## Chapter 3: Stoichiometry I

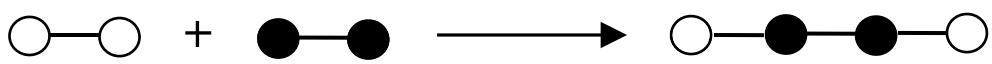
**3.2 *See Section 3.1.***

1. Assume the equation contains one formula unit of the most complicated substance, and bring the atoms of this substance into balance by adjusting the coefficients of the substances on the other side of the equation.

2. Continue by adjusting the coefficients of the reactants and products until the same numbers of each type of atom appear on both sides of the equation and the coefficients are given in terms of whole numbers.

3. Check to make sure your final equation contains the same numbers of atoms of each type on both sides.

**3.4 *See Section. 3.1***



$H_2$ $O_2$ $H_2O_2$

**3.6 *See Section 3.2.***

The number of objects in one mole has been determined experimentally to be $6.022 \times 10^{23}$ (when expressed to four significant figures) and is known as Avogadro's number.

**3.8 *See Section 3.2 and Table 3.2.***

The units for atomic, molecular and formula masses are atomic mass units (μ), and the units for molar masses are grams per mole (g/mol).

**3.10 *See Section 3.2.***

$$\text{moles} \xrightarrow{\text{Avogadro's number}} \text{atoms}$$

where the number of moles is multiplied by $\dfrac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}$ .

**3.12 *See Section 3.3, Figure 3.4, and Question 3.9.***

The masses of C and H are obtained as outlined in 3.9. These masses are subtracted from the total mass of sample burned to obtain the mass of O in the sample. The percentage O by mass is calculated using (mass of O/mass of sample) × 100%.

**3.14 *See Section 3.4.***

One mol of $N_2$ reacts with two mol $H_2$ to form one mole $N_2H_4$.

**3.16** ***See Section 3.4.***

"The reaction was carried out with the reactants present in stoichiometric amounts," means the reactants were present in the exact amounts necessary to react with one another and none was present in excess.

**3.18** ***See Section 3.4***

The statement, "$C_2H_4$ is the limiting reactant and oxygen is present in excess in a combustion reaction," means the amount of carbon dioxide and water that can be produced by the reaction is limited by the amount of $C_2H_4$ present. It also means the $C_2H_4$ is completely consumed and an excess of oxygen remains when the reaction is complete.

**3.20** ***See Section. 3.1***

(a) $2SO_2 + O_2 \rightarrow 2SO_3$

(b) sulfur trioxide

**3.22** ***See Section 3.1 and Examples 3.1, 3.2.***

(a) Unbalanced: $Mg_3N_2 + H_2O \rightarrow NH_3 + Mg(OH)_2$ Start with $Mg_3N_2$.

Step 1: $Mg_3N_2 + H_2O \rightarrow \underline{2}NH_3 + \underline{3}Mg(OH)_2$ Balances Mg, N.

Step 2: $Mg_3N_2 + \underline{6}H_2O \rightarrow 2NH_3 + 3Mg(OH)_2$ Balances H,O.

(b) Unbalanced: $Fe + O_2 \rightarrow Fe_2O_3$ Start with Fe.

Step 1: $\underline{2}Fe + O_2 \rightarrow Fe_2O_3$ Balances Fe.

Step 2: $2Fe + \underline{\tfrac{3}{2}}\ O_2 \rightarrow Fe_2O_3$ Balances O.

Step 3: $\underline{4}Fe + \underline{3}O_2 \rightarrow \underline{2}Fe_2O_3$ Gives whole number coefficients.

(c) Unbalanced: $Zn + H_3PO_4 \rightarrow H_2 + Zn_3(PO_4)_2$ Start with $H_3PO_4$.

Step 1: $Zn + \underline{2}H_3PO_4 \rightarrow H_2 + Zn_3(PO_4)_2$ Balances $PO_4$ units.

Step 2: $\underline{3}Zn + 2H_3PO_4 \rightarrow H_2 + Zn_3(PO_4)_2$ Balances Zn.

Step 3: $3Zn + 2H_3PO_4 \rightarrow 3H_2 + Zn_3(PO_4)_2$ Balances H.

**3.24** ***See Section. 3.1***

(a)? Al(s) + ? $O_2$(g) $\longrightarrow$ ? $Al_2O_3$(s) Select order: Al then O

2 Al(s) + ? $O_2$(g) $\longrightarrow$ 1 $Al_2O_3$(s) 2 Al

2 Al(s) + $\frac{3}{2}$ $O_2$(g) $\longrightarrow$ 1 $Al_2O_3$(s) 3 O

4 Al(s) + 3 $O_2$(g) $\longrightarrow$ 2 $Al_2O_3$(s)

(b)? $N_2$(g) + ? $H_2$(g) $\longrightarrow$ ? $NH_3$(g) Select order: H then N

? $N_2$(g) + 3 $H_2$(g) $\longrightarrow$ 2 $NH_3$(g) 6 H

1 $N_2$(g) + 3 $H_2$(g) $\longrightarrow$ 2 $NH_3$(g) 2 N

(c)? $C_6H_6(\ell)$ + ? $O_2(g) \longrightarrow$ ? $H_2O(\ell)$ + ? $CO_2(g)$ Select order: C, H then O

$C_6H_6(\ell)$ + ? $O_2(g) \longrightarrow$ ? $H_2O(\ell) + 6\ CO_2(g)$ 6 C

$C_6H_6(\ell)$ + ? $O_2(g) \longrightarrow 3\ H_2O(\ell) + 6\ CO_2(g)$ 6 H

$C_6H_6(\ell) + \frac{15}{2}\ O_2(g) \longrightarrow 3\ H_2O(\ell) + 6\ CO_2(g)$ 15 O

$2\ C_6H_6(\ell) + 15\ O_2(g) \longrightarrow 6\ H_2O(\ell) + 12\ CO_2(g)$

**3.26** ***See Section. 3.1.***

(a) HF (g) $\xrightarrow{H_2O} H^+ (aq) + F^- (aq)$

(b) $Li_2SO_4$ (s) $\xrightarrow{H_2O} 2\ Li^+ (aq) + SO_4^{2-} (aq)$

**3.28** ***See Section. 3.1.***

(a) HCl (g) $\xrightarrow{H_2O} H^+ (aq) + Cl^- (aq)$

(b) $Ba(NO_3)_2$ (s) $\xrightarrow{H_2O} Ba^{2+} (aq) + 2NO_3^- (aq)$

**3.30** ***See Section 3.1 and Example 3.3.***

(a) Unbalanced: $Mg(OH)_2 + HF \rightarrow MgF_2 + H_2O$ Start with $MgF_2$.

Step 1: $Mg(OH)_2 + \underline{2}HF \rightarrow MgF_2 + H_2O$ Balances F.

Step 2: $Mg(OH)_2 + 2HF \rightarrow MgF_2 + \underline{2}H_2O$ Balances H, O.

(b) Unbalanced: $HCl + NaOH \rightarrow H_2O + NaCl$ Start with NaCl.

Step 1: $HCl + NaOH \rightarrow H_2O + NaCl$ Balances H, O, Na, Cl

(c) Unbalanced: $H_2SO_4 + Sr(OH)_2 \rightarrow H_2O + SrSO_4$ Start with $H_2O$.

Step 1: $H_2SO_4 + Sr(OH)_2 \rightarrow \underline{2}H_2O + SrSO_4$ Balances H, O.

**3.32** ***See Section. 3.1.***

(a) Unbalanced: ? $C_4H_{10}(g)$ + ? $O_2(g) \longrightarrow$ ? $CO_2(g)$ + ? $H_2O(g)$

Balance C: $C_4H_{10}(g)$ + ? $O_2(g) \longrightarrow 4\ CO_2(g)$ + ? $H_2O(g)$ 4 C’s

Balance H: $C_4H_{10}(g)$ + ? $O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$ 10 H’s

Balance O: $C_4H_{10}(g) + \frac{13}{2}\ O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$ 13 O’s

Multiply by 2: $2\ C_4H_{10}(g) + 13\ O_2(g) \longrightarrow 8\ CO_2(g) + 10\ H_2O(g)$

(b) Unbalanced: ? $C_6H_{12}O_6(s)$ + ? $O_2(g) \longrightarrow$ ? $CO_2(g)$ + ? $H_2O(g)$

Balance C: $C_6H_{12}O_6(s)$ + ? $O_2(g) \longrightarrow 6\ CO_2(g)$ + ? $H_2O(g)$6 C’s

Balance H: $C_6H_{12}O_6(s)$ + ? $O_2(g) \longrightarrow 6\ CO_2(g) + 6\ H_2O(g)$ 12 H’s

Balance O: $C_6H_{12}O_6(s) + 6\ O_2(g) \longrightarrow 6\ CO_2(g) + 6\ H_2O(g)$ 18 O’s

(c) Unbalanced: ? $C_4H_8O(\ell)$ + ? $O_2(g) \longrightarrow$ ? $CO_2(g)$ + ? $H_2O(g)$

Balance C: $C_4H_8O(\ell)$ + ? $O_2(g) \longrightarrow 4\ CO_2(g)$ + ? $H_2O(g)$ — 4 C's

Balance H: $C_4H_8O(\ell)$ + ? $O_2(g) \longrightarrow 4\ CO_2(g) + 4\ H_2O(g)$ — 8 H's

Balance O: $C_4H_8O(\ell) + \frac{11}{2}\ O_2(g) \longrightarrow 4\ CO_2(g) + 4\ H_2O(g)$ — 12 O's

Multiply by 2: $2\ C_4H_8O(\ell) + 11\ O_2(g) \longrightarrow 8\ CO_2(g) + 8\ H_2O(g)$

**3.34** ***See Section 3.1 and Examples 3.3, 3.4.***

(a) Unbalanced: $C_8H_8 + O_2 \rightarrow CO_2 + H_2O$ — Start with $C_8H_8$.
Step 1: $C_8H_8 + O_2 \rightarrow \underline{8}CO_2 + \underline{4}H_2O$ — Balances C, H.
Step 2: $C_8H_8 + 10O_2 \rightarrow 8CO_2 + 4H_2O$ — Balances O.

(b) Unbalanced: $HCl + KOH \rightarrow H_2O + KCl$ — Start with KCl.
Step 1: $HCl + KOH \rightarrow H_2O + KCl$ — Balances H, O, K, Cl.

**3.36** ***See Section 3.1 and Example 3.4.***

Unbalanced: $CH_3OC(CH_3)_3 + O_2 \rightarrow CO_2 + H_2O$ — Start with $CH_3OC(CH_3)_3$.
Step 1: $CH_3OC(CH_3)_3 + O_2 \rightarrow \underline{5}CO_2 + \underline{6}H_2O$ — Balance C, H.
Step 2: $CH_3OC(CH_3)_3 + \frac{15}{2}O_2 \rightarrow 5CO_2 + 6H_2O$ — Balances O.
Step 3: $\underline{2}CH_3OC(CH_3)_3 + \underline{15}O_2 \rightarrow \underline{10}CO_2 + \underline{12}H_2O$ — Gives whole number coefficients.

**3.38** ***See Section 3.1 and Example 3.3.***

Unbalanced: $H_3PO_3 \rightarrow H_3PO_4 + PH_3$ — Start with $H_3PO_3$ and $H_3PO_4$.
Step 1: $\underline{4}H_3PO_3 \rightarrow \underline{3}H_3PO_4 + PH_3$ — Balances O, P, H.

**3.40** ***See Section 3.1 and Examples 3.1, 3.2.***

Unbalanced: $UO_2 + CCl_4 \rightarrow UCl_4 + COCl_2$ — Start with $CCl_4$.
Step 1: $UO_2 + \underline{2}CCl_4 \rightarrow UCl_4 + \underline{2}COCl_2$ — Balances Cl, C, O.

**3.42** ***See Section. 3.1.***

(a) $NH_4Cl$: Assume hydrogen has an an oxidation number of +1 and chlorine -1; therefore N has an oxidation number of -3.

(b) $N_2O$: Assume oxygen has an oxidation number of -2; therefore each nitrogen has an oxidation state of +1.

(c) Ag: All elements in their elemental state have an oxidation number of zero

(d) $AuI_3$: Assume each iodine has an oxidation number of -1; therefore gold must have an oxidation number of +3.

**3.44** ***See Section. 3.1.***

$Na_2O_2$: Assume that each sodium has an oxidation number of +1; therefore each oxygen must have an oxidation number of -1.

**3.46** ***See Section 3.1 and Example 3.5.***

Unbalanced: $Fe + O_2 \rightarrow Fe_2O_3$ Start with $O_2$.
Step 1: $Fe + \underline{3}O_2 \rightarrow \underline{2}Fe_2O_3$ Balances O.
Step 2: $\underline{4}Fe + 3O_2 \rightarrow 2Fe_2O_3$ Balances Fe.
Assigning an ON of -2 to oxygen in $Fe_2O_3$ gives $4Fe + 3O_2 \rightarrow 2FeO_3$
ON: 0 0 +3,-2
The ON of the Fe atoms increase, and the ON of the O atoms decrease.
The Fe metal is oxidized, and the O atoms in $O_2$ are reduced.

**3.48** ***See Section 3.1 and Example 3.5.***

Unbalanced: $P_4(s) + O_2(g) \rightarrow P_4O_{10}(s)$ Start with $O_2$.
Step 1: $P_4(s) + \underline{5}O_2(g) \rightarrow P_4O_{10}(s)$ Balances O.
Assigning an ON of –2 to O in $P_4O_{10}$ gives $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$
ON 0 0 +5,–2
The ON of the P atoms increases, and the ON of the O atoms decreases.
The P atoms in $P_4$ are oxidized, and the O atoms in $O_2$ are reduced.

**3.50** ***See Section 3.1 and Example 3.5.***

Unbalanced: $MnO_2 + HCl \rightarrow MnCl_2 + Cl_2 + H_2O$ Start with HCl.
Step 1: $MnO_2 + \underline{4}HCl \rightarrow MnCl_2 + Cl_2 + H_2O$ Balances Cl.
Step 2: $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$ Balances H, O.
Assigning an ON of –2 to O in $MnO_2$ and $H_2O$, an ON of +1 to H in HCl and $H_2O$ and an ON of –1 to Cl in HCl and $MnCl_2$ gives $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$
ON: +4,–2 +1,–1 +2,–1 0 +1,–2
The ON of some of the Cl atoms in HCl increases, and the ON of the Mn atoms decreases.
The Cl atoms in HCl that undergo an increase in ON are oxidized, and the Mn atoms in $MnO_2$ are reduced.

**3.52** ***See Section 3.2 and Example 3.6.***

(a) $? \text{ atoms Xe} = 0.0778 \text{ mol Xe} \times \dfrac{6.022 \times 10^{23} \text{ atoms Xe}}{1 \text{ mol Xe}} = \mathbf{4.69 \times 10^{23} \textbf{ atoms Xe}}$

(b) $? \text{ atoms K} = 1.45 \text{ mol K} \times \dfrac{6.022 \times 10^{23} \text{ atoms K}}{1 \text{ mol K}} \times \mathbf{8.73 \times 10^{23} \textbf{ atoms K}}$

(c) $? \text{ atoms Ti} = 55.8 \text{ mol Ti} \times \dfrac{6.022 \times 10^{23} \text{ atoms Ti}}{1 \text{ mol Ti}} = \mathbf{3.36 \times 10^{25} \textbf{ atoms Ti}}$

**3.54 *See Section 3.2 and Example 3.6.***

(a) ? molecules $Cl_2$ = 0.223 mol $Cl_2$ × $\frac{6.022 \times 10^{23} \text{ molecules } Cl_2}{1 \text{ mol } Cl_2}$ = **1.34 × 10$^{23}$ molecules $Cl_2$**

(b) ? molecules $N_2H_4$ = 14.7 mol $N_2H_4$ × $\frac{6.022 \times 10^{23} \text{ molecules } N_2H_4}{1 \text{ mol } N_2H_4}$ = **8.85 × 10$^{24}$ molecules $N_2H_4$**

(c) ? molecules $C_9H_{18}$ = 0.334 mol $C_9H_{18}$ × $\frac{6.022 \times 10^{23} \text{ molecules } C_9H_{18}}{1 \text{ mol } C_9H_{18}}$ = **2.01 × 10$^{23}$ molecules $C_9H_{18}$**

(d) ? molecules $CO_2$ = 1.22 mol $CO_2$ × $\frac{6.022 \times 10^{23} \text{ molecules } CO_2}{1 \text{ mol } CO_2}$ = **7.35 × 10$^{23}$ molecules $CO_2$**

**3.56 *See Section 3.2 and Example 3.6.***

(a) ? mol $Br_2$ = 1.33 × 10$^{26}$ molecules $Br_2$ × $\frac{1 \text{ mol } Br_2}{6.022 \times 10^{23} \text{ molecules } Br_2}$ = **221 mol $Br_2$**

(b) ? mol $C_5H_{12}$ = 7.71 × 10$^{26}$ molecules $C_5H_{12}$ × $\frac{1 \text{ mol } C_5H_{12}}{6.022 \times 10^{23} \text{ molecules } C_5H_{12}}$ = **1.28 × 10$^{3}$ mol $C_5H_{12}$**

(c) ? mol $B_2H_6$ = 2.34 × 10$^{23}$ molecules $B_2H_6$ × $\frac{1 \text{ mol } B_2H_6}{6.022 \times 10^{23} \text{ molecules } B_2H_6}$ = **0.389 mol $B_2H_6$**

(d) ? mol Ne = 7.76 × 10$^{23}$ atoms Ne × $\frac{1 \text{ mol Ne}}{6.022 \times 10^{23} \text{ atoms Ne}}$ = **1.29 mol Ne**

**3.58 *See Section 3.2 and Example 3.7.***

(a) Molecular mass for $N_2O_4$:
2[N] × 14.01 = 28.02
4[O] × 16.00 = 64.00
92.02; Hence, molar mass for $N_2O_4$ is **92.02 g/mol.**

(b) Formula mass for $Na_2SO_4$:
1[Na] × 22.99 = 45.98
2[S] × 32.07 = 32.07
4[O] × 16.00 = 64.00
142.05; Hence, molar mass for $Na_2SO_4$ is **142.05 g/mol.**

(c) Molecular mass for $C_6H_{10}O_2$:
6[C] × 12.01 = 72.06
10[H] × 1.01 = 10.10
2[O] × 16.00 = 32.0
114.16; Hence, molar mass for $C_6H_{10}O_2$ is **114.16 g/mol.**

**3.60 *See Section 3.2 and Example 3.7.***

(a) Molecular mass $N_2O_2$: 2[N] × 14.01 = 28.02

$2[O] \times 16.00 \quad \underline{= 32.00}$

60.02 Hence, molar mass for $N_2O_2$ is **60.02 g/mol.**

(b) Formula mass for $(NH_4)CO_3$: $2[N] \times 14.01 = 28.02$

$8[H] \times 1.01 = 8.08$

$1[C] \times 12.01 = 12.01$

$3[O] \times 16.00 \quad \underline{= 48.00}$

96.11 Hence, molar mass for $(NH_4)_2CO_3$ is **96.11 g/mol.**

(c) Molecular mass for $C_8H_{15}N$: $8[C] \times 12.01 = 96.08$

$15[H] \times 1.01 = 15.15$

$1[N] \times 14.01 \quad \underline{= 14.01}$

125.24 Hence, molar mass for $C_8H_{15}N$ is **125.24 g/mol.**

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|---|---|
| **3.62** | ***See Section. 3.2.*** |

(a) Molar mass $H_2SO_4$ = 2 × (1.0079 g) + (32.065 g) + 4 × (15.9994 g) = 98.078 g/mol

$$39.2\ g\ \mathbf{H_2SO_4} \times \frac{1\ mol\ H_2SO_4}{98.078\ g\ H_2SO_4} = 0.400\ mol\ H_2SO_4$$

(b) Molar mass $O_2$ = 2 × (15.9994 g) = 31.9988 g/mol

$$8.00\ g\ \mathbf{O}_2 \times \frac{1\ mol\ O_2}{31.9988\ g\ O_2} = 0.250\ mol\ O_2$$

(c) Molar mass $NH_3$ = 14.0067 g + 3 × (1.0079 g) = 17.0304 g/mol

$$10.7\ g\ \mathbf{N}H_3 \times \frac{1\ mol\ \mathbf{N}H_3}{17.0304\ g\ \mathbf{N}H_3} = 0.628\ mol\ \mathbf{N}H_3$$

| | |
|---|---|
| **3.64** | ***See Section 3.2 and Examples 3.7, 3.8.*** |

(a) ? g $CO_2$ = 78.4 mol $CO_2$ × $\frac{44.01\ \text{g CO}_2}{1\ \text{mol CO}_2}$ = $\mathbf{3.45 \times 10^3}$ **g** $\mathbf{CO_2}$

(b) ? mol $AgNO_3$ = 192 g $AgNO_3$ × $\frac{1\ \text{mol AgNO}_3}{169.9\ \text{g AgNO}_3}$ = **1.13 mol** $\mathbf{AgNO_3}$

(c) ? molecules $CH_4$ = 9.22 g $CH_4$ × $\frac{1\ \text{mol CH}_4}{16.05\ \text{g CH}_4} \times \frac{6.022 \times 10^{23}\ \text{molecules CH}_4}{1\ \text{mol CH}_4}$ = $\mathbf{3.46 \times 10^{23}}$ **molecules** $\mathbf{CH_4}$

| | |
|---|---|
| **3.66** | ***See Section 3.2 and Example 3.7.*** |

(a) ? g $N_2O_4$ = 7.55 mol $N_2O_4$ × $\frac{92.02\ \text{g N}_2\text{O}_4}{1\ \text{mol N}_2\text{O}_4}$ = **695 g** $\mathbf{N_2O_4}$

(b) ? g $CaCl_2$ = 9.2 mol $CaCl_2$ × $\frac{111.0\ \text{g CaCl}_2}{1\ \text{mol CaCl}_2}$ = $\mathbf{1.0 \times 10^3}$ **g** $\mathbf{CaCl_2}$

(c) $? \text{ CO} = 0.44 \text{ mol CO} \times \frac{28.0 \text{ g CO}}{1 \text{ mol CO}} = \mathbf{12 \text{ g CO}}$

| | |
|---|---|
| **3.68** | ***See Section 3.2 and Example 3.8.*** |

(a) $? \text{ molecules } H_2 = 3.4 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.02 \text{ g } H_2} \times \frac{6.022 \times 10^{23} \text{ molecules } H_2}{1 \text{ mol } H_2} = \mathbf{1.0 \times 10^{24} \text{ molecules } H_2}$

(b) $? \text{ atoms H} = 1.0 \times 10^{24} \text{ molecules } H_2 \times \frac{2 \text{ atoms H}}{1 \text{ molecule } H_2} = \mathbf{2.0 \times 10^{24} \text{ atoms H}}$

| | |
|---|---|
| **3.70** | ***See Section 3.2 and Examples 3.6, 3.7, 3.8.*** |

(a) $? \text{ mol } SO_3 = 3.31 \text{ g } SO_3 \times \frac{1 \text{ mol } SO_3}{80.07 \text{ g } SO_3} = \mathbf{0.413 \text{ mol } SO_3}$

(b) $? \text{ molecules } SO_3 = 0.413 \text{ mol } SO_3 \times \frac{6.022 \times 10^{23} \text{ molecules } SO_3}{1 \text{ mol } SO_3} = \mathbf{2.49 \times 10^{23} \text{ molecules } SO_3}$

(c) $? \text{ atoms S} = 2.49 \times 10^{23} \text{ molecules } SO_3 \times \frac{1 \text{ atoms S}}{1 \text{ molecule } SO_3} = \mathbf{2.49 \times 10^{23} \text{ atoms S}}$

$? \text{ atoms O} = 2.49 \times 10^{23} \text{ molecules } OS_3 \ \frac{3 \text{ atoms O}}{1 \text{ molecule } SO_3} = \mathbf{7.47 \times 10^{23} \text{ atoms O}}$

| | |
|---|---|
| **3.72** | ***See Sections. 1.4 and 3.2.*** |

$? \text{ moles } C_2H_5OH = 0.9 \text{ fl. oz.} \times \frac{29.56 \text{ mL}}{1 \text{ fl. oz.}} \times \frac{0.7894 \text{ g } C_2H_5OH}{1 \text{ mL } C_2H_5OH} \times \frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g } C_2H_5OH} = \mathbf{0.4559 \text{ mol}}$

$\mathbf{C_2H_5OH}$

| | |
|---|---|
| **3.74** | ***See Section 3.2 and Examples 3.6, 3.7, 3.8.*** |

(a) Molecular mass for $Ni(CO)_4$:1[Ni] × 58.69 = 58.69

| | |
|---|---|
| 1[Ni] × 58.69 | = 58.69 |
| 4[C] × 12.01 | = 48.04 |
| 4[O] × 16.00 | = 64.00 |
| | 170.73 |

Hence, molar mass for $Ni(CO)_4$ is **170.73 g/mol.**

(b) $? \text{ mol } Ni(CO)_4 = 3.22 \text{ g } Ni(CO)_4 \times \frac{1 \text{ mol } Ni(CO)_4}{170.73 \text{ g } Ni(CO)_4} = \mathbf{1.89 \times 10^{-2} \text{ mol } Ni(CO)_4}$

(c) $? \text{ molecules } Ni(CO)_4 = 5.67 \text{ g } Ni(CO)_4 \ \frac{1 \text{ mol } Ni(CO)_4}{170.73 \text{ g } Ni(CO)_4} \times \frac{6.022 \times 10^{23} \text{ molecules } Ni(CO)_4}{1 \text{ mol } Ni(CO)_4}$

$= \mathbf{2.00 \times 10^{22} \text{ molecules } Ni(CO)_4}$

(d) ? atoms C = 34 g $Ni(CO)_4$ × $\frac{1 \text{ mol } Ni(CO)_4}{170.73 \text{ g } Ni(CO)_4}$ × $\frac{6.022 \times 10^{23} \text{ molecules } Ni(CO)_4}{1 \text{ mol } Ni(CO)_4}$ ×

$\frac{4 \text{ atom C}}{1 \text{ molecule } Ni(CO)_4}$

= **4.8 × $10^{23}$ atoms C**

Note: $Ni(CO)_4$ is composed of a metal and nonmetals and would be predicted to be an ionic compound (2.6). However, the observation that it is a volatile compound rather than a crystalline solid suggests it is a molecular compound (2.7).

**3.76** ***See Section. 3.2.***

Formula for hydrogen peroxide: $H_2O_2$

? moles H atoms = 0.011 mol $H_2O_2$ × $\frac{2 \text{ mol H}}{1 \text{ mol } H_2O_2}$ = **0.022 mol H atoms**

**3.78** ***See Section 3.3 and Example 3.9.***

(a) Molecular mass for $C_6H_{12}$:

6[C] × 12.01 = 72.06 %C = $\frac{72.06 \text{ g C}}{84.18 \text{ g } C_6H_{12}}$ × 100% = **85.60% C**

12[H] × 1.01 = 12.12 %H = $\frac{12.12 \text{ g H}}{84.18 \text{ g } C_6H_{12}}$ × 100% = **14.40% H**

84.18

Molar mass for $C_6H_{12}$ is 84.18 g/mol.

(b) Molecular mass for $C_5H_{12}$:

5[C] × 12.01 = 60.05 %C = $\frac{60.05 \text{ g C}}{88.17 \text{ g } C_5H_{12}O}$ × 100% = **68.11% C**

12[H] × 1.01 = 12.12 %H = $\frac{12.12 \text{ g H}}{88.17 \text{ g } C_5H_{12}O}$ × 100% = **13.75% H**

1[O] × 1600 = 16.00 %O = $\frac{16.00 \text{ g O}}{88.17 \text{ g } C_5H_{12}O}$ = 100% = **18.15% O**

88.17

Molar mass for $C_5H_{12}O$ is 88.17 g/mol.

(c) Formula mass for $NiCl_2$:

1[Ni] × 58.69 = 58.69 %Ni = $\frac{58.69 \text{ g Ni}}{129.59 \text{ g } NiCl_2}$ × 100% = **45.29% Ni**

2[Cl] × 35.45 = 70.90 %Cl = $\frac{70.90 \text{ g Cl}}{129.59 \text{ g } NiCl_2}$ × 100% = **54.71% Cl**

129.59

Molar mass for $NiCl_2$ is 129.59 g/mol.

**3.80 *See Section. 3.3.***

Mass of CuS to provide 10.0g of Cu:

To calculate the weight percent of Cu in CuS, we need the respective atomic weights:
Cu = 63.546 S = 32.066 adding CuS = 95.612

The % of Cu in CuS is then: $\dfrac{63.546\text{ g Cu}}{95.612\text{ g CuS}} \times 100 = 66.46\%$ Cu

**3.82 *See Sections 2.8, 3.3, and Example 3.9.***

Formula mass for $MgCO_3$:1[Mg] × 24.30 %Mg = $\dfrac{24.30\text{ g Mg}}{84.31\text{ g MgCO}_3} \times 100\% =$ **28.83% Mg**

1[C] × 12.01 = 12.01 %C = $\dfrac{12.01\text{ g C}}{84.31\text{ g MgCO}_3} \times 100\% =$ **14.25% C**

3[O] × 16.00 $\underline{= 48.00}$ %O = $\dfrac{48.00\text{ g O}}{84.31\text{ g MgCO}_3} \times 100\% =$ **56.93% O**

84.31

Molar mass for $MgCO_3$ is 84.31 g/mol.

**3.84 *See Section 2.8, 3.3, and Example 3.9.***

(a) (b)

Formula mass for $CaCO_3$:

1[Ca] × 40.08 = 40.08 %Ca = $\dfrac{40.08\text{ g Ca}}{100.09\text{ g CaCO}_3} \times 100\% =$ **40.04% Ca**

1[C] × 12.01 = 12.01 %Ca = $\dfrac{12.01\text{ g C}}{100.09\text{ g CaCO}_3} \times 100\% =$ **12.00% C**

3[O] × 16.00 $\underline{= 48.00}$ %O = $\dfrac{48.00\text{ g O}}{100.09\text{ g CaCO}_3} \times 100\% =$ **47.96% O**

100.09

Molar mass for $CaCO_3$ is 100.09 g/mol.

**3.86 *See Section 3.3 and Example 3.9.***

Formula mass for $C_6H_4(OH)Cl$: 6[C] × 12.01 = 72.06 %C = $\dfrac{72.06\text{ g C}}{128.56\text{ g C}_6\text{H}_4\text{ (OH) Cl}} \times 100\%$ 56.05%C

5[H] × 1.01 = 5.05 %H = $\dfrac{5.05\text{ g H}}{128.56\text{ g C}_6\text{H}_4\text{ (OH) Cl}} \times 100\% = 3.93\%$H

1[O] × 16.00 %O = $\dfrac{16.00\text{ g O}}{128.56\text{ g C}_6\text{H}_4\text{ (OH) Cl}} \times 100\% = 12.45\%$O

1[Cl] × 35.45 $\underline{= 35.45}$ %Cl = $\dfrac{35.45\text{ g Cl}}{128.56\text{ g C}_6\text{H}_4\text{ (OH) Cl}} \times 100\% = 27.57\%$Cl

128.56

Molar mass for $C_6H_4(OH)Cl$ is 128.56 g/mol.

Analytical percent composition of 56.05% C, 3.93% H, and 27.57% Cl agrees with percent composition of $C_6H_4(OH)Cl$. **Compound is $C_6H_4(OH)Cl$.**

**3.88 *See Section 3.3.***

(a) Molecular mass for $C_4H_{10}O$:

$4[C] \times 12.01 = 48.04$

$10[H] \times 1.01 = 10.10$

$1[O] \times 16.00 = \underline{16.00}$

74.14 Hence, molar mass for $C_4H_{10}O$ is 74.14 g/mol.

$$? \text{ g C} = 1.80 \text{ g } C_4H_{10}O \times \frac{48.04 \text{ g C}}{74.14 \text{ g } C_4H_{10}O} = \mathbf{1.17\ g\ C}$$

(b) Formula mass for $Na_2CO_3$:

$2[Na] \times 22.99 = 45.98$

$1[C] \times 12.01 = 12.01$

$3[O] \times 16.00 = \underline{48.00}$

105.99 Hence, molar mass for $Na_2CO_3$ is 105.99 g/mol

$$? \text{ g C} = 0.00223 \text{ g } Na_2CO_3 \times \frac{12.01 \text{ g C}}{105.99 \text{ g } Na_2CO_3} = \mathbf{2.53 \times 10^{-4}\ g\ C}$$

(c) Molecular mass for $C_5H_{11}N$:

$5[C] \times 12.01 = 60.05$

$11[H] \times 1.01 = 11.11$

$1[N] \times 14.01 = \underline{14.01}$

85.17 Hence, molar mass for $C_5H_{11}N$ is 85.17 g/mol.

$$? \text{ g C} = 22.1 \text{ g } C_5H_{11}N \times \frac{60.05 \text{ g C}}{85.17 \text{ g } C_5H_{11}N} = \mathbf{15.6\ g\ C}$$

Note: See direct method used in working 3.73.

**3.90 *See Section 3.3.***

(a) $? \text{ g H} = 4.33 \text{ g } H_2O \times \frac{2.02 \text{ g H}}{18.02 \text{ g } H_2O} = \mathbf{0.485\ g\ H}$

(b) $? \text{ g H} = 1.22 \text{ g } C_2H_2 \times \frac{2.02 \text{ g H}}{26.04 \text{ g } C_2H_2} = \mathbf{0.0946\ g\ H}$

(c) $? \text{ g H} = 4.44 \text{ g } N_2H_4 \times \frac{4.04 \text{ g H}}{32.06 \text{ g } N_2H_4} \times \mathbf{0.560\ g\ H}$

**3.92 *See Section 3.3 and Example 3.10.***

$$? \text{ g C} = 4.06 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 1.11 \text{ g C} \quad \%C = \frac{1.11 \text{ g C}}{2.770 \text{ g sample}} \times 100\% = \mathbf{40.1\%\ C}$$

$$? \text{ g H} = 1.66 \text{ g } H_2O \times \frac{2.02 \text{ g H}}{18.02 \text{ g } H_2O} = 0.186 \text{ g H} \quad \%H = \frac{0.186 \text{ g H}}{2.770 \text{ g sample}} \times 100\% = \mathbf{6.7\%\ H}$$

$$? \text{ g O} = \text{g sample - g C - g H} \qquad \%O = \frac{1.47 \text{ g O}}{2.770 \text{ g sample}} \times 100\% = \mathbf{53.1\%\ O}$$

$$= 2.770 \text{ g} - 1.11 \text{ g} - 0.186 \text{ g} = 1.47 \text{ g O}$$

**3.94** ***See Section 3.3.***

$$? \text{ g C} = 1.04 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 0.284 \text{ g C} \qquad \%C = \frac{0.284 \text{ g C}}{0.513 \text{ g sample}} \times 100\% = \mathbf{55.4\%\ C}$$

$$? \text{ g H} = 0.704 \text{ g } H_2O \times \frac{2.02 \text{ g H}}{18.02 \text{ g } H_2O} = 0.0789 \text{ g H} \quad \%H = \frac{0.0789 \text{ g H}}{0.513 \text{ g sample}} \times 100\% = \mathbf{15.4\%\ H}$$

$$? \text{ g N} = \text{g sample} - \text{g C} - \text{gH} \qquad \%N = \frac{0.150 \text{ g N}}{0.513 \text{ g sample}} \times 100\% = \mathbf{29.2\%\ N}$$

$$= 0.513 \text{ g} - 0.284 \text{ g} - 0.0789 \text{ g} = 0.150 \text{ g N}$$

**3.96** ***See Section 3.3 and Example 3.11.***

$$? \text{ mol C} = 0.80 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 0.067 \text{ mol C} \quad \text{relative mol C} = \frac{0.067 \text{ mol C}}{0.067} = 1.00 \text{ mol C}$$

$$? \text{ mol H} = 0.20 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.20 \text{ mol H} \quad \text{relative mol H} = \frac{0.20 \text{ mol H}}{0.067} = 3.0 \text{ mol H}$$

The empirical formula is $\mathbf{CH_3}$**.**

**3.98** ***See Section 3.3 and Example 3.11.***

$$? \text{ mol N} = 0.152 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 0.0109 \text{ mol N} \quad \text{relative mol N} = \frac{0.0109 \text{ mol N}}{0.0109} = 1.00 \text{ mol N}$$

$$? \text{ mol O} = 0.348 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.0218 \text{ mol O} \qquad \text{relative mol O} = \frac{0.0218 \text{ mol O}}{0.0109} \ 2.00 \text{ mol O}$$

The empirical formula is $\mathbf{NO_2}$**.**

**3.100** ***See Section 3.3 and Example 3.11.***

Assume the sample has a mass of 100.00 g and therefore contains 52.7 g Se and 47.3 g Cl.

$$? \text{ mol Se} = 52.7 \text{ g Se} \times \frac{1 \text{ mol Se}}{78.96 \text{ g Se}} = 0.667 \text{ mol Se} \qquad \text{relative mol Se} = \frac{0.667 \text{ mol Se}}{0.667} = 1.00 \text{ mol Se}$$

$$? \text{ mol Cl} = 47.3 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.453 \text{ g Cl}} = 1.33 \text{ mol Cl} \qquad \text{relative mol Cl} = \frac{1.33 \text{ mol Cl}}{0.667} = 1.99 \text{ mol Cl}$$

The empirical formula is $\mathbf{SeCl_2}$.

**3.102** ***See Section 3.3 and Example 3.11.***

$? \text{ mol Cr} = 0.173 \text{ g Cr} \times \dfrac{1 \text{ mol Cr}}{51.996 \text{ g Cr}} = 0.00333 \text{ mol Cr}$ relative mol Cr = $\dfrac{0.00333 \text{ mol Cr}}{0.00333} = 1.00$ mol Cr

$? \text{ mol O} = 0.160 \text{ g O} \times \dfrac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.0100 \text{ mol O}$ relative mol O = $\dfrac{0.0100 \text{ mol O}}{0.00333} = 3.00$ mol O

The empirical formula is $\mathbf{CrO_3}$.

**3.104** ***See Section. 3.3.***

$$? \text{ g C} = 0.104 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.06635 \text{ mol C}$$

$$? \text{ mol H} = 1.22 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.1254 \text{ mol H}$$

$$? \text{ g C} = 0.06635 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.7969 \text{ g C}$$

$$? \text{ g H} = 0.1354 \text{ mol H} \times \frac{1.01 \text{ g H}}{1 \text{ mol H}} = 0.1365 \text{ g H}$$

$$? \text{ g O} = 1.20 \text{ g total} - \left[0.7969 \text{ g C} + 0.1365 \text{ g H}\right] = 0.2666 \text{ g O}$$

$$? \text{ mol O} = 0.2666 \text{ g O} \times \frac{1 \text{ mol O}}{15.99 \text{ g O}} = 0.01666 \text{ mol O}$$

| Element | Moles of Element | Divide by Smallest |
|---|---|---|
| C | 0.06635 | $\dfrac{0.06635}{0.01666} = 4.00$ |
| H | 0.1354 | $\dfrac{0.1354}{0.01666} = 8.00$ |
| O | 0.01666 | $\dfrac{0.01666}{0.01666} = 1.00$ |

The simplest formula is $\mathbf{C_4H_8O}$

**3.106** ***See Section 3.3 and Example 3.11.***

Assume the sample has a mass of 100.00 g and therefore contains 79.95 g C, 9.40 g H and 10.65 g O.

$? \text{ mol C} = 79.95 \text{ g C} \times \dfrac{1 \text{ mol C}}{12.011 \text{ g C}} = 6.656 \text{ mol C}$ relative mol C = $\dfrac{6.656 \text{ mol C}}{0.6656} = 10.00$ mol C

$? \text{ mol H} = 9.40 \text{ g H} \times \dfrac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.33 \text{ mol H}$ relative mol H = $\dfrac{9.33 \text{ mol H}}{0.6656} = 14.0$ mol H

$? \text{ mol O} = 10.65 \text{ g O} \times \dfrac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.6656 \text{ mol O}$ relative mol O = $\dfrac{0.6656 \text{ mol O}}{0.6656} = 1.000$ mol O

The empirical formula is $\mathbf{C_{10}H_{14}O}$.

**3.108 *See Section 3.3.***

$$? \text{ mol C} = 0.170 \text{ g} \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.003855 \text{ mol C}$$

$$? \text{ mol H} = 0.0348 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.003863 \text{ mol H}$$

$$? \text{ g C} = 0.003855 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.0464 \text{ g C}$$

$$? \text{ g H} = 0.003863 \text{ mol H} \times \frac{1.01 \text{ g H}}{1 \text{ mol H}} = 0.003894 \text{ g H}$$

$$\% \text{ C} = \frac{0.0464 \text{ g C}}{0.459 \text{ g sample}} \times 100 = \mathbf{10.1\% \ C}$$

$$\% \text{ H} = \frac{0.003894 \text{ g H}}{0.459 \text{ g sample}} \times 100 = \mathbf{0.84\% \ H}$$

**3.110 *See Section 3.3 and Example 3.11.***

$$? \text{ g C} = 3.94 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 1.08 \text{ g C} \quad ?\text{g H} = 1.89 \text{ g H}_2\text{O} \times \frac{2.02 \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 0.212 \text{ g H}$$

$$? \text{ g N in 2.18 g sample} = 2.18 \text{ g sample} \times \frac{0.235 \text{ g N}}{1.23 \text{ g sample}} = 0.417 \text{ g N}$$

? g = g sample – g C – g H – g N = 2.18 g – 1.08 g – 0.212 g – 0.417 g = 0.47 g O

$$? \text{ mol C} = 1.08 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 0.0899 \text{ mol C} \qquad \text{relative mol C} = \frac{0.0899 \text{ mol C}}{0.029} = 3.1 \text{ mol C}$$

$$? \text{ mol H} = 0.212 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.210 \text{ mol H} \qquad \text{relative mol H} = \frac{0.210 \text{ mol H}}{0.029} = 7.2 \text{ mol H}$$

$$? \text{ mol N} = 0.417 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.0298 \text{ mol N} \qquad \text{relative mol N} = \frac{0.0298 \text{ mol N}}{0.029} = 1.0 \text{ mol N}$$

$$? \text{ mol O} = 0.47 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 0.029 \text{ mol O} \qquad \text{relative mol O} = \frac{0.029 \text{ mol O}}{0.029} = 1.0 \text{ mol O}$$

The empirical formula is **$C_3H_7NO$.**

Note: An alternative method of solving this problem involves calculating % C from g C in 2.18 g sample, %H from g H in 2.18 g sample, % N from g N in 1.23 g sample, % O from 100.00% – % C – % H – % N and working with a 100.00 sample.

**3.112 *See Section 3.3 and Example 3.12.***

(a) Formula mass for HO:

$$\begin{array}{ll} 1[\text{H}] \times 1.0 & = 1.0 \\ 1[\text{O}] \times 16.0 & \underline{= 16.0} \\ 17.0 & \end{array}$$

Hence, molar mass for HO is 17.0 g/mol.

$$n = \frac{\text{molar mass compound}}{\text{molar mass HO}} = \frac{34 \text{ g / mol}}{17.0 \text{ g/ mol}} = 2$$

The molecular formula is **$H_2O_2$.**

**3.114** ***See Section 3.3 and Example 3.12.***

(a) Formula mass for $C_5H_{10}O$:

| | | |
|---|---|---|
| $5[C] \times 12.0$ | $= 60.0$ | |
| $10[H] \times 1.0$ | $= 10.0$ | |
| $1[O] \times 16.0$ | $\underline{= 16.0}$ | |
| 86.0 | | Hence, molar mass for $C_5H_{10}O$ is 86.0 g/mol. |

$$n = \frac{\text{molar mass compound}}{\text{molar mass } C_5H_{10}O} = \frac{258 \text{ g / mol}}{86.0 \text{ g / mol}} = 3$$

The molecular formula is $\mathbf{C_{15}H_{30}O_3}$**.**

(b) Formula mass for $PCl_3$:

| | | |
|---|---|---|
| $1[P] \times 31.0$ | $= 31.0$ | |
| $3[Cl] \times 35.5$ | $\underline{= 106.5}$ | |
| 137.5 | | Hence, molar mass for $PCl_3$ is 137.5 g/mol. |

$$n \frac{\text{molar mass compound}}{\text{molar mass } PCl_3} = \frac{137.3 \text{ g / mol}}{137.5 \text{ g / mol}} = 1$$

The molecular formula is $\mathbf{PCl_3}$

**3.116** ***See Section. 3.3.***

Empirical and Molecular formula for Mandelic Acid:

$$63.15 \text{ g C} \bullet \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 5.258 \text{ mol C}$$

$$5.30 \text{ g H} \bullet \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 5.28 \text{ mol H}$$

$$31.55 \text{ g O} \bullet \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 1.972 \text{ mol O}$$

Using the smallest number of atoms, we calculate the ratio of atoms:

$$\frac{5.258 \text{ mol C}}{1.972 \text{ mol O}} = \frac{2.666 \text{ mol C}}{1 \text{ mol O}} \text{ or } \frac{22/3 \text{ mol C}}{1 \text{ mol O}} \text{ or } \frac{8/3 \text{ mol C}}{1 \text{ mol O}}$$

So 3 mol O combine with 8 mol C and 8 mol H so the **empirical formula is** $\mathbf{C_8H_8O_3}$. The formula mass of $C_8H_8O_3$ is 152.15. Given the data that the molar mass is 152.15 g/mL, the **molecular formula for mandelic acid is** $\mathbf{C_8H_8O_3}$**.**

**3.118** ***See Section 3.3 and Example 3.12.***

Assume the sample has a mass of 100.0 g and therefore contains 40.0 g C, 6.71 g H and 53.3 g O.

$$? \text{ mol C} = 40.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.33 \text{ mol C} \quad \text{relative mol C} = \frac{3.33 \text{ mol C}}{3.33} = 1.00 \text{ mol C}$$

$$? \text{ mol H} = 6.71 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.66 \text{ mol H} \quad \text{relative mol H} = \frac{6.66 \text{ mol H}}{3.33} = 2.00 \text{ mol H}$$

$$? \text{ mol O} = 53.3 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} \times 3.33 \text{ mol O} \quad \text{relative mol O} = \frac{3.33 \text{ mol O}}{3.33} = 1.00 \text{ mol O}$$

The empirical formula is $CH_2O$.

Formula mass for $CH_2O$:

| | | |
|---|---|---|
| | $1[C] \times 12.0$ | $= 12.0$ |
| | $2[H] \times 1.0$ | $= 2.0$ |
| | $1[O] \times 16.0$ | $\underline{= 16.0}$ |
| | $= 30.0$ | Hence, molar mass for $CH_2O$ is 30.0 g/mol. |

$$n = \frac{\text{molar mass compound}}{\text{molar mass } CH_2O} = \frac{180 \text{ g / mol}}{30.0 \text{ g / mol}} = 6$$

The molecular formula for fructose is $\mathbf{C_6H_{12}O_6}$.

**3.120** ***See Sections 3.2, 3.4, and Examples 3.4, 3.13, 3.14.***

(a) Unbalanced: $C_4H_8O + O_2 \rightarrow CO_2 + H_2O$ Start with $C_4H_8O$.

Step 1: $C_4H_8O + O_2 \rightarrow \underline{4}CO_2 + \underline{4}H_2O$ Balances C, H.

Step 2: $C_4H_8O + \frac{11}{\underline{2}} O_2 \rightarrow 4CO_2 + 4H_4O$ Balances O.

Step 3: $\underline{2}C_4H_8O + \underline{11}O_2 \rightarrow \underline{8}CO_2 + \underline{8}H_2O$ Gives whole number coefficients.

(b) *Strategy: g $C_4H_8O \rightarrow$ mol $C_4H_8O \rightarrow$ mol $O_2 \rightarrow$ g $O_2$*

$$? \text{ g } O_2 = 5.33 \text{ g } C_4H_8O \times \frac{1 \text{ mol } C_4H_8O}{72.0 \text{ g } C_4H_8O} \times \frac{11 \text{ mol } O_2}{2 \text{ mol } C_4H_8O} \times \frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2} = \mathbf{13.0 \text{ g } O_2}$$

**3.122** ***See Sections 3.1, 3.4, and Examples 3.1, 3.2, 3.13.***

Unbalanced: $N_2 + H_2 \rightarrow NH_3$ Start with $NH_3$.

Step 1: $N_2 + H_2 \rightarrow \underline{2}NH_3$ Balances N.

Step 2: $N_2 + \underline{3}H_2 \rightarrow 2NH_3$ Balances H.

*Strategy: g $N_2 \rightarrow$ mol $N_2 \rightarrow$ mol $NH_3 \rightarrow$ g $NH_3$*

$$? \text{ g } NH_3 = 5.33 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.0 \text{ g } N_2} \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} \times \frac{17.0 \text{ g } NH_3}{1 \text{ mol } NH_3} = \mathbf{6.47 \text{ g } NH_3}$$

**3.124** ***See Section. 3.4.***

Use the stoichiometry of the balanced equation as a conversion factor to convert the moles of product to moles of reactant. The balanced equation says: 4 mol HCl are needed to make 1 mol $Cl_2$.

$$12.5 \; mol \; Cl_2 \times \frac{4 \; mol \; HCl}{1 \; mol \; Cl_2} = \mathbf{50.0 \text{ mol HCl}}$$

**3.126** ***See Sections 3.1, 3.4, and Examples 3.1, 3.2, 3.13, 3.14.***

Balanced: $HCl + NaOH \rightarrow NaCl + H_2O$

*Strategy: g $H_2O$ → mol $H_2O$ → mol NaCl → g NaCl*

$$? \text{ g NaCl} = 78.2 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O} \times \frac{1 \text{ mol NaCI}}{1 \text{ mol } H_2O} \times \frac{58.5 \text{ g NaCI}}{1 \text{ mol NaCI}} = \mathbf{254 \text{ g NaCl}}$$

**3.128** ***See Section. 3.5.***

(a) $2Sb + 3Cl_2 \rightarrow 2SbCl_3$

(b) According to the figure, there appears to be one Sb atom left over after the reaction is complete; therefore $Cl_2$ is the limiting reactant.

**3.130** ***See Section 3.5 and Example 3.16.***

Balanced: $Zn(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Zn(NO_3)_2(aq)$

*Strategy: g Zn → mol Zn → mol Ag → g Ag*

$$? \text{ g Ag based on Zn} = 3.22 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.4 \text{ g Zn}} \times \frac{2 \text{ mol Zn}}{1 \text{ mol Zn}} \times \frac{107.9 \text{ g Zg}}{1 \text{ mol Ag}} = 10.6 \text{ g Ag}$$

*Strategy: g $AgNO_3$ → mol $AgNO_3$ → mol Ag → g Ag*

$$? \text{ g Ag based on } AgNO_3 = 4.35 \text{ g } AgNO_3 \times \frac{1 \text{ mol } AgNO_3}{169.9 \text{ g } AgNO_3} \times \frac{2 \text{ mol Ag}}{2 \text{ mol } AgNO_3} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 2.76 \text{ g Ag}$$

$AgNO_3$ is the limiting reactant because it produces less Ag. The maximum amount of Ag which can be produced from 3.22 g Zn and 4.35 g $AgNO_3$ is **2.76 g.**

**3.132** ***See Sections 3.1, 3.5, and Examples 3.1, 3.2, 3.16.***

Unbalanced: $P_4 + O_2 \rightarrow P_4O_{10}$ Start with $P_4O_{10}$.

Step 1: $P_4 + \underline{5}O_2 \rightarrow P_4O_{10}$ Balances O.

*Strategy: g $P_4$ → mol $P_4$ → mol $P_4O_{10}$ → g $P_4O_{10}$*

$$? \text{ g } P_4O_{10} \text{ based on } P_4 = 2.2 \text{ g } P_4 \times \frac{1 \text{ mol } P_4}{124.0 \text{ g } P_4} \times \frac{1 \text{ mol } P_4O_{10}}{1 \text{ mol } P_4} \times \frac{284.0 \text{ g } P_4O_{10}}{1 \text{ mol } P_4O_{10}} = 5.0 \text{ g } P_4O_{10}$$

*Strategy: g $O_2$ → mol $O_2$ → mol $P_4O_{10}$ → g $P_4O_{10}$*

$$? \text{ g } P_4O_{10} \text{ based on } O_2 = 4.2 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.0 \text{ g } O_2} \times \frac{1 \text{ mol } P_4O_{10}}{5 \text{ mol } O_2} \times \frac{284.0 \text{ g } P_4O_{10}}{1 \text{ mol } P_4O_{10}} = 7.5 \text{ g } P_4O_{10}$$

$P_4$ is the limiting reactant because it produces less $P_4O_{10}$. The maximum amount of $P_4O_{10}$ that can be produced from 2.2 g $P_4$ and 4.2 g $O_2$ is 5.0 g. Hence, the theoretical yield is **5.0 g $P_4O_{10}$.**

**3.134** ***See Sections 3.1, 3.4, and Examples 3.1, 3.2, 3.13, 3.14.***

Unbalanced: $CS_2 + Cl_2 \rightarrow CCl_4 + S_2Cl_2$ Start with $Cl_2$.

Step 1: $CS_2 + \underline{3}Cl_2 \rightarrow CCl_4 + S_2Cl_2$ Balances $Cl_2$.

*Strategy: g $CS_2$ → mol $CS_2$ → mol $CCl_4$ → g $CCl_4$*

$$? \text{ g } CCl_4 = 43.1 \text{ g } CS_2 \times \frac{1 \text{ mol } CS_2}{76.2 \text{ g } CS_2} \times \frac{1 \text{ mol } CCI_4}{1 \text{ mol } CS_2} \times \frac{153.8 \text{ g } CCI_4}{1 \text{ mol } CCI_4} = 87.0 \text{ g } CCl_4$$

$$\text{percent yield } CCl_4 = \frac{\text{actual yield } CCl_4}{\text{theoretical yield } CCl_4} \times 100\% = \frac{45.2 \text{ g } CCl_4}{87.0 \text{ g } CCl_4} \times 100\% = 52.0\%$$

*Strategy: g $CS_2$ → mol $CS_2$ → mol $S_2Cl_2$ → g $S_2Cl_2$*

$$? \text{ g } S_2Cl_2 = 43.1 \text{ g } CS_2 \times \frac{1 \text{ mol } CS_2}{76.2 \text{ g } CS_2} \times \frac{1 \text{ mol } S_2Cl_2}{1 \text{ mol } CS_2} \times \frac{135.1 \text{ g } S_2Cl_2}{1 \text{ mol } S_2Cl_2} = 76.4 \text{ g } S_2Cl_2$$

$$\text{percent yield } S_2Cl_2 = \frac{\text{actual yield } S_2Cl_2}{\text{theoretical yield } S_2Cl_2} \times 100\% = \frac{41.3 \text{ g } S_2Cl_2}{76.4 \text{ g } S_2Cl_2} \times 100\% = \mathbf{54.1\%}$$

**3.136 *See Sections 3.1, 3.5, and Examples 3.1, 3.2, 3.17.***

Unbalanced: $P_4 + Cl_2 \rightarrow PCl_5$Start with $PCl_5$

Step 1: $P_4 + Cl_2 \rightarrow \underline{4}PCl_5$ Balances P.

Step 2: $P_4 + \underline{10}Cl_2 \rightarrow 4PCl_5$ Balances Cl.

*Strategy: g $P_4$ → mol $P_4$ → mol $PCl_5$ → g $PCl_5$*

$$? \text{ g } PCl_5 \text{ based on } P_4 = 2.3 \text{ g } P_4 \times \frac{1 \text{ mol } P_4}{124.0 \text{ g } P_4} \times \frac{4 \text{ mol } PCl_5}{1 \text{ mol } PCl_5} \times \frac{208.2 \text{ g } PCl_5}{1 \text{ mol } PCl_5} = 15 \text{ g } PCl_5$$

*Strategy: g $Cl_2$ → mol $Cl_2$ → mol $PCl_5$ → g $PCl_5$*

$$? \text{ g } PCl_5 \text{ based on } Cl_2 = 7.0 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.9 \text{ g } Cl_2} \times \frac{4 \text{ mol } PCl_5}{10 \text{ mol } Cl_2} \times \frac{208.2 \text{ g } PCl_5}{1 \text{ mol } PCl_5} = 8.2 \text{ g } PCl_5$$

$Cl_2$ is the limiting reactant because it produces less $PCl_5$. The maximum amount of $PCl_5$ that can be produced from 2.3 g $P_4$ and 7.0 g $Cl_2$ is 8.2 g. Hence, the theoretical yield is **8.2 g $PCl_5$.**

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{7.1 \text{ g } PCl_5}{8.2 \text{ g } PCl_5} \times 100\% = \mathbf{87\%.}$$

**3.138 *See Section. 3.5 and Example 3.20.***

The mole ratio comes from the balanced equation.

$$5.0 \times 10^3 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.0158 \text{ g } H_2} \times \frac{1 \text{ mol } CH_3OH}{2 \text{ mol } H_2} \times \frac{32.0417 \text{ g } CH_3OH}{1 \text{ mol } CH_3OH} = 4.0 \times 10^4 \text{ g } CH_3OH$$

The given mass of $CH_3OH$ is the actual yield. Use these two masses to calculate the percent yield.

$$\frac{3.5 \times 10^3 \ g\ CH_3OH\ actual}{4.0 \times 10^4 \ g\ CH_3OH\ theoretical} \times 100\% = \mathbf{8.8\%\ yield}$$

**3.140 *See Section. 3.5 and Example 3.20.***

Balanced equation: $2KClO_3$ $2KCl + 3O_2$

(a) Calculate the theoretical yield of $O_2$.

Plan: g $KClO_3$ $\overset{(1)}{Þ}$ mol $KClO_3$ $\overset{(2)}{Þ}$ mol $O_2$ $\overset{(3)}{Þ}$ g $O_2$ (theoretical)

(1) $? \text{ mol KClO}_3 = \dfrac{\text{g KClO}_3}{\text{FW KClO}_3} = \dfrac{3.75 \text{ g}}{122.55 \text{ g/mol}} = 0.0306 \text{ mol KClO}_3$

(2) $? \text{ mol O}_2 = 0.0306 \text{ mol KClO}_3 \text{ x } \dfrac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} = 0.0459 \text{ mol O}_2$

(3) ? g $O_2$ = mol $O_2$ x AW $O_2$ = 0.0459 mol x 32.0 g/mol = **1.47 g $O_2$** (theoretical yield)
Alternatively by dimensional analysis:

$? \text{ g O}_2 = 3.75 \text{ g KClO}_3 \text{ x } \dfrac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \text{ x } \dfrac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \text{ x } \dfrac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} = \mathbf{1.47 \text{ g O}_2}$

(b) Calculate the percent yield of $O_2$.

$\% \text{ yield} = \dfrac{\text{actual yield}}{\text{theoretical yield}} \text{ x } 100\% = \dfrac{1.05 \text{ g}}{1.47 \text{ g}} \text{ x } 100\% = \mathbf{71.4\%}$

**3.142 *See Section. 3.5 and Problem 3.141.***

(a) Unbalanced: $NO + O_2 \rightarrow NO_2$ Start with NO
Step 1: $\underline{2}NO + O_2 \rightarrow \underline{2}NO_2$ Balances N and O.
*Strategy: g NO → mol NO → mol $NO_2$*

$? \text{ g NO}_2 \text{ based on NO} = 75.0 \text{ g NO} \times \dfrac{1 \text{ mol NO}}{30.01 \text{ g NO}} \times \dfrac{2 \text{ mol NO}_2}{2 \text{ mol NO}} = 2.50 \text{ mol NO}_2$

*Strategy: g $O_2$ → mol $O_2$ → mol $NO_2$*

$? \text{ g NO}_2 \text{ based on O}_2 = 45.0 \text{ g O}_2 \times \dfrac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \dfrac{2 \text{ mol NO}_2}{1 \text{ mol O}_2} = 2.81 \text{ mol NO}_2$

**NO is the limiting reactant because it produces less $NO_2$.**

(b) $? \text{ mol O}_2 \text{ reacted} = 2.50 \text{ mol NO} \times \dfrac{1 \text{ mol O}_2}{2 \text{ mol NO}} = 1.25 \text{ mol O}_2$

$? \text{ mol O}_2 \text{ initial} = 45.0 \text{ g O}_2 \times \dfrac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 1.41 \text{ mol O}_2$

Number of moles of $O_2$ remaining = # mol initial - # mol reacted:

1.41 mol $O_2$ – 1.25 mol $O_2$ = 0.16 mol $O_2$

*Strategy: mol $O_2$ → g $O_2$*

$? \text{ g O}_2 = 0.16 \text{ mol O}_2 \times \dfrac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = \mathbf{5.12 \text{ g O}_2 \text{ remaining}}$

**3.144 *See Sections 3.1, 3.4, and Examples 3.1, 3.2, 3.13, 3.14.***

Balanced: $2 NaOH + X_2 \rightarrow NaX + NaXO + H_2O$
*Strategy: Let y equal molar mass of $X_2$ in g $X_2$ → mol $X_2$ → mol NaX → g NaX set up.*

$$? \text{ g NaX} = 3.11 \text{ g } X_2 \times \frac{1 \text{ mol } X_2}{y \text{ g } X_2} \times \frac{1 \text{ mol NaX}}{1 \text{ mol } X_2} \times \frac{(23.0 + 0.5\ y) \text{ g NaX}}{1 \text{ mole NaX}} = 2.00 \text{ g NaX}$$

Thus, (3.11)(23.0 + 0.5 y) = 2.00 y, 71.5 + 1.56 y = 2.00 y, and y = 162.
Since $Br_2$ has a molar mass of 159.8 g/mol, the halogen is **Br.**

**3.146** ***See Sections 2.6, 2.8, 3.3, and Example 3.9.***

Iron(III) sulfate is $Fe_2(SO_4)_3$.
Formula mass for $Fe_2(SO_4)_3$:

$$2[\text{Fe}] \times 55.8 = 111.6 \qquad \%\text{F} = \frac{111.6 \text{ g Fe}}{399.9 \text{ g } Fe_2(SO_4)_3} \times 100\% = \mathbf{27.9\%\ Fe}$$

$$3[\text{S}] \times 32.1 = 96.3 \qquad \%\text{S} = \frac{96.3 \text{ g S}}{399.9 \text{ g } Fe_2(SO_4)_3} \times 100\% = \mathbf{24.1\%\ S}$$

$$12[\text{O}] \times 16.0 = \frac{192.0}{399.9} \qquad \%\text{O} = \frac{192.0 \text{ g O}}{399.9 \text{ g } Fe_2(SO_4)_3} \times 100\% = \mathbf{48.0\%\ O}$$

Molar mass for $Fe_2(SO_4)_3$ is 399.9 g/mol.

**3.148** ***See Sections 3.1, 3.2, 3.3, and Examples 3.1, 3.2, 3.13.***

(a) Unbalanced: $In_2S_3 + O_2 \rightarrow In_2O_3 + SO_2$ Start with $In_2S_3$.
Step 1: $In_2S_3 + O_2 \rightarrow In_2O_3 + \underline{3}SO_2$ Balances S.

Step 2: $In_2S_3 + \frac{9}{2}O_2 \rightarrow In_2O_3 + 3SO_2$ Balances O.

Unbalanced: $In_2O_3 + CO \rightarrow In + CO_2$ Start with $In_2O_3$.
Step 1: $In_2O_3 + CO \rightarrow \underline{2}In + CO_2$ Balances In.
Step 2: $In_2O_3 + \underline{3}CO \rightarrow 2In + 3CO_2$ Balances O.

(b) *Strategy: kg $In_2S_3$ → g $In_2S_3$ → mol $In_2S_3$ → mol $In_2S_3$ → mol In → g In → kg In*

$$? \text{ kg In} = 35.7 \text{ kg } In_2S_3 \times \frac{10^3 \text{ g } In_2S_3}{1 \text{ kg } In_2S_3} \times \frac{1 \text{ mol } In_2S_3}{325.9 \text{ g } In_2S_3} \times \frac{2 \text{ mol } In_2O_3}{2 \text{ mol } In_2S_3} \times \frac{2 \text{ mol In}}{1 \text{ mol } In_2S_3}$$

$$\times \frac{114.8 \text{ g In}}{1 \text{ mol In}} \times \frac{1 \text{ kg In}}{10^3 \text{ g In}} = \mathbf{25.2\ kg\ In}$$

Alternatively, using the composition of $In_2S_3$ gives

$$? \text{ kg In} = 35.7 \text{ kg } In_2S_3 \times \frac{229.6 \text{ g kg In}}{325.9 \text{ kg } In_2S_3} = \mathbf{25.2\ kg\ In}$$

**3.150** ***See Section. 3.4.***

? mass of $H_2O$ = 3.650 g – 1.782 g = 1.868 g $H_2O$

*Strategy: g $H_2O$ → mol $H_2$ O*

$$1.868 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 0.104 \text{ mol } H_2O$$

*Strategy: g $MgSO_4$ → mol $MgSO_4$*

$$1.782 \text{ g } MgSO_4 \times \frac{1 \text{ mol MgSO}_4}{120.37 \text{ g MgSO}_4} = 0.0148 \text{ mol MgSO}_4$$

Note: By heating the sample all of the water was driven off; therefore the 1.782 g contains only $MgSO_4$.

$$\frac{0.104 \text{ mol H}_2\text{O}}{0.0148 \text{ mol MgSO}_4} \simeq 7 \text{ mol H}_2\text{O}$$

**There are 7 moles of $H_2O$ for every one mole of $MgSO_4$.**

| **3.152** ***See Sections 3.2 and 3.3.*** |
|---|

Assume the sample contains 100.0 g and therefore contains 25.5 g Cu, 12.8 g S, 57.7 g O and 4.0 g H.

*Strategy: g Cu → mol Cu → mol $CuSO_4$*

$$? \text{ mol CuSO}_4 = 25.5 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.5 \text{ g Cu}} \times \frac{1 \text{ mol CuSO}_4}{1 \text{ mol Cu}} = 0.402 \text{ mol CuSO}_4$$

*Strategy: g H → mol H → mol $H_2O$*

$$? \text{ mol H}_2\text{O} = 4.0 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} \times \frac{1 \text{ mol H}_2\text{O}}{2 \text{ mol H}} = 2.0 \text{ mol H}_2\text{O}$$

$$\frac{\text{mol H}_2\text{O}}{\text{mol CuSO}_4} = \frac{2.0 \text{ mol}}{0.402 \text{ mol}} = 5.0$$

The value of × in $CuSO_4 \cdot \times H_2O$ is **5.0**

| **3.154** ***See Section 3.3 and Examples 3.12.*** |
|---|

Assume the sample contains 100.00 g and therefore contains 71.56 g C, 6.71 g H, 4.91 g N and 16.82 g O.

$$? \text{ mol C} = 71.56 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 5.96 \text{ mol C} \quad \text{relative mol C} = \frac{5.96 \text{ mol C}}{0.350} = 17.0 \text{ mol C}$$

$$? \text{ mol H} = 6.71 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.66 \text{ mol H} \quad \text{relative mol H} = \frac{6.66 \text{ mol H}}{0.350} = 19.0 \text{ mol H}$$

$$? \text{ mol N} = 4.91 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 0.350 \text{ mol N} \quad \text{relative mol N} = \frac{0.350 \text{ mol N}}{0.350} = 1.00 \text{ mol N}$$

$$? \text{ mol O} = 16.82 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 1.05 \text{ mol O} \quad \text{relative mol O} = \frac{1.05 \text{ mol O}}{0.350} = 3.00 \text{ mol O}$$

The empirical formula is $C_{17}H_{19}NO_3$.
Empirical formula mass for $C_{17}H_{19}NO_3$:

| | | | |
|---|---|---|---|
| 17[C] × 12.0 | = | 204.0 | The molar mass of the empirical formula for |
| 19[H] × 1.0 | = | 19.0 | morphine is numerically equal to the molar mass |
| 1[N] × 14.0 | = | 14.0 | of morphine. Hence, the molecular formula for |

$$3[O] \times 16.0 \quad = \quad \underline{48.0}$$ morphine **is $C_{17}H_{19}NO_3$.**

$$= \quad 285.0$$

**3.156 *See Sections 3.1, 3.4, and Examples 3.1, 3.2, 3.15..***

Unbalanced: $NaWCl_6 \rightarrow Na_2WCl_6 + WCl_6$ Start with $Na_2$ $WCl_6$.

Step 1: $\underline{2}NaWCl_6 \rightarrow Na_2WCl_6 + WCl_6$ Balances Na, W, Cl.

*Strategy: g $NaWCl_6 \rightarrow$ mol $NaWCl_6 \rightarrow$ mol $WCl_6 \rightarrow$ g $WCl_6$*

$$? \text{ g } WCl_6 = 5.64 \text{ g } NaWCl_6 \times \frac{1 \text{ mol } NaWCI_6}{419.6 \text{ g } NaWCI_6} \times \frac{1 \text{ mol } WCI_6}{2 \text{ mol } NaWCI_6} \times \frac{396.6 \text{ g } WCI_6}{1 \text{ mol } WCI_6} = 2.67 \text{ g } WCl_6$$

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theorectical yield}} \times 100\% = \frac{1.52 \text{ g } WCI_6}{2.67 \text{ g } WCI_6} \times 100\% = \mathbf{56.9\%}$$

**3.158 *See Section 3.4.***

$$\% \text{ Cu} = \frac{\text{g Cu}}{\text{g sample}} \times 100\% = \frac{0.306 \text{ g Cu}}{1.20 \text{ g sample}} \times 100\% = 0.255\%$$

$$\times = \frac{\text{mol } H_2O}{\text{mol } CuSO_4}$$

$$? \text{ mol } CuSO_4 \text{ in sample} = 0.306 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{1 \text{ mol } CuSO_4}{1 \text{ mol Cu}} = 0.00482 \text{ mol Cu}$$

$$? \text{ g } CuSO_4 \text{ in sample} = 0.00482 \text{ mol } CuSO_4 \times \frac{159.62 \text{ g } CuSO_4}{1 \text{ mol } CuSO_4} = 0.769 \text{ g } CuSO_4$$

$$? \text{ g } H_2O \text{ in sample} = \text{g sample} - \text{g } CuSO_4 = 1.20 \text{ g} - 0.769 \text{ g} = 0.43 \text{ g } H_2O$$

$$? \text{ mol } H_2O \text{ in sample} = 0.43 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 0.024 \text{ mol } H_2O$$

$$\frac{\text{mol } H_2O}{\text{mol } CuSO_4} = \frac{0.024 \text{ mol}}{0.00482 \text{ mol}} = 5.0$$, so **the value of × in $CuSO_4 \cdot \times H_2O$ is 5.**

**3.160 *See Sections 3.1, 3.5, and Examples 3.1, 3.2, 3.16.***

| | | | | |
|---|---|---|---|---|
| Unbalanced: | $SCl_2 + NaF$ | $\rightarrow$ | $SF_4 + S_2Cl_2 + NaCl$ | Start with NaF. |
| Step 1: | $SCl_2 + \underline{4}NaF$ | $\rightarrow$ | $SF_4 + S_2Cl_2 + NaCl$ | Balances F. |
| Step 2: | $SCl_2 + 4NaF$ | $\rightarrow$ | $SF_4 + S_2Cl_2 + \underline{4}NaCl$ | Balances Na. |
| Step 3: | $\underline{3}SCl_2 + 4NaF$ | $\rightarrow$ | $SF_4 + S_2Cl_2 + 4NaCl$ | Balances S, Cl. |

*Strategy: g $SCl_2 \rightarrow$ mol $SCl_2 \rightarrow$ mol $SF_4 \rightarrow$ g $SF_4$*

$$? \text{ g } SF_4 \text{ based on } SCl_2 = 12.44 \text{ g } SCl_2 \times \frac{1 \text{ mol } SCI_2}{103.0 \text{ g } SCI_2} \times \frac{1 \text{ mol } SF_4}{3 \text{ mol } SCI_2} \times \frac{108.1 \text{ g } SF_4}{1 \text{ mol } SF_4} = 4.352 \text{ g } SF_4$$

*Strategy: g NaF $\rightarrow$ mol NaF $\rightarrow$ mol $SF_4 \rightarrow$ g $SF_4$*

$$? \text{ g } SF_4 \text{ based on NaF} = 10.11 \text{ g NaF} \times \frac{1 \text{ mol NaF}}{42.00 \text{ g NaF}} \times \frac{1 \text{ mol } SF_4}{4 \text{ mol NaF}} \times \frac{108.1 \text{ g } SF_4}{1 \text{ mol } SF_4} = 6.505 \text{ g } SF_4$$

$SCl_2$ is the limiting reactant because it produces less $SF_4$. The maximum amount of $SF_4$ that can be produced from 12.44 g $SCl_2$ and 10.11 g NaF is **4.352 g $SF_4$.**

**3.162** ***See Sections. 3.1 and 2.8.***

(a) Reactants: solid copper; concentrated nitric acid
Products: aqueous copper(II) nitrate; gaseous nitrogen dioxide; liquid water

(b)

| | | | | |
|---|---|---|---|---|
| Unbalanced: | $Cu + HNO_3$ | $\rightarrow$ | $Cu(NO_3)_2 + NO_2 + H_2O$ | Start with $HNO_3$ |
| Step 1: | $Cu + \underline{4}HNO_3$ | $\rightarrow$ | $Cu(NO_3)_2 + NO_2 + \underline{2}H_2O$ | Balances H |
| Step 2: | $Cu + HNO_3$ | $\rightarrow$ | $Cu(NO_3)_2 + \underline{2}NO_2 + H_2O$ | Balances N and O |

(c) Reactants: Cu = 0; $HNO_3$ (H = +1, N = +5, O = -2)
Products: $Cu(NO_3)_2$ (Cu = +2, N = +5, O = -2); $NO_2$ (N = +4, O = -2); $H_2O$ (H = +1, O = -2)

**Because the oxidation numbers of Cu and N change this is a redox reaction.**