At equilibrium,

$$K = \frac{[COCl_{2(g)}]}{[CO_{(g)}][Cl_{2(g)}]}$$

$$= 8.2 \times 10^{-2}$$

$$\frac{0.25 - x}{x^2} = 8.2 \times 10^{-2}$$

$$8.2 \times 10^{-2} x^2 = 0.25 - x$$

$$8.2 \times 10^{-2} x^2 + x - 0.25 = 0$$

$$x = \frac{-1 \pm \sqrt{1^2 - 4(8.2 \times 10^{-2})(-0.25)}}{2(8.2 \times 10^{-2})}$$

$$= 0.25 \text{ or } -12.4$$

The negative root is rejected because negative concentrations are impossible.

$$[CO_{(g)}] = [Cl_{2(g)}] = x$$

$$= 0.25 \text{ mol/L}$$

The equilibrium concentrations of carbon monoxide and chlorine are 0.25 mol/L.

8. $PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

Initial concentrations are

$$[PCl_{3(g)}] = \frac{0.500 \text{ mol}}{1.00 \text{ L}}$$

$$[PCl_{3(g)}] = 0.500 \text{ mol/L}$$

$$[Cl_{2(g)}] = \frac{0.500 \text{ mol}}{1.00 \text{ L}}$$

$$[Cl_{2(g)}] = 0.500 \text{ mol/L}$$

| ICE Table for the Decomposition of $PCl_{5(g)}$			
	$PCl_{5(g)} \rightleftharpoons$	$PCl_{3(g)} +$	$Cl_{2(g)}$
Initial concentration (mol/L)	0.00	0.500	0.500
Change in concentration (mol/L)	x	$-x$	$-x$
Equilibrium concentration (mol/L)	x	$0.500 - x$	$0.500 - x$

The equilibrium concentrations of phosphorus pentachloride, phosphorus trichloride, and chlorine are 0.0185 mol/L, 0.482 mol/L, and 0.482 mol/L, respectively.

7.6 THE SOLUBILITY PRODUCT CONSTANT

PRACTICE

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Understanding Concepts

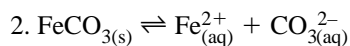1. $AgI_{(s)} \rightleftharpoons Ag^+_{(aq)} + I^-_{(aq)}$
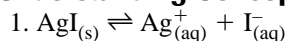
$$K_{sp} = [Ag^+][I^-]$$

$$K_{sp} = 1.5 \times 10^{-16}$$

$$[Ag^+] = [I^-]$$

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

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

The solubility of silver iodide is 1.2×10^{-8} mol/L.

2. $FeCO_{3(s)} \rightleftharpoons Fe^{2+}_{(aq)} + CO^{2-}_{3(aq)}$

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

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

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

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

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

$$[FeCO_{3(aq)}] = [Fe^{2+}_{(aq)}]$$

$$[FeCO_{3(aq)}] = 5.9 \times 10^{-6} \text{ mol/L}$$

The solubility of iron(II) carbonate is 5.9×10^{-6} mol/L.
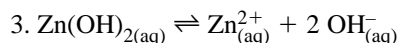
3. $Zn(OH)_{2(aq)} \rightleftharpoons Zn^{2+}_{(aq)} + 2\,OH^-_{(aq)}$

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

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

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

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

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

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

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

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

The solubility of zinc hydroxide is 2.2×10^{-6} mol/L.

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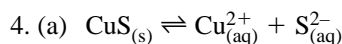
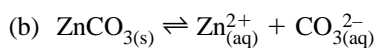| ICE Table for Calculating K_{sp} from Solubility			
	$CuS_{(s)} \rightleftharpoons$	$Cu^{2+}_{(aq)}$ +	$S^{2-}_{(aq)}$
Initial concentration (mol/L)		0	0
Change in concentration (mol/L)		+x	+x
Equilibrium concentration (mol/L)		x	x

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

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

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

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

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

(b) $ZnCO_{3(s)} \rightleftharpoons Zn^{2+}_{(aq)} + CO^{2-}_{3(aq)}$

ICE Table for Calculating K_{sp} from Solubility			
	$ZnCO_{3(s)} \rightleftharpoons$	$Zn^{2+}_{(aq)} +$	$CO^{2-}_{3(aq)}$
Initial concentration (mol/L)		0	0
Change in concentration (mol/L)		$+x$	$+x$
Equilibrium concentration (mol/L)		x	x

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

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

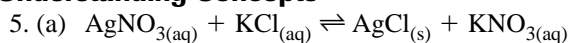$$= [3.87 \times 10^{-6}][3.87 \times 10^{-6}]$$

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

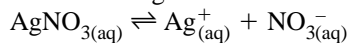The solubility product of zinc carbonate has a value of 1.50×10^{-11}.

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Understanding Concepts

5. (a) $AgNO_{3(aq)} + KCl_{(aq)} \rightleftharpoons AgCl_{(s)} + KNO_{3(aq)}$

Before mixing:
$AgNO_{3(aq)} \rightleftharpoons Ag^{+}_{(aq)} + NO^{-}_{3(aq)}$

$[AgNO_{3(aq)}] = [Ag^{+}_{(aq)}] = 0.010 \text{ mol/L}$

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

After mixing:
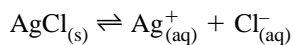$25.0 \text{ mL} + 25.0 \text{ mL} = 50.0 \text{ mL}$

Concentrations after mixing:

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

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

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

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

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

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

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

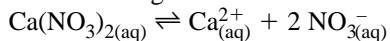$$Q = 1.25 \times 10^{-5}$$

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

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

(b) $Ca(NO_3)_{2(aq)} + 2\ KOH_{(aq)} \rightleftharpoons Ca(OH)_{2(s)} + 2\ KNO_{3(aq)}$

Before mixing:

$Ca(NO_3)_{2(aq)} \rightleftharpoons Ca^{2+}_{(aq)} + 2\ NO^-_{3(aq)}$
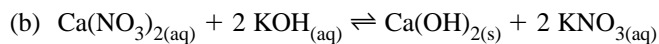
$[Ca(NO_3)_{2(aq)}] = [Ca^{2+}_{(aq)}] = 0.0010$ mol/L

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

Concentrations after mixing:

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

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

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

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

$Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2\ OH^-_{(aq)}$
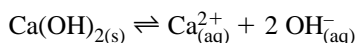
$\begin{aligned} Q &= [Ca^{2+}_{(aq)}][OH^-_{(aq)}]^2 \\ &= (0.0005\ \text{mol/L})(0.0010\ \text{mol/L})^2 \\ Q &= 5.0 \times 10^{-10} \\ K_{sp} &= 7.9 \times 10^{-6} \end{aligned}$

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

(c) $Pb(NO_3)_{2(aq)} + 2\ NaCl_{(aq)} \rightleftharpoons PbCl_{2(s)} + 2\ NaNO_{3(aq)}$

Before mixing:

$Pb(NO_3)_{2(aq)} \rightleftharpoons Pb^{2+}_{(aq)} + 2\ NO^-_{3(aq)}$
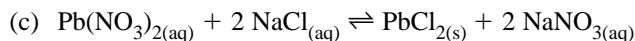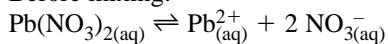
$[Pb(NO_3)_{2(aq)}] = [Pb^{2+}_{(aq)}] = 0.010$ mol/L

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

Concentrations after mixing:

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

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

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

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

$Pb(OH)_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2\ Cl^-_{(aq)}$
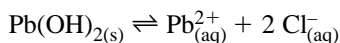
$\begin{aligned} Q &= [Pb^{2+}_{(aq)}][Cl^-_{(aq)}]^2 \\ &= (0.005\ \text{mol/L})(0.050\ \text{mol/L})^2 \\ Q &= 1.2 \times 10^{-5} \\ K_{sp} &= 1.2 \times 10^{-5} \end{aligned}$

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

Making Connections

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

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Understanding Concepts

7.

ICE Table for the Solubility of Silver Chloride in a Solution Containing $NaCl_{(aq)}$			
	$AgCl_{(s)} \rightleftharpoons$	$Ag^+_{(aq)} +$	$Cl^-_{(aq)}$
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 1.8×10^{-9} mol/L.

8.

ICE Table for the Solubility of $CaSO_4$ in a Solution Containing $Ca(NO_3)_{2(aq)}$			
	$CaSO_{4(s)} \rightleftharpoons$	$Ca^{2+}_{(aq)} +$	$SO_{4(aq)}^{2-}$
Initial concentration (mol/L)	—	0.010	0.00
Change in concentration (mol/L)	—	$+x$	$+x$
Equilibrium concentration (mol/L)	—	$0.010 + x$	x

$$K_{sp} = [Ca^{2+}_{(aq)}][SO_{4(aq)}^{2-}]$$
$$K_{sp} = 7.1 \times 10^{-5}$$

$0.010 + x \doteq 0.010$

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 7.1×10^{-3} mol/L.

9. (a)

ICE Table for the Solubility of Ag_2CrO_4 in Water			
	$Ag_2CrO_{4(s)} \rightleftharpoons$	$2\ Ag^+_{(aq)} +$	$CrO_{4(aq)}^{2-}$
Initial concentration (mol/L)			
Change in concentration (mol/L)	—	$+2x$	$+x$
Equilibrium concentration (mol/L)	—	$2x$	x

$$K_{sp} = [Ag^+_{(aq)}][CrO_{4(aq)}^{2-}]$$
$$K_{sp} = 1.1 \times 10^{-12}$$

$$(2x)^2 (x) = 1.1 \times 10^{-12}$$
$$4x^3 = 1.1 \times 10^{-12}$$
$$x = 6.5 \times 10^{-5}$$

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

(b)

ICE Table for the Solubility of Ag_2CrO_4 in a Solution of 0.1 mol/L Sodium Chromate			
	$Ag_2CrO_{4(s)} \rightleftharpoons$	$2\ Ag^+_{(aq)}$ +	$CrO_4^{2-}_{(aq)}$
Initial concentration (mol/L)		0.0	0.1
Change in concentration (mol/L)		+2x	+x
Equilibrium concentration (mol/L)		2x	x + 0.1

$$K_{sp} = [Ag^+_{(aq)}]^2[CrO_4^{2-}_{(aq)}]$$
$$K_{sp} = 1.1 \times 10^{-12}$$
$$(2x)^2(x + 0.10) = 1.1 \times 10^{-12}$$

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 1.7×10^{-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.

10. Soluble compounds containing either barium or sulfate ions would decrease the solubility of barium sulfate. Examples include barium nitrate, $Ba(NO_3)_2$, and potassium sulfate, K_2SO_4.

11. Soluble compounds containing either copper(II) or carbonate ions would decrease the solubility of copper(II) carbonate. Examples include copper(II) sulfate, $CuSO_4$, and sodium carbonate, Na_2CO_3.

12. 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

1. 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.

2. (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 $B_bC_{c(s)} \rightleftharpoons bB^+_{(aq)} + cC^-_{(aq)}$, the presence of a common ion $B^+_{(aq)}$ or $C^-_{(aq)}$ will shift the equilibrium to the left, reducing the solubility of the salt B_bC_c.

3. 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.

4.

ICE Table for the Solubility of Barium Sulfate			
	$BaSO_{4(s)} \rightleftharpoons$	$Ba^{2+}_{(aq)}$ +	$SO_4^{2-}_{(aq)}$
Initial concentration (mol/L)	—	0	0
Change in concentration (mol/L)	—	+x	+x
Equilibrium concentration (mol/L)	—	x	x

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

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

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

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

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

5.

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

$$K_{sp} = [Ag^+_{(aq)}][Br^-_{4(aq)}]$$

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

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

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

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

Solubility in g/100 mL …

mass of silver bromide in 1 L

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

$= 1.38 \times 10^{-4}$ 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 1.4×10^{-5} g/100 mL.

6.

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

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

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

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

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

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

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

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

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

7. If 2.4 g of TlCl dissolves in 100 mL, then 24 g must dissolve in 1.00 L.

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

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

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

$$K_{sp} = [Tl^+_{(aq)}][Cl^-_{(aq)}]$$
$$= [0.10][0.10]$$
$$K_{sp} = 1.0 \times 10^{-2}$$

The K_{sp} for thallium chloride is 1.0×10^{-2}.

8. If 0.0016 g of CaF_2 dissolves in 100 mL, then 0.016 g dissoves in 1.00 L.

$$C_{CaF_2} = \dfrac{\dfrac{0.016 \text{ g}}{78.08 \text{ g/mol}}}{1.00 \text{ L}}$$

$C_{CaF_2} = 2.05 \times 10^{-4} \text{ mol/L}$ \qquad\qquad (extra digits carried)

ICE Table for the Dissolving of Calcium Fluoride			
	$CaF_{2(aq)} \rightleftharpoons$	$Ca^{2+}_{(aq)} +$	$2 F^-_{(aq)}$
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)		2.05×10^{-4}	4.10×10^{-4}

$$K_{sp} = [Ca^{2+}_{(aq)}][F^-_{(aq)}]^2$$
$$= [2.05 \times 10^{-4}][4.10 \times 10^{-4}]^2$$
$$K_{sp} = 3.4 \times 10^{-11}$$

The K_{sp} for calcium fluoride is 3.4×10^{-11}.

9.

ICE Table for the Dissolving of Mercury(I) Chloride			
	$Hg_2Cl_{2(aq)} \rightleftharpoons$	$Hg_2^{2+}_{(aq)} +$	$2 Cl^-_{(aq)}$
Initial concentration (mol/L)		0	0
Change in concentration (mol/L)		$+x$	$+2x$
Equilibrium concentration (mol/L)		x	$2x$

$$K_{sp} = [Hg_2^{2+}_{(aq)}][Cl^-_{(aq)}]^2$$
$$(x)(2x)^2 = 1.5 \times 10^{-18}$$
$$4x^3 = 1.5 \times 10^{-18}$$

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

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

mass of Hg_2Cl_2 required to prepare 1.0 L of solution,

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

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

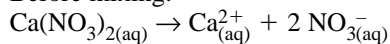mass of Hg_2Cl_2 required to prepare 500 mL of solution,

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

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

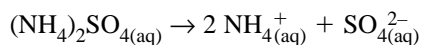The mass of mercury(I) chloride required is 1.7×10^{-4} g.

10. (a) $Ca(NO_3)_{2(aq)} + (NH_4)_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + 2\ NH_4NO_{3(aq)}$

Before mixing:
$Ca(NO_3)_{2(aq)} \rightarrow Ca^{2+}_{(aq)} + 2\ NO^-_{3(aq)}$

$[Ca(NO_3)_{2(aq)}] = [Ca^{2+}_{(aq)}] = 0.040$ mol/L

$(NH_4)_2SO_{4(aq)} \rightarrow 2\ NH^+_{4(aq)} + SO^{2-}_{4(aq)}$

$[(NH_4)_2SO_{4(aq)}] = [SO^{2-}_{4(aq)}] = 0.080$ mol/L

After mixing:
50 mL + 150 mL = 200 mL

Concentrations after mixing:

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

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

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

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

$CaSO_{4(s)} \rightleftharpoons Ca^{2+}_{(aq)} + SO^{2-}_{4(aq)}$

$\quad\quad Q = [Ca^{2+}_{(aq)}][SO^{2-}_{4(aq)}]$
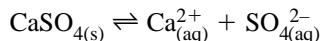
$\quad\quad\ \ = (0.010)(0.060)$

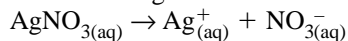$\quad\quad Q = 6.0 \times 10^{-4}$

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

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

(b) $AgNO_{3(aq)} + NH_4Cl_{(aq)} \rightarrow AgCl_{(s)} + NH_4NO_{3(aq)}$
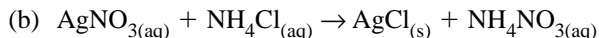
Before mixing:
$AgNO_{3(aq)} \rightarrow Ag^+_{(aq)} + NO^-_{3(aq)}$

$[AgNO_{3(aq)}] = [Ag^+_{(aq)}] = 2.2 \times 10^{-9}$ mol/L

$[NH_4Cl_{(aq)}] = [Cl^-_{(aq)}] = 0.050$ mol/L

After mixing:
50 mL + 50 mL = 100 mL
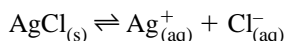
Concentrations after mixing:

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

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

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

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

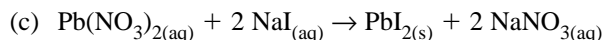$\text{AgCl}_{(s)} \rightleftharpoons \text{Ag}^+_{(aq)} + \text{Cl}^-_{(aq)}$

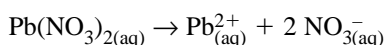$$Q = [\text{Ag}^+_{(aq)}][\text{Cl}^-_{(aq)}]$$
$$= (1.1 \times 10^{-9})(0.025)$$
$$Q = 2.8 \times 10^{-11}$$

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

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

(c) $\text{Pb(NO}_3)_{2(aq)} + 2\,\text{NaI}_{(aq)} \rightarrow \text{PbI}_{2(s)} + 2\,\text{NaNO}_{3(aq)}$

Before mixing:

$\text{Pb(NO}_3)_{2(aq)} \rightarrow \text{Pb}^{2+}_{(aq)} + 2\,\text{NO}^-_{3(aq)}$

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

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

After mixing:

100 mL + 50 mL = 150 mL
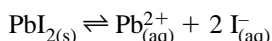
Concentrations after mixing:

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

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

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

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

$\text{PbI}_{2(s)} \rightleftharpoons \text{Pb}^{2+}_{(aq)} + 2\,\text{I}^-_{(aq)}$

$$Q = [\text{Pb}^{2+}_{(aq)}][\text{I}^-_{(aq)}]^2$$
$$= (1.4 \times 10^{-3})(2.0 \times 10^{-3})^2$$
$$Q = 5.6 \times 10^{-9}$$

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

Q is smaller than K_{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			
	$\text{PbI}_{2(s)} \rightleftharpoons$	$\text{Pb}^{2+}_{(aq)} +$	$2\,\text{I}^-_{(aq)}$
Initial concentration (mol/L)			0.1
Change in concentration (mol/L)	—	$+x$	$+2x$
Equilibrium concentration (mol/L)		x	$2x + 0.1$

$$K_{sp} = [\text{Pb}^{2+}_{(aq)}][\text{I}^-_{(aq)}]^2$$
$$K_{sp} = 8.5 \times 10^{-9}$$
$$(x)(2x + 0.10)^2 = 8.5 \times 10^{-9}$$

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 8.5×10^{-7} mol/L.

Applying Inquiry Skills

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

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

1.4×10^{-3} mol of zinc reacted.

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

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

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

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

$$\begin{aligned} K_{sp} &= [Pb^{2+}_{(aq)}][Cl^-_{(aq)}]^2 \\ &= [1.4 \times 10^{-2}][2.8 \times 10^{-2}]^2 \\ K_{sp} &= 1.1 \times 10^{-5} \end{aligned}$$

The K_{sp} for lead(II) chloride is 1.1×10^{-5}.

13. **Analysis**

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

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

2.5×10^{-3} mol of copper reacted.

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

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

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

5.0×10^{-3} mol of silver was produced.

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

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

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

(d) $\begin{aligned} K_{sp} &= [Ag^+_{(aq)}][C_2H_3O^-_{2(aq)}] \\ &= [5.0 \times 10^{-2}][5.0 \times 10^{-2}] \\ K_{sp} &= 2.5 \times 10^{-3} \end{aligned}$

The solubility product for silver acetate is 2.5×10^{-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_{sp} would be lower than expected.