1.64. (See Example 1–10.) A white crystalline compound that contains only iodine, potassium, and oxygen is purified. A mass of 87.41 g of pure compound is found on analysis to contain 15.97 g of potassium and 51.83 g of iodine. What is the empirical formula of the compound?

1-64 87.41 g of the compound contains 51.83 g of I, 15.97 g of K, and, by difference, 19.61 g of O. Convert each of these masses to chemical amount by dividing by the proper elemental molar mass: 0.4084 mol I, 0.4085 mol K and 1.2257 mol O. The ratio of these chemical amounts is 1.000 to 1.000 to 3.001, and therefore the empirical formula is KIO_3.

1.20. Iodine (I) and fluorine (F) form a series of binary compounds with the following compositions:

Compound	Mass % I	Mass % F
1	86.979	13.021
2	69.007	30.993
3	57.191	42.089
4	48.829	51.171

(a) Compute in each case the mass of fluorine that combines with 1.0000 g of iodine.

(b) Show that these compounds satisfy the law of multiple proportions (see Problem 19) by figuring out small whole-number ratios among the four answers in part (a).

1-20 (a) In each case the mass of fluorine that combines with 1.0000 g of iodine is simply the mass percentage of fluorine divided by the mass percentage of iodine. This calculation is simplified by assuming each sample of the compounds to have a total mass of 100.000 g. The masses contributed by each element in the compounds are then easily computed.

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

(b) The law of multiple proportions involves the ratio of the numbers in the last column. Simply divide all four by the smallest. The results are: 1.0000 for compound 1; 3.0002 for compound 2; 5.0002 for compound 3; and 7.0004 for compound 4. These equal the small whole numbers 1, 3, 5, and 7 within the precision of the data.

2.27. Sodium nitrate ($NaNO_3$) decomposes, upon moderate heating, to yield sodium nitrite ($NaNO_2$) and gaseous oxygen. A 3.4671 g sample of *impure* sodium nitrate was heated until the evolution of oxygen ceased. The sample residue was subsequently found to have a mass of 2.9073 g. What percentage of the original impure sample was actually sodium nitrate? Assume that none of the impurities evolved gases.

2-27. The problem is one where we seem to start with information about the sodium nitrate, but in fact that is what we need to determine. We have a final and a beginning mass. This lets us get to the mass of the *oxygen* produced.

Reaction stoichiometry: $2\ NaNO_3 \rightarrow O_2 + 2\ NaNO_2$

Mass of oxygen produced: 3.4671 g – 2.9073 g = 0.5598 g.

Mass of sodium nitrate reacted:

$$0.5598 \times \frac{1 \text{ mol } O_2}{31.9988 \text{ g } O_2} \times \frac{2 \text{ mol } NaNO_3}{1 \text{ mol } O_2} \times \frac{84.9947 \text{ g } NaNO_3}{1 \text{ mol } NaNO_3} = 2.974 \text{ g } NaNO_3$$

For the purity of the sodium nitrate:

$$\frac{2.974 \text{ g } NaNO_3 \text{ reacted}}{3.4671 \text{ g sample}} \times 100\% = \mathbf{85.77\%}$$

2.38. Compute the maximum mass of AgCl that can form when 35.453 g of NaCl and 107.87 g of $AgNO_3$ are mixed in water solution.

2-38 **The $AgNO_3$ and NaCl react in a 1 to 1 molar ratio to give AgCl, which precipitates, and $NaNO_3$, which stays in solution. The chemical amount of $AgNO_3$ is 0.63500 mol, and the chemical amount of NaCl is slightly less, 0.60663 mol. Hence the NaCl is the limiting reactant. The theoretical yield of AgCl is computed based on the amount of the limiting reactant:**

$$35.453 g\ NaCl \times \frac{1 mol\ NaCl}{58.4428 g\ NaCl} \times \frac{1 mol\ AgCl}{1 mol\ NaCl} \times \frac{143.323 g\ AgCl}{1 mol\ AgCl}$$

$$= 86.944 g\ AgCl$$

2.52. (See Example 2–11.) When the blue liquid dinitrogen trioxide (N_2O_3) is added to a solution of sodium hydroxide ($NaOH$), it reacts to give sodium nitrite ($NaNO_2$):

$$N_2O_3(\ell) + 2\,NaOH(aq) \longrightarrow 2\,NaNO_2(aq) + H_2O(\ell)$$

What is the concentration of sodium nitrite if 2.13 g of N_2O_3 is added to an excess of aqueous sodium hydroxide and the resulting solution is diluted to a total volume of 500.00 mL?

2.52

2 Referencing the balanced chemical equation given in the problem, the chemical amount of sodium nitrate may be calculated. The molar concentration of this material is then obtained by dividing the chemical amount by the total volume of the solution.

$$2.13g\ N_2O_3 \times \frac{1mol\ N_2O_3}{76.011g\ N_2O_3} \times \frac{2mol\ NaNO_2}{1mol\ N_2O_3} = 0.0560mol\ NaNO_2$$

$$c_{NaNO_2} = \frac{0.0560mol\ NaNO_2}{0.50000L} = 0.112M\ NaNO_2$$

5.86. Complete combustion (burning) of 2.00 volumes of a gaseous hydrocarbon to $CO_2(g)$ and $H_2O(g)$ required 9.00 volumes of pure $O_2(g)$. Both volumes were measured at the same temperature and pressure. The reaction yielded 0.135 g of H_2O and 0.330 g of CO_2. Determine the molecular formula of the hydrocarbon.

5-86 A hydrocarbon contains only the elements carbon and hydrogen. Let the molecular formula of the hydrocarbon be C_xH_y. The hydrocarbon and oxygen combine in a 2 to 9 ratio by volume. It follows that they combine in a 2 to 9 molar ratio as well, because the chemical amounts (of ideal gases) are in direct proportion to their volumes when T and P are the same. The equation for the combustion reaction is:

$$2C_xH_y(g) + 9O_2(g) \longrightarrow 2xCO_2(g) + yH_2O(g)$$

Balancing oxygen between the two sides of this equation requires the relationship 18 = 4x + y. The chemical amount of hydrogen in the sample of the hydrocarbon is:

$$n_H = 0.135g\ H_2O \times \frac{1mol\ H_2O}{18.0153g\ H_2O} \times \frac{2mol\ H}{1mol\ H_2O} = 0.0150mol\ H$$

and the chemical amount of carbon in the sample is:

$$n_C = 0.330g\ CO_2 \times \frac{1mol\ CO_2}{44.0098g\ CO_2} \times \frac{1mol\ C}{1mol\ CO_2} = 0.00750mol\ C$$

Clearly, the chemical amount of H is twice that of C, so regardless of the size of the sample, y = 2x. There are now two equations relating the unknowns x and y. Solving gives x = 3 and y = 6, so the hydrocarbon is C_3H_6.

5.94. In an optical bottle, beams of laser light replace the physical walls of conventional containers to confine gases. The laser beams are tightly focused. Suppose that they briefly (for 0.5 s) exert enough pressure to confine 500 sodium atoms in a volume of $1.0 \times 10^{-15}\ m^3$ at a temperature of 0.00024 K.

(a) Use the ideal gas law to compute the pressure exerted on the "walls" of the optical bottle.

(b) In this gas, the mean free path (the average di the sodium atoms travel between collisions) Compare this with the mean free path of the gaseous sodium at room conditions.

5-94 (a) Convert the volume of the trap to liters (using $1L = 10^{-3}\ m^3$) and the amount of its contents to moles. Then apply the ideal gas law:

$$P = n \times \frac{RT}{V} = \frac{500}{6.022137 \times 10^{23}\, mol^{-1}} \times \frac{(0.08206\ {}^{L\cdot atm}/_{mol\cdot K}) \cdot (0.00024K)}{1.0 \times 10^{-12}\, L}$$

$$= 1.6 \times 10^{-14}\, atm$$

(b) The mean free path of the sodium atoms in the optical trap (3.9 m) is much longer than that of gaseous sodium atoms at room conditions (about 10^{-7} m).