CHAPTER 3

STOICHIOMETRY

Atomic Masses and the Mass Spectrometer

23. 186.207 = 0.6260(186.956) + 0.3740(A), 186.207 – 117.0 = 0.3740(A)

$$A = \frac{69.2}{0.3740} = 185 \text{ amu (A = 184.95 amu without rounding to proper significant figures)}$$

24. A = 0.0140(203.973) + 0.2410(205.9745) + 0.2210(206.9759) + 0.5240(207.9766)

A = 2.86 + 49.64 + 45.74 + 109.0 = 207.2 amu; from the periodic table, the element is Pb.

25. Average atomic mass = A = 0.0800(45.95269) + 0.0730(46.951764) + 0.7380(47.947947) + 0.0550(48.947841) + 0.0540(49.944792) = 47.88 amu

This is element Ti (titanium).

26. Because we are not given the relative masses of the isotopes, we need to estimate the masses of the isotopes. A good estimate is to assume that only the protons and neutrons contribute to the overall mass of the atom and that the atomic mass of a proton and neutron are each 1.00 amu. So the masses are about: ^{54}Fe, 54.00 amu; ^{56}Fe, 56.00 amu; ^{57}Fe, 57.00 amu; ^{58}Fe, 58.00 amu. Using these masses and the abundances given in the mass spectrum, the calculated average atomic mass would be:

0.0585(54.00) + 0.9175(56.00) + 0.0212(57.00) + 0.0028(58.00) = 55.91 amu

The average atomic mass listed in the periodic table is 55.85 amu.

27. If silver is 51.82% ^{107}Ag, then the remainder is ^{109}Ag (48.18%). Determining the atomic mass (A) of ^{109}Ag:

$$107.868 = \frac{51.82(106.905) + 48.18(A)}{100}$$

10786.8 = 5540. + (48.18)A, A = 108.9 amu = atomic mass of ^{109}Ag

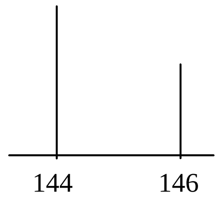
28. Let x = % of ^{151}Eu and y = % of ^{153}Eu, then $x + y = 100$ and $y = 100 - x$.

$$151.96 = \frac{x(150.9196) + (100 - x)(152.9209)}{100}$$

15196 = (150.9196)x + 15292.09 – (152.9209)x, –96 = –(2.0013)x

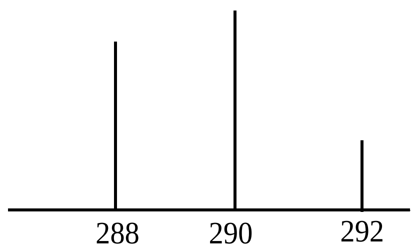
x = 48%; 48% ^{151}Eu and 100 – 48 = 52% ^{153}Eu

29. GaAs can be either $^{69}GaAs$ or $^{71}GaAs$. The mass spectrum for GaAs will have two peaks at 144 (= 69 + 75) and 146 (= 71 + 75) with intensities in the ratio of 60 : 40 or 3 : 2.



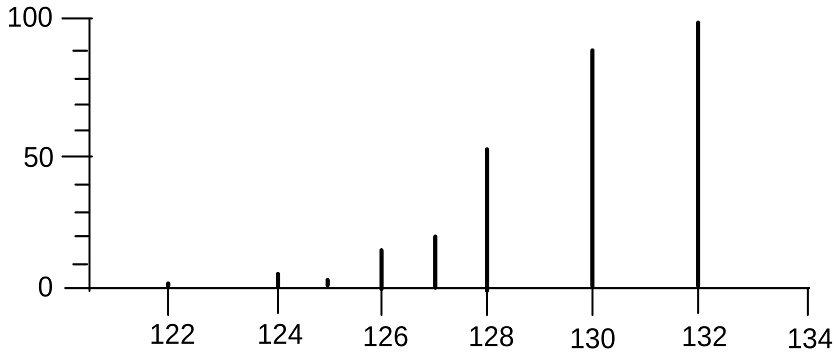
Ga_2As_2 can be $^{69}Ga_2As_2$, $^{69}Ga^{71}GaAs_2$, or $^{71}Ga_2As_2$. The mass spectrum will have three peaks at 288, 290, and 292 with intensities in the ratio of 36 : 48 : 16 or 9 : 12 : 4. We get this ratio from the following probability table:

	^{69}Ga (0.60)	^{71}Ga (0.40)
^{69}Ga (0.60)	0.36	0.24
^{71}Ga (0.40)	0.24	0.16



30.

Compound	Mass	Intensity	Scaled Intensity Largest Peak = 100
$H_2{}^{120}Te$	121.92	0.09	0.3
$H_2{}^{122}Te$	123.92	2.46	7.1
$H_2{}^{123}Te$	124.92	0.87	2.5
$H_2{}^{124}Te$	125.92	4.61	13.4
$H_2{}^{125}Te$	126.92	6.99	20.3
$H_2{}^{126}Te$	127.92	18.71	54.3
$H_2{}^{128}Te$	129.92	31.79	92.2
$H_2{}^{130}Te$	131.93	34.48	100.0



31. There are three peaks in the mass spectrum, each 2 mass units apart. This is consistent with two isotopes, differing in mass by two mass units. The peak at 157.84 corresponds to a Br_2 molecule composed of two atoms of the lighter isotope. This isotope has mass equal to 157.84/2, or 78.92. This corresponds to ^{79}Br. The second isotope is ^{81}Br with mass equal to 161.84/2 = 80.92. The peaks in the mass spectrum correspond to $^{79}Br_2$, $^{79}Br^{81}Br$, and $^{81}Br_2$ in order of increasing mass. The intensities of the highest and lowest masses tell us the two isotopes are present at about equal abundance. The actual abundance is 50.68% ^{79}Br and 49.32% ^{81}Br.

Moles and Molar Masses

32. $4.24 \text{ g } C_6H_6 \times \dfrac{1 \text{ mol}}{78.11 \text{ g}} = 5.43 \times 10^{-2} \text{ mol } C_6H_6$

$$5.43 \times 10^{-2} \text{ mol } C_6H_6 \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 3.27 \times 10^{22} \text{ molecules } C_6H_6$$

Each molecule of C_6H_6 contains 6 atoms C + 6 atoms H = 12 atoms total.

$$3.27 \times 10^{22} \text{ molecules } C_6H_6 \times \frac{12 \text{ atoms total}}{\text{molecule}} = 3.92 \times 10^{23} \text{ atoms total}$$

$$0.224 \text{ mol } H_2O \times \frac{18.02 \text{ g}}{\text{mol}} = 4.04 \text{ g } H_2O$$

$$0.224 \text{ mol } H_2O \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 1.35 \times 10^{23} \text{ molecules } H_2O$$

$$1.35 \times 10^{23} \text{ molecules } H_2O \times \frac{3 \text{ atoms total}}{\text{molecule}} = 4.05 \times 10^{23} \text{ atoms total}$$

$$2.71 \times 10^{22} \text{ molecules } CO_2 \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} = 4.50 \times 10^{-2} \text{ mol } CO_2$$

$$4.50 \times 10^{-2} \text{ mol } CO_2 \times \frac{44.01 \text{ g}}{\text{mol}} = 1.98 \text{ g } CO_2$$

$$2.71 \times 10^{22} \text{ molecules } CO_2 \times \frac{3 \text{ atoms total}}{\text{molecule } CO_2} = 8.13 \times 10^{22} \text{ atoms total}$$

$$3.35 \times 10^{22} \text{ atoms total} \times \frac{1 \text{ molecule}}{6 \text{ atoms total}} = 5.58 \times 10^{21} \text{ molecules } CH_3OH$$

$$5.58 \times 10^{21} \text{ molecules } CH_3OH \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} = 9.27 \times 10^{-3} \text{ mol } CH_3OH$$

$$9.27 \times 10^{-3} \text{ mol } CH_3OH \times \frac{32.04 \text{ g}}{\text{mol}} = 0.297 \text{ g } CH_3OH$$

33. $$4.0 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2} \times \frac{6.022 \times 10^{23} \text{ atoms H}}{1 \text{ mol H}} = 2.4 \times 10^{24} \text{ atoms}$$

$$4.0 \text{ g He} \times \frac{1 \text{ mol He}}{4.003 \text{ g He}} \times \frac{6.022 \times 10^{23} \text{ atoms He}}{1 \text{ mol He}} = 6.0 \times 10^{23} \text{ atoms}$$

$$1.0 \text{ mol } F_2 \times \frac{2 \text{ mol F}}{1 \text{ mol } F_2} \times \frac{6.022 \times 10^{23} \text{ atoms F}}{1 \text{ mol F}} = 1.2 \times 10^{24} \text{ atoms}$$

$$44.0 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{3 \text{ mol atoms } (1 \text{ C} + 2 \text{ O})}{1 \text{ mol } CO_2} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}}$$

$$= 1.81 \times 10^{24} \text{ atoms}$$

$$146 \text{ g } SF_6 \times \frac{1 \text{ mol } SF_6}{146.07 \text{ g } SF_6} \times \frac{7 \text{ mol atoms } (1 \text{ S} + 6 \text{ F})}{1 \text{ mol } SF_6} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}} = 4.21 \times 10^{24} \text{ atoms}$$

The order is: 4.0 g He < 1.0 mol F_2 < 44.0 g CO_2 < 4.0 g H_2 < 146 g SF_6

34. Molar mass of $C_6H_8O_6$ = 6(12.011) + 8(1.0079) + 6(15.999) = 176.123 g/mol

$$500.0 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol}}{176.12 \text{ g}} = 2.839 \times 10^{-3} \text{ mol}$$

$$2.839 \times 10^{-3} \text{ mol} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 1.710 \times 10^{21} \text{ molecules}$$

35. a. 9(12.011) + 8(1.0079) + 4(15.999) = 180.158 g/mol

b. $$500. \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol}}{180.16 \text{ g}} = 2.78 \times 10^{-3} \text{ mol}$$

$$2.78 \times 10^{-3}\ \text{mol} \times \frac{6.022 \times 10^{23}\ \text{molecules}}{\text{mol}} = 1.67 \times 10^{21}\ \text{molecules}$$

36. a. $$14\ \text{mol C} \times \frac{12.011\ \text{g}}{\text{mol C}} + 18\ \text{mol H} \times \frac{1.0079\ \text{g}}{\text{mol H}} + 2\ \text{mol N} \times \frac{14.007\ \text{g}}{\text{mol N}} + 5\ \text{mol O} \times \frac{15.999\ \text{g}}{\text{mol O}} = 294.305\ \text{g/mol}$$

b. $$10.0\ \text{g}\ C_{14}H_{18}N_2O_5 \times \frac{1\ \text{mol}\ C_{14}H_{18}N_2O_5}{294.3\ \text{g}\ C_{14}H_{18}N_2O_5} = 3.40 \times 10^{-2}\ \text{mol}\ C_{14}H_{18}N_2O_5$$

c. $$1.56\ \text{mol} \times \frac{294.3\ \text{g}}{\text{mol}} = 459\ \text{g}\ C_{14}H_{18}N_2O_5$$

d. $$5.0\ \text{mg} \times \frac{1\ \text{g}}{1000\ \text{mg}} \times \frac{1\ \text{mol}}{294.3\ \text{g}} \times \frac{6.02 \times 10^{23}\ \text{molecles}}{\text{mol}} = 1.0 \times 10^{19}\ \text{molecules}\ C_{14}H_{18}N_2O_5$$

e. $$1.2\ \text{g}\ C_{14}H_{18}N_2O_5 \times \frac{1\ \text{mol}\ C_{14}H_{18}N_2O_5}{294.3\ \text{g}\ C_{14}H_{18}N_2O_5} \times \frac{2\ \text{mol N}}{\text{mol}\ C_{14}H_{18}N_2O_5} \times \frac{6.02 \times 10^{23}\ \text{atoms N}}{\text{mol N}} = 4.9 \times 10^{21}\ \text{atoms N}$$

f. $$1.0 \times 10^{9}\ \text{molecules} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{atoms}} \times \frac{294.3\ \text{g}}{\text{mol}} = 4.9 \times 10^{-13}\ \text{g}$$

g. $$1\ \text{molecule} \times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}} \times \frac{294.305\ \text{g}}{\text{mol}} = 4.887 \times 10^{-22}\ \text{g}\ C_{14}H_{18}N_2O_5$$

37. a. $2(12.01) + 3(1.008) + 3(35.45) + 2(16.00) = 165.39$ g/mol

b. $$500.0\ \text{g} \times \frac{1\ \text{mol}}{165.39\ \text{g}} = 3.023\ \text{mol}\ C_2H_3Cl_3O_2$$

c. $$2.0 \times 10^{-2}\ \text{mol} \times \frac{165.39\ \text{g}}{\text{mol}} = 3.3\ \text{g}\ C_2H_3Cl_3O_2$$

d. $$5.0\ \text{g}\ C_2H_3Cl_3O_2 \times \frac{1\ \text{mol}}{165.39\ \text{g}} \times \frac{6.02 \times 10^{23}\ \text{molecules}}{\text{mol}} \times \frac{3\ \text{atoms Cl}}{\text{molecule}} = 5.5 \times 10^{22}\ \text{atoms of chlorine}$$

e. $$1.0\ \text{g Cl} \times \frac{1\ \text{mol Cl}}{35.45\ \text{g}} \times \frac{1\ \text{mol}\ C_2H_3Cl_3O_2}{3\ \text{mol Cl}} \times \frac{165.39\ \text{g}\ C_2H_3Cl_3O_2}{\text{mol}\ C_2H_3Cl_3O_2} = 1.6\ \text{g chloral hydrate}$$

f. $500 \text{ molecules} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{165.39 \text{ g}}{\text{mol}} = 1.373 \times 10^{-19} \text{ g}$

38. $1.0 \text{ lb flour} \times \frac{454 \text{ g flour}}{\text{lb flour}} \times \frac{30.0 \times 10^{-9} \text{ g } C_2H_4Br_2}{\text{g flour}} \times \frac{1 \text{ mol } C_2H_4Br_2}{187.9 \text{ g } C_2H_4Br_2}$

$$\times \frac{6.02 \times 10^{23} \text{ molecules}}{\text{mol } C_2H_4Br_2} = 4.4 \times 10^{16} \text{ molecules } C_2H_4Br_2$$

39. a. $20.0 \text{ mg } C_8H_{10}N_4O_2 \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol}}{194.20 \text{ g}} = 1.03 \times 10^{-4} \text{ mol } C_8H_{10}N_4O_2$

b. $2.72 \times 10^{21} \text{ molecules } C_2H_5OH \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} = 4.52 \times 10^{-3} \text{ mol } C_2H_5OH$

c. $1.50 \text{ g } CO_2 \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 3.41 \times 10^{-2} \text{ mol } CO_2$

40. a. A chemical formula gives atom ratios as well as mole ratios. We will use both ratios to illustrate how these conversion factors can be used.

Molar mass of $C_2H_5O_2N$ = 2(12.01) + 5(1.008) + 2(16.00) + 14.0l = 75.07 g/mol

$$5.00 \text{ g } C_2H_5O_2N \times \frac{1 \text{ mol } C_2H_5O_2N}{75.07 \text{ g } C_2H_5O_2N} \times \frac{6.022 \times 10^{23} \text{ molecules } C_2H_5O_2N}{\text{mol } C_2H_5O_2N}$$

$$\times \frac{1 \text{ atom N}}{\text{molecule } C_2H_5O_2N} = 4.01 \times 10^{22} \text{ atoms N}$$

b. Molar mass of Mg_3N_2 = 3(24.31) + 2(14.01) = 100.95 g/mol

$$5.00 \text{ g } Mg_3N_2 \times \frac{1 \text{ mol } Mg_3N_2}{100.95 \text{ g } Mg_3N_2} \times \frac{6.022 \times 10^{23} \text{ formula units } Mg_3N_2}{\text{mol } Mg_3N_2}$$

$$\times \frac{2 \text{ atoms}}{\text{mol } Mg_3N_2} = 5.97 \times 10^{22} \text{ atoms N}$$

c. Molar mass of $Ca(NO_3)_2$ = 40.08 + 2(14.01) + 6(16.00) = 164.10 g/mol

$$5.00 \text{ g } Ca(NO_3)_2 \times \frac{1 \text{ mol } Ca(NO_3)_2}{164.10 \text{ g } Ca(NO_3)_2} \times \frac{2 \text{ mol N}}{\text{mol } Ca(NO_3)_2}$$

$$\times \frac{6.022 \times 10^{23} \text{ atoms N}}{\text{mol N}} = 3.67 \times 10^{22} \text{ atoms N}$$

d. Molar mass of N_2O_4 = 2(14.01) + 4(16.00) = 92.02 g/mol

$$5.00\text{ g }N_2O_4 \times \frac{1\text{ mol }N_2O_4}{92.02\text{ g }N_2O_4} \times \frac{2\text{ mol N}}{\text{mol }N_2O_4} \times \frac{6.022 \times 10^{23}\text{ atoms N}}{\text{mol N}}$$

$$= 6.54 \times 10^{22}\text{ atoms N}$$

Percent Composition

41. NO: $\text{Mass \% N} = \frac{14.01\text{ g N}}{30.01\text{ g NO}} \times 100 = 46.68\%\text{ N}$

NO_2: $\text{Mass \% N} = \frac{14.01\text{ g N}}{46.01\text{ g }NO_2} \times 100 = 30.45\%\text{ N}$

N_2O: $\text{Mass \% N} = \frac{2(14.01)\text{ g N}}{44.02\text{ g }N_2O} \times 100 = 63.65\%\text{ N}$

From the calculated mass percents, only NO is 46.7% N by mass, so NO could be this species. Any other compound having NO as an empirical formula could also be the compound.

42. a. $C_8H_{10}N_4O_2$: Molar mass = 8(12.01) + 10(1.008) + 4(14.0l) + 2(16.00) = 194.20 g/mol

$$\text{Mass \% C} = \frac{8(12.01)\text{ g C}}{194.20\text{ g }C_8H_{10}N_4O_2} \times 100 = \frac{96.08}{194.20} \times 100 = 49.47\%\text{ C}$$

b. $C_{12}H_{22}O_{11}$: Molar mass = 12(12.01) + 22(1.008) + 11(16.00) = 342.30 g/mol

$$\text{Mass \% C} = \frac{12(12.01)\text{ g C}}{342.30\text{ g }C_{12}H_{22}O_{11}} \times 100 = 42.10\%\text{ C}$$

c. C_2H_5OH: Molar mass = 2(12.01) + 6(1.008) + 1(16.00) = 46.07 g/mol

$$\text{Mass \% C} = \frac{2(12.01)\text{ g C}}{46.07\text{ g }C_2H_5OH} \times 100 = 52.14\%\text{ C}$$

The order from lowest to highest mass percentage of carbon is:

sucrose ($C_{12}H_{22}O_{11}$) < caffeine ($C_8H_{10}N_4O_2$) < ethanol (C_2H_5OH)

43. Molar mass = 20(12.01) + 29(1.008) + 19.00 + 3(16.00) = 336.43 g/mol

$$\text{Mass \% C} = \frac{20(12.01)\text{ g C}}{336.43\text{ g compound}} \times 100 = 71.40\%\text{ C}$$

$$\text{Mass \% H} = \frac{29(1.008)\text{ g H}}{336.43\text{ g compound}} \times 100 = 8.689\%\text{ H}$$

$$\text{Mass \% F} = \frac{19.00\text{ g F}}{336.43\text{ g compound}} \times 100 = 5.648\%\text{ F}$$

Mass % O = 100.00 – (71.40 + 8.689 + 5.648) = 14.26% O or:

$$\text{Mass \% O} = \frac{3(16.00)\text{ g O}}{336.43\text{ g compound}} \times 100 = 14.27\%\text{ O}$$

44. a. $C_3H_4O_2$: Molar mass = 3(12.011) + 4(1.0079) + 2(15.999) = 36.033 + 4.0316 + 31.998 = 72.063 g/mol

$$\text{Mass \% C} = \frac{36.033\text{g C}}{72.063\text{g compound}} \times 100 = 50.002\%\text{ C}$$

$$\text{Mass \% H} = \frac{4.0316\text{g H}}{72.063\text{g compound}} \times 100 = 5.5945\%\text{ H}$$

Mass % O = 100.000 – (50.002 + 5.5945) = 44.404% O or:

$$\%\text{ O} = \frac{31.998\text{g}}{72.063\text{g}} \times 100 = 44.403\%\text{ O}$$

b. $C_4H_6O_2$: Molar mass = 4(12.011) + 6(1.0079) + 2(15.999) = 48.044 + 6.0474 + 31.998 = 86.089 g/mol

$$\text{Mass \% C} = \frac{48.044\text{g}}{86.089\text{g}} \times 100 = 55.807\%\text{ C};\quad \text{mass \% H} = \frac{6.0474\text{g}}{86.089\text{g}} \times 100 = 7.0246\%\text{ H}$$

Mass % O = 100.000 – (55.807 + 7.0246) = 37.168% O

c. C_3H_3N: Molar mass = 3(12.011) + 3(1.0079) + 1(14.007) = 36.033 + 3.0237 + 14.007 = 53.064 g/mol

$$\text{Mass \% C} = \frac{36.033\text{g}}{53.064\text{g}} \times 100 = 67.905\%\text{ C};\quad \text{mass \% H} = \frac{3.0237\text{g}}{53.064\text{g}} \times 100 = 5.6982\%\text{ H}$$

$$\text{Mass \% N} = \frac{14.007\text{g}}{53.064\text{g}} \times 100 = 26.396\%\text{ N}$$ or % N = 100.000 – (67.905 + 5.6982) = 26.397% N

45. In 1 mole of $YBa_2Cu_3O_7$, there are 1 mole of Y, 2 moles of Ba, 3 moles of Cu, and 7 moles of O.

$$\text{Molar mass} = 1\text{ mol Y}\left(\frac{88.91\text{ g Y}}{\text{mol Y}}\right) + 2\text{ mol Ba}\left(\frac{137.3\text{ g Ba}}{\text{mol Ba}}\right) + 3\text{ mol Cu}\left(\frac{63.55\text{ g Cu}}{\text{mol Cu}}\right) + 7\text{ mol O}\left(\frac{16.00\text{ g O}}{\text{mol O}}\right)$$

Molar mass = 88.91 + 274.6 + 190.65 + 112.00 = 666.2 g/mol

$$\text{Mass \% Y} = \frac{88.91\text{ g}}{666.2\text{ g}} \times 100 = 13.35\%\text{ Y};\ \ \text{mass \% Ba} = \frac{274.6\text{ g}}{666.2\text{ g}} \times 100 = 41.22\%\text{ Ba}$$

$$\text{Mass \% Cu} = \frac{190.65\text{ g}}{666.2\text{ g}} \times 100 = 28.62\%\text{ Cu};\ \ \text{mass \% O} = \frac{112.0\text{ g}}{666.2\text{ g}} \times 100 = 16.81\%\text{ O}$$

46. If we have 100.0 g of Portland cement, we have 50. g Ca_3SiO_5, 25 g Ca_2SiO_4, 12 g $Ca_3Al_2O_6$, 8.0 g Ca_2AlFeO_5, and 3.5 g $CaSO_4 \bullet 2H_2O$.

Mass percent Ca:

$$50.\text{ g } Ca_3SiO_5 \times \frac{1\text{ mol } Ca_3SiO_5}{228.33\text{ g } Ca_3SiO_5} \times \frac{3\text{ mol Ca}}{1\text{ mol } Ca_3SiO_5} \times \frac{40.08\text{ g Ca}}{1\text{ mol Ca}} = 26\text{ g Ca}$$

$$25\text{ g } Ca_2SiO_4 \times \frac{80.16\text{ g Ca}}{172.25\text{ g } Ca_2SiO_4} = 12\text{ g Ca}$$

$$12\text{ g } Ca_3Al_2O_6 \times \frac{120.24\text{ g Ca}}{270.20\text{ g } Ca_3Al_2O_6} = 5.3\text{ g Ca}$$

$$8.0\text{ g } Ca_2AlFeO_5 \times \frac{80.16\text{ g Ca}}{242.99\text{ g } Ca_2AlFeO_5} = 2.6\text{ g Ca}$$

$$3.5\text{ g } CaSO_4 \bullet 2H_2O \times \frac{40.08\text{ g Ca}}{172.18\text{ g } CaSO_4 \bullet 2\,H_2O} = 0.81\text{ g Ca}$$

Mass of Ca = 26 + 12 + 5.3 + 2.6 + 0.81 = 47 g Ca

$$\%\text{ Ca} = \frac{47\text{ g Ca}}{100.0\text{ g cement}} \times 100 = 47\%\text{ Ca}$$

Mass percent Al:

$$12\text{ g } Ca_3Al_2O_6 \times \frac{53.96\text{ g Al}}{270.20\text{ g } Ca_3Al_2O_6} = 2.4\text{ g Al}$$

$$8.0 \text{ g } Ca_2AlFeO_5 \times \frac{26.98 \text{ g Al}}{242.99 \text{ g } Ca_2AlFeO_5} = 0.89 \text{ g Al}$$

$$\% \text{ Al} = \frac{2.4 \text{ g} + 0.89 \text{ g}}{100.0 \text{ g}} \times 100 = 3.3\% \text{ Al}$$

Mass percent Fe:

$$8.0 \text{ g } Ca_2AlFeO_5 \times \frac{55.85 \text{ g Fe}}{242.99 \text{ g } Ca_2AlFeO_5} = 1.8 \text{ g Fe}; \ \% \text{ Fe} = \frac{1.8 \text{ g}}{100.0 \text{ g}} \times 100 = 1.8\% \text{ Fe}$$

47. There are 0.390 g Cu for every 100.000 g of fungal laccase. Let's assume 100.000 g fungal laccase.

$$\text{Mol fungal laccase} = 0.390 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{1 \text{ mol fungal laccase}}{4 \text{ mol Cu}} = 1.53 \times 10^{-3} \text{ mol}$$

$$\frac{x \text{ g fungal laccase}}{1 \text{ mol fungal laccase}} = \frac{100.000 \text{ g}}{1.53 \times 10^{-3} \text{ mol}}, \ x = \text{molar mass} = 6.54 \times 10^4 \text{ g/mol}$$

48. Assuming 100.00 g cyanocobalamin:

$$\text{mol cyanocobalamin} = 4.34 \text{ g Co} \times \frac{1 \text{ mol Co}}{58.93 \text{ g Co}} \times \frac{1 \text{ mol cyanocobalamin}}{\text{mol Co}}$$

$$= 7.36 \times 10^{-2} \text{ mol cyanocobalamin}$$

$$\frac{x \text{ g cyanocobalamin}}{1 \text{ mol cyanocobalamin}} = \frac{100.00 \text{ g}}{7.36 \times 10^{-2} \text{ mol}}, \ x = \text{molar mass} = 1360 \text{ g/mol}$$

Empirical and Molecular Formulas

49. a. SNH: Empirical formula mass = 32.07 + 14.01 + 1.008 = 47.09 g/mol

$$\frac{188.35 \text{ g}}{47.09 \text{ g}} = 4.000;$$ so the molecular formula is $(SNH)_4$ or $S_4N_4H_4$.

b. $NPCl_2$: Empirical formula mass = 14.01 + 30.97 + 2(35.45) = 115.88 g/mol

$$\frac{347.64 \text{ g}}{115.88 \text{ g}} = 3.0000;$$ molecular formula is $(NPCl_2)_3$ or $N_3P_3Cl_6$.

c. CoC_4O_4: 58.93 + 4(12.01) + 4(16.00) = 170.97 g/mol

$$\frac{341.94 \text{ g}}{170.97 \text{ g}} = 2.0000;$$ molecular formula: $Co_2C_8O_8$

d. SN: 32.07 + 14.01 = 46.08 g/mol; $\frac{184.32\text{ g}}{46.08\text{ g}}$ = 4.000; molecular formula: S_4N_4

50. a. The molecular formula is N_2O_4. The smallest whole number ratio of the atoms (the empirical formula) is NO_2.

b. Molecular formula: C_3H_6; empirical formula: CH_2

c. Molecular formula: P_4O_{10}; empirical formula: P_2O_5

d. Molecular formula: $C_6H_{12}O_6$; empirical formula: CH_2O

51. a. Molar mass of CH_2O = 1 mol C $\left(\frac{12.011\text{ g}}{\text{mol C}}\right)$ + 2 mol H $\left(\frac{1.0079\text{ g}}{\text{mol H}}\right)$

+ 1 mol O $\left(\frac{15.999\text{ g}}{\text{mol O}}\right)$ = 30.026 g/mol

% C = $\frac{12.011\text{ g C}}{30.026\text{ g CH}_2\text{O}}$ × 100 = 40.002% C; % H = $\frac{2.0158\text{ g H}}{30.026\text{ g CH}_2\text{O}}$ × 100 = 6.7135% H

% O = $\frac{15.999\text{ g O}}{30.026\text{ g CH}_2\text{O}}$ × 100 = 53.284% O or % O = 100.000 – (40.002 + 6.7135)

= 53.285%

b. Molar mass of $C_6H_{12}O_6$ = 6(12.011) + 12(1.0079) + 6(15.999) = 180.155 g/mol

% C = $\frac{72.066\text{ g C}}{180.155\text{ g C}_6\text{H}_{12}\text{O}_6}$ × 100 = 40.002%; % H = $\frac{12(1.0079)\text{ g}}{180.155\text{ g}}$ × 100 = 6.7136%

% O = 100.00 – (40.002 + 6.7136) = 53.284%

c. Molar Mass of $HC_2H_3O_2$ = 2(12.011) + 4(1.0079) + 2(15.999) = 60.052 g/mol

% C = $\frac{24.022\text{ g}}{60.052\text{ g}}$ × 100 = 40.002%; % H = $\frac{4.0316\text{ g}}{60.052\text{ g}}$ × 100 = 6.7135%

% O = 100.000 – (40.002 + 6.7135) = 53.285%

All three compounds have the same empirical formula, CH_2O, and different molecular formulas. The composition of all three in mass percent is also the same (within rounding differences). Therefore, elemental analysis will give us only the empirical formula.

52. Assuming 100.00 g of compound (mass oxygen = 100.00 g – 41.39 g C – 3.47 g H = 55.14 g O):

$$41.39 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.446 \text{ mol C}; \quad 3.47 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 3.44 \text{ mol H}$$

$$55.14 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 3.446 \text{ mol O}$$

All are the same mole values, so the empirical formula is CHO. The empirical formula mass is 12.01 + 1.008 + 16.00 = 29.02 g/mol.

$$\text{Molar mass} = \frac{15.0 \text{ g}}{0.129 \text{ mol}} = 116 \text{ g/mol}$$

$$\frac{\text{Molar mass}}{\text{Empirical mass}} = \frac{116}{29.02} = 4.00; \quad \text{molecular formula} = (CHO)_4 = C_4H_4O_4$$

53. Assuming 100.0 g of compound:

$$26.7 \text{ g P} \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 0.862 \text{ mol P}; \quad 12.1 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.864 \text{ mol N}$$

$$61.2 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 1.73 \text{ mol Cl}$$

$\frac{1.73}{0.862} = 2.01$; the empirical formula is $PNCl_2$.

The empirical formula mass is ≈ 31.0 + 14.0 + 2(35.5) = 116 g/mol.

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{580}{116} = 5.0; \text{ the molecular formula is } (PNCl_2)_5 = P_5N_5Cl_{10}.$$

54. Out of 100.00 g of adrenaline, there are:

$$56.79 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 4.728 \text{ mol C}; \quad 6.56 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.51 \text{ mol H}$$

$$28.37 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 1.773 \text{ mol O}; \quad 8.28 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.591 \text{ mol N}$$

Dividing each mole value by the smallest number:

$$\frac{4.728}{0.591} = 8.00; \quad \frac{6.51}{0.591} = 11.0; \quad \frac{1.773}{0.591} = 3.00; \quad \frac{0.591}{0.591} = 1.00$$

This gives adrenaline an empirical formula of $C_8H_{11}O_3N$.

55. Compound I: mass O = 0.6498 g Hg_xO_y – 0.6018 g Hg = 0.0480 g O

$$0.6018 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} = 3.000 \times 10^{-3} \text{ mol Hg}$$

$$0.0480 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.00 \times 10^{-3} \text{ mol O}$$

The mole ratio between Hg and O is 1 : 1, so the empirical formula of compound I is HgO.
Compound II: mass Hg = 0.4172 g Hg_xO_y – 0.016 g O = 0.401 g Hg

$$0.401 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} = 2.00 \times 10^{-3} \text{ mol Hg}; \quad 0.016 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.0 \times 10^{-3} \text{ mol O}$$

The mole ratio between Hg and O is 2 : 1, so the empirical formula is Hg_2O.

56. $$1.121 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 8.003 \times 10^{-2} \text{ mol N}; \; 0.161 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.60 \times 10^{-1} \text{ mol H}$$

$$0.480 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.00 \times 10^{-2} \text{ mol C}; \quad 0.640 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 4.00 \times 10^{-2} \text{ mol O}$$

Dividing all mole values by the smallest number:

$$\frac{8.003 \times 10^{-2}}{4.00 \times 10^{-2}} = 2.00; \quad \frac{1.60 \times 10^{-1}}{4.00 \times 10^{-2}} = 4.00; \quad \frac{4.00 \times 10^{-2}}{4.00 \times 10^{-2}} = 1.00$$

The empirical formula is N_2H_4CO.

57. First, we will determine composition in mass percent. We assume that all the carbon in the 0.213 g CO_2 came from the 0.157 g of the compound and that all the hydrogen in the 0.0310 g H_2O came from the 0.157 g of the compound.

$$0.213 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 0.0581 \text{ g C}; \quad \% \text{ C} = \frac{0.0581 \text{ g C}}{0.157 \text{ g compound}} \times 100 = 37.0\% \text{ C}$$

$$0.0310 \text{ g } H_2O \times \frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O} = 3.47 \times 10^{-3} \text{ g H}; \quad \% \text{ H} = \frac{3.47 \times 10^{-3} \text{ g}}{0.157 \text{ g}} \times 100 = 2.21\% \text{ H}$$

We get the mass percent of N from the second experiment:

$$0.0230 \text{ g } NH_3 \times \frac{14.01 \text{ g N}}{17.03 \text{ g } NH_3} = 1.89 \times 10^{-2} \text{ g N}$$

$$\% \text{ N} = \frac{1.89 \times 10^{-2} \text{ g}}{0.103 \text{ g}} \times 100 = 18.3\% \text{ N}$$

The mass percent of oxygen is obtained by difference:

$$\% \text{ O} = 100.00 - (37.0 + 2.21 + 18.3) = 42.5\% \text{ O}$$

So, out of 100.00 g of compound, there are:

$$37.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.08 \text{ mol C}; \quad 2.21 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 2.19 \text{ mol H}$$

$$18.3 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.31 \text{ mol N}; \quad 42.5 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.66 \text{ mol O}$$

Lastly, and often the hardest part, we need to find simple whole number ratios. Divide all mole values by the smallest number:

$$\frac{3.08}{1.31} = 2.35; \quad \frac{2.19}{1.31} = 1.67; \quad \frac{1.31}{1.31} = 1.00; \quad \frac{2.66}{1.31} = 2.03$$

Multiplying all these ratios by 3 gives an empirical formula of $C_7H_5N_3O_6$.

58. a. Only acrylonitrile contains nitrogen. If we have 100.00 g of polymer:

$$8.80 \text{ g N} \times \frac{1 \text{ mol } C_3H_3N}{14.01 \text{ g N}} \times \frac{53.06 \text{ g } C_3H_3N}{1 \text{ mol } C_3H_3N} = 33.3 \text{ g } C_3H_3N$$

$$\% \ C_3H_3N = \frac{33.3 \text{ g } C_3H_3N}{100.00 \text{ g polymer}} = 33.3\% \ C_3H_3N$$

Only butadiene in the polymer reacts with Br_2:

$$0.605 \text{ g } Br_2 \times \frac{1 \text{ mol } Br_2}{159.8 \text{ g } Br_2} \times \frac{1 \text{ mol } C_4H_6}{\text{mol } Br_2} \times \frac{54.09 \text{ g } C_4H_6}{\text{mol } C_4H_6} = 0.205 \text{ g } C_4H_6$$

$$\% \ C_4H_6 = \frac{0.205 \text{ g}}{1.20 \text{ g}} \times 100 = 17.1\% \ C_4H_6$$

b. If we have 100.0 g of polymer:

$$33.3 \text{ g } C_3H_3N \times \frac{1 \text{ mol } C_3H_3N}{53.06 \text{ g}} = 0.628 \text{ mol } C_3H_3N$$

$$17.1 \text{ g } C_4H_6 \times \frac{1 \text{ mol } C_4H_6}{54.09 \text{ g } C_4H_6} = 0.316 \text{ mol } C_4H_6$$

$$49.6 \text{ g } C_8H_8 \times \frac{1 \text{ mol } C_8H_8}{104.14 \text{ g } C_8H_8} = 0.476 \text{ mol } C_8H_8$$

Dividing by 0.316: $\frac{0.628}{0.316} = 1.99$; $\frac{0.316}{0.316} = 1.00$; $\frac{0.476}{0.316} = 1.51$

This is close to a mole ratio of 4 : 2 : 3. Thus there are 4 acrylonitrile to 2 butadiene to 3 styrene molecules in the polymer, or $(A_4B_2S_3)_n$.

59. First, we will determine composition by mass percent:

$$16.01 \text{ mg } CO_2 \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{12.011 \text{ g C}}{44.009 \text{ g } CO_2} \times \frac{1000 \text{ mg}}{\text{g}} = 4.369 \text{ mg C}$$

$$\% \text{ C} = \frac{4.369 \text{ mg C}}{10.68 \text{ mg compound}} \times 100 = 40.91\% \text{ C}$$

$$4.37 \text{ mg } H_2O \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O} \times \frac{1000 \text{ mg}}{\text{g}} = 0.489 \text{ mg H}$$

$$\% \text{ H} = \frac{0.489 \text{ mg}}{10.68 \text{ mg}} \times 100 = 4.58\% \text{ H}; \quad \% \text{ O} = 100.00 - (40.91 + 4.58) = 54.51\% \text{ O}$$

So, in 100.00 g of the compound, we have:

$$40.91 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.406 \text{ mol C}; \quad 4.58 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.54 \text{ mol H}$$

$$54.51 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 3.407 \text{ mol O}$$

Dividing by the smallest number: $\frac{4.54}{3.406} = 1.33 \approx \frac{4}{3}$; the empirical formula is $C_3H_4O_3$.

The empirical formula mass of $C_3H_4O_3$ is ≈ 3(12) + 4(1) + 3(16) = 88 g.

Because $\frac{176.1}{88} = 2.0$, the molecular formula is $C_6H_8O_6$.

60. $41.98 \text{ mg } CO_2 \times \frac{12.011 \text{ mg C}}{44.009 \text{ mg } CO_2} = 11.46 \text{ mg C}$; $\% \text{ C} = \frac{11.46 \text{ mg}}{19.81 \text{ mg}} \times 100 = 57.85\% \text{ C}$

$6.45 \text{ mg } H_2O \times \frac{2.016 \text{ mg H}}{18.02 \text{ mg } H_2O} = 0.722 \text{ mg H}$; $\% \text{ H} = \frac{0.722 \text{ mg}}{19.81 \text{ mg}} \times 100 = 3.64\% \text{ H}$

% O = 100.00 – (57.85 + 3.64) = 38.51% O

Out of 100.00 g terephthalic acid, there are:

$$57.85 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 4.816 \text{ mol C}; \quad 3.64 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 3.61 \text{ mol H}$$

$$38.51 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 2.407 \text{ mol O}$$

$$\frac{4.816}{2.407} = 2.001; \quad \frac{3.61}{2.407} = 1.50; \quad \frac{2.407}{2.407} = 1.000$$

The C : H : O mole ratio is 2 : 1.5 : 1 or 4 : 3 : 2. The empirical formula is $C_4H_3O_2$. Mass of $C_4H_3O_2 \approx 4(12) + 3(1) + 2(16) = 83$.

Molar mass = $\frac{41.5 \text{ g}}{0.250 \text{ mol}}$ = 166 g/mol; $\frac{166}{83}$ = 2.0; the molecular formula is $C_8H_6O_4$.

Balancing Chemical Equations

61. a. $16\ Cr(s) + 3\ S_8(s) \rightarrow 8\ Cr_2S_3(s)$

    b. $2\ NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$

    c. $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$

    d. $2\ Eu(s) + 6\ HF(g) \rightarrow 2\ EuF_3(s) + 3\ H_2(g)$

    e. $2\ C_6H_6(l) + 15\ O_2(g) \rightarrow 12\ CO_2(g) + 6\ H_2O(g)$

62. An important part to this problem is writing out correct formulas. If the formulas are incorrect, then the balanced reaction is incorrect.

    a. $C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$

    b. $3\ Pb(NO_3)_2(aq) + 2\ Na_3PO_4(aq) \rightarrow Pb_3(PO_4)_2(s) + 6\ NaNO_3(aq)$

63. Only one product is formed in this representation. This product has two Ys bonded to an X. The other substance present in the product mixture is just the excess of one of the reactants (Y). The best equation has smallest whole numbers. Here, answer c would be this smallest whole number equation ($X + 2\ Y \rightarrow XY_2$). Answers a and b have incorrect products listed, and for answer d, an equation only includes the reactants that go to produce the product; excess reactants are not shown in an equation.

64. a. $2\ KO_2(s) + 2\ H_2O(l) \rightarrow 2\ KOH(aq) + O_2(g) + H_2O_2(aq)$ or

$4\ KO_2(s) + 6\ H_2O(l) \rightarrow 4\ KOH(aq) + O_2(g) + 4\ H_2O_2(aq)$

b. $Fe_2O_3(s) + 6\ HNO_3(aq) \rightarrow 2\ Fe(NO_3)_3(aq) + 3\ H_2O(l)$

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

d. $PCl_5(l) + 4\ H_2O(l) \rightarrow H_3PO_4(aq) + 5\ HCl(g)$

e. $2\ CaO(s) + 5\ C(s) \rightarrow 2\ CaC_2(s) + CO_2(g)$

f. $2\ MoS_2(s) + 7\ O_2(g) \rightarrow 2\ MoO_3(s) + 4\ SO_2(g)$

g. $FeCO_3(s) + H_2CO_3(aq) \rightarrow Fe(HCO_3)_2(aq)$

65. When balancing reactions, start with elements that appear in only one of the reactants and one of the products, and then go on to balance the remaining elements.

a. $C_6H_{12}O_6(s) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Balance C atoms: $C_6H_{12}O_6 + O_2 \rightarrow 6\ CO_2 + H_2O$

Balance H atoms: $C_6H_{12}O_6 + O_2 \rightarrow 6\ CO_2 + 6\ H_2O$

Lastly, balance O atoms: $C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$

b. $Fe_2S_3(s) + HCl(g) \rightarrow FeCl_3(s) + H_2S(g)$

Balance Fe atoms: $Fe_2S_3 + HCl \rightarrow 2\ FeCl_3 + H_2S$

Balance S atoms: $Fe_2S_3 + HCl \rightarrow 2\ FeCl_3 + 3\ H_2S$

There are 6 H and 6 Cl on right, so balance with 6 HCl on left:

$Fe_2S_3(s) + 6\ HCl(g) \rightarrow 2\ FeCl_3(s) + 3\ H_2S(g)$

c. $CS_2(l) + NH_3(g) \rightarrow H_2S(g) + NH_4SCN(s)$

C and S are balanced; balance N:

$CS_2 + 2\ NH_3 \rightarrow H_2S + NH_4SCN$

H is also balanced. $CS_2(l) + 2\ NH_3(g) \rightarrow H_2S(g) + NH_4SCN(s)$.

66. a. $SiO_2(s) + C(s) \rightarrow Si(s) + CO(g)$; Si is balanced.

Balance oxygen atoms: $SiO_2 + C \rightarrow Si + 2\ CO$

Balance carbon atoms: $SiO_2(s) + 2\ C(s) \rightarrow Si(s) + 2\ CO(g)$

b. $SiCl_4(l) + Mg(s) \rightarrow Si(s) + MgCl_2(s)$; Si is balanced.

Balance Cl atoms: $SiCl_4 + Mg \rightarrow Si + 2\ MgCl_2$

Balance Mg atoms: $SiCl_4(l) + 2\ Mg(s) \rightarrow Si(s) + 2\ MgCl_2(s)$

c. $Na_2SiF_6(s) + Na(s) \rightarrow Si(s) + NaF(s)$; Si is balanced.

Balance F atoms: $Na_2SiF_6 + Na \rightarrow Si + 6\ NaF$

Balance Na atoms: $Na_2SiF_6(s) + 4\ Na(s) \rightarrow Si(s) + 6\ NaF(s)$

Reaction Stoichiometry

67. $2\ LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(aq) + H_2O(l)$

The total volume of air exhaled each minute for the 7 astronauts is 7 × 20. = 140 L/min.

$$25{,}000\text{ g LiOH} \times \frac{1\text{ mol LiOH}}{23.95\text{ g LiOH}} \times \frac{1\text{ mol }CO_2}{2\text{ mol LiOH}} \times \frac{44.01\text{ g }CO_2}{\text{mol }CO_2} \times \frac{100\text{ g air}}{4.0\text{ g }CO_2} \times$$

$$\frac{1\text{ mL air}}{0.0010\text{ g air}} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{1\text{ min}}{140\text{ L air}} \times \frac{1\text{ h}}{60\text{ min}} = 68\text{ h} = 2.8\text{ days}$$

68. $$1.0 \times 10^4\text{ kg waste} \times \frac{3.0\text{ kg }NH_4^+}{100\text{ kg waste}} \times \frac{1000\text{ g}}{\text{kg}} \times \frac{1\text{ mol }NH_4^+}{18.04\text{ g }NH_4^+} \times \frac{1\text{ mol }C_5H_7O_2N}{55\text{ mol }NH_4^+} \times$$

$$\frac{113.1\text{ g }C_5H_7O_2N}{\text{mol }C_5H_7O_2N} = 3.4 \times 10^4\text{ g tissue if all }NH_4^+\text{ converted}$$

Because only 95% of the NH_4^+ ions react:

mass of tissue = $(0.95)(3.4 \times 10^4\text{ g}) = 3.2 \times 10^4$ g or 32 kg bacterial tissue

69. $$1.0 \times 10^3\text{ g phosphorite} \times \frac{75\text{ g }Ca_3(PO_4)_2}{100\text{ g phosphorite}} \times \frac{1\text{ mol }Ca_3(PO_4)_2}{310.18\text{ g }Ca_3(PO_4)_2} \times$$

$$\frac{1\text{ mol }P_4}{2\text{ mol }Ca_3(PO_4)_2} \times \frac{123.88\text{ g }P_4}{\text{mol }P_4} = 150\text{ g }P_4$$

70. Total mass of copper used:

$$10{,}000\text{ boards} \times \frac{(8.0\text{ cm} \times 16.0\text{ cm} \times 0.060\text{ cm})}{\text{board}} \times \frac{8.96\text{ g}}{\text{cm}^3} = 6.9 \times 10^5\text{ g Cu}$$

Amount of Cu to be recovered = $0.80 \times (6.9 \times 10^5\text{ g}) = 5.5 \times 10^5$ g Cu

$$5.5 \times 10^5\text{ g Cu} \times \frac{1\text{ mol Cu}}{63.55\text{ g Cu}} \times \frac{1\text{ mol }Cu(NH_3)_4Cl_2}{\text{mol Cu}} \times \frac{202.6\text{ g }Cu(NH_3)_4Cl_2}{\text{mol }Cu(NH_3)_4Cl_2}$$

$$= 1.8 \times 10^6\text{ g }Cu(NH_3)_4Cl_2$$

$$5.5 \times 10^5 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{4 \text{ mol NH}_3}{\text{mol Cu}} \times \frac{17.03 \text{ g NH}_3}{\text{mol NH}_3} = 5.9 \times 10^5 \text{ g NH}_3$$

71. $$1.000 \text{ kg Al} \times \frac{1000 \text{ g Al}}{\text{kg Al}} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol NH}_4\text{ClO}_4}{3 \text{ mol Al}} \times \frac{117.49 \text{ g NH}_4\text{ClO}_4}{\text{mol NH}_4\text{ClO}_4} = 4355 \text{ g NH}_4\text{ClO}_4$$

72. $10\ KClO_3(s) + 3\ P_4(s) \rightarrow 3\ P_4O_{10}(s) + 10\ KCl(s)$

$$52.9 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{3 \text{ mol P}_4\text{O}_{10}}{10 \text{ mol KClO}_3} \times \frac{283.88 \text{ g P}_4\text{O}_{10}}{\text{mol P}_4\text{O}_{10}} = 36.8 \text{ g P}_4\text{O}_{10}$$

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

$$15.0 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 0.269 \text{ mol Fe};\ \ 0.269 \text{ mol Fe} \times \frac{2 \text{ mol Al}}{2 \text{ mol Fe}} \times \frac{26.98 \text{ g Al}}{\text{mol Al}} = 7.26 \text{ g Al}$$

$$0.269 \text{ mol Fe} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{2 \text{ mol Fe}} \times \frac{159.70 \text{ g Fe}_2\text{O}_3}{\text{mol Fe}_2\text{O}_3} = 21.5 \text{ g Fe}_2\text{O}_3$$

$$0.269 \text{ mol Fe} \times \frac{1 \text{ mol Al}_2\text{O}_3}{2 \text{ mol Fe}} \times \frac{101.96 \text{ g Al}_2\text{O}_3}{\text{mol Al}_2\text{O}_3} = 13.7 \text{ g Al}_2\text{O}_3$$

74. $$1.0 \times 10^6 \text{ kg HNO}_3 \times \frac{1000 \text{ g HNO}_3}{\text{kg HNO}_3} \times \frac{1 \text{ mol HNO}_3}{63.0 \text{ g HNO}_3} = 1.6 \times 10^7 \text{ mol HNO}_3$$

We need to get the relationship between moles of HNO_3 and moles of NH_3. We have to use all three equations:

$$\frac{2 \text{ mol HNO}_3}{3 \text{ mol NO}_2} \times \frac{2 \text{ mol NO}_2}{2 \text{ mol NO}} \times \frac{4 \text{ mol NO}}{4 \text{ mol NH}_3} = \frac{16 \text{ mol HNO}_3}{24 \text{ mol NH}_3}$$

Thus we can produce 16 mol HNO_3 for every 24 mol NH_3 that we begin with:

$$1.6 \times 10^7 \text{ mol HNO}_3 \times \frac{24 \text{ mol NH}_3}{16 \text{ mol HNO}_3} \times \frac{17.0 \text{ g NH}_3}{\text{mol NH}_3} = 4.1 \times 10^8 \text{ g or } 4.1 \times 10^5 \text{ kg NH}_3$$

This is an oversimplified answer. In practice, the NO produced in the third step is recycled back continuously into the process in the second step. If this is taken into consideration, then the conversion factor between mol NH_3 and mol HNO_3 turns out to be 1 : 1; that is, 1 mol of NH_3 produces 1 mol of HNO_3. Taking into consideration that NO is recycled back gives an answer of 2.7×10^5 kg NH_3 reacted.

Limiting Reactants and Percent Yield

75. One method to solve limiting-reagent problems is to assume that each reactant is limiting and then calculate how much product could be produced from each reactant. The reactant that produces the smallest amount of product will run out first and is the limiting reagent.

$$5.00 \times 10^6 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{2 \text{ mol HCN}}{2 \text{ mol } NH_3} = 2.94 \times 10^5 \text{ mol HCN}$$

$$5.00 \times 10^6 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{2 \text{ mol HCN}}{3 \text{ mol } O_2} = 1.04 \times 10^5 \text{ mol HCN}$$

$$5.00 \times 10^6 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4} \times \frac{2 \text{ mol HCN}}{2 \text{ mol } CH_4} = 3.12 \times 10^5 \text{ mol HCN}$$

O_2 is limiting because it produces the smallest amount of HCN. Although more product could be produced from NH_3 and CH_4, only enough O_2 is present to produce 1.04×10^5 mol HCN. The mass of HCN that can be produced is:

$$1.04 \times 10^5 \text{ mol HCN} \times \frac{27.03 \text{ g HCN}}{\text{mol HCN}} = 2.81 \times 10^6 \text{ g HCN}$$

$$5.00 \times 10^6 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{6 \text{ mol } H_2O}{3 \text{ mol } O_2} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 5.63 \times 10^6 \text{ g } H_2O$$

76. $2\ C_3H_6(g) + 2\ NH_3(g) + 3\ O_2(g) \rightarrow 2\ C_3H_3N(g) + 6\ H_2O(g)$

a. We will solve this limiting reagent problem using the same method as described in Exercise 3.75.

$$1.00 \times 10^3 \text{ g } C_3H_6 \times \frac{1 \text{ mol } C_3H_6}{42.08 \text{ g } C_3H_6} \times \frac{2 \text{ mol } C_3H_3N}{2 \text{ mol } C_3H_6} = 23.8 \text{ mol } C_3H_3N$$

$$1.50 \times 10^3 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{2 \text{ mol } C_3H_3N}{2 \text{ mol } NH_3} = 88.1 \text{ mol } C_3H_3N$$

$$2.00 \times 10^3 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{2 \text{ mol } C_3H_3N}{3 \text{ mol } O_2} = 41.7 \text{ mol } C_3H_3N$$

C_3H_6 is limiting because it produces the smallest amount of product, and the mass of acrylonitrile that can be produced is:

$$23.8 \text{ mol} \times \frac{53.06 \text{ g } C_3H_3N}{\text{mol}} = 1.26 \times 10^3 \text{ g } C_3H_3N$$

b. $23.8 \text{ mol } C_3H_3N \times \frac{6 \text{ mol } H_2O}{2 \text{ mol } C_3H_3N} \times \frac{18.02 \text{ g } H_2O}{\text{mol } H_2O} = 1.29 \times 10^3 \text{ g } H_2O$

Amount NH_3 needed:

$$23.8 \text{ mol } C_3H_3N \times \frac{2 \text{ mol } NH_3}{2 \text{ mol } C_3H_3N} \times \frac{17.03 \text{ g } NH_3}{\text{mol } NH_3} = 405 \text{ g } NH_3$$

Amount NH_3 in excess = 1.50×10^3 g – 405 g = 1.10×10^3 g NH_3

Amount O_2 needed:

$$23.8 \text{ mol } C_3H_3N \times \frac{3 \text{ mol } O_2}{2 \text{ mol } C_3H_3N} \times \frac{32.00 \text{ g } O_2}{\text{mol } O_2} = 1.14 \times 10^3 \text{ g } O_2$$

Amount O_2 in excess = 2.00×10^3 g – 1.14×10^3 g = 860 g O_2

1.10×10^3 g NH_3 and 860 g O_2 are in excess.

77. The product formed in the reaction is NO_2; the other species present in the product picture is excess O_2. Therefore, NO is the limiting reactant. In the pictures, 6 NO molecules react with 3 O_2 molecules to form 6 NO_2 molecules.

$$6\ NO(g) + 3\ O_2(g) \rightarrow 6\ NO_2(g)$$

For smallest whole numbers, the balanced reaction is:

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

78. In the following table we have listed three rows of information. The “Initial” row is the number of molecules present initially, the “Change” row is the number of molecules that react to reach completion, and the “Final” row is the number of molecules present at completion. To determine the limiting reactant, let’s calculate how much of one reactant is necessary to react with the other.

$$10 \text{ molecules } O_2 \times \frac{4 \text{ molecules } NH_3}{5 \text{ molecules } O_2} = 8 \text{ molecules } NH_3 \text{ to react with all the } O_2$$

Because we have 10 molecules of NH_3 and only 8 molecules of NH_3 are necessary to react with all the O_2, O_2 is limiting.

	$4\ NH_3(g)$	+	$5\ O_2(g)$	$\rightarrow$	$4\ NO(g)$	+	$6\ H_2O(g)$
Initial	10 molecules		10 molecules		0		0
Change	-8 molecules		-10 molecules		+8 molecules		+12 molecules
Final	2 molecules		0		8 molecules		12 molecules

The total number of molecules present after completion = 2 molecules NH_3 + 0 molecules O_2 + 8 molecules NO + 12 molecules H_2O = 22 molecules.

79. $1.50 \text{ g BaO}_2 \times \dfrac{1 \text{ mol BaO}_2}{169.3 \text{ g BaO}_2} = 8.86 \times 10^{-3} \text{ mol BaO}_2$

$$25.0 \text{ mL} \times \frac{0.0272 \text{ g HCl}}{\text{mL}} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} = 1.87 \times 10^{-2} \text{ mol HCl}$$

The required mole ratio from the balanced reaction is 2 mol HCl to 1 mol BaO_2. The actual mole ratio is:

$$\frac{1.87 \times 10^{-2} \text{ mol HCl}}{8.86 \times 10^{-3} \text{ mol BaO}_2} = 2.11$$

Because the actual mole ratio is larger than the required mole ratio, the denominator (BaO_2) is the limiting reagent.

$$8.86 \times 10^{-3} \text{ mol BaO}_2 \times \frac{1 \text{ mol H}_2\text{O}_2}{\text{mol BaO}_2} \times \frac{34.02 \text{ g H}_2\text{O}_2}{\text{mol H}_2\text{O}_2} = 0.301 \text{ g H}_2\text{O}_2$$

The amount of HCl reacted is:

$$8.86 \times 10^{-3} \text{ mol BaO}_2 \times \frac{2 \text{ mol HCl}}{\text{mol BaO}_2} = 1.77 \times 10^{-2} \text{ mol HCl}$$

Excess mol HCl = 1.87×10^{-2} mol – 1.77×10^{-2} mol = 1.0×10^{-3} mol HCl

$$\text{Mass of excess HCl} = 1.0 \times 10^{-3} \text{ mol HCl} \times \frac{36.46 \text{ g HCl}}{\text{mol HCl}} = 3.6 \times 10^{-2} \text{ g HCl}$$

80. $25.0 \text{ g Ag}_2\text{O} \times \dfrac{1 \text{ mol}}{231.8 \text{ g}} = 0.108 \text{ mol Ag}_2\text{O}$

$$50.0 \text{ g C}_{10}\text{H}_{10}\text{N}_4\text{SO}_2 \times \frac{1 \text{ mol}}{250.29 \text{ g}} = 0.200 \text{ mol C}_{10}\text{H}_{10}\text{N}_4\text{SO}_2$$

$$\frac{\text{mol C}_{10}\text{H}_{10}\text{N}_4\text{SO}_2}{\text{mol Ag}_2\text{O}} \text{ (actual)} = \frac{0.200}{0.108} = 1.85$$

The actual mole ratio is less than the required mole ratio (2), so $C_{10}H_{10}N_4SO_2$ is limiting.

$$0.200 \text{ mol C}_{10}\text{H}_{10}\text{N}_4\text{SO}_2 \times \frac{2 \text{ mol AgC}_{10}\text{H}_9\text{N}_4\text{SO}_2}{2 \text{ mol C}_{10}\text{H}_{10}\text{N}_4\text{SO}_2} \times \frac{357.18\text{g}}{\text{mol AgC}_{10}\text{H}_9\text{N}_4\text{SO}_2}$$

$$= 71.4 \text{ g AgC}_{10}\text{H}_9\text{N}_4\text{SO}_2 \text{ produced}$$

81. $P_4(s) + 6\ F_2(g) \rightarrow 4\ PF_3(g)$; the theoretical yield of PF_3 is:

$$120.\ \text{g } PF_3 \text{ (actual)} \times \frac{100.0\ \text{g } PF_3 \text{ (theoretical)}}{78.1\ \text{g } PF_3 \text{ (actual)}} = 154\ \text{g } PF_3 \text{ (theoretical)}$$

$$154\ \text{g } PF_3 \times \frac{1\ \text{mol } PF_3}{87.97\ \text{g } PF_3} \times \frac{6\ \text{mol } F_2}{4\ \text{mol } PF_3} \times \frac{38.00\ \text{g } F_2}{\text{mol } F_2} = 99.8\ \text{g } F_2$$

99.8 g F_2 are needed to actually produce 120. g of PF_3 if the percent yield is 78.1%.

82. a. From the reaction stoichiometry we would expect to produce 4 mol of acetaminophen for every 4 mol of $C_6H_5O_3N$ reacted. The actual yield is 3 mol of acetaminophen compared with a theoretical yield of 4 mol of acetaminophen. Solving for percent yield by mass (where M = molar mass acetaminophen):

$$\text{percent yield} = \frac{3\ \text{mol} \times M}{4\ \text{mol} \times M} \times 100 = 75\%$$

b. The product of the percent yields of the individual steps must equal the overall yield, 75%.

(0.87)(0.98)(x) = 0.75, x = 0.88; step III has a percent yield of 88%.

83. $$2.50\ \text{metric tons } Cu_3FeS_3 \times \frac{1000\ \text{kg}}{\text{metric ton}} \times \frac{1000\ \text{g}}{\text{kg}} \times \frac{1\ \text{mol } Cu_3FeS_3}{342.71\ \text{g}} \times \frac{3\ \text{mol Cu}}{1\ \text{mol } Cu_3FeS_3} \times \frac{63.55\ \text{g}}{\text{mol Cu}} = 1.39 \times 10^6\ \text{g Cu (theoretical)}$$

$$1.39 \times 10^6\ \text{g Cu (theoretical)} \times \frac{86.3\ \text{g Cu (actual)}}{100.\ \text{g Cu (theoretical)}} = 1.20 \times 10^6\ \text{g Cu} = 1.20 \times 10^3\ \text{kg Cu}$$

$$= 1.20\ \text{metric tons Cu (actual)}$$

84. a. $$1142\ \text{g } C_6H_5Cl \times \frac{1\ \text{mol } C_6H_5Cl}{112.55\ \text{g } C_6H_5Cl} = 10.15\ \text{mol } C_6H_5Cl$$

$$485\ \text{g } C_2HOCl_3 \times \frac{1\ \text{mol } C_2HOCl_3}{147.38\ \text{g } C_2HOCl_3} = 3.29\ \text{mol } C_2HOCl_3$$

From the balanced equation, the required mole ratio is $\frac{2\ \text{mol } C_6H_5Cl}{1\ \text{mol } C_2HOCl_3} = 2$. The actual mole ratio present is $\frac{10.15\ \text{mol } C_6H_5Cl}{3.29\ \text{mol } C_2HOCl_3} = 3.09$. The actual mole ratio is greater than the required mole ratio, so the denominator of the actual mole ratio (C_2HOCl_3) is limiting.

$$3.29 \text{ mol } C_2HOCl_3 \times \frac{1 \text{ mol } C_{14}H_9Cl_5}{\text{mol } C_2HOCl_3} \times \frac{354.46 \text{ g } C_{14}H_9Cl_5}{\text{mol } C_{14}H_9Cl_5} = 1170 \text{ g } C_{14}H_9Cl_5 \text{ (DDT)}$$

b. C_2HOCl_3 is limiting, and C_6H_5Cl is in excess.

c. $$3.29 \text{ mol } C_2HOCl_3 \times \frac{2 \text{ mol } C_6H_5Cl}{\text{mol } C_2HOCl_3} \times \frac{112.55 \text{ g } C_6H_5Cl}{\text{mol } C_6H_5Cl} = 741 \text{ g } C_6H_5Cl \text{ reacted}$$

1142 g – 741 g = 401 g C_6H_5Cl in excess

d. $$\text{Percent yield} = \frac{200.0 \text{ g DDT}}{1170 \text{ g DDT}} \times 100 = 17.1\%$$

Additional Exercises

85. The volume of a gas is proportional to the number of molecules of gas. Thus the formulas are:

I: NH_3 II: N_2H_4 III: HN_3

The mass ratios are:

I: $\frac{4.634 \text{ g N}}{\text{g H}}$ II: $\frac{6.949 \text{ g N}}{\text{g H}}$ III: $\frac{41.7 \text{ g N}}{\text{g H}}$

If we set the atomic mass of H equal to 1.008, then the atomic mass, A, for nitrogen is:

I: 14.01 II: 14.01 III. 14.0

For example, for compound I: $\frac{A}{3(1.008)} = \frac{4.634}{1}$, A = 14.01

86. Assuming 100.00 g of tetrodotoxin:

$$41.38 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.445 \text{ mol C}; \; 13.16 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 0.9395 \text{ mol N}$$

$$5.37 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.33 \text{ mol H}; \; 40.09 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 2.506 \text{ mol O}$$

Divide by the smallest number:

$$\frac{3.445}{0.9395} = 3.667; \; \frac{5.33}{0.9395} = 5.67; \; \frac{2.506}{0.9395} = 2.667$$

To get whole numbers for each element, multiply through by 3.

Empirical formula: $(C_{3.667}H_{5.67}NO_{2.667})_3 = C_{11}H_{17}N_3O_8$; the mass of the empirical formula is 319.3 g/mol.

Molar mass tetrodotoxin = $\dfrac{1.59\times10^{-21}\text{ g}}{3\text{ molecules}\times\dfrac{1\text{ mol}}{6.022\times10^{23}\text{ molecules}}}$ = 319 g/mol

Because the empirical mass and molar mass are the same, the molecular formula is the same as the empirical formula, $C_{11}H_{17}N_3O_8$.

$$165\text{ lb}\times\frac{1\text{ kg}}{2.2046\text{ lb}}\times\frac{10.\ \mu\text{g}}{\text{kg}}\times\frac{1\times10^{-6}\text{ g}}{\mu\text{g}}\times\frac{1\text{ mol}}{319.3\text{ g}}\times\frac{6.022\times10^{23}\text{ molecules}}{1\text{ mol}}$$

= 1.4×10^{18} molecules tetrodotoxin is the LD_{50} dosage

87. $1.375\text{ g AgI}\times\dfrac{1\text{ mol AgI}}{234.8\text{ g AgI}} = 5.856\times10^{-3}$ mol AgI = 5.856×10^{-3} mol I

$1.375\text{ g AgI}\times\dfrac{126.9\text{ g I}}{234.8\text{ g AgI}}$ = 0.7431 g I; XI_2 contains 0.7431 g I and 0.257 g X.

$5.856\times10^{-3}\text{ mol I}\times\dfrac{1\text{ mol X}}{2\text{ mol I}} = 2.928\times10^{-3}$ mol X

Molar mass = $\dfrac{0.257\text{ g X}}{2.928\times10^{-3}\text{ mol X}} = \dfrac{87.8\text{ g}}{\text{mol}}$; atomic mass = 87.8 amu (X is Sr.)

88. We would see the peaks corresponding to:

$^{10}B^{35}Cl_3$ [mass ≈ 10 + 3(35) = 115 amu], $^{10}B^{35}Cl_2{}^{37}Cl$ (117), $^{10}B^{35}Cl^{37}Cl_2$ (119), $^{10}B^{37}Cl_3$ (121), $^{11}B^{35}Cl_3$ (116), $^{11}B^{35}Cl_2{}^{37}Cl$ (118), $^{11}B^{35}Cl^{37}Cl_2$ (120), $^{11}B^{37}Cl_3$ (122)

We would see a total of eight peaks at approximate masses of 115, 116, 117, 118, 119, 120, 121, and 122.

89. Assuming 1 mole of vitamin A (286.4 g vitamin A):

mol C = 286.4 g vitamin A × $\dfrac{0.8396\text{ g C}}{\text{g vitamin A}}\times\dfrac{1\text{ mol C}}{12.011\text{ g C}}$ = 20.00 mol C

mol H = 286.4 g vitamin A × $\dfrac{0.1056\text{ g H}}{\text{g vitamin A}}\times\dfrac{1\text{ mol H}}{1.0079\text{ g H}}$ = 30.01 mol H

Because 1 mole of vitamin A contains 20 mol C and 30 mol H, the molecular formula of vitamin A is $C_{20}H_{30}E$. To determine E, lets calculate the molar mass of E:

286.4 g = 20(12.01) + 30(1.008) + molar mass E, molar mass E = 16.0 g/mol

From the periodic table, E = oxygen, and the molecular formula of vitamin A is $C_{20}H_{30}O$.

90. X_2Z: 40.0% X and 60.0% Z by mass; $\frac{\text{mol X}}{\text{mol Z}} = 2 = \frac{40.0/A_x}{60.0/A_z} = \frac{(40.0)A_z}{(60.0)A_x}$ or $A_z = 3A_x$ where A = molar mass

For XZ_2, molar mass = $A_x + 2A_z = A_x + 2(3A_x) = 7A_x$.

Mass % X = $\frac{A_x}{7A_x} \times 100 = 14.3\%$ X; % Z = 100.0 – 14.3 = 85.7% Z

91. $453 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{2 \text{ mol Fe}} \times \frac{159.70 \text{ g Fe}_2\text{O}_3}{\text{mol Fe}_2\text{O}_3} = 648 \text{ g Fe}_2\text{O}_3$

Mass % Fe_2O_3 = $\frac{648 \text{ g Fe}_2\text{O}_3}{752 \text{ g ore}} \times 100 = 86.2\%$

92. $^{12}C_2{}^{1}H_6$: 2(12.000000) + 6(1.007825) = 30.046950 amu

$^{12}C^{1}H_2{}^{16}O$: 1(12.000000) + 2(1.007825) + 1(15.994915) = 30.010565 amu

$^{14}N^{16}O$: 1(14.003074) + 1(15.994915) = 29.997989 amu

The peak results from $^{12}C^{1}H_2{}^{16}O$.

93. $\frac{^{85}\text{Rb atoms}}{^{87}\text{Rb atoms}} = 2.591$; If we had exactly 100 atoms, x = number of ^{85}Rb atoms and 100 – x = number of ^{87}Rb atoms.

$\frac{x}{100 - x} = 2.591$, $x = 259.1 - (2.591)x$, $x = \frac{259.1}{3.591} = 72.15$; 72.15% ^{85}Rb

0.7215(84.9117) + 0.2785(A) = 85.4678, A = $\frac{85.4678 - 61.26}{0.2785} = 86.92$ amu

94. a. At 40.0 g of Na added, Cl_2 and Na both run out at the same time (both are limiting reactants). Past 40.0 g of Na added, Cl_2 is limiting, and because the amount of Cl_2 present in each experiment was the same quantity, no more NaCl can be produced. Before 40.0 g of Na added, Na was limiting. As more Na was added (up to 40.0 g Na), more NaCl was produced.

b. $20.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{2 \text{ mol NaCl}}{2 \text{ mol Na}} \times \frac{58.44 \text{ g NaCl}}{\text{mol NaCl}} = 50.8 \text{ g NaCl}$

c. At 40.0 g Na added, both Cl_2 and Na are present in stoichiometric amounts.

$40.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol Na}} \times \frac{70.90 \text{ g Cl}_2}{\text{mol Cl}_2} = 61.7 \text{ g Cl}_2$

61.7 g Cl_2 was present at 40.0 g Na added, and from the problem, the same 61.7 g Cl_2 was present in each experiment.

d. At 50.0 g Na added, Cl_2 is limiting:

$$61.7 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2} \times \frac{2 \text{ mol NaCl}}{1 \text{ mol } Cl_2} \times \frac{58.44 \text{ g NaCl}}{\text{mol NaCl}} = 101.7 \text{ g} = 102 \text{ g NaCl}$$

e. $20.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{1 \text{ mol } Cl_2}{2 \text{ mol Na}} \times \frac{70.90 \text{ g } Cl_2}{\text{mol } Cl_2} = 30.8 \text{ g } Cl_2$ reacted

Excess Cl_2 = 61.7 g Cl_2 initially – 30.8 g Cl_2 reacted = 30.9 g Cl_2 in excess

Note: We know that 40.0 g Na is the point where Na and the 61.7 g of Cl_2 run out at the same time. So if 20.0 g of Na are reacted, one-half of the Cl_2 that was present at 40.0 g Na reacted will be in excess. The previous calculation confirms this.

For 50.0 g Na reacted, Cl_2 is limiting and 40.0 g Na will react as determined previously.

Excess Na = 50.0 g Na initially – 40.0 g Na reacted = 10.0 g Na in excess.

95. $17.3 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 17.2 \text{ mol H}$; $82.7 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.89 \text{ mol C}$

$\frac{17.2}{6.89} = 2.50$; the empirical formula is C_2H_5.

The empirical formula mass is ~29 g, so two times the empirical formula would put the compound in the correct range of the molar mass. Molecular formula = $(C_2H_5)_2 = C_4H_{10}$

$$2.59 \times 10^{23} \text{ atoms H} \times \frac{1 \text{ molecule } C_4H_{10}}{10 \text{ atoms H}} \times \frac{1 \text{ mol } C_4H_{10}}{6.022 \times 10^{23} \text{ molecules}}$$

$$= 4.30 \times 10^{-2} \text{ mol } C_4H_{10}$$

$$4.30 \times 10^{-2} \text{ mol } C_4H_{10} \times \frac{58.12 \text{ g}}{\text{mol } C_4H_{10}} = 2.50 \text{ g } C_4H_{10}$$

96. Assuming 100.00 g E_3H_8:

$$\text{Mol E} = 8.73 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} \times \frac{3 \text{ mol E}}{8 \text{ mol H}} = 3.25 \text{ mol E}$$

$\frac{x \text{ g E}}{1 \text{ mol E}} = \frac{91.27 \text{ g E}}{3.25 \text{ mol E}}$, x = molar mass of E = 28.1 g/mol; atomic mass of E = 28.1 amu

97. Mass of H_2O = 0.755 g $CuSO_4 \bullet xH_2O$ – 0.483 g $CuSO_4$ = 0.272 g H_2O

$$0.483 \text{ g } CuSO_4 \times \frac{1 \text{ mol } CuSO_4}{159.62 \text{ g } CuSO_4} = 0.00303 \text{ mol } CuSO_4$$

$$0.272 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 0.0151 \text{ mol } H_2O$$

$$\frac{0.0151 \text{ mol } H_2O}{0.00303 \text{ mol } CuSO_4} = \frac{4.98 \text{ mol } H_2O}{1 \text{ mol } CuSO_4}$$; compound formula = $CuSO_4 \bullet 5H_2O$, $x = 5$

98. In 1 hour, the 1000. kg of wet cereal contains 580 kg H_2O and 420 kg of cereal. We want the final product to contain 20.% H_2O. Let x = mass of H_2O in final product.

$$\frac{x}{420 + x} = 0.20, \ x = 84 + (0.20)x, \ x = 105 \approx 110 \text{ kg } H_2O$$

The amount of water to be removed is 580 − 110 = 470 kg/h.

99. $$1.20 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g}} \times \frac{1 \text{ mol C}}{\text{mol } CO_2} \times \frac{1 \text{ mol } C_{24}H_{30}N_3O}{24 \text{ mol C}} \times \frac{376.51 \text{ g}}{\text{mol } C_{24}H_{30}N_3O}$$

$$= 0.428 \text{ g } C_{24}H_{30}N_3O$$

$$\frac{0.428 \text{ g } C_{24}H_{30}N_3O}{1.00 \text{ g sample}} \times 100 = 42.8\% \ C_{24}H_{30}N_3O \text{ (LSD)}$$

100. $Ca_3(PO_4)_2(s) + 3\ H_2SO_4(aq) \rightarrow 3\ CaSO_4(s) + 2\ H_3PO_4(aq)$

$$1.0 \times 10^3 \text{ g } Ca_3(PO_4)_2 \times \frac{1 \text{ mol } Ca_3(PO_4)_2}{310.2 \text{ g } Ca_3(PO_4)_2} = 3.2 \text{ mol } Ca_3(PO_4)_2$$

$$1.0 \times 10^3 \text{ g conc. } H_2SO_4 \times \frac{98 \text{ g } H_2SO_4}{100 \text{ g conc. } H_2SO_4} \times \frac{1 \text{ mol } H_2SO_4}{98.1 \text{ g } H_2SO_4} = 10. \text{ mol } H_2SO_4$$

The required mole ratio from the balanced equation is 3 mol H_2SO_4 to 1 mol $Ca_3(PO_4)_2$. The actual ratio is $\frac{10. \text{ mol } H_2SO_4}{3.2 \text{ mol } Ca_3(PO_4)_2} = 3.1$.

This is larger than the required mole ratio, so $Ca_3(PO_4)_2$ is the limiting reagent.

$$3.2 \text{ mol } Ca_3(PO_4)_2 \times \frac{3 \text{ mol } CaSO_4}{\text{mol } Ca_3(PO_4)_2} \times \frac{136.2 \text{ g } CaSO_4}{\text{mol } CaSO_4} = 1300 \text{ g } CaSO_4 \text{ produced}$$

$$3.2 \text{ mol } Ca_3(PO_4)_2 \times \frac{2 \text{ mol } H_3PO_4}{\text{mol } Ca_3(PO_4)_2} \times \frac{98.0 \text{ g } H_3PO_4}{\text{mol } H_3PO_4} = 630 \text{ g } H_3PO_4 \text{ produced}$$

101. Molar mass X_2 = $\dfrac{0.105\text{ g}}{8.92 \times 10^{20}\text{ molecules} \times \dfrac{1\text{ mol}}{6.022 \times 10^{23}\text{ molecules}}}$ = 70.9 g/mol

The mass of X = 1/2(70.9 g/mol) = 35.5 g/mol. This is the element chlorine.

Assuming 100.00 g of MX_3 compound:

$$54.47\text{ g Cl} \times \frac{1\text{ mol}}{35.45\text{ g}} = 1.537\text{ mol Cl}$$
$$1.537\text{ mol Cl} \times \frac{1\text{ mol M}}{3\text{ mol Cl}} = 0.5123\text{ mol M}$$

Molar mass of M = $\dfrac{45.53\text{ g M}}{0.5123\text{ mol M}}$ = 88.87 g/mol M

M is the element yttrium (Y), and the name of YCl_3 is yttrium(III) chloride.

The balanced equation is $2\ Y + 3\ Cl_2 \rightarrow 2\ YCl_3$.

Assuming Cl_2 is limiting:

$$1.00\text{ g } Cl_2 \times \frac{1\text{ mol } Cl_2}{70.90\text{ g } Cl_2} \times \frac{2\text{ mol } YCl_3}{3\text{ mol } Cl_2} \times \frac{195.26\text{ g } YCl_3}{1\text{ mol } YCl_3} = 1.84\text{ g } YCl_3$$

Assuming Y is limiting:

$$1.00\text{ g Y} \times \frac{1\text{ mol Y}}{88.91\text{ g Y}} \times \frac{2\text{ mol } YCl_3}{2\text{ mol Y}} \times \frac{195.26\text{ g } YCl_3}{1\text{ mol } YCl_3} = 2.20\text{ g } YCl_3$$

Because Cl_2, when it all reacts, produces the smaller amount of product, Cl_2 is the limiting reagent, and the theoretical yield is 1.84 g YCl_3.

102. The reaction is $BaX_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ HX(aq)$.

$0.124\text{ g } BaSO_4 \times \dfrac{137.3\text{ g Ba}}{233.4\text{ g } BaSO_4}$ = 0.0729 g Ba; % Ba = $\dfrac{0.0729\text{ g Ba}}{0.158\text{ g } BaX_2} \times 100$ = 46.1%

The formula is BaX_2 (from positions of the elements in the periodic table), and 100.0 g of compound contains 46.1 g Ba and 53.9 g of the unknown halogen. There must also be:

$$46.1\text{ g Ba} \times \frac{1\text{ mol Ba}}{137.3\text{ g Ba}} \times \frac{2\text{ mol X}}{\text{mol Ba}} = 0.672\text{ mol of the halogen in 53.9 g of halogen}$$

Therefore, the molar mass of the halogen is $\dfrac{53.9\text{ g}}{0.672\text{ mol}}$ = 80.2 g/mol.

This molar mass is close to that of bromine. Thus the formula of the compound is $BaBr_2$.

103. Consider the case of aluminum plus oxygen. Aluminum forms Al^{3+} ions; oxygen forms O^{2-} anions. The simplest compound of the two elements is Al_2O_3. Similarly, we would expect the formula of a Group 6A element with Al to be Al_2X_3. Assuming this, out of 100.00 g of compound, there are 18.56 g Al and 81.44 g of the unknown element, X. Let's use this information to determine the molar mass of X, which will allow us to identify X from the periodic table.

$$18.56 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol X}}{2 \text{ mol Al}} = 1.032 \text{ mol X}$$

81.44 g of X must contain 1.032 mol of X.

$$\text{Molar mass of X} = \frac{81.44 \text{ g X}}{1.032 \text{ mol X}} = 78.91 \text{ g/mol}$$

From the periodic table, the unknown element is selenium, and the formula is Al_2Se_3.

104. Empirical formula mass = 12.01 + 1.008 = 13.02 g/mol; because 104.14/13.02 = 7.998 ≈ 8, the molecular formula for styrene is $(CH)_8 = C_8H_8$.

$$2.00 \text{ g } C_8H_8 \times \frac{1 \text{ mol } C_8H_8}{104.14 \text{ g } C_8H_8} \times \frac{8 \text{ mol H}}{\text{mol } C_8H_8} \times \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol H}} = 9.25 \times 10^{22} \text{ atoms H}$$

105. $2\ NaNO_3(s) \rightarrow 2\ NaNO_2(s) + O_2(g)$; the amount of $NaNO_3$ in the impure sample is:

$$0.2864 \text{ g } NaNO_2 \times \frac{1 \text{ mol } NaNO_2}{69.00 \text{ g } NaNO_2} \times \frac{2 \text{ mol } NaNO_3}{2 \text{ mol } NaNO_2} \times \frac{85.00 \text{ g } NaNO_3}{\text{mol } NaNO_3}$$

$$= 0.3528 \text{ g } NaNO_3$$

$$\text{Mass percent } NaNO_3 = \frac{0.3528 \text{ g } NaNO_3}{0.4230 \text{ g sample}} \times 100 = 83.40\%$$

106. $PaO_2 + O_2 \rightarrow Pa_xO_y$ (unbalanced)

$$0.200 \text{ g } PaO_2 \times \frac{231 \text{ g Pa}}{263 \text{ g } PaO_2} = 0.1757 \text{ g Pa}$$ (We will carry an extra significant figure.)

$0.2081 \text{ g } Pa_xO_y - 0.1757 \text{ g Pa} = 0.0324 \text{ g O}$; $0.0324 \text{ g O} \times \dfrac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.025 \times 10^{-3} \text{ mol O}$

$$0.1757 \text{ g Pa} \times \frac{1 \text{ mol Pa}}{231 \text{ g Pa}} = 7.61 \times 10^{-4} \text{ mol Pa}$$

$$\frac{\text{Mol O}}{\text{Mol Pa}} = \frac{2.025 \times 10^{-3} \text{ mol O}}{7.61 \times 10^{-4} \text{ mol Pa}} = 2.66 \approx 2\frac{2}{3} = \frac{8 \text{ mol O}}{3 \text{ mol Pa}}$$; empirical formula: Pa_3O_8

Challenge Problems

107. The balanced equations are:

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g) \text{ and } 4\ NH_3(g) + 7\ O_2(g) \rightarrow 4\ NO_2(g) + 6\ H_2O(g)$$

Let $4x$ = number of moles of NO formed, and let $4y$ = number of moles of NO_2 formed. Then:

$$4x\ NH_3 + 5x\ O_2 \rightarrow 4x\ NO + 6x\ H_2O \text{ and } 4y\ NH_3 + 7y\ O_2 \rightarrow 4y\ NO_2 + 6y\ H_2O$$

All the NH_3 reacted, so $4x + 4y = 2.00$.

10.00 – 6.75 = 3.25 mol O_2 reacted, so $5x + 7y = 3.25$.

Solving by the method of simultaneous equations:

$$\begin{array}{rl} 20x + 28y = & 13.0 \\ -\underline{20x - 20y = -10.0} & \\ 8y = & 3.0, \end{array} \quad y = 0.38;\ \ 4x + 4 \times 0.38 = 2.00,\ \ x = 0.12$$

Mol NO = $4x$ = 4 × 0.12 = 0.48 mol NO formed

108. $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(l)$; $C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$
30.07 g/mol 44.09 g/mol

Let x = mass C_2H_6, so 9.780 – x = mass C_3H_8. Use the balanced reaction to set up an equation for the moles of O_2 required.

$$\frac{x}{30.07} \times \frac{7}{2} + \frac{9.780 - x}{44.09} \times \frac{5}{1} = 1.120 \text{ mol } O_2$$

Solving: x = 3.7 g C_2H_6; $\frac{3.7 \text{ g}}{9.780 \text{ g}} \times 100 = 38\%$ C_2H_6 by mass

109. For a gas, density and molar mass are directly proportional to each other.

Molar mass XH_n = 2.393(32.00) = $\frac{76.58 \text{ g}}{\text{mol}}$

$$0.803 \text{ g } H_2O \times \frac{2 \text{ mol H}}{18.02 \text{ g } H_2O} = 8.91 \times 10^{-2} \text{ mol H}$$

$$\frac{8.91 \times 10^{-2} \text{ mol H}}{2.23 \times 10^{-2} \text{ mol } XH_n} = \frac{4 \text{ mol H}}{\text{mol } XH_n}$$

Molar mass X = 76.58 – 4(1.008 g) = 72.55 g/mol; the element is Ge.

110. $a\ N_2H_4 + b\ NH_3 + (10.00 - 4.062)\ O_2 \rightarrow c\ NO_2 + d\ H_2O$

Setting up four equations to solve for the four unknowns:

$2a + b = c$ (N mol balance)

$2c + d = 2(10.00 - 4.062)$ (O mol balance)

$4a + 3b = 2d$ (H mol balance)

$a(32.05) + b(17.03) = 61.00$ (mass balance)

Solving the simultaneous equations gives $a = 1.12 = 1.1$ mol N_2H_4.

$$\frac{1.1\ \text{mol}\ N_2H_4 \times 32.05\ \text{g/mol}\ N_2H_4}{61.00\ \text{g}} \times 100 = 58\%\ N_2H_4$$

111. Let x = mass KCl and y = mass KNO_3. Assuming 100.0 g of mixture, $x + y = 100.0$ g.

Molar mass KCl = 74.55 g/mol; molar mass KNO_3 = 101.11 g/mol

Mol KCl = $\frac{x}{74.55}$; mol KNO_3 = $\frac{y}{101.11}$

Knowing that the mixture is 43.2% K, then in the 100.0 g mixture:

$$39.10\left(\frac{x}{74.55} + \frac{y}{101.11}\right) = 43.2$$

We have two equations and two unknowns:

$$(0.5245)x + (0.3867)y = 43.2$$
$$x + y = 100.0$$

Solving, $x = 32.9$ g KCl; $\frac{32.9\ \text{g}}{100.0\ \text{g}} \times 100 = 32.9\%$ KCl

112. We know that water is a product, so one of the elements in the compound is hydrogen.

$X_aH_b + O_2 \rightarrow H_2O + ?$

To balance the H atoms, the mole ratio between X_aH_b and $H_2O = \frac{2}{b}$.

mol compound = $\frac{1.39\ \text{g}}{62.09\ \text{g/mol}} = 0.0224$ mol; mol $H_2O = \frac{1.21\ \text{g}}{18.02\ \text{g/mol}} = 0.0671$ mol

$\frac{2}{b} = \frac{0.0224}{0.0671}$, $b = 6$; X_aH_6 has a molar mass of 62.09 g/mol.

62.09 = a(molar mass of X) + 6(1.008), a(molar mass of X) = 56.04

Some possible identities for X could be Fe (a = 1), Si (a = 2), N (a = 4), and Li (a = 8). N fits the data best so N_4H_6 is the most likely formula.

113. When the discharge voltage is low, the ions present are in the form of molecules. When the discharge voltage is increased, the bonds in the molecules are broken, and the ions present are in the form of individual atoms. Therefore, the high discharge data indicate that the ions $^{16}O^+$, $^{18}O^+$, and $^{40}Ar^+$ are present. The only combination of these individual ions that can explain the mass data at low discharge is $^{16}O^{16}O^+$ (mass = 32), $^{16}O^{18}O^+$ (mass = 34), and $^{40}Ar^+$ (mass = 40). Therefore, the gas mixture contains $^{16}O^{16}O$, $^{16}O^{18}O$, and ^{40}Ar. To determine the percent composition of each isotope, we use the relative intensity data from the high discharge data to determine the percentage that each isotope contributes to the total relative intensity. For ^{40}Ar:

$$\frac{1.0000}{0.7500 + 0.0015 + 1.0000} \times 100 = \frac{1.0000}{1.7515} \times 100 = 57.094\%\ ^{40}Ar$$

For ^{16}O: $\frac{0.7500}{1.7515} \times 100 = 42.82\%\ ^{16}O$; for ^{18}O: $\frac{0.0015}{1.7515} \times 100 = 8.6 \times 10^{-2}\%\ ^{18}O$

Note: ^{18}F instead of ^{18}O could also explain the data. However, OF(g) is not a stable compound. This is why ^{18}O is the best choice because O_2(g) does form.

114. $Fe(s) + \frac{1}{2} O_2(g) \rightarrow FeO(s)$; $2\ Fe(s) + \frac{3}{2} O_2(g) \rightarrow Fe_2O_3(s)$

$20.00\ g\ Fe \times \frac{1\ mol\ Fe}{55.85\ g} = 0.3581\ mol$

$(11.20 - 3.24)\ g\ O_2 \times \frac{1\ mol\ O_2}{32.00\ g} = 0.2488\ mol\ O_2$ consumed (1 extra sig. fig.)

Assuming x mol of FeO is produced from x mol of Fe, so that 0.3581 – x mol of Fe reacts to form Fe_2O_3:

$$x\ Fe + \frac{1}{2} x\ O_2 \rightarrow x\ FeO$$

$$(0.3581 - x)\ mol\ Fe + \frac{3}{2}\left(\frac{0.3581 - x}{2}\right) mol\ O_2 \rightarrow \left(\frac{0.3581 - x}{2}\right) mol\ Fe_2O_3$$

Setting up an equation for total moles of O_2 consumed:

$$\frac{1}{2}x + \frac{3}{4}(0.3581 - x) = 0.2488\ mol\ O_2, \quad x = 0.0791 = 0.079\ mol\ FeO$$

$0.079 \text{ mol FeO} \times \frac{71.85 \text{ g FeO}}{\text{mol}} = 5.7 \text{ g FeO produced}$

Mol Fe_2O_3 produced = $\frac{0.3581 - 0.079}{2} = 0.140 \text{ mol } Fe_2O_3$

$0.140 \text{ mol } Fe_2O_3 \times \frac{159.70 \text{ g } Fe_2O_3}{\text{mol}} = 22.4 \text{ g } Fe_2O_3$ produced

115. $4.000 \text{ g } M_2S_3 \rightarrow 3.723 \text{ g } MO_2$

There must be twice as many moles of MO_2 as moles of M_2S_3 in order to balance M in the reaction. Setting up an equation for 2(mol M_2S_3) = mol MO_2 where A = molar mass M:

$$2\left(\frac{4.000 \text{ g}}{2A + 3(32.07)}\right) = \frac{3.723 \text{ g}}{A + 2(16.00)}, \quad \frac{8.000}{2A + 96.21} = \frac{3.723}{A + 32.00}$$

(8.000)A + 256.0 = (7.446)A + 358.2, (0.554)A = 102.2, A = 184 g/mol; atomic mass = 184 amu

116. The two relevant equations are:

$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ and $Mg(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Let x = mass Mg, so 10.00 – x = mass Zn. From the balanced equations, moles H_2 = moles Zn + moles Mg.

mol H_2 = $0.5171 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.0158 \text{ g } H_2} = 0.2565 \text{ mol } H_2$

$0.2565 = \frac{x}{24.31} + \frac{10.00 - x}{65.38}$; solving, x = 4.008 g Mg.

$\frac{4.008 \text{ g}}{10.00 \text{ g}} \times 100 = 40.08\%$ Mg

117. 10.00 g XCl_2 + excess $Cl_2 \rightarrow$ 12.55 g XCl_4; 2.55 g Cl reacted with XCl_2 to form XCl_4. XCl_4 contains 2.55 g Cl and 10.00 g XCl_2. From mole ratios, 10.00 g XCl_2 must also contain 2.55 g Cl; mass X in XCl_2 = 10.00 – 2.55 = 7.45 g X.

$$2.55 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} \times \frac{1 \text{ mol } XCl_2}{2 \text{ mol Cl}} \times \frac{1 \text{ mol X}}{\text{mol } XCl_2} = 3.60 \times 10^{-2} \text{ mol X}$$

So, 3.60×10^{-2} mol X has a mass equal to 7.45 g X. The molar mass of X is:

$$\frac{7.45 \text{ g X}}{3.60 \times 10^{-2} \text{ mol X}} = 207 \text{ g/mol X};$$ atomic mass = 207 amu, so X is Pb.

118. The balanced equation is 2 Sc(s) + 2x HCl(aq) $\longrightarrow$ 2 $ScCl_x$(aq)+ x H_2(g).

The mol ratio of Sc to H_2 = $\frac{2}{x}$.

mol Sc = 2.25 g Sc × $\frac{1\ \text{mol Sc}}{44.96\ \text{g Sc}}$ = 0.0500 mol Sc

mol H_2 = 0.1502 g H_2 × $\frac{1\ \text{mol } H_2}{2.0158\ \text{g } H_2}$ = 0.07451 mol H_2

$\frac{2}{x} = \frac{0.0500}{0.07451}$, $x = 3$; the formula is $ScCl_3$.

119. $LaH_{2.90}$ is the formula. If only La^{3+} is present, LaH_3 would be the formula. If only La^{2+} is present, LaH_2 would be the formula. Let x = mol La^{2+} and y = mol La^{3+}:

$(La^{2+})_x(La^{3+})_yH_{(2x + 3y)}$ where $x + y = 1.00$ and $2x + 3y = 2.90$

Solving by simultaneous equations:

$$\begin{array}{rcl} 2x + 3y &=& 2.90 \\ \underline{-2x - 2y} &=& \underline{-2.00} \\ y &=& 0.90 \end{array} \quad \text{and } x = 0.10$$

$LaH_{2.90}$ contains $\frac{1}{10}$ La^{2+}, or 10.% La^{2+}, and $\frac{9}{10}$ La^{3+}, or 90.% La^{3+}.

120. $C_xH_yO_z$ + oxygen $\rightarrow$ x CO_2 + y/2 H_2O

$$\text{Mass \% C in aspirin} = \frac{2.20\ \text{g } CO_2 \times \frac{1\ \text{mol } CO_2}{44.01\ \text{g } CO_2} \times \frac{1\ \text{mol C}}{\text{mol } CO_2} \times \frac{12.01\ \text{g C}}{\text{mol C}}}{1.00\ \text{g aspirin}} = 60.0\%\ \text{C}$$

$$\text{Mass \% H in aspirin} = \frac{0.400\ \text{g } H_2O \times \frac{1\ \text{mol } H_2O}{18.02\ \text{g } H_2O} \times \frac{2\ \text{mol H}}{\text{mol } H_2O} \times \frac{1.008\ \text{g H}}{\text{mol H}}}{1.00\ \text{g aspirin}} = 4.48\%\ \text{H}$$

Mass % O = 100.00 – (60.0 + 4.48) = 35.5% O

Assuming 100.00 g aspirin:

60.0 g C × $\frac{1\ \text{mol C}}{12.01\ \text{g C}}$ = 5.00 mol C; 4.48 g H × $\frac{1\ \text{mol H}}{1.008\ \text{g H}}$ = 4.44 mol H

35.5 g O × $\frac{1\ \text{mol O}}{16.00\ \text{g O}}$ = 2.22 mol O

Dividing by the smallest number: $\frac{5.00}{2.22}$ = 2.25; $\frac{4.44}{2.22}$ = 2.00

Empirical formula: $(C_{2.25}H_{2.00}O)_4 = C_9H_8O_4$. Empirical mass ≈ 9(12) + 8(1) + 4(16) = 180 g/mol; this is in the 170–190 g/mol range, so the molecular formula is also $C_9H_8O_4$.

Balance the aspirin synthesis reaction to determine the formula for salicylic acid.

$C_aH_bO_c + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2$, $C_aH_bO_c$ = salicylic acid = $C_7H_6O_3$

121. The balanced equations are:

$$C(s) + 1/2\ O_2(g) \rightarrow CO(g) \text{ and } C(s) + O_2(g) \rightarrow CO_2(g)$$

If we have 100.0 mol of products, then we have 72.0 mol CO_2, 16.0 mol CO, and 12.0 mol O_2. The initial moles of C equals 72.0 (from CO_2) + 16.00 (from CO) = 88.0 mol C and the initial moles of O_2 equals 72.0 (from CO_2) + 16.0/2 (from CO) + 12.0 (unreacted O_2) = 92.0 mol O_2. The initial reaction mixture contained:

$$\frac{92.0 \text{ mol } O_2}{88.0 \text{ mol C}} = 1.05 \text{ mol } O_2\text{/mol C}$$

122. Let M = unknown element

Mass % M = $\frac{\text{mass M}}{\text{total mass compound}} \times 100 = \frac{2.077}{3.708} \times 100$ = 56.01% M

100.00 – 56.01 = 43.99% O

Assuming 100.00 g compound:

$$43.99 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 2.750 \text{ mol O}$$

If MO is the formula of the oxide, then M has a molar mass of $\frac{56.01 \text{ g M}}{2.750 \text{ mol M}}$ = 20.37 g/mol. This is too low for the molar mass. We must have fewer moles of M than moles O present in the formula. Some possibilities are MO_2, M_2O_3, MO_3, etc. It is a guessing game as to which to try. Let's assume an MO_2 formula. Then the molar mass of M is:

$$\frac{56.01 \text{ g M}}{2.750 \text{ mol O} \times \frac{1 \text{ mol M}}{2 \text{ mol O}}} = 40.73 \text{ g/mol}$$

This is close to calcium, but calcium forms an oxide having the CaO formula, not CaO_2.

If MO_3 is assumed to be the formula, then the molar mass of M calculates to be 61.10 g/mol which is too large. Therefore, the mol O to mol M ratio must be between 2 and 3. Some reasonable possibilities are 2.25, 2.33, 2.5, 2.67, and 2.75 (these are reasonable because they will lead to whole number formulas). Trying a mol O to mol M ratio of 2.5 to 1 gives a molar mass of:

$$\frac{56.01\text{ g M}}{2.750\text{ mol O} \times \frac{1\text{ mol M}}{2.5\text{ mol O}}} = 50.92\text{ g/mol}$$

This is the molar mass of vanadium and V_2O_5 is a reasonable formula for an oxide of vanadium. The other choices for the O : M mole ratios between 2 and 3 do not give as reasonable results. Therefore, M is vanadium, and the formula is V_2O_5.

Marathon Problems

123. To solve the limiting-reagent problem, we must determine the formulas of all the compounds so that we can get a balanced reaction.

a. 40 million trillion = $(40 \times 10^6) \times 10^{12} = 4.000 \times 10^{19}$ (assuming 4 sig. figs.)

$$4.000 \times 10^{19}\text{ molecules A} \times \frac{1\text{ mol A}}{6.022 \times 10^{23}\text{ molecules A}} = 6.642 \times 10^{-5}\text{ mol A}$$

$$\text{Molar mass of A} = \frac{4.26 \times 10^{-3}\text{ g A}}{6.642 \times 10^{-5}\text{ mol A}} = 64.1\text{ g/mol}$$

Mass of carbon in 1 mol of A is:

$$64.1\text{ g A} \times \frac{37.5\text{ g C}}{100.0\text{ g A}} = 24.0\text{ g carbon} = 2\text{ mol carbon in substance A}$$

The remainder of the molar mass (64.1 g – 24.0 g = 40.1 g) is due to the alkaline earth metal. From the periodic table, calcium has a molar mass of 40.08 g/mol. The formula of substance A is CaC_2.

b. 5.36 g H + 42.5 g O = 47.9 g; substance B only contains H and O. Determining the empirical formula of B:

$$5.36\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 5.32\text{ mol H};\quad \frac{5.32}{2.66} = 2.00$$

$$42.5\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 2.66\text{ mol O};\quad \frac{2.66}{2.66} = 1.00$$

Empirical formula: H_2O; the molecular formula of substance B could be H_2O, H_4O_2, H_6O_3, etc. The most reasonable choice is water (H_2O) for substance B.

c. Substance C + $O_2 \rightarrow CO_2 + H_2O$; substance C must contain carbon and hydrogen and may contain oxygen. Determining the mass of carbon and hydrogen in substance C:

$$33.8 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{\text{mol } CO_2} \times \frac{12.01 \text{ g C}}{\text{mol C}} = 9.22 \text{ g carbon}$$

$$6.92 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{\text{mol } H_2O} \times \frac{1.008 \text{ g H}}{\text{mol H}} = 0.774 \text{ g hydrogen}$$

9.22 g carbon + 0.774 g hydrogen = 9.99 g; because substance C initially weighed 10.0 g, there is no oxygen present in substance C. Determining the empirical formula for substance C:

$$9.22 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.768 \text{ mol carbon}$$

$$0.774 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.768 \text{ mol hydrogen}$$

Mol C/mol H = 1.00; the empirical formula is CH which has an empirical formula mass ≈ 13. Because the mass spectrum data indicate a molar mass of 26 g/mol, the molecular formula for substance C is C_2H_2.

d. Substance D is $Ca(OH)_2$. Now we can answer the question. The balanced equation is:

$$CaC_2(s) + 2\, H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

$$45.0 \text{ g } CaC_2 \times \frac{1 \text{ mol } CaC_2}{64.10 \text{ g } CaC_2} = 0.702 \text{ mol } CaC_2$$

$$23.0 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 1.28 \text{ mol } H_2O; \quad \frac{\text{mol } H_2O}{\text{mol } CaC_2} = \frac{1.28}{0.702} = 1.82$$

Because the actual mole ratio present is smaller than the required 2 : 1 mole ratio from the balanced equation, H_2O is limiting.

$$1.28 \text{ mol } H_2O \times \frac{1 \text{ mol } C_2H_2}{2 \text{ mol } H_2O} \times \frac{26.04 \text{ g } C_2H_2}{\text{mol } C_2H_2} = 16.7 \text{ g } C_2H_2 = \text{mass of product C}$$

124. a. i. If the molar mass of A is greater than the molar mass of B, then we cannot determine the limiting reactant because, while we have a fewer number of moles of A, we also need fewer moles of A (from the balanced reaction).

ii. If the molar mass of B is greater than the molar mass of A, then B is the limiting reactant because we have a fewer number of moles of B and we need more B (from the balanced reaction).

b. A + 5 B $\rightarrow$ 3 CO_2 + 4 H_2O

To conserve mass: 44.01 + 5(B) = 3(44.01) + 4(18.02); solving: B = 32.0 g/mol

Because B is diatomic, the best choice for B is O_2.

c. We can solve this without mass percent data simply by balancing the equation:

A + 5 O_2 $\rightarrow$ 3 CO_2 + 4 H_2O

A must be C_3H_8 (which has a similar molar mass to CO_2). This is also the empirical formula.

Note: $\frac{3(12.01)}{3(12.01) + 8(1.008)} \times 100 = 81.71\%$ C. So this checks.