} < 4.1 \times 10^{-4} \text{ mol/L}$$

Statement: The solubility of copper(I) chloride is greater than the solubility of mercury(I) chloride.

Tutorial 3 Practice, page 468

1. Given: $[\text{CuNO}_3(\text{aq})] = 0.015 \text{ mol/L}$; $[\text{KI}(\text{aq})] = 0.075 \text{ mol/L}$

Required: To predict whether a precipitate will form and identify any precipitate

Solution:

Step 1. Identify all of the ions that will be present in the mixture.

$\text{CuNO}_3(\text{aq})$ will contribute aqueous ions of copper, $\text{Cu}^+(\text{aq})$, and nitrate, $\text{NO}_3^-(\text{aq})$.

$\text{KI}(\text{aq})$ will contribute aqueous ions of potassium, $\text{K}^+(\text{aq})$, and iodide, $\text{I}^-(\text{aq})$.

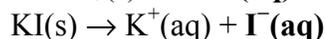
All four ions will be in the mixture: $\text{Cu}^+(\text{aq})$, $\text{NO}_3^-(\text{aq})$, $\text{K}^+(\text{aq})$, and $\text{I}^-(\text{aq})$.

Step 2. Use the solubility rules in Table 3 on page 465 of the Student Book to predict whether a precipitate may form.

Table 3 indicates that $\text{KNO}_3(\text{s})$ is highly soluble and $\text{CuI}(\text{s})$ has low solubility. Therefore, a precipitate of solid copper(I) iodide, $\text{CuI}(\text{s})$, may form.

Step 3. Determine the concentrations of the ions that may react to form a precipitate.

First, write the balanced dissolution equations for $\text{CuNO}_3(\text{aq})$ and $\text{KI}(\text{aq})$.



These balanced equations indicate that both dissolution reactions have a 1:1:1 molar ratio. Therefore, the concentrations of both ions are equal to the given concentrations of the ionic compounds. Therefore,

$$[\text{Cu}^+(\text{aq})] = [\text{CuNO}_3(\text{aq})]$$

$$[\text{Cu}^+(\text{aq})] = 0.015 \text{ mol/L}$$

and

$$[\text{I}^-(\text{aq})] = [\text{KI}(\text{aq})]$$

$$[\text{I}^-(\text{aq})] = 0.075 \text{ mol/L}$$

Step 4. Determine the trial ion product equation of the precipitate that may form. $\text{CuI}(\text{s})$ may precipitate. First, write a balanced equation for the dissolution equilibrium system of $\text{CuI}(\text{s})$.



From the balanced equation, the trial ion product equation is

$$Q = [\text{Cu}^+(\text{aq})][\text{I}^-(\text{aq})]$$

Step 5. Calculate Q by substituting the concentrations of the dissolved ions in the mixture into the trial ion product equation.

$$Q = [\text{Cu}^+(\text{aq})][\text{I}^-(\text{aq})]$$

$$= (0.015)(0.075)$$

$$Q = 1.1 \times 10^{-3}$$

Step 6. Compare the K_{sp} and Q values of the compound that may precipitate, and then predict whether the position of this dissolution equilibrium system will shift and whether a precipitate will form.

From Table 1 in Appendix B4,

$$K_{\text{sp}} \text{ of } \text{CuI}(\text{s}) = 1.3 \times 10^{-12}$$

$$Q = 1.1 \times 10^{-3}$$

Q is greater than K_{sp} , so the dissolution equilibrium system will shift to the left and a precipitate of $\text{CuI}(\text{s})$ will form.

Statement: A precipitate of solid copper(I) iodide, CuI(s) , will form when the two solutions are mixed.

2. Given: $V_{\text{Pb(NO}_3)_2} = 100.0 \text{ mL}$; $[\text{Pb(NO}_3)_2(\text{aq})] = 0.0500 \text{ mol/L}$; $V_{\text{NaI}} = 200.0 \text{ mL}$;

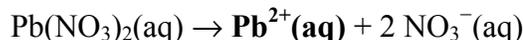
$[\text{NaI}(\text{aq})] = 0.100 \text{ mol/L}$

Required: To predict whether lead(II) iodide, $\text{PbI}_2(\text{s})$, will precipitate

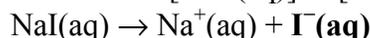
Solution:

Step 1. The initial $\text{Pb(NO}_3)_2(\text{aq})$ solution and initial $\text{NaI}(\text{aq})$ solution will contribute these four ions to the mixture: $\text{Pb}^{2+}(\text{aq})$, $\text{NO}_3^-(\text{aq})$, $\text{Na}^+(\text{aq})$, and $\text{I}^-(\text{aq})$.

Step 2. Determine the concentrations of the ions that may react to form a precipitate.



The ratio of $[\text{Pb}^{2+}(\text{aq})]$ to $[\text{Pb(NO}_3)_2(\text{aq})]$ is 1:1.



The ratio of $[\text{I}^-(\text{aq})]$ to $[\text{NaI}(\text{aq})]$ is 1:1. Therefore,

$$[\text{Pb}^{2+}(\text{aq})] = 0.0500 \text{ mol/L}$$

$$[\text{I}^-(\text{aq})] = 0.100 \text{ mol/L}$$

These concentrations are the concentration of $\text{Pb}^{2+}(\text{aq})$ ions in the initial $\text{Pb(NO}_3)_2(\text{aq})$ solution and the concentration of $\text{I}^-(\text{aq})$ in the initial $\text{NaI}(\text{aq})$ solution *before* these solutions are mixed. When the initial solutions are mixed, the volume of the final mixture, V_f , increases and is the sum of the volumes of the two initial solutions.

$$V_f = V_{\text{Pb(NO}_3)_2} + V_{\text{NaI}}$$

$$= 100.0 \text{ mL} + 200.0 \text{ mL}$$

$$V_f = 300.0 \text{ mL}$$

Use the dilution formula to calculate the concentrations of $\text{Pb}^{2+}(\text{aq})$ and $\text{I}^-(\text{aq})$ in the final mixture.

$$c_c V_c = c_d V_d$$

$$c_d = \frac{c_c V_c}{V_d}$$

where c_c is the initial, concentrated, ion concentration and c_d is the final, more dilute, ion concentration. V_c and V_d are the corresponding volumes.

Substitute the appropriate values for $\text{Pb}^{2+}(\text{aq})$ into the dilution equation.

$$[\text{Pb}^{2+}(\text{aq})]_{\text{final}} = (0.0500 \text{ mol/L}) \left(\frac{100.0 \text{ mL}}{300.0 \text{ mL}} \right)$$

$$[\text{Pb}^{2+}(\text{aq})]_{\text{final}} = 0.0167 \text{ mol/L}$$

Then, substitute the appropriate values for $\text{I}^-(\text{aq})$ into the dilution equation.

$$[\text{I}^-(\text{aq})]_{\text{final}} = (0.100 \text{ mol/L}) \left(\frac{200.0 \text{ mL}}{300.0 \text{ mL}} \right)$$

$$[\text{I}^-(\text{aq})]_{\text{final}} = 0.0667 \text{ mol/L}$$

Step 3. Determine the trial ion product equation of the precipitate that may form.



$$Q = [\text{Pb}^{2+}(\text{aq})][\text{I}^{-}(\text{aq})]^2$$

Step 4. Calculate Q .

$$Q = [\text{Pb}^{2+}(\text{aq})][\text{I}^{-}(\text{aq})]^2$$

$$= (0.0167)(0.0667)^2$$

$$Q = 7.4 \times 10^{-5}$$

Step 5. Compare the K_{sp} and Q values of the compound that may precipitate, and then predict whether the position of this dissolution equilibrium system will shift and whether a precipitate will form.

From Table 1 in Appendix B4,

$$K_{\text{sp}} \text{ of } \text{PbI}_2(\text{s}) = 8.5 \times 10^{-9}$$

$$Q = 7.4 \times 10^{-5}$$

Since Q is greater than K_{sp} , the dissolution equilibrium system shifts to the left and precipitation occurs.

Statement: A precipitate of solid lead(II) iodide, $\text{PbI}_2(\text{s})$, will form when these solutions are mixed.

3. Given: $V_{\text{CaCl}_2} = 250.0 \text{ mL}$; $[\text{CaCl}_2(\text{aq})] = 0.0025 \text{ mol/L}$; $V_{\text{KOH}} = 300.0 \text{ mL}$;

$[\text{KOH}(\text{aq})] = 0.0015 \text{ mol/L}$

Required: To predict whether calcium hydroxide, $\text{Ca}(\text{OH})_2(\text{s})$, will precipitate

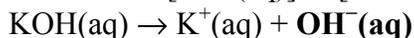
Solution:

Step 1. The initial $\text{CaCl}_2(\text{aq})$ solution and initial $\text{KOH}(\text{aq})$ solution will contribute these four ions to the mixture: $\text{Ca}^{2+}(\text{aq})$, $\text{Cl}^{-}(\text{aq})$, $\text{K}^{+}(\text{aq})$, and $\text{OH}^{-}(\text{aq})$.

Step 2. Determine the concentrations of the ions that may react to form a precipitate.



The ratio of $[\text{Ca}^{2+}(\text{aq})]$ to $[\text{CaCl}_2(\text{aq})]$ is 1:1.



The ratio of $[\text{OH}^{-}(\text{aq})]$ to $[\text{KOH}(\text{aq})]$ is 1:1.

Therefore,

$$[\text{Ca}^{2+}(\text{aq})] = 0.0025 \text{ mol/L} \text{ and } [\text{OH}^{-}(\text{aq})] = 0.0015 \text{ mol/L}$$

These concentrations are the concentration of $\text{Ca}^{2+}(\text{aq})$ ions in the initial $\text{CaCl}_2(\text{aq})$ solution and the concentration of $\text{OH}^{-}(\text{aq})$ in the initial $\text{KOH}(\text{aq})$ solution *before* these solutions are mixed. When the initial solutions are mixed, the volume of the final mixture, V_{f} , increases and is the sum of the volumes of the two initial solutions.

$$V_{\text{f}} = V_{\text{CaCl}_2} + V_{\text{KOH}}$$

$$= 250.0 \text{ mL} + 300.0 \text{ mL}$$

$$V_{\text{f}} = 550.0 \text{ mL}$$

Use the dilution formula to calculate the concentrations of $\text{Ca}^{2+}(\text{aq})$ and $\text{OH}^{-}(\text{aq})$ in the final mixture.

$$c_c V_c = c_d V_d$$

$$c_d = \frac{c_c V_c}{V_d}$$

where c_c is the initial, concentrated, ion concentration and c_d is the final, more dilute, ion concentration. V_c and V_d are the corresponding volumes.

Substitute the appropriate values for $\text{Ca}^{2+}(\text{aq})$ into the dilution equation.

$$[\text{Ca}^{2+}(\text{aq})]_{\text{final}} = (0.0025 \text{ mol/L}) \left(\frac{250.0 \text{ mL}}{550.0 \text{ mL}} \right)$$

$$[\text{Ca}^{2+}(\text{aq})]_{\text{final}} = 0.0011 \text{ mol/L}$$

Then, substitute the appropriate values for $\text{OH}^{-}(\text{aq})$ into the dilution equation.

$$[\text{OH}^{-}(\text{aq})]_{\text{final}} = (0.0015 \text{ mol/L}) \left(\frac{300.0 \text{ mL}}{550.0 \text{ mL}} \right)$$

$$[\text{OH}^{-}(\text{aq})]_{\text{final}} = 0.00082 \text{ mol/L}$$

Step 3. Determine the trial ion product equation of the precipitate that may form.



$$Q = [\text{Ca}^{2+}(\text{aq})][\text{OH}^{-}(\text{aq})]^2$$

Step 4. Calculate Q .

$$Q = [\text{Ca}^{2+}(\text{aq})][\text{OH}^{-}(\text{aq})]^2$$

$$= (0.0011)(0.00082)^2$$

$$Q = 7.4 \times 10^{-10}$$

Step 5. Compare the K_{sp} and Q values of the compound that may precipitate, and then predict whether the position of this dissolution equilibrium system will shift and whether a precipitate will form.

From Table 1 in Appendix B4,

$$K_{\text{sp}} \text{ of } \text{Ca}(\text{OH})_2(\text{s}) = 5.5 \times 10^{-6}$$

$$Q = 7.4 \times 10^{-10}$$

Since Q is less than K_{sp} , no precipitation occurs.

Statement: A calcium hydroxide precipitate will not form when these solutions are mixed.

4. Given: $V_{\text{CaCl}_2} = 100.0 \text{ mL}$; $[\text{CaCl}_2(\text{aq})] = 0.100 \text{ mol/L}$; $V_{\text{Na}_2\text{SO}_4} = 100.0 \text{ mL}$;

$$[\text{Na}_2\text{SO}_4(\text{aq})] = 0.0400 \text{ mol/L}$$

Required: To predict whether a precipitate will form and identify any precipitate

Solution:

Step 1. The initial $\text{CaCl}_2(\text{aq})$ solution and initial $\text{Na}_2\text{SO}_4(\text{aq})$ solution will contribute these four ions to the mixture: $\text{Ca}^{2+}(\text{aq})$, $\text{Cl}^{-}(\text{aq})$, $\text{Na}^{+}(\text{aq})$, and $\text{SO}_4^{2-}(\text{aq})$.

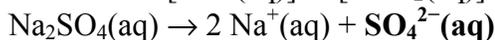
Step 2. Use the solubility rules in Table 3 on page 465 to predict whether a precipitate may form.

Table 3 indicates that NaCl is highly soluble and CaSO₄(s) has low solubility. Therefore, a precipitate of solid calcium sulfate, CaSO₄(s), may form.

Step 3. Determine the concentrations of the ions that may react to form a precipitate.



The ratio of [Ca²⁺(aq)] to [CaCl₂(aq)] is 1:1.



The ratio of [SO₄²⁻(aq)] to [Na₂SO₄(aq)] is 1:1.

Therefore, [Ca²⁺(aq)] = 0.100 mol/L and [SO₄²⁻(aq)] = 0.0400 mol/L.

These concentrations are the concentration of Ca²⁺(aq) ions in the initial CaCl₂(aq) solution and the concentration of SO₄²⁻(aq) in the initial Na₂SO₄(aq) solution *before* these solutions are mixed.

When the initial solutions are mixed, the volume of the final mixture, V_f , increases and is the sum of the volumes of the two initial solutions.

$$\begin{aligned} V_f &= V_{\text{CaCl}_2} + V_{\text{Na}_2\text{SO}_4} \\ &= 100.0 \text{ mL} + 100.0 \text{ mL} \end{aligned}$$

$$V_f = 200.0 \text{ mL}$$

Use the dilution formula to calculate the concentrations of Ca²⁺(aq) and SO₄²⁻(aq) in the final mixture.

$$c_c V_c = c_d V_d$$

$$c_d = \frac{c_c V_c}{V_d}$$

where c_c is the initial, concentrated, ion concentration and c_d is the final, more dilute, ion concentration. V_c and V_d are the corresponding volumes.

Substitute the appropriate values for Ca²⁺(aq) into the dilution equation.

$$[\text{Ca}^{2+}(\text{aq})]_{\text{final}} = (0.100 \text{ mol/L}) \left(\frac{100.0 \text{ mL}}{200.0 \text{ mL}} \right)$$

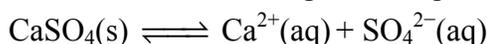
$$[\text{Ca}^{2+}(\text{aq})]_{\text{final}} = 0.0500 \text{ mol/L}$$

Then, substitute the appropriate values for SO₄²⁻(aq) into the dilution equation.

$$[\text{SO}_4^{2-}(\text{aq})]_{\text{final}} = (0.0400 \text{ mol/L}) \left(\frac{100.0 \text{ mL}}{200.0 \text{ mL}} \right)$$

$$[\text{SO}_4^{2-}(\text{aq})]_{\text{final}} = 0.0200 \text{ mol/L}$$

Step 4. Determine the trial ion product equation of the precipitate that may form.



$$Q = [\text{Ca}^{2+}(\text{aq})][\text{SO}_4^{2-}(\text{aq})]$$

Step 5. Calculate Q .

$$Q = [\text{Ca}^{2+}(\text{aq})][\text{SO}_4^{2-}(\text{aq})]$$

$$= (0.0500)(0.0200)$$

$$Q = 1.00 \times 10^{-3}$$

Step 6. Compare the K_{sp} and Q values of the compound that may precipitate, and then predict whether the position of this dissolution equilibrium system will shift and whether a precipitate will form.

From Table 1 in Appendix B4,

$$K_{sp} \text{ of } \text{CaSO}_4(\text{s}) = 7.1 \times 10^{-5}$$

$$Q = 1.00 \times 10^{-3}$$

Since Q is greater than K_{sp} , the dissolution equilibrium system shifts to the left and precipitation occurs.

Statement: A precipitate of solid calcium sulfate, $\text{CaSO}_4(\text{s})$, will form when these solutions are mixed.

Tutorial 4 Practice, page 470

1. Silver acetate, $\text{AgC}_2\text{H}_3\text{O}_2(\text{s})$, will have low solubility in pure water, $\text{H}_2\text{O}(\text{l})$, because from Table 3 on page 465 of the Student Book, silver acetate has low solubility in water. In 0.001 mol/L silver nitrate, $\text{AgNO}_3(\text{aq})$, silver acetate will have lower solubility than in water due to the common ion effect of silver ions.

In 0.01 mol/L silver nitrate, silver acetate will have lower solubility than in 0.001 mol/L silver nitrate due to increased common ion effect of silver ions.

In 0.001 mol/L sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2(\text{aq})$, the solubility of silver acetate will be similar to its solubility in 0.001 mol/L silver nitrate due to the common ion effect of acetate ions.

In 0.01 mol/L sodium nitrate, $\text{NaNO}_3(\text{aq})$, the solubility of silver acetate will be similar to its solubility in water because there are no common ions:

2. (a) **Given:** K_{sp} of $\text{AgCl}(\text{s}) = 1.8 \times 10^{-10}$

Required: molar solubility of $\text{AgCl}(\text{s})$ in pure water

Solution:

Step 1. Write the balanced equation for the solution equilibrium.



Step 2. Write the solubility product constant equation.

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

Step 3. Determine the equilibrium concentrations in mol/L of the dissolved ions. Use an ICE table. Since the concentration of solid AgCl remains constant, it is not included.

	$\text{AgCl}(\text{s})$	\rightleftharpoons	$\text{Ag}^+(\text{aq})$	+	$\text{Cl}^-(\text{aq})$
I	—		0.00		0.00
C	—		+x		+x
E	—		x		x

Step 4. Substitute the equilibrium concentration expressions from the ICE table into the K_{sp} equation and solve for x .

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

$$K_{sp} = (x)(x)$$

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

$$x = \sqrt{1.8 \times 10^{-10}}$$

$$x = 1.3 \times 10^{-5}$$

Statement: The molar solubility of $\text{AgCl}(\text{s})$ in pure water is 1.3×10^{-5} mol/L.

(b) Given: K_{sp} of $\text{AgCl}(\text{s}) = 1.8 \times 10^{-10}$

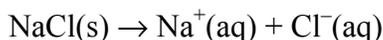
Required: molar solubility of $\text{AgCl}(\text{s})$ in 0.10 mol/L $\text{NaCl}(\text{aq})$

Solution:

Step 1. Identify the common ion.

The ion present in both $\text{AgCl}(\text{s})$ and $\text{NaCl}(\text{s})$ is the chloride ion, $\text{Cl}^-(\text{aq})$.

Step 2. Determine the concentration in mol/L of the common ion, $\text{Cl}^-(\text{aq})$, in the $\text{NaCl}(\text{aq})$ solution before $\text{AgCl}(\text{s})$ was added. Table 3 on page 465 indicates that $\text{NaCl}(\text{s})$ has high solubility and dissolves in water to form $\text{Na}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions.



Since the ions dissolve in a 1:1 molar ratio, the concentrations of sodium and chloride ions are equal to the concentration of the $\text{NaCl}(\text{aq})$ solution.

$$[\text{Na}^+(\text{aq})] = 0.10 \text{ mol/L before adding } \text{AgCl}(\text{s})$$

$$[\text{Cl}^-(\text{aq})] = 0.10 \text{ mol/L before adding } \text{AgCl}(\text{s})$$

Step 3. Determine the concentration in mol/L of the common ion once the solid has dissolved.

Table 3 indicates that $\text{AgCl}(\text{s})$ has a low solubility in water at 25 °C. Therefore, it will establish a dynamic equilibrium in solution.



The $\text{AgCl}(\text{s})$ is being dissolved in solution that already contains some $\text{Cl}^-(\text{aq})$ ions. Therefore, the $\text{AgCl}(\text{s})$ equilibrium will shift to the left relative to its position in pure water and therefore less $\text{AgCl}(\text{s})$ will dissolve in the $\text{NaCl}(\text{aq})$ solution than in pure water. The final concentration of $\text{Cl}^-(\text{aq})$ ions will be equal to the initial concentration (0.10 mol/L) plus the change in concentration caused by adding $\text{AgCl}(\text{s})$ to the solution. Summarize these changes in an ICE table.

	$\text{AgCl}(\text{s})$	\rightleftharpoons	$\text{Ag}^+(\text{aq})$	+	$\text{Cl}^-(\text{aq})$
I	—		0.00		0.10
C	—		+ x		+ x
E	—		x		$x + 0.10$

Step 4. Write the equation for the solubility product, K_{sp} , and substitute in the value of K_{sp} for $\text{AgCl}(s)$ from Table 1 in Appendix B4, and the equilibrium concentration expressions from the ICE table. Solve for x .

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

$$1.8 \times 10^{-10} = (x)(x + 0.10)$$

$$x = \frac{1.8 \times 10^{-10}}{x + 0.10}$$

Since $\text{AgCl}(s)$ has a very low solubility, the amount of $\text{Cl}^-(\text{aq})$ ions that the $\text{AgCl}(s)$ will add to the solution will be very small compared to the 0.10 mol/L provided by the $\text{NaCl}(\text{aq})$ solution. Therefore, make the simplifying assumption that $x + 0.10 \approx 0.10$.

$$x \approx \frac{1.8 \times 10^{-10}}{0.10}$$

$$x \approx 1.8 \times 10^{-9}$$

Since x represents the amount of silver ions that can dissolve in this equilibrium system and there is a 1:1 ratio of silver ions to solid silver chloride, then x is also the solubility of $\text{AgCl}(s)$. Since this value is 8 orders of magnitude smaller than the initial concentration of $\text{Cl}^-(\text{aq})$, the simplifying assumption was warranted.

Statement: The molar solubility of silver chloride in a 0.10 mol/L aqueous solution of sodium chloride is 1.8×10^{-9} mol/L.

(c) The answers to (a) and (b) make sense because the common chloride ion reduces the solubility of silver chloride.

3. Given: 1.00 L of 0.200 mol/L $\text{LiCl}(\text{aq})$

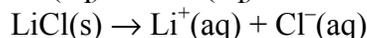
Required: amount of $\text{PbCl}_2(s)$ that can be dissolved

Solution:

Step 1. Identify the common ion.

The ion present in both $\text{PbCl}_2(s)$ and $\text{LiCl}(s)$ is the chloride ion, $\text{Cl}^-(\text{aq})$.

Step 2. Determine the concentration in mol/L of the common ion, $\text{Cl}^-(\text{aq})$, in the $\text{LiCl}(\text{aq})$ solution before $\text{PbCl}_2(s)$ was added. Table 3 on page 465 of the Student Book indicates that $\text{LiCl}(s)$ has high solubility and dissolves in water to form $\text{Li}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions.



Since the ions dissolve in a 1:1 molar ratio, the concentrations of lithium and chloride ions are equal to the concentration of the $\text{LiCl}(\text{aq})$ solution.

$$[\text{Li}^+(\text{aq})] = 0.200 \text{ mol/L before adding } \text{PbCl}_2(s)$$

$$[\text{Cl}^-(\text{aq})] = 0.200 \text{ mol/L before adding } \text{PbCl}_2(s)$$

Step 3. Determine the concentration in mol/L of the common ion once the solid has dissolved.

Table 3 indicates that $\text{PbCl}_2(\text{s})$ has a low solubility in water at 25°C . Therefore, it will establish a dynamic equilibrium in solution.



The $\text{PbCl}_2(\text{s})$ is being dissolved in solution that already contains some $\text{Cl}^{-}(\text{aq})$ ions. Therefore, the $\text{PbCl}_2(\text{s})$ equilibrium will shift to the left relative to its position in pure water and therefore less $\text{PbCl}_2(\text{s})$ will dissolve in the $\text{LiCl}(\text{aq})$ solution than in pure water. The final concentration of $\text{Cl}^{-}(\text{aq})$ ions will be equal to the initial concentration (0.200 mol/L) plus the change in concentration caused by adding $\text{PbCl}_2(\text{s})$ to the solution. Summarize these changes in an ICE table.

	$\text{PbCl}_2(\text{s})$	\rightleftharpoons	$\text{Pb}^{2+}(\text{aq})$	+	$2 \text{Cl}^{-}(\text{aq})$
I	---		0.00		0.200
C	---		+x		+2x
E	---		x		2x + 0.200

Step 4. Write the equation for the solubility product, K_{sp} .

$$K_{\text{sp}} = [\text{Pb}^{2+}(\text{aq})][\text{Cl}^{-}(\text{aq})]^2$$

From Table 1 in Appendix B4,

$$K_{\text{sp}} \text{ of } \text{PbCl}_2(\text{s}) = 1.2 \times 10^{-5}$$

Substitute into the equation for the solubility product the value of K_{sp} for $\text{AgCl}(\text{s})$ and the equilibrium concentration expressions from the ICE table. Solve for x .

$$K_{\text{sp}} = [\text{Pb}^{2+}(\text{aq})][\text{Cl}^{-}(\text{aq})]^2$$

$$1.2 \times 10^{-5} = (x)(2x + 0.200)^2$$

$$x = \frac{1.2 \times 10^{-5}}{(2x + 0.200)^2}$$

Since for chloride ions, the solubility of $\text{PbCl}_2(\text{s})$ is low, the amount of $\text{Cl}^{-}(\text{aq})$ ions that the $\text{PbCl}_2(\text{s})$ will add to the solution will be very small compared to the 0.200 mol/L provided by the $\text{LiCl}(\text{aq})$ solution. Therefore, make the simplifying assumption that $2x + 0.200 \approx 0.200$.

$$x \approx \frac{1.2 \times 10^{-5}}{(0.200)^2}$$

$$x \approx 3.0 \times 10^{-4} \text{ mol/L}$$

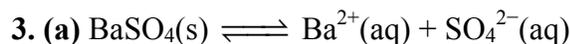
Since x represents the amount of lead ions that can dissolve in this equilibrium system and there is a 1:1 ratio of lead ions to solid lead(II) chloride, then x is also the solubility of $\text{PbCl}_2(\text{s})$. Since this value is 3 orders of magnitude smaller than the initial concentration of $\text{Cl}^{-}(\text{aq})$, the simplifying assumption was warranted.

Statement: Approximately 3.0×10^{-4} mol/L of solid lead(II) chloride will dissolve in this solution. Since the volume of the solution is 1.00 L, 3.0×10^{-4} mol $\text{PbCl}_2(\text{s})$ will dissolve.

Section 7.6 Questions, page 471

1. The molar solubility is the amount of material that will dissolve in 1 mol of solvent. The product solubility constant, K_{sp} , is the equilibrium constant for the dissolution reaction.

2. A common ion is an ion that is present in two different solutes. Its presence decreases the solubility of the less-soluble substance because it shifts the equilibrium toward the solid.



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

4. Both solid silver chloride, $\text{AgCl}(\text{s})$, and solid silver bromide, $\text{AgBr}(\text{s})$, ionize in a 1:1 ratio, so the compound with the smaller K_{sp} value, $\text{AgBr}(\text{s})$, is less soluble in water.

5. **Given:** solubility of $\text{AgI}(\text{s}) = 2.14 \times 10^{-7}$ g/100 mL at 25 °C

Required: K_{sp} of $\text{AgI}(\text{s})$ at 25 °C

Solution:

Step 1. Write the balanced equation.



Step 2. Write the solubility product constant equation.

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

Step 3. Determine the equilibrium concentrations in mol/L of the dissolved ions. Use an ICE table. Since the concentration of $\text{AgI}(\text{s})$ remains constant, it is not included.

	$\text{AgI}(\text{s})$	\rightleftharpoons	$\text{Ag}^+(\text{aq})$	+	$\text{I}^-(\text{aq})$
I	---		0.00		0.00
C	---		+x		+x
E	---		x		x

Step 4. Calculate all concentrations in mol/L. To convert given mass solubility of solid silver iodide to molar solubility, first calculate the molar mass of silver iodide. Then, use the molar mass to convert mass solubility to molar solubility.

$$\text{molar mass of AgI} = 107.87 \text{ g/mol} + 126.90 \text{ g/mol}$$

$$\text{molar mass of AgI} = 234.77 \text{ g/mol (two extra digits carried)}$$

$$\text{molar solubility of AgI}(\text{s}) = \frac{2.14 \times 10^{-7} \text{ g}}{100 \text{ mL}} \times \frac{1 \text{ mol}}{234.77 \text{ g}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$\text{molar solubility of AgI}(\text{s}) = 9.115 \times 10^{-9} \text{ mol/L (one extra digit carried)}$$

Step 5. Write the K_{sp} equation using the expressions in the ICE table, and then substitute the molar solubility of solid silver iodide for x .

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

$$= (x)(x)$$

$$= (9.115 \times 10^{-9})(9.115 \times 10^{-9})$$

$$K_{sp} = 8.31 \times 10^{-17}$$

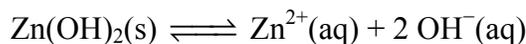
Statement: The solubility product constant of solid silver iodide is 8.31×10^{-17} .

6. Given: (from Table 1 in Appendix B4) K_{sp} of $Zn(OH)_2(s)$ at $25\text{ }^\circ\text{C} = 7.7 \times 10^{-17}$

Required: molar solubility of $Zn(OH)_2(s)$ at $25\text{ }^\circ\text{C}$

Solution:

Step 1. Write the balanced chemical equation.



Step 2. Write the solubility product constant equation.

$$K_{sp} = [Zn^{2+}(aq)][OH^-(aq)]^2$$

Step 3. Determine the equilibrium concentrations in mol/L of the dissolved ions. Use an ICE table. Since the concentration of $Zn(OH)_2(s)$ remains constant, it is not included.

	$Zn(OH)_2(s)$	\rightleftharpoons	$Zn^{2+}(aq)$	+	$2 OH^-(aq)$
I	—		0.00		0.00
C	—		+x		+2x
E	—		x		2x

Step 4. Substitute into the equation for the solubility product the value of K_{sp} for $Zn(OH)_2(s)$ and the equilibrium concentration expressions from the ICE table. Solve for x .

$$K_{sp} = [Zn^{2+}(aq)][OH^-(aq)]^2$$

$$K_{sp} = (x)(2x)^2$$

$$7.7 \times 10^{-17} = 4x^3$$

$$x^3 = 1.925 \times 10^{-17}$$

$$x = \sqrt[3]{1.925 \times 10^{-17}}$$

$$x = 2.7 \times 10^{-6}$$

$[Zn^{2+}(aq)] =$ molar solubility of $Zn(OH)_2(s)$

molar solubility of $Zn(OH)_2(s) = 2.7 \times 10^{-6}$ mol/L

Statement: The molar solubility of solid zinc hydroxide is 2.7×10^{-6} mol/L.

7. Given: $[Ca(NO_3)_2(aq)] = 0.01$ mol/L; $[Na_3PO_4(aq)] = 0.025$ mol/L

Required: To predict whether a precipitate will form and identify any precipitate

Solution:

Step 1. Identify all of the ions that will be present in the mixture.

$Ca(NO_3)_2(aq)$ will contribute aqueous ions of calcium, $Ca^{2+}(aq)$, and nitrate, $NO_3^-(aq)$.

$Na_3PO_4(aq)$ will contribute aqueous ions of sodium, $Na^+(aq)$, and phosphate, $PO_4^{3-}(aq)$.

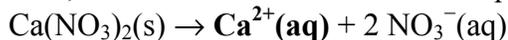
All four ions will be in the mixture: $Ca^{2+}(aq)$, $NO_3^-(aq)$, $Na^+(aq)$, and $PO_4^{3-}(aq)$.

Step 2. Use the solubility rules in Table 3 on page 465 to predict whether a precipitate may form.

Table 3 indicates that $NaNO_3(s)$ is highly soluble and $Ca_3(PO_4)_2(s)$ has low solubility. Therefore, a precipitate of solid calcium phosphate may form.

Step 3. Determine the concentrations of the ions that may react to form a precipitate.

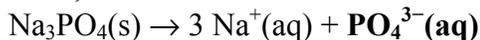
First, write the balanced dissolution equation for $\text{Ca}(\text{NO}_3)_2(\text{aq})$.



From the balanced equation, $\text{Ca}(\text{NO}_3)_2(\text{s})$ and $\text{Ca}^{2+}(\text{aq})$ have a 1:1 molar ratio.

Therefore, the concentrations of the solute and the ions are equal. Therefore, $[\text{Ca}^{2+}(\text{aq})] = 0.01 \text{ mol/L}$

Then, write the balanced dissolution equation for $\text{Na}_3\text{PO}_4(\text{aq})$.



From the balanced equation, $\text{Na}_3\text{PO}_4(\text{s})$ and $\text{PO}_4^{3-}(\text{aq})$ have a 1:1 molar ratio.

Therefore, the concentrations of the solute and the ions are equal. Therefore, $[\text{PO}_4^{3-}(\text{aq})] = 0.025 \text{ mol/L}$

Step 4. Determine the trial ion product equation of the precipitate that may form.

$\text{Ca}_3(\text{PO}_4)_2(\text{s})$ may precipitate. First, write a balanced equation for the dissolution equilibrium system of $\text{Ca}_3(\text{PO}_4)_2(\text{s})$.



From the balanced equation, the trial ion product equation is

$$Q = [\text{Ca}^{2+}(\text{aq})]^3 [\text{PO}_4^{3-}(\text{aq})]^2$$

Step 5. Calculate Q by substituting the concentrations of the dissolved ions in the mixture into the trial ion product equation.

$$Q = [\text{Ca}^{2+}(\text{aq})]^3 [\text{PO}_4^{3-}(\text{aq})]^2$$

$$= (0.01)^3 (0.025)^2$$

$$Q = 6.2 \times 10^{-10}$$

Step 6. Compare the K_{sp} and Q values of the compound that may precipitate, and then predict whether the position of this dissolution equilibrium system will shift and whether a precipitate will form.

From Table 1 in Appendix B4,

$$K_{\text{sp}} \text{ of } \text{Ca}_3(\text{PO}_4)_2(\text{s}) = 2.1 \times 10^{-33}$$

$$Q = 6.2 \times 10^{-10}$$

Q is greater than K_{sp} , so the dissolution equilibrium system will shift to the left and a precipitate of $\text{Ca}_3(\text{PO}_4)_2(\text{s})$ will form.

Statement: A precipitate of solid calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2(\text{s})$, will form when the two solutions are mixed.

8. Given: $V_{\text{Mg}(\text{NO}_3)_2} = 100.0 \text{ mL}$; $[\text{Mg}(\text{NO}_3)_2(\text{aq})] = 4.0 \times 10^{-4} \text{ mol/L}$;

$V_{\text{NaOH}} = 100.0 \text{ mL}$; $[\text{NaOH}(\text{aq})] = 2.0 \times 10^{-3} \text{ mol/L}$

Required: To predict whether a precipitate will form and identify any precipitate

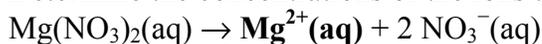
Solution:

Step 1. The initial $\text{Mg}(\text{NO}_3)_2(\text{aq})$ solution and initial $\text{NaOH}(\text{aq})$ solution will contribute these four ions to the mixture: $\text{Mg}^{2+}(\text{aq})$, $\text{NO}_3^{-}(\text{aq})$, $\text{Na}^{+}(\text{aq})$, and $\text{OH}^{-}(\text{aq})$.

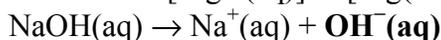
Step 2. Use the solubility rules in Table 3 on page 465 to predict whether a precipitate may form.

Table 3 indicates that NaNO_3 is highly soluble and $\text{Mg}(\text{OH})_2$ has low solubility. Therefore, a precipitate of solid magnesium hydroxide, $\text{Mg}(\text{OH})_2(\text{s})$, may form.

Step 3. Determine the concentrations of the ions that may react to form a precipitate.



The ratio of $[\text{Mg}^{2+}(\text{aq})]$ to $[\text{Mg}(\text{NO}_3)_2(\text{aq})]$ is 1:1.



The ratio of $[\text{OH}^{-}(\text{aq})]$ to $[\text{NaOH}(\text{aq})]$ is 1:1.

Therefore, $[\text{Mg}^{2+}(\text{aq})] = 4.0 \times 10^{-4} \text{ mol/L}$ and $[\text{OH}^{-}(\text{aq})] = 2.0 \times 10^{-3} \text{ mol/L}$.

These concentrations are the concentration of $\text{Mg}^{2+}(\text{aq})$ ions in the initial $\text{Mg}(\text{NO}_3)_2(\text{aq})$ solution and the concentration of $\text{OH}^{-}(\text{aq})$ in the initial $\text{NaOH}(\text{aq})$ solution *before* these solutions are mixed. When the initial solutions are mixed, the volume of the final mixture, V_f , increases and is the sum of the volumes of the two initial solutions.

$$\begin{aligned} V_f &= V_{\text{Mg}(\text{NO}_3)_2} + V_{\text{NaOH}} \\ &= 100.0 \text{ mL} + 100.0 \text{ mL} \end{aligned}$$

$$V_f = 200.0 \text{ mL}$$

Use the dilution formula to calculate the concentrations of $\text{Mg}^{2+}(\text{aq})$ and $\text{OH}^{-}(\text{aq})$ in the final mixture.

$$c_c V_c = c_d V_d$$

$$c_d = \frac{c_c V_c}{V_d}$$

where c_c is the initial, concentrated, ion concentration and c_d is the final, more dilute, ion concentration. V_c and V_d are the corresponding volumes.

Substitute the appropriate values for $\text{Mg}^{2+}(\text{aq})$ into the dilution equation.

$$[\text{Mg}^{2+}(\text{aq})]_{\text{final}} = (4.0 \times 10^{-4} \text{ mol/L}) \left(\frac{100.0 \text{ mL}}{200.0 \text{ mL}} \right)$$

$$[\text{Mg}^{2+}(\text{aq})]_{\text{final}} = 2.0 \times 10^{-4} \text{ mol/L}$$

Then, substitute the appropriate values for $\text{OH}^{-}(\text{aq})$ into the dilution equation.

$$[\text{OH}^{-}(\text{aq})]_{\text{final}} = (2.0 \times 10^{-3} \text{ mol/L}) \left(\frac{100.0 \text{ mL}}{200.0 \text{ mL}} \right)$$

$$[\text{OH}^{-}(\text{aq})]_{\text{final}} = 1.0 \times 10^{-3} \text{ mol/L}$$

Step 4. Determine the trial ion product equation of the precipitate that may form.



$$Q = [\text{Mg}^{2+}(\text{aq})][\text{OH}^{-}(\text{aq})]^2$$

Step 5. Calculate Q .

$$\begin{aligned} Q &= [\text{Mg}^{2+}(\text{aq})][\text{OH}^{-}(\text{aq})]^2 \\ &= (2.0 \times 10^{-4})(1.0 \times 10^{-3})^2 \end{aligned}$$

$$Q = 2.0 \times 10^{-10}$$

Step 6. Compare the K_{sp} and Q values of the compound that may precipitate, and then predict whether the position of this dissolution equilibrium system will shift and whether a precipitate will form.

From Table 1 in Appendix B4,

$$K_{sp} \text{ of } \text{Mg}(\text{OH})_2 = 5.6 \times 10^{-12}$$

$$Q = 2.0 \times 10^{-10}$$

Since Q is greater than K_{sp} , the dissolution equilibrium system shifts to the left and a precipitate will form.

Statement: A precipitate of solid magnesium hydroxide, $\text{Mg}(\text{OH})_2(\text{s})$, will form when the two solutions are mixed.

9. Given: K_{sp} of $\text{Ag}_2\text{CrO}_4(\text{s}) = 1.1 \times 10^{-12}$

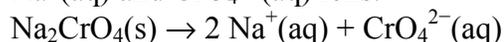
Required: molar solubility of $\text{Ag}_2\text{CrO}_4(\text{s})$ in 0.10 mol/L $\text{Na}_2\text{CrO}_4(\text{aq})$

Solution:

Step 1. Identify the common ion.

The ion present in both $\text{Ag}_2\text{CrO}_4(\text{s})$ and $\text{Na}_2\text{CrO}_4(\text{s})$ is the chromate ion, $\text{CrO}_4^{2-}(\text{aq})$.

Step 2. Determine the concentration in mol/L of the common ion, $\text{CrO}_4^{2-}(\text{aq})$, in the $\text{Na}_2\text{CrO}_4(\text{aq})$ solution before $\text{Ag}_2\text{CrO}_4(\text{s})$ was added. Table 3 on page 465 indicates that $\text{Na}_2\text{CrO}_4(\text{s})$ has high solubility and dissolves in water to form $\text{Na}^+(\text{aq})$ and $\text{CrO}_4^{2-}(\text{aq})$ ions.



Chromate in $\text{Na}_2\text{CrO}_4(\text{s})$ dissolves in a 1:1 ratio, so $[\text{CrO}_4^{2-}(\text{aq})] = 0.10 \text{ mol/L}$ before adding $\text{Ag}_2\text{CrO}_4(\text{s})$.

Step 3. Determine the concentration in mol/L of the common ion once the solid has dissolved.

Table 3 indicates that $\text{Ag}_2\text{CrO}_4(\text{s})$ has a low solubility in water at 25 °C.

Therefore, it will establish a dynamic equilibrium in solution.



The $\text{Ag}_2\text{CrO}_4(\text{s})$ is being dissolved in solution that already contains some $\text{CrO}_4^{2-}(\text{aq})$ ions. Therefore, the $\text{Ag}_2\text{CrO}_4(\text{s})$ equilibrium will shift to the left relative to its position in pure water and therefore less $\text{Ag}_2\text{CrO}_4(\text{s})$ will dissolve in the $\text{Na}_2\text{CrO}_4(\text{aq})$ solution than in pure water. The final concentration of $\text{CrO}_4^{2-}(\text{aq})$ ions will be equal to the initial concentration (0.10 mol/L) plus the change in concentration caused by adding $\text{Ag}_2\text{CrO}_4(\text{s})$ to the solution. Summarize these changes in an ICE table.

	$\text{Ag}_2\text{CrO}_4(\text{s})$	\rightleftharpoons	$2 \text{Ag}^+(\text{aq})$	+	$\text{CrO}_4^{2-}(\text{aq})$
I	—		0.00		0.10
C	—		+2x		+x
E	—		2x		x + 0.10

Step 4. Write the equation for the solubility product, K_{sp} , and substitute in the value of K_{sp} for $\text{AgCrO}_4(\text{s})$ from Table 1 in Appendix B4, and the equilibrium concentration expressions from the ICE table. Solve for x .

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

$$1.1 \times 10^{-12} = (2x)^2(x + 0.10)$$

$$1.1 \times 10^{-12} = 4x^2(x + 0.10)$$

$$x^2 = \frac{1.1 \times 10^{-12}}{4(x + 0.10)}$$

$$x = \sqrt{\frac{1.1 \times 10^{-12}}{4(x + 0.10)}}$$

For chromate ions, $\text{AgCrO}_4(\text{s})$ solubility is low, so the amount of $\text{CrO}_4^{2-}(\text{aq})$ ions that the $\text{AgCrO}_4(\text{s})$ will add to the solution will be very small compared to the 0.10 mol/L provided by the $\text{Na}_2\text{CrO}_4(\text{aq})$ solution. Therefore, make the simplifying assumption that $x + 0.10 \approx 0.10$.

$$x \approx \sqrt{\frac{1.1 \times 10^{-12}}{4(0.10)}}$$

$$x \approx 1.7 \times 10^{-6}$$

Statement: The molar solubility of silver chromate in a 0.10 mol/L aqueous solution of sodium chromate is 1.7×10^{-6} mol/L.

10. The solubility product is the product of the ion concentrations of a saturated solution. The trial ion product is the product of ion concentrations of a real or proposed solution.

11. (a) Answers may vary. Sample answer: Solutions that could be used to precipitate out the phosphate from the water are solutions of the salts of a non-Group 1 metal. Two examples of such salts are $\text{CaCl}_2(\text{aq})$ or $\text{Sr}(\text{NO}_3)_2(\text{aq})$.

(b) As predicted by Le Châtelier's principle, the addition of the cation causes a stress to the equilibrium system that is relieved by shifting the equilibrium toward the solid phosphate salt.

12. Answers may vary. Sample answer: The formation of stalactites and stalagmites in limestone caves is an example of solubility equilibrium in the natural environment. Calcium carbonate from limestone dissolves in ground water and forms an equilibrium between the dissociated and solid calcium carbonate. There is also an equilibrium between carbon dioxide, carbonate ions, and hydrogen carbonate ions. As carbon dioxide leaves the solution into the air, this equilibrium shifts in favour of increased carbonate ion concentration. When the carbonate ion concentration increases, equilibrium with calcium carbonate shifts in favour of solid calcium carbonate, which precipitates to form a solid. Solid calcium carbonate deposits hanging down from the tops of caves are stalactites. Where the solution drips to the ground and the precipitate forms there the deposits are called stalagmites.