(e) **Procedure**

1. Use a graduated cylinder to add 10.0 mL of hard water to a 250-mL beaker.
2. Use another graduated cylinder to add about 10 mL of sodium oxalate solution to the beaker.
3. Allow the precipitate to settle.
4. Add another 10-mL aliquot of sodium oxalate solution to the beaker, observing whether or not more precipitate forms as the added solution mixes with the clear upper portion of the solution in the beaker.
5. Repeat step 4 until no more precipitate forms.
6. Measure and record the mass of a piece of filter paper.
7. Filter, wash, and dry the precipitate.
8. Measure and record the mass of the filter paper plus dry precipitate.
9. Dispose of all materials as instructed.

Making Connections

6. (a) Raw materials, like aluminum oxide, are purchased according to the amount needed for reaction.
   (b) Products, like aluminum, are priced so that the amount produced will be profitable.
   (c) Pollutant amounts will be calculated and decisions on processes to control them will be made accordingly.
   (d) Amounts of all materials must be calculated to allow costing of things like packaging, disposal, and shipping.

5.5 CALCULATING LIMITING AND EXCESS REAGENTS

PRACTICE

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Understanding Concepts

1. (a)

$$\begin{array}{lllll} 2\ CH_{4(g)} & + \quad 3\ O_{2(g)} & \rightarrow & 2\ CO_{(g)} & +\quad 4\ H_2O_{(g)} \\ m & & & & \\ 16.05\ g/mol & \quad\ \ 3.0\ mol & & & \end{array}$$

$$n_{CH_4} = 3.0\ mol \times \frac{2}{3}$$

$$n_{CH_4} = 2.0\ mol$$

$$m_{CH_4} = 2.0\ \cancel{mol} \times \frac{16.05\ g}{1\ \cancel{mol}}$$

$$m_{CH_4} = 32\ g$$

or

$$m_{CH_4} = 3\ \cancel{mol}\ \cancel{O_2} \times \frac{2\ \cancel{mol}\ \cancel{CH_4}}{3\ \cancel{mol}\ \cancel{O_2}} \times \frac{16.05\ g\ CH_4}{1\ \cancel{mol}\ \cancel{CH_4}}$$

$$m_{CH_4} = 32\ g$$

The mass of methane that will react is 32 g.

(b)

$$\begin{array}{lllll} CH_{4(g)} & + \quad 2\ O_{2(g)} & \rightarrow & CO_{2(g)} & +\quad 2\ H_2O_{(g)} \\ m & & & & \\ 16.05\ g/mol & \quad\ \ 3.0\ mol & & & \end{array}$$

$$n_{CH_4} = 3.0\ mol \times \frac{1}{2}$$

$$n_{CH_4} = 1.5\ mol$$

$$m_{CH_4} = 1.5\ \cancel{mol} \times \frac{16.05\ g}{1\ \cancel{mol}}$$

$$m_{CH_4} = 24\ g$$

or

$$m_{CH_4} = 3\ \cancel{mol}\ \cancel{O_2} \times \frac{1\ \cancel{mol}\ \cancel{CH_4}}{2\ \cancel{mol}\ \cancel{O_2}} \times \frac{16.05\ g\ CH_4}{1\ \cancel{mol}\ \cancel{CH_4}}$$

$$m_{CH_4} = 24\ \text{g}$$

The mass of methane that will react is 24 g.

(c) In a closed garage the oxygen in the air will be the limiting reagent.

2. $$\begin{array}{lllll} 8\ Zn_{(s)} & + & S_{8(s)} & \rightarrow & 8\ ZnS_{(s)} \\ 5.00\ \text{g} & & m & & \\ 65.38\ \text{g/mol} & & 256.48\ \text{g/mol} & & \end{array}$$

$$n_{Zn} = 5.00\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{65.38\ \cancel{\text{g}}}$$

$$n_{Zn} = 0.0765\ \text{mol}$$

$$n_{S_8} = 0.0765\ \text{mol} \times \frac{1}{8}$$

$$n_{S_8} = 0.00956\ \text{mol}$$

$$m_{S_8} = 0.00956\ \cancel{\text{mol}} \times \frac{256.48\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$m_{S_8} = 2.45\ \text{g}$$

or

$$m_{S_8} = 5.00\ \cancel{\text{g}}\ \cancel{Zn} \times \frac{1\ \cancel{\text{mol}}\ \cancel{Zn}}{65.38\ \cancel{\text{g}}\ \cancel{Zn}} \times \frac{1\ \cancel{\text{mol}}\ \cancel{S_8}}{8\ \cancel{\text{mol}}\ \cancel{Zn}} \times \frac{256.48\ \text{g}\ S_8}{1\ \cancel{\text{mol}}\ \cancel{S_8}}$$

$$m_{S_8} = 2.45\ \text{g}$$

excess $m_{S_8} = 2.45\ \text{g} \times 110\% = 2.70\ \text{g}$

The mass of sulfur, to be in reasonable excess, should be 2.70 g.

3. $$\begin{array}{lllll} 2\ Na_{(s)} & + & Cl_{2(g)} & \rightarrow & 2\ NaCl_{(s)} \\ 0.75\ \text{g} & & m & & \\ 22.99\ \text{g/mol} & & 70.90\ \text{g/mol} & & \end{array}$$

$$n_{Na} = 0.75\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{22.99\ \cancel{\text{g}}}$$

$$n_{Na} = 0.033\ \text{mol}$$

$$n_{Cl_2} = 0.033\ \text{mol} \times \frac{1}{2}$$

$$n_{Cl_2} = 0.016\ \text{mol}$$

$$m_{Cl_2} = 0.016\ \cancel{\text{mol}} \times \frac{70.90\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$m_{Cl_2} = 1.2\ \text{g}$$

or

$$m_{Cl_2} = 0.75\ \cancel{\text{g}}\ \cancel{Na} \times \frac{1\ \cancel{\text{mol}}\ \cancel{Na}}{22.99\ \cancel{\text{g}}\ \cancel{Na}} \times \frac{1\ \cancel{\text{mol}}\ \cancel{Cl_2}}{2\ \cancel{\text{mol}}\ \cancel{Na}} \times \frac{70.90\ \text{g}\ Cl_2}{1\ \cancel{\text{mol}}\ \cancel{Cl_2}}$$

$$m_{Cl_2} = 1.2\ \text{g}$$

excess $m_{Cl_2} = 1.2\ \text{g} \times 110\% = 1.3\ \text{g}$

The mass of chlorine, to be in reasonable excess, should be 1.3 g.

4. $$\begin{array}{lllllll} Ca(OH)_{2(s)} & + & 2\ HCl_{(aq)} & \rightarrow & CaCl_{2(aq)} & + & 2\ H_2O_{(l)} \\ m & & 2.17\ \text{g} & & & & \\ 74.10\ \text{g/mol} & & 36.46\ \text{g/mol} & & & & \end{array}$$

$$n_{HCl} = 2.17\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{36.46\ \cancel{\text{g}}}$$

$$n_{HCl} = 0.0595\ \text{mol}$$

$$n_{Ca(OH)_2} = 0.0595\ \text{mol} \times \frac{1}{2}$$

$$n_{Ca(OH)_2} = 0.0298\ \text{mol}$$

$$m_{Ca(OH)_2} = 0.0298\ \cancel{\text{mol}} \times \frac{74.10\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$m_{Ca(OH)_2} = 2.21 \text{ g}$$

or

$$m_{Ca(OH)_2} = 2.17 \text{ } \cancel{\text{g}} \text{ } \cancel{\text{HCl}} \times \frac{1 \text{ } \cancel{\text{mol}} \text{ } \cancel{\text{HCl}}}{36.46 \text{ } \cancel{\text{g}} \text{ } \cancel{\text{HCl}}} \times \frac{1 \text{ } \cancel{\text{mol}} \text{ } \cancel{\text{Ca(OH)}_2}}{2 \text{ } \cancel{\text{mol}} \text{ } \cancel{\text{HCl}}} \times \frac{74.10 \text{ g Ca(OH)}_2}{1 \text{ } \cancel{\text{mol}} \text{ } \cancel{\text{Ca(OH)}_2}}$$

$$m_{Ca(OH)_2} = 2.21 \text{ g}$$

$$\text{excess } m_{Ca(OH)_2} = 2.21 \text{ g} \times 110\% = 2.43 \text{ g}$$

The mass of calcium hydroxide, to be in reasonable excess, should be 2.43 g.

PRACTICE

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Understanding Concepts

5. Prediction cannot be made from the amount of reagent in excess because not all of the amount will react.

6. (a)

$$\begin{array}{lllllll} Zn_{(s)} & + & CuSO_{4(aq)} & \rightarrow & Cu_{(s)} & + & ZnSO_{4(aq)} \\ 0.42 \text{ mol} & & 0.22 \text{ mol} & & & & \end{array}$$

Since this reactant mole ratio is 1:1, it is obvious by inspection that zinc is in excess, by an amount of (0.42 – 0.22) mol = 0.20 mol.

(b)

$$\begin{array}{lllllll} Cl_{2(g)} & + & 2 \text{ } NaI_{(aq)} & \rightarrow & I_{2(s)} & + & 2 \text{ } NaCl_{(aq)} \\ 10 \text{ mmol} & & 10 \text{ mmol} & & & & \end{array}$$

Since this reactant mole ratio is 1:2, 10 mmol of $NaI_{(aq)}$ will require 10 mmol × 1/2 = 5.0 mmol of $Cl_{2(g)}$ for reaction. The $Cl_{2(g)}$ is in excess by an amount of (10 – 5.0) mmol = 5 mmol.

(c)

$$\begin{array}{lllllll} AlCl_{3(aq)} & + & 3 \text{ } NaOH_{(aq)} & \rightarrow & Al(OH)_{3(s)} & + & 3 \text{ } NaCl_{(aq)} \\ 20 \text{ g} & & 19 \text{ g} & & & & \\ 133.33 \text{ g/mol} & & 40.00 \text{ g/mol} & & & & \end{array}$$

$$n_{AlCl_3} = 20 \text{ } \cancel{\text{g}} \times \frac{1 \text{ mol}}{133.33 \text{ } \cancel{\text{g}}}$$

$$n_{AlCl_3} = 0.150 \text{ mol}$$

$$n_{NaOH} = 19 \text{ } \cancel{\text{g}} \times \frac{1 \text{ mol}}{40.00 \text{ } \cancel{\text{g}}}$$

$$n_{NaOH} = 0.475 \text{ mol}$$

Since this reactant mole ratio is 1:3, 0.150 mol of $AlCl_{3(aq)}$ will require 0.150 mol × 3/1 = 0.450 mol of $NaOH_{(aq)}$ for reaction.

The $NaOH_{(aq)}$ is in excess, by an amount of (0.475 − 0.450) mol = 0.025 mol, or 25 mmol.

7. (a)

$$\begin{array}{lllll} 8 \text{ } Zn_{(s)} & + & S_{8(s)} & \rightarrow & 8 \text{ } ZnS_{(s)} \\ 5.00 \text{ g} & & 3.00 \text{ g} & & \\ 65.38 \text{ g/mol} & & 256.48 \text{ g/mol} & & \end{array}$$

$$n_{Zn} = 5.00 \text{ } \cancel{\text{g}} \times \frac{1 \text{ mol}}{65.38 \text{ } \cancel{\text{g}}}$$

$$n_{Zn} = 0.0765 \text{ mol}$$

$$n_{S_8} = 3.00 \text{ } \cancel{\text{g}} \times \frac{1 \text{ mol}}{256.48 \text{ } \cancel{\text{g}}}$$

$$n_{S_8} = 0.0117 \text{ mol}$$

Since this reactant mole ratio is 8:1, 0.0765 mol of $Zn_{(s)}$ will require 0.0765 mol × 1/8 = 0.00956 mol of $S_{8(s)}$ for reaction. The $S_{8(s)}$ is in excess; so the $Zn_{(s)}$ is the limiting reagent.

(b)

$$\begin{array}{lllll} 8 \text{ } Zn_{(s)} & + & S_{8(s)} & \rightarrow & 8 \text{ } ZnS_{(s)} \\ 5.00 \text{ g} & & & & m \\ 65.38 \text{ g/mol} & & & & 97.44 \text{ g/mol} \end{array}$$

$$\begin{aligned}
n_{Zn} &= 5.00\ \cancel{g} \times \frac{1\ mol}{65.38\ \cancel{g}}\\
n_{Zn} &= 0.0765\ mol\\
n_{ZnS} &= 0.0765\ mol \times \frac{8}{8}\\
n_{ZnS} &= 0.0765\ mol\\
m_{ZnS} &= 0.0765\ \cancel{mol} \times \frac{97.44\ g}{1\ \cancel{mol}}\\
m_{ZnS} &= 7.45\ g
\end{aligned}$$

or

$$\begin{aligned}
m_{ZnS} &= 5.00\ \cancel{g}\ \cancel{Zn} \times \frac{1\ \cancel{mol}\ \cancel{Zn}}{65.38\ \cancel{g}\ \cancel{Zn}} \times \frac{8\ \cancel{mol}\ \cancel{S_8}}{8\ \cancel{mol}\ \cancel{Zn}} \times \frac{97.44\ g\ S_8}{1\ \cancel{mol}\ \cancel{S_8}}\\
m_{ZnS} &= 7.45\ g
\end{aligned}$$

The mass of zinc sulfide produced would be 7.45 g.

8. (a)

$$\begin{array}{lllllll}
C_3H_{8(g)} & + & 5\ O_{2(g)} & \rightarrow & 3\ CO_{2(g)} & + & 4\ H_2O_{(g)}\\
100.7\ g & & 367.4\ g & & & & \\
44.11\ g/mol & & 32.00\ g/mol & & & &
\end{array}$$

$$\begin{aligned}
n_{C_3H_8} &= 100.7\ \cancel{g} \times \frac{1\ mol}{44.11\ \cancel{g}}\\
n_{C_3H_8} &= 2.28\ mol\\
n_{O_2} &= 367.4\ \cancel{g} \times \frac{1\ mol}{32.00\ \cancel{g}}\\
n_{O_2} &= 11.5\ mol
\end{aligned}$$

Since this reactant mole ratio is 1:5, 2.28 mol of $C_3H_{8(g)}$ will require 2.28 mol × 5/1 = 11.4 mol of $O_{2(g)}$ for reaction. The $O_{2(g)}$ is in excess; so the $C_3H_{8(g)}$ is the limiting reagent.

(b)

$$\begin{array}{lllllll}
C_3H_{8(g)} & + & 5\ O_{2(g)} & \rightarrow & 3\ CO_{2(g)} & + & 4\ H_2O_{(g)}\\
100.7\ g & & & & & & m\\
44.11\ g/mol & & & & & & 18.02\ g/mol
\end{array}$$

$$\begin{aligned}
n_{C_3H_8} &= 100.7\ \cancel{g} \times \frac{1\ mol}{44.11\ \cancel{g}}\\
n_{C_3H_8} &= 2.283\ mol\\
n_{H_2O} &= 2.283\ mol \times \frac{4}{1}\\
n_{H_2O} &= 9.132\ mol\\
m_{H_2O} &= 9.132\ \cancel{mol} \times \frac{18.02\ g}{1\ \cancel{mol}}\\
m_{H_2O} &= 164.6\ g
\end{aligned}$$

or

$$\begin{aligned}
m_{H_2O} &= 100.7\ \cancel{g}\ \cancel{C_3H_8} \times \frac{1\ \cancel{mol}\ \cancel{C_3H_8}}{44.11\ \cancel{g}\ \cancel{C_3H_8}} \times \frac{4\ \cancel{mol}\ \cancel{H_2O}}{1\ \cancel{mol}\ \cancel{C_3H_8}} \times \frac{18.02\ g\ H_2O}{1\ \cancel{mol}\ \cancel{H_2O}}\\
m_{H_2O} &= 164.6\ g
\end{aligned}$$

The mass of water formed is 164.6 g.

9. (a)

$$\begin{array}{lllll}
2\ SO_{2(g)} & + & O_{2(g)} & \rightarrow & 2\ SO_{3(g)}\\
4.55\ kg & & 2.88\ kg & & \\
64.06\ g/mol & & 32.00\ g/mol & &
\end{array}$$

$$\begin{aligned}
n_{SO_2} &= 4.55\ k\cancel{g} \times \frac{1\ mol}{64.06\ \cancel{g}}\\
n_{SO_2} &= 0.0710\ kmol\\
n_{O_2} &= 2.88\ k\cancel{g} \times \frac{1\ mol}{32.00\ \cancel{g}}
\end{aligned}$$

$$n_{O_2} = 0.0900 \text{ kmol}$$

Since this reactant mole ratio is 2:1, 0.0710 kmol of $SO_{2(g)}$ will require 0.0710 kmol × 1/2 = 0.0355 kmol of $O_{2(g)}$ for reaction. The $O_{2(g)}$ is in excess; so the $SO_{2(g)}$ is the limiting reagent.

(b)

$$\begin{array}{lllll} 2\,SO_{2(g)} & + & O_{2(g)} & \rightarrow & 2\,SO_{3(g)} \\ 4.55 \text{ kg} & & & & m \\ 64.06 \text{ g/mol} & & & & 80.06 \text{ g/mol} \end{array}$$

$$\begin{aligned} n_{SO_2} &= 4.55 \text{ k}\cancel{\text{g}} \times \frac{1 \text{ mol}}{64.06 \cancel{\text{g}}} \\ n_{SO_2} &= 0.0710 \text{ kmol} \\ n_{SO_3} &= 0.0710 \text{ kmol } \times \frac{2}{2} \\ n_{SO_3} &= 0.0710 \text{ kmol} \\ m_{SO_3} &= 0.0710 \text{ k}\cancel{\text{mol}} \times \frac{80.06 \text{ g}}{1 \cancel{\text{mol}}} \\ m_{SO_3} &= 5.69 \text{ kg} \end{aligned}$$

or

$$\begin{aligned} m_{SO_3} &= 4.55 \text{ k}\cancel{\text{g}} \cancel{SO_2} \times \frac{1 \cancel{\text{mol}} \cancel{SO_2}}{64.06 \cancel{\text{g}} \cancel{SO_2}} \times \frac{2 \cancel{\text{mol}} \cancel{SO_3}}{2 \cancel{\text{mol}} \cancel{SO_2}} \times \frac{80.06 \text{ g } SO_3}{1 \cancel{\text{mol}} \cancel{SO_3}} \\ m_{SO_3} &= 5.69 \text{ kg} \end{aligned}$$

The mass of sulfur trioxide produced is 5.69 kg.

10. (a)

$$\begin{array}{lllllll} Fe_2O_{3(s)} & + & 3\,CO_{(g)} & \rightarrow & 2\,Fe_{(s)} & + & 3\,CO_{2(g)} \\ 74.2 \text{ kg} & & 40.3 \text{ kg} & & & & \\ 159.70 \text{ g/mol} & & 28.01 \text{ g/mol} & & & & \end{array}$$

$$\begin{aligned} n_{Fe_2O_3} &= 74.2 \text{ k}\cancel{\text{g}} \times \frac{1 \text{ mol}}{159.70 \cancel{\text{g}}} \\ n_{Fe_2O_3} &= 0.465 \text{ kmol} \\ n_{CO} &= 40.3 \text{ k}\cancel{\text{g}} \times \frac{1 \text{ mol}}{28.01 \cancel{\text{g}}} \\ n_{CO} &= 1.44 \text{ kmol} \end{aligned}$$

Since this reactant mole ratio is 1:3, 0.465 kmol of $Fe_2O_{3(s)}$ will require 0.465 kmol × 3/1 = 1.39 kmol of $CO_{(g)}$ for reaction. The $CO_{(g)}$ is in excess; so the $Fe_2O_{3(s)}$ is the limiting reagent.

(b)

$$\begin{array}{lllllll} Fe_2O_{3(s)} & + & 3\,CO_{(g)} & \rightarrow & 2\,Fe_{(s)} & + & 3\,CO_{2(g)} \\ 74.2 \text{ kg} & & & & m & & \\ 159.70 \text{ g/mol} & & & & 55.85 \text{ g/mol} & & \end{array}$$

$$\begin{aligned} n_{Fe_2O_3} &= 74.2 \text{ k}\cancel{\text{g}} \times \frac{1 \text{ mol}}{159.70 \cancel{\text{g}}} \\ n_{Fe_2O_3} &= 0.465 \text{ kmol} \\ n_{Fe} &= 0.465 \text{ kmol} \times \frac{2}{1} \\ n_{Fe} &= 0.929 \text{ kmol} \\ m_{Fe} &= 0.929 \text{ k}\cancel{\text{mol}} \times \frac{55.85 \text{ g}}{1 \cancel{\text{mol}}} \\ m_{Fe} &= 51.9 \text{ kg} \end{aligned}$$

or

$$\begin{aligned} m_{Fe} &= 74.2 \text{ k}\cancel{\text{g}} \cancel{Fe_2O_3} \times \frac{1 \cancel{\text{mol}} \cancel{Fe_2O_3}}{159.70 \cancel{\text{g}} \cancel{Fe_2O_3}} \times \frac{2 \cancel{\text{mol}} \cancel{Fe}}{1 \cancel{\text{mol}} \cancel{Fe_2O_3}} \times \frac{55.85 \text{ g Fe}}{1 \cancel{\text{mol}} \cancel{Fe}} \\ m_{Fe} &= 51.9 \text{ kg} \end{aligned}$$

The mass of iron produced would be 51.9 kg.

Applying Inquiry Skills

11. **Experimental Design**
    A solid reaction product can be removed (by filtration, for instance) from the reaction system. If some more of the (presumed) excess reagent is added, and no reaction occurs, then the limiting reagent is completely reacted and the other reagent was in excess.

PRACTICE

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Applying Inquiry Skills

12. (a) If aluminum is the limiting reagent, the silver metal solid will have completely disappeared when the reaction stops.
    (b) If copper(II) sulfate is the limiting reagent, the silver metal solid will not have completely disappeared when the reaction stops.
13. (a) If magnesium is the limiting reagent, the silver metal solid will have completely disappeared when the reaction stops.
    (b) If hydrochloric acid is the limiting reagent, the silver metal solid will not have completely disappeared when the reaction stops.

SECTION 5.5 QUESTIONS

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Understanding Concepts

1. The limiting reagent is determined by the reaction mole ratio; so the mass may actually be much greater than the excess reagent, as long as the amount is such that it is completely reacted.
2. The reaction only occurs until the limiting reagent is consumed, so the amount of the limiting reagent determines all other amounts involved in the reaction.
3. One (or more) reactants must be in excess to ensure that the limiting reagent will be completely reacted.
4. (a)

$$\begin{array}{lllll} Zn_{(s)} & + \quad 2\ HCl_{(aq)} & \rightarrow & H_{2(g)} & + \quad ZnCl_{2(aq)} \\ 0.35\ \text{mol} & \quad\ \ 0.60\ \text{mol} & & & \end{array}$$

Since this reactant mole ratio is 1:2, 0.60 mol of $HCl_{(aq)}$ will require 0.60 mol × 1/2 = 0.30 mol of $Zn_{(s)}$ for reaction. The $Zn_{(s)}$ is in excess; so the $HCl_{(aq)}$ is the limiting reagent, and will be completely consumed.

(b) The $Zn_{(s)}$ is in excess, by an amount of (0.35 − 0.30) mol = 0.05 mol.
excess m_{Zn} = 0.05 mol × 65.38 g/1 mol = 3 g
The mass of zinc remaining after the reaction ends will be 3 g.

5. (a)

$$\begin{array}{llll} 2\ KBr_{(aq)} & + \quad Pb(NO_3)_{2(aq)} & \rightarrow \ 2\ KNO_{3(aq)} & + \quad PbBr_{2(s)} \\ m & \quad\ \ 3.50\ \text{g} & & \\ 119.00\ \text{g/mol} & \quad\ \ 331.22\ \text{g/mol} & & \end{array}$$

$$n_{Pb(NO_3)_2} = 3.50\ \cancel{g} \times \frac{1\ \text{mol}}{331.22\ \cancel{g}}$$

$$n_{Pb(NO_3)_2} = 0.0106\ \text{mol}$$

$$n_{KBr} = 0.0106\ \text{mol} \times \frac{1}{2}$$

$$n_{KBr} = 0.00528\ \text{mol}$$

$$m_{KBr} = 0.00528\ \cancel{\text{mol}} \times \frac{119.00\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$m_{KBr} = 0.629\ \text{g}$$

or

$$m_{KBr} = 3.50\ \cancel{g}\ Pb(\cancel{NO_3})_2 \times \frac{1\ \cancel{\text{mol}}\ Pb(\cancel{NO_3})_2}{331.22\ \cancel{g}\ Pb(\cancel{NO_3})_2} \times \frac{1\ \cancel{\text{mol}}\ \cancel{KBr}}{2\ \cancel{\text{mol}}\ Pb(\cancel{NO_3})_2} \times \frac{119.00\ \text{g KBr}}{1\ \cancel{\text{mol}}\ \cancel{KBr}}$$

$$m_{KBr} = 0.629\ \text{g}$$

$$\text{excess } m_{KBr} = 0.629\text{ g} \times 110\% = 0.692\text{ g}$$

The mass of potassium bromide, to be in reasonable excess, should be 692 mg.

(b)
$$\begin{array}{lllllll} 2\,KBr_{(aq)} & + & Pb(NO_3)_{2(aq)} & \rightarrow & 2\,KNO_{3(aq)} & + & PbBr_{2(s)} \\ & & 3.50\text{ g} & & m & & \\ & & 331.22\text{ g/mol} & & 367.00\text{ g/mol} & & \end{array}$$

$$n_{Pb(NO_3)_2} = 3.50\,\cancel{\text{g}} \times \frac{1\text{ mol}}{331.22\,\cancel{\text{g}}}$$

$$n_{Pb(NO_3)_2} = 0.0106\text{ mol}$$

$$n_{PbBr_2} = 0.0106\text{ mol} \times \frac{1}{1}$$

$$n_{PbBr_2} = 0.0106\text{ mol}$$

$$m_{PbBr_2} = 0.0106\,\cancel{\text{mol}} \times \frac{119.00\text{ g}}{1\,\cancel{\text{mol}}}$$

$$m_{PbBr_2} = 0.629\text{ g}$$

or

$$m_{PbBr_2} = 3.50\,\cancel{\text{g}}\,\text{Pb}\cancel{(\text{NO}_3)_2} \times \frac{1\,\cancel{\text{mol}}\,\text{Pb}\cancel{(\text{NO}_3)_2}}{331.22\,\cancel{\text{g}}\,\text{Pb}\cancel{(\text{NO}_3)_2}} \times \frac{1\,\cancel{\text{mol}}\,\cancel{\text{PbBr}_2}}{1\,\cancel{\text{mol}}\,\text{Pb}\cancel{(\text{NO}_3)_2}} \times \frac{119.00\text{ g PbBr}_2}{1\,\cancel{\text{mol}}\,\cancel{\text{PbBr}_2}}$$

$$m_{PbBr_2} = 0.629\text{ g}$$

The mass of lead(II) bromide produced would be 0.629 g.

6. (a)
$$\begin{array}{lllllll} Cu_{(s)} & + & 2\,AgNO_{3(aq)} & \rightarrow & 2\,Ag_{(s)} & + & Cu(NO_3)_{2(aq)} \\ 10.0\text{ g} & & 20.0\text{ g} & & & & \\ 63.55\text{ g/mol} & & 169.88\text{ g/mol} & & & & \end{array}$$

$$n_{Cu} = 10.0\,\cancel{\text{g}} \times \frac{1\text{ mol}}{63.55\,\cancel{\text{g}}}$$

$$n_{Cu} = 0.157\text{ mol}$$

$$n_{AgNO_3} = 20.0\,\cancel{\text{g}} \times \frac{1\text{ mol}}{169.88\,\cancel{\text{g}}}$$

$$n_{AgNO_3} = 0.118\text{ mol}$$

Since this reactant mole ratio is 1:2, 0.118 mol of $AgNO_{3(aq)}$ will require 0.118 mol × 1/2 = 0.0589 mol of $Cu_{(s)}$ for reaction. The $Cu_{(s)}$ is in excess; so the $AgNO_{3(aq)}$ is the limiting reagent.

(b)
$$\begin{array}{lllllll} Cu_{(s)} & + & 2\,AgNO_{3(aq)} & \rightarrow & 2\,Ag_{(s)} & + & Cu(NO_3)_{2(aq)} \\ & & 20.0\text{ g} & & m & & \\ & & 169.88\text{ g/mol} & & 107.87\text{ g/mol} & & \end{array}$$

$$n_{AgNO_3} = 20.0\,\cancel{\text{g}} \times \frac{1\text{ mol}}{169.88\,\cancel{\text{g}}}$$

$$n_{AgNO_3} = 0.118\text{ mol}$$

$$n_{Ag} = 0.118\text{ mol} \times \frac{2}{2}$$

$$n_{Ag} = 0.118\text{ mol}$$

$$m_{Ag} = 0.118\,\cancel{\text{mol}} \times \frac{107.87\text{ g}}{1\,\cancel{\text{mol}}}$$

$$m_{Ag} = 12.7\text{ g}$$

or

$$m_{Ag} = 20.0\,\cancel{\text{g}}\,\cancel{\text{AgNO}_3} \times \frac{1\,\cancel{\text{mol}}\,\cancel{\text{AgNO}_3}}{169.88\,\cancel{\text{g}}\,\cancel{\text{AgNO}_3}} \times \frac{2\,\cancel{\text{mol}}\,\cancel{\text{Ag}}}{2\,\cancel{\text{mol}}\,\cancel{\text{AgNO}_3}} \times \frac{107.87\text{ g Ag}}{1\,\cancel{\text{mol}}\,\cancel{\text{Ag}}}$$

$$m_{Ag} = 12.7\text{ g}$$

The mass of silver crystals produced would be 12.7 g.

7. (a)

$$\begin{array}{lllll} \mathrm{FeCl_{3(aq)}} & + \quad 3\ \mathrm{NaOH_{(aq)}} & \rightarrow & \mathrm{Fe(OH)_{3(s)}} & + \quad 3\ \mathrm{NaCl_{(aq)}} \\ 26.8\ \mathrm{g} & \quad\ \ 21.5\ \mathrm{g} & & & \\ 162.20\ \mathrm{g/mol} & \quad\ \ 40.00\ \mathrm{g/mol} & & & \end{array}$$

$$n_{\mathrm{NaOH}} = 21.5\ \cancel{\mathrm{g}} \times \frac{1\ \mathrm{mol}}{40.00\ \cancel{\mathrm{g}}}$$

$$n_{\mathrm{NaOH}} = 0.538\ \mathrm{mol}$$

$$n_{\mathrm{FeCl_3}} = 26.8\ \cancel{\mathrm{g}} \times \frac{1\ \mathrm{mol}}{162.20\ \cancel{\mathrm{g}}}$$

$$n_{\mathrm{FeCl_3}} = 0.165\ \mathrm{mol}$$

Since this reactant mole ratio is 1:3, 0.165 mol of $FeCl_{3(aq)}$ will require 0.165 mol × 3/1 = 0.496 mol of $NaOH_{(aq)}$ for reaction. The $NaOH_{(aq)}$ is in excess; so the $FeCl_{3(aq)}$ is the limiting reagent.

(b)

$$\begin{array}{llll} \mathrm{FeCl_{3(aq)}} \ + \ 3\ \mathrm{NaOH_{(aq)}} & \rightarrow & \mathrm{Fe(OH)_{3(s)}} \ + & 3\ \mathrm{NaCl_{(aq)}} \\ 26.8\ \mathrm{g} & & m & m \\ 162.20\ \mathrm{g/mol} & & 106.88\ \mathrm{g/mol} & 58.44\ \mathrm{g/mol} \end{array}$$

$$n_{\mathrm{FeCl_3}} = 26.8\ \cancel{\mathrm{g}} \times \frac{1\ \mathrm{mol}}{162.20\ \cancel{\mathrm{g}}}$$

$$n_{\mathrm{FeCl_3}} = 0.165\ \mathrm{mol}$$

$$n_{\mathrm{Fe(OH)_3}} = 0.165\ \mathrm{mol} \times \frac{1}{1}$$

$$n_{\mathrm{Fe(OH)_3}} = 0.165\ \mathrm{mol}$$

$$m_{\mathrm{Fe(OH)_3}} = 0.165\ \cancel{\mathrm{mol}} \times \frac{106.88\ \mathrm{g}}{1\ \cancel{\mathrm{mol}}}$$

$$m_{\mathrm{Fe(OH)_3}} = 17.7\ \mathrm{g}$$

or

$$m_{\mathrm{Fe(OH)_3}} = 26.8\ \cancel{\mathrm{g}}\ \cancel{\mathrm{FeCl_3}} \times \frac{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{FeCl_3}}}{162.20\ \cancel{\mathrm{g}}\ \cancel{\mathrm{FeCl_3}}} \times \frac{1\ \cancel{\mathrm{mol}}\ \mathrm{Fe}\cancel{\mathrm{(OH)_3}}}{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{FeCl_3}}} \times \frac{106.88\ \mathrm{g\ Fe(OH)_3}}{1\ \cancel{\mathrm{mol}}\ \mathrm{Fe}\cancel{\mathrm{(OH)_3}}}$$

$$m_{\mathrm{Fe(OH)_3}} = 17.7\ \mathrm{g}$$

The mass of iron(III) hydroxide produced would be 17.7 g

and

$$n_{\mathrm{NaCl}} = 0.165\ \mathrm{mol} \times \frac{3}{1}$$

$$n_{\mathrm{NaCl}} = 0.496\ \mathrm{mol}$$

$$m_{\mathrm{NaCl}} = 0.496\ \cancel{\mathrm{mol}} \times \frac{58.44\ \mathrm{g}}{1\ \cancel{\mathrm{mol}}}$$

$$m_{\mathrm{NaCl}} = 29.0\ \mathrm{g}$$

or

$$m_{\mathrm{NaCl}} = 26.8\ \cancel{\mathrm{g}}\ \cancel{\mathrm{FeCl_3}} \times \frac{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{FeCl_3}}}{162.20\ \cancel{\mathrm{g}}\ \cancel{\mathrm{FeCl_3}}} \times \frac{3\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{NaCl}}}{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{FeCl_3}}} \times \frac{58.44\ \mathrm{g\ NaCl}}{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{NaCl}}}$$

$$m_{\mathrm{NaCl}} = 29.0\ \mathrm{g}$$

The mass of sodium chloride produced would be 29.0 g.

Applying Inquiry Skills

8. Determining if enough excess reagent has been added is the same as determining if all of the limiting reagent has reacted. This can be done in any situation where a product can be removed from the reaction. For example, in a reaction of zinc and hydrochloric acid, where hydrogen bubbles escape, if more zinc is added and no bubbles form, then all of the hydrochloric acid must have reacted.