CHAPTER 4 THREE MAJOR CLASSES OF CHEMICAL REACTIONS

END–OF–CHAPTER PROBLEMS

4.1 Plan: Review the discussion on the polar nature of water.
Solution:
Water is polar because the distribution of its bonding electrons is unequal, resulting in polar bonds, and the shape of the molecule (bent) is unsymmetrical.

4.2 Plan: Solutions that conduct an electric current contain electrolytes.
Solution:
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.

4.3 Plan: Review the discussion on ionic compounds in water.
Solution:
The ions on the surface of the solid attract the water molecules (cations attract the "negative" ends and anions attract the "positive" ends 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.

4.4 Plan: Recall that ionic compounds dissociate into their ions when dissolved in water. Examine the charges of the ions in each scene and the ratio of cations to anions.
Solution:
a) $CaCl_2$ dissociates to produce one Ca^{2+} ion for every two Cl^{-} ions. **Scene B** contains four 2+ ions and twice that number of 1– ions.
b) Li_2SO_4 dissociates to produce two Li^{+} ions for every one SO_4^{2-} ion. **Scene C** contains eight 1+ ions and half as many 2– ions.
c) NH_4Br dissociates to produce one NH_4^{+} ion for every one Br^{-} ion. **Scene A** contains equal numbers of 1+ and 1– ions.

4.5 Plan: Write the formula for magnesium nitrate and note the ratio of magnesium ions to nitrate ions.
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. **Scene B** best represents a volume of magnesium nitrate solution. Only Scene B has twice as many nitrate ions (red circles) as magnesium ions (blue circles).

4.6 Plan: Review the discussion of ionic compounds in water.
Solution:
In some ionic compounds, 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 compounds will be insoluble in water.

4.7 Plan: Review the discussion of covalent compounds in water.
Solution:
Some covalent compounds that contain the **hydrogen atom** dissociate into ions when dissolved in water. These compounds form **acidic** solutions in water; three examples are **HCl**, **HNO_3**, and **HBr**.

4.8 Plan: Compounds that are soluble in water tend to be ionic compounds or covalent compounds that have polar bonds. Many ionic compounds are soluble in water because the attractive force between the oppositely charged ions in an ionic compound are replaced with an attractive force between the polar water molecule and the ions when the compound is dissolved in water. Covalent compounds with polar bonds are often soluble in water since the polar bonds of the covalent compound interact with those in water.

Solution:
a) Benzene, a covalent compound, is likely to be **insoluble** in water because it is nonpolar and water is polar.
b) Sodium hydroxide ($NaOH$) is an ionic compound and is therefore likely to be **soluble** in water.
c) Ethanol (CH_3CH_2OH) will likely be **soluble** in water because it contains a polar –OH bond like water.
d) Potassium acetate ($KC_2H_3O_2$) is an ionic compound and will likely be **soluble** in water.

4.9 Plan: Compounds that are soluble in water tend to be ionic compounds or covalent compounds that have polar bonds. Many ionic compounds are soluble in water because the attractive force between the oppositely charged ions in an ionic compound are replaced with an attractive force between the polar water molecule and the ions when the compound is dissolved in water. Covalent compounds with polar bonds are often soluble in water since the polar bonds of the covalent compound interact with those in water.
Solution:
a) Lithium nitrate is an ionic compound and is expected to be **soluble** in water.
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 be expected to be **insoluble** 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**.

4.10 Plan: Substances whose aqueous solutions conduct an electric current are electrolytes such as ionic compounds, acids, and bases.
Solution:
a) Cesium bromide, CsBr, is a soluble ionic compound, and a solution of this salt in water contains Cs^+ and $Br^–$ ions. Its solution **conducts** an electric current.
b) HI is a strong acid that dissociates completely in water. Its aqueous solution contains H^+ and $I^–$ ions, so it **conducts** an electric current.

4.11 Plan: Substances whose aqueous solutions conduct an electric current are electrolytes such as ionic compounds, acids, and bases.
Solution:
a) Potassium sulfate, K_2SO_4, is an ionic compound that is soluble in water, producing K^+ and $SO_4^{2–}$ ions. Its solution **conducts** an electric current.
b) Sucrose is neither an ionic compound, an acid, nor a base, so it would be a nonelectrolyte (even though it's soluble in water). Its solution does **not conduct** an electric current.

4.12 Plan: To determine the total moles of ions released, write an equation that shows the compound dissociating into ions with the correct molar ratios. Convert mass and formula units to moles of compound and use the molar ratio to convert moles of compound to moles of ions.
Solution:
a) Each mole of K_3PO_4 forms 3 moles of K^+ ions and 1 mole of $PO_4^{3–}$ ions, or a total of 4 moles of ions:
$K_3PO_4(s) \rightarrow 3K^+(aq) + PO_4^{3–}(aq)$

$$\text{Moles of ions} = \left(0.75 \text{ mol } K_3PO_4\right)\left(\frac{4 \text{ mol ions}}{1 \text{ mol } K_3PO_4}\right) = \mathbf{3.0 \text{ mol of ions}}.$$

b) Each mole of $NiBr_2 \bullet 3H_2O$ forms 1mole of Ni^{2+} ions and 2 moles of $Br^–$ ions, or a total of 3 moles of ions:
$NiBr_2 \bullet 3H_2O(s) \rightarrow Ni^{2+}(aq) + 2Br^–(aq)$. The waters of hydration become part of the larger bulk of water. Convert mass to moles using the molar mass.

$$\text{Moles of ions} = \left(6.88 \text{ x } 10^{-3} \text{ g } NiBr_2 \bullet 3H_2O\right)\left(\frac{1 \text{ mol } NiBr_2 \bullet 3H_2O}{272.54 \text{ g } NiBr_2 \bullet 3H_2O}\right)\left(\frac{3 \text{ mol ions}}{1 \text{ mol } NiBr_2 \bullet 3H_2O}\right)$$
$$= 7.5732\text{x}10^{-5} = \mathbf{7.57x10^{-5} \text{ mol of ions}}$$

c) Each mole of $FeCl_3$ forms 1mole of Fe^{3+} ions and 3 moles of Cl^- ions, or a total of 4 moles of ions: $FeCl_3(s) \rightarrow Fe^{3+}(aq) + 3Cl^-(aq)$. Recall that a mole contains $6.022x10^{23}$ entities, so a mole of $FeCl_3$ contains $6.022x10^{23}$ units of $FeCl_3$, more easily expressed as formula units.

$$\text{Moles of ions} = \left(2.23x10^{22}\text{ FU FeCl}_3\right)\left(\frac{1\text{ mol FeCl}_3}{6.022x10^{23}\text{FU FeCl}_3}\right)\left(\frac{4\text{ mol ions}}{1\text{ mol FeCl}_3}\right)$$

$= 0.148124 =$ **0.148 mol of ions**

4.13 Plan: To determine the total moles of ions released, write an equation that shows the compound dissociating into ions with the correct molar ratios. Convert mass and formula units to moles of compound and use the molar ratio to convert moles of compound to moles of ions.

Solution:

a) Each mole of Na_2HPO_4 forms 2 moles of Na^+ ions and 1 mole of HPO_4^{2-} ions, or a total of 3 moles of ions: $Na_2HPO_4(s) \rightarrow 2Na^+(aq) + HPO_4^{2-}(aq)$.

$$\text{Moles of ions} = \left(0.734\text{ mol Na}_2\text{HPO}_4\right)\left(\frac{3\text{ mol ions}}{1\text{ mol Na}_2\text{HPO}_4}\right) = 2.202 = \mathbf{2.20\text{ mol of ions}}$$

b) Each mole of $CuSO_4•5H_2O$ forms 1 mole of Cu^{2+} ions and 1 mole of SO_4^{2-} ions, or a total of 2 moles of ions: $CuSO_4•5H_2O(s) \rightarrow Cu^{+2}(aq) + SO_4^{2-}(aq)$. The waters of hydration become part of the larger bulk of water. Convert mass to moles using the molar mass.

$$\text{Moles of ions} = \left(3.86\text{ g CuSO}_4•5\text{H}_2\text{O}\right)\left(\frac{1\text{ mol CuSO}_4•5\text{H}_2\text{O}}{249.70\text{ g CuSO}_4•5\text{H}_2\text{O}}\right)\left(\frac{2\text{ mol ions}}{1\text{ mol CuSO}_4•5\text{H}_2\text{O}}\right)$$

$= 3.0907x10^{-2} =$ **$3.09x10^{-2}$ mol of ions**

c) Each mole of $NiCl_2$ forms 1 mole of Ni^{2+} ions and 2 moles of Cl^- ions, or a total of 3 moles of ions: $NiCl_2(s) \rightarrow Ni^{2+}(aq) + 2Cl^-(aq)$. Recall that a mole contains $6.022x10^{23}$ entities, so a mole of $NiCl_2$ contains $6.022x10^{23}$ units of $NiCl_2$, more easily expressed as formula units.

$$\text{Moles of ions} = \left(8.66x10^{20}\text{FU NiCl}_2\right)\left(\frac{1\text{ mol NiCl}_2}{6.022x10^{23}\text{FU NiCl}_2}\right)\left(\frac{3\text{ mol ions}}{1\text{ mol NiCl}_2}\right)$$

$= 4.31418x10^{-3} =$ **$4.31x10^{-3}$ mol of ions**

4.14 Plan: To determine the total moles of ions released, write an equation that shows the compound dissociating into ions with the correct molar ratios. Convert the information given to moles of compound and use the molar ratio to convert moles of compound to moles of ions. Avogadro's number is used to convert moles of ions to numbers of ions.

Solution:

a) Each mole of $AlCl_3$ forms 1mole of Al^{3+} ions and 3 moles of Cl^- ions: $AlCl_3(s) \rightarrow Al^{3+}(aq) + 3Cl^-(aq)$. Molarity and volume must be converted to moles of $AlCl_3$.

$$\text{Moles of AlCl}_3 = \left(130.\text{ mL}\right)\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.45\text{ mol AlCl}_3}{\text{L}}\right) = 0.0585\text{ mol AlCl}_3$$

$$\text{Moles of Al}^{3+} = \left(0.0585\text{ mol AlCl}_3\right)\left(\frac{1\text{ mol Al}^{3+}}{1\text{ mol AlCl}_3}\right) = 0.0585 = \mathbf{0.058\text{ mol Al}^{3+}}$$

$$\text{Number of Al}^{3+}\text{ ions} = \left(0.0585\text{ mol Al}^{3+}\right)\left(\frac{6.022x10^{23}\text{ Al}^{3+}}{1\text{ mol Al}^{3+}}\right) = 3.52287x\,10^{22} = \mathbf{3.5x10^{22}\text{ Al}^{3+}\text{ ions}}$$

$$\text{Moles of Cl}^- = \left(0.0585\text{ mol AlCl}_3\right)\left(\frac{3\text{ mol Cl}^-}{1\text{ mol AlCl}_3}\right) = 0.1755 = \mathbf{0.18\text{ mol Cl}^-}$$

$$\text{Number of Cl}^-\text{ ions} = \left(0.1755\text{ mol Cl}^-\right)\left(\frac{6.022x10^{23}\text{ Cl}^-}{1\text{ mol Cl}^-}\right) = 1.05686x10^{23} = \mathbf{1.1x10^{23}\text{ Cl}^-\text{ ions}}$$

b) Each mole of Li_2SO_4 forms 2 moles of Li^+ ions and 1 mole of SO_4^{2-} ions: $Li_2SO_4(s) \rightarrow 2Li^+(aq) + SO_4^{2-}(aq)$.

$$\text{Moles of } Li_2SO_4 = (9.80\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{2.59\text{ g }Li_2SO_4}{1\text{ L}}\right)\left(\frac{1\text{ mol }Li_2SO_4}{109.95\text{ g }Li_2SO_4}\right) = 2.3085\text{x}10^{-4}\text{ mol }Li_2SO_4$$

$$\text{Moles of } Li^+ = (2.3085\text{x}10^{-4}\text{ mol }Li_2SO_4)\left(\frac{2\text{ mol }Li^+}{1\text{ mol }Li_2SO_4}\right) = 4.6170\text{x}10^{-4} = \mathbf{4.62x10^{-4}\text{ mol }Li^+}$$

$$\text{Number of } Li^+ \text{ ions} = (4.6170\text{x}10^{-4}\text{ mol }Li^+)\left(\frac{6.022\text{ x }10^{23}\,Li^+}{1\text{ mol }Li^+}\right) = 2.7804\text{x}10^{20} = \mathbf{2.78x10^{20}\ Li^+\text{ ions}}$$

$$\text{Moles of } SO_4^{2-} = (2.3085\text{x}10^{-4}\text{ mol }Li_2SO_4)\left(\frac{1\text{ mol }SO_4^{2-}}{1\text{ mol }Li_2SO_4}\right) = 2.3085\text{x}10^{-4} = \mathbf{2.31x10^{-4}\text{ mol }SO_4^{2-}}$$

$$\text{Number of } SO_4^{2-} \text{ ions} = (2.3085\text{x}10^{-4}\text{ mol }SO_4^{2-})\left(\frac{6.022\text{ x }10^{23}\,SO_4^{2-}}{1\text{ mol }SO_4^{2-}}\right)$$
$$= 1.39018\text{x}10^{20} = \mathbf{1.39x10^{20}\ SO_4^{2-}\text{ ions}}$$

c) Each mole of KBr forms 1 mole of K^+ ions and 1 mole of Br^- ions: $KBr(s) \rightarrow K^+(aq) + Br^-(aq)$.

$$\text{Moles of KBr} = (245\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{3.68\text{x}10^{22}\text{ FU KBr}}{\text{L}}\right)\left(\frac{1\text{ mol KBr}}{6.022\text{x}10^{23}\text{ FU KBr}}\right) = 0.01497\text{ mol KBr}$$

$$\text{Moles of } K^+ = (0.01497\text{ mol KBr})\left(\frac{1\text{ mol }K^+}{1\text{ mol KBr}}\right) = 0.01497 = \mathbf{1.50x10^{-2}\text{ mol }K^+}$$

$$\text{Number of } K^+ \text{ ions} = (0.01497\text{ mol }K^+)\left(\frac{6.022\text{x}10^{23}\,K^+}{1\text{ mol }K^+}\right) = 9.016\text{x}10^{21} = \mathbf{9.02x10^{21}\ K^+\text{ ions}}$$

$$\text{Moles of } Br^- = (0.01497\text{ mol KBr})\left(\frac{1\text{ mol }Br^-}{1\text{ mol KBr}}\right) = 0.01497 = \mathbf{1.50x10^{-2}\text{ mol }Br^-}$$

$$\text{Number of } Br^- \text{ ions} = (0.01497\text{ mol }Br^-)\left(\frac{6.022\text{x}10^{23}\,Br^-}{1\text{ mol }Br^-}\right) = 9.016\text{x}10^{21} = \mathbf{9.02x10^{21}\ Br^-\text{ ions}}$$

4.15 Plan: To determine the total moles of ions released, write an equation that shows the compound dissociating into ions with the correct molar ratios. Convert the information given to moles of compound and use the molar ratio to convert moles of compound to moles of ions. Avogadro's number is used to convert moles of ions to numbers of ions.

Solution:

a) Each mole of $MgCl_2$ forms 1 mole of Mg^{2+} ions and 2 moles of Cl^- ions: $MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$.

$$\text{Moles of } MgCl_2 = (88.\text{mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{1.75\text{ mol }MgCl_2}{\text{L}}\right) = 0.154\text{ mol }MgCl_2$$

$$\text{Moles of } Mg^{2+} = (0.154\text{ mol }MgCl_2)\left(\frac{1\text{ mol }Mg^{2+}}{1\text{ mol }MgCl_2}\right) = 0.154 = \mathbf{0.15\text{ mol }Mg^{2+}}$$

$$\text{Number of } Mg^{2+} \text{ ions} = (0.154\text{ mol }Mg^{2+})\left(\frac{6.022\text{x}10^{23}\,Mg^{2+}}{1\text{ mol }Mg^{2+}}\right) = 9.27388\text{x}10^{22} = \mathbf{9.3x10^{22}\ Mg^{2+}\text{ ions}}$$

$$\text{Moles of } Cl^- = (0.154\text{ mol }MgCl_2)\left(\frac{2\text{ mol }Cl^-}{1\text{ mol }MgCl_2}\right) = 0.308 = \mathbf{0.31\text{ mol }Cl^-}$$

Number of Cl^- ions = $(0.308 \text{ mol } Cl^-)\left(\frac{6.022\text{x}10^{23}\ Cl^-}{1 \text{ mol } Cl^-}\right)$ = 1.854776x10^{23} = **1.9x10^{23} Cl^- ions**

b) Each mole of $Al_2(SO_4)_3$ forms 2 moles of Al^{3+} ions and 3 moles of SO_4^{2-} ions:
$Al_2(SO_4)_3(s) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$.

Moles of $Al_2(SO_4)_3$ = $(321 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right)\left(\frac{0.22 \text{ g } Al_2(SO_4)_3}{1 \text{ L}}\right)\left(\frac{1 \text{ mol } Al_2(SO_4)_3}{342.17 \text{ g } Al_2(SO_4)_3}\right)$

= 2.06389x10^{-4} mol $Al_2(SO_4)_3$

Moles of Al^{3+} = $(2.06389\text{x}10^{-4} \text{ mol } Al_2(SO_4)_3)\left(\frac{2 \text{ mol } Al^{3+}}{1 \text{ mol } Al_2(SO_4)_3}\right)$ = 4.12777x10^{-4} = **4.1x10^{-4} mol Al^{3+}**

Number of Al^{3+} ions = $(4.12777\text{x}10^{-4} \text{ mol } Al^{3+})\left(\frac{6.022\text{x}10^{23}\ Al^{3+}}{1 \text{ mol } Al^{3+}}\right)$ = 2.4857x10^{20} = **2.5x10^{20} Al^{3+} ions**

Moles of SO_4^{2-} = $(2.06389\text{x}10^{-4} \text{ mol } Al_2(SO_4)_3)\left(\frac{3 \text{ mol } SO_4^{2-}}{1 \text{ mol } Al_2(SO_4)_3}\right)$ = 6.191659x10^{-4} = **6.2x10^{-4} mol SO_4^{2}**

Number of SO_4^{2-} ions = $(6.191659 \text{ x } 10^{-4} \text{ mol } SO_4^{2-})\left(\frac{6.022 \text{ x } 10^{23}\ SO_4^{2-}}{1 \text{ mol } SO_4^{2-}}\right)$

= 3.7286x10^{20} = **3.7x10^{20} SO_4^{2-} ions**

c) Each mole of $CsNO_3$ forms 1 mole of Cs^+ ions and 1 mole of NO_3^- ions: $CsNO_3(s) \rightarrow Cs^+(aq) + NO_3^-(aq)$

Moles of $CsNO_3$ = $(1.65 \text{ L})\left(\frac{8.83\text{x}10^{21} \text{ FU } CsNO_3}{\text{L}}\right)\left(\frac{1 \text{ mol } CsNO_3}{6.022\text{x}10^{23} \text{ FU } CsNO_3}\right)$ = 0.024194 mol $CsNO_3$

Moles of Cs^+ = $(0.024194 \text{ mol } CsNO_3)\left(\frac{1 \text{ mol } Cs^+}{1 \text{ mol } CsNO_3}\right)$ = 0.024194 = **0.0242 mol Cs^+**

Number of Cs^+ ions = $(0.024194 \text{ mol } Cs^+)\left(\frac{6.022\text{x}10^{23}\ Cs^+}{1 \text{ mol } Cs^+}\right)$ = 1.45695x10^{22} = **1.46x10^{22} Cs^+ ions**

Moles of NO_3^- = $(0.024194 \text{ mol } CsNO_3)\left(\frac{1 \text{ mol } NO_3^-}{1 \text{ mol } CsNO_3}\right)$ = 0.024194 = **0.0242 mol NO_3^-**

Number of NO_3^- ions = $(0.024194 \text{ mol } NO_3^-)\left(\frac{6.022\text{x}10^{23}\ NO_3^-}{1 \text{ mol } NO_3^-}\right)$ = 1.45695x10^{22} = **1.46x10^{22} NO_3^- ions**

4.16 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. Calculate the moles of acid by multiplying the molarity (moles/L) by the volume in liters.

Solution:

a) $HClO_4(aq) \rightarrow H^+(aq) + ClO_4^-(aq)$

Moles H^+ = mol $HClO_4$ = $(1.40 \text{ L})\left(\frac{0.25 \text{ mol}}{1 \text{ L}}\right)$ = **0.35 mol H^+**

b) $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

Moles H^+ = mol HNO_3 = $(6.8 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right)\left(\frac{0.92 \text{ mol}}{1 \text{ L}}\right)$ = 6.256x10^{-3} = **6.3x10^{-3} mol H^+**

c) $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

Moles H^+ = mol HCl = $(2.6 \text{ L})\left(\frac{0.085 \text{ mol}}{1 \text{ L}}\right)$ = 0.221 = **0.22 mol H^+**

4.17 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 HBr, HI, and HNO_3 each produce one mole of H^+ upon dissociation, so moles H^+ = moles acid. Calculate the moles of acid by multiplying the molarity (moles/L) by the volume in liters.

Solution:

a) $HBr(aq) \rightarrow H^+(aq) + Br^-(aq)$

$$\text{Moles } H^+ = \text{mol HBr} = (1.4\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.75\text{ mol}}{1\text{ L}}\right) = 1.05\text{x}10^{-3} = \mathbf{1.0x10^{-3}\ mol\ H^+}$$

b) $HI(aq) \rightarrow H^+(aq) + I^-(aq)$

$$\text{Moles } H^+ = \text{mol HI} = (2.47\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{1.98\text{ mol}}{1\text{ L}}\right) = 4.8906\text{x}10^{-3} = \mathbf{4.89x10^{-3}\ mol\ H^+}$$

c) $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

$$\text{Moles } H^+ = \text{mol } HNO_3 = (395\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.270\text{ mol}}{1\text{ L}}\right) = 0.10665 = \mathbf{0.107\ mol\ H^+}$$

4.18 Plan: Convert the mass of the seawater in kg to g and use the density to convert the mass of the seawater to volume in L. Convert mass of each compound to moles of compound and then use the molar ratio in the dissociation of the compound to find the moles of each ion. The molarity of each ion is the moles of ion divided by the volume of the seawater. To find the total molarity of the alkali metal ions [Group 1A(1)], add the moles of the alkali metal ions and divide by the volume of the seawater. Perform the same calculation to find the total molarity of the alkaline earth metal ions [Group 2A(2)] and the anions (the negatively charged ions).

Solution:

a) The volume of the seawater is needed.

$$\text{Volume (L) of seawater} = (1.00\text{ kg})\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{\text{cm}^3}{1.025\text{ g}}\right)\left(\frac{1\text{ mL}}{1\text{ cm}^3}\right)\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right) = 0.97560976\text{ L}$$

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

NaCl:

Each mole of NaCl forms 1 mole of Na^+ ions and 1 mole of Cl^- ions: $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$

$$\text{Moles of NaCl} = (26.5\text{ g NaCl})\left(\frac{1\text{ mol NaCl}}{58.44\text{ g NaCl}}\right) = 0.4534565\text{ mol NaCl}$$

$$\text{Moles of } Na^+ = (0.4534565\text{ mol NaCl})\left(\frac{1\text{ mol } Na^+}{1\text{ mol NaCl}}\right) = 0.4534565\text{ mol } Na^+$$

$$\text{Moles of } Cl^- = (0.4534565\text{ mol NaCl})\left(\frac{1\text{ mol } Cl^-}{1\text{ mol NaCl}}\right) = 0.4534565\text{ mol } Cl^-$$

$MgCl_2$:

Each mole of $MgCl_2$ forms 1 mole of Mg^{2+} ions and 2 moles of Cl^- ions: $MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$

$$\text{Moles of } MgCl_2 = (2.40\text{ g } MgCl_2)\left(\frac{1\text{ mol } MgCl_2}{95.21\text{ g } MgCl_2}\right) = 0.025207\text{ mol } MgCl_2$$

$$\text{Moles of } Mg^{2+} = (0.025207\text{ mol } MgCl_2)\left(\frac{1\text{ mol } Mg^{2+}}{1\text{ mol } MgCl_2}\right) = 0.025207\text{ mol } Mg^{2+}$$

$$\text{Moles of } Cl^- = (0.025207\text{ mol } MgCl_2)\left(\frac{2\text{ mol } Cl^-}{1\text{ mol } MgCl_2}\right) = 0.050415\text{ mol } Cl^-$$

$MgSO_4$:

Each mole of $MgSO_4$ forms 1 mole of Mg^{2+} ions and 1 mole of SO_4^{2-} ions: $MgSO_4(s) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq)$

Moles of $MgSO_4$ = $(3.35 \text{ g } MgSO_4)\left(\frac{1 \text{ mol } MgSO_4}{120.38 \text{ g } MgSO_4}\right)$ = 0.0278285 mol $MgSO_4$

Moles of Mg^{2+} = $(0.0278285 \text{ mol } MgSO_4)\left(\frac{1 \text{ mol } Mg^{2+}}{1 \text{ mol } MgSO_4}\right)$ = 0.0278285 mol Mg^{2+}

Moles of SO_4^{2-} = $(0.0278285 \text{ mol } MgSO_4)\left(\frac{1 \text{ mol } SO_4^{2-}}{1 \text{ mol } MgSO_4}\right)$ = 0.0278285 mol SO_4^{2-}

$CaCl_2$:
Each mole of $CaCl_2$ forms 1 mole of Ca^{2+} ions and 2 moles of Cl^- ions: $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$

Moles of $CaCl_2$ = $(1.20 \text{ g } CaCl_2)\left(\frac{1 \text{ mol } CaCl_2}{110.98 \text{ g } CaCl_2}\right)\left(\frac{1 \text{ mol } Ca^{2+}}{1 \text{ mol } CaCl_2}\right)$ = 0.0108128 mol $CaCl_2$

Moles of Ca^{2+} = $(0.0108128 \text{ mol } CaCl_2)\left(\frac{1 \text{ mol } Ca^{2+}}{1 \text{ mol } CaCl_2}\right)$ = 0.0108128 mol Ca^{2+}

Moles of Cl^- = $(0.0108128 \text{ mol } CaCl_2)\left(\frac{2 \text{ mol } Cl^-}{1 \text{ mol } CaCl_2}\right)$ = 0.0216255 mol Cl^-

KCl:
Each mole of KCl forms 1 mole of K^+ ions and 1 mole of Cl^- ions: $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$

Moles of KCl = $(1.05 \text{ g KCl})\left(\frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}}\right)$ = 0.0140845 mol KCl

Moles of K^+ = $(0.0140845 \text{ mol KCl})\left(\frac{1 \text{ mol } K^+}{1 \text{ mol KCl}}\right)$ = 0.0140845 mol K^+

Moles of Cl^- = $(0.0140845 \text{ mol KCl})\left(\frac{1 \text{ mol } Cl^-}{1 \text{ mol KCl}}\right)$ = 0.0140845 mol Cl^-

$NaHCO_3$:
Each mole of $NaHCO_3$ forms 1 mole of Na^+ ions and 1 mole of HCO_3^- ions: $NaHCO_3(s) \rightarrow Na^+(aq) + HCO_3^-(aq)$

Moles of $NaHCO_3$ = $(0.315 \text{ g } NaHCO_3)\left(\frac{1 \text{ mol } NaHCO_3}{84.01 \text{ g } NaHCO_3}\right)$ = 0.00374955 mol $NaHCO_3$

Moles of Na^+ = $(0.00374955 \text{ mol } NaHCO_3)\left(\frac{1 \text{ mol } Na^+}{1 \text{ mol } NaHCO_3}\right)$ = 0.00374955 mol Na^+

Moles of HCO_3^- = $(0.00374955 \text{ mol } NaHCO_3)\left(\frac{1 \text{ mol } HCO_3^-}{1 \text{ mol } NaHCO_3}\right)$ = 0.00374955 mol HCO_3^-

NaBr
Each mole of NaBr forms 1 mole of Na^+ ions and 1 mole of Br^- ions: $NaBr(s) \rightarrow Na^+(aq) + Br^-(aq)$

Moles of NaBr = $(0.098 \text{ g NaBr})\left(\frac{1 \text{ mol NaBr}}{102.89 \text{ g NaBr}}\right)$ = 0.0009524735 mol NaBr

Moles of Na^+ = $(0.0009524735 \text{ mol NaBr})\left(\frac{1 \text{ mol } Na^+}{1 \text{ mol NaBr}}\right)$ = 0.0009524735 mol Na^+

Moles of Br^- = $(0.0009524735 \text{ mol NaBr})\left(\frac{1 \text{ mol } Br^-}{1 \text{ mol NaBr}}\right)$ = 0.0009524735 mol Br^-

Total moles of each ion:

Cl^-: 0.4534565 + 0.050415 + 0.0216255 + 0.0140845 = 0.5395815 mol Cl^-
Na^+: 0.4534565 + 0.00374955 + 0.0009524735 = 0.458158523 mol Na^+
Mg^{2+}: 0.025207 + 0.0278285 = 0.0530355 mol Mg^{2+}
SO_4^{2-}: 0.0278285 mol SO_4^{2-}
Ca^{2+}: 0.0108128 mol Ca^{2+}
K^+: 0.0140845 mol K^+
HCO_3^-: 0.00374955 mol HCO_3^-
Br^-: 0.0009524735 mol Br^-

Dividing each of the numbers of moles by the volume (0.97560976 L) and rounding to the proper number of significant figures gives the molarities.

$$M = \frac{\text{mol}}{\text{L}}$$

$$M\ Cl^- = \frac{0.5395815 \text{ mol } Cl^-}{0.97560976 \text{ L}} = 0.55307 = \mathbf{0.553\ \mathit{M}\ Cl^-}$$

$$M\ Na^+ = \frac{0.45815823 \text{ mol } Na^+}{0.97560976 \text{ L}} = 0.469612 = \mathbf{0.470\ \mathit{M}\ Na^+}$$

$$M\ Mg^{2+} = \frac{0.0530355 \text{ mol } Mg^{2+}}{0.97560976 \text{ L}} = 0.054361 = \mathbf{0.0544\ \mathit{M}\ Mg^{2+}}$$

$$M\ SO_4^{2-} = \frac{0.0278285 \text{ mol } SO_4^{2-}}{0.97560976 \text{ L}} = 0.028524 = \mathbf{0.0285\ \mathit{M}\ SO_4^{2-}}$$

$$M\ Ca^{2+} = \frac{0.0108128 \text{ mol } Ca^{2+}}{0.97560976 \text{ L}} = 0.011083 = \mathbf{0.0111\ \mathit{M}\ Ca^{2+}}$$

$$M\ K^+ = \frac{0.0140845 \text{ mol } K^+}{0.97560976 \text{ L}} = 0.014437 = \mathbf{0.0144\ \mathit{M}\ K^+}$$

$$M\ HCO_3^- = \frac{0.00374955 \text{ mol } HCO_3^-}{0.97560976 \text{ L}} = 0.003843 = \mathbf{0.00384\ \mathit{M}\ HCO_3^-}$$

$$M\ Br^- = \frac{0.0009524735 \text{ mol } Br^-}{0.97560976 \text{ L}} = 0.0009763 = \mathbf{0.00098\ \mathit{M}\ Br^-}$$

b) The alkali metal cations are Na^+ and K^+. Add the molarities of the individual ions.

0.469612 M Na^+ + 0.014437 M K^+ = 0.484049 = **0.484 *M*** total for alkali metal cations

c) The alkaline earth metal cations are Mg^{2+} and Ca^{2+}. Add the molarities of the individual ions.

0.054361 M Mg^{2+} + 0.011083 M Ca^{2+} = 0.065444 = **0.0654 *M*** total for alkaline earth cations

d) The anions are Cl^-, SO_4^{2-}, HCO_3^-, and Br^-. Add the molarities of the individual ions.

0.55307 M Cl^- + 0.028524 M SO_4^{2-} + 0.003843 M HCO_3^- + 0.0009763 M Br^-
= 0.5864133 = **0.586 *M*** total for anions

4.19 Plan: Use the molarity and volume of the ions to find the moles of each ion. Multiply the moles of each ion by that ion's charge to find the total moles of charge. Since sodium ions have a +1 charge, the total moles of charge equals the moles of sodium ions.

Solution:

$$\text{Moles of } Ca^{2+} = \left(1.0 \times 10^3 \text{ L}\right)\left(\frac{0.015 \text{ mol } Ca^{2+}}{\text{L}}\right) = 15 \text{ mol } Ca^{2+}$$

$$\text{Moles of charge from } Ca^{2+} = \left(15 \text{ mol } Ca^{2+}\right)\left(\frac{2 \text{ mol charge}}{1 \text{ mol } Ca^{2+}}\right) = 30. \text{ mol charge from } Ca^{2+}$$

$$\text{Moles of } Fe^{3+} = \left(1.0\text{x}10^3 \text{ L}\right)\left(\frac{0.0010 \text{ mol } Fe^{3+}}{\text{L}}\right) = 1.0 \text{ mol } Fe^{3+}$$

Moles of charge from Fe^{3+} = $\left(1.0 \text{ mol } Fe^{3+}\right)\left(\frac{3 \text{ mol charge}}{1 \text{ mol } Fe^{3+}}\right)$ = 3.0 mol charge from Fe^{3+}

Total moles of charge = 30. mol + 3.0 mol = 33 mol charge

Moles Na^+ = $\left(33 \text{ mol charge}\right)\left(\frac{1 \text{ mol } Na^+}{1 \text{ mol charge}}\right)$ = **33 mol Na^+**

4.20 Plan: Write the total ionic and net ionic equations for the reaction given. The total ionic equation shows all soluble ionic substances dissociated into ions. The net ionic equation eliminates the spectator ions. New equations may be written by replacing the spectator ions in the given equation by other spectator ions.
Solution:
The reaction given has the following total ionic and net ionic equations:
Total ionic equation: $Ba^{2+}(aq) + \underline{2NO_3}^-(aq) + \underline{2Na^+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + \underline{2Na^+}(aq) + \underline{2NO_3}^-(aq)$
The spectator ions are underlined and are omitted:
Net ionic equation: $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$
New equations will contain a soluble barium compound and a soluble carbonate compound.
The "new" equations are:
Molecular: $BaCl_2(aq) + K_2CO_3(aq) \rightarrow BaCO_3(s) + 2KCl(aq)$
Total ionic: $Ba^{2+}(aq) + 2Cl^-(aq) + 2K^+(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + 2K^+(aq) + 2Cl^-(aq)$
Molecular: $BaBr_2(aq) + (NH_4)_2CO_3(aq) \rightarrow BaCO_3(s) + 2NH_4Br(aq)$
Total ionic: $Ba^{2+}(aq) + 2Br^-(aq) + 2NH_4^+(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + 2NH_4^+(aq) + 2Br^-(aq)$

4.21 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.

4.22 Plan: Write the new cation-anion combinations as the products of the reaction and use the solubility rules to determine if any of the new combinations are insoluble. The spectator ions are the ions that are present in the soluble ionic compound.
Solution:
a) $Ca(NO_3)_2(aq) + 2NaCl(aq) \rightarrow CaCl_2(aq) + 2NaNO_3(aq)$
Since the possible products ($CaCl_2$ and $NaNO_3$) are both soluble, no reaction would take place.
b) $2KCl(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbCl_2(s)$
According to the solubility rules, KNO_3 is soluble but $PbCl_2$ is insoluble so a precipitation reaction takes place. The K^+ and NO_3^- would be spectator ions, because their salt is soluble.

4.23 Plan: Use the solubility rules to predict the products of this reaction. Ions not involved in the precipitate are spectator ions and are not included in the net ionic equation.
Solution:
Assuming that the left beaker is $AgNO_3$ (because it has gray Ag^+ ions) and the right must be NaCl, then the NO_3^- is blue, the Na^+ is brown, and the Cl^- is green. (Cl^- must be green since it is present with Ag^+ in the precipitate in the beaker on the right.)
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: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

4.24 Plan: Write the new cation-anion combinations as the products of the reaction and use the solubility rules to determine if any of the new combinations are insoluble. The total ionic equation shows all soluble ionic substances dissociated into ions. The spectator ions are the ions that are present in the soluble ionic compound. The spectator ions are omitted from the net ionic equation.
Solution:
a) Molecular: $Hg_2(NO_3)_2(aq) + 2KI(aq) \rightarrow Hg_2I_2(s) + 2KNO_3(aq)$
Total ionic: $Hg_2^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow Hg_2I_2(s) + 2K^+(aq) + 2NO_3^-(aq)$
Net ionic: $Hg_2^{2+}(aq) + 2I^-(aq) \rightarrow Hg_2I_2(s)$
Spectator ions are K^+ and NO_3^-.

b) Molecular: $FeSO_4(aq) + Sr(OH)_2(aq) \rightarrow Fe(OH)_2(s) + SrSO_4(s)$

Total ionic: $Fe^{2+}(aq) + SO_4^{2-}(aq) + Sr^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s) + SrSO_4(s)$

Net ionic: This is the same as the total ionic equation because there are no spectator ions.

4.25 Plan: Write the new cation-anion combinations as the products of the reaction and use the solubility rules to determine if any of the new combinations are insoluble. The total ionic equation shows all soluble ionic substances dissociated into ions. The spectator ions are the ions that are present in the soluble ionic compound. The spectator ions are omitted from the net ionic equation.

Solution:

a) Molecular: $3CaCl_2(aq) + 2Cs_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6CsCl(aq)$

Total ionic: $3Ca^{2+}(aq) + 6Cl^-(aq) + 6Cs^+(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s) + 6Cs^+(aq) + 6Cl^-(aq)$

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: $2Na^+(aq) + S^{2-}(aq) + Zn^{2+}(aq) + SO_4^{2-}(aq) \rightarrow ZnS(s) + 2Na^+(aq) + SO_4^{2-}(aq)$

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

Spectator ions are Na^+ and SO_4^{2-}.

4.26 Plan: 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. Any ions not involved in a precipitate are spectator ions and are omitted from the net ionic equation.

Solution:

a) $NaNO_3(aq) + CuSO_4(aq) \rightarrow Na_2SO_4(aq) + Cu(NO_3)_2(aq)$

No precipitate will form. The ions Na^+ and SO_4^{2-} will not form an insoluble salt according to the first solubility rule which states that 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. There is no reaction.

b) A precipitate will form because silver ions, Ag^+, and bromide ions, Br^-, will combine to form a solid salt, silver bromide, AgBr. The ammonium and nitrate ions do not form a precipitate.

Molecular: $NH_4Br(aq) + AgNO_3(aq) \rightarrow AgBr(s) + NH_4NO_3(aq)$

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

Net ionic: $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$

4.27 Plan: 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. Any ions not involved in a precipitate are spectator ions and are omitted from the net ionic equation.

Solution:

a) Barium carbonate ($BaCO_3$) precipitates since the solubility rules state that all common carbonates are insoluble.

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

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

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

b) Aluminum phosphate ($AlPO_4$) precipitates since most common phosphates are insoluble; the sodium nitrate is soluble.

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

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

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

4.28 Plan: 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 $Pb(NO_3)_2$ are required to produce 0.628 g of PbI_2. The molarity is calculated by dividing moles of $Pb(NO_3)_2$ by its volume in liters.

Solution:

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

Moles of $Pb(NO_3)_2$ = $\left(0.628 \text{ g } PbI_2\right)\left(\frac{1 \text{ mol } PbI_2}{461.0 \text{ g } PbI_2}\right)\left(\frac{1 \text{ mol } Pb(NO_3)_2}{1 \text{ mol } PbI_2}\right)$ = 0.001362256 mol $Pb(NO_3)_2$

Moles of Pb^{2+} = moles of $Pb(NO_3)_2$ = 0.001362256 mol Pb^{2+}

Molarity of Pb^{2+} = $\frac{\text{moles } Pb^{2+}}{\text{volume of } Pb^{2+}} = \frac{0.001362256 \text{ mol}}{38.5 \text{ mL}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$ = 0.035383 = **0.0354 *M* Pb^{2+}**

4.29 Plan: Write a balanced equation for the chemical reaction described in the problem. By applying the solubility rules to the two possible products (KNO_3 and AgCl), determine that AgCl is the precipitate. By using molar relationships, determine how many moles of $AgNO_3$ are required to produce 0.842 g of AgCl. The molarity is calculated by dividing moles of $AgNO_3$ by its volume in liters.

Solution:

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

Moles of $AgNO_3$ = $\left(0.842 \text{ g AgCl}\right)\left(\frac{1 \text{ mol AgCl}}{143.4 \text{ g AgCl}}\right)\left(\frac{1 \text{ mol } AgNO_3}{1 \text{ mol AgCl}}\right)$ = 0.0058717 mol $AgNO_3$

Moles of Ag^+ = moles of $AgNO_3$ = 0.0058717 mol Ag^+

Molarity of Ag^+ = $\frac{\text{moles } Ag^+}{\text{volume of } Ag^+} = \frac{0.0058717 \text{ mol}}{25.0 \text{ mL}}\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$ = 0.2348675 = **0.235 *M* Ag^+**

4.30 Plan: 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. Any ions not involved in a precipitate are spectator ions and are omitted from the net ionic equation. Use the molar ratio in the balanced net ionic equation to calculate the mass of product.

Solution:

a) The yellow spheres cannot be ClO_4^- or NO_3^- as these ions form only soluble compounds. So the yellow sphere must be SO_4^{2-}. The only sulfate compounds possible that would be insoluble are Ag_2SO_4 and $PbSO_4$. The precipitate has a 1:1 ratio between its ions. Ag_2SO_4 has a 2:1 ratio between its ions. Therefore the blue spheres are Pb^{2+} and the yellow spheres are SO_4^{2-}. The precipitate is thus **$PbSO_4$**.

b) The net ionic equation is $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$.

c) Mass (g) of $PbSO_4$ = $\left(10 \text{ } Pb^{2+} \text{ spheres}\right)\left(\frac{5.0\text{x}10^{-4} \text{ mol } Pb^{2+}}{1 \text{ } Pb^{2+} \text{ sphere}}\right)\left(\frac{1 \text{ mol } PbSO_4}{1 \text{ mol } Pb^{2+}}\right)\left(\frac{303.3 \text{ g } PbSO_4}{1 \text{ mol } PbSO_4}\right)$

= 1.5165 = **1.5 g $PbSO_4$**

4.31 Plan: 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. Any ions not involved in a precipitate are spectator ions and are omitted from the net ionic equation. Use the molar ratio in the balanced net ionic equation to calculate the mass of product.

Solution:

a) There are 9 purple spheres representing cations and 7 green spheres representing anions. In the precipitate, there are 8 purple spheres (cations) and 4 green spheres (anions), indicating a 2:1 ratio between cation and anion in the compound. Only Reaction 3 produces a precipitate (Ag_2SO_4) fitting this description:

$Li_2SO_4(aq) + 2AgNO_3(aq) \rightarrow 2LiNO_3(aq) + Ag_2SO_4(s)$

Reaction 1 does not produce a precipitate since all common nitrate and chloride compounds are soluble. Reaction 2 does not produce a precipitate since all common perchlorate and chloride compounds are soluble. Reaction 4 produces a precipitate, $PbBr_2$, but it has a cation:anion ratio of 1:2, instead of 2:1.

Total ionic equation for Reaction 3 =

$$2Li^+(aq) + SO_4^{2-}(aq) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow 2Li^+(aq) + 2NO_3^-(aq) + Ag_2SO_4(s)$$

Net ionic equation = $2Ag^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s)$

b) There are 4 unreacted spheres of ions.

Number of ions = $\left(4 \text{ spheres}\right)\left(\frac{2.5\text{x}10^{-3} \text{ mol ions}}{1 \text{ sphere}}\right)\left(\frac{6.022\text{x}10^{23} \text{ ions}}{1 \text{ mol ions}}\right)$ = $6.022\text{x}10^{21}$ = **$6.0\text{x}10^{21}$ ions**

c) Mass (g) of solid =

$$\left(4 \text{ spheres of } SO_4^{2-} \text{ ions}\right)\left(\frac{2.5x10^{-3} \text{ mol } SO_4^{2-} \text{ ions}}{1 \text{ sphere}}\right)\left(\frac{1 \text{ mol } Ag_2SO_4}{1 \text{ mol } SO_4^{2-}}\right)\left(\frac{311.9 \text{ g } Ag_2SO_4}{1 \text{ mol } Ag_2SO_4}\right)$$

$= 3.119 =$ **3.1 g solid**

4.32 Plan: Write a balanced equation for the reaction. Find the moles of $AgNO_3$ by multiplying the molarity and volume of the $AgNO_3$ solution; use the molar ratio in the balanced equation to find the moles of Cl^- present in the 25.00 mL sample. Then, 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 the mass of Cl^- by the mass of the sample volume and multiplying by 100.

Solution:

The balanced equation is $AgNO_3(aq) + Cl^-(aq) \rightarrow AgCl(s) + NO_3^-(aq)$.

$$\text{Moles of } AgNO_3 = \left(53.63 \text{ mL}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.2970 \text{ mol } AgNO_3}{\text{L}}\right) = 0.01592811 \text{ mol } AgNO_3$$

$$\text{Mass (g) of } Cl^- = \left(0.01592811 \text{ mol } AgNO_3\right)\left(\frac{1 \text{ mol } Cl^-}{1 \text{ mol } AgNO_3}\right)\left(\frac{35.45 \text{ g Cl}}{1 \text{ mol } Cl^-}\right) = 0.56465 \text{ g } Cl^-$$

$$\text{Mass (g) of seawater sample} = \left(25.00 \text{ mL}\right)\left(\frac{1.024 \text{ g}}{\text{mL}}\right) = 25.60 \text{ g sample}$$

$$\text{Mass \% } Cl^- = \frac{\text{mass } Cl^-}{\text{mass sample}} \text{ x } 100\% = \frac{0.56465 \text{ g } Cl^-}{25.60 \text{ g sample}} \text{ x } 100\% = 2.20566 = \mathbf{2.206\% \ Cl^-}$$

4.33 Plan: Write the reaction between aluminum sulfate and sodium hydroxide and check the solubility rules to determine the precipitate. Spectator ions are omitted from the net ionic equation. Find the moles of sodium hydroxide by multiplying its molarity by its volume in liters; find the moles of aluminum sulfate by converting grams per liter to moles per liter and multiplying by the volume of that solution. To determine which reactant is limiting, calculate the amount of precipitate formed from each reactant, assuming an excess of the other reactant, using the molar ratio from the balanced equation. The smaller amount of precipitate is the answer.

Solution:

a) According to the solubility rules, most common sulfate compounds are soluble, but most common hydroxides are insoluble. Aluminum hydroxide is the precipitate.

Total ionic equation: $Al_2(SO_4)_3(aq) + 6NaOH(aq) \rightarrow 3Na_2SO_4(aq) + 2Al(OH)_3(s)$

Net ionic equation: $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

$$\text{b) Moles of } Al_2(SO_4)_3 = \left(627 \text{ mL}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{15.8 \text{ g } Al_2(SO_4)_3}{\text{L}}\right)\left(\frac{1 \text{ mol } Al_2(SO_4)_3}{342.17 \text{ g } Al_2(SO_4)_3}\right)$$

$$= 0.028952 \text{ mol } Al_2(SO_4)_3$$

$$\text{Mass (g) of } Al(OH)_3 \text{ from } Al_2(SO_4)_3 = \left(0.028952 \text{ mol } Al_2(SO_4)_3\right)\left(\frac{2 \text{ mol } Al(OH)_3}{1 \text{ mol } Al_2(SO_4)_3}\right)\left(\frac{78.00 \text{ g } Al(OH)_3}{1 \text{ mol } Al(OH)_3}\right)$$

$$= 4.5166 \text{ g } Al(OH)_3$$

$$\text{Moles of NaOH} = \left(185.5 \text{ mL}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.533 \text{ mol NaOH}}{\text{L}}\right) = 0.0988715 \text{ mol NaOH}$$

$$\text{Mass (g) of } Al(OH)_3 \text{ from NaOH} = \left(0.0988715 \text{ mol NaOH}\right)\left(\frac{2 \text{ mol } Al(OH)_3}{6 \text{ mol NaOH}}\right)\left(\frac{78.00 \text{ g } Al(OH)_3}{1 \text{ mol } Al(OH)_3}\right)$$

$= 2.570659 =$ **2.57 g $Al(OH)_3$**

NaOH is the limiting reagent.

4.34 Plan: Write the chemical reaction between the two reactants. Then write the total ionic equation in which all soluble ionic substances are dissociated into ions. Omit spectator ions in the net ionic equation.

Solution:
The molecular equation is $H_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow SrSO_4(s) + 2H_2O(l)$
The total ionic equation is:

$$2H^+(aq) + SO_4^{2-}(aq) + Sr^{2+}(aq) + 2OH^-(aq) \rightarrow SrSO_4(s) + 2H_2O(l)$$

According to the solubility rules, $SrSO_4$ is insoluble and therefore does not dissociate into ions.
Since there are no spectator ions, the total and net ionic equations are the same.

4.35 Plan: Review the section on acid-base reactions.
Solution:
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) Strong acids and bases dissociate 100% into ions in aqueous solution.

4.36 Plan: Review the section on acid-base reactions.
Solution:
a) There are many possibilities including: acetic acid ($HC_2H_3O_2$), chlorous acid ($HClO_2$), and nitrous acid (HNO_2). All acids are weak except for the six strong acids listed in the text.
b) NH_3
c) Strong acids and bases dissociate 100% into ions and are therefore strong electrolytes; weak acids and bases dissociate much less than this (typically less than 10%) in aqueous solution and are therefore weak electrolytes. 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.

4.37 Plan: Since strong acids and bases dissociate completely in water, these substances can be written as ions in a total ionic equation; since weak acids and bases dissociate into ions only to a small extent, these substances appear undissociated in total ionic equations.
Solution:
a) Acetic acid is a weak acid and sodium hydroxide is a strong base:
Molecular equation: $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$
Total ionic equation: $CH_3COOH\ (aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + CH_3COO^-(aq) + H_2O(l)$
Net ionic equation (remove the spectator ion Na^+): $CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)$
Hydrochloric acid is a strong acid:
Molecular equation: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
Total ionic equation: $H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$
Net ionic equation (remove the spectator ions Na^+ and Cl^-): $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
The difference in the net ionic equation is due to the fact that CH_3COOH is a weak acid and dissociates very little while HCl is a strong acid and dissociates completely.
b) When acetic acid dissociates in water, most of the species in the solution is un-ionized acid, $CH_3COOH(aq)$; the amounts of its ions, H^+ and CH_3COO^-, are equal but very small: $\mathbf{[CH_3COOH] >> [H^+] = [CH_3COO^-]}$.

4.38 Plan: Remember that strong acids and bases can be written as ions in the total ionic equation but weak acids and bases cannot be written as ions. Omit spectator ions from the net ionic equation.
Solution:
a) KOH is a strong base and HBr is a strong acid; both may be written in dissociated form. KBr is a soluble compound since all Group 1A(1) compounds are soluble.
Molecular equation: $KOH(aq) + HBr(aq) \rightarrow KBr(aq) + H_2O(l)$
Total ionic equation: $K^+(aq) + OH^-(aq) + H^+(aq) + Br^-(aq) \rightarrow K^+(aq) + Br^-(aq) + H_2O(l)$
Net ionic equation: $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
The spectator ions are $K^+(aq)$ and $Br^-(aq)$.
b) 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 (as ions). NH_4Cl is a soluble compound, because all ammonium compounds are soluble.
Molecular equation: $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$
Total ionic equation: $NH_3(aq) + H^+(aq) + Cl^-(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$
Net ionic equation: $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$
Cl^- is the only spectator ion.

4.39 Plan: Remember that strong acids and bases can be written as ions in the total ionic equation but weak acids and bases cannot be written as ions. Omit spectator ions from the net ionic equation.

Solution:

a) CsOH is a strong base and HNO_3 is a strong acid; both may be written in dissociated form. $CsNO_3$ is a soluble compound since all nitrate compounds are soluble.

Molecular equation: $CsOH(aq) + HNO_3(aq) \rightarrow CsNO_3(aq) + H_2O(l)$

Total ionic equation: $Cs^+(aq) + OH^-(aq) + H^+(aq) + NO_3^-(aq) \rightarrow Cs^+(aq) + NO_3^-(aq) + H_2O(l)$

Net ionic equation: $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

Spectator ions are Cs^+ and NO_3^-.

b) $HC_2H_3O_2$ is a weak acid and is written in the molecular form. $Ca(OH)_2$ is a strong base and is written in the dissociated form (as ions). $Ca(C_2H_3O_2)_2$ is a soluble compound, because all acetate compounds are soluble.

Molecular equation: $Ca(OH)_2(aq) + 2HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + 2H_2O(l)$

Total ionic equation: $Ca^{2+}(aq) + 2OH^-(aq) + 2HC_2H_3O_2\ (aq) \rightarrow Ca^{2+}(aq) + 2C_2H_3O_2^-(aq) + 2H_2O(l)$

Net ionic equation: $OH^-(aq) + HC_2H_3O_2(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

Spectator ion is Ca^{2+}.

4.40 Plan: Write an acid-base reaction between $CaCO_3$ and HCl. Remember that HCl is a strong acid.

Solution:

Calcium carbonate dissolves in HCl(*aq*) because the carbonate ion, a base, reacts with the acid to form H_2CO_3 which decomposes into $CO_2(g)$ and $H_2O(l)$.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2CO_3(aq)$$

Total ionic equation:

$$CaCO_3(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + H_2O(l) + CO_2(g)$$

Net ionic equation:

$$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$$

4.41 Plan: Write an acid-base reaction between $Zn(OH)_2$ and HNO_3. Remember that HNO_3 is a strong acid.

Solution:

Zinc hydroxide dissolves in HCl(*aq*) because the hydroxide ion, a base, reacts with the acid to form soluble zinc nitrate and water.

$$Zn(OH)_2(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2H_2O(aq)$$

Total ionic equation:

$$Zn(OH)_2(s) + 2H^+(aq) + 2NO_3^-(aq) \rightarrow Zn^{2+}(aq) + 2NO_3^-(aq) + 2H_2O(l)$$

Net ionic equation:

$$Zn(OH)_2(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + 2H_2O(l)$$

4.42 Plan: Write a balanced equation. Find the moles of KOH from the molarity and volume information and use the molar ratio in the balanced equation to find the moles of acid present. Divide the moles of acid by its volume to determine the molarity.

Solution:

The reaction is: $KOH(aq) + CH_3COOH(aq) \rightarrow CH_3COOK(aq) + H_2O(l)$

$$\text{Moles of KOH} = (25.98\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.1180\text{ mol KOH}}{\text{L}}\right) = 0.00306564\text{ mol KOH}$$

$$\text{Moles of CH}_3\text{COOH} = (0.00306564\text{ mol KOH})\left(\frac{1\text{mol CH}_3\text{COOH}}{1\text{ mol KOH}}\right) = 0.00306564\text{ mol CH}_3\text{COOH}$$

$$\text{Molarity of CH}_3\text{COOH} = \left(\frac{0.00306564\text{ mol CH}_3\text{COOH}}{52.50\text{ mL}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right) = 0.05839314 = \mathbf{0.05839\ \mathit{M}\ CH_3COOH}$$

4.43 Plan: Write a balanced equation. Find the moles of NaOH from the molarity and volume information and use the molar ratio in the balanced equation to find the moles of acid present. Divide the moles of acid by its volume to determine the molarity.

Solution:
The reaction is: $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

$$\text{Moles of NaOH} = (26.25\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.1850\text{ mol NaOH}}{\text{L}}\right) = 0.00485625\text{ mol NaOH}$$

$$\text{Moles of } H_2SO_4 = (0.00485625\text{ mol NaOH})\left(\frac{1\text{mol } H_2SO_4}{2\text{ mol NaOH}}\right) = 0.002428125\text{ mol } H_2SO_4$$

$$\text{Molarity of } H_2SO_4 = \left(\frac{0.002428125\text{ mol } H_2SO_4}{25.00\text{ mL}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right) = 0.097125 = \mathbf{0.09712}\ \boldsymbol{M}\ \mathbf{H_2SO_4}$$

4.44 Plan: Write a balanced equation. Find the moles of H_2SO_4 from the molarity and volume information and use the molar ratio in the balanced equation to find the moles of $NaHCO_3$ required to react with that amount of H_2SO_4. Divide the moles of $NaHCO_3$ by its molarity to find the volume.
Solution:
The reaction is: $2NaHCO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l) + 2CO_2(g)$

$$\text{Moles of } H_2SO_4 = (88\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{2.6\text{ mol } H_2SO_4}{\text{L}}\right) = 0.2288\text{ mol } H_2SO_4$$

$$\text{Moles of } NaHCO_3 = (0.2288\text{ mol } H_2SO_4)\left(\frac{2\text{ mol } NaHCO_3}{1\text{ mol } H_2SO_4}\right) = 0.4576\text{ mol } NaHCO_3$$

$$\text{Volume (mL) of } NaHCO_3 = (0.4576\text{ mol } NaHCO_3)\left(\frac{1\text{ L}}{1.6\text{ mol } NaHCO_3}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

$$= 286 = \mathbf{2.9 \times 10^2\ mL\ NaHCO_3}$$

4.45 Plan: Balance the reaction. Convert the amount of UO_2 from kg to g to moles; use the molar ratio in the balanced reaction to find the moles of HF required to react with the moles of UO_2. Divide moles of HF by its molarity to calculate the volume.
Solution:
The reaction is: $UO_2(s) + 4HF(aq) \rightarrow UF_4(s) + 2H_2O(l)$

$$\text{Moles of } UO_2 = (2.15\text{ kg } UO_2)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol } UO_2}{270.0\text{ g } UO_2}\right) = 7.96296\text{ mol } UO_2$$

$$\text{Moles of HF} = (7.96296\text{ mol } UO_2)\left(\frac{4\text{ mol HF}}{1\text{ mol } UO_2}\right) = 31.85184\text{ mol HF}$$

$$\text{Volume (L) of HF} = (31.85184\text{ mol HF})\left(\frac{1\text{ L}}{2.40\text{ mol HF}}\right) = 13.2716 = \mathbf{13.3\ L\ HF}$$

4.46 Plan: Write balanced equations for the reaction of NaOH with oxalic acid, benzoic acid, and HCl. Find the moles of added NaOH from the molarity and volume information; then use the molarity and volume information for HCl to find the moles of HCl required to react with the excess NaOH. Use the molar ratio in the NaOH/HCl reaction to find the moles of excess NaOH. The moles of NaOH required to titrate the acid samples is the difference of the added NaOH and the excess NaOH. Let x = mass of benzoic acid and 0.3471 – x = mass of oxalic acid. Convert the mass of each acid to moles using the molar mass and use the molar ratios in the balanced reactions to find the amounts of each acid. Mass percent is calculated by dividing the mass of benzoic acid by the mass of the sample and multiplying by 100.
Solution:
Oxalic acid is $H_2C_2O_4$ and benzoic acid is $HC_7H_5O_2$.
The reactions are: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
$2NaOH(aq) + H_2C_2O_4(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(l)$
$NaOH(aq) + HC_7H_5O_2(aq) \rightarrow NaC_7H_5O_2(aq) + H_2O(l)$

$$\text{Moles of NaOH added} = (100.0 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.1000 \text{ mol NaOH}}{1 \text{ L}}\right) = 0.01000 \text{ mol NaOH}$$

$$\text{Moles of added HCl} = (20.00 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.2000 \text{ mol HCl}}{1 \text{ L}}\right) = 0.004000 \text{ mol HCl}$$

$$\text{Moles of excess NaOH} = (0.004000 \text{ mol HCl})\left(\frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}}\right) = 0.004000 \text{ mol NaOH}$$

Moles of NaOH required to titrate sample = moles NaOH added – moles excess NaOH
= 0.01000 mol – 0.004000 mol = 0.006000 mol NaOH

Let x = mass of $HC_7H_5O_2$ and 0.3471 – x = mass of $H_2C_2O_4$

Moles of NaOH required to titrate $HC_7H_5O_2$ =

$$(\text{x g } HC_7H_5O_2)\left(\frac{1 \text{ mol } HC_7H_5O_2}{122.12 \text{ g } HC_7H_5O_2}\right)\left(\frac{1 \text{ mol NaOH}}{1 \text{ mol } HC_7H_5O_2}\right) = 0.008189\text{x}$$

$$\text{Moles of NaOH required to titrate } H_2C_2O_4 = ((0.3471 - \text{x}) \text{ g } H_2C_2O_4)\left(\frac{1 \text{ mol } H_2C_2O_4}{90.04 \text{ g } H_2C_2O_4}\right)\left(\frac{2 \text{ mol NaOH}}{1 \text{ mol } H_2C_2O_4}\right)$$

= 0.007710 – 0.02221x

Moles of NaOH required to titrate sample = 0.006000 mol = 0.008189x + (0.007710 – 0.02221x)

0.006000 = 0.007710 – 0.014021x

–0.001710 = –0.014021x

0.12196 = x = mass of $HC_7H_5O_2$

$$\text{Mass \% of } HC_7H_5O_2 = \frac{\text{mass of } HC_7H_5O_2}{\text{mass of sample}}(100) = \frac{0.12196 \text{ g}}{0.3471 \text{ g}}(100) = 35.1368 = \mathbf{35.14\%}$$

4.47 Plan: Write balanced reactions between HNO_3 and each of the bases. Find the moles of HNO_3 from its molarity and volume. Let x = mass of $Al(OH)_3$ and 0.4826 – x = mass of $Mg(OH)_2$. Convert the mass of each base to moles using the molar mass and use the molar ratios in the balanced reactions to find the amounts of each base. Mass percent is calculated by dividing the mass of $Al(OH)_3$ by the mass of the sample and multiplying by 100.

Solution:

The reactions are: $3HNO_3(aq) + Al(OH)_3(aq) \rightarrow Al(NO_3)_3(aq) + 3H_2O(l)$

$2HNO_3(aq) + Mg(OH)_2(aq) \rightarrow Mg(NO_3)_2(aq) + 2H_2O(l)$

$$\text{Moles of } HNO_3 = (17.30 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{1.000 \text{ mol } HNO_3}{1 \text{ L}}\right) = 0.0173 \text{ mol } HNO_3$$

Let x = mass of $Al(OH)_3$ and 0.4826 – x = mass of $Mg(OH)_2$

Moles of HNO_3 required to titrate $Al(OH)_3$ =

$$(\text{x g } Al(OH)_3)\left(\frac{1 \text{ mol } Al(OH)_3}{78.00 \text{ g } Al(OH)_3}\right)\left(\frac{3 \text{ mol } HNO_3}{1 \text{ mol } Al(OH)_3}\right) = 0.038462\text{x}$$

Moles of HNO_3 required to titrate $Mg(OH)_2$ =

$$((0.4826 - \text{x}) \text{ g } Mg(OH)_2)\left(\frac{1 \text{ mol } Mg(OH)_2}{58.33 \text{ g } Mg(OH)_2}\right)\left(\frac{2 \text{ mol } HNO_3}{1 \text{ mol } Mg(OH)_2}\right)$$

= 0.01655 – 0.03429x

Moles of HNO_3 required to titrate sample = 0.0173 mol = 0.038462x + (0.01655 – 0.03429x)

0.0173 = 0.004172x + 0.01655

0.17977 g = x = mass of $Al(OH)_3$

$$\text{Mass \% of } Al(OH)_3 = \frac{\text{mass of } Al(OH)_3}{\text{mass of sample}}(100) = \frac{0.17977 \text{ g}}{0.4826 \text{ g}}(100) = 37.2503 = \mathbf{37.25\%}$$

4.48 Plan: Recall that oxidation is the loss of electrons and reduction is the gain of electrons.
Solution:
The electrons that a substance gains during reduction must come from somewhere. So there must be an oxidation in which electrons are lost, to provide the electrons gained during reduction.

4.49 Plan: An oxidizing agent gains electrons and therefore has an atom whose oxidation number decreases during the reaction. Use the Rules for Assigning an Oxidation Number to assign S in H_2SO_4 an O.N. and see if this oxidation number changes during the reaction. An acid transfers a proton during reaction.
Solution:
a) In H_2SO_4, hydrogen has an O.N. of +1, for a total of +2; oxygen has an O.N. of –2 for a total of –8. The S has an O.N. of +6. In SO_2, the O.N. of oxygen is –2 for a total of –4 and S has an O.N. of +4. So the S has been reduced from +6 to +4 and is an **oxidizing agent**. Iodine is oxidized during the reaction.
b) The oxidation number of S is +6 in H_2SO_4; in $BaSO_4$, Ba has an O.N. of +2, the four oxygen atoms have a total O.N. of –8, and S is again +6. Since the oxidation number of S (or any of the other atoms) did not change, this is not a redox reaction. H_2SO_4 transfers a proton to $F^–$ to produce HF, so it acts as an **acid**.

4.50 Plan: Consult the Rules for Assigning an Oxidation Number. The sum of the O.N. values for the atoms in a molecule equals zero, while the sum of the O.N. values for the atoms in an ion equals the ion's charge.
Solution:
a) NH_2OH. Hydrogen has an O.N. of +1, for a total of +3 for the three hydrogen atoms. Oxygen has an O.N. of –2. The O.N. of N must be –1 since [(–1) + (+3) + (–2)] = 0. **N = –1**
b) N_2F_4. The O.N. of each fluorine is –1 for a total of –4; the sum of the O.N.s for the two N atoms must be +4, so each N has an O.N. of +2. **N = +2**
c) NH_4^+. The O.N. of each hydrogen is +1 for a total of +4; the O.N. of nitrogen must be –3 since the overall sum of the O.N.s must be +1: [(–3) + (+4)] = +1 **N = –3**
d) HNO_2. The O.N. of hydrogen is +1 and that of each oxygen is –2 for a total of –4 from the oxygens. The O.N. of nitrogen must be +3 since [(+1) + (+3) + (–4)] = 0. **N = +3**

4.51 Plan: Consult the Rules for Assigning an Oxidation Number. The sum of the O.N. values for the atoms in a molecule equals zero.
Solution:
a) $SOCl_2$. The O.N. of oxygen is –2 and that of each chlorine is –1 for a total of –2 for the two chlorine atoms. The O.N. of sulfur must be +4 since [(+4) + (–2) + (–2)] = 0. **S = +4**
b) H_2S_2. The O.N. of each hydrogen is +1, for a total of +2. The sum of the O.N.s of the two sulfur atoms must equal –2, so the O.N. of each S atom is –1. **S = –1**
c) H_2SO_3. The O.N. of each hydrogen atom is +1 for a total of +2; the O.N. of each oxygen atom is –2 for a total of –6. The O.N. of the sulfur must be +4 since [(+2) + (+4) + (–6)] = 0. **S = +4**
d) Na_2S. The O.N. of each sodium [Group 1A(1)] is +1, for a total of +2. The O.N. of sulfur is –2. **S = –2**

4.52 Plan: Consult the Rules for Assigning an Oxidation Number. The sum of the O.N. values for the atoms in a molecule equals zero, while the sum of the O.N. values for the atoms in an ion equals the ion's charge.
Solution:
a) AsH_3. H is combined with a nonmetal, so its O.N. is +1 (Rule 3). Three H atoms have a sum of +3. To have a sum of 0 for the molecule, As has an O.N. of –3. **As = –3**
b) $H_2AsO_4^–$. The O.N. of H in this compound is +1, for a total of +2. The O.N. of each oxygen is –2, for a total of –8. As has an O.N. of **+5** since [(+2) + (+5) + (–8)] = –1, the charge of the ion. **As = +5**
c) $AsCl_3$. Each chlorine has an O.N. of –1, for a total of –3. The O.N. of As is +3. **As = +3**

4.53 Plan: Consult the Rules for Assigning an Oxidation Number. The sum of the O.N. values for the atoms in a molecule equals zero, while the sum of the O.N. values for the atoms in an ion equals the ion's charge.
Solution:
a) $H_2P_2O_7^{2–}$. The O.N.of each hydrogen is +1, for a total of +2; the O.N. of each oxygen is –2, for a total of –14. The sum of the O.N.s of the two phosphorus atoms must be +10 since [(+2) + (+10) + (–14)] = –2, the charge of the ion. Each of the two phosphorus atoms has an O.N. of +5. **P = +5**
b) PH_4^+. The O.N. of each hydrogen is +1, for a total of +4. The O.N. of P is –3 since [(–3) + (+4)] = +1, the charge of the ion. **P = –3**

c) PCl_5. The O.N. of each Cl is –1, for a total of –5. The O.N. of P is therefore +5. **P = +5**

4.54 Plan: Consult the Rules for Assigning an Oxidation Number. The sum of the O.N. values for the atoms in a molecule equals zero, while the sum of the O.N. values for the atoms in an ion equals the ion's charge.
Solution:
a) MnO_4^{2-}. The O.N. of each oxygen is –2, for a total of –8; the O.N. of Mn must be +6 since [(+6) + (–8)] = –2, the charge of the ion. **Mn = +6**
b) Mn_2O_3. The O.N. of each oxygen is –2, for a total of –6; the sum of the O.N.s of the two Mn atoms must be +6. The O.N. of each manganese is +3. **Mn = +3**
c) $KMnO_4$. The O.N. of potassium is +1 and the O.N. of each oxygen is –2, for a total of –8. The O.N. of Mn is +7 since [(+1) + (+7) + (–8)] = 0. **Mn = +7**

4.55 Plan: Consult the Rules for Assigning an Oxidation Number. The sum of the O.N. values for the atoms in a molecule equals zero, while the sum of the O.N. values for the atoms in an ion equals the ion's charge.
Solution:
a) CrO_3. The O.N. of each oxygen atom is –2, for a total of –6. The O.N. of chromium must be +6. **Cr = +6**
b) $Cr_2O_7^{2-}$. The O.N. of each oxygen is –2, for a total of –14. The sum of the O.N.s of the two chromium atoms must be +12 since [(+12) + (–14)] = –2, the charge of the ion. Each of the two chromium atoms has an O.N. of +6. **Cr = +6**
c) $Cr_2(SO_4)_3$. It is convenient to treat the polyatomic ion SO_4^{2-} as a unit with a –2 charge, for a total of –6 for the three sulfate ions. The sum of the two chromium atoms must be +6 and the O.N. of each chromium atom is +3. **Cr = +3**

4.56 Plan: First, assign oxidation numbers to all atoms following the rules. The reactant that is the reducing agent contains an atom that is oxidized (O.N. increases from the left side to the right side of the equation). The reactant that is the oxidizing agent contains an atom that is reduced (O.N. decreases from the left side to the right side of the equation). Recognize that the agent is the compound that contains the atom that is oxidized or reduced, not just the atom itself.
Solution:
a) +2 +6 –8   –8   –4   +2
+1 +3 –2   +7 –2   +1   +2   +4 –2   +1 –2
$5H_2C_2O_4(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$
Mn in MnO_4^- changes from +7 to +2 (reduction). Therefore, **MnO_4^- is the oxidizing agent**. C in $H_2C_2O_4$ changes from +3 to +4 (oxidation), so **$H_2C_2O_4$ is the reducing agent**.
b)   –6   +2
0   +1   +5 –2   +2   +2 –2   +1 –2
$3Cu(s) + 8H^+(aq) + 2NO_3^-(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)$
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**.

4.57 Plan: First, assign oxidation numbers to all atoms following the rules. The reactant that is the reducing agent contains an atom that is oxidized (O.N. increases from the left side to the right side of the equation). The reactant that is the oxidizing agent contains an atom that is reduced (O.N. decreases from the left side to the right side of the equation). Recognize that the agent is the compound that contains the atom that is oxidized or reduced, not just the atom itself.
Solution:
0   +1   +2   0
a) $Sn(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$
Sn changes from 0 to +2 (is oxidized) so **Sn is the reducing agent**. H changes from +1 to 0 (is reduced) so **H^+ is the oxidizing agent.**
b)   +2 –2   +2
+1   +1 –1   +2   +3   +1 –2
$2H^+(aq) + H_2O_2(aq) + 2Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + 2H_2O(l)$
Oxygen changes from –1 in H_2O_2 to –2 in H_2O (is reduced) so **H_2O_2 is the oxidizing agent**. Fe changes from +2
to +3 (is oxidized) so **Fe^{2+} is the reducing agent**.

4.58 Plan: First, assign oxidation numbers to all atoms following the rules. The reactant that is the reducing agent contains an atom that is oxidized (O.N. increases from the left side to the right side of the equation). The reactant that is the oxidizing agent contains an atom that is reduced (O.N. decreases from the left side to the right side of the equation). Recognize that the agent is the compound that contains the atom that is oxidized or reduced, not just the atom itself.

Solution:

a) –6 –6 –4 +2

+1 –1 0 +5 –2 +4 –1 +4 –2 +1 –2

$8H^+(aq) + 6Cl^-(aq) + Sn(s) + 4NO_3^-(aq) \rightarrow SnCl_6^{2-}(aq) + 4NO_2(g) + 4H_2O(l)$

Nitrogen changes from an O.N. of +5 in NO_3^- to +4 in NO_2 (is reduced) so **NO_3^- is the oxidizing agent.** Sn changes from an O.N. of 0 to an O.N. of +4 in $SnCl_6^{2-}$ (is oxidized) so **Sn is the reducing agent**.

b) –8 +2

+7 –2 –1 +1 0 +2 +1 –2

$2MnO_4^-(aq) + 10Cl^-(aq) + 16H^+(aq) \rightarrow 5Cl_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$

Manganese changes from an O.N. of +7 in MnO_4^- to an O.N. of +2 in Mn^{2+} (is reduced) so **MnO_4^- is the oxidizing** agent. Chlorine changes its O.N. from –1 in Cl^- to 0 as the element Cl_2 (is oxidized) so **Cl^- is the reducing agent**.

4.59 Plan: First, assign oxidation numbers to all atoms following the rules. The reactant that is the reducing agent contains an atom that is oxidized (O.N. increases from the left side to the right side of the equation). The reactant that is the oxidizing agent contains an atom that is reduced (O.N. decreases from the left side to the right side of the equation). Recognize that the agent is the compound that contains the atom that is oxidized or reduced, not just the atom itself.

Solution:

a) +12 –14 –6 –8 +2

+1 +6 –2 +4 –2 +3 +6 –2 +1 –2

$8H^+(aq) + Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq) + 2H_2O(l)$

Chromium changes from an O.N. of +6 in $Cr_2O_7^{2-}$ to +3 in Cr^{3+} (is reduced) so **$Cr_2O_7^{2-}$ is the oxidizing agent**.
Sulfur changes from an O.N. of +4 in SO_3^{2-} to +6 in SO_4^{2-} (is oxidized) so **SO_3^{2-} is the reducing agent**.

b) –6 +2 –8 +4 +3

+5 –2 0 –2 +1 +1 –2 +2 –2 +1 –3 +1

$NO_3^-(aq) + 4Zn(s) + 7OH^-(aq) + 6H_2O(l) \rightarrow 4Zn(OH)_4^{2-}(aq) + NH_3(aq)$

Nitrogen changes from an O.N. of +5 in NO_3^- to an O.N. of –3 in NH_3 (is reduced) so **NO_3^- is the oxidizing agent**.

Zinc changes from an O.N. of 0 to an O.N. of +2 in $Zn(OH)_4^{2-}$ (is oxidized) so **Zn is the reducing agent**.

4.60 Plan: Find the moles of $Cr_2O_7^{2-}$ from the molarity and volume information. Use the molar ratio in the balanced equation to find the moles of C_2H_5OH and multiply the moles of C_2H_5OH by its molar mass to determine the mass of C_2H_5OH present. Mass percent is calculated by dividing the mass of C_2H_5OH by the mass of the sample and multiplying by 100.

Solution:

$$\text{Moles of } Cr_2O_7^{2-} = (35.46 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.05961 \text{ mol } Cr_2O_7^{2-}}{1\text{ L}}\right) = 0.0021138 \text{ mol } Cr_2O_7^{2-}$$

$$\text{Moles of } C_2H_5OH = (0.0021138 \text{ mol } Cr_2O_7^{2-})\left(\frac{1 \text{ mol } C_2H_5OH}{2 \text{ mol } Cr_2O_7^{2-}}\right) = 0.0010569 \text{ mol } C_2H_5OH$$

$$\text{Mass (g) of } C_2H_5OH = (0.0010569 \text{ mol } C_2H_5OH)\left(\frac{46.07 \text{ g } C_2H_5OH}{1 \text{ mol } C_2H_5OH}\right) = 0.0486914 \text{ g } C_2H_5OH$$

$$\text{Mass percent of } C_2H_5OH = \frac{\text{mass of } C_2H_5OH}{\text{mass of sample}}(100) = \frac{0.0486914 \text{ g } C_2H_5OH}{28.00 \text{ g sample}}(100)$$

$$= 0.173898 = \mathbf{0.1739\%\ C_2H_5OH}$$

4.61 Plan: The three types of redox reactions are combination, decomposition, and displacement. In a combination reaction, two or more reactants form one product, so the number of substances decreases. In a decomposition reaction, one reactant forms two or more products, so the number of substances increases. In a displacement reaction, the number of substances is the same, but atoms exchange places.
Solution:
a) **decomposition** b) **combination** c) **displacement**

4.62 Plan: Recall that a reactant breaks down into two or more products in a decomposition reaction, while reactants combine to form a product in a combination reaction.
Solution:
By definition, elements cannot decompose into anything simpler, so they could not be reactants in a decomposition reaction.

4.63 Plan: Review the types of redox reaction discussed in this section.
Solution:
Combination, decomposition, and displacement reactions generally produce only one compound; **combustion reactions**, however, often produce both carbon dioxide and water.

4.64 Plan: In a combination reaction, two or more reactants form one product. In a decomposition reaction, one reactant forms two or more products. In a displacement reaction, atoms or ions exchange places. Balance the reactions by inspection.
Solution:
a) $2Sb(s) + 3Cl_2(g) \rightarrow 2SbCl_3(s)$
Combination: two reactants combine to form one product.
b) $2AsH_3(g) \rightarrow 2As(s) + 3H_2(g)$
Decomposition: one reactant breaks into two products.
c) $Zn(s) + Fe(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Fe(s)$
Displacement: one Zn displaces one Fe atom.

4.65 Plan: In a combination reaction, two or more reactants form one product. In a decomposition reaction, one reactant forms two or more products. In a displacement reaction, atoms or ions exchange places. Balance the reactions by inspection.
Solution:
a) $Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
Displacement: one Mg displaces two H atoms.
b) $Cr(NO_3)_3(aq) + Al(s) \rightarrow Al(NO_3)_3(aq) + Cr(s)$
Displacement: one Al displaces one Cr atom.
c) $PF_3(g) + F_2(g) \rightarrow PF_5(g)$
Combination: two reactants combine to form one product.

4.66 Plan: In a combination reaction, two or more reactants form one product. Two elements as reactants often results in a combination reaction. In a decomposition reaction, one reactant forms two or more products; one reactant only often indicates a decomposition reaction. In a displacement reaction, atoms or ions exchange places. An element and a compound as reactants often indicate a displacement reaction. Balance the reactions by inspection.
Solution:
a) The combination of two nonmetals gives a covalent compound.
$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
b) Some compounds undergo thermal decomposition to simpler substances.
$2NaClO_3(s) \xrightarrow{\Delta} 2NaCl(s) + 3O_2(g)$
c) This is a displacement reaction. Active metals like Ba can displace hydrogen from water.
$Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$

4.67 Plan: In a combination reaction, two or more reactants form one product. Two elements as reactants often results in a combination reaction. In a decomposition reaction, one reactant forms two or more products; one reactant only often indicates a decomposition reaction. In a displacement reaction, atoms or ions exchange places. An element and a compound as reactants often indicate a displacement reaction. Balance the reactions by inspection.
Solution:
a) This is a displacement reaction in which iron displaces hydrogen.

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

b) The combination of two nonmetals gives a covalent compound.

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

c) Some compounds undergo decomposition to their elements during electrolysis in which electrical energy is absorbed.

$$BaCl_2(l) \xrightarrow{\text{electricity}} Ba(l) + Cl_2(g)$$

4.68 Plan: In a combination reaction, two or more reactants form one product. Two elements as reactants often results in a combination reaction. In a decomposition reaction, one reactant forms two or more products; one reactant only often indicates a decomposition reaction. In a displacement reaction, atoms or ions exchange places. An element and a compound as reactants often indicate a displacement reaction. Balance the reactions by inspection.
Solution:
a) Cs, a metal, and I_2, a nonmetal, combine to form the binary ionic compound, CsI.

$$2Cs(s) + I_2(s) \rightarrow 2CsI(s)$$

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 $Mn^0(s)$.

$$2Al(s) + 3MnSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Mn(s)$$

c) This is a combination reaction in which sulfur dioxide, SO_2, a nonmetal oxide, combines with oxygen, O_2, to form the higher oxide, SO_3.

$$2SO_2(g) + O_2(g) \xrightarrow{\Delta} 2SO_3(g)$$

It is not clear from the problem, but energy must be added to force this reaction to proceed.
d) Butane is a four carbon hydrocarbon with the formula C_4H_{10}. 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.

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

e) Total ionic equation in which soluble species are shown dissociated into ions:

$$2Al(s) + 3Mn^{2+}(aq) + 3SO_4^{2-}(aq) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq) + 3Mn(s)$$

Net ionic equation in which the spectator ions are omitted:

$$2Al(s) + 3Mn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Mn(s)$$

Note that the molar coefficients are not simplified because the number of electrons lost (6 e^-) must equal the electrons gained (6 e^-).

4.69 Plan: In a combination reaction, two or more reactants form one product. Two elements as reactants often results in a combination reaction. In a decomposition reaction, one reactant forms two or more products; one reactant only often indicates a decomposition reaction. In a displacement reaction, atoms or ions exchange places. An element and a compound as reactants often indicate a displacement reaction. Balance the reactions by inspection.
Solution:
a) Pentane is a five carbon hydrocarbon with the formula C_5H_{12}. 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.

$$C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$$

b) Phosphorus trichloride, PCl_3, is a nonmetal halide that combines with additional halogen, to form the higher halide, PCl_5.

$$PCl_3(l) + Cl_2(g) \rightarrow PCl_5(s)$$

c) This is a displacement reaction. Active metals like Zn can displace hydrogen from acid.

$Zn(s) + 2HBr(aq) \rightarrow ZnBr_2(aq) + H_2(g)$

d) This is a displacement reaction in which bromine displaces iodine. A halogen higher in the periodic table can displace a halogen that is lower.

$2KI(aq) + Br_2(l) \rightarrow 2KBr(aq) + I_2(s)$

e) Total ionic equation in which soluble species are shown dissociated into ions:

$2K^+(aq) + 2I^-(aq) + Br_2(l) \rightarrow 2K^+(aq) + 2Br^-(aq) + I_2(s)$

Net ionic equation in which the spectator ions are omitted:

$2I^-(aq) + Br_2(l) \rightarrow I_2(s) + 2Br^-(aq)$

4.70 Plan: Write a balanced equation that shows the decomposition of HgO to its elements. Convert the mass of HgO to moles and use the molar ratio from the balanced equation to find the moles and then the mass of O_2. Perform the same calculation to find the mass of the other product.

Solution:

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

$$\text{Moles of HgO} = (4.27\ \text{kg HgO})\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol HgO}}{216.6\ \text{g HgO}}\right) = 19.71376\ \text{mol HgO}$$

$$\text{Moles of } O_2 = (19.71376\ \text{mol HgO})\left(\frac{1\ \text{mol } O_2}{2\ \text{mol HgO}}\right) = 9.85688\ \text{mol } O_2$$

$$\text{Mass (g) of } O_2 = (9.85688\ \text{mol } O_2)\left(\frac{32.00\ \text{g } O_2}{1\ \text{mol } O_2}\right) = 315.420 = \mathbf{315\ g\ O_2}$$

The other product is **mercury**.

$$\text{Moles of Hg} = (19.71376\ \text{mol HgO})\left(\frac{2\ \text{mol Hg}}{2\ \text{mol HgO}}\right) = 19.71376\ \text{mol Hg}$$

$$\text{Mass (kg) Hg} = (19.71376\ \text{mol Hg})\left(\frac{200.6\ \text{g Hg}}{1\ \text{mol Hg}}\right)\left(\frac{1\ \text{kg}}{10^3\ \text{g}}\right) = 3.95458 = \mathbf{3.95\ kg\ Hg}$$

4.71 Plan: Write a balanced equation that shows the decomposition of calcium chloride to its elements. Convert the mass of $CaCl_2$ to moles and use the molar ratio from the balanced equation to find the moles and then the mass of Cl_2. Perform the same calculation to find the mass of the other product.

Solution:

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

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

$$\text{Moles of } CaCl_2 = (874\ \text{g } CaCl_2)\left(\frac{1\ \text{mol } CaCl_2}{110.98\ \text{g } CaCl_2}\right) = 7.87529\ \text{mol } CaCl_2$$

$$\text{Moles of } Cl_2 = (7.87529\ \text{mol } CaCl_2)\left(\frac{1\ \text{mol } Cl_2}{1\ \text{mol } CaCl_2}\right) = 7.87529\ \text{mol } Cl_2$$

$$\text{Mass (g) of } Cl_2 = (7.87529\ \text{mol } Cl_2)\left(\frac{70.90\ \text{g } Cl_2}{1\ \text{mol } Cl_2}\right) = 558.358 = \mathbf{558\ g\ Cl_2}$$

The other product is **calcium**.

$$\text{Moles of Ca} = (7.87529\ \text{mol } CaCl_2)\left(\frac{1\ \text{mol Ca}}{1\ \text{mol } CaCl_2}\right) = 7.87529\ \text{mol Ca}$$

$$\text{Mass (g) of Ca} = (7.87529\ \text{mol Ca})\left(\frac{40.08\ \text{g Ca}}{1\ \text{mol Ca}}\right) = 315.64 = \mathbf{316\ g\ Ca}$$

4.72 Plan: To determine the reactant in excess, write the balanced equation (metal + $O_2 \rightarrow$ metal oxide), convert reactant masses to moles, and use molar ratios to see which reactant makes the smaller ("limiting") amount of product. Use the limiting reactant to calculate the amount of product formed. Use the molar ratio to find the amount of excess reactant required to react with the limiting reactant; the amount of excess reactant that remains is the initial amount of excess reactant minus the amount required for the reaction.

Solution:

The balanced equation is $4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$.

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

Moles of Li_2O if O_2 limiting = $(6.50\text{ g }O_2)\left(\frac{1\text{ mol }O_2}{32.00\text{ g }O_2}\right)\left(\frac{2\text{ mol }Li_2O}{1\text{ mol }O_2}\right)$ = 0.40625 mol Li_2O

Li is the limiting reactant since it produces the smaller amount of product; **O_2 is in excess**.

b) Using Li as the limiting reagent, 0.1166979 = **0.117 mol Li_2O** is formed.

c) Li is limiting, thus there will be none remaining (**0 g Li**).

Mass (g) of Li_2O = $(0.1166979\text{ mol }Li_2O)\left(\frac{29.88\text{ g }Li_2O}{1\text{ mol }Li_2O}\right)$ = 3.4869 = **3.49 g Li_2O**

Mass (g) of O_2 reacted = $(1.62\text{ g Li})\left(\frac{1\text{ mol Li}}{6.941\text{ g Li}}\right)\left(\frac{1\text{ mol }O_2}{4\text{ mol Li}}\right)\left(\frac{32.00\text{ g }O_2}{1\text{ mol }O_2}\right)$ = 1.867166 g O_2

Remaining O_2 = initial amount – amount reacted = 6.50 g O_2 – 1.867166 g O_2 = 4.632834 = **4.63 g O_2**

4.73 Plan: To determine the reactant in excess, write the balanced equation (metal + $N_2 \rightarrow$ metal nitride), convert reactant masses to moles, and use molar ratios to see which reactant makes the smaller ("limiting") amount of product. Use the limiting reactant to calculate the amount of product formed. Use the molar ratio to find the amount of excess reactant required to react with the limiting reactant; the amount of excess reactant that remains is the initial amount of excess reactant minus the amount required for the reaction.

Solution:

The balanced equation is $3Mg(s) + N_2(g) \xrightarrow{\Delta} Mg_3N_2(s)$.

a) Moles of Mg_3N_2 if Mg is limiting = $(2.22\text{ g Mg})\left(\frac{1\text{ mol Mg}}{24.31\text{ g Mg}}\right)\left(\frac{1\text{ mol }Mg_3N_2}{3\text{ mol Mg}}\right)$ = 0.030440 mol Mg_3N_2

Moles of Mg_3N_2 if N_2 is limiting = $(3.75\text{ g }N_2)\left(\frac{1\text{ mol }N_2}{28.02\text{ g }N_2}\right)\left(\frac{1\text{ mol }Mg_3N_2}{1\text{ mol }N_2}\right)$ = 0.13383 mol Mg_3N_2

Mg is the limiting reactant since it produces the smaller amount of product; **N_2 is present in excess**.

b) Using Mg as the limiting reactant, 0.030440 = **0.0304 mol Mg_3N_2** is formed.

c) There will be **0 Mg** remaining since it is the limiting reagent and will be completely consumed.

Mass (g) of Mg_3N_2 = $(0.030440\text{ mol }Mg_3N_2)\left(\frac{100.95\text{ g }Mg_3N_2}{1\text{ mol }Mg_3N_2}\right)$ = 3.07292 = **3.07 g Mg_3N_2**

Mass (g) of N_2 reacted = $(2.22\text{ g Mg})\left(\frac{1\text{ mol Mg}}{24.31\text{ g Mg}}\right)\left(\frac{1\text{ mol }N_2}{3\text{ mol Mg}}\right)\left(\frac{28.02\text{ g }N_2}{1\text{ mol }N_2}\right)$ = 0.852933 g N_2

Remaining N_2 = initial amount – amount reacted = 3.75 g N_2 – 0.852933 g N_2 = 2.897067 = **2.90 g N_2**

4.74 Plan: Since mass must be conserved, the original amount of mixture – amount of remaining solid = mass of carbon dioxide produced. Write a balanced equation and use molar ratios to convert from the mass of CO_2 produced to the amount of $CaCO_3$ reacted. Mass percent is calculated by dividing the mass of $CaCO_3$ by the mass of the sample and multiplying by 100.

Solution:

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

Mass (g) of CO_2 produced = mass of mixture – mass of remaining solid = 0.693 g – 0.508 g = 0.185 g CO_2

Mass (g) of $CaCO_3$ = $(0.185 \text{ g } CO_2)\left(\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}\right)\left(\frac{1 \text{ mol } CaCO_3}{1 \text{ mol } CO_2}\right)\left(\frac{100.09 \text{ g } CaCO_3}{1 \text{ mol } CaCO_3}\right)$ = 0.420737 g $CaCO_3$

Mass % $CaCO_3$ = $\frac{\text{mass of } CaCO_3}{\text{mass of sample}}(100\%)$ = $\frac{0.420737 \text{ g } CaCO_3}{0.693 \text{ g sample}}(100\%)$ = 60.7124 = **60.7% $CaCO_3$**

4.75 Plan: Write the balanced equation for the displacement reaction, convert reactant masses to moles, and use molar ratios to see which reactant makes the smaller ("limiting") amount of product. Use the limiting reactant to calculate the amount of product formed.

Solution:

The balanced reaction is $2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(l) + Al_2O_3(s)$.

Moles of Fe if Al is limiting = $(1.50 \text{ kg Al})\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol Al}}{26.98 \text{ g Al}}\right)\left(\frac{2 \text{ mol Fe}}{2 \text{ mol Al}}\right)$ = 55.59674 mol Fe

Moles of Fe if Fe_2O_3 is limiting = $(25.0 \text{ mol } Fe_2O_3)\left(\frac{2 \text{ mol Fe}}{1 \text{ mol } Fe_2O_3}\right)$ = 50.0 mol Fe

Fe_2O_3 is the limiting reactant since it produces the smaller amount of Fe; 50.0 moles of Fe forms.

Mass (g) of Fe = $(50.0 \text{ mol Fe})\left(\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}\right)$ = 2792.5 = **2790 g Fe**

4.76 Plan: Ferrous ion is Fe^{2+}. Write a reaction to show the conversion of Fe to Fe^{2+}. Convert the mass of Fe in a 125-g serving to the mass of Fe in a 737-g sample. Use molar mass to convert mass of Fe to moles of Fe and use Avogadro's number to convert moles of Fe to moles of ions.

Solution:

a) Fe oxidizes to Fe^{2+} with a loss of 2 electrons. The H^+ in the acidic food is reduced to H_2 with a gain of 2 electrons. The balanced reaction is:

$$\begin{array}{lcccc} & Fe(s) + & 2H^+(aq) \rightarrow & Fe^{2+}(aq) + & H_2(g) \\ \text{O.N.:} & 0 & +1 & +2 & 0 \end{array}$$

b) Mass (g) of Fe in the jar of tomato sauce = $(737 \text{ g sauce})\left(\frac{49 \text{ mg Fe}}{125 \text{ g sauce}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)$ = 0.288904 g Fe

Number of Fe^{2+} ions = $(0.288904 \text{ g Fe})\left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right)\left(\frac{1 \text{ mol } Fe^{2+}}{1 \text{ mol Fe}}\right)\left(\frac{6.022 \times 10^{23} Fe^{2+} \text{ions}}{1 \text{ mol } Fe^{2+}}\right)$

= 3.11509×10^{21} = **3.1×10^{21} Fe^{2+} ions** per jar of sauce

4.77 Plan: Convert the mass of glucose to moles and use the molar ratios from the balanced equation to find the moles of ethanol and CO_2. The amount of ethanol is converted from moles to grams using its molar mass. The amount of CO_2 is converted from moles to volume in liters using the conversion factor given.

Solution:

Moles of C_2H_5OH = $(100. \text{ g } C_6H_{12}O_6)\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.16 \text{ g } C_6H_{12}O_6}\right)\left(\frac{2 \text{ mol } C_2H_5OH}{1 \text{ mol } C_6H_{12}O_6}\right)$ = 1.11012 mol C_2H_5OH

Mass (g) of C_2H_5OH = $(1.11012 \text{ mol } C_2H_5OH)\left(\frac{46.07 \text{ g } C_2H_5OH}{1 \text{ mol } C_2H_5OH}\right)$ = 51.143 = **51.1 g C_2H_5OH**

Moles of CO_2 = $(100. \text{ g } C_6H_{12}O_6)\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.16 \text{ g } C_6H_{12}O_6}\right)\left(\frac{2 \text{ mol } CO_2}{1 \text{ mol } C_6H_{12}O_6}\right)$ = 1.11012 mol CO_2

Volume (L) of CO_2 = $(1.11012 \text{ mol } CO_2)\left(\frac{22.4 \text{ L } CO_2}{1 \text{ mol } CO_2}\right)$ = 24.8667 = **24.9 L CO_2**

4.78 Plan: Find the moles of $KMnO_4$ from the molarity and volume information. Use the molar ratio in the balanced equation to find the moles and then mass of iron. Mass percent is calculated by dividing the mass of iron by the mass of the sample and multiplying by 100.

Solution:

$$\text{Moles of } KMnO_4 = (39.32\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{0.03190\ \text{mol } MnO_4^-}{\text{L}}\right) = 0.00125431\ \text{mol } KMnO_4$$

$$\text{Mass (g) of Fe} = \left(0.00125431\ \text{mol } MnO_4^-\right)\left(\frac{5\ \text{mol } Fe^{2+}}{1\ \text{mol } MnO_4^-}\right)\left(\frac{55.85\ \text{g Fe}}{1\ \text{mol } Fe^{2+}}\right) = 0.350266\ \text{g Fe}$$

$$\text{Mass \% of Fe} = \frac{\text{mass of Fe}}{\text{mass of sample}}(100) = \frac{0.350266\ \text{g}}{1.1081\ \text{g}}(100) = 31.6096 = \mathbf{31.61\%\ Fe}$$

4.79 Plan: Write balanced equations for the two acid-base reactions. Find the moles of H_2SO_4 from the molarity and volume information and use the molar ratio in the balanced equation for the reaction of H_2SO_4 and NaOH to find the moles of NaOH used in the titration. Divide the moles of NaOH by its volume to determine its molarity. Then find the moles of NaOH used in the titration of HCl by multiplying the NaOH molarity by its volume; use the molar ratio in this reaction to find moles of HCl. Dividing moles of HCl by its volume gives its molarity.

Solution:

Write the balanced chemical equations:

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

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

$$\text{Moles of } H_2SO_4 = (50.0\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{0.0782\ \text{mol } H_2SO_4}{\text{L}}\right) = 0.00391\ \text{mol } H_2SO_4$$

$$\text{Moles of NaOH} = \left(0.00391\ \text{mol } H_2SO_4\right)\left(\frac{2\ \text{mol NaOH}}{1\ \text{mol } H_2SO_4}\right) = 0.00782\ \text{mol NaOH}$$

$$\text{Molarity of NaOH} = \left(\frac{0.00782}{18.4\ \text{mL}}\right)\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right) = \mathbf{0.425\ \mathit{M}\ NaOH}$$

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

$$\text{Moles of NaOH} = (27.5\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{0.425\ \text{mol NaOH}}{\text{L}}\right) = 0.0116875\ \text{mol NaOH}$$

$$\text{Moles of HCl} = (0.0116875\ \text{mol NaOH})\left(\frac{1\ \text{mol HCl}}{1\ \text{mol NaOH}}\right) = 0.0116875\ \text{mol HCl}$$

$$\text{Molarity of HCl} = (0.0116875\ \text{mol HCl})\left(\frac{1}{100.\text{mL}}\right)\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right) = 0.116875 = \mathbf{0.117\ \mathit{M}\ HCl}$$

4.80 Plan: Recall that the total ionic equation shows all soluble ionic substances dissociated into ions and the net ionic equation omits the spectator ions. Use the molar ratio in the balanced reaction to find the moles of acid and base. Divide the moles of acid and base by the volume to obtain the molarity.

Solution:

a) Molecular: $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

Total ionic: $2H^+(aq) + SO_4^{2-}(aq) + 2Na^+(aq) + 2OH^-(aq) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq) + 2H_2O(l)$

Net ionic: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(Na^+ and SO_4^{2-} are spectator ions.)

$$\text{b) Moles of } H_2SO_4 = (2\ \text{orange spheres})\left(\frac{0.010\ \text{mol } SO_4^{2-}}{1\ \text{orange sphere}}\right)\left(\frac{1\ \text{mol } H_2SO_4}{1\ \text{mol } SO_4^{2-}}\right) = \mathbf{0.020\ mol\ H_2SO_4}$$

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$$\text{Moles of NaOH} = (0.020 \text{ mol } H_2SO_4)\left(\frac{2 \text{ mol NaOH}}{1 \text{ mol } H_2SO_4}\right) = \mathbf{0.040 \text{ mol NaOH}}$$

c) $$\text{Molarity of } H_2SO_4 = (0.020 \text{ mol } H_2SO_4)\left(\frac{1}{25\text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3}\text{ L}}\right) = \mathbf{0.80}\ \boldsymbol{M}\ \mathbf{H_2SO_4}$$

$$\text{Molarity of NaOH} = (0.040 \text{ mol NaOH})\left(\frac{1}{25\text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3}\text{ L}}\right) = \mathbf{1.6}\ \boldsymbol{M}\ \mathbf{NaOH}$$

4.81 Plan: Write balanced chemical equations for the acid-base titration reactions. To find the concentration of HA, find the moles of NaOH used for its titration by multiplying the molarity of NaOH by the volume used in the titration and using the molar ratio to find the moles of HA; dividing moles of HA by its volume gives the molarity. Multiply the molarity of HA by the volume of HA in the acid mixture to find the moles of HA in the mixture. Use the molar ratio to find the volume of NaOH required to titrate this amount of HA. The total volume of NaOH used in the titration of the mixture minus the volume required to titrate HA is the volume of NaOH required to titrate HB. Use this volume and molarity of NaOH and the molar ratio to find the moles and then molarity of HB. The volume of HB in the acid mixture is the total volume minus the volume of HA.

Solution:

The balanced chemical equations for HA or HB with sodium hydroxide are the same. For HA it is:

$$HA(aq) + NaOH(aq) \rightarrow NaA(aq) + H_2O(l)$$

To find the concentration of HA:

$$\text{Moles of NaOH} = (87.3 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right)\left(\frac{0.0906 \text{ mol NaOH}}{\text{L}}\right) = 0.007909 \text{ mol NaOH}$$

$$\text{Moles of HA} = (0.007909 \text{ mol NaOH})\left(\frac{1 \text{ mol HA}}{1 \text{ mol NaOH}}\right) = 0.007909 \text{ mol HA}$$

$$\text{Molarity of HA} = (0.007909 \text{ mol HA})\left(\frac{1}{43.5 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3}\text{ L}}\right) = 0.1818248 = \mathbf{0.182}\ \boldsymbol{M}\ \mathbf{HA}$$

The titration of the acid mixture involves the reaction of NaOH with both of the acids.

$$\text{Moles of HA in the acid mixture} = (37.2 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right)\left(\frac{0.1818248 \text{ mol HA}}{\text{L}}\right) = 0.0067639 \text{ mol HA}$$

Volume (mL) of NaOH required to titrate HA =

$$(0.0067639 \text{ mol HA})\left(\frac{1 \text{ mol NaOH}}{1 \text{ mol HA}}\right)\left(\frac{1 \text{ L}}{0.0906 \text{ mol NaOH}}\right)\left(\frac{1 \text{ mL}}{10^{-3}\text{ L}}\right) = 74.6565 \text{ mL NaOH}$$

Volume of NaOH required to titrate HB = total NaOH volume – volume of NaOH required to titrate HA
= 96.4 mL – 74.6565 mL = 21.7435 mL NaOH

$$\text{Moles of HB} = (21.7435 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right)\left(\frac{0.0906 \text{ mol NaOH}}{\text{L}}\right)\left(\frac{1 \text{ mol HB}}{1 \text{ mol NaOH}}\right) = 0.00196996 \text{ mol HB}$$

Volume (mL) of HB = Volume of mixture – volume of HA = 50.0 mL – 37.2 mL = 12.8 mL

$$\text{Molarity of HB} = (0.00196996 \text{ mol HB})\left(\frac{1}{12.8 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3}\text{ L}}\right) = 0.153903 = \mathbf{0.154}\ \boldsymbol{M}\ \mathbf{HB}$$

4.82 Plan: For part a), assign oxidation numbers to each element; the oxidizing agent has an atom whose oxidation number decreases while the reducing agent has an atom whose oxidation number increases. For part b), use the molar ratios, beginning with step 3, to find the moles of NO_2, then moles of NO, then moles of NH_3 required to produce the given mass of HNO_3.

Solution:

a) *Step 1*

+3 +2

–3 +1 0 +2 –2 +1 –2

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$

N is oxidized from –3 in NH_3 to +2 in NO; O is reduced from 0 in O_2 to to –2 in NO.

Oxidizing agent = O_2 **Reducing agent = NH_3**

Step 2 –4

+2 –2 0 +4 –2

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

N is oxidized from +2 in NO to +4 in NO_2; O is reduced from 0 in O_2 to –2 in NO_2.

Oxidizing agent = O_2 **Reducing agent = NO**

Step 3

–4 +2 –6

+4 –2 +1 –2 +1+5 –2 +2 –2

$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(l) + NO(g)$

N is oxidized from +4 in NO_2 to +5 in HNO_3; N is reduced from +4 in NO_2 to +2 in NO.

Oxidizing agent = NO_2 **Reducing agent = NO_2**

b) Moles of NO_2 = $\left(3.0\text{x}10^4 \text{ kg } HNO_3\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } HNO_3}{63.02 \text{ g } HNO_3}\right)\left(\frac{3 \text{ mol } NO_2}{2 \text{ mol } HNO_3}\right) = 7.14059\text{x}10^5 \text{ mol } NO_2$

Moles of NO = $\left(7.14059\text{x}10^5 \text{ mol } NO_2\right)\left(\frac{2 \text{ mol NO}}{2 \text{ mol } NO_2}\right) = 7.14059\text{x}10^5 \text{ mol NO}$

Moles of NH_3 = $\left(7.14059\text{x}10^5 \text{ mol NO}\right)\left(\frac{4 \text{ mol } NH_3}{4 \text{ mol NO}}\right) = 7.14059\text{x}10^5 \text{ mol } NH_3$

Mass (kg) of NH_3 = $\left(7.14059\text{x}10^5 \text{ mol } NH_3\right)\left(\frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = 1.21604\text{x}10^4 = \mathbf{1.2x10^4 \text{ kg } NH_3}$

4.83 Plan: Write the formulas of the reactants; create cation-anion combinations other than the original reactants and determine if they are insoluble. A precipitate forms if reactant ions can form combinations that are insoluble, as determined by the solubility rules in Table 4.1. Any ions not involved in a precipitate are spectator ions and are omitted from the net ionic equation. For the acid-base reactions, strong acids and bases dissociate completely in water and can be written as ions in a total ionic equation; weak acids and bases dissociate into ions only to a small extent, so these substances appear undissociated in total ionic equations.

Solution:

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) $Ba(OH)_2(aq) + 2\ HCN(aq) \rightarrow Ba(CN)_2(aq) + 2\ H_2O(l)$

$OH^-(aq) + HCN(aq) \rightarrow CN^-(aq) + H_2O(l)$

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

4.84 Plan: In part a), use the density of the alloy to find the volume of a 0.263-g sample of alloy. That volume is the sum of the volume of Mg and Al in the alloy. Letting x = mass of Mg and 0.263 – x = mass of Al, find the volume of each metal and set that equal to the total volume of the alloy. In part b), write balanced displacement reactions in which Mg and Al displace hydrogen from the HCl to produce H_2. Use the molar ratios to find the masses of Mg and Al that must be present to produce the given amount of H_2. In part c), write balanced reactions for the formation of MgO and Al_2O_3 and use molar ratios to find the masses of Mg and Al that must be present in the sample to produce the given amount of oxide.

Solution:

a) Let x = mass of Mg and 0.263 – x = mass of Al

$$\text{Volume (cm}^3\text{) of alloy} = (0.263\text{ g alloy})\left(\frac{1\text{ cm}^3}{2.40\text{ g alloy}}\right) = 0.10958\text{ cm}^3$$

Volume of alloy = volume of Mg + volume of Al

$$0.10958\text{ cm}^3 = (\text{x g Mg})\left(\frac{1\text{ cm}^3\text{ Mg}}{1.74\text{ g Mg}}\right) + ((0.263 - \text{x})\text{ g Al})\left(\frac{1\text{ cm}^3\text{ Al}}{2.70\text{ g Al}}\right)$$

$0.10958\text{ cm}^3 = 0.574713\text{x} + 0.097407 - 0.37037\text{x}$

$0.012173 = 0.204343\text{x}$

$\text{x} = 0.05957\text{ g Mg}$

$$\text{Mass percent Mg} = \frac{\text{mass of Mg}}{\text{mass of alloy sample}}(100) = \frac{0.05957\text{ g Mg}}{0.263\text{ g sample alloy}}(100) = 22.6502 = \mathbf{22.7\%\ Mg}$$

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

$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

Let x = mass of Mg and 0.263 – x = mass of Al

Moles of H_2 produced = moles of H_2 from Mg + moles of H_2 from Al

$$1.38\text{x}10^{-2}\text{ mol H}_2 = (\text{x g Mg})\left(\frac{1\text{ mol Mg}}{24.31\text{ g Mg}}\right)\left(\frac{1\text{ mol H}_2}{1\text{ mol Mg}}\right) + ((0.263 - \text{x})\text{ g Al})\left(\frac{1\text{ mol Al}}{26.98\text{ g Al}}\right)\left(\frac{3\text{ mol H}_2}{2\text{ mol Al}}\right)$$

$1.38\text{x}10^{-2}\text{ mol H}_2 = 0.041135\text{x} + 0.014622 - 0.055597\text{x}$

$8.22\text{x}10^{-4} = 0.014462\text{x}$

$\text{x} = 0.05684\text{ g Mg}$

$$\text{Mass percent Mg} = \frac{\text{mass of Mg}}{\text{mass of alloy sample}}(100) = \frac{0.05684\text{ g Mg}}{0.263\text{ g sample alloy}}(100) = 21.6122 = \mathbf{21.6\%\ Mg}$$

c) $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

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

Let x = mass of Mg and 0.263 – x = mass of Al

Mass of oxide produced = mass of MgO from Mg + mass of Al_2O_3 from Al

0.483 g oxide =

$$(\text{x g Mg})\left(\frac{1\text{ mol Mg}}{24.31\text{ g Mg}}\right)\left(\frac{2\text{ mol MgO}}{2\text{ mol Mg}}\right)\left(\frac{40.31\text{ g MgO}}{1\text{ mol MgO}}\right) + ((0.263 - \text{x})\text{ g Al})\left(\frac{1\text{ mol Al}}{26.98\text{ g Al}}\right)\left(\frac{2\text{ mol Al}_2\text{O}_3}{4\text{ mol Al}}\right)\left(\frac{101.96\text{ g Al}_2\text{O}_3}{1\text{ mol Al}_2\text{O}_3}\right)$$

$0.483\text{ g} = 1.6582\text{x} + 0.49695 - 1.88955\text{x}$

$0.01395 = 0.23135\text{x}$

$\text{x} = 0.060298\text{ g Mg}$

$$\text{Mass percent Mg} = \frac{\text{mass of Mg}}{\text{mass of alloy sample}}(100) = \frac{0.060298\text{ g Mg}}{0.263\text{ g sample alloy}}(100) = 22.927 = \mathbf{22.9\%\ Mg}$$

4.85 Plan: Write a balanced equation and use the molar ratio between Na_2O_2 and CO_2 to convert the amount of Na_2O_2 given to the amount of CO_2 that reacts with that amount. Convert that amount of CO_2 to liters of air.

Solution:

The reaction is: $2Na_2O_2(s) + 2CO_2(g) \rightarrow 2Na_2CO_3(s) + O_2(g)$.

$$\text{Mass (g) of } CO_2 = (80.0\text{ g } Na_2O_2)\left(\frac{1\text{ mol } Na_2O_2}{77.98\text{ g } Na_2O_2}\right)\left(\frac{2\text{ mol } CO_2}{2\text{ mol } Na_2O_2}\right)\left(\frac{44.01\text{ g } CO_2}{1\text{ mol } CO_2}\right) = 45.1500\text{ g } CO_2$$

$$\text{Volume (L) of air} = (45.150\text{ g } CO_2)\left(\frac{\text{L air}}{0.0720\text{ g } CO_2}\right) = 627.08 = \mathbf{627\ L\ air}$$

4.86 Plan: Convert the given volume of seawater to units of mL and use the density of seawater to find the mass of that volume of seawater. Use the given % by mass of Mg in seawater to find the mass of Mg.
Solution:

$$\text{Volume (mL) of seawater} = (1.00\text{ km}^3)\left(\frac{10^3\text{ m}}{1\text{ km}}\right)^3\left(\frac{1\text{ cm}}{10^{-2}\text{ m}}\right)^3\left(\frac{1\text{ mL}}{1\text{ cm}^3}\right) = 1.00\text{x}10^{15}\text{ mL}$$

$$\text{Mass (g) of seawater} = (1.00\text{x}10^{15}\text{ mL})\left(\frac{1.04\text{ g}}{1\text{ mL}}\right) = 1.04\text{x}10^{15}\text{ g}$$

$$\text{Mass (kg) Mg} = (1.04\text{x}10^{15}\text{ mL})\left(\frac{0.13\%\text{ Mg}}{100\%}\right)\left(\frac{1\text{ kg}}{10^3\text{ g}}\right) = 1.3520\text{ x }10^9 = \mathbf{1.4\ x\ 10^9\ kg\ Mg}$$

4.87 Plan: To determine the reactant in excess, convert reactant masses to moles, and use molar ratios to see which reactant makes the smaller ("limiting") amount of product. Use the limiting reactant to calculate the amount of product formed. Use the molar ratio to find the amount of excess reactant required to react with the limiting reactant; the amount of excess reactant that remains is the initial amount of excess reactant minus the amount required for the reaction. Multiply moles of products and excess reactant by Avogadro's number to obtain number of molecules.
Solution:

$$\text{a) Moles of } C_2H_5Cl \text{ if } C_2H_4 \text{ is limiting} = (0.100\text{ kg } C_2H_4)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol } C_2H_4}{28.05\text{ g } C_2H_4}\right)\left(\frac{1\text{ mol } C_2H_5Cl}{1\text{ mol } C_2H_4}\right)$$
$$= 3.56506\text{ mol } C_2H_5Cl$$

$$\text{Moles of } C_2H_5Cl \text{ if HCl is limiting} = (0.100\text{ kg HCl})\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol HCl}}{36.46\text{ g HCl}}\right)\left(\frac{1\text{ mol } C_2H_5Cl}{1\text{ mol HCl}}\right)$$
$$= 2.74273\text{ mol } C_2H_5Cl$$

The HCl is limiting.
Moles HCl remaining = 0 mol

$$\text{Moles of } C_2H_4 \text{ initially present} = (0.100\text{ kg } C_2H_4)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol } C_2H_4}{28.05\text{ g } C_2H_4}\right) = 3.56506\text{ mol } C_2H_4$$

$$\text{Moles of } C_2H_4 \text{ that react} = (0.100\text{ kg HCl})\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol HCl}}{36.46\text{ g HCl}}\right)\left(\frac{1\text{ mol } C_2H_4}{1\text{ mol HCl}}\right) = 2.74273\text{ mol } C_2H_4$$

Moles of C_2H_4 remaining = initial moles – reacted moles = 3.56506 mol – 2.74273 mol = 0.82233 mol C_2H_4
Moles of C_2H_5Cl formed = 2.74273 mol C_2H_5Cl
Total moles of gas = moles HCl + moles C_2H_4 + moles C_2H_5Cl = 0 mol + 0.82233 mol + 2.74273 mol
= 3.56506 mol

$$\text{Molecules of gas} = (3.56506\text{ mol gas})\left(\frac{6.022\text{x}10^{23}\text{ molecules}}{1\text{ mol gas}}\right) = 2.146879\text{x}10^{24} = \mathbf{2.15x10^{24}\ molecules}$$

b) This will still be based on the HCl as the limiting reactant.

$$\text{Initial moles of HCl} = (0.100\text{ kg HCl})\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol HCl}}{36.46\text{ g HCl}}\right) = 2.74273\text{ mol HCl}$$

Moles of HCl remaining = intial moles/2 = (2.74273 mol HCl)/2 = 1.371365 mol HCl

Moles of C_2H_4 reacting with half of HCl = $(1.371365 \text{ mol HCl})\left(\frac{1 \text{ mol } C_2H_4}{1 \text{ mol HCl}}\right)$ = 1.371365 mol C_2H_4

Moles of C_2H_4 remaining = initial moles – reacted moles = 3.56506 mol – 1.371365 mol = 2.193695 mol C_2H_4

Moles of C_2H_5Cl formed = $(1.371365 \text{ mol HCl})\left(\frac{1 \text{ mol } C_2H_5Cl}{1 \text{ mol HCl}}\right)$ = 1.371365 mol C_2H_5Cl

Total moles of gas = moles HCl + moles C_2H_4 + moles C_2H_5Cl = 1.371365 mol + 2.193695 mol + 1.371365 mol
= 4.936425 = **4.94 mol total**

4.88 Plan: Write balanced equations for the reaction of CO_2 with the various metal hydroxides. Convert the mass of metal hydroxide to moles by dividing by the molar mass; use the mole ratio in the balanced equation to find the moles and then mass of CO_2 required to react with the metal hydroxide.

Solution:

The reactions are $2\ LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(l)$

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

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

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

= 3215.762 = **3.22×10^3 g CO_2**

b) Mass CO_2 absorbed by 1.00 g LiOH:

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

= 0.918789 = **0.919 g CO_2**

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

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

= 0.75450 = **0.754 g CO_2**

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

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

= 0.846346 = **0.846 g CO_2**

4.89 Plan: Balance the equation to obtain the correct molar ratios. Use the mass percents to find the mass of each reactant in a 1.00 g sample, convert the mass of each reactant to moles, and use the molar ratios to find the limiting reactant and the amount of CO_2 produced. Convert moles of CO_2 produced to volume using the given conversion factor.

Solution:

a) Here is a suggested method for 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.

$Ca(H_2PO_4)_2(s) + 2NaHCO_3(s) \xrightarrow{\Delta} 2CO_2(g) + 2H_2O(g) + CaHPO_4(s) + Na_2HPO_4(s)$

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.

Mass (g) of $NaHCO_3$ = $(1.00 \text{ g})\left(\frac{31.0\%}{100\%}\right)$ = 0.31 g $NaHCO_3$

Mass (g) of $Ca(H_2PO_4)_2 = (1.00\ g)\left(\frac{35.0\%}{100\%}\right)$ = 0.35 g $Ca(H_2PO_4)_2$

Moles of CO_2 if $NaHCO_3$ is limiting = $(0.31\ g\ NaHCO_3)\left(\frac{1\ mol\ NaHCO_3}{84.01\ g\ NaHCO_3}\right)\left(\frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3}\right)$

$= 3.690x10^{-3}$ mol CO_2

Moles of CO_2 if $Ca(H_2PO_4)_2$ is limiting = $(0.35 g\ Ca(H_2PO_4)_2)\left(\frac{1\ mol\ Ca(H_2PO_4)_2}{234.05\ g\ Ca(H_2PO_4)_2}\right)\left(\frac{2\ mol\ CO_2}{1\ mol\ Ca(H_2PO_4)_2}\right)$

$= 2.9908x10^{-3}$ mol CO_2

Since $Ca(H_2PO_4)_2$ produces the smaller amount of product, it is the limiting reactant and **$3.0x10^{-3}$ mol CO_2** will be produced.

b) Volume (L) of $CO_2 = (2.9908x10^{-3}\ mol\ CO_2)\left(\frac{37.0\ L}{1\ mol\ CO_2}\right)$ = 0.1106596 = **0.11 L CO_2**

4.90 Plan: Write a balanced acid-base reaction. Find the total moles of NaOH used by multiplying its molarity and volume in liters and use the molar ratio in the reaction to find the moles of HNO_3. Divide moles of HNO_3 by its volume to obtain the molarity. Use the molarity and volume information to find the moles of NaOH initially added and the moles of HNO_3 initially present. The difference of these two values is the moles of excess NaOH.

Solution:

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.

Moles of NaOH = $(23.22\ mL)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.0502\ mol\ NaOH}{L}\right)$ = 0.0011656 mol NaOH

Moles of HNO_3 = $(0.0011656\ mol\ NaOH)\left(\frac{1\ mol\ HNO_3}{1\ mol\ NaOH}\right)$ = 0.0011656 mol HNO_3

Molarity of HNO_3 = $(0.0011656\ mol\ HNO_3)\left(\frac{1}{80.00\ mL}\right)\left(\frac{1\ mL}{10^{-3}\ L}\right)$ = 0.01457055 = **0.0146 *M*** HNO_3

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

Moles of NaOH = $(20.00\ mL)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.0502\ mol\ NaOH}{L}\right)$ = $1.004x10^{-3}$ mol NaOH

Moles of HNO_3 = $(50.00\ mL)\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.01457055\ mol\ HNO_3}{L}\right)$ = $7.285275x10^{-4}$ mol HNO_3

Moles of NaOH required to titrate $7.285275x10^{-4}$ mol HNO_3 = $7.285275x10^{-4}$ mol NaOH

Moles excess NaOH = moles of added NaOH – moles of NaOH required for reaction

= $1.004x10^{-3}$ mol NaOH – $7.285275x10^{-4}$ mol NaOH

= $2.754725x10^{-4}$ = **$2.8x10^{-4}$ mol NaOH**

4.91 Plan: To determine the empirical formula, find the moles of each element present and divide by the smallest number of moles to get the smallest ratio of atoms. To find the molecular formula, divide the molar mass by the mass of the empirical formula to find the factor by which to multiple the empirical formula. Write the balanced acid-base reaction for part c) and use the molar ratio in that reaction to find the mass of bismuth(III) hydroxide.

Solution:

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.

$$\text{Moles of C} = (0.1880 \text{ g } CO_2)\left(\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}\right)\left(\frac{1 \text{ mol C}}{1 \text{ mol } CO_2}\right) = 4.271756\text{x}10^{-3} \text{ mol C}$$

$$\text{Moles of H} = (0.02750 \text{ g } H_2O)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)\left(\frac{2 \text{ mol H}}{1 \text{ mol } H_2O}\right) = 3.052164\text{x}10^{-3} \text{ mol H}$$

$$\text{Moles of Bi} = (0.1422 \text{ g } Bi_2O_3)\left(\frac{1 \text{ mol } Bi_2O_3}{466.0 \text{ g } Bi_2O_3}\right)\left(\frac{2 \text{ mol Bi}}{1 \text{ mol } Bi_2O_3}\right) = 6.103004\text{x}10^{-4} \text{ mol Bi}$$

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.

$$\text{Mass (g) of C} = (4.271756\text{x}10^{-3} \text{ mol C})\left(\frac{12.01 \text{ g C}}{1 \text{ mol C}}\right) = 0.0513038 \text{ g C}$$

$$\text{Mass (g) of H} = (3.052164\text{x}10^{-3} \text{ mol H})\left(\frac{1.008 \text{ g H}}{1 \text{ mol H}}\right) = 0.0030766 \text{ g H}$$

$$\text{Mass (g) of Bi} = (6.103004\text{x}10^{-4} \text{ mol Bi})\left(\frac{209.0 \text{ g Bi}}{1 \text{ mol Bi}}\right) = 0.127553 \text{ g Bi}$$

Mass (g) of O = mass of sample – (mass C + mass H + mass Bi)

= 0.22105 g sample – (0.0513038 g C + 0.0030766 g H + 0.127553 g Bi) = 0.0391166 g O

$$\text{Moles of O} = (0.0391166 \text{ g O})\left(\frac{1 \text{ mol O}}{16.00 \text{ g O}}\right) = 2.44482\text{x}10^{-4} \text{ mol O}$$

Divide each of the moles by the smallest value (moles Bi).

$$C = \frac{4.271756\text{x}10^{-3}}{6.103004\text{x}10^{-4}} = 7 \qquad H = \frac{3.052164\text{x}10^{-3}}{6.103004\text{x}10^{-4}} = 5$$

$$O = \frac{2.4448\text{x}10^{-3}}{6.103004\text{x}10^{-4}} = 4 \qquad Bi = \frac{6.103004\text{x}10^{-4}}{6.103004\text{x}10^{-4}} = 1$$

Empirical formula = $\mathbf{C_7H_5O_4Bi}$

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 $C_7H_5O_4Bi$ = $\mathbf{C_{21}H_{15}O_{12}Bi_3}$.

c) $Bi(OH)_3(s) + 3HC_7H_5O_3(aq) \rightarrow Bi(C_7H_5O_3)_3(s) + 3H_2O(l)$

$$\text{d) Moles of } C_{21}H_{15}O_{12}Bi_3 = (0.600 \text{ mg } C_{21}H_{15}O_{12}Bi_3)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mol } C_{21}H_{15}O_{12}Bi_3}{1086 \text{ g } C_{21}H_{15}O_{12}Bi_3}\right)$$

$$= 5.52486\text{x}10^{-4} \text{ mol } C_{21}H_{15}O_{12}Bi_3$$

Mass (mg) of $Bi(OH)_3$ =

$$(5.52486\text{x}10^{-7} \text{ mol } C_{21}H_{15}O_{12}Bi_3)\left(\frac{3 \text{ mol Bi}}{1 \text{ mol } C_{21}H_{15}O_{12}Bi_3}\right)\left(\frac{1 \text{ mol } Bi(OH)_3}{1 \text{ mol Bi}}\right)\left(\frac{260.0 \text{ g } Bi(OH)_3}{1 \text{ mol } Bi(OH)_3}\right)\left(\frac{1 \text{ mg}}{10^{-3} \text{ g}}\right)\left(\frac{100\%}{88.0\%}\right)$$

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

4.92 Plan: Use the solubility rules to predict the products of this reaction. For the total ionic equation, write all soluble ionic substances as dissociated ions. Ions not involved in the precipitate are spectator ions and are not included in the net ionic equation. Find the moles of dissolved ions and divide each by the volume in liters to find the concentration. The volume of the final solution is the sum of the volumes of the two reactant solutions.

Solution:

a) According to the solubility rules, all chloride compounds are soluble and most common carbonate compounds are insoluble. $CaCO_3$ is the precipitate.

Molecular equation: $Na_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2NaCl(aq)$

Total ionic equation: $2Na^+(aq) + CO_3^{2-}(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \rightarrow CaCO_3(s) + 2Na^+(aq) + 2Cl^-(aq)$

Net ionic equation: $CO_3^{2-}(aq) + Ca^{2+}(aq) \rightarrow CaCO_3(s)$

b) Ca^{2+} and CO_3^{2-} combine in a 1:1 ratio in $CaCO_3$. There are two spheres of Ca^{2+} and three spheres of CO_3^{2-}

ion. Since there are fewer spheres of Ca^{2+}, Ca^{2+} is the limiting reactant.

$$\text{Mass of } CaCO_3 = (2\ Ca^{2+}\text{spheres})\left(\frac{0.050\ \text{mol } Ca^{2+}}{1\ \text{sphere}}\right)\left(\frac{1\ \text{mol } CaCO_3}{1\ \text{mol } Ca^{2+}}\right)\left(\frac{100.09\ \text{g } CaCO_3}{1\ \text{mol } CaCO_3}\right)$$

$$= 10.009 = \mathbf{10.\ g\ CaCO_3}$$

c) Original moles:

$$\text{Moles of } Na^+ = (6\ Na^+\text{spheres})\left(\frac{0.050\ \text{mol } Na^+}{1\ \text{sphere}}\right) = 0.30\ \text{mol } Na^+$$

$$\text{Moles of } CO_3^{2-} = (3\ CO_3^{2-}\text{spheres})\left(\frac{0.050\ \text{mol } CO_3^{2-}}{1\ \text{sphere}}\right) = 0.15\ \text{mol } CO_3^{2-}$$

$$\text{Moles of } Ca^{2+} = (2\ Ca^{2+}\text{spheres})\left(\frac{0.050\ \text{mol } Ca^{2+}}{1\ \text{sphere}}\right) = 0.10\ \text{mol } Ca^{2+}$$

$$\text{Moles of } Cl^- = (4\ Cl^-\text{spheres})\left(\frac{0.050\ \text{mol } Cl^-}{1\ \text{sphere}}\right) = 0.20\ \text{mol } Cl^-$$

The moles of Na^+ and Cl^- do not change. The moles of Ca^{2+} goes to zero, and removes 0.10 mol of CO_3^{2-}.
Moles of remaining CO_3^{2-} = 0.15 mol CO_3^{2-} – 0.10 mol = 0.050 mol CO_3^{2-}
Volume of final solution = 250. mL + 250. mL = 500. mL

$$\text{Molarity of } Na^+ = \frac{0.30\ \text{mol}}{500.\ \text{mL}}\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right) = \mathbf{0.60}\ \boldsymbol{M}\ \mathbf{Na^+}$$

$$\text{Molarity of } Cl^- = \frac{0.20\ \text{mol}}{500.\ \text{mL}}\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right) = \mathbf{0.40}\ \boldsymbol{M}\ \mathbf{Cl^-}$$

$$\text{Molarity of } CO_3^{2-} = \frac{0.050\ \text{mol}}{500.\ \text{mL}}\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right) = \mathbf{0.10}\ \boldsymbol{M}\ \mathbf{CO_3^{2-}}$$

4.93 Plan: Write balanced equations. Use the density to convert volume of fuel to mass of fuel and then use the molar ratios to convert mass of each fuel to the mass of oxygen required for the reaction. Use the conversion factor given to convert mass of oxygen to volume of oxygen.

Solution:

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

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

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

b) The mass of each fuel must be found:

$$\text{Mass (g) of gasoline} = (1.00\ \text{L})\left(\frac{90\%}{100\%}\right)\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right)\left(\frac{0.742\ \text{g}}{1\ \text{mL}}\right) = 667.8\ \text{g gasoline}$$

$$\text{Mass (g) of ethanol} = (1.00\ \text{L})\left(\frac{10\%}{100\%}\right)\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right)\left(\frac{0.789\ \text{g}}{1\ \text{mL}}\right) = 78.9\ \text{g ethanol}$$

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

$$= 2338.64\ \text{g } O_2$$

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

$$= 164.41\ \text{g } O_2$$

Total mass (g) of O_2 = 2338.64 g O_2 + 164.41 g O_2 = 2503.05 = $\mathbf{2.50x10^3\ g\ O_2}$

c) Volume (L) of O_2 = $(2503.05 \text{ g } O_2)\left(\frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2}\right)\left(\frac{22.4 \text{ L}}{1 \text{ mol } O_2}\right)$ = 1752.135 = **$1.75x10^3$ L O_2**

d) Volume (L) of air = $(1752.135 \text{ L } O_2)\left(\frac{100\%}{20.9\%}\right)$ = 8383.42 = **$8.38x10^3$ L air**

4.94 Plan: Write balanced reactions for the complete combustion of gasoline and for the incomplete combustion. Use molar ratios to find the moles of CO_2 and moles of CO produced. Obtain the number of molecules of each gas by multiplying moles by Avogadro's number.

Solution:

a) Complete combustion: 1. $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

Incomplete combustion: 2. $2C_8H_{18}(l) + 17O_2(g) \rightarrow 16CO(g) + 18H_2O(g)$

Assuming a 100-g sample of gasoline, 95%, or 95.0 g, will react by equation 1, and 5.0%, or 5.0 g, will react by equation 2.

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

= $4.00693x10^{24}$ molecules CO_2

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

= $2.10891x10^{23}$ molecules CO

Ratio of CO_2 to CO molecules = $\frac{4.00693\text{x}10^{24} \text{ } CO_2 \text{ molecules}}{2.10891\text{x}10^{23} \text{ CO molecules}}$ = 18.99998 = **19**

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

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

= 292.83 g CO_2

Mass (g) of CO = $(5.0 \text{ g } C_8H_{18})\left(\frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}}\right)\left(\frac{16 \text{ mol CO}}{2 \text{ mol } C_8H_{18}}\right)\left(\frac{28.01 \text{ g CO}}{1 \text{ mol CO}}\right)$ = 9.8091 g CO

Mass ratio of CO_2 to CO = $\frac{292.83 \text{ g } CO_2}{9.8091 \text{ g CO}}$ = 29.85289 = **30**

c) Let x = fraction of CO_2 and y = fraction of CO. For a 1/1 mass ratio of CO_2 to CO, $\frac{(x)(44.01)}{y(28.01)}$ = 1, where 44.01 g/mol is the molar mass of CO_2 and 28.01 g/mol is the molar mass of CO. x + y = 1 or y = 1 – x

Substituting: $\frac{(x)(44.01)}{(1-x)(28.01)} = 1$

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.

4.95 Plan: From the molarity and volume of the base NaOH, find the moles of NaOH and use the molar ratios from the two balanced equations to convert the moles of NaOH to moles of HBr to moles of vitamin C. Use the molar mass of vitamin C to convert moles to grams.

Solution:

Moles of NaOH = $(43.20 \text{ mL NaOH})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{0.1350 \text{ mol NaOH}}{1 \text{ L}}\right)$ = 0.005832 mol NaOH

$$\text{Mass (g) of vitamin C} = (0.005832 \text{ mol NaOH})\left(\frac{1 \text{ mol HBr}}{1 \text{ mol NaOH}}\right)\left(\frac{1 \text{ mol } C_6H_8O_6}{2 \text{ mol HBr}}\right)\left(\frac{176.12 \text{ g } C_6H_8O_6}{1 \text{ mol } C_6H_8O_6}\right)\left(\frac{1 \text{ mg}}{10^{-3} \text{ g}}\right)$$

$= 513.5659 = 513.6$ mg $C_6H_8O_6$

Yes, the tablets have the quantity advertised.

4.96 Plan: Remember that oxidation numbers change in a redox reaction. For the calculations, use the molarity and volume of HCl to find the moles of HCl and use the molar ratios from the balanced equation to convert moles of HCl to moles and then grams of the desired substance.

Solution:

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.

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

$= 199{,}625 =$ **2.00×10^5 g Fe_2O_3**

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

$= 405{,}500 =$ **4.06×10^5 g $FeCl_3$**

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

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

$= 209{,}437.5 =$ **2.09×10^5 g Fe**

$$\text{Mass } FeCl_2 = (2.50 \text{ x } 10^3 \text{ L})\left(\frac{3.00 \text{ mol HCl}}{\text{L}}\right)\left(\frac{1 \text{ mol } FeCl_2}{2 \text{ mol HCl}}\right)\left(\frac{126.75 \text{ g } FeCl_2}{1 \text{ mol } FeCl_2}\right)$$

$= 475{,}312.5 =$ **4.75×10^5 g $FeCl_2$**

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).

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

$= 2.0313$ g $FeCl_3$ (unrounded)

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

$= 0.635452$ g $FeCl_2$ (unrounded)

Ratio = (0.635452 g $FeCl_2$)/(2.0313 g $FeCl_3$) = 0.312830 = **0.313**

4.97 Plan: For part a), assign oxidation numbers to each element. The reactant that is the reducing agent contains an atom that is oxidized (O.N. increases from the left side to the right side of the equation). The reactant that is the oxidizing agent contains an atom that is reduced (O.N. decreases from the left side to the right side of the equation). Use the molar ratios in the balanced equation to convert mass of ammonium perchlorate to moles of product and to moles of Al required in the reaction. Use the density values to convert masses to volumes.

Solution:

a) +4 –8 +6 –6 –3 +2

–3 +1+7 –2 0 +3 –2 +3 –1 +1 –2 +2 –2

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

The O.N. of chlorine decreases from +7 in NH_4ClO_4 to –1 in $AlCl_3$ and is reduced; the O.N. of Al increases from 0 in Al to +3 in the products and is oxidized. The **oxidizing agent is ammonium perchlorate and the reducing agent is aluminum**.

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

= 1276.70 = $\mathbf{1.28x10^3}$ **mol gas**

c) Initial volume:

Volume (L) of NH_4ClO_4 = $\left(50.0 \text{ kg}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ cc}}{1.95 \text{ g}}\right)\left(\frac{1 \text{ mL}}{1 \text{ cc}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$ = 25.6410 L

Mass of Al = $\left(50.0 \text{ kg } NH_4ClO_4\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } NH_4ClO_4}{117.49 \text{ g } NH_4ClO_4}\right)\left(\frac{3 \text{ mol Al}}{3 \text{ mol } NH_4ClO_4}\right)\left(\frac{26.98 \text{ g Al}}{1 \text{ mol Al}}\right)$ = 11481.828 g Al

Volume (L) of Al = $\left(11481.828 \text{ g Al}\right)\left(\frac{1 \text{ cc}}{2.70 \text{ g Al}}\right)\left(\frac{1 \text{ mL}}{1 \text{ cc}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$ = 4.2525 L

Initial volume = 25.6410 L + 4.2525 L = 29.8935 L

Final volume:

Mass (g) of Al_2O_3 = $\left(50.0 \text{ kg } NH_4ClO_4\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } NH_4ClO_4}{117.49 \text{ g } NH_4ClO_4}\right)\left(\frac{1 \text{ mol } Al_2O_3}{3 \text{ mol } NH_4ClO_4}\right)\left(\frac{101.96 \text{ g } Al_2O_3}{1 \text{ mol } Al_2O_3}\right)$

= 14463.64 g Al_2O_3

Volume (L) of Al_2O_3 = $\left(14463.674 \text{ g } Al_2O_3\right)\left(\frac{1 \text{ cc}}{3.97 \text{ g } Al_2O_3}\right)\left(\frac{1 \text{ mL}}{1 \text{ cc}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$ = 3.6432 L

Mass (g) of $AlCl_3$ = $\left(50.0 \text{ kg } NH_4ClO_4\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } NH_4ClO_4}{117.49 \text{ g } NH_4ClO_4}\right)\left(\frac{1 \text{ mol } AlCl_3}{3 \text{ mol } NH_4ClO_4}\right)\left(\frac{133.33 \text{ g } AlCl_3}{1 \text{ mol } AlCl_3}\right)$

= 18913.67 g $AlCl_3$

Volume (L) of $AlCl_3$ = $\left(18913.67 \text{ g } AlCl_3\right)\left(\frac{1 \text{ cc}}{2.44 \text{ g } AlCl_3}\right)\left(\frac{1 \text{ mL}}{1 \text{ cc}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$ = 7.7515 L

Volume (L) of gas = $\left(1276.70 \text{ mol gas}\right)\left(\frac{22.4 \text{ L}}{1 \text{ mol gas}}\right)$ = 28598.08 L

Final volume = 3.6432 L + 7.7515 L + 28598.08 L = 28609.4747 L

Volume change = Final volume – initial volume = (28609.4747 L) – (29.8935 L) = 28579.5812 = $\mathbf{2.86x10^4}$ **L**

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