CHAPTER 3 STOICHIOMETRY OF FORMULAS AND EQUATIONS

FOLLOW–UP PROBLEMS

3.1A  Plan: The mass of carbon must be changed from mg to g. The molar mass of carbon can then be used to determine the number of moles.

Solution:

\[
\text{Moles of carbon} = 315 \text{ mg C} \left( \frac{10^{-3} \text{ g}}{1 \text{ mg}} \right) \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 2.6228 \times 10^{-2} = 2.62 \times 10^{-2} \text{ mol C}
\]

Road map:

- Mass (mg) of C
- \(10^3 \text{ mg} = 1 \text{ g}\)
- Mass (g) of C
- Divide by \(M\) (g/mol)
- Amount (moles) of C

3.1B  Plan: The number of moles of aluminum must be changed to g. Then the mass of aluminum per can can be used to calculate the number of soda cans that can be made from 52 mol of Al.

Solution:

\[
\text{Number of soda cans} = 52 \text{ mol Al} \left( \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \right) \left( \frac{1 \text{ soda can}}{14 \text{ g Al}} \right) = 100.21 = 100 \text{ soda cans}
\]

Road map:

- Amount (mol) of Al
- Multiply by \(M\) (g/mol)
  (1 mol Al = 26.98 g Al)
- Mass (g) of Al
- 14 g Al = 1 soda can
- Number of cans

3.2A  Plan: Avogadro’s number is needed to convert the number of nitrogen molecules to moles. Since nitrogen molecules are diatomic (composed of two N atoms), the moles of molecules must be multiplied by 2 to obtain moles of atoms.

Solution:

\[
\text{Moles of N atoms} = \left( 9.72 \times 10^{21} \text{ N}_2 \text{ molecules} \right) \left( \frac{1 \text{ mol N}_2}{6.022 \times 10^{23} \text{ N}_2 \text{ molecules}} \right) \left( \frac{2 \text{ N atoms}}{1 \text{ mol N}_2} \right)
\]

\[
= 3.228 \times 10^{-2} = 3.23 \times 10^{-2} \text{ mol N}
\]
3.2B  **Plan:** Avogadro’s number is needed to convert the number of moles of He to atoms.
**Solution:**

Number of He atoms = 325 mol He \( \left( \frac{6.022 \times 10^{23} \text{ He atoms}}{1 \text{ mol He}} \right) \) = 1.9572 \( \times 10^{26} \) = **1.96 \( \times 10^{26} \) He atoms**

3.3A  **Plan:** Avogadro’s number is needed to convert the number of atoms to moles. The molar mass of manganese can then be used to determine the number of grams.
**Solution:**

Mass (g) of Mn = \( \left( 3.22 \times 10^{20} \text{ Mn atoms} \right) \left( \frac{1 \text{ mol Mn}}{6.022 \times 10^{23} \text{ Mn atoms}} \right) \left( \frac{54.94 \text{ g Mn}}{1 \text{ mol Mn}} \right) \)  
= 2.9377 \( \times 10^{-2} \) = **2.94 \( \times 10^{-2} \) g Mn**

3.3B  **Plan:** Use the molar mass of copper to calculate the number of moles of copper present in a penny. Avogadro’s number is then needed to convert the number of moles of Cu to Cu atoms.
**Solution:**

Number of Cu atoms = 0.0625 g Cu \( \left( \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \right) \left( \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}} \right) \)  
= 5.9225 \( \times 10^{20} \) = **5.92 \( \times 10^{20} \) Cu atoms**
3.4A  Plan: Avogadro’s number is used to change the number of formula units to moles. Moles may be changed to mass using the molar mass of sodium fluoride, which is calculated from its formula.

Solution:
The formula of sodium fluoride is NaF.
\[ M_{\text{NaF}} = (1 \times M_{\text{Na}}) + (1 \times M_{\text{F}}) = 22.99 \text{ g/mol} + 19.00 \text{ g/mol} = 41.99 \text{ g/mol} \]

\[ \text{Mass (g)} \text{ of NaF} = \left( \frac{1.19 \times 10^{19} \text{ NaF formula units}}{6.022 \times 10^{23} \text{ NaF formula units}} \right) \left( \frac{41.99 \text{ g NaF}}{1 \text{ mol NaF}} \right) = 8.29759 \times 10^{-4} = 8.30 \times 10^{-4} \text{ g NaF} \]

3.4B  Plan: Convert the mass of calcium chloride from pounds to g. Use the molar mass to calculate the number of moles of calcium chloride in the sample. Finally, use Avogadro’s number to change the number of moles to formula units.

Solution:
\[ M_{\text{CaCl}_2} = (1 \times M_{\text{Ca}}) + (2 \times M_{\text{Cl}}) = 40.08 \text{ g/mol} + 2(35.45 \text{ g/mol}) = 110.98 \text{ g/mol} \]

\[ \text{Number of formula units of CaCl}_2 = 400 \text{ lb} \left( \frac{453.6 \text{ g}}{1 \text{ lb}} \right) \left( \frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2} \right) \left( \frac{6.022 \times 10^{23} \text{ formula units CaCl}_2}{1 \text{ mol CaCl}_2} \right) = 9.8453 \times 10^{26} = 1 \times 10^{27} \text{ formula units CaCl}_2 \]
Multiply by Avogadro’s number
(1 mol CaCl₂ = 6.022 x 10²³ CaCl₂ formula units)

No. of formula units of CaCl₂

3.5A Plan: Avogadro’s number is used to change the number of molecules to moles. Moles may be changed to mass by multiplying by the molar mass. The molar mass of tetraphosphorus decoxide is obtained from its chemical formula. Each molecule has four phosphorus atoms, so the total number of atoms is four times the number of molecules.

Solution:
a) Tetra = 4, and deca = 10 to give P₄O₁₀.

The molar mass, \( M \), is the sum of the atomic weights, expressed in g/mol:
\[
P = 4(30.97) = 123.88 \text{ g/mol} \\
O = 10(16.00) = 160.00 \text{ g/mol} \\
\]
\[
M = 123.88 \text{ g/mol} + 160.00 \text{ g/mol} = 283.88 \text{ g/mol of P₄O₁₀} \\
\]
Mass (g) of P₄O₁₀ = \( \frac{1 \text{ mol P₄O₁₀}}{6.022 \times 10²³ \text{ molecules P₄O₁₀}} \times \frac{283.88 \text{ g}}{1 \text{ mol}} \)
\[= \frac{6.022 \times 10²³ \text{ molecules P₄O₁₀}}{1 \text{ mol}} \times 283.88 \text{ g/mol} \]
\[= 21.9203 = 21.9 \text{ g P₄O₁₀} \]
b) Number of P atoms = \( \frac{1 \text{ mol P₄O₁₀}}{6.022 \times 10²³ \text{ molecules P₄O₁₀}} \times \frac{4 \text{ atoms P}}{1 \text{ P₄O₁₀ molecule}} \)
\[= \frac{1 \text{ mol P₄O₁₀}}{6.022 \times 10²³ \text{ molecules P₄O₁₀}} \times \frac{4 \text{ atoms P}}{1 \text{ P₄O₁₀ molecule}} \]
\[= 1.86 \times 10²³ \text{ P atoms} \]

3.5B Plan: The mass of calcium phosphate is converted to moles of calcium phosphate by dividing by the molar mass. Avogadro’s number is used to change the number of moles to formula units. Each formula unit has two phosphate ions, so the total number of phosphate ions is two times the number of formula units.

Solution:
a) The formula of calcium phosphate is Ca₃(PO₄)₂.

The molar mass, \( M \), is the sum of the atomic weights, expressed in g/mol:
\[
M = (3 \times \text{M of Ca}) + (2 \times \text{M of P}) + (8 \times \text{M of O}) \\
= (3 \times 40.08 \text{ g/mol Ca}) + (2 \times 30.97 \text{ g/mol P}) + (8 \times 16.00 \text{ g/mol O}) \\
= 310.18 \text{ g/mol Ca₃(PO₄)₂} \\
\]
No. of formula units Ca₃(PO₄)₂ = \( \frac{75.5 \text{ g Ca₃(PO₄)₂}}{310.18 \text{ g Ca₃(PO₄)₂}} \)
\[= 1.4658 \times 10²³ = 1.47 \times 10²³ \text{ formula units Ca₃(PO₄)₂} \]
b) No. of phosphate (PO₄³⁻) ions = \( \frac{2 \text{ PO₄³⁻ ions}}{1 \text{ formula unit Ca₃(PO₄)₂}} \)
\[= 2.94 \times 10²³ \text{ phosphate ions} \]

3.6A Plan: Calculate the molar mass of glucose. The total mass of carbon in the compound divided by the molar mass of the compound, multiplied by 100% gives the mass percent of C.

Solution:
The formula for glucose is C₆H₁₂O₆. There are 6 atoms of C per each formula.
Molar mass of C₆H₁₂O₆ = (6 x \( \text{M of C} \)) + (12 x \( \text{M of H} \)) + (6 x \( \text{M of O} \))
\[
= (6 \times 12.01 \text{ g/mol}) + (12 \times 1.008 \text{ g/mol}) + (6 \times 16.00 \text{ g/mol}) \\
= 180.16 \text{ g/mol} \\
\]
3.6B Plan: Calculate the molar mass of CCl$_3$F. The total mass of chlorine in the compound divided by the molar mass of the compound, multiplied by 100% gives the mass percent of Cl.

Solution:
The formula is CCl$_3$F. There are 3 atoms of Cl per each formula.

Molar mass of CCl$_3$F = (1 x $M$ of C) + (3 x $M$ of Cl) + (1 x $M$ of F)

= (1 x 12.01 g/mol) + (3 x 35.45 g/mol) + (1 x 19.00 g/mol)

= 137.36 g/mol

Mass % of Cl = \( \frac{\text{total mass of Cl}}{\text{molar mass of CCl}_3F} \times 100 \)

= \( \frac{3 \times 35.45 \text{ g/mol}}{137.36 \text{ g/mol}} \times 100 \) = 77.42% Cl

3.7A Plan: Multiply the mass of the sample by the mass fraction of C found in the preceding problem.

Solution:

Mass (g) of C = 16.55 g C$_6$H$_{12}$O$_6$ \( \times \) \( \frac{72.06 \text{ g C}}{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6} \) = 6.6196 = 6.620 g C

3.7B Plan: Multiply the mass of the sample by the mass fraction of Cl found in the preceding problem.

Solution:

Mass (g) of Cl = 112 g CCl$_3$F \( \times \) \( \frac{106.35 \text{ g Cl}}{137.36 \text{ g CCl}_3F} \) = 86.6900 = 86.7 g Cl

3.8A Plan: We are given fractional amounts of the elements as subscripts. Convert the fractional amounts to whole numbers by dividing each number by the smaller number and then multiplying by the smallest integer that will turn both subscripts into integers.

Solution:

Divide each subscript by the smaller value, 0.170: B$_{0.170}$O$_{0.255}$ = B$_1$O$_{1.5}$

Multiply the subscripts by 2 to obtain integers: B$_1$x$_2$O$_1$.5$x_2$ = B$_2$O$_3$

3.8B Plan: We are given fractional amounts of the elements as subscripts. Convert the fractional amounts to whole numbers by dividing each number by the smaller number and then multiplying by the smallest integer that will turn both subscripts into integers.

Solution:

Divide each subscript by the smaller value, 6.80: C$_{6.80}$H$_{18.1}$ = C$_1$H$_{2.67}$

Multiply the subscripts by 3 to obtain integers: C$_{1.3}$H$_{2.66}$x$_3$ = C$_4$H$_8$

3.9A Plan: Calculate the number of moles of each element in the sample by dividing by the molar mass of the corresponding element. The calculated numbers of moles are the fractional amounts of the elements and can be used as subscripts in a chemical formula. Convert the fractional amounts to whole numbers by dividing each number by the smallest subscripted number.

Solution:

Moles of H = 1.23 g H \( \times \) \( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \) = 1.22 mol H

Moles of P = 12.64 g P \( \times \) \( \frac{1 \text{ mol P}}{30.97 \text{ g P}} \) = 0.408 mol P

Moles of O = 26.12 g O \( \times \) \( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \) = 1.63 mol O

Divide each subscript by the smaller value, 0.408: H$_{1.22}$ P$_{0.408}$ O$_{1.63}$ = H$_3$PO$_4$
3.9B  **Plan:** The moles of sulfur may be calculated by dividing the mass of sulfur by the molar mass of sulfur. The moles of sulfur and the chemical formula will give the moles of M. The mass of M divided by the moles of M will give the molar mass of M. The molar mass of M can identify the element.

**Solution:**

\[
\text{Moles of } S = \left( \frac{2.88 \text{ g } S}{32.06 \text{ g } S} \right) \left( \frac{1 \text{ mol } S}{32.06 \text{ g } S} \right) = 0.0898 \text{ mol } S
\]

\[
\text{Moles of } M = \left( \frac{0.0898 \text{ mol } S}{2 \text{ mol } S} \right) \left( \frac{2 \text{ mol } M}{3 \text{ mol } S} \right) = 0.0599 \text{ mol } M
\]

\[
\text{Molar mass of } M = \frac{3.12 \text{ g } M}{0.0599 \text{ mol } M} = 52.0868 = 52.1 \text{ g/mol}
\]

The element is Cr (52.00 g/mol); M is **Chromium** and M₂S₃ is **chromium(III) sulfide**.

3.10A  **Plan:** If we assume there are 100 grams of this compound, then the masses of carbon and hydrogen, in grams, are numerically equivalent to the percentages. Divide the atomic mass of each element by its molar mass to obtain the moles of each element. Dividing each of the moles by the smaller value gives the simplest ratio of C and H. The smallest multiplier to convert the ratios to whole numbers gives the empirical formula. To obtain the molecular formula, divide the given molar mass of the compound by the molar mass of the empirical formula to find the whole-number by which the empirical formula is multiplied.

**Solution:**

Assuming 100 g of compound gives 95.21 g C and 4.79 g H:

\[
\text{Moles of } C = \left( \frac{95.21 \text{ g } C}{12.01 \text{ g } C} \right) \left( \frac{1 \text{ mol } C}{12.01 \text{ g } C} \right) = 7.92756 \text{ mol } C
\]

\[
\text{Mole of } H = \left( \frac{4.79 \text{ g } H}{1.008 \text{ g } H} \right) \left( \frac{1 \text{ mol } H}{1.008 \text{ g } H} \right) = 4.75198 \text{ mol } H
\]

Divide each of the moles by 4.75198, the smaller value:

\[
\text{C}_{\frac{7.92756}{4.75198}}\text{H}_{\frac{4.75198}{4.75198}} = \text{C}_{1.6683}\text{H}_{1}
\]

The value 1.668 is 5/3, so the moles of C and H must each be multiplied by 3. If it is not obvious that the value is near 5/3, use a trial and error procedure whereby the value is multiplied by the successively larger integer until a value near an integer results. This gives C₅H₃ as the empirical formula. The molar mass of this formula is:

\[
\frac{(5 \times 12.01 \text{ g/mol}) + (3 \times 1.008 \text{ g/mol})}{63.074 \text{ g/mol}} = 252.30 \text{ g/mol}
\]

Thus, the empirical formula must be multiplied by 4 to give 4(C₅H₃) = C₂₀H₁₂ as the molecular formula of benzo[a]pyrene.

3.10B  **Plan:** If we assume there are 100 grams of this compound, then the masses of carbon, hydrogen, nitrogen, and oxygen, in grams, are numerically equivalent to the percentages. Divide the atomic mass of each element by its molar mass to obtain the moles of each element. Dividing each of the moles by the smaller value gives the simplest ratio of C, H, N, and O. To obtain the molecular formula, divide the given molar mass of the compound by the molar mass of the empirical formula to find the whole-number by which the empirical formula is multiplied.

**Solution:**

Assuming 100 g of compound gives 49.47 g C, 5.19 g H, 28.86 g N, and 16.48 g O:

\[
\text{Moles of } C = \left( \frac{49.47 \text{ g } C}{12.01 \text{ g } C} \right) \left( \frac{1 \text{ mol } C}{12.01 \text{ g } C} \right) = 4.119 \text{ mol } C
\]

\[
\text{Moles of } H = \left( \frac{5.19 \text{ g } H}{1.008 \text{ g } H} \right) \left( \frac{1 \text{ mol } H}{1.008 \text{ g } H} \right) = 5.15 \text{ mol } H
\]

\[
\text{Moles of } N = \left( \frac{28.86 \text{ g } N}{14.01 \text{ g } N} \right) \left( \frac{1 \text{ mol } N}{14.01 \text{ g } N} \right) = 2.060 \text{ mol } N
\]

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Moles of O = 16.48 g O \left( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 1.030 \text{ mol O}

Divide each subscript by the smaller value, 1.030:

\begin{align*}
\text{C} & \quad \div \quad 1.030 = 4.119 \\
\text{H} & \quad \div \quad 1.030 = 5.15 \\
\text{N} & \quad \div \quad 1.030 = 2.060 \\
\text{O} & \quad \div \quad 1.030 = 1.030
\end{align*}

This gives \( \text{C}_4\text{H}_5\text{N}_2\text{O} \) as the empirical formula. The molar mass of this formula is:

\( (4 \times 12.01 \text{ g/mol}) + (5 \times 1.008 \text{ g/mol}) + (2 \times 14.01 \text{ g/mol}) + (1 \times 16.00 \text{ g/mol}) = 97.10 \text{ g/mol} \)

The molar mass of caffeine is 194.2 g/mol, which is larger than the empirical formula mass of 97.10 g/mol, so the molecular formula must be a whole-number multiple of the empirical formula.

\[
\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}} = \frac{194.2 \text{ g/mol}}{97.10 \text{ g/mol}} = 2
\]

Thus, the empirical formula must be multiplied by 2 to give \( 2(\text{C}_4\text{H}_5\text{N}_2\text{O}) = \text{C}_8\text{H}_{10}\text{N}_4\text{O}_2 \) as the molecular formula of caffeine.

3.11A Plan: The carbon in the sample is converted to carbon dioxide, the hydrogen is converted to water, and the remaining material is chlorine. The grams of carbon dioxide and the grams of water are both converted to moles. One mole of carbon dioxide gives one mole of carbon, while one mole of water gives two moles of hydrogen. Using the molar masses of carbon and hydrogen, the grams of each of these elements in the original sample may be determined. The original mass of sample minus the masses of carbon and hydrogen gives the mass of chlorine. The mass of chlorine and the molar mass of chlorine will give the moles of chlorine. Once the moles of each of the elements have been calculated, divide by the smallest value, and, if necessary, multiply by the smallest number required to give a set of whole numbers for the empirical formula. Compare the molar mass of the empirical formula to the molar mass given in the problem to find the molecular formula.

Solution:
Determine the moles and the masses of carbon and hydrogen produced by combustion of the sample.

\[
\begin{align*}
0.451 \text{ g CO}_2 & \left( \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) = 0.010248 \text{ mol C} \\
0.0617 \text{ g H}_2\text{O} & \left( \frac{2 \text{ mol H}}{18.016 \text{ g H}_2\text{O}} \right) = 0.0068495 \text{ mol H}
\end{align*}
\]

The mass of chlorine is given by: 0.250 g sample – (0.12307 g C + 0.006904 g H) = 0.120 g Cl

The moles of chlorine are:

\[
0.120 \text{ g Cl} \left( \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} \right) = 0.0033850 \text{ mol Cl}. \text{ This is the smallest number of moles.}
\]

Divide each mole value by the lowest value, 0.0033850:

\[
\begin{align*}
\text{C} & \quad 0.010248 \div 0.0033850 = 3 \\
\text{H} & \quad 0.0068495 \div 0.0033850 = 2.06 \\
\text{Cl} & \quad 0.0033850 = 1
\end{align*}
\]

The empirical formula has the following molar mass:

\[
(3 \times 12.01 \text{ g/mol}) + (2 \times 1.008 \text{ g/mol}) + (35.45 \text{ g/mol}) = 73.496 \text{ g/mol C}_3\text{H}_2\text{Cl}
\]

\[
\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}} = \frac{146.99 \text{ g/mol}}{73.496 \text{ g/mol}} = 2
\]

Thus, the molecular formula is two times the empirical formula, \( 2(\text{C}_3\text{H}_2\text{Cl}) = \text{C}_6\text{H}_4\text{Cl}_2 \).

3.11B Plan: The carbon in the sample is converted to carbon dioxide, the hydrogen is converted to water, and the remaining material is oxygen. The grams of carbon dioxide and the grams of water are both converted to moles. One mole of carbon dioxide gives one mole of carbon, while one mole of water gives two moles of hydrogen. Using the molar masses of carbon and hydrogen, the grams of each of these elements in the original sample may be determined. The original mass of sample minus the masses of carbon and hydrogen gives the mass of oxygen. The mass of oxygen and the molar mass of oxygen will give the moles of oxygen. Once the moles of each of the elements have been calculated, divide by the smallest value, and, if necessary, multiply by the smallest number required to give a set of whole numbers for the empirical formula. Compare the molar mass of the empirical formula to the molar mass given in the problem to find the molecular formula.

Solution:
Determine the moles and the masses of carbon and hydrogen produced by combustion of the sample.
3.516 g CO₂ \( \left( \frac{1 \text{ mol CO₂}}{44.01 \text{ g CO₂}} \right) \) \( \times \) \( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \) = 0.07989 mol C \( \times \) \( \frac{1 \text{ mol CO₂}}{44.01 \text{ g CO₂}} \) = 0.9595 g C

1.007 g H₂O \( \left( \frac{1 \text{ mol H₂O}}{18.02 \text{ g H₂O}} \right) \) \( \times \) \( \frac{2 \text{ mol H}}{1 \text{ mol H₂O}} \) = 0.1118 mol H \( \times \) \( \frac{1 \text{ mol H₂O}}{18.02 \text{ g H₂O}} \) = 0.1127 g H

The mass of oxygen is given by: 1.200 g sample – (0.9595 g C + 0.1127 g H) = 0.128 g O

The moles of C and H are calculated above. The moles of oxygen are:

\[ \frac{0.128 \text{ g O}}{16.00 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}}} = 0.00800 \text{ mol O}. \]

This gives C₁₀H₁₄O as the empirical formula. The molar mass of this formula is:

\[ 10 \times 12.01 \text{ g/mol} + 14 \times 1.008 \text{ g/mol} + 16.00 \text{ g/mol} = 150.21 \text{ g/mol}. \]

The molar mass of the steroid is 300.42 g/mol, which is larger than the empirical formula mass of 150.21 g/mol, so the molecular formula must be a whole-number multiple of the empirical formula.

\[ \text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}} = \frac{300.42 \text{ g/mol}}{150.21 \text{ g/mol}} = 2. \]

Thus, the empirical formula must be multiplied by 2 to give \( 2 \times (\text{C}_{10}\text{H}_{14}\text{O}) = \text{C}_{20}\text{H}_{28}\text{O}_2 \) as the molecular formula of the steroid.

3.12A Plan: In each part it is necessary to determine the chemical formulas, including the physical states, for both the reactants and products. The formulas are then placed on the appropriate sides of the reaction arrow. The equation is then balanced.

Solution:

a) Sodium is a metal (solid) that reacts with water (liquid) to produce hydrogen (gas) and a solution of sodium hydroxide (aqueous). Sodium is Na; water is H₂O; hydrogen is H₂; and sodium hydroxide is NaOH.

\[ \text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{(g)} + \text{NaOH(aq)} \]

Balancing will precede one element at a time. One way to balance hydrogen gives:

\[ 2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{(g)} + 2\text{NaOH(aq)} \]

Next, the sodium will be balanced:

\[ 2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{(g)} + 2\text{NaOH(aq)} \]

On inspection, we see that the oxygen is already balanced.

b) Aqueous nitric acid reacts with calcium carbonate (solid) to produce carbon dioxide (gas), water (liquid), and aqueous calcium nitrate. Nitric acid is HNO₃; calcium carbonate is CaCO₃; carbon dioxide is CO₂; water is H₂O; and calcium nitrate is Ca(NO₃)₂. The starting equation is

\[ \text{HNO}_3\text{(aq)} + \text{CaCO}_3\text{(s)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)} + \text{Ca(NO}_3\text{)}_2\text{(aq)} \]

Initially, Ca and C are balanced. Proceeding to another element (such as N, or better yet the group of elements in NO₃⁻) gives the following partially balanced equation:

\[ 2\text{HNO}_3\text{(aq)} + \text{CaCO}_3\text{(s)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)} + \text{Ca(NO}_3\text{)}_2\text{(aq)} \]

Now, all the elements are balanced.

c) We are told all the substances involved are gases. The reactants are phosphorus trichloride and hydrogen fluoride, while the products are phosphorus trifluoride and hydrogen chloride. Phosphorus trichloride is PCl₃; hydrogen fluoride is HF; and hydrogen chloride is HCl. The initial equation is:

\[ \text{PCl}_3\text{(g)} + \text{HF(g)} \rightarrow \text{PF}_3\text{(g)} + \text{HCl(g)} \]

Initially, P and H are balanced. Proceed to another element (either F or Cl); if we will choose Cl, it balances as:

\[ \text{PCl}_3\text{(g)} + \text{HF(g)} \rightarrow \text{PF}_3\text{(g)} + 3\text{HCl(g)} \]

The balancing of the Cl unbalances the H, this should be corrected by balancing the H as:

\[ \text{PCl}_3\text{(g)} + 3\text{HF(g)} \rightarrow \text{PF}_3\text{(g)} + 3\text{HCl(g)} \]

Now, all the elements are balanced.
3.12B  **Plan:** In each part it is necessary to determine the chemical formulas, including the physical states, for both the reactants and products. The formulas are then placed on the appropriate sides of the reaction arrow. The equation is then balanced.

**Solution:**

a) We are told that nitroglycerine is a liquid reactant, and that all the products are gases. The formula for nitroglycerine is given. Carbon dioxide is CO2; water is H2O; nitrogen is N2; and oxygen is O2. The initial equation is:

\[
C_3H_5N_3O_9(l) \rightarrow CO_2(g) + H_2O(g) + N_2(g) + O_2(g)
\]

Counting the atoms shows no atoms are balanced.

One element should be picked and balanced. Any element except oxygen will work. Oxygen will not work in this case because it appears more than once on one side of the reaction arrow. We will start with carbon. Balancing C gives:

\[
C_3H_5N_3O_9(l) \rightarrow 3CO_2(g) + H_2O(g) + N_2(g) + O_2(g)
\]

Now balancing the hydrogen gives:

\[
C_3H_5N_3O_9(l) \rightarrow 3CO_2(g) + 5/2H_2O(g) + N_2(g) + O_2(g)
\]

Similarly, if we balance N we get:

\[
C_3H_5N_3O_9(l) \rightarrow 3CO_2(g) + 5/2H_2O(g) + 3/2N_2(g) + O_2(g)
\]

Clear the fractions by multiplying everything except the unbalanced oxygen by 2:

\[
2C_3H_5N_3O_9(l) \rightarrow 6CO_2(g) + 5H_2O(g) + 3N_2(g) + O_2(g)
\]

This leaves oxygen to balance. Balancing oxygen gives:

\[
2C_3H_5N_3O_9(l) \rightarrow 6CO_2(g) + 5H_2O(g) + 3N_2(g) + 1/2O_2(g)
\]

Again clearing fractions by multiplying everything by 2 gives:

\[
4C_3H_5N_3O_9(l) \rightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)
\]

Now all the elements are balanced.

b) Potassium superoxide (KO2) is a solid. Carbon dioxide (CO2) and oxygen (O2) are gases. Potassium carbonate (K₂CO₃) is a solid. The initial equation is:

\[
KO_2(s) + CO_2(g) \rightarrow O_2(g) + K_2CO_3(s)
\]

Counting the atoms indicates that the carbons are balanced, but none of the other atoms are balanced.

One element should be picked and balanced. Any element except oxygen will work (oxygen will be more challenging to balance because it appears more than once on each side of the reaction arrow). Because the carbons are balanced, we will start with potassium. Balancing potassium gives:

\[
2KO_2(s) + CO_2(g) \rightarrow O_2(g) + K_2CO_3(s)
\]

Now all elements except for oxygen are balanced. Balancing oxygen by adding a coefficient in front of the O₂ gives:

\[
2KO_2(s) + CO_2(g) \rightarrow 3/2O_2(g) + K_2CO_3(s)
\]

Clearing the fractions by multiplying everything by 2 gives:

\[
4KO_2(s) + 2CO_2(g) \rightarrow 3O_2(g) + 2K_2CO_3(s)
\]

Now all the elements are balanced.

c) Iron(III) oxide (Fe₂O₃) is a solid, as is iron metal (Fe). Carbon monoxide (CO) and carbon dioxide (CO₂) are gases. The initial equation is:

\[
Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)
\]

Counting the atoms indicates that the carbons are balanced, but none of the other atoms are balanced.

One element should be picked and balanced. Because oxygen appears in more than one compound on one side of the reaction arrow, it is best not to start with that element. Because the carbons are balanced, we will start with iron. Balancing iron gives:

\[
Fe_2O_3(s) + CO(g) \rightarrow 2Fe(s) + CO_2(g)
\]

Now all the atoms but oxygen are balanced. There are 4 oxygen atoms on the left hand side of the reaction arrow and 2 oxygen atoms on the right hand side of the reaction arrow. In order to balance the oxygen, we want to change the coefficients in front of the carbon-containing compounds (if we changed the coefficient in front of the iron(III) oxide, the iron atoms would no longer be balanced). To maintain the balance of carbons, the coefficients in front of the carbon monoxide and the carbon dioxide must be the same. On the left hand side of the equation, there are 3 oxygens in Fe₂O₃ plus 1X oxygen atoms from the CO (where X is the coefficient in the balanced equation). On the right hand side of the equation, there are 2X oxygen atoms. The number of oxygen atoms on both sides of the equation should be the same:

\[
3 + 1X = 2X
\]
Balancing oxygen by adding a coefficient of 3 in front of the CO and CO₂ gives:

\[
\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})
\]

Now all the elements are balanced.

3.13A Plan: Count the number of each type of atom in each molecule to write the formulas of the reactants and products.
Solution:
\[
6\text{CO}(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g})
\]
or,
\[
2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})
\]

3.13B Plan: Count the number of each type of atom in each molecule to write the formulas of the reactants and products.
Solution:
\[
6\text{H}_2(\text{g}) + 2\text{N}_2(\text{g}) \rightarrow 4\text{NH}_3(\text{g})
\]
or,
\[
3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})
\]

3.14A Plan: The reaction, like all reactions, needs a balanced chemical equation. The balanced equation gives the molar ratio between the moles of iron and moles of iron(III) oxide.
Solution:
The names and formulas of the substances involved are: iron(III) oxide, Fe₂O₃, and aluminum, Al, as reactants, and aluminum oxide, Al₂O₃, and iron, Fe, as products. The iron is formed as a liquid; all other substances are solids. The equation begins as:

\[
\text{Fe}_2\text{O}_3(\text{s}) + \text{Al}(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + \text{Fe}(\text{l})
\]

There are 2 Fe, 3 O, and 1 Al on the reactant side and 1 Fe, 3 O, and 2 Al on the product side.
Balancing aluminum: \[
\text{Fe}_2\text{O}_3(\text{s}) + 2\text{Al}(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + \text{Fe}(\text{l})
\]
Balancing iron: \[
\text{Fe}_2\text{O}_3(\text{s}) + 2\text{Al}(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + 2\text{Fe}(\text{l})
\]

Moles of Fe₂O₃ = \(3.60 \times 10^3 \text{ mol Fe} \left(\frac{1 \text{ mol Fe}_2\text{O}_3}{2 \text{ mol Fe}}\right) = 1.80 \times 10^3 \text{ mol Fe}_2\text{O}_3\)

Road map:

<table>
<thead>
<tr>
<th>Amount (moles) of Fe₂O₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>Molar ratio</td>
</tr>
<tr>
<td>(2 mol Fe = 1 mol Fe₂O₃)</td>
</tr>
</tbody>
</table>

3.14B Plan: The reaction, like all reactions, needs a balanced chemical equation. The balanced equation gives the molar ratio between the moles of aluminum and moles of silver sulfide.
Solution:
The names and formulas of the substances involved are: silver sulfide, Ag₂S, and aluminum, Al, as reactants; and aluminum sulfide, Al₂S₃, and silver, Ag, as products. All reactants and compounds are solids. The equation begins as:

\[
\text{Ag}_2\text{S}(\text{s}) + \text{Al}(\text{s}) \rightarrow \text{Al}_2\text{S}_3(\text{s}) + \text{Ag}(\text{s})
\]

There are 2 Ag, 1 S, and 1 Al on the reactant side and 2 Al, 3 S, and 1 Ag on the product side.
Balancing sulfur: \[
3\text{Ag}_2\text{S}(\text{s}) + \text{Al}(\text{s}) \rightarrow \text{Al}_2\text{S}_3(\text{s}) + \text{Ag}(\text{s})
\]
Balancing silver: \[
3\text{Ag}_2\text{S}(\text{s}) + \text{Al}(\text{s}) \rightarrow \text{Al}_2\text{S}_3(\text{s}) + 6\text{Ag}(\text{s})
\]
Balancing aluminum: \[
3\text{Ag}_2\text{S}(\text{s}) + 2\text{Al}(\text{s}) \rightarrow \text{Al}_2\text{S}_3(\text{s}) + 6\text{Ag}(\text{s})
\]

Moles of Al = 0.253 \text{ mol Ag}_2\text{S} \left(\frac{2 \text{ mol Al}}{3 \text{ mol Ag}_2\text{S}}\right) = 0.1687 = 0.169 \text{ mol Al}
3.15A  
**Plan:** Divide the formula units of aluminum oxide by Avogadro’s number to obtain moles of compound. The balanced equation gives the molar ratio between moles of iron(III) oxide and moles of iron.

**Solution:**

\[
\text{Moles of Fe} = \left(1.85 \times 10^{25} \text{ Fe}_2\text{O}_3 \text{ formula units}\right) \left(\frac{1 \text{ mol Fe}_2\text{O}_3}{6.022 \times 10^{23} \text{ Fe}_2\text{O}_3 \text{ formula units}}\right) \left(\frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3}\right) = 61.4414 = 61.4 \text{ mol Fe}
\]

Road map:

- Amount (moles) of Ag₂S
- Molar ratio (3 mol Ag₂S = 2 mol Al)
- Amount (moles) of Al

3.15B  
**Plan:** Divide the mass of silver sulfide by its molar mass to obtain moles of the compound. The balanced equation gives the molar ratio between moles of silver sulfide and moles of silver.

**Solution:**

\[
\text{Moles of Ag} = \left(\frac{32.6 \text{ g Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}}\right) \left(\frac{6 \text{ mol Ag}}{3 \text{ mol Ag}_2\text{S}}\right) = 0.2630 = 0.263 \text{ mol Ag}
\]

Road map:

- No. of Fe₂O₃ formula units
- Divide by Avogadro’s number (6.022 x 10²³ Fe₂O₃ formula units = 1 mol Fe₂O₃)
- Amount (moles) of Fe₂O₃
- Molar ratio (1 mol Fe₂O₃ = 2 mol Fe)
- Amount (moles) of Fe

3.16A  
**Plan:** The mass of aluminum oxide must be converted to moles by dividing by its molar mass. The balanced chemical equation (follow-up problem 3.14A) shows there are two moles of aluminum for every mole of aluminum oxide. Multiply the moles of aluminum by Avogadro’s number to obtain atoms of Al.
Solution:
Atoms of Al = \left( \frac{1.00 \text{ g Al}_2\text{O}_3}{101.96 \text{ g Al}_2\text{O}_3} \right) \left( \frac{2 \text{ mol Al}}{1 \text{ mol Al}_2\text{O}_3} \right) \left( \frac{6.022 \times 10^{23} \text{ atoms Al}}{1 \text{ mol Al}} \right)
= 1.18125 \times 10^{22} = 1.18 \times 10^{22} \text{ atoms Al}

Road map:

- Mass (g) of Al\(_2\)O\(_3\)
- Divide by \(\mathcal{M}\) (g/mol)
- Amount (moles) of Al\(_2\)O\(_3\)
- Molar ratio
- Amount (moles) of Al
- Multiply by Avogadro’s number
- Number of Al atoms

3.16B Plan: The mass of aluminum sulfide must be converted to moles by dividing by its molar mass. The balanced chemical equation (follow-up problem 3.14B) shows there are two moles of aluminum for every mole of aluminum sulfide. Multiply the moles of aluminum by its molar mass to obtain the mass (g) of aluminum.

Solution:
Mass (g) of Al = 12.1 \text{ g Al}_2\text{S}_3 \left( \frac{1 \text{ mol Al}}{150.14 \text{ g Al}_2\text{S}_3} \right) \left( \frac{2 \text{ mol Al}}{1 \text{ mol Al}_2\text{S}_3} \right) \left( \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \right) = 4.3487 = 4.35 \text{ g Al}

Road map:

- Mass (g) of Al\(_2\)S\(_3\)
- Divide by \(\mathcal{M}\) (g/mol)
- (150.14 g Al\(_2\)S\(_3\) = 1 mol Al\(_2\)S\(_3\))
- Amount (moles) of Al\(_2\)S\(_3\)
- Molar ratio
- (1 mol Al\(_2\)S\(_3\) = 2 mol Al)
- Amount (moles) of Al
- Multiply by \(\mathcal{M}\) (g/mol)
- (1 mol Al = 26.98 g Al)
- Mass (g) of Al

3.17A Plan: Write the balanced chemical equation for each step. Add the equations, canceling common substances.

Solution:
Step 1 \(2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g)\)
Step 2 \(\text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq)\)

3-12
Adjust the coefficients since 2 moles of SO$_3$ are produced in Step 1 but only 1 mole of SO$_3$ is consumed in Step 2. We have to double all of the coefficients in Step 2 so that the amount of SO$_3$ formed in Step 1 is used in Step 2.

Step 1 \[ 2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g) \]
Step 2 \[ 2\text{SO}_3(g) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2\text{SO}_4(aq) \]

Add the two equations and cancel common substances.

Step 1 \[ 2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g) \]
Step 2 \[ 2\text{SO}_3(g) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2\text{SO}_4(aq) \]

Or \[ 2\text{SO}_2(g) + \text{O}_2(g) + 2\text{SO}_3(g) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2\text{SO}_4(aq) \]

3.17B Plan: Write the balanced chemical equation for each step. Add the equations, canceling common substances.

Solution:

Step 1 \[ \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \]
Step 2 \[ \text{NO}(g) + 2\text{O}_3(g) \rightarrow 2\text{NO}_2(g) + 2\text{O}_2(g) \]

Adjust the coefficients since 2 moles of NO are produced in Step 1 but only 1 mole of NO is consumed in Step 2. We have to double all of the coefficients in Step 2 so that the amount of NO formed in Step 1 is used in Step 2.

Step 1 \[ \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \]
Step 2 \[ 2\text{NO}(g) + 2\text{O}_3(g) \rightarrow 2\text{NO}_2(g) + 2\text{O}_2(g) \]

Add the two equations and cancel common substances.

Step 1 \[ \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \]
Step 2 \[ 2\text{NO}(g) + 2\text{O}_3(g) \rightarrow 2\text{NO}_2(g) + 2\text{O}_2(g) \]

\[ \text{N}_2(g) + 2\text{O}_3(g) \rightarrow 2\text{NO}_2(g) + \text{O}_2(g) \]

Or

\[ \text{N}_2(g) + \text{O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g) \]

3.18A Plan: Count the molecules of each type, and find the simplest ratio. The simplest ratio leads to a balanced chemical equation. The substance with no remaining particles is the limiting reagent.

Solution:

4 AB molecules react with 3 B$_2$ molecules to produce 4 molecules of AB$_2$, with 1 B$_2$ molecule remaining unreacted. The balanced chemical equation is

\[ 4\text{AB}(g) + 2\text{B}_2(g) \rightarrow 4\text{AB}_2(g) \text{ or } 2\text{AB}(g) + \text{B}_2(g) \rightarrow \text{AB}_2(g) \]

The limiting reagent is AB since there is a B$_2$ molecule left over (excess).

3.18B Plan: Write a balanced equation for the reaction. Use the molar ratios in the balanced equation to find the amount (molecules) of SO$_3$ produced when each reactant is consumed. The reactant that gives the smaller amount of product is the limiting reagent.

Solution:

5 SO$_2$ molecules react with 2 O$_2$ molecules to produce molecules of SO$_3$. The balanced chemical equation is

\[ 2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g) \]

Amount (molecules) of SO$_3$ produced from the SO$_2$ = 5 molecules SO$_2$ \( \left( \frac{2 \text{molecules SO}_3}{2 \text{molecules SO}_2} \right) = 5 \text{ molecules SO}_3 \)

Amount (molecules) of SO$_3$ produced from the O$_2$ = 2 molecules SO$_2$ \( \left( \frac{2 \text{molecules SO}_3}{1 \text{molecule O}_2} \right) = 4 \text{ molecules SO}_3 \)

O$_2$ is the limiting reagent since it produces less SO$_3$ than the SO$_2$ does.

3.19A Plan: Use the molar ratios in the balanced equation to find the amount of AB$_2$ produced when 1.5 moles of each reactant is consumed. The smaller amount of product formed is the actual amount.

Solution:

Moles of AB$_2$ from AB = \( (1.5 \text{ mol AB}) \left( \frac{2 \text{ mol AB}_2}{2 \text{ mol AB}} \right) = 1.5 \text{ mol AB}_2 \)
Moles of AB₂ from B₂ = \( \left(1.5 \text{ mol B₂} \right) \left( \frac{2 \text{ mol AB₂}}{1 \text{ mol B₂}} \right) = 3.0 \text{ mol AB₂} \)

Thus AB is the limiting reagent and only 1.5 mol of AB₂ will form.

3.19B Plan: Use the molar ratios in the balanced equation to find the amount of SO₃ produced when 4.2 moles of SO₂ are consumed and, separately, the amount of SO₃ produced when 3.6 moles of O₂ are consumed. The smaller amount of product formed is the actual amount.

Solution:
The balanced chemical equation is:
\[ 2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g) \]

Amount (mol) of SO₃ produced from the SO₂ = 4.2 mol SO₂ \( \times \frac{2 \text{ mol SO₃}}{2 \text{ mol SO₂}} \) = 4.2 mol SO₃

Amount (mol) of SO₃ produced from the O₂ = 3.6 mol SO₂ \( \times \frac{2 \text{ mol SO₃}}{1 \text{ mol O₂}} \) = 7.2 mol SO₃

4.2 mol of SO₃ (the smaller amount) will be produced.

3.20A Plan: First, determine the formulas of the materials in the reaction and write a balanced chemical equation. Using the molar mass of each reactant, determine the moles of each reactant. Use molar ratios from the balanced equation to determine the moles of aluminum sulfide that may be produced from each reactant. The reactant that generates the smaller number of moles is limiting. Change the moles of aluminum sulfide from the limiting reactant to the grams of product using the molar mass of aluminum sulfide. To find the excess reactant amount, find the amount of excess reactant required to react with the limiting reagent and subtract that amount from the amount given in the problem.

Solution:
The balanced equation is: \( 2\text{Al} (s) + 3\text{S} (s) \rightarrow \text{Al}_2\text{S}_3 (s) \)

Determine the moles of product from each reactant:

Moles of Al₂S₃ from Al = \( (10.0 \text{ g Al}) \left( \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \right) \left( \frac{1 \text{ mol Al}_2\text{S}_3}{2 \text{ mol Al}} \right) = 0.18532 \text{ mol Al}_2\text{S}_3 \)

Moles of Al₂S₃ from S = \( (15.0 \text{ g S}) \left( \frac{1 \text{ mol S}}{32.06 \text{ g S}} \right) \left( \frac{1 \text{ mol Al}_2\text{S}_3}{3 \text{ mol S}} \right) = 0.155958 \text{ mol Al}_2\text{S}_3 \)

Sulfur produces less product so it is the limiting reactant.

Mass (g) of Al₂S₃ = \( (0.155958 \text{ mol Al}_2\text{S}_3) \left( \frac{150.14 \text{ g Al}_2\text{S}_3}{1 \text{ mol Al}_2\text{S}_3} \right) = 23.4155 = 23.4 \text{ g Al}_2\text{S}_3 \)

The mass of aluminum used in the reaction is now determined:

Mass (g) of Al = \( (15.0 \text{ g S}) \left( \frac{1 \text{ mol S}}{32.06 \text{ g S}} \right) \left( \frac{2 \text{ mol Al}}{3 \text{ mol Al}} \right) \left( \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \right) = 8.4155 \text{ g Al used} \)

Subtracting the mass of aluminum used from the initial aluminum gives the mass remaining.

Excess Al = Initial mass of Al – mass of Al reacted = 10.0 g – 8.4155 g = 1.5845 = 1.6 g Al

3.20B Plan: First, determine the formulas of the materials in the reaction and write a balanced chemical equation. Using the molar mass of each reactant, determine the moles of each reactant. Use molar ratios and the molar mass of carbon dioxide from the balanced equation to determine the mass of carbon dioxide that may be produced from each reactant. The reactant that generates the smaller mass of carbon dioxide is limiting. To find the excess reactant amount, find the amount of excess reactant required to react with the limiting reagent and subtract that amount from the amount given in the problem.

Solution:
The balanced equation is: \( 2\text{C}_4\text{H}_{10}(g) + 13\text{O}_2(g) \rightarrow 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g) \)

Determine the mass of product formed from each reactant:

Mass (g) of CO₂ from C₄H₁₀ = 4.65 g C₄H₁₀ \( \left( \frac{1 \text{ mol C₄H₁₀}}{58.12 \text{ g C₄H₁₀}} \right) \left( \frac{8 \text{ mol CO₂}}{2 \text{ mol C₄H₁₀}} \right) \left( \frac{44.01 \text{ g CO₂}}{1 \text{ mol CO₂}} \right) = 14.0844 = 14.1 \text{ g CO₂} \)

Mass (g) of CO₂ from O₂ = 10.0 g O₂ \( \left( \frac{1 \text{ mol O₂}}{32.00 \text{ g O₂}} \right) \left( \frac{8 \text{ mol CO₂}}{13 \text{ mol O₂}} \right) \left( \frac{44.01 \text{ g CO₂}}{1 \text{ mol CO₂}} \right) = 8.4635 = 8.46 \text{ g CO₂} \)
Oxygen produces the smallest amount of product, so it is the limiting reagent, and 8.46 g of CO₂ are produced.

The mass of butane used in the reaction is now determined:

\[
\text{Mass (g) of } C_4H_{10} = 10.0 \text{ g O}_2 \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2}\right) \left(\frac{2 \text{ mol } C_4H_{10}}{1 \text{ mol O}_2}\right) \left(\frac{58.12 \text{ g } C_4H_{10}}{1 \text{ mol } C_4H_{10}}\right) = 2.7942 = 2.79 \text{ g } C_4H_{10} \text{ used}
\]

Subtracting the mass of butane used from the initial butane gives the mass remaining.

Excess butane = Initial mass of butane – mass of butane reacted = 4.65 g – 2.79 g = 1.86 g butane

3.21A Plan: Determine the formulas, and then balance the chemical equation. The mass of marble is converted to moles, the molar ratio (from the balanced equation) gives the moles of CO₂, and finally the theoretical yield of CO₂ is determined from the moles of CO₂ and its molar mass. To calculate percent yield, divide the given actual yield of CO₂ by the theoretical yield, and multiply by 100.

Solution:
The balanced equation: \(CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)\)

Find the theoretical yield of carbon dioxide.

\[
\text{Mass (g) of } CO_2 = \left(10.0 \text{ g } CaCO_3 \right) \left(\frac{1 \text{ mol } CaCO_3}{100.09 \text{ g } CaCO_3}\right) \left(\frac{1 \text{ mol } CO_2}{1 \text{ mol } CaCO_3}\right) \left(\frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right) = 4.39704 \text{ g } CO_2
\]

The percent yield:

\[
\text{% yield of } CO_2 = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) (100\%) = \left(\frac{3.65 \text{ g } CO_2}{4.39704 \text{ g } CO_2}\right) (100\%) = 83.0104 = 83.0\%
\]

3.21B Plan: Determine the formulas, and then balance the chemical equation. The mass of sodium chloride is converted to moles, the molar ratio (from the balanced equation) gives the moles of sodium carbonate, and finally the theoretical yield of sodium carbonate is determined from the moles of sodium carbonate and its molar mass. To calculate percent yield, divide the given actual yield of sodium carbonate by the theoretical yield, and multiply by 100.

Solution:
The balanced equation: \(2NaCl + CaCO_3 \rightarrow CaCl_2 + Na_2CO_3\)

Find the theoretical yield of sodium carbonate.

\[
\text{Mass (g) of } Na_2CO_3 = \left(112 \text{ g NaCl}\right) \left(\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}}\right) \left(\frac{1 \text{ mol } Na_2CO_3}{2 \text{ mol NaCl}}\right) \left(\frac{105.99 \text{ g } Na_2CO_3}{1 \text{ mol } Na_2CO_3}\right) = 101.5647 = 102 \text{ g } Na_2CO_3
\]

The percent yield:

\[
\text{% yield of } Na_2CO_3 = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) (100\%) = \left(\frac{92.6 \text{ g } Na_2CO_3}{102 \text{ g } Na_2CO_3}\right) (100\%) = 90.7843 = 90.8\%
\]

END–OF–CHAPTER PROBLEMS

3.1 Plan: The atomic mass of an element expressed in amu is numerically the same as the mass of 1 mole of the element expressed in grams. We know the moles of each element and have to find the mass (in g). To convert moles of element to grams of element, multiply the number of moles by the molar mass of the element.

Solution:

\[
\text{Mass Al (g) = } \left(3 \text{ mol Al}\right) \left(\frac{26.98 \text{ g Al}}{1 \text{ mol Al}}\right) = 80.94 \text{ g Al}
\]

\[
\text{Mass Cl (g) = } \left(2 \text{ mol Cl}\right) \left(\frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}}\right) = 70.90 \text{ g Cl}
\]
3.2 **Plan:** The molecular formula of sucrose tells us that 1 mole of sucrose contains 12 moles of carbon atoms. Multiply the moles of sucrose by 12 to obtain moles of carbon atoms; multiply the moles of carbon atoms by Avogadro’s number to convert from moles to atoms.

**Solution:**

a) Moles of C atoms = \( \left( 1 \text{ mol } C_{12}H_{22}O_{11} \right) \left( \frac{12 \text{ mol C}}{1 \text{ mol } C_{12}H_{22}O_{11}} \right) = 12 \text{ mol C} \)

b) C atoms = \( \left( 2 \text{ mol } C_{12}H_{22}O_{11} \right) \left( \frac{12 \text{ mol C}}{1 \text{ mol } C_{12}H_{22}O_{11}} \right) \left( \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol C}} \right) = 1.445 \times 10^{25} \text{ C atoms} \)

3.3 **Plan:** Review the list of elements that exist as diatomic or polyatomic molecules.

**Solution:**

“1 mol of chlorine” could be interpreted as a mole of chlorine atoms or a mole of chlorine molecules, \( \text{Cl}_2 \). Specify which to avoid confusion. The same problem is possible with other diatomic or polyatomic molecules, e.g., \( \text{F}_2 \), \( \text{Br}_2 \), \( \text{I}_2 \), \( \text{H}_2 \), \( \text{O}_2 \), \( \text{N}_2 \), \( \text{S}_8 \), and \( \text{P}_4 \). For these elements, as for chlorine, it is not clear if atoms or molecules are being discussed.

3.4 The molecular mass is the sum of the atomic masses of the atoms or ions in a molecule. The molar mass is the mass of 1 mole of a chemical entity. Both will have the same numeric value for a given chemical substance but molecular mass will have the units of amu and molar mass will have the units of g/mol.

3.5 A mole of a particular substance represents a fixed number of chemical entities and has a fixed mass. Therefore the mole gives us an easy way to determine the number of particles (atoms, molecules, etc.) in a sample by weighing it. The mole maintains the same mass relationship between macroscopic samples as exist between individual chemical entities. It relates the number of chemical entities (atoms, molecules, ions, electrons) to the mass.

3.6 **Plan:** The mass of the compound is given. Divide the given mass by the molar mass of the compound to convert from mass of compound to number of moles of compound. The molecular formula of the compound tells us that 1 mole of compound contains 2 moles of phosphorus atoms. Use the ratio between P atoms and \( \text{P}_4 \) molecules (4:1) to convert moles of phosphorus atoms to moles of phosphorus molecules. Finally, multiply moles of \( \text{P}_4 \) molecules by Avogadro’s number to find the number of molecules.

**Solution:**

**Roadmap**

- Mass (g) of \( \text{Ca}_3(\text{PO}_4)_2 \)
- Divide by \( M \) (g/mol)
- Amount (mol) of \( \text{Ca}_3(\text{PO}_4)_2 \)
- Molar ratio between \( \text{Ca}_3(\text{PO}_4)_2 \) and P atoms
- Amount (moles) of P atoms
- Molar ratio between P atoms and \( \text{P}_4 \) molecules
- Amount (moles) of \( \text{P}_4 \) molecules
- Multiply by \( 6.022 \times 10^{23} \) formula units/mol
- Number of \( \text{P}_4 \) molecules
3.7 Plan: The relative atomic masses of each element can be found by counting the number of atoms of each element and comparing the overall masses of the two samples.

Solution:
a) The element on the left (green) has the higher molar mass because only 5 green balls are necessary to counterbalance the mass of 6 yellow balls. Since the green ball is heavier, its atomic mass is larger, and therefore its molar mass is larger.

b) The element on the left (red) has more atoms per gram. This figure requires more thought because the number of red and blue balls is unequal and their masses are unequal. If each pan contained 3 balls, then the red balls would be lighter. The presence of 6 red balls means that they are that much lighter. Because the red ball is lighter, more red atoms are required to make 1 g.

c) The element on the left (orange) has fewer atoms per gram. The orange balls are heavier, and it takes fewer orange balls to make 1 g.

d) Neither element has more atoms per mole. Both the left and right elements have the same number of atoms per mole. The number of atoms per mole (6.022x10²³) is constant and so is the same for every element.

3.8 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

Solution:
a) \[ M = (1 \times M\text{ of Sr}) + (2 \times M\text{ of O}) + (2 \times M\text{ of H}) \]
   \[ = (1 \times 87.62 \text{ g/mol Sr}) + (2 \times 16.00 \text{ g/mol O}) + (2 \times 1.008 \text{ g/mol H}) \]
   \[ = 121.64 \text{ g/mol of Sr(OH)_2} \]

b) \[ M = (2 \times M\text{ of N}) + (3 \times M\text{ of O}) \]
   \[ = (2 \times 14.01 \text{ g/mol N}) + (3 \times 16.00 \text{ g/mol O}) \]
   \[ = 76.02 \text{ g/mol of N}_2\text{O}_3 \]

c) \[ M = (1 \times M\text{ of Na}) + (1 \times M\text{ of Cl}) + (3 \times M\text{ of O}) \]
   \[ = (1 \times 22.99 \text{ g/mol Na}) + (1 \times 35.45 \text{ g/mol Cl}) + (3 \times 16.00 \text{ g/mol O}) \]
   \[ = 106.44 \text{ g/mol of NaClO}_3 \]

d) \[ M = (2 \times M\text{ of Cr}) + (3 \times M\text{ of O}) \]
   \[ = (2 \times 52.00 \text{ g/mol Cr}) + (3 \times 16.00 \text{ g/mol O}) \]
   \[ = 152.00 \text{ g/mol of Cr}_2\text{O}_3 \]

3.9 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

Solution:
a) \[ M = (3 \times M\text{ of N}) + (12 \times M\text{ of H}) + (1 \times M\text{ of P}) + (4 \times M\text{ of O}) \]
   \[ = (3 \times 14.01 \text{ g/mol N}) + (12 \times 1.008 \text{ g/mol H}) + (1 \times 30.97 \text{ g/mol P}) + (4 \times 16.00 \text{ g/mol O}) \]
   \[ = 149.10 \text{ g/mol of } (\text{NH}_4)\text{PO}_4 \]

b) \[ M = (1 \times M\text{ of C}) + (2 \times M\text{ of H}) + (2 \times M\text{ of Cl}) \]
   \[ = (1 \times 12.01 \text{ g/mol C}) + (2 \times 1.008 \text{ g/mol H}) + (2 \times 35.45 \text{ g/mol Cl}) \]
   \[ = 84.93 \text{ g/mol of CH}_2\text{Cl}_2 \]

c) \[ M = (1 \times M\text{ of Cu}) + (1 \times M\text{ of S}) + (9 \times M\text{ of O}) + (10 \times M\text{ of H}) \]
   \[ = (1 \times 63.55 \text{ g/mol Cu}) + (1 \times 32.06 \text{ g/mol S}) + (9 \times 16.00 \text{ g/mol O}) + (10 \times 1.008 \text{ g/mol H}) \]
   \[ = 249.69 \text{ g/mol of } \text{CuSO}_4\cdot5\text{H}_2\text{O} \]

d) \[ M = (1 \times M\text{ of Br}) + (3 \times M\text{ of F}) \]
   \[ = (1 \times 79.90 \text{ g/mol Br}) + (3 \times 19.00 \text{ g/mol F}) \]
   \[ = 136.90 \text{ g/mol of BrF}_3 \]

3.10 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of
the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole
of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

**Solution:**

a) \( M = (1 \times M_{\text{Sn}}) + (1 \times M_{\text{O}}) \)
\[= (1 \times 118.7 \text{ g/mol Sn}) + (1 \times 16.00 \text{ g/mol O}) \]
\[= 134.7 \text{ g/mol of SnO} \]

b) \( M = (1 \times M_{\text{Ba}}) + (2 \times M_{\text{F}}) \)
\[= (1 \times 137.3 \text{ g/mol Ba}) + (2 \times 19.00 \text{ g/mol F}) \]
\[= 175.3 \text{ g/mol of BaF}_2 \]

c) \( M = (2 \times M_{\text{Al}}) + (3 \times M_{\text{S}}) + (12 \times M_{\text{O}}) \)
\[= (2 \times 26.98 \text{ g/mol Al}) + (3 \times 32.06 \text{ g/mol S}) + (12 \times 16.00 \text{ g/mol O}) \]
\[= 342.14 \text{ g/mol of Al}_2(\text{SO}_4)_3 \]

d) \( M = (1 \times M_{\text{Mn}}) + (2 \times M_{\text{Cl}}) \)
\[= (1 \times 54.94 \text{ g/mol Mn}) + (2 \times 35.45 \text{ g/mol Cl}) \]
\[= 125.84 \text{ g/mol of MnCl}_2 \]

3.11 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of
the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole
of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

**Solution:**

a) \( M = (2 \times M_{\text{N}}) + (4 \times M_{\text{O}}) \)
\[= (2 \times 14.01 \text{ g/mol N}) + (4 \times 16.00 \text{ g/mol O}) \]
\[= 92.02 \text{ g/mol of N}_2\text{O}_4 \]

b) \( M = (4 \times M_{\text{C}}) + (10 \times M_{\text{H}}) + (1 \times M_{\text{O}}) \)
\[= (4 \times 12.01 \text{ g/mol C}) + (10 \times 1.008 \text{ g/mol H}) + (1 \times 16.00 \text{ g/mol O}) \]
\[= 74.12 \text{ g/mol of C}_4\text{H}_9\text{OH} \]

c) \( M = (1 \times M_{\text{Mg}}) + (1 \times M_{\text{S}}) + (11 \times M_{\text{O}}) + (14 \times M_{\text{H}}) \)
\[= (1 \times 24.31 \text{ g/mol Mg}) + (1 \times 32.06 \text{ g/mol S}) + (11 \times 16.00 \text{ g/mol O}) + (14 \times 1.008 \text{ g/mol H}) \]
\[= 246.48 \text{ g/mol of MgSO}_4\cdot7\text{H}_2\text{O} \]

d) \( M = (1 \times M_{\text{Ca}}) + (4 \times M_{\text{C}}) + (6 \times M_{\text{H}}) + (4 \times M_{\text{O}}) \)
\[= (1 \times 40.08 \text{ g/mol Ca}) + (4 \times 12.01 \text{ g/mol C}) + (6 \times 1.008 \text{ g/mol H}) + (4 \times 16.00 \text{ g/mol O}) \]
\[= 158.17 \text{ g/mol of Ca(C}_2\text{H}_3\text{O}_2)_2 \]

3.12 Plan: Determine the molar mass of each substance; then perform the appropriate molar conversions.

To find the mass in part a), multiply the number of moles by the molar mass of Zn. In part b), first multiply by
Avogadro’s number to obtain the number of F2 molecules. The molecular formula tells us that there are 2 F atoms
in each molecule of F2; use the 2:1 ratio to convert F2 molecules to F atoms. In part c), convert mass of Ca to
moles of Ca by dividing by the molar mass of Ca. Then multiply by Avogadro’s number to obtain the number of
Ca atoms.

**Solution:**

a) \( (0.346 \text{ mol Zn}) \left( \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} \right) = 22.6 \text{ g Zn} \)

b) \( (2.62 \text{ mol F}_2) \left( \frac{6.022 \times 10^{23} \text{ F}_2 \text{ molecules}}{1 \text{ mol F}_2} \right) \left( \frac{2 \text{ atoms}}{1 \text{ F}_2 \text{ molecule}} \right) = 3.16 \times 10^{24} \text{ F atoms} \)

c) \( 28.5 \text{ g Ca} \left( \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) \left( \frac{6.022 \times 10^{23} \text{ Ca atoms}}{1 \text{ mol Ca}} \right) = 4.28 \times 10^{23} \text{ Ca atoms} \)

3.13 Plan: Determine the molar mass of each substance; then perform the appropriate molar conversions. In part a),
convert mg units to g units by dividing by 10^3; then convert mass of Mn to moles of Mn by dividing by the molar
mass of Mn. In part b) convert number of Cu atoms to moles of Cu by dividing by Avogadro’s number. In part c)
divide by Avogadro’s number to convert number of Li atoms to moles of Li; then multiply by the molar mass of
Li to find the mass.

**Solution:**
a) \((62.0 \text{ mg Mn}) \left( \frac{1 \text{ g Mn}}{10^3 \text{ mg Mn}} \right) \left( \frac{1 \text{ mol Mn}}{54.94 \text{ g Mn}} \right) = 1.13 \times 10^{-3} \text{ mol Mn} \)

b) \((1.36 \times 10^{22} \text{ Cu atoms}) \left( \frac{1 \text{ mol Cu}}{6.022 \times 10^{23} \text{ Cu atoms}} \right) = 0.0226 \text{ mol Cu} \)

c) \((8.05 \times 10^{24} \text{ Li atoms}) \left( \frac{1 \text{ mol Li}}{6.022 \times 10^{23} \text{ Li atoms}} \right) \left( \frac{6.941 \text{ g Li}}{1 \text{ mol Li}} \right) = 92.8 \text{ g Li} \)

### 3.14 Plan:
Determine the molar mass of each substance; then perform the appropriate molar conversions.

To find the mass in part a), multiply the number of moles by the molar mass of the substance. In part b), first convert mass of compound to moles of compound by dividing by the molar mass of the compound. The molecular formula of the compound tells us that 1 mole of compound contains 6 moles of oxygen atoms; use the 1:6 ratio to convert moles of compound to moles of oxygen atoms. In part c), convert mass of compound to moles of compound by dividing by the molar mass of the compound. Since 1 mole of compound contains 6 moles of oxygen atoms, multiply the moles of compound by 6 to obtain moles of oxygen atoms; then multiply by Avogadro’s number to obtain the number of oxygen atoms.

### Solution:

a) \(\mu\) of KMnO₄ = \((1 \times \mu\) of K) + \((1 \times \mu\) of Mn) + \((4 \times \mu\) of O)
\[ = (1 \times 39.10 \text{ g/mol K}) + (1 \times 54.94 \text{ g/mol Mn}) + (4 \times 16.00 \text{ g/mol O}) = 158.04 \text{ g/mol of KMnO}_4 \]

Mass of KMnO₄ = \(\left( \frac{0.68 \text{ mol KMnO}_4}{1 \text{ mol KMnO}_4} \right) \left( \frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} \right) = 107.467 = 1.1 \times 10^2 \text{ g KMnO}_4 \)

b) \(\mu\) of Ba(NO₃)₂ = \((1 \times \mu\) of Ba) + \((2 \times \mu\) of N) + \((6 \times \mu\) of O)
\[ = (1 \times 137.3 \text{ g/mol Ba}) + (2 \times 14.01 \text{ g/mol N}) + (6 \times 16.00 \text{ g/mol O}) = 261.3 \text{ g/mol Ba(NO}_3)_2 \]

Moles of Ba(NO₃)₂ = \(\left( \frac{8.18 \text{ g Ba(NO}_3)_2}{261.3 \text{ g Ba(NO}_3)_2} \right) = 0.031305 \text{ mol Ba(NO}_3)_2 \)

Moles of O atoms = \(\left( \frac{6 \text{ mol O atoms}}{1 \text{ mol Ba(NO}_3)_2} \right) = 0.18783 = 0.188 \text{ mol O atoms} \)

c) \(\mu\) of CaSO₄•2H₂O = \((1 \times \mu\) of Ca) + \((1 \times \mu\) of S) + \((6 \times \mu\) of O) + \((4 \times \mu\) of H)
\[ = (1 \times 40.08 \text{ g/mol Ca}) + (1 \times 32.06 \text{ g/mol S}) + (6 \times 16.00 \text{ g/mol O}) + (4 \times 1.008 \text{ g/mol H}) = 172.17 \text{ g/mol} \)

(Note that the waters of hydration are included in the molar mass.)

Moles of CaSO₄•2H₂O = \(\left( \frac{7.3 \times 10^{-3} \text{ g CaSO}_4 \cdot 2\text{H}_2\text{O}}{172.17 \text{ g CaSO}_4 \cdot 2\text{H}_2\text{O}} \right) \left( \frac{1 \text{ mol CaSO}_4 \cdot 2\text{H}_2\text{O}}{172.17 \text{ g CaSO}_4 \cdot 2\text{H}_2\text{O}} \right) = 4.239995 \times 10^{-5} \text{ mol} \)

Moles of O atoms = \(\left( \frac{4.239995 \times 10^{-5} \text{ mol CaSO}_4 \cdot 2\text{H}_2\text{O}}{1 \text{ mol CaSO}_4 \cdot 2\text{H}_2\text{O}} \right) \left( \frac{6 \text{ mol O atoms}}{1 \text{ mol CaSO}_4 \cdot 2\text{H}_2\text{O}} \right) = 2.543997 \times 10^{-5} \text{ mol O atoms} \)

Number of O atoms = \(\left( \frac{2.543997 \times 10^{-5} \text{ mol O atoms}}{1 \text{ mol O atoms}} \right) \left( \frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O atoms}} \right) = 1.5320 \times 10^{20} = 1.5 \times 10^{20} \text{ O atoms} \)

### 3.15 Plan:
Determine the molar mass of each substance, then perform the appropriate molar conversions.

To find the mass in part a), divide the number of molecules by Avogadro’s number to find moles of compound and then multiply the mole amount by the molar mass in grams; convert from mass in g to mass in kg. In part b), first convert mass of compound to moles of compound by dividing by the molar mass of the compound. The molecular formula of the compound tells us that 1 mole of compound contains 2 moles of chlorine atoms; use the 1:2 ratio to convert moles of compound to moles of chlorine atoms. In part c), convert mass of compound to moles of compound by dividing by the molar mass of the compound. Since 1 mole of compound contains 2 moles of H⁻ ions, multiply the moles of compound by 2 to obtain moles of H⁻ ions; then multiply by Avogadro’s number to obtain the number of H⁻ ions.

### Solution:

a) \((62.0 \text{ mg Mn}) \left( \frac{1 \text{ g Mn}}{10^3 \text{ mg Mn}} \right) \left( \frac{1 \text{ mol Mn}}{54.94 \text{ g Mn}} \right) = 1.13 \times 10^{-3} \text{ mol Mn} \)

b) \((1.36 \times 10^{22} \text{ Cu atoms}) \left( \frac{1 \text{ mol Cu}}{6.022 \times 10^{23} \text{ Cu atoms}} \right) = 0.0226 \text{ mol Cu} \)

c) \((8.05 \times 10^{24} \text{ Li atoms}) \left( \frac{1 \text{ mol Li}}{6.022 \times 10^{23} \text{ Li atoms}} \right) \left( \frac{6.941 \text{ g Li}}{1 \text{ mol Li}} \right) = 92.8 \text{ g Li} \)
a) \( M \) of \( \text{NO}_2 = (1 \times M \text{ of N}) + (2 \times M \text{ of O}) \\
= (1 \times 14.01 \text{ g/mol N}) + (2 \times 16.00 \text{ g/mol O}) = 46.01 \text{g/mol of NO}_2 \\
Moles \text{ of } \text{NO}_2 = \left( \frac{1 \text{ mol NO}_2}{6.022 \times 10^{23} \text{ molecules NO}_2} \right) \times 7.63866 \times 10^{-3} \text{ mol NO}_2 \\
Mass (\text{kg}) \text{ of } \text{NO}_2 = \left( \frac{3.51455 \times 10^{-4} \text{ mol NO}_2}{1 \text{ mol NO}_2} \right) \times 1 \text{ kg/} (\frac{151.00 \text{ g NO}_2}{1 \text{ mol NO}_2}) \\
= 3.51455 \times 10^{-4} \text{ kg NO}_2 \\
b) \( M \) of \( \text{C}_2\text{H}_4\text{Cl}_2 = (2 \times M \text{ of C}) + (4 \times M \text{ of H}) + (2 \times M \text{ of Cl}) \\
= (2 \times 12.01 \text{g/mol C}) + (4 \times 1.008 \text{ g/mol H}) + (2 \times 35.45 \text{ g/mol Cl}) = 98.95 \text{g/mol of } \text{C}_2\text{H}_4\text{Cl}_2 \\
Moles \text{ of } \text{C}_2\text{H}_4\text{Cl}_2 = \left( \frac{1 \text{ mol } \text{C}_2\text{H}_4\text{Cl}_2}{98.95 \text{ g } \text{C}_2\text{H}_4\text{Cl}_2} \right) \times 6.21526 \times 10^{-4} \text{ mol } \text{C}_2\text{H}_4\text{Cl}_2 \\
Moles \text{ of Cl atoms} = \left( \frac{1.2431 \times 10^{-3} \text{ mol Cl atoms}}{1 \text{ mol } \text{C}_2\text{H}_4\text{Cl}_2} \right) \times 2.0615 \times 10^{-4} \text{ mol } \text{Cl atoms} \\
= 1.24 \times 10^{-3} \text{ mol Cl atoms} \\
c) \( M \) of \( \text{SrH}_2 = (1 \times M \text{ of Sr}) + (2 \times M \text{ of H}) = (1 \times 87.62 \text{ g/mol Sr}) + (2 \times 1.008 \text{ g/mol H}) = 89.64 \text{g/mol of } \text{SrH}_2 \\
Moles \text{ of } \text{SrH}_2 = \left( \frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) \times 8.24 \times 10^{-5} \text{ mol SrH}_2 \\
Moles \text{ of H}^- \text{ ions} = \left( \frac{0.1298528 \text{ mol H}^- \text{ ions}}{1 \text{ mol } \text{SrH}_2} \right) \times 2.0615 \times 10^{-4} \text{ mol } \text{H}^- \text{ ions} \\
= 0.1298528 \text{ mol H}^- \text{ ions} \\
Number \text{ of H}^- \text{ ions} = \left( \frac{7.82 \times 10^{22} \text{ H}^- \text{ ions}}{1 \text{ mol } \text{H}^-} \right) \times 0.1298528 \text{ mol H}^- \text{ ions} \\
= 7.82 \times 10^{22} \text{ H}^- \text{ ions} 

3.16 Plan: Determine the molar mass of each substance; then perform the appropriate molar conversions. To find the mass in part a), multiply the number of moles by the molar mass of the substance. In part b), first convert the mass of compound in kg to mass in g and divide by the molar mass of the compound to find moles of compound. In part c), convert mass of compound in mg to mass in g and divide by the molar mass of the compound to find moles of compound. Since 1 mole of compound contains 2 moles of nitrogen atoms, multiply the moles of compound by 2 to obtain moles of nitrogen atoms; then multiply by Avogadro’s number to obtain the number of nitrogen atoms.

Solution:
a) \( M \) of \( \text{MnSO}_4 = (1 \times M \text{ of Mn}) + (1 \times M \text{ of S}) + (4 \times M \text{ of O}) \\
= (1 \times 54.94 \text{ g/mol Mn}) + (1 \times 32.06 \text{ g/mol S}) + (4 \times 16.00 \text{ g/mol O}) = 151.00 \text{g/mol of MnSO}_4 \\
Mass (\text{g}) \text{ of } \text{MnSO}_4 = \left( \frac{151.00 \text{ g MnSO}_4}{1 \text{ mol MnSO}_4} \right) \times 9.7244 \text{ g MnSO}_4 \approx 9.72 \text{ g MnSO}_4 \\
b) \( M \) of \( \text{Fe(ClO}_4)_3 = (1 \times M \text{ of Fe}) + (3 \times M \text{ of Cl}) + (12 \times M \text{ of O}) \\
= (1 \times 55.85 \text{ g/mol Fe}) + (3 \times 35.45 \text{ g/mol Cl}) + (12 \times 16.00 \text{ g/mol O}) = 354.20 \text{g/mol of Fe(ClO}_4)_3 \\
Mass (\text{g}) \text{ of } \text{Fe(ClO}_4)_3 = \left( \frac{354.20 \text{ g Fe(ClO}_4)_3}{1 \text{ mol Fe(ClO}_4)_3} \right) \times 44.6076 \text{ g Fe(ClO}_4)_3 \approx 44.6 \text{ mol Fe(ClO}_4)_3 \\
c) \( M \) of \( \text{NH}_4\text{NO}_3 = (2 \times M \text{ of N}) + (4 \times M \text{ of H}) + (2 \times M \text{ of O}) \\
= (2 \times 14.01 \text{ g/mol N}) + (4 \times 1.008 \text{ g/mol H}) + (2 \times 16.00 \text{ g/mol O}) = 64.05 \text{g/mol NH}_4\text{NO}_3 \\
Mass (\text{g}) \text{ of } \text{NH}_4\text{NO}_3 = \left( \frac{64.05 \text{ g NH}_4\text{NO}_3}{1 \text{ mol NH}_4\text{NO}_3} \right) \times 0.0926 \text{ g NH}_4\text{NO}_3 \approx 0.0926 \text{ g NH}_4\text{NO}_3 

3-20
Moles of NH₄NO₂ = (0.0926 g NH₄NO₂) \( \frac{1 \text{ mol NH₄NO₂}}{64.05 \text{ g NH₄NO₂}} \) = 1.44575 \times 10^{-3} \text{ mol NH₄NO₂}

Moles of N atoms = \( \frac{1.44575 \times 10^{-3} \text{ mol NH₄NO₂} \times 2 \text{ mol N}}{1 \text{ mol NH₄NO₂}} \) = 2.8915 \times 10^{-3} \text{ mol N}

Number of N atoms = \( \frac{2.8915 \times 10^{-3} \text{ mol N} \times 6.022 \times 10^{23} \text{ N atoms}}{1 \text{ mol N}} \) = 1.74126 \times 10^{21} = 1.74 \times 10^{21} \text{ N atoms}

3.17 Plan: Determine the molar mass of each substance; then perform the appropriate molar conversions. In part a), divide the mass by the molar mass of the compound to find moles of compound. Since 1 mole of compound contains 3 moles of ions (1 mole of Sr²⁺ and 2 moles of F⁻), multiply the moles of compound by 3 to obtain moles of ions and then multiply by Avogadro’s number to obtain the number of ions. In part b), multiply the number of moles by the molar mass of the substance to find the mass in g and then convert to kg. In part c), divide the number of formula units by Avogadro’s number to find moles; multiply the number of moles by the molar mass to obtain the mass in g and then convert to mg.

Solution:

a) \( \text{M} \) of SrF₂ = (1 x \( \text{M} \) of Sr) + (2 x \( \text{M} \) of F) = (1 x 87.62 g/mol Sr) + (2 x 19.00 g/mol F) = 125.62 g/mol of SrF₂

Moles of SrF₂ = \( \frac{1 \text{ mol SrF₂}}{125.62 \text{ g SrF₂}} \) = 0.303296 mol SrF₂

Moles of ions = \( \frac{3 \text{ mol ions}}{1 \text{ mol SrF₂}} \) = 0.909888 mol ions

Number of ions = \( \frac{6.022 \times 10^{23} \text{ ions}}{1 \text{ mol ions}} \) = 5.47935 \times 10^{23} = 5.48 \times 10^{23} \text{ ions}

b) \( \text{M} \) of CuCl₂•2H₂O = (1 x \( \text{M} \) of Cu) + (2 x \( \text{M} \) of Cl) + (4 x \( \text{M} \) of H) + (2 x \( \text{M} \) of O) = (1 x 63.55 g/mol Cu) + (2 x 35.45 g/mol Cl) + (4 x 1.008 g/mol H) + (2 x 16.00 g/mol O) = 170.48 g/mol of CuCl₂•2H₂O

(Note that the waters of hydration are included in the molar mass.)

Mass (g) of CuCl₂•2H₂O = \( \frac{170.48 \text{ g CuCl₂•2H₂O}}{1 \text{ mol CuCl₂•2H₂O}} \) = 610.32 g CuCl₂•2H₂O

Mass (kg) of CuCl₂•2H₂O = 610.32 \text{ g CuCl₂•2H₂O} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 0.610 \text{ kg CuCl₂•2H₂O}

c) \( \text{M} \) of Bi(NO₃)₃•5H₂O = (1 x \( \text{M} \) of Bi) + (3 x \( \text{M} \) of N) + (10 x \( \text{M} \) of H) + (14 x \( \text{M} \) of O) = (1 x 209.0 g/mol Bi) + (3 x 14.01 g/mol N) + (10 x 1.008 g/mol H) + (14 x 16.00 g/mol O) = 485.11 g/mol of Bi(NO₃)₃•5H₂O

(Note that the waters of hydration are included in the molar mass.)

Moles of Bi(NO₃)₃•5H₂O = \( \frac{1 \text{ mol Bi(NO₃)₃•5H₂O}}{6.022 \times 10^{23} \text{ FU}} \) = 0.047825 mol Bi(NO₃)₃•5H₂O

Mass (g) of Bi(NO₃)₃•5H₂O = 0.047825 \text{ mol Bi(NO₃)₃•5H₂O} \times \frac{485.1 \text{ g Bi(NO₃)₃•5H₂O}}{1 \text{ mol Bi(NO₃)₃•5H₂O}} = 23.1999 g

Mass (mg) of Bi(NO₃)₃•5H₂O = 23.1999 \text{ g Bi(NO₃)₃•5H₂O} \times \frac{1 \text{ mg}}{10^{-3} \text{ g}} = 23199.9 = 2.32 \times 10^4 \text{ mg Bi(NO₃)₃•5H₂O}

3.18 Plan: The formula of each compound must be determined from its name. The molar mass for each formula.
comes from the formula and atomic masses from the periodic table. Determine the molar mass of each substance, then perform the appropriate molar conversions. In part a), multiply the moles by the molar mass of the compound to find the mass of the sample. In part b), divide the number of molecules by Avogadro’s number to find moles; multiply the number of moles by the molar mass to obtain the mass. In part c), divide the mass by the molar mass to find moles of compound and multiply moles by Avogadro’s number to find the number of formula units. In part d), use the fact that each formula unit contains 1 Na ion, 1 perchlorate ion, 1 Cl atom, and 4 O atoms.

Solution:

a) Carbonate is a polyatomic anion with the formula, \( \text{CO}_3^{2-} \). Copper(I) indicates \( \text{Cu}^+ \). The correct formula for this ionic compound is \( \text{Cu}_2\text{CO}_3 \).

\[
\mathcal{M} \text{ of Cu}_2\text{CO}_3 = (2 \times \mathcal{M} \text{ of Cu}) + (1 \times \mathcal{M} \text{ of C}) + (3 \times \mathcal{M} \text{ of O})
\]
\[
= (2 \times 63.55 \text{ g/mol Cu}) + (1 \times 12.01 \text{ g/mol C}) + (3 \times 16.00 \text{ g/mol O}) = 187.11 \text{ g/mol of Cu}_2\text{CO}_3
\]

Mass \((\text{g})\) of \( \text{Cu}_2\text{CO}_3 \) = \[
8.35 \text{ mol Cu}_2\text{CO}_3 \left( \frac{187.11 \text{ g Cu}_2\text{CO}_3}{1 \text{ mol Cu}_2\text{CO}_3} \right) = 1562.4 = 1.56 \times 10^3 \text{ g Cu}_2\text{CO}_3
\]

b) Dinitrogen pentaoxide has the formula \( \text{N}_2\text{O}_5 \). Di- indicates 2 N atoms and penta- indicates 5 O atoms.

\[
\mathcal{M} \text{ of N}_2\text{O}_5 = (2 \times \mathcal{M} \text{ of N}) + (5 \times \mathcal{M} \text{ of O})
\]
\[
= (2 \times 14.01 \text{ g/mol N}) + (5 \times 16.00 \text{ g/mol O}) = 108.02 \text{ g/mol of N}_2\text{O}_5
\]

Moles of \( \text{N}_2\text{O}_5 \) = \[
20 \times 25 \left( \frac{1 \text{ mol N}_2\text{O}_5}{6.022 \times 10^{23} \text{ N}_2\text{O}_5 \text{ molecules}} \right) = 6.7087 \times 10^{-4} \text{ mol N}_2\text{O}_5
\]

Mass \((\text{g})\) of \( \text{N}_2\text{O}_5 \) = \[
108.02 \text{ g N}_2\text{O}_5 \left( \frac{1 \text{ mol N}_2\text{O}_5}{6.022 \times 10^{23} \text{ N}_2\text{O}_5 \text{ molecules}} \right) = 0.072467 = 0.0725 \text{ g N}_2\text{O}_5
\]

c) The correct formula for this ionic compound is \( \text{NaClO}_4 \); Na has a charge of +1 (Group 1 ion) and the perchlorate ion is \( \text{ClO}_4^- \).

\[
\mathcal{M} \text{ of NaClO}_4 = (1 \times \mathcal{M} \text{ of Na}) + (1 \times \mathcal{M} \text{ of Cl}) + (4 \times \mathcal{M} \text{ of O})
\]
\[
= (1 \times 22.99 \text{ g/mol Na}) + (1 \times 35.45 \text{ g/mol Cl}) + (4 \times 16.00 \text{ g/mol O}) = 122.44 \text{ g/mol of NaClO}_4
\]

Moles of \( \text{NaClO}_4 \) = \[
78.9 \text{ g NaClO}_4 \left( \frac{1 \text{ mol NaClO}_4}{122.44 \text{ g NaClO}_4} \right) = 0.644397 = 0.644 \text{ mol NaClO}_4
\]

\( \text{FU} \) = formula units

\[
\text{FU of NaClO}_4 = \left( \frac{0.644397 \text{ mol NaClO}_4}{1 \text{ mol NaClO}_4} \right) \times \left( \frac{6.022 \times 10^{23} \text{ FU NaClO}_4}{1 \text{ mol NaClO}_4} \right) = 3.88 \times 10^{23} \text{ FU NaClO}_4
\]

d) Number of \( \text{Na}^+ \) ions = \[
3.88 \times 10^{23} \text{ FU NaClO}_4 \left( \frac{1 \text{ Na}^+ \text{ ion}}{1 \text{ FU NaClO}_4} \right) = 3.88 \times 10^{23} \text{ Na}^+ \text{ ions}
\]

Number of \( \text{ClO}_4^- \) ions = \[
3.88 \times 10^{23} \text{ FU NaClO}_4 \left( \frac{1 \text{ ClO}_4^- \text{ ion}}{1 \text{ FU NaClO}_4} \right) = 3.88 \times 10^{23} \text{ ClO}_4^- \text{ ions}
\]

Number of Cl atoms = \[
3.88 \times 10^{23} \text{ FU NaClO}_4 \left( \frac{4 \text{ Cl atoms}}{1 \text{ FU NaClO}_4} \right) = 3.88 \times 10^{23} \text{ Cl atoms}
\]

Number of O atoms = \[
3.88 \times 10^{23} \text{ FU NaClO}_4 \left( \frac{1 \text{ O atom}}{1 \text{ FU NaClO}_4} \right) = 1.55 \times 10^{24} \text{ O atoms}
\]
Solution:
a) Sulfate is a polyatomic anion with the formula, $SO_4^{2-}$. Chromium(III) indicates $Cr^{3+}$. Decahydrate indicates 10 water molecules (“waters of hydration”). The correct formula for this ionic compound is $Cr_2(SO_4)_3\cdot10H_2O$.

$$M \text{ of } Cr_2(SO_4)_3\cdot10H_2O = (2 \times M \text{ of } Cr) + (3 \times M \text{ of } S) + (22 \times M \text{ of } O) + (20 \times M \text{ of } H)$$

$$= (2 \times 52.00 \text{ g/mol } Cr) + (3 \times 32.06 \text{ g/mol } S) + (22 \times 16.00 \text{ g/mol } O) + (20 \times 1.008 \text{ g/mol } H)$$

$$= 572.34 \text{ g/mol of } Cr_2(SO_4)_3\cdot10H_2O$$

Mass (g) of $Cr_2(SO_4)_3\cdot10H_2O = \left(\frac{8.42 \text{ mol } Cr_2(SO_4)_3\cdot10H_2O}{572.34 \text{ g/mol }}\right) \times 8.42 \text{ mol } Cr_2(SO_4)_3\cdot10H_2O = 4819.103 = 4.82 \times 10^3 \text{ g } Cr_2(SO_4)_3\cdot10H_2O$

b) Dichlorine heptaoxide has the formula $Cl_2O_7$. Di- indicates 2 Cl atoms and hepta- indicates 7 O atoms.

$$M \text{ of } Cl_2O_7 = (2 \times M \text{ of } Cl) + (7 \times M \text{ of } O)$$

$$= (2 \times 35.45 \text{ g/mol } Cl) + (7 \times 16.00 \text{ g/mol } O) = 182.9 \text{ g/mol of } Cl_2O_7$$

Moles of $Cl_2O_7 = \left(\frac{1.83 \times 10^{24} \text{ molecules } Cl_2O_7}{1 \text{ mol } Cl_2O_7}\right) = 3.038858 \text{ mol } Cl_2O_7$

Mass (g) of $Cl_2O_7 = \left(\frac{1 \text{ mol } Cl_2O_7}{6.022 \times 10^{23} \text{ molecules } Cl_2O_7}\right) \times 1 \text{ mol } Cl_2O_7 = 555.807 = 5.56 \times 10^2 \text{ g } Cl_2O_7$

c) The correct formula for this ionic compound is $Li_2SO_4$; Li has a charge of +1 (Group 1 ion) and the sulfate ion is $SO_4^{2-}$.

$$M \text{ of } Li_2SO_4 = (2 \times M \text{ of } Li) + (1 \times M \text{ of } S) + (4 \times M \text{ of } O)$$

$$= (2 \times 6.941 \text{ g/mol } Li) + (1 \times 32.06 \text{ g/mol } S) + (4 \times 16.00 \text{ g/mol } O) = 109.94 \text{ g/mol of } Li_2SO_4$$

Moles of $Li_2SO_4 = \left(\frac{6.2 \text{ g } Li_2SO_4}{109.94 \text{ g } Li_2SO_4}\right) = 0.056394 = 0.056 \text{ mol } Li_2SO_4$

$$F.U. \text{ of } Li_2SO_4 = \left(\frac{6.022 \times 10^{23} \text{ FU } Li_2SO_4}{1 \text{ mol } Li_2SO_4}\right) = 3.3960 \times 10^{22} \text{ FU } Li_2SO_4$$

d) Number of $Li^+$ ions = $\left(\frac{2 \text{ Li}^+ \text{ ions}}{1 \text{ F.U. } Li_2SO_4}\right) = 6.7920 \times 10^{22} = 6.8 \times 10^{22} \text{ Li}^+ \text{ ions}$

Number of $SO_4^{2-}$ ions = $\left(\frac{1 \text{ SO}_4^{2-} \text{ ion}}{1 \text{ F.U. } Li_2SO_4}\right) = 3.3960 \times 10^{22} = 3.4 \times 10^{22} \text{ SO}_4^{2-} \text{ ions}$

Number of $S$ atoms = $\left(\frac{1 \text{ S atom}}{1 \text{ F.U. } Li_2SO_4}\right) = 3.3960 \times 10^{22} = 3.4 \times 10^{22} \text{ S atoms}$

Number of $O$ atoms = $\left(\frac{4 \text{ O atoms}}{1 \text{ F.U. } Li_2SO_4}\right) = 1.3584 \times 10^{23} = 1.4 \times 10^{23} \text{ O atoms}$

3.20 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass percent = \(\frac{\text{total mass of element}}{\text{molar mass of compound}}\) \(\times 100\%\).

Solution:
a) Ammonium bicarbonate is an ionic compound consisting of ammonium ions, $NH_4^+$ and bicarbonate ions, $HCO_3^-$. The formula of the compound is $NH_4HCO_3$.

$$M \text{ of } NH_4HCO_3 = (1 \times M \text{ of } N) + (5 \times M \text{ of } H) + (1 \times M \text{ of } C) + (3 \times M \text{ of } O)$$

$$= (1 \times 14.01 \text{ g/mol } N) + (5 \times 1.008 \text{ g/mol } H) + (1 \times 12.01 \text{ g/mol } C) + (3 \times 16.00 \text{ g/mol } O)$$

$$= 79.06 \text{ g/mol of } NH_4HCO_3$$

There are 5 moles of H in 1 mole of $NH_4HCO_3$. 3-23
Mass (g) of H = \left( 5 \text{ mol H} \right) \left( \frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) = 5.040 \text{ g H}

Mass percent = \frac{\text{total mass H}}{\text{molar mass of compound}} \times 100 = \frac{5.040 \text{ g H}}{79.06 \text{ g NH}_4\text{HCO}_3} \times 100 = 6.374905 \approx 6.375\% \text{ H}

b) Sodium dihydrogen phosphate heptahydrate is a salt that consists of sodium ions, Na\(^+\), dihydrogen phosphate ions, H\(_2\)PO\(_4\)\(^-\), and seven waters of hydration. The formula is NaH\(_2\)PO\(_4\)•7H\(_2\)O. Note that the waters of hydration are included in the molar mass.

\[ \mathcal{M} \text{ of NaH}_2\text{PO}_4\cdot7\text{H}_2\text{O} = (1 \times \mathcal{M} \text{ of Na}) + (16 \times \mathcal{M} \text{ of H}) + (1 \times \mathcal{M} \text{ of P}) + (11 \times \mathcal{M} \text{ of O}) \]
\[ = (1 \times 22.99 \text{ g/mol Na}) + (16 \times 1.008 \text{ g/mol H}) + (1 \times 30.97 \text{ g/mol P}) + (11 \times 16.00 \text{ g/mol O}) \]
\[ = 246.09 \text{ g/mol NaH}_2\text{PO}_4\cdot7\text{H}_2\text{O} \]

There are 11 moles of O in 1 mole of NaH\(_2\)PO\(_4\)•7H\(_2\)O.

Mass (g) of O = \left( 11 \text{ mol O} \right) \left( \frac{16.00 \text{ g O}}{1 \text{ mol O}} \right) = 176.00 \text{ g O}

Mass percent = \frac{\text{total mass O}}{\text{molar mass of compound}} \times 100 = \frac{176.00 \text{ g O}}{246.09 \text{ g NaH}_2\text{PO}_4\cdot7\text{H}_2\text{O}} \times 100 = 71.52\% \text{ O}

3.21 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass percent = \frac{\text{total mass of element}}{\text{molar mass of compound}} \times 100.

Solution:

a) Strontium periodate is an ionic compound consisting of strontium ions, Sr\(^{2+}\) and periodate ions, IO\(_4\)\(^-\).

The formula of the compound is Sr(IO\(_4\))\(_2\).

\[ \mathcal{M} \text{ of Sr(IO}_4\text{)}\(_2\) = (1 \times \mathcal{M} \text{ of Sr}) + (2 \times \mathcal{M} \text{ of I}) + (8 \times \mathcal{M} \text{ of O}) \]
\[ = (1 \times 87.62 \text{ g/mol Sr}) + (2 \times 126.9 \text{ g/mol I}) + (8 \times 16.00 \text{ g/mol O}) \]
\[ = 469.4 \text{ g/mol of Sr(IO}_4\text{)}\(_2\) \]

There are 2 moles of I in 1 mole of Sr(IO\(_4\))\(_2\).

Mass (g) of I = \left( 2 \text{ mol I} \right) \left( \frac{126.9 \text{ g I}}{1 \text{ mol I}} \right) = 253.8 \text{ g I}

Mass percent = \frac{\text{total mass I}}{\text{molar mass of compound}} \times 100 = \frac{253.8 \text{ g I}}{469.4 \text{ g Sr(IO}_4\text{)}\(_2\)} \times 100 = 54.07\% \text{ I}

b) Potassium permanganate is an ionic compound consisting of potassium ions, K\(^+\) and permanganate ions, MnO\(_4\)\(^-\).

The formula of the compound is KMnO\(_4\).

\[ \mathcal{M} \text{ of KMnO}_4 = (1 \times \mathcal{M} \text{ of K}) + (1 \times \mathcal{M} \text{ of Mn}) + (4 \times \mathcal{M} \text{ of O}) \]
\[ = (1 \times 39.10 \text{ g/mol K}) + (1 \times 54.94 \text{ g/mol Mn}) + (4 \times 16.00 \text{ g/mol O}) \]
\[ = 158.04 \text{ g/mol of KMnO}_4 \]

There is 1 mole of Mn in 1 mole of KMnO\(_4\).

Mass (g) of Mn = \left( 1 \text{ mol Mn} \right) \left( \frac{54.94 \text{ g Mn}}{1 \text{ mol Mn}} \right) = 54.94 \text{ g Mn}

Mass percent = \frac{\text{total mass Mn}}{\text{molar mass of compound}} \times 100 = \frac{54.94 \text{ g Mn}}{158.04 \text{ g KMnO}_4} \times 100 = 34.76\% \text{ Mn}

3.22 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass fraction = \frac{\text{total mass of element}}{\text{molar mass of compound}}.

Solution: 

3-24
a) Cesium acetate is an ionic compound consisting of Cs\(^+\) cations and C\(_2\)H\(_3\)O\(_2\)\(^-\) anions. (Note that the formula for acetate ions can be written as either C\(_2\)H\(_3\)O\(_2\)\(^-\) or CH\(_3\)COO\(^-\).) The formula of the compound is CsC\(_2\)H\(_3\)O\(_2\).

\[
\mathcal{M} \text{ of } \text{CsC}_2\text{H}_3\text{O}_2 = (1 \times M \text{ of Cs}) + (2 \times M \text{ of C}) + (3 \times M \text{ of H}) + (2 \times M \text{ of O})
\]
\[
= (1 \times 132.9 \text{ g/mol Cs}) + (2 \times 12.01 \text{ g/mol C}) + (3 \times 1.008 \text{ g/mol H}) + (2 \times 16.00 \text{ g/mol O})
\]
\[
= 191.9 \text{ g/mol of CsC}_2\text{H}_3\text{O}_2
\]

There are 2 moles of C in 1 mole of CsC\(_2\)H\(_3\)O\(_2\).

\[
\text{Mass (g) of C} = (2 \text{ mol C}) \left( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 24.02 \text{ g C}
\]

\[
\text{Mass fraction} = \frac{\text{total mass of C}}{\text{molar mass of compound}} = \frac{24.02 \text{ g C}}{191.9 \text{ g CsC}_2\text{H}_3\text{O}_2} = 0.1252 \text{ mass fraction C}
\]

b) Uranyl sulfate trihydrate is a salt that consists of uranyl ions, UO\(_2^{2+}\), sulfate ions, SO\(_4^{2-}\), and three waters of hydration. The formula is UO\(_2\)SO\(_4\)·3H\(_2\)O. Note that the waters of hydration are included in the molar mass.

\[
\mathcal{M} \text{ of } \text{UO}_2\text{SO}_4\cdot3\text{H}_2\text{O} = (1 \times M \text{ of U}) + (9 \times M \text{ of O}) + (1 \times M \text{ of S}) + (6 \times M \text{ of H})
\]
\[
= (1 \times 238.0 \text{ g/mol U}) + (9 \times 16.00 \text{ g/mol O}) + (1 \times 32.06 \text{ g/mol S}) + (6 \times 1.008 \text{ g/mol H})
\]
\[
= 420.1 \text{ g/mol of UO}_2\text{SO}_4\cdot3\text{H}_2\text{O}
\]

There are 9 moles of O in 1 mole of UO\(_2\)SO\(_4\)·3H\(_2\)O.

\[
\text{Mass (g) of O} = (9 \text{ mol O}) \left( \frac{16.00 \text{ g O}}{1 \text{ mol O}} \right) = 144.0 \text{ g O}
\]

\[
\text{Mass fraction} = \frac{\text{total mass of O}}{\text{molar mass of compound}} = \frac{144.0 \text{ g O}}{420.1 \text{ g UO}_2\text{SO}_4\cdot3\text{H}_2\text{O}} = 0.3428 \text{ mass fraction O}
\]

3.23 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass fraction = \(\frac{\text{total mass of element}}{\text{molar mass of compound}}\).

Solution:

a) Calcium chlorate is an ionic compound consisting of Ca\(^{2+}\) cations and ClO\(_3^{-}\) anions. The formula of the compound is Ca(ClO\(_3\))\(_2\).

\[
\mathcal{M} \text{ of } \text{Ca(ClO}_3)_2 = (1 \times M \text{ of Ca}) + (2 \times M \text{ of Cl}) + (6 \times M \text{ of O})
\]
\[
= (1 \times 40.08 \text{ g/mol Ca}) + (2 \times 35.45 \text{ g/mol Cl}) + (6 \times 16.00 \text{ g/mol O})
\]
\[
= 206.98 \text{ g/mol of Ca(ClO}_3)_2
\]

There are 2 moles of Cl in 1 mole of Ca(ClO\(_3\))\(_2\).

\[
\text{Mass (g) of Cl} = (2 \text{ mol Cl}) \left( \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} \right) = 70.92 \text{ g Cl}
\]

\[
\text{Mass fraction} = \frac{\text{total mass of Cl}}{\text{molar mass of compound}} = \frac{70.90 \text{ g Cl}}{206.98 \text{ g Ca(ClO}_3)_2} = 0.3425 \text{ mass fraction Cl}
\]

b) Dinitrogen trioxide has the formula N\(_2\)O\(_3\). Di- indicates 2 N atoms and tri- indicates 3 O atoms.

\[
\mathcal{M} \text{ of } \text{N}_2\text{O}_3 = (2 \times M \text{ of N}) + (3 \times M \text{ of O})
\]
\[
= (2 \times 14.01 \text{ g/mol N}) + (3 \times 16.00 \text{ g/mol O}) = 76.02 \text{ g/mol of N}_2\text{O}_3
\]

There are 2 moles of N in 1 mole of N\(_2\)O\(_3\).

\[
\text{Mass (g) of N} = (2 \text{ mol N}) \left( \frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 28.02 \text{ g N}
\]

\[
\text{Mass fraction} = \frac{\text{total mass of N}}{\text{molar mass of compound}} = \frac{28.02 \text{ g N}}{76.02 \text{ g N}_2\text{O}_3} = 0.3686 \text{ mass fraction N}
\]

3.24 Plan: Divide the mass given by the molar mass of O\(_2\) to find moles. Since 1 mole of oxygen molecules contains 2 moles of oxygen atoms, multiply the moles by 2 to obtain moles of atoms and then multiply by Avogadro’s number to obtain the number of atoms.

Solution:
Moles of O₂ = \( \left( \frac{1 \text{ mol O₂}}{32.00 \text{ g O₂}} \right) \left( \frac{38.0 \text{ g O₂}}{1 \text{ mol O₂}} \right) = 1.1875 \text{ mol O₂} \)

Moles of O atoms = \( \left( \frac{2 \text{ mol O atoms}}{1 \text{ mol O₂}} \right) \left( \frac{1.1875 \text{ mol O₂}}{1 \text{ mol O₂}} \right) = 2.375 \text{ mol O atoms} \)

Number of O atoms = \( \left( \frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O atoms}} \right) \left( \frac{2.375 \text{ mol O atoms}}{1 \text{ mol O atoms}} \right) = 1.43 \times 10^{24} \text{ O atoms} \)

**Plan:** Determine the formula of cisplatin from the figure, and then calculate the molar mass from the formula. Divide the mass given by the molar mass to find moles of cisplatin. Since 1 mole of cisplatin contains 6 moles of hydrogen atoms, multiply the moles given by 6 to obtain moles of hydrogen and then multiply by Avogadro’s number to obtain the number of atoms.

**Solution:**
The formula for cisplatin is Pt(Cl)\(_2\)(NH\(_3\))\(_2\).

\[ M_{\text{Pt(Cl)\(_2\)(NH\(_3\))\(_2\)}} = (1 \times \text{M of Pt}) + (2 \times \text{M of Cl}) + (2 \times \text{M of N}) + (6 \times \text{M of H}) \]

\[ = (1 \times 195.1 \text{ g/mol Pt}) + (2 \times 35.45 \text{ g/mol Cl}) + (2 \times 14.01 \text{ g/mol N}) + (6 \times 1.008 \text{ g/mol H}) \]

\[ = 300.1 \text{ g/mol of Pt(Cl)\(_2\)(NH\(_3\))\(_2\)} \]

a) Moles of cisplatin = \( \left( \frac{285.3 \text{ g cisplatin}}{300.1 \text{ g cisplatin}} \right) = 0.9506831 = 0.9507 \text{ mol cisplatin} \)

b) Moles of H atoms = \( \left( \frac{6 \text{ mol H}}{1 \text{ mol cisplatin}} \right) \left( \frac{0.98 \text{ mol cisplatin}}{1 \text{ mol cisplatin}} \right) = 5.88 \text{ mol H atoms} \)

Number of H atoms = \( \left( \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H atoms}} \right) \left( \frac{5.88 \text{ mol H atoms}}{1 \text{ mol H atoms}} \right) = 3.540936 \times 10^{24} = 3.5 \times 10^{24} \text{ H atoms} \)

**Plan:** Determine the formula of allyl sulfide from the figure, and then calculate the molar mass from the formula. In part a), multiply the given amount in moles by the molar mass to find the mass of the sample. In part b), divide the given mass by the molar mass to find moles of compound. Since 1 mole of compound contains 6 moles of carbon atoms, multiply the moles of compound by 6 to obtain moles of carbon and then multiply by Avogadro’s number to obtain the number of atoms.

**Solution:**
The formula, from the figure, is (C\(_3\)H\(_5\))\(_2\)S.

\[ M_{(C\(_3\)H\(_5\))\(_2\)S} = (6 \times \text{M of C}) + (10 \times \text{M of H}) + (1 \times \text{M of S}) \]

\[ = (6 \times 12.01 \text{ g/mol C}) + (10 \times 1.008 \text{ g/mol H}) + (1 \times 32.06 \text{ g/mol S}) \]

\[ = 114.20 \text{ g/mol of (C\(_3\)H\(_5\))\(_2\)S} \]

a) Mass (g) of allyl sulfide = \( \left( \frac{2.63 \text{ mol allyl sulfide}}{1 \text{ mol allyl sulfide}} \right) \left( \frac{114.20 \text{ g allyl sulfide}}{1 \text{ mol allyl sulfide}} \right) = 300.3460 = 300. \text{ g allyl sulfide} \)

b) Moles of allyl sulfide = \( \left( \frac{35.7 \text{ g (C\(_3\)H\(_5\))\(_2\)S}}{114.20 \text{ g (C\(_3\)H\(_5\))\(_2\)S}} \right) \left( \frac{1 \text{ mol (C\(_3\)H\(_5\))\(_2\)S}}{2 \text{ mol (C\(_3\)H\(_5\))\(_2\)S}} \right) = 0.312609 \text{ mol allyl sulfide} \)

Moles of C atoms = \( \left( \frac{0.312609 \text{ mol (C\(_3\)H\(_5\))\(_2\)S}}{1 \text{ mol (C\(_3\)H\(_5\))\(_2\)S}} \right) \left( \frac{6 \text{ mol C}}{2 \text{ mol (C\(_3\)H\(_5\))\(_2\)S}} \right) = 1.8757 \text{ mol C atoms} \)

Number of C atoms = \( \left( \frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C atoms}} \right) \left( \frac{1.8757 \text{ mol C atoms}}{1 \text{ mol C atoms}} \right) = 1.129546 \times 10^{24} = 1.13 \times 10^{24} \text{ C atoms} \)

**Plan:** Determine the molar mass of rust. Convert mass in kg to mass in g and divide by the molar mass to find the moles of rust. Since each mole of rust contains 1 mole of Fe\(_2\)O\(_3\), multiply the moles of rust by 1 to obtain moles of Fe\(_2\)O\(_3\). Multiply the moles of Fe\(_2\)O\(_3\) by 2 to obtain moles of Fe (1:2 Fe\(_2\)O\(_3\):Fe mole ratio) and multiply by the molar mass of Fe to convert to mass.

**Solution:**
a) \( M \) of \( \text{Fe}_2\text{O}_3\cdot4\text{H}_2\text{O} = (2 \times M \text{ of Fe}) + (7 \times M \text{ of O}) + (8 \times M \text{ of H}) \)
\[
= (2 \times 55.85 \text{ g/mol Fe}) + (7 \times 16.00 \text{ g/mol O}) + (8 \times 1.008 \text{ g/mol H}) = 231.76 \text{ g/mol}
\]
Mass (g) of rust = \[45.2 \text{ kg rust} \times \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 4.52 \times 10^4 \text{ g} \]
Moles of rust = \[\left( 4.52 \times 10^4 \text{ g rust} \right) \left( \frac{1 \text{ mol rust}}{231.76 \text{ g rust}} \right) = 195.029 = 195 \text{ mol rust} \]
b) The formula shows that there is 1 mole of \( \text{Fe}_2\text{O}_3 \) for every mole of rust, so there are also 195 mol of \( \text{Fe}_2\text{O}_3 \).
c) Moles of iron = \[\left( 195.029 \text{ mol } \text{Fe}_2\text{O}_3 \right) \left( \frac{2 \text{ mol Fe}}{1 \text{ mol } \text{Fe}_2\text{O}_3} \right) = 390.058 \text{ mol Fe} \]
Mass (g) of iron = \[\left( 390.058 \text{ mol Fe} \right) \left( \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \right) = 2.18 \times 10^4 \text{ g Fe} \]

3.28 Plan: Determine the molar mass of propane. Divide the given mass by the molar mass to find the moles. Since each mole of propane contains 3 moles of carbon, multiply the moles of propane by 3 to obtain moles of C atoms. Multiply the moles of C by its molar mass to obtain mass of carbon.
Solution:
a) The formula of propane is \( \text{C}_3\text{H}_8 \).
\( M \) of \( \text{C}_3\text{H}_8 \) = \( (3 \times \text{M} \text{ of C}) + (8 \times \text{M} \text{ of H}) = (3 \times 12.01 \text{ g/mol C}) + (8 \times 1.008 \text{ g/mol H}) = 44.09 \text{ g/mol} \)
Moles of \( \text{C}_3\text{H}_8 \) = \( \left( 85.5 \text{ g } \text{C}_3\text{H}_8 \right) \left( \frac{1 \text{ mol } \text{C}_3\text{H}_8}{44.09 \text{ g } \text{C}_3\text{H}_8} \right) = 1.939215 = 1.94 \text{ mol } \text{C}_3\text{H}_8 \)
b) Moles of C = \( \left( 1.939215 \text{ mol } \text{C}_3\text{H}_8 \right) \left( \frac{3 \text{ mol } \text{C}}{1 \text{ mol } \text{C}_3\text{H}_8} \right) = 5.817645 \text{ mol } \text{C} \)
Mass (g) of C = \( \left( 5.817645 \text{ mol } \text{C} \right) \left( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 69.86992 = 69.9 \text{ g } \text{C} \)

3.29 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of nitrogen present. Multiply the number of moles of nitrogen by its molar mass to find the total mass of nitrogen in 1 mole of compound. Divide the total mass of nitrogen by the molar mass of compound and multiply by 100 to determine mass percent. Mass percent = \( \frac{\text{mol N} \times \text{(molar mass N)}}{\text{molar mass of compound}} \times 100 \). Then rank the values in order of decreasing mass percent N.
Solution:

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Molar Mass (g/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium nitrate</td>
<td>KNO\textsubscript{3}</td>
<td>101.11</td>
</tr>
<tr>
<td>Ammonium nitrate</td>
<td>NH\textsubscript{4}NO\textsubscript{3}</td>
<td>80.05</td>
</tr>
<tr>
<td>Ammonium sulfate</td>
<td>(NH\textsubscript{4})\textsubscript{2}SO\textsubscript{4}</td>
<td>132.14</td>
</tr>
<tr>
<td>Urea</td>
<td>CO(NH\textsubscript{2})\textsubscript{2}</td>
<td>60.06</td>
</tr>
</tbody>
</table>

Mass % N in potassium nitrate = \( \frac{1 \text{ mol N} \times (14.01 \text{ g/mol N})}{101.11 \text{ g/mol}} \times 100 = 13.86\% \text{ N} \)
Mass % N in ammonium nitrate = \( \frac{2 \text{ mol N} \times (14.01 \text{ g/mol N})}{80.05 \text{ g/mol}} \times 100 = 35.00\% \text{ N} \)
Mass % N in ammonium sulfate = \( \frac{2 \text{ mol N} \times (14.01 \text{ g/mol N})}{132.14 \text{ g/mol}} \times 100 = 21.20\% \text{ N} \)
Mass % N in urea = \( \frac{2 \text{ mol N} \times (14.01 \text{ g/mol N})}{60.06 \text{ g/mol}} \times 100 = 46.65\% \text{ N} \)
Rank is $\text{CO(NH}_2\text{)}_2 > \text{NH}_4\text{NO}_3 > (\text{NH}_4\text{)}_2\text{SO}_4 > \text{KNO}_3$

### 3.30 Plan
The volume must be converted from cubic feet to cubic centimeters. The volume and the density will give the mass of galena which is then divided by molar mass to obtain moles. Part b) requires a conversion from cubic decimeters to cubic centimeters. The density allows a change from volume in cubic centimeters to mass which is then divided by the molar mass to obtain moles; the amount in moles is multiplied by Avogadro’s number to obtain formula units of PbS which is also the number of Pb atoms due to the 1:1 PbS:Pb mole ratio.

**Solution:**
Lead(II) sulfide is composed of $\text{Pb}^{2+}$ and $\text{S}^{2-}$ ions and has a formula of PbS.

$\mathcal{M}$ of PbS = $(1 \times \mathcal{M}$ of Pb) + $(1 \times \mathcal{M}$ of S) = $(1 \times 207.2 \text{ g/mol Pb}) + (1 \times 32.06 \text{ g/mol S}) = 239.3 \text{ g/mol}$

**a)** Volume (cm$^3$) = $(1.00 \text{ ft}^3 \text{ PbS}) \left( \frac{(12 \text{ in})^3}{(1 \text{ ft})^3} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in}} \right)^3 = 28316.85 \text{ cm}^3$

Mass (g) of PbS = $(28316.85 \text{ cm}^3 \text{ PbS}) \left( \frac{7.46 \text{ g PbS}}{1 \text{ cm}^3} \right) = 211243.7 \text{ g PbS}$

Moles of PbS = $(211243.7 \text{ g PbS}) \left( \frac{1 \text{ mol PbS}}{239.3 \text{ g PbS}} \right) = 882.7568 = \text{883 mol PbS}$

**b)** Volume (cm$^3$) = $(1.00 \text{ dm}^3 \text{ PbS}) \left( \frac{(0.1 \text{ m})^3}{(1 \text{ dm})^3} \right) \left( \frac{1 \text{ cm}}{10^{-2} \text{ m}} \right)^3 = 1.00 \times 10^3 \text{ cm}^3$

Mass (g) of PbS = $(1.00 \times 10^3 \text{ cm}^3 \text{ PbS}) \left( \frac{7.46 \text{ g PbS}}{1 \text{ cm}^3} \right) = 7460 \text{ g PbS}$

Moles of PbS = $(7460 \text{ g PbS}) \left( \frac{1 \text{ mol PbS}}{239.3 \text{ g PbS}} \right) = 31.17426 \text{ mol PbS}$

Moles of Pb = $(31.17426 \text{ mol PbS}) \left( \frac{1 \text{ mol Pb}}{1 \text{ mol PbS}} \right) = 31.17426 \text{ mol Pb}$

Number of lead atoms = $(31.17426 \text{ mol Pb}) \left( \frac{6.022 \times 10^{23} \text{ Pb atoms}}{1 \text{ mol Pb}} \right) = 1.87731 \times 10^{25} = 1.88 \times 10^{25} \text{ Pb atoms}$

### 3.31 Plan
If the molecular formula for hemoglobin (Hb) were known, the number of Fe$^{2+}$ ions in a molecule of hemoglobin could be calculated. It is possible to calculate the mass of iron from the percentage of iron and the molar mass of the compound. Assuming you have 1 mole of hemoglobin, take 0.33% of its molar mass as the mass of Fe in that 1 mole. Divide the mass of Fe by its molar mass to find moles of Fe in 1 mole of hemoglobin which is also the number of ions in 1 molecule.

**Solution:**
Mass of Fe = $(0.33\% \text{ Fe}) \left( \frac{6.8 \times 10^3 \text{ g mol}}{100 \% \text{ Hb}} \right) = 224.4 \text{ g Fe}$

Moles of Fe = $(224.4 \text{ g Fe}) \left( \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) = 4.0179 = 4.0 \text{ mol Fe}^{2+}/\text{mol Hb}$

Thus, there are $4 \text{ Fe}^{2+}/\text{molecule Hb}$.

### 3.32 Plan
Review the definitions of empirical and molecular formulas.

**Solution:**
An empirical formula describes the type and simplest ratio of the atoms of each element present in a compound, whereas a molecular formula describes the type and actual number of atoms of each element in a molecule of the compound. The empirical formula and the molecular formula can be the same. For example, the compound formaldehyde has the molecular formula, \( \text{CH}_2\text{O} \). The carbon, hydrogen, and oxygen atoms are present in the ratio of 1:2:1. The ratio of elements cannot be further reduced, so formaldehyde’s empirical formula and molecular formula are the same. Acetic acid has the molecular formula, \( \text{C}_2\text{H}_4\text{O}_2 \). The carbon, hydrogen, and oxygen atoms are present in the ratio of 2:4:2, which can be reduced to 1:2:1. Therefore, acetic acid’s empirical formula is \( \text{CH}_2\text{O} \), which is different from its molecular formula. Note that the empirical formula does not uniquely identify a compound, because acetic acid and formaldehyde share the same empirical formula but are different compounds.

3.33 1. Compositional data may be given as the mass of each element present in a sample of compound.
2. Compositional data may be provided as mass percents of each element in the compound.
3. Compositional data obtained through combustion analysis provides the mass of C and H in a compound.

3.34 Plan: Remember that the molecular formula tells the actual number of moles of each element in one mole of compound.
Solution:
a) No, this information does not allow you to obtain the molecular formula. You can obtain the empirical formula from the number of moles of each type of atom in a compound, but not the molecular formula.
b) Yes, you can obtain the molecular formula from the mass percentages and the total number of atoms.
Plan:
1) Assume a 100.0 g sample and convert masses (from the mass % of each element) to moles using molar mass.
2) Identify the element with the lowest number of moles and use this number to divide into the number of moles for each element. You now have at least one elemental mole ratio (the one with the smallest number of moles) equal to 1.00 and the remaining mole ratios that are larger than one.
3) Examine the numbers to determine if they are whole numbers. If not, multiply each number by a whole-number factor to get whole numbers for each element. You will have to use some judgment to decide when to round. Write the empirical formula using these whole numbers.
4) Check the total number of atoms in the empirical formula. If it equals the total number of atoms given then the empirical formula is also the molecular formula. If not, then divide the total number of atoms given by the total number of atoms in the empirical formula. This should give a whole number. Multiply the number of atoms of each element in the empirical formula by this whole number to get the molecular formula. If you do not get a whole number when you divide, return to step 3 and revise how you multiplied and rounded to get whole numbers for each element.

Roadmap:

| Mass (g) of each element (express mass percent directly as grams) |
| Divide by \( M \) (g/mol) |
| Amount (mol) of each element |
| Use numbers of moles as subscripts |
| Preliminary empirical formula |
| Change to integer subscripts |
| Empirical formula |
| Divide total number of atoms in molecule by the number of atoms in the |
empirical formula and multiply the empirical formula by that factor

Molecular formula

c) Yes, you can determine the molecular formula from the mass percent and the number of atoms of one element in a compound. Plan:
1) Follow steps 1–3 in part b).
2) Compare the number of atoms given for the one element to the number in the empirical formula. Determine the factor the number in the empirical formula must be multiplied by to obtain the given number of atoms for that element. Multiply the empirical formula by this number to get the molecular formula.

Roadmap:
(Same first three steps as in b).

Empirical formula

Divide the number of atoms of the one element in the molecule by the number of atoms of that element in the empirical formula and multiply the empirical formula by that factor

Molecular formula

d) No, the mass % will only lead to the empirical formula.
e) Yes, a structural formula shows all the atoms in the compound. Plan: Count the number of atoms of each type of element and record as the number for the molecular formula.

Roadmap:
Count the number of atoms of each element and use these numbers as subscripts

Molecular formula

3.35 MgCl₂ is an empirical formula, since ionic compounds such as MgCl₂ do not contain molecules.

3.36 Plan: Examine the number of atoms of each type in the compound. Divide all atom numbers by the common factor that results in the lowest whole-number values. Add the molar masses of the atoms to obtain the empirical formula mass.
Solution:
a) C₂H₄ has a ratio of 2 carbon atoms to 4 hydrogen atoms, or 2:4. This ratio can be reduced to 1:2, so that the empirical formula is CH₂. The empirical formula mass is 12.01 g/mol C + 2(1.008 g/mol H) = 14.03 g/mol.
b) The ratio of atoms is 2:6:2, or 1:3:1. The empirical formula is CH₃O and its empirical formula mass is 12.01 g/mol C + 3(1.008 g/mol H) + 16.00 g/mol O = 31.03 g/mol.
c) Since, the ratio of elements cannot be further reduced, the molecular formula and empirical formula are the same, N₂O₅. The formula mass is 2(14.01 g/mol N) + 5(16.00 g/mol O) = 108.02 g/mol.
d) The ratio of elements is 3 atoms of barium to 2 atoms of phosphorus to 8 atoms of oxygen, or 3:2:8. This ratio cannot be further reduced, so the empirical formula is also Ba₃(PO₄)₂, with a formula mass of 3(137.3 g/mol Ba) + 2(30.97 g/mol P) + 8(16.00 g/mol O) = 601.8 g/mol.
e) The ratio of atoms is 4:16, or 1:4. The empirical formula is TeI₄, and the formula mass is 127.6 g/mol Te + 4(126.9 g/mol I) = 635.2 g/mol.

3.37 Plan: Examine the number of atoms of each type in the compound. Divide all atom numbers by the common factor that results in the lowest whole-number values. Add the molar masses of the atoms to obtain the empirical
a) C₄H₈ has a ratio of 4 carbon atoms to 8 hydrogen atoms, or 4:8. This ratio can be reduced to 1:2, so that the empirical formula is CH₂. The empirical formula mass is 12.01 g/mol C + 2(1.008 g/mol H) = 14.03 g/mol.

b) C₃H₆O₃ has a ratio of atoms of 3:6:3, or 1:2:1. The empirical formula is CH₂O and its empirical formula mass is 12.01 g/mol C + 2(1.008 g/mol H) + 16.00 g/mol O = 30.03 g/mol.

c) P₄O₁₀ has a ratio of 4 P atoms to 10 O atoms, or 4:10. This ratio can be reduced to 2:5, so that the empirical formula is P₂O₅. The empirical formula mass is 2(30.97 g/mol P) + 5(16.00 g/mol O) = 141.94 g/mol.

d) Ga₂(SO₄)₃ has a ratio of 2 atoms of gallium to 3 atoms of sulfur to 12 atoms of oxygen, or 2:3:12. This ratio cannot be further reduced, so the empirical formula is also Ga₂(SO₄)₃, with a formula mass of 2(69.72 g/mol Ga) + 3(32.06 g/mol S) + 12(16.00 g/mol O) = 427.6 g/mol.

e) Al₂Br₆ has a ratio of atoms of 2:6, or 1:3. The empirical formula is AlBr₃, and the formula mass is 26.98 g/mol Al + 3(79.90 g/mol Br) = 266.7 g/mol.

3.38 Plan: Use the chemical symbols and count the atoms of each type to obtain the molecular formula. Divide the molecular formula by the largest common factor to give the empirical formula. Use nomenclature rules to derive the name. This compound is composed of two nonmetals. The naming rules for binary covalent compounds indicate that the element with the lower group number is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.

Solution: The compound has 2 sulfur atoms and 2 chlorine atoms and a molecular formula of S₂Cl₂. The compound’s name is disulfur dichloride. Sulfur is named first since it has the lower group number. The prefix di- is used for both elements since there are 2 atoms of each element. The empirical formula is S₂Cl₂ or SCl₂.

\[ M \text{ of } S_2Cl_2 = (2 \times M \text{ of S}) + (2 \times M \text{ of Cl}) = (2 \times 32.06 \text{ g/mol S}) + (2 \times 35.45 \text{ g/mol Cl}) = 135.02 \text{ g/mol} \]

3.39 Plan: Use the chemical symbols and count the atoms of each type to obtain the molecular formula. Divide the molecular formula by the largest common factor to give the empirical formula. Use nomenclature rules to derive the name. This compound is composed of two nonmetals. The naming rules for binary covalent compounds indicate that the element with the lower group number is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.

Solution: The compound has 4 phosphorus atoms and 6 oxygen atoms and a molecular formula of P₄O₆. The compound’s name is tetraphosphorus hexaoxide. Phosphorus is named first since it has the lower group number. The prefix tetra- is for phosphorus since there are 4 P atoms and hexa- is used for oxygen since there are 6 O atoms. The empirical formula is P₂O₃ or P₂O₅.

\[ M \text{ of } P_4O_6 = (4 \times M \text{ of P}) + (6 \times M \text{ of O}) = (4 \times 30.97 \text{ g/mol P}) + (6 \times 16.00 \text{ g/mol O}) = 219.88 \text{ g/mol} \]

3.40 Plan: Determine the molar mass of each empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution: Only approximate whole-number values are needed.

a) CH₂ has empirical mass equal to 12.01 g/mol C + 2(1.008 g/mol C) = 14.03 g/mol

\[ \text{Whole-number multiple } = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{42.08 \text{ g/mol}}{14.03 \text{ g/mol}} \right) = 3 \]

Multiplying the subscripts in CH₂ by 3 gives C₃H₆.

b) NH₂ has empirical mass equal to 14.01 g/mol N + 2(1.008 g/mol H) = 16.03 g/mol
Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{32.05 \text{ g/mol}}{16.03 \text{ g/mol}} \right) = 2

Multiplying the subscripts in NH₂ by 2 gives N₂H₄.

c) NO₂ has empirical mass equal to 14.01 g/mol N + 2(16.00 g/mol O) = 46.01 g/mol

Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{92.02 \text{ g/mol}}{46.01 \text{ g/mol}} \right) = 2

Multiplying the subscripts in NO₂ by 2 gives N₂O₄.

d) CHN has empirical mass equal to 12.01 g/mol C + 1.008 g/mol H + 14.01 g/mol N = 27.03 g/mol

Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{135.14 \text{ g/mol}}{27.03 \text{ g/mol}} \right) = 5

Multiplying the subscripts in CHN by 5 gives C₅H₅N₅.

3.41 Plan: Determine the molar mass of each empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution:

Only approximate whole-number values are needed.

a) CH has empirical mass equal to 12.01 g/mol C + 1.008 g/mol H = 13.02 g/mol

Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{78.11 \text{ g/mol}}{13.02 \text{ g/mol}} \right) = 6

Multiplying the subscripts in CH by 6 gives C₆H₆.

b) C₃H₆O₂ has empirical mass equal to 3(12.01 g/mol C) + 6(1.008 g/mol H) + 2(16.00 g/mol O) = 74.08 g/mol

Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{74.08 \text{ g/mol}}{74.08 \text{ g/mol}} \right) = 1

Multiplying the subscripts in C₃H₆O₂ by 1 gives C₃H₆O₂.

c) HgCl has empirical mass equal to 200.6 g/mol Hg + 35.45 g/mol Cl = 236.0 g/mol

Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{472.1 \text{ g/mol}}{236.0 \text{ g/mol}} \right) = 2

Multiplying the subscripts in HgCl by 2 gives Hg₂Cl₂.

d) C₇H₄O₂ has empirical mass equal to 7(12.01 g/mol C) + 4(1.008 g/mol H) + 2(16.00 g/mol O) = 120.10 g/mol

Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left( \frac{240.20 \text{ g/mol}}{120.10 \text{ g/mol}} \right) = 2

Multiplying the subscripts in C₇H₄O₂ by 2 gives C₁₄H₈O₄.

3.42 Plan: The empirical formula is the smallest whole-number ratio of the atoms or moles in a formula. All data must be converted to moles of an element by dividing mass by the molar mass. Divide each mole number by the smallest mole number to convert the mole ratios to whole numbers.

Solution:

a) 0.063 mol Cl and 0.22 mol O: preliminary formula is Cl₁₀₆₃O₇₀₂₂

Converting to integer subscripts (dividing all by the smallest subscript):

Cl₁₀₆₃O₀₂₂ \rightarrow Cl₁O₃.₅

The formula is Cl₁O₃.₅, which in whole numbers (x 2) is Cl₂O₇.

b) Find moles of elements by dividing by molar mass:

Moles of Si = (2.45 g Si) \left( \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} \right) = 0.08722 \text{ mol Si}
Moles of Cl = \( \frac{12.4 \text{ g Cl}}{35.45 \text{ g Cl}} \) = 0.349788 mol Cl

Preliminary formula is Si\(_{0.08722}\)Cl\(_{0.349788}\)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{Si_{0.08722}Cl_{0.349788}}{0.08722} \rightarrow SiCl_4
\]

The empirical formula is SiCl\(_4\).

c) Assume a 100 g sample and convert the masses to moles by dividing by the molar mass:

Moles of C = (100 g) \( \left( \frac{27.3 \text{ parts C by mass}}{100 \text{ parts by mass}} \right) \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) \) = 2.2731 mol C

Moles of O = (100 g) \( \left( \frac{72.7 \text{ parts O by mass}}{100 \text{ parts by mass}} \right) \left( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) \) = 4.5438 mol O

Preliminary formula is C\(_{2.2731}\)O\(_{4.5438}\)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{C_{2.2731}O_{4.5438}}{2.2731} \rightarrow C_{1}O_{2}
\]

The empirical formula is CO\(_2\).

3.43 Plan: The empirical formula is the smallest whole-number ratio of the atoms or moles in a formula. All data must be converted to moles of an element by dividing mass by the molar mass. Divide each mole number by the smallest mole number to convert the mole ratios to whole numbers.

Solution:

a) 0.039 mol Fe and 0.052 mol O: preliminary formula is Fe\(_{0.039}\)O\(_{0.052}\)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{Fe_{0.039}O_{0.052}}{0.039} \rightarrow Fe_{1}O_{1.33}
\]

The formula is Fe\(_{1}O_{1.33}\), which in whole numbers (x 3) is Fe\(_3\)O\(_4\).

b) Find moles of elements by dividing by molar mass:

Moles of P = (0.903 g P) \( \left( \frac{1 \text{ mol P}}{30.97 \text{ g P}} \right) \) = 0.029157 mol P

Moles of Br = (6.99 g Br) \( \left( \frac{1 \text{ mol Br}}{79.90 \text{ g Br}} \right) \) = 0.087484 mol Br

Preliminary formula is P\(_{0.029157}\)Br\(_{0.087484}\)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{P_{0.029157}Br_{0.087484}}{0.029157} \rightarrow P_{1}Br_{3}
\]

The empirical formula is PBr\(_3\).

c) Assume a 100 g sample and convert the masses to moles by dividing by the molar mass:

79.9% C and 100 – 79.9 = 20.1% H

Moles of C = (100 g) \( \left( \frac{79.9 \text{ parts C by mass}}{100 \text{ parts by mass}} \right) \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) \) = 6.6528 mol C

Moles of H = (100 g) \( \left( \frac{20.1 \text{ parts H by mass}}{100 \text{ parts by mass}} \right) \left( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) \) = 19.940 mol H

Preliminary formula is C\(_{6.6528}\)H\(_{19.940}\)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{C_{6.6528}H_{19.940}}{6.6528} \rightarrow C_{1}H_{3}
\]

The empirical formula is CH\(_3\).
3.44  **Plan:** The percent oxygen is 100% minus the percent nitrogen. Assume 100 grams of sample, and then the moles of each element may be found by dividing the mass of each element by its molar mass. Divide each of the moles by the smaller value, and convert to whole numbers to get the empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

**Solution:**

a) \( \% \text{ O} = 100\% - \% \text{ N} = 100\% - 30.45\% \text{ N} = 69.55\% \text{ O} \)

Assume a 100 g sample and convert the masses to moles by dividing by the molar mass:

\[
\text{Moles of N} = \left(100 \text{ g}\right) \left(\frac{30.45 \text{ parts N by mass}}{100 \text{ parts by mass}}\right) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}}\right) = 2.1734 \text{ mol N}
\]

\[
\text{Moles of O} = \left(100 \text{ g}\right) \left(\frac{69.55 \text{ parts O by mass}}{100 \text{ parts by mass}}\right) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}}\right) = 4.3469 \text{ mol O}
\]

Preliminary formula is \( \text{N}_2.1734\text{O}_{4.3469} \)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{2.1734}{0.7440} \rightarrow \text{N}_1\text{O}_3
\]

The empirical formula is \( \text{NO}_2 \).

b) Formula mass of empirical formula = 14.01 g/mol N + 2(16.00 g/mol O) = 46.01 g/mol

\[
\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{90 \text{ g/mol}}{46.01 \text{ g/mol}}\right) = 2
\]

Multiplying the subscripts in \( \text{NO}_2 \) by 2 gives \( \text{N}_2\text{O}_4 \) as the molecular formula.

Note: Only an approximate value of the molar mass is needed.

3.45  **Plan:** The percent silicon is 100% minus the percent chlorine. Assume 100 grams of sample, and then the moles of each element may be found by dividing the mass of each element by its molar mass. Divide each of the moles by the smaller value, and convert to whole numbers to get the empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

**Solution:**

a) \( \% \text{ Si} = 100\% - \% \text{ Cl} = 100\% - 79.1\% \text{ Cl} = 20.9\% \text{ Si} \)

Assume a 100 g sample and convert the masses to moles by dividing by the molar mass:

\[
\text{Moles of Si} = \left(100 \text{ g}\right) \left(\frac{20.9 \text{ parts Si by mass}}{100 \text{ parts by mass}}\right) \left(\frac{1 \text{ mol Si}}{28.09 \text{ g Si}}\right) = 0.7440 \text{ mol Si}
\]

\[
\text{Moles of Cl} = \left(100 \text{ g}\right) \left(\frac{79.1 \text{ parts Cl by mass}}{100 \text{ parts by mass}}\right) \left(\frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}}\right) = 2.2313 \text{ mol Cl}
\]

Preliminary formula is \( \text{Si}_{0.7440}\text{Cl}_{2.2313} \)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\text{Si}_{0.7440} \rightarrow \text{Si}_{1}
\]

\[
\text{Cl}_{2.2313} \rightarrow \text{Cl}_3
\]

The empirical formula is \( \text{SiCl}_3 \).

b) Formula mass of empirical formula = 28.09 g/mol Si + 3(35.45 g/mol Cl) = 134.44 g/mol

\[
\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{269 \text{ g/mol}}{134.44 \text{ g/mol}}\right) = 2
\]

Multiplying the subscripts in \( \text{SiCl}_3 \) by 2 gives \( \text{Si}_2\text{Cl}_6 \) as the molecular formula.

3.46  **Plan:** The moles of the metal are known, and the moles of fluorine atoms may be found in part a) from the M:F mole ratio in the compound formula. In part b), convert moles of F atoms to mass and subtract the mass of F from
the mass of MF₂ to find the mass of M. In part c), divide the mass of M by moles of M to determine the molar mass of M which can be used to identify the element.

Solution:

a) Determine the moles of fluorine.

\[
\text{Moles of F} = \left(0.600 \text{ mol M}\right) \left( \frac{2 \text{ mol F}}{1 \text{ mol M}} \right) = 1.20 \text{ mol F}
\]

b) Determine the mass of M.

\[
\text{Mass of F} = \left(1.20 \text{ mol F}\right) \left( \frac{19.00 \text{ g F}}{1 \text{ mol F}} \right) = 22.8 \text{ g F}
\]

\[
\text{Mass (g) of M} = \text{MF}_2(\text{g}) - \text{F(г)} = 46.8 \text{ g} - 22.8 \text{ g} = 24.0 \text{ g M}
\]

c) The molar mass is needed to identify the element.

\[
\text{Molar mass of M} = \frac{24.0 \text{ g M}}{0.600 \text{ mol M}} = 40.0 \text{ g/mol}
\]

The metal with the closest molar mass to 40.0 g/mol is calcium.

3.47 Plan: The moles of the metal oxide are known, and the moles of oxygen atoms may be found in part a) from the compound:oxygen mole ratio in the compound formula. In part b), convert moles of O atoms to mass and subtract the mass of O from the mass of M₂O₃ to find the mass of M. In part c), find moles of M from the compound:M mole ratio and divide the mass of M by moles of M to determine the molar mass of M which can be used to identify the element.

Solution:

a) Determine the moles of oxygen.

\[
\text{Moles of O} = \left(0.370 \text{ mol M}_2\text{O}_3\right) \left( \frac{3 \text{ mol O}}{1 \text{ mol M}_2\text{O}_3} \right) = 1.11 \text{ mol O}
\]

b) Determine the mass of M.

\[
\text{Mass of O} = \left(1.11 \text{ mol O}\right) \left( \frac{16.00 \text{ g O}}{1 \text{ mol O}} \right) = 17.76 \text{ g O}
\]

\[
\text{Mass (g) of M} = \text{M}_2\text{O}_3(\text{g}) - \text{O(г)} = 55.4 \text{ g (M + O)} - 17.76 = 37.64 = 37.6 \text{ g M}
\]

c) First, the number of moles of M must be calculated.

\[
\text{Moles M} = \left(0.370 \text{ mol M}_2\text{O}_3\right) \left( \frac{2 \text{ mol M}}{1 \text{ mol M}_2\text{O}_3} \right) = 0.740 \text{ mol M}
\]

The molar mass is needed to identify the element.

\[
\text{Molar mass of M} = \frac{37.6 \text{ g M}}{0.740 \text{ mol M}} = 50.86 \text{ g/mol}
\]

The metal with the closest molar mass to 50.9 g/mol is vanadium.

3.48 Plan: The empirical formula is the smallest whole-number ratio of the atoms or moles in a formula. Divide each mmole number by the smallest mmole value to convert the mmole ratios to whole numbers. Since all the values are given in millimoles, there is no need to convert to moles.

Solution:

Preliminary formula is C₆.16H₈.56N₁.23

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{C_{6.16}}{1.23} \frac{H_{8.56}}{1.23} \frac{N_{1.23}}{1.23} \rightarrow C_5H_7N_1
\]

The empirical formula is C₅H₇N₁.

3.49 Plan: The empirical formula is the smallest whole-number ratio of the atoms or moles in a formula. Assume 100 grams of cortisol so the percentages are numerically equivalent to the masses of each element. Convert each of the
masses to moles by dividing by the molar mass of each element involved. Divide each mole number by the smallest mole number to convert the mole ratios to whole numbers. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution:

Moles of C = \( \frac{69.6 \text{ g C}}{12.01 \text{ g C/mol}} \) = 5.7952 mol C

Moles of H = \( \frac{8.34 \text{ g H}}{1.008 \text{ g H/mol}} \) = 8.2738 mol H

Moles of O = \( \frac{22.1 \text{ g O}}{16.00 \text{ g O/mol}} \) = 1.38125 mol O

Preliminary formula is \( \text{C}_{5.7952}\text{H}_{8.2738}\text{O}_{1.38125} \)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\begin{align*}
\text{C} & \rightarrow 4.20 \\
\text{H} & \rightarrow 6.27 \\
\text{O} & \rightarrow 1.38
\end{align*}
\]

The carbon value is not close enough to a whole number to round the value. The smallest number that 4.20 may be multiplied by to get close to a whole number is 5. (You may wish to prove this to yourself.) All three ratios need to be multiplied by five: \( 5(\text{C}_{5.7952}\text{H}_{8.2738}\text{O}_{1.38125}) = \text{C}_{21}\text{H}_{30}\text{O}_{5} \).

The empirical formula mass is \( \frac{21(12.01 \text{ g/mol C}) + 30(1.008 \text{ g/mol H}) + 5(16.00 \text{ g/mol O})}{1} = 362.45 \text{ g/mol} \)

Whole-number multiple = \( \frac{362.47 \text{ g/mol}}{362.45 \text{ g/mol}} = 1 \)

The empirical formula mass and the molar mass given are the same, so the empirical and the molecular formulas are the same. The molecular formula is \( \text{C}_{21}\text{H}_{30}\text{O}_{5} \).

Plan: Determine the molecular formula from the figure, and the molar mass from the molecular formula. The formula gives the relative numbers of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound.

Mass percent = \( \frac{\text{total mass of element}}{\text{molar mass of compound}} \times 100 \). 

Solution:

Molecular formula = \( \text{C}_8\text{H}_9\text{NO}_2 \)

Molar mass = \( 8(12.01 \text{ g/mol C}) + 9(1.008 \text{ g/mol H}) + 1(14.01 \text{ g/mol N}) + 2(16.00 \text{ g/mol O}) = 151.16 \text{ g/mol} \)

There are 8 moles of C in 1 mole of \( \text{C}_8\text{H}_9\text{NO}_2 \).

Mass (g) of C = \( 8 \text{ mol C} \left( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 96.08 \text{ g C} \)

Mass percent C = \( \frac{96.08 \text{ g C}}{151.16 \text{ g C}_8\text{H}_9\text{NO}_2} \times 100 = 63.56\% \)

There are 9 moles of H in 1 mole of \( \text{C}_8\text{H}_9\text{NO}_2 \).

Mass (g) of H = \( 9 \text{ mol H} \left( \frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) = 9.072 \text{ g H} \)

Mass percent H = \( \frac{9.072 \text{ g H}}{151.16 \text{ g C}_8\text{H}_9\text{NO}_2} \times 100 = 6.002\% \)

There is 1 mole of N in 1 mole of \( \text{C}_8\text{H}_9\text{NO}_2 \).

Mass (g) of N = \( 1 \text{ mol N} \left( \frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 14.01 \text{ g N} \)

Mass percent N = \( \frac{14.01 \text{ g N}}{151.16 \text{ g C}_8\text{H}_9\text{NO}_2} \times 100 = 9.268\% \)
There are 2 moles of O in 1 mole of C₈H₉NO₂.

\[
\text{Mass (g) of } O = \left(2 \text{ mol H}\right)\left(16.00 \text{ g O} \frac{1 \text{ mol O}}{1 \text{ mol O}}\right) = 32.00 \text{ g O}
\]

Mass percent \(\frac{\text{total mass O}}{\text{molar mass of compound}} \times 100\% = \frac{32.00 \text{ g O}}{151.16 \text{ g C₈H₉NO₂}} \times 100\% = 21.1696 = 21.17\% \text{ O}

3.51 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO₂, and all of the hydrogen present in the sample is found in the hydrogen of H₂O. Convert the mass of CO₂ to moles and use the ratio between CO₂ and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H₂O. The moles of oxygen are more difficult to find, because additional O₂ was added to cause the combustion reaction. Subtracting the masses of C and H from the mass of the sample gives the mass of O. Convert the mass of O to moles of O. Take the moles of C, H, and O and divide by the smallest value to convert to whole numbers to get the empirical formula. Determine the empirical formula mass and compare it to the molar mass given in the problem to see how the empirical and molecular formulas are related. Finally, determine the molecular formula.

Solution:

\[
\text{Moles of } C = \left(0.449 \text{ g CO}_2\right)\left(1 \text{ mol CO}_2 \frac{1 \text{ mol C}}{44.01 \text{ g CO}_2}\right) = 0.010202 \text{ mol C}
\]

\[
\text{Mass (g) of } C = \left(0.010202 \text{ mol C}\right)\left(12.01 \text{ g C} \frac{1 \text{ mol C}}{1 \text{ mol CO}_2}\right) = 0.122526 \text{ g C}
\]

\[
\text{Moles of } H = \left(0.184 \text{ g H}_2O\right)\left(1 \text{ mol H} \frac{2 \text{ mol H}}{1 \text{ mol H}_2O}\right) = 0.020422 \text{ mol H}
\]

\[
\text{Mass (g) of } H = \left(0.020422 \text{ mol H}\right)\left(1.008 \text{ g H} \frac{1 \text{ mol H}}{1 \text{ mol H}_2O}\right) = 0.020585 \text{ g H}
\]

\[
\text{Mass (g) of } O = \text{Sample mass} - \left(\text{mass of C + mass of H}\right) = 0.1595 \text{ g} - \left(0.122526 \text{ g C} + 0.020585 \text{ g H}\right) = 0.016389 \text{ g O}
\]

\[
\text{Moles of } O = \left(0.016389 \text{ g O}\right)\left(1 \text{ mol O} \frac{16.00 \text{ g O}}{1 \text{ mol O}}\right) = 0.0010243 \text{ mol O}
\]

Preliminary formula = C₀.₀₁₀₂₀₂H₀.₀₂₀₄₂₂O₀.₀₀₁₀₂₄₃

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\text{C}_0.010202 \text{H}_0.020422 \text{O}_0.0010243 \rightarrow \text{C}_{10}\text{H}_{20}\text{O}_1
\]

Empirical formula = C₁₀H₂₀O

Empirical formula mass = 10(12.01 g/mol C) + 20(1.008 g/mol H) + 1(16.00 g/mol O) = 156.26 g/mol

The empirical formula mass is the same as the given molar mass so the empirical and molecular formulas are the same. The molecular formula is C₁₀H₂₀O.

3.52 A balanced chemical equation describes:
1) The identities of the reactants and products.
2) The molar (and molecular) ratios by which reactants form products.
3) The physical states of all substances in the reaction.

3.53 In a balanced equation, the total mass of the reactants is equal to the total mass of the products formed in the reaction. Thus, the law of mass conversation is obeyed.

3.54 Students I and II are incorrect. Both students changed a given formula. Only coefficients should be changed when balancing; subscripts cannot be changed. Student I failed to identify the product correctly, writing AlCl₂ instead of AlCl₃. Student II used atomic chlorine instead of molecular chlorine as a reactant. Student III followed the correct process, changing only coefficients.
3.55 Plan: Examine the diagram and label each formula. We will use A for red atoms and B for green atoms.

Solution:
The reaction shows A₂ and B₂ diatomic molecules forming AB molecules. Equal numbers of A₂ and B₂ combine to give twice as many molecules of AB. Thus, the reaction is A₂ + B₂ → 2AB. This is the balanced equation in b.

3.56 Plan: Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.

Solution:
a) _Cu(s) + ___ S₈(s) → ___ CuS(s)
Balance the S first, because there is an obvious deficiency of S on the right side of the equation. The 8 S atoms in S₈ require the coefficient 8 in front of CuS:
___ Cu(s) + ___ S₈(s) → 8CuS(s)
Then balance the Cu. The 16 Cu atoms in CuS require the coefficient 16 in front of Cu:
16Cu(s) + S₈(s) → 8CuS(s)
b) __ Cu(NO₃)₂(aq) + __ H₂O(l) → __ H₃PO₄(l)
Balance the P first, because there is an obvious deficiency of P on the right side of the equation. The 4 P atoms in P₄O₁₀ require a coefficient of 4 in front of H₃PO₄:
__ P₄O₁₀(s) + __ H₂O(l) → 4H₃PO₄(l)
Balance the H next, because H is present in only one reactant and only one product. The 12 H atoms in 4H₃PO₄ on the right require a coefficient of 6 in front of H₂O:
__ P₄O₁₀(s) + 6H₂O(l) → 4H₃PO₄(l)
Balance the O last, because it appears in both reactants and is harder to balance. There are 16 O atoms on each side:
P₄O₁₀(s) + 6H₂O(l) → 4H₃PO₄(l)
c) __ B₂O₃(g) + __ NaOH(aq) → __ Na₃BO₃(aq) + __ H₂O(l)
Balance oxygen last because it is present in more than one place on each side of the reaction. The 2 B atoms in B₂O₃ on the left require a coefficient of 2 in front of Na₃BO₃ on the right:
__ B₂O₃(g) + __ NaOH(aq) → 2Na₃BO₃(aq) + __ H₂O(l)
The 6 Na atoms in 2Na₃BO₃ on the right require a coefficient of 6 in front of NaOH on the left:
__ B₂O₃(g) + 6NaOH(aq) → 2Na₃BO₃(aq) + __ H₂O(l)
The 6 H atoms in 6H₂O on the left require a coefficient of 3 in front of H₂O on the right:
__ B₂O₃(g) + 6NaOH(aq) → 2Na₃BO₃(aq) + 3H₂O(l)
The oxygen is now balanced with 9 O atoms on each side:
B₂O₃(s) + 6NaOH(aq) → 2Na₃BO₃(aq) + 3H₂O(l)
d) __ CH₃NH₂(g) + __ O₂(g) → __ CO₂(g) + __ H₂O(g) + __ N₂(g)
There are 2 N atoms on the right in N₂ so a coefficient of 2 is required in front of CH₃NH₂ on the left:
2CH₃NH₂(g) + __ O₂(g) → __ CO₂(g) + __ H₂O(g) + 2N₂(g)
There are now 10 H atoms in 2CH₃NH₂ on the left so a coefficient of 5 is required in front of H₂O on the right:
2CH₃NH₂(g) + __ O₂(g) → __ CO₂(g) + 5H₂O(g) + 2N₂(g)
The 2 C atoms on the left require a coefficient of 2 in front of CO₂ on the right:
2CH₃NH₂(g) + __ O₂(g) → 2CO₂(g) + 5H₂O(g) + 2N₂(g)
The 9 O atoms on the right (4 O atoms in 2CO₂ plus 5 in 5H₂O) require a coefficient of 9/2 in front of O₂ on the left:
2CH₃NH₂(g) + 9/2O₂(g) → 2CO₂(g) + 5H₂O(g) + 2N₂(g)
Multiply all coefficients by 2 to obtain whole numbers:
4CH₃NH₂(g) + 9O₂(g) → 4CO₂(g) + 10H₂O(g) + 4N₂(g)

3.57 Plan: Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.

Solution:
a) __ Cu(NO₃)₂(aq) + ___ KOH(aq) → __ Cu(OH)₂(s) + ___ KNO₃(aq)
The 2 N atoms in Cu(NO₃)₂ on the left require a coefficient of 2 in front of KNO₃ on the right:
__ Cu(NO₃)₂(aq) + ___ KOH(aq) → __ Cu(OH)₂(s) + 2KNO₃(aq)
The 6 C atoms in C₆H₁₀O₅ on the left require a coefficient of 6 in front of CO₂ on the right:
\[ \text{C}_6\text{H}_{10}\text{O}_5(s) + 6\text{CO}_2(g) \rightarrow 6\text{CO}_2(g) + 3\text{H}_2\text{O}(l) \]
There are 8 O atoms and 2 H atoms on each side:
\[ \text{Cu(NO}_3\text{)}_2(aq) + 2\text{KOH}(aq) \rightarrow \text{Cu(OH)}_2(s) + 2\text{KNO}_3(aq) \]
\[ \text{b) } \text{BCl}_3(g) + 3\text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{BO}_3(s) + 3\text{HCl(g)} \]
The 3 Cl atoms in BCl₃ on the left require a coefficient of 3 in front of HCl on the right:
\[ \text{b) } \text{BCl}_3(g) + 3\text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{BO}_3(s) + 3\text{HCl(g)} \]
The 6 H atoms in 2Sc(OH)₃ on the right require a coefficient of 3 in front of H₂O on the left:
\[ \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
There are 3 O atoms and 1 B atom on each side:
\[ \text{CaSiO}_3(s) \]
\[ \text{d) } \text{(CN)}_2(g) + 4\text{H}_2\text{O}(l) \rightarrow \text{(CN)}_2(g) + 4\text{H}_2\text{O}(l) \]
\[ \text{H}_2\text{C}_2\text{O}_4(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) \]
\[ \text{b) } \text{H}_3\text{PO}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{HPO}_4(aq) + 2\text{H}_2\text{O}(l) \]
\[ \text{c) } \text{Na}_2\text{HPO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{Na}_2\text{HPO}_4(aq) + \text{H}_2\text{O}(l) \]
\[ \text{d) } \text{CaSiO}_3(s) \]
\[ \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{FeO}(s) + \text{O}_2(g) \]
\[ \text{Cu(OH)}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Cu(OH)}_2(s) + 2\text{H}_2\text{O}(l) \]
\[ \text{2H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(l) + \text{O}_2(g) \]
\[ \text{Cu(NO}_3\text{)}_2(aq) + 2\text{KOH}(aq) \rightarrow \text{Cu(OH)}_2(s) + 2\text{KNO}_3(aq) \]
\[ \text{b) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{c) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{d) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq) \]
\[ \text{CO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{CO}_3(aq) \]
\[ \text{SO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq) \]
\[ \text{Cu(OH)}_2(s) \rightarrow \text{Cu}_2\text{O}_2(s) + \text{H}_2\text{O}(l) \]
\[ \text{3.58 Plan: Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.} \]
\[ \text{Solution:} \]
\[ \text{Cu(NO}_3\text{)}_2(aq) + 2\text{KOH}(aq) \rightarrow \text{Cu(OH)}_2(s) + 2\text{KNO}_3(aq) \]
\[ \text{b) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{c) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{d) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{3.58 Plan: Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.} \]
\[ \text{Solution:} \]
\[ \text{Cu(NO}_3\text{)}_2(aq) + 2\text{KOH}(aq) \rightarrow \text{Cu(OH)}_2(s) + 2\text{KNO}_3(aq) \]
\[ \text{b) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{c) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{d) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{3.58 Plan: Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.} \]
\[ \text{Solution:} \]
\[ \text{Cu(NO}_3\text{)}_2(aq) + 2\text{KOH}(aq) \rightarrow \text{Cu(OH)}_2(s) + 2\text{KNO}_3(aq) \]
\[ \text{b) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{c) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{d) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{3.58 Plan: Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.} \]
\[ \text{Solution:} \]
\[ \text{Cu(NO}_3\text{)}_2(aq) + 2\text{KOH}(aq) \rightarrow \text{Cu(OH)}_2(s) + 2\text{KNO}_3(aq) \]
\[ \text{b) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{c) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{d) } \text{Sc}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l) \rightarrow 2\text{Sc(OH)}_3(s) \]
\[ \text{C}_6\text{H}_{10}\text{O}_5 (s) + 9\text{O}_2 (g) \rightarrow 6\text{CO}_2 (g) + 5\text{H}_2\text{O} (g) \]

There are 17 O atoms on the right (12 in 6CO2 and 5 in 5H2O); there are 5 O atoms in C6H10O5 so a coefficient of 2 in front of O2 on the left will bring the total of O atoms to 17:

\[ \text{C}_6\text{H}_{10}\text{O}_5 (s) + 6\text{O}_2 (g) \rightarrow 6\text{CO}_2 (g) + 5\text{H}_2\text{O} (g) \]

**3.59 Plan:** Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.

**Solution:**

a) \[ \underline{\text{As}_4\text{S}_6 (s)} + 9\text{O}_2 (g) \rightarrow \underline{\text{As}_4\text{O}_6 (s)} + 6\text{SO}_2 (g) \]

The 4 As atoms in \( \text{As}_4\text{S}_6 \) on the left require a coefficient of 4 in front of \( \text{As}_4\text{O}_6 \) on the right:

\[ \underline{4\text{As}_4\text{O}_6 (s)} + 9\text{O}_2 (g) \rightarrow \underline{4\text{As}_4\text{O}_6 (s)} + 6\text{SO}_2 (g) \]

There are 17 O atoms on the right (6 in \( \text{As}_4\text{O}_6 \) and 12 in 6SO2); there are 5 O atoms in \( \text{C}_6\text{H}_{10}\text{O}_5 \) so a coefficient of 6 in front of O2 will bring the total of O atoms on the right:

\[ \underline{4\text{As}_4\text{O}_6 (s)} + 12\text{O}_2 (g) \rightarrow \underline{4\text{As}_4\text{O}_6 (s)} + 6\text{SO}_2 (g) \]

b) \[ \underline{\text{Ca}_3(\text{PO}_4)_2 (s)} + \underline{6\text{SiO}_2 (s)} \rightarrow \underline{\text{Si}_6\text{O}_18 (s)} + \underline{6\text{Ca}_3(\text{PO}_4)_2 (s)} \]

The 3 P atoms in \( \text{Ca}_3(\text{PO}_4)_2 \) on the left require a coefficient of 2 in front of \( \text{Ca}_3(\text{PO}_4)_2 \) on the right:

\[ \underline{2\text{Ca}_3(\text{PO}_4)_2 (s)} + \underline{6\text{SiO}_2 (s)} \rightarrow \underline{\text{Si}_6\text{O}_18 (s)} + \underline{\text{Ca}_3(\text{PO}_4)_2 (s)} \]

The 4 O atoms in \( \text{Ca}_3(\text{PO}_4)_2 \) on the right require a coefficient of 4 in front of \( \text{H}_2\text{O} \) on the left:

\[ \underline{\text{Ca}_3(\text{PO}_4)_2 (s)} + 4\text{H}_2\text{O} (g) \rightarrow \underline{\text{Ca}_3(\text{PO}_4)_2 (s)} + \underline{6\text{H}_2\text{O} (g)} \]

The 6 Si atoms in \( \text{Si}_6\text{O}_18 \) on the right require a coefficient of 6 in front of \( \text{SiO}_2 \) on the left:

\[ \underline{6\text{SiO}_2 (s)} \rightarrow \underline{\text{Si}_6\text{O}_18 (s)} \]

There are 28 O atoms on the left (16 in \( 2\text{Ca}_3(\text{PO}_4)_2 \) and 12 in \( 6\text{SiO}_2 \)); there are 18 O atoms on the right in 6CaSiO3 so a coefficient of 10 in front of CO on the right will bring the total O atoms to 18 on the right:

\[ \underline{2\text{Ca}_3(\text{PO}_4)_2 (s)} + 6\text{SiO}_2 (s) + 10\text{CO} (g) \rightarrow \underline{\text{Si}_6\text{O}_18 (s)} + \underline{6\text{CaSiO}_3 (s)} + \underline{10\text{CO} (g)} \]

The 10 C atoms in 10CO on the right require a coefficient of 10 in front of C on the left:

\[ \underline{2\text{Ca}_3(\text{PO}_4)_2 (s)} + 6\text{SiO}_2 (s) + 10\text{C} (l) \rightarrow \underline{\text{Si}_6\text{O}_18 (s)} + 
\underline{6\text{CaSiO}_3 (s)} + 
\underline{10\text{C} (l)} \]

The 3 Fe atoms in \( \text{Fe}_2\text{O}_3 \) on the right require a coefficient of 3 in front of Fe on the left:

\[ \underline{3\text{Fe} (s)} + \underline{3\text{H}_2\text{O} (g)} \rightarrow \underline{\text{Fe}_2\text{O}_3 (s)} + \underline{3\text{H}_2 (g)} \]

The 8 H atoms on the left in 4H2O require a coefficient of 4 in front of H2 on the right:

\[ \underline{3\text{Fe} (s)} + 4\text{H}_2\text{O} (g) \rightarrow \underline{\text{Fe}_2\text{O}_3 (s)} + \underline{4\text{H}_2 (g)} \]

The 12 S atoms on the right (4 in S2N4 and 8 in S8) require a coefficient of 6 in front of S2Cl2 on the left:

\[ \underline{6\text{S}_2\text{Cl}_2 (l)} \rightarrow \underline{4\text{S}_2\text{N}_4 (s)} + \underline{8\text{S}_8 (s)} + \underline{\text{NH}_4\text{Cl} (s)} \]

The 16 N atoms on the right (4 in S2N4 and 12 in 12NH4Cl) require a coefficient of 16 in front of NH3 on the left:

\[ \underline{6\text{S}_2\text{Cl}_2 (l)} + \underline{16\text{NH}_3 (g)} \rightarrow \underline{4\text{S}_2\text{N}_4 (s)} + \underline{8\text{S}_8 (s)} + \underline{12\text{NH}_4\text{Cl} (s)} \]

**3.60 Plan:** The names must first be converted to chemical formulas. Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used. Remember that oxygen is diatomic.

**Solution:**

a) Gallium (a solid) and oxygen (a gas) are reactants and solid gallium(III) oxide is the only product:

\[ \underline{\text{Ga} (s)} + 3\underline{\text{O}_2 (g)} \rightarrow \underline{\text{Ga}_2\text{O}_3 (s)} \]

A coefficient of 2 in front of Ga on the left is needed to balance the 2 Ga atoms in \( \text{Ga}_2\text{O}_3 \):

\[ 2\underline{\text{Ga} (s)} + 3\underline{\text{O}_2 (g)} \rightarrow \underline{\text{Ga}_2\text{O}_3 (s)} \]

The 3 O atoms in \( \text{Ga}_2\text{O}_3 \) on the right require a coefficient of 3/2 in front of O2 on the left:

\[ 2\underline{\text{Ga} (s)} + 3/2\underline{\text{O}_2 (g)} \rightarrow \underline{\text{Ga}_2\text{O}_3 (s)} \]

Multiply all coefficients by 2 to obtain whole numbers:
b) Liquid hexane and oxygen gas are the reactants while carbon dioxide gas and gaseous water are the products:

\[ \text{C}_6\text{H}_{14}(l) + \text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 7\text{H}_2\text{O}(g) \]

The 6 C atoms in C6H14 on the left require a coefficient of 6 in front of CO2 on the right:

\[ \text{C}_6\text{H}_{14}(l) + \text{O}_2(g) \rightarrow 6\text{CO}_2(g) + \text{H}_2\text{O}(g) \]

The 14 H atoms in C6H14 on the left require a coefficient of 7 in front of H2O on the right:

\[ \text{C}_6\text{H}_{14}(l) + \text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 7\text{H}_2\text{O}(g) \]

The 19 O atoms on the right (12 in 6CO2 and 7 in 7H2O) require a coefficient of 19/2 in front of O2 on the left:

\[ \frac{19}{2}\text{C}_6\text{H}_{14}(l) + \text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 7\text{H}_2\text{O}(g) \]

Multiply all coefficients by 2 to obtain whole numbers:

\[ 2\text{C}_6\text{H}_{14}(l) + 3\text{O}_2(g) \rightarrow 12\text{CO}_2(g) + 14\text{H}_2\text{O}(g) \]

c) Aqueous solutions of calcium chloride and sodium phosphate are the reactants; solid calcium phosphate and an aqueous solution of sodium chloride are the products:

\[ \text{CaCl}_2(aq) + \text{Na}_3\text{PO}_4(aq) \rightarrow \text{Ca}_3(\text{PO}_4)_2(s) + 3\text{NaCl}(aq) \]

The 3 Ca atoms in Ca3(PO4)2 on the right require a coefficient of 3 in front of CaCl2 on the left:

\[ 3\text{CaCl}_2(aq) + \text{Na}_3\text{PO}_4(aq) \rightarrow \text{Ca}_3(\text{PO}_4)_2(s) + 3\text{NaCl}(aq) \]

The 6 Cl atoms in 3CaCl2 on the left require a coefficient of 6 in front of NaCl on the right:

\[ 3\text{CaCl}_2(aq) + \text{Na}_3\text{PO}_4(aq) \rightarrow \text{Ca}_3(\text{PO}_4)_2(s) + 6\text{NaCl}(aq) \]

The 6 Na atoms in 6NaCl on the right require a coefficient of 2 in front of Na3PO4 on the left:

\[ 3\text{CaCl}_2(aq) + 2\text{Na}_3\text{PO}_4(aq) \rightarrow \text{Ca}_3(\text{PO}_4)_2(s) + 6\text{NaCl}(aq) \]

There are now 2 P atoms on each side:

\[ 3\text{CaCl}_2(aq) + 2\text{Na}_3\text{PO}_4(aq) \rightarrow \text{Ca}_3(\text{PO}_4)_2(s) + 6\text{NaCl}(aq) \]

3.61 Plan: The names must first be converted to chemical formulas. Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used. Remember that oxygen is diatomic.

Solution:

a) Aqueous solutions of lead(II) nitrate and potassium iodide are the reactants; solid lead(II) iodide and an aqueous solution of potassium nitrate are the products:

\[ \text{Pb(NO}_3)_2(aq) + \text{KI}(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq) \]

There are 2 N atoms in Pb(NO3)2 on the left so a coefficient of 2 is required in front of KNO3 on the right:

\[ \text{Pb(NO}_3)_2(aq) + 2\text{KI}(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq) \]

There are now 2 Na atoms on each side:

\[ \text{Pb(NO}_3)_2(aq) + 2\text{KI}(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq) \]

b) Liquid disilicon hexachloride and water are the reactants and solid silicon dioxide, hydrogen chloride gas and hydrogen gas are the products:

\[ \text{Si}_2\text{Cl}_6(l) + \text{H}_2\text{O}(l) \rightarrow 2\text{SiO}_2(s) + \text{HCl}(g) + \text{H}_2(g) \]

The 2 Si atoms in Si2Cl6 on the left require a coefficient of 2 in front of SiO2 on the right:

\[ \text{Si}_2\text{Cl}_6(l) + \text{H}_2\text{O}(l) \rightarrow 2\text{SiO}_2(s) + \text{HCl}(g) + \text{H}_2(g) \]

The 6 Cl atoms in Si2Cl6 on the left require a coefficient of 6 in front of HCl on the right:

\[ \text{Si}_2\text{Cl}_6(l) + \text{H}_2\text{O}(l) \rightarrow 2\text{SiO}_2(s) + 6\text{HCl}(g) + \text{H}_2(g) \]

The 4 O atoms in 2SiO2 on the right require a coefficient of 4 in front of H2O on the left:

\[ \text{Si}_2\text{Cl}_6(l) + 4\text{H}_2\text{O}(l) \rightarrow 2\text{SiO}_2(s) + 6\text{HCl}(g) + \text{H}_2(g) \]

There are 8 H atoms in 4H2O on the left; there are 8 H atoms on the right (6 in 6HCl and 2 in H2):

\[ \text{Si}_2\text{Cl}_6(l) + 4\text{H}_2\text{O}(l) \rightarrow 2\text{SiO}_2(s) + 6\text{HCl}(g) + \text{H}_2(g) \]

c) Nitrogen dioxide and water are the reactants and an aqueous solution of nitric acid and nitrogen monoxide gas are the products:

\[ \text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_3(aq) + \text{NO}(g) \]

Start with hydrogen it occurs in only one reactant and one product:

The 2 H atoms in H2O on the left require a coefficient of 2 in front of HNO3 on the right:

\[ \text{NO}_2(g) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g) \]
\[\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g)\]

The 3 N atoms on the right (2 in 2HNO₃ and 1 in NO) require a coefficient of 3 in front of NO₂ on the left;

\[3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g)\]

There are now 7 O atoms on each side:

\[3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g)\]

3.62 **Plan:** The names must first be converted to chemical formulas. Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used. Remember that oxygen is diatomic. Add the three reactions to obtain the overall equation. Substances on each side of the arrow cancel out.

**Solution:**

a) **Step 1**  \(\text{Cl}(g) + \text{O}_3(g) \rightarrow \text{ClO}(g) + \text{O}_2(g)\)

**Step 2**  \(2\text{ClO}(g) \rightarrow \text{ClOOCl}(g)\)

**Step 3**  \(\text{ClOOCl}(g) \rightarrow 2\text{Cl}(g) + \text{O}_2(g)\)

b) Add the 3 steps to obtain the overall balanced equation after multiplying Step 1 by 2 so that ClO(g) and Cl(g), intermediate products, can be eliminated from the overall equation.

**Step 1**  \(2\text{Cl}(g) + 2\text{O}_3(g) \rightarrow 2\text{ClO}(g) + 2\text{O}_2(g)\)

**Step 2**  \(2\text{ClO}(g) \rightarrow \text{ClOOCl}(g)\)

**Step 3**  \(\text{ClOOCl}(g) \rightarrow 2\text{Cl}(g) + \text{O}_2(g)\)

\[2\text{Cl}(g) + 2\text{O}_3(g) + 2\text{ClO}(g) + \text{ClOOCl}(g) \rightarrow 2\text{ClO}(g) + 2\text{O}_2(g) + \text{ClOOCl}(g) + 2\text{Cl}(g) + \text{O}_2(g)\]

\[2\text{O}_3(g) \rightarrow 3\text{O}_2(g)\]

3.63 **Plan:** The *stoichiometrically equivalent molar ratio* is the ratio of the coefficients in the balanced equation. This can be used as a conversion factor to calculate amounts of reactants or products in a chemical reaction.

3.64 **Plan:** Write a balanced chemical reaction to obtain the molar ratio between the reactants. Compare the number of particles of each reactant with the molar ratio to find the limiting reactant. Use the limiting reactant to calculate the number of product molecules that will form.

**Solution:**

a) The reaction is \(A_2 + B_2 \rightarrow AB_3\) or, correctly balanced, \(A_2 + 3B_2 \rightarrow 2AB_3\). The mole ratio between \(A_2\) and \(B_2\) is 1:3. Three times as many \(B_2\) molecules are required as you have of \(A_2\) molecules. With 3 \(A_2\) molecules present, 3 x 3 = 9 \(B_2\) molecules would be required. Since you have only 6 \(B_2\) molecules, \(B_2\) is the limiting **reagent**.

b) The balanced equation shows that 2\(AB_3\) molecules are produced for every 3 \(B_2\) molecules that react. Use the 3:2 mole ratio between the limiting reactant, \(B_2\), and \(AB_3\):

\[
\text{Number of molecules of product} = \left(\frac{2 \text{ \(AB_3\) molecules}}{3 \text{ \(B_2\) molecules}}\right) \times 6 \text{ \(B_2\) molecules} = 4 \text{ \(AB_3\) molecules}
\]

3.65 **Plan:** The percent yield is the ratio of the actual to the theoretical value. Both yields can be expressed as a mass or mole comparison. The percent yield will be the same since mass and moles are directly proportional.

3.66 **Plan:** First, write a balanced chemical equation. Since \(A\) is the limiting reagent (\(B\) is in excess), \(A\) is used to determine the amount of \(C\) formed, using the mole ratio between reactant \(A\) and product \(C\).

**Solution:**

Plan: The balanced equation is \(aA + bB \rightarrow cC\). Divide the mass of \(A\) by its molar mass to obtain moles of \(A\). Use the molar ratio from the balanced equation to find the moles of \(C\). Multiply moles of \(C\) by its molar mass to obtain mass of \(C\).

**Roadmap:**

\[
\text{Mass (g) of A} \div \text{M (g/mol)}
\]
3.67 Plan: First, write a balanced chemical equation. Since the amounts of both reactants are given, the limiting reactant must be determined.

Solution:
Plan: The balanced equation is \( \text{dD} + \text{eE} \rightarrow \text{fF} \). Divide the mass of each reactant by its molar mass to obtain moles of each reactant. Use the appropriate molar ratios from the balanced equation to find the moles of \( \text{F} \) obtained from Reactant \( \text{D} \) and Reactant \( \text{E} \). The smaller amount of \( \text{F} \) is the amount produced. Multiply moles of \( \text{F} \) by its molar mass to obtain mass of \( \text{F} \).

Roadmap:

<table>
<thead>
<tr>
<th>Mass (g) of D</th>
<th>Mass (g) of E</th>
</tr>
</thead>
<tbody>
<tr>
<td>Divide by ( \mathcal{M} ) (g/mol)</td>
<td>Divide by ( \mathcal{M} ) (g/mol)</td>
</tr>
<tr>
<td>Amount (mol) of D</td>
<td>Amount (mol) of E</td>
</tr>
<tr>
<td>Molar ratio between D and F</td>
<td>Molar ratio between E and F</td>
</tr>
<tr>
<td>Amount (moles) of F</td>
<td>Amount (moles) of F</td>
</tr>
</tbody>
</table>

Choose lower number of moles of \( \text{F} \) and multiply by \( \mathcal{M} \) (g/mol)

3.68 Plan: Always check to see if the initial equation is balanced. If the equation is not balanced, it should be balanced before proceeding. Use the mole ratio from the balanced chemical equation to determine the moles of \( \text{Cl}_2 \) produced. The equation shows that 1 mole of \( \text{Cl}_2 \) is produced for every 4 moles of \( \text{HCl} \) that react. Multiply the moles of \( \text{Cl}_2 \) produced by the molar mass to convert to mass in grams.

Solution:
\( 4\text{HCl}(aq) + \text{MnO}_2(s) \rightarrow \text{MnCl}_2(aq) + 2\text{H}_2\text{O}(g) + \text{Cl}_2(g) \)

a) Moles of \( \text{Cl}_2 = \left( \frac{1.82 \text{ mol HCl}}{4 \text{ mol HCl}} \right) = 0.455 \text{ mol Cl}_2 \)

b) Mass (g) of \( \text{Cl}_2 = \left( \frac{0.455 \text{ mol Cl}_2}{1 \text{ mol Cl}_2} \right) \left( \frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} \right) = 32.2595 = 32.3 \text{ g Cl}_2 \)
3.69 **Plan:** Always check to see if the initial equation is balanced. If the equation is not balanced, it should be balanced before proceeding. Divide the amount of reactant in grams by its molar mass to determine moles of reactant. Use the mole ratio from the balanced chemical equation to determine the moles of Bi produced. The equation shows that 2 moles of Bi are produced for every 1 mole of Bi$_2$O$_3$ that reacts.

**Solution:**

\[
\text{Bi}_2\text{O}_3(s) + 3\text{C(s)} \rightarrow 2\text{Bi(s)} + 3\text{CO(g)}
\]

a) Moles of Bi$_2$O$_3$ = \( \frac{283 \text{ g Bi}_2\text{O}_3}{466.0 \text{ g Bi}_2\text{O}_3} \times 0.607296 = 0.607 \text{ mol Bi}_2\text{O}_3 \)

b) Moles of Bi = \( \frac{2 \text{ mol Bi}}{1 \text{ mol Bi}_2\text{O}_3} \times 0.607 \text{ mol Bi}_2\text{O}_3 = 1.21 \text{ mol Bi} \)

3.70 **Plan:** Convert the kilograms of oxygen to grams of oxygen and then moles of oxygen by dividing by its molar mass. Use the moles of oxygen and the mole ratio from the balanced chemical equation to determine the moles of KNO$_3$ required. Multiply the moles of KNO$_3$ by its molar mass to obtain the mass in grams.

**Solution:**

a) Mass (g) of O$_2$ = \( 56.6 \text{ kg O}_2 \times \frac{10^5 \text{ g}}{1 \text{ kg}} = 5.66 \times 10^4 \text{ g O}_2 \)

Moles of O$_2$ = \( \frac{5.66 \times 10^4 \text{ g O}_2}{32.00 \text{ g O}_2} \times 1.76875 \times 10^3 \text{ mol O}_2 \)

Moles of KNO$_3$ = \( 1.76875 \text{ mol O}_2 \times \frac{4 \text{ mol KNO}_3}{5 \text{ mol O}_2} = 1415 = 1.42 \times 10^3 \text{ mol KNO}_3 \)

b) Mass (g) of KNO$_3$ = \( 1415 \text{ mol KNO}_3 \times \frac{101.11 \text{ g KNO}_3}{1 \text{ mol KNO}_3} = 143070.65 = 1.43 \times 10^5 \text{ g KNO}_3 \)

Combining all steps gives:

Mass (g) of KNO$_3$ = \( 56.6 \text{ kg O}_2 \times \frac{10^5 \text{ g O}_2}{1 \text{ kg}} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{4 \text{ mol KNO}_3}{5 \text{ mol O}_2} \times \frac{101.11 \text{ g KNO}_3}{1 \text{ mol KNO}_3} \)

\[= 143070.65 = 1.43 \times 10^5 \text{ g KNO}_3 \]

3.71 **Plan:** Convert mass of Cr$_2$S$_3$ to moles by dividing by its molar mass. Use the mole ratio between Cr$_2$S$_3$ and Cr$_2$O$_3$ from the balanced chemical equation to determine the moles of Cr$_2$O$_3$ required. Multiply the moles of Cr$_2$O$_3$ by its molar mass to obtain the mass in grams.

**Solution:**

a) Moles of Cr$_2$S$_3$ = \( \frac{421 \text{ g Cr}_2\text{S}_3}{200.18 \text{ g Cr}_2\text{S}_3} = 2.103107 \text{ mol Cr}_2\text{S}_3 \)

Moles of Cr$_2$O$_3$ = \( 2.103107 \text{ mol Cr}_2\text{S}_3 \times \frac{1 \text{ mol Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{S}_3} = 2.103107 = 2.10 \text{ mol Cr}_2\text{O}_3 \)

b) Mass (g) of Cr$_2$O$_3$ = \( 2.103107 \text{ mol Cr}_2\text{O}_3 \times \frac{152.00 \text{ g Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{O}_3} = 319.672 = 3.20 \times 10^2 \text{ g Cr}_2\text{O}_3 \)

Combining all steps gives:

Mass (g) of Cr$_2$O$_3$ = \( 421 \text{ g Cr}_2\text{S}_3 \times \frac{1 \text{ mol Cr}_2\text{S}_3}{200.18 \text{ g Cr}_2\text{S}_3} \times \frac{1 \text{ mol Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{S}_3} \times \frac{152.00 \text{ g Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{O}_3} \)

\[= 319.672 = 3.20 \times 10^2 \text{ g Cr}_2\text{O}_3 \]

3.72 **Plan:** First, balance the equation. Convert the grams of diborane to moles of diborane by dividing by its molar mass. Use mole ratios from the balanced chemical equation to determine the moles of the products. Multiply the mole amount of each product by its molar mass to obtain mass in grams.
Solution:
The balanced equation is: $\text{B}_2\text{H}_6(\text{g}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_3\text{BO}_3(\text{s}) + 6\text{H}_2(\text{g})$.

Moles of $\text{B}_2\text{H}_6 = \left( \frac{43.82 \text{ g} \text{B}_2\text{H}_6}{27.67 \text{ g} \text{B}_2\text{H}_6} \right) \frac{1 \text{ mol} \text{B}_2\text{H}_6}{1 \text{ mol} \text{B}_2\text{H}_6} = 1.583665 \text{ mol B}_2\text{H}_6$

Moles of $\text{H}_3\text{BO}_3 = \left( 1.583665 \text{ mol} \text{B}_2\text{H}_6 \right) \frac{2 \text{ mol} \text{H}_3\text{BO}_3}{1 \text{ mol} \text{B}_2\text{H}_6} = 3.16733 \text{ mol H}_3\text{BO}_3$

Mass (g) of $\text{H}_3\text{BO}_3 = \left( \frac{143.82 \text{ g} \text{H}_3\text{BO}_3}{2 \text{ mol} \text{H}_3\text{BO}_3} \right) \frac{1 \text{ mol} \text{H}_3\text{BO}_3}{1 \text{ mol} \text{H}_3\text{BO}_3} = 195.83597 = 195.8 \text{ g H}_3\text{BO}_3$

Combining all steps gives:

Mass (g) of $\text{H}_3\text{BO}_3 = \left( \frac{43.82 \text{ g} \text{B}_2\text{H}_6}{2 \text{ mol} \text{B}_2\text{H}_6} \right) \frac{1 \text{ mol} \text{B}_2\text{H}_6}{1 \text{ mol} \text{B}_2\text{H}_6} \frac{1 \text{ mol} \text{H}_3\text{BO}_3}{1 \text{ mol} \text{H}_3\text{BO}_3} \frac{61.83 \text{ g} \text{H}_3\text{BO}_3}{2 \text{ mol} \text{H}_3\text{BO}_3} = 195.83597 = 195.8 \text{ g H}_3\text{BO}_3$

Moles of $\text{H}_2 = \left( 1.583665 \text{ mol} \text{B}_2\text{H}_6 \right) \frac{6 \text{ mol} \text{H}_2}{1 \text{ mol} \text{B}_2\text{H}_6} = 9.50199 \text{ mol H}_2$

Mass (g) of $\text{H}_2 = \left( \frac{2.016 \text{ g} \text{H}_2}{1 \text{ mol} \text{H}_2} \right) \frac{1 \text{ mol} \text{H}_2}{1 \text{ mol} \text{H}_2} = 19.16 \text{ g H}_2$

Combining all steps gives:

Mass (g) of $\text{H}_2 = \left( \frac{43.82 \text{ g} \text{B}_2\text{H}_6}{2 \text{ mol} \text{B}_2\text{H}_6} \right) \frac{1 \text{ mol} \text{B}_2\text{H}_6}{1 \text{ mol} \text{B}_2\text{H}_6} \frac{6 \text{ mol} \text{H}_2}{1 \text{ mol} \text{B}_2\text{H}_6} \frac{2.016 \text{ g} \text{H}_2}{1 \text{ mol} \text{H}_2} \frac{1 \text{ mol} \text{H}_2}{1 \text{ mol} \text{H}_2} = 19.16 \text{ g H}_2$

Plan: First, balance the equation. Convert the grams of silver sulfide to moles of silver sulfide by dividing by its molar mass. Use mole ratios from the balanced chemical equation to determine the moles of the products. Multiply the mole amount of each product by its molar mass to obtain mass in grams.

Solution:

First, balance the equation: $\text{Ag}_2\text{S}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{AgCl}(\text{s}) + \text{H}_2\text{S}(\text{g})$

Moles of $\text{Ag}_2\text{S} = \left( \frac{174 \text{ g} \text{Ag}_2\text{S}}{247.9 \text{ g} \text{Ag}_2\text{S}} \right) \frac{1 \text{ mol} \text{Ag}_2\text{S}}{1 \text{ mol} \text{Ag}_2\text{S}} = 0.7018959 \text{ mol Ag}_2\text{S}$

Moles of $\text{AgCl} = \left( 0.7018959 \text{ mol} \text{Ag}_2\text{S} \right) \frac{2 \text{ mol} \text{AgCl}}{1 \text{ mol} \text{Ag}_2\text{S}} = 1.403792 \text{ mol AgCl}$

Mass (g) of $\text{AgCl} = \left( \frac{143.4 \text{ g} \text{AgCl}}{1 \text{ mol} \text{AgCl}} \right) \frac{1 \text{ mol} \text{AgCl}}{1 \text{ mol} \text{AgCl}} = 201.304 = 201 \text{ g AgCl}$

Combining all steps gives:

Mass (g) $\text{AgCl} = \left( \frac{174 \text{ g} \text{Ag}_2\text{S}}{247.9 \text{ g} \text{Ag}_2\text{S}} \right) \frac{1 \text{ mol} \text{Ag}_2\text{S}}{1 \text{ mol} \text{Ag}_2\text{S}} \frac{2 \text{ mol} \text{AgCl}}{1 \text{ mol} \text{Ag}_2\text{S}} \frac{143.4 \text{ g} \text{AgCl}}{1 \text{ mol} \text{AgCl}} \frac{1 \text{ mol} \text{AgCl}}{1 \text{ mol} \text{AgCl}} = 201.304 = 201 \text{ g AgCl}$

Moles of $\text{H}_2\text{S} = \left( 0.7018959 \text{ mol} \text{Ag}_2\text{S} \right) \frac{1 \text{ mol} \text{H}_2\text{S}}{1 \text{ mol} \text{Ag}_2\text{S}} = 0.7018959 \text{ mol H}_2\text{S}$

Mass (g) of $\text{H}_2\text{S} = \left( \frac{34.08 \text{ g} \text{H}_2\text{S}}{1 \text{ mol} \text{H}_2\text{S}} \right) \frac{1 \text{ mol} \text{H}_2\text{S}}{1 \text{ mol} \text{H}_2\text{S}} = 23.9206 = 23.9 \text{ g H}_2\text{S}$

Combining all steps gives:

Mass (g) of $\text{H}_2\text{S} = \left( \frac{174 \text{ g} \text{Ag}_2\text{S}}{247.9 \text{ g} \text{Ag}_2\text{S}} \right) \frac{1 \text{ mol} \text{Ag}_2\text{S}}{1 \text{ mol} \text{Ag}_2\text{S}} \frac{1 \text{ mol} \text{H}_2\text{S}}{1 \text{ mol} \text{Ag}_2\text{S}} \frac{34.08 \text{ g} \text{H}_2\text{S}}{1 \text{ mol} \text{H}_2\text{S}} \frac{1 \text{ mol} \text{H}_2\text{S}}{1 \text{ mol} \text{H}_2\text{S}} = 23.9206 = 23.9 \text{ g H}_2\text{S}$

Plan: Write the balanced equation by first writing the formulas for the reactants and products. Convert the mass of phosphorus to moles by dividing by the molar mass, use the mole ratio between phosphorus and chlorine from the
balanced chemical equation to obtain moles of chlorine, and finally divide the moles of chlorine by its molar mass to obtain amount in grams.

**Solution:**
Reactants: formula for phosphorus is given as P₄ and formula for chlorine gas is Cl₂ (chlorine occurs as a diatomic molecule). Product: formula for phosphorus pentachloride (the name indicates one phosphorus atom and five chlorine atoms) is PCl₅.

Equation: P₄ + Cl₂ → PCl₅

Balancing the equation: P₄ + 10Cl₂ → 4PCl₅

Moles of P₄ = \( \left( \frac{\text{1 mol P₄}}{123.88 \text{ g P₄}} \right) \left( \frac{\text{4 mol P₄}}{\text{455 g P₄}} \right) \) = 3.67291 mol P₄

Moles of Cl₂ = \( \left( \frac{\text{3.67291 mol P₄}}{\text{1 mol P₄}} \right) \left( \frac{\text{10 mol Cl₂}}{\text{1 mol P₄}} \right) \frac{\text{4 mol P₄}}{\text{455 g P₄}} \) = 36.7291 mol Cl₂

Mass (g) of Cl₂ = \( \left( \frac{\text{70.90 g Cl₂}}{\text{1 mol Cl₂}} \right) \left( \frac{\text{36.7291 mol Cl₂}}{\text{1 mol Cl₂}} \right) \) = 2604.09 = \( 2.60 \times 10^3 \) g Cl₂

Combining all steps gives:

Mass (g) of Cl₂ = \( \left( \frac{\text{70.90 g Cl₂}}{\text{1 mol Cl₂}} \right) \left( \frac{\text{36.7291 mol Cl₂}}{\text{1 mol Cl₂}} \right) \frac{\text{4 mol P₄}}{\text{455 g P₄}} \frac{\text{1 mol Cl₂}}{\text{1 mol P₄}} \) = 2604.09267 = \( 2.60 \times 10^3 \) g Cl₂

3.75 **Plan:** Write the balanced equation by first writing the formulas for the reactants and products. Convert the mass of sulfur to moles by dividing by the molar mass, use the mole ratio between sulfur and fluorine from the balanced chemical equation to obtain moles of fluorine, and finally divide the moles of fluorine by its molar mass to obtain amount in grams.

**Solution:**
Reactants: formula for sulfur is given as S₈ and formula for fluorine gas is F₂ (fluorine occurs as a diatomic molecule). Product: formula for sulfur hexafluoride (the name indicates one sulfur atom and six fluoride atoms) is SCl₆.

Equation: S₈ + F₂ → SF₆

Balancing the equation: S₈(s) + 24F₂(g) → 8SCl₆(s)

Moles of S₈ = \( \left( \frac{\text{17.8 g S₈}}{\text{256.48 g S₈}} \right) \left( \frac{\text{1 mol S₈}}{\text{24 mol F₂}} \right) \) = 0.0694011 mol S₈

Moles of F₂ = \( \left( \frac{\text{0.0694011 mol S₈}}{\text{1 mol S₈}} \right) \left( \frac{\text{24 mol F₂}}{\text{1 mol S₈}} \right) \) = 1.665626 mol F₂

Mass (g) of F₂ = \( \left( \frac{\text{38.00 g F₂}}{\text{1 mol F₂}} \right) \left( \frac{\text{1.665626 mol F₂}}{\text{1 mol F₂}} \right) \) = 63.294 = 63.3 g F₂

Combining all steps gives:

Mass (g) of F₂ = \( \left( \frac{\text{38.00 g F₂}}{\text{1 mol F₂}} \right) \left( \frac{\text{1.665626 mol F₂}}{\text{1 mol F₂}} \right) \frac{\text{17.8 g S₈}}{\text{256.48 g S₈}} \frac{\text{1 mol S₈}}{\text{24 mol F₂}} \) = 63.29382 = 63.3 g F₂

3.76 **Plan:** Begin by writing the chemical formulas of the reactants and products in each step. Next, balance each of the equations. Combine the equations for the separate steps by adjusting the equations so the intermediate (iodine monochloride) cancels. Finally, change the mass of product from kg to grams to moles by dividing by the molar mass and use the mole ratio between iodine and product to find the moles of iodine. Multiply moles by the molar mass of iodine to obtain mass of iodine.

**Solution:**
a) **Step 1** I₂(s) + Cl₂(g) → 2ICl(s)

**Step 2** ICl(s) + Cl₂(g) → ICl₃(s)
b) Multiply the coefficients of the second equation by 2, so that ICl(s), an intermediate product, can be eliminated from the overall equation.
\[ \text{I}_2(s) + \text{Cl}_2(g) \rightarrow 2\text{ICl}(s) \]

\[ 2\text{ICl}(s) + 2\text{Cl}_2(g) \rightarrow 2\text{ICl}_3(s) \]

\[ \text{I}_2(s) + \text{Cl}_2(g) + 2\text{ICl}(s) + 2\text{Cl}_2(g) \rightarrow 2\text{ICl}(s) + 2\text{ICl}_3(s) \]

Overall equation: \[ \text{I}_2(s) + 3\text{Cl}_2(g) \rightarrow 2\text{ICl}_3(s) \]

\[ \text{c)} \text{ Mass (g) of ICl}_3 = \left( 2.45 \text{ kg ICl}_3 \right) \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 2450 \text{ g ICl}_3 \]

Moles of ICl\(_3\) = \( \left( 2450 \text{ g ICl}_3 \right) \left( \frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) = 10.506 \text{ mol ICl}_3 \)

Moles of I\(_2\) = \( \left( 10.506 \text{ mol ICl}_3 \right) \left( \frac{1 \text{ mol I}_2}{2 \text{ mol ICl}_3} \right) = 5.253 \text{ mol I}_2 \)

Mass (g) of I\(_2\) = \( \left( 5.253 \text{ mol I}_2 \right) \left( \frac{253.8 \text{ g I}_2}{1 \text{ mol I}_2} \right) = 1333.211 \text{ g I}_2 \)

Combining all steps gives:

\[ \text{c)} \text{ Mass (g) of I}_{2} = \left( 2450 \text{ g ICl}_3 \right) \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 2450 \text{ g ICl}_3 \]

\[ \text{c) Mass (g) of ICl}_3 = \left( 2.45 \text{ kg ICl}_3 \right) \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 2450 \text{ g ICl}_3 \]

\[ \text{Moles of ICl}_3 = \left( 2450 \text{ g ICl}_3 \right) \left( \frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) = 10.506 \text{ mol ICl}_3 \]

\[ \text{Moles of I}_2 = \left( 10.506 \text{ mol ICl}_3 \right) \left( \frac{1 \text{ mol I}_2}{2 \text{ mol ICl}_3} \right) = 5.253 \text{ mol I}_2 \]

\[ \text{Mass (g) of I}_2 = \left( 5.253 \text{ mol I}_2 \right) \left( \frac{253.8 \text{ g I}_2}{1 \text{ mol I}_2} \right) = 1333.211 \text{ g I}_2 \]

3.77 Plan: Begin by writing the chemical formulas of the reactants and products in each step. Next, balance each of the equations. Combine the equations for the separate steps and cancel the intermediate (lead(II) oxide). Finally, change the amount of lead from metric tons to grams to moles by dividing by the molar mass and use the mole ratio between lead and sulfur dioxide to find the moles of sulfur dioxide. Multiply moles by the molar mass of sulfur dioxide to obtain mass and convert to metric tons.

Solution:

a) Step 1 \[ 2\text{PbS}(s) + 3\text{O}_2(g) \rightarrow 2\text{PbO}(s) + 2\text{SO}_2(g) \]

Step 2 \[ 2\text{PbO}(s) + \text{PbS}(s) \rightarrow 3\text{Pb}(l) + \text{SO}_2(g) \]

b) Combine the two reactions:

\[ 2\text{PbS}(s) + 3\text{O}_2(g) \rightarrow 2\text{PbO}(s) + 2\text{SO}_2(g) \]

\[ 2\text{PbO}(s) + \text{PbS}(s) \rightarrow 3\text{Pb}(l) + \text{SO}_2(g) \]

Overall equation: \[ 3\text{PbS}(s) + 3\text{O}_2(g) \rightarrow 3\text{Pb}(l) + 3\text{SO}_2(g) \]

\[ \text{PbS(s) + O}_2(g) \rightarrow \text{Pb(l) + SO}_2(g) \]

\[ \text{c) 1 metric ton = 1000 kg} \]

Mass (g) of Pb = \( \left( 1 \text{ ton Pb} \right) \left( \frac{10^3 \text{ kg}}{1 \text{ ton}} \right) = 1.000 \times 10^6 \text{ g Pb} \)

Moles of Pb = \( \left( 1 \times 10^6 \text{ g Pb} \right) \left( \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \right) = 4826.255 \text{ mol Pb} \)

Moles of SO\(_2\) = \( \left( 4826.255 \text{ mol Pb} \right) \left( \frac{1 \text{ mol SO}_2}{1 \text{ mol Pb}} \right) = 4826.255 \text{ mol SO}_2 \)

Metric ton SO\(_2\) = \( \left( 4826.255 \text{ mol SO}_2 \right) \left( \frac{64.06 \text{ g SO}_2}{1 \text{ mol SO}_2} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ ton}}{10^3 \text{ kg}} \right) = 0.309217 = 0.3092 \text{ ton SO}_2 \)

Combining all steps gives:
Metric ton SO\(_2\) = (1 ton Pb) \[\left( \frac{10^3 \text{ kg}}{1 \text{ ton}} \right) \times \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \times \left( \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \right) \times \left( \frac{1 \text{ mol SO}_2}{1 \text{ mol Pb}} \right) \times \left( \frac{64.06 \text{ g SO}_2}{1 \text{ mol SO}_2} \right) \times \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \times \left( \frac{1 \text{ ton}}{10 \text{ kg}} \right) = 0.309217 = 0.3092 \text{ ton SO}_2\]

3.78 Plan: Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of CaO formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of CaO is the limiting reactant. Convert the moles of CaO obtained from the limiting reactant to grams using the molar mass.

Solution:

\[2\text{Ca(s)} + \text{O}_2(g) \rightarrow 2\text{CaO(s)}\]

a) Moles of Ca = \(\left( \frac{4.20 \text{ g Ca}}{40.08 \text{ g Ca}} \right) \times \left( \frac{1 \text{ mol Ca}}{1 \text{ mol Ca}} \right) = 0.104790 \text{ mol Ca}\)

Moles of CaO from Ca = \(\left( \frac{2 \text{ mol CaO}}{2 \text{ mol Ca}} \right) \times \left( 0.104790 \text{ mol Ca} \right) = 0.104790 = 0.105 \text{ mol CaO}\)

b) Moles of O\(_2\) = \(\left( \frac{2.80 \text{ g O}_2}{32.00 \text{ g O}_2} \right) \times \left( \frac{1 \text{ mol O}_2}{1 \text{ mol O}_2} \right) = 0.0875 \text{ mol O}_2\)

Moles of CaO from O\(_2\) = \(\left( \frac{2 \text{ mol CaO}}{1 \text{ mol O}_2} \right) \times \left( 0.0875 \text{ mol O}_2 \right) = 0.17500 = 0.175 \text{ mol CaO}\)

c) Calcium is the limiting reactant since it will form less calcium oxide.

d) The mass of CaO formed is determined by the limiting reactant, Ca.

Mass (g) of CaO = \(\left( \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} \right) \times \left( 0.104790 \text{ mol CaO} \right) = 5.8766 = 5.88 \text{ g CaO}\)

Combining all steps gives:

Mass (g) of CaO = \(\left( \frac{4.20 \text{ g Ca}}{40.08 \text{ g Ca}} \right) \times \left( \frac{1 \text{ mol Ca}}{1 \text{ mol Ca}} \right) \times \left( \frac{2 \text{ mol CaO}}{2 \text{ mol Ca}} \right) \times \left( \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} \right) = 5.8766 = 5.88 \text{ g CaO}\)

3.79 Plan: Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of H\(_2\) formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of H\(_2\) is the limiting reactant. Convert the moles of H\(_2\) obtained from the limiting reactant to grams using the molar mass.

Solution:

\[\text{SrH}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Sr(OH)}_2(s) + 2\text{H}_2(g)\]

a) Moles of SrH\(_2\) = \(\left( 5.70 \text{ g SrH}_2 \right) \times \left( \frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) = 0.0635877 \text{ mol SrH}_2\)

Moles of H\(_2\) from SrH\(_2\) = \(\left( \frac{2 \text{ mol H}_2}{1 \text{ mol SrH}_2} \right) \times \left( 0.0635877 \text{ mol SrH}_2 \right) = 0.127175 = 0.127 \text{ mol H}_2\)

b) Mass (g) of H\(_2\)O = \(\left( 4.75 \text{ g H}_2\text{O} \right) \times \left( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 0.263596 \text{ mol H}_2\text{O}\)

Moles of H\(_2\) from H\(_2\)O = \(\left( \frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \right) \times \left( 0.263596 \text{ mol H}_2\text{O} \right) = 0.263596 = 0.264 \text{ mol H}_2\)

c) SrH\(_2\) is the limiting reagent since it will yield fewer moles of hydrogen gas.

d) The mass of H\(_2\) formed is determined by the limiting reactant, SrH\(_2\).

Mass (g) of H\(_2\) = \(\left( 0.127175 \text{ mol H}_2 \right) \times \left( \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) = 0.256385 = 0.256 \text{ g H}_2\)

Combining all steps gives:
Mass (g) of H₂ = \( (5.70 \text{ g SrH}_2) \left( \frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) \left( \frac{2 \text{ mol H}_2}{1 \text{ mol SrH}_2} \right) \left( \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) = 0.256385 = \boxed{0.256 \text{ g H}_2} \)

3.80 Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of HIO₃ formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of HIO₃ formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:
The balanced chemical equation for this reaction is:
\[
2\text{ICl}_3 + 3\text{H}_2\text{O} \to \text{ICl} + \text{HIO}_3 + 5\text{HCl}
\]

\textit{Hint:} Balance the equation by starting with oxygen. The other elements are in multiple reactants and/or products and are harder to balance initially.

Finding the moles of HIO₃ from the moles of ICl₃ (if H₂O is limiting):
\[
\text{Moles of ICl}_3 = (635 \text{ g ICl}_3) \left( \frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) = 2.722985 \text{ mol ICl}_3
\]

\[
\text{Moles of HIO}_3 \text{ from ICl}_3 = (2.722985 \text{ mol ICl}_3) \left( \frac{1 \text{ mol HIO}_3}{2 \text{ mol ICl}_3} \right) = 1.361492 = 1.36 \text{ mol HIO}_3
\]

Finding the moles of HIO₃ from the moles of H₂O (if ICl₃ is limiting):
\[
\text{Moles of H}_2\text{O} = (118.5 \text{ g H}_2\text{O}) \left( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 6.57603 \text{ mol H}_2\text{O}
\]

\[
\text{Moles HIO}_3 \text{ from H}_2\text{O} = (6.57603 \text{ mol H}_2\text{O}) \left( \frac{1 \text{ mol HIO}_3}{3 \text{ mol H}_2\text{O}} \right) = 2.19201 = 2.19 \text{ mol HIO}_3
\]

ICl₃ is the limiting reagent and will produce \(1.36\ \text{mol HIO}_3\).

Mass (g) of HIO₃ = \( (1.361492 \text{ mol HIO}_3) \left( \frac{175.9 \text{ g HIO}_3}{1 \text{ mol HIO}_3} \right) = 239.486 = \boxed{239 \text{ g HIO}_3}\)

Combining all steps gives:
\[
\text{Mass (g) of HIO}_3 = (635 \text{ g ICl}_3) \left( \frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) \left( \frac{1 \text{ mol HIO}_3}{2 \text{ mol ICl}_3} \right) \left( \frac{175.9 \text{ g HIO}_3}{1 \text{ mol HIO}_3} \right) = 239.486 = \boxed{239 \text{ g HIO}_3}\]

The remaining mass of the excess reagent can be calculated from the amount of H₂O combining with the limiting reagent.

Moles of H₂O required to react with 635 g ICl₃ = \( (2.722985 \text{ mol ICl}_3) \left( \frac{3 \text{ mol H}_2\text{O}}{2 \text{ mol ICl}_3} \right) = 4.0844775 \text{ mol H}_2\text{O} \)

Mass (g) of H₂O required to react with 635 g ICl₃ = \( (4.0844775 \text{ mol H}_2\text{O}) \left( \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 73.6023 = 73.6 \text{ g H}_2\text{O reacted} \)

Remaining H₂O = 118.5 g - 73.6 g = \boxed{44.9 \text{ g H}_2\text{O}}

3.81 Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of H₂S formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of H₂S formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:
The balanced chemical equation for this reaction is:
Al₂S₃ + 6H₂O → 2Al(OH)₃ + 3H₂S

Finding the moles of H₂S from the moles of Al₂S₃ (if H₂O is limiting):

Moles of Al₂S₃ = \( \frac{158 \text{ g Al}_2\text{S}_3}{150.14 \text{ g Al}_2\text{S}_3} = 1.05235 \text{ mol Al}_2\text{S}_3 \)

Moles of H₂S from Al₂S₃ = \( \frac{3 \text{ mol H}_2\text{S}}{1 \text{ mol Al}_2\text{S}_3} \times 1.05235 \text{ mol Al}_2\text{S}_3 = 3.15705 = 3.16 \text{ mol H}_2\text{S} \)

Finding the moles of H₂S from the moles of H₂O (if Al₂S₃ is limiting):

Moles of H₂O = \( \frac{131 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 7.26970 \text{ mol H}_2\text{O} \)

Moles of H₂S from H₂O = \( \frac{3 \text{ mol H}_2\text{S}}{6 \text{ mol H}_2\text{O}} \times 7.26970 \text{ mol H}_2\text{O} = 3.63485 = 3.63 \text{ mol H}_2\text{S} \)

Al₂S₃ is the limiting reagent and \( \textbf{3.16 mol of H}_2\text{S} \) will form.

Mass (g) of H₂S = \( \frac{34.08 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \times 3.15705 \text{ mol H}_2\text{S} = 107.592 = 108 \text{ g H}_2\text{S} \)

Combining all steps gives:

\( \text{Grams H}_2\text{S} = \frac{34.08 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \times 1.05235 \text{ mol Al}_2\text{S}_3 \times \frac{150.14 \text{ g Al}_2\text{S}_3}{1 \text{ mol Al}_2\text{S}_3} = 107.592 = 108 \text{ g H}_2\text{S} \)

The remaining mass of the excess reagent can be calculated from the amount of H₂O combining with the limiting reagent.

Moles of H₂O required to react with 158 g of Al₂S₃ = \( \frac{6 \text{ mol H}_2\text{O}}{1 \text{ mol Al}_2\text{S}_3} \times 1.05235 \text{ mol Al}_2\text{S}_3 = 6.31410 \text{ mol H}_2\text{O} \)

Mass (g) of H₂O required to react with 158 g of Al₂S₃ = \( \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times 6.31410 \text{ mol H}_2\text{O} = 113.780 \text{ g H}_2\text{O} \)

Remaining H₂O = 131 g H₂O - 113.780 g H₂O = 17.220 = 17 g H₂O

3.82 Plan: Write the balanced equation; the formula for carbon is C, the formula for oxygen is O₂ and the formula for carbon dioxide is CO₂. To determine which reactant is limiting, calculate the amount of CO₂ formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reactant. Use the limiting reactant and the mole ratio from the balanced chemical equation to determine the amount of CO₂ formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced equation is: C(s) + O₂(g) → CO₂(g)

Finding the moles of CO₂ from the moles of carbon (if O₂ is limiting):

Moles of CO₂ from C = \( \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \times 0.100 \text{ mol C} = 0.100 \text{ mol CO}_2 \)

Finding the moles of CO₂ from the moles of oxygen (if C is limiting):

Moles of O₂ = \( \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times 8.00 \text{ g O}_2 = 0.250 \text{ mol O}_2 \)

Moles of CO₂ from O₂ = \( \frac{1 \text{ mol CO}_2}{1 \text{ mol O}_2} \times 0.250 \text{ mol O}_2 = 0.25000 = 0.250 \text{ mol CO}_2 \)

Carbon is the limiting reactant and will be used to determine the amount of CO₂ that will form.

Mass (g) of CO₂ = \( \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \times 0.100 \text{ mol CO}_2 = 4.401 = 4.40 \text{ g CO}_2 \)
Since carbon is limiting, the \( \text{O}_2 \) is in excess. The amount remaining depends on how much combines with the limiting reagent.

Moles of \( \text{O}_2 \) required to react with 0.100 mol of \( \text{C} \):
\[
\left( 0.100 \text{ mol C} \right) \left( \frac{1 \text{ mol O}_2}{1 \text{ mol C}} \right) = 0.100 \text{ mol O}_2
\]

Mass (g) of \( \text{O}_2 \) required to react with 0.100 mol of \( \text{C} \):
\[
\left( 0.100 \text{ mol O}_2 \right) \left( \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 3.20 \text{ g O}_2
\]

Remaining \( \text{O}_2 \) = 8.00 g – 3.20 g = \( \bf{4.80} \text{ g O}_2 \)

3.83 Plan: Write the balanced equation; the formula for hydrogen is \( \text{H}_2 \), the formula for oxygen is \( \text{O}_2 \), and the formula for water is \( \text{H}_2\text{O} \). To determine which reactant is limiting, calculate the amount of \( \text{H}_2\text{O} \) formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of \( \text{H}_2\text{O} \) formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:
The balanced equation is:
\[
2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l)
\]
Finding the moles of \( \text{H}_2\text{O} \) from the moles of hydrogen (if \( \text{O}_2 \) is limiting):
\[
\text{Moles of } \text{H}_2 = (0.0375 \text{ g H}_2) \left( \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \right) = 0.01860 \text{ mol H}_2
\]
\[
\text{Moles of } \text{H}_2\text{O} \text{ from } \text{H}_2 = (0.01860 \text{ mol H}_2) \left( \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \right) = 0.01860 = 0.0186 \text{ mol H}_2\text{O}
\]
Finding the moles of \( \text{H}_2\text{O} \) from the moles of oxygen (if \( \text{H}_2 \) is limiting):
\[
\text{Mole of } \text{H}_2\text{O} \text{ from } \text{O}_2 = (0.0185 \text{ mol O}_2) \left( \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \right) = 0.0370 \text{ mol H}_2\text{O}
\]
The hydrogen is the limiting reactant, and will be used to determine the amount of water that will form.
\[
\text{Mass (g) of } \text{H}_2\text{O} = (0.01860 \text{ mol H}_2\text{O}) \left( \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 0.335172 = \bf{0.335} \text{ g H}_2\text{O}
\]
Since the hydrogen is limiting; the oxygen must be the excess reactant. The amount of excess reactant is determined from the limiting reactant.
\[
\text{Moles of } \text{O}_2 \text{ required to react with } 0.0375 \text{ g of } \text{H}_2 = (0.01860 \text{ mol H}_2) \left( \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \right) = 0.00930 \text{ mol O}_2
\]
\[
\text{Mass (g) of } \text{O}_2 \text{ required to react with } 0.0375 \text{ g of } \text{H}_2 = (0.00930 \text{ mol O}_2) \left( \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 0.2976 \text{ g O}_2
\]
\[
\text{Mass of } \text{O}_2 \text{ supplied} = (0.0185 \text{ mol O}_2) \left( \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 0.5920 \text{ g O}_2
\]
\[
\text{Remaining } \text{O}_2 = 0.5920 \text{ g } - 0.2976 \text{ g } = 0.2944 = \bf{0.294} \text{ g O}_2
\]

3.84 Plan: The question asks for the mass of each substance present at the end of the reaction. “Substance” refers to both reactants and products. Solve this problem using multiple steps. Recognizing that this is a limiting reactant problem, first write a balanced chemical equation. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, \( \text{AlCl}_3 \) is used. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of both products formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:
The balanced chemical equation is:
Al(NO₂)₃(aq) + 3NH₄Cl(aq) → AlCl₃(aq) + 3N₂(g) + 6H₂O(l)

Now determine the limiting reagent. We will use the moles of AlCl₃ produced to determine which is limiting.

Finding the moles of AlCl₃ from the moles of Al(NO₂)₃ (if NH₄Cl is limiting):

Moles of Al(NO₂)₃ = \( \frac{72.5 \text{ g Al(NO}_2\text{)}_3}{165.01 \text{ g Al(NO}_2\text{)}_3} \times 1 \text{ mol Al(NO}_2\text{)}_3 = 0.439367 \text{ mol Al(NO}_2\text{)}_3 \)

Moles of AlCl₃ from Al(NO₂)₃ = \( \frac{0.439367 \text{ mol Al(NO}_2\text{)}_3}{1 \text{ mol AlCl}_3} \times 1 \text{ mol AlCl}_3 = 0.439 \text{ mol AlCl}_3 \)

Finding the moles of AlCl₃ from the moles of NH₄Cl (if Al(NO₂)₃ is limiting):

Moles of NH₄Cl = \( \frac{58.6 \text{ g NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} \times 1 \text{ mol NH}_4\text{Cl} = 1.09553 \text{ mol NH}_4\text{Cl} \)

Moles of AlCl₃ from NH₄Cl = \( \frac{1.09553 \text{ mol NH}_4\text{Cl}}{3 \text{ mol NH}_4\text{Cl}} \times 1 \text{ mol AlCl}_3 = 0.365177 = 0.365 \text{ mol AlCl}_3 \)

Ammonium chloride is the limiting reactant, and it is used for all subsequent calculations.

Mass of substances after the reaction:

Al(NO₂)₃:

Mass (g) of Al(NO₂)₃ (the excess reactant) required to react with 58.6 g of NH₄Cl =

\( \frac{1.09553 \text{ mol NH}_4\text{Cl}}{3 \text{ mol NH}_4\text{Cl}} \times \frac{1 \text{ mol Al(NO}_2\text{)}_3}{1 \text{ mol Al(NO}_2\text{)}_3} \times \frac{165.01 \text{ g Al(NO}_2\text{)}_3}{1 \text{ mol Al(NO}_2\text{)}_3} = 60.2579 = 60.3 \text{ g Al(NO}_2\text{)}_3 \)

Al(NO₂)₃ remaining: 72.5 g – 60.3 g = 12.2 g Al(NO₂)₃

NH₄Cl: None left since it is the limiting reagent.

AlCl₃:

Mass (g) of AlCl₃ = \( \frac{0.365177 \text{ mol AlCl}_3}{1 \text{ mol AlCl}_3} \times 133.33 \text{ g AlCl}_3 = 48.689 = 48.7 \text{ g AlCl}_3 \)

N₂:

Mass (g) of N₂ = \( \frac{3 \text{ mol N}_2}{3 \text{ mol NH}_4\text{Cl}} \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 30.697 = 30.7 \text{ g N}_2 \)

H₂O:

Mass (g) of H₂O = \( \frac{6 \text{ mol H}_2\text{O}}{3 \text{ mol NH}_4\text{Cl}} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 39.483 = 39.5 \text{ g H}_2\text{O} \)

3.85 Plan: The question asks for the mass of each substance present at the end of the reaction. “Substance” refers to both reactants and products. Solve this problem using multiple steps. Recognizing that this is a limiting reactant problem, first write a balanced chemical equation. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, CaF₂ is used. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of both products formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:
The balanced chemical equation is:

\( \text{Ca(NO}_3\text{)}_2(s) + 2\text{NH}_3\text{F}(s) \rightarrow \text{CaF}_2(s) + 2\text{N}_2\text{O}(g) + 4\text{H}_2\text{O}(g) \)

Now determine the limiting reagent. We will use the moles of CaF₂ produced to determine which is limiting.

Finding the moles of CaF₂ from the moles of Ca(NO₃)₂ (if NH₄F is limiting):

Moles of Ca(NO₃)₂ = \( \frac{16.8 \text{ g Ca(NO}_3\text{)}_2}{164.10 \text{ g Ca(NO}_3\text{)}_2} \times 1 \text{ mol Ca(NO}_3\text{)}_2 = 0.1023766 \text{ mol Ca(NO}_3\text{)}_2 \)
Moles of CaF₂ from Ca(NO₃)₂ = \(\left(0.1023766 \text{ mol Ca(NO}_3\text{)}_2\right) \times \left(\frac{1 \text{ mol CaF}_2}{1 \text{ mol Ca(NO}_3\text{)}_2}\right) = 0.1023766 \text{ mol CaF}_2 = 0.102 \text{ mol CaF}_2\)

Finding the moles of CaF₂ from the moles of NH₄F (if Ca(NO₃)₂ is limiting):

\[
\text{Moles of NH}_4\text{F} = \left(17.50 \text{ g NH}_4\text{F}\right) \times \left(\frac{1 \text{ mol NH}_4\text{F}}{37.04 \text{ g NH}_4\text{F}}\right) = 0.47246 \text{ mol NH}_4\text{F}
\]

\[
\text{Moles of CaF}_2 \text{ from NH}_4\text{F} = \left(0.47246 \text{ mol NH}_4\text{F}\right) \times \left(\frac{1 \text{ mol CaF}_2}{2 \text{ mol NH}_4\text{F}}\right) = 0.23623 \text{ mol CaF}_2 = 0.236 \text{ mol CaF}_2
\]

Calcium nitrate is the limiting reactant, and it is used for all subsequent calculations.

Mass of substances after the reaction:

\[\text{Ca(NO}_3\text{)}_2: \text{None (It is the limiting reactant.)}\]

\[\text{NH}_4\text{F: Mass (g) of NH}_4\text{F (the excess reactant) required to react with 16.8 g of Ca(NO}_3\text{)}_2 = \left(0.1023766 \text{ mol Ca(NO}_3\text{)}_2\right) \times \left(\frac{2 \text{ mol NH}_4\text{F}}{1 \text{ mol Ca(NO}_3\text{)}_2}\right) \times \left(\frac{37.04 \text{ g NH}_4\text{F}}{1 \text{ mol NH}_4\text{F}}\right) = 7.58406 \text{ g Ca(NO}_3\text{)}_2\]

\[
\text{NH}_4\text{F remaining:} \ 17.50 \text{ g} - 7.58 \text{ g} = 9.9159 \text{ g} \ (9.92 \text{ g NH}_4\text{F})
\]

\[\text{CaF}_2: \text{Mass (g) of CaF}_2 = \left(0.1023766 \text{ mol Ca(NO}_3\text{)}_2\right) \times \left(\frac{1 \text{ mol CaF}_2}{1 \text{ mol Ca(NO}_3\text{)}_2}\right) \times \left(\frac{78.08 \text{ g CaF}_2}{1 \text{ mol CaF}_2}\right) = 7.99356 \text{ g CaF}_2 = 7.99 \text{ g CaF}_2
\]

\[\text{N}_2\text{O: Mass (g) of N}_2\text{O} = \left(0.1023766 \text{ mol Ca(NO}_3\text{)}_2\right) \times \left(\frac{2 \text{ mol N}_2\text{O}}{1 \text{ mol Ca(NO}_3\text{)}_2}\right) \times \left(\frac{44.02 \text{ g N}_2\text{O}}{1 \text{ mol N}_2\text{O}}\right) = 9.0132 \text{ g N}_2\text{O} = 9.01 \text{ g N}_2\text{O}
\]

\[\text{H}_2\text{O: Mass (g) of H}_2\text{O} = \left(0.1023766 \text{ mol Ca(NO}_3\text{)}_2\right) \times \left(\frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol Ca(NO}_3\text{)}_2}\right) \times \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}}\right) = 7.3793 \text{ g H}_2\text{O} = 7.38 \text{ g H}_2\text{O}
\]

3.86 Plan: Express the yield of each step as a fraction of 1.00; multiply the fraction of the first step by that of the second step and then multiply by 100 to get the overall percent yield.

Solution:

\[73\% = 0.73; 68\% = 0.68\]

\[(0.73 \times 0.68) \times 100 = 49.64 = 50\%\]

3.87 Plan: Express the yield of each step as a fraction of 1.00; multiply the fraction of the first step by that of the second step and then multiply by 100 to get the overall percent yield.

Solution:

\[48\% = 0.48; 73\% = 0.73\]

\[(0.48 \times 0.73) \times 100 = 35.04 = 35\%\]

3.88 Plan: Write and balance the chemical equation using the formulas of the substances. Determine the theoretical yield of the reaction from the mass of tungsten(VI) oxide. To do that, convert the mass of tungsten(VI) oxide to moles by dividing by its molar mass and then use the mole ratio between tungsten(VI) oxide and water to determine the moles and then mass of water that should be produced. Use the density of water to determine the actual yield of water in grams. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield.

Solution:

The balanced chemical equation is:

\[\text{WO}_3(s) + 3\text{H}_2(g) \rightarrow \text{W(s)} + 3\text{H}_2\text{O(l)}\]

Determining the theoretical yield of \(\text{H}_2\text{O}\):
3.89 **Plan:** Write and balance the chemical equation using the formulas of the substances. Determine the theoretical yield of the reaction from the mass of phosphorus trichloride. To do that, convert the mass of phosphorus trichloride to moles by dividing by its molar mass and then use the mole ratio between phosphorus trichloride and HCl to determine the moles and then mass of HCl that should be produced. The actual yield of the HCl is given. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield.

**Solution:**

The balanced chemical equation is:

$$\text{PCl}_3(\ell) + 3\text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{PO}_3(\text{aq}) + 3\text{HCl}(\text{g})$$

Determine the theoretical yield of HCl:

Moles of PCl₃ = \(\frac{200. \text{ g PCl}_3}{137.32 \text{ g PCl}_3}\) = 1.456452 mol PCl₃

Mass (g) of HCl (theoretical yield) = \(\frac{3 \text{ mol HCl}}{1 \text{ mol PCl}_3}\) \(\frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}}\) = 159.3067 g HCl

Actual yield (g) of HCl is given as 128 g HCl.

Calculate the percent yield:

% yield = \(\frac{\text{actual Yield}}{\text{theoretical Yield}}\) x 100% = \(\frac{128 \text{ g HCl}}{159.3067 \text{ g HCl}}\) x 100% = 80.3481586 = **80.3%**

3.90 **Plan:** Write the balanced chemical equation. Since quantities of two reactants are given, we must determine which is the limiting reactant. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, CH₃Cl is used. Only 75.0% of the calculated amounts of products actually form, so the actual yield is 75% of the theoretical yield.

**Solution:**

The balanced equation is: $\text{CH}_4(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{CH}_3\text{Cl}(\text{g}) + \text{HCl}(\text{g})$

Determining the limiting reactant:

Finding the moles of CH₃Cl from the moles of CH₄ (if Cl₂ is limiting):

Moles of CH₄ = \(\frac{20.5 \text{ g CH}_4}{16.04 \text{ g CH}_4}\) = 1.278055 mol CH₄

Moles of CH₃Cl from CH₄ = \(\frac{1.278055 \text{ mol CH}_4}{1 \text{ mol CH}_4}\) = 1.278055 mol CH₃Cl

Finding the moles of CH₃Cl from the moles of Cl₂ (if CH₄ is limiting):

Moles of Cl₂ = \(\frac{45.0 \text{ g Cl}_2}{70.90 \text{ g Cl}_2}\) = 0.634697 mol Cl₂
Moles of CH₃Cl from Cl₂ = \(0.634697 \text{ mol Cl}_2\left(\frac{1 \text{ mol CH}_3\text{Cl}}{1 \text{ mol Cl}_2}\right) = 0.634697 \text{ mol CH}_3\text{Cl}\)

Chlorine is the limiting reactant and is used to determine the theoretical yield of CH₃Cl:

Mass (g) of CH₃Cl (theoretical yield) = \(0.634697 \text{ mol CH}_3\text{Cl} \left(\frac{50.48 \text{ g CH}_3\text{Cl}}{1 \text{ mol CH}_3\text{Cl}}\right) = 32.0395 \text{ g CH}_3\text{Cl}\)

\[
\text{% yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \times 100\% \\
\text{Actual yield (g) of CH}_3\text{Cl} = \text{% yield} \left(\frac{\text{theoretical yield}}{100\%}\right) = \frac{75\%}{100\%} (32.0395 \text{ g CH}_3\text{Cl}) = 24.02962 = 24.0 \text{ g CH}_3\text{Cl}
\]

3.91 Plan: Write the balanced chemical equation. Since quantities of two reactants are given, we must determine which is the limiting reactant. To determine which reactant is limiting, calculate the amount of product formed from each reactant, assuming an excess of the other reactant. Only 93.0% of the calculated amount of product actually forms, so the actual yield is 93.0% of the theoretical yield.

Solution:
The balanced equation is: \(3\text{Ca(s) + N}_2(g) \rightarrow \text{Ca}_3\text{N}_2(s)\)

Determining the limiting reactant:
Finding the moles of \(\text{Ca}_3\text{N}_2\) from the moles of \(\text{Ca}\) (if \(\text{N}_2\) is limiting):

\[
\text{Moles of Ca} = \left(\frac{56.6 \text{ g Ca}}{40.08 \text{ g Ca}}\right) \left(\frac{1 \text{ mol Ca}}{1 \text{ mol Ca}}\right) = 1.412176 \text{ mol Ca}
\]

\[
\text{Moles of Ca}_3\text{N}_2 \text{ from Ca} = \left(1.412176 \text{ mol Ca}\right) \left(\frac{1 \text{ mol Ca}_3\text{N}_2}{3 \text{ mol Ca}}\right) = 0.470725 \text{ mol Ca}_3\text{N}_2
\]

Finding the moles of \(\text{Ca}_3\text{N}_2\) from the moles of \(\text{N}_2\) (if \(\text{Ca}\) is limiting):

\[
\text{Moles of N}_2 = \left(\frac{30.5 \text{ g N}_2}{28.02 \text{ g N}_2}\right) \left(\frac{1 \text{ mol N}_2}{1 \text{ mol N}_2}\right) = 1.08851 \text{ mol N}_2
\]

\[
\text{Moles of Ca}_3\text{N}_2 \text{ from N}_2 = \left(1.08851 \text{ mol N}_2\right) \left(\frac{1 \text{ mol Ca}_3\text{N}_2}{1 \text{ mol N}_2}\right) = 1.08851 \text{ mol Ca}_3\text{N}_2
\]

Ca is the limiting reactant and is used to determine the theoretical yield of \(\text{Ca}_3\text{N}_2\):

Mass (g) of \(\text{Ca}_3\text{N}_2\) (theoretical yield) = \(0.470725 \text{ mol Ca}_3\text{N}_2 \left(\frac{148.26 \text{ g Ca}_3\text{N}_2}{1 \text{ mol Ca}_3\text{N}_2}\right) = 69.7897 \text{ g Ca}_3\text{N}_2
\]

\[
\text{% yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \times 100\% \\
\text{Actual yield (g) of Ca}_3\text{N}_2 = \text{% yield} \left(\frac{\text{theoretical yield}}{100\%}\right) = \frac{93\%}{100\%} (69.7897 \text{ g Ca}_3\text{N}_2) = 64.9044 = 64.9 \text{ g Ca}_3\text{N}_2
\]

3.92 Plan: Write the balanced chemical equation; the formula for fluorine is \(\text{F}_2\), the formula for carbon tetrafluoride is \(\text{CF}_4\), and the formula for nitrogen trifluoride is \(\text{NF}_3\). To determine which reactant is limiting, calculate the amount of \(\text{CF}_4\) formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the mass of \(\text{CF}_4\) formed.

Solution:
The balanced chemical equation is:

\((\text{CN})_2(g) + 7\text{F}_2(g) \rightarrow 2\text{CF}_4(g) + 2\text{NF}_3(g)\)

Determining the limiting reactant:
Finding the moles of \(\text{CF}_4\) from the moles of \((\text{CN})_2\) (if \(\text{F}_2\) is limiting):

\[
\text{Moles of } (\text{CN})_2 = \left(\frac{70.1 \text{ g (CN)}_2}{26.99 \text{ g (CN)}_2}\right) \left(\frac{1 \text{ mol (CN)}_2}{1 \text{ mol (CN)}_2}\right) = 2.55492 \text{ mol (CN)}_2
\]

\[
\text{Moles of CF}_4 \text{ from (CN)}_2 = \left(2.55492 \text{ mol (CN)}_2\right) \left(\frac{2 \text{ mol CF}_4}{2 \text{ mol (CN)}_2}\right) = 2.55492 \text{ mol CF}_4
\]
Moles of CF$_4$ from (CN)$_2$ = \((60.0 \text{ g (CN)$_2$}) \left(\frac{1 \text{ mol (CN)$_2$}}{52.04 \text{ g (CN)$_2$}}\right) \left(\frac{2 \text{ mol CF}_4}{1 \text{ mol (CN)$_2$}}\right)\) = 2.30592 mol CF$_4$

Finding the moles of CF$_4$ from the moles of F$_2$ (if (CN)$_2$ is limiting):

Moles of CF$_4$ from F$_2$ = \((60.0 \text{ g F}_2) \left(\frac{1 \text{ mol F}_2}{38.00 \text{ g F}_2}\right) \left(\frac{2 \text{ mol CF}_4}{7 \text{ mol F}_2}\right)\) = 0.4511278 mol CF$_4$

F$_2$ is the limiting reactant, and will be used to calculate the amount of CF$_4$ produced.

Mass (g) of CF$_4$ = \((60.0 \text{ g F}_2) \left(\frac{1 \text{ mol F}_2}{38.00 \text{ g F}_2}\right) \left(\frac{2 \text{ mol CF}_4}{7 \text{ mol F}_2}\right) \left(\frac{88.01 \text{ g CF}_4}{1 \text{ mol CF}_4}\right)\) = 39.70376 g = 39.7 g CF$_4$

3.93 Plan: Write and balance the chemical reaction. Remember that both chlorine and oxygen exist as diatomic molecules. Use the mole ratio between oxygen and dichlorine monoxide to find the moles of dichlorine monoxide that reacted. Multiply the amount in moles by Avogadro’s number to convert to number of molecules.

Solution:

a) Both oxygen and chlorine are diatomic. Scene A best represents the product mixture as there are O$_2$ and Cl$_2$ molecules in Scene A. Scene B shows oxygen and chlorine atoms and Scene C shows atoms and molecules. Oxygen and chlorine atoms are NOT products of this reaction.

b) The balanced reaction is 2Cl$_2$O(g) → 2Cl$_2$(g) + O$_2$(g).

c) There is a 2:1 mole ratio between Cl$_2$ and O$_2$. In Scene A, there are 6 green molecules and 3 red molecules. Since twice as many Cl$_2$ molecules are produced as there are O$_2$ molecules produced, the red molecules are the O$_2$ molecules.

Moles of Cl$_2$O = \(3 \left(\frac{2 \text{ O atoms}}{1 \text{ O}_2 \text{ molecule}}\right) \left(\frac{0.050 \text{ mol O atoms}}{1 \text{ O atom}}\right) \left(\frac{1 \text{ mol O}_2 \text{ molecule}}{2 \text{ mol O atoms}}\right) = 0.30 \text{ mol Cl}_2\text{O}\)

Molecules of Cl$_2$O = \((0.30 \text{ mol Cl}_2\text{O}) \left(\frac{6.022 \times 10^{23} \text{ Cl}_2\text{O molecules}}{1 \text{ mol Cl}_2\text{O}}\right) = 1.8066 \times 10^{23} \text{ Cl}_2\text{O molecules}\)

3.94 Plan: Write a balanced equation for the reaction. Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of nitrogen monoxide formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of product is the limiting reactant. Convert the moles of nitrogen monoxide obtained from the limiting reactant to grams using the molar mass.

Solution:

The balanced chemical equation is:

\[4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)\]

Determining the limiting reactant:

Finding the moles of NO from the amount of NH$_3$ (if O$_2$ is limiting):

Moles of NO from NH$_3$ = \(\left(485 \text{ g NH}_3\right) \left(\frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3}\right) \left(\frac{4 \text{ mol NO}}{4 \text{ mol NH}_3}\right) = 28.47915 \text{ mol NO}\)

Finding the moles of NO from the amount of O$_2$ (if NH$_3$ is limiting):

Moles of NO from O$_2$ = \(\left(792 \text{ g O}_2\right) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2}\right) \left(\frac{4 \text{ mol NO}}{5 \text{ mol O}_2}\right) = 19.8 \text{ mol NO}\)

O$_2$ is the limiting reactant, and will be used to calculate the amount of NO formed:

Mass (g) of NO = \(\left(19.8 \text{ mol NO}\right) \left(\frac{30.01 \text{ g NO}}{1 \text{ mol NO}}\right) = 594.198 = 594 \text{ g NO}\)

Combining all of the steps gives:
**3.95 Plan:** Write a balanced equation. Use the density of butane to convert the given volume of butane to mass and divide by the molar mass of butane to convert mass to moles. Use the mole ratio between butane and oxygen to find the moles and then mass of oxygen required for the reaction. The mole ratio between butane and water is used to find the moles of water produced and the mole ratio between butane and carbon dioxide is used to find the moles of carbon dioxide produced. The total moles of product are multiplied by Avogadro’s number to find the number of product molecules.

**Solution:**

The balanced chemical equation is:

\[ 2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g) \]

a) Moles of \( C_4H_{10} \) = \( \left( \frac{5.50 \text{ mL}}{1 \text{ mL}} \right) \left( \frac{0.579 \text{ g } C_4H_{10}}{1 \text{ mL } C_4H_{10}} \right) \left( \frac{1 \text{ mol } C_4H_{10}}{58.12 \text{ g } C_4H_{10}} \right) = 0.054792 \text{ mol } C_4H_{10} \]

Mass (g) of \( O_2 \) = \( \left( \frac{0.054792 \text{ mol } C_4H_{10}}{2 \text{ mol } C_4H_{10}} \right) \left( \frac{13 \text{ mol } O_2}{1 \text{ mol } C_4H_{10}} \right) \left( \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} \right) = 11.3967 \text{ g } O_2 \]

b) Moles of \( H_2O \) = \( \left( \frac{0.054792 \text{ mol } C_4H_{10}}{2 \text{ mol } C_4H_{10}} \right) \left( \frac{10 \text{ mol } H_2O}{2 \text{ mol } C_4H_{10}} \right) = 0.27396 \text{ mol } H_2O \]

c) Moles of \( CO_2 \) = \( \left( \frac{0.054792 \text{ mol } C_4H_{10}}{2 \text{ mol } C_4H_{10}} \right) \left( \frac{8 \text{ mol } CO_2}{2 \text{ mol } C_4H_{10}} \right) = 0.219168 \text{ mol } CO_2 \]

Total moles = 0.27396 mol \( H_2O \) + 0.219168 mol \( CO_2 \) = 0.493128 mol

Total molecules = \( \left( \frac{0.493128 \text{ mol}}{1 \text{ mol}} \right) \left( \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \right) = 2.96962 \times 10^{23} \text{ molecules} \)

**3.96 Plan:** Write a balanced equation for the reaction. Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of \( NaBH_4 \) formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of product is the limiting reactant. Convert the moles of \( NaBH_4 \) obtained from the limiting reactant to grams using the molar mass. This is the theoretical yield of \( NaBH_4 \). Since there is a yield of 88.5%, the amount of \( NaBH_4 \) actually obtained will be 88.5% of the theoretical yield.

**Solution:**

The balanced chemical equation is:

\[ 2NaH(s) + B_2H_6(g) \rightarrow 2NaBH_4(s) \]

Determining the limiting reactant:

Finding the moles of \( NaBH_4 \) from the amount of \( NaH \) (if \( B_2H_6 \) is limiting):

Moles of \( NaBH_4 \) from \( NaH \) = \( \left( \frac{7.98 \text{ g } NaH}{24.00 \text{ g } NaH} \right) \left( \frac{1 \text{ mol } NaH}{2 \text{ mol } NaH} \right) = 0.3325 \text{ mol } NaBH_4 \)

Finding the moles of \( NaBH_4 \) from the amount of \( B_2H_6 \) (if \( NaH \) is limiting):

Moles of \( NaBH_4 \) from \( B_2H_6 \) = \( \left( \frac{8.16 \text{ g } B_2H_6}{27.67 \text{ g } B_2H_6} \right) \left( \frac{1 \text{ mol } B_2H_6}{1 \text{ mol } B_2H_6} \right) = 0.58981 \text{ mol } NaBH_4 \)

\( NaH \) is the limiting reactant, and will be used to calculate the theoretical yield of \( NaBH_4 \).

Mass (g) of \( NaBH_4 \) = \( \left( \frac{0.3325 \text{ mol } NaBH_4}{1 \text{ mol } NaBH_4} \right) \left( \frac{37.83 \text{ g } NaBH_4}{1 \text{ mol } NaBH_4} \right) = 12.5785 \text{ g } NaBH_4 \)

% yield = \( \frac{\text{actual Yield}}{\text{theoretical Yield}} \times 100\% \)

3-57
Mass (g) of NaBH₄ = \left( \frac{\text{% yield}}{100\%} \right) \left( \text{theoretical yield} \right) = \left( \frac{88.5\%}{100\%} \right) \left( 12.5785 \text{ g NaBH}_4 \right) = 11.13197 = 11.1 \text{ g NaBH}_4

Combining all steps gives:

\text{Mass (g) of NaBH}_4 = \left( 7.98 \text{ g NaH} \right) \left( \frac{1 \text{ mol NaH}}{24.00 \text{ g NaH}} \right) \left( \frac{2 \text{ mol NaBH}_4}{2 \text{ mol NaH}} \right) \left( \frac{37.83 \text{ g NaBH}_4}{1 \text{ mol NaBH}_4} \right) \left( \frac{88.5\%}{100\%} \right)

= 11.13197 = 11.1 \text{ g NaBH}_4

Plan: An atomic mass unit (amu) is exactly 1/12 the mass of a $^{12}\text{C}$ atom and Avogadro’s number of $^{12}\text{C}$ atoms has a mass of 12 g.

Solution:

a) Mass in grams = \left( 1 \text{ amu} \right) \left( \frac{1 \text{ atom}^{12}\text{C}}{12 \text{ amu}} \right) \left( \frac{1 \text{ mol}^{12}\text{C}}{6.022 \times 10^{23} \text{ atoms}^{12}\text{C}} \right) \left( \frac{12 \text{ g}^{12}\text{C}}{1 \text{ mol}^{12}\text{C}} \right) = 1.6605779 \times 10^{-24} = 1.661 \times 10^{-24} \text{ g}

Remember that all the values in this calculation except Avogadro’s number are exact numbers.

b) The g/amu ratio is $1.661 \times 10^{-24} \text{ g/amu}$.

Plan: First determine the empirical formula. Convert the mass of each element to moles by dividing the mass of each element by its molar mass. Divide each of the moles by the smaller value, and convert to whole numbers to get the empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution:

Moles of S = \left( 2.288 \text{ g S} \right) \left( \frac{1 \text{ mol S}}{32.06 \text{ g S}} \right) = 0.0713662 \text{ mol S}

Moles of N = \left( 1.000 \text{ g N} \right) \left( \frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 0.0713776 \text{ mol N}

Preliminary formula is S₀.₀₇₁₃₄₃₉N₀.₀₇₁₃₆₆₂

Converting to integer subscripts (dividing all by the smallest subscript):

$S_{0.0713662}O_{0.0713776} \rightarrow S_in_1$

The empirical formula is SN.

Formula mass of empirical formula = 32.06 g/mol S + 14.01 g/mol N = 46.07 g/mol

Whole-number multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \frac{184.27 \text{ g/mol}}{46.07 \text{ g/mol}} = 4

Multiplying the subscripts in SN by 4 gives $S_4N_4$ as the molecular formula.

Plan: The first step is to write and balance the chemical equation for the reaction. Convert the mass of each reactant to moles by dividing by the molar mass, remembering that the mass of the phosphoric acid reactant is 85% of the given mass of phosphoric acid solution. To determine which reactant is limiting, calculate the amount of hydroxyapatite formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of hydroxyapatite that forms.

Solution:

a) The balanced equation is:

$5\text{Ca(OH)}_2(\text{aq}) + 3\text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{Ca}_5(\text{PO}_4)_3(\text{OH})(\text{s}) + 9\text{H}_2\text{O}(\text{l})$

b) Finding the moles of $\text{Ca}_5(\text{PO}_4)_3(\text{OH})$ from the moles of $\text{Ca(OH)}_2$ (if $\text{H}_3\text{PO}_4$ is limiting):

Moles of $\text{Ca(OH)}_2$ = \left( 100. \text{ g Ca(OH)}_2 \right) \left( \frac{1 \text{ mol Ca(OH)}_2}{74.10 \text{ g Ca(OH)}_2} \right) = 1.349528 \text{ mol Ca(OH)}_2
Moles of Ca₅(PO₄)₃(OH) from Ca(OH)₂ = \left( \frac{1 \text{ mol Ca(OH)}_2}{5 \text{ mol Ca(OH)}_2} \right) \left( \frac{1 \text{ mol Ca₅(PO₄)}_3(\text{OH})}{1 \text{ mol Ca(OH)}_2} \right) = 0.2699056 \text{ mol Ca₅(PO₄)}_3(\text{OH})

Finding the moles of Ca₅(PO₄)₃(OH) from the moles of H₃PO₄ (if Ca(OH)₂ is limiting):

Moles of H₃PO₄ = \left( \frac{85 \text{ g H₃PO₄}}{100 \text{ g H₃PO₄ solution}} \right) \left( \frac{1 \text{ mol H₃PO₄}}{97.99 \text{ g H₃PO₄}} \right) = 0.867435 \text{ mol H₃PO₄}

Moles of Ca₅(PO₄)₃(OH) from H₃PO₄ = \left( \frac{0.867435 \text{ mol H₃PO₄}}{3 \text{ mol Ca₅(PO₄)₃(OH)}} \right) = 0.289145 \text{ mol Ca₅(PO₄)₃(OH)}

Ca(OH)₂ is the limiting reactant, and will be used to calculate the amount of Ca₅(PO₄)₃(OH) produced.

Mass (g) of Ca₅(PO₄)₃(OH) = \left( \frac{502.32 \text{ g Ca₅(PO₄)₃(OH)}}{1 \text{ mol Ca₅(PO₄)₃(OH)}} \right) \left( \frac{0.2699056 \text{ mol Ca₅(PO₄)₃(OH)}}{1 \text{ mol Ca(OH)₂}} \right) = 135.57898 = 140 \text{ g Ca₅(PO₄)₃(OH)}

3.100 Plan: The moles of narceine and the moles of water are required. We can assume any mass of narceine hydrate (we will use 100 g), and use this mass to determine the moles of hydrate. The moles of water in the hydrate is obtained by taking 10.8% of the 100 g mass of hydrate and converting the mass to moles of water. Divide the moles of water by the moles of hydrate to find the value of x.

Solution:
Assuming a 100 g sample of narceine hydrate:

Moles of narceine hydrate = \left( \frac{100 \text{ g narceine hydrate}}{499.52 \text{ g narceine hydrate}} \right) = 0.20019 \text{ mol narceine hydrate}

Mass (g) of H₂O = \left( \frac{10.8\% \text{ H₂O}}{100\% \text{ narceine hydrate}} \right) = 10.8 \text{ g H₂O}

Moles of H₂O = \left( \frac{10.8 \text{ g H₂O}}{18.02 \text{ g H₂O}} \right) = 0.59933 \text{ mol H₂O}

\[ x = \frac{\text{moles of H₂O}}{\text{moles of hydrate}} = \frac{0.59933 \text{ mol}}{0.20019 \text{ mol}} = 3 \]

Thus, there are three water molecules per mole of hydrate. The formula for narceine hydrate is narceine•3H₂O.

3.101 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative numbers of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass percent = \frac{\text{mol of element}}{\text{mol of compound}} \times 100. List the compounds from the highest %H to the lowest.

Solution:

<table>
<thead>
<tr>
<th>Name</th>
<th>Chemical formula</th>
<th>Molar mass (g/mol)</th>
<th>Mass percent H = \frac{\text{mol of element}}{\text{mol of compound}} \times 100</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ethane</td>
<td>C₂H₆</td>
<td>30.07</td>
<td>\frac{6 \text{ mol}(1.008 \text{ g/mol})}{30.07 \text{ g}} \times 100 = 20.11% H</td>
</tr>
<tr>
<td>Propane</td>
<td>C₃H₈</td>
<td>44.09</td>
<td>\frac{8 \text{ mol}(1.008 \text{ g/mol})}{44.09 \text{ g}} \times 100 = 18.29% H</td>
</tr>
<tr>
<td>Benzene</td>
<td>C₆H₆</td>
<td>78.11</td>
<td>\frac{6 \text{ mol}(1.008 \text{ g/mol})}{78.11 \text{ g}} \times 100 = 7.743% H</td>
</tr>
</tbody>
</table>
Ethanol       C₂H₅OH             46.07 \[\frac{6 \text{ mol}(1.008 \text{ g/mol})}{46.07 \text{ g}}(100) = 13.13\% \text{ H}\]
Cetyl palmitate       C₃₂H₆₄O₂              480.83 \[\frac{64 \text{ mol}(1.008 \text{ g/mol})}{480.83 \text{ g}}(100) = 13.42\% \text{ H}\]

The hydrogen percentage decreases in the following order:
Ethane > Propane > Cetyl palmitate > Ethanol > Benzene

3.102 Plan: First determine the empirical formula. Assume 100 grams of sample, and then the moles of each element may be found by dividing the mass of each element by its molar mass. Divide each of the moles by the smallest value, and convert to whole numbers to get the empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution:
Assume a 100 g sample and convert the masses to moles by dividing by the molar mass:

Moles of C = \[\left(\frac{68.2 \text{ parts C by mass}}{100 \text{ parts by mass}}\right)\left(\frac{1 \text{ mol C}}{12.01 \text{ g C}}\right) = 5.6786 \text{ mol C}\]
Moles of H = \[\left(\frac{6.86 \text{ parts H by mass}}{100 \text{ parts by mass}}\right)\left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 6.8056 \text{ mol H}\]
Moles of N = \[\left(\frac{15.9 \text{ parts N by mass}}{100 \text{ parts by mass}}\right)\left(\frac{1 \text{ mol N}}{14.01 \text{ g N}}\right) = 1.1349 \text{ mol N}\]
Moles of O = \[\left(\frac{9.08 \text{ parts O by mass}}{100 \text{ parts by mass}}\right)\left(\frac{1 \text{ mol O}}{16.00 \text{ g O}}\right) = 0.5675 \text{ mol O}\]

Preliminary formula is \(\text{C}_{5.6786}\text{H}_{6.8056}\text{N}_{1.1349}\text{O}_{0.5675}\)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\text{C}_{\frac{5.6786}{0.5675}}\text{H}_{\frac{6.8056}{0.5675}}\text{N}_{\frac{1.1349}{0.5675}}\text{O}_{\frac{0.5675}{0.5675}} \rightarrow \text{C}_{10}\text{H}_{12}\text{N}_{2}\text{O}
\]

The empirical formula is \(\text{C}_{10}\text{H}_{12}\text{N}_{2}\text{O}\).

Formula mass of empirical formula = \(10(12.01 \text{ g/mol C}) + 12(1.008 \text{ g/mol H}) + 2(14.01 \text{ g/mol N}) + 1(16.00 \text{ g/mol O}) = 176.22 \text{ g/mol}\)

Whole-number multiple = \(\frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \frac{176 \text{ g/mol}}{176.22 \text{ g/mol}} = 1\)

The empirical formula mass and the molar mass are the same, thus, the molecular and empirical formulas are the same. The molecular formula is \(\text{C}_{10}\text{H}_{12}\text{N}_{2}\text{O}\).

3.103 a) Before 1961 the amu was defined as 1/16 the average mass of an atom of naturally occurring oxygen and Avogadro’s number was fixed by the definition of the mole as about 16 g of oxygen. After 1961, Avogadro’s number was fixed by the definition of the mole as exactly 12 g of carbon-12. So Avogadro’s number changed to a smaller value in 1961. However, the difference is quite small.
b) Yes, the definition of the mole changed.
c) Yes, the mass of a mole of substance changed.
d) Avogadro’s number to 3 significant figures will not change since the difference in values is so small.

3.104 Plan: The names must first be converted to chemical formulas. Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used. Remember that oxygen, chlorine, and hydrogen are diatomic.

Solution:
a) All of the substances are gases.
\[\text{H}_2\text{S(g)} + \text{O}_2(g) \xrightarrow{\Delta} \text{SO}_2(g) + \text{H}_2\text{O(g)}\]
There are 2 O atoms in O2 on the left and 3 O atoms in SO2 and H2O on the right; place a coefficient of 2 in front of H2O on the right and a coefficient of 2 in front of O2 on the left for a total of 4 oxygen atoms on each side:

\[ \text{H}_2\text{S}(g) + 2\text{O}_2(g) \xrightarrow{\Delta} \text{SO}_2(g) + 2\text{H}_2\text{O}(g) \]

Now the 4 H atoms in 2H2O on the right require a coefficient of 2 in front of H2S on the left:

\[ 2\text{H}_2\text{S}(g) + 2\text{O}_2(g) \xrightarrow{\Delta} \text{SO}_2(g) + 2\text{H}_2\text{O}(g) \]

The 2 S atoms in 2H2S on the left require a coefficient of 2 in front of SO2 on the right:

\[ 2\text{H}_2\text{S}(g) + 2\text{O}_2(g) \xrightarrow{\Delta} 2\text{SO}_2(g) + 2\text{H}_2\text{O}(g) \]

Now the O atoms are no longer balanced; the 6 O atoms on the right (4 in 2SO2 and 2 in 2H2O) require a coefficient of 6 in front of O2 on the left:

\[ 2\text{H}_2\text{S}(g) + 3\text{O}_2(g) \xrightarrow{\Delta} 2\text{SO}_2(g) + 2\text{H}_2\text{O}(g) \]

b) All of the substances are solid (crystalline).

\[ \text{KClO}_3(s) \xrightarrow{\Delta} \text{KCl}(s) + \text{KClO}_4(s) \]

There are 3 O atoms in KClO3 on the left and 4 O atoms in KClO4 on the right. Place a coefficient of 4 in front of KClO3 and a coefficient of 3 in front of KClO4 for a total of 12 O atoms on each side. The K and Cl atoms are balanced with 4 K atoms and 4 Cl atoms on each side:

\[ 4\text{KClO}_3(s) \xrightarrow{\Delta} \text{KCl}(s) + 3\text{KClO}_4(s) \]

c) Hydrogen and water vapor are gases; iron and iron(III) oxide are solids.

\[ \text{H}_2(g) + \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + \text{H}_2\text{O}(g) \]

The 2 Fe atoms in Fe2O3 on the left require a coefficient of 2 in front of Fe on the right:

\[ \text{H}_2(g) + \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + \text{H}_2\text{O}(g) \]

The 3 O atoms in Fe2O3 on the left require a coefficient of 3 in front of H2O on the right:

\[ \text{H}_2(g) + \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + 3\text{H}_2\text{O}(g) \]

The 6 H atoms in 3H2O on the right require a coefficient of 3 in front of H2 on the left:

\[ 3\text{H}_2(g) + \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + 3\text{H}_2\text{O}(g) \]

d) All of the substances are gases; combustion required oxygen as a reactant.

\[ \text{C}_2\text{H}_6(g) + \text{O}_2(g) \xrightarrow{\Delta} 2\text{CO}_2(g) + \text{H}_2\text{O}(g) \]

The 2 C atoms in C2H6 on the left require a coefficient of 2 in front of CO2 on the right:

\[ \text{C}_2\text{H}_6(g) + \text{O}_2(g) \xrightarrow{\Delta} 2\text{CO}_2(g) + \text{H}_2\text{O}(g) \]

The 6 H atoms in C2H6 on the left require a coefficient of 3 in front of H2O on the right:

\[ \text{C}_2\text{H}_6(g) + \text{O}_2(g) \xrightarrow{\Delta} 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g) \]

The 7 O atoms on the right (4 in 2CO2 and 3 in 3H2O) require a coefficient of 7/2 in front of O2 on the left:

\[ \text{C}_2\text{H}_6(g) + \frac{7}{2}\text{O}_2(g) \xrightarrow{\Delta} 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g) \]

Double all coefficients to get whole number coefficients:

\[ 2\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \xrightarrow{\Delta} 4\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \]

e) Iron(II) chloride and iron(III) fluoride are solids and the other substances are gases.

\[ \text{FeCl}_2(s) + \text{ClF}_3(g) \rightarrow \text{FeF}_3(s) + \text{Cl}_2(g) \]

There are 3 Cl atoms on the left (2 in FeCl2 and 1 in ClF3) and 2 Cl atoms in Cl2 on the right. Place a coefficient of 2 in front of Cl2 and a coefficient of 2 in front of ClF3 on the left for a total of 4 Cl atoms on each side:

\[ \text{FeCl}_2(s) + 2\text{ClF}_3(g) \rightarrow \text{FeF}_3(s) + 2\text{Cl}_2(g) \]

The 6 F atoms in 2ClF3 require a coefficient of 2 in front of FeF3 on the right:

\[ \text{FeCl}_2(s) + 2\text{ClF}_3(g) \rightarrow 2\text{FeF}_3(s) + 2\text{Cl}_2(g) \]

The 2 Fe atoms in FeF3 on the right require a coefficient of 2 in front of FeCl2 on the left:

\[ 2\text{FeCl}_2(s) + 2\text{ClF}_3(g) \rightarrow 2\text{FeF}_3(s) + 2\text{Cl}_2(g) \]

Now the Cl atoms are not balanced with 6 on the left (4 in 2FeCl2 and 2 in 2ClF3) and 4 in 2Cl2 on the right; place a coefficient of 3 in front of Cl2 on the right:

\[ 2\text{FeCl}_2(s) + 2\text{ClF}_3(g) \rightarrow 2\text{FeF}_3(s) + 3\text{Cl}_2(g) \]

3.105 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of
the carbon present in the sample is found in the carbon of CO2, and all of the hydrogen present in the sample is found in the hydrogen of H2O. Convert the mass of CO2 to moles and use the ratio between CO2 and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H2O. Divide the moles of C and H by the smaller value to convert to whole numbers to get the empirical formula.

Solution:

Isobutylene + O2 → CO2 + H2O

Moles of C = \( \frac{(2.657 \times 10^{-2} \text{ mol CO}_2) \times (44.01 \text{ g CO}_2)}{1 \text{ mol CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.06037 \text{ mol C} \)

Moles of H = \( \frac{(1.089 \times 10^{-2} \text{ mol H}_2\text{O}) \times (18.02 \text{ g H}_2\text{O})}{2 \text{ mol H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.1209 \text{ mol H} \)

Preliminary formula = C\(^{0.06037}\)H\(^{0.1209}\)

Converting to integer subscripts (dividing all by the smallest subscript):

\[
\frac{0.06037}{0.06037} \quad \frac{0.1209}{0.06037} \rightarrow C^1H^2
\]

This gives an empirical formula of \( \text{CH}_2 \).

### 3.106 Plan:

The key to solving this problem is determining the overall balanced equation. Each individual step must be set up and balanced first. The separate equations can then be combined to get the overall equation. The mass of iron is converted to moles of iron by dividing by the molar mass, and the mole ratio from the balanced equation is used to find the moles and then the mass of CO required to produce that number of moles of iron.

Solution:

a) In the first step, ferric oxide (ferric denotes Fe\(^{3+}\)) reacts with carbon monoxide to form Fe\(_3\)O\(_4\) and carbon dioxide:

\[
3\text{Fe}_2\text{O}_3(s) + \text{CO}(g) \rightarrow 2\text{Fe}_3\text{O}_4(s) + \text{CO}_2(g)
\]

In the second step, \( \text{Fe}_3\text{O}_4(s) + \text{CO}(g) \rightarrow 3\text{FeO}(s) + \text{CO}_2(g) \) (2)

In the third step, ferrous oxide reacts with more carbon monoxide to form molten iron:

\[
\text{FeO}(s) + \text{CO}(g) \rightarrow \text{Fe(l)} + \text{CO}_2(g)
\]

Common factors are needed to allow these equations to be combined. The intermediate products are \( \text{Fe}_3\text{O}_4 \) and \( \text{FeO} \), so multiply equation (2) by 2 to cancel \( \text{Fe}_3\text{O}_4 \) and equation (3) by 6 to cancel \( \text{FeO} \):

\[
3\text{Fe}_2\text{O}_3(s) + 9\text{CO}(g) \rightarrow 6\text{Fe}(l) + 9\text{CO}_2(g)
\]

Then divide by 3 to obtain the smallest integer coefficients:

\[
\text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g)
\]

b) A metric ton is equal to 1000 kg.

Converting 45.0 metric tons of Fe to mass in grams:

\[
\text{Mass (g) of Fe = } \left(45.0 \text{ ton Fe}\right) \left(\frac{10^3 \text{ kg}}{1 \text{ ton}}\right) \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right) = 8.05730 \times 10^5 \text{ mol Fe}
\]

\[
\text{Mass (g) of CO = } \left(8.05730 \times 10^5 \text{ mol Fe}\right) \left(\frac{2 \text{ mol CO}}{1 \text{ mol Fe}}\right) \left(\frac{28.01 \text{ g CO}}{1 \text{ mol CO}}\right) = 3.39 \times 10^7 \text{ g CO}
\]
The balanced chemical equation is:
$$\text{C}_7\text{H}_8(\ell) + 9\text{O}_2(\text{g}) \rightarrow 7\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$$

a) Moles of C\textsubscript{7}H\textsubscript{8} = \[ \left( \frac{0.867 \text{ g C}_7\text{H}_8}{92.13 \text{ g C}_7\text{H}_8} \right) \left( \frac{1 \text{ mol C}_7\text{H}_8}{1 \text{ mL C}_7\text{H}_8} \right) = 0.01882123 \text{ mol C}_7\text{H}_8 \]

Mass (g) oxygen = \[ \left( \frac{9 \text{ mol O}_2}{1 \text{ mol C}_7\text{H}_8} \right) \left( \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 54.20514 = 54.2 \text{ g O}_2 \]

b) Total moles of gas = \[ \left( \frac{11 \text{ mol product gas}}{1 \text{ mol C}_7\text{H}_8} \right) \left( \frac{1 \text{ mol C}_7\text{H}_8}{18.82123 \text{ mol C}_7\text{H}_8} \right) = 2.07034 = 2.07 \text{ mol of gas} \]

The 11 mol of gas is an exact, not measured, number, so it does not affect the significant figures.

c) Moles of H\textsubscript{2}O = \[ \left( \frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_7\text{H}_8} \right) \left( \frac{0.1882123 \text{ mol C}_7\text{H}_8}{1 \text{ mol C}_7\text{H}_8} \right) = 0.7528492 \text{ mol H}_2\text{O} \]

Molecules of H\textsubscript{2}O = \[ \left( \frac{6.022 \times 10^{23} \text{ H}_2\text{O molecules}}{1 \text{ mol H}_2\text{O}} \right) \left( \frac{0.7528492 \text{ mol H}_2\text{O}}{1 \text{ mol C}_7\text{H}_8} \right) = 4.53366 \times 10^{23} = 4.53 \times 10^{23} \text{ molecules H}_2\text{O} \]

3.108 Plan: If 100.0 g of dinitrogen tetroxide reacts with 100.0 g of hydrazine (N\textsubscript{2}H\textsubscript{4}), what is the theoretical yield of nitrogen if no side reaction takes place? First, we need to identify the limiting reactant. To determine which reactant is limiting, calculate the amount of nitrogen formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of nitrogen. Then determine the amount of limiting reactant required to produce 10.0 grams of NO. Reduce the amount of limiting reactant by the amount used to produce NO. The reduced amount of limiting reactant is then used to calculate an “actual yield.” The “actual” and theoretical yields will give the maximum percent yield.

Solution:
The balanced reaction is 2N\textsubscript{2}H\textsubscript{4}(l) + N\textsubscript{2}O\textsubscript{4}(l) → 3N\textsubscript{2}(g) + 4H\textsubscript{2}O(g)

Determining the limiting reactant:
Finding the moles of N\textsubscript{2} from the amount of N\textsubscript{2}O\textsubscript{4} (if N\textsubscript{2}H\textsubscript{4} is limiting):
Moles of N\textsubscript{2} from N\textsubscript{2}O\textsubscript{4} = \[ \left( \frac{1 \text{ mol N}_2 \text{O}_4}{92.02 \text{ g N}_2 \text{O}_4} \right) \left( \frac{3 \text{ mol N}_2}{1 \text{ mol N}_2 \text{O}_4} \right) = 3.26016 \text{ mol N}_2 \]

Finding the moles of N\textsubscript{2} from the amount of N\textsubscript{2}H\textsubscript{4} (if N\textsubscript{2}O\textsubscript{4} is limiting):
N\textsubscript{2} from N\textsubscript{2}H\textsubscript{4} = \[ \left( \frac{1 \text{ mol N}_2 \text{H}_4}{32.05 \text{ g N}_2 \text{H}_4} \right) \left( \frac{3 \text{ mol N}_2}{2 \text{ mol N}_2 \text{H}_4} \right) = 4.68019 \text{ mol N}_2 \]

N\textsubscript{2}O\textsubscript{4} is the limiting reactant.

Theoretical yield of N\textsubscript{2} = \[ \left( \frac{100.0 \text{ g N}_2 \text{O}_4}{92.02 \text{ g N}_2 \text{O}_4} \right) \left( \frac{3 \text{ mol N}_2}{1 \text{ mol N}_2 \text{O}_4} \right) \left( \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} \right) = 91.3497 \text{ g N}_2 \]

How much of the limiting reactant is used to produce 10.0 g NO?
N\textsubscript{2}H\textsubscript{4}(l) + 2N\textsubscript{2}O\textsubscript{4}(l) → 6NO(g) + 2H\textsubscript{2}O(g)

Mass (g) of N\textsubscript{2}O\textsubscript{4} used = \[ \left( \frac{10.0 \text{ g NO}}{30.01 \text{ g NO}} \right) \left( \frac{2 \text{ mol N}_2 \text{O}_4}{6 \text{ mol NO}} \right) \left( \frac{92.02 \text{ g N}_2 \text{O}_4}{1 \text{ mol N}_2 \text{O}_4} \right) = 10.221 \text{ g N}_2 \text{O}_4 \]

Amount of N\textsubscript{2}O\textsubscript{4} available to produce N\textsubscript{2} = 100.0 g N\textsubscript{2}O\textsubscript{4} – mass of N\textsubscript{2}O\textsubscript{4} required to produce 10.0 g NO
= 100.0 g – 10.221 g = 89.779 g N\textsubscript{2}O\textsubscript{4}

Determine the “actual yield” of N\textsubscript{2} from 89.779 g N\textsubscript{2}O\textsubscript{4}:

“Actual yield” of N\textsubscript{2} = \[ \left( \frac{89.779 \text{ g N}_2 \text{O}_4}{92.02 \text{ g N}_2 \text{O}_4} \right) \left( \frac{3 \text{ mol N}_2}{1 \text{ mol N}_2 \text{O}_4} \right) \left( \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} \right) \]

3-63
3.109 Plan: Write a balanced chemical equation, using X to represent the halogen element. Convert the mass of strontium sulfate to moles and use the mole ratio in the balanced equation to find the moles of strontium halide required to produce that number of moles of strontium sulfate. Divide the mass of strontium halide by the moles of strontium halide to determine its molar mass. Subtract out the molar mass of strontium to obtain the molar mass of X₂. Divide the molar mass of X₂ by 2 to determine the molar mass of X. The molar mass of X can be used to identify X, using the periodic table and examining Group 7A(17). Once the identity of X is known, the formula of the strontium halide can be written.

Solution:

\[
\text{SrX}_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{SrSO}_4(s) + 2\text{HX}(aq)
\]

\[
\begin{align*}
\text{Moles of } \text{SrSO}_4 &= \left( \frac{0.755 \text{ g } \text{SrSO}_4}{183.68 \text{ g } \text{SrSO}_4} \right) = 0.0041104 \text{ mol } \text{SrSO}_4 \\
\text{Moles of } \text{SrX}_2 &= \left( \frac{0.0041104 \text{ mol } \text{SrSO}_4}{1 \text{ mol } \text{SrSO}_4} \right) = 0.0041104 \text{ mol } \text{SrX}_2
\end{align*}
\]

The 0.652 g sample of SrX₂ = 0.0041104 mol SrX₂

\[
\begin{align*}
\text{SrX}_2 &= \left( \frac{0.652 \text{ g } \text{SrX}_2}{0.0041104 \text{ mol } \text{SrX}_2} \right) = 158.62 \text{ g/mol} = \text{molar mass} \\
\text{Molar mass of } X_2 &= \text{molar mass of } \text{SrX}_2 - \text{molar mass of } \text{Sr} = 158.62 \text{ g} - 87.62 \text{ g} = 71.00 \text{ g} = X_2 \\
\text{Molar mass of } X &= 71.00 \text{ g}/2 = 35.50 = 35.5 \text{ g/mol} = \text{Cl}
\end{align*}
\]

The original halide formula is SrCl₂.

3.110 Plan: Identify the product molecules and write the balanced equation. To determine the limiting reactant for part b), examine the product circle to see which reactant remains in excess and which reactant was totally consumed. For part c), use the mole ratios in the balanced equation to determine the number of moles of product formed by each reactant, assuming the other reactant is in excess. The reactant that produces fewer moles of product is the limiting reactant. Use the mole ratio between the two reactants to determine the moles of excess reactant required to react with the limiting reactant. The difference between the initial moles of excess reactant and the moles required for reaction is the moles of excess reactant that remain.

Solution:

a) The contents of the boxes give:

\[
\text{AB}_2 + \text{B}_2 \rightarrow \text{AB}_3
\]

Balancing the reaction gives:

\[
2\text{AB}_2 + \text{B}_2 \rightarrow 2\text{AB}_3
\]

b) Two B₂ molecules remain after reaction so B₂ is in excess. All of the AB₂ molecules have reacted so AB₂ is the limiting reactant.

c) Finding the moles of AB₁ from the moles of AB₂ (if B₂ is limiting):

\[
\text{Moles of } \text{AB}_3 \text{ from } \text{AB}_2 = \left( 5.0 \text{ mol } \text{AB}_2 \right) \left( \frac{2 \text{ mol } \text{AB}_3}{2 \text{ mol } \text{AB}_2} \right) = 5.0 \text{ mol } \text{AB}_3
\]

Finding the moles of AB₁ from the moles of B₂ (if AB₂ is limiting):

\[
\text{Moles of } \text{AB}_3 \text{ from } \text{B}_2 = \left( 3.0 \text{ mol } \text{B}_2 \right) \left( \frac{2 \text{ mol } \text{AB}_3}{1 \text{ mol } \text{B}_2} \right) = 6.0 \text{ mol } \text{AB}_3
\]

AB₂ is the limiting reagent and 5.0 mol of AB₃ is formed.

d) Moles of B₂ that react with 5.0 mol AB₂ = \left( 5.0 \text{ mol } \text{AB}_2 \right) \left( \frac{1 \text{ mol } \text{B}_2}{2 \text{ mol } \text{AB}_2} \right) = 2.5 \text{ mol } \text{B}_2

The unreacted B₂ is 3.0 mol – 2.5 mol = 0.5 mol B₂.

3.111 Plan: Write the formulas in the form CₓHᵧO₂. Reduce the formulas to obtain the empirical formulas.
Add the atomic masses in that empirical formula to obtain the molecular mass.

Solution:

Compound A: $\text{C}_4\text{H}_{10}\text{O}_2 = \text{C}_2\text{H}_5\text{O}$  
Compound B: $\text{C}_2\text{H}_4\text{O}$  
Compound C: $\text{C}_4\text{H}_8\text{O}_2 = \text{C}_2\text{H}_4\text{O}$  
Compound D: $\text{C}_6\text{H}_{12}\text{O}_3 = \text{C}_2\text{H}_4\text{O}$  
Compound E: $\text{C}_5\text{H}_8\text{O}_2$

Compounds B, C and D all have the same empirical formula, $\text{C}_2\text{H}_4\text{O}$. The molecular mass of this formula is $(2 \times 12.01 \text{ g/mol C}) + (4 \times 1.008 \text{ g/mol H}) + (1 \times 16.00 \text{ g/mol O}) = 44.05 \text{ g/mol}$.

Plan: Write the balanced chemical equation. Since quantities of two reactants are given, we must determine which is the limiting reactant. To determine which reactant is limiting, calculate the amount of product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent to determine the theoretical yield of product. The actual yield is given. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield.

Solution:

Determine the balanced chemical equation:

$$\text{ZrOCl}_2\cdot8\text{H}_2\text{O}(\text{s}) + 4\text{H}_2\text{C}_2\text{O}_4\cdot2\text{H}_2\text{O}(\text{s}) + 4\text{KOH}(\text{aq}) \rightarrow \text{K}_2\text{Zr(C}_2\text{O}_4)_3(\text{H}_2\text{C}_2\text{O}_4)\cdot\text{H}_2\text{O}(\text{s}) + 2\text{KCl}(\text{aq}) + 20\text{H}_2\text{O}(\text{l})$$

Determining the limiting reactant:

Finding the moles of product from the amount of ZrOCl$_2$•8H$_2$O (if H$_2$C$_2$O$_4$•2H$_2$O is limiting):

$$\text{Moles of product from ZrOCl}_2\cdot8\text{H}_2\text{O} = \left(\frac{1 \text{ mol ZrOCl}_2\cdot8\text{H}_2\text{O}}{322.25 \text{ g ZrOCl}_2\cdot8\text{H}_2\text{O}}\right) \left(\frac{1 \text{ mol product}}{1 \text{ mol ZrOCl}_2\cdot8\text{H}_2\text{O}}\right) = 0.00521334 \text{ mol product}$$

Finding the moles of product from the amount of H$_2$C$_2$O$_4$•2H$_2$O (if ZrOCl$_2$•8H$_2$O is limiting):

$$\text{Moles of product from H}_2\text{C}_2\text{O}_4\cdot2\text{H}_2\text{O} = \left(\frac{1 \text{ mol H}_2\text{C}_2\text{O}_4\cdot2\text{H}_2\text{O}}{126.07 \text{ g H}_2\text{C}_2\text{O}_4\cdot2\text{H}_2\text{O}}\right) \left(\frac{1 \text{ mol product}}{4 \text{ mol H}_2\text{C}_2\text{O}_4\cdot2\text{H}_2\text{O}}\right) = 0.0103117 \text{ mol product}$$

It is not necessary to find the moles of product from KOH because KOH is stated to be in excess. The ZrOCl$_2$•8H$_2$O is the limiting reactant, and will be used to calculate the theoretical yield:

$$\text{Mass (g) of product} = \left(\frac{0.00521334 \text{ mol product}}{1 \text{ mol product}}\right) \left(\frac{541.53 \text{ g product}}{1 \text{ mol product}}\right) = 2.82318 \text{ g product}$$

Calculating the percent yield:

$$\text{Percent yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \times 100\% = \left(\frac{1.25 \text{ g}}{2.82318 \text{ g}}\right) \times 100\% = 44.276 = 44.3\% \text{ yield}$$

Plan: Since 85% of ions in seawater are from NaCl, take 85% of the mass percent of dissolved ions (4.0%) to find the mass % of NaCl in part a). To find the mass % of Na$^+$ and Cl$^-$ individually in part b), use the ratio of the mass of the two ions to the mass of NaCl. To find the molarity in part c), use the mass of NaCl in 100 g of seawater; convert mass of NaCl to moles and mass of seawater to volume in liters, using the density. Molarity = moles of NaCl/L of seawater.

Solution:

a) $(4.0\% \text{ ions}) \left(\frac{85\% \text{ NaCl}}{100\% \text{ ions}}\right) = 3.4\% \text{ NaCl}$

b) $\text{% Na}^+ \text{ ions} = (3.4\% \text{ NaCl}) \left(\frac{22.99 \text{ g Na}^+}{58.44 \text{ g NaCl}}\right) = 1.3375 = 1.3\% \text{ Na}^+ \text{ ions}$

$\text{% Cl}^- \text{ ions} = (3.4\% \text{ NaCl}) \left(\frac{35.45 \text{ g Cl}^-}{58.44 \text{ g NaCl}}\right) = 2.062 = 2.1\% \text{ Cl}^- \text{ ions}$

3.114 a) False, a mole of one substance has the same number of units as a mole of any other substance. b) True
c) **False**, a limiting-reactant problem is present when the quantity of available material is given for more than one reactant.

d) **False**, The empirical and molecular formulas of a compound can be the same, but they are often different (they are different when the formula masses of the molecular and empirical formulas are different).

3.115 **Plan:** To convert mass to moles, divide the mass by the molar mass of the substance. To convert moles to mass, divide by the molar mass. To obtain number of particles, multiply moles by Avogadro’s number. Divide a number of particles by Avogadro’s number to obtain moles.

**Solution:**
a) Since 1 mole of any substance contains Avogadro’s number of entities, equal amounts of moles of various substances contain equal numbers of entities. The number of entities (O₃ molecules) in 0.4 mol of O₃ is equal to the number of entities (O atoms) in 0.4 mol of O atoms.

b) O₃ has a molar mass of 3(16.0 g/mol O) = 48.0 g/mol; O has a molar mass of 1(16.0 g/mol O) = 16.0 g/mol. Since O₃ has a larger molar mass than O, **0.4 mol of O₃** has a greater mass than 0.4 mol of O.

c) Moles of N₂O₄ = (4.0 g N₂O₄) \( \left( \frac{1 \text{ mol N}_2\text{O}_4}{92.02 \text{ g N}_2\text{O}_4} \right) = 0.043 \text{ mol N}_2\text{O}_4 \)

Moles of SO₂ = (3.3 g SO₂) \( \left( \frac{1 \text{ mol SO}_2}{64.06 \text{ g SO}_2} \right) = 0.052 \text{ mol SO}_2 \)

SO₂ is the larger quantity in terms of moles.

d) Mass (g) of C₂H₄ = (0.6 mol C₂H₄) \( \left( \frac{28.05 \text{ g C}_2\text{H}_4}{1 \text{ mol C}_2\text{H}_4} \right) = 17 \text{ g C}_2\text{H}_4 \)

Mass (g) of F₂ = (0.6 mol F₂) \( \left( \frac{38.00 \text{ g F}_2}{1 \text{ mol F}_2} \right) = 23 \text{ g F}_2 \)

F₂ is the greater quantity in terms of mass.

Note that if each of these values is properly rounded to one significant figure, the answers are identical.

e) Total moles of ions in 2.3 mol NaClO₃ = (2.3 mol NaClO₃) \( \left( \frac{2 \text{ mol ions}}{1 \text{ mol NaClO}_3} \right) = 4.6 \text{ mol ions} \)

Total moles of ions in 2.2 mol MgCl₂ = (2.2 mol MgCl₂) \( \left( \frac{3 \text{ mol ions}}{1 \text{ mol MgCl}_2} \right) = 6.6 \text{ mol ions} \)

MgCl₂ is the greater quantity in terms of total moles of ions.

f) The compound with the lower molar mass will have more molecules in a given mass. H₂O (18.02 g/mol) has a lower molar mass than H₂O₂ (34.02 g/mol). **1.0 g H₂O** has more molecules than 1.0 g H₂O₂.

g) Moles of NaBr = (0.500 L NaBr) \( \left( \frac{0.500 \text{ mol}}{1 \text{ L}} \right) = 0.250 \text{ mol NaBr} \)

Moles of Na⁺ = (0.250 mol NaBr) \( \left( \frac{1 \text{ mol Na}^+}{1 \text{ mol NaBr}} \right) = 0.250 \text{ mol Na}^+ \)

Moles of NaCl = (0.0146 kg NaCl) \( \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \right) = 0.250 \text{ mol NaCl} \)

Moles of Na⁺ = (0.250 mol NaCl) \( \left( \frac{2 \text{ mol ions}}{1 \text{ mol NaCl}} \right) = 0.250 \text{ mol Na}^+ \)

The two quantities are equal.

h) The heavier atoms, ⁵²³U, will give a greater total mass since there is an equal number of particles of both.

3.116 **Plan:** Write a balanced equation. The coefficients in the balanced equation give the number of molecules or moles of each reactant and product. Moles are converted to amount in grams by multiplying by the molar masses.

**Solution:**
P₄S₃(s) + 8O₂(g) → P₄O₁₀(s) + 3SO₂(g)
a) 1 molecule of \( \text{P}_4\text{S}_3 \) reacts with 8 molecules of \( \text{O}_2 \) to produce 1 molecule of \( \text{P}_4\text{O}_{10} \) and 3 molecules of \( \text{SO}_2 \).
b) 1 mol of \( \text{P}_4\text{S}_3 \) reacts with 8 mol of \( \text{O}_2 \) to produce 1 mol of \( \text{P}_4\text{O}_{10} \) and 3 mol of \( \text{SO}_2 \).
c) 220.09 g of \( \text{P}_4\text{S}_3 \) react with \( 8(32.00 \text{ g/mol O}) = 256.00 \text{ g of O}_2 \) to produce 283.88 g of \( \text{P}_4\text{O}_{10} \) and \( 3(64.06 \text{ g/mol SO}_2) = 192.18 \text{ g of SO}_2 \).

3.117 Plan: Write a balanced equation. Use the actual yield (105 kg) and the percent yield (98.8%) to find the theoretical yield of hydrogen. Use the mole ratio between hydrogen and water in the balanced equation to obtain the amount of hydrogen required to produce that theoretical yield of water.

Solution:
The balanced equation is \( 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) \)

% yield = \( \frac{\text{actual Yield}}{\text{theoretical Yield}} \times 100\% \)

Theoretical yield (g) of \( \text{H}_2\text{O} \) = \( \frac{\text{actual yield}}{\% \text{ yield}} \) \( (100) = \frac{105 \text{ kg}}{98.8\%} = 106.2753 \text{ kg H}_2\text{O} \)

Mass (g) of \( \text{H}_2 \) = \( \left(106.2753 \text{ kg H}_2\text{O}\right) \left(\frac{2 \text{ mol H}_2}{18.02 \text{ g H}_2\text{O}}\right) \left(\frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2}\right) \left(\frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2}\right) \)

= 1.1886 \times 10^4 = 1.19 \times 10^4 \text{ g H}_2

3.118 Plan: For part a), convert 3.0 kg/s to kg/y and multiply by 50 billion years. For part b), the mass of nitrogen is 75.5% of the mass of the atmosphere. Divide the mass of \( \text{N}_2 \) in grams by its molar mass to obtain moles.

Solution:
a) Mass of atmosphere = \( (3.0 \text{ kg/s} \times 60 \text{ s/1 min} \times 60 \text{ min/1 h} \times 24 \text{ h/1 d} \times 365.25 \text{ d/1 y} \times 50 \times 10^9 \text{ y}) \)

= 4.73364 \times 10^{18} = 4.7 \times 10^{18} \text{ kg atmosphere}

b) Mass (g) of \( \text{N}_2 \) = \( \frac{(75.5\% \text{ atmosphere})(103 \text{ g \text{N}_2}}{1 \text{ kg atmosphere}}) = 3.573898 \times 10^{21} \text{ g N}_2

Moles of \( \text{N}_2 \) = \( \frac{(3.573898 \times 10^{21} \text{ g N}_2)}{28.02 \text{ g N}_2} \) = 1.27548 \times 10^{20} = 1.3 \times 10^{20} \text{ mol N}_2

3.119 Plan: Divide the given mass of a substance by its molar mass to obtain moles; multiply the given moles of a substance by its molar mass to obtain mass in grams. Number of particles is obtained by multiplying an amount in moles by Avogadro’s number. Density is used to convert mass to volume.

Solution:
a) Moles of \( \text{NH}_4\text{Br} \) = \( \frac{(0.588 \text{ g NH}_4\text{Br})(1 \text{ mol NH}_4\text{Br})}{97.94 \text{ g NH}_4\text{Br}} \) = 0.0060037 = 0.00600 \text{ mol NH}_4\text{Br}

b) Moles of \( \text{KNO}_3 \) = \( \frac{(88.5 \text{ g KNO}_3)(1 \text{ mol KNO}_3)}{101.11 \text{ g KNO}_3} \) = 0.875284 \text{ mol KNO}_3

Number of \( \text{K}^+ \) ions = \( \frac{(0.875284 \text{ mol KNO}_3)(1 \text{ mol K}^+)}{1 \text{ mol KNO}_3} \left(\frac{6.022 \times 10^{23} \text{ K}^+ \text{ ions}}{1 \text{ mol K}^+}\right) \)

= 5.27096 \times 10^{23} = 5.27 \times 10^{23} \text{ K}^+ \text{ ions}

c) Mass (g) of \( \text{C}_3\text{H}_8\text{O}_3 \) = \( \frac{(5.85 \text{ mol C}_3\text{H}_8\text{O}_3)(92.09 \text{ g C}_3\text{H}_8\text{O}_3)}{1 \text{ mol C}_3\text{H}_8\text{O}_3} \) = 538.7265 = 539 \text{ g C}_3\text{H}_8\text{O}_3
d) Mass (g) of CHCl₃ = \left( \frac{3.85 \text{ mol CHCl}_3}{1 \text{ mol CHCl}_3} \right) \left( \frac{119.37 \text{ g CHCl}_3}{1 \text{ mol CHCl}_3} \right) = 340.2045 \text{ g CHCl}_3

Volume (mL) of CHCl₃ = \left( \frac{340.2045 \text{ g CHCl}_3}{1.48 \text{ g CHCl}_3} \right) = 229.868 = 230 \text{ mL CHCl}_3

e) Moles of Na⁺ = \left( \frac{2.11 \text{ mol Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3} \right) \left( \frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{CO}_3} \right) = 4.22 \text{ mol Na}^+

Number of Na⁺ = \left( \frac{4.22 \text{ mol Na}^+}{1 \text{ mol Na}^+} \right) \left( \frac{6.022 \times 10^{23} \text{ Na}^+ \text{ ions}}{1 \text{ mol Na}^+} \right) = 2.54 \times 10^{24} \text{ Na}^+ \text{ ions}

f) Moles of Cd atoms = \left( \frac{25.0 \mu \text{g Cd}}{1 \mu \text{g}} \right) \left( \frac{1 \text{ mol Cd}}{112.4 \text{ g Cd}} \right) = 2.224199 \times 10^{-7} \text{ mol Cd atoms}

Number of Cd atoms = \left( \frac{2.224199 \times 10^{-7} \text{ mol Cd}}{1 \text{ mol Cd}} \right) \left( \frac{6.022 \times 10^{23} \text{ Cd atoms}}{1 \text{ mol Cd}} \right) = 1.34 \times 10^{17} \text{ Cd atoms}

g) Number of F atoms = \left( \frac{0.0015 \text{ mol F}}{1 \text{ mol F}} \right) \left( \frac{2 \text{ mol F}}{1 \text{ mol F}} \right) \left( \frac{6.022 \times 10^{23} \text{ F atoms}}{1 \text{ mol F}} \right) = 1.8 \times 10^{21} \text{ F atoms}

3.120 Neither A nor B has any XY₃ molecules. Both C and D have XY₃ molecules. D shows both XY₃ and XY molecules. Only C has a single XY₃ product, thus the answer is C.

3.121 Plan: Deal with the methane and propane separately, and combine the results. Balanced equations are needed for each hydrocarbon. The total mass and the percentages will give the mass of each hydrocarbon. The mass of each hydrocarbon is changed to moles, and through the balanced chemical equation the amount of CO₂ produced by each gas may be found. Summing the amounts of CO₂ gives the total from the mixture. For part b), let x and 252 – x represent the masses of CH₄ and C₃H₈, respectively.

Solution:
a) The balanced chemical equations are:
   Methane: \(\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l)\)
   Propane: \(\text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(l)\)

Mass (g) of CO₂ from each:
   Methane: \(\left( 200.0 \text{ g mixture} \right) \left( \frac{25.0\%}{100\%} \right) \left( \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \right) \left( \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \right) \left( \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 137.188 \text{ g CO}_2\)
   Propane: \(\left( 200.0 \text{ g mixture} \right) \left( \frac{75.0\%}{100\%} \right) \left( \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \right) \left( \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \right) \left( \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 449.183 \text{ g CO}_2\)

Total CO₂ = 137.188 g + 449.183 g = 586.371 g = 586 g CO₂

b) Since the mass of CH₄ + the mass of C₃H₈ = 252 g, let \(x\) = mass of CH₄ in the mixture and 252 – \(x\) = mass of C₃H₈ in the mixture. Use mole ratios to calculate the amount of CO₂ formed from x amount of CH₄ and the amount of CO₂ formed from 252 – \(x\) amount of C₃H₈. The total mass of CO₂ produced = 748 g.

The total moles of CO₂ produced = \(\left( 748 \text{ g CO}_2 \right) \left( \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) = 16.996 \text{ mol CO}_2\)

\[16.996 \text{ mol CO}_2 = \left( \frac{x \text{ g CH}_4}{16.04 \text{ g CH}_4} \right) \left( \frac{1 \text{ mol CO}_2}{16.04 \text{ g CH}_4} \right) + \left( \frac{(252 - x) \text{ g C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \right) \left( \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \right)\]

\[16.996 \text{ mol CO}_2 = \frac{x}{16.04} \text{ mol CO}_2 + \frac{3(252 - x)}{44.09} \text{ mol CO}_2\]
16.996 mol CO₂ = \frac{x}{16.04} \text{ mol CO₂} + \frac{756 - 3x}{44.09} \text{ mol CO₂}

16.996 mol CO₂ = 0.06234x \text{ mol CO₂} + (17.147 - 0.06804x \text{ mol CO₂})

x = 26.49 \text{ g CH₄}

252 - x = 252 \text{ g} - 26.49 \text{ g} = 225.51 \text{ g C₃H₈}

Mass % CH₄ = \frac{26.49 \text{ g CH₄}}{252 \text{ g mixture}} (100) = \text{ 10.5% CH₄}

Mass % C₃H₈ = \frac{225.51 \text{ g C₃H₈}}{252 \text{ g mixture}} (100) = \text{ 89.5% C₃H₈}

3.122 Plan: If we assume a 100-gram sample of fertilizer, then the 30:10:10 percentages become the masses, in grams, of N, P₂O₅, and K₂O. These masses may be changed to moles of substance, and then to moles of each element. To get the desired x:y:1.0 ratio, divide the moles of each element by the moles of potassium.

Solution:
A 100-gram sample of 30:10:10 fertilizer contains 30 g N, 10 g P₂O₅, and 10 g K₂O.

Moles of N = \left( \frac{30 \text{ g N}}{14.01 \text{ g N}} \right) = 2.1413 \text{ mol N}

Moles of P = \left( \frac{10 \text{ g P₂O₅}}{141.94 \text{ g P₂O₅}} \right) = 0.14090 \text{ mol P}

Moles of K = \left( \frac{10 \text{ g K₂O}}{94.20 \text{ g K₂O}} \right) = 0.21231 \text{ mol K}

This gives a N:P:K ratio of 2.1413:0.14090:0.21231
The ratio must be divided by the moles of K and rounded.
\frac{2.1413 \text{ mol N}}{0.21231} = 10.086
\frac{0.14090 \text{ mol P}}{0.21231} = 0.66365
\frac{0.21231 \text{ mol K}}{0.21231} = 1

0.21231:0.66365:1.000 or 10:0.66:1.0

3.123 Plan: If we assume a 100-gram sample of fertilizer, then the 10:10:10 percentages become the masses, in grams, of N, P₂O₅, and K₂O. These masses may be changed to moles of substance, and then to moles of each element. Use the mole ratio between N and ammonium sulfate, P and ammonium hydrogen phosphate, and K and potassium chloride to find the mass of each compound required to provide the needed amount of the respective element. Divide the mass of each compound by the total mass of sample, 100 g, and multiply by 100 for mass %.

Solution:
Assume a 100 g sample. 10:10:10 indicates 10 g N, 10 g P₂O₅ and 10 g K₂O.

Moles of N = \left( \frac{10 \text{ g N}}{14.01 \text{ g N}} \right) = 0.713776 \text{ mol N}

Moles of P = \left( \frac{10 \text{ g P₂O₅}}{141.94 \text{ g P₂O₅}} \right) = 0.14090 \text{ mol P}

Moles of K = \left( \frac{10 \text{ g K₂O}}{94.20 \text{ g K₂O}} \right) = 0.21231 \text{ mol K}

To obtain 0.713776 mol N from (NH₄)₂SO₄:
\left( \frac{0.713776 \text{ mol N}}{2 \text{ mol N}} \right) = 47.1592 \text{ g (NH₄)₂SO₄}

Mass % N = \frac{47.1592 \text{ g (NH₄)₂SO₄}}{100 \text{ g mixture}} (100) = \text{ 47.16% N}

Mass % P = \frac{13.214 \text{ g (NH₄)₂SO₄}}{100 \text{ g mixture}} (100) = \text{ 13.21% P}

Mass % K = \frac{17.147 \text{ g (NH₄)₂SO₄}}{100 \text{ g mixture}} (100) = \text{ 17.15% K}
Mass % \((\text{NH}_4)_2\text{SO}_4 = \frac{\text{mass of (NH}_4)_2\text{SO}_4}{\text{mass of mixture}} \times 100 \) = 47.1592% = \textbf{47.2\% (NH}_4)_2\text{SO}_4

To obtain 0.14090 mol P from \((\text{NH}_4)_2\text{HPO}_4:

\frac{0.14090 \text{ mol P}}{1 \text{ mol (NH}_4)_2\text{HPO}_4} \times \frac{132.06 \text{ g (NH}_4)_2\text{HPO}_4}{1 \text{ mol (NH}_4)_2\text{HPO}_4} = 18.6073 \text{ g (NH}_4)_2\text{HPO}_4

Mass % \((\text{NH}_4)_2\text{HPO}_4 = \frac{\text{mass of (NH}_4)_2\text{HPO}_4}{\text{mass of mixture}} \times 100 \) = 18.6073% = \textbf{18.6\% (NH}_4)_2\text{HPO}_4

To obtain 0.21231 mol K from KCl:

\frac{0.21231 \text{ mol K}}{1 \text{ mol KCl}} \times \frac{74.55 \text{ g KCl}}{1 \text{ mol KCl}} = 15.8277 \text{ g KCl}

Mass % KCl = \frac{\text{mass of KCl}}{\text{mass of mixture}} \times 100 \) = 15.8277% = \textbf{15.8\% KCl}

3.124 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO₂, and all of the hydrogen present in the sample is found in the hydrogen of H₂O. Convert the mass of CO₂ to moles and use the ratio between CO₂ and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H₂O. Subtracting the masses of C and H from the mass of the sample gives the mass of Fe. Convert the mass of Fe to moles of Fe. Take the moles of C, H, and Fe and divide by the smallest value to convert to whole numbers to get the empirical formula.

Solution:

Ferrocene + \( ?\text{O}_2(\text{g}) \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)

0.9437 g 2.233 g 0.457 g

Moles of C = \( \frac{2.233 \text{ g CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} = 0.050738 \text{ mol C} \)

Mass (g) of C = \( \frac{0.050738 \text{ mol C}}{1 \text{ mol C}} \times 12.01 \text{ g C} = 0.60936 \text{ g C} \)

Moles of H = \( \frac{0.457 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.050721 \text{ mol H} \)

Mass (g) of H = \( \frac{0.050721 \text{ mol H}}{1 \text{ mol H}} \times 1.008 \text{ g H} = 0.051127 \text{ g H} \)

Mass (g) of Fe = Sample mass – (mass of C + mass of H) = 0.9437 g – (0.60936 g C + 0.051127 g H) = 0.283213 g Fe

Moles of Fe = \( \frac{0.283213 \text{ g Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}}{1 \text{ mol Fe}} = 0.005071 \text{ mol Fe} \)

Preliminary formula = \( \text{C}_{0.050738}\text{H}_{0.050721}\text{Fe}_{0.005071} \)

Converting to integer subscripts (dividing all by the smallest subscript):

\( \text{C}_{0.050738}\text{H}_{0.050721}\text{Fe}_{0.005071} \rightarrow \text{C}_{10}\text{H}_{10}\text{Fe}_1 \)

Empirical formula = \( \text{C}_{10}\text{H}_{10}\text{Fe} \)

3.125 Plan: Assume 100 grams of mixture. This means the mass of each compound, in grams, is the same as its percentage. Find the mass of C from CO and from CO₂ and add these masses together. For mass %, divide the total mass of C by the mass of the mixture (100 g) and multiply by 100.

Solution:

100 g of mixture = 35 g CO and 65 g CO₂.
Mass (g) of C from CO = (35.0 g CO) \left( \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \right) \left( \frac{1 \text{ mol C}}{1 \text{ mol CO}} \right) \left( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 15.007 \text{ g C}

Mass (g) of C from CO₂ = (65.0 g CO₂) \left( \frac{1 \text{ mol CO₂}}{44.01 \text{ g CO₂}} \right) \left( \frac{1 \text{ mol C}}{1 \text{ mol CO₂}} \right) \left( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 17.738 \text{ g C}

Total mass (g) of C = 15.007 g + 17.738 g = 32.745 g

Mass % C = \frac{\text{mass of C}}{\text{mass of mixture}} \times 100 = \frac{32.745 \text{ g C}}{100 \text{ g mixture}} \times 100 = 32.745\% C

3.126 Plan: Write a balanced equation for the reaction. Count the molecules of each reactant to obtain the moles of each reactant present. Use the mole ratios in the equation to calculate the amount of product formed. Only 87.0% of the calculated amount of product actually forms, so the actual yield is 87.0% of the theoretical yield.

Solution:
The balanced equation is SiH₄ + N₂F₄ → SiF₄ + N₂ + 2H₂.

Moles of SiH₄ = \left( \frac{1.25 \times 10⁻² \text{ mol}}{1 \text{ molecule}} \right) = 0.0375 \text{ mol SiH₄}

Moles of N₂F₄ = \left( \frac{1.25 \times 10⁻² \text{ mol}}{1 \text{ molecule}} \right) = 0.0375 \text{ mol N₂F₄}

Since there is an equal amount of each reactant and the ratio between each reactant and SiF₄ is 1:1, neither reactant is in excess and either may be used to calculate the amount of SiF₄ produced.

Mass (g) of SiF₄ = \left( \frac{0.0375 \text{ mol SiH₄}}{1 \text{ mol SiH₄}} \right) \left( \frac{104.09 \text{ g SiF₄}}{1 \text{ mol SiF₄}} \right) = 3.903375 \text{ g SiF₄}

% yield = \left( \frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100

Actual yield (g) of SiF₄ = \left( \frac{87\%}{100\%} \right) \left( 3.903375 \text{ g SiF₄} \right) = 3.3959 = 3.4 \text{ g SiF₄}

3.127 Plan: Determine the molecular formula from the figure. Once the molecular formula is known, use the periodic table to determine the molar mass. Convert the volume of lemon juice in part b) from qt to mL and use the density to convert from mL to mass in g. Take 6.82% of that mass to find the mass of citric acid and use the molar mass to convert to moles.

Solution:
a) The formula of citric acid obtained by counting the number of carbon atoms, oxygen atoms, and hydrogen atoms is C₆H₈O₇.

Molar mass = (6 \times 12.01 \text{ g/mol C}) + (8 \times 1.008 \text{ g/mol H}) + (7 \times 16.00 \text{ g/mol O}) = 192.12 \text{ g/mol}

b) Converting volume of lemon juice in qt to mL:

Volume (mL) of lemon juice = (1.50 qt) \left( \frac{1 \text{ L}}{1.057 \text{ qt}} \right) \left( \frac{1 \text{ mL}}{10⁻³ \text{ L}} \right) = 1419.111 \text{ mL}

Converting volume to mass in grams:

Mass (g) of lemon juice = (1419.111 mL) \left( \frac{1.09 \text{ g}}{\text{ mL}} \right) = 1546.831 \text{ g lemon juice}

Mass (g) of C₆H₈O₇ = (1546.831 g lemon juice) \left( \frac{6.82\% \text{ C₆H₈O₇}}{100\% \text{ lemon juice}} \right) = 105.494 \text{ g C₆H₈O₇}

Moles of C₆H₈O₇ = (105.494 g C₆H₈O₇) \left( \frac{1 \text{ mol C₆H₈O₇}}{192.12 \text{ g C₆H₈O₇}} \right) = 0.549104 = 0.549 \text{ mol C₆H₈O₇}
3.128 \textbf{Plan:} Determine the formulas of each reactant and product, then balance the individual equations. Remember that nitrogen and oxygen are diatomic. Combine the three smaller equations to give the overall equation, where some substances serve as intermediates and will cancel. Use the mole ratio between nitrogen and nitric acid in the overall equation to find the moles and then mass of nitric acid produced. The amount of nitrogen in metric tons must be converted to mass in grams to convert the mass of nitrogen to moles.

\textbf{Solution:}

\textit{a) Nitrogen and oxygen combine to form nitrogen monoxide:}
\[
N_2(g) + O_2(g) \rightarrow 2\text{NO}(g)
\]
\textit{Nitrogen monoxide reacts with oxygen to form nitrogen dioxide:}
\[
2\text{NO}(g) + O_2(g) \rightarrow 2\text{NO}_2(g)
\]
\textit{Nitrogen dioxide combines with water to form nitric acid and nitrogen monoxide:}
\[
3\text{NO}_2(g) + \text{H}_2\text{O}(g) \rightarrow 2\text{HNO}_3(\text{aq}) + \text{NO}(g)
\]

\textit{b) Combining the reactions may involve adjusting the equations in various ways to cancel out as many materials as possible other than the reactants added and the desired products.}
\[
2 \times (N_2(g) + O_2(g) \rightarrow 2\text{NO}(g)) = 2N_2(g) + 2O_2(g) \rightarrow 4\text{NO}(g)
\]
\[
3 \times (2\text{NO}(g) + O_2(g) \rightarrow 2\text{NO}_2(g)) = 6\text{NO}(g) + 3O_2(g) \rightarrow 6\text{NO}_2(g)
\]
\[
2 \times (3\text{NO}_2(g) + \text{H}_2\text{O}(g) \rightarrow 2\text{HNO}_3(\text{aq}) + \text{NO}(g)) = 6\text{NO}_2(g) + 2\text{H}_2\text{O}(g) \rightarrow 4\text{HNO}_3(\text{aq}) + 2\text{NO}(g)
\]

Multiplying the above equations as shown results in the 6 moles of NO on each side and the 6 moles of NO2 on each side canceling. Adding the equations gives:
\[
2N_2(g) + 5O_2(g) + 2\text{H}_2\text{O}(g) \rightarrow 4\text{HNO}_3(\text{aq})
\]

\textit{c) Mass (g) of N}_2 = \left[\frac{332}{10} \text{ kg} \right] \left[\frac{10 \text{ g}}{1 \text{ kg}}\right] = 1.35 \times 10^9 \text{ g N}_2
\]
\[
\text{Moles of N}_2 = \left[\frac{1 \text{ mol N}}{28.02 \text{ g N}_2}\right] \left[\frac{10^3 \text{ g N}_2}{1 \text{ kg}}\right] = 4.817987 \times 10^7 \text{ mol N}_2
\]
\[
\text{Mass (g) of HNO}_3 = \left[\frac{4 \text{ mol HNO}_3}{2 \text{ mol N}_2}\right] \left[\frac{63.02 \text{ g HNO}_3}{1 \text{ mol HNO}_3}\right] = 6.072591 \times 10^9 \text{ g HNO}_3
\]
\[
\text{Metric tons HNO}_3 = \left[\frac{6.072591 \times 10^9 \text{ g HNO}_3}{1 \text{ kg}}\right] \left[\frac{1 \text{ t}}{10^3 \text{ kg}}\right] = 6.072591 \times 10^3 = 6.07 \times 10^3 \text{ metric tons HNO}_3
\]

3.129 \textbf{Plan:} Write and balance the chemical reaction. Use the mole ratio to find the amount of product that should be produced and take 66\% of that amount to obtain the actual yield.

\textbf{Solution:}

\[2\text{NO}(g) + O_2(g) \rightarrow 2\text{NO}_2(g)\]

With 6 molecules of NO and 3 molecules of O2 reacting, 6 molecules of NO2 can be produced. If the reaction only has a 66\% yield, then (0.66)(6) = 4 molecules of NO2 will be produced. \textbf{Circle A} shows the formation of 4 molecules of NO2. Circle B also shows the formation of 4 molecules of NO2 but also has 2 unreacted molecules of NO and 1 unreacted molecule of O2. Since neither reactant is limiting, there will be no unreacted reactant remaining after the reaction is over.

3.130 \textbf{Plan:} For parts \textit{a) and b), convert the masses to moles. Take the moles and divide by the smallest value to convert to whole numbers to get the empirical formula. For part \textit{c), write the two balanced equations and use two equations as shown.}

\textbf{Solution:}

\textit{a) Moles of Pt} = \left[\frac{0.327 \text{ g Pt}}{195.1 \text{ g Pt}}\right] = 0.001676 \text{ mol Pt}
\[
\text{Mass (g) of F} = \text{mass of product} – \text{mass of Pt} = 0.519 \text{ g} – 0.327 \text{ g} = 0.192 \text{ g F}
\]
\[
\text{Moles of F} = \left[\frac{0.192 \text{ g F}}{19.00 \text{ g F}}\right] = 0.010105 \text{ mol F}
\]
\[
\text{Preliminary formula} = \text{Pt}_{0.001676}\text{F}_{0.010105}
\]
Converting to integer subscripts (dividing all by the smallest subscript):
\[
\frac{Pt_{0.001676}F_{0.001010}}{0.001676} \rightarrow PtF_6
\]
Empirical formula = \textbf{PtF}_6

b) Moles of PtF\(_6\) = \(\frac{0.265 \text{ g PtF}_6}{309.1 \text{ g PtF}_6} \times \frac{1 \text{ mol PtF}_6}{309.1 \text{ g PtF}_6}\) = 0.0008576 mol PtF\(_6\)

Mass of Xe = mass of product – mass of Xe = 0.378 g – 0.265 g = 0.113 g Xe

Moles of Xe = \(\frac{0.113 \text{ g Xe}}{131.3 \text{ g Xe}} \times \frac{1 \text{ mol Xe}}{131.3 \text{ g Xe}}\) = 0.0008606 mol Xe

Preliminary formula = \textbf{Xe}_{0.0008606}(\textbf{PtF}_6)_{0.0008576}

Converting to integer subscripts (dividing all by the smallest subscript):
\[
\frac{0.0008576}{0.0008606} \rightarrow \frac{0.0008576}{0.0008576}
\]
Empirical formula = \textbf{XePtF}_6

c) This problem can be solved as a system of two equations and two unknowns.

The two equations are:  The two unknowns are:
\[
\text{Xe}(g) + 2\text{F}_2(g) \rightarrow \text{XeF}_4(s) \quad x = \text{mol XeF}_4 \text{ produced}
\]
\[
\text{Xe}(g) + 3\text{F}_2(g) \rightarrow \text{XeF}_6(s) \quad y = \text{mol XeF}_6 \text{ produced}
\]
Moles of Xe consumed = 1.85x10\(^{-4}\) mol present – 9.00x10\(^{-6}\) mol excess = 1.76x10\(^{-4}\) mol Xe

Then  \(x + y = 1.76x10^{-4}\) mol Xe consumed
\[
2x + 3y = 5.00x10^{-4}\) mol \text{F}_2 \text{ consumed}
\]

Solve for \(x\) using the first equation and substitute the value of \(x\) into the second equation:
\[
x = 1.76x10^{-4} - y
\]
\[
2(1.76x10^{-4} - y) + 3y = 5.00x10^{-4}
\]
\[
3.52x10^{-4} - 2y + 3y = 5.00x10^{-4}
\]
\[
y = (5.00x10^{-4}) - (3.52x10^{-4}) = 1.48x10^{-4}\) mol XeF\(_6\)
\[
x = (1.76x10^{-4}) - (1.48x10^{-4}) = 2.8x10^{-5}\) mol XeF\(_4\)

Converting moles of each product to grams using the molar masses:
\[
\text{Mass (g) of XeF}_4 = \left(2.8x10^{-5}\) \text{ mol XeF}_4\right) \left(\frac{207.3 \text{ g XeF}_4}{1 \text{ mol XeF}_4}\right) = 5.8044x10^{-3}\) g XeF\(_4\)
\]
\[
\text{Mass (g) of XeF}_6 = \left(1.48x10^{-4}\) \text{ mol XeF}_6\right) \left(\frac{245.3 \text{ g XeF}_6}{1 \text{ mol XeF}_6}\right) = 3.63044x10^{-2}\) g XeF\(_6\)
\]

Calculate the percent of each compound using the total weight of the products:
\[
\left(5.8044 x 10^{-3} + 3.63044x10^{-2}\right) \text{g} = 0.0421088\) g
\]
\[
\text{Mass % XeF}_4 = \frac{\text{mass of XeF}_4}{\text{total mass}} \times 100 = \frac{5.8044x10^{-3}\) g XeF\(_4\)}{0.0421088\) g} \times 100 = 13.784 = 14% XeF\(_4\)
\]
\[
\text{Mass % XeF}_6 = \frac{\text{mass of XeF}_6}{\text{total mass}} \times 100 = \frac{3.63044x10^{-2}\) g XeF\(_6\)}{0.0421088\) g} \times 100 = 86.2157 = 86.2% XeF\(_6\)
\]

3.131 Plan: Use the mass percent to find the mass of heme in the sample; use the molar mass to convert the mass of heme to moles. Then find the mass of Fe in the sample by using the mole ratio between heme and iron. The mass of hemin is found by using the mole ratio between heme and hemoglobin.

Solution:

a) Mass (g) of heme = \(0.65\) g hemoglobin \(\left(\frac{6.0\% \text{ heme}}{100\% \text{ hemoglobin}}\right) = 0.039\) g heme

b) Moles of heme = \(0.039\) g heme \(\left(\frac{1 \text{ mol heme}}{616.49 \text{ g heme}}\right) = 6.32614\times10^{-5} = 6.3\times10^{-5}\) mol heme
c) Mass (g) of Fe = \( \left( 6.32614 \times 10^{-5} \text{ mol heme} \right) \left( \frac{1 \text{ mol Fe}}{1 \text{ mol heme}} \right) \left( \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \right) = 3.5331 \times 10^{-3} = 3.5 \times 10^{-3} \text{ g Fe} \)

d) Mass (g) of hemin = \( \left( 6.32614 \times 10^{-5} \text{ mol heme} \right) \left( \frac{1 \text{ mol hemin}}{1 \text{ mol heme}} \right) \left( \frac{651.94 \text{ g hemin}}{1 \text{ mol hemin}} \right) = 4.1243 \times 10^{-2} = 0.41 \text{ g hemin} \)

3.132 Plan: Find the Mn:O ratio in the two oxides. Write two equations to solve simultaneously; one equation shows that the sum of the ratio of Mn in the two oxides will equal the ratio of Mn in the sample and the other equation shows that the total amount of oxide in the sample is the sum of the amounts of the two oxides. The two equations will give the mole ratio of the two oxides. Convert moles of each oxide to mass to obtain the mass ratio of the two oxides from which the mass % of each can be calculated. Use that mass % of each to find the mass of each in the sample. For part b), the moles of Mn\(^{3+}\) come from the Mn\(_2\)O\(_3\) and the moles of Mn\(^{2+}\) come from the MnO.

Solution:

Mn:O ratio:

In sample: 1.00:1.42 or 0.704
In braunite: 2.00:3.00 or 0.667
In manganosite: 1.00:1.00 or 1.00

a) The total amount of ore is equal to the amount of braunite (B) + the amount of manganosite (M).

\( B + M = 1.00 \)
\( M = 1.00 - B \)

The amount of Mn is dependent on the sample’s composition.

\( M(1.00) + B(0.667) = 0.704 \)
\( (1.00 - B)(1.00) + B(0.667) = 0.704 \)
\( 1.00 - 1.00B + 0.667B = 0.704 \)
\( 0.296 = 0.333B \)
\( B = 0.888889 \text{ mol braunite} \)
\( M = 1.00 - B = 1.00 - 0.888889 = 0.111111 \text{ mol manganosite} \)

Mass (g) of braunite = \( (0.888889 \text{ mol}) \left( \frac{157.88 \text{ g}}{1 \text{ mol}} \right) = 140.338 \text{ g braunite} \)

Mass (g) of manganosite = \( (0.111111 \text{ mol}) \left( \frac{70.94 \text{ g}}{1 \text{ mol}} \right) = 7.88221 \text{ g manganosite} \)

There are 140.338 g of braunite for every 7.88221 g of manganosite. Finding mass % of each:

Mass % braunite = \( \frac{\text{mass of braunite}}{\text{mass of braunite} + \text{manganosite}} \times 100 = \frac{140.338 \text{ g}}{140.338 + 7.88221 \text{ g}} \times 100 = 94.6821\% \)

Mass % manganosite = \( \frac{\text{mass of manganosite}}{\text{mass of braunite} + \text{manganosite}} \times 100 = \frac{7.88221 \text{ g}}{140.338 + 7.88221 \text{ g}} \times 100 \)

= 5.3179%

In the 542.3 g sample:

Mass (g) of braunite = \( (542.3 \text{ g sample}) \left( \frac{94.6821 \text{ braunite}}{100\% \text{ sample}} \right) = 513.461 = 513 \text{ g braunite} \)

Mass (g) of manganosite = \( (542.3 \text{ g sample}) \left( \frac{5.3179\% \text{ manganosite}}{100\% \text{ sample}} \right) = 28.839 = 28.8 \text{ g manganosite} \)

b) Each mole of braunite, Mn\(_2\)O\(_3\), contains 2 moles of Mn\(^{3+}\) while each mole of manganosite, MnO, contains 1 mole of Mn\(^{2+}\).

Moles of Mn\(^{3+}\) = \( 2(0.888889 \text{ mol braunite}) = 1.777778 \text{ mol Mn}^{3+} \)
Moles of Mn\(^{2+}\) = \( 1(0.111111 \text{ mol manganosite}) = 0.111111 \text{ mol Mn}^{2+} \)

\( \frac{\text{Mn}^{3+}:\text{Mn}^{2+}}{0.111111 \text{ mol Mn}^{2+}} = 16.000 = 16.0 \)
Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of nitrogen present. Multiply the number of moles of nitrogen by its molar mass to find the total mass of nitrogen in 1 mole of compound. Mass percent = \( \frac{\text{total mass of element}}{\text{molar mass of compound}} \times 100 \). For part b), convert mass of ornithine to moles, use the mole ratio between ornithine and urea to find the moles of urea, and then use the ratio between moles of urea and nitrogen to find the moles and mass of nitrogen produced.

Solution:

a) Urea: \( \text{CH}_4\text{N}_2\text{O}, \ M = 60.06 \text{ g/mol} \)
There are 2 moles of N in 1 mole of \( \text{CH}_4\text{N}_2\text{O} \).

Mass (g) of N = \( \left( 2 \text{ mol N} \right) \left( \frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 28.02 \text{ g N} \)

\[ \text{Mass percent} = \frac{\text{total mass N}}{\text{molar mass of compound}} \times 100 = \frac{28.02 \text{ g N}}{60.06 \text{ g CH}_4\text{N}_2\text{O}} = 46.65\% \text{ N in urea} \]

Arginine: \( \text{C}_6\text{H}_{15}\text{N}_4\text{O}_2, \ M = 175.22 \text{ g/mol} \)
There are 4 moles of N in 1 mole of \( \text{C}_6\text{H}_{15}\text{N}_4\text{O}_2 \).

Mass (g) of N = \( \left( 4 \text{ mol N} \right) \left( \frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 56.04 \text{ g N} \)

\[ \text{Mass percent} = \frac{\text{total mass N}}{\text{molar mass of compound}} \times 100 = \frac{56.04 \text{ g N}}{175.22 \text{ g C}_6\text{H}_{15}\text{N}_4\text{O}_2} = 31.98\% \text{ N in arginine} \]

Ornithine: \( \text{C}_5\text{H}_{13}\text{N}_2\text{O}_2, \ M = 133.17 \text{ g/mol} \)
There are 2 moles of N in 1 mole of \( \text{C}_5\text{H}_{13}\text{N}_2\text{O}_2 \).

Mass (g) of N = \( \left( 2 \text{ mol N} \right) \left( \frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 28.02 \text{ g N} \)

\[ \text{Mass percent} = \frac{\text{total mass N}}{\text{molar mass of compound}} \times 100 = \frac{28.02 \text{ g N}}{133.17 \text{ g C}_5\text{H}_{13}\text{N}_2\text{O}_2} = 21.04\% \text{ N in ornithine} \]

b) Moles of urea = \( \left( 135.2 \text{ g C}_5\text{H}_{13}\text{N}_2\text{O}_2 \right) \left( \frac{1 \text{ mol C}_5\text{H}_{13}\text{N}_2\text{O}_2}{133.17 \text{ g C}_5\text{H}_{13}\text{N}_2\text{O}_2} \right) \left( \frac{1 \text{ mol CH}_4\text{N}_2\text{O}}{1 \text{ mol C}_5\text{H}_{13}\text{N}_2\text{O}_2} \right) = 1.015244 \text{ mol urea} \)

Mass (g) of nitrogen = \( \left( 1.015244 \text{ mol CH}_4\text{N}_2\text{O} \right) \left( \frac{2 \text{ mol N}}{1 \text{ mol CH}_4\text{N}_2\text{O}} \right) \left( \frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 28.447 = 28.45 \text{ g N} \)

Plan: To determine which reactant is limiting, calculate the amount of aspirin formed from each reactant, assuming an excess of the other reactant. Use the density of acetic anhydride to determine the amount of this reactant in grams. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of aspirin. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield. Use the formula for percent atom economy to determine that quantity.

Solution:

a) Finding the moles of aspirin from the moles of \( \text{C}_7\text{H}_6\text{O}_3 \) (if \( \text{(CH}_3\text{CO})_2\text{O} \) is limiting):

Moles of aspirin from \( \text{C}_7\text{H}_6\text{O}_3 = \left( 3.077 \text{ g C}_7\text{H}_6\text{O}_3 \right) \left( \frac{1 \text{ mol C}_7\text{H}_6\text{O}_3}{138.12 \text{ g C}_7\text{H}_6\text{O}_3} \right) \left( \frac{1 \text{ mol C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_7\text{H}_6\text{O}_3} \right) = 0.0222777 \text{ mol C}_9\text{H}_8\text{O}_4 \)

Finding the moles of aspirin from the moles of \( \text{C}_4\text{H}_6\text{O}_3 \) (if \( \text{C}_7\text{H}_6\text{O}_3 \) is limiting):

Mass (g) of \( \text{(CH}_3\text{CO})_2\text{O} = \left( 5.50 \text{ mL (CH}_3\text{CO})_2\text{O} \right) \left( \frac{1.080 \text{ g (CH}_3\text{CO})_2\text{O}}{1 \text{ mL}} \right) = 5.94 \text{ g (CH}_3\text{CO})_2\text{O} \)
Moles of aspirin from \((CH_3CO)_2O\) = \(\frac{1 \ mol \ (CH_3CO)_2O}{102.09 \ g \ (CH_3CO)_2O} \times \frac{1 \ mol \ C_7H_6O_3}{1 \ mol \ (CH_3CO)_2O}\)

\[= 0.058183955 \ mol \ C_7H_6O_3\]

The limiting reactant is \(C_7H_6O_3\).

b) First, calculate the theoretical yield from the limiting reagent:

Mass (g) of \(C_9H_8O_4\) = \(\frac{1 \ mol \ C_9H_8O_4}{1 \ mol \ C_9H_8O_4} \times \frac{180.15 \ g \ C_9H_8O_4}{1 \ mol \ C_9H_8O_4}\)

\[= 4.01333 \ g \ C_9H_8O_4\]

Percent yield = \(\frac{\text{actual Yield}}{\text{theoretical Yield}}\) x 100% = \(\frac{3.281 \ g}{4.01333 \ g}\) x 100% = 81.7526 = 81.75% yield

c) % atom economy = \(\frac{\text{no. of moles} \times \text{molar mass of desired products}}{\text{sum of (no. of moles} \times \text{molar mass}) \text{ for all products}}\) x 100%

\[= \frac{(1 \ mol)(180.15 \ g/mol)}{(1 \ mol)(180.15 \ g/mol) + (1 \ mol)(60.05 \ g/mol)}\] x 100% = 75.00% atom economy

3.135 Plan: Determine the molar mass of each product and use the equation for percent atom economy.
Solution:
Molar masses of product: \(N_2H_4\): 32.05 g/mol \(NaCl\): 58.44 g/mol \(H_2O\): 18.02 g/mol

% atom economy = \(\frac{\text{no. of moles} \times \text{molar mass of desired products}}{\text{sum of (no. of moles} \times \text{molar mass}) \text{ for all products}}\) x 100%

\[= \frac{1 \ mol \times 32.05 \ g/mol}{(1 \ mol \times 32.05 \ g/mol) + (1 \ mol \times 58.44 \ g/mol) + (1 \ mol \times 18.02 \ g/mol)}\] x 100% = 29.5364 = 29.54% atom economy

3.136 Plan: Use the mass percent to obtain the mass of lead(II) chromate needed. Use the mole ratio between \(PbCrO_4\) and \(K_2CrO_4\) and then the mole ratio between \(K_2CrO_4\) and \(FeCr_2O_4\) to find the moles and then mass of \(FeCr_2O_4\) needed to produce that mass of \(PbCrO_4\).
Solution:

Mass (g) of \(PbCrO_4\) in 1 kg of paint = \(1 \ kg \ paint \times \frac{10^3 \ g}{kg} \times \frac{0.511\%}{100\%}\) = 5.11 g \(PbCrO_4\)

Moles of \(K_2CrO_4\) = \(\frac{5.11 \ g \ PbCrO_4}{323.2 \ g \ PbCrO_4} \times \frac{1 \ mol \ K_2CrO_4}{1 \ mol \ PbCrO_4}\) = 0.015811 mol \(K_2CrO_4\)

Mass of \(FeCr_2O_4\) = \(\frac{0.015811 \ mol \ K_2CrO_4}{8 \ mol \ K_2CrO_4} \times \frac{223.85 \ g \ FeCr_2O_4}{1 \ mol \ FeCr_2O_4}\) = 1.769606 = 1.77 g \(FeCr_2O_4\)

3.137 Plan: Convert the mass of ethanol to moles, and use the mole ratio between ethanol and diethyl ether to determine the theoretical yield of diethyl ether. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield. The difference between the actual and theoretical yields is related to the quantity of ethanol that did not produce diethyl ether, forty-five percent of which produces ethylene instead. Use the mole ratio between ethanol and ethylene to find the mass of ethylene produced by the forty-five percent of ethanol that did not produce diethyl ether.
Solution:
a) The determination of the theoretical yield:
Mass (g) of diethyl ether = \( \frac{1 \text{ mol CH}_2\text{CH}_2\text{OH}}{46.07 \text{ g CH}_2\text{CH}_2\text{OH}} \times \frac{1 \text{ mol CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3}{2 \text{ mol CH}_2\text{CH}_2\text{OH}} \times \frac{74.12 \text{ g CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3}{1 \text{ mol CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3} \) 
= 40.2214 g diethyl ether

Determining the percent yield:

\[
\text{Percent yield} = \left( \frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\% = \left( \frac{35.9}{40.2214} \right) \times 100\% = 89.2560 = 89.3\% \text{ yield}
\]

b) To determine the amount of ethanol not producing diethyl ether, we will use the difference between the theoretical yield and actual yield to determine the amount of diethyl ether that did not form and hence, the amount of ethanol that did not produce the desired product. Forty-five percent of this amount will be used to determine the amount of ethylene formed.

Mass difference = theoretical yield – actual yield = 40.2214 g – 35.9 g = 4.3214 g diethyl ether that did not form

\[
\text{Mass (g) of ethanol not producing diethyl ether} = \left( \frac{1 \text{ mol CH}_2\text{CH}_2\text{OCH}_2\text{CH}_3}{74.12 \text{ g CH}_2\text{CH}_2\text{OCH}_2\text{CH}_3} \right) \times \left( \frac{2 \text{ mol CH}_3\text{CH}_2\text{OH}}{1 \text{ mol CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3} \right) \times \left( \frac{46.07 \text{ g CH}_3\text{CH}_2\text{OH}}{1 \text{ mol CH}_3\text{CH}_2\text{OH}} \right) \times \left( \frac{45.0\%}{100\%} \right) = 5.37202 \text{ g ethanol}
\]

Mass of ethanol producing ethylene = \( \left( \frac{1 \text{ mol CH}_3\text{CH}_2\text{OH}}{46.07 \text{ g CH}_3\text{CH}_2\text{OH}} \right) \times \left( \frac{1 \text{ mol C}_2\text{H}_4}{1 \text{ mol CH}_3\text{CH}_2\text{OH}} \right) \times \left( \frac{28.05 \text{ g C}_2\text{H}_4}{1 \text{ mol C}_2\text{H}_4} \right) \)

= 1.47185 = 1.47 g ethylene

3.138 Plan: First balance the given chemical equation. To determine which reactant is limiting, calculate the amount of ZnS formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of ZnS. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield. For part b), determine the mass of Zn that does not produce ZnS; use that amount of zinc and the mole ratio between Zn and ZnO in that reaction to determine the mass of ZnO produced. Find the moles of S8 in the reactant and the moles of S8 in the product ZnS. The difference between these two amounts is the moles of S8 in SO2.

Solution:

a) The balanced equation is 8Zn(s) + S8(s) → 8ZnS(s).

Finding the limiting reagent:

Finding the moles of ZnS from the moles of Zn (if S8 is limiting):

Moles of ZnS from Zn = \( \left( \frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} \right) \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \times \frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} \) = 1.27256 mol ZnS

Finding the moles of ZnS from the moles of S8 (if Zn is limiting):

Moles of ZnS from S8 = \( \left( \frac{1 \text{ mol S8}}{256.48 \text{ g S8}} \right) \times \frac{8 \text{ mol ZnS}}{1 \text{ mol S8}} \times \frac{1 \text{ mol S8}}{256.48 \text{ g S8}} \times \frac{8 \text{ mol ZnS}}{1 \text{ mol S8}} \) = 1.6344 mol ZnS

The zinc will produce less zinc sulfide, thus, zinc is the limiting reactant and will first be used to determine the theoretical yield and then the percent yield.

Theoretical yield (g) of ZnS = \( \left( \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \right) \times \frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} \times \frac{97.44 \text{ g ZnS}}{1 \text{ mol ZnS}} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \) = 123.9983 g ZnS (unrounded)

Percent yield = \( \left( \frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\% = \left( \frac{104.4 \text{ g ZnS}}{123.9983 \text{ g ZnS}} \right) \times 100\% = 84.1947 = 84.2\% \text{ yield} \)
b) The reactions with oxygen are:

\[ 2\text{Zn(s)} + \text{O}_2(g) \rightarrow 2\text{ZnO(s)} \]

\[ \text{S}_8(s) + 8\text{O}_2(g) \rightarrow 8\text{SO}_2(g) \]

The theoretical yield indicates that 84.2% of the zinc produced zinc sulfide so \((100 - 84.2\%) = 15.8\%\) of the zinc became zinc oxide. This allows the calculation of the amount of zinc oxide formed.

\[
\text{Mass (g) of Zn that does not produce ZnS} = \left( \frac{15.8\% \times 83.2\text{ g Zn}}{100\%} \right) = 13.1456\text{ g ZnS}
\]

\[
\text{Mass (g) of ZnO} = \left( \frac{1\text{ mol Zn}}{2\text{ mol ZnO}} \times \frac{65.38\text{ g ZnO}}{13.1456\text{ g Zn}} \right) = 16.4\text{ g ZnO}
\]

The calculation is slightly different for the sulfur. We need to determine the amount of sulfur not in zinc sulfide. The sulfur not in the zinc sulfide must be in sulfur dioxide. The amount of sulfur not in zinc sulfide will be converted to the mass of sulfur dioxide.

\[
\text{Moles of S}_8\text{ in original S}_8\text{ reactant} = \left( \frac{52.4\text{ g S}_8}{256.48\text{ g S}_8} \right) = 0.204304\text{ mol S}_8
\]

\[
\text{Moles of S}_8\text{ in ZnS product} = \left( \frac{104.4\text{ g ZnS}}{97.44\text{ g ZnS}} \times \frac{1\text{ mol S}_8}{8\text{ mol ZnS}} \right) = 0.133928\text{ mol S}_8
\]

\[
\text{Moles of S}_8\text{ in SO}_2 = 0.204304 - 0.133928 = 0.070376\text{ mol S}_8
\]

\[
\text{Mass (g) of SO}_2 = \left( \frac{8\text{ mol SO}_2}{1\text{ mol SO}_2} \times \frac{64.06\text{ g SO}_2}{0.070376\text{ mol S}_8} \right) = 36.1\text{ g SO}_2
\]

3.139 Plan: For part a), use the given solubility of the salt to find the mass that is soluble in the given volume of water. For part b), convert the mass of dissolved salt in part a) to moles of salt and then to moles of cocaine and then to mass of cocaine. Use the solubility of cocaine to find the volume of water needed to dissolve this mass of cocaine.

Solution:

a) \[
\text{Mass (g) of dissolved salt} = \left( 50.0\text{ mL H}_2\text{O} \times \frac{10^{-3}\text{ L H}_2\text{O}}{1\text{ mL H}_2\text{O}} \times \frac{2.50\text{ kg salt}}{1\text{ L H}_2\text{O}} \times \frac{10^3\text{ g}}{1\text{ kg}} \right) = 125\text{ g salt}
\]

b) \[
\text{Moles of dissolved salt} = \left( 125\text{ g salt} \times \frac{1\text{ mol salt}}{339.81\text{ g salt}} \right) = 0.367853\text{ mol salt}
\]

\[
\text{Mass (g) cocaine} = \left( 0.367853\text{ mol salt} \times \frac{1\text{ mol cocaine}}{1\text{ mol salt}} \times \frac{303.35\text{ g cocaine}}{1\text{ mol cocaine}} \right) = 111.588\text{ g cocaine}
\]

\[
\text{Volume (L) of water needed to dissolve the cocaine} = \left( 111.588\text{ g cocaine} \times \frac{1\text{ L}}{1.70\text{ g cocaine}} \right) = 65.64\text{ L}
\]

Additional water needed = total volume needed – original volume of water

\[
= 65.64\text{ L} - 0.0500\text{ L} = 65.59 = 65.6\text{ L H}_2\text{O}
\]

3.140 Plan: Use the given values of \(x\) to find the molar mass of each compound. To determine which reactant is limiting, calculate the amount of either product formed from each reactant, assuming an excess of the other reactants. The reactant that produces the smallest amount of product is the limiting reagent. To find the mass of excess reactants, find the mass of each excess reactant that is required to react with the limiting reagent and subtract that mass from the starting mass.

a) \(x = 0\)

\[
\text{La}_2\text{Sr}_3\text{CuO}_4 = 2(138.9\text{ g/mol La}) + 0(87.62\text{ g/mol Sr}) + 1(63.55\text{ g/mol Cu}) + 4(16.00\text{ g/mol O}) = 405.4\text{ g/mol}
\]

\(x = 1\)

\[
\text{La}_2\text{Sr}_3\text{CuO}_4 = 1(138.9\text{ g/mol La}) + 1(87.62\text{ g/mol Sr}) + 1(63.55\text{ g/mol Cu}) + 4(16.00\text{ g/mol O}) = 354.1\text{ g/mol}
\]

\(x = 0.163\)

\[
\text{La}_{0.163}\text{Sr}_{0.163}\text{CuO}_4 = \text{La}_{1.837}\text{Sr}_{0.163}\text{CuO}_4
\]

\[
= 1.837(138.9\text{ g/mol La}) + 0.163(87.62\text{ g/mol Sr}) + 1(63.55\text{ g/mol Cu}) + 4(16.00\text{ g/mol O})
\]
= 397.0 g/mol

b) Assuming x grams to be the “equal” mass leads to:

Moles of product from BaCO₃ = \( \left( \frac{x \text{ g BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left( \frac{1 \text{ mol BaCO}_3}{4 \text{ mol BaCO}_3} \right) \left( \frac{2 \text{ mol YBa}_2\text{Cu}_3\text{O}_7}{6 \text{ mol YBa}_2\text{Cu}_3\text{O}_7} \right) \)

= 0.002534x mol product

Moles of product from CuO = \( \left( \frac{x \text{ g CuO}}{79.55 \text{ g CuO}} \right) \left( \frac{1 \text{ mol CuO}}{6 \text{ mol CuO}} \right) \left( \frac{2 \text{ mol YBa}_2\text{Cu}_3\text{O}_7}{1 \text{ mol YBa}_2\text{Cu}_3\text{O}_7} \right) \)

= 0.004190x mol product

Moles of product from Y₂O₃ = \( \left( \frac{x \text{ g Y}_2\text{O}_3}{225.82 \text{ g Y}_2\text{O}_3} \right) \left( \frac{1 \text{ mol Y}_2\text{O}_3}{1 \text{ mol Y}_2\text{O}_3} \right) \left( \frac{2 \text{ mol YBa}_2\text{Cu}_3\text{O}_7}{2 \text{ mol YBa}_2\text{Cu}_3\text{O}_7} \right) \)

= 0.008857x mol product

BaCO₃ is the limiting reactant.

c) These calculations are based on the limiting reactant.

BaCO₃ remaining = 0% (limiting reagent)

CuO remaining = x g CuO – \( \left( \frac{x \text{ g BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left( \frac{1 \text{ mol BaCO}_3}{4 \text{ mol BaCO}_3} \right) \left( \frac{6 \text{ mol CuO}}{1 \text{ mol CuO}} \right) \)

= 0.39521x g CuO

Percent CuO = \( \left( \frac{0.39521 \times x \text{ g}}{x \text{ g}} \right) \times 100\% = 39.521 = \text{39.52\% CuO remaining} \)

Y₂O₃ remaining = x g Y₂O₃ – \( \left( \frac{x \text{ g BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left( \frac{1 \text{ mol BaCO}_3}{4 \text{ mol BaCO}_3} \right) \left( \frac{225.82 \text{ g Y}_2\text{O}_3}{1 \text{ mol Y}_2\text{O}_3} \right) \)

= 0.713862x g Y₂O₃

Percent Y₂O₃ = \( \left( \frac{0.713862 \times x \text{ g}}{x \text{ g}} \right) \times 100\% = 71.3862 = \text{71.39\% Y}_2\text{O}_3 \text{ remaining} \)