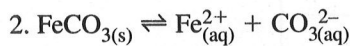


$$[\text{Ag}^+] = [\text{I}^-]$$

$$1.5 \times 10^{-16} = [\text{Ag}^+]^2$$

$$[\text{Ag}^+] = 1.2 \times 10^{-8} \text{ mol/L}$$

The solubility of silver iodide is  $1.2 \times 10^{-8} \text{ mol/L}$ .



$$K_{\text{sp}} = [\text{Fe}_{(\text{aq})}^{2+}][\text{CO}_{3(\text{aq})}^{2-}]$$

$$K_{\text{sp}} = 3.5 \times 10^{-11}$$

$$[\text{Fe}_{(\text{aq})}^{2+}] = [\text{CO}_{3(\text{aq})}^{2-}]$$

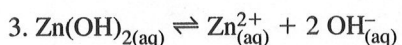
$$3.5 \times 10^{-11} = [\text{Fe}_{(\text{aq})}^{2+}]^2$$

$$[\text{Fe}_{(\text{aq})}^{2+}] = 5.9 \times 10^{-6} \text{ mol/L}$$

$$[\text{FeCO}_3(\text{aq})] = [\text{Fe}_{(\text{aq})}^{2+}]$$

$$[\text{FeCO}_3(\text{aq})] = 5.9 \times 10^{-6} \text{ mol/L}$$

The solubility of iron(II) carbonate is  $5.9 \times 10^{-6} \text{ mol/L}$ .



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

$$K_{\text{sp}} = 4.5 \times 10^{-17}$$

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

$$4.5 \times 10^{-17} = [\text{Zn}_{(\text{aq})}^{2+}][2 \text{Zn}_{(\text{aq})}^{2+}]^2$$

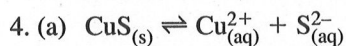
$$4.5 \times 10^{-17} = 4[\text{Zn}_{(\text{aq})}^{2+}]^3$$

$$[\text{Zn}_{(\text{aq})}^{2+}] = 2.2 \times 10^{-6}$$

$$[\text{Zn}(\text{OH})_2(\text{aq})] = [\text{Zn}_{(\text{aq})}^{2+}]$$

$$[\text{Zn}(\text{OH})_2(\text{aq})] = 2.2 \times 10^{-6} \text{ mol/L}$$

The solubility of zinc hydroxide is  $2.2 \times 10^{-6} \text{ mol/L}$ .



ICE Table for Calculating $K_{\text{sp}}$ from Solubility			
	$\text{CuS}(\text{s}) \rightleftharpoons$	$\text{Cu}_{(\text{aq})}^{2+} +$	$\text{S}_{(\text{aq})}^{2-}$
Initial concentration (mol/L)		0	0
Change in concentration (mol/L)		+x	+x
Equilibrium concentration (mol/L)		x	x

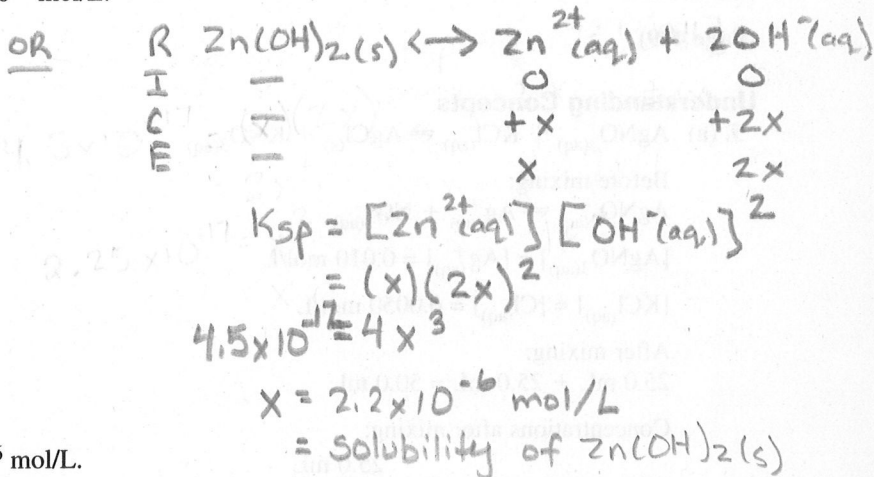
$$x = 8.89 \times 10^{-19} \text{ mol/L}$$

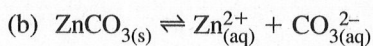
$$K_{\text{sp}} = [\text{Cu}_{(\text{aq})}^{2+}][\text{S}_{(\text{aq})}^{2-}]$$

$$= [8.89 \times 10^{-19}][8.89 \times 10^{-19}]$$

$$K_{\text{sp}} = 7.90 \times 10^{-37}$$

The solubility product of copper(II) sulfide has a value of  $7.90 \times 10^{-37}$ .





ICE Table for Calculating $K_{sp}$ from Solubility			
	$\text{ZnCO}_3(s) \rightleftharpoons$	$\text{Zn}^{2+}_{(aq)} +$	$\text{CO}_3^{2-}_{(aq)}$
Initial concentration (mol/L)		0	0
Change in concentration (mol/L)		+x	+x
Equilibrium concentration (mol/L)		x	x

$$x = 3.87 \times 10^{-6} \text{ mol/L}$$

$$K_{sp} = [\text{Zn}^{2+}_{(aq)}][\text{CO}_3^{2-}_{(aq)}]$$

$$= [3.87 \times 10^{-6}][3.87 \times 10^{-6}]$$

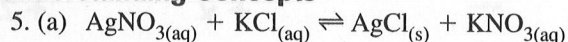
$$K_{sp} = 1.50 \times 10^{-11}$$

The solubility product of zinc carbonate has a value of  $1.50 \times 10^{-11}$ .

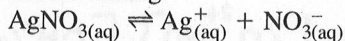
## PRACTICE

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### Understanding Concepts



Before mixing:



$$[\text{AgNO}_3(aq)] = [\text{Ag}^+_{(aq)}] = 0.010 \text{ mol/L}$$

$$[\text{KCl}(aq)] = [\text{Cl}^-_{(aq)}] = 0.0050 \text{ mol/L}$$

After mixing:

$$25.0 \text{ mL} + 25.0 \text{ mL} = 50.0 \text{ mL}$$

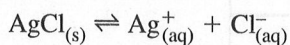
Concentrations after mixing:

$$[\text{Ag}^+_{(aq)}] = 0.010 \text{ mol/L} \times \frac{25.0 \text{ mL}}{50.0 \text{ mL}}$$

$$[\text{Ag}^+_{(aq)}] = 0.005 \text{ mol/L}$$

$$[\text{Cl}^-_{(aq)}] = 0.0050 \text{ mol/L} \times \frac{25.0 \text{ mL}}{50.0 \text{ mL}}$$

$$[\text{Cl}^-_{(aq)}] = 0.0025 \text{ mol/L}$$



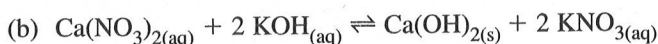
$$Q = [\text{Ag}^+_{(aq)}][\text{Cl}^-_{(aq)}]$$

$$= (0.005 \text{ mol/L})(0.0025 \text{ mol/L})$$

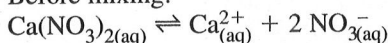
$$Q = 1.25 \times 10^{-5}$$

$$K_{sp} = 1.8 \times 10^{-10}$$

$Q$  is larger than  $K_{sp}$ . Therefore, a precipitate does form.



Before mixing:



$$[\text{Ca}(\text{NO}_3)_2(\text{aq})] = [\text{Ca}^{2+}_{(\text{aq})}] = 0.0010 \text{ mol/L}$$

$$[\text{KOH}(\text{aq})] = [\text{OH}^{-}_{(\text{aq})}] = 0.0020 \text{ mol/L}$$

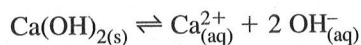
Concentrations after mixing:

$$[\text{Ca}^{2+}_{(\text{aq})}] = \frac{0.0010 \text{ mol/L}}{2}$$

$$[\text{Ca}^{2+}_{(\text{aq})}] = 0.0005 \text{ mol/L}$$

$$[\text{OH}^{-}_{(\text{aq})}] = \frac{0.0020 \text{ mol/L}}{2}$$

$$[\text{OH}^{-}_{(\text{aq})}] = 0.0010 \text{ mol/L}$$



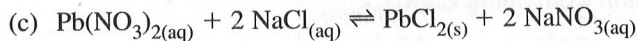
$$Q = [\text{Ca}^{2+}_{(\text{aq})}][\text{OH}^{-}_{(\text{aq})}]^2$$

$$= (0.0005 \text{ mol/L})(0.0010 \text{ mol/L})^2$$

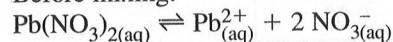
$$Q = 5.0 \times 10^{-10}$$

$$K_{\text{sp}} = 7.9 \times 10^{-6}$$

$Q$  is <sup>smaller</sup> ~~larger~~ than  $K_{\text{sp}}$ . Therefore, a precipitate does <sup>not</sup> form.



Before mixing:



$$[\text{Pb}(\text{NO}_3)_2(\text{aq})] = [\text{Pb}^{2+}_{(\text{aq})}] = 0.010 \text{ mol/L}$$

$$[\text{NaCl}(\text{aq})] = [\text{Cl}^{-}_{(\text{aq})}] = 0.10 \text{ mol/L}$$

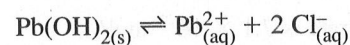
Concentrations after mixing:

$$[\text{Pb}^{2+}_{(\text{aq})}] = \frac{0.010 \text{ mol/L}}{2}$$

$$[\text{Pb}^{2+}_{(\text{aq})}] = 0.005 \text{ mol/L}$$

$$[\text{Cl}^{-}_{(\text{aq})}] = \frac{0.10 \text{ mol/L}}{2}$$

$$[\text{Cl}^{-}_{(\text{aq})}] = 0.050 \text{ mol/L}$$



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

$$= (0.005 \text{ mol/L})(0.050 \text{ mol/L})^2$$

$$Q = 1.2 \times 10^{-5}$$

$$K_{\text{sp}} = 1.2 \times 10^{-5}$$

$Q = K_{\text{sp}}$ . Therefore, a precipitate just begins to form.

### Making Connections

6. The concentration of barium ions in the solution is so low ( $1.0 \times 10^{-5} \text{ mol/L}$ ) that the suspension can be safely consumed.

## PRACTICE

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### Understanding Concepts

7.

ICE Table for the Solubility of Silver Chloride in a Solution Containing NaCl <sub>(aq)</sub>			
	AgCl <sub>(s)</sub> ⇌	Ag <sup>+</sup> <sub>(aq)</sub> +	Cl <sup>-</sup> <sub>(aq)</sub>
Initial concentration (mol/L)	—	0.0	0.10
Change in concentration (mol/L)	—	+x	+x
Equilibrium concentration (mol/L)	—	x	0.10 + x

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

$$K_{sp} = 1.8 \times 10^{-10}$$

$$0.10 + x \doteq 0.10$$

Therefore,

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

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

The solubility of silver chloride in a 0.10 mol/L solution of sodium chloride is  $1.8 \times 10^{-9}$  mol/L.

8.

ICE Table for the Solubility of CaSO <sub>4</sub> in a Solution Containing Ca(NO <sub>3</sub> ) <sub>2(aq)</sub>			
	CaSO <sub>4(s)</sub> ⇌	Ca <sup>2+</sup> <sub>(aq)</sub> +	SO <sub>4</sub> <sup>2-</sup> <sub>(aq)</sub>
Initial concentration (mol/L)	—	0.010	0.00
Change in concentration (mol/L)	—	+x	+x
Equilibrium concentration (mol/L)	—	0.010 + x	x

$$K_{sp} = [Ca^{2+}_{(aq)}][SO_4^{2-}_{(aq)}]$$

$$K_{sp} = 7.1 \times 10^{-5}$$

$$0.010 + x \doteq 0.010$$

Therefore,

$$(0.010)(x) = 7.1 \times 10^{-5}$$

$$x = 7.1 \times 10^{-3}$$

The solubility of calcium sulfate in a 0.010 mol/L solution of calcium nitrate is  $7.1 \times 10^{-3}$  mol/L.

9. (a)

ICE Table for the Solubility of Ag <sub>2</sub> CrO <sub>4</sub> in Water			
	Ag <sub>2</sub> CrO <sub>4(s)</sub> ⇌	2 Ag <sup>+</sup> <sub>(aq)</sub> +	CrO <sub>4</sub> <sup>2-</sup> <sub>(aq)</sub>
Initial concentration (mol/L)	—	—	—
Change in concentration (mol/L)	—	+2x	+x
Equilibrium concentration (mol/L)	—	2x	x

$$K_{sp} = [Ag^+_{(aq)}]^2[CrO_4^{2-}_{(aq)}]$$

$$K_{sp} = 1.1 \times 10^{-12}$$



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

$$4x^3 = 1.1 \times 10^{-12}$$

$$x = 6.5 \times 10^{-5}$$

The solubility of silver chromate is  $6.5 \times 10^{-5}$  mol/L.

(b)

ICE Table for the Solubility of $\text{Ag}_2\text{CrO}_4$ in a Solution of 0.1 mol/L Sodium Chromate			
	$\text{Ag}_2\text{CrO}_{4(s)} \rightleftharpoons$	$2 \text{Ag}_{(aq)}^+$	$\text{CrO}_{4(aq)}^{2-}$
Initial concentration (mol/L)		0.0	0.1
Change in concentration (mol/L)		+2x	+x
Equilibrium concentration (mol/L)		2x	x + 0.1

$$K_{sp} = [\text{Ag}_{(aq)}^+]^2[\text{CrO}_{4(aq)}^{2-}]$$

$$K_{sp} = 1.1 \times 10^{-12}$$

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

Assume  $x + 0.10 \doteq 0.10$

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

$$4x^2 = 1.1 \times 10^{-11}$$

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

The solubility of silver chromate is  $1.7 \times 10^{-6}$  mol/L.

(c) The solubility of silver chromate in a solution containing the chromate ion should be lower than its solubility in water. Presence of chromate already in solution shifts the equilibrium to the left, decreasing the amount of silver chromate that dissolves.

- Soluble compounds containing either barium or sulfate ions would decrease the solubility of barium sulfate. Examples include barium nitrate,  $\text{Ba}(\text{NO}_3)_2$ , and potassium sulfate,  $\text{K}_2\text{SO}_4$ .
- Soluble compounds containing either copper(II) or carbonate ions would decrease the solubility of copper(II) carbonate. Examples include copper(II) sulfate,  $\text{CuSO}_4$ , and sodium carbonate,  $\text{Na}_2\text{CO}_3$ .
- The reaction of silver ions with thiosulfate removes silver ions from the equilibrium, prompting the forward reaction to predominate, consuming silver chloride.

## SECTION 7.6 QUESTIONS

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### Understanding Concepts

- The solubility of a salt is the amount of salt that dissolves in a given amount of solvent to give a saturated solution. Solubility product constant is the product of the molar concentrations of the ions in the saturated solution.
- (a) The common ion effect is a reduction in the solubility of a salt caused by the presence of another salt having a common ion.  
(b) For the general dissociation equilibrium equation  $\text{B}_b\text{C}_c(s) \rightleftharpoons b\text{B}_{(aq)}^+ + c\text{C}_{(aq)}^-$ , the presence of a common ion  $\text{B}_{(aq)}^+$  or  $\text{C}_{(aq)}^-$  will shift the equilibrium to the left, reducing the solubility of the salt  $\text{B}_b\text{C}_c$ .
- When the solutions combine, there are more ions in solution than necessary to achieve saturation, resulting in an unstable supersaturated solution. The excess ions precipitate from solution, leaving behind a saturated solution.
- 

ICE Table for the Solubility of Barium Sulfate			
	$\text{BaSO}_{4(s)} \rightleftharpoons$	$\text{Ba}_{(aq)}^{2+} +$	$\text{SO}_{4(aq)}^{2-}$
Initial concentration (mol/L)	—	0	0
Change in concentration (mol/L)	—	+x	+x
Equilibrium concentration (mol/L)	—	x	x

$$K_{sp} = [\text{Ba}_{(aq)}^{2+}][\text{SO}_{4(aq)}^{2-}]$$

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

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

$$x = 1.0 \times 10^{-5}$$

The molar solubility of barium sulfate is  $1.0 \times 10^{-5}$  mol/L.

5.

ICE Table for the Solubility of Silver Bromide			
	$\text{AgBr}_{(s)} \rightleftharpoons$	$\text{Ag}_{(aq)}^{+} +$	$\text{Br}_{(aq)}^{-}$
Initial concentration (mol/L)	—	0	0
Change in concentration (mol/L)	—	+x	+x
Equilibrium concentration (mol/L)	—	x	x

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

$$K_{sp} = 5.4 \times 10^{-13}$$

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

$$x = 7.35 \times 10^{-7} \quad (\text{extra digits carried})$$

molar solubility of silver bromide =  $7.35 \times 10^{-7}$  mol/L

Solubility in g/100 mL ...

mass of silver bromide in 1 L

$$= 7.35 \times 10^{-7} \text{ mol/L} \times 187.77 \text{ g/mol}$$

$$= 1.38 \times 10^{-4} \text{ g/L}$$

mass of silver bromide in 100 mL

$$m = 1.38 \times 10^{-4} \text{ g} \times \frac{100 \text{ mL}}{1000 \text{ mL}}$$

$$m = 1.4 \times 10^{-5} \text{ g/100 mL}$$

The solubility of silver bromide is  $1.4 \times 10^{-5}$  g/100 mL.

6.

ICE Table for the Solubility of Strontium Fluoride			
	$\text{SrF}_{2(s)} \rightleftharpoons$	$\text{Sr}_{(aq)}^{2+} +$	$2 \text{F}_{(aq)}^{-}$
Initial concentration (mol/L)	—	0	0
Change in concentration (mol/L)	—	+x	+2x
Equilibrium concentration (mol/L)	—	x	2x

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

$$K_{sp} = 4.3 \times 10^{-9}$$

$$(x)(2x)^2 = 4.3 \times 10^{-9}$$

$$4x^3 = 4.3 \times 10^{-9}$$

$$x = 1.0 \times 10^{-3}$$

$$[\text{F}_{(aq)}^{-}] = 2 \times 1.0 \times 10^{-3}$$

$$[\text{F}_{(aq)}^{-}] = 2.0 \times 10^{-3} \text{ mol/L}$$

The fluoride ion concentration in a saturated solution of strontium fluoride is  $2.0 \times 10^{-3}$  mol/L.

7. If 2.4 g of  $\text{TlCl}$  dissolves in 100 mL, then 24 g must dissolve in 1.00 L.

$$C_{\text{TlCl}} = \frac{\left(\frac{24 \text{ g}}{239.83 \text{ g/mol}}\right)}{1.00 \text{ L}}$$

$$C_{\text{TlCl}} = 0.10 \text{ mol/L}$$

ICE Table for the Dissolving of Thallium(I) Chloride			
	$\text{TlCl}_{(s)} \rightleftharpoons$	$\text{Tl}_{(aq)}^+$	$\text{Cl}_{(aq)}^-$
Initial concentration (mol/L)	—	0	0
Change in concentration (mol/L)	—	+0.10	+0.10
Equilibrium concentration (mol/L)	—	0.10	0.10

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

$$= [0.10][0.10]$$

$$K_{\text{sp}} = 1.0 \times 10^{-2}$$

The  $K_{\text{sp}}$  for thallium chloride is  $1.0 \times 10^{-2}$ .

8. If 0.0016 g of  $\text{CaF}_2$  dissolves in 100 mL, then 0.016 g dissolves in 1.00 L.

$$C_{\text{CaF}_2} = \frac{0.016 \text{ g}}{78.08 \text{ g/mol}} \times \frac{1000 \text{ mL}}{100 \text{ mL}}$$

$$C_{\text{CaF}_2} = 2.05 \times 10^{-4} \text{ mol/L}$$

(extra digits carried)

ICE Table for the Dissolving of Calcium Fluoride			
	$\text{CaF}_{2(aq)} \rightleftharpoons$	$\text{Ca}_{(aq)}^{2+}$	$2 \text{F}_{(aq)}^-$
Initial concentration (mol/L)		0	0
Change in concentration (mol/L)		$+2.05 \times 10^{-4}$	$+4.10 \times 10^{-4}$
Equilibrium concentration (mol/L)		$2.05 \times 10^{-4}$	$4.10 \times 10^{-4}$

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

$$= [2.05 \times 10^{-4}][4.10 \times 10^{-4}]^2$$

$$K_{\text{sp}} = 3.4 \times 10^{-11}$$

The  $K_{\text{sp}}$  for calcium fluoride is  $3.4 \times 10^{-11}$ .

9.

ICE Table for the Dissolving of Mercury(I) Chloride			
	$\text{Hg}_2\text{Cl}_{2(aq)} \rightleftharpoons$	$\text{Hg}_{2(aq)}^{2+}$	$2 \text{Cl}_{(aq)}^-$
Initial concentration (mol/L)		0	0
Change in concentration (mol/L)		+x	+2x
Equilibrium concentration (mol/L)		x	2x

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

$$(x)(2x)^2 = 1.5 \times 10^{-18}$$

$$4x^3 = 1.5 \times 10^{-18}$$



$$x = 7.21 \times 10^{-7} \quad (\text{extra digits carried})$$

The molar solubility of mercury(I) chloride is  $7.21 \times 10^{-7}$  mol/L.

mass of  $\text{Hg}_2\text{Cl}_2$  required to prepare 1.0 L of solution,

$$m = 7.21 \times 10^{-7} \text{ mol/L} \times 472.08 \text{ g/mol}$$

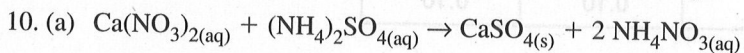
$$m = 3.40 \times 10^{-4} \text{ g}$$

mass of  $\text{Hg}_2\text{Cl}_2$  required to prepare 500 mL of solution,

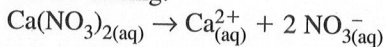
$$m = \frac{3.40 \times 10^{-4} \text{ g}}{2}$$

$$m = 1.7 \times 10^{-4} \text{ g}$$

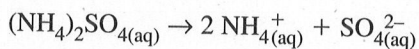
The mass of mercury(I) chloride required is  $1.7 \times 10^{-4}$  g.



Before mixing:



$$[\text{Ca}(\text{NO}_3)_2(\text{aq})] = [\text{Ca}_{(\text{aq})}^{2+}] = 0.040 \text{ mol/L}$$



$$[(\text{NH}_4)_2\text{SO}_4(\text{aq})] = [\text{SO}_{4(\text{aq})}^{2-}] = 0.080 \text{ mol/L}$$

After mixing:

$$50 \text{ mL} + 150 \text{ mL} = 200 \text{ mL}$$

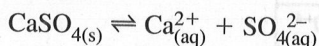
Concentrations after mixing:

$$[\text{Ca}_{(\text{aq})}^{2+}] = 0.040 \text{ mol/L} \times \frac{50 \text{ mL}}{200 \text{ mL}}$$

$$[\text{Ca}_{(\text{aq})}^{2+}] = 0.010 \text{ mol/L}$$

$$[\text{SO}_{4(\text{aq})}^{2-}] = 0.080 \text{ mol/L} \times \frac{150 \text{ mL}}{200 \text{ mL}}$$

$$[\text{SO}_{4(\text{aq})}^{2-}] = 0.060 \text{ mol/L}$$



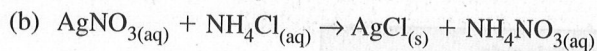
$$Q = [\text{Ca}_{(\text{aq})}^{2+}][\text{SO}_{4(\text{aq})}^{2-}]$$

$$= (0.010)(0.060)$$

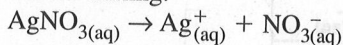
$$Q = 6.0 \times 10^{-4}$$

$$K_{\text{sp}} = 7.1 \times 10^{-5}$$

$Q$  is larger than  $K_{\text{sp}}$ . Therefore, a precipitate does form.



Before mixing:



$$[\text{AgNO}_3(\text{aq})] = [\text{Ag}_{(\text{aq})}^+] = 2.2 \times 10^{-9} \text{ mol/L}$$

$$[\text{NH}_4\text{Cl}(\text{aq})] = [\text{Cl}_{(\text{aq})}^-] = 0.050 \text{ mol/L}$$

After mixing:

$$50 \text{ mL} + 50 \text{ mL} = 100 \text{ mL}$$

Concentrations after mixing:

$$[\text{Ag}_{(\text{aq})}^+] = 2.2 \times 10^{-9} \text{ mol/L} \times \frac{50 \text{ mL}}{100 \text{ mL}}$$



$$[\text{Ag}^+_{(\text{aq})}] = 1.1 \times 10^{-9} \text{ mol/L}$$

$$[\text{Cl}^-_{(\text{aq})}] = 0.050 \text{ mol/L} \times \frac{50 \text{ mL}}{100 \text{ mL}}$$

$$[\text{Cl}^-_{(\text{aq})}] = 0.025 \text{ mol/L}$$



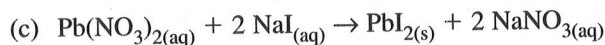
$$Q = [\text{Ag}^+_{(\text{aq})}][\text{Cl}^-_{(\text{aq})}]$$

$$= (1.1 \times 10^{-9})(0.025)$$

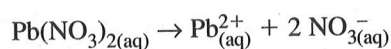
$$Q = 2.8 \times 10^{-11}$$

$$K_{\text{sp}} = 1.8 \times 10^{-10}$$

$Q$  is smaller than  $K_{\text{sp}}$ . Therefore, a precipitate does not form.



Before mixing:



$$[\text{Pb}(\text{NO}_3)_2_{(\text{aq})}] = [\text{Pb}^{2+}_{(\text{aq})}] = 2.1 \times 10^{-3} \text{ mol/L}$$

$$[\text{NaI}_{(\text{aq})}] = [\text{I}^-_{(\text{aq})}] = 0.006 \text{ mol/L}$$

After mixing:

$$100 \text{ mL} + 50 \text{ mL} = 150 \text{ mL}$$

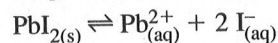
Concentrations after mixing:

$$[\text{Pb}^{2+}_{(\text{aq})}] = 2.1 \times 10^{-3} \text{ mol/L} \times \frac{100 \text{ mL}}{150 \text{ mL}}$$

$$[\text{Pb}^{2+}_{(\text{aq})}] = 1.4 \times 10^{-3} \text{ mol/L}$$

$$[\text{I}^-_{(\text{aq})}] = 0.006 \text{ mol/L} \times \frac{50 \text{ mL}}{150 \text{ mL}}$$

$$[\text{I}^-_{(\text{aq})}] = 2.0 \times 10^{-3} \text{ mol/L}$$



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

$$= (1.4 \times 10^{-3})(2.0 \times 10^{-3})^2$$

$$Q = 5.6 \times 10^{-9}$$

$$K_{\text{sp}} = 8.5 \times 10^{-9}$$

$Q$  is smaller than  $K_{\text{sp}}$ . Therefore, a precipitate does not form.

11.

ICE Table for the Solubility of Lead Iodide in a Solution of 0.1 mol/L Sodium Iodide			
	$\text{PbI}_{2(\text{s})} \rightleftharpoons$	$\text{Pb}^{2+}_{(\text{aq})} +$	$2 \text{I}^-_{(\text{aq})}$
Initial concentration (mol/L)			0.1
Change in concentration (mol/L)	—	+x	+2x
Equilibrium concentration (mol/L)		x	2x + 0.1

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

$$K_{\text{sp}} = 8.5 \times 10^{-9}$$

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

$$\text{Assume } 2x + 0.10 \doteq 0.10$$

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

$$x = 8.5 \times 10^{-7}$$

The solubility of lead iodide in a 0.10 mol/L solution of sodium iodide is  $8.5 \times 10^{-7}$  mol/L.

### Applying Inquiry Skills

$$12. (a) n_{\text{Zn}} = \frac{0.094 \text{ g}}{65.38 \text{ g/mol}}$$

$$n_{\text{Zn}} = 1.4 \times 10^{-3} \text{ mol}$$

$1.4 \times 10^{-3}$  mol of zinc reacted.

$$(b) C_{\text{Pb}^{2+}} = \frac{1.4 \times 10^{-3} \text{ mol}}{0.100 \text{ L}}$$

$$C_{\text{Pb}^{2+}} = 1.4 \times 10^{-2} \text{ mol/L}$$

The molar concentration of the lead ions is  $1.4 \times 10^{-2}$  mol/L.

$$(c) \text{ If } [\text{Pb}_{(\text{aq})}^{2+}] = 1.4 \times 10^{-2} \text{ mol/L, then } [\text{Cl}_{(\text{aq})}^{-}] = 2.8 \times 10^{-2} \text{ mol/L.}$$

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

$$= [1.4 \times 10^{-2}][2.8 \times 10^{-2}]^2$$

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

The  $K_{\text{sp}}$  for lead(II) chloride is  $1.1 \times 10^{-5}$ .

### 13. Analysis

$$(a) n_{\text{Cu}} = \frac{0.16 \text{ g}}{63.55 \text{ g/mol}}$$

$$n_{\text{Cu}} = 2.5 \times 10^{-3} \text{ mol}$$

$2.5 \times 10^{-3}$  mol of copper reacted.

$$(b) n_{\text{Ag}} = 2n_{\text{Cu}}$$

$$= 2 \times 2.5 \times 10^{-3}$$

$$n_{\text{Ag}} = 5.0 \times 10^{-3} \text{ mol}$$

$5.0 \times 10^{-3}$  mol of silver was produced.

$$(c) C_{\text{Ag}^{+}} = \frac{5.0 \times 10^{-3} \text{ mol}}{0.100 \text{ L}}$$

$$C_{\text{Ag}^{+}} = 5.0 \times 10^{-2} \text{ mol/L}$$

The molar concentration of the silver ions is  $5.0 \times 10^{-2}$  mol/L.

$$(d) K_{\text{sp}} = [\text{Ag}_{(\text{aq})}^{+}][\text{C}_2\text{H}_3\text{O}_2_{(\text{aq})}^{-}]$$

$$= [5.0 \times 10^{-2}][5.0 \times 10^{-2}]$$

$$K_{\text{sp}} = 2.5 \times 10^{-3}$$

The solubility product for silver acetate is  $2.5 \times 10^{-3}$ .

### Evaluation

- (e) The mass of copper observed would be larger than intended, suggesting that not as much silver reacted with it. The silver and acetate concentrations, as a result, would be lower, resulting in a lower value for the solubility product.

The silver residue made the copper heavier. Because of this error, the experimenter would erroneously conclude that less copper metal reacted, implying that there is less silver in solution. The final value for  $K_{\text{sp}}$  would be lower than expected.