Chapter 1

The Nature and Conceptual Basis of Chemistry

1.2 "Absolute" means pure, so absolute alcohol is a substance; milk is a mixture; copper wire is a substance. Rust is a mixture (the reason for this answer is discussed on text page 7 with respect to salt / sodium chloride). Barium bromide is a substance. Concrete, baking soda, and baking powder are all mixtures. Absolute alcohol and barium bromide are compounds; copper wire is an element. All of the mixtures are heterogeneous.

1.4 Proving that a material is *not* an element requires finding a method, any one method, that successfully breaks it down into simpler substances. Proving that a material *is* an element requires the proof of a negative: that there is no way to break it down. Mistaken reports of new elements arise when a compound or mixture resists breakdown by all of the methods tried but subsequently proves decomposable by new methods.

1.6 The ratio of the mass of tellurium to the mass of hafnium in this compound is

$$\frac{m_{\text{Te}}}{m_{\text{Hf}}} = \frac{31.5 \text{ g Te}}{25.0 \text{ g Hf}} = \frac{1.26 \text{ g Te}}{\text{g Hf}}$$

Because the compound from the rock is identical, it contains Te and Hf in the same ratio.

$$m_{\text{Te}} = 0.125 \text{ g Hf} \times \left(1.26 \frac{\text{g Te}}{\text{g Hf}}\right) = 0.158 \text{ g Te}$$

The compound may of course contain other elements.

1.8 **a)** The mass of fluorine that combines with 1.0000 g of iodine in these compounds is the mass percentage of fluorine divided by the mass percentage of iodine. This is well shown by considering samples of the compounds that have masses of 100.000 g. The masses contributed by each element in the compounds are then very easily computed. The ratios in the last column of the following table, which are formed by the indicated divisions, are the answers.

Compound 1	13.021 g F / 86.979 g I	0.14970 g F / g I
Compound 2	30.993 g F / 69.007 g I	0.44913 g F / g I
Compound 3	42.809 g F / 57.191 g I	0.74853 g F / g I
Compound 4	51.171 g F / 48.829 g I	1.04796 g F / g I

b) The law of multiple proportions involves the ratio of these ratios. Divide all four of the answers in part a) by the smallest of the answers. The results are: 1.0000 for compound 1; 3.0002 for compound 2; 5.0002 for compound 3; 7.0004 for compound 4. These equal the small whole numbers 1, 3, 5, and 7 within the precision of the data.

1.10 As in problem **1.8**, calculate the masses of chlorine per gram of tungsten in the four compounds:

Compound 1	27.83 g Cl / 72.17 g W	0.3856 g Cl / g W
Compound 2	43.55 g Cl / 56.45 g W	0.7715 g Cl / g W
Compound 3	49.09 g Cl / 50.91 g W	0.9643 g Cl / g W
Compound 4	53.64 g Cl / 46.36 g W	1.1570 g Cl/ g W

The ratios of each mass of chlorine to the smallest mass of chlorine are

0.3856/0.3856 = 1.0000 = 2 : 2 0.7715/0.3856 = 2.0008 = 4 : 2
0.9643/0.3856 = 2.5008 = 5 : 2 1.1570/0.3856 = 3.0005 =6 : 2

The formulas are WCl_2, WCl_4, WCl_5, and WCl_6.

1.12 The *only* products are gaseous N_2 and gaseous H_2. From the formula of the starting compound there are twice as many molecules of H_2 as of N_2 in the products. The law of combining volumes (or, in this case, the law of "uncombining" volumes) then assures that the volume of hydrogen is twice the volume of nitrogen as long as the temperature and pressure remain unchanged. The answer is 27.4 mL.

1.14 The balanced chemical equation for this reaction is

$$2CH_3OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$$

2.0 L of CO_2 and 4.0 L of H_2O are produced from 2.0 L of CH_2OH, according to the law of combining volumes (and under the assumption that the reaction goes to completion as written).

1.16 To solve for the *y*-component of the velocity of the electron, start with Equation 1.4:

$$\frac{1}{2}a_y t^2 = \frac{1}{2}\left(\frac{e}{m_e}\right)\left(\frac{l}{|v_x|}\right)^2 E$$

Now substitute $a_y = \Delta v_y / \Delta t$:

$$\frac{1}{2}\left(\frac{\Delta v_y}{\Delta t}\right)t^2 = \frac{1}{2}\left(\frac{e}{m_e}\right)\left(\frac{l}{|v_x|}\right)^2 E$$

Define the time at which the electron just enters the left side of the deflection plate as t_0 and arbitrarily set this time equal to zero. At the entry point, the electron's velocity solely lies along the *x*-axis, therefore the initial *y*-component of the velocity, $v_{y,0}$, is also equal to zero. When the electron exits the plate region at time *t*, its velocity has a non-zero *y*-component, v_y. Taking this into consideration, we can write the *y*-component of the acceleration as:

$$a_y = \frac{\Delta v_y}{\Delta t} = \frac{v_y - v_{y,0}}{t - t_0} = \frac{v_y - 0}{t - 0} = \frac{v_y}{t}$$

Substituting, we get:

$$\frac{1}{2}\left(\frac{v_y}{t}\right)t^2 = \frac{1}{2}\left(\frac{e}{m_e}\right)\left(\frac{l}{|v_x|}\right)^2 E$$

$$v_y t = \left(\frac{e}{m_e}\right)\left(\frac{l}{|v_x|}\right)^2 E$$

$$v_y = \left(\frac{e}{m_e}\right)\left(\frac{l}{|v_x|}\right)^2 \frac{E}{t}$$

If we approximate the time spent in the plate region as:

$$t = \frac{l}{|v_x|}$$

where we take $|v_x|$ as the initial electron velocity, one last substitution yields:

$$v_y = \left(\frac{e}{m_e}\right)\left(\frac{l}{|v_x|}\right)^2 \frac{E}{\left(l/|v_x|\right)}$$
$$= \left(\frac{e}{m_e}\right)\left(\frac{l}{|v_x|}\right)E$$

Solve for v_y by substituting the appropriate values:

$v_y = (1.602 \times 10^{-19}\ \text{C} / 9.109 \times 10^{-31}\ \text{kg}) \times (0.10\ \text{m} / |8.00 \times 10^6\ \text{m s}^{-1}|) \times 200\ \text{V m}^{-1} = 4.40 \times 10^5\ \text{m s}^{-1}$

Because the deflection is in the y direction is negative, the y-component of the velocity must also be negative, therefore $v_y = -4.40 \times 10^5\ \text{m s}^{-1}$.

Strictly speaking, this is an approximation. The time spent in the plate region is actually just a little longer because of the curved path the electron takes. To check the change in the x-component of the velocity, we can use the relationship:

$$v^2 = v_x^2 + v_y^2$$
$$v_x^2 = v^2 - v_y^2$$
$$v_x = \pm\sqrt{v^2 - v_y^2}$$

Letting $v = 8.00 \times 10^6\ \text{m s}^{-1}$ (the electron velocity before it enters the plate region) and $v_y = -4.40 \times 10^5\ \text{m s}^{-1}$, we get $v_x = 7.99 \times 10^6\ \text{m s}^{-1}$. This is less than a 1% error, so our assumptions above are not bad!

Now we address the ratio of S/L and its relationship to v_x and v_y if v_y is calculated at the right edge of the plates. Once the electron leaves the plate region, it is no longer subject to a force acting on it. As such, the electron moves in straight line motion with constant velocity components v_x and v_y. Taking t_f to be the time of flight between the right edge of the plate and the fluorescent screen, the electron will travel a distance of $v_x t_f$ in the x-direction and a distance of $v_y t_f$ in the y-direction. With respect to the distances shown in Figure 1.11, the amount of travel in the x-direction between the right edge of the plate and the screen is $L - l/2$. In the y-direction, the distance is $S - D$. We thus have the following relationships:

$$v_x t_f = L - \frac{l}{2}$$
$$v_y t_f = S - D$$

Rearranging:

$$v_x t_f + \frac{l}{2} = L$$
$$v_y t_f + D = S$$

Taking the ratio of S/L:

$$\frac{S}{L} = \frac{v_y t_f + D}{v_x t_f + l/2}$$

1.18 The atomic mass of naturally occurring neon is found by multiplying each isotope's fractional abundance by its mass and summing over all the isotopes

$$A = A_1 p_1 + A_2 p_2 + \cdots + A_n p_n$$
$$A_{\text{Ne}} = (0.9000)(19.99212) + (0.0027)(20.99316) + (0.0973)(21.99132) = 20.19$$

1.20 This problem resembles problem **1.16**, except that the atomic mass of one of the five isotopes of Zr is not known and the weighted-average atomic mass of natural zirconium (91.224) is known. Obtain the natural abundance of the isotope of interest by subtraction

$$p\left({}^{90}\text{Zr}\right) = 1 - 0.1127 - 0.1717 - 0.1733 - 0.0278 = 0.5145$$

Let the relative mass of this isotope be A_{90}. Then

$$91.224 = 0.5145A_{90} + 0.1127(90.9056) + 0.1717(91.9050) + 0.1733(93.9063) + 0.0278(95.9083)$$
$$A_{90} = 89.91$$

1.22 **a)** Promethium has an atomic number of 61; the ratio of the number of neutrons to protons in ^{145}Pm is (145 – 61)/61 = 1.377.

b) A neutral atom of Pm has 61 electrons.

1.24 The ${}^{266}_{109}\text{Mt}$ atom has 109 protons, 109 electrons, and 157 neutrons.

1.26 Avogadro's number of ^{12}C atoms has a mass of exactly 12 g. By use of the ratio given in the problem, Avogadro's number of F atoms must have a mass of 18.998403 g. For 100 million atoms of fluorine

$$m_{\text{F}} = 10^8 \text{atoms F} \times \left(\frac{18.998403 \text{ g F}}{6.0221420 \times 10^{23} \text{atom F}}\right) = 3.1547584 \times 10^{-15} \text{g F}$$

1.28 The problem is about handling the nesting of parentheses in chemical formulas when computing molecular masses, formula masses, and molar masses. The answers: **a)** 177.382; **b)** 598.156; **c)** 254.2; **d)** 98.079; **e)** 450.446. These are relative masses and so have no units.

1.30 There are Avogadro's number of gold atoms in a mole of gold, each with a diameter of 2.88×10^{-10} m. The length of the line is $\left(6.022 \times 10^{23}\right)\left(2.88 \times 10^{-10}\text{m}\right) = 1.73 \times 10^{14}\text{m}$.

1.32 Express the amounts of each sample in the same unit of mass. In the case of the SF_4, convert from the given number of moles to grams using the molar mass. In the cases of the Cl_2O_7 and Ar, convert from the given number of particles to chemical amount and from there to mass. The amount of CH_4 is already in grams. The results are

$$\text{SF}_4(115 \text{ g}) < \text{CH}_4(117 \text{ g}) < \text{Cl}_2\text{O}_7(264 \text{ g}) < \text{Ar}(2770 \text{ g})$$

1.34

$$10.0 \text{ cm}^3\text{Au} \times \left(\frac{19.32 \text{ g Au}}{1 \text{ cm}^3\text{Au}}\right) \times \left(\frac{1 \text{ troy ounce}}{31.1035 \text{ g Au}}\right) \times \left(\frac{\$400}{1 \text{ troy ounce}}\right) = \$2.4846 \times 10^3$$

The cost to three significant figures is \$2.480.

1.36

$$N_{\text{Si}} = 415\text{cm}^3 \times \frac{0.00278 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ mol Si}_2\text{H}_6}{62.219 \text{ g Si}_2\text{H}_6} \times \frac{6.022 \times 10^{23} \text{ molecule Si}_2\text{H}_6}{1 \text{ mol Si}_2\text{H}_6}$$
$$\times \frac{2 \text{ atoms Si}}{1 \text{ molecule Si}} = 2.23 \times 10^{22} \text{atom Si}$$

1.38 **a)** As suggested in the hint, look at *differences* in charge. List the droplets in order of increasing charge and compute the amount by which each differs in charge from its predecessor

Droplet No.	Charge/ 10^{-19} C	Difference/ 10^{-19} C	No. e^- on droplet
1	6.563	—	4
2	8.204	1.641	5
3	11.50	3.296	7
4	13.13	1.63	8
5	16.48	3.35	10
6	18.08	1.60	11
7	19.71	1.63	12
8	22.89	3.18	14
9	26.18	3.29	16

Moving down the list, the charges on the droplets increase by either 1 or 2 times 1.64×10^{-19} C. This suggests that a fundamental unit of charge exists and is approximately equal to 1.64×10^{-19} C. Dividing this quantity into the nine observed charges gives results that are all very close to whole numbers. Take these whole numbers to equal the number of the fundamental units of charge (electrons) on each droplet.

b) Divide the observed charge on each droplet by the apparent number of electrons on that droplet (that is, divide column 2 by column 4 in the preceding):

1. 1.6407×10^{-19}C
2. 1.6408×10^{-19}C
3. 1.6429×10^{-19}C
4. 1.6413×10^{-19}C
5. 1.6480×10^{-19}C
6. 1.6436×10^{-19}C
7. 1.6425×10^{-19}C
8. 1.6350×10^{-19}C
9. 1.6363×10^{-19}C

The average of the nine values is 1.641×10^{-19}C.

c) In the preceding, the least difference among the charges on the nine droplets has been taken as the "quantum of electrical charge." The actual quantum of charge however might be some fraction ($\frac{1}{2}, \frac{1}{3}, \frac{1}{4}$, etc.) of this value. One can confirm that 1.641×10^{-19}C is truly the charge on a single electron only by extensive search for droplets having lesser charge.

1.40

$$\text{Density of neutron star} = \frac{\text{mass}}{\text{volume}} = \frac{6.0\times10^{56}\times1.675\times10^{-24}\text{g}}{\left(\frac{4}{3}\right)\pi\left(20\times10^{5}\text{cm}\right)^{3}} = 3.0\times10^{13}\text{g cm}^{-3}$$

$$\text{Mass of } ^{232}\text{Th nucleus} = 142\left(1.675\times10^{-24}\text{ kg}\right) + 90\left(1.673\times10^{-24}\text{ g}\right) = 3.884\times10^{-22}\text{ g}$$

$$\text{Density of } ^{232}\text{Th nucleus} = \frac{3.884\times10^{-22}\text{g}}{\left(\frac{4}{3}\right)\pi\left(9.1\times10^{-13}\text{cm}\right)^{3}} = 1.2\times10^{14}\text{g cm}^{-3}$$

This is four times larger than the density of a neutron star.

1.42 Let A_1 be the *fractional* abundance of ^{85}Rb and let A_1 be the fractional abundance of ^{87}Rb.

Then

$$A_1\,(84.9117) + A_2(86.9092) = 85.4678 \text{ and } A_1 + A_2 = 1$$

Solving gives $A_1 = 0.7216$. The percentage of ^{85}Rb is 72.16%; ^{87}Rb is 27.84%.

Chapter 2

Chemical Formulas, Chemical Equations, and Reaction Yields

2.2 Find the mass fractions by dividing the mass of each element present in a mole of acetaminophen (N-acetyl-*p*-aminophenol, $C_8H_9NO_2$) by the molar mass

$$f_C = \frac{8 \times 12.011 \text{ g mol}^{-1}}{151.165 \text{ g mol}^{-1}} = 0.6356 \qquad f_H = \frac{9 \times 1.00794 \text{ g mol}^{-1}}{151.165 \text{ g mol}^{-1}} = 0.06001$$

$$f_N = \frac{1 \times 14.0067 \text{ g mol}^{-1}}{151.165 \text{ g mol}^{-1}} = 0.09266 \qquad f_O = \frac{2 \times 15.9994 \text{ g mol}^{-1}}{151.165 \text{ g mol}^{-1}} = 0.2117$$

The mass percentages equal the mass fractions multiplied by 100%.

2.4 Save work by *estimating* the fluorine content of each compound. Thus, HF is certainly the compound richest in fluorine by mass because the only other atom in its formula is the very light hydrogen atom—there is only 1 unit of non-fluorine mass per fluorine atom. The non-fluorine mass per fluorine atom in C_6HF_5 is about (6 × 12)/5 ≈ 14; in BrF it is 79.9; in UF_6 it is 238/6 ≈ 40. The desired order is therefore BrF < UF_6 < C_6HF_5 < HF.

2.6 The pharmacist mixes 286 g of Na_2CO_3 with 150 g of $C_2H_5NO_2$, using water as a mixing agent. After all the water is driven away, the mixture weighs 436 g. The mass of carbon from the Na_2CO_3 is (12.011/105.988) × 286 g; the mass of carbon from the $C_2H_5NO_2$ is (2 × 12.011/75.067) × 150 g where the 105.988 and 75.067 are the respective molecular masses of the compounds. The mass of carbon in the mixture is the sum of these two masses. It equals 80.412 g. The mass percentage of carbon is this mass divided by 436 g and multiplied by 100%. It is 18.4%.

2.8 Imagine 100.0 g of bromoform. The mass of bromine is 94.85 g, the mass of hydrogen is 0.40 g, and the mass of carbon is 4.75 g. Convert each of these masses to chemical amount by dividing by the molar mass of the element: there are 1.18705 mol of Br, 0.39685 mol of H, and 0.39547 mol of C. (Nonsignificant figures appear in these intermediate values for the sake of greater precision in the final result.) The three chemical amounts stand in the ratio of 2.99 to 1.003 to 1. Within the precision of the data this ratio is 3 to 1 to 1. The empirical formula is Br_3HC.

2.10 Imagine 100.0000 g of the compound. This sample contains 1.6907 g of O and 98.3093 g of Cs. The chemical amount of oxygen is its mass divided by its molar mass; it equals 0.10567 mol. The chemical amount of cesium is 0.73969 mol. The ratio of the chemical amounts is 7.000 to 1, making the empirical formula Cs_7O.

2.12 The empirical formulas of the five compounds are

A. CO_2, B. CO_1, C. C_4O_3, D. C_3O_2, E. C_5O_2

All five of these compositions exist. The first two are well-known. The third is mellitic anhydride (molecular formula $C_{12}O_9$). The last two are carbon suboxides having molecular formulas identical to their empirical formulas.

2.14 Compare the chemical amounts of the three elements in the compounds

$$n_{\text{Ca}} = n_{\text{CaO}} = \frac{2.389\text{ g}}{56.0774\text{ g mol}^{-1}} = 0.04260\text{ mol}$$

$$n_{\text{C}} = n_{\text{CO}_2} = \frac{1.876\text{ g}}{44.010\text{ g mol}^{-1}} = 0.04263\text{ mol}$$

$$n_{\text{N}} = n_{\text{NO}_2} = \frac{3.921\text{ g}}{46.0055\text{ g mol}^{-1}} = 0.08523\text{ mol}$$

Divide each of these amounts by the smallest. The resulting ratios are 1 : 1.001 : 2.001. This means that the empirical formula is $CaCN_2$

2.16 Two gases held at the same conditions contain the same number of molecules per unit volume. This means that the ratio of molar masses of these two gases equals the ratio of their densities:

$$\frac{\text{M}_{\text{P}_n}}{\text{M}_{\text{N}_2}} = \frac{2.7\text{ g L}^{-1}}{0.62\text{ g L}^{-1}} \text{ hence } \text{M}_{\text{P}_n} = \left(\frac{2.7}{0.62}\right)\left(28.014\text{ g mol}^{-1}\right) = 1.2\times10^2\text{ g mol}^{-1}$$

The molar mass of atomic P is only 30.97 g mol^{-1}. The P_4 molecule however has a molar mass of 124 g mol^{-1}, which equals 1.2×10^2 g mol^{-1} within the precision of the data (two significant digits). Conclude that there are four P atoms per molecule under these conditions.

2.18 **a)** The relative molecular mass of the binary compound is (2.53) (28.013) = 70.9.

b) Each molecule of the compound has one N atom because

$$\frac{8.21\text{ g}}{70.9\text{ g mol}^{-1}} = 0.116 \text{ mol compound contains } \frac{1.62\text{ g}}{14.007\text{ g mol}^{-1}} = 0.116\text{ mol N}$$

c) The maximum possible value of the relative atomic mass of the other element is 70.9 – 14.0 = 56.9.

d) Divide 56.9 by the positive integers: 56.9/2 = 28.5; 56.9/3 = 19.0; 56.9/4 = 24.2, and so forth. Compare the results to the known relative atomic masses. The closet fit is F (relative atomic mass 19.0).

e) NF_3

2.20 **a)** $2\text{Al} + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2$

b) $4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$

c) $2\text{Fe} + \text{O}_2 + 2\text{H}_2\text{O} \rightarrow 2\text{Fe}(\text{OH})_2$

d) $2\text{HSbCl}_4 + 3\text{H}_2\text{S} \rightarrow \text{Sb}_2\text{S}_3 + 8\text{HCl}$

e) $2\text{Al} + \text{Cr}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2\text{Cr}$

f) $\text{XeF}_4 + 2\text{H}_2\text{O} \rightarrow \text{Xe} + \text{O}_2 + 4\text{HF}$

g) $(\text{NH}_4)_2\text{Cr}_2\text{O}_7 \rightarrow \text{N}_2 + \text{Cr}_2\text{O}_3 + 4\text{H}_2\text{O}$

h) $\text{NaBH}_4 + 2\text{H}_2\text{O} \rightarrow \text{NaBO}_2 + 4\text{H}_2$

2.22 **a)**

$$m_{\text{Ca(OH)}_2} = \left(\frac{1.000\text{ g CaCO}_3}{100.09\text{ g mol}^{-1}}\right)\left(\frac{1\text{ mol Ca(OH)}_2}{1\text{ mol CaCO}_3}\right)\times 74.094\text{ g mol}^{-1} = 0.7403\text{ g}$$

b)

$$m_{CO_2} = \left(\frac{1.000 \text{ g } C_3H_8}{44.096 \text{ g mol}^{-1}}\right)\left(\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}\right) \times 44.010 \text{ g mol}^{-1} = 2.994 \text{ g}$$

c)

$$m_{Mg_2P_2O_7} = \left(\frac{1.000 \text{ g } MgNH_4PO_4}{137.315 \text{ g mol}^{-1}}\right)\left(\frac{1 \text{ mol } Mg_2P_2O_7}{2 \text{ mol } MgNH_4PO_4}\right) \times 222.55 \text{ g mol}^{-1} = 0.8104 \text{ g}$$

2.24 Whatever the reaction or series of reactions that gives $Pt_2C_{10}H_{18}N_2S_2O_6$, the chemical amount that forms cannot exceed half the chemical amount of platinum because the compound contains two Pt per molecule. Thus

$$m_{\text{cmpd(max.)}} = 1.406 \text{ g Pt} \times \frac{1 \text{ mol Pt}}{195.08 \text{ g Pt}} \times \frac{1 \text{ mol } Pt_2C_{10}H_{18}N_2S_2O_6}{2 \text{ mol Pt}}$$
$$\times \frac{716.55 \text{ g } Pt_2C_{10}H_{18}N_2S_2O_6}{1 \text{ mol } Pt_2C_{10}H_{18}N_2S_2O_6} = 2.582 \text{ g } Pt_2C_{10}H_{18}N_2S_2O_6$$

2.26 Use the balanced equation $Si_4H_{10} + \frac{13}{2}O_2 \rightarrow 4SiO_2 + 5H_2O$

$$m_{SiO_2} = 25.0 \text{ cm}^3 \text{ } Si_4H_{10} \times \frac{0.825 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ mol}}{122.42 \text{ g}} \times \frac{4 \text{ mol } SiO_2}{1 \text{ mol } Si_4H_{10}} \times \frac{60.0843 \text{ g } SiO_2}{1 \text{ mol } SiO_2}$$
$$= 40.5 \text{ g } SiO_2$$

2.28

$$m_{CS_2} = 67.2 \text{ g S} \times \frac{1 \text{ mol S}}{32.066 \text{ g S}} \times \frac{1 \text{ mol } CS_2}{4 \text{ mol S}} \times \frac{76.143 \text{ g } CS_2}{1 \text{ mol } CS_2} = 39.9 \text{ g } CS_2$$

2.30

$$m_{Ca_3(PO_4)_2} = 69.8 \text{ g } P_4 \times \frac{1 \text{ mol}}{123.895 \text{ g}} \times \frac{2 \text{ mol } Ca^3(PO_4)_2}{1 \text{ mol } P_4} \times \frac{310.18 \text{ g}}{1 \text{ mol}} = 349 \text{ g } Ca_3(PO_4)_2$$
$$m_{CaSiO_3} = 69.8 \text{ g } P_4 \times \frac{1 \text{ mol}}{123.895 \text{ g}} \times \frac{6 \text{ mol } CaSiO_3}{1 \text{ mol } P_4} \times \frac{116.16 \text{ g}}{1 \text{ mol}} = 393 \text{ g } CaSiO_3$$

2.32 **a)** Let M_A represent the molar mass of element A. Then $M_A + 3(126.90)$ g mol^{-1} is the molar mass of AI_3 and $M_A + 3(35.453)$ g mol^{-1} is the molar mass of ACl_3. Use the stoichiometry of the equation

to express the mass of ACl_3 obtained in terms of the mass of AI_3 treated

$$0.3776 \text{ g} = 0.8000 \text{ g } AI_3 \times \frac{1 \text{ mol } AI_3}{(M_A + 380.7) \text{ g } AI_3} \times \frac{1 \text{ mol } ACl_3}{1 \text{ mol } AI_3} \times \frac{(M_A + 106.36) \text{ g } ACl_3}{1 \text{ mol } ACI_3}$$

Which becomes

$$0.8000(M_A + 106.36) = 0.3776(M_A + 380.7) \quad \text{from which} \quad M_A = 138.9 \text{ g mol}^{-1}$$

b) The element is lanthanum, La.

2.34 Express the amount of hydrogen that is evolved as a chemical amount

$$n_{H_2} = \frac{0.738 \text{ g}}{2.016 \text{ g mol}^{-1}} = 0.366 \text{ mol } H_2$$

Let the mass of iron in the original mixture be represented by x. Then

$$n_{Fe} = \frac{x \text{ g Fe}}{55.845 \text{ g mol}^{-1}} \quad \text{and} \quad n_{Al} = \frac{(9.62 - x)\text{g Al}}{26.982 \text{ g mol}^{-1}}$$

The 0.366 mol of H_2 is evolved by these molar amounts of Fe and Al in reaction with excess HCl according to the equations given in the problem. Determine the molar amount of H_2 that each metal produces and add them

$$\frac{x \text{ g Fe}}{55.845 \text{ g mol}^{-1}}\left(\frac{1 \text{ mol } H_2}{1 \text{ mol Fe}}\right) + \frac{(9.62 - x) \text{ g Al}}{26.982 \text{ g mol}^{-1}}\left(\frac{3 \text{ mol } H_2}{2 \text{ mol Al}}\right) = 0.366 \text{ mol } H_2$$

$$-0.03769\, x + 0.5348 = 0.366 \quad \text{from which } x = 4.48 \text{ g Fe}$$

2.36 **a)** $NH_3 + CH_4 \rightarrow HCN + 3\ H_2$

b) The reaction system contains 700.0 g at the start, distributed as follows

$$\frac{500.0 \text{ g } CH_4}{16.043 \text{ g mol}^{-1}} = 31.17 \text{ mol } CH_4 \quad \text{and} \quad \frac{200.0 \text{ g } NH_3}{17.031 \text{ g mol}^{-1}} = 11.74 \text{ mol } NH_3$$

NH_3 is the limiting reactant. After the reaction goes to completion the system contains

$$m_{HCN} = 11.74 \text{ mol } NH_3 \times \frac{1 \text{ mol HCN}}{1 \text{ mol } NH_3} \times \frac{27.026 \text{ g}}{\text{mol HCN}} = 317.3 \text{ g HCN}$$

$$m_{H_2} = 11.74 \text{ mol } NH_3 \times \frac{3 \text{ mol } H_2}{1 \text{ mol } NH_3} \times \frac{2.0158 \text{ g}}{\text{mol } H_2} = 71.0 \text{ g } H_2$$

$$m_{CH_4} = (31.17 - 11.74 \text{ mol } CH_4) \times \frac{16.043 \text{ g}}{\text{mol } CH_4} = 311.7 \text{ g } CH_4$$

The sum of these three masses is 700.0 g, the same total mass as before the reaction.

2.38 The theoretical yield of $TiCl_4$ is

$$m_{TiCl_4} = 7.39 \text{ kg } TiO_2 \times \frac{1 \text{ kmol } TiO_2}{79.88 \text{ kg } TiO_2} \times \frac{1 \text{ kmol } TiCl_4}{1 \text{ kmol } TiCl_2} \times \frac{189.69 \text{ kg } TiCl_4}{1 \text{ kmol } TiCl_4} = 17.55 \text{ kg}$$

The percentage yield is 14.24 kg/17.55 kg × 100% = 81.1%.

2.40 Do the calculations to five significant figures, the precision of the tabulated atomic mass of tungsten. There is 0.43134 mol of W in a 100.000 g sample of the white oxide, and 1.29395 mol of O. The ratio of these numbers is 2.9998—the empirical formula is WO_3. 100.000 g of the blue oxide contains 0.43975 mol W and 1.19709 mol of O. The ratio of these two amounts is 2.7222. This turns out to equal the ratio of 49 to 18, within the precision of the data. Hence the formula $W_{18}O_{49}$ is a correct answer. The blue oxide is really a nonstoichiometric compound, however.

2.42 Consider a 100.00 g sample of this binary compound. It contains 78.06 g of Ni and 21.94 g of O. This is 78.06 g/58.69 g mol^{-1} = 1.330 mol of Ni and 21.94 g/15.9994 g mol^{-1} = 1.371 mol of O. The ratio of these two chemical amounts is 1.031 to 1. If the data are truly precise to four significant figures, the compound is almost certainly a nonstoichiometric compound. The "almost" appears because " $Ni_{1000}O_{1031}$ " is a conceivable stoichiometric formulation. These subscripts are whole numbers, but they are hardly small whole numbers.

2.44 **a)** The balanced equations for the conversion of cyanuric acid to isocyanuric acid and the reaction of isocyanuric acid with nitrogen dioxide are

$$\boxed{C_3N_3(OH)_3 \longrightarrow 3\ HNCO} \quad \text{and} \quad \boxed{8\ HNCO + 6\ NO_2 \longrightarrow 7\ N_2 + 8\ CO_2 + 4\ H_2O}$$

Balancing by inspection in the second equation works but requires some care. Assign 1 as the coefficient of HNCO. Then, focus on C and H because these elements occur in only one compound on each side of the equation. If HNCO on the left has a coe_cient of 1, CO_2 on the right must have a coefficient of 1 and H_2O on the right must have a coefficient of $\frac{1}{2}$ to achieve balance in C and H. These two coefficients on the right imply a total of $2\frac{1}{2}$ mol of O on the right. The "1 HNCO" on the left supplies only 1 mol of O, and its coeffcient must not be changed. The NO_2 must supply the other $\frac{3}{2}$ mol of O. To do this, its coefficient must be $\frac{3}{4}$. The left side now has $1 + \frac{3}{4} = \frac{7}{4}$ mol of N. The coe_cient of N_2 on the right must therefore equal $\frac{7}{8}$ because $\frac{7}{8} \times 2 = \frac{7}{4}$ (the 2 comes from thesubscript in N_2). Multiplying all five coeffcients by 8 eliminates fractional coefficients.

b) Use a series of unit-factors constructed from the molar masses of the compounds and the coefficients of the two balanced equations

2.46 The only product of the reaction that contains nitrogen is m-toluidine; the only reactant that contains nitrogen is 3′-methylphthalanilic acid. It follows that the mass of nitrogen in the 3′-methylphthalanilic acid must equal the mass of nitrogen in the m-toluidine. The m-toluidine (empirical formula C_7H_9N) is 13.1% nitrogen by mass (calculated as in problem 2.1). The 5.23 g of m-toluidine therefore contains 0.685 g of nitrogen. The 3′-methylphthalanilic acid contains 5.49% nitrogen by mass (as given in the problem). The issue thus becomes _nding the mass of 3′-methylphthalanilic acid that contains 0.685 g of nitrogen. Let this mass equal x. Then $0{:}0549x = 0{:}685$ g. Solving gives x equalto 12.5 g. This analysis is equivalent to the following:

$$m = 5.23\text{ g toluidine} \times \left(\frac{13.1\text{ g N}}{100\text{ g toluidine}}\right)\left(\frac{100\text{ g 3′-methyl...}}{5.49\text{ g N}}\right) = \boxed{12.5\text{ g 3′-methyl...}}$$

2.48 **a)** Write an unbalanced equation to represent what the statement of the problem reveals about the process

$$C_{12}H_{22}O_{11} + O_2 \longrightarrow C_6H_8O_7 + H_2O$$

Balance this equation as to carbon by inserting the coe_cient 2 in front of the citric acid. Then balance the H atoms by putting a 3 in front of the water (of the 22 H's on the left, 16 appear in the citric acid, and the rest appear in the water). Next, consider the oxygen. The right side has (2 × 7) + (3 × 1) = 17 O's. On the left side, the sucrose furnishes 11 O's so the remaining 6 must come from 3 molecules of oxygen. The balanced equation is

$$\boxed{C_{12}H_{22}O_{11} + 3\ O_2 \longrightarrow 2\ C_6H_8O_7 + 3\ H_2O}$$

b) The balanced equation provides the information to write the second unit-factor in the following

$$m_{C_6H_8O_7} = 15.0\text{ kg sucrose} \times \left(\frac{1\text{ kmol sucrose}}{342.3\text{ kg sucrose}}\right)\left(\frac{2\text{ kmol citric acid}}{1\text{ kmol sucrose}}\right)\left(\frac{192.12\text{ kg citric acid}}{1\text{ kmol citric acid}}\right)$$
$$= \boxed{16.8\text{ kg citric acid}}$$

Tip. Save some effort by creating and using unit-factors such as "1 kilomole sucrose / 342.3 kg sucrose." Also, only *part* of the balanced equation is needed, namely the 2 : 1 molar ratio of citric acid to sucrose. The O_2 and H_2O could have been left out.

2.50 **a)** Compute the number of moles of XBr_2 that is present and recognize that two moles of AgBr appear for every one mole of XBr_2 present in the 5.000 g sample. This fact appears in the second unit-factor in the following

$$n_{XBr_2} = 1.0198 \text{ g AgBr} \times \left(\frac{1 \text{ mol AgBr}}{187.77 \text{ g AgBr}}\right)\left(\frac{1 \text{ mol } XBr_2}{2 \text{ mol AgBr}}\right) = 0.002716 \text{ mol } XBr_2$$

The molar mass M of any substance equals its mass divided by its chemical amount

$$M_{XBr_2} = \frac{m_{XBr_2}}{n_{XBr_2}} = \frac{0.5000 \text{ g}}{0.002716 \text{ mol}} = 184.1 \text{ g mol}^{-1}$$

b) The relative atomic mass of element X equals the relative molecular mass of the compound minus the contribution of the bromine:

relative atomic mass of X = 184.1 − 2(79.9) = $\boxed{24:3}$

Checking the atomic masses in the periodic table shows that X is magnesium, $\boxed{\text{Mg}}$.

2.52 In step 1 of the Solvay process, 1 mol of NH_3 (along with 1 mol of H_2O) combines with 1 mol of CO_2, holding it for attack by NaCl. This attack (step 2) gives 1 mol of $NaHCO_3$ while driving off NH_4Cl as a by-product. Heating the 1 mol of $NaHCO_3$ (step 3) then gives $\frac{1}{2}$ mol of the product Na_2CO_3. For each mole of NH_3 that is put in, $\frac{1}{2}$ mol of Na_2CO_3 comes out. The following set-up uses this fact. It also uses the fact that a metric ton (1000 kg) is a megagram (1 Mg, a million grams) and the fact that a megamole (Mmol) is 10^6 (one million) moles.

$$m_{Na_2CO_3} = 1 \text{ metric ton } NH_3 \times \left(\frac{1 \text{ Mg}}{1 \text{ metric ton}}\right)\left(\frac{1 \text{ Mmol } NH_3}{17.03 \text{ Mg } NH_3}\right)\left(\frac{1/2 \text{ Mmol } Na_2CO_3}{1 \text{ Mmol } NH_3}\right)$$
$$\times \left(\frac{105.99 \text{ Mg } Na_2CO_3}{1 \text{ Mmol } Na_2CO_3}\right)\left(\frac{1 \text{ metric ton}}{1 \text{ Mg}}\right) = \boxed{3.11 \text{ metric ton } Na_2CO_3}$$

2.54 Assume that the limestone raw material is pure calcium carbonate ($CaCO_3$). Add the three steps listed in the problem. The CaO cancels out between the _rst and second steps and the CaC_2 cancels out between the second and third. The result

$$CaCO_3 + 3\ C + 2\ H_2O \longrightarrow C_2H_2 + Ca(OH)_2 + CO + CO_2$$

is balanced. It indicates that the over-all process generates 1 mol of C_2H_2 for every 1 mol of $CaCO_3$ that is put in. Use this fact as a unit-factor to obtain the theoretical yield of (C_2H_2) (acetylene). It is not necessary to compute the theoretical yields of CaO (lime) and CaC_2 (calcium carbide) formed and subsequently consumed on the way to the final product. The following set-up uses two additional facts: a metric ton is 10^6 g, also called a megagram (Mg), and a megamole (Mmol) is 10^6 moles.

$$m_{C_2H_2} = 10.0 \text{ Mg } CaCO_3 \times \left(\frac{1 \text{ Mmol } CaCO_3}{100.1 \text{ Mg } CaCO_3}\right)\left(\frac{1 \text{ Mmol } C_2H_2}{1 \text{ Mmol } CaCO_3}\right)\left(\frac{26.03 \text{ Mg } C_2H_2}{1 \text{ Mmol } C_2H_2}\right)$$
$$= 2.60 \text{ Mg } C_2H_2$$

The percent yield equals the actual yield divided by the theoretical yield and multiplied by 100%:

$$\text{percent yield } C_2H_2 = \frac{2.32 \text{ Mg } C_2H_2}{2.60 \text{ Mg } C_2H_2} \times 100\% = 89.2\%$$

Tip. "Overall" (sum of steps) chemical equations sometimes fool people. The one used here provides a correct molar relationship between the $CaCO_3$ and C_2H_2. It does *not* indicate that $CaCO_3$ reacts directly with C and H_2O.