CHAPTER 4 THE MAJOR CLASSES OF CHEMICAL REACTIONS

4.1   The distribution of its bonding electrons and the shape of the molecule are both unsymmetrical.

4.2   Ionic and polar covalent compounds are most likely to be soluble in water. Because water is polar, the partial charges in water molecules are able to interact with the charges, either ionic or dipole-induced, in other substances.

4.3   Ions must be present in an aqueous solution for it to conduct an electric current. Ions come from ionic compounds or from other electrolytes such as acids and bases.

4.4   The ions on the surface of the solid attract the water molecules (cations attract the "negative" end and anions attract the "positive" end of the water molecules). The interaction of the solvent with the ions overcomes the attraction of the oppositely charged ions for one another, and they are released into the solution.

4.5   a) **2** — This is the only solution containing a divalent cation (Ca^{2+}).
        b) **3** — This is the only solution containing a divalent anion (SO_4^{2-}).
        c) **1** — This is the only solution where both the cation (NH_4^+) and the anion (Br^-) are univalent.

4.6   The box in (**2**) best represents a volume of magnesium nitrate solution. Upon dissolving the salt in water, magnesium nitrate, $Mg(NO_3)_2$, would dissociate to form one Mg^{2+} ion for every two NO_3^- ions, thus forming twice as many nitrate ions. Only box (2) has twice as many nitrate ions (red circles) as magnesium ions (blue circles).

4.7   In some cases, the force of the attraction between the ions is so strong that it cannot be overcome by the interaction of the ions with the water molecules. These materials will be insoluble in water.

4.8   The interaction with water depends on the structure of the molecule. If the interaction is good, the substance will be soluble; otherwise, the substance will not be very soluble. For example, if a covalent molecule contains polar groups, they will interact well with the polar solvent water. A few covalent molecules, such as HCl, produce ions when dissolving in water.

4.9   In general, covalent compounds which produce ions in aqueous solution interact with the water molecules to form either $H^+(aq)$ or $OH^-(aq)$, producing an acidic or basic (alkaline) solution, respectively. These compounds contain combined hydrogen atoms. Examples include hydrochloric acid (HCl), acetic acid ($HC_2H_3O_2$), and ammonia (NH_3).

4.10   a) Benzene is likely to be insoluble in water because it is non-polar and water is polar.
        b) Sodium hydroxide, an ionic compound, is likely to be soluble in water since the ions from sodium hydroxide will be held in solution through ion-dipole attractions with water.
        c) Ethanol (CH_3CH_2OH) will likely be soluble in water because the alcohol group (-OH) will hydrogen bond with the water.
        d) Potassium acetate, an ionic compound, will likely dissolve in water to form sodium ions and acetate ions that are held in solution through ion-dipole attractions to water.

4.11   a) Lithium nitrate, an ionic compound, would be expected to be soluble in water, and the solubility rules confirm this.
        b) Glycine (H_2NCH_2COOH) is a covalent compound, but it contains polar N-H and O-H bonds. This would make the molecule interact well with polar water molecules, and make it likely that it would be soluble.

c) Pentane (C_5H_{12}) has no bonds of significant polarity, so it would not be expected to be soluble in the polar solvent water.
d) Ethylene glycol ($HOCH_2CH_2OH$) molecules contain polar O-H bonds, similar to water, so it would be expected to be soluble.

4.12   a) An aqueous solution that contains ions conducts electricity. CsI is a soluble ionic compound, and a solution of this salt in water contains Cs^+ and I^- ions. Its solution conducts electricity.
b) HBr is a strong acid that dissociates completely in water. Its aqueous solution contains H^+ and Br^- ions, so it conducts electricity.

4.13   a) Yes; KOH is a strong base.
b) No; glucose is neither a salt, acid, nor base, so it would be a nonelectrolyte (even though it's soluble in water).

4.14   Plan: To determine the total moles of ions released, write a dissolution equation showing the correct molar ratios, and convert the given amounts to moles if necessary.
Solution:
a) Each mole of NH_4Cl dissolves in water to form 1 mole of NH_4^+ ions and 1 mole of Cl^- ions, or a total of 2 moles of ions.

$$\left(0.37 \text{ mol NH}_4\text{Cl}\right)\left(\frac{2 \text{ mol ions}}{1 \text{ mol NH}_4\text{Cl}}\right) = \textbf{0.74 mol of ions}$$

b) Each mole of $Ba(OH)_2 \cdot 8H_2O$ forms 1 mole of barium ions (Ba^{2+}) and 2 mol of hydroxide ions (OH^-), or a total of 3 mol of ions.

$$\left(35.4 \text{ g Ba(OH)}_2 \cdot 8\text{H}_2\text{O}\right)\left(\frac{1 \text{ mol Ba(OH)}_2 \cdot 8\text{H}_2\text{O}}{315.4 \text{ g Ba(OH)}_2 \cdot 8\text{H}_2\text{O}}\right)\left(\frac{3 \text{ mol ions}}{1 \text{ mol Ba(OH)}_2 \cdot 8\text{H}_2\text{O}}\right)$$
$$= 0.336715 = \textbf{0.337 mol of ions}$$

c) Each mole of LiCl produces 2 moles of ions (1 mole of lithium ions and 1 mole of chloride ions).

$$\left(3.55 \times 10^{18} \text{ FU LiCl}\right)\left(\frac{1 \text{ mol LiCl}}{6.022 \times 10^{23} \text{ FU LiCl}}\right)\left(\frac{2 \text{ mol ions}}{1 \text{ mol LiCl}}\right)$$
$$= 1.17901 \times 10^{-5} = \textbf{1.18} \times \textbf{10}^{-5} \textbf{ mol of ions}$$

4.15   a) Each mole of Rb_2SO_4 dissolves in water to form 2 moles of Rb^+ ions and 1 mole of SO_4^{2-} ions, or a total of 3 moles of ions.

$$\left(0.805 \text{ mol Rb}_2\text{SO}_4\right)\left(\frac{3 \text{ mol ions}}{1 \text{ mol Rb}_2\text{SO}_4}\right) = 2.415 = 2.42 \textbf{ mol of ions}$$

b) Each mole of $Ca(NO_3)_2$ forms 1 mole of calcium ions (Ca^{2+}) and 2 moles of nitrate ions (NO_3^-), or a total of 3 moles of ions.

$$\left(3.85 \times 10^{-3} \text{ g Ca(NO}_3)_2\right)\left(\frac{1 \text{ mol Ca(NO}_3)_2}{164.10 \text{ g Ca(NO}_3)_2}\right)\left(\frac{3 \text{ mol ions}}{1 \text{ mol Ca(NO}_3)_2}\right)$$
$$= 7.03839 \times 10^{-5} = \textbf{7.04} \times \textbf{10}^{-5} \textbf{ mol of ions}$$

c) Each moles of $Sr(HCO_3)_2$ produces 3 moles of ions (1 mole of strontium ions and two moles of bicarbonate ions).

$$\left(4.03 \times 10^{19} \text{ F.U. Sr(HCO}_3)_2\right)\left(\frac{1 \text{ mol Sr(HCO}_3)_2}{6.022 \times 10^{23} \text{ FU Sr(HCO}_3)_2}\right)\left(\frac{3 \text{ mol ions}}{1 \text{ mol Sr(HCO}_3)_2}\right)$$
$$= 2.0076 \times 10^{-4} = \textbf{2.01} \times \textbf{10}^{-4} \textbf{ mol of ions}$$

4.16   Plan: To determine the total moles of ions released, write a dissolution equation showing the correct molar ratios, and convert the given amounts to moles if necessary.

Solution:
a) Recall that phosphate, PO_4^{3-}, is a polyatomic anion and does not dissociate further in water.

     $K_3PO_4(s) \rightarrow 3\ K^+(aq) + PO_4^{3-}(aq)$
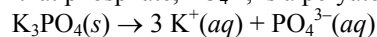
     4 moles of ions are released when one mole of K_3PO_4 dissolves, so the total number of moles released is
(0.83 mol K_3PO_4) (4 mol ions/mol K_3PO_4) = 3.32 = **3.3 mol of ions**

b) $NiBr_2 \cdot 3H_2O(s) \rightarrow Ni^{2+}(aq) + 2\ Br^-(aq)$

     Three moles of ions are released when 1 mole of $NiBr_2 \cdot 3H_2O$ dissolves.
The waters of hydration become part of the larger bulk of water. Convert the grams of $NiBr_2 \cdot 3H_2O$ to moles using the molar mass (be sure to include the mass of the water):

$$\left(8.11 \times 10^{-3}\ g\ NiBr_2 \cdot 3H_2O\right)\left(\frac{1\ mol\ NiBr_2 \cdot 3H_2O}{272.54\ g\ NiBr_2 \cdot 3H_2O}\right)\left(\frac{3\ mol\ ions}{1\ mol\ NiBr_2 \cdot 3H_2O}\right)$$

$$= 8.9271 \times 10^{-5} = \textbf{8.93} \times \textbf{10}^{-5}\ \textbf{mol of ions}$$

c) $FeCl_3(s) \rightarrow Fe^{3+}(aq) + 3\ Cl^-(aq)$

     Recall that a mole contains 6.022×10^{23} entities, so a mole of $FeCl_3$ contains 6.022×10^{23} units of $FeCl_3$, (more easily expressed as formula units). Since the problem specifies only 1.23×10^{21} formula units, we know that the amount is some fraction of a mole.

$$\left(1.23 \times 10^{21}\ FU\ FeCl_3\right)\left(\frac{1\ mol\ FeCl_3}{6.022 \times 10^{23}\ FU\ FeCl_3}\right)\left(\frac{4\ mol\ ions}{1\ mol\ FeCl_3}\right)$$

$$= 8.17004 \times 10^{-3} = \textbf{8.17} \times \textbf{10}^{-3}\ \textbf{mol of ions}$$

4.17   a) $Na_2HPO_4(s) \rightarrow 2\ Na^+(aq) + HPO_4^{2-}$ 3 ions produced
           (0.734 mol Na_2HPO_4) (3 mol ions/mol Na_2HPO_4) = 2.202 = **2.20 mol of ions**

       b) $CuSO_4 \cdot 5H_2O(s) \rightarrow Cu^{+2}(aq) + SO_4^{2-}(aq)$ 2 ions produced

$$\left(3.86\ g\ CuSO_4 \cdot 5H_2O\right)\left(\frac{1\ mol\ CuSO_4 \cdot 5H_2O}{217.63\ g\ CuSO_4 \cdot 5H_2O}\right)\left(\frac{2\ mol\ ions}{1\ mol\ CuSO_4 \cdot 5H_2O}\right)$$

$$= 3.5473 \times 10^{-2} = \textbf{3.55} \times \textbf{10}^{-2}\ \textbf{mol of ions}$$

       c) $NiCl_2(s) \rightarrow Ni^{2+}(aq) + 2\ Cl^-(aq)$ 3 ions produced

$$\left(8.66 \times 10^{20}\ FU\ NiCl_2\right)\left(\frac{1\ mol\ NiCl_2}{6.022 \times 10^{23}\ FU\ NiCl_2}\right)\left(\frac{3\ mol\ ions}{1\ mol\ NiCl_2}\right)$$

$$= 4.31418 \times 10^{-3} = \textbf{4.31} \times \textbf{10}^{-3}\ \textbf{mol of ions}$$

4.18   a) $AlCl_3(s) \rightarrow Al^{3+}(aq) + 3\ Cl^-(aq)$

$$\text{Moles } Al^{3+} = \left(100.\ mL\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{2.45\ mol\ AlCl_3}{L}\right)\left(\frac{1\ mol\ Al^{3+}}{1\ mol\ AlCl_3}\right) = \textbf{0.245 mol } \textbf{Al}^{3+}$$

$$\text{Moles } Cl^- = \left(100.\ mL\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{2.45\ mol\ AlCl_3}{L}\right)\left(\frac{3\ mol\ Cl^-}{1\ mol\ AlCl_3}\right) = \textbf{0.735 mol } \textbf{Cl}^-$$

$$Al^{3+}\ ions = \left(0.245\ mol\ Al^{3+}\right)\left(\frac{6.022 \times 10^{23}\ Al^{3+}}{1\ mol\ Al^{3+}}\right) = 1.47539 \times 10^{23} = \textbf{1.48} \times \textbf{10}^{23}\ \textbf{Al}^{3+}\ \textbf{ions}$$

$$Cl^-\ ions = \left(0.735\ mol\ Cl^-\right)\left(\frac{6.022 \times 10^{23}\ Cl^-}{1\ mol\ Cl^-}\right) = 4.42617 \times 10^{23} = \textbf{4.43} \times \textbf{10}^{23}\ \textbf{Cl}^-\ \textbf{ions}$$

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

$$\text{Moles } Li^+ = \left(1.80\ L\right)\left(\frac{2.59\ g\ Li_2SO_4}{1\ L}\right)\left(\frac{1\ mol\ Li_2SO_4}{109.95\ g\ Li_2SO_4}\right)\left(\frac{2\ mol\ Li^+}{1\ mol\ Li_2SO_4}\right) = 0.08480 = \textbf{0.0848 mol } \textbf{Li}^+$$

$$\text{Moles SO}_4{}^{2-} = (1.80 \text{ L})\left(\frac{2.59 \text{ g Li}_2\text{SO}_4}{1 \text{ L}}\right)\left(\frac{1 \text{ mol Li}_2\text{SO}_4}{109.95 \text{ g Li}_2\text{SO}_4}\right)\left(\frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol Li}_2\text{SO}_4}\right) = 0.04240 = \textbf{0.0424 mol SO}_4{}^{2-}$$

$$\text{Li}^+ \text{ ions} = (0.084802 \text{ mol Li}^+)\left(\frac{6.022 \text{ x } 10^{23} \text{ Li}^+}{1 \text{ mol Li}^+}\right) = 5.106787 \text{ x } 10^{22} = \textbf{5.11 x } \textbf{10}^{22} \textbf{ Li}^+ \textbf{ ions}$$

$$\text{SO}_4{}^{2-} \text{ ions} = (0.042401 \text{ mol SO}_4^{2-})\left(\frac{6.022 \text{ x } 10^{23} \text{ SO}_4^{2-}}{1 \text{ mol SO}_4^{2-}}\right) = 2.55339 \text{ x } 10^{22} = \textbf{2.55 x } \textbf{10}^{22} \textbf{ SO}_4{}^{2-} \textbf{ ions}$$

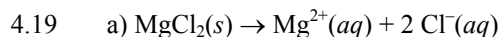c) $\text{KBr}(s) \rightarrow \text{K}^+(aq) + \text{Br}^-(aq)$

$$\text{K}^+ \text{ ions} = (225 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{1.68 \text{ x } 10^{22} \text{ FU KBr}}{\text{L}}\right)\left(\frac{1 \text{ K}^+}{1 \text{ FU KBr}}\right) = \textbf{3.78 x } \textbf{10}^{21} \textbf{ K}^+ \textbf{ ions}$$

$$\text{Br}^- \text{ ions} = (225 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{1.68 \text{ x } 10^{22} \text{ FU KBr}}{\text{L}}\right)\left(\frac{1 \text{ Br}^-}{1 \text{ FU KBr}}\right) = \textbf{3.78 x } \textbf{10}^{21} \textbf{ Br}^- \textbf{ ions}$$

$$\text{Moles K}^+ = (3.78 \text{ x } 10^{21} \text{ K}^+)\left(\frac{1 \text{ mol K}^+}{6.022 \text{ x } 10^{23} \text{ K}^+}\right) = 6.27698 \text{ x } 10^{-3} = \textbf{6.28 x } \textbf{10}^{-3} \textbf{ mol K}^+$$

$$\text{Moles Br}^- = (3.78 \text{ x } 10^{21} \text{ Br}^-)\left(\frac{1 \text{ mol Br}^-}{6.022 \text{ x } 10^{23} \text{ Br}^-}\right) = 6.27698 \text{ x } 10^{-3} = \textbf{6.28 x } \textbf{10}^{-3} \textbf{ mol Br}^-$$

4.19   a) $\text{MgCl}_2(s) \rightarrow \text{Mg}^{2+}(aq) + 2 \text{ Cl}^-(aq)$

$$\text{Moles Mg}^{2+} = (88.\text{mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{1.75 \text{ mol MgCl}_2}{\text{L}}\right)\left(\frac{1 \text{ mol Mg}^{2+}}{1 \text{ mol MgCl}_2}\right) = 0.154 = \textbf{0.15 mol Mg}^{2+}$$

$$\text{Moles Cl}^- = (88.\text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{1.75 \text{ mol MgCl}_2}{\text{L}}\right)\left(\frac{2 \text{ mol Cl}^-}{1 \text{ mol MgCl}_2}\right) = 0.308 = \textbf{0.31 mol Cl}^-$$

$$\text{Mg}^{2+} \text{ ions} = (0.154 \text{ mol Mg}^{2+})\left(\frac{6.022 \text{ x } 10^{23} \text{ Mg}^{2+}}{1 \text{ mol Mg}^{2+}}\right) = 9.27388 \text{ x } 10^{22} = \textbf{9.3 x } \textbf{10}^{23} \textbf{ Mg}^{2+} \textbf{ ions}$$

$$\text{Cl}^- \text{ ions} = (0.308 \text{ mol Cl}^-)\left(\frac{6.022 \text{ x } 10^{23} \text{ Cl}^-}{1 \text{ mol Cl}^-}\right) = 1.854776 \text{ x } 10^{23} = \textbf{1.9 x } \textbf{10}^{23} \textbf{ Cl}^- \textbf{ ions}$$

b) $\text{Al}_2(\text{SO}_4)_3(s) \rightarrow 2 \text{ Al}^{3+}(aq) + 3 \text{ SO}_4{}^{2-}(aq)$
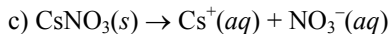
$$\text{Moles Al}^{3+} = (321 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.22 \text{ g Al}_2(\text{SO}_4)_3}{1 \text{ L}}\right)\left(\frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{342.17 \text{ g Al}_2(\text{SO}_4)_3}\right)\left(\frac{2 \text{ mol Al}^{3+}}{1 \text{ mol Al}_2(\text{SO}_4)_3}\right)$$
$$= 4.12777 \text{ x } 10^{-4} = \textbf{4.1 x } \textbf{10}^{-4} \textbf{ mol Al}^{3+}$$

$$\text{Moles SO}_4{}^{2-} = (321 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.22 \text{ g Al}_2(\text{SO}_4)_3}{1 \text{ L}}\right)\left(\frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{342.17 \text{ g Al}_2(\text{SO}_4)_3}\right)\left(\frac{3 \text{ mol SO}_4^{2-}}{1 \text{ mol Al}_2(\text{SO}_4)_3}\right)$$
$$= 6.191659 = \textbf{6.2 x } \textbf{10}^{-4} \textbf{ mol SO}_4{}^{2-}$$

$$\text{Al}^{3+} \text{ ions} = (4.12777 \text{ x } 10^{-4} \text{ mol Al}^{3+})\left(\frac{6.022 \text{ x } 10^{23} \text{ Al}^{3+}}{1 \text{ mol Al}^{3+}}\right) = 2.4857 \text{ x } 10^{20} = \textbf{2.5 x } \textbf{10}^{20} \textbf{ Al}^{3+} \textbf{ ions}$$

$$\text{SO}_4{}^{2-} \text{ ions} = (6.191659 \text{ x } 10^{-4} \text{ mol SO}_4^{2-})\left(\frac{6.022 \text{ x } 10^{23} \text{ SO}_4^{2-}}{1 \text{ mol SO}_4^{2-}}\right) = 3.7286 \text{ x } 10^{20} = \textbf{3.7 x } \textbf{10}^{20} \textbf{ SO}_4{}^{2-} \textbf{ ions}$$

c) $CsNO_3(s) \rightarrow Cs^+(aq) + NO_3^-(aq)$

$$Cs^+ \text{ ions} = (1.65 \text{ L}) \left(\frac{8.83 \times 10^{21} \text{ FU CsNO}_3}{\text{L}} \right) \left(\frac{1 \text{ Cs}^+}{1 \text{ FU CsNO}_3} \right)$$

$$= 1.45695 \times 10^{22} = \mathbf{1.46 \times 10^{22} \text{ Cs}^+ \text{ ions}}$$

$$NO_3^- \text{ ions} = (1.65 \text{ L}) \left(\frac{8.83 \times 10^{21} \text{ FU CsNO}_3}{\text{L}} \right) \left(\frac{1 \text{ NO}_3^-}{1 \text{ FU CsNO}_3} \right)$$

$$= 1.45695 \times 10^{22} = \mathbf{1.46 \times 10^{22} \text{ NO}_3^- \text{ ions}}$$

$$\text{Moles Cs}^+ = (1.45695 \times 10^{22} \text{ Cs}^+) \left(\frac{1 \text{ mol Cs}^+}{6.022 \times 10^{23} \text{ Cs}^+} \right) = 0.0241938 = \mathbf{0.0242 \text{ mol Cs}^+}$$

$$\text{Moles NO}_3^- = (1.45695 \times 10^{22} \text{ NO}_3^-) \left(\frac{1 \text{ mol NO}_3^-}{6.022 \times 10^{23} \text{ NO}_3^-} \right) = 0.0241938 = \mathbf{0.0242 \text{ mol NO}_3^-}$$

4.20   Plan: The acids in this problem are all strong acids, so you can assume that all acid molecules dissociate completely to yield H^+ ions and associated anions. One mole of $HClO_4$, HNO_3 and HCl each produce one mole of H^+ upon dissociation, so moles H^+ = moles acid. Molarity is expressed as moles/L instead of as M.
Solution:
a) Moles H^+ = mol $HClO_4$ = (1.40 L) (0.25 mol/L) = **0.35 mol H$^+$**
b) Moles H^+ = mol HNO_3 = (1.8 mL) (10^{-3} L/1 mL) (0.72 mol/L) = 1.296 x 10^{-3} = **1.3 x 10^{-3} mol H$^+$**
c) Moles H^+ = mol HCl = (7.6 L) (0.056 mol/L) = 0.4256 = **0.43 mol H$^+$**

4.21   The acids in this problem are all strong acids, so you can assume that all acid molecules dissociate completely to yield H^+ ions and associated anions. One mole of HBr, HI and HNO_3 each produce one mole of H^+ upon dissociation, so moles H^+ = moles acid. Molarity is expressed as moles/L instead of as M.
a) Moles H^+ = mol HBr = (1.4 mL) (10^{-3} L/1mL) (0.75 mol/L) = 1.05 x 10^{-3} = **1.0 x 10^{-3} mol H$^+$**
b) Moles H^+ = mol HI = (2.47 mL) (10^{-3}L/1 mL) (1.98 mol/L) = 4.8906 x 10^{-3} = **4.89 x 10^{-3} mol H$^+$**
c) Moles H^+ = mol HNO_3 = (395 mL) (10^{-3} L/1 mL) (0.270 mol/L) = 0.10665 = **0.107 mol H$^+$**

4.22   In this problem, there is no intermediate rounding. The ions are not calculated separately for those compounds where the number of cations equals the number of anions. The significant figures in each intermediate answer are underlined.
a) The volume of the seawater is needed.

$$\text{Volume (seawater)} = (1.00 \text{ kg}) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{\text{cm}^3}{1.04 \text{ g}} \right) \left(\frac{1 \text{ mL}}{1 \text{ cm}^3} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.\underline{961}538 \text{ L}$$

The moles of each ion are needed. If an ion comes from more than one source, the total moles are needed.
NaCl

$$\text{Moles Na}^+ = \text{mol Cl}^- = (26.5 \text{ g NaCl}) \left(\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \right) \left(\frac{1 \text{ mol ion}}{1 \text{ mol NaCl}} \right) = 0.\underline{4534}565 \text{ mol}$$

$MgCl_2$

$$\text{Moles Mg}^{2+} = (2.40 \text{ g MgCl}_2) \left(\frac{1 \text{ mol MgCl}_2}{95.21 \text{ g MgCl}_2} \right) \left(\frac{1 \text{ mol Mg}^{2+}}{1 \text{ mol MgCl}_2} \right) = 0.0\underline{2520}7 \text{ mol Mg}^{2+}$$

$$\text{Moles Cl}^- = (2.40 \text{ g MgCl}_2) \left(\frac{1 \text{ mol MgCl}_2}{95.21 \text{ g MgCl}_2} \right) \left(\frac{2 \text{ mol Cl}^-}{1 \text{ mol MgCl}_2} \right) = 0.0\underline{5041}5 \text{ mol Cl}^-$$

$MgSO_4$

$$\text{Moles Mg}^{2+} = \text{mol SO}_4^{2-} = (3.35 \text{ g MgSO}_4) \left(\frac{1 \text{ mol MgSO}_4}{120.38 \text{ g MgSO}_4} \right) \left(\frac{1 \text{ mol ion}}{1 \text{ mol MgSO}_4} \right) = 0.0\underline{2782}85 \text{ mol}$$

CaCl$_2$

$$\text{Moles Ca}^{2+} = \left(1.20 \text{ g CaCl}_2\right)\left(\frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2}\right)\left(\frac{1 \text{ mol Ca}^{2+}}{1 \text{ mol CaCl}_2}\right) = 0.0\underline{108}128 \text{ mol Ca}^{2+}$$

$$\text{Moles Cl}^- = \left(1.20 \text{ g CaCl}_2\right)\left(\frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2}\right)\left(\frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2}\right) = 0.0\underline{216}255 \text{ mol Cl}^-$$

KCl

$$\text{Moles K}^+ = \text{mol Cl}^- = \left(1.05 \text{ g KCl}\right)\left(\frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}}\right)\left(\frac{1 \text{ mol ion}}{1 \text{ mol KCl}}\right) = 0.0\underline{140}845 \text{ mol}$$

NaHCO$_3$

$$\text{Moles Na}^+ = \text{mol HCO}_3^- = \left(0.315 \text{ g NaHCO}_3\right)\left(\frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3}\right)\left(\frac{1 \text{ mol ion}}{1 \text{ mol NaHCO}_3}\right) = 0.00\underline{374}955 \text{ mol}$$

NaBr

$$\text{Moles Na}^+ = \text{mol Br}^- = \left(0.098 \text{ g NaBr}\right)\left(\frac{1 \text{ mol NaBr}}{102.89 \text{ g NaBr}}\right)\left(\frac{1 \text{ mol ion}}{1 \text{ mol NaBr}}\right) = 0.000\underline{952}4735 \text{ mol}$$

Total moles of each ion:
Cl$^-$: $0.\underline{453}4565 + 0.050415 + 0.0216255 + 0.0140845 = 0.\underline{539}5815$ mol Cl$^-$
Na$^+$: $0.\underline{453}4565 + 0.00374955 + 0.0009524735 = 0.458158523$ mol Na$^+$
Mg^{2+}: $0.0\underline{25}207 + 0.0278285 = 0.0530355$ mol Mg^{2+}
SO$_4^{2-}$: $0.0\underline{278}285$ mol SO$_4^{2-}$
Ca^{2+}: $0.0\underline{108}128$ mol Ca^{2+}
K$^+$: $0.0\underline{140}845$ mol K$^+$
HCO$_3^-$: $0.00\underline{374}955$ mol HCO$_3^-$
Br$^-$: $0.000\underline{952}4735$ mol Br$^-$
Dividing each of the numbers of moles by the volume (0.$\underline{961}$538 L) and rounding to the proper number of significant figures gives the molarities.
$M = $ (mol ion) / L solution
M Cl$^-$ = (0.$\underline{539}$5815 mol) / (0.$\underline{961}$538 L) = 0.$\underline{561}$165 = **0.561 M Cl$^-$**
M Na$^+$ = (0.$\underline{458}$158523 mol) / (0.$\underline{961}$538 L) = 0.$\underline{476}$485 = **0.476 M Na$^+$**
M Mg^{2+} = (0.0$\underline{530}$355 mol) / (0.$\underline{961}$538 L) = 0.0551569 = **0.0552 M Mg^{2+}**
M SO$_4^{2-}$ = (0.0$\underline{278}$285 mol) / (0.$\underline{961}$538 L) = 0.02894165 = **0.0289 M SO$_4^{2-}$**
M Ca^{2+} = (0.0$\underline{108}$128 mol) / (0.$\underline{961}$538 L) = 0.0$\underline{112}$453 = **0.0112 M Ca^{2+}**
M K$^+$ = (0.0$\underline{140}$845 mol) / (0.$\underline{961}$538 L) = 0.0$\underline{146}$479 = **0.0146 M K$^+$**
M HCO$_3^-$ = (0.00$\underline{374}$955 mol) / (0.$\underline{961}$538 L) = 0.00$\underline{389}$953 = **0.00390 M HCO$_3^-$**
M Br$^-$ = (0.000$\underline{952}$4735 mol) / (0.$\underline{961}$538 L) = 0.00099057 = **0.00099 M Br$^-$**
b) The alkali metal cations are Na$^+$ and K$^+$. Simply sum the molarities of the individual ions.
     0.$\underline{476}$485 M Na$^+$ + 0.0146479 M K$^+$ = 0.$\underline{491}$1329 = **0.491 M** total for alkali metal cations
c) The alkaline earth metal cations are Mg^{2+} and Ca^{2+}. Simply sum the molarities of the individual ions.
     0.0$\underline{551}$569 M Mg^{2+} + 0.0$\underline{112}$453 M Ca^{2+} = 0.0$\underline{664}$022 = **0.0664 M** total for alkaline-earth cations
d) The anions are Cl$^-$, SO$_4^{2-}$, HCO$_3^-$, and Br$^-$. Simply sum the molarities of the individual ions.
     0.$\underline{561}$165 M Cl$^-$ + 0.02894165 M SO$_4^{2-}$ + 0.00$\underline{389}$953 M HCO$_3^-$ + 0.00$\underline{099}$057 M Br$^-$
          = 0.$\underline{594}$99675 = **0.595 M** total for anions

4.23   The moles of the calcium ions and the iron ions are needed. The moles of each of the ions to be replaced must be multiplied by the charge to get the total moles of charge. Since sodium has a +1 charge the total moles of charge equals the moles of sodium ions.
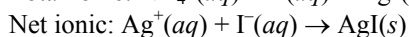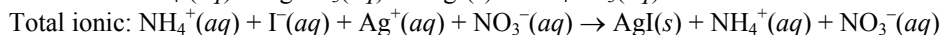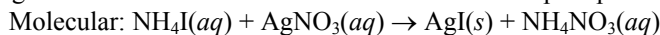
$$\left(1.0 \times 10^3 \text{ L}\right)\left(\frac{0.015 \text{ mol Ca}^{2+}}{\text{L}}\right)\left(\frac{2 \text{ mol charge}}{1 \text{ mol Ca}^{2+}}\right) = 30. \text{ mol charge from Ca}^{2+}$$

$$\left(1.0 \times 10^3 \text{ L}\right)\left(\frac{0.0010 \text{ mol Fe}^{3+}}{\text{L}}\right)\left(\frac{3 \text{ mol charge}}{1 \text{ mol Fe}^{3+}}\right) = 3.0 \text{ mol charge from Fe}^{3+}$$
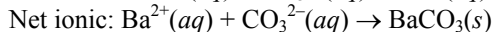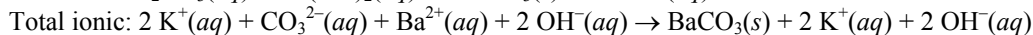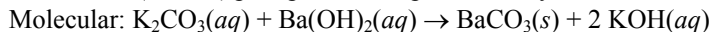
$$\text{Moles Na}^+ = \left((30. + 3.0)\text{mol charge}\right)\left(\frac{1\text{ mol Na}^+}{1\text{ mol charge}}\right) = \textbf{33 mol Na}^+$$

4.24   Ions in solution that do not participate in the reaction do not appear in a net ionic equation. These spectator ions remain as dissolved ions throughout the reaction. These ions are only present to balance charge.

4.25   The reaction given has the following total ionic and net ionic equations:
Total ionic $Ba^{2+}(aq) + 2\ NO_3^-(aq) + 2\ Na^+(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + 2\ Na^+(aq) + 2\ NO_3^-(aq)$
Net ionic $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$
The "new" equations are:
Molecular: $BaCl_2(aq) + K_2CO_3(aq) \rightarrow BaCO_3(s) + 2\ KCl(aq)$
Total ionic: $Ba^{2+}(aq) + 2\ Cl^-(aq) + 2\ K^+(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + 2\ K^+(aq) + 2\ Cl^-(aq)$
Molecular: $BaBr_2(aq) + (NH_4)_2CO_3(aq) \rightarrow BaCO_3(s) + 2\ NH_4Br(aq)$
Total ionic: $Ba^{2+}(aq) + 2\ Br^-(aq) + s\ NH_4^+(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + 2\ NH_4^+(aq) + 2\ Br^-(aq)$

4.26   If the electrostatic attraction between the ions is greater than the attraction of the ions for water molecules, the ions will form a precipitate. This is the basis for the solubility rules.

4.27   a) Since the possible products ($CaCl_2$ and KNO_3) are both soluble, no reaction would take place.
b) According to the table, KNO_3 is soluble but $PbCl_2$ is insoluble. The K^+ and NO_3^- would be spectator ions, because their salt is soluble.

4.28   Assuming that the left beaker is $AgNO_3$ (because it has gray Ag^+ ion) and the right must be NaCl, then the NO_3^- is blue, the Na^+ is brown, and the Cl^- is green.
Molecular equation: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
Total ionic equation: $Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$
Net ionic equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

4.29   <u>Plan:</u> Check to see if any of the ion pairs are not soluble according to the solubility rules.
a) Molecular: $Hg_2(NO_3)_2(aq) + 2\ KI(aq) \rightarrow Hg_2I_2(s) + 2\ KNO_3(aq)$
        Total ionic: $Hg_2^{2+}(aq) + 2\ NO_3^-(aq) + 2\ K^+(aq) + 2\ I^-(aq) \rightarrow Hg_2I_2(s) + 2\ K^+(aq) + 2\ NO_3^-(aq)$
        Net ionic: $Hg_2^{2+}(aq) + 2\ I^-(aq) \rightarrow Hg_2I_2(s)$
        Spectator ions are K^+ and NO_3^-.
b) Molecular: $FeSO_4(aq) + Ba(OH)_2(aq) \rightarrow Fe(OH)_2(s) + BaSO_4(s)$
        Total ionic: $Fe^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2\ OH^-(aq) \rightarrow Fe(OH)_2(s) + BaSO_4(s)$
        Net ionic: This is the same as the total ionic equation, because there are no spectator ions.

4.30   Check to see if any of the ion pairs are not soluble according to the solubility rules.
a) Molecular: $3\ CaCl_2(aq) + 2\ Cs_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6\ CsCl(aq)$
        Total ionic: $3\ Ca^{2+}(aq) + 6\ Cl^-(aq) + 6\ Cs^+(aq) + 2\ PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s) + 6\ Cs^+(aq) + 6\ Cl^-(aq)$
        Net ionic: $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$
        Spectator ions are Cs^+ and Cl^-.
b) Molecular: $Na_2S(aq) + ZnSO_4(aq) \rightarrow ZnS(s) + Na_2SO_4(aq)$
        Total ionic: $2\ Na^+(aq) + S^{2-}(aq) + Zn^{2+}(aq) + SO_4^{2-}(aq) \rightarrow ZnS(s) + 2\ Na^+(aq) + SO_4^{2-}(aq)$
        Net ionic: $S^{2-}(aq) + Zn^{2+}(aq) \rightarrow ZnS(s)$
        Spectator ions are Na^+ and SO_4^{2-}.

4.31   <u>Plan:</u> A precipitate forms if reactant ions can form combinations that are insoluble, as determined by the solubility rules in Table 4.1. Create cation-anion combinations other than the original reactants and determine if they are insoluble.
<u>Solution:</u>
a) No precipitate will form. The ions Na^+ and SO_4^{2-} will not form an insoluble salt according to solubility rule #1: *All common compounds of Group 1A ions are soluble*. The ions Cu^{2+} and NO_3^- will not form an insoluble salt according to the solubility rule #2: *All common nitrates are soluble*.
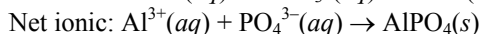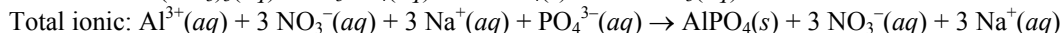
b) A precipitate will form because silver ions, Ag^+, and iodide ions, I^-, will combine to form a solid salt, silver iodide, AgI. The ammonium and nitrate ions do not form a precipitate.

        Molecular: $NH_4I(aq) + AgNO_3(aq) \rightarrow AgI(s) + NH_4NO_3(aq)$

        Total ionic: $NH_4^+(aq) + I^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgI(s) + NH_4^+(aq) + NO_3^-(aq)$

        Net ionic: $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

4.32   a) Barium carbonate ($BaCO_3$) precipitates; the potassium hydroxide is a strong base.

        Molecular: $K_2CO_3(aq) + Ba(OH)_2(aq) \rightarrow BaCO_3(s) + 2\ KOH(aq)$

        Total ionic: $2\ K^+(aq) + CO_3^{2-}(aq) + Ba^{2+}(aq) + 2\ OH^-(aq) \rightarrow BaCO_3(s) + 2\ K^+(aq) + 2\ OH^-(aq)$

        Net ionic: $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$

     b) Aluminum phosphate ($AlPO_4$) precipitates; the sodium nitrate is soluble.

        Molecular: $Al(NO_3)_3(aq) + Na_3PO_4(aq) \rightarrow AlPO_4(s) + 3\ NaNO_3(aq)$

        Total ionic: $Al^{3+}(aq) + 3\ NO_3^-(aq) + 3\ Na^+(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s) + 3\ NO_3^-(aq) + 3\ Na^+(aq)$

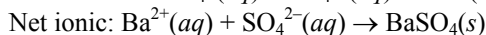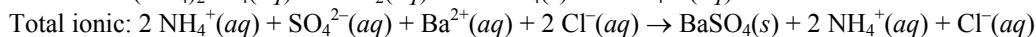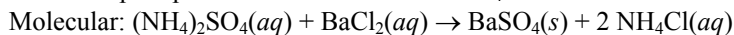   Net ionic: $Al^{3+}(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s)$

4.33   <u>Plan:</u> A precipitate forms if reactant ions can form combinations that are insoluble, as determined by the solubility rules in Table 4.1. Create cation-anion combinations other than the original reactants and determine if they are insoluble.
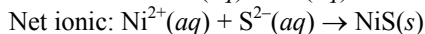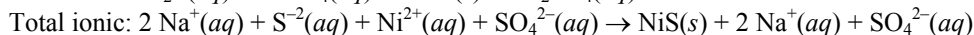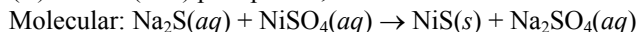
      <u>Solution:</u>

     a) New cation-anion combinations are potassium nitrate and iron(II) chloride. The rules state that all common nitrates and chlorides (with some exceptions) are soluble, so no precipitate forms.

     b) New cation-anion combinations are ammonium chloride and barium sulfate. Again, the rules state that most chlorides are soluble; however, another rule sta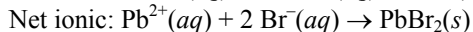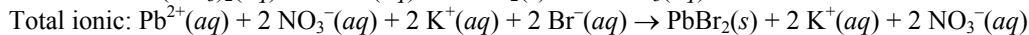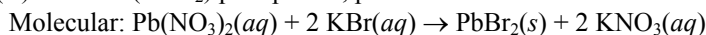tes that sulfate compounds containing barium are insoluble. Barium sulfate is a precipitate and its formula is $BaSO_4$.

        Molecular: $(NH_4)_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2\ NH_4Cl(aq)$

        Total ionic: $2\ NH_4^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2\ Cl^-(aq) \rightarrow BaSO_4(s) + 2\ NH_4^+(aq) + Cl^-(aq)$

        Net ionic: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

4.34   a) Nickel(II) sulfide (NiS) precipitates, sodium sulfate is soluble.

        Molecular: $Na_2S(aq) + NiSO_4(aq) \rightarrow NiS(s) + Na_2SO_4(aq)$

        Total ionic: $2\ Na^+(aq) + S^{-2}(aq) + Ni^{2+}(aq) + SO_4^{2-}(aq) \rightarrow NiS(s) + 2\ Na^+(aq) + SO_4^{2-}(aq)$

        Net ionic: $Ni^{2+}(aq) + S^{2-}(aq) \rightarrow NiS(s)$

     b) Lead(II) bromide ($PbBr_2$) precipitates, potassium nitrate is soluble.

        Molecular: $Pb(NO_3)_2(aq) + 2\ KBr(aq) \rightarrow PbBr_2(s) + 2\ KNO_3(aq)$

        Total ionic: $Pb^{2+}(aq) + 2\ NO_3^-(aq) + 2\ K^+(aq) + 2\ Br^-(aq) \rightarrow PbBr_2(s) + 2\ K^+(aq) + 2\ NO_3^-(aq)$

        Net ionic: $Pb^{2+}(aq) + 2\ Br^-(aq) \rightarrow PbBr_2(s)$

4.35   <u>Plan:</u> Write a balanced equation for the chemical reaction described in the problem. By applying the solubility rules to the two possible products ($NaNO_3$ and PbI_2), determine that PbI_2 is the precipitate. By using molar relationships, determine how many moles of $Pb(NO_3)_2$ (and thus Pb^{2+} ion) are required to produce 0.628 g of PbI_2.

      <u>Solution:</u>

     The reaction is: $Pb(NO_3)_2(aq) + 2\ NaI(aq) \rightarrow PbI_2(s) + 2\ NaNO_3(aq)$.

$$M\ Pb^{2+} = \left(0.628\text{ g PbI}_2\right)\left(\frac{1\text{ mol PbI}_2}{461.0\text{ g PbI}_2}\right)\left(\frac{1\text{ mol Pb}^{2+}}{1\text{ mol PbI}_2}\right)\left(\frac{1}{35.0\text{ mL}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

$$= 0.0389216 = \mathbf{0.0389}\ \boldsymbol{M}\ \mathbf{Pb^{2+}}$$

4.36   The reaction is $AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$.

$$M\ Ag^+ = \left(0.842\text{ g AgCl}\right)\left(\frac{1\text{ mol AgCl}}{143.4\text{ g AgCl}}\right)\left(\frac{1\text{ mol Ag}^+}{1\text{ mol AgCl}}\right)\left(\frac{1}{25.0\text{ mL}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

$$= 0.2348675 = \mathbf{0.235}\ \boldsymbol{M}\ \mathbf{Ag^+}$$

4.37   <u>Plan:</u> The balanced equation for this reaction is $AgNO_3(aq) + Cl^-(aq) \rightarrow AgCl(s) + NO_3^-(aq)$.
First, determine the moles of Cl^- present in the 25.00 mL sample. Second, convert moles of Cl^- into grams, and convert the sample volume into grams using the given density. The mass percent of Cl^- is found by dividing one result by the other.
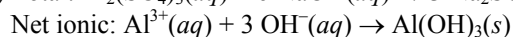<u>Solution:</u>

$$(43.63 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.3020 \text{ mol AgNO}_3}{L}\right)\left(\frac{1 \text{ mol Cl}^-}{1 \text{ mol AgNO}_3}\right)\left(\frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}^-}\right) = 0.467098 \text{ g Cl (unrounded)}$$

$$\text{Mass of sample} = (25.00 \text{ mL})\left(\frac{1.04 \text{ g}}{\text{mL}}\right) = 26.0 \text{ g sample}$$

$$\text{Mass \% Cl} = \frac{\text{Mass Cl}}{\text{Mass Sample}} \times 100\% = \frac{0.467098 \text{ g Cl}}{26.0 \text{ g Sample}} \times 100\% = 1.79653 = \textbf{1.80\% Cl}$$

4.38   a) Total: $Al_2(SO_4)_3(aq) + 6 \text{ NaOH}(aq) \rightarrow 3 \text{ Na}_2SO_4(aq) + 2 \text{ Al(OH)}_3(s)$
       Net ionic: $Al^{3+}(aq) + 3 \text{ OH}^-(aq) \rightarrow Al(OH)_3(s)$
    b) Determine the limiting reactant:

$$\text{Moles Al}_2(SO_4)_3 = (627 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{15.8 \text{ g Al}_2(SO_4)_3}{L}\right)\left(\frac{1 \text{ mol Al}_2(SO_4)_3}{342.17 \text{ g Al}_2(SO_4)_3}\right)$$

$$= 0.028952 \text{ mol Al}_2(SO_4)_3 \text{ (unrounded)}$$
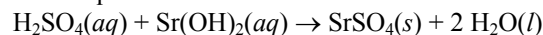
$$\text{Moles NaOH} = (185.5 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.533 \text{ mol NaOH}}{L}\right) = 0.0988715 \text{ mol NaOH (unrounded)}$$
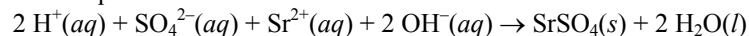
       NaOH is the limiting reactant.
       Finishing with limiting reactant

$$\text{Grams Al(OH)}_3 = (0.0988715 \text{ mol NaOH})\left(\frac{2 \text{ mol Al(OH)}_3}{6 \text{ mol NaOH}}\right)\left(\frac{78.00 \text{ g Al(OH)}_3}{1 \text{ mol Al(OH)}_3}\right) = 2.570659 = \textbf{2.57 g Al(OH)}_3$$

4.39   The molecular equation is:
          $H_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow SrSO_4(s) + 2 H_2O(l)$
    The total ionic equation is:
          $2 H^+(aq) + SO_4^{2-}(aq) + Sr^{2+}(aq) + 2 OH^-(aq) \rightarrow SrSO_4(s) + 2 H_2O(l)$
    Since there are no spectator ions, the total and net ionic equations are the same.

4.40 $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

4.41   a) Any three of HCl, HBr, HI, HNO_3, H_2SO_4, or $HClO_4$
    b) Any three of NaOH, KOH, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$
    c) They dissociate 100% into ions in aqueous solution.

4.42   a) There are many possibilities including: acetic acid, chlorous acid, nitrous acid, ...
    b) NH_3
    c) Strong acids or bases dissociate 100%; weak acids or bases dissociate much less than this (typically less than 10%) in aqueous solution. The electrical conductivity of a solution of a strong acid or base would be much higher than that of a weak acid or base of equal concentration.

4.43   a) The formation of a gas, $SO_2(g)$, and formation of water drive this reaction to completion, because both products remove reactants from solution.
    b) The formation of a precipitate, $Ba_3(PO_4)_2(s)$, will cause this reaction to go to completion. This reaction is one between an acid and a base, so the formation of water molecules through the combination of H^+ and OH^- ions also drives the reaction.

4.44   Molecular: $HC_2H_3O_2(aq) + NaOH(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l)$
Net ionic: $HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$
Molecular: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
Net ionic: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
The difference is because CH_3COOH is a weak acid and HCl is a strong acid.

4.45   a) Molecular equation: $KOH(aq) + HI(aq) \rightarrow KI(aq) + H_2O(l)$
Total ionic equation: $K^+(aq) + OH^-(aq) + H^+(aq) + I^-(aq) \rightarrow K^+(aq) + I^-(aq) + H_2O(l)$
Net ionic equation: $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
The spectator ions are $K^+(aq)$ and $I^-(aq)$
b) Molecular equation: $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$
Total ionic equation: $NH_3(aq) + H^+(aq) + Cl^-(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$
NH_3 is a weak base and is written in the molecular form. HCl is a strong acid and is written in the dissociated
form. NH_4Cl is a soluble compound, because all ammonium compounds are soluble.
Net ionic equation: $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$
Cl^- is the only spectator ion.

4.46   a) Molecular: $CsOH(aq) + HNO_3(aq) \rightarrow CsNO_3(aq) + H_2O(l)$
    Total ionic: $Cs^+(aq) + OH^-(aq) + H^+(aq) + NO_3^-(aq) \rightarrow Cs^+(aq) + NO_3^-(aq) + H_2O(l)$
    Net ionic: $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
    Spectator ions are Cs^+ and NO_3^-.
b) Molecular: $Ca(OH)_2(aq) + 2\,HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + 2\,H_2O(l)$
    Total ionic: $Ca^{2+}(aq) + 2\,OH^-(aq) + 2\,HC_2H_3O_2\,(aq) \rightarrow Ca^{2+}(aq) + 2\,C_2H_3O_2^-(aq) + 2\,H_2O(l)$
    Net ionic: $OH^-(aq) + HC_2H_3O_2(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$
    Spectator ion is Ca^{2+}.

4.47   Calcium carbonate dissolves in $HCl(aq)$ because the carbonate ion, a base, reacts with the acid to form $CO_2(g)$.
Total ionic equation:
    $CaCO_3(s) + H^+(aq) + 2\,Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2\,Cl^-(aq) + H_2O(l) + CO_2(g)$
Net ionic equation:
    $CaCO_3(s) + H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

4.48   The HNO_3 reacts with the bound OH^-, producing H_2O:
Total ionic equation:
    $Zn(OH)_2(s) + 2\,H^+(aq) + 2\,NO_3^-(aq) \rightarrow Zn^{2+}(aq) + 2\,NO_3^-(aq) + 2\,H_2O(l)$
Net ionic equation:
    $Zn(OH)_2(s) + 2\,H^+(aq) \rightarrow Zn^{2+}(aq) + 2\,H_2O(l)$

4.49   The reaction is: $KOH(aq) + CH_3COOH(aq) \rightarrow KCH_3COO(aq) + H_2O(l)$
$$M = \left(\frac{0.1180 \text{ mol KOH}}{L}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)(25.98 \text{ mL})\left(\frac{1 \text{mol CH}_3\text{COOH}}{1 \text{ mol KOH}}\right)\left(\frac{1}{52.50 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$
    $= 0.05839314 = \textbf{0.05839 } \textbf{\textit{M}} \textbf{ CH}_3\textbf{COOH}$

4.50   The reaction is: $2\,NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\,H_2O(l)$
$$M = \left(\frac{0.1750 \text{ mol NaOH}}{L}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)(36.25 \text{ mL})\left(\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}\right)\left(\frac{1}{25.00 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$
    $= 0.126875 = \textbf{0.1269 } \textbf{\textit{M}} \textbf{ H}_2\textbf{SO}_4$

4.51   The reaction is: 2 NaHCO$_3$(aq) + H$_2$SO$_4$(aq) → Na$_2$SO$_4$(aq) + 2 H$_2$O(l) + 2 CO$_2$(g)

$$V = \left(\frac{2.6 \text{ mol H}_2\text{SO}_4}{L}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)(78 \text{ mL})\left(\frac{2 \text{ mol NaHCO}_3}{1 \text{ mol H}_2\text{SO}_4}\right)\left(\frac{1 \text{ L}}{1.5 \text{ mol NaHCO}_3}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

     = 270.4 = **2.7 x 10^2 mL NaHCO$_3$**

4.52   The reaction is: NaOH(aq) + HCl(aq) → NaCl(aq) + H$_2$O(l)

$$M = \left(\frac{0.1528 \text{ mol HCl}}{L}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)(25.00 \text{ mL})\left(\frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}}\right)\left(\frac{1}{(39.21 - 2.24)\text{mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

     = 0.103327 = **0.1033 M NaOH**

4.53   The reaction is: UO$_2$(s) + 4 HF(aq) → UF$_4$(s) + 2 H$_2$O(l)

$$V = (2.25 \text{ kg UO}_2)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol UO}_2}{270.0 \text{ g UO}_2}\right)\left(\frac{4 \text{ mol HF}}{1 \text{ mol UO}_2}\right)\left(\frac{1 \text{ L}}{2.50 \text{ mol HF}}\right) = 13.3333 = \textbf{13.3 L HF}$$

4.54   a) Since H is bonded to a nonmetal, O.N. (H) = +1. Then O.N. (S) = **-2**.
   b) Assume O.N. (O) = -2. Then O.N. (S) = **+4**.

4.55   **No**, O.N. (N) = -3, and O.N. (H) = +1, O.N. (Cl) = -1 throughout. Since no elements changes oxidation number, this cannot be a redox reaction.

4.56   An oxidizing agent causes something else to be oxidized; i.e., to lose electrons. The oxidizing agent accepts these electrons and is reduced.

4.57   Electrons cannot appear and disappear as if by magic; they must come from somewhere and go to somewhere. Therefore, something must be oxidized and something else be reduced.

4.58   a) The S in SO$_4^{2-}$ (i.e., H$_2$SO$_4$) has O.N. = +6, and in SO$_2$, O.N. (s) = +4, so the S has been reduced (and the I$^-$ oxidized), so the H$_2$SO$_4$ is an oxidizing agent.
   b) The oxidation numbers remain constant throughout; H$_2$SO$_4$ transfers a proton to F$^-$ to produce HF, so it acts as an acid.

4.59   **Oxidizing agent = NO$_2$; Reducing agent = NH$_3$;** the NO$_2$ is the oxidizing agent (O.N. (N) goes from +4 in NO$_2$ to 0 in N$_2$), and the NH$_3$ is the reducing agent (O.N.(N) goes from -3 to 0).

4.60   Refer to the rules in Table 4.3. Remember that oxidation number (O.N.) is not the same thing as ionic charge. The O.N. is the charge an atom would have **if** electrons were transferred completely. The compounds in this problem contain covalent bonds, or a combination of ionic and covalent bonds (Na$_2$C$_2$O$_4$). Covalent bonds do not involve the complete transfer of electrons.
   a) CF$_2$Cl$_2$. The rules dictate that F and Cl each have an O.N. of -1; two F and two Cl yield a sum of -4, so C must have a +4 O.N. **C = +4**
   b) Na$_2$C$_2$O$_4$. The rule dictates that Na has a +1 O.N.; rule 5 dictates that O has a -2 O.N.; two Na and four O yield a sum of -6 [(+2) + (-8)]. Therefore, the total of the O.N.'s on the two C atoms is +6 and each C is +3. **C = +3**
   c) HCO$_3^-$. H is combined with nonmetals and has an O.N. of +1; O has an O.N. of -2. To have an overall oxidation state equal to -1, C must be +4 because (+1) + (+4) + (-6) = -1. **C = +4**
   d) C$_2$H$_6$. Each H has an O.N. of +1; six H gives +6. The sum of O.N.'s for the two C atoms must be -6, so each C is -3. **C = -3**

4.61   a) -1 b) +3 c) +5 d) -1

4.62 a) NH_2OH: (O.N. for N) + 3(+1 for H) + 1(-2 for O) = 0   O.N. for N = **-1**
      b) N_2H_4:   2(O.N. for N) + 4(+1 for H) = 0   O.N. for N = **-2**
      c) NH_4^+:   (O.N. for N) + 4(+1 for H) = +1   O.N. for N = **-3**
      d) HNO_2:   (O.N. for N) + 1(+1 for H) + 2(-2 for O) = 0 O.N. for N = **+3**

4.63 a) +4   b) -1   c) +4   d) -2

4.64 a) AsH_3. H is combined with a nonmetal, so its O.N. is +1 (Rule 3). The O.N. for As is -3.
      b) H_3AsO_4. The H's in this formula are acidic hydrogens. When an acid dissociates in water, it forms H_3O^+ or H^+
      ion, therefore, the O.N. of H in this compound is +1, or +3 for 3 H's. Oxygen's O.N. is -2, with total O.N. of -8 (4
      times -2), so As needs to have an O.N. of +5.
      c) $AsCl_3$. Cl has an O.N. of -1, total of -3, so As must have an O.N. of +3.
      a) **As = -3**   b) **As = +5**   c) **As = +3**

4.65 a) +5   b) -3   c) +5

4.66 a) MnO_4^{2-}:   (O.N. for Mn) + 4(-2 for O) = -2   O.N. for Mn = **+6**
      b) Mn_2O_3:   {2(O.N. for Mn)} + 3(-2 for O) = 0   O.N. for Mn = **+3**
      c) $KMnO_4$:   1(+1 for K) + (O.N. for Mn) + 4(-2 for O) = 0   O.N. for Mn = **+7**

4.67 a) +6   b) +6   c) +3

4.68 Oxidizing agent: substance that accepts the electrons released by the substance that is oxidized.
      Reducing agent: substance that provides the electrons accepted by the substance that is reduced.
      First, assign oxidation numbers to all atoms. Second, recognize that the agent is the compound that contains the
      atom that is gaining or losing electrons, not just the atom itself.
      a) 5 $H_2C_2O_4(aq)$ + 2 $MnO_4^-(aq)$ + 6 $H^+(aq)$ → 2 $Mn^{2+}(aq)$ + 10 $CO_2(g)$ + 8 $H_2O(l)$
      H = +1   Mn = +7   H = +1   Mn = +2   C = +4   H = +1
      C = +3   O = -2   O = -2   O = -2
      O = -2
      Hydrogen and oxygen do not change oxidation state. The Mn changes from +7 to +2 (reduction). Therefore,
      MnO_4^- is the oxidizing agent. C changes from +3 to +4 (oxidation), so **$H_2C_2O_4$ is the reducing agent**.
      b) 3 Cu(s) + 8 $H^+(aq)$ + 2 $NO_3^-(aq)$ → 3 $Cu^{2+}(aq)$ + 2 NO(g) + 4 $H_2O(l)$
      Cu = 0   H = +1   N = +5   Cu = +2   N = +2   H = +1
      O = -2   O = -2   O = -2
      Cu changes from 0 to +2 (is oxidized) and **Cu is the reducing agent**. N changes from +5 (in NO_3^-) to +2 (in NO)
      and is reduced, so **NO_3^- is the oxidizing agent**.

4.69 a) Oxidizing agent = **H^+**   Reducing agent = **Sn**
      b) Oxidizing agent = **H_2O_2**   Reducing agent = **Fe^{2+}**

4.70 a) **Oxidizing agent is NO_3^-** because nitrogen changes from +5 O.N. in NO_3^- to +4 O.N. in NO_2. **Reducing agent
      is Sn** because its O.N. changes from 0 as the element to +4 in $SnCl_6^{2-}$.
      b) **Oxidizing agent is MnO_4^-** because manganese changes from +7 O.N. in MnO_4^- to +2 O.N. in Mn^{2+}. **Reducing
      agent is Cl^-** because its O.N. changes from -1 in Cl^- to 0 as the element to Cl_2.

4.71 a) Oxidizing agent = **$Cr_2O_7^{2-}$**   Reducing agent = **SO_3^{2-}**
      b) Oxidizing agent = **NO_3^-**   Reducing agent = **Zn**

4.72 S is in Group 6A(16), so its highest possible O.N. is +6 and its lowest possible O.N. is 6 - 8 = -2.
      a) The lowest O.N. for S [(Group 6A(16)] is 6 - 8 = -2, which occurs in S^{2-}. Therefore, when S^{2-} reacts in an
      oxidation-reduction reaction, S can only increase its O.N. (oxidize), so S^{2-} can only function as a reducing agent.
      b) The highest O.N. for S is +6, which occurs in SO_4^{2-}. Therefore, when SO_4^{2-} reacts in an oxidation-reduction
      reaction, the S can only decrease its O.N. (reduce), so SO_4^{2-} can only function as an oxidizing agent.
      c) The O.N. of S in SO_2 is +4, so it can increase or decrease its O.N. Therefore, SO_2 can function as either an
      oxidizing or reducing agent.

4.73 N is Group 5A (15), so its highest possible O.N. is +5 and its lowest possible O.N. is 5 - 8 = -3.
   a) In N^{3-}, O.N. (N) = -3, so it can only be oxidized (i.e., act as a reducing agent).
   b) In NO_3^-, O.N. (N) = +5, so it can only be reduced (i.e., act as an oxidizing agent).
   c) In NO_2^-, O.N. (N) = +3, so it can be either oxidized or reduced (i.e., can act as a reducing or oxidizing agent).

4.74 a) <u>Step: 1</u>: Assign oxidation numbers to all elements.
   <u>Step: 2</u>: Identify oxidized and reduced species.
   a) The O.N. of Cr decreases from +6 in K_2CrO_4 to +3 in $Cr(NO_3)_3$, so chromium is gaining electrons and is reduced. The O.N. of Fe increases from +2 in $Fe(NO_3)_2$ to +3 in $Fe(NO_3)_3$, so iron is losing electrons and is oxidized.
   <u>Step: 3</u>: Compute number of electrons lost in oxidation and gained in reduction.
   <u>Step: 4</u>: Multiply by factors to make e^- lost equal e^- gained.
   Cr gained 3 electrons so Fe should be multiplied by 3 to lose 3 electrons
   <u>Step: 5</u>: Finish balancing by inspection.
   8 $HNO_3(aq)$ + $K_2CrO_4(aq)$ + 3 $Fe(NO_3)_2(aq)$ → 2 $KNO_3(aq)$ + 3 $Fe(NO_3)_3(aq)$ + $Cr(NO_3)_3(aq)$ + 4 $H_2O(l)$
   Oxidizing agent: K_2CrO_4 Reducing agent: $Fe(NO_3)_2$
   b) 8 $HNO_3(aq)$ + 3 $C_2H_6O(l)$ + $K_2Cr_2O_7(aq)$ → 2 $KNO_3(aq)$ + 3 $C_2H_4O(l)$ + 7 $H_2O(l)$ + 2 $Cr(NO_3)_3(aq)$
   Oxidizing agent: $K_2Cr_2O_7$ Reducing agent: C_2H_6O
   c) 6 $HCl(aq)$ + 2 $NH_4Cl(aq)$ + $K_2Cr_2O_7(aq)$ → 2 $KCl(aq)$ + 2 $CrCl_3(aq)$ + $N_2(g)$ + 7 $H_2O(l)$
   Oxidizing agent: $K_2Cr_2O_7$ Reducing agent: NH_4Cl
   d) $KClO_3(aq)$ + 6 $HBr(aq)$ → 3 $Br_2(l)$ + 3 $H_2O(l)$ + $KCl(aq)$
   Oxidizing agent: $KClO_3$ Reducing agent: HBr

4.75 a) 2 $HCl(aq)$ + 2 $FeCl_2(aq)$ + $H_2O_2(aq)$ → 2 $FeCl_3(aq)$ + 2 $H_2O(l)$
   H_2O_2 is the oxidizing agent (O.N.(O) goes from -1 to -2) and **$FeCl_2$ is the reducing agent** (O.N.(Fe) goes from +2 to +3).
   b) $I_2(s)$ + 2 $Na_2S_2O_3(aq)$ → $Na_2S_4O_6(aq)$ + 2 $NaI(aq)$
   I_2 is the oxidizing agent (O.N.(I) goes from 0 to -1) and **$Na_2S_2O_3$ is the reducing agent** (O.N.(S) goes from +2 to +2.5).
   c) 8 $HNO_3(aq)$ + 6 $KI(aq)$ → 2 $NO(g)$ + 3 $I_2(s)$ + 4 $H_2O(l)$ + 6 $KNO_3(aq)$
   HNO_3 is the oxidizing agent (O.N.(N) goes from +5 to +2) and **KI is the reducing agent** (O.N.(I) goes from -1 to 0).
   d) 3 $PbO(s)$ + 2 $NH_3(aq)$ → $N_2(g)$ + 3 $H_2O(l)$ + 3 $Pb(s)$
   PbO is the oxidizing agent (O.N.(Pb) goes from +2 to 0) and **NH_3 is the reducing agent** (O.N.(N) goes from -3 to 0).

4.76 <u>Plan:</u> Determine the moles of permanganate ion present. The moles of permanganate ion along with the balanced chemical equation give the moles of hydrogen peroxide. The molar mass of hydrogen peroxide is required to finish the problem. The unrounded numbers from each step are carried to the next step to minimize rounding errors.
   <u>Solution:</u>

   a) Moles $MnO_4^- = \left(\dfrac{0.105 \text{ mol } MnO_4^-}{L} \right)\left(\dfrac{10^{-3} \text{ L}}{1 \text{ mL}} \right)(43.2 \text{ mL}) = 4.536 \times 10^{-3} = $ **4.54 x 10^{-3} mol MnO_4^-**

   b) Moles $H_2O_2 = \left(4.536 \times 10^{-3} \text{ } MnO_4^- \right)\left(\dfrac{5 \text{ mol } H_2O_2}{2 \text{ mol } MnO_4^-} \right) = 0.01134 = $ **0.0113 mol H_2O_2**

   c) Mass $H_2O_2 = \left(0.01134 \text{ mol } H_2O_2 \right)\left(\dfrac{34.02 \text{ g } H_2O_2}{1 \text{ mol } H_2O_2} \right) = 0.3857868 = $ **0.386 g H_2O_2**

   d) Mass percent $H_2O_2 = \left(\dfrac{0.3857868 \text{ g } H_2O_2}{13.8 \text{ g Sample}} \right) \times 100\% = 2.79556 = $ **2.80% H_2O_2**

   e) **H_2O_2** (The hydrogen peroxide reduces the manganese from +7 to +2.)

4.77   Mass $C_2H_5OH = \left(\dfrac{0.05961\ mol\ Cr_2O_7^{2-}}{1\ L}\right)\left(\dfrac{10^{-3}\ L}{1\ mL}\right)(35.46\ mL)\left(\dfrac{1\ mol\ C_2H_5OH}{2\ mol\ Cr_2O_7^{2-}}\right)\left(\dfrac{46.07\ g\ C_2H_5OH}{1\ mol\ C_2H_5OH}\right)$

        = 0.0486907 g C_2H_5OH (unrounded)

        Mass percent $C_2H_5OH = \left(\dfrac{0.0486907\ g\ C_2H_5OH}{28.00\ g\ Sample}\right)$ x 100% = 0.173895 = **0.1739% C$_2$H$_5$OH**

4.78   a) **decomposition**   b) **combination**   c) **displacement**

4.79   By definition, elements cannot decompose into anything simpler, so they could not be reactants in a decomposition reaction.

4.80   **Combination** and **displacement**; possibly **decomposition**

4.81   A common example of a combination/redox reaction is the combination of a metal and nonmetal to form an ionic salt, such as 2 Mg(s) + O$_2$(g) → 2 MgO(s). A common example of a combination/non-redox reaction is the combination of a metal oxide and water to form an acid, such as: CaO(s) + H$_2$O(l) → Ca(OH)$_2$(aq).

4.82   **Yes**, since the oxidation number of O$_2$ will change from zero (in the element) to some negative value (typically -2) in the product(s) during the reaction.

4.83   a) Ca(s) + 2 H$_2$O(l) → Ca(OH)$_2$(aq) + H$_2$(g)
        Displacement: one Ca atom displaces 2 H atoms.
    b) 2 NaNO$_3$(s) → 2 NaNO$_2$(s) + O$_2$(g)
        Decomposition: one reactant breaks into two products.
    c) C$_2$H$_2$(g) + 2 H$_2$(g) → C$_2$H$_6$(g)
        Combination: reactants combine to form one product.

4.84   a) 2 HI(g) → H$_2$(g) + I$_2$(g)   **decomposition**
    b) Zn(s) + 2 AgNO$_3$(aq) → Zn(NO$_3$)$_2$(aq) + 2 Ag(s)   **displacement**
    c) 2 NO(g) + O$_2$(g) → N$_2$O$_4$(l)   **combination**

4.85   Recall the definitions of each type of reaction:
    Combination: X + Y → Z;   decomposition: Z → X + Y
    Single displacement: X + YZ → XZ + Y   double displacement: WX + YZ → WZ + YX
    a) 2 Sb(s) + 3 Cl$_2$(g) → 2 SbCl$_3$(s)   **combination**
    b) 2 AsH$_3$(g) → 2 As(s) + 3 H$_2$(g)   **decomposition**
    c) 3 Mn(s) + 2 Fe(NO$_3$)$_3$(aq) → 3 Mn(NO$_3$)$_2$(aq) + 2 Fe(s)   **displacement**

4.86   a) Mg(s) + 2 H$_2$O(g) → Mg(OH)$_2$(s) + H$_2$(g)   **displacement**
    b) Cr(NO$_3$)$_3$(aq) + Al(s) → Al(NO$_3$)$_3$(aq) + Cr(s)   **displacement**
    c) PF$_3$(g) + F$_2$(g) → PF$_5$(g)   **combination**

4.87   a) The combination between a metal and a nonmetal gives a binary ionic compound.
        Ca(s) + Br$_2$(l) → CaBr$_2$(s)
    b) Many metal oxides release oxygen gas upon thermal decomposition.

        2 Ag$_2$O(s) $\xrightarrow{\Delta}$ 4 Ag(s) + O$_2$(g)
    c) This is a single displacement reaction. Mn is a more reactive metal and displaces Cu^{2+} from solution.
        Mn(s) + Cu(NO$_3$)$_2$(aq) → Mn(NO$_3$)$_2$(aq) + Cu(s)

4.88   a) $Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

b) $2\ LiCl(l) \xrightarrow{\text{elect}} 2\ Li(l) + Cl_2(g)$

c) $SnCl_2(aq) + Co(s) \rightarrow CoCl_2(aq) + Sn(s)$

4.89   a) $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$

b) $2\ NaClO_3(s) \xrightarrow{\Delta} 2\ NaCl(s) + 3\ O_2(g)$

c) $Ba(s) + 2\ H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$

4.90   a) $Fe(s) + 2\ HClO_4(aq) \rightarrow Fe(ClO_4)_2(aq) + H_2(g)$

b) $S_8(s) + 8\ O_2(g) \xrightarrow{\Delta} 8\ SO_2(g)$

c) $BaCl_2(l) \xrightarrow{\text{elect}} Ba(s) + Cl_2(g)$

4.91   a) Cs, a metal, and I_2, a nonmetal, react to form the binary ionic compound, CsI.
$$2\ Cs(s) + I_2(s) \rightarrow 2\ CsI(s)$$
b) Al is a stronger reducing agent than Mn and is able to displace Mn from solution, i.e., cause the reduction from $Mn^{2+}(aq)$ to $Mn^0(s)$.
$$2\ Al(s) + 3\ MnSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3\ Mn(s)$$
c) Sulfur dioxide, SO_2, is a nonmetal oxide that reacts with oxygen, O_2, to form the higher oxide, SO_3.
$$2\ SO_2(g) + O_2(g) \xrightarrow{\Delta} 2\ SO_3(g)$$
It is not clear from the problem, but energy must be added to force this reaction to proceed.
d) Propane is a three-carbon hydrocarbon with the formula C_3H_8. It burns in the presence of oxygen, O_2, to form carbon dioxide gas and water vapor. Although this is a redox reaction that could be balanced using the oxidation number method, it is easier to balance by considering only atoms on either side of the equation. First, balance carbon and hydrogen (because they only appear in one species on each side of the equation), and then balance oxygen.
$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$
e) Total ionic equation:
$$2\ Al(s) + 3\ Mn^{2+}(aq) + 3\ SO_4{}^{2-}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ SO_4{}^{2-}(aq) + 3\ Mn(s)$$
Net ionic equation:
$$2\ Al(s) + 3\ Mn^{2+}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ Mn(s)$$
Note that the molar coefficients are not simplified because the number of electrons lost (6 e⁻) must equal the electrons gained (6 e⁻).

4.92   a) $C_5H_{12}(l) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(g)$
b) $PCl_3(l) + Cl_2(g) \rightarrow PCl_5(s)$
c) $Zn(s) + 2\ HBr(aq) \rightarrow ZnBr_2(aq) + H_2(g)$
d) $2\ KI(aq) + Br_2(l) \rightarrow 2\ KBr(aq) + I_2(s)$
e) $2\ I^-(aq) + Br_2(l) \rightarrow I_2(s) + 2\ Br^-(aq)$

4.93   The balanced chemical equation is $2\ HgO(s) \xrightarrow{\Delta} 2\ Hg(l) + O_2(g)$

$$\text{Mass } O_2 = \left(4.27\ \text{kg HgO}\right)\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol HgO}}{216.6\ \text{g HgO}}\right)\left(\frac{1\ \text{mol } O_2}{2\ \text{mol HgO}}\right)\left(\frac{32.00\ \text{g } O_2}{1\ \text{mol } O_2}\right) = 315.420 = \mathbf{315\ g\ O_2}$$

The other product is **mercury**.

$$\text{Mass Hg} = \left(4.27\ \text{kg HgO}\right)\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol HgO}}{216.6\ \text{g HgO}}\right)\left(\frac{2\ \text{mol Hg}}{2\ \text{mol HgO}}\right)\left(\frac{200.6\ \text{g Hg}}{1\ \text{mol Hg}}\right)\left(\frac{1\ \text{kg}}{10^3\ \text{g}}\right)$$

$$= 3.95458 = \mathbf{3.95\ kg\ Hg}$$

4.94   The balanced chemical equation is $CaCl_2(l) \xrightarrow{\text{elect}} Ca(s) + Cl_2(g)$

Note, the reaction cannot be done in the presence of water as elemental calcium would displace the hydrogen from the water.

$$\text{Mass } Cl_2 = \left(874 \text{ g } CaCl_2\right)\left(\frac{1 \text{ mol } CaCl_2}{110.98 \text{ g } CaCl_2}\right)\left(\frac{1 \text{ mol } Cl_2}{1 \text{ mol } CaCl_2}\right)\left(\frac{70.90 \text{ g } Cl_2}{1 \text{ mol } Cl_2}\right) = 558.358 = \textbf{558 g } Cl_2$$

The other product is **calcium**.

$$\text{Mass } Ca = \left(874 \text{ g } CaCl_2\right)\left(\frac{1 \text{ mol } CaCl_2}{110.98 \text{ g } CaCl_2}\right)\left(\frac{1 \text{ mol } Ca}{1 \text{ mol } CaCl_2}\right)\left(\frac{40.08 \text{ g } Ca}{1 \text{ mol } Ca}\right) = 315.64 = \textbf{316 g } Ca$$

4.95   To determine the reactant in excess, write the balanced equation (metal + O_2 → metal oxide), convert reactant masses to moles, and use molar ratios to see which reactant makes the smaller ("limiting") amount of product.

$4 Li(s) + O_2(g) \rightarrow 2 Li_2O(s)$
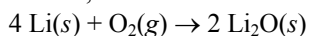
a) Moles Li_2O if Li limiting $= \left(1.62 \text{ g Li}\right)\left(\frac{1 \text{ mol Li}}{6.941 \text{ g Li}}\right)\left(\frac{2 \text{ mol } Li_2O}{4 \text{ mol Li}}\right) = 0.1166979 \text{ mol } Li_2O$ (unrounded)

Moles Li_2O if O_2 limiting $= \left(6.00 \text{ g } O_2\right)\left(\frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2}\right)\left(\frac{2 \text{ mol } Li_2O}{1 \text{ mol } O_2}\right) = 0.375 \text{ mol } Li_2O$ (unrounded)

Li is the limiting reactant; O_2 **is in excess**.
b) $0.1166979 = \textbf{0.117 mol } Li_2O$
c) Li is limiting, thus there will be none remaining (**0 g Li**).

$$\text{Grams } Li_2O = \left(1.62 \text{ g Li}\right)\left(\frac{1 \text{ mol Li}}{6.941 \text{ g Li}}\right)\left(\frac{2 \text{ mol } Li_2O}{4 \text{ mol Li}}\right)\left(\frac{29.88 \text{ g } Li_2O}{1 \text{ mol } Li_2O}\right) = 3.4869 = \textbf{3.49 g } Li_2O$$

$$\text{Grams } O_2 \text{ used} = \left(1.62 \text{ g Li}\right)\left(\frac{1 \text{ mol Li}}{6.941 \text{ g Li}}\right)\left(\frac{1 \text{ mol } O_2}{4 \text{ mol Li}}\right)\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}\right) = 1.867166 \text{ g } O_2 \text{ (unrounded)}$$

The beginning part of these calculations is repeated from the limiting reactant calculation 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.
Remaining O_2 = 6.00 g O_2 - 1.867166 g O_2 = 4.13283 = **4.13 g O_2**

4.96 $3 Mg(s) + N_2(g) \xrightarrow{\Delta} Mg_3N_2(s)$

a) Mole Mg_3N_2 (from Mg) $= \left(2.22 \text{ g Mg}\right)\left(\frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}}\right)\left(\frac{1 \text{ mol } Mg_3N_2}{3 \text{ mol Mg}}\right) = 0.030440 = 0.0304 \text{ mol } Mg_3N_2$

Mole Mg_3N_2 (from N_2) $= \left(3.75 \text{ g } N_2\right)\left(\frac{1 \text{ mol } N_2}{28.02 \text{ g } N_2}\right)\left(\frac{1 \text{ mol } Mg_3N_2}{1 \text{ mol } N_2}\right) = 0.13383 = 0.134 \text{ mol } Mg_3N_2$

Mg is the limiting reactant; N_2 **is present in excess**.
b) **0.0304 mol Mg_3N_2**

c) Mass $Mg_3N_2 = \left(2.22 \text{ g Mg}\right)\left(\frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}}\right)\left(\frac{1 \text{ mol } Mg_3N_2}{3 \text{ mol Mg}}\right)\left(\frac{100.95 \text{ g } Mg_3N_2}{1 \text{ mol } Mg_3N_2}\right)$

$= 3.07293 = \textbf{3.07 g } Mg_3N_2$
There will be **no Mg** remaining.

Mass $N_2 = 3.75 \text{ g } N_2 - \left(2.22 \text{ g Mg}\right)\left(\frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}}\right)\left(\frac{1 \text{ mol } N_2}{3 \text{ mol Mg}}\right)\left(\frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2}\right)$

$= 2.897067 = \textbf{2.90 g } N_2$ present in excess.

4.97 $2 KClO_3(s) \xrightarrow{\Delta} 2 KCl(s) + 3 O_2(g)$

Mass $KClO_3 = \left((0.900 - 0.700)g\ O_2\right)\left(\dfrac{1\ mol\ O_2}{32.00\ g\ O_2}\right)\left(\dfrac{2\ mol\ KClO_3}{3\ mol\ O_2}\right)\left(\dfrac{122.55\ g\ KClO_3}{1\ mol\ KClO_3}\right)$

$= 0.510625\ g\ KClO_3$ (unrounded)

Mass % $KClO_3 = \left(\dfrac{0.510625\ g\ KClO_3}{0.900\ g\ Sample}\right) \times 100\% = 56.7361 = \mathbf{56.7\%\ KClO_3}$
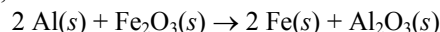
4.98 $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

Mass $CaCO_3 = \left((0.693 - 0.508)g\ CO_2\right)\left(\dfrac{1\ mol\ CO_2}{44.01\ g\ CO_2}\right)\left(\dfrac{1\ mol\ CaCO_3}{1\ mol\ CO_2}\right)\left(\dfrac{100.09\ g\ CaCO_3}{1\ mol\ CaCO_3}\right)$

$= 0.420737\ g\ CaCO_3$ (unrounded)

Mass % $CaCO_3 = \left(\dfrac{0.420737\ g\ CaCO_3}{0.693\ g\ Sample}\right) \times 100\% = 60.7124 = \mathbf{60.7\%\ CaCO_3}$

4.99   To find the mass of Fe, write a balanced equation for the reaction, determine whether Al or Fe_2O_3 is the limiting reactant, and convert to mass.

$2 Al(s) + Fe_2O_3(s) \rightarrow 2 Fe(s) + Al_2O_3(s)$

When the masses of both reactants are given, you must determine which reactant is limiting.

Mole Fe (from Al) $= \left(1.00\ kg\ Al\right)\left(\dfrac{10^3\ g}{1\ kg}\right)\left(\dfrac{1\ mol\ Al}{26.98\ g\ Al}\right)\left(\dfrac{2\ mol\ Fe}{2\ mol\ Al}\right) = 37.064\ mol\ Fe$ (unrounded)

Mole Fe (from Fe_2O_3) $= \left(2.00\ mol\ Fe_2O_3\right)\left(\dfrac{2\ mol\ Fe}{1\ mol\ Fe_2O_3}\right) = 4.00\ mol\ Fe$

Fe_2O_3 is limiting, so 4.00 moles of Fe forms.

Mass $= \left(4.00\ mol\ Fe\right)\left(\dfrac{55.85\ g\ Fe}{1\ mol\ Fe}\right) = 223.4 = 223\ g\ Fe$

Though not required by the problem, this could be converted to **0.223 kg**.

4.100   In ionic compounds, iron has two common oxidation states, +2 and +3. First, write the balanced equations for the formation and decomposition of compound A. Then, determine which reactant is limiting and from the amount of the limiting reactant calculate how much compound B will form.

Compound A is iron chloride with iron in the higher, +3, oxidation state. Thus, the formula for compound A is $FeCl_3$ and the correct name is iron(III) chloride. The balanced equation for formation of $FeCl_3$ is:

$2 Fe(s) + 3 Cl_2(g) \rightarrow 2 FeCl_3(s)$
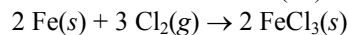
The $FeCl_3$ decomposes to $FeCl_2$ with iron in the +2 oxidation state. The balanced equation for the decomposition of $FeCl_3$ is:

$2 FeCl_3(s) \rightarrow 2 FeCl_2(s) + Cl_2(g)$
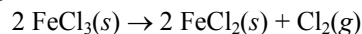
To find out whether 50.6 g Fe or 83.8 g Cl_2 limits the amount of product, calculate the number of moles of iron(III) chloride that could form based on each reactant.

Mole $FeCl_3$ (from Fe) $= \left(50.6\ g\ Fe\right)\left(\dfrac{1\ mol\ Fe}{55.85\ g\ Fe}\right)\left(\dfrac{2\ mol\ FeCl_3}{2\ mol\ Fe}\right) = 0.905998\ FeCl_3$ (unrounded)

Mole $FeCl_3$ (from Cl_2) $= \left(83.8\ g\ Cl_2\right)\left(\dfrac{1\ mol\ Cl_2}{70.90\ g\ Cl_2}\right)\left(\dfrac{2\ mol\ FeCl_3}{3\ mol\ Cl_2}\right) = 0.787964\ FeCl_3$ (unrounded)

Since fewer moles of $FeCl_3$ are produced from the available amount of chlorine, the chlorine is limiting and we calculate the amount of compound B from the amount of chlorine gas.

$$\text{Mass FeCl}_2 = (83.8 \text{ g Cl}_2)\left(\frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2}\right)\left(\frac{2 \text{ mol FeCl}_3}{3 \text{ mol Cl}_2}\right)\left(\frac{2 \text{ mol FeCl}_2}{2 \text{ mol FeCl}_3}\right)\left(\frac{126.75 \text{ g FeCl}_2}{1 \text{ mol FeCl}_2}\right)$$

$$= 99.874 = \mathbf{99.9 \text{ g FeCl}_2}$$

4.101 The equilibrium state is dynamic, because the forward and reverse processes continue even after apparent change has ceased.

4.102 The system must be closed so the gaseous product cannot escape.

4.103 Acetic acid molecules are constantly and randomly colliding with each other and with water molecules. Occasionally, an H^+ is transferred from an acetic acid molecule to a water molecule because of one of these collisions. At first, this process occurs at a larger rate than the reverse transfer of an H^+ from a H_3O^+ to a CH_3COO^-. As the concentration of H^+ and CH_3COO^- build up, the rate of transfer of H^+ between them equals the rate of transfer of H^+ between CH_3COOH and H_2O and equilibrium is reached when about 2% of the CH_3COOH molecules have ionized.

4.104 On a molecular scale, chemical reactions are dynamic. If NO and Br_2 are placed in a container, molecules of NO and molecules of Br_2 will react to form NOBr. Some of the NOBr molecules will decompose and the resulting NO and Br_2 molecules will recombine with different NO and Br_2 molecules. In this sense, the reaction is dynamic because the original NO and Br_2 pairings do not remain permanently attached to each other. Eventually, the rate of the forward reaction (combination of NO and Br_2) will equal the rate of the reverse reaction (decomposition of NOBr) at which point the reaction is said to have reached equilibrium. If you could take "snapshot" pictures of the molecules at equilibrium, you would see a constant number of reactant (NO, Br_2) and product molecules (NOBr), but the pairings would not stay the same.

4.105 The reaction is $N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)$.
a) First, the limiting reactant must be found.

$$\text{Mole NH}_3 \text{ (from H}_2) = (10.0 \text{ g H}_2)\left(\frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2}\right)\left(\frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2}\right) = 3.306878 \text{ mol NH}_3 \text{ (unrounded)}$$

$$\text{Mole NH}_3 \text{ (from N}_2) = (20.0 \text{ g N}_2)\left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2}\right)\left(\frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2}\right) = 1.42755 \text{ mol NH}_3 \text{ (unrounded)}$$

N_2 is the limiting reactant.

$$\text{Mass NH}_3 = (20.0 \text{ g N}_2)\left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2}\right)\left(\frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2}\right)\left(\frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3}\right) = 24.311 \text{ g NH}_3 \text{ (unrounded)}$$
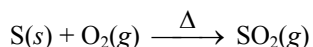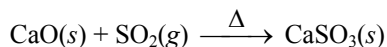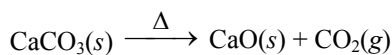
% yield = (15.0 g actual)/(24.311 g theoretical) x 100% = 61.700 = **61.7%**
b) Initially, 4.96 moles of H_2 and 0.714 moles of N_2 were present. 15.0 g (= 0.881 moles) NH_3 was formed, requiring that 0.440 moles of N_2 and 1.32 moles of H_2 be consumed. This leaves **0.274 moles N_2** and **3.64 moles H_2** at equilibrium.

4.106 a) $Fe(s) + 2 H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$
O.N.:   0   +1   +2   0

b) Fe^{2+} ions $= (737 \text{ g Sauce})\left(\frac{49 \text{ mg Fe}}{125 \text{ g Sauce}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right)\left(\frac{1 \text{ mol Fe}^{2+}}{1 \text{ mol Fe}}\right)\left(\frac{6.022 \times 10^{23} \text{ Fe}^{2+}\text{ions}}{1 \text{ mol Fe}^{2+}}\right)$

$= 3.11509 \times 10^{21} = \mathbf{3.1 \times 10^{21} \text{ Fe}^{2+} \text{ ions}}$ per jar of sauce

4.107 The reactions are:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

$$CaO(s) + SO_2(g) \xrightarrow{\Delta} CaSO_3(s)$$

$$S(s) + O_2(g) \xrightarrow{\Delta} SO_2(g)$$

These reactions sum to: $CaCO_3(s) + S(s) + O_2(g) \rightarrow CaSO_3(s) + CO_2(g)$

Mass $CaCO_3 =$

$$\left(8.5 \times 10^4 \text{ kg Coal}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{0.33\%}{100\%}\right)\left(\frac{1 \text{ mol S}}{32.07 \text{ g S}}\right)\left(\frac{1 \text{ mol } CaCO_3}{1 \text{ mol S}}\right)\left(\frac{100.09 \text{ g } CaCO_3}{1 \text{ mol } CaCO_3}\right)\left(\frac{100\%}{70\%}\right)$$

$$= 1.25062 \times 10^6 = \textbf{1.3} \times \textbf{10}^6 \text{ g } \textbf{CaCO}_3 \text{ needed}$$

Note: The percent ratio yields the mass of sulfur in the compound.

4.108 Mass of $C_2H_5OH = \left(10.0 \text{ g } C_6H_{12}O_6\right)\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.16 \text{ g } C_6H_{12}O_6}\right)\left(\frac{2 \text{ mol } C_2H_5OH}{1 \text{ mol } C_6H_{12}O_6}\right)\left(\frac{46.07 \text{ g } C_2H_5OH}{1 \text{ mol } C_2H_5OH}\right)$

$$= 5.1143 = \textbf{5.11 g } \textbf{C}_2\textbf{H}_5\textbf{OH}$$

Volume $CO_2 = \left(10.0 \text{ g } C_6H_{12}O_6\right)\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.16 \text{ g } C_6H_{12}O_6}\right)\left(\frac{2 \text{ mol } CO_2}{1 \text{ mol } C_6H_{12}O_6}\right)\left(\frac{22.4 \text{ L } CO_2}{1 \text{ mol } CO_2}\right)$

$$= 2.4866785 = \textbf{2.49 L } \textbf{CO}_2$$

4.109 Mass Fe $= \left(39.32 \text{ mL}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.03190 \text{ mol } MnO_4^-}{\text{L}}\right)\left(\frac{5 \text{ mol } Fe^{2+}}{1 \text{ mol } MnO_4^-}\right)\left(\frac{55.85 \text{ g Fe}}{1 \text{ mol } Fe^{2+}}\right)$

$$= 0.3502655 \text{ g Fe (unrounded)}$$

Mass % Fe $= [(0.3502655 \text{ g Fe}) / (1.1081 \text{ g Sample})] \times 100\% = 31.60956 = \textbf{31.61\% Fe}$

4.110 a) A solution of $Na_2C_2O_4$ (soluble salt) is added to a solution containing 1.9348 g of road salt (NaCl and $CaCl_2$, both soluble salts) and an insoluble precipitate is formed (CaC_2O_4). The Na^+ and Cl^- ions are spectators in this reaction because all Na^+ salts are soluble and most Cl^- salts are soluble.

$$Ca^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow CaC_2O_4(s)$$
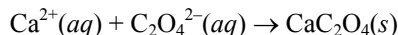
b) You may recognize this reaction as a redox titration, because the permanganate ion, MnO_4^-, is a common oxidizing agent. The MnO_4^- oxidizes the oxalate ion, $C_2O_4^{2-}$ to CO_2. Mn changes from +7 to +2 (reduction) and C changes from +3 to +4 (oxidation). The equation that describes this process is:
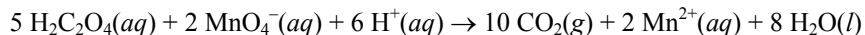
$$H_2C_2O_4(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + CO_2(g)$$
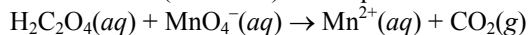
$H_2C_2O_4$ is a weak acid, so it cannot be written in a fully dissociated form ($2 \text{ H}^+(aq) + C_2O_4^{2-}(aq)$). $KMnO_4$ is a soluble salt, so it can be written in its dissociated form. $K^+(aq)$ is omitted because it is a spectator ion. We will balance the equation using the oxidation number method (you can verify by doing the half reaction method) and first assigning oxidation numbers to all elements in the reaction.

$$H_2C_2O_4(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 2 \text{ } CO_2(g)$$

Oxidation numbers: +1 +3 -2   +7 -2   +2   +4 -2
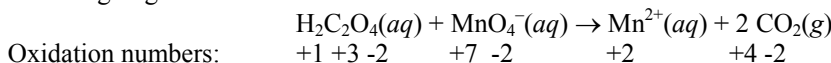
Identify the oxidized and reduced species and multiply one or both species by the appropriate factors to make the electrons lost equal the electrons gained.

$$5 \text{ } H_2C_2O_4(aq) + 2 \text{ } MnO_4^-(aq) \rightarrow 2 \text{ } Mn^{2+}(aq) + 10 \text{ } CO_2(g)$$
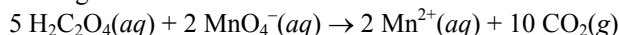
Adding water and $H^+(aq)$ to finish balancing the equation is appropriate since the reaction takes place in acidic medium. Add 8 $H_2O(l)$ to right side of equation to balance the oxygen and then add 6 $H^+(aq)$ to the left to balance hydrogen.

$$5 \text{ } H_2C_2O_4(aq) + 2 \text{ } MnO_4^-(aq) + 6 \text{ } H^+(aq) \rightarrow 10 \text{ } CO_2(g) + 2 \text{ } Mn^{2+}(aq) + 8 \text{ } H_2O(l)$$
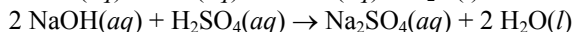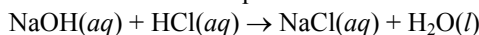
c) Oxidizing agent = $\textbf{KMnO}_4$

d) Reducing agent = $\textbf{H}_2\textbf{C}_2\textbf{O}_4$ (remember that the *agent* refers to the whole compound, not a particular element within a compound)

e) The balanced equations provide the accurate molar ratios between species. Working backwards, we know that
(Moles of MnO_4^-) is related to moles of $C_2O_4^{2-}$ at the endpoint (5:2)
(Moles of $C_2O_4^{2-}$) is related to moles of CaC_2O_4 (1:1)
(Moles of CaC_2O_4) is related to moles of $CaCl_2$ (1:1)

Mass $CaCl_2 = \left(37.68 \text{ mL}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.1019 \text{ mol } KMnO_4}{1 \text{ L}}\right)\left(\frac{5 \text{ mol } H_2C_2O_4}{2 \text{ mol } KMnO_4}\right)\left(\frac{1 \text{ mol } CaCl_2}{1 \text{ mol } H_2C_2O_4}\right)\left(\frac{110.98 \text{ g } CaCl_2}{1 \text{ mol } CaCl_2}\right)$

$$= 1.06529 \text{ g } CaCl_2 \text{ (unrounded)}$$

Mass % $CaCl_2 = [(1.06529 \text{ g } CaCl_2) / (1.9348 \text{ g Sample})] \times 100\% = 55.0594 = \textbf{55.06\% CaCl}_2$

4.111 Write the balanced chemical equations:
  $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
  $2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$

Determine the NaOH concentration from the reaction of NaOH with H_2SO_4.

$$M\ NaOH = \left(50.0\ mL\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.0782\ mol\ H_2SO_4}{L}\right)\left(\frac{2\ mol\ NaOH}{1\ mol\ H_2SO_4}\right)\left(\frac{1}{18.4\ mL}\right)\left(\frac{1\ mL}{10^{-3}\ L}\right)$$

  $= \mathbf{0.425\ \textit{M}\ NaOH}$

Use the NaOH concentration and the reaction of HCl with NaOH to determine HCl concentration.

$$M\ HCl = \left(27.5\ mL\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.425\ mol\ NaOH}{L}\right)\left(\frac{1\ mol\ HCl}{1\ mol\ NaOH}\right)\left(\frac{1}{100.\ mL}\right)\left(\frac{1\ mL}{10^{-3}\ L}\right)$$

  $= 0.116875 = \mathbf{0.117\ \textit{M}\ HCl}$

4.112 a)   Molecular: $H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$
       Total ionic: $2\ H^+(aq) + SO_4^{2-}(aq) + 2\ Na^+(aq) + 2\ OH^-(aq) \rightarrow 2\ Na^+(aq) + SO_4^{2-}(aq) + 2\ H_2O(l)$
       Net ionic: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
   b)   **0.020 mol of H_2SO_4** and **0.040 mol of NaOH** reacted.
   c) $M\ H_2SO_4 = (0.020\ mol\ H_2SO_4/25mL)\ (1\ mL\ /\ 10^{-3}\ L) = \mathbf{0.80\ \textit{M}\ H_2SO_4}$
       $M\ NaOH = (0.040\ mol\ NaOH/25\ mL)\ (1\ mL\ /\ 10^{-3}\ L) = \mathbf{1.6\ \textit{M}\ NaOH}$

4.113 The balanced equation for this reaction is:
  $CaMg(CO_3)_2(s) + 4\ HCl(aq) \rightarrow Ca^{2+}(aq) + Mg^{2+}(aq) + 2\ H_2O(l) + 2\ CO_2(g) + 4\ Cl^-(aq)$

To find mass %, convert the solution to moles of HCl, use the molar ratio to find moles of dolomite, convert to mass and divide by the mass of soil.

$$Mass\ CaMg(CO_3)_2 = \left(33.56\ mL\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.2516\ mol\ HCl}{1\ L}\right)\left(\frac{1\ mol\ CaMg(CO_3)_2}{4\ mol\ HCl}\right)\left(\frac{184.41\ g\ CaMg(CO_3)_2}{1\ mol\ CaMg(CO_3)_2}\right)$$

  $= 0.389275\ g\ CaMg(CO_3)_2$ (unrounded)

Mass % $CaMg(CO_3)_2 = [(0.389275\ g\ CaMg(CO_3)_2)\ /\ (12.86\ s\ Soil)]\ x\ 100\% = 3.02702 = \mathbf{3.027\%\ CaMg(CO_3)_2}$

4.114 The balanced chemical equations for HA or HB with sodium hydroxide are the same. For HA it is:
  $HA(aq) + NaOH(aq) \rightarrow NaA(aq) + H_2O(l)$
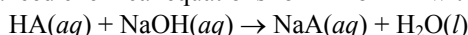
The concentration of first HA solution is:

$$M\ HA = \left(87.3\ mL\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.0906\ mol\ NaOH}{L}\right)\left(\frac{1\ mol\ HA}{1\ mol\ NaOH}\right)\left(\frac{1}{43.5\ mL}\right)\left(\frac{1\ mL}{10^{-3}\ L}\right)$$

  $= 0.1818248 = \mathbf{0.182\ \textit{M}\ HA}$

(To finish the problem, it must be assumed that both HA solutions have the same concentration. This is not necessarily true based on the information given in the problem.)

The titration of the acid mixture involves the reaction of NaOH with both of the acids. The amount reacting with HA may be related to the first titration as:

Volume NaOH reacting with second HA = (87.3 mL NaOH) (37.2 mL HA / 43.5 mL HA)
    = 74.65655 mL (unrounded)

Thus, (96.4 -74.65655) mL = 21.74345 mL NaOH reacted with HB. Calculating the molarity of HB is now similar to calculating the molarity of HA:

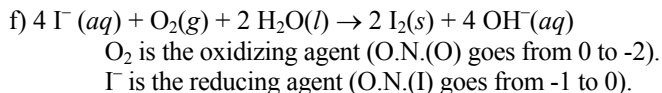$$M\ HB = \left(21.74345\ mL\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.0906\ mol\ NaOH}{L}\right)\left(\frac{1\ mol\ HB}{1\ mol\ NaOH}\right)\left(\frac{1}{(50.0-37.2)\ mL}\right)\left(\frac{1\ mL}{10^{-3}\ L}\right)$$

  $= 0.153902857 = \mathbf{0.154\ \textit{M}\ HB}$

4.115   a) Assigning oxidation states to each element indicates which are the oxidizing and reducing agents.
Step 1: N oxidized from -3 to +2 by O_2, and O is reduced from 0 to -2.
Oxidizing agent = O_2 Reducing agent = NH_3
Step 2: N oxidized from +2 to +4 by O_2, and O is reduced from 0 to -2.
Oxidizing agent = O_2 Reducing agent = NO
Step 3: N oxidized from +4 to +5 by NO_2, and N is reduced from +4 to +2.
Oxidizing agent = NO_2 Reducing agent = NO_2
b) This problem could be done as a stepwise set of calculations, however, if the equations are combined into an overall reaction the answer may be found in one calculation:

$$12\ NH_3(g) + 21\ O_2(g) \rightarrow 14\ H_2O(l) + 8\ HNO_3(l) + 4\ NO(g)$$

$$\text{Mass } NH_3 = \left(3.0 \text{ x } 10^4 \text{ kg } HNO_3\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } HNO_3}{63.02 \text{ g } HNO_3}\right)\left(\frac{12 \text{ mol } NH_3}{8 \text{ mol } HNO_3}\right)\left(\frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)$$

$$= 1.21604 \text{ x } 10^4 = \textbf{1.2 x } 10^4 \textbf{ kg } NH_3$$

This assumes that the NO produced in step 3 is not recycled.

4.116   a) $MnS(s) + 2\ HBr(aq) \rightarrow MnBr_2(aq) + H_2S(g)$
$MnS(s) + 2\ H^+(aq) \rightarrow Mn^{2+}(aq) + H_2S(g)$
b) $K_2CO_3(aq) + Sr(NO_3)_2(aq) \rightarrow SrCO_3(s) + 2\ KNO_3(aq)$
$CO_3^{2-}(aq) + Sr^{2+}(aq) \rightarrow SrCO_3(s)$
c) $KNO_2(aq) + HCl(aq) \rightarrow HNO_2(aq) + KCl(aq)$
$NO_2^-(aq) + H^+(aq) \rightarrow HNO_2(aq)$
d) $Ca(OH)_2(aq) + HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
e) $Ba(C_2H_3O_2)_2(aq) + FeSO_4(aq) \rightarrow BaSO_4(s) + Fe(C_2H_3O_2)_2(aq)$
$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
f) $ZnCO_3(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l) + CO_2(g)$
$ZnCO_3(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l) + CO_2(g)$
g) $Cu(NO_3)_2(aq) + H_2S(aq) \rightarrow CuS(s) + 2\ HNO_3(aq)$
$Cu^{2+}(aq) + H_2S(aq) \rightarrow CuS(s) + 2\ H^+(aq)$
h) $Mg(OH)_2(s) + 2\ HClO_3(aq) \rightarrow Mg(ClO_3)_2(aq) + 2\ H_2O(l)$
$Mg(OH)_2(s) + 2\ H^+(aq) \rightarrow Mg^{2+}(aq) + 2\ H_2O(l)$
i) $KCl(aq) + (NH_4)_3PO_4(aq) \rightarrow$ No Reaction
j) $Ba(OH)_2(aq) + 2\ HCN(aq) \rightarrow Ba(CN)_2(aq) + 2\ H_2O(l)$
$OH^-(aq) + HCN(aq) \rightarrow CN^-(aq) + H_2O(l)$

4.117   a) $4\ KOH(aq) + 3\ H_2O_2(aq) + 2\ Cr(OH)_3(s) \rightarrow 2\ K_2CrO_4(aq) + 8\ H_2O(l)$
H_2O_2 is the oxidizing agent (O.N.(O) goes from -1 to -2).
$Cr(OH)_3$ is the reducing agent (O.N.(Cr) goes from +3 to +6).
b) $4\ MnO_4^-(aq) + 3\ ClO_2^-(aq) + 2\ H_2O(l) \rightarrow 4\ MnO_2(s) + 3\ ClO_4^-(aq) + 4\ OH^-(aq)$
MnO_4^- is the oxidizing agent (O.N.(Mn) goes from +7 to +4).
ClO_2^- is the reducing agent (O.N.(Cl) goes from +3 to +7).
c) $2\ KMnO_4(aq) + 3\ Na_2SO_3(aq) + H_2O(l) \rightarrow 2\ MnO_2(s) + 3\ Na_2SO_4(aq) + 2\ KOH(aq)$
$KMnO_4$ is the oxidizing agent (O.N.(Mn) goes from +7 to +4).
Na_2SO_3 is the reducing agent (O.N.(S) goes from +4 to +6).
d) $2\ CrO_4^{2-}(aq) + 3\ HSnO_2^-(aq) + H_2O(l) \rightarrow 2\ CrO_2^-(aq) + 3\ HSnO_3^-(aq) + 2\ OH^-(aq)$
CrO_4^{2-} is the oxidizing agent (O.N.(Cr) goes from +6 to +3).
$HSnO_2^-$ is the reducing agent (O.N.(Sn) goes from +2 to +4).
e) $2\ KMnO_4(aq) + 3\ NaNO_2(aq) + H_2O(l) \rightarrow 2\ MnO_2(s) + 3\ NaNO_3(aq) + 2\ KOH(aq)$
$KMnO_4$ is the oxidizing agent (O.N.(Mn) goes from +7 to +4).
$NaNO_2$ is the reducing agent (O.N.(N) goes from +3 to +5).

f) $4 \, I^-(aq) + O_2(g) + 2 \, H_2O(l) \rightarrow 2 \, I_2(s) + 4 \, OH^-(aq)$
O_2 is the oxidizing agent (O.N.(O) goes from 0 to -2).
I^- is the reducing agent (O.N.(I) goes from -1 to 0).
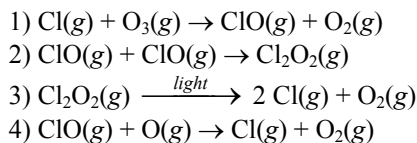
4.118   The reactions are:
1) $Cl(g) + O_3(g) \rightarrow ClO(g) + O_2(g)$
2) $ClO(g) + ClO(g) \rightarrow Cl_2O_2(g)$
3) $Cl_2O_2(g) \xrightarrow{light} 2 \, Cl(g) + O_2(g)$
4) $ClO(g) + O(g) \rightarrow Cl(g) + O_2(g)$
a) In reaction 1, Cl is oxidized (0 to +1), and O is reduced (0 to -1).
   In reaction 2, no oxidation or reduction occurs.
   In reaction 3, Cl is reduced (+1 to 0) and O is oxidized (-1 to 0).
   In reaction 4, Cl is reduced (+1 to 0) and O (from ClO) is oxidized (-1 to 0).
   The oxidation-reduction reactions are **1, 3, and 4**.
b) Combining the equations requires multiplying equation 1 by 2, and then adding the three equations to give:
   $2 \, O_3(g) \rightarrow 3 \, O_2(g)$
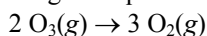
4.119   The reaction is: $2 \, Na_2O_2(s) + 2 \, CO_2(g) \rightarrow 2 \, Na_2CO_3(s) + O_2(g)$.

$$\text{Volume} = \left(80.0 \text{ g Na}_2\text{O}_2\right)\left(\frac{1 \text{ mol Na}_2\text{O}_2}{77.98 \text{ g Na}_2\text{O}_2}\right)\left(\frac{2 \text{ mol CO}_2}{2 \text{ mol Na}_2\text{O}_2}\right)\left(\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2}\right)\left(\frac{\text{L Air}}{0.0720 \text{ g CO}_2}\right)$$

$= 627.08 = $ **627 L Air**

4.120 $$\text{Mass Mg} = \left(1.00 \text{ km}^3\right)\left(\frac{10^3 \text{ m}}{1 \text{ km}}\right)^3\left(\frac{1 \text{ cm}}{10^{-2} \text{ m}}\right)^3\left(\frac{1 \text{ mL}}{1 \text{ cm}^3}\right)\left(\frac{1.04 \text{ g}}{1 \text{ mL}}\right)\left(\frac{0.13\% \text{ Mg}}{100\%}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)$$

$= 1.3520 \times 10^9 = $ **1.4×10^9 kg Mg**

4.121   A 1.00 kg piece of glass of composition 75% SiO_2, 15% Na_2O, and 10% CaO would contain
   (1.00 kg) (75% / 100%) = **0.75 kg SiO_2**
   (1.00 kg) (15% / 100%) = 0.15 kg Na_2O
   (1.00 kg) (10% / 100%) = 0.10 kg CaO
In this example, the SiO_2 is added directly while the sodium oxide comes from decomposition of sodium carbonate
and the calcium oxide from decomposition of calcium carbonate:
   $Na_2CO_3(s) \rightarrow Na_2O(s) + CO_2(g)$
   $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
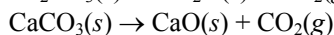
$$\text{Mass Na}_2\text{CO}_3 = \left(0.15 \text{ kg Na}_2\text{O}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol Na}_2\text{O}}{61.98 \text{ g Na}_2\text{O}}\right)\left(\frac{1 \text{ mol Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{O}}\right)\left(\frac{105.99 \text{ g Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)$$

$= 0.25651 = $ **0.26 kg Na_2CO_3**

$$\text{Mass CaCO}_3 = \left(0.10 \text{ kg CaO}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol CaO}}{56.08 \text{ g CaO}}\right)\left(\frac{1 \text{ mol CaCO}_3}{1 \text{ mol CaO}}\right)\left(\frac{100.09 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)$$

$= 0.178477 = $ **0.18 kg $CaCO_3$**
Combine 0.75 kg SiO_2, 0.26 kg Na_2CO_3, and 0.18 kg $CaCO_3$ and heat to make 1.00 kg glass.

4.122   a) $$\text{Moles C}_2\text{H}_4 = \left(0.100 \text{ kg C}_2\text{H}_4\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol C}_2\text{H}_4}{28.05 \text{ g C}_2\text{H}_4}\right) = 3.56506 \text{ mol C}_2\text{H}_4 \text{ (unrounded)}$$

$$\text{Moles HCl} = \left(0.100 \text{ kg HCl}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}}\right) = 2.74273 \text{ mol HCl (unrounded)}$$
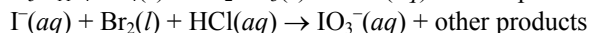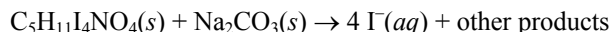
The HCl is limiting.

Moles HCl remaining = 0 mol

Moles C_2H_4 remaining = 3.56506 mol C_2H_4 - (2.74273 mol HCl) (1mol C_2H_4 / 1 mol HCl)

   = 0.822328 = 0.82 mol C_2H_4

Moles C_2H_5Cl formed = (2.74273 mol HCl) (1 mol C_2H_5Cl / 1 mol HCl) = 2.74273 = 2.74 mol C_2H_5Cl

Molecules of gas = Total mol x Avogadro's Number

   = [(0 + 0.822328 + 2.74273) mol](6.022 x 10^{23} molecules / mol)

    = 2.1468779 x 10^{24} = **2.15 x 10^{24} molecules**
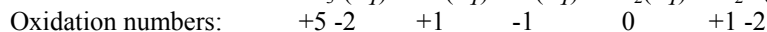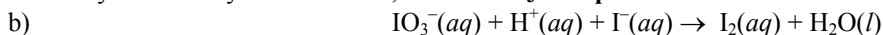
b) This will still be based on the HCl.

Moles HCl = (2.74273 mol HCl) / 2 = 1.371365 mol HCl (unrounded)

Moles C_2H_4 = 3.56506 mol C_2H_4 - [(1.371365 mol HCl) (1 mol C_2H_4 / 1 mol HCl)]

   = 2.193695 mol C_2H_4 (unrounded)

Moles C_2H_5Cl = (1.371365 mol HCl) (1 mol C_2H_5Cl / 1 mol HCl)

   = 1.371365 mol C_2H_5Cl (unrounded)

Total moles = (1.371365 + 2.193695 + 1.371365) mol = 4.936425 = **4.94 mol total**

4.123 Let x = grams of NaCl, then 35.0 - x = g of $MgSO_4$.

Moles NaCl = x grams (1 mol NaCl / 58.44 g NaCl) = x / 58.44

Moles $MgSO_4$ = (35 - x) grams (1 mol $MgSO_4$ / 120.38 g $MgSO_4$) = (35 - x) / 120.38

Given that moles of NaCl is twice as large as moles of $MgSO_4$

x / 58.44 = 2 (35 - x) / 120.38

x / 58.44 = (70 - 2 x) / 120.38

120.38 x = 58.44 (70 - 2 x) = 4090.8 - 116.88 x

(120.38 + 116.88) x = 237.26 x = 4090.8

x = 4090.8 / 237.26 = 17.2418 = **17.2 g NaCl**

35.0 - x = 35.0 - 17.2418 = 17.7582 = **17.8 g $MgSO_4$**

4.124 a) There is not enough information to write complete chemical equations, but the following equations can be written:

   $C_5H_{11}I_4NO_4(s)$ + $Na_2CO_3(s)$ → 4 $I^-(aq)$ + other products

   $I^-(aq)$ + $Br_2(l)$ + HCl(aq) → $IO_3^-(aq)$ + other products

For every mole of thyroxine reacted, **4 moles IO_3^- are produced**.

b) $IO_3^-(aq)$ + $H^+(aq)$ + $I^-(aq)$ → $I_2(aq)$ + $H_2O(l)$

   Oxidation numbers: +5 -2 +1 -1   0 +1 -2

This is a difficult equation to balance because the iodine species are both reducing and oxidizing. Start balancing the equation by placing a coefficient of 5 in front of $I^-(aq)$, so the electrons lost equal the electrons gained. Do <u>not</u> place a 5 in front of $I_2(aq)$, because not all of the $I_2(aq)$ comes from oxidation of $I^-(aq)$. Some of the $I_2(aq)$ comes from the reduction of $IO_3^-(aq)$. Place a coefficient of 3 in front of $I_2(aq)$ to correctly balance iodine. The reaction is now balanced from a redox standpoint, so finish balancing the reaction by balancing the oxygen and hydrogen.

   $IO_3^-(aq)$ + 6 $H^+(aq)$ + 5 $I^-(aq)$ → 3 $I_2(aq)$ + 3 $H_2O(l)$
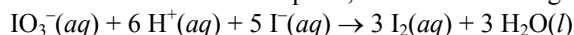
   IO_3^- is the oxidizing agent, and **I^- is the reducing agent**.

If 3 moles of I_2 are produced per mole of IO_3^-, and 4 moles of IO_3^- are produced per mole of thyroxine, then **12 moles of I_2 are produced per mole of thyroxine**.

c) Using Thy to represent thyroxine

The balanced equation for this reaction is $I_2(aq)$ + 2 $S_2O_3^{2-}(aq)$ → 2 $I^-(aq)$ + $S_4O_6^{2-}(aq)$.

$$\text{Mass Thy} = (17.23 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.1000 \text{ mol } S_2O_3^{2-}}{L}\right)\left(\frac{1 \text{ mol } I_2}{2 \text{ mol } S_2O_3^{2-}}\right)\left(\frac{1 \text{ mol Thy}}{12 \text{ mol } I_2}\right)\left(\frac{776.8 \text{ g Thy}}{1 \text{ mol Thy}}\right)$$

   = 0.055767766 g Thy (unrounded)

Mass % Thy = [(0.055767766 g Thy) / (0.4332 g extract)] x 100%

   = 12.8734 = **12.87% thyroxine**

4.125 The reactions are 2 LiOH(s) + $CO_2(g)$ → $Li_2CO_3(s)$ + $H_2O(l)$

$Mg(OH)_2(s)$ + $CO_2(g)$ → $MgCO_3(s)$ + $H_2O(l)$

2 $Al(OH)_3(s)$ + 3 $CO_2(g)$ → $Al_2(CO_3)_3(s)$ + 3 $H_2O(l)$

a) Mass $CO_2 = (3.50 \text{ kg LiOH})\left(\dfrac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\dfrac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}}\right)\left(\dfrac{1 \text{ mol } CO_2}{2 \text{ mol LiOH}}\right)\left(\dfrac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right)$

$= 3215.762 = \textbf{3.22 x } \mathbf{10^3} \textbf{ g } \mathbf{CO_2}$

b) Mass CO_2 absorbed by 1.00 g LiOH

$\text{Mass} = (1.00 \text{ g LiOH})\left(\dfrac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}}\right)\left(\dfrac{1 \text{ mol } CO_2}{2 \text{ mol LiOH}}\right)\left(\dfrac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right)$

$= 0.918789 = \textbf{0.919 g } \mathbf{CO_2}$

Mass CO_2 absorbed by 1.00 g $Mg(OH)_2$

$\text{Mass} = (1.00 \text{ g } Mg(OH)_2)\left(\dfrac{1 \text{ mol } Mg(OH)_2}{58.33 \text{ g } Mg(OH)_2}\right)\left(\dfrac{1 \text{ mol } CO_2}{1 \text{ mol } Mg(OH)_2}\right)\left(\dfrac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right)$

$= 0.75450 = \textbf{0.754 g } \mathbf{CO_2}$

Mass CO_2 absorbed by 1.00 g $Al(OH)_3$

$\text{Mass} = (1.00 \text{ g } Al(OH)_3)\left(\dfrac{1 \text{ mol } Al(OH)_3}{78.00 \text{ g } Al(OH)_3}\right)\left(\dfrac{3 \text{ mol } CO_2}{2 \text{ mol } Al(OH)_3}\right)\left(\dfrac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right)$

$= 0.846346 = \textbf{0.846 g } \mathbf{CO_2}$

4.126 Balance the equation to obtain the correct molar ratios. This is not a redox reaction as none of the O.N.'s change. Here is a suggested method for approaching balancing the equation.
— Since PO_4^{2-} remains as a unit on both sides of the equation, treat it as a unit when balancing.
— On first inspection, one can see that Na needs to be balanced by adding a "2" in front of $NaHCO_3$. This then affects the balance of C, so add a "2" in front of CO_2.
— Hydrogen is not balanced, so change the coefficient of water to "2," as this will have the least impact on the other species.
— Verify that the other species are balanced.

$$Ca(H_2PO_4)_2(s) + 2\, NaHCO_3(s) \xrightarrow{\Delta} 2\, CO_2(g) + 2\, H_2O(g) + CaHPO_4(s) + Na_2HPO_4(s)$$
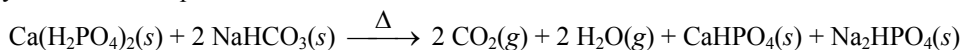
a) Determine whether $Ca(H_2PO_4)_2$ or $NaHCO_3$ limits the production of CO_2. In each case calculate the moles of CO_2 that might form.

$\text{Mole } CO_2 \,(NaHCO_3) = (1.00 \text{ g})\left(\dfrac{31\%}{100\%}\right)\left(\dfrac{1 \text{ mol } NaHCO_3}{84.01 \text{ g } NaHCO_3}\right)\left(\dfrac{2 \text{ mol } CO_2}{2 \text{ mol } NaHCO_3}\right)$

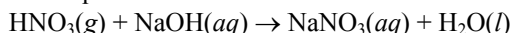$= 3.690 \text{ x } 10^{-3} \text{ mol } CO_2 \text{ (unrounded)}$

$\text{Mole } CO_2 \,(Ca(H_2PO_4)_2) = (1.00 \text{ g})\left(\dfrac{35\%}{100\%}\right)\left(\dfrac{1 \text{ mol } Ca(H_2PO_4)_2}{234.05 \text{ g } Ca(H_2PO_4)_2}\right)\left(\dfrac{2 \text{ mol } CO_2}{1 \text{ mol } Ca(H_2PO_4)_2}\right)$

$= 2.9908 \text{ x } 10^{-3} \text{ mol } CO_2 \text{ (unrounded)}$

Since $Ca(H_2PO_4)_2$ is limiting, $\mathbf{3.0 \text{ x } 10^{-3} \text{ mol } CO_2}$ will be produced.

b) Volume $CO_2 = (2.9908 \text{ x } 10^{-3} \text{ mol } CO_2)(37.0 \text{ L/mol } CO_2) = 0.1106596 = \textbf{0.11 L } \mathbf{CO_2}$

4.127 The chemical equation is:

$$HNO_3(g) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

a) It takes a total of $(20.00 + 3.22)$ mL $= 23.22$ mL NaOH to titrate a total of $(50.00 + 30.00)$ mL $= 80.00$ mL of acid.

$M \text{ HCl} = (23.22 \text{ mL})\left(\dfrac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\dfrac{0.0502 \text{ mol NaOH}}{L}\right)\left(\dfrac{1 \text{ mol } HNO_3}{1 \text{ mol NaOH}}\right)\left(\dfrac{1}{80.00 \text{ mL}}\right)\left(\dfrac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$

$= 0.01457055 = \textbf{0.0146 } \textbf{\textit{M}} \textbf{ HCl}$

b) First calculate the moles of the acid and base present. The difference will give the excess NaOH.

$\text{Moles NaOH} = (20.00 \text{ mL})\left(\dfrac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\dfrac{0.0502 \text{ mol NaOH}}{L}\right) = 1.004 \text{ x } 10^{-3} \text{ mol NaOH (unrounded)}$

$$\text{Moles HCl} = \left(50.00 \text{ mL}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.01457055 \text{ mol HCl}}{\text{L}}\right)$$

$$= 7.285275 \times 10^{-4} \text{ mol HCl (unrounded)}$$

Moles excess NaOH = 1.004×10^{-3} mol NaOH - (7.285275×10^{-4} mol HCl) (1 mol NaOH / 1 mol HCl)

$$= 2.754725 \times 10^{-4} = 2.8 \times 10^{-4} \text{ mol NaOH}$$

4.128   a) Determine the moles of each element present. The sample was burned in an unknown amount of O_2, therefore, the moles of oxygen must be found by a different method.

$$\text{Moles C} = \left(0.1880 \text{ g CO}_2\right)\left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2}\right)\left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2}\right) = 4.271756 \times 10^{-3} \text{ mol C (unrounded)}$$

$$\text{Moles H} = \left(0.02750 \text{ g H}_2\text{O}\right)\left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}}\right)\left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}}\right) = 3.052164 \times 10^{-3} \text{ mol H (unrounded)}$$

$$\text{Moles Bi} = \left(0.1422 \text{ g Bi}_2\text{O}_3\right)\left(\frac{1 \text{ mol Bi}_2\text{O}_3}{466.0 \text{ g Bi}_2\text{O}_3}\right)\left(\frac{2 \text{ mol Bi}}{1 \text{ mol Bi}_2\text{O}_3}\right) = 6.103004 \times 10^{-4} \text{ mol Bi (unrounded)}$$

Subtracting the mass of each element present from the mass of the sample will give the mass of oxygen originally present in the sample. This mass is used to find the moles of oxygen.
0.22105 g sample - [12.01 g/mol (4.271756×10^{-3} mol C)] - [1.008 g/mol (3.052164×10^{-3} mol H)]
        - [(209.0 g/mol (6.103004×10^{-4} mol Bi)] = 0.0391168 g O (unrounded)
Moles O = (0.0391168 g O) (1 mol O / 16.00 g O) = 2.44480×10^{-3} mol O (unrounded)
Divide each of the moles by the smallest value (moles Bi).
C = (4.271756×10^{-3} mol) / (6.103004×10^{-4} mol) = 7
H = (3.052164×10^{-3} mol) / (6.103004×10^{-4} mol) = 5
O = (2.44480×10^{-3} mol) / (6.103004×10^{-4} mol) = 4
Bi = (6.103004×10^{-4} mol) / (6.103004×10^{-4} mol) = 1
        Empirical formula = **$C_7H_5O_4Bi$**
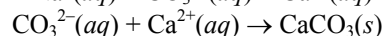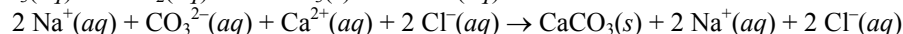b) The empirical formula mass is 362 g/mol. Therefore, there are 1086 / 362 = 3 empirical formula units per molecular formula making the molecular formula = 3 x $C_7H_5O_4Bi$ = **$C_{21}H_{15}O_{12}Bi_3$**.
c) Bi(OH)$_3$(s) + 3 HC$_7$H$_5$O$_3$(aq) → Bi(C$_7$H$_5$O$_3$)$_3$(s) + 3 H$_2$O(l)

$$\text{d) } \left(0.600 \text{ mg}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mol Active}}{1086 \text{ g}}\right)\left(\frac{3 \text{ mol Bi}}{1 \text{ mol Active}}\right)\left(\frac{1 \text{ mol Bi(OH)}_3}{1 \text{ mol Bi}}\right)\left(\frac{260.0 \text{ g Bi(OH)}_3}{1 \text{ mol Bi(OH)}_3}\right)\left(\frac{1 \text{ mg}}{10^{-3} \text{ g}}\right)\left(\frac{100\%}{88.0\%}\right)$$

$$= 0.48970 = \textbf{0.490 mg Bi(OH)}_3$$

4.129   a) Na$_2$CO$_3$(aq) + CaCl$_2$(aq) → CaCO$_3$(s) + 2 NaCl(aq)
        2 Na$^+$(aq) + CO$_3$$^{2-}$(aq) + Ca^{2+}(aq) + 2 Cl$^-$(aq) → CaCO$_3$(s) + 2 Na$^+$(aq) + 2 Cl$^-$(aq)
        CO$_3$$^{2-}$(aq) + Ca^{2+}(aq) → CaCO$_3$(s)
b) Mass = 2 (0.050 mol ion) (1 mol CaCO$_3$ / 1 mol ion) (100.09 g CaCO$_3$ / 1 mol CaCO$_3$)
        = 10.009 = **10. g CaCO$_3$**
c) Original moles: Na$^+$ = 6 (0.050) mol   CO$_3$$^{2-}$ = 3 (0.050) mol
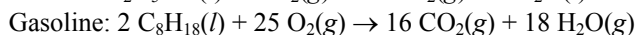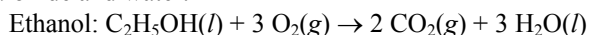   Ca^{2+} = 2 (0.050) mol   Cl$^-$ = 4 (0.050) mol
The moles of Na$^+$ and Cl$^-$ do not change. The moles of Ca^{2+} goes to zero, and removes 2 (0.050) mol of CO$_3$$^{2-}$.

$$M \text{ Na}^+ = \frac{6\left(0.050 \text{ mol Na}^+\right)}{\left(250. + 250.\right)\text{mL}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right) = \textbf{0.60 } M \textbf{ Na}^+$$

$$M \text{ Cl}^- = \frac{4\left(0.050 \text{ mol Cl}^-\right)}{\left(250. + 250.\right)\text{mL}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right) = \textbf{0.40 } M \textbf{ Cl}^-$$

$$M \text{ CO}_3{}^{2-} = \frac{\left(3-2\right)\left(0.050 \text{ mol CO}_3{}^{2-}\right)}{\left(250. + 250.\right)\text{mL}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right) = \textbf{0.10 } M \textbf{ CO}_3{}^{2-}$$

4.130 a) Complete combustion of hydrocarbons involves heating the hydrocarbon in the presence of oxygen to produce carbon dioxide and water.

Ethanol: $C_2H_5OH(l) + 3 O_2(g) \rightarrow 2 CO_2(g) + 3 H_2O(l)$

Gasoline: $2 C_8H_{18}(l) + 25 O_2(g) \rightarrow 16 CO_2(g) + 18 H_2O(g)$

b) The amounts of each fuel must be found:

Gasoline = (1.00 L) (90.0% / 100%) (1 mL / 10^{-3} L) (0.742 g /mL) = 667.8 g gasoline (unrounded)

Ethanol = (1.00 L) (10.0% / 100%) (1 mL / 10^{-3} L) (0.789 g /mL) = 78.9 g ethanol

$$\text{Mass } O_2 \text{ (gasoline)} = \left(667.8 \text{ g } C_8H_{18}\right)\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{25 \text{ mol } O_2}{2 \text{ mol } C_8H_{18}}\right)\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}\right)$$

$= 2338.64$ g O_2 (unrounded)

$$\text{Mass } O_2 \text{ (ethanol)} = \left(78.9 \text{ g } C_2H_5OH\right)\left(\frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g } C_2H_5OH}\right)\left(\frac{3 \text{ mol } O_2}{1 \text{ mol } C_2H_5OH}\right)\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}\right)$$

$= 164.41$ g O_2 (unrounded)

Total O_2 = 2338.64 g O_2 + 164.41 g O_2 = 2503.05 = **2.50 x 10^3 g O_2**

c) (2503.05 g O_2) (1 mol O_2 / 32.00 g) (22.4 L / mol O_2 = 1752.135 = **1.75 x 10^3 L**

d) (1752.135 L O_2) (100% / 20.9%) = 8383.42 = **8.38 x 10^3 L air**

4.131 1) $2 C_8H_{18}(l) + 25 O_2(g) \rightarrow 16 CO_2(g) + 18 H_2O(g)$

2) $2 C_8H_{18}(l) + 17 O_2(g) \rightarrow 16 CO(g) + 18 H_2O(g)$
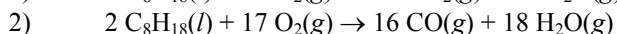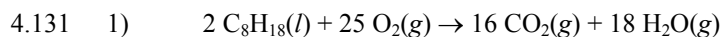
a) Assume 100 grams of gasoline, thus 95.0 grams will react by equation 1, and 5.0 grams will react by equation 2.

$$\text{Molecules } CO_2 = \left(95.0 \text{ g } C_8H_{18}\right)\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}}\right)\left(\frac{6.022 \times 10^{23} CO_2}{1 \text{ mol } CO_2}\right)$$

$= 4.00693 \times 10^{24}$ molecules CO_2 (unrounded)

$$\text{Molecules } CO = \left(5.0 \text{ g } C_8H_{18}\right)\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{16 \text{ mol } CO}{2 \text{ mol } C_8H_{18}}\right)\left(\frac{6.022 \times 10^{23} CO}{1 \text{ mol } CO}\right)$$

$= 2.10891 \times 10^{23}$ molecules CO (unrounded)

Ratio = $(4.00693 \times 10^{24}$ molecules $CO_2) / (2.10891 \times 10^{23}$ molecules CO$) = 18.99998 = $ **19**

b) Again, we may assume 100 grams of gasoline.

$$\text{Mass } CO_2 = \left(95.0 \text{ g } C_8H_{18}\right)\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}}\right)\left(\frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right)$$

$= 292.83$ g CO_2 (unrounded)

$$\text{Mass } CO = \left(5.0 \text{ g } C_8H_{18}\right)\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{16 \text{ mol } CO}{2 \text{ mol } C_8H_{18}}\right)\left(\frac{28.01 \text{ g } CO}{1 \text{ mol } CO}\right)$$

$= 9.8091$ g CO (unrounded)

Ratio = $(292.83$ g $CO_2) / (9.8091$ g CO$) = 29.852889 = $ **30**

c) For the mass ratio to be equal $\dfrac{(x)(44.01)}{y(28.01)} = 1$

Where x and y are % so, $x + y = 1$ or $y = 1 - x$
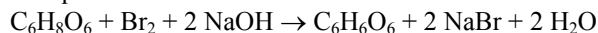
Substituting $\dfrac{(x)(44.01)}{(1 - x)(28.01)} = 1$

44.01x =   28.01 - 28.01x

72.02x =   28.01

x   =   0.39 and y = 1 - 0.39 = 0.61

Thus, **61%** of the gasoline must form CO.

4.132   Combine the two equations to get an overall equation (this will require multiplying the second equation by 2). The overall equation is:

$$C_6H_8O_6 + Br_2 + 2\ NaOH \rightarrow C_6H_6O_6 + 2\ NaBr + 2\ H_2O$$

This leads to:

$$\text{Mass vitamin C} = (43.20\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{0.1350\ \text{mol NaOH}}{\text{L}}\right)\left(\frac{1\ \text{mol } C_6H_8O_6}{2\ \text{mol NaOH}}\right)\left(\frac{176.12\ \text{g } C_6H_8O_6}{1\ \text{mol } C_6H_8O_6}\right)\left(\frac{1\ \text{mg}}{10^{-3}\ \text{g}}\right)$$

$$= 513.5659 = 513.6\ \text{mg } C_6H_8O_6$$

Yes, the tablets have the quantity advertised.

4.133   a) Mass $MgCl_2$ = $(12.4\ \text{g NaCl})\left(\dfrac{1\ \text{mol NaCl}}{58.44\ \text{g NaCl}}\right)\left(\dfrac{2\ \text{mol ions}}{1\ \text{mol NaCl}}\right)\left(\dfrac{1\ \text{mol } MgCl_2}{3\ \text{mol ions}}\right)\left(\dfrac{95.21\ \text{g } MgCl_2}{1\ \text{mol } MgCl_2}\right)$

$$= 13.46799 = \textbf{13.5 g } \textbf{MgCl}_2$$

b) Mass CaS = $(12.4\ \text{g NaCl})\left(\dfrac{1\ \text{mol NaCl}}{58.44\ \text{g NaCl}}\right)\left(\dfrac{2\ \text{mol ions}}{1\ \text{mol NaCl}}\right)\left(\dfrac{1\ \text{mol CaS}}{2\ \text{mol ions}}\right)\left(\dfrac{72.15\ \text{g CaS}}{1\ \text{mol CaS}}\right)$

$$= 15.3090 = \textbf{15.3 g CaS}$$

c) The CaS solution dissolves the most protein. All three solutions have equal ion concentrations, but CaS will form two divalent ions.

4.134   a) The **second reaction is a redox process** because the O.N. of iron changes from 0 to +2 (it oxidizes) while the O.N. of hydrogen changes from +1 to 0 (it reduces).
b) Determine the moles of HCl present and use the balanced chemical equation to determine the appropriate quantities.

$$\text{Mass } Fe_2O_3 = (2.50 \times 10^3\ \text{L})\left(\frac{3.00\ \text{mol HCl}}{\text{L}}\right)\left(\frac{1\ \text{mol } Fe_2O_3}{6\ \text{mol HCl}}\right)\left(\frac{159.70\ \text{g } Fe_2O_3}{1\ \text{mol } Fe_2O_3}\right)$$

$$= 199625 = \textbf{2.00} \times \textbf{10}^5 \textbf{ g } \textbf{Fe}_2\textbf{O}_3$$

$$\text{Mass } FeCl_3 = (2.50 \times 10^3\ \text{L})\left(\frac{3.00\ \text{mol HCl}}{\text{L}}\right)\left(\frac{2\ \text{mol } FeCl_3}{6\ \text{mol HCl}}\right)\left(\frac{162.20\ \text{g } FeCl_3}{1\ \text{mol } FeCl_3}\right)$$

$$= 405500 = \textbf{4.06} \times \textbf{10}^5 \textbf{ g } \textbf{FeCl}_3$$

c) Use reaction 2 like reaction 1 was used in part b.

$$\text{Mass Fe} = (2.50 \times 10^3\ \text{L})\left(\frac{3.00\ \text{mol HCl}}{\text{L}}\right)\left(\frac{1\ \text{mol Fe}}{2\ \text{mol HCl}}\right)\left(\frac{55.85\ \text{g Fe}}{1\ \text{mol Fe}}\right)$$

$$= 209437.5 = \textbf{2.09} \times \textbf{10}^5 \textbf{ g Fe}$$
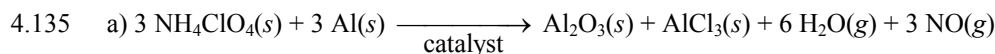
$$\text{Mass } FeCl_2 = (2.50 \times 10^3\ \text{L})\left(\frac{3.00\ \text{mol HCl}}{\text{L}}\right)\left(\frac{1\ \text{mol } FeCl_2}{2\ \text{mol HCl}}\right)\left(\frac{126.75\ \text{g } FeCl_2}{1\ \text{mol } FeCl_2}\right)$$

$$= 475312.5 = \textbf{4.75} \times \textbf{10}^5 \textbf{ g } \textbf{FeCl}_2$$

d) Use 1.00 g Fe_2O_3 to determine the mass of $FeCl_3$ formed (reaction 1), and 0.280 g Fe to determine the mass of $FeCl_2$ formed (reaction 2).

$$\text{Mass } FeCl_3 = (1.00\ \text{g } Fe_2O_3)\left(\frac{1\ \text{mol } Fe_2O_3}{159.70\ \text{g } Fe_2O_3}\right)\left(\frac{2\ \text{mol } FeCl_3}{1\ \text{mol } Fe_2O_3}\right)\left(\frac{162.20\ \text{g } FeCl_3}{1\ \text{mol } FeCl_3}\right)$$

$$= 2.0313\ \text{g } FeCl_3 \text{ (unrounded)}$$

$$\text{Mass } FeCl_2 = (0.280\ \text{g Fe})\left(\frac{1\ \text{mol Fe}}{55.85\ \text{g Fe}}\right)\left(\frac{1\ \text{mol } FeCl_2}{1\ \text{mol Fe}}\right)\left(\frac{126.75\ \text{g } FeCl_2}{1\ \text{mol } FeCl_2}\right)$$

$$= 0.635452\ \text{g } FeCl_2 \text{ (unrounded)}$$

Ratio = $(0.635452\ \text{g } FeCl_2) / (2.0313\ \text{g } FeCl_3) = 0.312830 = \textbf{0.313}$

4.135   a) $3\ NH_4ClO_4(s) + 3\ Al(s) \xrightarrow[\text{catalyst}]{} Al_2O_3(s) + AlCl_3(s) + 6\ H_2O(g) + 3\ NO(g)$

Oxidizing agent = ammonium perchlorate reducing agent = aluminum

b) Moles of gas $= \left(50.0\ \text{kg}\ NH_4ClO_4\right)\left(\dfrac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\dfrac{1\ \text{mol}\ NH_4ClO_4}{117.49\ \text{g}\ NH_4ClO_4}\right)\left(\dfrac{9\ \text{mol Gas}}{3\ \text{mol}\ NH_4ClO_4}\right)$

$= 1276.70 = \mathbf{1.28\ x\ 10^3\ mol\ gas}$

c) Initial volumes:

$V\ NH_4ClO_4 = \left(50.0\ \text{kg}\right)\left(\dfrac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\dfrac{1\ \text{cc}}{1.95\ \text{g}}\right)\left(\dfrac{1\ \text{mL}}{1\ \text{cc}}\right)\left(\dfrac{10^{-3}\ \text{L}}{1\ \text{mL}}\right) = 25.6541\ \text{L (unrounded)}$

$V\ Al =$

$\left(50.0\ \text{kg}\ NH_4ClO_4\right)\left(\dfrac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\dfrac{1\ \text{mol}\ NH_4ClO_4}{117.49\ \text{g}\ NH_4ClO_4}\right)\left(\dfrac{3\ \text{mol Al}}{3\ \text{mol}\ NH_4ClO_4}\right)\left(\dfrac{26.98\text{g Al}}{1\ \text{mol Al}}\right)\left(\dfrac{1\ \text{cc}}{2.70\ \text{g Al}}\right)\left(\dfrac{1\ \text{mL}}{1\ \text{cc}}\right)\left(\dfrac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)$

$= 4.2525\ \text{L (unrounded)}$

Initial volume $= 25.6541\ \text{L} + 4.2525\ \text{L} = 29.9066\ \text{L (unrounded)}$

Final volume:

Volume $Al_2O_3 =$

$\left(50.0\ \text{kg}\ NH_4ClO_4\right)\left(\dfrac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\dfrac{1\ \text{mol}\ NH_4ClO_4}{117.49\ \text{g}\ NH_4ClO_4}\right)\left(\dfrac{1\ \text{mol}\ Al_2O_3}{3\ \text{mol}\ NH_4ClO_4}\right)\left(\dfrac{101.96\ \text{g}\ Al_2O_3}{1\ \text{mol}\ Al_2O_3}\right)\left(\dfrac{1\ \text{cc}}{3.97\ \text{g}\ Al_2O_3}\right)\left(\dfrac{1\ \text{mL}}{1\ \text{cc}}\right)\left(\dfrac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)$

$= 3.6432\ \text{L (unrounded)}$

Volume $AlCl_3 =$

$\left(50.0\ \text{kg}\ NH_4ClO_4\right)\left(\dfrac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\dfrac{1\ \text{mol}\ NH_4ClO_4}{117.49\ \text{g}\ NH_4ClO_4}\right)\left(\dfrac{1\ \text{mol}\ AlCl_3}{3\ \text{mol}\ NH_4ClO_4}\right)\left(\dfrac{133.33\ \text{g}\ AlCl_3}{1\ \text{mol}\ AlCl_3}\right)\left(\dfrac{1\ \text{cc}}{2.44\ \text{g}\ AlCl_3}\right)\left(\dfrac{1\ \text{mL}}{1\ \text{cc}}\right)\left(\dfrac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)$

$= 7.7515\ \text{L (unrounded)}$

V gas $= (1276.70\ \text{mol gas})(22.4\ \text{L / mol gas}) = 28598.08\ \text{L (unrounded)}$

Final volume $= 3.6432\ \text{L} + 7.7515\ \text{L} + 28598.08\ \text{L} = 28609.4747\ \text{L (unrounded)}$

Volume change $= (28609.4747\ \text{L}) - (29.9066\ \text{L}) = 28579.5681 = \mathbf{2.86\ x\ 10^4\ L}$

The volumes of all solids (before and after) are insignificant.

4.136   1) $Cu(s) + Br_2(aq) \rightarrow CuBr_2(aq)$
            $Cu(s) + Br_2(aq) \rightarrow Cu^{2+}(aq) + 2\ Br^-(aq)$
        2) $CuBr_2(aq) + 2\ NaOH(aq) \rightarrow Cu(OH)_2(s) + 2\ NaBr(aq)$
            $Cu^{2+}(aq) + 2\ OH^-(aq) \rightarrow Cu(OH)_2(s)$
        3) $Cu(OH)_2(s) \rightarrow CuO(s) + H_2O(l)$
            $Cu(OH)_2(s) \rightarrow CuO(s) + H_2O(l)$
        4) $CuO(s) + 2\ HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O(l)$
            $CuO(s) + 2\ H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$
        5) $3\ Cu(NO_3)_2(aq) + 2\ Na_3PO_4(aq) \rightarrow Cu_3(PO_4)_2(s) + 6\ NaNO_3(aq)$
            $3\ Cu^{2+}(aq) + 2\ PO_4^{3-}(aq) \rightarrow Cu_3(PO_4)_2(s)$
        6) $Cu_3(PO_4)_2(s) + 3\ H_2SO_4(aq) \rightarrow 3\ CuSO_4(aq) + 2\ H_3PO_4(aq)$
            $Cu_3(PO_4)_2(s) + 6\ H^+(aq) \rightarrow 3\ Cu^{2+}(aq) + 2\ H_3PO_4(aq)$
        7) $CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$
            $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$