CHAPTER 8

Chemical Composition

INTRODUCTION

Before beginning a project of any kind, it is always important to know the quantity of material needed to finish the project. It is usually possible to count the number of individual items you will need. In chemistry, it is difficult to count the number of atoms or molecules needed because the individual particles are too small, and there are too many of them. This chapter will show you how you can count the number of particles by weighing them.

CHAPTER DISCUSSION

The Mole

One crucial concept in this chapter is that of the mole. Make sure you understand why it is so important. There are two main ideas you need to consider:

1. We can count objects by weighing a sample of the objects provided we know the average mass of the objects.
2. Relative masses of two or more different objects stay the same but can be in larger units if we have the same number of objects.

Section 8.1 in your text provides a very good discussion of this first point. Make sure to read this. Talk to an instructor if you have difficulty with it. Let's look at the second point more carefully.

Suppose we have two blocks, a red block and a yellow block. The red block weighs 1.0 ounce, and the yellow block weighs 4.0 ounces. Now suppose we have 16 of each block. What is the mass of each sample? The sample of red blocks weighs 16.0 ounces, and the sample of yellow blocks weighs 64.0 ounces. But note that 16.0 ounces is also 1.0 pound, thus 64.0 ounces is 4.0 pounds. Note the relative masses of the blocks:

	One Block	Sixteen Blocks
red	1.0 ounce	1.0 pound
yellow	4.0 ounces	4.0 pounds

The relative masses stay the same (1:4) but the units are changed. Why is this important?

Recall from Chapter 4 that the periodic table gives us the relative masses of the elements. But what are the units of these? Actually, there need not be any unit at all, the units could be anything; that is, the table gives us relative masses much like the 1:4 ratio we see in the examples with the blocks. But what units would be useful for us? The standard unit of mass that we will use is the gram. However, the average hydrogen atom, for example, has a mass of 1.66057×10^{-24} g. This is much too small for us to deal with. We would like to keep the relative mass of hydrogen at 1.008 (as it is on the periodic table) but in units of grams. How many hydrogen atoms are there in a 1.008-g sample of hydrogen?

$$1.008 \text{ g H} \times \left(\frac{1 \text{ atom H}}{1.66057 \times 10^{-24} \text{ g H}}\right) = 6.022 \times 10^{23} \text{ H atoms}$$

Thus, if we have 6.022×10^{23} atoms of hydrogen the sample will have a mass of 1.008 g. Along the same lines, if we have 6.022×10^{23} atoms of carbon, the sample will have a mass of 12.01 g. If we have 6.022×10^{23} atoms of oxygen, the sample will have a mass of 16.00 g. That is, 6.022×10^{23}, which is called a mole, is the number that converts the units on the periodic table (called amu or atomic mass units) to grams.

The mole is just a number (like a dozen is 12), and it allows us to convert between how many atoms (or molecules) we have and the mass of the sample. In the example with the blocks, this number was 16 (to convert ounces to grams). To convert amu to grams, we use the mole.

Formulas and Mass Percent

Would you say ammonia (NH_3) is mostly nitrogen or mostly hydrogen? Your answer depends on if you are looking at the number of atoms or the mass of the atoms. In terms of numbers of atoms, ammonia is ¾ hydrogen (there are four atoms making up an ammonia molecule, and three of them are hydrogen). But it is often important to know the composition by mass of a compound. Chemists have instruments that give them percent by mass data, and they use this to determine the formulas of compounds. How can we do this?

Suppose we have 1.0 mole of ammonia molecules. The molar mass is 17.034 g (N = 14.01 g/mol, and each hydrogen is 1.008 g/mol, so the molar mass of NH_3 is 14.01 + 3(1.008) = 17.034 g). How much of this mass is hydrogen? How much of this mass is nitrogen?

Since there are three hydrogen atoms per ammonia molecule, there are three moles of hydrogen atoms per mole of ammonia molecule. Thus, the total mass of hydrogen should be 3(1.008) or 3.024 g, and there is 14.01 g of nitrogen. The mass percent of each element is:

$$\text{Hydrogen: } \frac{3.024 \text{ g}}{17.034 \text{ g}} \times 100\% = 17.75\% \text{ hydrogen by mass}$$

$$\text{Nitrogen: } \frac{14.01 \text{ g}}{17.034 \text{ g}} \times 100\% = 82.25\% \text{ nitrogen by mass}$$

Note that even though there are more hydrogen atoms than nitrogen atoms, the percent by mass of hydrogen in this case is lower than that of nitrogen.

Try the following example before reading on:

What is the mass percent of hydrogen and of nitrogen for the compound N_2H_6?

To solve this, you can find the molar mass of the compound, which is 34.068, and the total mass of hydrogen (6.048 g) and nitrogen (28.02 g). If you are having difficulty getting these numbers, read through your text or talk with your instructor.

Thus, the mass percent of each element is:

$$\text{Hydrogen: } \frac{6.028 \text{ g}}{34.068 \text{ g}} \times 100\% = 17.75\% \text{ hydrogen by mass}$$

$$\text{Nitrogen: } \frac{28.02 \text{ g}}{34.068 \text{ g}} \times 100\% = 82.25\% \text{ nitrogen by mass}$$

Why is this significant? Note that this is the same percent by mass as ammonia (NH_3). Chemists generally use percent-by-mass data to determine the formula of a compound. But what if we know a compound is 17.75% hydrogen by mass and 82.25% nitrogen by mass? Is the formula NH_3? Or N_2H_6? Actually, you should prove to yourself that any formula that has 3 times as many hydrogen atoms as nitrogen atoms (N_xH_{3x}) will be 17.75% hydrogen and 82.25% nitrogen by mass. So what are we to do?

This is why we need to know the molar mass of a compound to know the actual formula for the compound (the molecular formula). Given just the percent-by-mass data allows us to determine only the ratio of atoms in the molecule (in this case 1:3) and determine the empirical formula. Let's consider an example.

> A compound consisting of carbon, hydrogen, and oxygen is 40.00% carbon by mass and 6.71% hydrogen by mass. What is the empirical formula of the compound?

In this case, we are given the percent-by-mass data and are asked to determine the formula for the compound. How should we think about this?

We are looking for the formula (empirical) of this compound. Recall that the formula gives us the ratio of atoms in the compound. The general formula for this compound is

$$C_xH_yO_z$$

where *x*, *y*, and *z* represent the number of atoms in one molecule of the compound (or, in the case of the empirical formula, the lowest whole-number ratio). Our problem, then, is to determine these numbers. Thus, we have been given mass-percent data for the atoms and need to determine the number of atoms. In essence, we will have to change a mass to a number. How do we do this? Well, we know, for example, that 1 mole (a number, 6.022×10^{23}) of carbon atoms has a mass of 12.01 g. Thus, we will use the molar mass of the atoms to make this conversion.

Before we go on, notice that we have done quite a bit of thinking about this problem before doing any calculations. This is generally a good approach with chemistry problems. That is, think about the setup of the problem before worrying about the specifics. Do not immediately try to plug numbers into a given equation. Think about the problem first.

Now we know that we will need to convert the mass data to numbers using the mole. But what is the mass of carbon? Hydrogen? Oxygen? We are only given the mass-percent data. This requires us to think about what is meant by mass percent.

Stating that the compound is 40.00% carbon by mass means that for every 100.00 g of compound, 40.00 g is carbon.

It is not stating that we have 40.00 g of carbon necessarily, but that 40.00 g out of every 100.00 g of compound is carbon (if we had 200.00 g of compound, we would have 80.00 g of carbon, for example). Since we are trying to determine the number ratio for the atoms, we need to know only the mass ratio, not the actual masses. The easiest way to do this is to assume we have 100.00 g of the compound (we could assume any mass here; make sure you understand why). Therefore, the masses of the atoms are:

mass of carbon:	40.00 g
mass of hydrogen:	6.71 g
mass of oxygen:	53.29 g

Note that these masses add up to 100.00 g (which is how we can determine the value for oxygen). Now we can convert these masses to numbers (moles) by using the molar masses of each element. Do this before reading on.

The mole ratios you should have calculated are

moles of carbon:	3.33
moles of hydrogen:	6.66
moles of oxygen:	3.33

If you cannot get these numbers, review this in your text, work with a friend, or see your instructor.

We now have the mole ratios. Does this mean the empirical formula is $C_{3.33}H_{6.66}O_{3.33}$? No. We must represent the formulas with whole numbers (we cannot have fractions of atoms). Note, however, that 3.33:6.66:3.33 can be written as 1:2:1. Since we used mass-ratio data, we calculated atom-ratio data. That is, we do not actually have 3.33 moles of carbon (or 2 moles for that matter) but we have a 1:2:1 ratio of atoms of carbon, hydrogen, and oxygen, respectively. Thus, we can write the empirical formula as

$$C_1H_2O_1$$

But this is not necessarily the actual (molecular) formula of the compound. You should prove to yourself, for example, that the formula $C_2H_4O_2$ has the same mass percent of each atom as $C_1H_2O_1$. To determine the actual formula we need to know the molar mass of the compound.

For example, if we are told that the molar mass of this compound is about 180 g/mol, how do we determine the molecular formula? We know that the answer has to have the general formula $C_xH_{2x}O_x$. Recall that the molar mass is the sum of the masses of the atoms. Thus, $180 = 12.01(x) + 1.008(2x) + 16.00(x)$, or $180 = 30.026x$. Solving for x, we get 6. Therefore, the molecular formula is $C_6H_{12}O_6$.

Again, note the amount of thought that went into solving this problem. Do not expect to simply plug numbers into equations. By thinking about the underlying concepts involved, you will find that you can solve quite difficult and novel problems. This is a major goal.

LEARNING REVIEW

1. A hardware store employee determined that the average mass of a certain size nail was 2.35 g.
   a. How many nails are there in 1057.5 g nails?
   b. If a customer needs 1500 nails, what mass of nails should the employee weigh out?
2. Ten individual screws have masses of 10.23 g, 10.19 g, 10.24 g, 10.23 g, 10.26 g, 10.23 g, 10.28 g, 10.30 g, 10.25 g, and 10.26 g. What is the average mass of a screw?
3. The average mass of a hydrogen atom is 1.008 amu. How many hydrogen atoms are there in a sample that has a mass of 25,527.6 amu?
4. The average mass of a sodium atom is 22.99 amu. What is the mass in amu of a sample of sodium atoms that contains 3.29×10^3 sodium atoms?
5. A sample with a mass of 4.100×10^5 amu is 25.00% carbon, and 75.00% hydrogen by mass. How many atoms of carbon and hydrogen are in the sample? The average mass of a carbon atom is 12.01 amu and of a hydrogen atom is 1.008 amu.
6. The average mass of a neon atom is 20.18 amu.
   a. How many grams of neon are found in a mole of neon?
   b. How many atoms of neon are in a mole of neon?
7. What is the value of Avogadro's number, and how is it defined?

8. Use the average mass values found inside the front cover of your textbook to solve the problems below.

   a. A helium balloon contains 5.38×10^{22} helium atoms. How many grams of helium are in the balloon?

   b. A piece of iron was found to contain 3.25 mol Fe. How many grams are in the sample?

   c. A sample of liquid bromine contains 65.00 g Br atoms. How many bromine atoms are in the sample?

   d. A sample of zinc contains 0.78 mol Zn. How many zinc atoms are in this sample?

9. What is meant by the term "molar mass"?

10. Calculate the molar mass of the following substances.

    a. Fe_2O_3

    b. NH_3

    c. C_2H_5OH

    d. CO_2

    e. N_2O_5

11. Calculate the molar mass of these ionic compounds.

    a. HCl

    b. $MgBr_2$

    c. $Pb(OH)_2$

    d. $Cu(NO_3)_2$

    e. KCl

    f. Na_2SO_4

12. Acetone, which has a formula of C_3H_6O, is used as a solvent in some fingernail polish removers. How many moles of acetone are in 5.00 g of acetone?

13. How many grams of potassium sulfate are in 0.623 mol potassium sulfate?

14. Calculate the mass fraction of nitrogen in N_2O_5.

15. Calculate the mass percent of each element in the following substances.

    a. CH_3NH_2

    b. H_2SO_4

16. Explain the difference between the empirical formula and the molecular formula of a compound.

17. The molecular formula of the gas, acetylene, is C_2H_2. What is the empirical formula?

18. When 2.500 g of an oxide of mercury, Hg_xO_y, is decomposed into the elements by heating, 2.405 g of mercury is produced. Calculate the empirical formula for this compound.

Hg_xO_y	$\rightarrow$	x Hg	+	y O
2.500 g		2.405 g		? g

19. A compound was analyzed and found to contain only carbon, hydrogen and chlorine. A 6.380-g sample of the compound contained 2.927 g carbon and 0.5729 g hydrogen. What is the empirical formula of the compound?

20. The compound benzamide has the following percent composition. What is the empirical formula?

$$C = 69.40\% \quad H = 5.825\% \quad N = 11.57\% \quad O = 13.21\%$$

21. The empirical formula for a compound used in the past as a green paint pigment is $C_2H_3As_3Cu_2O_8$. The molar mass is 1013.71 g. What is the molecular formula?

22. A sugar that is broken down by the body to produce energy has the following percent composition.

$$C = 39.99\% \quad H = 6.713\% \quad O = 53.29\%$$

The molar mass is 210.18 g. What is the molecular formula?

ANSWERS TO THE LEARNING REVIEW

1.

a. This problem relies on the principle of counting by weighing. The question "how many nails?" can be answered because we are given the average mass of 1 nail.

$$\frac{1 \text{ nail}}{2.35 \text{ g}} \times 1057.5 \text{ g} = 450. \text{ nails}$$

b. If we know the mass of 1 nail equals 2.35 g, then the mass of 1500 nails is a multiple of 2.35 g.

$$\frac{2.35 \text{ g}}{1 \text{ nail}} \times 1500. \text{ nails} = 3530 \text{ g}$$

2. Average mass can be determined by adding the masses of each individual screw, then dividing by the number of screws measured.

$$10.23 \text{ g} + 10.19 \text{ g} + 10.24 \text{ g} + 10.23 \text{ g} + 10.26 \text{ g} + 10.23 \text{ g} + 10.28 \text{ g} + 10.30 \text{ g} + 10.25 \text{ g} + 10.26 \text{ g} = 102.47 \text{ g}$$

The total mass of all 10 screws is 102.47 g.

$$\frac{102.47 \text{ g}}{10 \text{ screws}} = 10.25 \text{ g/screw}$$

The average mass of a screw is 10.25 g.

3. This problem is an example of counting by weighing. We are given the average mass of one hydrogen atom, and asked for the number of hydrogen atoms in some other mass of hydrogen.

$$\frac{1 \text{ hydrogen atom}}{1.008 \text{ amu}} \times 25{,}527.6 \text{ amu} = 25{,}330 \text{ hydrogen atoms}$$

4. If we know the average mass of an atom, we can calculate the mass of any quantity of atoms.

$$\frac{22.99 \text{ amu}}{1 \text{ sodium atom}} \times 3.29 \times 10^3 \text{ sodium atoms} = 75{,}600 \text{ amu}$$

5. The total mass of the sample is 4.100×10^5 amu. Of this mass, 25.00% comes from carbon atoms. So the mass contributed by carbon is

$$4.100 \times 10^5 \text{ amu} \times 0.2500 = 1.025 \times 10^5 \text{ amu}$$

The mass contributed by hydrogen is the original mass minus the mass contributed by carbon

$$4.100 \times 10^5 \text{ amu} - 1.025 \times 10^5 \text{ amu} = 3.075 \times 10^5 \text{ amu}$$

Now that we know the total mass of each kind of atom, we can use the average mass of one atom to count the number of atoms present.

$$\frac{1 \text{ hydrogen atom}}{1.008 \text{ amu}} \times 3.075 \times 10^5 \text{ amu} = 3.051 \times 10^5 \text{ hydrogen atoms}$$

$$\frac{1 \text{ carbon atom}}{12.01 \text{ amu}} \times 1.025 \times 10^5 \text{ amu} = 8.535 \times 10^3 \text{ carbon atoms}$$

6.

a. A mole of any element always contains a mass in grams equal to the average atomic mass of that element. So there are 20.18 g Ne in 1 mol Ne.

b. A mole of atoms of any element always contains 6.022×10^{23} atoms.

7. Avogadro's number is the number equal to the number of atoms in 12.01 grams of carbon. Chemists have accurately determined this number to be 6.022×10^{23} atoms.

8. These problems use conversions between moles and grams or between moles and number of atoms. For each element, you must write a different conversion factor for moles to grams depending upon the average mass for that element.

a. This problem requires first determining the moles of He and then converting moles to grams.

$$5.38 \times 10^{22} \text{ He atoms} \times \frac{1 \text{ mol He}}{6.022 \times 10^{23} \text{ He atoms}} \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 0.357 \text{ g He}$$

b. $3.25 \text{ mol Fe} \times \dfrac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 182 \text{ g Fe}$

c. This is a two-step problem requiring that you first calculate the number of moles of Br, then the number of Br atoms.

$$65.00 \text{ g Br} \times \frac{1 \text{ mol Br}}{79.90 \text{ g Br}} \times \frac{6.022 \times 10^{23} \text{ Br atoms}}{1 \text{ mol Br}} = 4.899 \times 10^{23} \text{ Br atoms}$$

d. $0.78 \text{ mol Zn} \times \dfrac{6.022 \times 10^{23} \text{ Zn atoms}}{1 \text{ mol Zn}} = 4.7 \times 10^{23} \text{ Zn atoms}$

9. Molar mass is the number of grams found in 1 mole of a substance. The molar mass is calculated by adding together the masses of each atom in the substance.

10.

a. Fe_2O_3 contains two Fe atoms and three O atoms.

$(2 \times 55.85 \text{ g Fe}) + (3 \times 16.00 \text{ g O}) = 159.7 \text{ g}$

b. NH_3 contains one nitrogen atom and three hydrogen atoms.

$(1 \times 14.01\ g\ N) + (3 \times 1.008\ g\ H) = 17.03\ g$

c. C_2H_5OH contains two C atoms, six H atoms and one O atom.

$(2 \times 12.01\ g\ C) + (6 \times 1.008\ g\ H) + (1 \times 16.00\ g\ O) = 46.07\ g$

d. CO_2 contains one C atom and two O atoms.

$(1 \times 12.01\ g\ C) + (2 \times 16.00\ g\ O) = 44.01\ g$

e. N_2O_5 contains two N atoms and five O atoms

$(2 \times 14.02\ g\ N) + (5 \times 16.00\ g\ O) = 108.0\ g$

11.

a. $1.008\ g\ H + 35.45\ g\ Cl = 36.46\ g$

b. $24.31\ g\ Mg + (2 \times 79.90\ g\ Br) = 184.1\ g$

c. $207.19\ g\ Pb + (2 \times 16.00\ g\ O) + (2 \times 1.008\ H) = 241.2\ g$

d. $63.55\ g\ Cu + (2 \times 14.01\ g\ N) + (6 \times 16.00\ g\ O) = 187.6\ g$

e. $39.10\ g\ K + 35.45\ g\ Cl = 74.55\ g$

f. $(2 \times 22.99\ g\ Na) + 32.07\ g\ S + (4 \times 16\ g\ O) = 142.1\ g$

12. To solve this problem, we need to know how many grams of acetone are in one mole of aceto e. The number of grams of acetone equal to one mole of acetone is the molar mass.

$$\text{molar mass acetone} = (3 \times 12.01\ g) + (6 \times 1.008\ g) + 16.00\ g = 58.08\ g$$

$$5.00\ \text{g acetone} \times \frac{1\ \text{mol acetone}}{58.08\ \text{g acetone}} = 0.0861\ \text{mol acetone}$$

13. molar mass $K_2SO_4 = (2 \times 39.10\ g) + 32.07\ g + (4 \times 16.00\ g) = 174.3\ g$

$$0.623\ \text{mol}\ K_2SO_4 \times \frac{174.3\ g\ K_2SO_4}{1\ \text{mol}\ K_2SO_4} = 109\ g\ K_2SO_4$$

14. Mass fraction is equal to the mass of the desired element, in this case nitrogen, divided by the molar mass.

$$\frac{28.02\ g\ N}{108.0\ g\ \text{total}} = 0.2594$$

15.

a. molar mass of CH_3NH_2 is $12.01\ g\ C + (5 \times 1.008\ g\ H) + 14.01\ g\ N = 31.06\ g$ total

$$\frac{12.01\ g\ C}{31.06\ g\ \text{total}} \times 100 = 38.67\%\ C$$

$$\frac{5.040\ g\ H}{31.06\ g\ \text{total}} \times 100 = 16.23\%\ H$$

$$\frac{14.01\ g\ N}{31.06\ g\ \text{total}} \times 100 = 45.11\%\ N$$

b. molar mass of H_2SO_4 is $(2 \times 1.008 \text{ g H}) + 32.07 \text{ g S} + (4 \times 16.00 \text{ g O}) = 98.09 \text{ g total}$

$$\frac{2.016 \text{ g H}}{98.09 \text{ g total}} \times 100 = 2.055\% \text{ H}$$

$$\frac{32.07 \text{ g S}}{98.09 \text{ g total}} \times 100 = 32.69\% \text{ S}$$

$$\frac{64.00 \text{ g O}}{98.09 \text{ g total}} \times 100 = 65.25\% \text{ O}$$

16. The empirical formula gives only the relative number of atoms, that is, a ratio of each kind of atom. The molecular formula tells exactly how many of each kind of atom are present in the molecule.

17. For every two atoms of carbon in acetylene there are two atoms of hydrogen. The ratio of carbon atoms to hydrogen atoms is 1:1. So the empirical formula of acetylene is CH.

18. Since we know that mercury and the oxygen combined weighed 2.500 g before the reaction took place and that the mass of the mercury is 2.405 g, then the mass of oxygen must be $2.500 - 2.405 = 0.095$ g. We can now convert grams of mercury and grams of oxygen to moles using the atomic masses of these elements.

$$2.405 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.59 \text{ g Hg}} = 0.01199 \text{ mol Hg}$$

$$0.095 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.0059 \text{ mol O}$$

The ratio of mercury atoms to oxygen atoms is, $\frac{0.01199}{0.0059} = 2.03$ to 1.

So, there are twice as many Hg atoms as O atoms, and the empirical formula is Hg_2O.

19. When 6.380 g of a compound that contained only carbon, hydrogen and chlorine was analyzed, it was found to contain 2.927 g carbon and 0.5729 g hydrogen. The mass of chlorine must be equal to the total mass minus the mass of carbon plus hydrogen.

$$\text{mass of chlorine} = 6.380 - (2.927 \text{ g C} + 0.5729 \text{ g H}) = 2.880 \text{ g}$$

The moles of each kind of atom are determined from the average atomic mass.

$$2.927 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.2437 \text{ mol C}$$

$$0.5729 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.5684 \text{ mol H}$$

$$2.880 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.08124 \text{ mol Cl}$$

Express the mole ratios in whole numbers by dividing each number of moles by the smallest number of moles.

$$\frac{0.2437 \text{ mol C}}{0.08124} = 3.000 \text{ mol C}$$

$$\frac{0.5684 \text{ mol H}}{0.08124} = 7.000 \text{ mol H}$$

$$\frac{0.08124 \text{ mol Cl}}{0.08124} = 1.000 \text{ mol Cl}$$

The empirical formula is C_3H_7Cl.

20. This problem provides only percent composition data for the compound benzamide. It does not provide an analysis in grams for each of the elements present. We need to know how many grams of each element are present in a sample of benzamide so we can calculate the moles of each element. We can convert percent composition data to grams of each element. Assume that we have 100.0 g of benzamide. Of that sample, 69.40% is carbon. So for a 100.0 g sample, 69.40 g are carbon, 5.825 g are hydrogen, 11.57 g are nitrogen, and 13.21 g are oxygen. We can calculate the number of moles of each element.

$$69.40 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.779 \text{ mol C}$$

$$5.825 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.779 \text{ mol H}$$

$$11.57 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.8258 \text{ mol N}$$

$$13.21 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.8256 \text{ mol O}$$

Now divide each number of moles by the smallest number of moles to convert the number of moles to whole numbers.

$$\frac{5.779 \text{ mol C}}{0.08256} = 7.000 \text{ mol C}$$

$$\frac{5.779 \text{ mol H}}{0.08256} = 7.000 \text{ mol H}$$

$$\frac{0.08258 \text{ mol N}}{0.08256} = 1.000 \text{ mol N}$$

$$\frac{0.8256 \text{ mol O}}{0.08256} = 1.000 \text{ mol O}$$

The empirical formula is C_7H_7NO.

21. If you are given both the molar mass and the empirical formula, determining the molecular formula is straightforward. If we multiply all the atoms in the empirical formula by some number, we will have a correct molecular formula. So the molecular formula is a multiple of the empirical formula. We can determine what this multiple is by comparing the molar mass of the molecular formula with the molar mass of the empirical formula.

$$\text{molar mass empirical formula} = (2 \times 12.01 \text{ g C}) + (3 \times 1.008 \text{ g H}) + (3 \times 74.92 \text{ g As}) + (2 \times 63.55 \text{ g Cu}) + (8 \times 16.00 \text{ g O}) = 506.9 \text{ g}$$

The molar mass of the empirical formula is 506.9 g, and we know that the molar mass of the molecular formula is 1013.7 g. There are two empirical formulas in the molecular formula.

$$\frac{1013.7 \text{ g in molecular formula}}{506.9 \text{ g in empirical formula}} = 2.000$$

So, the molecular formula is 2 times the empirical formula. The molecular formula is $2(C_2H_3As_3Cu_2O_8)$ or $C_4H_6As_6Cu_4O_{16}$.

22. In this problem, we are asked to find the molecular formula given the molar mass and the percent composition. To determine the molecular formula, we must first find the empirical formula.

$$39.99 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.330 \text{ mol C}$$

$$6.713 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.660 \text{ mol H}$$

$$53.29 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.331 \text{ mol O}$$

Divide each molar quantity by the smallest number of moles to convert the number of moles to a whole number.

$$\frac{3.330 \text{ mol C}}{3.330} = 1.000$$

$$\frac{6.660 \text{ mol H}}{3.330} = 2.000$$

$$\frac{3.331 \text{ mol O}}{3.330} = 1.000$$

The empirical formula is CH_2O. The molar mass of the molecule is 210.18 g. So we need to know the molar mass of the empirical formula.

molar mass empirical formula = 12.01 g C + (2 × 1.008 g H) + 16.00 g O = 30.03 g

How many empirical formulas are there in one molecular formula? We can tell by dividing the molar mass of the molecular formula by the molar mass of the empirical formula.

$$\frac{210.18 \text{ g in molecular formula}}{30.03 \text{ g in empirical formula}} = 6.999$$

The molecular formula is 7 times the empirical formula.

molecular formula = 7 × (CH_2O) or $C_7H_{14}O_7$