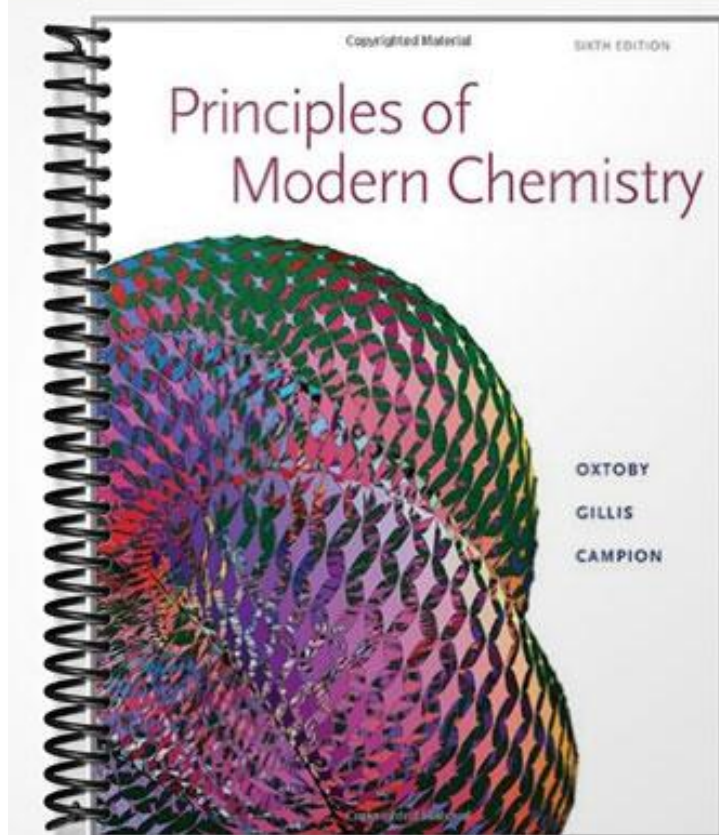


# SOLUTIONS MANUAL



Instructor's Manual TO ACCOMPANY  
*Principles of Modern Chemistry*, 6th edition

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# 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

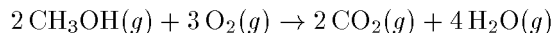
The ratios of each mass of chlorine to the smallest mass of chlorine are

$$\begin{aligned} 0.3856/0.3856 &= 1.0000 = 2 : 2 & 0.7715/0.3856 &= 2.0008 = 4 : 2 \\ 0.9643/0.3856 &= 2.5008 = 5 : 2 & 1.1570/0.3856 &= 3.0005 = 6 : 2 \end{aligned}$$

The formulas are  $\text{WCl}_2$ ,  $\text{WCl}_4$ ,  $\text{WCl}_5$ , and  $\text{WCl}_6$ .

1.12 The *only* products are gaseous  $\text{N}_2$  and gaseous  $\text{H}_2$ . From the formula of the starting compound there are twice as many molecules of  $\text{H}_2$  as of  $\text{N}_2$  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



2.0 L of  $\text{CO}_2$  and 4.0 L of  $\text{H}_2\text{O}$  are produced from 2.0 L of  $\text{CH}_3\text{OH}$ , according to the law of combining volumes (and under the assumption that the reaction goes to completion as written).

1.16 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

$$\begin{aligned} A &= A_1p_1 + A_2p_2 + \cdots + A_np_n \\ A_{\text{Ne}} &= (0.9000)(19.99212) + (0.0027)(20.99316) + (0.0973)(21.99132) = 20.19 \end{aligned}$$

1.18 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

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

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

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

1.20 a) Promethium has an atomic number of 61; the ratio of the number of neutrons to protons in  $^{145}\text{Pm}$  is  $(145 - 61)/61 = 1.377$ .

b) A neutral atom of Pm has 61 electrons.

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

1.24 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  $1.64 \times 10^{-19}$  C. This suggests that a fundamental unit of charge exists and is approximately equal to  $1.64 \times 10^{-19}$  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):

$$\begin{array}{lll}
 1. & 1.6407 \times 10^{-19} \text{ C} & 4. & 1.6413 \times 10^{-19} \text{ C} & 7. & 1.6425 \times 10^{-19} \text{ C} \\
 2. & 1.6408 \times 10^{-19} \text{ C} & 5. & 1.6480 \times 10^{-19} \text{ C} & 8. & 1.6350 \times 10^{-19} \text{ C} \\
 3. & 1.6429 \times 10^{-19} \text{ C} & 6. & 1.6436 \times 10^{-19} \text{ C} & 9. & 1.6363 \times 10^{-19} \text{ C}
 \end{array}$$

The average of the nine values is  $1.641 \times 10^{-19}$  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  $1.641 \times 10^{-19}$  C is truly the charge on a single electron only by extensive search for droplets having lesser charge.

1.26

$$\text{Density of neutron star} = \frac{\text{mass}}{\text{volume}} = \frac{6.0 \times 10^{56} \times 1.675 \times 10^{-24} \text{ g}}{\left(\frac{4}{3}\right)\pi(20 \times 10^5 \text{ cm})^3} = 3.0 \times 10^{13} \text{ g cm}^{-3}$$

$$\text{Mass of } {}^{232}\text{Th nucleus} = 142(1.675 \times 10^{-24} \text{ kg}) + 90(1.673 \times 10^{-24} \text{ g}) = 3.884 \times 10^{-22} \text{ g}$$

$$\text{Density of } {}^{232}\text{Th nucleus} = \frac{3.884 \times 10^{-22} \text{ g}}{\left(\frac{4}{3}\right)\pi(9.1 \times 10^{-13} \text{ cm})^3} = 1.2 \times 10^{14} \text{ g cm}^{-3}$$

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

1.28 Let  $A_1$  be the *fractional* abundance of  ${}^{85}\text{Rb}$  and let  $A_2$  be the fractional abundance of  ${}^{87}\text{Rb}$ . Then

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

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



## Chapter 2

# Chemical Formulas, Chemical Equations, and Reaction Yields

- 2.2** Avogadro's number of  $^{12}\text{C}$  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

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

- 2.4** 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.
- 2.6** There are Avogadro's number of gold atoms in a mole of gold, each with a diameter of  $2.88 \times 10^{-10}$  m. The length of the line is  $(6.022 \times 10^{23})(2.88 \times 10^{-10} \text{ m}) = 1.73 \times 10^{14}$  m.
- 2.8** Express the amounts of each sample in the same unit of mass. In the case of the  $\text{SF}_4$ , convert from the given number of moles to grams using the molar mass. In the cases of the  $\text{Cl}_2\text{O}_7$  and Ar, convert from the given number of particles to chemical amount and from there to mass. The amount of  $\text{CH}_4$  is already in grams. The results are

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

### 2.10

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

The cost to three significant figures is \$2480.

### 2.12

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

- 2.14** 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

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

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

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

- 2.16** 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  $(6 \times 12)/5 \approx 14$ ; in BrF it is 79.9; in  $UF_6$  it is  $238/6 \approx 40$ . The desired order is therefore  $BrF < UF_6 < C_6HF_5 < HF$ .
- 2.18** 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  $(12.011/105.988) \times 286$  g; the mass of carbon from the  $C_2H_5NO_2$  is  $(2 \times 12.011/75.067) \times 150$  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%.
- 2.20** 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  $Br_3HC$ .
- 2.22** 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  $Cs_7O$ .
- 2.24** The empirical formulas of the five compounds are

A.  $CO_2$ , B. CO, 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.

- 2.26** Compare the chemical amounts of the three elements in the compounds

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

$$n_C = n_{CO_2} = \frac{1.876 \text{ g}}{44.010 \text{ g mol}^{-1}} = 0.04263 \text{ mol}$$

$$n_N = n_{NO_2} = \frac{3.921 \text{ g}}{46.0055 \text{ g mol}^{-1}} = 0.08523 \text{ mol}$$



**2.38** Use the balanced equation  $\text{Si}_4\text{H}_{10} + \frac{13}{2}\text{O}_2 \rightarrow 4\text{SiO}_2 + 5\text{H}_2\text{O}$

$$\begin{aligned} m_{\text{SiO}_2} &= 25.0 \text{ cm}^3 \text{ Si}_4\text{H}_{10} \times \frac{0.825 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ mol}}{122.42 \text{ g}} \times \frac{4 \text{ mol SiO}_2}{1 \text{ mol Si}_4\text{H}_{10}} \times \frac{60.0843 \text{ g SiO}_2}{1 \text{ mol SiO}_2} \\ &= 40.5 \text{ g SiO}_2 \end{aligned}$$

**2.40**

$$m_{\text{CS}_2} = 67.2 \text{ g S} \times \frac{1 \text{ mol S}}{32.066 \text{ g S}} \times \frac{1 \text{ mol CS}_2}{4 \text{ mol S}} \times \frac{76.143 \text{ g CS}_2}{1 \text{ mol CS}_2} = 39.9 \text{ g CS}_2$$

**2.42**

$$\begin{aligned} m_{\text{Ca}_3(\text{PO}_4)_2} &= 69.8 \text{ g P}_4 \times \frac{1 \text{ mol}}{123.895 \text{ g}} \times \frac{2 \text{ mol Ca}_3(\text{PO}_4)_2}{1 \text{ mol P}_4} \times \frac{310.18 \text{ g}}{1 \text{ mol}} = 349 \text{ g Ca}_3(\text{PO}_4)_2 \\ m_{\text{CaSiO}_3} &= 69.8 \text{ g P}_4 \times \frac{1 \text{ mol}}{123.895 \text{ g}} \times \frac{6 \text{ mol CaSiO}_3}{1 \text{ mol P}_4} \times \frac{116.16 \text{ g}}{1 \text{ mol}} = 393 \text{ g CaSiO}_3 \end{aligned}$$

**2.44** a) Let  $\mathcal{M}_A$  represent the molar mass of element A. Then  $\mathcal{M}_A + 3(126.90) \text{ g mol}^{-1}$  is the molar mass of  $\text{AlI}_3$  and  $\mathcal{M}_A + 3(35.453) \text{ g mol}^{-1}$  is the molar mass of  $\text{AlCl}_3$ . Use the stoichiometry of the equation to express the mass of  $\text{AlCl}_3$  obtained in terms of the mass of  $\text{AlI}_3$  treated

$$0.3776 \text{ g} = 0.8000 \text{ g AlI}_3 \times \frac{1 \text{ mol AlI}_3}{(\mathcal{M}_A + 380.7) \text{ g AlI}_3} \times \frac{1 \text{ mol AlCl}_3}{1 \text{ mol AlI}_3} \times \frac{(\mathcal{M}_A + 106.36) \text{ g AlCl}_3}{1 \text{ mol AlCl}_3}$$

Which becomes

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

b) The element is lanthanum, La.

**2.46** Express the amount of hydrogen that is evolved as a chemical amount

$$n_{\text{H}_2} = \frac{0.738 \text{ g}}{2.016 \text{ g mol}^{-1}} = 0.366 \text{ mol H}_2$$

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

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

The 0.366 mol of  $\text{H}_2$  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  $\text{H}_2$  that each metal produces and add them

$$\begin{aligned} \frac{x \text{ g Fe}}{55.845 \text{ g mol}^{-1}} \left( \frac{1 \text{ mol H}_2}{1 \text{ mol Fe}} \right) + \frac{(9.62 - x) \text{ g Al}}{26.982 \text{ g mol}^{-1}} \left( \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \right) &= 0.366 \text{ mol H}_2 \\ -0.03769x + 0.5348 &= 0.366 \quad \text{from which } x = 4.48 \text{ g Fe} \end{aligned}$$

**2.48** a)  $\text{NH}_3 + \text{CH}_4 \rightarrow \text{HCN} + 3\text{H}_2$

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

$$\frac{500.0 \text{ g CH}_4}{16.043 \text{ g mol}^{-1}} = 31.17 \text{ mol CH}_4 \quad \text{and} \quad \frac{200.0 \text{ g NH}_3}{17.031 \text{ g mol}^{-1}} = 11.74 \text{ mol NH}_3$$

$\text{NH}_3$  is the limiting reactant. After the reaction goes to completion the system contains

$$m_{\text{HCN}} = 11.74 \text{ mol NH}_3 \times \frac{1 \text{ mol HCN}}{1 \text{ mol NH}_3} \times \frac{27.026 \text{ g}}{\text{mol HCN}} = 317.3 \text{ g HCN}$$

$$m_{\text{H}_2} = 11.74 \text{ mol NH}_3 \times \frac{3 \text{ mol H}_2}{1 \text{ mol NH}_3} \times \frac{2.0158 \text{ g}}{\text{mol H}_2} = 71.0 \text{ g H}_2$$

$$m_{\text{CH}_4} = (31.17 - 11.74 \text{ mol CH}_4) \times \frac{16.043 \text{ g}}{\text{mol CH}_4} = 311.7 \text{ g CH}_4$$

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

**2.50** The theoretical yield of  $\text{TiCl}_4$  is

$$m_{\text{TiCl}_4} = 7.39 \text{ kg TiO}_2 \times \frac{1 \text{ kmol TiO}_2}{79.88 \text{ kg TiO}_2} \times \frac{1 \text{ kmol TiCl}_4}{1 \text{ kmol TiO}_2} \times \frac{189.69 \text{ kg TiCl}_4}{1 \text{ kmol TiCl}_4} = 17.55 \text{ kg}$$

The percentage yield is  $14.24 \text{ kg}/17.55 \text{ kg} \times 100\% = 81.1\%$ .

**2.52** 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  $\text{WO}_3$ . 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  $\text{W}_{18}\text{O}_{49}$  is a correct answer. The blue oxide is really a nonstoichiometric compound, however.

**2.54** 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  $78.06 \text{ g}/58.69 \text{ g mol}^{-1} = 1.330 \text{ mol}$  of Ni and  $21.94 \text{ g}/15.9994 \text{ g mol}^{-1} = 1.371 \text{ mol}$  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 “ $\text{Ni}_{1000}\text{O}_{1031}$ ” is a conceivable stoichiometric formulation. These subscripts are whole numbers, but they are hardly small whole numbers.

**2.56** The empirical formula of the compound is BCl based on

$$n_{\text{B}} = \frac{0.664 \text{ g B}}{10.811 \text{ g mol}^{-1}} = 0.0614 \text{ mol B} \quad n_{\text{Cl}} = \frac{(2.842 \text{ g} - 0.664 \text{ g}) \text{ Cl}}{35.453 \text{ g mol}^{-1}} = 0.0614 \text{ mol Cl}$$

To obtain the molecular formula, write

$$n_{\text{Cl}_2} = 0.0614 \text{ mol Cl} \times \left( \frac{1 \text{ mol Cl}_2}{2 \text{ mol Cl}} \right) = 0.0307 \text{ mol Cl}_2$$

From Avogadro's hypothesis, the chemical amount of the gaseous compound in the original sample was

$$n_{\text{compound}} = 0.0307 \text{ mol Cl}_2 \times \left( \frac{0.153 \text{ L compound}}{0.688 \text{ L Cl}_2} \right) = 0.00683 \text{ mol compound}$$

This allows computation of the molar mass

$$\mathcal{M}_{\text{compound}} = \frac{2.842 \text{ g}}{0.00683 \text{ mol}} = 416.3 \text{ g mol}^{-1}$$

The molar mass corresponding to the empirical formula BCl is only  $46.264 \text{ g mol}^{-1}$ . The ratio  $416.3/46.264$  equals 9.00, so the molecular formula is  $\text{B}_9\text{Cl}_9$ .

**2.58** The balanced chemical equation reads



The right-hand side of the equation is short 9 atom of C, which must be part of the phenylalanine molecule. Balancing amounts of H, N and O in the same way establishes the molecular formula of phenylalanine to be  $\text{C}_9\text{H}_{11}\text{NO}_2$ .

**2.60** Assume that all the carbon is evolved as  $\text{CH}_4$ . Then 1 mol  $\text{Al}_4\text{C}_3$  generates 3 mol  $\text{CH}_4$

$$m_{\text{CH}_4} = \frac{63.2 \text{ g Al}_4\text{C}_3}{143.96 \text{ g mol}^{-1}} \times \frac{3 \text{ mol CH}_4}{1 \text{ mol Al}_4\text{C}_3} \times \frac{16.043 \text{ g CH}_4}{1 \text{ mol CH}_4} = 21.1 \text{ g}$$

**2.62** Let  $x$  represent the mass of  $\text{SrCO}_3$  in the sample. Then  $0.800 - x$  is the mass of  $\text{BaCO}_3$ .

$$n_{\text{SrCO}_3} = \frac{x \text{ g}}{147.63 \text{ g mol}^{-1}} \quad \text{and} \quad n_{\text{BaCO}_3} = \frac{(0.800 - x) \text{ g}}{197.34 \text{ g mol}^{-1}}$$

$$\text{and} \quad n_{\text{CO}_2} = \frac{0.211 \text{ g CO}_2}{44.010 \text{ g mol}^{-1}} = 0.004794 \text{ mol}$$

Each mole of  $\text{SrCO}_3$  and  $\text{BaCO}_3$  generates one mole of  $\text{CO}_2$ . This means

$$\left( \frac{x}{147.63} + \frac{0.800 - x}{197.34} \right) \text{ mol} = 0.004794 \text{ mol} \quad \text{which gives} \quad x = 0.434 \text{ g}$$

The percentage of  $\text{SrCO}_3$  in the sample was  $0.434 \text{ g}/0.800 \text{ g} \times 100\% = 54.2\%$ .

**2.64** Assume that your car gets 20 miles per U.S. gallon. Then reducing driving by 20 miles per week would save you 1 gallon a week, or 52 gallons a year. In the metric system this is

$$V = 52 \text{ gallons} \times \left( \frac{3.785 \text{ L}}{1 \text{ U.S. gallon}} \right) = 200 \text{ L gasoline}$$

Gasoline floats on water, so its density is less than the density of water (which is  $1.0 \text{ g cm}^{-3}$ ). Take the density of gasoline to be  $0.9 \text{ g cm}^{-3}$ , or  $0.9 \text{ kg L}^{-1}$ . Your annual savings is then about 180 kg of gasoline. Approximate the chemical formula of gasoline as  $\text{C}_8\text{H}_{18}$  (octane). Then

$$n_{\text{gasoline}} \approx 180 \times 10^3 \text{ g C}_8\text{H}_{18} \times \left( \frac{1 \text{ mol}}{114.2 \text{ g C}_8\text{H}_{18}} \right) = 1.6 \times 10^3 \text{ mol}$$

Combustion of octane release 8 mol of  $\text{CO}_2$  per mole. The mass of  $\text{CO}_2$  not released is

$$m_{\text{CO}_2} \approx 1.6 \times 10^3 \text{ mol C}_8\text{H}_{18} \times \left( \frac{8 \text{ mol CO}_2}{1 \text{ mol C}_8\text{H}_{18}} \right) \times \left( \frac{44 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 5.7 \times 10^5 \text{ g} \approx 570 \text{ kg}$$

$$\approx 570 \text{ kg} \left( \frac{1 \text{ lb}}{0.4536 \text{ kg}} \right) = 1300 \text{ lb CO}_2$$

This is on the order of 1000 pounds, the amount that is mentioned in the problem. If your car gets 30 miles per gallon, then the same weekly 20 mile cut in its use would reduce your  $\text{CO}_2$  emissions by 2/3 of 1300 pounds.

- 2.66** Compute the theoretical yield of  $\text{KClO}_4$  for comparison to the observed yield of 3.00 g. Balance the chemical equation:  $4 \text{KClO}_3(s) \rightarrow 3 \text{KClO}_4(s) + \text{KCl}(s)$ . Then

$$\begin{aligned} m_{\text{KClO}_4} &= 4.00 \text{ g KClO}_3 \times \left( \frac{1 \text{ mol KClO}_3}{122.549 \text{ g KClO}_3} \right) \left( \frac{3 \text{ mol KClO}_4}{4 \text{ mol KClO}_3} \right) \left( \frac{138.549 \text{ g KClO}_4}{1 \text{ mol KClO}_4} \right) \\ &= 3.392 \text{ g KClO}_4 \end{aligned}$$

The percentage yield is  $(3.00/3.392) \times 100\% = 88.4\%$ .

- 2.68** The yield of the product is less than the theoretical, and this fact must be reckoned with. A very good way is to use the percentage yield to construct an additional unit-factor (the second factor in the following)

$$\begin{aligned} m_{\text{Si}} &= 125 \text{ g Si}_3\text{N}_4 \text{ isolated} \times \left( \frac{100 \text{ g Si}_3\text{N}_4 \text{ formed}}{95.0 \text{ g Si}_3\text{N}_4 \text{ isolated}} \right) \times \left( \frac{1 \text{ mol Si}_3\text{N}_4}{140.285 \text{ g Si}_3\text{N}_4 \text{ formed}} \right) \\ &\quad \times \left( \frac{3 \text{ mol Si}}{1 \text{ mol Si}_3\text{N}_4} \right) \times \left( \frac{28.086 \text{ g Si}}{1 \text{ mol Si}} \right) = 79.0 \text{ g Si} \end{aligned}$$