

CHAPTER 3 STOICHIOMETRY OF FORMULAS AND EQUATIONS

- 3.1 Cl 35.45 amu \equiv 35.45 g/mol Cl
Mass Cl = (3 mol Cl) x (35.45 g Cl/1 mol Cl) = **106.4 g Cl**
Al 26.98 amu \equiv 26.98 g/mol Al
Mass Al = (2 mol Al) x (26.98 g Al/1 mol Al) = **53.96 g Al**

3.2 Plan: The formulas are based on the mole ratios of the constituents. Avogadro's number allows the change from moles to atoms.

Solution:

a) Moles of C atoms = (1 mol Sucrose) $\left(\frac{12 \text{ mol C}}{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}} \right) = \mathbf{12 \text{ mol C}}$

b) C atoms = (1 mol C₁₂H₂₂O₁₁) $\left(\frac{12 \text{ mol C}}{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}} \right) \left(\frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C}} \right) = \mathbf{7.226 \times 10^{24} \text{ C atoms}}$

3.3 "1 mol of nitrogen" could be interpreted as a mole of nitrogen atoms or a mole of nitrogen molecules. Specify which to avoid confusion. The same problem is possible with other diatomic or polyatomic molecules, e.g., Cl₂, Br₂, S₈, and P₄. For these elements, as for nitrogen, 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 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 P₄ molecules =

$$(2.5 \text{ g Ca}_3(\text{PO}_4)_2) \left(\frac{1 \text{ mol Ca}_3(\text{PO}_4)_2}{\text{Molar Mass}} \right) \left(\frac{2 \text{ mol P}}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \right) \left(\frac{1 \text{ mol P}_4}{4 \text{ mol P}} \right) \left(\frac{\text{Avogadro's Number P}_4 \text{ molecules}}{1 \text{ mol P}_4} \right)$$

3.7 Plan: It is possible to relate the relative atomic masses by counting the number of atoms.

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 six red balls means that they are that much lighter. Because the red ball is lighter, more red atoms are required to make 1 gram.

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 gram.

d) Both the left and right elements have the same number of atoms per mole. The number of atoms per mole (6.022×10^{23}) 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 mass of the element times the number present in the formula gives the molar mass.

Solution:

a) The molar mass, \mathcal{M} , is the sum of the atomic weights, expressed in g/mol:

$$\begin{aligned}\text{Sr} &= &&= 87.62 \text{ g Sr/mol Sr(OH)}_2 \\ \text{O} &= (2 \text{ mol O}) (16.00 \text{ g O/mol O}) &&= 32.00 \text{ g O/mol Sr(OH)}_2 \\ \text{H} &= (2 \text{ mol H}) (1.008 \text{ g H/mol H}) &&= 2.016 \text{ g H/mol Sr(OH)}_2 \\ &&&= 121.64 \text{ g/mol of Sr(OH)}_2\end{aligned}$$

b) $\mathcal{M} = (2 \text{ mol N}) (14.01 \text{ g N/mol N}) + (1 \text{ mol O}) (16.00 \text{ g O/mol O}) = \mathbf{44.02 \text{ g/mol of N}_2\text{O}}$

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

d) $\mathcal{M} = (2 \text{ mol Cr}) (52.00 \text{ g Cr/mol Cr}) + (3 \text{ mol O}) (16.00 \text{ g O/mol O}) = \mathbf{152.00 \text{ g/mol of Cr}_2\text{O}_3}$

3.9 a) $(\text{NH}_4)_3\text{PO}_4 = 3(14.01) + 12(1.008) + 30.97 + 4(16.00) = \mathbf{149.10 \text{ g/mol}}$

b) $\text{CH}_2\text{Cl}_2 = 12.01 + 2(1.008) + 2(35.45) = \mathbf{84.93 \text{ g/mol}}$

c) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 63.55 + 32.07 + 9(16.00) + 10(1.008) = \mathbf{249.70 \text{ g/mol}}$

d) $\text{BrF}_5 = 79.90 + 5(19.00) = \mathbf{174.90 \text{ g/mol}}$

3.10 Plan: Locate each of the elements on the periodic table and record its atomic mass. The mass of the element times the number present in the formula gives the molar mass.

Solution:

a) $\mathcal{M} = (1 \text{ mol Sn}) (118.7 \text{ g Sn/mol Sn}) + (2 \text{ mol O}) (16.00 \text{ g O/mol O}) = \mathbf{150.7 \text{ g/mol of SnO}_2}$

b) $\mathcal{M} = (1 \text{ mol Ba}) (137.3 \text{ g Ba/mol Ba}) + (2 \text{ mol F}) (19.00 \text{ g F/mol F}) = \mathbf{175.3 \text{ g/mol of BaF}_2}$

c) $\mathcal{M} = (2 \text{ mol Al}) (26.98 \text{ g Al/mol Al}) + (3 \text{ mol S}) (32.07 \text{ g S/mol S}) + (12 \text{ mol O}) (16.00 \text{ g O/mol O})$
 $= \mathbf{342.17 \text{ g/mol of Al}_2(\text{SO}_4)_3}$

d) $\mathcal{M} = (1 \text{ mol Mn}) (54.94 \text{ g Mn/mol Mn}) + (2 \text{ mol Cl}) (35.45 \text{ g Cl/mol Cl}) = \mathbf{125.84 \text{ g/mol of MnCl}_2}$

3.11 a) $\text{N}_2\text{O}_4 = 2(14.01) + 4(16.00) = \mathbf{92.02 \text{ g/mol}}$

b) $\text{C}_8\text{H}_{10} = 8(12.01) + 10(1.008) = \mathbf{106.16 \text{ g/mol}}$

c) $\text{MgSO}_4 \cdot 7\text{H}_2\text{O} = 24.31 + 32.07 + 11(16.00) + 14(1.008) = \mathbf{246.49 \text{ g/mol}}$

d) $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2 = 40.08 + 4(12.01) + 6(1.008) + 4(16.00) = \mathbf{158.17 \text{ g/mol}}$

3.12 Plan: The mass of a substance and its number of moles are related through the conversion factor of \mathcal{M} , the molar mass expressed in g/mol. The moles of a substance and the number of entities per mole are related by the conversion factor, Avogadro's number.

Solution:

a) \mathcal{M} of $\text{KMnO}_4 = 39.10 + 54.94 + (4 \times 16.00) = 158.04 \text{ g/mol of KMnO}_4$

$$\text{Mass of KMnO}_4 = (0.57 \text{ mol KMnO}_4) \left(\frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} \right) = 90.08 = \mathbf{9.0 \times 10^1 \text{ g KMnO}_4}$$

b) \mathcal{M} of $\text{Mg}(\text{NO}_3)_2 = 24.31 + (2 \times 14.01) + (6 \times 16.00) = 148.33 \text{ g/mol Mg}(\text{NO}_3)_2$

$$\begin{aligned}\text{Moles of O atoms} &= (8.18 \text{ g Mg}(\text{NO}_3)_2) \left(\frac{1 \text{ mol Mg}(\text{NO}_3)_2}{148.33 \text{ g Mg}(\text{NO}_3)_2} \right) \left(\frac{6 \text{ mol O atoms}}{1 \text{ mol Mg}(\text{NO}_3)_2} \right) \\ &= 0.33088 = \mathbf{0.331 \text{ mol O atoms}}\end{aligned}$$

c) \mathcal{M} of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 63.55 + 32.07 + (4 \times 16.00) + (5 \times 18.016) = 249.70 \text{ g/mol}$

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

$$\begin{aligned}\text{O atoms} &= (8.1 \times 10^{-3} \text{ g Cu Cmpd}) \left(\frac{1 \text{ mol Cu Cmpd}}{249.70 \text{ g Cu Cmpd}} \right) \left(\frac{9 \text{ mol O atoms}}{1 \text{ mol Cu Cmpd}} \right) \left(\frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O atoms}} \right) \\ &= 1.7581 \times 10^{20} = \mathbf{1.8 \times 10^{20} \text{ O atoms}}\end{aligned}$$

3.13 a) $\text{Mass NO}_2 = (3.8 \times 10^{20} \text{ molecules NO}_2) \left(\frac{1 \text{ mol NO}_2}{6.022 \times 10^{23} \text{ molecules NO}_2} \right) \left(\frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2} \right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right)$
 $= 2.9033 \times 10^{-5} = \mathbf{2.9 \times 10^{-5} \text{ kg NO}_2}$

$$\begin{aligned} \text{b) Moles Cl atoms} &= (0.0425 \text{ g C}_2\text{H}_4\text{Cl}_2) \left(\frac{1 \text{ mol C}_2\text{H}_4\text{Cl}_2}{98.95 \text{ g C}_2\text{H}_4\text{Cl}_2} \right) \left(\frac{2 \text{ mol Cl atoms}}{1 \text{ mol C}_2\text{H}_4\text{Cl}_2} \right) \\ &= 8.5902 \times 10^{-4} = \mathbf{8.59 \times 10^{-4} \text{ mol Cl atoms}} \end{aligned}$$

$$\begin{aligned} \text{c) Number of H}^- &= (4.92 \text{ g SrH}_2) \left(\frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) \left(\frac{2 \text{ mol H}^-}{1 \text{ mol SrH}_2} \right) \left(\frac{6.022 \times 10^{23} \text{ H}^- \text{ ions}}{1 \text{ mol H}^-} \right) \\ &= 6.610495 \times 10^{22} = \mathbf{6.61 \times 10^{22} \text{ H}^- \text{ ions}} \end{aligned}$$

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

Solution:

$$\begin{aligned} \text{a) } \mathcal{M} \text{ of MnSO}_4 &= (54.94 \text{ g Mn/mol Mn}) + (32.07 \text{ g S/mol S}) + [(4 \text{ mol O}) (16.00 \text{ g O/mol O})] \\ &= 151.01 \text{ g/mol of MnSO}_4 \end{aligned}$$

$$\text{Mass of MnSO}_4 = 0.64 \text{ mol MnSO}_4 \times 151.01 \text{ g MnSO}_4/\text{mol MnSO}_4 = 96.65 = \mathbf{97 \text{ g MnSO}_4}$$

$$\begin{aligned} \text{b) } \mathcal{M} \text{ of Fe(ClO}_4)_3 &= (55.85 \text{ g Fe/mol Fe}) + [(3 \text{ mol Cl}) (35.45 \text{ g Cl/mol Cl})] + [(12 \text{ mol O}) (16.00 \text{ g O/mol O})] \\ &= 354.20 \text{ g/mol of Fe(ClO}_4)_3 \end{aligned}$$

$$\begin{aligned} \text{Moles Fe(ClO}_4)_3 &= (15.8 \text{ g Fe(ClO}_4)_3) / [(354.20 \text{ g Fe(ClO}_4)_3/\text{mol Fe(ClO}_4)_3)] \\ &= 0.044608 = \mathbf{4.46 \times 10^{-2} \text{ mol Fe(ClO}_4)_3} \end{aligned}$$

$$\begin{aligned} \text{c) } \mathcal{M} \text{ of NH}_4\text{NO}_2 &= [(2 \text{ mol N}) (14.01 \text{ g N/mol N})] + [(4 \text{ mol H}) (1.008 \text{ g H/mol H})] \\ &+ [(2 \text{ mol O}) (16.00 \text{ g O/mol O})] = 64.05 \text{ g/mol of NH}_4\text{NO}_2 \end{aligned}$$

$$\text{N atoms} = 92.6 \text{ g NH}_4\text{NO}_2 \left(\frac{1 \text{ mol}}{64.05 \text{ g}} \right) \left(\frac{2 \text{ mol N}}{1 \text{ mol NH}_4\text{NO}_2} \right) \left(\frac{6.022 \times 10^{23} \text{ N atoms}}{1 \text{ mol N}} \right)$$

3.15 FU = formula unit

$$\begin{aligned} \text{a) Ions} &= (38.1 \text{ g CaF}_2) \left(\frac{1 \text{ mol CaF}_2}{78.08 \text{ g CaF}_2} \right) \left(\frac{6.022 \times 10^{23} \text{ FU CaF}_2}{1 \text{ mol CaF}_2} \right) \left(\frac{3 \text{ ions}}{1 \text{ FU CaF}_2} \right) \\ &= 8.81550 \times 10^{23} = \mathbf{8.82 \times 10^{23} \text{ ions}} \end{aligned}$$

$$\begin{aligned} \text{b) Mass CuCl}_2 \cdot 2\text{H}_2\text{O} &= (3.58 \text{ mol CuCl}_2 \cdot 2\text{H}_2\text{O}) \left(\frac{170.48 \text{ g CuCl}_2 \cdot 2\text{H}_2\text{O}}{1 \text{ mol CuCl}_2 \cdot 2\text{H}_2\text{O}} \right) \left(\frac{1 \text{ mg}}{10^{-3} \text{ g}} \right) \\ &= 610318.4 = \mathbf{6.10 \times 10^5 \text{ mg CuCl}_2 \cdot 2\text{H}_2\text{O}} \end{aligned}$$

$$\begin{aligned} \text{c) Mass Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O} &= (2.88 \times 10^{22} \text{ FU}) \left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ FU}} \right) \left(\frac{485.1 \text{ g Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}}{1 \text{ FU Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}} \right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right) \\ &= 0.0231997 = \mathbf{0.0232 \text{ kg Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}} \end{aligned}$$

3.16 Plan: The formula of each compound must be determined from its name. The molar mass for each formula comes from the formula and the periodic table entries. Avogadro's number is also necessary.

Solution:

a) Carbonate is a polyatomic anion with the formula, CO_3^{2-} . The correct formula for this ionic compound is Cu_2CO_3 .

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

$$\text{Mass Cu}_2\text{CO}_3 = (8.41 \text{ mol Cu}_2\text{CO}_3) \left(\frac{187.11 \text{ g Cu}_2\text{CO}_3}{1 \text{ mol Cu}_2\text{CO}_3} \right) = 1573.595 = \mathbf{1.57 \times 10^3 \text{ g Cu}_2\text{CO}_3}$$

b) Dinitrogen pentoxide has the formula N_2O_5 .

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

$$\begin{aligned} \text{Mass N}_2\text{O}_5 &= (2.04 \times 10^{21} \text{ N}_2\text{O}_5 \text{ molecules}) \left(\frac{1 \text{ mol N}_2\text{O}_5}{6.022 \times 10^{23} \text{ N}_2\text{O}_5 \text{ molecules}} \right) \left(\frac{108.02 \text{ g N}_2\text{O}_5}{1 \text{ mol N}_2\text{O}_5} \right) \\ &= 0.365926 = \mathbf{0.366 \text{ g N}_2\text{O}_5} \end{aligned}$$

c) The correct formula for this ionic compound is NaClO_4 . There are Avogadro's number of entities (in this case, formula units) in a mole of this compound.

$$\mathcal{M} \text{ of } \text{NaClO}_4 = 22.99 + 35.45 + (4 \times 16.00) = 122.44 \text{ g/mol}$$

$$\text{Moles } \text{NaClO}_4 = (57.9 \text{ g } \text{NaClO}_2) \left(\frac{1 \text{ mol } \text{NaClO}_4}{122.44 \text{ g } \text{NaClO}_4} \right) = 0.47288 = \mathbf{0.473 \text{ mol } \text{NaClO}_4}$$

FU = formula units

$$\text{FU } \text{NaClO}_4 = (57.9 \text{ g } \text{NaClO}_2) \left(\frac{1 \text{ mol } \text{NaClO}_4}{122.44 \text{ g } \text{NaClO}_4} \right) \left(\frac{6.022 \times 10^{23} \text{ FU } \text{NaClO}_4}{1 \text{ mol } \text{NaClO}_4} \right) =$$

$$2.8477115 \times 10^{23} = \mathbf{2.85 \times 10^{23} \text{ FU } \text{NaClO}_4}$$

d) The number of ions or atoms is calculated from the formula units given in part c. Note the unrounded initially calculated value is used to avoid intermediate rounding.

$$2.8477115 \times 10^{23} \text{ mol } \text{NaClO}_4 \left(\frac{1 \text{ Na}^+ \text{ ion}}{1 \text{ FU } \text{NaClO}_4} \right) = \mathbf{2.85 \times 10^{23} \text{ Na}^+ \text{ ions}}$$

$$2.8477115 \times 10^{23} \text{ mol } \text{NaClO}_4 \left(\frac{1 \text{ ClO}_4^- \text{ ion}}{1 \text{ FU } \text{NaClO}_4} \right) = \mathbf{2.85 \times 10^{23} \text{ ClO}_4^- \text{ ions}}$$

$$2.8477115 \times 10^{23} \text{ mol } \text{NaClO}_4 \left(\frac{1 \text{ Cl atom}}{1 \text{ FU } \text{NaClO}_4} \right) = \mathbf{2.85 \times 10^{23} \text{ Cl atoms}}$$

$$2.8477115 \times 10^{23} \text{ mol } \text{NaClO}_4 \left(\frac{4 \text{ O atoms}}{1 \text{ FU } \text{NaClO}_4} \right) = \mathbf{1.14 \times 10^{24} \text{ O atoms}}$$

3.17 a) The formula is $\text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O}$, and the molar mass is 572.4 g/mol.

$$\text{Grams } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O} = (3.52 \text{ mol } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O}) \left(\frac{572.4 \text{ g}}{\text{mol}} \right)$$

$$= 2014.848 = \mathbf{2.01 \times 10^3 \text{ g } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O}}$$

b) The formula is Cl_2O_7 , and the molar mass is 182.9 g/mol.

$$\text{Grams } \text{Cl}_2\text{O}_7 = (9.64 \times 10^{24} \text{ molecules } \text{Cl}_2\text{O}_7) \left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \right) \left(\frac{182.9 \text{ g } \text{Cl}_2\text{O}_7}{1 \text{ mol}} \right)$$

$$= 2927.858 = \mathbf{2.93 \times 10^3 \text{ g } \text{Cl}_2\text{O}_7}$$

c) The formula is Li_2SO_4 , and the molar mass is 109.95 g/mol.

$$\text{Mol } \text{Li}_2\text{SO}_4 = (56.2 \text{ g } \text{Li}_2\text{SO}_4) \left(\frac{1 \text{ mol } \text{Li}_2\text{SO}_4}{109.95 \text{ g } \text{Li}_2\text{SO}_4} \right) = 0.5111414 = \mathbf{0.511 \text{ mol } \text{Li}_2\text{SO}_4}$$

$$\text{FU } \text{Li}_2\text{SO}_4 = (56.2 \text{ g } \text{Li}_2\text{SO}_4) \left(\frac{1 \text{ mol } \text{Li}_2\text{SO}_4}{109.95 \text{ g } \text{Li}_2\text{SO}_4} \right) \left(\frac{6.022 \times 10^{23} \text{ FU}}{1 \text{ mol } \text{Li}_2\text{SO}_4} \right)$$

$$= 3.07809 \times 10^{23} = \mathbf{3.08 \times 10^{23} \text{ FU } \text{Li}_2\text{SO}_4}$$

d) Note the unrounded initially calculated FU value is used to avoid intermediate rounding.

$$3.07809 \times 10^{23} \text{ FU } \text{Li}_2\text{SO}_4 \left(\frac{2 \text{ Li}^+ \text{ ions}}{1 \text{ FU } \text{Li}_2\text{SO}_4} \right) = 6.156187 \times 10^{23} = \mathbf{6.16 \times 10^{23} \text{ Li}^+ \text{ ions}}$$

$$3.07809 \times 10^{23} \text{ FU } \text{Li}_2\text{SO}_4 \left(\frac{1 \text{ SO}_4^{2-} \text{ ion}}{1 \text{ FU } \text{Li}_2\text{SO}_4} \right) = 3.07809 \times 10^{23} = \mathbf{3.08 \times 10^{23} \text{ SO}_4^{2-} \text{ ions}}$$

$$3.07809 \times 10^{23} \text{ FU } \text{Li}_2\text{SO}_4 \left(\frac{1 \text{ S atom}}{1 \text{ FU } \text{Li}_2\text{SO}_4} \right) = 3.07809 \times 10^{23} = \mathbf{3.08 \times 10^{23} \text{ S atoms}}$$

$$3.07809 \times 10^{23} \text{ FU Li}_2\text{SO}_4 \left(\frac{4 \text{ O atoms}}{1 \text{ FU Li}_2\text{SO}_4} \right) = 1.231237 \times 10^{24} = \mathbf{1.23 \times 10^{24} \text{ O atoms}}$$

3.18 Plan: Determine the formula and the molar mass of each compound. The formula gives the number of atoms of each type of element present. Masses come from the periodic table.

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 .

$$\mathcal{M} \text{ of } \text{NH}_4\text{HCO}_3 = (14.01 \text{ g/mol}) + (5 \times 1.008 \text{ g/mol}) + (12.01 \text{ g/mol}) + (3 \times 16.00 \text{ g/mol}) \\ = 79.06 \text{ g/mol } \text{NH}_4\text{HCO}_3$$

In 1 mole of ammonium bicarbonate, with a mass of 79.06 g, there are 5 H atoms with a mass of 5.040 g.

$$\frac{(5 \text{ mol H})(1.008 \text{ g/mol H})}{79.06 \text{ g/mol } \text{NH}_4\text{HCO}_3} \times 100\% = 6.374905 = \mathbf{6.375\% \text{ H}}$$

b) Sodium dihydrogen phosphate heptahydrate is a salt that consists of sodium ions, Na^+ , dihydrogen phosphate ions, H_2PO_4^- , and seven waters of hydration. The formula is $\text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O}$. Note that the waters of hydration are included in the molar mass.

$$\mathcal{M} \text{ of } \text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O} = (22.99 \text{ g/mol}) + (16 \times 1.008 \text{ g/mol}) + (30.97 \text{ g/mol}) + (11 \times 16.00 \text{ g/mol}) \\ = 246.09 \text{ g/mol } \text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O}$$

In each mole of $\text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O}$ (with mass of 246.09 g), there are 11 x 16.00 g/mol or 176.00 g of oxygen.

$$\frac{(11 \text{ mol O})(16.00 \text{ g/mol O})}{246.09 \text{ g/mol}} \times 100\% = 71.51855 = \mathbf{71.52\% \text{ O}}$$

3.19 a) $\text{Sr}(\text{IO}_4)_2$ 437.42 g/mol

$$\frac{(2 \text{ mol I})(126.9 \text{ g/mol I})}{437.42 \text{ g/mol}} \times 100\% = 54.0690 = \mathbf{54.07\% \text{ I}}$$

b) KMnO_4 158.04 g/mol

$$\frac{(1 \text{ mol Mn})(54.94 \text{ g/mol Mn})}{158.04 \text{ g/mol}} \times 100\% = 34.76335 = \mathbf{34.76\% \text{ Mn}}$$

3.20 Plan: Mass fraction is related to the mass percentage, however the mass fraction is expressed in decimal rather than percentage form.

Solution:

a) Cesium acetate is an ionic compound consisting of Cs^+ cations and $\text{C}_2\text{H}_3\text{O}_2^-$ anions. (Note that the formula for acetate ions can be written as either $\text{C}_2\text{H}_3\text{O}_2^-$ or CH_3COO^- .) The formula of the compound is $\text{CsC}_2\text{H}_3\text{O}_2$. One mole of $\text{CsC}_2\text{H}_3\text{O}_2$ weighs 191.9 g:

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

$$\text{Mass fraction of C} = \frac{(2 \text{ mol C})(12.01 \text{ g/mol C})}{191.9 \text{ g/mol}} = 0.125169 = \mathbf{0.1252 \text{ mass fraction C}}$$

b) The formula for this compound is $\text{UO}_2\text{SO}_4 \cdot 3\text{H}_2\text{O}$.

$$\mathcal{M} \text{ of } \text{UO}_2\text{SO}_4 \cdot 3\text{H}_2\text{O} = 238.0 + (9 \times 16.00) + 32.07 + (6 \times 1.008) = 420.1 \text{ g/mol}$$

$$\text{Mass fraction of O} = \frac{(9 \text{ mol O})(16.00 \text{ g/mol O})}{420.1 \text{ g/mol}} = 0.3427755 = \mathbf{0.3428 \text{ mass fraction O}}$$

3.21 a) The formula is $\text{Ca}(\text{ClO}_3)_2$, and the molar mass is 206.98 g/mol.

$$\text{Mass fraction of Cl} = \frac{(2 \text{ mol Cl})(35.45 \text{ g/mol Cl})}{206.98 \text{ g/mol}} = 0.342545 = \mathbf{0.3425 \text{ mass fraction Cl}}$$

b) The formula is P_4O_6 , and the molar mass is 219.88 g/mol.

$$\text{Mass fraction of P} = \frac{(4 \text{ mol P})(30.97 \text{ g/mol P})}{219.88 \text{ g/mol}} = 0.5633982 = \mathbf{0.5634 \text{ mass fraction P}}$$

$$\begin{aligned} 3.22 \quad \text{O atoms} &= (38.0 \text{ g O}_2) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \left(\frac{6.022 \times 10^{23} \text{ O}_2 \text{ molecules}}{1 \text{ mol O}_2} \right) \left(\frac{2 \text{ O atoms}}{1 \text{ O}_2 \text{ molecule}} \right) \\ &= 1.430225 \times 10^{24} = \mathbf{1.43 \times 10^{24} \text{ O atoms}} \end{aligned}$$

3.23 Plan: Determine the formula of cisplatin from the figure, and then calculate the molar mass from the formula. The molar mass is necessary for the subsequent calculations.

Solution:

The formula for cisplatin is $Pt(Cl)_2(NH_3)_2$

$$\mathcal{M} \text{ of } Pt(Cl)_2(NH_3)_2 = 195.1 + (2 \times 35.45) + (2 \times 14.01) + (6 \times 1.008) = 300.1 \text{ g/mol}$$

$$a) \text{ Moles cisplatin} = (285.3 \text{ g cisplatin}) \left(\frac{1 \text{ mol cisplatin}}{300.1 \text{ g cisplatin}} \right) = 0.9506831 = \mathbf{0.9507 \text{ mol cisplatin}}$$

$$\begin{aligned} b) \text{ H atoms} &= (0.98 \text{ mol cisplatin}) \left(\frac{6 \text{ mol H}}{1 \text{ mol cisplatin}} \right) \left(\frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}} \right) \\ &= 3.540936 \times 10^{24} = \mathbf{3.5 \times 10^{24} \text{ H atoms}} \end{aligned}$$

3.24 The formula, from the figure, is $(C_3H_5)_2S$, and the molar mass is 114.21 g/mol.

$$\begin{aligned} a) \text{ Grams allyl sulfide} &= (1.63 \text{ mol allyl sulfide}) \left(\frac{114.21 \text{ g allyl sulfide}}{1 \text{ mol allyl sulfide}} \right) \\ &= 186.1623 = \mathbf{186 \text{ g allyl sulfide}} \end{aligned}$$

$$\begin{aligned} b) \text{ C atoms} &= (4.77 \text{ g } (C_3H_5)_2S) \left(\frac{1 \text{ mol } (C_3H_5)_2S}{114.21 \text{ g } (C_3H_5)_2S} \right) \left(\frac{6 \text{ mol C}}{1 \text{ mol } (C_3H_5)_2S} \right) \left(\frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C}} \right) \\ &= 1.5090591 \times 10^{23} = \mathbf{1.51 \times 10^{23} \text{ C atoms}} \end{aligned}$$

3.25 Plan: Determine the molar mass of rust. Use the molar mass to find the moles of rust. The moles of rust may be related to the grams of iron through the mole ratio.

Solution:

$$a) \mathcal{M} \text{ of } Fe_2O_3 \cdot 4H_2O = (2 \times 55.85) + (3 \times 16.00) + (4 \times 18.016) = 231.76 \text{ g/mol}$$

$$\text{Moles of compound} = (65.2 \text{ kg rust}) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol rust}}{231.76 \text{ g rust}} \right) = 281.3255 = \mathbf{281 \text{ mol rust}}$$

b) The formula shows that there is one mole of Fe_2O_3 for every mole of rust, so there are also **281 mol of Fe_2O_3** .

c) Calculate grams of iron by determining the mole ratio and converting to grams.

$$\begin{aligned} \text{Grams of iron} &= (65.2 \text{ kg rust}) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol rust}}{231.76 \text{ g rust}} \right) \left(\frac{1 \text{ mol } Fe_2O_3}{1 \text{ mol rust}} \right) \left(\frac{2 \text{ mol Fe}}{1 \text{ mol } Fe_2O_3} \right) \left(\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \right) \\ &= 31424.059 = \mathbf{3.14 \times 10^4 \text{ g Fe}} \end{aligned}$$

$$3.26 \quad a) \text{ Moles } C_3H_8 = (75.3 \text{ g } C_3H_8) \left(\frac{1 \text{ mol } C_3H_8}{44.09 \text{ g } C_3H_8} \right) = 1.707870 = 1.71 \text{ mol } C_3H_8$$

$$b) \text{ Grams C} = (75.3 \text{ g } C_3H_8) \left(\frac{1 \text{ mol } C_3H_8}{44.09 \text{ g } C_3H_8} \right) \left(\frac{3 \text{ mol C}}{1 \text{ mol } C_3H_8} \right) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 61.534566 = \mathbf{61.5 \text{ g C}}$$

- 3.27 Plan: Determine the formulas for the compounds where needed. Determine the molar mass of each formula. Calculate the percent nitrogen by dividing the mass of nitrogen in a mole of compound by the molar mass of the compound, and multiply the result by 100%. Then rank the values.

Solution:

Name	Formula	Molar Mass (g/mol)
Potassium nitrate	KNO ₃	101.11
Ammonium nitrate	NH ₄ NO ₃	80.05
Ammonium sulfate	(NH ₄) ₂ SO ₄	132.12
Urea	CO(NH ₂) ₂	60.06

Calculating the nitrogen percentages:

$$\text{Potassium nitrate} \quad \frac{(1 \text{ mol N})(14.01 \text{ g/mol N})}{101.11 \text{ g/mol}} \times 100 = 13.856196 = \mathbf{13.86\% \text{ N}}$$

$$\text{Ammonium nitrate} \quad \frac{(2 \text{ mol N})(14.01 \text{ g/mol N})}{80.05 \text{ g/mol}} \times 100 = 35.003123 = \mathbf{35.00\% \text{ N}}$$

$$\text{Ammonium sulfate} \quad \frac{(2 \text{ mol N})(14.01 \text{ g/mol N})}{132.12 \text{ g/mol}} \times 100 = 21.20799 = \mathbf{21.21\% \text{ N}}$$

$$\text{Urea} \quad \frac{(2 \text{ mol N})(14.01 \text{ g/mol N})}{60.06 \text{ g/mol}} \times 100 = 46.6533 = \mathbf{46.65\% \text{ N}}$$

Rank is CO(NH₂)₂ > NH₄NO₃ > (NH₄)₂SO₄ > KNO₃

- 3.28 Plan: The volume must be converted from cubic feet to cubic centimeters (or vice versa). The volume and the density will give you mass, and the mass with the molar mass gives you moles. Part (b) requires a conversion from cubic decimeters, instead of cubic feet, to cubic centimeters. The density allows you to change these cubic centimeters to mass, the molar mass allows you to find moles, and finally Avogadro's number allows you to make the last step.

Solution:

The molar mass of galena is 239.3 g/mol.

$$\begin{aligned} \text{a) Moles PbS} &= (1.00 \text{ ft}^3 \text{ PbS}) \left(\frac{12 \text{ in}}{1 \text{ ft}} \right)^3 \left(\frac{2.54 \text{ cm}}{1 \text{ in}} \right)^3 \left(\frac{7.46 \text{ g PbS}}{1 \text{ cm}^3} \right) \left(\frac{1 \text{ mol PbS}}{239.3 \text{ g PbS}} \right) \\ &= 882.7566886 = \mathbf{883 \text{ mol PbS}} \end{aligned}$$

$$\begin{aligned} \text{b) } (1.00 \text{ dm}^3 \text{ PbS}) &\left(\frac{0.1 \text{ m}}{1 \text{ dm}} \right)^3 \left(\frac{1 \text{ cm}}{10^{-2} \text{ m}} \right)^3 \left(\frac{7.46 \text{ g PbS}}{1 \text{ cm}^3} \right) \left(\frac{1 \text{ mol PbS}}{239.3 \text{ g PbS}} \right) \left(\frac{1 \text{ mol Pb}}{1 \text{ mol PbS}} \right) \left(\frac{6.022 \times 10^{23} \text{ Pb atoms}}{1 \text{ mol Pb}} \right) \\ &= 1.87731 \times 10^{25} = \mathbf{1.88 \times 10^{25} \text{ Pb atoms}} \end{aligned}$$

- 3.29 Plan: If the molecular formula for hemoglobin (Hb) was known, the number of Fe²⁺ 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. From the mass of iron, the moles of iron per mole of hemoglobin may be found.

Solution:

$$\left(\frac{0.33\% \text{ Fe}}{100\% \text{ Hb}} \right) \left(\frac{6.8 \times 10^4 \text{ g}}{\text{mol}} \right) \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) = 4.0179 = 4.0 \text{ mol Fe}^{2+}/\text{mol Hb}$$

Thus, there are **4 Fe²⁺/molecule Hb**.

- 3.30 Determine the empirical formula from the mass % of the elements.
Determine the E.F. from elemental composition.
Determine the E.F. from combustion analysis.
- 3.31 a) No, 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.

Solution:

- 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.
- 4) Write the empirical formula using the whole numbers from step 3.
- 5) 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.

c) Yes, you can determine the molecular formula from the mass percent and the number of atoms of one element in a compound. Solution plan:

- 1) Follow steps 1-4 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.

d) No, the mass % will only lead to the empirical formula.

e) Yes, a structural formula shows all the atoms in the compound. Solution plan: Count the number of atoms of each type of element and record as the number for the molecular formula.

3.32 MgCl_2 is an empirical formula, since ionic compounds such as MgCl_2 do not contain molecules.

3.33 Plan: Examine the number of atoms of each type in the compound. Divide all atom numbers by any common factor. The final answers must be the lowest whole-number values.

Solution:

a) C_2H_4 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 CH_2 . The empirical formula mass is $12.01 + 2(1.008) = \mathbf{14.03 \text{ g/mol}}$.

b) The ratio of atoms is 2:6:2, or 1:3:1. The empirical formula is CH_3O and its empirical formula mass is $12.01 + 3(1.008) + 16.00 = \mathbf{31.03 \text{ g/mol}}$

c) Since, the ratio of elements cannot be further reduced, the molecular formula and empirical formula are the same, N_2O_5 . The formula mass is $2(14.01) + 5(16.00) = \mathbf{108.02 \text{ g/mol}}$.

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 $\text{Ba}_3(\text{PO}_4)_2$, with a formula mass of $3(137.3) + 2(30.97) + 8(16.00) = \mathbf{601.8 \text{ g/mol}}$.

e) The empirical formula is TeI_4 , and the formula mass is $127.6 + 4(126.9) = \mathbf{635.2 \text{ g/mol}}$.

- | | | | |
|------|---|--|---|
| 3.34 | a) $\text{C}_4\text{H}_8 = \text{M.F.}$ | $\text{CH}_2 = \text{E.F.}$ | E.F. mass = $\mathbf{14.03 \text{ g/mol}}$ |
| | b) $\text{C}_3\text{H}_6\text{O}_3 = \text{M.F.}$ | $\text{CH}_2\text{O} = \text{E.F.}$ | E.F. mass = $\mathbf{30.03 \text{ g/mol}}$ |
| | c) $\text{P}_4\text{O}_{10} = \text{M.F.}$ | $\text{P}_2\text{O}_5 = \text{E.F.}$ | E.F. mass = $\mathbf{141.94 \text{ g/mol}}$ |
| | d) $\text{Ga}_2(\text{SO}_4)_3 = \text{M.F.}$ | $\text{Ga}_2(\text{SO}_4)_3 = \text{M.F.} = \text{E.F.}$ | E.F. mass = $\mathbf{427.6 \text{ g/mol}}$ |
| | e) $\text{Al}_2\text{Br}_6 = \text{M.F.}$ | $\text{AlBr}_3 = \text{E.F.}$ | E.F. mass = $\mathbf{266.7 \text{ g/mol}}$ |

3.35 Plan: Determine the molar mass of each empirical formula. The molar mass of each compound divided by its empirical formula mass gives the number of times the empirical formula is within the molecule. Multiply the empirical formula by the number of times the empirical formula appears to get the molecular formula.

Solution:

Only approximate whole number values are needed.

a) CH₂ has empirical mass equal to 14.03 g/mol

$$\left(\frac{42.08 \text{ g/mol}}{14.03 \text{ g/mol}} \right) = 3$$

Multiplying the subscripts in CH₂ by 3 gives **C₃H₆**

b) NH₂ has empirical mass equal to 16.03 g/mol

$$\left(\frac{32.05 \text{ g/mol}}{16.03 \text{ g/mol}} \right) = 2$$

Multiplying the subscripts in NH₂ by 2 gives **N₂H₄**

c) NO₂ has empirical mass equal to 46.01 g/mol

$$\left(\frac{92.02 \text{ g/mol}}{46.01 \text{ g/mol}} \right) = 2$$

Multiplying the subscripts in NO₂ by 2 gives **N₂O₄**

d) CHN has empirical mass equal to 27.03 g/mol

$$\left(\frac{135.14 \text{ g/mol}}{27.03 \text{ g/mol}} \right) = 5$$

Multiplying the subscripts in CHN by 5 gives **C₅H₅N₅**

3.36 a) CH has empirical mass equal to = 13.02 g/mol

$$\left(\frac{78.11 \text{ g/mol}}{13.02 \text{ g/mol}} \right) = 6 \quad \text{M.F.} = 6 \times \text{E.F.} = \mathbf{C_6H_6}$$

b) C₃H₆O₂ has empirical mass equal to = 74.08 g/mol

$$\left(\frac{74.08 \text{ g/mol}}{74.08 \text{ g/mol}} \right) = 1 \quad \text{M.F.} = 1 \times \text{E.F.} = \mathbf{C_3H_6O_2}$$

c) HgCl has empirical mass equal to = 236.0 g/mol

$$\left(\frac{472.1 \text{ g/mol}}{236.0 \text{ g/mol}} \right) = 2 \quad \text{M.F.} = 2 \times \text{E.F.} = \mathbf{Hg_2Cl_2}$$

d) C₇H₄O₂ has empirical mass equal to = 120.10 g/mol

$$\left(\frac{240.20 \text{ g/mol}}{120.10 \text{ g/mol}} \right) = 2 \quad \text{M.F.} = 2 \times \text{E.F.} = \mathbf{C_{14}H_8O_4}$$

3.37 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. Using the smallest number of moles present, convert the mole ratios to whole numbers.

Solution:

$$\text{a) } \left(\frac{0.063 \text{ mol Cl}}{0.063 \text{ mol Cl}} \right) = 1 \quad \left(\frac{0.22 \text{ mol O}}{0.063 \text{ mol Cl}} \right) = 3.5$$

The formula is Cl₁O_{3.5}, which in whole numbers (x 2) is **Cl₂O₇**

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

$$\left(\frac{0.08722 \text{ mol Si}}{0.08722 \text{ mol Si}} \right) = 1 \quad \left(\frac{0.349788 \text{ mol Cl}}{0.08722 \text{ mol Si}} \right) = 4$$

The empirical formula is **SiCl₄**.

c) Assume a 100 g sample and convert the masses to moles.

$$(100 \text{ g}) \left(\frac{27.3\% \text{ C}}{100\%} \right) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 2.2731 \text{ mol C} \quad (100 \text{ g}) \left(\frac{72.7\% \text{ O}}{100\%} \right) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 4.5438 \text{ mol O}$$

$$\left(\frac{2.2731 \text{ mol C}}{2.2731 \text{ mol C}} \right) = 1 \quad \left(\frac{4.5438 \text{ mol O}}{2.2731 \text{ mol C}} \right) = 2$$

The empirical formula is **CO₂**.

$$3.38 \quad a) \left(\frac{0.039 \text{ mol Fe}}{0.039 \text{ mol Fe}} \right) = 1 \quad \left(\frac{0.052 \text{ mol O}}{0.039 \text{ mol Fe}} \right) = 1.3333$$

The formula is Fe₁O_{1.3333}, which in whole numbers (x 3) is **Fe₃O₄**

$$b) (0.903 \text{ g P}) \left(\frac{1 \text{ mol P}}{30.97 \text{ g P}} \right) = 0.029157 \text{ mol P} \quad (6.99 \text{ g Br}) \left(\frac{1 \text{ mol Br}}{79.90 \text{ g Br}} \right) = 0.087484 \text{ mol Br}$$

$$\left(\frac{0.029157 \text{ mol P}}{0.029157 \text{ mol P}} \right) = 1 \quad \left(\frac{0.087484 \text{ mol Br}}{0.029157 \text{ mol P}} \right) = 3$$

The empirical formula is **PBr₃**.

c) Assume a 100 g sample and convert the masses to moles.

$$(100 \text{ g}) \left(\frac{79.9\% \text{ C}}{100\%} \right) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 6.6528 \text{ mol C}$$

$$(100 \text{ g}) \left(\frac{(100 - 79.9)\% \text{ H}}{100\%} \right) \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 19.940 \text{ mol H}$$

$$\left(\frac{6.6528 \text{ mol C}}{6.6528 \text{ mol C}} \right) = 1 \quad \left(\frac{19.940 \text{ mol H}}{6.6528 \text{ mol C}} \right) = 3$$

The empirical formula is **CH₃**.

3.39 **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. Divide each of the moles by the smaller value, and convert to whole numbers to get the empirical formula. The empirical formula mass and the given mass lead to the molecular formula.

Solution:

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

For a 100-gram sample, the mass, in grams, is numerically identical to the mass percent.

$$\text{Moles N} = (30.45 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 2.1734 \text{ mol N}$$

$$\text{Moles O} = (69.55 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 4.3469 \text{ mol O}$$

$$\left(\frac{2.1734 \text{ mol N}}{2.1734 \text{ mol N}} \right) = 1 \quad \left(\frac{4.3469 \text{ mol O}}{2.1734 \text{ mol N}} \right) = 2$$

The empirical formula is **NO₂**.

b) Formula mass of empirical formula = $14.01 \text{ g N/mol} + 2(16.00 \text{ g O/mol}) = 46.01 \text{ g/mol}$
Molar mass / Empirical formula mass = $(90 \text{ g/mol}) / (46 \text{ g/mol}) = 2$
Thus, the molecular formula is twice the empirical formula. Note: only an approximate value is needed.
The molecular formula is N_2O_4 .

3.40 The percent silicon is 100% minus the percent chlorine. Assume 100 grams of sample, and then the moles of each element may be found. Divide each of the moles by the smaller value, and convert to whole numbers to get the empirical formula. The empirical formula mass and the given mass lead to the molecular formula.

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

For a 100-gram sample, the mass, in grams, is numerically identical to the mass percent.

$$\text{Moles Si} = (20.9 \text{ g Si}) \left(\frac{1 \text{ mol Si}}{28.09 \text{ g Si}} \right) = 0.7440 \text{ mol Si}$$

$$\text{Moles Cl} = (79.1 \text{ g Cl}) \left(\frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} \right) = 2.2313 \text{ mol Cl}$$

$$\left(\frac{0.7440 \text{ mol Si}}{0.7440 \text{ mol Si}} \right) = 1 \qquad \left(\frac{2.2313 \text{ mol Cl}}{0.7440 \text{ mol Si}} \right) = 3$$

The empirical formula is SiCl_3 .

b) Formula mass of empirical formula = $28.09 \text{ g Si/mol} + 3(35.45 \text{ g Cl/mol}) = 134.44 \text{ g/mol}$

Molar mass / Empirical formula mass = $(269 \text{ g/mol}) / (134.44 \text{ g/mol}) = 2$

Thus, the molecular formula is twice the empirical formula. Note: only an approximate value is needed.

The molecular formula is Si_2Cl_6 .

3.41 Plan: The balanced equation for this reaction is: $\text{M}(s) + \text{F}_2(g) \rightarrow \text{MF}_2(s)$ since fluorine, like other halogens, exists as a diatomic molecule. The moles of the metal are known, and the moles of everything else may be found from these moles using the balanced chemical equation.

Solution:

a) Determine the moles of fluorine.

$$\text{Moles F} = (0.600 \text{ mol M}) \left(\frac{2 \text{ mol F}}{1 \text{ mol M}} \right) = \mathbf{1.20 \text{ mol F}}$$

b) The grams of M are the grams of MF_2 minus the grams of F present.

$$\text{Grams M} = 46.8 \text{ g (M + F)} - (1.20 \text{ mol F}) \left(\frac{19.00 \text{ g F}}{1 \text{ mol F}} \right) = \mathbf{24.0 \text{ g M}}$$

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

Molar mass of M = $24.0 \text{ g M} / 0.600 \text{ mol M} = 40.0 \text{ g/mol}$

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

3.42 a) moles O = $(0.370 \text{ mol M}_2\text{O}_3) \left(\frac{3 \text{ mol O}}{1 \text{ mol M}_2\text{O}_3} \right) = \mathbf{1.11 \text{ mol O}}$

b) The grams of M are the grams of M_2O_3 minus the grams of O present.

$$\text{Grams M} = 55.4 \text{ g (M + O)} - (1.11 \text{ mol O}) \left(\frac{16.00 \text{ g O}}{1 \text{ mol O}} \right) = 37.64 = \mathbf{37.6 \text{ g M}}$$

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

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

The molar mass is needed to identify the element. (Use the unrounded mass of M to avoid intermittent rounding errors.)

Molar mass of M = $37.64 \text{ g M} / 0.740 \text{ mol M} = 50.86 \text{ g/mol}$

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

- 3.43 Divide all the millimoles by the smallest value and convert to whole numbers to get the empirical formula. Since all the values are given in millimoles, there is no need to convert to another unit.

$$\left(\frac{6.16 \text{ mmol C}}{1.23 \text{ mmol N}}\right) = 5 \quad \left(\frac{8.56 \text{ mmol H}}{1.23 \text{ mmol N}}\right) = 7 \quad \left(\frac{1.23 \text{ mmol N}}{1.23 \text{ mmol N}}\right) = 1$$

All the ratios are close to whole numbers so no other adjustments are necessary.

The empirical formula is $\text{C}_5\text{H}_7\text{N}$.

- 3.44 Plan: 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 using the molar mass of each element involved. Divide all moles by the lowest number of moles and convert to whole numbers to determine the empirical formula. The empirical formula mass and the given molar mass will then relate the empirical formula to the molecular formula.

Solution:

$$\text{Moles C} = (69.6 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}}\right) = 5.7952 \text{ mol C}$$

$$\text{Moles H} = (8.34 \text{ g H}) \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}}\right) = 8.2738 \text{ mol H}$$

$$\text{Moles O} = (22.1 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}}\right) = 1.38125 \text{ mol O}$$

$$\left(\frac{5.7952 \text{ mol C}}{1.38125 \text{ mol O}}\right) = 4.20 \quad \left(\frac{8.2738 \text{ mol H}}{1.38125 \text{ mol O}}\right) = 6.00 \quad \left(\frac{1.38125 \text{ mol O}}{1.38125 \text{ mol O}}\right) = 1.00$$

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 to get the empirical formula of $\text{C}_{21}\text{H}_{30}\text{O}_5$.

The empirical formula mass is:

$$21 (12.01 \text{ g C/mol}) + 30 (1.008 \text{ g H/mol}) + 5 (16.00 \text{ g O/mol}) = 362.45 \text{ g/mol}$$

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 $\text{C}_{21}\text{H}_{30}\text{O}_5$.

- 3.45 Determine the molecular formula from the figure, and the molar mass from the molecular formula. Divide the amount of each element by the molar mass and multiply the answer by 100%.

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

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

$$\% \text{ C} = [8(12.01 \text{ g C/mol})/151.16 \text{ g/mol}] \times 100\% = 63.5618 = \mathbf{63.56\% \text{ C}}$$

$$\% \text{ H} = [9(1.008 \text{ g H/mol})/151.16 \text{ g/mol}] \times 100\% = 6.00159 = \mathbf{6.002\% \text{ H}}$$

$$\% \text{ N} = [14.01 \text{ g N/mol}/151.16 \text{ g/mol}] \times 100\% = 9.2683 = \mathbf{9.268\% \text{ N}}$$

$$\% \text{ O} = [2(16.00 \text{ g O/mol})/151.16 \text{ g/mol}] \times 100\% = 21.1696 = \mathbf{21.17\% \text{ O}}$$

- 3.46 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 CO_2 , and all of the hydrogen present in the sample is found in the hydrogen of H_2O . The moles of oxygen are more difficult to find, because additional O_2 was added to cause the combustion reaction. The masses of CO_2 and H_2O are used to find both the mass of C and H and the moles of C and H. 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, and convert to a whole number 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: (There is no intermediate rounding.)

Initial mole determination:

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

$$\text{Moles H} = (0.184 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.020422 \text{ mol H}$$

Now determine the masses of C and H:

$$\text{Grams C} = (0.010202 \text{ mol C}) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 0.122526 \text{ g C}$$

$$\text{Grams H} = (0.020422 \text{ mol H}) \left(\frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) = 0.020585 \text{ g H}$$

Determine the mass and then the moles of O:

$$0.1595 \text{ g (C, H, and O)} - 0.122526 \text{ g C} - 0.020585 \text{ g H} = 0.016389 \text{ g O}$$

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

Divide by the smallest number of moles: (Rounding is acceptable for these answers.)

$$\left(\frac{0.010202 \text{ mol C}}{0.0010243 \text{ mol O}} \right) = 10 \quad \left(\frac{0.020422 \text{ mol H}}{0.0010243 \text{ mol O}} \right) = 20 \quad \left(\frac{0.0010243 \text{ mol O}}{0.0010243 \text{ mol O}} \right) = 1$$

Empirical formula = C₁₀H₂₀O

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

The empirical formula mass matches the given molar mass so the empirical and molecular formulas are the same.

The molecular formula is **C₁₀H₂₀O**.

3.47 A balanced chemical equation describes:

- The identities of the reactants and products
- The molar (and molecular) ratios by which reactants form products
- The physical states of all substances in the reaction

3.48 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 conservation of mass is obeyed.

3.49 Students I and II are incorrect. Both students changed a given formula. Only coefficients should be changed when balancing. Student I failed to identify the product correctly, and Student II used atomic chlorine instead of molecular chlorine as a reactant. Student III followed the correct process.

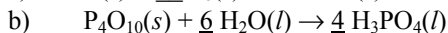
3.50 Plan: Exam the diagram and label each formula. We will use A for red atoms and B for green atoms.

Solution:

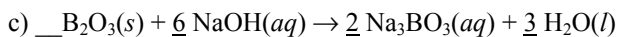
The reaction shows A₂ and B₂ molecules forming AB molecules. Equal numbers of A₂ and B₂ combine to give twice as many molecules of AB. Thus, the reaction is A₂ + B₂ → 2 AB. This is the answer to part b.

3.51 Plan: Balancing is a trial and error procedure. Do one blank/one element at a time.

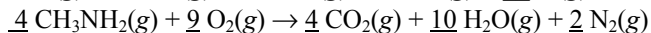
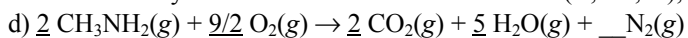
Solution:



Hint: Balance the P first, because there is an obvious deficiency of P on the right side of the equation. Balance the H next, because H is present in only one reactant and only one product. Balance the O last, because it appears in both reactants and is harder to balance.



Hint: Oxygen is again the hardest element to balance because it is present in more than one place on each side of the reaction. If you balance the easier elements first (B, Na, H), the oxygen will automatically be balanced.



Hint: You should balance odd/even numbers of oxygen using the “half” method, and then multiply all coefficients by two.

- 3.52 a) $\text{Cu}(\text{NO}_3)_2(\text{aq}) + 2 \text{KOH}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) + 2 \text{KNO}_3(\text{aq})$
 b) $\text{BCl}_3(\text{g}) + 3 \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{BO}_3(\text{s}) + 3 \text{HCl}(\text{g})$
 c) $\text{CaSiO}_3(\text{s}) + 6 \text{HF}(\text{g}) \rightarrow \text{SiF}_4(\text{g}) + \text{CaF}_2(\text{s}) + 3 \text{H}_2\text{O}(\text{l})$
 d) $(\text{CN})_2(\text{g}) + 4 \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{C}_2\text{O}_4(\text{aq}) + 2 \text{NH}_3(\text{g})$

3.53 Plan: The balancing is a trial and error procedure. Do one blank/one element at a time.

Solution:

- a) $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g})$
 b) $\text{Sc}_2\text{O}_3(\text{s}) + 3 \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{Sc}(\text{OH})_3(\text{s})$
 c) $\text{H}_3\text{PO}_4(\text{aq}) + 2 \text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{HPO}_4(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$
 d) $\text{C}_6\text{H}_{10}\text{O}_5(\text{s}) + 6 \text{O}_2(\text{g}) \rightarrow 6 \text{CO}_2(\text{g}) + 5 \text{H}_2\text{O}(\text{g})$
- 3.54 a) $\text{As}_4\text{S}_6(\text{s}) + 9 \text{O}_2(\text{g}) \rightarrow \text{As}_4\text{O}_6(\text{s}) + 6 \text{SO}_2(\text{g})$
 b) $2 \text{Ca}_3(\text{PO}_4)_2(\text{s}) + 6 \text{SiO}_2(\text{s}) + 10 \text{C}(\text{s}) \rightarrow \text{P}_4(\text{g}) + 6 \text{CaSiO}_3(\text{l}) + 10 \text{CO}(\text{g})$
 c) $3 \text{Fe}(\text{s}) + 4 \text{H}_2\text{O}(\text{g}) \rightarrow \text{Fe}_3\text{O}_4(\text{s}) + 4 \text{H}_2(\text{g})$
 d) $6 \text{S}_2\text{Cl}_2(\text{l}) + 16 \text{NH}_3(\text{g}) \rightarrow \text{S}_4\text{N}_4(\text{s}) + \text{S}_8(\text{s}) + 12 \text{NH}_4\text{Cl}(\text{s})$

3.55 Plan: The names must first be converted to chemical formulas. The balancing is a trial and error procedure. Do one blank/one element at a time.

Solution:

- a) $4 \text{Ga}(\text{s}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{Ga}_2\text{O}_3(\text{s})$
 b) $2 \text{C}_6\text{H}_{14}(\text{l}) + 19 \text{O}_2(\text{g}) \rightarrow 12 \text{CO}_2(\text{g}) + 14 \text{H}_2\text{O}(\text{g})$
 c) $3 \text{CaCl}_2(\text{aq}) + 2 \text{Na}_3\text{PO}_4(\text{aq}) \rightarrow \text{Ca}_3(\text{PO}_4)_2(\text{s}) + 6 \text{NaCl}(\text{aq})$
- 3.56 a) $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2 \text{KI}(\text{aq}) \rightarrow \text{PbI}_2(\text{s}) + 2 \text{KNO}_3(\text{aq})$
 b) $\text{Si}_2\text{Cl}_6(\text{l}) + 4 \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{SiO}_2(\text{s}) + 6 \text{HCl}(\text{g}) + \text{H}_2(\text{g})$
 c) $3 \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{HNO}_3(\text{aq}) + \text{NO}(\text{g})$

3.57 The *stoichiometrically equivalent molar ratio* is the ratio of the coefficients in the balanced equation. The molar ratio is equivalent to the mass ratios in mass units. This can be used to calculate masses of reactants or products in a chemical reaction.

3.58 First, write a balanced equation: $a \text{A} + b \text{B} \rightarrow c \text{C}$

$$\text{Mass C} = (5 \text{ g A}) \left(\frac{1 \text{ mol A}}{\text{Molar Mass A}} \right) \left(\frac{c \text{ Mol C}}{a \text{ Mol A}} \right) \left(\frac{\text{Molar Mass C}}{\text{Mol C}} \right) = \text{g C}$$

3.59 Plan: First, write a balanced chemical equation. Determine if D or E is the limiting reagent. Use the limiting reagent to determine the amount of F formed.

Solution:

The balanced equation is: $d \text{D} + e \text{E} \rightarrow f \text{F}$

From the relationship $\text{moles} = \text{g}/\mathcal{M}$, calculate the moles of D and E present. By comparing the ratio of moles present to that needed in the balanced equation, determine the limiting reagent (LR). Based on the number of moles of limiting reagent, proceed as follows.

$$\text{Mass F} = (\text{g LR}) \left(\frac{1 \text{ mol LR}}{\text{Molar Mass LR}} \right) \left(\frac{f \text{ Mol F}}{\text{Mol LR}} \right) \left(\frac{\text{Molar Mass F}}{\text{Mol F}} \right) = \text{g F}$$

3.60 The percent yield is the ratio of actual to 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.

3.61 Plan: Use the mole ratio from the balanced chemical equation to determine the moles produced. Use the moles with the molar mass to determine the grams produced.

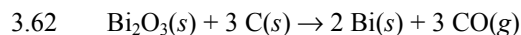
Solution:

$$\text{a) Moles Cl}_2 = (1.82 \text{ mol HCl}) \left(\frac{1 \text{ mol Cl}_2}{4 \text{ mol HCl}} \right) = \mathbf{0.455 \text{ mol Cl}_2}$$

$$\text{b) Grams Cl}_2 = (1.82 \text{ mol HCl}) \left(\frac{1 \text{ mol Cl}_2}{4 \text{ mol HCl}} \right) \left(\frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} \right) = 32.2595 = \mathbf{32.3 \text{ g Cl}_2}$$

The beginning of the calculation is repeated to emphasize that the second part of the problem is simply an extension of the first part. There is no need to repeat the entire calculation, as only the final step times the answer of the first part will give the final answer to this part.

Hint: Always check to see if the initial equation is balanced. If the equation is not balanced, it should be balanced before proceeding.



$$\text{a) Moles Bi}_2\text{O}_3 = (352 \text{ g Bi}_2\text{O}_3) \left(\frac{1 \text{ mol Bi}_2\text{O}_3}{466.0 \text{ g Bi}_2\text{O}_3} \right) = 0.7553648 = \mathbf{0.755 \text{ mol Bi}_2\text{O}_3}$$

$$\text{b) Moles Bi} = (352 \text{ g Bi}_2\text{O}_3) \left(\frac{1 \text{ mol Bi}_2\text{O}_3}{466.0 \text{ g Bi}_2\text{O}_3} \right) \left(\frac{2 \text{ mol Bi}}{1 \text{ mol Bi}_2\text{O}_3} \right) = 1.5107 = \mathbf{1.51 \text{ mol Bi}}$$

3.63 Plan: Convert the kilograms of oxygen to the moles of oxygen. Use the moles of oxygen and the mole ratios from the balanced chemical equation to determine the moles of KNO_3 . The moles of KNO_3 and its molar mass will give the grams.

Solution:

$$\text{a) Moles KNO}_3 = (88.6 \text{ kg O}_2) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \left(\frac{4 \text{ mol KNO}_3}{5 \text{ mol O}_2} \right) = 2215 = \mathbf{2.22 \times 10^3 \text{ mol KNO}_3}$$

$$\begin{aligned} \text{b) Grams KNO}_3 &= (88.6 \text{ kg O}_2) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \left(\frac{4 \text{ mol KNO}_3}{5 \text{ mol O}_2} \right) \left(\frac{101.11 \text{ g KNO}_3}{1 \text{ mol KNO}_3} \right) \\ &= 223958.65 = \mathbf{2.24 \times 10^5 \text{ g KNO}_3} \end{aligned}$$

The beginning of the calculation is repeated to emphasize that the second part of the problem is simply an extension of the first part. There is no need to repeat the entire calculation, as only the final step times the answer of the first part will give the final answer to this part.

3.64 a) Moles $\text{Cr}_2\text{O}_3 = (421 \text{ g Cr}_2\text{S}_3) \left(\frac{1 \text{ mol Cr}_2\text{S}_3}{200.21 \text{ g Cr}_2\text{S}_3} \right) \left(\frac{1 \text{ mol Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{S}_3} \right) = 2.10279 = \mathbf{2.10 \text{ mol Cr}_2\text{O}_3}$

$$\begin{aligned} \text{b) Grams Cr}_2\text{O}_3 &= (421 \text{ g Cr}_2\text{S}_3) \left(\frac{1 \text{ mol Cr}_2\text{S}_3}{200.21 \text{ g Cr}_2\text{S}_3} \right) \left(\frac{1 \text{ mol Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{S}_3} \right) \left(\frac{152.00 \text{ g Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{O}_3} \right) \\ &= 319.624 = \mathbf{3.20 \times 10^2 \text{ g Cr}_2\text{O}_3} \end{aligned}$$

3.65 Plan: First, balance the equation. Convert the grams of diborane to moles of diborane using its molar mass. Use mole ratios from the balanced chemical equation to determine the moles of the products. Use the moles and molar mass of each product to determine the mass formed.

Solution:

The balanced equation is: $\text{B}_2\text{H}_6(g) + 6 \text{ H}_2\text{O}(l) \rightarrow 2 \text{ H}_3\text{BO}_3(s) + 6 \text{ H}_2(g)$.

$$\begin{aligned} \text{Mass H}_3\text{BO}_3 &= (33.61 \text{ g B}_2\text{H}_6) \left(\frac{1 \text{ mol B}_2\text{H}_6}{27.67 \text{ g B}_2\text{H}_6} \right) \left(\frac{2 \text{ mol H}_3\text{BO}_3}{1 \text{ mol B}_2\text{H}_6} \right) \left(\frac{61.83 \text{ g H}_3\text{BO}_3}{1 \text{ mol H}_3\text{BO}_3} \right) \\ &= 150.206 = \mathbf{150.2 \text{ g H}_3\text{BO}_3} \end{aligned}$$

$$\begin{aligned} \text{Mass H}_2 &= (33.61 \text{ g B}_2\text{H}_6) \left(\frac{1 \text{ mol B}_2\text{H}_6}{27.67 \text{ g B}_2\text{H}_6} \right) \left(\frac{6 \text{ mol H}_2}{1 \text{ mol B}_2\text{H}_6} \right) \left(\frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) \\ &= 14.69268 = \mathbf{14.69 \text{ g H}_2} \end{aligned}$$

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

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

$$\text{Grams H}_2\text{S} = 174 \text{ g Ag}_2\text{S} \left(\frac{1 \text{ mol Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}} \right) \left(\frac{1 \text{ mol H}_2\text{S}}{1 \text{ mol Ag}_2\text{S}} \right) \left(\frac{34.09 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \right) = 23.9276 = \mathbf{23.9 \text{ g H}_2\text{S}}$$

3.67 Plan: Write the balanced equation by first writing the formulas for the reactants and products. Reactants: formula for phosphorus is given as P_4 and formula for chlorine gas is Cl_2 (chlorine occurs as a diatomic molecule). Products: formula for phosphorus pentachloride — the name indicates one phosphorus atom and five chlorine atoms to give the formula PCl_5 . Convert the mass of phosphorus to grams, use the mole ratio from the balanced chemical equation, and finally use the molar mass of chlorine to get the mass of chlorine.

Solution:

Formulas give the equation: $\text{P}_4 + \text{Cl}_2 \rightarrow \text{PCl}_5$

Balancing the equation: $\text{P}_4 + 10 \text{Cl}_2 \rightarrow 4 \text{PCl}_5$

$$\text{Grams Cl}_2 = (355 \text{ g P}_4) \left(\frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} \right) \left(\frac{10 \text{ mol Cl}_2}{1 \text{ mol P}_4} \right) \left(\frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} \right) = 2031.76 = \mathbf{2.03 \times 10^3 \text{ g Cl}_2}$$

3.68 First, balance the equation: $\text{S}_8(s) + 24 \text{F}_2(g) \rightarrow 8 \text{SF}_6(s)$

$$\text{Grams F}_2 = (17.8 \text{ g S}_8) \left(\frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \right) \left(\frac{24 \text{ mol F}_2}{1 \text{ mol S}_8} \right) \left(\frac{38.00 \text{ g F}_2}{1 \text{ mol F}_2} \right) = 63.27409 = \mathbf{63.3 \text{ g F}_2}$$

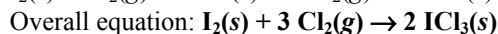
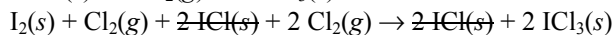
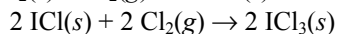
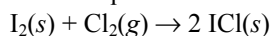
3.69 Plan: Begin by writing the chemical formulas of the reactants and products in each step. Next, balance each of the equations. Combine the equations for the separate steps by adjusting the equations so the intermediate (iodine monochloride) cancels. Finally, change the mass of product to mole and use the mole ratio and molar mass of iodine to determine the mass of iodine.

Solution:

a) First step: $\text{I}_2(s) + \text{Cl}_2(g) \rightarrow 2 \text{ICl}(s)$

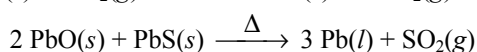
Second step: $\text{ICl}(s) + \text{Cl}_2(g) \rightarrow \text{ICl}_3(s)$

b) Multiply the coefficients of the second equation by 2, so that $\text{ICl}(s)$, an intermediate product, can be eliminated from the overall equation.



$$\begin{aligned} \text{c) Grams I}_2 &= (31.4 \text{ kg ICl}_3) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) \left(\frac{1 \text{ mol I}_2}{2 \text{ mol ICl}_3} \right) \left(\frac{253.8 \text{ g I}_2}{1 \text{ mol I}_2} \right) \\ &= 17086.88 = \mathbf{1.71 \times 10^4 \text{ g I}_2} \end{aligned}$$

3.70 a) $2 \text{PbS}(s) + 3 \text{O}_2(g) \xrightarrow{\Delta} 2 \text{PbO}(s) + 2 \text{SO}_2(g)$



b) $\text{PbS}(s) + \text{O}_2(g) \xrightarrow{\Delta} \text{Pb}(l) + \text{SO}_2(g)$

c) 1 metric ton = tonne = 1000 kg

$$\begin{aligned} \text{Tonne SO}_2 &= (\text{tonne Pb}) \left(\frac{10^3 \text{ kg}}{1 \text{ tonne}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \right) \left(\frac{1 \text{ mol SO}_2}{1 \text{ mol Pb}} \right) \left(\frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} \right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left(\frac{1 \text{ tonne}}{10^3 \text{ kg}} \right) \\ &= 0.309218 = \mathbf{0.3092 \text{ tonne SO}_2} \end{aligned}$$

3.71 Plan: Convert the given masses to moles and use the mole ratio from the balanced chemical equation to find the moles of CaO that will form. The reactant that produces the least moles of CaO is the limiting reactant. Convert the moles of CaO from the limiting reactant to grams using the molar mass.

Solution:

a) Moles CaO from Ca = $(4.20 \text{ g Ca}) \left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) \left(\frac{2 \text{ mol CaO}}{2 \text{ mol Ca}} \right) = 0.104790 = \mathbf{0.105 \text{ mol CaO}}$

b) Moles CaO from O₂ = $(2.80 \text{ g O}_2) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \left(\frac{2 \text{ mol CaO}}{1 \text{ mol O}_2} \right) = 0.17500 = \mathbf{0.175 \text{ mol CaO}}$

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

d) Grams CaO = $(4.20 \text{ g Ca}) \left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) \left(\frac{2 \text{ mol CaO}}{2 \text{ mol Ca}} \right) \left(\frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} \right) = 5.8766 = \mathbf{5.88 \text{ g CaO}}$

3.72 $\text{SrH}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l}) \rightarrow \text{Sr}(\text{OH})_2(\text{s}) + 2 \text{H}_2(\text{g})$

a) Moles H₂ from SrH₂ = $(5.63 \text{ g SrH}_2) \left(\frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) \left(\frac{2 \text{ mol H}_2}{1 \text{ mol SrH}_2} \right) = 0.125614 = \mathbf{0.126 \text{ mol H}_2}$

b) Moles H₂ from H₂O = $(4.80 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \right) = 0.26637 = \mathbf{0.266 \text{ mol H}_2}$

c) SrH₂ is the limiting reagent since it will yield fewer moles of hydrogen gas.

d) Grams H₂ = $(5.63 \text{ g SrH}_2) \left(\frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) \left(\frac{2 \text{ mol H}_2}{1 \text{ mol SrH}_2} \right) \left(\frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) = 0.2532369 = \mathbf{0.253 \text{ g H}_2}$

3.73 Plan: First, balance the chemical equation. Determine which of the reactants is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of material formed and the amount of the other reactant used. The difference between the amount of reactant used and the initial reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation for this reaction is:



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

Next, find the limiting reactant by using the molar ratio to find the smaller number of moles of HIO₃ that can be produced from each reactant given and excess of the other:

Moles HIO₃ from ICl₃ = $(685 \text{ g ICl}_3) \left(\frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) \left(\frac{1 \text{ mol HIO}_3}{2 \text{ mol ICl}_3} \right) = 1.468696 = 1.47 \text{ mol HIO}_3$

Moles HIO₃ from H₂O = $(117.4 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{1 \text{ mol HIO}_3}{3 \text{ mol H}_2\text{O}} \right) = 2.17166 = 2.17 \text{ mol HIO}_3$

ICl₃ is the limiting reagent and will produce **1.47 mol HIO₃**. Use the limiting reagent to find the grams of HIO₃ formed.

Grams HIO₃ = $(685 \text{ g ICl}_3) \left(\frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) \left(\frac{1 \text{ mol HIO}_3}{2 \text{ mol ICl}_3} \right) \left(\frac{175.9 \text{ g HIO}_3}{1 \text{ mol HIO}_3} \right) = 258.355 = \mathbf{258 \text{ g HIO}_3}$

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

$$\begin{aligned} \text{Remaining H}_2\text{O} &= 117.4 \text{ g H}_2\text{O} - (685 \text{ g ICl}_3) \left(\frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) \left(\frac{3 \text{ mol H}_2\text{O}}{2 \text{ mol ICl}_3} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) \\ &= 38.002 = \mathbf{38.0 \text{ g H}_2\text{O}} \end{aligned}$$

- 3.74 First balance the equation: $\text{Al}_2\text{S}_3 + 6 \text{ H}_2\text{O}(l) \rightarrow 2 \text{ Al}(\text{OH})_3(s) + 3 \text{ H}_2\text{S}(g)$
Determine the limiting reactant:

$$\text{Moles H}_2\text{S from Al}_2\text{S}_3 = (158 \text{ g Al}_2\text{S}_3) \left(\frac{1 \text{ mol Al}_2\text{S}_3}{150.17 \text{ g Al}_2\text{S}_3} \right) \left(\frac{3 \text{ mol H}_2\text{S}}{1 \text{ mol Al}_2\text{S}_3} \right) = 3.15642 = 3.16 \text{ mol H}_2\text{S}$$

$$\text{Moles H}_2\text{S from H}_2\text{O} = (131 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{3 \text{ mol H}_2\text{S}}{6 \text{ mol H}_2\text{O}} \right) = 3.63485 = 3.63 \text{ mol H}_2\text{S}$$

Al₂S₃ is the limiting reagent and **3.16 mol of H₂S** will form.

Using the limiting reagent, we can calculate the mass of H₂S that will form.

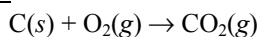
$$\text{Grams H}_2\text{S} = (158 \text{ g Al}_2\text{S}_3) \left(\frac{1 \text{ mol Al}_2\text{S}_3}{150.17 \text{ g Al}_2\text{S}_3} \right) \left(\frac{3 \text{ mol H}_2\text{S}}{1 \text{ mol Al}_2\text{S}_3} \right) \left(\frac{34.09 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \right) = 107.602 = \mathbf{108 \text{ g H}_2\text{S}}$$

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

$$\begin{aligned} \text{Remaining H}_2\text{O} &= 131 \text{ g H}_2\text{O} - (158 \text{ g Al}_2\text{S}_3) \left(\frac{1 \text{ mol Al}_2\text{S}_3}{150.17 \text{ g Al}_2\text{S}_3} \right) \left(\frac{6 \text{ mol H}_2\text{O}}{1 \text{ mol Al}_2\text{S}_3} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) \\ &= 17.24 = \mathbf{17 \text{ g H}_2\text{O}} \end{aligned}$$

- 3.75 Plan: Write the balanced equation: formula for carbon is C, formula for oxygen is O₂ and formula for carbon dioxide is CO₂. Determine the limiting reagent by seeing which reactant will yield the smaller amount of product. The limiting reagent is used for all subsequent calculations.

Solution:



$$\text{Moles CO}_2 \text{ from C} = (0.100 \text{ mol C}) \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \right) = 0.100 \text{ mol CO}_2$$

$$\text{Moles CO}_2 \text{ from O}_2 = (8.00 \text{ g O}_2) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol O}_2} \right) = 2.5000 = 2.50 \text{ mol CO}_2$$

The C is the limiting reactant and will be used to determine the amount of CO₂ that will form.

$$\text{Grams CO}_2 = (0.100 \text{ mol C}) \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \right) \left(\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 4.401 = \mathbf{4.40 \text{ g CO}_2}$$

Since the C is limiting, the **O₂ is in excess**. The amount remaining depends on how much combines with the limiting reagent.

$$\text{Remaining O}_2 = 8.00 \text{ g O}_2 - (0.100 \text{ mol C}) \left(\frac{1 \text{ mol O}_2}{1 \text{ mol C}} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 4.8000 = \mathbf{4.80 \text{ g O}_2}$$

- 3.76 First, balance the equation: $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{l})$
Determine the limiting reagent:

$$\text{Mole H}_2\text{O from H}_2 = (0.0359 \text{ g H}_2) \left(\frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \right) \left(\frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \right) = 0.0178075 = 0.0178 \text{ mol H}_2\text{O}$$

$$\text{Mole H}_2\text{O from O}_2 = (0.0175 \text{ mol O}_2) \left(\frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \right) = 0.0350 \text{ mol H}_2\text{O}$$

The hydrogen is the limiting reactant, and will be used to determine the amount of water that will form.

$$\text{Grams H}_2\text{O} = (0.0359 \text{ g H}_2) \left(\frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \right) \left(\frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 0.32089 = \mathbf{0.321 \text{ g H}_2\text{O}}$$

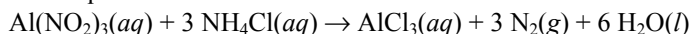
Since the hydrogen is limiting; the **oxygen must be excess reactant**. The amount of excess reactant is determined from the limiting reactant.

$$\begin{aligned} \text{Remaining O}_2 &= (0.0175 \text{ mol O}_2) \left(\frac{32.00 \text{ mol O}_2}{1 \text{ mol O}_2} \right) - (0.0359 \text{ g H}_2) \left(\frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \right) \left(\frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) \\ &= 0.275079 = \mathbf{0.28 \text{ g O}_2} \end{aligned}$$

- 3.77 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. Using the molar relationships from the balanced equation, determine which reactant is limiting. Any product can be used to predict the limiting reactant; in this case, AlCl_3 is used. Additional significant figures are retained until the last step.

Solution:

The balanced chemical equation is:



Now determine the limiting reagent. We will use the moles of AlCl_3 produced to determine which is limiting.

$$\begin{aligned} \text{Mole AlCl}_3 \text{ from Al}(\text{NO}_2)_3 &= (62.5 \text{ g Al}(\text{NO}_2)_3) \left(\frac{1 \text{ mol Al}(\text{NO}_2)_3}{165.01 \text{ g Al}(\text{NO}_2)_3} \right) \left(\frac{1 \text{ mol AlCl}_3}{1 \text{ mol Al}(\text{NO}_2)_3} \right) \\ &= 0.37876 = 0.379 \text{ mol AlCl}_3 \end{aligned}$$

$$\text{Mole AlCl}_3 \text{ from NH}_4\text{Cl} = (54.6 \text{ g NH}_4\text{Cl}) \left(\frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} \right) \left(\frac{1 \text{ mol AlCl}_3}{3 \text{ mol NH}_4\text{Cl}} \right) = 0.34025 = 0.340 \text{ mol AlCl}_3$$

Ammonium chloride is the limiting reactant, and it is important for all subsequent calculations.

Mass of substances after the reaction:

$\text{Al}(\text{NO}_2)_3$:

$$\begin{aligned} 62.5 \text{ g Al}(\text{NO}_2)_3 - (54.6 \text{ g NH}_4\text{Cl}) \left(\frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} \right) \left(\frac{1 \text{ mol Al}(\text{NO}_2)_3}{3 \text{ mol NH}_4\text{Cl}} \right) \left(\frac{165.01 \text{ g Al}(\text{NO}_2)_3}{1 \text{ mol Al}(\text{NO}_2)_3} \right) \\ = 6.35526 = 6.4 \text{ g Al}(\text{NO}_2)_3 \end{aligned}$$

(It is the limiting reactant.)

AlCl_3 :

$$(54.6 \text{ g NH}_4\text{Cl}) \left(\frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} \right) \left(\frac{1 \text{ mol AlCl}_3}{3 \text{ mol NH}_4\text{Cl}} \right) \left(\frac{133.33 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} \right) = 45.3656 = \mathbf{45.4 \text{ g AlCl}_3}$$

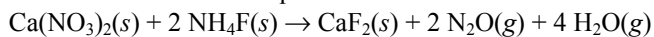
N_2 :

$$(54.6 \text{ g NH}_4\text{Cl}) \left(\frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} \right) \left(\frac{3 \text{ mol N}_2}{3 \text{ mol NH}_4\text{Cl}} \right) \left(\frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} \right) = 28.601 = \mathbf{28.6 \text{ g N}_2}$$

H_2O :

$$(54.6 \text{ g NH}_4\text{Cl}) \left(\frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} \right) \left(\frac{6 \text{ mol H}_2\text{O}}{3 \text{ mol NH}_4\text{Cl}} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 36.7879 = \mathbf{36.8 \text{ g H}_2\text{O}}$$

3.78 The balanced chemical equation is:



Now determine the limiting reagent. We will use the moles of CaF_2 produced to determine which is limiting.

$$\begin{aligned} \text{Mole CaF}_2 \text{ from Ca}(\text{NO}_3)_2 &= (16.8 \text{ g Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol Ca}(\text{NO}_3)_2}{164.10 \text{ g Ca}(\text{NO}_3)_2} \right) \left(\frac{1 \text{ mol CaF}_2}{1 \text{ mol Ca}(\text{NO}_3)_2} \right) \\ &= 0.1023766 = 0.102 \text{ mol CaF}_2 \end{aligned}$$

$$\text{Mole CaF}_2 \text{ from NH}_4\text{F} = (17.50 \text{ g NH}_4\text{F}) \left(\frac{1 \text{ mol NH}_4\text{F}}{37.04 \text{ g NH}_4\text{F}} \right) \left(\frac{1 \text{ mol CaF}_2}{2 \text{ mol NH}_4\text{F}} \right) = 0.236231 = 0.236 \text{ mol CaF}_2$$

Calcium nitrate is the limiting reactant, and it is important for all subsequent calculations.

Mass of substances after the reaction:

$\text{Ca}(\text{NO}_3)_2$: **None** (It is the limiting reactant.)

NH_4F :

$$\begin{aligned} 17.50 \text{ g NH}_4\text{F} - (16.8 \text{ g Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol Ca}(\text{NO}_3)_2}{164.10 \text{ g Ca}(\text{NO}_3)_2} \right) \left(\frac{2 \text{ mol NH}_4\text{F}}{1 \text{ mol Ca}(\text{NO}_3)_2} \right) \left(\frac{37.04 \text{ g NH}_4\text{F}}{1 \text{ mol NH}_4\text{F}} \right) \\ = 9.9159 = \mathbf{9.92 \text{ g NH}_4\text{F}} \end{aligned}$$

CaF_2 :

$$(16.8 \text{ g Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol Ca}(\text{NO}_3)_2}{164.10 \text{ g Ca}(\text{NO}_3)_2} \right) \left(\frac{1 \text{ mol CaF}_2}{1 \text{ mol Ca}(\text{NO}_3)_2} \right) \left(\frac{78.08 \text{ g CaF}_2}{1 \text{ mol CaF}_2} \right) = 7.99356 = \mathbf{7.99 \text{ g CaF}_2}$$

N_2O :

$$(16.8 \text{ g Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol Ca}(\text{NO}_3)_2}{164.10 \text{ g Ca}(\text{NO}_3)_2} \right) \left(\frac{2 \text{ mol N}_2\text{O}}{1 \text{ mol Ca}(\text{NO}_3)_2} \right) \left(\frac{44.02 \text{ g N}_2\text{O}}{1 \text{ mol N}_2\text{O}} \right) = 9.0132 = \mathbf{9.01 \text{ g N}_2\text{O}}$$

H_2O :

$$(16.8 \text{ g Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol Ca}(\text{NO}_3)_2}{164.10 \text{ g Ca}(\text{NO}_3)_2} \right) \left(\frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol Ca}(\text{NO}_3)_2} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 7.3793 = \mathbf{7.38 \text{ g H}_2\text{O}}$$

3.79 Plan: Multiply the yield of the first step by that of the second step to get the overall yield.

Solution:

It is simpler to use the decimal equivalents of the percent yields, and then convert to percent using 100%.

$$(0.82) (0.65) (100\%) = 53.3 = \mathbf{53\%}$$

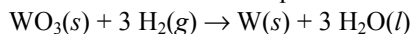
3.80 It is simpler to use the decimal equivalents of the percent yields, and then convert to percent using 100%.

$$(0.48) (0.73) (100\%) = 35.04 = \mathbf{35\%}$$

3.81 Plan: Balance the chemical equation using the formulas of the substances. Determine the yield (theoretical yield) for the reaction from the mass of tungsten(VI) oxide. The actual yield divided by the yield just calculated (with the result multiplied by 100%) gives the percent yield.

Solution: (Rounding to the correct number of significant figures will be postponed until the final result.)

The balanced chemical equation is:



Theoretical yield of H_2O :

$$(41.5 \text{ g WO}_3) \left(\frac{1 \text{ mol WO}_3}{231.9 \text{ g WO}_3} \right) \left(\frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol WO}_3} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 9.67439 \text{ g H}_2\text{O}$$

Actual yield, in grams, of H_2O :

$$(9.50 \text{ mL H}_2\text{O}) \left(\frac{1.00 \text{ g H}_2\text{O}}{1 \text{ mL H}_2\text{O}} \right) = 9.50 \text{ g H}_2\text{O}$$

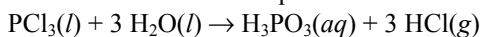
Calculate the percent yield:

$$\left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\% = \left(\frac{9.50 \text{ g H}_2\text{O}}{9.67439 \text{ g H}_2\text{O}} \right) \times 100\% = 98.1974 = \mathbf{98.2\%}$$

- 3.82 Balance the chemical equation using the formulas of the substances. Determine the yield (theoretical yield) for the reaction from the mass of tungsten(VI) oxide. The actual yield divided by the yield just calculated (with the result multiplied by 100%) gives the percent yield.

(Rounding to the correct number of significant figures will be postponed until the final result.)

The balanced chemical equation is:



Theoretical yield of H₂O:

$$(200. \text{ g PCl}_3) \left(\frac{1 \text{ mol PCl}_3}{137.32 \text{ g PCl}_3} \right) \left(\frac{3 \text{ mol HCl}}{1 \text{ mol PCl}_3} \right) \left(\frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} \right) = 159.3067 \text{ g HCl (Theoretical yield)}$$

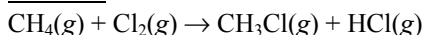
The actual yield is given (128 g HCl).

Calculate the percent yield:

$$\left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\% = \left(\frac{128 \text{ g HCl}}{159.3067 \text{ g HCl}} \right) \times 100\% = 83.3481586 = \mathbf{83.3\%}$$

- 3.83 Plan: Write the balanced chemical equation. Since quantities of reactants are present, we must determine which is limiting. Only 80.0% of the calculated amounts of products will form. (Rounding to the correct number of significant figures will be postponed until the final result.)

Solution:



$$\text{Mole HCl from CH}_4 = (18.5 \text{ g CH}_4) \left(\frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \right) \left(\frac{1 \text{ mol CH}_3\text{Cl}}{1 \text{ mol CH}_4} \right) = 1.153367 \text{ mol CH}_3\text{Cl}$$

$$\text{Mole HCl from Cl}_2 = (43.0 \text{ g Cl}_2) \left(\frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \right) \left(\frac{1 \text{ mol CH}_3\text{Cl}}{1 \text{ mol Cl}_2} \right) = 0.606488 \text{ mol CH}_3\text{Cl}$$

Chlorine is the limiting reactant.

$$\begin{aligned} \text{Grams CH}_3\text{Cl} &= (43.0 \text{ g Cl}_2) \left(\frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \right) \left(\frac{1 \text{ mol CH}_3\text{Cl}}{1 \text{ mol Cl}_2} \right) \left(\frac{50.48 \text{ g CH}_3\text{Cl}}{1 \text{ mol CH}_3\text{Cl}} \right) \left(\frac{80.0\%}{100\%} \right) \\ &= 24.4924 = \mathbf{24.5 \text{ g CH}_3\text{Cl}} \end{aligned}$$

The beginning of the calculation is repeated to emphasize that the second part of the problem is simply an extension of the first part. There is no need to repeat the entire calculation as only the final step(s) times the answer of the first part will give the final answer to this part.

- 3.84 First, balance the chemical equation: $3 \text{Ca}(s) + \text{N}_2(g) \rightarrow \text{Ca}_3\text{N}_2(s)$

$$\text{Mole Ca}_3\text{N}_2 \text{ from Ca} = (56.6 \text{ g Ca}) \left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) \left(\frac{1 \text{ mol Ca}_3\text{N}_2}{3 \text{ mol Ca}} \right) = 0.470725 \text{ mol Ca}_3\text{N}_2$$

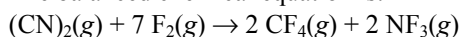
$$\text{Mole Ca}_3\text{N}_2 \text{ from N}_2 = (30.5 \text{ g N}_2) \left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} \right) \left(\frac{1 \text{ mol Ca}_3\text{N}_2}{1 \text{ mol N}_2} \right) = 1.0885 \text{ mol Ca}_3\text{N}_2$$

Ca is the limiting reactant.

$$\begin{aligned} \text{Grams Ca}_3\text{N}_2 &= (56.6 \text{ g Ca}) \left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) \left(\frac{1 \text{ mol Ca}_3\text{N}_2}{3 \text{ mol Ca}} \right) \left(\frac{148.26 \text{ g Ca}_3\text{N}_2}{1 \text{ mol Ca}_3\text{N}_2} \right) \left(\frac{93.0\%}{100\%} \right) \\ &= 64.90444 = \mathbf{64.9 \text{ g Ca}_3\text{N}_2} \end{aligned}$$

3.85 Plan: The first step is to determine the chemical formulas so a balanced chemical equation can be written. The limiting reactant must be determined. Finally, the mass of CF_4 is determined from the limiting reactant.

Solution: (Rounding to the correct number of significant figures will be postponed until the final result.)
The balanced chemical equation is:



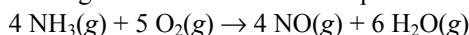
$$\text{Mole CF}_4 \text{ from } (\text{CN})_2 = (80.0 \text{ g } (\text{CN})_2) \left(\frac{1 \text{ mol } (\text{CN})_2}{52.04 \text{ g } (\text{CN})_2} \right) \left(\frac{2 \text{ mol CF}_4}{1 \text{ mol } (\text{CN})_2} \right) = 3.074558 \text{ mol CF}_4$$

$$\text{Mole CF}_4 \text{ from } \text{F}_2 = (80.0 \text{ g } \text{F}_2) \left(\frac{1 \text{ mol } \text{F}_2}{38.00 \text{ g } \text{F}_2} \right) \left(\frac{2 \text{ mol CF}_4}{7 \text{ mol } \text{F}_2} \right) = 0.60150 \text{ mol CF}_4$$

F_2 is the limiting reactant, and will be used to calculate the yield.

$$\text{Grams CF}_4 = (80.0 \text{ g } \text{F}_2) \left(\frac{1 \text{ mol } \text{F}_2}{38.00 \text{ g } \text{F}_2} \right) \left(\frac{2 \text{ mol CF}_4}{7 \text{ mol } \text{F}_2} \right) \left(\frac{88.01 \text{ g CF}_4}{1 \text{ mol CF}_4} \right) = 52.9383 = \mathbf{52.9 \text{ g CF}_4}$$

3.86 Start by deriving the balanced chemical equation:



Determine the limiting reactant. (Rounding to the correct number of significant figures will be postponed until the final result.)

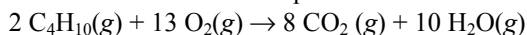
$$\text{Mole NO from } \text{NH}_3 = (466 \text{ g } \text{NH}_3) \left(\frac{1 \text{ mol } \text{NH}_3}{17.03 \text{ g } \text{NH}_3} \right) \left(\frac{4 \text{ mol NO}}{4 \text{ mol } \text{NH}_3} \right) = 27.36348 \text{ mol NO}$$

$$\text{Mole NO from } \text{O}_2 = (812 \text{ g } \text{O}_2) \left(\frac{1 \text{ mol } \text{O}_2}{32.00 \text{ g } \text{O}_2} \right) \left(\frac{4 \text{ mol NO}}{5 \text{ mol } \text{O}_2} \right) = 20.3 \text{ mol NO}$$

O_2 is the limiting reactant, and will be used to calculate the yield.

$$\text{Grams NO} = (812 \text{ g } \text{O}_2) \left(\frac{1 \text{ mol } \text{O}_2}{32.00 \text{ g } \text{O}_2} \right) \left(\frac{4 \text{ mol NO}}{5 \text{ mol } \text{O}_2} \right) \left(\frac{30.01 \text{ g NO}}{1 \text{ mol NO}} \right) = 609.203 = \mathbf{609 \text{ g NO}}$$

3.87 Determine the balanced chemical equation:



$$\begin{aligned} \text{a) Grams O}_2 &= (6.50 \text{ mL } \text{C}_4\text{H}_{10}) \left(\frac{0.579 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mol } \text{C}_4\text{H}_{10}}{58.12 \text{ g } \text{C}_4\text{H}_{10}} \right) \left(\frac{13 \text{ mol O}_2}{2 \text{ mol } \text{C}_4\text{H}_{10}} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) \\ &= 13.4688 = \mathbf{13.5 \text{ g O}_2} \end{aligned}$$

$$\begin{aligned} \text{b) Moles CO}_2 &= (6.50 \text{ mL } \text{C}_4\text{H}_{10}) \left(\frac{0.579 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mol } \text{C}_4\text{H}_{10}}{58.12 \text{ g } \text{C}_4\text{H}_{10}} \right) \left(\frac{8 \text{ mol CO}_2}{2 \text{ mol } \text{C}_4\text{H}_{10}} \right) \\ &= 0.259016 = \mathbf{0.259 \text{ mol CO}_2} \end{aligned}$$

$$\begin{aligned} \text{c) Total moles} &= (6.50 \text{ mL } \text{C}_4\text{H}_{10}) \left(\frac{0.579 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mol } \text{C}_4\text{H}_{10}}{58.12 \text{ g } \text{C}_4\text{H}_{10}} \right) \left(\frac{8 \text{ mol CO}_2}{2 \text{ mol } \text{C}_4\text{H}_{10}} \right) + \\ &\quad (6.50 \text{ mL } \text{C}_4\text{H}_{10}) \left(\frac{0.579 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mol } \text{C}_4\text{H}_{10}}{58.12 \text{ g } \text{C}_4\text{H}_{10}} \right) \left(\frac{10 \text{ mol H}_2\text{O}}{2 \text{ mol } \text{C}_4\text{H}_{10}} \right) = 0.5827858 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Total molecules} &= (0.5827858 \text{ mol}) \left(\frac{6.002 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \right) \\ &= 3.5095 \times 10^{23} = \mathbf{3.51 \times 10^{23} \text{ molecules}} \end{aligned}$$

- 3.88 Balancing the chemical equation gives:
 $2 \text{NaH}(s) + \text{B}_2\text{H}_6(g) \rightarrow 2 \text{NaBH}_4(s)$
 Determine the limiting reactant. (Rounding to the correct number of significant figures will be postponed until the final result.)

$$\text{Mole NaBH}_4 \text{ from NaH} = (7.88 \text{ g NaH}) \left(\frac{1 \text{ mol NaH}}{24.00 \text{ g NaH}} \right) \left(\frac{2 \text{ mol NaBH}_4}{2 \text{ mol NaH}} \right) = 0.32833 \text{ mol NaBH}_4$$

$$\text{Mole NaBH}_4 \text{ from B}_2\text{H}_6 = (8.12 \text{ g B}_2\text{H}_6) \left(\frac{1 \text{ mol B}_2\text{H}_6}{27.73 \text{ g B}_2\text{H}_6} \right) \left(\frac{2 \text{ mol NaBH}_4}{1 \text{ mol B}_2\text{H}_6} \right) = 0.585647 \text{ mol NaBH}_4$$

NaH is the limiting reactant, and will be used to calculate the yield.

$$(7.88 \text{ g NaH}) \left(\frac{1 \text{ mol NaH}}{24.00 \text{ g NaH}} \right) \left(\frac{2 \text{ mol NaBH}_4}{2 \text{ mol NaH}} \right) \left(\frac{37.83 \text{ g NaBH}_4}{1 \text{ mol NaBH}_4} \right) \left(\frac{95.5\%}{100\%} \right) = 11.8619 = \mathbf{11.9 \text{ g NaBH}_4}$$

- 3.89 Plan: The spheres represent particles of solute and the amount of *solute* per given volume of *solution* determines its concentration.

Solution:

- Box C has more solute added because it contains 2 more spheres than Box A contains.
- Box B has more solvent because solvent molecules have displaced two solute molecules.
- Box C has a higher molarity, because it has more moles of solute per volume of solution.
- Box B has a lower concentration (and molarity), because it has fewer moles of solute per volume of solution.

3.90
$$\text{Molarity} = \frac{\text{moles Solute}}{\text{Liters Solution}}$$

The amount of solute = concentration x volume of solution

Moles solute = molarity x liters of solution

Mass solute = molarity x liters of solution x molar mass

- 3.91 Volumes may **not** be additive when two different solutions are mixed, so the final volume may be slightly different from 1000.0 mL. The correct method would state, "Take 100.0 mL of the 10.0 M solution and add water until the total volume is 1000 mL."

- 3.92 Plan: In all cases, the definition of molarity (moles/liters) will be important. The molar mass is important in some cases. The chemical formulas are necessary.

Solution:

$$\begin{aligned} \text{a) Grams Ca(C}_2\text{H}_3\text{O}_2)_2 &= (175.8 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.207 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2}{1 \text{ L}} \right) \left(\frac{158.17 \text{ g Ca(C}_2\text{H}_3\text{O}_2)_2}{1 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2} \right) \\ &= 5.7559 = 5.76 \text{ g Ca(C}_2\text{H}_3\text{O}_2)_2 \end{aligned}$$

$$\text{b) Molarity KI} = \left(\frac{21.1 \text{ g KI}}{500. \text{ mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1 \text{ mol KI}}{166.0 \text{ g KI}} \right) = 0.254217 = \mathbf{0.254 \text{ M KI}}$$

$$\text{c) Moles NaCN} = \left(\frac{0.850 \text{ mol NaCN}}{1 \text{ L}} \right) (145.6 \text{ L}) = 123.76 = \mathbf{124 \text{ mol NaCN}}$$

- 3.93 Molarity is a convenient way of expressing concentration; and, it is useful to use the definition (mole/L) in place of *M*.

$$\begin{aligned} \text{a) Liters KOH solution} &= \left(\frac{1 \text{ L}}{2.26 \text{ mol KOH}} \right) \left(\frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} \right) 8.42 \text{ g KOH} \\ &= 0.066399 = \mathbf{0.0664 \text{ L KOH solution}} \end{aligned}$$

$$\text{b) Number Cu}^{2+} \text{ ions} = \left(\frac{2.3 \text{ mol CuCl}_2}{\text{L}} \right) \left(\frac{1 \text{ mol Cu}^{2+}}{1 \text{ mol CuCl}_2} \right) (52 \text{ L}) \left(\frac{6.022 \times 10^{23} \text{ Cu}^{2+} \text{ ions}}{1 \text{ mol Cu}^{2+}} \right)$$

$$= 7.2023 \times 10^{25} = \mathbf{7.2 \times 10^{25} \text{ Cu}^{2+} \text{ ions}}$$

$$\text{c) } M \text{ glucose} = \left(\frac{135 \text{ mmol glucose}}{275 \text{ mL}} \right) = 0.490909 = \mathbf{0.491 M \text{ glucose}}$$

3.94 Plan: It will help to rewrite M as its definition (mole/L). (a) You will need to convert milliliters to liters, and determine the molar mass of potassium sulfate. (b) The simplest way will be to convert the milligrams to millimoles. Molarity may not only be expressed as moles/L, but also as mmoles/mL. (c) Convert the milliliters to liters, and find the moles of solute. It will be necessary to use Avogadro's number to determine the number of ions present.

Solution:

$$\text{a) Grams potassium sulfate} = \left(\frac{5.62 \times 10^{-2} \text{ mol K}_2\text{SO}_4}{\text{L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) (475 \text{ mL}) \left(\frac{174.27 \text{ g K}_2\text{SO}_4}{1 \text{ mol K}_2\text{SO}_4} \right)$$

$$= 4.6521 = \mathbf{4.65 \text{ g K}_2\text{SO}_4}$$

$$\text{b) Molarity calcium chloride} = \left(\frac{6.55 \text{ mg CaCl}_2}{1 \text{ mL}} \right) \left(\frac{1 \text{ mmol CaCl}_2}{110.98 \text{ mg CaCl}_2} \right) = 0.05902 = \mathbf{0.0590 M \text{ CaCl}_2}$$

If you believe that molarity must be moles/liters then the calculation becomes:

$$\left(\frac{6.55 \text{ mg CaCl}_2}{1 \text{ mL}} \right) \left(\frac{10^{-3} \text{ g}}{1 \text{ mg}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2} \right) = 0.05902 = \mathbf{0.0590 M \text{ CaCl}_2}$$

Notice that the two central terms cancel each other.

$$\text{c) Number of Mg}^{2+} \text{ ions} = \left(\frac{0.184 \text{ mol MgBr}_2}{1 \text{ L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) (1 \text{ mL}) \left(\frac{1 \text{ mol Mg}^{2+}}{1 \text{ mol MgBr}_2} \right) \left(\frac{6.022 \times 10^{23} \text{ Mg}^{2+} \text{ ions}}{1 \text{ mol Mg}^{2+}} \right)$$

$$= 1.1080 \times 10^{20} = \mathbf{1.11 \times 10^{20} \text{ Mg}^{2+} \text{ ions}}$$

$$3.95 \text{ a) Molarity of silver nitrate} = \left(\frac{46.0 \text{ g AgNO}_3}{335 \text{ mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1 \text{ mol AgNO}_3}{169.9 \text{ g AgNO}_3} \right) = 0.80820 = \mathbf{0.808 M \text{ AgNO}_3}$$

$$\text{b) Liters manganese(II) sulfate solution} = (57.0 \text{ g MnSO}_4) \left(\frac{1 \text{ mol MnSO}_4}{151.01 \text{ g MnSO}_4} \right) \left(\frac{1 \text{ L}}{0.385 \text{ mol MnSO}_4} \right)$$

$$= 0.98041 = \mathbf{0.980 L \text{ MnSO}_4 \text{ solution}}$$

$$\text{c) Milliliters ATP solution} = \left(\frac{1 \text{ L}}{6.44 \times 10^{-2} \text{ mol ATP}} \right) \left(\frac{10^{-3} \text{ mol ATP}}{1 \text{ mmol ATP}} \right) (1.68 \text{ mmol ATP}) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)$$

$$= 26.0867 = \mathbf{26.1 \text{ mL ATP solution}}$$

This problem can be simplified by using $M = \text{mmol/mL}$ instead of $M = \text{mol/L}$:

$$\left(\frac{1 \text{ mL}}{6.44 \times 10^{-2} \text{ mmol ATP}} \right) (1.68 \text{ mmol ATP}) = 26.0867 = \mathbf{26.1 \text{ mL ATP solution}}$$

3.96 Plan: These are dilution problems. Dilution problems can be solved by converting to moles and using the new volume, however, it is much easier to use $M_1V_1 = M_2V_2$. Part (c) may be done as two dilution problems or as a mole problem. The dilution equation does not require a volume in liters; it only requires that the volume units match.

Solution:

$$\text{a) } M_1 = 0.250 \text{ M KCl} \quad V_1 = 37.00 \text{ mL} \quad M_2 = ? \quad V_2 = 150.00 \text{ mL}$$

$$M_2 = M_1V_1 / V_2 = (0.250 \text{ M KCl}) (37.00 \text{ mL}) / (150.00 \text{ mL}) = 0.061667 = \mathbf{0.0617 M \text{ KCl}}$$

b) $M_1 = 0.0706 \text{ M } (\text{NH}_4)_2\text{SO}_4$ $V_1 = 25.71 \text{ mL}$ $M_2 = ?$ $V_2 = 500.00 \text{ mL}$
 $M_2 = M_1 V_1 / V_2 = (0.0706 \text{ M } (\text{NH}_4)_2\text{SO}_4) (25.71 \text{ mL}) / (500.00 \text{ mL}) = 0.003630 = \mathbf{0.00363 \text{ M } (\text{NH}_4)_2\text{SO}_4}$

c) When working this as a mole problem it is necessary to find the individual number of moles of sodium ions in each separate solution. (Rounding to the proper number of significant figures will only be done for the final answer.)

$$\text{Moles Na}^+ \text{ from NaCl solution} = \left(\frac{0.288 \text{ mol NaCl}}{1 \text{ L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) (3.58 \text{ mL}) \left(\frac{1 \text{ mol Na}^+}{1 \text{ mol NaCl}} \right)$$

$$= 0.00103104 \text{ mol Na}^+$$

$$\text{Moles Na}^+ \text{ from Na}_2\text{SO}_4 \text{ solution} = \left(\frac{6.51 \times 10^{-3} \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) (500. \text{ mL}) \left(\frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{SO}_4} \right)$$

$$= 0.006510 \text{ mol Na}^+$$

$$\text{Molarity of Na}^+ = \left(\frac{(0.00103104 + 0.006510) \text{ mol Na}^+}{(3.58 + 500.) \text{ mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 0.01497486 = \mathbf{0.0150 \text{ M Na}^+ \text{ ions}}$$

3.97 These are dilution problems using the equation $M_1 V_1 = M_2 V_2$.

a) $M_1 = 2.050 \text{ M Cu}(\text{NO}_3)_2$ $V_1 = ?$ $M_2 = 0.8543 \text{ M Cu}(\text{NO}_3)_2$ $V_2 = 750.0 \text{ mL}$
 $V_1 = M_2 V_2 / M_1 = (0.8543 \text{ M}) (750.0 \text{ mL}) / (2.050 \text{ M}) = 312.5488 = \mathbf{312.5 \text{ mL}}$

b) 1.03 M CaCl_2 gives: $M \text{ Cl}^- = \left(\frac{1.03 \text{ mol CaCl}_2}{1 \text{ L}} \right) \left(\frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2} \right) = 2.06 \text{ M Cl}^- \text{ ions}$

$$M_1 = 2.06 \text{ M Cl}^- \quad V_1 = ? \quad M_2 = 2.66 \times 10^{-2} \text{ M Cl}^- \text{ ions} \quad V_2 = 350. \text{ mL}$$

$$V_1 = M_2 V_2 / M_1 = (2.66 \times 10^{-2} \text{ M}) (350. \text{ mL}) / (2.06 \text{ M}) = 4.5194 = \mathbf{4.52 \text{ mL}}$$

c) $M_1 = 0.155 \text{ M Li}_2\text{CO}_3$ $V_1 = 18.0 \text{ mL}$ $M_2 = 0.0700 \text{ M Li}_2\text{CO}_3$ $V_2 = ?$
 $V_2 = M_1 V_1 / M_2 = (0.155 \text{ M}) (18.0 \text{ mL}) / (0.0700 \text{ M}) = 39.8571 = \mathbf{39.9 \text{ mL}}$

3.98 Plan: You will need to pay particular attention to the units throughout this problem. The 70.0% by mass translates to 70.0 g solute/100 g solution. The molar mass of nitric acid is also needed.

Solution:

a) Mass HNO_3 per liter = $\left(\frac{1.41 \text{ g Solution}}{1 \text{ mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{70.0 \text{ g HNO}_3}{100 \text{ g Solution}} \right) = \mathbf{987 \text{ g HNO}_3 / \text{L}}$

b) Molarity of $\text{HNO}_3 = \left(\frac{1.41 \text{ g Solution}}{1 \text{ mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{70.0 \text{ g HNO}_3}{100 \text{ g Solution}} \right) \left(\frac{1 \text{ mol HNO}_3}{63.02 \text{ g HNO}_3} \right)$
 $= 15.6617 = \mathbf{15.7 \text{ M HNO}_3}$

3.99 a) moles per milliliter = $\left(\frac{18.3 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = \mathbf{1.83 \times 10^{-2} \text{ mol H}_2\text{SO}_4 / \text{mL}}$

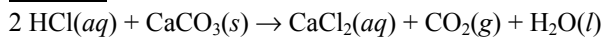
b) To calculate the mass percent, the mass of H_2SO_4 in each milliliter (1.84 g) of solution is needed.

$$\text{Mass percent} = \left(\frac{18.3 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{98.09 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} \right) (100\%) \left(\frac{1 \text{ mL}}{1.84 \text{ g Solution}} \right)$$

$$= 97.5569 = \mathbf{97.6\% \text{ H}_2\text{SO}_4 \text{ by mass}}$$

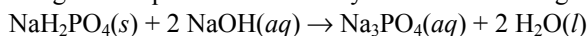
3.100 Plan: Convert the mass of calcium carbonate to moles, and use the balanced chemical equation to find the moles of hydrochloric acid required. The moles of acid along with the molarity of the acid will give the volume required. The molarity of the solution is given in the calculation as mol/L.

Solution:



$$\begin{aligned} \text{Volume required} &= (16.2 \text{ g CaCO}_3) \left(\frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \right) \left(\frac{2 \text{ mol HCl}}{1 \text{ mol CaCO}_3} \right) \left(\frac{1 \text{ L}}{0.383 \text{ mol HCl}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \\ &= 845.1923 = \mathbf{845 \text{ mL HCl solution}} \end{aligned}$$

- 3.101 Convert the molarity of sodium hydroxide to moles, and use the balanced chemical equation to find the moles of sodium dihydrogen phosphate required. The moles of sodium dihydrogen phosphate, along with its molar mass, will give the grams required. The molarity of the solution is given in the calculation as mol/L.



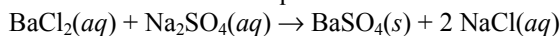
Mass of sodium dihydrogen phosphate =

$$\begin{aligned} (38.74 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.275 \text{ mol NaOH}}{1 \text{ L}} \right) \left(\frac{1 \text{ mol NaH}_2\text{PO}_4}{2 \text{ mol NaOH}} \right) \left(\frac{119.98 \text{ g NaH}_2\text{PO}_4}{1 \text{ mol NaH}_2\text{PO}_4} \right) \\ = 0.639103 = \mathbf{0.639 \text{ g NaH}_2\text{PO}_4} \end{aligned}$$

- 3.102 Plan: The first step is to write and balance the chemical equation for the reaction. Use the molarity and volume of each of the reactants to determine the moles of each as a prelude to determining which is the limiting reactant. Use the limiting reactant to determine the mass of barium sulfate that will form.

Solution:

The balanced chemical equation is:



The mole and limiting reactant calculations are:

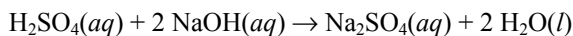
$$\text{Moles BaSO}_4 \text{ from BaCl}_2 = (25.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.160 \text{ mol BaCl}_2}{1 \text{ L}} \right) \left(\frac{1 \text{ mol BaSO}_4}{1 \text{ mol BaCl}_2} \right) = 0.00400 \text{ mol BaSO}_4$$

$$\text{Moles BaSO}_4 \text{ from Na}_2\text{SO}_4 = (68.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.055 \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{1 \text{ mol BaSO}_4}{1 \text{ mol Na}_2\text{SO}_4} \right) = 0.00374 \text{ mol BaSO}_4$$

Sodium sulfate is the limiting reactant.

$$\begin{aligned} \text{Grams BaSO}_4 &= (68.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.055 \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{1 \text{ mol BaSO}_4}{1 \text{ mol Na}_2\text{SO}_4} \right) \left(\frac{233.4 \text{ g BaSO}_4}{1 \text{ mol BaSO}_4} \right) \\ &= 0.872916 = \mathbf{0.87 \text{ g BaSO}_4} \end{aligned}$$

- 3.103 The first step is to write and balance the chemical equation for the reaction. Use the molarity and volume of each of the reactants to determine the moles of each as a prelude to determining which is the limiting reactant. Use the limiting reactant to determine the mass of the other substance consumed.



We can use either product to determine the limiting reactant. We will use sodium sulfate.

$$\begin{aligned} \text{Moles Na}_2\text{SO}_4 \text{ from H}_2\text{SO}_4 &= (350.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.210 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{1 \text{ mol Na}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} \right) \\ &= 0.0735 \text{ mol Na}_2\text{SO}_4 \end{aligned}$$

$$\text{Moles Na}_2\text{SO}_4 \text{ from NaOH} = (0.500 \text{ L}) \left(\frac{0.196 \text{ mol NaOH}}{1 \text{ L}} \right) \left(\frac{1 \text{ mol Na}_2\text{SO}_4}{2 \text{ mol NaOH}} \right) = 0.0490 \text{ mol Na}_2\text{SO}_4$$

NaOH is the limiting reactant. Finish the problem using the limiting reactant.

Moles H₂SO₄ remaining = Initial moles - Moles reacting with NaOH

$$\begin{aligned} &= (350.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.210 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \right) - (0.500 \text{ L}) \left(\frac{0.196 \text{ mol NaOH}}{1 \text{ L}} \right) \left(\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \right) \\ &= \mathbf{0.0245 \text{ mol H}_2\text{SO}_4} \end{aligned}$$

$$3.104 \quad \text{Molarity of sodium hypochlorite} = \left(\frac{20.5 \text{ g NaClO}}{375 \text{ mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1 \text{ mol NaClO}}{74.44 \text{ g NaClO}} \right)$$

$$= 0.73437 = \mathbf{0.734 \text{ M NaClO}}$$

3.105 Plan: The first part of the problem is a simple dilution problem ($M_1V_1 = M_2V_2$). The second part requires the molar mass of the HCl along with the molarity.

Solution:

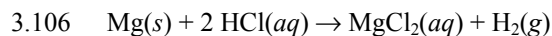
$$\text{a) } M_1 = 11.7 \text{ M} \quad V_1 = ? \quad M_2 = 3.5 \text{ M} \quad V_2 = 5.0 \text{ gal}$$

$$V_1 = M_2V_2 / M_1 = (3.5 \text{ M})(5.0 \text{ gal}) / (11.7 \text{ M}) = 1.4957 \text{ gallons (unrounded)}$$

Instructions: Be sure to wear goggles to protect your eyes! Pour approximately 3.0 gallons of water into the container. Add slowly and with mixing 1.5 gallon of concentrated HCl into the water. Dilute to 5.0 gallons with water.

$$\text{b) Volume needed} = (9.55 \text{ g HCl}) \left(\frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \right) \left(\frac{1 \text{ L}}{11.7 \text{ mol HCl}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)$$

$$= 22.38725 = \mathbf{22.4 \text{ mL muriatic acid solution}}$$



$$\text{Moles of HCl reacting with Mg} = \left(\frac{0.750 \text{ mol HCl}}{1 \text{ L}} \right) (0.100 \text{ L}) - 0.0125 \text{ mol HCl} = 0.0625 \text{ mol HCl}$$

$$\text{Grams of Mg reacting} = (0.0625 \text{ mol HCl}) \left(\frac{1 \text{ mol Mg}}{2 \text{ mol HCl}} \right) \left(\frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} \right) = 0.7596875 \text{ g Mg}$$

$$\text{Mass percent Mg} = \left(\frac{0.7596875 \text{ g Mg}}{1.32 \text{ g Sample}} \right) (100\%) = 57.552 = \mathbf{57.6\% \text{ Mg}}$$

3.107 a) An atomic mass unit (amu) is exactly 1/12 the mass of a ^{12}C atom.

$$\text{Mass in grams} = (1 \text{ amu}) \left(\frac{1 \text{ atom } ^{12}\text{C}}{12 \text{ amu}} \right) \left(\frac{1 \text{ mol } ^{12}\text{C}}{6.022 \times 10^{23} \text{ atoms } ^{12}\text{C}} \right) \left(\frac{12 \text{ g } ^{12}\text{C}}{1 \text{ mol } ^{12}\text{C}} \right)$$

$$= 1.6605778 \times 10^{-24} = \mathbf{1.661 \times 10^{-24} \text{ g}}$$

Remember that all the values in this calculation except Avogadro's number are exact numbers

b) Either use $\mathbf{1.661 \times 10^{-24} \text{ g / amu}}$ or $\mathbf{1 \text{ g} / 6.022 \times 10^{23} \text{ amu}}$.

3.108 First, determine the empirical formula:

$$(2.288 \text{ g S}) \left(\frac{1 \text{ mol S}}{32.07 \text{ g S}} \right) = 0.0713439 \text{ mol S}$$

$$(1.000 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 0.0713776 \text{ mol N}$$

Dividing both moles by the smaller value shows the empirical formula to be SN (empirical formula mass = 46.08 g/mol).

The ratio of the given molar mass to the empirical formula mass is $(184.27 \text{ g/mol}) / (46.08 \text{ g/mol}) = 4$

The molecular formula is four times the empirical formula, giving $(\text{SN})_4$ or $\mathbf{S_4N_4}$.

3.109 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 mass of water present and convert the mass to moles of the hydrate. The mass of water will be converted to moles. Finally, the ratio of the moles of hydrate to moles of water will give the amount of water present.

Solution:

$$\text{Moles narceine hydrate} = (100 \text{ g narceine hydrate}) \left(\frac{1 \text{ mol narceine hydrate}}{499.52 \text{ g narceine hydrate}} \right)$$

$$= 0.20019 \text{ mol narceine hydrate}$$

$$\text{Moles H}_2\text{O} = (100 \text{ g narceine hydrate}) \left(\frac{10.8\% \text{ H}_2\text{O}}{100\% \text{ narceine hydrate}} \right) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right)$$

$$= 0.59933 \text{ mol H}_2\text{O}$$

The ratio of water to hydrate is: $(0.59933 \text{ mol}) / (0.20019 \text{ mol}) = 3$

Thus, there are three water molecules per mole of hydrate. The formula for narceine hydrate is **narceine•3H₂O**

- 3.110 Plan: Determine the formula, then the molar mass of each compound. Determine the mass of hydrogen in each formula. The mass of hydrogen divided by the molar mass of the compound (with the result multiplied by 100%) will give the mass percent hydrogen. Ranking, based on the percents, is easy.

Solution:

Name	Chemical formula	Molar mass (g/mol)	Mass percent H [(mass H) / (molar mass)] x 100%
Ethane	C ₂ H ₆	30.07	[(6 x 1.008) / (30.07)] x 100% = 20.11% H
Propane	C ₃ H ₈	44.09	[(8 x 1.008) / (44.09)] x 100% = 18.29% H
Cetyl palmitate	C ₃₂ H ₆₄ O ₂	480.83	[(64 x 1.008) / (480.83)] x 100% = 13.42% H
Ethanol	C ₂ H ₅ OH	46.07	[(6 x 1.008) / (46.07)] x 100% = 13.13% H
Benzene	C ₆ H ₆	78.11	[(6 x 1.008) / (78.11)] x 100% = 7.743% H

The hydrogen percentage decreases in the following order:

Ethane > Propane > Cetyl palmitate > Ethanol > Benzene

- 3.111 First, determine the empirical formula and empirical formula mass. The empirical formula mass and the molar mass give the molecular formula for the compound.
Assume 100 grams of sample, thus the percentages are the mass, in grams, of each element.

$$\text{Mole C} = (68.2 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 5.6786 \text{ mol C}$$

$$\text{Mole H} = (6.86 \text{ g H}) \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 6.8056 \text{ mol H}$$

$$\text{Mole N} = (15.9 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 1.1349 \text{ mol N}$$

$$\text{Mole O} = (9.08 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 0.5675 \text{ mol O}$$

Divide each mole value by the smallest value (mole O).

$$\text{C: } 5.6786 \text{ mol} / 0.5675 \text{ mol} = 10.0$$

$$\text{H: } 6.8056 \text{ mol} / 0.5675 \text{ mol} = 12.0$$

$$\text{N: } 1.1349 \text{ mol} / 0.5675 \text{ mol} = 2.00$$

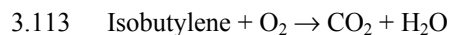
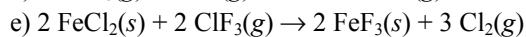
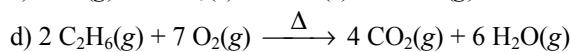
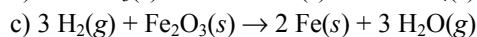
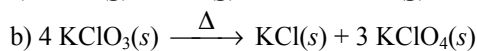
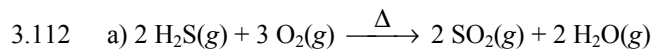
$$\text{O: } 0.5675 \text{ mol} / 0.5675 \text{ mol} = 1.00$$

This gives the empirical formula: C₁₀H₁₂N₂O with a formula mass =

$$10(12.01 \text{ g C/mol}) + 12(1.008 \text{ g H/mol}) + 2(14.01 \text{ g N/mol}) + 1(16.00 \text{ g O/mol}) = 176.22 \text{ g/mol}$$

The empirical formula mass and the molar mass are the same, thus, the molecular and empirical formulas are the same.

Molecular formula: **C₁₀H₁₂N₂O**



$$\text{Mole C} = (2.657 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) = 0.06037 \text{ mol C}$$

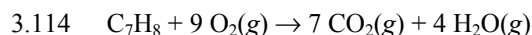
$$\text{Mole H} = (1.089 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.1209 \text{ mol H}$$

Divide each mole value by the smaller value (mole C).

$$\text{C: } 0.06037 \text{ mol} / 0.06037 \text{ mol} = 1.00$$

$$\text{H: } 0.1209 \text{ mol} / 0.06037 \text{ mol} = 2.00$$

This gives an empirical formula of **CH₂**.



$$\text{Moles C}_7\text{H}_8 = (15.0 \text{ mL}) \left(\frac{0.867 \text{ g C}_7\text{H}_8}{1 \text{ mL}} \right) \left(\frac{1 \text{ mol C}_7\text{H}_8}{92.13 \text{ g C}_7\text{H}_8} \right) = 0.14116 \text{ mol C}_7\text{H}_8 \text{ (unrounded)}$$

$$\text{a) Grams oxygen} = (0.14116 \text{ mol C}_7\text{H}_8) \left(\frac{9 \text{ mol O}_2}{1 \text{ mol C}_7\text{H}_8} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 40.654 = \mathbf{40.7 \text{ g O}_2}$$

$$\text{b) Total moles of gas} = (0.14116 \text{ mol C}_7\text{H}_8) \left(\frac{11 \text{ mol gas}}{1 \text{ mol C}_7\text{H}_8} \right) = 1.55276 = \mathbf{1.55 \text{ mol of gas}}$$

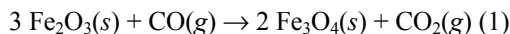
The 11 mol of gas is an exact, not measured, number, so it does not affect the significant figures.

$$\begin{aligned} \text{c) Molecules of water} &= (0.14116 \text{ mol C}_7\text{H}_8) \left(\frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_7\text{H}_8} \right) \left(\frac{6.022 \times 10^{23} \text{ H}_2\text{O molecules}}{1 \text{ mol H}_2\text{O}} \right) \\ &= 3.40026 \times 10^{23} = \mathbf{3.40 \times 10^{23} \text{ molecules H}_2\text{O}} \end{aligned}$$

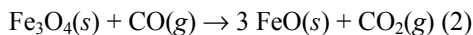
3.115 Plan: The key to solving this problem is determining the overall balanced equation. Each individual step must be set up and balanced first. The separate equations can then be combined to get the overall equation. The mass of iron may be converted to moles of iron, and the mole ratio and molar mass of carbon monoxide used to determine the mass of carbon monoxide.

Solution:

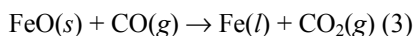
a) In the first step, ferric oxide (ferric denotes Fe^{3+}) reacts with carbon monoxide to form Fe_3O_4 and carbon dioxide:



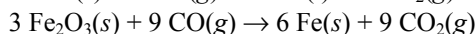
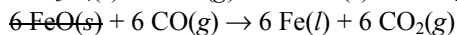
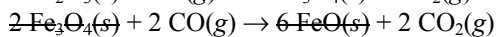
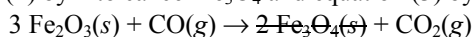
In the second step, Fe_3O_4 reacts with more carbon monoxide to form ferrous oxide:



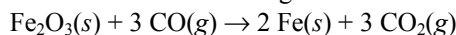
In the third step, ferrous oxide reacts with more carbon monoxide to form molten iron:



Common factors are needed to allow these equations to be combined. The intermediate products are Fe_3O_4 and FeO , so multiply equation (2) by 2 to cancel Fe_3O_4 and equation (3) by 6 to cancel FeO :



Then divide by 3 to obtain the smallest integer coefficients:



b) A metric ton is a tonne, and is equal to 1000 kg.

$$\begin{aligned} \text{Grams carbon monoxide} &= (40.0 \text{ tonnes Fe}) \left(\frac{10^3 \text{ kg}}{1 \text{ tonne}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) \left(\frac{3 \text{ mol CO}}{2 \text{ mol Fe}} \right) \left(\frac{28.01 \text{ g CO}}{1 \text{ mol CO}} \right) \\ &= 3.00913 \times 10^7 = \mathbf{3.01 \times 10^7 \text{ g CO}} \end{aligned}$$

3.116 Plan: If 100.0 g of dinitrogen tetroxide reacts with 100.0 g of hydrazine (N_2H_4), what is the theoretical yield of nitrogen if no side reaction takes place? First, we need to identify the limiting reactant. The limiting reactant can be used to calculate the theoretical yield. 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:

Determining the limiting reactant:

$$\text{N}_2 \text{ from } \text{N}_2\text{O}_4 = (100.0 \text{ g } \text{N}_2\text{O}_4) \left(\frac{1 \text{ mol } \text{N}_2\text{O}_4}{92.02 \text{ g } \text{N}_2\text{O}_4} \right) \left(\frac{3 \text{ mol } \text{N}_2}{1 \text{ mol } \text{N}_2\text{O}_4} \right) = 3.26016 \text{ mol } \text{N}_2$$

N_2O_4 is the limiting reactant.

$$\begin{aligned} \text{Theoretical yield of } \text{N}_2 &= (100.0 \text{ g } \text{N}_2\text{O}_4) \left(\frac{1 \text{ mol } \text{N}_2\text{O}_4}{92.02 \text{ g } \text{N}_2\text{O}_4} \right) \left(\frac{3 \text{ mol } \text{N}_2}{1 \text{ mol } \text{N}_2\text{O}_4} \right) \left(\frac{28.02 \text{ g } \text{N}_2}{1 \text{ mol } \text{N}_2} \right) \\ &= 91.3497 \text{ g } \text{N}_2 \text{ (unrounded)} \end{aligned}$$

How much limiting reactant used to produce 10.0 g NO?

$$\text{Grams } \text{N}_2\text{O}_4 \text{ used} = (10.0 \text{ g NO}) \left(\frac{1 \text{ mol NO}}{30.01 \text{ g NO}} \right) \left(\frac{2 \text{ mol } \text{N}_2\text{O}_4}{6 \text{ mol NO}} \right) \left(\frac{92.02 \text{ g } \text{N}_2\text{O}_4}{1 \text{ mol } \text{N}_2\text{O}_4} \right) = 10.221 \text{ g } \text{N}_2\text{O}_4 \text{ (unrounded)}$$

Determine the “actual yield.”

$$\begin{aligned} \text{“Actual yield” of } \text{N}_2 &= ((100.0 - 10.221) \text{ g } \text{N}_2\text{O}_4) \left(\frac{1 \text{ mol } \text{N}_2\text{O}_4}{92.02 \text{ g } \text{N}_2\text{O}_4} \right) \left(\frac{3 \text{ mol } \text{N}_2}{1 \text{ mol } \text{N}_2\text{O}_4} \right) \left(\frac{28.02 \text{ g } \text{N}_2}{1 \text{ mol } \text{N}_2} \right) \\ &= 82.01285 \text{ g } \text{N}_2 \text{ (unrounded)} \end{aligned}$$

Theoretical yield = [“Actual yield” / theoretical yield] x 100%

$$[(82.01285 \text{ g } \text{N}_2) / (91.3497 \text{ g } \text{N}_2)] \times 100\% = 89.7790 = \mathbf{89.8\%}$$

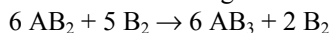
3.117 M = molarity V = volume usually in liters
dil = diluted solution conc = concentrated solution

The equation works because the quantity (moles) of solute remains the same.

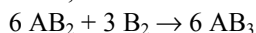
3.118 Plan: Count the number of each type of molecule in the reactant box and in the product box. Subtract any molecules of excess reagent (molecules appearing in both boxes). The remaining material is the overall equation. This will need to be simplified if there is a common factor among the substances in the equation. The balanced chemical equation is necessary for the remainder of the problem.

Solution:

a) The contents of the boxes give:



B_2 is in excess, so two molecules need to be removed from each side. This gives:



Three is a common factor among the coefficients, and all coefficients need to be divided by this value to give the final balanced equation:



b) B_2 was in excess, thus AB_2 is the limiting reactant.

c) From 3.0 mol of B_2 , it is possible to make 6.0 mol of AB_3 ($3.0 \times 2/1$). From 5.0 mol of AB_2 it is possible to make 5.0 mol of AB_3 ($5.0 \times 2/2$). Thus, AB_2 is limiting and the maximum is **5.0 mol of AB_3** .

d) A total of 5.0 mol AB₃ would react with 5.0 x 1/2 = 2.5 mol of B₂.
The un-reacted B₂ is 3.0 mol - 2.5 mol = **0.5 mol B₂**.

3.119 These are all dilution problems and the equation $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$. In parts c and d further calculations are necessary.

a) $M_{\text{conc}} = 18.0 M$ $V_{\text{conc}} = ?$ $M_{\text{dil}} = 0.309 M$ $V_{\text{dil}} = 2.00 L$
 $(18.0 M)(V_{\text{conc}}) = (0.309 M)(2.00 L)$

$$V_{\text{conc}} = \frac{(0.309 M)(2.00 L)}{(18.0 M)} = 0.034333 = \mathbf{0.0343 L}$$

b) $M_{\text{conc}} = 0.225 M$ $V_{\text{conc}} = 80.6 \text{ mL}$ $M_{\text{dil}} = ?$ $V_{\text{dil}} = 0.250 L$
 $(0.225 M)(80.6 \text{ mL}) = (M_{\text{dil}})(0.250 L)$

$$M_{\text{dil}} = \frac{(0.225 M)(80.6 \text{ mL})\left(\frac{10^{-3} L}{1 \text{ mL}}\right)}{(0.250 L)} = 0.072540 = \mathbf{0.0725 M}$$

c) $M_{\text{conc}} = 0.0262 M$ $V_{\text{conc}} = 0.150 L$ $M_{\text{dil}} = 0.0100 M$ $V_{\text{dil}} = ?$
 $(0.0262 M)(0.150 L) = (0.0100 M)(V_{\text{dil}})$

$$V_{\text{dil}} = \frac{(0.0262 M)(0.150 L)}{(0.0100 M)} = 0.393 L$$

Final volume - initial volume = volume added.

$$0.393 L - 0.150 L = \mathbf{0.243 L} \text{ of water}$$

d) $M_{\text{conc}} = 0.745 M$ $V_{\text{conc}} = 64.0 \text{ mL}$ $M_{\text{dil}} = ?$ $V_{\text{dil}} = 0.100 L$
 $(0.745 M)(64.0 \text{ mL}) = (M_{\text{dil}})(0.100 L)$

$$M_{\text{dil}} = \frac{(0.745 M)(64.0 \text{ mL})\left(\frac{10^{-3} L}{1 \text{ mL}}\right)}{(0.100 L)} = 0.4768 M \text{ (unrounded)}$$

The concentration must now be converted to grams per milliliter.

$$\left(\frac{0.4768 \text{ mol Ca(NO}_3)_2}{L}\right)\left(\frac{10^{-3} L}{1 \text{ mL}}\right)\left(\frac{164.10 \text{ g Ca(NO}_3)_2}{1 \text{ mol Ca(NO}_3)_2}\right) = 0.07824 = \mathbf{0.0782 g Ca(NO}_3)_2/\text{mL}$$

3.120 Plan: The molar mass of the compound must be calculated. The molar mass and the numbers in the problem (with a few unit conversions) are then placed in the definition of molarity.

Solution:

$$\text{Molar Mass } C_{12}H_{11}N_5 = 12 (12.01 \text{ g C/mol}) + 11 (1.008 \text{ g H/mol}) + 5 (14.01 \text{ g N/mol}) = 225.26 \text{ g/mol}$$

$$M = \left(\frac{0.030 \text{ mg } C_{12}H_{11}N_5}{150. \text{ mL}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mL}}{10^{-3} L}\right)\left(\frac{1 \text{ mol } C_{12}H_{11}N_5}{225.26 \text{ g } C_{12}H_{11}N_5}\right)$$

$$= 8.878629 \times 10^{-7} = \mathbf{8.9 \times 10^{-7} M C_{12}H_{11}N_5}$$

Note: By setting up the central conversions as shown, it is obvious that they are irrelevant.

3.121 The formula and molar mass are needed in addition to the information in the problem. The formula is: C₄H₆O₂, and the molar mass is 86.09 g/mol.

$$\text{Moles/month} = \left(\frac{3.74 \times 10^8 \text{ lb } C_4H_6O_2}{1 \text{ yr}}\right)\left(\frac{1 \text{ yr}}{12 \text{ months}}\right)\left(\frac{1 \text{ kg}}{2.205 \text{ lb}}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } C_4H_6O_2}{86.09 \text{ g } C_4H_6O_2}\right)$$

$$= 1.64183 \times 10^8 = \mathbf{1.64 \times 10^8 \text{ mol } C_4H_6O_2 / \text{ month}}$$

3.122 a) $(4.0\% \text{ ions})\left(\frac{85\% \text{ NaCl}}{100\% \text{ ions}}\right) = \mathbf{3.4\% \text{ NaCl}}$

$$\text{b) \% Na}^+ \text{ ions} = (3.4\% \text{ NaCl}) \left(\frac{22.99 \text{ g Na}^+}{58.44 \text{ g NaCl}} \right) = 1.3375 = \mathbf{1.3\% \text{ Na}^+ \text{ ions}}$$

$$\% \text{ Cl}^- \text{ ions} = (3.4\% \text{ NaCl}) \left(\frac{35.45 \text{ g Cl}^-}{58.44 \text{ g NaCl}} \right) = 2.062 = \mathbf{2.1\% \text{ Cl}^- \text{ ions}}$$

$$\text{c) } M \text{ NaCl} = \left(\frac{3.4 \text{ g NaCl}}{100 \text{ g Seawater}} \right) \left(\frac{1.025 \text{ g Seawater}}{1 \text{ mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \right) = 0.596338 = \mathbf{0.60 \text{ M NaCl}}$$

- 3.123 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 presented when the quantity of available material is given for more than one reactant.
 d) **False**, to prepare 1.00 L of 3.00 M NaCl, weigh 175.5 g of NaCl and add enough water to make 1.00 L
 e) **True**
- 3.124 **Plan:** Count the total number of spheres in each box. The number in box A divided by the volume change in each part will give the number we are looking for and allow us to match boxes.
Solution:
 The number in each box is: A = 12, B = 6, C = 4, and D = 3.
 a) When the volume is tripled, there should be $12/3 = 4$ spheres in a box. This is box **C**.
 b) When the volume is doubled, there should be $12/2 = 6$ spheres in a box. This is box **B**.
 c) When the volume is quadrupled, there should be $12/4 = 3$ spheres in a box. This is box **D**.
- 3.125 a) **Equal** Equal numbers of particles
 b) **O₃** 0.4 mol O₃ has a greater mass
 c) **SO₂** (4.0 g N₂O₄) (1 mol/92.02 g N₂O₄) = 0.043 mol N₂O₄
 (3.3 g SO₂) (1 mol/64.07 g SO₂) = 0.052 mol SO₂
 d) **F₂** (0.6 mol C₂H₄) (28.05 g/1 mol C₂H₄) = 17g
 (0.6 mol F₂) (38.00 g/mol F₂) = 23 g
- Note that if each of these values is properly rounded to one significant figure, the answers are identical.
- e) **MgCl₂** (2.2 mol MgCl₂) (3 ions/mol MgCl₂) = 6.6 mol
 (2.3 mol NaClO₃) (2 ions/mol NaClO₃) = 4.6
- f) **H₂O** The compound with the lower molar mass will have more molecules in a given mass.
- g) **Equal** (0.500 mol NaBr/L) (0.500L) (1 mol Na⁺/mol NaBr) = 0.250 mol Na⁺
 (0.0146 kg NaCl) (10³ g/kg) (1 mol NaCl/58.44 g) (1 mol Na⁺/mol NaCl) = 0.250 mol Na⁺
 Multiply the moles by Avogadro's number to find the number of ions. This step is unnecessary since the moles are equal.
- h) **²³⁸U** The heavier atoms will give a greater total mass.
- 3.126 $\text{P}_4\text{S}_3(\text{s}) + 8 \text{O}_2(\text{g}) \rightarrow \text{P}_4\text{O}_{10}(\text{s}) + 3 \text{SO}_2(\text{g})$
 a) 1 molecule of P₄S₃ reacts with 8 molecules of O₂ to produce 1 molecule of P₄O₁₀ and 3 molecules of SO₂.
 b) 1 mol P₄S₃ reacts with 8 mol O₂ to produce 1 mol P₄O₁₀ and 3 mol SO₃.
 c) 220.09 g of P₄S₃ react with 256.00 g of O₂ to produce 383.33 g of P₄O₁₀ and 192.21 g of SO₃.
- 3.127 $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g})$
 Mass H₂ = (85.0 kg H₂O) $\left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \right) \left(\frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) \left(\frac{100\%}{98.8\%} \right)$
 = 9.62493 x 10³ = 9.62 x 10³ g H₂

- 3.128 Plan: This problem may be done as two dilution problems with the two final molarities added, or, as done here, it may be done by calculating, then adding the moles and dividing by the total volume.

Solution:

$$M \text{ KBr} = \frac{\text{Total Moles KBr}}{\text{Total Volume}} = \frac{\text{Moles KBr from Solution 1} + \text{Moles KBr from Solution 2}}{\text{Volume Solution 1} + \text{Volume Solution 2}}$$

$$M \text{ KBr} = \frac{\left(\frac{0.053 \text{ mol KBr}}{1 \text{ L}}\right)(0.200 \text{ L}) + \left(\frac{0.078 \text{ mol KBr}}{1 \text{ L}}\right)(0.550 \text{ L})}{0.200 \text{ L} + 0.550 \text{ L}} = 0.071333 = \mathbf{0.071 \text{ M KBr}}$$

- 3.129 a) Mass of atmosphere = $\left(\frac{3.0 \text{ kg}}{\text{s}}\right)\left(\frac{60 \text{ s}}{1 \text{ min}}\right)\left(\frac{60 \text{ min}}{1 \text{ hr}}\right)\left(\frac{24 \text{ hr}}{1 \text{ d}}\right)\left(\frac{365.25 \text{ d}}{1 \text{ y}}\right)(50 \times 10^9 \text{ y})$
 $= 4.73364 \times 10^{18} = \mathbf{4.7 \times 10^{18} \text{ kg atmosphere}}$
 b) Begin with the unrounded total mass of the atmosphere.

$$\text{Moles N}_2 = (4.73364 \times 10^{18} \text{ kg Air})\left(\frac{75.5\%}{100\%}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2}\right)$$

$$= 1.27548 \times 10^{20} = \mathbf{1.3 \times 10^{20} \text{ mol N}_2}$$

- 3.130 a) Moles $\text{NH}_4\text{Br} = (0.588 \text{ g NH}_4\text{Br})\left(\frac{1 \text{ mol NH}_4\text{Br}}{97.94 \text{ g NH}_4\text{Br}}\right) = 0.0060037 = \mathbf{0.00600 \text{ mol NH}_4\text{Br}}$

b) Number of K^+ ions = $(68.5 \text{ g KNO}_3)\left(\frac{1 \text{ mol KNO}_3}{101.11 \text{ g KNO}_3}\right)\left(\frac{1 \text{ mol K}^+}{1 \text{ mol KNO}_3}\right)\left(\frac{6.022 \times 10^{23} \text{ K}^+ \text{ ions}}{1 \text{ mol K}^+}\right)$
 $= 4.07978 \times 10^{23} = \mathbf{4.08 \times 10^{23} \text{ K}^+ \text{ ions}}$

c) Mass $\text{C}_3\text{H}_8\text{O}_3 = (5.85 \text{ mol C}_3\text{H}_8\text{O}_3)\left(\frac{92.09 \text{ g C}_3\text{H}_8\text{O}_3}{1 \text{ mol C}_3\text{H}_8\text{O}_3}\right) = 538.7265 = \mathbf{539 \text{ g C}_3\text{H}_8\text{O}_3}$

d) Volume $\text{CHCl}_3 = (2.55 \text{ mol CHCl}_3)\left(\frac{119.37 \text{ g CHCl}_3}{1 \text{ mol CHCl}_3}\right)\left(\frac{\text{mL}}{1.48 \text{ g CHCl}_3}\right) = 205.671 = \mathbf{206 \text{ mL CHCl}_3}$

e) Number of $\text{Na}^+ = (2.11 \text{ mol Na}_2\text{CO}_3)\left(\frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{CO}_3}\right)\left(\frac{6.022 \times 10^{23} \text{ Na}^+ \text{ ions}}{1 \text{ mol Na}^+}\right)$
 $= 2.54128 \times 10^{24} = \mathbf{2.54 \times 10^{24} \text{ Na}^+ \text{ ions}}$

f) Number of Cd atoms = $(10.0 \mu\text{g Cd})\left(\frac{10^{-6} \text{ g}}{1 \mu\text{g}}\right)\left(\frac{1 \text{ mol Cd}}{112.4 \text{ g Cd}}\right)\left(\frac{6.022 \times 10^{23} \text{ Cd atoms}}{1 \text{ mol Cd}}\right)$
 $= 5.35765 \times 10^{16} = \mathbf{5.36 \times 10^{16} \text{ Cd atoms}}$

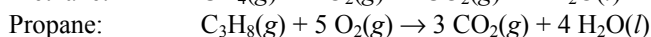
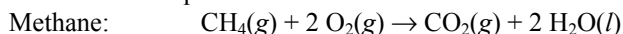
g) Number of F atoms = $(0.0015 \text{ mol F}_2)\left(\frac{2 \text{ mol F}}{1 \text{ mol F}_2}\right)\left(\frac{6.022 \times 10^{23} \text{ F atoms}}{1 \text{ mol F}}\right)$
 $= 1.8066 \times 10^{21} = \mathbf{1.8 \times 10^{21} \text{ F atoms}}$

- 3.131 Neither A nor B has any XY_3 molecules. Both C and D have XY_3 molecules. D forms both XY_3 and XY molecules. Only C has a single XY_3 product, thus the answer is C.

- 3.132 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 CO_2 produced by each gas may be found. Summing the amounts of CO_2 gives the total from the mixture.

Solution:

The balanced chemical equations are:



Mass of CO_2 from each:

$$\text{Methane: } (200. \text{g Mixture}) \left(\frac{25.0\%}{100\%} \right) \left(\frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \right) \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \right) \left(\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 137.188 \text{ g CO}_2$$

$$\text{Propane: } (200. \text{g Mixture}) \left(\frac{75.0\%}{100\%} \right) \left(\frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \right) \left(\frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \right) \left(\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 449.183 \text{ g CO}_2$$

$$\text{Total CO}_2 = 137.188 \text{ g} + 449.183 \text{ g} = 586.371 = \mathbf{586 \text{ g CO}_2}$$

$$3.133 \quad \text{Total moles of HCl in final solution} = \left(\frac{0.893 \text{ mol HCl}}{1 \text{ L}} \right) ((1.20 + 3.37) \text{ L}) = 4.08101 \text{ mol HCl (unrounded)}$$

$$\text{Moles HCl in solution of known concentration} = \left(\frac{0.325 \text{ mol HCl}}{1 \text{ L}} \right) (1.20 \text{ L}) = 0.390 \text{ mol HCl}$$

$$\text{Moles HCl added from solution with unknown concentration} = \text{Total moles} - \text{moles from known solution} \\ = (4.08101 - 0.390) \text{ mol HCl} = 3.69101 \text{ mol HCl (unrounded)}$$

$$\text{Molarity of unknown solution} = (3.69101 \text{ mol HCl}) / (3.37 \text{ L}) = 1.095255 = \mathbf{1.10 \text{ M HCl}}$$

- 3.134 Plan: If we assume a 100-gram sample of fertilizer, then the 30:10:10 percentages become the masses, in grams, of N, P_2O_5 , and K_2O . 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 P_2O_5 , and 10 g K_2O .

$$\text{Moles N} = (30 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 2.1413 \text{ mol N (unrounded)}$$

$$\text{Moles P} = (10 \text{ g P}_2\text{O}_5) \left(\frac{1 \text{ mol P}_2\text{O}_5}{141.94 \text{ g P}_2\text{O}_5} \right) \left(\frac{2 \text{ mol P}}{1 \text{ mol P}_2\text{O}_5} \right) = 0.14090 \text{ mol P (unrounded)}$$

$$\text{Moles K} = (10 \text{ g K}_2\text{O}) \left(\frac{1 \text{ mol K}_2\text{O}}{94.20 \text{ g K}_2\text{O}} \right) \left(\frac{2 \text{ mol K}}{1 \text{ mol K}_2\text{O}} \right) = 0.21231 \text{ mol K (unrounded)}$$

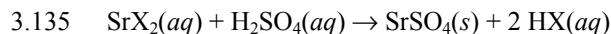
This gives a ratio of 2.1413:0.14090:0.21231

The ratio must be divided by the moles of K and rounded.

$$(2.1413/0.21231):(0.14090/0.21231):(0.21231/0.21231)$$

$$10.086:0.66365:1.000$$

$$10:0.66:1.0$$



$$0.652 \text{ g} \qquad \qquad \qquad 0.755 \text{ g}$$

Let M = the atomic mass of X.

$$\text{Moles SrX}_2 = (0.755 \text{ g SrSO}_4) \left(\frac{1 \text{ mol SrSO}_4}{183.69 \text{ g SrSO}_4} \right) \left(\frac{1 \text{ mol SrX}_2}{1 \text{ mol SrSO}_4} \right) = 0.004110185 \text{ mol SrX}_2 \text{ (unrounded)}$$

$$(0.004110185 \text{ mol SrX}_2) \left(\frac{(87.62 + 2M) \text{ g SrX}_2}{1 \text{ mol SrX}_2} \right) = 0.652 \text{ g SrX}_2$$

$$(0.004110185 (87.62 + 2M)) = 0.652$$

$$0.360134 + 0.00822037 M = 0.652$$

$$M = 35.505 = 35.5 \text{ g/mol} = \text{Cl} \quad \text{The original halide formula is } \mathbf{SrCl}_2.$$

3.136 Assuming 100 g of mixture, we have 40.0 g CH₄ + 60.0 g C₂H₆.

$$\text{C from CH}_4 = (40.0 \text{ g CH}_4) \left(\frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CH}_4} \right) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 29.950 \text{ g C (unrounded)}$$

$$\text{C from C}_2\text{H}_6 = (60.0 \text{ g C}_2\text{H}_6) \left(\frac{1 \text{ mol C}_2\text{H}_6}{30.07 \text{ g C}_2\text{H}_6} \right) \left(\frac{2 \text{ mol C}}{1 \text{ mol C}_2\text{H}_6} \right) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 47.928 \text{ g C (unrounded)}$$

$$\text{Mass percent C} = \left(\frac{(29.950 + 47.928) \text{ g}}{100 \text{ g Sample}} \right) \times 100\% = 77.878 = \mathbf{77.9\% C}$$

3.137 Plan: Assume 100 grams of mixture. This means the mass of each compound, in grams, is the same as its percentage.

Solution:

$$\text{C from CO} = (35.0 \text{ g CO}) \left(\frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}} \right) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 15.007 \text{ g C (unrounded)}$$

$$\text{C from CO}_2 = (65.0 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 17.738 \text{ g C (unrounded)}$$

$$\text{Mass percent C} = \left(\frac{(15.007 + 17.738) \text{ g}}{100 \text{ g Sample}} \right) \times 100\% = 32.745 = \mathbf{32.7\% C}$$

3.138 Ferrocene + ? O₂(g) → CO₂ + H₂O

0.9437 g 2.233 g 0.457 g

$$\text{Moles C} = (2.233 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) = 0.050738 \text{ mol C (unrounded)}$$

$$\text{Moles H} = (0.457 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.050721 \text{ mol H (unrounded)}$$

Mass of Fe = mass of sample - mass of C - mass of H
 = 0.9437 g Fe, C, H - (0.050738 mol C) (12.01 g C/mol C) - (0.050721 mol H) (1.008 g H/mol H)
 = (0.9437 - 0.60936 - 0.05113) g = 0.28321 g Fe (unrounded)

Moles Fe = (0.28321 g Fe) (1 mol Fe/ 55.85 g Fe) = 0.0050709 mol Fe

Divide all moles by the smallest (Fe):

C: (0.050738 mol / 0.0050709 mol) = 10.0
 H: (0.050721 mol / 0.0050709 mol) = 10.0
 Fe: (0.0050709 mol / 0.0050709 mol) = 1.0

Thus, the empirical formula is: **C₁₀H₁₀Fe**

3.139 Plan: Determine the molecular formula from the figure. Once the molecular formula is known, use the periodic table to determine the molar mass. With the molar mass and the additional information in part (b), the moles can be found.

Solution:

a) The formula of citric acid obtained by counting the number of carbon atoms, oxygen atoms, and hydrogen atoms is **C₆H₈O₇**.

$$\text{Molar mass} = (6 \times 12.01) + (8 \times 1.008) + (7 \times 16.00) = \mathbf{192.12 \text{ g/mol}}$$

b) Determine the mass of citric acid in the lemon juice, and then use the molar mass to find the moles.

$$\text{Moles C}_6\text{H}_8\text{O}_7 = (1.50 \text{ qt}) \left(\frac{1 \text{ L}}{1.057 \text{ qt}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1.09 \text{ g}}{\text{mL}} \right) \left(\frac{6.82\%}{100\%} \right) \left(\frac{1 \text{ mol C}_6\text{H}_8\text{O}_7}{192.12 \text{ g acid}} \right)$$

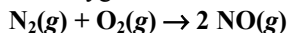
$$= 0.549104 = \mathbf{0.549 \text{ mol C}_6\text{H}_8\text{O}_7}$$

- 3.140 Plan: Determine the formulas of each reactant and product, then balance the individual equations. Combine the three smaller equations to give the overall equation, where some substances serve as intermediates and will cancel. The amount of nitrogen plus the balanced overall chemical equation will give the amount of nitric acid formed.

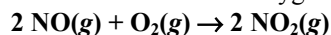
Solution:

a) Derive the formulas and balance the equations.

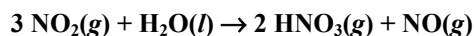
Nitrogen and oxygen combine to form nitrogen monoxide:



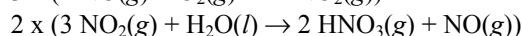
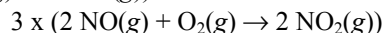
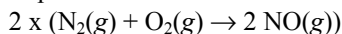
Nitrogen monoxide reacts with oxygen to form nitrogen dioxide:



Nitrogen dioxide combines with water to form nitric acid and nitrogen monoxide:



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.



Multiplying the above equations, summing the results, and canceling any substance appearing on both sides gives:



c) Metric tons HNO_3 =

$$\begin{aligned} & \left(1.25 \times 10^3 \text{ t N}_2\right) \left(\frac{10^3 \text{ kg}}{1 \text{ t}}\right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2}\right) \left(\frac{4 \text{ mol HNO}_3}{2 \text{ mol N}_2}\right) \left(\frac{63.02 \text{ g HNO}_3}{1 \text{ mol HNO}_3}\right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) \left(\frac{1 \text{ t}}{10^3 \text{ kg}}\right) \\ & = 5.622769 \times 10^3 = \mathbf{5.62 \times 10^3 \text{ metric tons HNO}_3} \end{aligned}$$

- 3.141 a) The moles of anhydrous alum and of water are needed. Dividing the moles of water by the moles of the anhydrous salt gives the value of x.

Mass of water driven off on heating = 0.5404 g - 0.2941 g = 0.2463 g H_2O

$$\text{Moles water} = (0.2463 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}}\right) = 0.013668 \text{ mol H}_2\text{O} \text{ (unrounded)}$$

$$\begin{aligned} \text{Moles anhydrous} &= (0.2941 \text{ g KAl(SO}_4)_2) \left(\frac{1 \text{ mol KAl(SO}_4)_2}{258.22 \text{ g KAl(SO}_4)_2}\right) \\ &= 0.00113895 \text{ mol anhydrous (unrounded)} \end{aligned}$$

$$x = (0.013668 \text{ mol H}_2\text{O}) / (0.00113895 \text{ mol anhydrous}) = \mathbf{12.00} = x$$

Complete formula: **$\text{KAl(SO}_4)_2 \cdot 12\text{H}_2\text{O}$**

b) There is a 1:1 mole ratio relating the Al to alum. Using the formula determined in part (a), the molar mass of alum is 474.46 g/mol. The theoretical yield (unrounded) is:

$$\text{Grams Alum} = (0.7500 \text{ g Al}) \left(\frac{1 \text{ mol Al}}{26.98 \text{ g Al}}\right) \left(\frac{1 \text{ mol Alum}}{1 \text{ mol Al}}\right) \left(\frac{474.46 \text{ g Alum}}{1 \text{ mol Alum}}\right) = 13.18921 \text{ g Alum}$$

$$\text{Percent yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\right) \times 100\% = \left(\frac{8.500 \text{ g}}{13.18921 \text{ g}}\right) \times 100\% = 64.4466 = \mathbf{64.45\% \text{ yield}}$$

$$3.142 \text{ Moles C} = (2.838 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2}\right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2}\right) = 0.0644853 \text{ mol C (unrounded)}$$

$$\text{Moles H} = (0.8122 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}}\right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}}\right) = 0.090144 \text{ mol H (unrounded)}$$

The iron calculation needs an adjustment because of the differing sample size. This adjustment involves multiplying the iron by the sample size of the other sample, and dividing by the size of the iron sample: (1.5173g / 0.3355 g).

$$\text{Moles Fe} = (0.0758 \text{ g Fe}_2\text{O}_3) \left(\frac{1 \text{ mol Fe}_2\text{O}_3}{159.70 \text{ g Fe}_2\text{O}_3} \right) \left(\frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \right) \left(\frac{1.5173 \text{ g}}{0.3355 \text{ g}} \right) = 0.004293 \text{ mol Fe (unrounded)}$$

To determine the amount of oxygen the above moles need to be converted to grams of each element and subtracted from the original sample mass (1.5173 g).

$$\text{Mass C} = (0.0644853 \text{ mol C}) (12.01 \text{ g C/mol}) = 0.774468 \text{ g C}$$

$$\text{Mass H} = (0.090144 \text{ mol H}) (1.008 \text{ g H/mol}) = 0.090865 \text{ g H}$$

$$\text{Mass Fe} = (0.004293 \text{ mol Fe}) (55.85 \text{ g Fe/mol}) = 0.239764 \text{ g Fe}$$

$$\text{Mass of O} = (1.5173 \text{ g Fe, C, H, O}) - (0.774468 \text{ g C}) - (0.090865 \text{ g H}) - (0.239764 \text{ g Fe}) = 0.412203 \text{ g O}$$

$$\text{Moles of O} = (0.412203 \text{ g O} (1 \text{ mol O} / 16.00 \text{ g O})) = 0.0257627 \text{ mol O}$$

Divide each of the moles by the smallest value (Fe):

$$\text{C: } (0.0644853 \text{ mol}) / (0.004293 \text{ mol}) = 15.0$$

$$\text{H: } (0.090144 \text{ mol}) / (0.004293 \text{ mol}) = 21.0$$

$$\text{Fe: } (0.004293 \text{ mol}) / (0.004293 \text{ mol}) = 1.0$$

$$\text{O: } (0.0257627 \text{ mol}) / (0.004293 \text{ mol}) = 6.0$$

These values give an empirical formula = **C₁₅H₂₁O₆Fe**

- 3.143 a) Moles of each element are needed. These may be determined from the mass of each element. The mass of Pt is given and the mass of F is (0.519 - 0.327) g F = 0.192 g F. Calculating the moles of each element gives:

$$\text{Moles Pt} = (0.327 \text{ g Pt}) (1 \text{ mol Pt} / 195.1 \text{ g Pt}) = 0.001676 \text{ mol Pt (unrounded)}$$

$$\text{Moles F} = (0.192 \text{ g F}) (1 \text{ mol F} / 19.00 \text{ g F}) = 0.010105 \text{ mol F (unrounded)}$$

Dividing each by the smaller value (Pt) gives:

$$\text{Pt: } (0.001676 \text{ mol}) / (0.001676 \text{ mol}) = 1.0$$

$$\text{F: } (0.010105 \text{ mol}) / (0.001676 \text{ mol}) = 6.0$$

Empirical formula = **PtF₆**

- b) Determine the moles of Xe and of PtF₆.

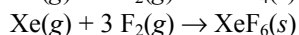
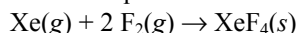
$$\text{Moles PtF}_6 = (0.265 \text{ g PtF}_6) (1 \text{ mol PtF}_6 / 309.1 \text{ g PtF}_6) = 0.0008573 \text{ mol PtF}_6 \text{ (unrounded)}$$

$$\text{Moles Xe} = [(0.378 - 0.265) \text{ g Xe}] (1 \text{ mol Xe} / 131.3 \text{ g Xe}) = 0.0008606 \text{ mol Xe (unrounded)}$$

The moles are essentially identical so dividing by the smaller value gives 1 in each case. The empirical formula is: **XePtF₆**.

- c) This problem can be solved as a system of two equations and two unknowns.

The two equations are:



$$\text{Moles of Xe consumed} = 1.85 \times 10^{-4} \text{ mol present} - 9.00 \times 10^{-6} \text{ mol excess} = 1.76 \times 10^{-4} \text{ mol Xe}$$

$$\text{Then } x + y = 1.76 \times 10^{-4} \text{ mol Xe consumed}$$

$$2x + 3y = 5.00 \times 10^{-4} \text{ mol F}_2 \text{ consumed}$$

Solve for x using the first equation and substitute the value of x into the second equation:

$$x = 1.76 \times 10^{-4} - y$$

$$2(1.76 \times 10^{-4} - y) + 3y = 5.00 \times 10^{-4}$$

$$(3.52 \times 10^{-4}) - 2y + 3y = 5.00 \times 10^{-4}$$

$$y = (5.00 \times 10^{-4}) - (3.52 \times 10^{-4}) = 1.48 \times 10^{-4} \text{ mol XeF}_6$$

$$x = (1.76 \times 10^{-4}) - (1.48 \times 10^{-4}) = 2.8 \times 10^{-5} \text{ mol XeF}_4$$

Convert moles of each product to grams using the molar masses (unrounded values):

$$\text{Mass XeF}_4 = (2.8 \times 10^{-5} \text{ mol XeF}_4) \times (207.3 \text{ g/mol}) = 5.8044 \times 10^{-3} \text{ g XeF}_4$$

$$\text{Mass XeF}_6 = (1.48 \times 10^{-4} \text{ mol XeF}_6) \times (245.3 \text{ g/mol}) = 3.63044 \times 10^{-2} \text{ g XeF}_6$$

Calculate the percent of each compound using the total weight of the products:

$$(5.8044 \times 10^{-3} + 3.63044 \times 10^{-2}) \text{ g} = 0.0421088 \text{ g}$$

$$\% \text{ XeF}_4 = [(5.8044 \times 10^{-3} \text{ g}) / (0.0421088 \text{ g})] \times 100\% = 13.784 = \mathbf{14\% \text{ XeF}_4}$$

$$\% \text{ XeF}_6 = [(3.63044 \times 10^{-2} \text{ g}) / (0.0421088 \text{ g})] \times 100\% = 86.2157 = \mathbf{86.2\% \text{ XeF}_6}$$

3.144 Plan: This problem involves running through a series of unit conversions.

Solution:

$$\text{a) Grams of heme} = (0.45 \text{ g hemoglobin}) \left(\frac{6.0\% \text{ heme}}{100\% \text{ hemoglobin}} \right) = 0.0270 = \mathbf{0.027 \text{ g heme}}$$

$$\text{b) Mole of heme} = (0.027 \text{ g heme}) \left(\frac{1 \text{ mol heme}}{616.49 \text{ g heme}} \right) = 4.37966 \times 10^{-5} = \mathbf{4.4 \times 10^{-5} \text{ mol heme}}$$

$$\begin{aligned} \text{c) Grams of Fe} &= (4.37966 \times 10^{-5} \text{ mol heme}) \left(\frac{1 \text{ mol Fe}}{1 \text{ mol heme}} \right) \left(\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \right) \\ &= 2.44604 \times 10^{-3} = \mathbf{2.4 \times 10^{-3} \text{ g Fe}} \end{aligned}$$

$$\begin{aligned} \text{d) Grams of hemin} &= (4.37966 \times 10^{-5} \text{ mol Heme}) \left(\frac{1 \text{ mol hemin}}{1 \text{ mol heme}} \right) \left(\frac{651.94 \text{ g hemin}}{1 \text{ mol hemin}} \right) \\ &= 2.85528 \times 10^{-2} = \mathbf{2.9 \times 10^{-2} \text{ g hemin}} \end{aligned}$$

3.145 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.

$$M(1.00) + B(0.667) = 0.704$$

$$(1.00 - B)(1.00) + B(0.667) = 0.704$$

$$B = 0.888889 \text{ mol braunite}$$

$$M = 0.111111 \text{ mol manganosite}$$

$$\text{Mass of braunite} = 0.888889 \text{ mol} \left(\frac{157.88 \text{ g}}{1 \text{ mol}} \right) = 140.338 \text{ g braunite}$$

$$\text{Mass of manganosite (g)} = 0.111111 \text{ mol} \left(\frac{70.94 \text{ g}}{1 \text{ mol}} \right) = 7.88221 \text{ g manganosite}$$

$$\text{Mass percent braunite} = [140.338 \text{ g} / (140.338 \text{ g} + 7.88221 \text{ g})] \times 100\% = 94.6821\%$$

$$\text{Mass percent manganosite} = [7.88221 \text{ g} / (140.338 \text{ g} + 7.88221 \text{ g})] \times 100\% = 5.3179\%$$

$$\text{Mass of braunite} = (542.3 \text{ g}) (94.6821\% / 100\%) = 513.461 = \mathbf{514 \text{ g braunite}}$$

$$\text{Mass of manganosite} = (542.3 \text{ g}) (5.3179\% / 100\%) = 28.839 = \mathbf{28.8 \text{ g manganosite}}$$

$$\text{b) Mn}^{3+}: \text{Mn}^{2+} = \frac{2(0.88889)}{(0.11111)} = 16.000 = \mathbf{16.0}$$

3.146 First, a balanced chemical equation is needed: $4 \text{ FeS}_2(s) + 11 \text{ O}_2(g) \rightarrow 2 \text{ Fe}_2\text{O}_3(s) + 8 \text{ SO}_2(g)$

$$\begin{aligned} \text{Grams FeS}_2 &= (1.00 \text{ kg SO}_2) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol SO}_2}{64.07 \text{ g SO}_2} \right) \left(\frac{4 \text{ mol FeS}_2}{8 \text{ mol SO}_2} \right) \left(\frac{119.99 \text{ g FeS}_2}{1 \text{ mol FeS}_2} \right) \\ &= 936.3977 = \mathbf{936 \text{ g FeS}_2} \end{aligned}$$

$$\text{Grams O}_2 = (1.00 \text{ kg SO}_2) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol SO}_2}{64.07 \text{ g SO}_2} \right) \left(\frac{11 \text{ mol O}_2}{8 \text{ mol SO}_2} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 686.749 = \mathbf{687 \text{ g O}_2}$$

$$\begin{aligned} \text{Grams Fe}_2\text{O}_3 &= (1.00 \text{ kg SO}_2) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol SO}_2}{64.07 \text{ g SO}_2} \right) \left(\frac{2 \text{ mol Fe}_2\text{O}_3}{8 \text{ mol SO}_2} \right) \left(\frac{159.70 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} \right) \\ &= 623.147 = \mathbf{623 \text{ g Fe}_2\text{O}_3} \end{aligned}$$

- 3.147 Plan: Determine the molecular formula and the molar mass of each of the compounds. From the amount of nitrogen present and the molar mass, the percent nitrogen may be determined. The moles need to be determined for part (b).

Solution:

a) To find mass percent of nitrogen, first determine molecular formula, then the molar mass of each compound. Mass percent is then calculated from the mass of nitrogen in the compound divided by the molar mass of the compound, and multiply by 100%.

Urea: $\text{CH}_4\text{N}_2\text{O}$, $\mathcal{M} = 60.06 \text{ g/mol}$

$$\% \text{ N} = [(2 \times 14.01 \text{ g/mol N}) / 60.06 \text{ g/mol CH}_4\text{N}_2\text{O}] \times 100\% = 46.6580 = \mathbf{46.66\% \text{ N in urea}}$$

Arginine: $\text{C}_6\text{H}_{15}\text{N}_4\text{O}_2$, $\mathcal{M} = 175.22 \text{ g/mol}$

$$\% \text{ N} = [(4 \times 14.01 \text{ g/mol N}) / 175.22 \text{ g/mol C}_6\text{H}_{15}\text{N}_4\text{O}_2] \times 100\% \\ = 31.98265 = \mathbf{31.98\% \text{ N in arginine}}$$

Ornithine: $\text{C}_5\text{H}_{13}\text{N}_2\text{O}_2$, $\mathcal{M} = 133.17 \text{ g/mol}$

$$\% \text{ N} = [(2 \times 14.01 \text{ g/mol N}) / 133.17 \text{ g/mol C}_5\text{H}_{13}\text{N}_2\text{O}_2] \times 100\% \\ = 21.04077 = \mathbf{21.04\% \text{ N in ornithine}}$$

b) Grams of N =

$$(143.2 \text{ g C}_5\text{H}_{13}\text{N}_2\text{O}_2) \left(\frac{1 \text{ mol C}_5\text{H}_{13}\text{N}_2\text{O}_2}{133.17 \text{ g C}_5\text{H}_{13}\text{N}_2\text{O}_2} \right) \left(\frac{1 \text{ mol CH}_4\text{N}_2\text{O}}{1 \text{ mol C}_5\text{H}_{13}\text{N}_2\text{O}_2} \right) \left(\frac{2 \text{ mol N}}{1 \text{ mol CH}_4\text{N}_2\text{O}} \right) \left(\frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) \\ = 30.130390 = \mathbf{30.13 \text{ g N}}$$

- 3.148 a) Mole aspirin from $\text{C}_7\text{H}_6\text{O}_3 = (3.027 \text{ g C}_7\text{H}_6\text{O}_3) \left(\frac{1 \text{ mol C}_7\text{H}_6\text{O}_3}{138.12 \text{ g C}_7\text{H}_6\text{O}_3} \right) \left(\frac{1 \text{ mol C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_7\text{H}_6\text{O}_3} \right)$
 $= 0.0219157 \text{ mol C}_9\text{H}_8\text{O}_4$ (unrounded)

$$\text{Mole aspirin from C}_4\text{H}_6\text{O}_3 = (6.00 \text{ mL C}_4\text{H}_6\text{O}_3) \left(\frac{1.080 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mol C}_4\text{H}_6\text{O}_3}{102.09 \text{ g C}_4\text{H}_6\text{O}_3} \right) \left(\frac{1 \text{ mol C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_4\text{H}_6\text{O}_3} \right) \\ = 0.064734 \text{ mol C}_9\text{H}_8\text{O}_4$$
 (unrounded)

The limiting reactant is $\text{C}_7\text{H}_6\text{O}_3$

b) First, calculate the theoretical yield from the limiting reagent:

$$\text{Grams C}_9\text{H}_8\text{O}_4 = (3.027 \text{ g C}_7\text{H}_6\text{O}_3) \left(\frac{1 \text{ mol C}_7\text{H}_6\text{O}_3}{138.12 \text{ g C}_7\text{H}_6\text{O}_3} \right) \left(\frac{1 \text{ mol C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_7\text{H}_6\text{O}_3} \right) \left(\frac{180.15 \text{ g C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_9\text{H}_8\text{O}_4} \right) \\ = 3.9481 \text{ g C}_9\text{H}_8\text{O}_4$$
 (unrounded)

$$\text{Percent yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\% = \left(\frac{3.261 \text{ g}}{3.9481 \text{ g}} \right) \times 100\% = 82.5967 = \mathbf{82.60\% \text{ yield}}$$

$$\text{c) \% atom economy} = \frac{(\text{Number of Moles})(\text{Molar Mass of Desired Product})}{(\text{Sum of Number of Moles})(\text{Molar Mass for all Products})} \times 100\% \\ = \frac{(1 \text{ mol})(180.15 \text{ g/mol})}{(180.15 + 60.05) \text{ g/mol}} \times 100\% = \mathbf{75.00\% \text{ atom economy}}$$

- 3.149 Plan: Determine the molar mass of each product and plug into atom percent equation.

Solution:

Molar masses of product: N_2H_4 : 32.05 g/mol NaCl : 58.44 g/mol H_2O : 18.02 g/mol

$$\% \text{ atom economy} = \frac{(\text{Number of Moles})(\text{Molar Mass of Desired Product})}{(\text{Sum of Number of Moles})(\text{Molar Mass for all Products})} \times 100\% \\ = \frac{(1 \text{ mol})(32.05 \text{ g/mol})}{((32.05 + 58.44 + 18.02) \text{ g/mol})} \times 100\% = 29.5364 = \mathbf{29.54\% \text{ atom economy}}$$

3.150 mass $\text{FeCr}_2\text{O}_4 =$

$$\begin{aligned} & (\text{kg Paint}) \left(\frac{10^3 \text{ g}}{\text{kg}} \right) \left(\frac{0.511\%}{100\%} \right) \left(\frac{1 \text{ mol PbCrO}_4}{323.2 \text{ g PbCrO}_4} \right) \left(\frac{1 \text{ mol Cr}}{1 \text{ mol PbCrO}_4} \right) \left(\frac{1 \text{ mol FeCr}_2\text{O}_4}{2 \text{ mol Cr}} \right) \left(\frac{223.85 \text{ g FeCr}_2\text{O}_4}{1 \text{ mol FeCr}_2\text{O}_4} \right) \\ & = 1.769606 = \mathbf{1.77 \text{ g FeCr}_2\text{O}_4} \end{aligned}$$

3.151 Plan: Convert the mass of ethanol to moles, and use moles to determine the theoretical yield of diethyl ether. This theoretical yield, along with the given (actual) yield, is used to determine the percent yield. The difference between the actual and theoretical yields is related to the quantity of ethanol that did not produce diethyl ether. Fifty percent of the mass difference is converted to moles of diethyl ether, then to moles of ethanol and ethylene, and finally to the mass of ethylene.

Solution:

a) The determination of the theoretical yield:

Grams diethyl ether =

$$\begin{aligned} & (50.0 \text{ g CH}_3\text{CH}_2\text{OH}) \left(\frac{1 \text{ mol CH}_3\text{CH}_2\text{OH}}{46.07 \text{ g CH}_3\text{CH}_2\text{OH}} \right) \left(\frac{1 \text{ mol CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3}{2 \text{ mol CH}_3\text{CH}_2\text{OH}} \right) \left(\frac{74.12 \text{ g CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3}{1 \text{ mol CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3} \right) \\ & = 40.2214 \text{ g diethyl ether (unrounded)} \end{aligned}$$

Determining the percent yield:

$$\text{Percent yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\% = \left(\frac{33.9 \text{ g}}{40.2214 \text{ g}} \right) \times 100\% = 84.2835 = \mathbf{84.3\% \text{ yield}}$$

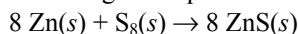
b) To determine the amount of ethanol not producing diethyl ether, we will use the difference between the theoretical yield and actual yield (unrounded) to determine the amount of diethyl ether that did not form and hence, the amount of ethanol that did not produce the desired product. Fifty percent of this amount will be used to determine the amount of ethylene formed.

Mass difference = $40.2214 \text{ g} - 33.9 \text{ g} = 6.3214 \text{ g}$ diethyl ether that did not form

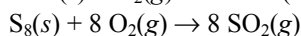
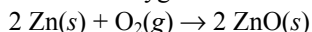
Grams ethylene =

$$\begin{aligned} & (6.3214 \text{ g (C}_2\text{H}_5)_2\text{O}) \left(\frac{1 \text{ mol (C}_2\text{H}_5)_2\text{O}}{74.12 \text{ g (C}_2\text{H}_5)_2\text{O}} \right) \left(\frac{2 \text{ mol C}_2\text{H}_6\text{O}}{1 \text{ mol (C}_2\text{H}_5)_2\text{O}} \right) \left(\frac{1 \text{ mol C}_2\text{H}_4}{1 \text{ mol C}_2\text{H}_6\text{O}} \right) \left(\frac{28.05 \text{ g C}_2\text{H}_4}{1 \text{ mol C}_2\text{H}_4} \right) \left(\frac{50.0\%}{100\%} \right) \\ & = 2.39227 = \mathbf{2.39 \text{ g ethylene}} \end{aligned}$$

3.152 First, balance the given equation:



The reactions with oxygen are:



a) Since the amounts of both starting materials are given, we must first find the limiting reagent:

$$\text{Moles of ZnS from Zn} = (85.2 \text{ g Zn}) \left(\frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \right) \left(\frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} \right) = 1.30295 \text{ mol ZnS (unrounded)}$$

$$\text{Moles of ZnS from S}_8 = (52.4 \text{ g S}_8) \left(\frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \right) \left(\frac{8 \text{ mol ZnS}}{1 \text{ mol S}_8} \right) = 1.6339 \text{ mol ZnS (unrounded)}$$

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.

$$\text{Theoretical yield} = (85.2 \text{ g Zn}) \left(\frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \right) \left(\frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} \right) \left(\frac{97.46 \text{ g ZnS}}{1 \text{ mol ZnS}} \right) = 126.986 \text{ g ZnS (unrounded)}$$

$$\text{Percent yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\% = \left(\frac{105.4 \text{ g}}{126.986 \text{ g}} \right) \times 100\% = 83.001 = \mathbf{83.0\% \text{ yield}}$$

b) The theoretical yield indicates that 83.0% of the zinc produced zinc sulfide so $(100 - 83.0)\% = 17.0\%$ of the zinc became zinc oxide. This allows the calculation of the amount of zinc oxide formed.

$$\begin{aligned} \text{Grams zinc oxide} &= (85.2 \text{ g Zn}) \left(\frac{17.0\%}{100\%} \right) \left(\frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \right) \left(\frac{2 \text{ mol ZnO}}{2 \text{ mol Zn}} \right) \left(\frac{81.39 \text{ g ZnO}}{1 \text{ mol ZnO}} \right) \\ &= 18.0280 = \mathbf{18.0 \text{ g ZnO}} \end{aligned}$$

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.

$$\begin{aligned} \text{Mass SO}_2 &= \left[52.4 \text{ g S}_8 \left(\frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \right) - (105.4 \text{ g ZnS}) \left(\frac{1 \text{ mol ZnS}}{97.46 \text{ g ZnS}} \right) \left(\frac{1 \text{ mol S}_8}{8 \text{ mol ZnS}} \right) \right] \left(\frac{8 \text{ mol SO}_2}{1 \text{ mol S}_8} \right) \left(\frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} \right) \\ &= 35.3959 = \mathbf{35 \text{ g SO}_2} \end{aligned}$$

The result of the subtraction in the calculation left only two significant figures.

3.153 Plan: Careful tracking of the units will carry you through this problem.

Solution:

$$\text{a) Dissolved salt} = \left(\frac{2.50 \text{ kg Salt}}{\text{L}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) (50.0 \text{ mL}) = \mathbf{125 \text{ g salt}}$$

b) Convert the grams from part (a) to moles of salt, then to moles of cocaine, and finally to grams of cocaine. A calculation similar to that in part (a) will give the volume required.

$$\begin{aligned} \text{Mass cocaine} &= (125 \text{ g Salt}) \left(\frac{1 \text{ mol Salt}}{339.81 \text{ g Salt}} \right) \left(\frac{1 \text{ mol Cocaine}}{1 \text{ mol Salt}} \right) \left(\frac{303.35 \text{ g Cocaine}}{1 \text{ mol Cocaine}} \right) \\ &= 111.588 \text{ g cocaine (unrounded)} \end{aligned}$$

$$\begin{aligned} \text{Additional water needed} &= (111.588 \text{ g Cocaine}) \left(\frac{1 \text{ L}}{1.70 \text{ g Cocaine}} \right) - (50.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \\ &= 65.59 = \mathbf{65.6 \text{ L H}_2\text{O}} \end{aligned}$$

3.154 a) $x = 0$

$$\text{La}_2\text{Sr}_0\text{CuO}_4 = 2 (138.9 \text{ g La/mol}) + 0 (87.62 \text{ g Sr/mol}) + 1 (63.55 \text{ g Cu/mol}) + 4 (16.00 \text{ g O/mol}) = \mathbf{405.4 \text{ g/mol}}$$

$x = 1$

$$\text{La}_1\text{Sr}_1\text{CuO}_4 = 1 (138.9 \text{ g La/mol}) + 1 (87.62 \text{ g Sr/mol}) + 1 (63.55 \text{ g Cu/mol}) + 4 (16.00 \text{ g O/mol}) = \mathbf{354.1 \text{ g/mol}}$$

$x = 0.163$

$$\text{La}_{(2-0.163)}\text{Sr}_{0.163}\text{CuO}_4$$

$$\begin{aligned} &= (2 - 0.163) (138.9 \text{ g La/mol}) + (0.163) (87.62 \text{ g Sr/mol}) + 1 (63.55 \text{ g Cu/mol}) + 4 (16.00 \text{ g O/mol}) \\ &= \mathbf{397.0 \text{ g/mol}} \end{aligned}$$

b) Assuming x grams to be the "equal" mass leads to:

$$\text{Moles product from BaCO}_3 = (x \text{ g BaCO}_3) \left(\frac{1 \text{ mol BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left(\frac{2 \text{ mol Product}}{4 \text{ mol BaCO}_3} \right) = 0.002534x \text{ mol product}$$

$$\text{Moles product from CuO} = (x \text{ g CuO}) \left(\frac{1 \text{ mol CuO}}{79.55 \text{ g CuO}} \right) \left(\frac{2 \text{ mol Product}}{6 \text{ mol CuO}} \right) = 0.004190x \text{ mol product}$$

$$\text{Moles product from Y}_2\text{O}_3 = (x \text{ g Y}_2\text{O}_3) \left(\frac{1 \text{ mol Y}_2\text{O}_3}{225.82 \text{ g Y}_2\text{O}_3} \right) \left(\frac{2 \text{ mol Product}}{1 \text{ mol Y}_2\text{O}_3} \right) = 0.008857x \text{ mol product}$$

BaCO₃ is the limiting reactant.

c) These calculations are based on the limiting reactant.

BaCO₃ remaining = 0% (Limiting reagent)

$$\text{CuO remaining} = x \text{ g CuO} - (x \text{ g BaCO}_3) \left(\frac{1 \text{ mol BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left(\frac{6 \text{ mol CuO}}{4 \text{ mol BaCO}_3} \right) \left(\frac{79.55 \text{ g CuO}}{1 \text{ mol CuO}} \right)$$

$$= 0.39521x \text{ g CuO (unrounded)}$$

$$\text{Percent CuO} = \left(\frac{0.39521x \text{ g}}{x \text{ g}} \right) \times 100\% = 39.521 = \mathbf{39.52\% \text{ CuO remaining}}$$

$$\text{Y}_2\text{O}_3 \text{ remaining} = x \text{ g Y}_2\text{O}_3 - (x \text{ g BaCO}_3) \left(\frac{1 \text{ mol BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left(\frac{1 \text{ mol Y}_2\text{O}_3}{4 \text{ mol BaCO}_3} \right) \left(\frac{225.82 \text{ g Y}_2\text{O}_3}{1 \text{ mol Y}_2\text{O}_3} \right)$$

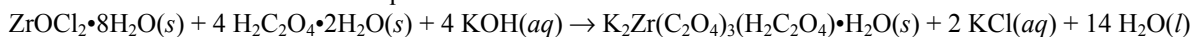
$$= 0.713862x \text{ g Y}_2\text{O}_3 \text{ (unrounded)}$$

$$\text{Percent Y}_2\text{O}_3 = \left(\frac{0.713862x \text{ g}}{x \text{ g}} \right) \times 100\% = 71.3862 = \mathbf{71.38\% \text{ Y}_2\text{O}_3 \text{ remaining}}$$

3.155 Plan: The balanced chemical equation is needed. From the balanced chemical equation and the masses we can calculate the limiting reagent. The limiting reagent will be used to calculate the theoretical yield, and finally the percent yield.

Solution:

Determine the balanced chemical equation:



Determine the limiting reactant:

$$\text{Moles product from ZrOCl}_2 \cdot 8\text{H}_2\text{O} = (1.60 \text{ g Zr compd}) \left(\frac{1 \text{ mol Zr compd}}{322.25 \text{ g Zr compd}} \right) \left(\frac{1 \text{ mol Product}}{1 \text{ mol Zr compd}} \right)$$

$$= 0.0049651 \text{ mol product (unrounded)}$$

Moles product from H₂C₂O₄•2H₂O =

$$(5.20 \text{ g H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}}{126.07 \text{ g H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}} \right) \left(\frac{1 \text{ mol Product}}{4 \text{ mol H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}} \right)$$

$$= 0.0103117 \text{ mol product (unrounded)}$$

Moles product from KOH = irrelevant because KOH is stated to be in excess.

The ZrOCl₂•8H₂O is limiting, and will be used to calculate the theoretical yield:

$$\text{Grams product} = (1.60 \text{ g Zr compd}) \left(\frac{1 \text{ mol Zr compd}}{322.25 \text{ g Zr compd}} \right) \left(\frac{1 \text{ mol Product}}{1 \text{ mol Zr compd}} \right) \left(\frac{541.53 \text{ g Product}}{1 \text{ mol Product}} \right)$$

$$= 2.68874 \text{ g product (unrounded)}$$

Finally, calculating the percent yield:

$$\text{Percent yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\% = \left(\frac{1.20 \text{ g}}{2.68874 \text{ g}} \right) \times 100\% = 44.630 = \mathbf{44.6\% \text{ yield}}$$

3.156 Convert the mass of carbon to moles and carry through the steps using these moles.

$$\text{a) Number of atoms} = (0.001 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) \left(\frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C}} \right) = 5.01415 \times 10^{19} = \mathbf{5 \times 10^{19} \text{ C atoms}}$$

$$\text{b) Mass PbO} = (0.001 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) \left(\frac{3 \text{ mol PbO}}{1 \text{ mol C}} \right) \left(\frac{223.2 \text{ g PbO}}{1 \text{ mol PbO}} \right) = 0.05575 = \mathbf{0.06 \text{ g PbO}}$$

3.157 Alcohols with $C_4H_{10}O$:

