Observations

(e) mass of 250-mL beaker = 102.41 g
mass of 250-mL beaker + $SrCl_{2(s)}$ = 104.41 g
mass of 150-mL beaker = 68.83 g
mass of 150-mL beaker + $CuSO_4 \cdot 5H_2O_{(s)}$ = 71.39 g
mass of filter paper = 0.90 g
mass of filter paper and precipitate = 3.24 g
The blue copper(II) sulfate solution mixed with the strontium chloride solution to produce a white precipitate and a blue solution.

Analysis

(f) mass of precipitate = 3.24 g – 0.90 g = 2.34 g
(g) According to the observations collected in this experiment, the mass of the strontium sulfate precipitate from the reaction of strontium chloride and copper(II) sulfate solutions was determined to be 2.34 g.

Evaluation

(h) The major source of error is likely caused by the number of mass measurements made. Making more measurements would help to reduce the error.
(i) The filtration design is adequate, and there are no obvious improvements to be made. Adequate care was taken in filtering and washing the precipitate.

$$\% \text{ difference} = \frac{2.34 \text{ g} - 2.32 \text{ g}}{2.32 \text{ g}} \times 100$$

$$\% \text{ difference} = 0.86\%$$

Based on the very low percent difference, the Prediction is valid.
(j) The stoichiometric method is valid because the Prediction was verified.

2.12 PERCENTAGE YIELD

PRACTICE

(Pages 158–159)

Understanding Concepts

1. $m_{C_7H_6O_3}$ used $= 2.00$ g $C_7H_6O_3$

actual yield $= 1.65$ g $C_8H_8O_3$

Balanced equation	$C_7H_6O_{3(s)}$ +	$CH_3OH_{(l)}$ →	$C_8H_8O_{3(l)}$ +	$H_2O_{(l)}$
Given mass (g)	2.00		1.65	
Molar mass (g/mol)	138.13	32.05	152.16	18.02

$$n_{C_7H_6O_3} = 2.00\ \cancel{g\ C_7H_6O_3} \times \frac{1\ mol\ C_7H_6O_3}{138.13\ \cancel{g\ C_7H_6O_3}}$$

$$n_{C_7H_6O_3} = 0.01448\ mol\ C_7H_6O_3$$

$$n_{C_8H_8O_3} = 0.01448\ \cancel{mol\ C_7H_6O_3} \times \frac{1\ mol\ C_8H_8O_3}{1\ \cancel{mol\ C_7H_6O_3}}$$

$$n_{C_8H_8O_3} = 0.01448\ mol\ C_8H_8O_3$$

$$m_{C_8H_8O_3} = 0.01448\ \cancel{mol\ C_8H_8O_3} \times \frac{152.16\ g\ C_8H_8O_3}{1\ \cancel{mol\ C_8H_8O_3}}$$

$$m_{C_8H_8O_3} = 2.203\ g\ C_8H_8O_3$$

The combined calculation is as follows:

$$m_{C_8H_8O_3} = 2.00\ \cancel{g\ C_7H_6O_3} \times \frac{1\ \cancel{mol\ C_7H_6O_3}}{138.13\ \cancel{g\ C_7H_6O_3}} \times \frac{1\ \cancel{mol\ C_8H_8O_3}}{1\ \cancel{mol\ C_7H_6O_3}} \times \frac{152.16\ g\ C_8H_8O_3}{1\ \cancel{mol\ C_8H_8O_3}}$$

$$m_{C_8H_8O_3} = 2.203\ g\ C_8H_8O_3$$

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{1.65\ \cancel{g}}{2.203\ \cancel{g}} \times 100\%$$

$$\text{percentage yield} = 74.9\%$$

The percentage yield of wintergreen is 74.9%.

2. (a) The balanced chemical equation is
$Al_{(s)} + Br_{2(l)} \rightarrow AlBr_{2(s)}$

(b) $m_{Br_2} = 53.7\ g\ Br$

actual yield = 48.4 g $AlBr_3$

Balanced chemical equation	$2\ Al_{(s)}$	+ $3\ Br_{2(l)}$ $\rightarrow$	$2\ AlBr_{3(s)}$
Given mass (g)		53.7	48.4
Molar mass (g/mol)	26.98	159.80	266.68

$$n_{Br_2} = 53.7\ \cancel{g\ Br_2} \times \frac{1\ mol\ Br_2}{159.80\ \cancel{g\ Br_2}}$$
$$n_{Br_2} = 0.336\ mol\ Br_2$$

$$n_{AlBr_3} = 0.336\ \cancel{mol\ Br_2} \times \frac{2\ mol\ AlBr_3}{3\ \cancel{mol\ Br_2}}$$
$$n_{AlBr_3} = 0.224\ mol\ AlBr_3$$

$$m_{AlBr_3} = 0.224\ \cancel{mol\ AlBr_3} \times \frac{266.68\ g\ AlBr_3}{1\ \cancel{mol\ AlBr_3}}$$
$$m_{AlBr_3} = 59.74\ g\ AlBr_3$$

The combined calculation is as follows:

$$m_{AlBr_3} = 53.7\ \cancel{g\ Br_2} \times \frac{1\ \cancel{mol\ Br_2}}{159.80\ \cancel{g\ Br_2}} \times \frac{2\ \cancel{mol\ AlBr_3}}{3\ \cancel{mol\ Br_2}} \times \frac{266.68\ g\ AlBr_3}{1\ \cancel{mol\ AlBr_3}}$$
$$m_{AlBr_3} = 59.74\ g\ AlBr_3$$

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$
$$= \frac{48.4\ \cancel{g}}{59.74\ \cancel{g}} \times 100\%$$
$$\text{percentage yield} = 81.0\%$$

The percentage yield of aluminum bromide is 81.0%.

3. $m_{HCl} = 0.999\ g\ HCl$

$\text{actual yield} = 1.541\ g\ ZnCl_2$

Balanced equation	$Zn_{(s)}$ +	$2\ HCl_{(aq)}$ →	$ZnCl_{2(aq)}$ +	$H_{2(g)}$
Given mass (g)		0.999	1.541	
Molar mass (g/mol)		36.46	136.28	

$$n_{HCl} = 0.999\ \cancel{g\ HCl} \times \frac{1\ mol\ HCl}{36.46\ \cancel{g\ HCl}}$$
$$n_{HCl} = 0.0274\ mol\ HCl$$

$$n_{ZnCl_2} = 0.0274\ \cancel{mol\ HCl} \times \frac{1\ mol\ ZnCl_2}{2\ \cancel{mol\ HCl}}$$
$$n_{ZnCl_2} = 0.0137\ mol\ ZnCl_2$$

$$m_{ZnCl_2} = 0.0137\ \cancel{mol\ ZnCl_2} \times \frac{136.28\ g\ ZnCl_2}{1\ \cancel{mol\ ZnCl_2}}$$
$$m_{ZnCl_2} = 1.867\ g\ ZnCl_2$$

The combined calculation is as follows:

$$m_{ZnCl_2} = 0.999\ \cancel{g\ HCl} \times \frac{1\ \cancel{mol\ HCl}}{36.46\ \cancel{g\ HCl}} \times \frac{1\ \cancel{mol\ ZnCl_2}}{2\ \cancel{mol\ HCl}} \times \frac{136.28\ g\ ZnCl_2}{1\ \cancel{mol\ ZnCl_2}}$$

$$m_{ZnCl_2} = 1.867\ g\ ZnCl_2$$

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{1.541\ \cancel{g\ ZnCl_2}}{1.867\ \cancel{g\ ZnCl_2}} \times 100\%$$

$$\text{percentage yield} = 82.5\%$$

The percentage yield of zinc chloride is 82.5%.

4. (a) $m_{C_7H_6O_3} = 213.0\ g\ C_7H_6O_3$

$\text{actual yield} = 189.3\ g\ C_9H_8O_4$

Balanced equation	$C_7H_6O_{3(s)}$ +	$C_4H_6O_{3(aq)}$ $\rightarrow$	$C_9H_8O_{4\,(s)}$ +	$HC_2H_3O_{2(aq)}$
Given mass (g)	213.0		189.3	
Molar mass (g/mol)	138.13		180.17	

$$n_{C_7H_6O_3} = 213.0\ \cancel{g\ C_7H_6O_3} \times \frac{1\ mol\ C_7H_6O_3}{138.13\ \cancel{g\ C_7H_6O_3}}$$

$$n_{C_7H_6O_3} = 1.542\ mol\ C_7H_6O_3$$

$$n_{C_9H_8O_4} = 1.542\ \cancel{mol\ C_7H_6O_3} \times \frac{1\ mol\ C_9H_8O_4}{1\ \cancel{mol\ C_7H_6O_3}}$$

$$n_{C_9H_8O_4} = 1.542\ mol\ C_9H_8O_4$$

$$m_{C_9H_8O_4} = 1.542\ \cancel{mol\ C_9H_8O_4} \times \frac{180.17\ g\ C_9H_8O_4}{1\ \cancel{mol\ C_9H_8O_4}}$$

$$m_{C_9H_8O_4} = 277.8\ g\ C_9H_8O_4$$

The theoretical yield of aspirin is 277.8 g.

The combined calculation is as follows:

$$m_{C_9H_8O_4} = 213.0\ \cancel{g\ C_7H_8O_3} \times \frac{1\ \cancel{mol C_7H_6O_3}}{138.13\ \cancel{gC_7H_6O_3}} \times \frac{1\ \cancel{mol\ C_9H_8O_4}}{1\ \cancel{mol\ C_7H_6O_3}} \times \frac{180.17\ g\ C_9H_8O_4}{1\ \cancel{mol\ C_9H_8O_4}}$$

$$m_{C_9H_8O_4} = 277.8\ g\ C_9H_8O_4$$

The theoretical yield of aspirin is 277.8 g.

(b) $$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$
$$= \frac{189.3\ \cancel{g}}{277.8\ \cancel{g}} \times 100\%$$
$$\text{percentage yield} = 68.14\%$$

The percentage yield of aspirin is 68.14%.

SECTION 2.12 QUESTIONS

(Page 159)

Understanding Concepts

1. Actual yield is the mass of a product measured when a chemical reaction is carried out. Theoretical yield is the mass of a product calculated using the information provided by a balanced chemical equation. It is the mass that is expected to be obtained if there is no experimental error.
2. No, the actual yield cannot truly be greater than the theoretical yield. The actual yield may be greater than the theoretical due to experimental error, such as errors in determining the mass or including by-products as part of the final mass.
3. $m_{AgNO_3} = 5.00\text{ g } AgNO_3$

   $\text{actual yield} = 5.03\text{ g } AgBr$

   (a) The balanced chemical equation for the reaction is:

   $$AgNO_{3(s)} + NaBr_{(s)} \rightarrow AgBr_{(s)} + NaNO_{3(s)}$$

   (b)

Balanced equation	$AgNO_{3(s)}$	+ $NaBr_{(s)}$ →	$AgBr_{(s)}$	+ $NaNO_{3(s)}$
Given mass (g)	5.00		5.03	
Molar mass (g/mol)	169.88		187.77	

$$n_{AgNO_3} = 5.00\ \cancel{\text{g } AgNO_3} \times \frac{1\text{ mol } AgNO_3}{169.88\ \cancel{\text{g } AgNO_3}}$$
$$n_{AgNO_3} = 0.02943\text{ mol } AgNO_3$$

$$n_{AgBr} = 0.02943\ \cancel{\text{mol } AgNO_3} \times \frac{1\text{ mol } AgBr}{1\ \cancel{\text{mol } AgNO_3}}$$
$$n_{AgBr} = 0.02943\text{ mol } AgBr$$

$$m_{AgBr} = 0.02943\ \cancel{\text{mol } AgBr} \times \frac{187.77\text{ g } AgBr}{1\ \cancel{\text{mol } AgBr}}$$
$$m_{AgBr} = 5.53\text{ g } AgBr$$

The theoretical yield of silver bromide is 5.53 g.

The combined calculation is as follows:

$$m_{AgBr} = 5.00\ \cancel{\text{g } AgNO_3} \times \frac{1\ \cancel{\text{mol } AgNO_3}}{169.88\ \cancel{\text{g } AgNO_3}} \times \frac{1\ \cancel{\text{mol } AgBr}}{1\ \cancel{\text{mol } AgNO_3}} \times \frac{187.77\text{ g } AgBr}{1\ \cancel{\text{mol } AgBr}}$$
$$m_{AgBr} = 5.53\text{ g } AgBr$$

The theoretical yield of silver bromide is 5.53 g.

(c) The actual yield of silver bromide is given as 5.03 g.

(d) $$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$
$$= \frac{5.03\ \cancel{\text{g}}\ \text{AgBr}}{5.53\ \cancel{\text{g}}\ \text{AgBr}} \times 100\%$$
$$\text{percentage yield} = 91.0\%$$
The percentage yield of silver bromide is 91.0%.

4. $m_{FeS} = 16.1\ \text{g FeS}$
$\text{actual yield} = 14.1\ \text{g}\ Fe_2O_3$

Balanced equation	$4\ FeS_{(s)}$	$+\ 7\ O_{2(g)}$	$\rightarrow\ 2\ Fe_2O_{3(s)}$	$+\ 4\ SO_{2(g)}$
Given mass (g)	16.1		14.1	
Molar mass (g/mol)	87.91		159.70	

$$n_{FeS} = 16.1\ \cancel{\text{g FeS}} \times \frac{1\ \text{mol FeS}}{87.91\ \cancel{\text{g FeS}}}$$
$$n_{FeS} = 0.1831\ \text{mol FeS}$$

$$n_{Fe_2O_3} = 0.1831\ \cancel{\text{mol FeS}} \times \frac{2\ \text{mol}\ Fe_2O_3}{4\ \cancel{\text{mol FeS}}}$$
$$n_{Fe_2O_3} = 0.09155\ \text{mol}\ Fe_2O_3$$

$$m_{Fe_2O_3} = 0.09155\ \cancel{\text{mol}\ Fe_2O_3} \times \frac{159.70\ \text{g}\ Fe_2O_3}{1\ \cancel{\text{mol}\ Fe_2O_3}}$$
$$m_{Fe_2O_3} = 14.62\ \text{g}\ Fe_2O_3$$

The combined calculation is as follows:
$$m_{Fe_2O_3} = 16.1\ \cancel{\text{g FeS}} \times \frac{1\ \cancel{\text{mol FeS}}}{87.91\ \cancel{\text{g FeS}}} \times \frac{2\ \cancel{\text{mol}\ Fe_2O_3}}{4\ \cancel{\text{mol FeS}}} \times \frac{159.70\ \text{g}\ Fe_2O_3}{1\ \cancel{\text{mol}\ Fe_2O_3}}$$
$$m_{Fe_2O_3} = 14.62\ \text{g}\ Fe_2O_3$$

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$
$$= \frac{14.1\ \cancel{\text{g}}}{14.62\ \cancel{\text{g}}} \times 100\%$$
$$\text{percentage yield} = 96.4\%$$
The percentage yield of iron(III) oxide is 96.4%.

Applying Inquiry Skills

5. The percentage yield may be increased by measuring the mass of the precipitate in the evaporating dish instead of using weighing paper. (The mass of the evaporating dish would have to be known and subtracted from the total mass of the evaporating dish with the precipitate.) Aqueous reactants are lost because they stick to the surfaces of their containers. Limiting the number of container transfers may also increase percentage yield.

Making Connections

6. (a) The typical percentage yield of the carmine extraction process is 23% of 62% pure carmine. Carmine is extracted by boiling the insects in water, followed by filtration, precipitation, and washing and drying the final product.
   (b) Cochineal processing plants have boosted local employment in Peru. While they earn only about 10% of the revenue generated from cochineal processing, an estimated 50 000 people harvest the insects by hand, dry them in the sun, and sell them to carmine processors in Lima, Peru's capital city. It is estimated that as many as 400 000 rural families depend on this industry for their livelihood.

2.13 INVESTIGATION: PERCENTAGE YIELD OF A CHEMICAL REACTION

(Page 160)

Prediction

(a) mass of copper(II) chloride dihydrate used = 2.00 g

$$m_{Cu} = 2.00\ \cancel{\text{g CuCl}_2\text{•2H}_2\text{O}} \times \frac{1\ \cancel{\text{mol CuCl}_2\text{•2H}_2\text{O}}}{170.49\ \cancel{\text{g CuCl}_2\text{•2H}_2\text{O}}} \times \frac{3\ \cancel{\text{mol Cu}}}{3\ \cancel{\text{mol CuCl}_2\text{•2H}_2\text{O}}} \times \frac{63.55\text{ g Cu}}{1\ \cancel{\text{mol Cu}}}$$

$$m_{Cu} = 0.745\text{ g Cu}$$

The theoretical yield of copper is 0.745 g.
It is predicted that the actual yield will be approximately 10% lower than the theoretical yield, or 0.671 g.

Qualitative Observations

A reddish residue of copper formed immediately when the foil was added to the solution.
Vigorous bubbling was observed on the aluminum surface as the reaction started.
Some of the copper changed to a green colour after it was dried overnight.
The aluminum foil broke up into many small pieces as the reaction occurred.

Quantitative Observations

mass of copper(II) chloride dihydrate = 2.00 g
mass of 150-mL beaker and copper residue = 86.38 g
mass of empty 150-mL beaker = 85.97 g
mass of copper residue collected = 86.38 g – 85.97 g
mass of copper residue collected = 0.410 g

Analysis

(b) There were many small pieces of aluminum left in the beaker when the reaction had come to a halt. Therefore, aluminum is the excess reagent and copper(II) chloride dihydrate is the limiting reagent.
(c) 0.410 g of copper was formed when 2.00 g of copper(II) chloride dihydrate reacted with excess aluminum.
(d)

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{0.410\ \cancel{\text{g}}}{0.745\ \cancel{\text{g}}} \times 100\%$$

$$\text{percentage yield} = 55.0\%$$

The percentage yield of copper is 55.0%.