CHAPTER 3 STOICHIOMETRY OF FORMULAS AND EQUATIONS

END–OF–CHAPTER PROBLEMS

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

Solution:

Al 26.98 amu ≡ 26.98 g/mol Al

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

Cl 35.45 amu ≡ 35.45 g/mol Cl

$$\text{Mass Cl (g)} = (2 \text{ mol Cl})\left(\frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}}\right) = \mathbf{70.90 \text{ g Cl}}$$

3.2 Plan: The molecular formula of sucrose tells us that 1 mole of sucrose contains 12 moles of carbon atoms. Multiply the moles of sucrose by 12 to obtain moles of carbon atoms; multiply the moles of carbon atoms by Avogadro's number to convert from moles to atoms.

Solution:

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

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

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

Solution:

"1 mol of chlorine" could be interpreted as a mole of chlorine atoms or a mole of chlorine molecules, Cl_2. Specify which to avoid confusion. The same problem is possible with other diatomic or polyatomic molecules, e.g., F_2, Br_2, I_2, H_2, O_2, N_2, S_8, and P_4. For these elements, as for chlorine, it is not clear if atoms or molecules are being discussed.

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

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

3.6 Plan: The relative atomic masses of each element can be found by counting the number of atoms of each element and comparing the overall masses of the two samples.

Solution:

a) Balance A: The element on the **left** (green) has the higher molar mass because only 5 green balls are necessary to counterbalance the mass of 6 yellow balls. Since the green ball is heavier, its atomic mass is larger, and therefore its molar mass is larger. Balance B: The element on the **right** (blue) has the higher molar mass since 3 blue balls are heavier than 6 red balls. Since the blue ball is heavier, its atomic mass is larger, and therefore its

molar mass is larger. Balance C: The element on the **left** (orange) has the higher molar mass because 5 orange balls are heavier than 5 purple balls. Since the orange ball is heavier, its atomic mass is larger, and therefore its molar mass is larger. Balance D: The element on the **left** (gray) has the higher molar mass because only 5 gray balls are necessary to counterbalance the mass of 7 red balls. Since the gray ball is heavier, its atomic mass is larger, and therefore its molar mass is larger.
b) The elements on the **right** in Balances A, C, and D have more atoms per gram. The element on the right in each of these balances is lighter. Because these elements are lighter, more atoms are required to make 1 g. In Balance B, the element on the **left** is lighter and would therefore require more atoms to make 1 g.
c) The elements on the **left** in Balances A, C, and D have fewer atoms per gram. Atoms of these elements are heavier, and it takes fewer balls to make 1 g. In Balance B, the element on the **right** is heavier and therefore has fewer atoms per gram.
d) **Neither** element on any of the balances has more atoms per mole. Both the left and right elements have the same number of atoms per mole. The number of atoms per mole ($6.022x10^{23}$) is constant and so is the same for every element.

3.7 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.
Solution:
a) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Sr) + (2 x $\mathcal{M}$ of O) + (2 x $\mathcal{M}$ of H)
= (1 x 87.62 g/mol Sr) + (2 x 16.00 g/mol O) + (2 x 1.008 g/mol H)
= 121.64 g/mol of $Sr(OH)_2$
b) $\mathcal{M}$ = (2 x $\mathcal{M}$ of N) + (3 x $\mathcal{M}$ of O)
= (2 x 14.01 g/mol N) + (3 x 16.00 g/mol O)
= 76.02 g/mol of N_2O_3
c) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Na) + (1 x $\mathcal{M}$ of Cl) + (3 x $\mathcal{M}$ of O)
= (1 x 22.99 g/mol Na) + (1 x 35.45 g/mol Cl) + (3 x 16.00 g/mol O)
= 106.44 g/mol of $NaClO_3$
d) $\mathcal{M}$ = (2 x $\mathcal{M}$ of Cr) + (3 x $\mathcal{M}$ of O)
= (2 x 52.00 g/mol Cr) + (3 x 16.00 g/mol O)
= 152.00 g/mol of Cr_2O_3

3.8 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.
Solution:
a) $\mathcal{M}$ = (3 x $\mathcal{M}$ of N) + (12 x $\mathcal{M}$ of H) + (1 x $\mathcal{M}$ of P) + (4 x $\mathcal{M}$ of O)
= (3 x 14.01 g/mol N) + (12 x 1.008 g/mol H) + (1 x 30.97 g/mol P) + (4 x 16.00 g/mol O)
= 149.10 g/mol of $(NH_4)_3PO_4$
b) $\mathcal{M}$ = (1 x $\mathcal{M}$ of C) + (2 x $\mathcal{M}$ of H) + (2 x $\mathcal{M}$ of Cl)
= (1 x 12.01 g/mol C) + (2 x 1.008 g/mol H) + (2 x 35.45 g/mol Cl)
= 84.93 g/mol of CH_2Cl_2
c) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Cu) + (1 x $\mathcal{M}$ of S) + (9 x $\mathcal{M}$ of O) + (10 x $\mathcal{M}$ of H)
= (1 x 63.55 g/mol Cu) + (1 x 32.07 g/mol S) + (9 x 16.00 g/mol O) + (10 x 1.008 g/mol H)
= 249.70 g/mol of $CuSO_4•5H_2O$
d) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Br) + (3 x $\mathcal{M}$ of F)
= (1 x 79.90 g/mol Br) + (3 x 19.00 g/mol F)
= 136.90 g/mol of BrF_3

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

Solution:
a) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Sn) + (1 x $\mathcal{M}$ of O)
= (1 x 118.7 g/mol Sn) + (1 x 16.00 g/mol O)
= **134.7 g/mol of SnO**
b) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Ba) + (2 x $\mathcal{M}$ of F)
= (1 x 137.3 g/mol Ba) + (2 x 19.00 g/mol F)
= **175.3 g/mol of BaF_2**
c) $\mathcal{M}$ = (2 x $\mathcal{M}$ of Al) + (3 x $\mathcal{M}$ of S) + (12 x $\mathcal{M}$ of O)
= (2 x 26.98 g/mol Al) + (3 x 32.07 g/mol S) + (12 x 16.00 g/mol O)
= **342.17 g/mol of $Al_2(SO_4)_3$**
d) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Mn) + (2 x $\mathcal{M}$ of Cl)
= (1 x 54.94 g/mol Mn) + (2 x 35.45 g/mol Cl)
= **125.84 g/mol of $MnCl_2$**

3.10 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.
Solution:
a) $\mathcal{M}$ = (2 x $\mathcal{M}$ of N) + (4 x $\mathcal{M}$ of O)
= (2 x 14.01 g/mol N) + (4 x 16.00 g/mol O)
= **92.02 g/mol of N_2O_4**
b) $\mathcal{M}$ = (4 x $\mathcal{M}$ of C) + (10 x $\mathcal{M}$ of H) + (1 x $\mathcal{M}$ of O)
= (4 x 12.01 g/mol C) + (10 x 1.008 g/mol H) + (1 x 16.00 g/mol O)
= **74.12 g/mol of C_4H_9OH**
c) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Mg) + (1 x $\mathcal{M}$ of S) + (11 x $\mathcal{M}$ of O) + (14 x $\mathcal{M}$ of H)
= (1 x 24.31 g/mol Mg) + (1 x 32.07 g/mol S) + (11 x 16.00 g/mol O) + (14 x 1.008 g/mol H)
= **246.49 g/mol of $MgSO_4{\bullet}7H_2O$**
d) $\mathcal{M}$ = (1 x $\mathcal{M}$ of Ca) + (4 x $\mathcal{M}$ of C) + (6 x $\mathcal{M}$ of H) + (4 x $\mathcal{M}$ of O)
= (1 x 40.08 g/mol Ca) + (4 x 12.01 g/mol C) + (6 x 1.008 g/mol H) + (4 x 16.00 g/mol O)
= **158.17 g/mol of $Ca(C_2H_3O_2)_2$**

3.11 Plan: Determine the molar mass of each substance, then perform the appropriate molar conversions. To find the mass in part a), multiply the number of moles by the molar mass of the substance. In part b), first convert mass of compound to moles of compound by dividing by the molar mass of the compound. The molecular formula of the compound tells us that 1 mole of compound contains 6 moles of oxygen atoms; use the 1:6 ratio to convert moles of compound to moles of oxygen atoms. In part c), convert mass of compound to moles of compound by dividing by the molar mass of the compound. Since 1 mole of compound contains 6 moles of oxygen atoms, multiply the moles of compound by 6 to obtain moles of oxygen atoms; then multiply by Avogadro's number to obtain the number of oxygen atoms.
Solution:
a) $\mathcal{M}$ of $KMnO_4$ = (1 x $\mathcal{M}$ of K) + (1 x $\mathcal{M}$ of Mn) + (4 x $\mathcal{M}$ of O)
= (1 x 39.10 g/mol K) + (1 x 54.94 g/mol Mn) + (4 x 16.00 g/mol O) = 158.04 g/mol of $KMnO_4$

$$\text{Mass of } KMnO_4 = \left(0.68 \text{ mol } KMnO_4\right)\left(\frac{158.04 \text{ g } KMnO_4}{1 \text{ mol } KMnO_4}\right) = 107.467 = \mathbf{1.1x10^2 \text{ g } KMnO_4}$$

b) $\mathcal{M}$ of $Ba(NO_3)_2$ = (1 x $\mathcal{M}$ of Ba) + (2 x $\mathcal{M}$ of N) + (6 x $\mathcal{M}$ of O)
= (1 x 137.3 g/mol Ba) + (2 x 14.01 g/mol N) + (6 x 16.00 g/mol O) = 261.3 g/mol $Ba(NO_3)_2$

$$\text{Moles of } Ba(NO_3)_2 = \left(8.18 \text{ g } Ba(NO_3)_2\right)\left(\frac{1 \text{ mol } Ba(NO_3)_2}{261.3 \text{ g } Ba(NO_3)_2}\right) = 0.031305 \text{ mol } Ba(NO_3)_2$$

$$\text{Moles of O atoms} = \left(0.031305 \text{ mol } Ba(NO_3)_2\right)\left(\frac{6 \text{ mol O atoms}}{1 \text{ mol } Ba(NO_3)_2}\right) = 0.18783 = \mathbf{0.188 \text{ mol O atoms}}$$

c) $\mathcal{M}$ of $CaSO_4{\bullet}2H_2O$ = (1 x $\mathcal{M}$ of Ca) + (1 x $\mathcal{M}$ of S) + (6 x $\mathcal{M}$ of O) + (4 x $\mathcal{M}$ of H)
= (1 x 40.08 g/mol Ca) + (1 x 32.07 g/mol S) + (6 x 16.00 g/mol O) + (4 x 1.008 g/mol H)
= 172.18 g/mol

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

$$\text{Moles of } CaSO_4\text{•}2H_2O = \left(7.3\text{x}10^{-3}\text{ g } CaSO_4\text{•}2H_2O\right)\left(\frac{1\text{ mol } CaSO_4\text{•}2H_2O}{172.18\text{ g } CaSO_4\text{•}2H_2O}\right) = 4.239749\text{x}10^{-5}\text{ mol}$$

$$\text{Moles of O atoms} = \left(4.239749\text{x}10^{-5}\text{ mol } CaSO_4\text{•}2H_2O\right)\left(\frac{6\text{ mol O atoms}}{1\text{ mol } CaSO_4\text{•}2H_2O}\right)$$

$$= 2.54385\text{x}10^{-5}\text{ mol O atoms}$$

$$\text{Number of O atoms} = \left(2.54385\text{x}10^{-4}\text{ mol O atoms}\right)\left(\frac{6.022\text{x}10^{23}\text{ O atoms}}{1\text{ mol O atoms}}\right)$$

$$= 1.5319\text{x}10^{20} = \mathbf{1.5x10^{20}\ O\ atoms}$$

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

Solution:

a) $\boldsymbol{\mathcal{M}}$ of NO_2 = (1 x $\boldsymbol{\mathcal{M}}$ of N) + (2 x $\boldsymbol{\mathcal{M}}$ of O)
= (1 x 14.01 g/mol N) + (2 x 16.00 g/mol O) = 46.01g/mol of NO_2

$$\text{Moles of } NO_2 = \left(4.6\text{x}10^{21}\text{ molecules } NO_2\right)\left(\frac{1\text{ mol } NO_2}{6.022\text{x}10^{23}\text{ molecules } NO_2}\right) = 7.63866\text{x}10^{-3}\text{ mol } NO_2$$

$$\text{Mass (kg) of } NO_2 = \left(7.63866\text{x}10^{-3}\text{ mol } NO_2\right)\left(\frac{46.01\text{ g } NO_2}{1\text{ mol } NO_2}\right)\left(\frac{1\text{ kg}}{10^3\text{ g}}\right) = 3.51455\text{x}10^{-4} = \mathbf{3.5x10^{-4}\ kg\ NO_2}$$

b) $\boldsymbol{\mathcal{M}}$ of $C_2H_4Cl_2$ = (2 x $\boldsymbol{\mathcal{M}}$ of C) + (4 x $\boldsymbol{\mathcal{M}}$ of H) + (2 x $\boldsymbol{\mathcal{M}}$ of Cl)
= (2 x 12.01g/mol C) + (4 x 1.008 g/mol H) + (2 x 35.45 g/mol Cl) = 98.95 g/mol of $C_2H_4Cl_2$

$$\text{Moles of } C_2H_4Cl_2 = \left(0.0615\text{ g } C_2H_4Cl_2\right)\left(\frac{1\text{ mol } C_2H_4Cl_2}{98.95\text{ g } C_2H_4Cl_2}\right) = 6.21526\text{x}10^{-4}\text{ mol } C_2H_4Cl_2$$

$$\text{Moles of Cl atoms} = \left(6.21526\text{x}10^{-4}\text{ mol } C_2H_4Cl_2\right)\left(\frac{2\text{ mol Cl atoms}}{1\text{ mol } C_2H_4Cl_2}\right) = 1.2431\text{x}10^{-3}$$

$$= \mathbf{1.24x10^{-3}\ mol\ Cl\ atoms}$$

c) $\boldsymbol{\mathcal{M}}$ of SrH_2 = (1 x $\boldsymbol{\mathcal{M}}$ of Sr) + (2 x $\boldsymbol{\mathcal{M}}$ of H) = (1 x 87.62 g/mol Sr) + (2 x 1.008 g/mol H) = 89.64 g/mol of SrH_2

$$\text{Moles of } SrH_2 = \left(5.82\text{ g } SrH_2\right)\left(\frac{1\text{ mol } SrH_2}{89.64\text{ g } SrH_2}\right) = 0.0649264\text{ mol } SrH_2$$

$$\text{Moles of } H^-\text{ ions} = \left(0.0649264\text{ mol } SrH_2\right)\left(\frac{2\text{ mol } H^-}{1\text{ mol } SrH_2}\right) = 0.1298528\text{ mol } H^-\text{ ions}$$

$$\text{Number of } H^-\text{ ions} = \left(0.1298528\text{ mol } H^-\text{ ions}\right)\left(\frac{6.022\text{x}10^{23}\ H^-\text{ ions}}{1\text{ mol } H^-}\right) = 7.81974\text{x}10^{22} = \mathbf{7.82x10^{22}\ H^-\ ions}$$

3.13 Plan: Determine the molar mass of each substance, then perform the appropriate molar conversions. To find the mass in part a), multiply the number of moles by the molar mass of the substance. In part b), first convert the mass of compound in kg to mass in g and divide by the molar mass of the compound to find moles of compound. In part c), convert mass of compound in mg to mass in g and divide by the molar mass of the compound to find

moles of compound. Since 1 mole of compound contains 2 moles of nitrogen atoms, multiply the moles of compound by 2 to obtain moles of nitrogen atoms; then multiply by Avogadro's number to obtain the number of nitrogen atoms.
Solution:
a) $\mathcal{M}$ of $MnSO_4$ = (1 x $\mathcal{M}$ of Mn) + (1 x $\mathcal{M}$ of S) + (4 x $\mathcal{M}$ of O)
= (1 x 54.94 g/mol Mn) + (1 x 32.07 g/mol S) + (4 x 16.00 g/mol O) = 151.01 g/mol of $MnSO_4$

$$\text{Mass (g) of } MnSO_4 = \left(6.44x10^{-2} \text{ mol } MnSO_4\right)\left(\frac{151.01 \text{ g } MnSO_4}{1 \text{ mol } MnSO_4}\right) = 9.725044 = \mathbf{9.73\ g\ MnSO_4}$$

b) $\mathcal{M}$ of $Fe(ClO_4)_3$ = (1 x $\mathcal{M}$ of Fe) + (3 x $\mathcal{M}$ of Cl) + (12 x $\mathcal{M}$ of O)
= (1 x 55.85 g/mol Fe) + (3 x 35.45 g/mol S) + (12 x 16.00 g/mol O)
= 354.20 g/mol of $Fe(ClO_4)_3$

$$\text{Mass (g) of } Fe(ClO_4)_3 = \left(15.8 \text{ kg } Fe(ClO_4)_3\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 1.58 \text{ x } 10^4 \text{ kg } Fe(ClO_4)_3$$

$$\text{Moles of } Fe(ClO_4)_3 = \left(1.58x10^4 \text{ g } Fe(ClO_4)_3\right)\left(\frac{1 \text{ mol } Fe(ClO_4)_3}{354.20 \text{ g } Fe(ClO_4)_3}\right) = 44.6076 = \mathbf{44.6\ mol\ Fe(ClO_4)_3}$$

c) $\mathcal{M}$ of NH_4NO_2 = (2 x $\mathcal{M}$ of N) + (4 x $\mathcal{M}$ of H) + (2 x $\mathcal{M}$ of O)
= (2 x 14.01 g/mol N) + (4 x 1.008 g/mol H) + (2 x 16.00 g/mol O) = 64.05 g/mol NH_4NO_2

$$\text{Mass (g) of } NH_4NO_2 = \left(92.6 \text{ mg } NH_4NO_2\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right) = 0.0926 \text{ g } NH_4NO_2$$

$$\text{Moles of } NH_4NO_2 = \left(0.0926 \text{ g } NH_4NO_2\right)\left(\frac{1 \text{ mol } NH_4NO_2}{64.05 \text{ g } NH_4NO_2}\right) = 1.44575x10^{-3} \text{ mol } NH_4NO_2$$

$$\text{Moles of N atoms} = \left(1.44575x10^{-3} \text{ mol } NH_4NO_2\right)\left(\frac{2 \text{ mol N atoms}}{1 \text{ mol } NH_4NO_2}\right) = 2.8915x10^{-3} \text{ mol N atoms}$$

$$\text{Number of N atoms} = \left(2.8915x10^{-3} \text{ mol N atoms}\right)\left(\frac{6.022 \text{ x } 10^{23} \text{ N atoms}}{1 \text{ mol N atoms}}\right)$$
$$= 1.74126 \text{ x } 10^{21} = \mathbf{1.74\ x\ 10^{21}\ N\ atoms}$$

3.14 Plan: Determine the molar mass of each substance, then perform the appropriate molar conversions. In part a), divide the mass by the molar mass of the compound to find moles of compound. Since 1 mole of compound contains 3 moles of ions (1 mole of Sr^{2+} and 2 moles of F^-), multiply the moles of compound by 3 to obtain moles of ions and then multiply by Avogadro's number to obtain the number of ions. In part b), multiply the number of moles by the molar mass of the substance to find the mass in g and then convert to kg. In part c), divide the number of formula units by Avogadro's number to find moles; multiply the number of moles by the molar mass to obtain the mass in g and then convert to mg.
Solution:
a) $\mathcal{M}$ of SrF_2 = (1 x $\mathcal{M}$ of Sr) + (2 x $\mathcal{M}$ of F)
= (1 x 87.62 g/mol Sr) + (2 x 19.00 g/mol F) = 125.62 g/mol of SrF_2

$$\text{Moles of } SrF_2 = \left(38.1 \text{ g } SrF_2\right)\left(\frac{1 \text{ mol } SrF_2}{125.62 \text{ g } SrF_2}\right) = 0.303296 \text{ mol } SrF_2$$

$$\text{Moles of ions} = \left(0.303296 \text{ mol } SrF_2\right)\left(\frac{3 \text{ mol ions}}{1 \text{ mol } SrF_2}\right) = 0.909888 \text{ mol ions}$$

$$\text{Number of ions} = \left(0.909888 \text{ mol ions}\right)\left(\frac{6.022 \text{ x } 10^{23} \text{ ions}}{1 \text{ mol ions}}\right) = 5.47935x10^{23} = \mathbf{5.48x10^{23}\ ions}$$

b) $\mathcal{M}$ of $CuCl_2 \cdot 2H_2O$ = (1 x $\mathcal{M}$ of Cu) + (2 x $\mathcal{M}$ of Cl) + (4 x $\mathcal{M}$ of H) + (2 x $\mathcal{M}$ of O)
= (1 x 63.55 g/mol Cu) + (2 x 35.45 g/mol Cl) + (4 x 1.008 g/mol H) + (2 x 16.00 g/mol O)
= 170.48 g/mol of $CuCl_2 \cdot 2H_2O$

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

Mass (g) of $CuCl_2•2H_2O$ = $\left(3.58 \text{ mol } CuCl_2•2H_2O\right)\left(\frac{170.48 \text{ g } CuCl_2•2H_2O}{1 \text{ mol } CuCl_2•2H_2O}\right)$ = 610.32 g $CuCl_2•2H_2O$

Mass (kg) of $CuCl_2•2H_2O$ = $\left(610.32 \text{ g } CuCl_2•2H_2O\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)$ = 0.61032 = **0.610 kg $CuCl_2•2H_2O$**

c) $\mathcal{M}$ of $Bi(NO_3)_3•5H_2O$ = (1 x $\mathcal{M}$ of Bi) + (3 x $\mathcal{M}$ of N) + (10 x $\mathcal{M}$ of H) + (14 x $\mathcal{M}$ of O)
= (1 x 209.0 g/mol Bi) + (3 x 14.01 g/mol N) + (10 x 1.008 g/mol H)
+ (14 x 16.00 g/mol H) = 485.11 g/mol of
$Bi(NO_3)_3•5H_2O$

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

Moles of $Bi(NO_3)_3•5H_2O$ = $\left(2.88 \text{x} 10^{22} \text{ FU}\right)\left(\frac{1 \text{ mol}}{6.022 \text{x} 10^{23} \text{ FU}}\right)$ = 0.047825 mol $Bi(NO_3)_3•5H_2O$

Mass (g) of $Bi(NO_3)_3•5H_2O$ = $\left(0.047825 \text{ mol } Bi(NO_3)_3•5H_2O\right)\left(\frac{485.1 \text{ g } Bi(NO_3)_3•5H_2O}{1 \text{ mol } Bi(NO_3)_3•5H_2O}\right)$ = 23.1999 g

Mass (mg) of $Bi(NO_3)_3•5H_2O$ = $\left(23.1999 \text{ g } Bi(NO_3)_3•5H_2O\right)\left(\frac{1 \text{ mg}}{10^{-3} \text{ g}}\right)$
= 23199.9 = **$2.32 \text{x} 10^4$ mg $Bi(NO_3)_3•5H_2O$**

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

Solution:

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

$\mathcal{M}$ of Cu_2CO_3 = (2 x $\mathcal{M}$ of Cu) + (1 x $\mathcal{M}$ of C) + (3 x $\mathcal{M}$ of O)
= (2 x 63.55 g/mol Cu) + (1 x 12.01 g/mol C) + (3 x 16.00 g/mol O) = 187.11 g/mol of Cu_2CO_3

Mass (g) of Cu_2CO_3 = $\left(8.35 \text{ mol } Cu_2CO_3\right)\left(\frac{187.11 \text{ g } Cu_2CO_3}{1 \text{ mol } Cu_2CO_3}\right)$ = 1562.4 = **$1.56 \text{x} 10^3$ g Cu_2CO_3**

b) Dinitrogen pentaoxide has the formula N_2O_5. Di- indicates 2 N atoms and penta- indicates 5 O atoms.

$\mathcal{M}$ of N_2O_5 = (2 x $\mathcal{M}$ of N) + (5 x $\mathcal{M}$ of O)
= (2 x 14.01 g/mol N) + (5 x 16.00 g/mol O) = 108.02 g/mol of N_2O_5

Moles of N_2O_5 = $\left(4.04 \text{x} 10^{20} \text{ } N_2O_5 \text{ molecules}\right)\left(\frac{1 \text{ mol } N_2O_5}{6.022 \text{x} 10^{23} \text{ } N_2O_5 \text{ molecules}}\right)$ = $6.7087 \text{x} 10^{-4}$ mol N_2O_5

Mass (g) of N_2O_5 = $\left(6.7087 \text{x} 10^{-4} \text{ mol } N_2O_5\right)\left(\frac{108.02 \text{ g } N_2O_5}{1 \text{ mol } N_2O_5}\right)$ = 0.072467 = **0.0725 g N_2O_5**

c) The correct formula for this ionic compound is $NaClO_4$; Na has a charge of +1 (Group 1 ion) and the perchlorate ion is ClO_4^-.

$\mathcal{M}$ of $NaClO_4$ = (1x $\mathcal{M}$ of Na) + (1 x $\mathcal{M}$ of Cl) + (4 x $\mathcal{M}$ of O)
= (1 x 22.99 g/mol Na) + (1 x 35.45 g/mol Cl) + (4 x 16.00 g/mol O) = 122.44 g/mol of $NaClO_4$

Moles of $NaClO_4$ = $\left(78.9 \text{ g } NaClO_4\right)\left(\frac{1 \text{ mol } NaClO_4}{122.44 \text{ g } NaClO_4}\right)$ = 0.644397 = **0.644 mol $NaClO_4$**

FU = formula units

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

$$= 3.88056\text{x}10^{23} = \mathbf{3.88x10^{23}\ FU\ NaClO_4}$$

d) Number of Na^+ ions $= (3.88056\text{x}10^{23} \text{ FU NaClO}_4)\left(\frac{1 \text{ Na}^+ \text{ ion}}{1 \text{ FU NaClO}_4}\right)$ = $\mathbf{3.88x10^{23}\ Na^+\ ions}$

Number of ClO_4^- ions $= (3.88056\text{x}10^{23} \text{ FU NaClO}_4)\left(\frac{1 \text{ ClO}_4^- \text{ ion}}{1 \text{ FU NaClO}_4}\right)$ = $\mathbf{3.88x10^{23}\ ClO_4^-\ ions}$

Number of Cl atoms $= (3.88056\text{x}10^{23} \text{ FU NaClO}_4)\left(\frac{1 \text{ Cl atom}}{1 \text{ FU NaClO}_4}\right)$ = $\mathbf{3.88x10^{23}}$ **Cl atoms**

Number of O atoms $= (3.88056\text{x}10^{23} \text{ FU NaClO}_4)\left(\frac{4 \text{ O atoms}}{1 \text{ FU NaClO}_4}\right)$ = $\mathbf{1.55x10^{24}}$ **O atoms**

3.16 Plan: The formula of each compound must be determined from its name. The molar mass for each formula comes from the formula and atomic masses from the periodic table. Determine the molar mass of each substance, then perform the appropriate molar conversions. In part a), multiply the moles by the molar mass of the compound to find the mass of the sample. In part b), divide the number of molecules by Avogadro's number to find moles; multiply the number of moles by the molar mass to obtain the mass. In part c), divide the mass by the molar mass to find moles of compound and multiply moles by Avogadro's number to find the number of formula units. In part d), use the fact that each formula unit contains 2 Li ions, 1 sulfate ion, 1 S atom, and 4 O atoms.

Solution:

a) Sulfate is a polyatomic anion with the formula, SO_4^{2-}. Chromium(III) indicates Cr^{3+}. Decahydrate indicates 10 water molecules ("waters of hydration"). The correct formula for this ionic compound is $Cr_2(SO_4)_3 \bullet 10H_2O$.

$\mathcal{M}$ of $Cr_2(SO_4)_3 \bullet 10H_2O$ = (2 x $\mathcal{M}$ of Cr) + (3 x $\mathcal{M}$ of S) + (22 x $\mathcal{M}$ of O) + (20 x $\mathcal{M}$ of H)

= (2 x 52.00 g/mol Cr) + (3 x 32.07 g/mol S) + (22 x 16.00 g/mol O) + (20 x 1.008 g/mol H)

= 572.4 g/mol of $Cr_2(SO_4)_3 \bullet 10H_2O$

$$\text{Mass (g) of Cr}_2(\text{SO}_4)_3\bullet 10\text{H}_2\text{O} = (8.42 \text{ mol Cr}_2(\text{SO}_4)_3\bullet 10\text{H}_2\text{O})\left(\frac{572.4 \text{ g}}{\text{mol}}\right)$$

$$= 4819.608 = \mathbf{4.82x10^3\ g\ Cr_2(SO_4)_3\bullet 10H_2O}$$

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

$\mathcal{M}$ of Cl_2O_7 = (2 x $\mathcal{M}$ of Cl) + (7 x $\mathcal{M}$ of O)

= (2 x 35.45 g/mol Cl) + (7 x 16.00 g/mol O) = 182.9 g/mol of Cl_2O_7

$$\text{Moles of Cl}_2\text{O}_7 = (1.83 \times 10^{24} \text{ molecules Cl}_2\text{O}_7)\left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}}\right) = 3.038858 \text{ mol Cl}_2\text{O}_7$$

$$\text{Mass (g) of Cl}_2\text{O}_7 = (3.038858 \text{ mol Cl}_2\text{O}_7)\left(\frac{182.9 \text{ g Cl}_2\text{O}_7}{1 \text{ mol}}\right) = 555.807 = \mathbf{5.56x10^2\ g\ Cl_2O_7}$$

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

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

= (2 x 6.941 g/mol Li) + (1 x 32.07 g/mol S) + (4 x 16.00 g/mol O) = 109.95 g/mol of Li_2SO_4

$$\text{Moles of Li}_2\text{SO}_4 = (6.2 \text{ g Li}_2\text{SO}_4)\left(\frac{1 \text{ mol Li}_2\text{SO}_4}{109.95 \text{ g Li}_2\text{SO}_4}\right) = 0.056389 = \mathbf{0.056\ mol\ Li_2SO_4}$$

$$\text{FU of Li}_2\text{SO}_4 = (0.056389 \text{ mol Li}_2\text{SO}_4)\left(\frac{6.022 \times 10^{23} \text{ FU}}{1 \text{ mol Li}_2\text{SO}_4}\right) = 3.3957\text{x}10^{22} = \mathbf{3.4x10^{22}\ FU\ Li_2SO_4}$$

d) Number of Li^+ ions = $\left(3.3957\text{x}10^{22}\text{ FU }Li_2SO_4\right)\left(\frac{2\ Li^+\text{ ions}}{1\text{ FU }Li_2SO_4}\right)$ = $6.7914\text{x}10^{22}$ = **$6.8\text{x}10^{22}$ Li^+ ions**

Number of SO_4^{2-} ions = $\left(3.3957\text{x}10^{22}\text{ FU }Li_2SO_4\right)\left(\frac{1\ SO_4^{2-}\text{ ion}}{1\text{ FU }Li_2SO_4}\right)$ = $3.3957\text{x}10^{22}$ = **$3.4\text{x}10^{22}$ SO_4^{2-} ions**

Number of S atoms = $\left(3.3957\text{x}10^{22}\text{ FU }Li_2SO_4\right)\left(\frac{1\text{ S atom}}{1\text{ FU }Li_2SO_4}\right)$ = $3.3957\text{x}10^{22}$ = **$3.4\text{x}10^{22}$ S atoms**

Number of O atoms = $\left(3.3957\text{x}10^{22}\text{ FU }Li_2SO_4\right)\left(\frac{4\text{ O atoms}}{1\text{ FU }Li_2SO_4}\right)$ = $1.3583\text{x}10^{23}$ = **$1.4\text{x}10^{23}$ O atoms**

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

Solution:
a) Ammonium bicarbonate is an ionic compound consisting of ammonium ions, NH_4^+ and bicarbonate ions, HCO_3^-. The formula of the compound is NH_4HCO_3.
$\mathcal{M}$ of NH_4HCO_3 = (1 x $\mathcal{M}$ of N) + (5 x $\mathcal{M}$ of H) + (1 x $\mathcal{M}$ of C) + (3 x $\mathcal{M}$ of O)
= (1 x 14.01 g/mol N) + (5 x 1.008 g/mol H) + (1 x 12.01 g/mol C) + (3 x 16.00 g/mol O)
= 79.06 g/mol of NH_4HCO_3
There are 5 moles of H in 1 mole of NH_4HCO_3.

Mass (g) of H = $\left(5\text{ mol H}\right)\left(\frac{1.008\text{ g H}}{1\text{ mol H}}\right)$ = 5.040 g H

Mass percent = $\frac{\text{total mass H}}{\text{molar mass of compound}}(100)$ = $\frac{5.040\text{ g H}}{79.06\text{ g }NH_4HCO_3}(100)$ = 6.374905 = **6.375% H**

b) Sodium dihydrogen phosphate heptahydrate is a salt that consists of sodium ions, Na^+, dihydrogen phosphate ions, $H_2PO_4^-$, and seven waters of hydration. The formula is $NaH_2PO_4{\bullet}7H_2O$. Note that the waters of hydration are included in the molar mass.
$\mathcal{M}$ of $NaH_2PO_4{\bullet}7H_2O$ = (1 x $\mathcal{M}$ of Na) + (16 x $\mathcal{M}$ of H) + (1 x $\mathcal{M}$ of P) + (11 x $\mathcal{M}$ of O)
= (1 x 22.99 g/mol Na) + (16 x 1.008 g/mol H) + (1 x 30.97 g/mol P) + (11 x 16.00 g/mol O)
= 246.09 g/mol $NaH_2PO_4{\bullet}7H_2O$
There are 11 moles of O in 1 mole of $NaH_2PO_4{\bullet}7H_2O$.

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

Mass percent = $\frac{\text{total mass O}}{\text{molar mass of compound}}(100)$ = $\frac{176.00\text{ g O}}{246.09\text{ g }NaH_2PO_4{\bullet}7H_2O}(100)$
= 71.51855 = **71.52% O**

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

Solution:
a) Strontium periodate is an ionic compound consisting of strontium ions, Sr^{2+} and periodate ions, IO_4^-. The formula of the compound is $Sr(IO_4)_2$.
$\mathcal{M}$ of $Sr(IO_4)_2$ = (1 x $\mathcal{M}$ of Sr) + (2 x $\mathcal{M}$ of I) + (8 x $\mathcal{M}$ of O)
= (1 x 87.62 g/mol Sr) + (2 x 126.9 g/mol I) + (8 x 16.00 g/mol O)
= 469.4 g/mol of $Sr(IO_4)_2$
There are 2 moles of I in 1 mole of $Sr(IO_4)_2$.

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

$$\text{Mass percent} = \frac{\text{total mass I}}{\text{molar mass of compound}}(100) = \frac{253.8\text{ g I}}{469.4\text{ g Sr(IO}_4)_2}(100) = 54.0690 = \mathbf{54.07\%\ I}$$

b) Potassium permanganate is an ionic compound consisting of potassium ions, K^+ and permanganate ions, MnO_4^-. The formula of the compound is $KMnO_4$.

$\mathcal{M}$ of $KMnO_4$ = (1 x $\mathcal{M}$ of K) + (1 x $\mathcal{M}$ of Mn) + (4 x $\mathcal{M}$ of O)
= (1 x 39.10 g/mol K) + (1 x 54.94 g/mol Mn) + (4 x 16.00 g/mol O)
= 158.04 g/mol of $KMnO_4$

There is 1 mole of Mn in 1 mole of $KMnO_4$.

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

$$\text{Mass percent} = \frac{\text{total mass Mn}}{\text{molar mass of compound}}(100) = \frac{54.94\text{ g Mn}}{158.04\text{ g KMnO}_4}(100) = 34.76335 = \mathbf{34.76\%\ Mn}$$

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

Solution:

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

$\mathcal{M}$ of $Pt(Cl)_2(NH_3)_2$ = (1 x $\mathcal{M}$ of Pt) + (2 x $\mathcal{M}$ of Cl) + (2 x $\mathcal{M}$ of N) + (6 x $\mathcal{M}$ of H)
= (1 x 195.1 g/mol Pt) + (2 x 35.45 g/mol Cl) + (2 x 14.01 g/mol N) + (6 x 1.008 g/mol H)
= 300.1 g/mol of $Pt(Cl)_2(NH_3)_2$

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

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

$$\text{Number of H atoms} = (5.88\text{ mol H atoms})\left(\frac{6.022\times10^{23}\text{ H atoms}}{1\text{ mol H atoms}}\right) = 3.540936\text{x}10^{24} = \mathbf{3.5x10^{24}\ H\ atoms}$$

3.20 Plan: Determine the molar mass of propane. Divide the given mass by the molar mass to find the moles. Since each mole of propane contains 3 moles of carbon, multiply the moles of propane by 3 to obtain moles of C atoms. Multiply the moles of C by its molar mass to obtain mass of carbon.

Solution:

a) The formula of propane is C_3H_8.

$\mathcal{M}$ of C_3H_8 = (3 x $\mathcal{M}$ of C) + (8 x $\mathcal{M}$ of H) = (3 x 12.01 g/mol C) + (8x1.008 g/mol H) = 44.09 g/mol

$$\text{Moles of C}_3\text{H}_8 = (85.5\text{ g C}_3\text{H}_8)\left(\frac{1\text{ mol C}_3\text{H}_8}{44.09\text{ g C}_3\text{H}_8}\right) = 1.939215 = \mathbf{1.94\ mol\ C_3H_8}$$

$$\text{b) Moles of C} = (1.939215\text{ mol C}_3\text{H}_8)\left(\frac{3\text{ mol C}}{1\text{ mol C}_3\text{H}_8}\right) = 5.817645\text{ mol C}$$

$$\text{Mass (g) of C} = (5.817645\text{ mol C})\left(\frac{12.01\text{ g C}}{1\text{ mol C}}\right) = 69.86992 = \mathbf{69.9\ g\ C}$$

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

Solution:

Name	Formula	Molar Mass (g/mol)
Potassium nitrate	KNO_3	101.11
Ammonium nitrate	NH_4NO_3	80.05
Ammonium sulfate	$(NH_4)_2SO_4$	132.15
Urea	$CO(NH_2)_2$	60.06

Mass % N in potassium nitrate = $\frac{(1\text{ mol N})(14.01\text{ g/mol N})}{101.11\text{ g/mol}}$ x 100 = 13.856196 = **13.86% N**

Mass % N in ammonium nitrate = $\frac{(2\text{ mol N})(14.01\text{ g/mol N})}{80.05\text{ g/mol}}$ x 100 = 35.003123 = **35.00% N**

Mass % N in ammonium sulfate = $\frac{(2\text{ mol N})(14.01\text{ g/mol N})}{132.15\text{ g/mol}}$ x 100 = 21.20318 = **21.20% N**

Mass % N in urea = $\frac{(2\text{ mol N})(14.01\text{ g/mol N})}{60.06\text{ g/mol}}$ x 100 = 46.6533 = **46.65% N**

Rank is $\mathbf{CO(NH_2)_2 > NH_4NO_3 > (NH_4)_2SO_4 > KNO_3}$

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

Solution:

Lead(II) sulfide is composed of Pb^{2+} and S^{2-} ions and has a formula of PbS.

$\mathcal{M}$ of PbS = (1 x $\mathcal{M}$ of Pb) + (1 x $\mathcal{M}$ of S) = (1 x 207.2 g/mol Pb) + (1 x 32.07 g/mol S) = 239.3 g/mol

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

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

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

b) Volume (cm^3) = $\left(1.00\text{ dm}^3\text{ PbS}\right)\left(\frac{(0.1\text{ m})^3}{(1\text{ dm})^3}\right)\left(\frac{(1\text{ cm})^3}{(10^{-2}\text{ m})^3}\right)$ = 1.00x10^3 cm^3

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

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

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

Number of lead atoms =

$$(31.17426 \text{ mol Pb})\left(\frac{6.022\text{x}10^{23} \text{ Pb atoms}}{1 \text{ mol Pb}}\right) = 1.87731\text{x}10^{25} = \mathbf{1.88x10^{25}\ Pb\ atoms}$$

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

Solution:

$$\text{Mass of Fe} = \left(\frac{0.33\% \text{ Fe}}{100\% \text{ Hb}}\right)\left(\frac{6.8\text{x}10^4 \text{ g}}{\text{mol}}\right) = 224.4 \text{ g Fe}$$

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

Thus, there are **4 Fe^{2+}/molecule Hb**.

3.24 Plan: Remember that the molecular formula tells the *actual* number of moles of each element in one mole of compound.

Solution:

a) No, this information does not allow you to obtain the molecular formula. You can obtain the empirical formula from the number of moles of each type of atom in a compound, but not the molecular formula.

b) Yes, you can obtain the molecular formula from the mass percentages and the total number of atoms.

Plan:

1) Assume a 100.0 g sample and convert masses (from the mass % of each element) to moles using molar mass.

2) Identify the element with the lowest number of moles and use this number to divide into the number of moles for each element. You now have at least one elemental mole ratio (the one with the smallest number of moles) equal to 1.00 and the remaining mole ratios that are larger than one.

3) Examine the numbers to determine if they are whole numbers. If not, multiply each number by a whole-number factor to get whole numbers for each element. You will have to use some judgment to decide when to round. Write the empirical formula using these whole numbers.

4) Check the total number of atoms in the empirical formula. If it equals the total number of atoms given then the empirical formula is also the molecular formula. If not, then divide the total number of atoms given by the total number of atoms in the empirical formula. This should give a whole number. Multiply the number of atoms of each element in the empirical formula by this whole number to get the molecular formula. If you do not get a whole number when you divide, return to step 3 and revise how you multiplied and rounded to get whole numbers for each element.

Roadmap:

Mass (g) of each element (express mass percent directly as grams)

↓ Divide by $\mathcal{M}$ (g/mol)

Amount (mol) of each element

↓ Use numbers of moles as subscripts

Preliminary empirical formula

↓ Change to integer subscripts

Empirical formula

↓ Divide total number of atoms in molecule by the number of atoms in the empirical formula and multiply the empirical formula by that factor

Molecular formula

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

1) Follow steps 1–3 in part b).

2) Compare the number of atoms given for the one element to the number in the empirical formula. Determine the factor the number in the empirical formula must be multiplied by to obtain the given number of atoms for that element. Multiply the empirical formula by this number to get the molecular formula.

Roadmap:

(Same first three steps as in b).

Empirical formula

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

Molecular formula

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

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

Roadmap:

Structural formula

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

Molecular formula

3.25 <u>Plan:</u> Examine the number of atoms of each type in the compound. Divide all atom numbers by the common factor that results in the lowest whole-number values. Add the molar masses of the atoms to obtain the empirical formula mass.

<u>Solution:</u>

a) C_2H_4 has a ratio of 2 carbon atoms to 4 hydrogen atoms, or 2:4. This ratio can be reduced to 1:2, so that the empirical formula is $\mathbf{CH_2}$. The empirical formula mass is 12.01 g/mol C + 2(1.008 g/mol H) = **14.03 g/mol**.

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

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

d) The ratio of elements is 3 atoms of barium to 2 atoms of phosphorus to 8 atoms of oxygen, or 3:2:8. This ratio cannot be further reduced, so the empirical formula is also $\mathbf{Ba_3(PO_4)_2}$, with a formula mass of 3(137.3 g/mol Ba) + 2(30.97 g/mol P) + 8(16.00 g/mol O) = **601.8 g/mol**.

e) The ratio of atoms is 4:16, or 1:4. The empirical formula is $\mathbf{TeI_4}$, and the formula mass is 127.6 g/mol Te + 4(126.9 g/mol I) = **635.2 g/mol**.

3.26 Plan: Examine the number of atoms of each type in the compound. Divide all atom numbers by the common factor that results in the lowest whole-number values. Add the molar masses of the atoms to obtain the empirical formula mass.

Solution:

a) C_4H_8 has a ratio of 4 carbon atoms to 8 hydrogen atoms, or 4:8. This ratio can be reduced to 1:2, so that the empirical formula is $\mathbf{CH_2}$. The empirical formula mass is 12.01 g/mol C + 2(1.008 g/mol H) = **14.03 g/mol**.

b) $C_3H_6O_3$ has a ratio of atoms of 3:6:3, or 1:2:1. The empirical formula is $\mathbf{CH_2O}$ and its empirical formula mass is 12.01 g/mol C + 2(1.008 g/mol H) + 16.00 g/mol O = **30.03 g/mol**.

c) P_4O_{10} has a ratio of 4 P atoms to 10 O atoms, or 4:10. This ratio can be reduced to 2:5, so that the empirical formula is $\mathbf{P_2O_5}$. The empirical formula mass is 2(30.97 g/mol P) + 5(16.00 g/mol O) = **141.94 g/mol**.

d) $Ga_2(SO_4)_3$ has a ratio of 2 atoms of gallium to 3 atoms of sulfur to 12 atoms of oxygen, or 2:3:12. This ratio cannot be further reduced, so the empirical formula is also $\mathbf{Ga_2(SO_4)_3}$, with a formula mass of 2(69.72 g/mol Ga) + 3(32.07 g/mol S) + 12(16.00 g/mol O) = **427.6 g/mol**.

e) Al_2Br_6 has a ratio of atoms of 2:6, or 1:3. The empirical formula is $\mathbf{AlBr_3}$, and the formula mass is 26.98 g/mol Al + 3(79.90 g/mol Br) = **266.7 g/mol**.

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

Solution:

Only approximate whole-number values are needed.

a) CH_2 has empirical mass equal to 12.01 g/mol C + 2(1.008 g/mol C) = 14.03 g/mol

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

Multiplying the subscripts in CH_2 by 3 gives $\mathbf{C_3H_6}$.

b) NH_2 has empirical mass equal to 14.01 g/mol N + 2(1.008 g/mol H) = 16.03 g/mol

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{32.05 \text{ g/mol}}{16.03 \text{ g/mol}}\right) = 2$$

Multiplying the subscripts in NH_2 by 2 gives $\mathbf{N_2H_4}$.

c) NO_2 has empirical mass equal to 14.01 g/mol N + 2(16.00 g/mol O) = 46.01 g/mol

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

Multiplying the subscripts in NO_2 by 2 gives $\mathbf{N_2O_4}$.

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

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

Multiplying the subscripts in CHN by 5 gives $\mathbf{C_5H_5N_5}$.

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

Solution:

Only approximate whole-number values are needed.

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

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

Multiplying the subscripts in CH by 6 gives $\mathbf{C_6H_6}$.

b) $C_3H_6O_2$ has empirical mass equal to 3(12.01 g/mol C) + 6(1.008 g/mol H) + 2(16.00 g/mol O) = 74.08 g/mol

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

Multiplying the subscripts in $C_3H_6O_2$ by 1 gives $\mathbf{C_3H_6O_2}$.

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

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

Multiplying the subscripts in HgCl by 2 gives $\mathbf{Hg_2Cl_2}$.

d) $C_7H_4O_2$ has empirical mass equal to 7(12.01 g/mol C) + 4(1.008 g/mol H) + 2(16.00 g/mol O) = 120.10 g/mol

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

Multiplying the subscripts in $C_7H_4O_2$ by 2 gives $\mathbf{C_{14}H_8O_4}$.

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

Solution:

a) 0.063 mol Cl and 0.22 mol O: preliminary formula is $Cl_{0.063}O_{0.22}$

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

$$Cl_{\frac{0.063}{0.063}}O_{\frac{0.22}{0.063}} \rightarrow Cl_1O_{3.5}$$

The formula is $Cl_1O_{3.5}$, which in whole numbers (x 2) is $\mathbf{Cl_2O_7}$.

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

$$\text{Moles of Si} = (2.45 \text{ g Si})\left(\frac{1 \text{ mol Si}}{28.09 \text{ g Si}}\right) = 0.08722 \text{ mol Si}$$

$$\text{Moles of Cl} = (12.4 \text{ g Cl})\left(\frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}}\right) = 0.349788 \text{ mol Cl}$$

Preliminary formula is $Si_{0.08722}Cl_{0.349788}$

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

$$Si_{\frac{0.08722}{0.08722}}Cl_{\frac{0.349788}{0.349788}} \rightarrow Si_1Cl_4$$

The empirical formula is $\mathbf{SiCl_4}$.

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

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

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

Preliminary formula is $C_{2.2731}O_{4.5438}$

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

$$C_{\frac{2.2731}{2.2731}}O_{\frac{4.5438}{2.2731}} \rightarrow C_1O_2$$

The empirical formula is $\mathbf{CO_2}$.

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

Solution:

a) 0.039 mol Fe and 0.052 mol O: preliminary formula is $Fe_{0.039}O_{0.052}$

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

$$Fe_{\frac{0.039}{0.039}}O_{\frac{0.052}{0.039}} \rightarrow Fe_1O_{1.33}$$

The formula is $Fe_1O_{1.33}$, which in whole numbers (x 3) is $\mathbf{Fe_3O_4}$.

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

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

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

Preliminary formula is $P_{0.029157}Br_{0.087484}$

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

$$P_{\frac{0.029157}{0.029157}}Br_{\frac{0.087484}{0.029157}} \rightarrow P_1Br_3$$

The empirical formula is $\mathbf{PBr_3}$.

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

79.9% C and 100 – 79.9 = 20.1% H

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

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

Preliminary formula is $C_{6.6528}H_{19.940}$

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

$$C_{\frac{6.6528}{6.6528}}H_{\frac{19.940}{6.6528}} \rightarrow C_1H_3$$

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

3.31 Plan: The moles of the metal are known, and the moles of fluorine atoms may be found in part a) from the M:F mole ratio in the compound formula. In part b), convert moles of F atoms to mass and subtract the mass of F from the mass of MF_2 to find the mass of M. In part c), divide the mass of M by moles of M to determine the molar mass of M which can be used to identify the element.

Solution:

a) Determine the moles of fluorine.

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

b) Determine the mass of M.

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

Mass (g) of M = MF_2(g) – F(g) = 46.8 g – 22.8 g = **24.0 g M**

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

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

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

3.32 Plan: The moles of the metal oxide are known, and the moles of oxygen atoms may be found in part a) from the compound:oxygen mole ratio in the compound formula. In part b), convert moles of O atoms to mass and subtract the mass of O from the mass of M_2O_3 to find the mass of M. In part c), find moles of M from the compound:M mole ratio and divide the mass of M by moles of M to determine the molar mass of M which can be used to identify the element.

Solution:

a) Determine the moles of oxygen.

$$\text{Moles of O} = (0.370 \text{ mol } M_2O_3)\left(\frac{3 \text{ mol O}}{1 \text{ mol } M_2O_3}\right) = \mathbf{1.11\ mol\ O}$$

b) Determine the mass of M.

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

Mass(g) of M = M_2O_3(g) – O(g) = 55.4 g (M + O) – 17.76 = 37.64 = **37.6 g M**

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

$$\text{Moles M} = (0.370 \text{ mol } M_2O_3)\left(\frac{2 \text{ mol M}}{1 \text{ mol } M_2O_3}\right) = 0.740 \text{ mol M}$$

The molar mass is needed to identify the element.

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

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

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

Solution:

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

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

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

Preliminary formula is $C_{5.7952}H_{8.2738}O_{1.38125}$

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

$$C_{\frac{5.7952}{1.38125}}H_{\frac{8.2738}{1.38125}}O_{\frac{1.38125}{1.38125}} \rightarrow C_{4.2}H_6O_1$$

The carbon value is not close enough to a whole number to round the value. The smallest number that 4.20 may be multiplied by to get close to a whole number is 5. (You may wish to prove this to yourself.) All three ratios need to be multiplied by five: $5(C_{4.2}H_6O_1) = C_{21}H_{30}O_5$.

The empirical formula mass is = 21(12.01 g/mol C) + 30(1.008 g/mol H) + 5(16.00 g/mol O) = 362.45 g/mol

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{362.47 \text{ g/mol}}{362.45 \text{ g/mol}}\right) = 1$$

The empirical formula mass and the molar mass given are the same, so the empirical and the molecular formulas are the same. The molecular formula is $\mathbf{C_{21}H_{30}O_5}$.

3.34 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO_2, and all of the hydrogen present in the sample is found in the hydrogen of H_2O. Convert the mass of CO_2 to moles and use the ratio between CO_2 and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H_2O. The moles of oxygen are more difficult to find, because additional O_2 was added to cause the combustion reaction. Subtracting the masses of C and H from the mass of the sample gives the mass of O. Convert the mass of O to moles of O. Take the moles of C, H, and O and divide by the smallest value to convert to whole numbers to get the empirical formula.

Determine the empirical formula mass and compare it to the molar mass given in the problem to see how the empirical and molecular formulas are related. Finally, determine the molecular formula.
Solution:

$$\text{Moles of C} = \left(0.449\text{ g } CO_2\right)\left(\frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2}\right)\left(\frac{1\text{ mol C}}{1\text{ mol } CO_2}\right) = 0.010202\text{ mol C}$$

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

$$\text{Moles of H} = \left(0.184\text{ g } H_2O\right)\left(\frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O}\right)\left(\frac{2\text{ mol H}}{1\text{ mol } H_2O}\right) = 0.020422\text{ mol H}$$

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

Mass (g) of O = Sample mass – (mass of C + mass of H)
= 0.1595 g – (0.122526 g C + 0.020585 g H) = 0.016389 g O

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

Preliminary formula = $C_{0.010202}H_{0.020422}O_{0.0010243}$
Converting to integer subscripts (dividing all by the smallest subscript):

$$C_{\frac{0.010202}{0.0010243}}H_{\frac{0.020422}{0.0010243}}O_{\frac{0.0010243}{0.0010243}} \rightarrow C_{10}H_{20}O_1$$

Empirical formula = $C_{10}H_{20}O$
Empirical formula mass = 10(12.01 g/mol C) + 20(1.008 g/mol H) + 1(16.00 g/mol O) = 156.26 g/mol
The empirical formula mass is the same as the given molar mass so the empirical and molecular formulas are the same. The molecular formula is $\mathbf{C_{10}H_{20}O}$.

3.35 Students I and II are incorrect. Both students changed a given formula. Only coefficients should be changed when balancing; subscripts cannot be changed. Student I failed to identify the product correctly, writing $AlCl_2$ instead of $AlCl_3$. Student II used atomic chlorine instead of molecular chlorine as a reactant. Student III followed the correct process, changing only coefficients.

3.36 Plan: Examine the diagram and label each formula. We will use A for red atoms and B for green atoms.
Solution:
The reaction shows A_2 and B_2 diatomic molecules forming AB molecules. Equal numbers of A_2 and B_2 combine to give twice as many molecules of AB. Thus, the reaction is $A_2 + B_2 \rightarrow 2\text{ AB}$. This is the balanced equation in **b**.

3.37 Plan: Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.
Solution:
a) __$Cu(s)$ + __ $S_8(s) \rightarrow$ __$Cu_2S(s)$
Balance the S first, because there is an obvious deficiency of S on the right side of the equation. The 8 S atoms in S_8 require the coefficient 8 in front of Cu_2S:
__$Cu(s)$ + __$S_8(s) \rightarrow \underline{8}Cu_2S(s)$
Then balance the Cu. The 16 Cu atoms in Cu_2S require the coefficient 16 in front of Cu:
$\mathbf{16Cu(s) + S_8(s) \rightarrow 8Cu_2S(s)}$
b) __$P_4O_{10}(s)$ + __$H_2O(l) \rightarrow$ __$H_3PO_4(l)$
Balance the P first, because there is an obvious deficiency of P on the right side of the equation. The 4 P atoms in P_4O_{10} require a coefficient of 4 in front of H_3PO_4:
___$P_4O_{10}(s)$ + __$H_2O(l) \rightarrow \underline{4}H_3PO_4(l)$
Balance the H next, because H is present in only one reactant and only one product. The 12 H atoms in $4H_3PO_4$ on the right require a coefficient of 6 in front of H_2O:
___ $P_4O_{10}(s) + \underline{6}H_2O(l) \rightarrow \underline{4}H_3PO_4(l)$

Balance the O last, because it appears in both reactants and is harder to balance. There are 16 O atoms on each side:

$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(l)$

c) $\_\_B_2O_3(s) + \_\_\ NaOH(aq) \rightarrow \_\_Na_3BO_3(aq) + \_\_H_2O(l)$

Balance oxygen last because it is present in more than one place on each side of the reaction. The 2 B atoms in B_2O_3 on the left require a coefficient of 2 in front of Na_3BO_3 on the right:

$\_\_B_2O_3(s) + \_\_NaOH(aq) \rightarrow \underline{2}Na_3BO_3(aq) + \_\_H_2O(l)$

The 6 Na atoms in $2Na_3BO_3$ on the right require a coefficient of 6 in front of NaOH on the left:

$\_\_B_2O_3(s) + \underline{6}NaOH(aq) \rightarrow \underline{2}Na_3BO_3(aq) + \_\_H_2O(l)$

The 6 H atoms in 6NaOH on the left require a coefficent of 3 in front of H_2O on the right:

$\_\_B_2O_3(s) + \underline{6}NaOH(aq) \rightarrow \underline{2}Na_3BO_3(aq) + \underline{3}H_2O(l)$

The oxygen is now balanced with 9 O atoms on each side:

$B_2O_3(s) + 6NaOH(aq) \rightarrow 2Na_3BO_3(aq) + 3H_2O(l)$

d) $\_\_CH_3NH_2(g) + \_\_O_2(g) \rightarrow \_\_CO_2(g) + \_\_H_2O(g) + \_\_N_2(g)$

There are 2 N atoms on the right in N_2 so a coefficient of 2 is required in front of CH_3NH_2 on the left:

$\underline{2}CH_3NH_2(g) + \_\_O_2(g) \rightarrow \_\_CO_2(g) + \_\_H_2O(g) + \_\_N_2(g)$

There are now 10 H atoms in $2CH_3NH_2$ on the left so a coefficient of 5 is required in front of H_2O on the right:

$\underline{2}CH_3NH_2(g) + \_\_O_2(g) \rightarrow \_\_CO_2(g) + \underline{5}H_2O(g) + \_\_N_2(g)$

The 2 C atoms on the left require a coefficient of 2 in front of CO_2 on the right:

$\underline{2}CH_3NH_2(g) + \_\_O_2(g) \rightarrow \underline{2}CO_2(g) + \underline{5}H_2O(g) + \_\_N_2(g)$

The 9 O atoms on the right (4 O atoms in $2CO_2$ plus 5 in $5H_2O$) require a coefficient of 9/2 in front of O_2 on the left:

$\underline{2}CH_3NH_2(g) + \underline{9/2}O_2(g) \rightarrow \underline{2}CO_2(g) + \underline{5}H_2O(g) + \_\_N_2(g)$

Multiply all coefficients by 2 to obtain whole numbers:

$4CH_3NH_2(g) + 9O_2(g) \rightarrow 4CO_2(g) + 10H_2O(g) + 2N_2(g)$

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

Solution:

a) $\_\_Cu(NO_3)_2(aq) + \_\_KOH(aq) \rightarrow \_\_Cu(OH)_2(s) + \_\_\ KNO_3(aq)$

The 2 N atoms in $Cu(NO_3)_2$ on the left require a coefficient of 2 in front of KNO_3 on the right:

$\_\_Cu(NO_3)_2(aq) + \_\_\ KOH(aq) \rightarrow \_\_Cu(OH)_2(s) + \underline{2}KNO_3(aq)$

The 2 K atoms in $2KNO_3$ on the right require a coefficient of 2 in front of KOH on the left:

$\_\_Cu(NO_3)_2(aq) + \underline{2}KOH(aq) \rightarrow \_\_Cu(OH)_2(s) + \underline{2}KNO_3(aq)$

There are 8 O atoms and 2 H atoms on each side:

$Cu(NO_3)_2(aq) + 2KOH(aq) \rightarrow Cu(OH)_2(s) + 2KNO_3(aq)$

b) $\_\_BCl_3(g) + \_\_H_2O(l) \rightarrow \_\_H_3BO_3(s) + \_\_HCl(g)$

The 3 Cl atoms in BCl_3 on the left require a coefficient of 3 in front of HCl on the right:

$\_\_BCl_3(g) + \_\_H_2O(l) \rightarrow \_\_H_3BO_3(s) + \underline{3}HCl(g)$

The 6 H atoms on the right (3 in H_3BO_3 and 3 in HCl) require a coefficient of 3 in front of H_2O on the left:

$\_\_BCl_3(g) + \underline{3}H_2O(l) \rightarrow \_\_H_3BO_3(s) + \underline{3}HCl(g)$

There are 3 O atoms and 1 B atom on each side:

$BCl_3(g) + 3H_2O(l) \rightarrow H_3BO_3(s) + 3HCl(g)$

c) $\_\_CaSiO_3(s) + \_\_HF(g) \rightarrow \_\_SiF_4(g) + \_\_CaF_2(s) + \_\_H_2O(l)$

The 6 F atoms on the right (4 in SiF_4 and 2 in CaF_2) require a coefficient of 6 in front of HF on the left:

$\_\_CaSiO_3(s) + \underline{6}HF(g) \rightarrow \_\_SiF_4(g) + \_\_CaF_2(s) + \_\_H_2O(l)$

The 6 H atoms in 6HF on the left require a coefficient of 3 in front of H_2O on the right:

$\_\_CaSiO_3(s) + \underline{6}HF(g) \rightarrow \_\_SiF_4(g) + \_\_CaF_2(s) + \underline{3}H_2O(l)$

There are 1 Ca atom, 1 Si atom, and 3 O atoms on each side:

$CaSiO_3(s) + 6HF(g) \rightarrow SiF_4(g) + CaF_2(s) + 3H_2O(l)$

d) $\_\_\ (CN)_2(g) + \_\_H_2O(l) \rightarrow \_\_H_2C_2O_4(aq) + \_\_NH_3(g)$

The 2 N atoms in $(CN)_2$ on the left requires a coefficient of 2 in front of NH_3 on the left:

$\_\_\ (CN)_2(g) + \_\_H_2O(l) \rightarrow \_\_H_2C_2O_4(aq) + \underline{2}NH_3(g)$

The 4 O atoms in $H_2C_2O_4$ on the right requires a coefficient of 4 in front of H_2O on the right:
__ $(CN)_2(g)$ + $\underline{4}H_2O(l) \rightarrow$ __$H_2C_2O_4(aq)$ + $\underline{2}NH_3(g)$
There are 2 C atoms and 8 H atoms on each side:
$\mathbf{(CN)_2(\mathit{g}) + 4H_2O(\mathit{l}) \rightarrow H_2C_2O_4(\mathit{aq}) + 2NH_3(\mathit{g})}$

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

Solution:

a) Gallium (a solid) and oxygen (a gas) are reactants and solid gallium(III) oxide is the only product:
__$Ga(s)$ + __$O_2(g) \rightarrow$ __$Ga_2O_3(s)$
A coefficient of 2 in front of Ga on the left is needed to balance the 2 Ga atoms in Ga_2O_3:
$\underline{2}Ga(s)$ + __$O_2(g) \rightarrow$ __$Ga_2O_3(s)$
The 3 O atoms in Ga_2O_3 on the right require a coefficient of 3/2 in front of O_2 on the left:
$\underline{2}Ga(s) + \underline{3/2}O_2(g) \rightarrow$ __$Ga_2O_3(s)$
Multiply all coefficients by 2 to obtain whole numbers:
$\mathbf{4Ga(\mathit{s}) + 3O_2(\mathit{g}) \rightarrow 2Ga_2O_3(\mathit{s})}$

b) Liquid hexane and oxygen gas are the reactants while carbon dioxide gas and gaseous water are the products:
__$C_6H_{14}(l)$ +__$O_2(g) \rightarrow$ __$CO_2(g)$ + __$H_2O(g)$
The 6 C atoms in C_6H_{14} on the left require a coefficient of 6 in front of CO_2 on the right:
__$C_6H_{14}(l)$ +__$O_2(g) \rightarrow \underline{6}CO_2(g)$ + __$H_2O(g)$
The 14 H atoms in C_6H_{14} on the left require a coefficient of 7 in front of H_2O on the right:
__$C_6H_{14}(l)$ +__$O_2(g) \rightarrow \underline{6}CO_2(g) + \underline{7}H_2O(g)$
The 19 O atoms on the right (12 in $6CO_2$ and 7 in $7H_2O$) require a coefficient of 19/2 in front of O_2 on the left:
Multiply all coefficients by 2 to obtain whole numbers:
$\mathbf{2C_6H_{14}(\mathit{l}) + 19O_2(\mathit{g}) \rightarrow 12CO_2(\mathit{g}) + 14H_2O(\mathit{g})}$

c) Aqueous solutions of calcium chloride and sodium phosphate are the reactants; solid calcium phosphate and an aqueous solution of sodium chloride are the products:
__$CaCl_2(aq)$ + __$Na_3PO_4(aq) \rightarrow$ __$Ca_3(PO_4)_2(s)$ + __$NaCl(aq)$
The 3 Ca atoms in $Ca_3(PO_4)_2$ on the right require a coefficient of 3 in front of $CaCl_2$ on the left:
$\underline{3}CaCl_2(aq)$ + __$Na_3PO_4(aq) \rightarrow$ __$Ca_3(PO_4)_2(s)$ + __$NaCl(aq)$
The 6 Cl atoms in $3CaCl_2$ on the left require a coefficient of 6 in front of NaCl on the right:
$\underline{3}CaCl_2(aq)$ + __$Na_3PO_4(aq) \rightarrow$ __$Ca_3(PO_4)_2(s) + \underline{6}NaCl(aq)$
The 6 Na atoms in 6NaCl on the right require a coefficient of 2 in front of Na_3PO_4 on the left:
$\underline{3}CaCl_2(aq) + \underline{2}Na_3PO_4(aq) \rightarrow$ __$Ca_3(PO_4)_2(s) + \underline{6}NaCl(aq)$
There are now 2 P atoms on each side:
$\mathbf{3CaCl_2(\mathit{aq}) + 2Na_3PO_4(\mathit{aq}) \rightarrow Ca_3(PO_4)_2(\mathit{s}) + 6NaCl(\mathit{aq})}$

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

Solution:

a) Aqueous solutions of lead(II) nitrate and potassium iodide are the reactants; solid lead(II) iodide and an aqueous solution of potassium nitrate are the products:
__$Pb(NO_3)_2(aq)$ + __$KI(aq) \rightarrow$ __$PbI_2(s)$ + __$KNO_3(aq)$
There are 2 N atoms in $Pb(NO_3)_2$ on the left so a coefficient of 2 is required in front of KNO_3 on the right:
__$Pb(NO_3)_2(aq)$ + __$KI(aq) \rightarrow$ __$PbI_2(s) + \underline{2}KNO_3(aq)$
The 2 K atoms in $2KNO_3$ and the 2 I atoms in PbI_2 on the right require a coefficient of 2 in front of KI on the left:
$Pb(NO_3)_2(aq) + \underline{2}KI(aq) \rightarrow PbI_2(s) + \underline{2}KNO_3(aq)$
There are now 6 O atoms on each side:
$\mathbf{Pb(NO_3)_2(\mathit{aq}) + 2KI(\mathit{aq}) \rightarrow PbI_2(\mathit{s}) + 2KNO_3(\mathit{aq})}$

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

__Si_2Cl_6(*l*) + __H_2O(*l*) → __SiO_2(*s*) + __HCl(*g*) + __H_2(*g*)

The 2 Si atoms in Si_2Cl_6 on the left require a coefficient of 2 in front of SiO_2 on the right:

__Si_2Cl_6(*l*) + __H_2O(*l*) → 2SiO_2(*s*) + __HCl(*g*) + __H_2(*g*)

The 6 Cl atoms in Si_2Cl_6 on the left require a coefficient of 6 in front of HCl on the right:

__Si_2Cl_6(*l*) + __H_2O(*l*) → 2SiO_2(*s*) + 6HCl(*g*) + __H_2(*g*)

The 4 O atoms in 2SiO_2 on the right require a coefficient of 4 in front of H_2O on the left.

__Si_2Cl_6(*l*) + 4H_2O(*l*) → 2SiO_2(*s*) + 6HCl(*g*) + __H_2(*g*)

There are 8 H atoms in 4H_2O on the left; there are 8 H atoms on the right (6 in 6HCl and 2 in H_2):

Si_2Cl_6(*l*) + 4H_2O(*l*) → 2SiO_2(*s*) + 6HCl(*g*) + H_2(*g*)

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

__NO_2(*g*) + __H_2O(*l*) → __HNO_3(*aq*) + __NO(*g*)

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

The 2 H atoms in H_2O on the left require a coefficient of 2 in front of HNO_3 on the right:

__NO_2(*g*) + __H_2O(*l*) → 2HNO_3(*aq*) + __NO(*g*)

The 3 N atoms on the right (2 in 2HNO_3 and 1 in NO) require a coefficient of 3 in front of NO_2 on the left;

3NO_2(*g*) + __H_2O(*l*) → 2HNO_3(*aq*) + __NO(*g*)

There are now 7 O atoms on each side:

3NO_2(*g*) + H_2O(*l*) → 2HNO_3(*aq*) + NO(*g*)

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

Solution:

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

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

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

3.42 Plan: Convert the kilograms of oxygen to grams of oxygen and then moles of oxygen by dividing by its molar mass. Use the moles of oxygen and the mole ratio from the balanced chemical equation to determine the moles of KNO_3 required. Multiply the moles of KNO_3 by its molar mass to obtain the mass in grams.

Solution:

$$\text{a) Mass (g) of } O_2 = (56.6\ \text{kg}\ O_2)\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right) = 5.66\text{x}10^4\ \text{g}\ O_2$$

$$\text{Moles of } O_2 = (5.66\text{x}10^4\ \text{g}\ O_2)\left(\frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2}\right) = 1.76875\text{x}10^3\ \text{mol}\ O_2$$

$$\text{Moles of } KNO_3 = (1.76875\ \text{mol}\ O_2)\left(\frac{4\ \text{mol}\ KNO_3}{5\ \text{mol}\ O_2}\right) = 1415 = \mathbf{1.42x10^3\ mol\ KNO_3}$$

$$\text{b) Mass (g) of } KNO_3 = (1415\ \text{mol}\ KNO_3)\left(\frac{101.11\ \text{g}\ KNO_3}{1\ \text{mol}\ KNO_3}\right) = 143070.65 = \mathbf{1.43x10^5\ g\ KNO_3}$$

Combining all steps gives:

$$\text{Mass (g) of } KNO_3 = (56.6\ \text{kg}\ O_2)\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2}\right)\left(\frac{4\ \text{mol}\ KNO_3}{5\ \text{mol}\ O_2}\right)\left(\frac{101.11\ \text{g}\ KNO_3}{1\ \text{mol}\ KNO_3}\right)$$

$$= 143070.65 = \mathbf{1.43x10^5\ g\ KNO_3}$$

3.43 Plan: Convert mass of Cr_2S_3 to moles by dividing by its molar mass. Use the mole ratio between Cr_2S_3 and Cr_2O_3 from the balanced chemical equation to determine the moles of Cr_2O_3 required. Multiply the moles of Cr_2O_3 by its molar mass to obtain the mass in grams.
Solution:

a) $$\text{Moles of } Cr_2S_3 = (421 \text{ g } Cr_2S_3)\left(\frac{1 \text{ mol } Cr_2S_3}{200.21 \text{ g } Cr_2S_3}\right) = 2.102792 \text{ mol } Cr_2S_3$$

$$\text{Moles of } Cr_2O_3 = (2.102792 \text{ mol } Cr_2S_3)\left(\frac{1 \text{ mol } Cr_2O_3}{1 \text{ mol } Cr_2S_3}\right) = 2.102792 = \mathbf{2.10 \text{ mol } Cr_2O_3}$$

b) $$\text{Mass (g) of } Cr_2O_3 = (2.102792 \text{ mol } Cr_2O_3)\left(\frac{152.00 \text{ g } Cr_2O_3}{1 \text{ mol } Cr_2O_3}\right) = 319.624 = \mathbf{3.20x10^2 \text{ g } Cr_2O_3}$$

Combining all steps gives:

$$\text{Mass (g) of } Cr_2O_3 = (421 \text{ g } Cr_2S_3)\left(\frac{1 \text{ mol } Cr_2S_3}{200.21 \text{ g } Cr_2S_3}\right)\left(\frac{1 \text{ mol } Cr_2O_3}{1 \text{ mol } Cr_2S_3}\right)\left(\frac{152.00 \text{ g } Cr_2O_3}{1 \text{ mol } Cr_2O_3}\right)$$
$$= 319.624 = \mathbf{3.20x10^2 \text{ g } Cr_2O_3}$$

3.44 Plan: First, balance the equation. Convert the grams of diborane to moles of diborane by dividing by its molar mass. Use mole ratios from the balanced chemical equation to determine the moles of the products. Multiply the mole amount of each product by its molar mass to obtain mass in grams.
Solution:
The balanced equation is: $B_2H_6(g) + 6H_2O(l) \rightarrow 2H_3BO_3(s) + 6H_2(g)$.

$$\text{Moles of } B_2H_6 = (43.82 \text{ g } B_2H_6)\left(\frac{1 \text{ mol } B_2H_6}{27.67 \text{ g } B_2H_6}\right) = 1.583665 \text{ mol } B_2H_6$$

$$\text{Moles of } H_3BO_3 = (1.583665 \text{ mol } B_2H_6)\left(\frac{2 \text{ mol } H_3BO_3}{1 \text{ mol } B_2H_6}\right) = 3.16733 \text{ mol } H_3BO_3$$

$$\text{Mass (g) of } H_3BO_3 = (3.16733 \text{ mol } H_3BO_3)\left(\frac{61.83 \text{ g } H_3BO_3}{1 \text{ mol } H_3BO_3}\right) = 195.83597 = \mathbf{195.8 \text{ g } H_3BO_3}$$

Combining all steps gives:

$$\text{Mass (g) of } H_3BO_3 = (43.82 \text{ g } B_2H_6)\left(\frac{1 \text{ mol } B_2H_6}{27.67 \text{ g } B_2H_6}\right)\left(\frac{2 \text{ mol } H_3BO_3}{1 \text{ mol } B_2H_6}\right)\left(\frac{61.83 \text{ g } H_3BO_3}{1 \text{ mol } H_3BO_3}\right)$$
$$= 195.83597 = \mathbf{195.8 \text{ g } H_3BO_3}$$

$$\text{Moles of } H_2 = (1.583665 \text{ mol } B_2H_6)\left(\frac{6 \text{ mol } H_2}{1 \text{ mol } B_2H_6}\right) = 9.50199 \text{ mol } H_2$$

$$\text{Mass (g) of } H_2 = (9.50199 \text{ mol } H_2)\left(\frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}\right) = 19.15901 \text{ g } H_2 = \mathbf{19.16 \text{ g } H_2}$$

Combining all steps gives:

$$\text{Mass (g) of } H_2 = (43.82 \text{ g } B_2H_6)\left(\frac{1 \text{ mol } B_2H_6}{27.67 \text{ g } B_2H_6}\right)\left(\frac{6 \text{ mol } H_2}{1 \text{ mol } B_2H_6}\right)\left(\frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}\right) = 19.15601 = \mathbf{19.16 \text{ g } H_2}$$

3.45 Plan: First, balance the equation. Convert the grams of silver sulfide to moles of silver sulfide by dividing by its molar mass. Use mole ratios from the balanced chemical equation to determine the moles of the products. Multiply the mole amount of each product by its molar mass to obtain mass in grams.
Solution:
First, balance the equation: $Ag_2S(s) + 2\ HCl(aq) \rightarrow 2\ AgCl(s) + H_2S(g)$

$$\text{Moles of } Ag_2S = (174\text{ g } Ag_2S)\left(\frac{1\text{ mol } Ag_2S}{247.9\text{ g } Ag_2S}\right) = 0.7018959\text{ mol } Ag_2S$$

$$\text{Moles of AgCl} = (0.7018959\text{ mol } Ag_2S)\left(\frac{2\text{ mol AgCl}}{1\text{ mol } Ag_2S}\right) = 1.403792\text{ mol AgCl}$$

$$\text{Mass (g) of AgCl} = (1.403792\text{ mol } Ag_2S)\left(\frac{143.4\text{ g AgCl}}{1\text{ mol AgCl}}\right) = 201.304 = \mathbf{201\text{ g AgCl}}$$

Combining all steps gives:

$$\text{Mass (g) AgCl} = (174\text{ g } Ag_2S)\left(\frac{1\text{ mol } Ag_2S}{247.9\text{ g } Ag_2S}\right)\left(\frac{2\text{ mol AgCl}}{1\text{ mol } Ag_2S}\right)\left(\frac{143.4\text{ g AgCl}}{1\text{ mol AgCl}}\right) = 201.304 = \mathbf{201\text{ g AgCl}}$$

$$\text{Moles of } H_2S = (0.7018959\text{ mol } Ag_2S)\left(\frac{1\text{ mol } H_2S}{1\text{ mol } Ag_2S}\right) = 0.7018959\text{ mol } H_2S$$

$$\text{Mass (g) of } H_2S = 0.7018959\text{ mol } H_2S\left(\frac{34.09\text{ g } H_2S}{1\text{ mol } H_2S}\right) = 23.9276 = \mathbf{23.9\text{ g } H_2S}$$

Combining all steps gives:

$$\text{Mass (g) of } H_2S = 174\text{ g } Ag_2S\left(\frac{1\text{ mol } Ag_2S}{247.9\text{ g } Ag_2S}\right)\left(\frac{1\text{ mol } H_2S}{1\text{ mol } Ag_2S}\right)\left(\frac{34.09\text{ g } H_2S}{1\text{ mol } H_2S}\right) = 23.9276 = \mathbf{23.9\text{ g } H_2S}$$

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

Solution:

Reactants: formula for phosphorus is given as P_4 and formula for chlorine gas is Cl_2 (chlorine occurs as a diatomic molecule). Product: formula for phosphorus pentachloride (the name indicates one phosphorus atom and five chlorine atoms) is PCl_5.

Equation: $P_4 + Cl_2 \rightarrow PCl_5$

Balancing the equation: $P_4 + 10Cl_2 \rightarrow 4PCl_5$

$$\text{Moles of } P_4 = (455\text{ g } P_4)\left(\frac{1\text{ mol } P_4}{123.88\text{ g } P_4}\right) = 3.67291\text{ mol } P_4$$

$$\text{Moles of } Cl_2 = (3.67291\text{ mol } P_4)\left(\frac{10\text{ mol } Cl_2}{1\text{ mol } P_4}\right) = 36.7291\text{ mol } Cl_2$$

$$\text{Mass (g) of } Cl_2 = (36.7291\text{ mol } Cl_2)\left(\frac{70.90\text{ g } Cl_2}{1\text{ mol } Cl_2}\right) = 2604.09 = \mathbf{2.60x10^3\text{ g } Cl_2}$$

Combining all steps gives:

$$\text{Mass (g) of } Cl_2 = (455\text{ g } P_4)\left(\frac{1\text{ mol } P_4}{123.88\text{ g } P_4}\right)\left(\frac{10\text{ mol } Cl_2}{1\text{ mol } P_4}\right)\left(\frac{70.90\text{ g } Cl_2}{1\text{ mol } Cl_2}\right) = 2604.09267 = \mathbf{2.60x10^3\text{ g } Cl_2}$$

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

Solution:

Reactants: formula for sulfur is given as S_8 and formula for fluorine gas is F_2 (fluorine occurs as a diatomic molecule). Product: formula for sulfur hexafluoride (the name indicates one sulfur atom and six fluoride atoms) is SCl_6.

Equation: $S_8 + F_2 \rightarrow SF_6$

Balancing the equation: $S_8(s) + 24F_2(g) \rightarrow 8SF_6(s)$

Moles of S_8 = $\left(17.8 \text{ g } S_8\right)\left(\frac{1 \text{ mol } S_8}{256.56 \text{ g } S_8}\right)$ = 0.0693795 mol S_8

Moles of F_2 = $\left(0.0693795 \text{ mol } S_8\right)\left(\frac{24 \text{ mol } F_2}{1 \text{ mol } S_8}\right)$ = 1.665108 mol F_2

Mass (g) of F_2 = $\left(1.665108 \text{ mol } F_2\right)\left(\frac{38.00 \text{ g } F_2}{1 \text{ mol } F_2}\right)$ = 63.274 = **63.3 g F_2**

Combining all steps gives:

Mass (g) of F_2 = $\left(17.8 \text{ g } S_8\right)\left(\frac{1 \text{ mol } S_8}{256.56 \text{ g } S_8}\right)\left(\frac{24 \text{ mol } F_2}{1 \text{ mol } S_8}\right)\left(\frac{38.00 \text{ g } F_2}{1 \text{ mol } F_2}\right)$ = 63.27409 = **63.3 g F_2**

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

Solution:

$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

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

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

b) Moles of O_2 = $\left(2.80 \text{ g } O_2\right)\left(\frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2}\right)$ = 0.0875 mol O_2

Moles of CaO from O_2 = $\left(0.0875 \text{ mol } O_2\right)\left(\frac{2 \text{ mol CaO}}{1 \text{ mol } O_2}\right)$ = 0.17500 = **0.175 mol CaO**

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

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

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

Combining all steps gives:

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

3.49 Plan: Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of H_2 formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of H_2 is the limiting reactant. Convert the moles of H_2 obtained from the limiting reactant to grams using the molar mass.

Solution:

$SrH_2(s) + 2H_2O(l) \rightarrow Sr(OH)_2(s) + 2H_2(g)$

a) Moles of SrH_2 = $\left(5.70 \text{ g } SrH_2\right)\left(\frac{1 \text{ mol } SrH_2}{89.64 \text{ g } SrH_2}\right)$ = 0.0635877 mol SrH_2

Moles of H_2 from SrH_2 = $\left(0.0635877 \text{ mol } SrH_2\right)\left(\frac{2 \text{ mol } H_2}{1 \text{ mol } SrH_2}\right)$ = 0.127175 = **0.127 mol H_2**

b) Mass (g) of H_2O = $\left(4.75 \text{ g } H_2O\right)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)$ = 0.263596 mol H_2O

Moles of H_2 from H_2O = $\left(0.263596 \text{ mol } H_2O\right)\left(\frac{2 \text{ mol } H_2}{2 \text{ mol } H_2O}\right)$ = 0.263596 = **0.264 mol H_2**

c) **SrH_2** is the limiting reagent since it will yield fewer moles of hydrogen gas.
d) The mass of H_2 formed is determined by the limiting reactant, SrH_2.

Mass (g) of H_2 = $\left(0.127175 \text{ mol } H_2\right)\left(\frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}\right)$ = 0.256385 = **0.256 g H_2**

Combining all steps gives:

Mass (g) of H_2 = $\left(5.70 \text{ g } SrH_2\right)\left(\frac{1 \text{ mol } SrH_2}{89.64 \text{ g } SrH_2}\right)\left(\frac{2 \text{ mol } H_2}{1 \text{ mol } SrH_2}\right)\left(\frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}\right)$ = 0.256385 = **0.256 g H_2**

3.50 Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of HIO_3 formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of HIO_3 formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation for this reaction is:

$$2ICl_3 + 3H_2O \rightarrow ICl + HIO_3 + 5HCl$$

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

Finding the moles of HIO_3 from the moles of ICl_3 (if H_2O is limiting):

Moles of ICl_3 = $\left(635 \text{ g } ICl_3\right)\left(\frac{1 \text{ mol } ICl_3}{233.2 \text{ g } ICl_3}\right)$ = 2.722985 mol ICl_3

Moles of HIO_3 from ICl_3 = $\left(2.722985 \text{ mol } ICl_3\right)\left(\frac{1 \text{ mol } HIO_3}{2 \text{ mol } ICl_3}\right)$ = 1.361492 = 1.36 mol HIO_3

Finding the moles of HIO_3 from the moles of H_2O (if ICl_3 is limiting):

Moles of H_2O = $\left(118.5 \text{ g } H_2O\right)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)$ = 6.57603 mol H_2O

Moles HIO_3 from H_2O = $\left(6.57603 \text{ mol } H_2O\right)\left(\frac{1 \text{ mol } HIO_3}{3 \text{ mol } H_2O}\right)$ = 2.19201 = 2.19 mol HIO_3

ICl_3 is the limiting reagent and will produce **1.36 mol HIO_3**.

Mass (g) of HIO_3 = $\left(1.361492 \text{ mol } HIO_3\right)\left(\frac{175.9 \text{ g } HIO_3}{1 \text{ mol } HIO_3}\right)$ = 239.486 = **239 g HIO_3**

Combining all steps gives:

Mass (g) of HIO_3 = $\left(635 \text{ g } ICl_3\right)\left(\frac{1 \text{ mol } ICl_3}{233.2 \text{ g } ICl_3}\right)\left(\frac{1 \text{ mol } HIO_3}{2 \text{ mol } ICl_3}\right)\left(\frac{175.9 \text{ g } HIO_3}{1 \text{ mol } HIO_3}\right)$ = 239.486 = **239 g HIO_3**

The remaining mass of the excess reagent can be calculated from the amount of H_2O combining with the limiting reagent.

Moles of H_2O required to react with 635 g ICl_3 = $\left(2.722985 \text{ mol } ICl_3\right)\left(\frac{3 \text{ mol } H_2O}{2 \text{ mol } ICl_3}\right)$ = 4.0844775 mol H_2O

Mass (g) of H_2O required to react with 635 g ICl_3 = $\left(4.0844775 \text{ mol } H_2O\right)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)$

= 73.6023 = 73.6 g H_2O reacted

Remaining H_2O = 118.5 g – 73.6 g = **44.9 g H_2O**

3.51 Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of H_2S formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of H_2S formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation for this reaction is:

$Al_2S_3 + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2S$

Finding the moles of H_2S from the moles of Al_2S_3 (if H_2O is limiting):

Moles of Al_2S_3 = $\left(158 \text{ g } Al_2S_3\right)\left(\frac{1 \text{ mol } Al_2S_3}{150.17 \text{ g } Al_2S_3}\right)$ = 1.05214 mol Al_2S_3

Moles of H_2S from Al_2S_3 = $\left(1.05214 \text{ mol } Al_2S_3\right)\left(\frac{3 \text{ mol } H_2S}{1 \text{ mol } Al_2S_3}\right)$ = 3.15642 = 3.16 mol H_2S

Finding the moles of H_2S from the moles of H_2O (if Al_2S_3 is limiting):

Moles of H_2O = $\left(131 \text{ g } H_2O\right)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)$ = 7.26970 mol H_2O

Moles of H_2S from H_2O = $\left(7.26970 \text{ mol } H_2O\right)\left(\frac{3 \text{ mol } H_2S}{6 \text{ mol } H_2O}\right)$ = 3.63485 = 3.63 mol H_2S

Al_2S_3 is the limiting reagent and **3.16 mol of H_2S** will form.

Mass (g) of H_2S = $\left(3.15642 \text{ mol } H_2S\right)\left(\frac{34.09 \text{ g } H_2S}{1 \text{ mol } H_2S}\right)$ = 107.602 = **108 g H_2S**

Combining all steps gives:

Grams H_2S = $\left(158 \text{ g } Al_2S_3\right)\left(\frac{1 \text{ mol } Al_2S_3}{150.17 \text{ g } Al_2S_3}\right)\left(\frac{3 \text{ mol } H_2S}{1 \text{ mol } Al_2S_3}\right)\left(\frac{34.09 \text{ g } H_2S}{1 \text{ mol } H_2S}\right)$ = 107.602 = **108 g H_2S**

The remaining mass of the excess reagent can be calculated from the amount of H_2O combining with the limiting reagent.

Moles of H_2O required to react with 158 g of Al_2S_3 = $\left(1.05214 \text{ mol } Al_2S_3\right)\left(\frac{6 \text{ mol } H_2O}{1 \text{ mol } Al_2S_3}\right)$ = 6.31284 mol H_2O

Mass (g) of H_2O required to react with 158 g of Al_2S_3 = $\left(6.31284 \text{ mol } H_2O\right)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)$ = 113.757 g H_2O

Remaining H_2O = 131 g H_2O – 113.757 g H_2O = 17.243 = **17 g H_2O**

3.52 Plan: Write the balanced equation; the formula for carbon is C, the formula for oxygen is O_2, and the formula for carbon dioxide is CO_2. To determine which reactant is limiting, calculate the amount of CO_2 formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of CO_2 formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced equation is: $C(s) + O_2(g) \rightarrow CO_2(g)$

Finding the moles of CO_2 from the moles of carbon (if O_2 is limiting):

Moles of CO_2 from C = $(0.100 \text{ mol C})\left(\frac{1 \text{ mol } CO_2}{1 \text{ mol C}}\right)$ = 0.100 mol CO_2

Finding the moles of CO_2 from the moles of oxygen (if C is limiting):

Moles of O_2 = $(8.00 \text{ g } O_2)\left(\frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2}\right)$ = 0.250 mol O_2

Moles of CO_2 from O_2 = $(0.250 \text{ mol} O_2)\left(\frac{1 \text{ mol } CO_2}{1 \text{ mol } O_2}\right)$ = 0.25000 = 0.250 mol CO_2

Carbon is the limiting reactant and will be used to determine the amount of CO_2 that will form.

Mass (g) of CO_2 = $(0.100 \text{ mol } CO_2)\left(\frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right)$ = 4.401 = **4.40 g CO_2**

Since carbon is limiting, the **O_2 is in excess**. The amount remaining depends on how much combines with the limiting reagent.

Moles of O_2 required to react with 0.100 mol of C = $(0.100 \text{ mol C})\left(\frac{1 \text{ mol } O_2}{1 \text{ mol C}}\right)$ = 0.100 mol O_2

Mass (g) of O_2 required to react with 0.100 mol of C = $(0.100 \text{ mol } O_2)\left(\frac{32.00 \text{ mol } O_2}{1 \text{ mol } O_2}\right)$ = 3.20 g O_2

Remaining O_2 = 8.00 g – 3.20 g = **4.80 g O_2**

3.53 Plan: Write the balanced equation; the formula for hydrogen is H_2, the formula for oxygen is O_2, and the formula for water is H_2O. To determine which reactant is limiting, calculate the amount of H_2O formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of H_2O formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced equation is: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Finding the moles of H_2O from the moles of hydrogen (if O_2 is limiting):

Moles of H_2 = $(0.0375 \text{ g } H_2)\left(\frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2}\right)$ = 0.01860 mol H_2

Moles of H_2O from H_2 = $(0.01860 \text{ mol } H_2)\left(\frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2}\right)$ = 0.01860 = 0.0186 mol H_2O

Finding the moles of H_2O from the moles of oxygen (if H_2 is limiting):

Mole of H_2O from O_2 = $(0.0185 \text{ mol } O_2)\left(\frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2}\right)$ = 0.0370 mol H_2O

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

Mass (g) of H_2O = $(0.01860 \text{ mol } H_2O)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)$ = 0.335172 = **0.335 g H_2O**

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

Moles of O_2 required to react with 0.0375 g of H_2 = $(0.01860 \text{ mol } H_2)\left(\frac{1 \text{ mol } O_2}{2 \text{ mol } H_2}\right)$ = 0.00930 mol O_2

Mass (g) of O_2 required to react with 0.0375 g of H_2 = $(0.00930 \text{ mol } O_2)\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}\right)$ = 0.2976 g O_2

Mass of O_2 supplied = $(0.0185 \text{ mol } O_2)\left(\frac{32.00 \text{ mol } O_2}{1 \text{ mol } O_2}\right)$ = 0.5920 g O_2

Remaining O_2 = 0.5920 g – 0.2976 g = 0.2944 = **0.294 g O_2**

3.54 Plan: The question asks for the mass of each substance present at the end of the reaction. "Substance" refers to both reactants and products. Solve this problem using multiple steps. Recognizing that this is a limiting reactant problem, first write a balanced chemical equation. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, $AlCl_3$ is used. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of both products formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation is:

$$Al(NO_2)_3(aq) + 3NH_4Cl(aq) \rightarrow AlCl_3(aq) + 3N_2(g) + 6H_2O(l)$$

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

Finding the moles of $AlCl_3$ from the moles of $Al(NO_2)_3$ (if NH_4Cl is limiting):

Moles of $Al(NO_2)_3$ = $(72.5 \text{ g } Al(NO_2)_3)\left(\frac{1 \text{ mol } Al(NO_2)_3}{165.01 \text{ g } Al(NO_2)_3}\right)$ = 0.439367 mol $Al(NO_2)_3$

Moles of $AlCl_3$ from $Al(NO_2)_3$ = $(0.439367 \text{ mol } Al(NO_2)_3)\left(\frac{1 \text{ mol } AlCl_3}{1 \text{ mol } Al(NO_2)_3}\right)$ = 0.439367 = 0.439 mol $AlCl_3$

Finding the moles of $AlCl_3$ from the moles of NH_4Cl (if $Al(NO_2)_3$ is limiting):

Moles of NH_4Cl = $(58.6 \text{ g } NH_4Cl)\left(\frac{1 \text{ mol } NH_4Cl}{53.49 \text{ g } NH_4Cl}\right)$ = 1.09553 mol NH_4Cl

Moles of $AlCl_3$ from NH_4Cl = $(1.09553 \text{ mol } NH_4Cl)\left(\frac{1 \text{ mol } AlCl_3}{3 \text{ mol } NH_4Cl}\right)$ = 0.365177 = 0.365 mol $AlCl_3$

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

Mass of substances after the reaction:

$Al(NO_2)_3$:

Mass (g) of $Al(NO_2)_3$ (the excess reactant) required to react with 58.6 g of NH_4Cl =

$$(1.09553 \text{ mol } NH_4Cl)\left(\frac{1 \text{ mol } Al(NO_2)_3}{3 \text{ mol } NH_4Cl}\right)\left(\frac{165.01 \text{ g } Al(NO_2)_3}{1 \text{ mol } Al(NO_2)_3}\right) = 60.2579 = 60.3 \text{ g } Al(NO_2)_3$$

$Al(NO_2)_3$ remaining: 72.5 g – 60.3 g = **12.2 g $Al(NO_2)_3$**

NH_4Cl: **None left** since it is the limiting reagent.

$AlCl_3$:

Mass (g) of $AlCl_3$ = $(0.365177 \text{ mol } AlCl_3)\left(\frac{133.33 \text{ g } AlCl_3}{1 \text{ mol } AlCl_3}\right)$ = 48.689 = **48.7 g $AlCl_3$**

N_2:

Mass (g) of N_2 = $(1.09553 \text{ mol } NH_4Cl)\left(\frac{3 \text{ mol } N_2}{3 \text{ mol } NH_4Cl}\right)\left(\frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2}\right)$ = 30.697 = **30.7 g N_2**

H_2O:

Mass (g) of H_2O = $(1.09553 \text{ mol } NH_4Cl)\left(\frac{6 \text{ mol } H_2O}{3 \text{ mol } NH_4Cl}\right)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)$ = 39.483 = **39.5 g H_2O**

3.55 Plan: The question asks for the mass of each substance present at the end of the reaction. “Substance” refers to both reactants and products. Solve this problem using multiple steps. Recognizing that this is a limiting reactant problem, first write a balanced chemical equation. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, CaF_2 is used. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of both products formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation is:

$Ca(NO_3)_2(s) + 2NH_4F(s) \rightarrow CaF_2(s) + 2N_2O(g) + 4H_2O(g)$

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

Finding the moles of CaF_2 from the moles of $Ca(NO_3)_2$ (if NH_4F is limiting):

$$\text{Moles of } Ca(NO_3)_2 = \left(16.8 \text{ g } Ca(NO_3)_2\right)\left(\frac{1 \text{ mol } Ca(NO_3)_2}{164.10 \text{ g } Ca(NO_3)_2}\right) = 0.1023766 \text{ mol } Ca(NO_3)_2$$

$$\text{Moles of } CaF_2 \text{ from } Ca(NO_3)_2 = \left(0.1023766 \text{ mol } Ca(NO_3)_2\right)\left(\frac{1 \text{ mol } CaF_2}{1 \text{ mol } Ca(NO_3)_2}\right)$$

$$= 0.1023766 = 0.102 \text{ mol } CaF_2$$

Finding the moles of CaF_2 from the moles of NH_4F (if $Ca(NO_3)_2$ is limiting):

$$\text{Moles of } NH_4F = \left(17.50 \text{ g } NH_4F\right)\left(\frac{1 \text{ mol } NH_4F}{37.04 \text{ g } NH_4F}\right) = 0.47246 \text{ mol } NH_4F$$

$$\text{Moles of } CaF_2 \text{ from } NH_4F = \left(0.47246 \text{ mol } NH_4F\right)\left(\frac{1 \text{ mol } CaF_2}{2 \text{ mol } NH_4F}\right) = 0.23623 = 0.236 \text{ mol } CaF_2$$

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

Mass of substances after the reaction:

$Ca(NO_3)_2$: **None** (It is the limiting reactant.)

NH_4F:

Mass (g) of NH_4F (the excess reactant) required to react with 16.8 g of $Ca(NO_3)_2$ =

$$\left(0.1023766 \text{ mol } Ca(NO_3)_2\right)\left(\frac{2 \text{ mol } NH_4F}{1 \text{ mol } Ca(NO_3)_2}\right)\left(\frac{37.04 \text{ g } NH_4F}{1 \text{ mol } NH_4F}\right) = 7.58406 \text{ g } Ca(NO_3)_2$$

NH_4F remaining: 17.50 g – 7.58 g = 9.9159 = **9.92 g NH_4F**

CaF_2:

$$\text{Mass (g) of } CaF_2 = \left(0.1023766 \text{ mol } Ca(NO_3)_2\right)\left(\frac{1 \text{ mol } CaF_2}{1 \text{ mol } Ca(NO_3)_2}\right)\left(\frac{78.08 \text{ g } CaF_2}{1 \text{ mol } CaF_2}\right) = 7.99356 = \mathbf{7.99 \text{ g } CaF_2}$$

N_2O:

$$\text{Mass (g) of } N_2O = \left(0.1023766 \text{ mol } Ca(NO_3)_2\right)\left(\frac{2 \text{ mol } N_2O}{1 \text{ mol } Ca(NO_3)_2}\right)\left(\frac{44.02 \text{ g } N_2O}{1 \text{ mol } N_2O}\right) = 9.0132 = \mathbf{9.01 \text{ g } N_2O}$$

H_2O:

$$\text{Mass (g) of } H_2O = \left(0.1023766 \text{ mol } Ca(NO_3)_2\right)\left(\frac{4 \text{ mol } H_2O}{1 \text{ mol } Ca(NO_3)_2}\right)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right) = 7.3793 = \mathbf{7.38 \text{ g } H_2O}$$

3.56 Plan: Express the yield of each step as a fraction of 1.00; multiply the fraction of the first step by that of the second step and then multiply by 100 to get the overall percent yield.

Solution:

73% = 0.73; 68% = 0.68

(0.73 x 0.68) x 100 = 49.64 = **50.%**

3.57 Plan: Express the yield of each step as a fraction of 1.00; multiply the fraction of the first step by that of the second step and then multiply by 100 to get the overall percent yield.
Solution:
48% = 0.48; 73% = 0.73
(0.48 x 0.73) x 100 = 35.04 = **35%**

3.58 Plan: Write and balance the chemical equation using the formulas of the substances. Determine the theoretical yield of the reaction from the mass of tungsten(VI) oxide. To do that, convert the mass of tungsten(VI) oxide to moles by dividing by its molar mass and then use the mole ratio between tungsten(VI) oxide and water to determine the moles and then mass of water that should be produced. Use the density of water to determine the actual yield of water in grams. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield.
Solution:
The balanced chemical equation is:
$WO_3(s) + 3H_2(g) \rightarrow W(s) + 3H_2O(l)$
Determining the theoretical yield of H_2O:

$$\text{Moles of } WO_3 = \left(45.5\text{ g } WO_3\right)\left(\frac{1\text{ mol } WO_3}{231.9\text{ g } WO_3}\right) = 0.1962053\text{ mol } WO_3$$

$$\text{Mass (g) of } H_2O \text{ (theoretical yield)} = \left(0.1962053\text{ mol } WO_3\right)\left(\frac{3\text{ mol } H_2O}{1\text{ mol } WO_3}\right)\left(\frac{18.02\text{ g } H_2O}{1\text{ mol } H_2O}\right) = 10.60686\text{ g } H_2O$$

Determining the actual yield of H_2O:

$$\text{Mass (g) of } H_2O \text{ (actual yield)} = \left(9.60\text{ mL } H_2O\right)\left(\frac{1.00\text{ g } H_2O}{1\text{ mL } H_2O}\right) = 9.60\text{ g } H_2O$$

$$\%\text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \text{ x } 100\% = \left(\frac{9.60\text{ g } H_2O}{10.60686\text{ g } H_2O}\right) \text{ x } 100\% = 90.5075 = \mathbf{90.5\%}$$

3.59 Plan: Write and balance the chemical equation using the formulas of the substances. Determine the theoretical yield of the reaction from the mass of phosphorus trichloride. To do that, convert the mass of phosphorus trichloride to moles by dividing by its molar mass and then use the mole ratio between phosphorus trichloride and HCl to determine the moles and then mass of HCl that should be produced. The actual yield of the HCl is given. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield.
Solution:
The balanced chemical equation is:
$PCl_3(l) + 3H_2O(l) \rightarrow H_3PO_3(aq) + 3HCl(g)$
Determining the theoretical yield of HCl:

$$\text{Moles of } PCl_3 = \left(200.\text{ g } PCl_3\right)\left(\frac{1\text{ mol } PCl_3}{137.32\text{ g } PCl_3}\right) = 1.456452\text{ mol } PCl_3$$

$$\text{Mass (g) of HCl (theoretical yield)} = \left(1.456452\text{ mol } PCl_3\right)\left(\frac{3\text{ mol HCl}}{1\text{ mol } PCl_3}\right)\left(\frac{36.46\text{ g HCl}}{1\text{ mol HCl}}\right) = 159.3067\text{ g HCl}$$

Actual yield (g) of HCl is given as 128 g HCl.
Calculate the percent yield:

$$\%\text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \text{ x } 100\% = \left(\frac{128\text{ g HCl}}{159.3067\text{ g HCl}}\right) \text{ x } 100\% = 80.3481586 = \mathbf{80.3\%}$$

3.60 Plan: Write the balanced chemical equation. Since quantities of two reactants are given, we must determine which is the limiting reactant. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, CH_3Cl is used. Only 75.0% of the calculated amounts of products actually form, so the actual yield is 75% of the theoretical yield.

Solution:

The balanced equation is: $CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$

Determining the limiting reactant:

Finding the moles of CH_3Cl from the moles of CH_4 (if Cl_2 is limiting):

Moles of $CH_4 = \left(20.5 \text{ g } CH_4\right)\left(\frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4}\right) = 1.278055 \text{ mol } CH_4$

Moles of CH_3Cl from $CH_4 = \left(1.278055 \text{ mol } CH_4\right)\left(\frac{1 \text{ mol } CH_3Cl}{1 \text{ mol } CH_4}\right) = 1.278055 \text{ mol } CH_3Cl$

Finding the moles of CH_3Cl from the moles of Cl_2 (if CH_4 is limiting):

Moles of $Cl_2 = \left(45.0 \text{ g } Cl_2\right)\left(\frac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2}\right) = 0.634697 \text{ mol } Cl_2$

Moles of CH_3Cl from $Cl_2 = \left(0.634697 \text{ mol } Cl_2\right)\left(\frac{1 \text{ mol } CH_3Cl}{1 \text{ mol } Cl_2}\right) = 0.634697 \text{ mol } CH_3Cl$

Chlorine is the limiting reactant and is used to determine the theoretical yield of CH_3Cl:

Mass (g) of CH_3Cl (theoretical yield) = $\left(0.634697 \text{ mol } CH_3Cl\right)\left(\frac{50.48 \text{ g } CH_3Cl}{1 \text{ mol } CH_3Cl}\right) = 32.0395 \text{ g } CH_3Cl$

% yield = $\left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \text{x } 100\%$

Actual yield (g) of $CH_3Cl = \frac{\% \text{ yield}}{100\%}\left(\text{theoretical yield}\right) = \frac{75\%}{100\%}\left(32.0395 \text{ g } CH_3Cl\right) = 24.02962 = \mathbf{24.0 \text{ g } CH_3Cl}$

3.61 Plan: Write the balanced chemical equation. Since quantities of two reactants are given, we must determine which is the limiting reactant. To determine which reactant is limiting, calculate the amount of product formed from each reactant, assuming an excess of the other reactant. Only 93.0% of the calculated amount of product actually forms, so the actual yield is 93.0% of the theoretical yield.

Solution:

The balanced equation is: $3Ca(s) + N_2(g) \rightarrow Ca_3N_2(s)$

Determining the limiting reactant:

Finding the moles of Ca_3N_2 from the moles of Ca (if N_2 is limiting):

Moles of Ca = $\left(56.6 \text{ g Ca}\right)\left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}}\right) = 1.412176 \text{ mol Ca}$

Moles of Ca_3N_2 from Ca = $\left(1.412176 \text{ mol Ca}\right)\left(\frac{1 \text{ mol } Ca_3N_2}{3 \text{ mol Ca}}\right) = 0.470725 \text{ mol } Ca_3N_2$

Finding the moles of Ca_3N_2 from the moles of N_2 (if Ca is limiting):

Moles of $N_2 = \left(30.5 \text{ g } N_2\right)\left(\frac{1 \text{ mol } N_2}{28.02 \text{ g } N_2}\right) = 1.08851 \text{ mol } N_2$

Moles of Ca_3N_2 from $N_2 = \left(1.08851 \text{ mol } N_2\right)\left(\frac{1 \text{ mol } Ca_3N_2}{1 \text{ mol } N_2}\right) = 1.08851 \text{ mol } Ca_3N_2$

Ca is the limiting reactant and is used to determine the theoretical yield of Ca_3N_2.

Mass (g) of Ca_3N_2 (theoretical yield) = $\left(0.470725 \text{ mol } Ca_3N_2\right)\left(\frac{148.26 \text{ g } Ca_3N_2}{1 \text{ mol } Ca_3N_2}\right) = 69.7897 \text{ g } Ca_3N_2$

$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \text{ x } 100\%$$

$$\text{Actual yield (g) of } Ca_3N_2 = \frac{\%\text{ yield}}{100\%}(\text{theoretical yield}) = \frac{93\%}{100\%}(69.7897 \text{ g } Ca_3N_2) = 64.9044 = \mathbf{64.9\ g\ Ca_3N_2}$$

3.62 Plan: Write the balanced equation; the formula for fluorine is F_2, the formula for carbon tetrafluoride is CF_4, and the formula for nitrogen trifluoride is NF_3. To determine which reactant is limiting, calculate the amount of CF_4 formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the mass of CF_4 formed.

Solution:

The balanced chemical equation is:

$(CN)_2(g) + 7F_2(g) \rightarrow 2CF_4(g) + 2NF_3(g)$

Determining the limiting reactant:

Finding the moles of CF_4 from the moles of $(CN)_2$ (if F_2 is limiting):

$$\text{Moles of } CF_4 \text{ from } (CN)_2 = (60.0 \text{ g } (CN)_2)\left(\frac{1 \text{ mol } (CN)_2}{52.04 \text{ g } (CN)_2}\right)\left(\frac{2 \text{ mol } CF_4}{1 \text{ mol } (CN)_2}\right) = 2.30592 \text{ mol } CF_4$$

Finding the moles of CF_4 from the moles of F_2 (if $(CN)_2$ is limiting):

$$\text{Moles of } CF_4 \text{ from } F_2 = (60.0 \text{ g } F_2)\left(\frac{1 \text{ mol } F_2}{38.00 \text{ g } F_2}\right)\left(\frac{2 \text{ mol } CF_4}{7 \text{ mol } F_2}\right) = 0.4511278 \text{ mol } CF_4$$

F_2 is the limiting reactant, and will be used to calculate the amount of CF_4 produced.

$$\text{Mass (g) of } CF_4 = (60.0 \text{ g } F_2)\left(\frac{1 \text{ mol } F_2}{38.00 \text{ g } F_2}\right)\left(\frac{2 \text{ mol } CF_4}{7 \text{ mol } F_2}\right)\left(\frac{88.01 \text{ g } CF_4}{1 \text{ mol } CF_4}\right) = 39.70376 = \mathbf{39.7\ g\ CF_4}$$

3.63 Plan: Write and balance the chemical reaction. Remember that both chlorine and oxygen exist as diatomic molecules. Use the mole ratio between oxygen and dichlorine monoxide to find the moles of dichlorine monoxide that reacted. Multiply the amount in moles by Avogadro's number to convert to number of molecules.

Solution:

a) Both oxygen and chlorine are diatomic. **Scene A** best represents the product mixture as there are O_2 and Cl_2 molecules in Scene A. Scene B shows oxygen and chlorine atoms and Scene C shows atoms and molecules. Oxygen and chlorine atoms are NOT products of this reaction.

b) The balanced reaction is $\mathbf{2Cl_2O(g) \rightarrow 2Cl_2(g) + O_2(g)}$.

c) There is a 2:1 mole ratio between Cl_2 and O_2. In Scene A, there are 6 green molecules and 3 red molecules. Since twice as many Cl_2 molecules are produced as there are O_2 molecules produced, the red molecules are the O_2 molecules.

$$\text{Moles of } Cl_2O = (3 \text{ } O_2 \text{ molecules})\left(\frac{2 \text{ O atoms}}{1 \text{ } O_2 \text{ molecule}}\right)\left(\frac{0.050 \text{ mol O atoms}}{1 \text{ O atom}}\right)\left(\frac{1 \text{ mol } O_2 \text{ molecules}}{2 \text{ mol O atoms}}\right)\left(\frac{2 \text{ mol } Cl_2O}{1 \text{ mol } O_2}\right)$$

$$= 0.30 \text{ mol } Cl_2O$$

$$\text{Molecules of } Cl_2O = (0.30 \text{ mol } Cl_2O)\left(\frac{6.022\text{x}10^{23} \text{ } Cl_2O \text{ molecules}}{1 \text{ mol } Cl_2O}\right)$$

$$= 1.8066\text{x}10^{23} = \mathbf{1.8x10^{23}\ Cl_2O\ molecules}$$

3.64 Plan: Write a balanced equation. Use the density of butane to convert the given volume of butane to mass and divide by the molar mass of butane to convert mass to moles. Use the mole ratio between butane and oxygen to find the moles and then mass of oxygen required for the reaction. The mole ratio between butane and water is used to find the moles of water produced and the mole ratio between butane and carbon dioxide is used to find the moles of carbon dioxide produced. The total moles of product are multiplied by Avogadro's number to find the number of product molecules.

Solution:
The balanced chemical equation is:
$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

a) Moles of C_4H_{10} = $\left(5.50 \text{ mL } C_4H_{10}\right)\left(\frac{0.579 \text{ g } C_4H_{10}}{1 \text{ mL } C_4H_{10}}\right)\left(\frac{1 \text{ mol } C_4H_{10}}{58.12 \text{ g } C_4H_{10}}\right)$ = 0.054792 mol C_4H_{10}

Mass (g) of O_2 = $\left(0.054792 \text{ mol } C_4H_{10}\right)\left(\frac{13 \text{ mol } O_2}{2 \text{ mol } C_4H_{10}}\right)\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}\right)$ = 11.3967 = **11.4 g O_2**

b) Moles of H_2O = $\left(0.054792 \text{ mol } C_4H_{10}\right)\left(\frac{10 \text{ mol } H_2O}{2 \text{ mol } C_4H_{10}}\right)$ = 0.27396 = **0.274 mol H_2O**

c) Moles of CO_2 = $\left(0.054792 \text{ mol } C_4H_{10}\right)\left(\frac{8 \text{ mol } CO_2}{2 \text{ mol } C_4H_{10}}\right)$ = 0.219168 mol CO_2

Total moles = 0.27396 mol H_2O + 0.219168 mol CO_2 = 0.493128 mol

Total molecules = $\left(0.493128 \text{ mol}\right)\left(\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}}\right)$ = 2.96962×10^{23} = **2.97×10^{23} molecules**

3.65 Plan: Write a balanced equation for the reaction. Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of $NaBH_4$ formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of product is the limiting reactant. Convert the moles of $NaBH_4$ obtained from the limiting reactant to grams using the molar mass. This is the theoretical yield of $NaBH_4$. Since there is a yield of 88.5%, the amount of $NaBH_4$ actually obtained will be 88.5% of the theoretical yield.

Solution:
The balanced chemical equation is:
$2NaH(s) + B_2H_6(g) \rightarrow 2NaBH_4(s)$
Determining the limiting reactant:
Finding the moles of $NaBH_4$ from the amount of NaH (if B_2H_6 is limiting):

Moles of $NaBH_4$ from NaH = $\left(7.98 \text{ g NaH}\right)\left(\frac{1 \text{ mol NaH}}{24.00 \text{ g NaH}}\right)\left(\frac{2 \text{ mol } NaBH_4}{2 \text{ mol NaH}}\right)$ = 0.3325 mol $NaBH_4$

Finding the moles of $NaBH_4$ from the amount of B_2H_6 (if NaH is limiting):

Moles of $NaBH_4$ from B_2H_6 = $\left(8.16 \text{ g } B_2H_6\right)\left(\frac{1 \text{ mol } B_2H_6}{27.67 \text{ g } B_2H_6}\right)\left(\frac{2 \text{ mol } NaBH_4}{1 \text{ mol } B_2H_6}\right)$ = 0.58981 mol $NaBH_4$

NaH is the limiting reactant, and will be used to calculate the theoretical yield of $NaBH_4$.

Mass (g) of $NaBH_4$ = $\left(0.3325 \text{ mol } NaBH_4\right)\left(\frac{37.83 \text{ g } NaBH_4}{1 \text{ mol } NaBH_4}\right)$ = 12.5785 g $NaBH_4$

% yield = $\left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right)$ x 100%

Mass (g) of $NaBH_4$ = $\left(\frac{\text{\% yield}}{100\%}\right)\left(\text{theoretical yield}\right)$ = $\left(\frac{88.5\%}{100\%}\right)\left(12.5785 \text{ g } NaHB_4\right)$ = 11.13197 = **11.1 g $NaBH_4$**

Combining all steps gives:

Mass (g) of $NaBH_4$ = $\left(7.98 \text{ g NaH}\right)\left(\frac{1 \text{ mol NaH}}{24.00 \text{ g NaH}}\right)\left(\frac{2 \text{ mol } NaBH_4}{2 \text{ mol NaH}}\right)\left(\frac{37.83 \text{ g } NaBH_4}{1 \text{ mol } NaBH_4}\right)\left(\frac{88.5\%}{100\%}\right)$

= 11.13197 = **11.1 g $NaBH_4$**

3.66 Plan: Recall that molarity = moles of solute/volume (L) of solution. Here you can use the number of particles in place of moles of solute.
Solution:
a) **Solution B** has the highest molarity as it has the largest number of particles, 12, in a volume of 50 mL.
b) **Solutions A and F** both have 8 particles in a volume of 50 mL and thus the same molarity. **Solutions C, D, and E** all have 4 particles in a volume of 50 mL and thus have the same molarity.
c) Mixing Solutions A and C results in 8 + 4 = 12 particles in a volume of 100 mL. That is a **lower molarity** than that of Solution B which has 12 particles in a volume of 50 mL or 24 particles in a volume of 100 mL.
d) Adding 50 mL to Solution D would result in 4 particles in a total volume of 100 mL; adding 75 mL to Solution F would result in 4 particles in a volume of 100 mL. The molarity of each solution would be the **same**.
e) Solution A has 8 particles in a volume of 50 mL while Solution E has the equivalent of 4 particles in a volume of 50 mL. The molarity of Solution E is half that of Solution A. Therefore **half of the volume, 12.5 mL**, of Solution E must be evaporated. When 12.5 mL of solvent is evaporated from Solution E, the result will be 2 particles in 12.5 mL or 8 particles in 50 mL as in Solution A.

3.67 Plan: The spheres represent particles of solute and the amount of *solute* per given volume of *solution* determines its concentration. Molarity = moles of solute/volume (L) of solution.
Solution:
a) **Box C** has more solute added because it contains 2 more spheres than Box A contains.
b) **Box B** has more solvent because solvent molecules have displaced two solute molecules.
c) **Box C** has a higher molarity, because it has more moles of solute per volume of solution.
d) **Box B** has a lower concentration (and molarity), because it has fewer moles of solute per volume of solution.

3.68 Plan: In all cases, use the known quantities and the definition of molarity $\left(M = \frac{\text{moles solute}}{\text{L of solution}}\right)$ to find the unknown quantity. Volume must be expressed in liters. The molar mass is used to convert moles to grams. The chemical formulas must be written to determine the molar mass. (a) You will need to convert milliliters to liters, multiply by the molarity to find moles, and convert moles to mass in grams. (b) Convert mass of solute to moles and volume from mL to liters. Divide the moles by the volume. (c) Multiply the molarity by the volume.
Solution:
a) Calculating moles of solute in solution:

$$\text{Moles of Ca(C}_2\text{H}_3\text{O}_2)_2 = (185.8 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right)\left(\frac{0.267 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2}{1 \text{ L}}\right) = 0.0496086 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2$$

Converting from moles of solute to grams:

$$\text{Mass (g) of Ca(C}_2\text{H}_3\text{O}_2)_2 = (0.0496086 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2)\left(\frac{158.17 \text{ g Ca(C}_2\text{H}_3\text{O}_2)_2}{1 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2}\right)$$

$$= 7.84659 = \mathbf{7.85\ g\ Ca(C_2H_3O_2)_2}$$

b) Converting grams of solute to moles:

$$\text{Moles of KI} = (21.1 \text{ g KI})\left(\frac{1 \text{ mol KI}}{166.0 \text{ g KI}}\right) = 0.127108 \text{ moles KI}$$

$$\text{Volume (L)} = (500.\text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right) = 0.500 \text{ L}$$

$$\text{Molarity of KI} = \frac{0.127108 \text{ mol KI}}{0.500 \text{ L}} = 0.254216 = \mathbf{0.254\ \mathit{M}\ KI}$$

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

3.69 Plan: In all cases, use the known quantities and the definition of molarity $\left(M = \frac{\text{moles solute}}{\text{L of solution}}\right)$ to find the unknown quantity. Volume must be expressed in liters. The molar mass is used to convert moles to grams. The chemical formulas must be written to determine the molar mass. (a) You will need to convert mass of solute to moles and divide by the molarity to obtain volume in liters, which is then converted to milliliters. (b) Multiply the volume by the molarity to obtain moles of solute. Use Avogadro's number to determine the number of ions present. (c) Divide mmoles by milliliters; molarity may not only be expressed as moles/L, but also as mmoles/mL.

Solution:

a) Converting mass of solute to moles:

$$\text{Moles of KOH} = (8.42\text{ g KOH})\left(\frac{1\text{ mol KOH}}{56.11\text{ g KOH}}\right) = 0.15006\text{ mol KOH}$$

$$\text{Volume (L) of KOH solution} = (0.15006\text{ mol KOH})\left(\frac{1\text{ L}}{2.26\text{ mol}}\right) = 0.066398\text{ L KOH solution}$$

$$\text{Volume (mL) of KOH solution} = (0.066398\text{ L KOH})\left(\frac{1\text{ L}}{10^{-3}\text{ mL}}\right) = 66.39823 = \mathbf{66.4\text{ mL KOH solution}}$$

$$\text{b) Moles of } CuCl_2 = (52\text{ L})\left(\frac{2.3\text{ mol }CuCl_2}{\text{L}}\right) = 119.6\text{ mol }CuCl_2$$

$$\text{Moles of } Cu^{2+}\text{ ions} = (119.6\text{ mol }CuCl_2)\left(\frac{1\text{ mol }Cu^{2+}}{1\text{ mol }CuCl_2}\right) = 119.6\text{ mol }Cu^{2+}\text{ ions}$$

Converting moles of ions to number of ions:

$$\text{Number of } Cu^{2+}\text{ ions} = (119.6\text{ mol }Cu^{2+}\text{ ions})\left(\frac{6.022\times10^{23}\,Cu^{2+}\text{ ions}}{1\text{ mol }Cu^{2+}\text{ ions}}\right) = 7.2023\text{x}10^{25} = \mathbf{7.2x10^{25}\ Cu^{2+}\ ions}$$

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

Note: Since 1 mmol is 10^{-3} mol and 1 mL is 10^{-3} L, we can use these units instead of converting to mol and L since molarity is a ratio of mol/L. Molarity may not only be expressed as moles/L, but also as mmoles/mL.

3.70 Plan: These are dilution problems. Dilution problems can be solved by converting to moles and using the new volume; however, it is much easier to use $M_1V_1 = M_2V_2$. The dilution equation does not require a volume in liters; it only requires that the volume units match. In part c), it is necessary to find the moles of sodium ions in each separate solution, add these two mole amounts, and divide by the total volume of the two solutions.

Solution:

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

$M_1V_1 = M_2V_2$

$$M_2 = \frac{M_1 \text{ x } V_1}{V_2} = \frac{(0.250\ M)(37.00\text{ mL})}{150.0\text{ mL}} = 0.061667 = \mathbf{0.0617}\ \boldsymbol{M}\ \mathbf{KCl}$$

b) $M_1 = 0.0706\ M\ (NH_4)_2SO_4$ $V_1 = 25.71$ mL $M_2 = ?$ $V_2 = 500.00$ mL

$M_1V_1 = M_2V_2$

$$M_2 = \frac{M_1 \text{ x } V_1}{V_2} = \frac{(0.0706\ M)(25.71\text{ mL})}{500.0\text{ mL}} = 0.003630 = \mathbf{0.00363}\ \boldsymbol{M}\ \mathbf{(NH_4)_2SO_4}$$

$$\text{c) Moles of } Na^{+}\text{ from NaCl solution} = (3.58\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.348\text{ mol NaCl}}{1\text{ L}}\right)\left(\frac{1\text{ mol }Na^{+}}{1\text{ mol NaCl}}\right)$$

$$= 0.00124584\text{ mol }Na^{+}$$

Moles of Na^+ from Na_2SO_4 solution = $(500.\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{6.81\text{x}10^{-2}\ \text{mol}\ Na_2SO_4}{1\ \text{L}}\right)\left(\frac{2\ \text{mol}\ Na^+}{1\ \text{mol}\ Na_2SO_4}\right)$

= 0.0681 mol Na^+

Total moles of Na^+ ions = 0.00124584 mol Na^+ ions + 0.0681 mol Na^+ ions = 0.06934584 mol Na^+ ions
Total volume = 3.58 mL + 500. mL = 503.58 mL = 0.50358 L

Molarity of Na^+ = $\frac{\text{total moles}\ Na^+\ \text{ions}}{\text{total volume}} = \frac{0.06934584\ \text{mol}\ Na^+\ \text{ions}}{0.50358\ \text{L}}$ = 0.1377057 = **0.138 *M* Na^+ ions**

3.71 Plan: These are dilution problems. Dilution problems can be solved by converting to moles and using the new volume; however, it is much easier to use $M_1V_1 = M_2V_2$. The dilution equation does not require a volume in liters; it only requires that the volume units match.
Solution:

a) M_1 = 2.050 *M* $Cu(NO_3)_2$ V_1 = ? M_2 = 0.8543 *M* $Cu(NO_3)_2$ V_2 = 750.0 mL

$M_1V_1 = M_2V_2$

$V_1 = \frac{M_2 \text{ x } V_2}{M_1} = \frac{(0.8543\ M)(750.0\ \text{mL})}{2.050\ M}$ = 312.5488 = **312.5 mL**

b) M_1 = 1.63 *M* $CaCl_2$ $M_1\ Cl^- = \left(\frac{1.63\ \text{mol}\ CaCl_2}{1\ \text{L}}\right)\left(\frac{2\ \text{mol}\ Cl^-}{1\ \text{mol}\ CaCl_2}\right)$ = 3.26 *M* Cl^- ions

M_1 = 3.26 *M* Cl^- V_1 = ? M_2 = 2.86x10^{-2} *M* Cl^- ions V_2 = 350. mL

$M_1V_1 = M_2V_2$

$V_1 = \frac{M_2 \text{ x } V_2}{M_1} = \frac{(2.86\text{x}10^{-2}\ M)(350.\ \text{mL})}{3.26\ M}$ = 3.07055= **3.07 mL**

c) M_1 = 0.155 *M* Li_2CO_3 V_1 = 18.0 mL M_2 = 0.0700 *M* Li_2CO_3 V_2 = ?

$M_1V_1 = M_2V_2$

$V_2 = \frac{M_1 \text{ x } V_1}{M_2} = \frac{(0.155\ M)(18.0\ \text{mL})}{(0.0700\ M)}$ = 39.8571 = **39.9 mL**

3.72 Plan: Use the density of the solution to find the mass of 1 L of solution. Volume in liters must be converted to volume in mL. The 70.0% by mass translates to 70.0 g solute/100 g solution and is used to find the mass of HNO_3 in 1 L of solution. Convert mass of HNO_3 to moles to obtain moles/L, molarity.
Solution:

a) Mass (g) of 1 L of solution = $(1\ \text{L solution})\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right)\left(\frac{1.41\ \text{g solution}}{1\ \text{mL}}\right)$ = 1410 g solution

Mass (g) of HNO_3 in 1 L of solution = $(1410\ \text{g solution})\left(\frac{70.0\ \text{g}\ HNO_3}{100\ \text{g solution}}\right)$ = **987 g HNO_3/L**

b) Moles of HNO_3 = $(987\ \text{g}\ HNO_3)\left(\frac{1\ \text{mol}\ HNO_3}{63.02\ \text{g}\ HNO_3}\right)$ = 15.6617 mol HNO_3

Molarity of HNO_3 = $\left(\frac{15.6617\ \text{mol}\ HNO_3}{1\ \text{L solution}}\right)$ = 15.6617 = **15.7 *M* HNO_3**

3.73 Plan: Use the molarity of the solution to find the moles of H_2SO_4 in 1 mL. Convert moles of H_2SO_4 to mass of H_2SO_4, divide that mass by the mass of 1 mL of solution, and multiply by 100 for mass percent. Use the density of the solution to find the mass of 1 mL of solution.
Solution:

a) Moles of H_2SO_4 in 1 mL = $\left(\frac{18.3\ \text{mol}\ H_2SO_4}{1\ \text{L}}\right)\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)$ = **1.83x10^{-2} mol H_2SO_4/mL**

b) Mass of H_2SO_4 in 1 mL = $\left(1.83x10^{-2} \text{ mol } H_2SO_4\right)\left(\frac{98.09 \text{ g } H_2SO_4}{1 \text{ mol } H_2SO_4}\right)$ = 1.79505 g H_2SO_4

Mass of 1 mL of solution = $\left(1 \text{ mL}\right)\left(\frac{1.84 \text{ g}}{1 \text{ mL}}\right)$ = 1.84 g solution

Mass percent = $\frac{\text{mass of } H_2SO_4}{\text{mass of solution}}(100)$ = $\frac{1.79505 \text{ g } H_2SO_4}{1.84 \text{ g solution}}(100)$ = 97.5571 = **97.6% H_2SO_4 by mass**

3.74 Plan: Convert the mass of calcium carbonate to moles, and use the mole ratio in the balanced chemical equation to find the moles of hydrochloric acid required to react with these moles of calcium carbonate. Use the molarity of HCl to find the volume that contains this number of moles.

Solution:

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

Converting from grams of $CaCO_3$ to moles:

Moles of $CaCO_3$ = $\left(16.2 \text{ g } CaCO_3\right)\left(\frac{1 \text{ mol } CaCO_3}{100.09 \text{ g } CaCO_3}\right)$ = 0.161854 mol $CaCO_3$

Converting from moles of $CaCO_3$ to moles of HCl:

Moles of HCl = $\left(0.161854 \text{ mol } CaCO_3\right)\left(\frac{2 \text{ mol HCl}}{1 \text{ mol } CaCO_3}\right)$ = 0.323708 mol HCl

Converting from moles of HCl to volume:

Volume (mL) of HCl = $\left(0.323708 \text{ mol HCl}\right)\left(\frac{1 \text{ L}}{0.383 \text{ mol HCl}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$ = 845.1906 = **845 mL HCl solution**

3.75 Plan: Convert the volume of NaOH solution to liters and multiply by the molarity of the solution to obtain moles of NaOH. Use the mole ratio in the balanced chemical equation to find the moles of NaH_2PO_4 required to react with these moles of NaOH. Finally, convert moles of NaH_2PO_4 to moles.

Solution:

$NaH_2PO_4(s) + 2NaOH(aq) \rightarrow Na_3PO_4(aq) + 2H_2O(l)$

Volume (L) = $\left(43.74 \text{ mL}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)$ = 0.04374 mL

Finding moles of NaOH:

Moles of NaOH = $\left(0.04374 \text{ L}\right)\left(\frac{0.285 \text{ mol NaOH}}{1 \text{ L}}\right)$ = 0.0124659 mol NaOH

Converting from moles of NaOH to moles of NaH_2PO_4:

Moles of NaH_2PO_4 = $\left(0.0124659 \text{ mol NaOH}\right)\left(\frac{1 \text{ mol } NaH_2PO_4}{2 \text{ mol NaOH}}\right)$ = 0.00623295 mol NaH_2PO_4

Converting from moles of NaH_2PO_4 to mass:

Mass (g) of NaH_2PO_4 = $\left(0.00623295 \text{ mol } NaH_2PO_4\right)\left(\frac{119.98 \text{ g } NaH_2PO_4}{1 \text{ mol } NaH_2PO_4}\right)$ = 0.747829 = **0.748 g NaH_2PO_4**

3.76 Plan: The first step is to write and balance the chemical equation for the reaction. Multiply the molarity and volume of each of the reactants to determine the moles of each. To determine which reactant is limiting, calculate the amount of barium sulfate formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the mass of barium sulfate formed.

Solution:

The balanced chemical equation is:

$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

$$\text{Moles of BaCl}_2 = (35.0\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.160\text{ mol BaCl}_2}{1\text{ L}}\right) = 0.00560\text{ mol BaCl}_2$$

Finding the moles of $BaSO_4$ from the moles of $BaCl_2$ (if Na_2SO_4 is limiting):

$$\text{Moles of BaSO}_4\text{ from BaCl}_2 = (0.00560\text{ moL BaCl}_2)\left(\frac{1\text{ mol BaSO}_4}{1\text{ mol BaCl}_2}\right) = 0.00560\text{ mol BaSO}_4$$

$$\text{Moles of Na}_2\text{SO}_4 = (58.0\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.065\text{ mol Na}_2\text{SO}_4}{1\text{ L}}\right) = 0.00377\text{ mol Na}_2\text{SO}_4$$

Finding the moles of $BaSO_4$ from the moles of Na_2SO_4 (if $BaCl_2$ is limiting):

$$\text{Moles BaSO}_4\text{ from Na}_2\text{SO}_4 = (0.00377\text{ moL Na}_2\text{SO}_4)\left(\frac{1\text{ mol BaSO}_4}{1\text{ mol Na}_2\text{SO}_4}\right) = 0.00377\text{ mol BaSO}_4$$

Sodium sulfate is the limiting reactant.
Converting from moles of $BaSO_4$ to mass:

$$\text{Mass (g) of BaSO}_4 = (0.0377\text{ moL BaSO}_4)\left(\frac{233.4\text{ g BaSO}_4}{1\text{ mol BaSO}_4}\right) = 0.879918 = \mathbf{0.88\text{ g BaSO}_4}$$

3.77 Plan: The first step is to write and balance the chemical equation for the reaction. Use the molarity and volume of each of the reactants to determine the moles of each. To determine which reactant is limiting, calculate the amount of either product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation is:

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

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

$$\text{Moles of H}_2\text{SO}_4 = (350.0\text{ mL})\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right)\left(\frac{0.210\text{ mol H}_2\text{SO}_4}{1\text{ L}}\right) = 0.0735\text{ mol H}_2\text{SO}_4$$

Finding the moles of Na_2SO_4 from the moles of H_2SO_4 (if NaOH is limiting):

$$\text{Moles of Na}_2\text{SO}_4\text{ from H}_2\text{SO}_4 = (0.0735\text{ moL H}_2\text{SO}_4)\left(\frac{1\text{ mol Na}_2\text{SO}_4}{1\text{ mol H}_2\text{SO}_4}\right) = 0.0735\text{ mol Na}_2\text{SO}_4$$

$$\text{Moles of NaOH} = (0.500\text{ L})\left(\frac{0.196\text{ mol NaOH}}{1\text{ L}}\right) = 0.0980\text{ mol NaOH}$$

Finding the moles of Na_2SO_4 from the moles of NaOH (if H_2SO_4 is limiting):

$$\text{Moles of Na}_2\text{SO}_4\text{ from NaOH} = (0.0980\text{ mol NaOH})\left(\frac{1\text{ mol Na}_2\text{SO}_4}{2\text{ mol NaOH}}\right) = 0.0490\text{ mol Na}_2\text{SO}_4$$

NaOH is the limiting reactant and will be used in the remainder of the calculations.

$$\text{Moles of H}_2\text{SO}_4\text{ that react with NaOH} = (0.0980\text{ mol NaOH})\left(\frac{1\text{ mol H}_2\text{SO}_4}{2\text{ mol NaOH}}\right) = 0.0490\text{ mol H}_2\text{SO}_4$$

Moles of H_2SO_4 remaining = initial moles – moles reacting with NaOH

= 0.0735 mol – 0.0490 mol = $\mathbf{0.0245\text{ mol H}_2\text{SO}_4}$

3.78 Plan: The first part of the problem is a simple dilution problem ($M_1V_1 = M_2V_2$). The volume in units of gallons can be used. In part b), convert mass of HCl to moles and use the molarity to find the volume that contains that number of moles.

Solution:

a) M_1 = 11.7 M V_1 = ? M_2 = 3.5 M V_2 = 3.0 gal

$$V_1 = \frac{M_2 \text{ x } V_2}{M_1} = \frac{(3.5\ M)(3.0\text{ gal})}{11.7\ M} = 0.897436\text{ gal}$$

Instructions: Be sure to wear goggles to protect your eyes! Pour approximately 2.0 gal of water into the container. Add slowly and with mixing 0.90 gal of 11.7 M HCl into the water. Dilute to 3.0 gal with water.

b) Converting from mass of HCl to moles of HCl:

$$\text{Moles of HCl} = (9.66\text{ g HCl})\left(\frac{1\text{ mol HCl}}{36.46\text{ g HCl}}\right) = 0.264948\text{ mol HCl}$$

Converting from moles of HCl to volume:

$$\text{Volume (mL) of solution} = (0.264948\text{ mol HCl})\left(\frac{1\text{ L}}{11.7\text{ mol HCl}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

= 22.64513 = **22.6 mL muriatic acid solution**

3.79 Plan: The moles of narceine and the moles of water are required. We can assume any mass of narceine hydrate (we will use 100 g), and use this mass to determine the moles of hydrate. The moles of water in the hydrate is obtained by taking 10.8% of the 100 g mass of hydrate and converting the mass to moles of water. Divide the moles of water by the moles of hydrate to find the value of x.

Solution:

Assuming a 100 g sample of narceine hydrate:

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

= 0.20019 mol narceine hydrate

$$\text{Mass (g) of } H_2O = (100\text{ g narceine hydrate})\left(\frac{10.8\%\ H_2O}{100\%\text{ narceine hydrate}}\right) = 10.8\text{ g } H_2O$$

$$\text{Moles of } H_2O = (10.8\text{ g } H_2O)\left(\frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O}\right) = 0.59933\text{ mol } H_2O$$

$$x = \frac{\text{moles of } H_2O}{\text{moles of hydrate}} = \frac{0.59933\text{ mol}}{0.20019\text{ mol}} = \mathbf{3}$$

Thus, there are three water molecules per mole of hydrate. The formula for narceine hydrate is **narceine•3H_2O**.

3.80 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative numbers of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass percent = $\frac{\text{total mass of element}}{\text{molar mass of compound}}(100)$. List the compounds from the highest %H to the lowest.

Solution:

Name	Chemical formula	Molar mass (g/mol)	Mass percent H = $\frac{\text{moles of H x molar mass}}{\text{molar mass of compound}}(100)$
Ethane	C_2H_6	30.07	$\frac{6\text{ mol}(1.008\text{ g/mol})}{30.07\text{ g}}(100)$ = 20.11% H
Propane	C_3H_8	44.09	$\frac{8\text{ mol}(1.008\text{ g/mol})}{44.09\text{ g}}(100)$ = 18.29% H
Benzene	C_6H_6	78.11	$\frac{6\text{ mol}(1.008\text{ g/mol})}{78.11\text{ g}}(100)$ = 7.743% H

Ethanol	C_2H_5OH	46.07	$\frac{6\text{ mol}(1.008\text{ g/mol})}{46.07\text{ g}}(100)$ = 13.13% H
Cetyl palmitate	$C_{32}H_{64}O_2$	480.83	$\frac{64\text{ mol}(1.008\text{ g/mol})}{480.83\text{ g}}(100)$ = 13.42% H

The hydrogen percentage decreases in the following order:

Ethane > Propane > Cetyl palmitate > Ethanol > Benzene

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

Solution:

a) All of the substances are gases.

$H_2S(g) + O_2(g) \xrightarrow{\Delta} SO_2(g) + H_2O(g)$

There are 2 O atoms in O_2 on the left and 3 O atoms in SO_2 and H_2O on the right; place a coefficient of 2 in front of H_2O on the right and a coefficient of 2 in front of O_2 on the left for a total of 4 oxygen atoms on each side:

$H_2S(g) + 2O_2(g) \xrightarrow{\Delta} SO_2(g) + 2H_2O(g)$

Now the 4 H atoms in $2H_2O$ on the right require a coefficient of 2 in front of H_2S on the left:

$2H_2S(g) + 2O_2(g) \xrightarrow{\Delta} SO_2(g) + 2H_2O(g)$

The 2 S atoms in $2H_2S$ on the left require a coefficient of 2 in front of SO_2 on the right:

$2H_2S(g) + 2O_2(g) \xrightarrow{\Delta} 2SO_2(g) + 2H_2O(g)$

Now the O atoms are no longer balanced; the 6 O atoms on the right (4 in $2SO_2$ and 2 in $2H_2O$) require a coefficient of 6 in front of O_2 on the left:

$\mathbf{2H_2S(g) + 3O_2(g) \xrightarrow{\Delta} 2SO_2(g) + 2H_2O(g)}$

b) All of the substances are solid (crystalline).

$KClO_3(s) \xrightarrow{\Delta} KCl(s) + KClO_4(s)$

There are 3 O atoms in $KClO_3$ on the left and 4 O atoms in $KClO_4$ on the right. Place a coefficient of 4 in front of $KClO_3$ and a coefficient of 3 in front of $KClO_4$ for a total of 12 O atoms on each side. The K and Cl atoms are balanced with 4 K atoms and 4 Cl atoms on each side:

$\mathbf{4KClO_3(s) \xrightarrow{\Delta} KCl(s) + 3KClO_4(s)}$

c) Hydrogen and water vapor are gases; iron and iron(III) oxide are solids.

$H_2(g) + Fe_2O_3(s) \rightarrow Fe(s) + H_2O(g)$

The 2 Fe atoms in Fe_2O_3 on the left require a coefficient of 2 in front of Fe on the right:

$H_2(g) + Fe_2O_3(s) \rightarrow 2Fe(s) + H_2O(g)$

The 3 O atoms in Fe_2O_3 on the left require a coefficient of 3 in front of H_2O on the right:

$H_2(g) + Fe_2O_3(s) \rightarrow 2Fe(s) + 3H_2O(g)$

The 6 H atoms in $3H_2O$ on the right require a coefficient of 3 in front of H_2 on the left:

$\mathbf{3H_2(g) + Fe_2O_3(s) \rightarrow 2Fe(s) + 3H_2O(g)}$

d) All of the substances are gases; combustion required oxygen as a reactant.

$C_2H_6(g) + O_2(g) \xrightarrow{\Delta} CO_2(g) + H_2O(g)$

The 2 C atoms in C_2H_6 on the left require a coefficient of 2 in front of CO_2 on the right:

$C_2H_6(g) + O_2(g) \xrightarrow{\Delta} 2CO_2(g) + H_2O(g)$

The 6 H atoms in C_2H_6 on the left require a coefficient of 3 in front of H_2O on the right:

$C_2H_6(g) + O_2(g) \xrightarrow{\Delta} 2CO_2(g) + 3H_2O(g)$

The 7 O atoms on the right (4 in $2CO_2$ and 3 in $3H_2O$) require a coefficient of 7/2 in front of O_2 on the left:

$C_2H_6(g) + 7/2O_2(g) \xrightarrow{\Delta} 2CO_2(g) + 3H_2O(g)$

Double all coefficients to get whole number coefficients:

$\mathbf{2C_2H_6(g) + 7O_2(g) \xrightarrow{\Delta} 4CO_2(g) + 6H_2O(g)}$

e) Iron(II) chloride and iron(III) fluoride are solids and the other substances are gases.

$FeCl_2(s) + ClF_3(g) \rightarrow FeF_3(s) + Cl_2(g)$

There are 3 Cl atoms on the left (2 in $FeCl_2$ and 1 in ClF_3) and 2 Cl atoms in Cl_2 on the right. Place a coefficient of 2 in front of Cl_2 and a coefficient of 2 in front of ClF_3 on the left for a total of 4 Cl atoms on each side:

$FeCl_2(s) + 2ClF_3(g) \rightarrow FeF_3(s) + 2Cl_2(g)$

The 6 F atoms in $2ClF_3$ require a coefficient of 2 in front of FeF_3 on the right:

$FeCl_2(s) + 2ClF_3(g) \rightarrow 2FeF_3(s) + 2Cl_2(g)$

The 2 Fe atoms in FeF_3 on the right require a coefficient of 2 in front of $FeCl_2$ on the left:

$2FeCl_2(s) + 2ClF_3(g) \rightarrow 2FeF_3(s) + 2Cl_2(g)$

Now the Cl atoms are not balanced with 6 on the left (4 in $2FeCl_2$ and 2 in $2ClF_3$) and 4 in $2Cl_2$ on the right;
place
a coefficient of 3 in front of Cl_2 on the right:

$\mathbf{2FeCl_2(s) + 2ClF_3(g) \rightarrow 2FeF_3(s) + 3Cl_2(g)}$

3.82 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO_2, and all of the hydrogen present in the sample is found in the hydrogen of H_2O. Convert the mass of CO_2 to moles and use the ratio between CO_2 and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H_2O. Divide the moles of C and H by the smaller value to convert to whole numbers to get the empirical formula.

Solution:

Isobutylene + $O_2 \rightarrow CO_2 + H_2O$

$$\text{Moles of C} = (2.657\ \text{g}\ CO_2)\left(\frac{1\ \text{mol}\ CO_2}{44.01\ \text{g}\ CO_2}\right)\left(\frac{1\ \text{mol C}}{1\ \text{mol}\ CO_2}\right) = 0.06037\ \text{mol C}$$

$$\text{Moles of H} = (1.089\ \text{g}\ H_2O)\left(\frac{1\ \text{mol}\ H_2O}{18.02\ \text{g}\ H_2O}\right)\left(\frac{2\ \text{mol H}}{1\ \text{mol}\ H_2O}\right) = 0.1209\ \text{mol H}$$

Preliminary formula = $C_{0.06037}H_{0.1209}$

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

$$C_{\frac{0.06037}{0.06037}}H_{\frac{0.1209}{0.06037}} \rightarrow C_1H_2$$

This gives an empirical formula of $\mathbf{CH_2}$.

3.83 Plan: Write a balanced equation. Use the density of toluene to convert the given volume of toluene to mass and divide by the molar mass of toluene to convert mass to moles. Use the mole ratio between toluene and oxygen to find the moles and then mass of oxygen required for the reaction. The mole ratio between toluene and the gaseous products are used to find the moles of product produced. The moles of water are multiplied by Avogadro's number to find the number of water molecules.

Solution:

The balanced chemical equation is:

$C_7H_8(l) + 9O_2(g) \rightarrow 7CO_2(g) + 4H_2O(g)$

$$\text{a) Moles of}\ C_7H_8 = (20.0\ \text{mL}\ C_7H_8)\left(\frac{0.867\ \text{g}\ C_7H_8}{1\ \text{mL}\ C_7H_8}\right)\left(\frac{1\ \text{mol}\ C_7H_8}{92.13\ \text{g}\ C_7H_8}\right) = 0.1882123\ \text{mol}\ C_7H_8$$

$$\text{Mass (g) oxygen} = (0.1882123\ \text{mol}\ C_7H_8)\left(\frac{9\ \text{mol}\ O_2}{1\ \text{mol}\ C_7H_8}\right)\left(\frac{32.00\ \text{g}\ O_2}{1\ \text{mol}\ O_2}\right) = 54.20514 = \mathbf{54.2\ g\ O_2}$$

$$\text{b) Total moles of gas} = (0.1882123\ \text{mol}\ C_7H_8)\left(\frac{11\ \text{mol product gas}}{1\ \text{mol}\ C_7H_8}\right) = 2.07034 = \mathbf{2.07\ mol\ of\ gas}$$

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

$$\text{c) Moles of}\ H_2O = (0.1882123\ \text{mol}\ C_7H_8)\left(\frac{4\ \text{mol}\ H_2O}{1\ \text{mol}\ C_7H_8}\right) = 0.7528492\ \text{mol}\ H_2O$$

$$\text{Molecules of } H_2O = (0.7528492 \text{ mol } H_2O)\left(\frac{6.022 \times 10^{23} \ H_2O \text{ molecules}}{1 \text{ mol } H_2O}\right)$$

$$= 4.53366 \times 10^{23} = \mathbf{4.53 \times 10^{23} \textbf{ molecules } H_2O}$$

3.84 Plan: If 100.0 g of dinitrogen tetroxide reacts with 100.0 g of hydrazine (N_2H_4), what is the theoretical yield of nitrogen if no side reaction takes place? First, we need to identify the limiting reactant. To determine which reactant is limiting, calculate the amount of nitrogen formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of nitrogen. Then determine the amount of limiting reactant required to produce 10.0 grams of NO. Reduce the amount of limiting reactant by the amount used to produce NO. The reduced amount of limiting reactant is then used to calculate an "actual yield." The "actual" and theoretical yields will give the maximum percent yield.

Solution:

The balanced reaction is $2N_2H_4(l) + N_2O_4(l) \rightarrow 3N_2(g) + 4H_2O(g)$

Determining the limiting reactant:

Finding the moles of N_2 from the amount of N_2O_4 (if N_2H_4 is limiting):

$$\text{Moles of } N_2 \text{ from } N_2O_4 = (100.0 \text{ g } N_2O_4)\left(\frac{1 \text{ mol } N_2O_4}{92.02 \text{ g } N_2O_4}\right)\left(\frac{3 \text{ mol } N_2}{1 \text{ mol } N_2O_4}\right) = 3.26016 \text{ mol } N_2$$

Finding the moles of N_2 from the amount of N_2H_4 (if N_2O_4 is limiting):

$$N_2 \text{ from } N_2H_4 = (100.0 \text{ g } N_2H_4)\left(\frac{1 \text{ mol } N_2H_4}{32.05 \text{ g } N_2H_4}\right)\left(\frac{3 \text{ mol } N_2}{2 \text{ mol } N_2H_4}\right) = 4.68019 \text{ mol } N_2$$

N_2O_4 is the limiting reactant.

$$\text{Theoretical yield of } N_2 = (100.0 \text{ g } N_2O_4)\left(\frac{1 \text{ mol } N_2O_4}{92.02 \text{ g } N_2O_4}\right)\left(\frac{3 \text{ mol } N_2}{1 \text{ mol } N_2O_4}\right)\left(\frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2}\right) = 91.3497 \text{ g } N_2$$

How much of the limiting reactant is used to produce 10.0 g NO?

$N_2H_4(l) + 2N_2O_4(l) \rightarrow 6NO(g) + 2H_2O(g)$

$$\text{Mass (g) of } N_2O_4 \text{ used} = (10.0 \text{ g NO})\left(\frac{1 \text{ mol NO}}{30.01 \text{ g NO}}\right)\left(\frac{2 \text{ mol } N_2O_4}{6 \text{ mol NO}}\right)\left(\frac{92.02 \text{ g } N_2O_4}{1 \text{ mol } N_2O_4}\right)$$

$$= 10.221 \text{ g } N_2O_4$$

Amount of N_2O_4 available to produce N_2 = 100.0 g N_2O_4 – mass of N_2O_4 required to produce 10.0 g NO

= 100.0 g – 10.221 g = 89.779 g N_2O_4

Determine the "actual yield" of N_2 from 89.779 g N_2O_4:

$$\text{“Actual yield” of } N_2 = (89.779 \text{ g } N_2O_4)\left(\frac{1 \text{ mol } N_2O_4}{92.02 \text{ g } N_2O_4}\right)\left(\frac{3 \text{ mol } N_2}{1 \text{ mol } N_2O_4}\right)\left(\frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2}\right)$$

$$= 82.01285 \text{ g } N_2$$

$$\text{Theoretical yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right)(100) = \left(\frac{82.01285}{91.3497}\right)(100) = 89.7790 = \mathbf{89.8\%}$$

3.85 Plan: Identify the product molecules and write the balanced equation. To determine the limiting reactant for part b), examine the product circle to see which reactant remains in excess and which reactant was totally consumed. For part c), use the mole ratios in the balanced equation to determine the number of moles of product formed by each reactant, assuming the other reactant is in excess. The reactant that produces fewer moles of product is the limiting reactant. Use the mole ratio between the two reactants to determine the moles of excess reactant required to react with the limiting reactant. The difference between the initial moles of excess reactant and the moles required for reaction is the moles of excess reactant that remain.

Solution:

a) The contents of the circles give:

$AB_2 + B_2 \rightarrow AB_3$

Balancing the reaction gives:

$$\mathbf{2AB_2 + B_2 \rightarrow 2AB_3}$$

b) Two B_2 molecules remain after reaction so B_2 is in excess. All of the AB_2 molecules have reacted so $\mathbf{AB_2}$ is the limiting reactant.

c) Finding the moles of AB_3 from the moles of AB_2 (if B_2 is limiting):

$$\text{Moles of AB}_3 \text{ from AB}_2 = (5.0 \text{ mol AB}_2)\left(\frac{2 \text{ mol AB}_3}{2 \text{ mol AB}_2}\right) = 5.0 \text{ mol AB}_3$$

Finding the moles of AB_3 from the moles of B_2 (if AB_2 is limiting):

$$\text{Moles of AB}_3 \text{ from B}_2 = (3.0 \text{ mol B}_2)\left(\frac{2 \text{ mol AB}_3}{1 \text{ mol B}_2}\right) = 6.0 \text{ mol AB}_3$$

AB_2 is the limiting reagent and **5.0 mol of AB_3** is formed.

$$\text{d) Moles of B}_2 \text{ that react with 5.0 mol AB}_2 = (5.0 \text{ mol AB}_2)\left(\frac{1 \text{ mol B}_2}{2 \text{ mol AB}_2}\right) = 2.5 \text{ mol B}_2$$

The unreacted B_2 is 3.0 mol – 2.5 mol = **0.5 mol B_2**.

3.86 Plan: Since 85% of ions in seawater are from NaCl, take 85% of the mass percent of dissolved ions (4.0%) to find the mass % of NaCl in part a). To find the mass % of Na^+ and Cl^- individually in part b), use the ratio of the mass of the two ions to the mass of NaCl. To find the molarity in part c), use the mass of NaCl in 100 g of seawater; convert mass of NaCl to moles and mass of seawater to volume in liters, using the density. Molarity = moles of NaCl/L of seawater.

Solution:

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

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

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

c) Since the mass % of NaCl is 3.4%, there are 3.4 g of NaCl in 100 g of seawater.

$$\text{Moles of NaCl} = (3.4 \text{ g NaCl})\left(\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}}\right) = 0.0581793 \text{mol NaCl}$$

$$\text{Volume (L) of 100 g of seawater} = (100 \text{ g seawater})\left(\frac{1 \text{ mL}}{1.025 \text{ g seawater}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.097561 \text{ L}$$

$$M \text{ NaCl} = \frac{\text{moles NaCl}}{\text{L seawater}} = \frac{0.0581793 \text{ mol}}{0.097561 \text{ L}} = 0.596338 = \mathbf{0.60 \text{ } \boldsymbol{M} \text{ NaCl}}$$

3.87 a) **False**, a mole of one substance has the same number of units as a mole of any other substance.

b) **True**

c) **False**, a limiting-reactant problem is present when the quantity of available material is given for more than one reactant.

d) **True**

3.88 Plan: Count the total number of spheres in each box. The number in box A divided by the volume change in each part will give the number we are looking for and allow us to match boxes.

Solution:

The number in each box is: A = 12, B = 6, C = 4, and D = 3.

a) When the volume is tripled, there should be 12/3 = 4 spheres in a box. This is box **C**.

b) When the volume is doubled, there should be 12/2 = 6 spheres in a box. This is box **B**.

c) When the volume is quadrupled, there should be 12/4 = 3 spheres in a box. This is box **D**.

3.89 Plan: To convert mass to moles, divide the mass by the molar mass of the substance. To convert moles to mass, divide by the molar mass. To obtain number of particles, multiply moles by Avogadro's number. Divide a number of particles by Avogadro's number to obtain moles.

Solution:

a) Since 1 mole of any substance contains Avogadro's number of entities, equal amounts of moles of various substances contain equal numbers of entities. The number of entities (O_3 molecules) in 0.4 mol of O_3 is **equal** to the number of entities (O atoms) in 0.4 mol of O atoms.

b) O_3 has a molar mass of 3(16.0 g/mol O) = 48.0 g/mol; O has a molar mass of 1(16.0 g/mol O) = 16.0 g/mol. Since O_3 has a larger molar mass than O, **0.4 mol of O_3** has a greater mass than 0.4 mol of O.

c) Moles of N_2O_4 = $\left(4.0\ \text{g}\ N_2O_4\right)\left(\frac{1\ \text{mol}\ N_2O_4}{92.02\ \text{g}\ N_2O_4}\right)$ = 0.043 mol N_2O_4

Moles of SO_2 = $\left(3.3\ \text{g}\ SO_2\right)\left(\frac{1\ \text{mol}\ SO_2}{64.07\ \text{g}\ SO_2}\right)$ = 0.052 mol SO_2

SO_2 is the larger quantity in terms of moles.

d) Mass (g) of C_2H_4 = $\left(0.6\ \text{mol}\ C_2H_4\right)\left(\frac{28.05\ \text{g}\ C_2H_4}{1\ \text{mol}\ C_2H_4}\right)$ = 17 g C_2H_4

Mass (g) of F_2 = $\left(0.6\ \text{mol}\ F_2\right)\left(\frac{38.00\ \text{g}\ F_2}{1\ \text{mol}\ F_2}\right)$ = 23 g F_2

F_2 is the greater quantity in terms of mass.

Note that if each of these values is properly rounded to one significant figure, the answers are identical.

e) Total moles of ions in 2.3 mol $NaClO_3$ = $\left(2.3\ \text{mol}\ NaClO_3\right)\left(\frac{2\ \text{mol ions}}{1\ \text{mol}\ NaClO_3}\right)$ = 4.6 mol ions

Total moles of ions in 2.2 mol $MgCl_2$ = $\left(2.2\ \text{mol}\ MgCl_2\right)\left(\frac{3\ \text{mol ions}}{1\ \text{mol}\ MgCl_2}\right)$ = 6.6 mol ions

$MgCl_2$ is the greater quantity in terms of total moles of ions.

f) The compound with the lower molar mass will have more molecules in a given mass. H_2O (18.02 g/mol) has a lower molar mass than H_2O_2 (34.02 g/mol). **1.0 g H_2O** has more molecules than 1.0 g H_2O_2.

g) Moles of NaBr = $\left(0.500\ \text{L NaBr}\right)\left(\frac{0.500\ \text{mol}}{1\ \text{L}}\right)$ = 0.250 mol NaBr

Moles of Na^+ = $\left(0.250\ \text{mol NaBr}\right)\left(\frac{1\ \text{mol}\ Na^+}{1\ \text{mol NaBr}}\right)$ = 0.250 mol Na^+

Moles of NaCl = $\left(0.0146\ \text{kg NaCl}\right)\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol NaCl}}{58.44\ \text{g NaCl}}\right)$ = 0.250 mol NaCl

Moles of Na^+ = $\left(0.250\ \text{mol NaCl}\right)\left(\frac{2\ \text{mol ions}}{1\ \text{mol NaCl}}\right)$ = 0.250 mol Na^+

The two quantities are **equal**.

h) The heavier atoms, **^{238}U**, will give a greater total mass since there is an equal number of particles of both.

3.90 Plan: Write a balanced equation. The coefficients in the balanced equation give the number of molecules or moles of each reactant and product. Moles are converted to amount in grams by multiplying by the molar masses.

Solution:

$P_4S_3(s) + 8O_2(g) \rightarrow P_4O_{10}(s) + 3SO_2(g)$

a) 1 molecule of P_4S_3 reacts with 8 molecules of O_2 to produce 1 molecule of P_4O_{10} and 3 molecules of SO_2.

b) 1 mol of P_4S_3 reacts with 8 mol of O_2 to produce 1 mol of P_4O_{10} and 3 mol of SO_2.

c) 220.09 g of P_4S_3 react with 8(32.00 g/mol O) = 256.00 g of O_2 to produce 283.88 g of P_4O_{10} and 3(64.07 g/mol SO_2) = 192.21 g of SO_2.

3.91 Plan: Write a balanced equation. Use the actual yield (105 kg) and the percent yield (98.8%) to find the theoretical yield of hydrogen. Use the mole ratio between hydrogen and water in the balanced equation to obtain the amount of hydrogen required to produce that theoretical yield of water.

Solution:

The balanced equation is $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

$$\% \text{ yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \text{x } 100\%$$

$$\text{Theoretical yield (g) of } H_2O = \frac{\text{actual yield}}{\%\text{ yield}}(100) = \frac{105 \text{ kg}}{98.8\%}(100) = 106.2753 \text{ kg } H_2O$$

$$\text{Mass (g) of } H_2 = (106.2753 \text{ kg } H_2O)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)\left(\frac{2 \text{ mol } H_2}{2 \text{ mol } H_2O}\right)\left(\frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2}\right)$$

$$= 1.18896\text{x}10^4 = \mathbf{1.19x10^4 \text{ g } H_2}$$

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

Solution:

$$M \text{ KBr} = \frac{\text{total moles KBr}}{\text{total volume}} = \frac{\text{moles KBr from solution 1 + moles KBr from solution 2}}{\text{volume solution 1 + volume solution 2}}$$

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

3.93 Plan: Divide the given mass of a substance by its molar mass to obtain moles; multiply the given moles of a substance by its molar mass to obtain mass in grams. Number of particles is obtained by multiplying an amount in moles by Avogadro's number. Density is used to convert mass to volume.

Solution:

a) $\text{Moles of } NH_4Br = (0.588 \text{ g } NH_4Br)\left(\frac{1 \text{ mol } NH_4Br}{97.94 \text{ g } NH_4Br}\right) = 0.0060037 = \mathbf{0.00600 \text{ mol } NH_4Br}$

b) $\text{Moles of } KNO_3 = (88.5 \text{ g } KNO_3)\left(\frac{1 \text{ mol } KNO_3}{101.11 \text{ g } KNO_3}\right) = 0.875284 \text{ mol } KNO_3$

$$\text{Number of } K^+ \text{ ions} = (0.875284 \text{ mol } KNO_3)\left(\frac{1 \text{ mol } K^+}{1 \text{ mol } KNO_3}\right)\left(\frac{6.022 \text{ x } 10^{23}\ K^+\text{ions}}{1 \text{ mol } K^+}\right)$$

$$= 5.27096\text{x}10^{23} = \mathbf{5.27x10^{23}\ K^+ \text{ ions}}$$

c) $\text{Mass (g) of } C_3H_8O_3 = (5.85 \text{ mol } C_3H_8O_3)\left(\frac{92.09 \text{ g } C_3H_8O_3}{1 \text{ mol } C_3H_8O_3}\right) = 538.7265 = \mathbf{539 \text{ g } C_3H_8O_3}$

d) $\text{Mass (g) of } CHCl_3 = (2.85 \text{ mol } CHCl_3)\left(\frac{119.37 \text{ g } CHCl_3}{1 \text{ mol } CHCl_3}\right) = 340.2045 \text{ g } CHCl_3$

$$\text{Volume (mL) of } CHCl_3 = (340.2045 \text{ g } CHCl_3)\left(\frac{\text{mL}}{1.48 \text{ g } CHCl_3}\right) = 229.868 = \mathbf{230. \text{ mL } CHCl_3}$$

e) $\text{Moles of } Na^+ = (2.11 \text{ mol } Na_2CO_3)\left(\frac{2 \text{ mol } Na^+}{1 \text{ mol } Na_2CO_3}\right) = 4.22 \text{ mol } Na^+$

$$\text{Number of } Na^+ = (4.22 \text{ mol } Na^+)\left(\frac{6.022\text{x}10^{23}\ Na^+ \text{ ions}}{1 \text{ mol } Na^+}\right) = 2.54128\text{x}10^{24} = \mathbf{2.54x10^{24}\ Na^+ \text{ ions}}$$

f) Moles of Cd atoms = $\left(25.0\mu g\ Cd\right)\left(\frac{10^{-6}\ g}{1\mu g}\right)\left(\frac{1\ mol\ Cd}{112.4\ g\ Cd}\right)$ = $2.224199x10^{-7}$ mol Cd atoms

Number of Cd atoms = $\left(2.224199x10^{-7}\ mol\ Cd\right)\left(\frac{6.022\,x\,10^{23}\ Cd\ atoms}{1\ mol\ Cd}\right)$

= $1.3394126x10^{17}$ = **$1.34x10^{17}$ Cd atoms**

g) Number of F atoms = $\left(0.0015\ mol\ F_2\right)\left(\frac{2\ mol\ F}{1\ mol\ F_2}\right)\left(\frac{6.022\,x\,10^{23}\ F\ atoms}{1\ mol\ F}\right)$

= $1.8066x10^{21}$ = **$1.8x10^{21}$ F atoms**

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

3.95 Plan: Deal with the methane and propane separately, and combine the results. Balanced equations are needed for each hydrocarbon. The total mass and the percentages will give the mass of each hydrocarbon. The mass of each hydrocarbon is changed to moles, and through the balanced chemical equation the amount of CO_2 produced by each gas may be found. Summing the amounts of CO_2 gives the total from the mixture. For part b), let x and 252 – x represent the masses of CH_4 and C_3H_8, respectively.

Solution:

a) The balanced chemical equations are:

Methane: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

Propane: $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

Mass (g) of CO_2 from each:

Methane: $\left(200.g\ Mixture\right)\left(\frac{25.0\%}{100\%}\right)\left(\frac{1\ mol\ CH_4}{16.04\ g\ CH_4}\right)\left(\frac{1\ mol\ CO_2}{1\ mol\ CH_4}\right)\left(\frac{44.01\ g\ CO_2}{1\ mol\ CO_2}\right)$ = 137.188 g CO_2

Propane: $\left(200.g\ Mixture\right)\left(\frac{75.0\%}{100\%}\right)\left(\frac{1\ mol\ C_3H_8}{44.09\ g\ C_3H_8}\right)\left(\frac{3\ mol\ CO_2}{1\ mol\ C_3H_8}\right)\left(\frac{44.01\ g\ CO_2}{1\ mol\ CO_2}\right)$ = 449.183 g CO_2

Total CO_2 = 137.188 g + 449.183 g = 586.371 = **586 g CO_2**

b) Since the mass of CH_4 + the mass of C_3H_8 = 252 g, let x = mass of CH_4 in the mixture and 252 – x = mass of C_3H_8 in the mixture. Use mole ratios to calculate the amount of CO_2 formed from x amount of CH_4 and the amount of CO_2 formed from 252 – x amount of C_3H_8.

The total mass of CO_2 produced = 748 g.

The total moles of CO_2 produced = $\left(748\ g\ CO_2\right)\left(\frac{1\ mol\ CO_2}{44.01\ g\ CO_2}\right)$ = 16.996 mol CO_2

16.996 mol CO_2 =

$$\left(x\ g\ CH_4\right)\left(\frac{1\ mol\ CH_4}{16.04\ g\ CH_4}\right)\left(\frac{1\ mol\ CO_2}{1\ mol\ CH_4}\right) + \left(252\ -\ x\ g\ C_3H_8\right)\left(\frac{1\ mol\ C_3H_8}{44.09\ g\ C_3H_8}\right)\left(\frac{3\ mol\ CO_2}{1\ mol\ C_3H_8}\right)$$

$$16.996\ mol\ CO_2 = \frac{x}{16.04}\ mol\ CO_2 + \frac{3(252\ -\ x)}{44.09}\ mol\ CO_2$$

$$16.996\ mol\ CO_2 = \frac{x}{16.04}\ mol\ CO_2 + \frac{756\ -\ 3x}{44.09}\ mol\ CO_2$$

16.996 mol CO_2 = 0.06234x mol CO_2 + (17.147 – 0.06804x mol CO_2)

16.996 = 17.147 – 0.0057x

x = 26.49 g CH_4 252 – x = 252 g – 26.49 g = 225.51 g C_3H_8

Mass % CH_4 = $\frac{mass\ of\ CH_4}{mass\ of\ mixture}(100) = \frac{26.49\ g\ CH_4}{252\ g\ mixture}(100)$ = **10.5% CH_4**

$$\text{Mass \% } C_3H_8 = \frac{\text{mass of } C_3H_8}{\text{mass of mixture}}(100) = \frac{225.51 \text{ g } C_3H_8}{252 \text{ g mixture}}(100) = \mathbf{89.5\% \ C_3H_8}$$

3.96 Plan: If we assume a 100-gram sample of fertilizer, then the 30:10:10 percentages become the masses, in grams, of N, P_2O_5, and K_2O. These masses may be changed to moles of substance, and then to moles of each element. To get the desired x:y:1.0 ratio, divide the moles of each element by the moles of potassium.

Solution:

A 100-gram sample of 30:10:10 fertilizer contains 30 g N, 10 g P_2O_5, and 10 g K_2O.

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

$$\text{Moles of P} = (10 \text{ g } P_2O_5)\left(\frac{1 \text{ mol } P_2O_5}{141.94 \text{ g } P_2O_5}\right)\left(\frac{2 \text{ mol P}}{1 \text{ mol } P_2O_5}\right) = 0.14090 \text{ mol P}$$

$$\text{Moles of K} = (10 \text{ g } K_2O)\left(\frac{1 \text{ mol } K_2O}{94.20 \text{ g } K_2O}\right)\left(\frac{2 \text{ mol K}}{1 \text{ mol } K_2O}\right) = 0.21231 \text{ mol K}$$

This gives a N:P:K ratio of 2.1413:0.14090:0.21231

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

$$\frac{2.1413 \text{ mol N}}{0.21231} = 10.086 \qquad \frac{0.14090 \text{ mol P}}{0.21231} = 0.66365 \qquad \frac{0.21231 \text{ mol K}}{0.21231} = 1$$

10.086:0.66365:1.000 or **10:0.66:1.0**

3.97 Plan: If we assume a 100-gram sample of fertilizer, then the 10:10:10 percentages become the masses, in grams, of N, P_2O_5, and K_2O. These masses may be changed to moles of substance, and then to moles of each element. Use the mole ratio between N and ammonium sulfate, P and ammonium hydrogen phsophate, and K and potassium chloride to find the mass of each compound required to provide the needed amount of the respective element. Divide the mass of each compound by the total mass of sample, 100 g, and multiply by 100 for mass %.

Solution:

Assume a 100 g sample. 10:10:10 indicates 10 g N, 10 g P_2O_5 and 10 g K_2O.

$$\text{Moles of N} = (10 \text{ g N})\left(\frac{1 \text{ mol N}}{14.01 \text{ g N}}\right) = 0.713776 \text{ mol N}$$

$$\text{Moles of P} = (10 \text{ g } P_2O_5)\left(\frac{1 \text{ mol } P_2O_5}{141.94 \text{ g } P_2O_5}\right)\left(\frac{2 \text{ mol P}}{1 \text{ mol } P_2O_5}\right) = 0.14090 \text{ mol P}$$

$$\text{Moles of K} = (10 \text{ g } K_2O)\left(\frac{1 \text{ mol } K_2O}{94.20 \text{ g } K_2O}\right)\left(\frac{2 \text{ mol K}}{1 \text{ mol } K_2O}\right) = 0.21231 \text{ mol K}$$

To obtain 0.713776 mol N from $(NH_4)_2SO_4$:

$$(0.713776 \text{ mol N})\left(\frac{1 \text{ mol } (NH_4)_2SO_4}{2 \text{ mol N}}\right)\left(\frac{132.15 \text{ g } (NH_4)_2SO_4}{1 \text{ mol } (NH_4)_2SO_4}\right) = 47.1627 \text{ g } (NH_4)_2SO_4$$

$$\text{Mass \% } (NH_4)_2SO_4 = \frac{\text{mass of } (NH_4)_2SO_4}{\text{mass of mixture}}(100) = \frac{47.1627 \text{ g } (NH_4)_2SO_4}{100 \text{ g mixture}}(100)$$

$$= 47.1627\% = \mathbf{47.2\% \ (NH_4)_2SO_4}$$

To obtain 0.14090 mol P from $(NH_4)_2HPO_4$:

$$(0.14090 \text{ mol P})\left(\frac{1 \text{ mol } (NH_4)_2HPO_4}{1 \text{ mol P}}\right)\left(\frac{132.06 \text{ g } (NH_4)_2HPO_4}{1 \text{ mol } (NH_4)_2HPO_4}\right) = 18.6073 \text{ g } (NH_4)_2HPO_4$$

$$\text{Mass \% } (NH_4)_2HPO_4 = \frac{\text{mass of } (NH_4)_2HPO_4}{\text{mass of mixture}}(100) = \frac{18.6073 \text{ g } (NH_4)_2HPO_4}{100 \text{ g mixture}}(100)$$

$$= 18.6073\% = \mathbf{18.6\% \ (NH_4)_2HPO_4}$$

To obtain 0.21231 mol K from KCl:

$$(0.21231 \text{ mol K})\left(\frac{1 \text{ mol KCl}}{1 \text{ mol K}}\right)\left(\frac{74.55 \text{ g KCl}}{1 \text{ mol KCl}}\right) = 15.8277 \text{ g KCl}$$

$$\text{Mass \% KCl} = \frac{\text{mass of KCl}}{\text{mass of mixture}}(100) = \frac{15.8277 \text{ g KCl}}{100 \text{ g mixture}}(100) = 15.8277\% = \mathbf{15.8\% \ KCl}$$

3.98 Plan: Write a balanced equation. Convert the mass of strontium sulfate produced to moles and use the mole ratio in the balanced equation to find the moles of strontium halide required to produce that amount of product. Divide the given mass of strontium halide by the moles of strontium halide to obtain its molar mass. Subtracting the molar mass of strontium from the molar mass of compound gives the molar mass of the halogen in the formula. The molar mass of the halogen is used to identify the halogen.

Solution:

$SrX_2(aq) + H_2SO_4(aq) \rightarrow SrSO_4(s) + 2\ HX(aq)$

0.652 g   0.755 g

$$\text{Moles } SrX_2 = (0.755 \text{ g } SrSO_4)\left(\frac{1 \text{ mol } SrSO_4}{183.69 \text{ g } SrSO_4}\right)\left(\frac{1 \text{ mol } SrX_2}{1 \text{ mol } SrSO_4}\right) = 0.004110186 \text{ mol } SrX_2$$

The 0.652 g sample of SrX_2 = 0.004110186 mol

$$SrX_2 = \frac{0.652 \text{ g}}{0.004110186 \text{ mol}} = 158.630 \text{ g/mol} = \text{molar mass}$$

Molar mass of X_2 = 158.630 g – molar mass of Sr

Molar mass of X_2 = 158.630 g – 87.62 g = 71.01 g = X_2

Molar mass of X = 71.01 g/2 = 35.505 = 35.5 g/mol = Cl   The original halide formula is $\mathbf{SrCl_2}$.

3.99 Plan: Assume 100 grams of mixture. This means the mass of each compound, in grams, is the same as its percentage. Find the mass of C from CO and from CO_2 and add these masses together. For mass %, divide the total mass of C by the mass of the mixture (100 g) and multiply by 100.

Solution:

100 g of mixture = 35 g CO and 65 g CO_2.

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

$$\text{Mass (g) of C from } CO_2 = (65.0 \text{ g } CO_2)\left(\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}\right)\left(\frac{1 \text{ mol C}}{1 \text{ mol } CO_2}\right)\left(\frac{12.01 \text{ g C}}{1 \text{ mol C}}\right) = 17.738 \text{ g C}$$

Total mass (g) of C = 15.007 g + 17.738 g = 32.745 g C

$$\text{Mass \% C} = \frac{\text{mass of C}}{\text{mass of mixture}}(100) = \frac{32.745 \text{ g C}}{100 \text{ g mixture}}(100) = 32.745 = \mathbf{32.7\% \ C}$$

3.100 Plan: Write a balanced equation for the reaction. Count the molecules of each reactant to obtain the moles of each reactant present. Use the mole ratios in the equation to calculate the amount of product formed. Only 87.0% of the calculated amount of product actually forms, so the actual yield is 87.0% of the theoretical yield.

Solution:

The balanced equation is $SiH_4 + N_2F_4 \rightarrow SiF_4 + N_2 + 2H_2$.

$$\text{Moles of } SiH_4 = (3 \text{ } SiH_4 \text{ molecules})\left(\frac{1.25\text{x}10^{-2} \text{ mol}}{1 \text{ molecule}}\right) = 0.0375 \text{ mol } SiH_4$$

$$\text{Moles of } N_2F_4 = (3 \text{ } N_2F_4 \text{ molecules})\left(\frac{1.25\text{x}10^{-2} \text{ mol}}{1 \text{ molecule}}\right) = 0.0375 \text{ mol } N_2F_4$$

Since there is an equal amount of each reactant and the ratio between each reactant and SiF_4 is 1:1, neither reactant is in excess and either may by used to calculate the amount of SiF_4 produced.

$$\text{Mass (g) of } SiF_4 = (0.0375 \text{ mol } SiH_4)\left(\frac{1 \text{ mol } SiF_4}{1 \text{ mol } SiH_4}\right)\left(\frac{104.09 \text{ g } SiF_4}{1 \text{ mol } SiF_4}\right) = 3.903375 \text{ g } SiF_4$$

$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \text{x } 100\%$$

$$\text{Actual yield (g) of } SiF_4 = \frac{\% \text{ yield}}{100\%}(\text{theoretical yield}) = \frac{87\%}{100\%}(3.903375 \text{ g } SiF_4) = 3.3959 = \mathbf{3.4 \text{ g } SiF_4}$$

3.101 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO_2, and all of the hydrogen present in the sample is found in the hydrogen of H_2O. Convert the mass of CO_2 to moles and use the ratio between CO_2 and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H_2O. Subtracting the masses of C and H from the mass of the sample gives the mass of Fe. Convert the mass of Fe to moles of Fe. Take the moles of C, H, and Fe and divide by the smallest value to convert to whole numbers to get the empirical formula.

Solution:

Ferrocene + $?O_2(g) \rightarrow CO_2$ + H_2O

0.9437 g 2.233 g 0.457 g

$$\text{Moles of C} = (2.233 \text{ g } CO_2)\left(\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}\right)\left(\frac{1 \text{ mol C}}{1 \text{ mol } CO_2}\right) = 0.050738 \text{ mol C}$$

$$\text{Mass (g) of C} = (0.050738 \text{ mol C})\left(\frac{12.01 \text{ g C}}{1 \text{ mol C}}\right) = 0.60936 \text{ g C}$$

$$\text{Moles of H} = (0.457 \text{ g } H_2O)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)\left(\frac{2 \text{ mol H}}{1 \text{ mol } H_2O}\right) = 0.050721 \text{ mol H}$$

$$\text{Mass (g) of H} = (0.050721 \text{ mol H})\left(\frac{1.008 \text{ g H}}{1 \text{ mol H}}\right) = 0.051127 \text{ g H}$$

Mass (g) of Fe = Sample mass – (mass of C + mass of H)

= 0.9437 g – (0.60936 g C + 0.052217 g H) = 0.283213 g Fe

$$\text{Moles of Fe} = (0.283213 \text{ g Fe})\left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right) = 0.005071 \text{ mol Fe}$$

Preliminary formula = $C_{0.050738}H_{0.050721}Fe_{0.005071}$

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

$$C_{\frac{0.050738}{0.005071}}H_{\frac{0.050721}{0.005071}}Fe_{\frac{0.005071}{0.005071}} \rightarrow C_{10}H_{10}Fe_1$$

Empirical formula = $C_{10}H_{10}Fe$

3.102 Plan: Determine the molecular formula from the figure. Once the molecular formula is known, use the periodic table to determine the molar mass. Convert the volume of lemon juice in part b) from qt to mL and use the density to convert from mL to mass in g. Take 6.82% of that mass to find the mass of citric acid and use the molar mass to convert to moles.

Solution:

a) The formula of citric acid obtained by counting the number of carbon atoms, oxygen atoms, and hydrogen atoms is $\mathbf{C_6H_8O_7}$.

Molar mass = (6 x 12.01 g/mol C) + (8 x 1.008 g/mol H) + (7 x 16.00 g/mol O) = **192.12 g/mol**

b) Converting volume of lemon juice in qt to mL:

$$\text{Volume (mL) of lemon juice} = (1.50 \text{ qt})\left(\frac{1 \text{ L}}{1.057 \text{ qt}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right) = 1419.111 \text{ mL}$$

Converting volume to mass in grams:

$$\text{Mass (g) of lemon juice} = (1419.111 \text{ mL})\left(\frac{1.09 \text{ g}}{\text{mL}}\right) = 1546.831 \text{ g lemon juice}$$

$$\text{Mass (g) of } C_6H_8O_7 = (1546.831 \text{ g lemon juice})\left(\frac{6.82\% \ C_6H_8O_7}{100\% \text{ lemon juice}}\right) = 105.494 \text{ g } C_6H_8O_7$$

$$\text{Moles of } C_6H_8O_7 = (105.494 \text{ g } C_6H_8O_7)\left(\frac{1 \text{ mol } C_6H_8O_7}{192.12 \text{ g } C_6H_8O_7}\right) = 0.549104 = \mathbf{0.549 \text{ mol } C_6H_8O_7}$$

3.103 Plan: For parts a) and b), convert the masses to moles. Take the moles and divide by the smallest value to convert to whole numbers to get the empirical formula. For part c), write the two balanced equations and use two equations as shown.

Solution:

$$\text{a) Moles of Pt} = (0.327 \text{ g Pt})\left(\frac{1 \text{ mol Pt}}{195.1 \text{ g Pt}}\right) = 0.001676 \text{ mol Pt}$$

Mass (g) of F = mass of product – mass of Pt = 0.519 g – 0.327 g = 0.192 g F

$$\text{Moles of F} = (0.192 \text{ g F})\left(\frac{1 \text{ mol F}}{19.00 \text{ g F}}\right) = 0.010105 \text{ mol F}$$

Preliminary formula = $Pt_{0.001676}F_{0.010105}$

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

$$Pt_{\frac{0.001676}{0.001676}}F_{\frac{0.0010105}{0.001676}} \rightarrow Pt_1F_6$$

Empirical formula = $\mathbf{PtF_6}$

$$\text{b) Moles of } PtF_6 = (0.265 \text{ g } PtF_6)\left(\frac{1 \text{ mol } PtF_6}{309.1 \text{ g } PtF_6}\right) = 0.0008576 \text{ mol } PtF_6$$

Mass of Xe = mass of product – mass of Xe = 0.378 g – 0.265 g = 0.113 g Xe

$$\text{Moles of Xe} = (0.113 \text{ g Xe})\left(\frac{1 \text{ mol Xe}}{131.3 \text{ g Xe}}\right) = 0.0008606 \text{ mol Xe}$$

Preliminary formula = $Xe_{0.0008606}(PtF_6)_{0.0008576}$

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

$$Xe_{\frac{0.0008606}{0.0008576}}(PtF_6)_{\frac{0.0008576}{0.0008576}} \rightarrow Xe_1(PtF_6)_1$$

Empirical formula = $\mathbf{XePtF_6}$

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

The two equations are:	The two unknowns are:
$Xe(g) + 2F_2(g) \rightarrow XeF_4(s)$	x = mol XeF_4 produced
$Xe(g) + 3F_2(g) \rightarrow XeF_6(s)$	y = mol XeF_6 produced

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

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

2x + 3y = $5.00\text{x}10^{-4}$ mol F_2 consumed

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

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

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

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

$y = (5.00\text{x}10^{-4}) - (3.52\text{x}10^{-4}) = 1.48\text{x}10^{-4}$ mol XeF_6

$x = (1.76\text{x}10^{-4}) - (1.48\text{x}10^{-4}) = 2.8\text{x}10^{-5}$ mol XeF_4

Converting moles of each product to grams using the molar masses:

$$\text{Mass (g) of } XeF_4 = (2.8\text{x}10^{-5} \text{ mol } XeF_4)\left(\frac{207.3 \text{ g } XeF_4}{1 \text{ mol } XeF_4}\right) = 5.8044\text{x}10^{-3} \text{ g } XeF_4$$

$$\text{Mass (g) of } XeF_6 = (1.48\text{x}10^{-4} \text{ mol } XeF_6)\left(\frac{245.3 \text{ g } XeF_6}{1 \text{ mol } XeF_6}\right) = 3.63044\text{x}10^{-2} \text{ g } XeF_6$$

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

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

$$\text{Mass \% XeF}_4 = \frac{\text{mass of XeF}_4}{\text{total mass}}(100) = \frac{5.8044\text{x}10^{-3}\text{ g XeF}_4}{0.0421088\text{ g}}(100) = 13.784 = \mathbf{14\%\ XeF_4}$$

$$\text{Mass \% XeF}_6 = \frac{\text{mass of XeF}_6}{\text{total mass}}(100) = \frac{3.63044\text{x}10^{-2}\text{ g XeF}_6}{0.0421088\text{ g}}(100) = 86.2157 = \mathbf{86.2\%\ XeF_6}$$

3.104 Plan: Use the mass percent to find the mass of heme in the sample; use the molar mass to convert the mass of heme to moles. Then find the mass of Fe in the sample by using the mole ratio between heme and iron. The mass of hemin is found by using the mole ratio between heme and hemoglobin.

Solution:

a) $\text{Mass (g) of heme} = (0.65\text{ g hemoglobin})\left(\frac{6.0\%\text{ heme}}{100\%\text{ hemoglobin}}\right) = \mathbf{0.039\ g\ heme}$

b) $\text{Moles of heme} = (0.039\text{ g heme})\left(\frac{1\text{ mol heme}}{616.49\text{ g heme}}\right) = 6.32614\text{x}10^{-5} = \mathbf{6.3x10^{-5}\ mol\ heme}$

c) $\text{Mass (g) of Fe} = (6.32614\text{x}10^{-5}\text{ mol heme})\left(\frac{1\text{ mol Fe}}{1\text{ mol heme}}\right)\left(\frac{55.85\text{ g Fe}}{1\text{ mol Fe}}\right)$

$= 3.5331\text{x}10^{-3} = \mathbf{3.5x10^{-3}\ g\ Fe}$

d) $\text{Mass (g) of hemin} = (6.32614\text{x}10^{-5}\text{ mol heme})\left(\frac{1\text{ mol hemin}}{1\text{ mol heme}}\right)\left(\frac{651.94\text{ g hemin}}{1\text{ mol hemin}}\right)$

$= 4.1243\text{x}10^{-2} = \mathbf{4.1x10^{-2}\ g\ hemin}$

3.105 Plan: Find the Mn:O ratio in the two oxides. Write two equations to solve simultaneously; one equation shows that the sum of the ratio of Mn in the two oxides will equal the ratio of Mn in the sample and the other equation shows that the total amount of oxide in the sample is the sum of the amounts of the two oxides. The two equations will give the mole ratio of the two oxides. Convert moles of each oxide to mass to obtain the mass ratio of the two oxides from which the mass % of each can be calculated. Use that mass % of each to find the mass of each in the sample. For part b), the moles of Mn^{3+} come from the Mn_2O_3 and the moles of Mn^{2+} come from the MnO.

Solution:

Mn:O ratio:

In sample:	1.00:1.42	or	0.704
In braunite:	2.00:3.00	or	0.667
In manganosite:	1.00:1.00	or	1.00

a) The total amount of ore is equal to the amount of braunite (B) + the amount of manganosite (M).

B + M = 1.00

M = 1.00 – B

The amount of Mn is dependent on the sample's composition.

M(1.00) + B(0.667) = 0.704

(1.00 – B)(1.00) + B(0.667) = 0.704

1.00 – 1.00B + 0.667B = 0.704

0.296 = 0.333B

B = 0.888889 mol braunite

M = 1.00 – B = 1.00 = 0.888889 = 0.111111 mol manganosite

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

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

There are 140.338 g of braunite for every 7.88221 g of manganosite. Finding mass % of each:

$$\text{Mass \% braunite} = \frac{\text{mass of braunite}}{\text{mass of braunite + manganosite}}(100) = \frac{140.338 \text{ g}}{140.338 + 7.88221 \text{ g}}(100) = 94.6821\%$$

$$\text{Mass \% manganosite} = \frac{\text{mass of manganosite}}{\text{mass of braunite + manganosite}}(100) = \frac{7.88221 \text{ g}}{140.338 + 7.88221 \text{ g}}(100)$$

$$= 5.3179\%$$

In the 542.3 g sample:

$$\text{Mass (g) of braunite} = (542.3 \text{ g sample})\left(\frac{94.6821 \text{ braunite}}{100\% \text{ sample}}\right) = 513.461 = \mathbf{513 \text{ g braunite}}$$

$$\text{Mass (g) of manganosite} = (542.3 \text{ g sample})\left(\frac{5.3179\% \text{ manganosite}}{100\% \text{ sample}}\right) = 28.839 = \mathbf{28.8 \text{ g manganosite}}$$

b) Each mole of braunite, Mn_2O_3, contains 2 moles of Mn^{3+} while each mole of manganosite, MnO, contains 1 mole of Mn^{2+}.

Moles of Mn^{3+} = 2(0.888889 mol braunite) = 1.777778 mol Mn^{3+}

Moles of Mn^{2+} = 1(0.111111 mol manganosite) = 0.111111 mol Mn^{2+}

$$Mn^{3+}:Mn^{2+} = \frac{1.777778 \text{ mol } Mn^{3+}}{0.111111 \text{ mol } Mn^{2+}} = 16.000 = \mathbf{16.0}$$

3.106 Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of hydroxyapatite formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of hydroxyapatite formed.

Solution:

a) $5Ca(OH)_2(aq) + 3H_3PO_4(aq) \rightarrow Ca_5(PO_4)_3(OH)(s) + 9H_2O(l)$

b) Find the limiting reagent.

$$\text{Moles of } Ca_5(PO_4)_3(OH) \text{ from } Ca(OH)_2 = (100.\text{ g } Ca(OH)_2)\left(\frac{1 \text{ mol } Ca(OH)_2}{74.10 \text{ g } Ca(OH)_2}\right)\left(\frac{1 \text{ mol } Ca_5(PO_4)_3(OH)}{5 \text{ mol } Ca(OH)_2}\right)$$

$$= 0.2699055 \text{ mol } Ca_5(PO_4)_3(OH)$$

Moles of $Ca_5(PO_4)_3(OH)$ from H_3PO_4 =

$$(100.\text{ g } H_3PO_4 \text{ solution})\left(\frac{85 \text{ g } H_3PO_4}{100.\text{ g } H_3PO_4 \text{ solution}}\right)\left(\frac{1 \text{ mol } H_3PO_4}{97.99 \text{ g } H_3PO_4}\right)\left(\frac{1 \text{ mol } Ca_5(PO_4)_3(OH)}{3 \text{ mol } H_3PO_4}\right)$$

$$= 0.2891452 \text{ mol } Ca_5(PO_4)_3(OH)$$

$Ca(OH)_2$ is the limiting reactant, and will be used to calculate the yield.

$$(100.\text{ g } Ca(OH)_2)\left(\frac{1 \text{ mol } Ca(OH)_2}{74.10 \text{ g } Ca(OH)_2}\right)\left(\frac{1 \text{ mol } Ca_5(PO_4)_3(OH)}{5 \text{ mol } Ca(OH)_2}\right)\left(\frac{502.32 \text{ g } Ca_5(PO_4)_3(OH)}{1 \text{ mol } Ca_5(PO_4)_3(OH)}\right)$$

$$= 135.57893 = \mathbf{140 \text{ g } Ca_5(PO_4)_3(OH)}$$

3.107 Plan: To determine which reactant is limiting, calculate the amount of aspirin formed from each reactant, assuming an excess of the other reactant. Use the density of acetic anhydride to determine the amount of this reactant in grams. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of aspirin. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield. Use the formula for percent atom economy to determine that quantity.

Solution:

a) Finding the moles of aspirin from the moles of $C_7H_6O_3$ (if $(CH_3CO)_2O$ is limiting):

$$\text{Moles of aspirin from } C_7H_6O_3 = (3.077 \text{ g } C_7H_6O_3)\left(\frac{1 \text{ mol } C_7H_6O_3}{138.12 \text{ g } C_7H_6O_3}\right)\left(\frac{1 \text{ mol } C_9H_8O_4}{1 \text{ mol } C_7H_6O_3}\right)$$

$$= 0.0222777 \text{ mol } C_9H_8O_4$$

Finding the moles of aspirin from the moles of $C_4H_6O_3$ (if $C_7H_6O_3$ is limiting):

Mass (g) of $(CH_3CO)_2O$ = $\left(5.50 \text{ mL } (CH_3CO)_2O\right)\left(\frac{1.080 \text{ g}}{1 \text{ mL}}\right)$ = 5.94 g $(CH_3CO)_2O$

Moles of aspirin from $(CH_3CO)_2O$ = $\left(5.94 \text{ g } (CH_3CO)_2O\right)\left(\frac{1 \text{ mol } (CH_3CO)_2O}{102.09 \text{ g } (CH_3CO)_2O}\right)\left(\frac{1 \text{ mol } C_9H_8O_4}{1 \text{ mol } (CH_3CO)_2O}\right)$

= 0.058183955 mol $C_9H_8O_4$

The limiting reactant is $\mathbf{C_7H_6O_3}$.

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

Mass (g) of $C_9H_8O_4$ = $\left(3.077 \text{ g } C_7H_6O_3\right)\left(\frac{1 \text{ mol } C_7H_6O_3}{138.12 \text{ g } C_7H_6O_3}\right)\left(\frac{1 \text{ mol } C_9H_8O_4}{1 \text{ mol } C_7H_6O_3}\right)\left(\frac{180.15 \text{ g } C_9H_8O_4}{1 \text{ mol } C_9H_8O_4}\right)$

= 4.01333 g $C_9H_8O_4$

Percent yield = $\left(\frac{\text{actual yield}}{\text{theoretical yield}}\right)$ x 100% = $\left(\frac{3.281 \text{ g}}{4.01333 \text{ g}}\right)$ x 100% = 81.7526 = **81.75% yield**

3.108 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of nitrogen present. Multiply the number of moles of nitrogen by its molar mass to find the total mass of nitrogen in 1 mole of compound. Mass percent = $\frac{\text{total mass of element}}{\text{molar mass of compound}}(100)$. For part b), convert mass of ornithine to moles, use the mole ratio between ornithine and urea to find the moles of urea, and then use the ratio between moles of urea and nitrogen to find the moles and mass of nitrogen produced.

Solution:

a) Urea: CH_4N_2O, $\mathcal{M}$ = 60.06 g/mol

There are 2 moles of N in 1 mole of CH_4N_2O.

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

Mass percent = $\frac{\text{total mass N}}{\text{molar mass of compound}}(100)$ = $\frac{28.02 \text{ g N}}{60.06 \text{ g } CH_4N_2O}(100)$ = 46.6533 = **46.65% N in urea**

Arginine: $C_6H_{15}N_4O_2$, $\mathcal{M}$ = 175.22 g/mol

There are 4 moles of N in 1 mole of $C_6H_{15}N_4O_2$.

Mass (g) of N = $\left(4 \text{ mol N}\right)\left(\frac{14.01 \text{ g N}}{1 \text{ mol N}}\right)$ = 56.04 g N

Mass percent = $\frac{\text{total mass N}}{\text{molar mass of compound}}(100)$ = $\frac{56.04 \text{ g N}}{175.22 \text{ g } C_6H_{15}N_4O_2}(100)$

= 31.98265 = **31.98% N in arginine**

Ornithine: $C_5H_{13}N_2O_2$, $\mathcal{M}$ = 133.17 g/mol

There are 2 moles of N in 1 mole of $C_5H_{13}N_2O_2$.

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

Mass percent = $\frac{\text{total mass N}}{\text{molar mass of compound}}(100)$ = $\frac{28.02 \text{ g N}}{133.17 \text{ g } C_5H_{13}N_2O_2}(100)$

= 21.04077 = **21.04% N in ornithine**

b) Moles of urea = $\left(135.2 \text{ g } C_5H_{13}N_2O_2\right)\left(\frac{1 \text{ mol } C_5H_{13}N_2O_2}{133.17 \text{ g } C_5H_{13}N_2O_2}\right)\left(\frac{1 \text{ mol } CH_4N_2O}{1 \text{ mol } C_5H_{13}N_2O_2}\right)$ = 1.015244 mol urea

Mass (g) of nitrogen = $\left(1.015244 \text{ mol } CH_4N_2O\right)\left(\frac{2 \text{ mol N}}{1 \text{ mol } CH_4N_2O}\right)\left(\frac{14.01 \text{ g N}}{1 \text{ mol N}}\right)$ = 28.447 = **28.45 g N**

3.109 Plan: Write and balance the chemical reaction. Use the mole ratio to find the amount of product that should be produced and take 66% of that amount to obtain the actual yield.

Solution:

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

With 6 molecules of NO and 3 molecules of O_2 reacting, 6 molecules of NO_2 can be produced.

If the reaction only has a 66% yield, then (0.66)(6) = 4 molecules of NO_2 will be produced**. Circle A** shows the formation of 4 molecules of NO_2. Circle B also shows the formation of 4 molecules of NO_2 but also has 2 unreacted molecules of NO and 1 unreacted molecule of O_2. Since neither reactant is limiting, there will be no unreacted reactant remaining after the reaction is over.

3.110 Plan: First balance the given chemical equation. To determine which reactant is limiting, calculate the amount of ZnS formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of ZnS. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield. For part b), determine the mass of Zn that does not produce ZnS; use that amount of zinc and the mole ratio between Zn and ZnO in that reaction to determine the mass of ZnO produced. Find the moles of S_8 in the reactant and the moles of S_8 in the product ZnS. The difference between these two amounts is the moles of S_8 in SO_2.

Solution:

a) The balanced equation is $8Zn(s) + S_8(s) \rightarrow 8ZnS(s)$.

Finding the limiting reagent:

Finding the moles of ZnS from the moles of Zn (if S_8 is limiting):

$$\text{Moles of ZnS from Zn} = (83.2 \text{ g Zn})\left(\frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}}\right)\left(\frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}}\right) = 1.27198 \text{ mol ZnS}$$

Finding the moles of ZnS from the moles of S_8 (if Zn is limiting):

$$\text{Moles of ZnS from } S_8 = (52.4 \text{ g } S_8)\left(\frac{1 \text{ mol } S_8}{256.56 \text{ g } S_8}\right)\left(\frac{8 \text{ mol ZnS}}{1 \text{ mol } S_8}\right) = 1.6339 \text{ mol ZnS}$$

The zinc will produce less zinc sulfide, thus, zinc is the limiting reactant and will first be used to determine the theoretical yield and then the percent yield.

$$\text{Theoretical yield (g) of ZnS} = (83.2 \text{ g Zn})\left(\frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}}\right)\left(\frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}}\right)\left(\frac{97.48 \text{ g ZnS}}{1 \text{ mol ZnS}}\right)$$

$$= 123.9923 \text{ g ZnS (unrounded)}$$

$$\text{Percent yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}}\right) \text{x } 100\% = \left(\frac{104.4 \text{ g}}{123.9923 \text{ g}}\right) \text{x } 100\% = 84.1988 = \mathbf{84.2\%\ yield}$$

b) The reactions with oxygen are:

$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$

$S_8(s) + 8O_2(g) \rightarrow 8SO_2(g)$

The theoretical yield indicates that 84.2% of the zinc produced zinc sulfide so (100 – 84.2)% = 15.8% of the zinc became zinc oxide. This allows the calculation of the amount of zinc oxide formed.

$$\text{Mass (g) of Zn that does not produce ZnS} = (83.2 \text{ g Zn})\left(\frac{15.8\%}{100\%}\right) = 13.1456 \text{ g ZnS}$$

$$\text{Mass (g) of ZnO} = (13.1456 \text{ g Zn})\left(\frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}}\right)\left(\frac{2 \text{ mol ZnO}}{2 \text{ mol Zn}}\right)\left(\frac{81.41 \text{ g ZnO}}{1 \text{ mol ZnO}}\right) = 16.3612 = \mathbf{16.4\ g\ ZnO}$$

The calculation is slightly different for the sulfur. We need to determine the amount of sulfur not in zinc sulfide. The sulfur not in the zinc sulfide must be in sulfur dioxide. The amount of sulfur not in zinc sulfide will be converted to the mass of sulfur dioxide.

$$\text{Moles of } S_8 \text{ in original } S_8 \text{ reactant} = (52.4 \text{ g } S_8)\left(\frac{1 \text{ mol } S_8}{256.56 \text{ g } S_8}\right) = 0.204241 \text{ mol } S_8$$

Moles of S_8 in ZnS product = $(104.4\ \text{g ZnS})\left(\frac{1\ \text{mol ZnS}}{97.48\ \text{g ZnS}}\right)\left(\frac{1\ \text{mol } S_8}{8\ \text{mol ZnS}}\right)$ = 0.133874 mol S_8

Moles of S_8 in SO_2 = 0.204241 – 0.133874 ml = 0.070367 mol S_8

Mass (g) of SO_2 = $(0.070367\ \text{mol } S_8)\left(\frac{8\ \text{mol } SO_2}{1\ \text{mol } S_8}\right)\left(\frac{64.07\ \text{g } SO_2}{1\ \text{mol } SO_2}\right)$ = 36.0673 = **36.1 g SO_2**

3.111 Plan: Use the given values of x to find the molar mass of each compound. . To determine which reactant is limiting, calculate the amount of either product formed from each reactant, assuming an excess of the other reactants. The reactant that produces the smallest amount of product is the limiting reagent. To find the mass of excess reactants, find the mass of each excess reactant that is required to react with the limiting reagent and subtract that mass from the starting mass.

a) x = 0

$La_2Sr_0CuO_4$ = 2(138.9 g/mol La) + 0(87.62 g/mol Sr) + 1(63.55 g/mol Cu) + 4(16.00 g/mol O) = **405.4 g/mol**

x = 1

$La_1Sr_1CuO_4$ = 1(138.9 g/mol La) + 1(87.62 g/mol Sr) + 1(63.55 g/mol Cu) + 4(16.00 g/mol O) = **354.1 g/mol**

x = 0.163

$La_{(2-0.163)}Sr_{0.163}CuO_4 = La_{1.837}Sr_{0.163}CuO_4$

= 1.837(138.9g/mol La) + 0.163 (87.62 g/mol Sr) + 1(63.55g/mol Cu) + 4(16.00g/mol O)

= **397.0 g/mol**

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

Moles of product from $BaCO_3$ = $(x\ \text{g } BaCO_3)\left(\frac{1\ \text{mol } BaCO_3}{197.3\ \text{g } BaCO_3}\right)\left(\frac{2\ \text{mol } YBa_2Cu_3O_7}{4\ \text{mol } BaCO_3}\right)$

= 0.002534x mol product

Moles of product from CuO = $(x\ \text{g CuO})\left(\frac{1\ \text{mol CuO}}{79.55\ \text{g CuO}}\right)\left(\frac{2\ \text{mol } YBa_2Cu_3O_7}{6\ \text{mol CuO}}\right)$ = 0.004190x mol product

Moles of product from Y_2O_3 = $(x\ \text{g } Y_2O_3)\left(\frac{1\ \text{mol } Y_2O_3}{225.82\ \text{g } Y_2O_3}\right)\left(\frac{2\ \text{mol } YBa_2Cu_3O_7}{1\ \text{mol } Y_2O_3}\right)$ = 0.008857x mol product

$BaCO_3$ is the limiting reactant.

c) These calculations are based on the limiting reactant.

$BaCO_3$ remaining = 0% (limiting reagent)

CuO remaining = x g CuO – $(x\ \text{g } BaCO_3)\left(\frac{1\ \text{mol } BaCO_3}{197.3\ \text{g } BaCO_3}\right)\left(\frac{6\ \text{mol CuO}}{4\ \text{mol } BaCO_3}\right)\left(\frac{79.55\ \text{g CuO}}{1\ \text{mol CuO}}\right)$

= 0.39521x g CuO

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

Y_2O_3 remaining = x g Y_2O_3 – $(x\ \text{g } BaCO_3)\left(\frac{1\ \text{mol } BaCO_3}{197.3\ \text{g } BaCO_3}\right)\left(\frac{1\ \text{mol } Y_2O_3}{4\ \text{mol } BaCO_3}\right)\left(\frac{225.82\ \text{g } Y_2O_3}{1\ \text{mol } Y_2O_3}\right)$

= 0.713862x g Y_2O_3

Percent Y_2O_3 = $\left(\frac{0.713862\ \text{x g}}{\text{x g}}\right)$ x 100% = 71.3862 = **71.39% Y_2O_3 remaining**