

# CHAPTER 2

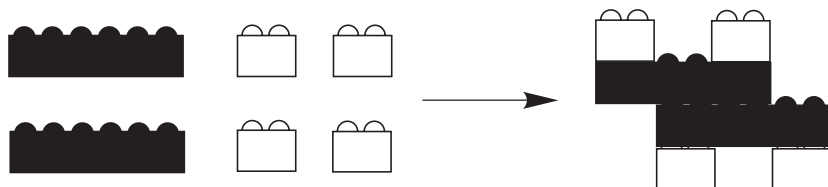
## COUNTING ATOMS AND MOLECULES

In this chapter, we will use the tools and techniques that we learned in the previous chapter, and we will begin to solve problems that involve the relationship between atoms and molecules. The first few sections provide the basic concepts and terminology that you will use in the rest of the chapter. The remaining sections are dedicated to common types of problems, representing the fundamental calculations that you will need to master (not only for your exams, but also to serve as a foundation for material that we will learn in subsequent chapters). The problems and techniques will slowly build in complexity, but the methods for solving these problems are not difficult. Once you get practice, you should find that these problems are all very straightforward.

### 2.1 EMPIRICAL AND MOLECULAR FORMULAS

We will now develop an analogy that we will use again and again throughout this book. It will help you understand what you are doing when you perform the calculations necessary to solve the problems in the rest of this chapter. In our analogy, we will think of atoms as Lego pieces. At first, this analogy might seem simple and unnecessary. But as you begin to deal with more challenging problems, the Lego analogy will help you to have a better understanding of the principles involved in solving problems. In the next chapter, when we discuss stoichiometry, we will rely heavily on the use of our Lego analogy.

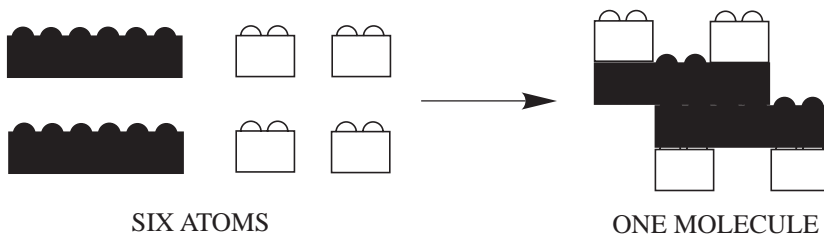
Lego pieces come in different sizes, shapes, and colors. The number of notches on the top of a Lego piece tells us how many other Lego pieces can connect with it to build a structure. We can imagine building very simple structures with our Lego pieces, for example:



Let's say we wanted to take some Lego pieces and actually build the structure shown above. The figure above tells us that we would need two black pieces and four white pieces. If we wanted to build 100 structures like the one above, we would need 200 black pieces and 400 white pieces. The more pieces we have, the more of these structures we can build, but we need to have the pieces in a certain ratio. Specifically, we need to have one black piece for every two white pieces. We can express this in a simple way:  $(\text{Black})_1(\text{White})_2$ . This expression does not tell us how the pieces are connected. In fact, it does not even tell us how many Lego pieces we need to build one structure. If I gave you one black piece and two white pieces, you would not be able to build even one structure.

Instead, we can choose an expression that *does* tell us how many pieces you need to build one structure. In our example, we would express it like this:  $(\text{Black})_2(\text{White})_4$ . This expression is similar to the one before, but now it tells the information we need to build one structure.

In our analogy, the individual Lego pieces represent atoms, and the structure represents a molecule:



So atoms are the building blocks of molecules, in the same way that Lego pieces are the building blocks of Lego structures.

We saw two ways to express the recipe for building one structure:  $(\text{Black})_2(\text{White})_4$  tells us the number of each type of piece that is needed to build one structure. And  $(\text{Black})_1(\text{White})_2$  just tells us the relative ratio of white and black pieces. The same is true when we use atoms to build molecules. We can express the molecules in a way that tells us exactly how many atoms are needed to build one molecule, or we can just show the ratio of atoms needed. A common example is  $\text{N}_2\text{O}_4$  vs.  $\text{NO}_2$ . The first formula tells us the molecule is made out of two nitrogen atoms and four oxygen atoms. The second formula only tells us the ratio of nitrogen to oxygen atoms, but does *not* tell us how many atoms are actually needed to build the molecule. The first formula ( $\text{N}_2\text{O}_4$ ) is called the *molecular formula*, and the second formula ( $\text{NO}_2$ ) is called the *empirical formula*. Notice that we do not use a subscript next to the nitrogen in  $\text{NO}_2$ . When there is no subscript, we assume that it means 1.

In some cases, the molecular formula *is* the empirical formula. For example, consider a water molecule, which is made with two hydrogen atoms and one oxy-

gen atom. The molecular formula is therefore  $\text{H}_2\text{O}$ , and the empirical formula is just  $\text{H}_2\text{O}$ .

EXERCISE 2.1. Consider the following molecular formula:



Is this also an empirical formula? If not, then write out the empirical formula.

Answer: To determine if this is an empirical formula, you need to see if the smallest whole numbers are used. In this case, you could use smaller whole numbers to express the ratio of atoms in the molecule. Both 4 and 10 can be divided by 2, giving  $\text{C}_2\text{H}_5$ . Therefore, the empirical formula is  $\text{C}_2\text{H}_5$ . Notice that an empirical formula uses the smallest *whole* numbers that express the ratio. So, your answer is not  $\text{CH}_{2.5}$  even though that expresses the correct ratio.

For each of the following molecular formulas below, write the empirical formula next to it. In some cases, the empirical formula might be the same as the molecular formula.

- |  |   |
|--|---|
| 2.2. $\text{C}_2\text{H}_6$ _____              | 2.3. $\text{C}_3\text{H}_8$ _____         |
| 2.4. $\text{CO}_2$ _____                       | 2.5. $\text{H}_2\text{SO}_4$ _____        |
| 2.6. $\text{C}_6\text{H}_{12}\text{O}_6$ _____ | 2.7. $\text{Hg}_2\text{I}_2$ _____        |
| 2.8. $\text{Al}_2\text{O}_3$ _____             | 2.9. $\text{C}_5\text{H}_5\text{N}$ _____ |

Empirical formulas are mostly an historical artifact. In the old days, chemists determined the formulas of compounds through combustion analysis. This technique allowed for chemists to determine empirical formulas, rather than molecular formulas. But in applying modern-day technology, we have all kinds of fancy techniques (NMR, Mass Spec, X-Ray Diffraction, etc.) that allow us to easily determine the molecular formula. It is somewhat of a mystery why textbooks continue to teach empirical formulas. For whatever reason, empirical formulas have not yet left the textbooks, so you will need to know what they are. But for the most part, you will not use empirical formulas as you move along in this course. The focus will be on molecular formulas.

In the rest of this chapter, most of the problems will start by giving a molecular formula. You will soon see that the molecular formula is a very valuable piece of information, because you can usually extract the information you need from the molecular formula. It is a common mistake for students to miss this point. When a problem says: “22.4 g of  $\text{Al}_2\text{O}_3$  . . .” you actually have actually been given *two* ← QU1 pieces of information: 22.4 g *and*  $\text{Al}_2\text{O}_3$ . If you fail to see the second piece of information, you will not have enough information to solve the problem. We will see many examples of this soon. But first, we have to go over a few more basics.

## 2.2 MOLECULES AND MOLES

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Let's revisit our Lego analogy, in an effort to keep things as simple as possible. Imagine that you need to build a lot of identical Lego structures, like the structure we saw in the previous section:



Your task is to build as many of these structures as possible. So you call a Lego distribution warehouse, and you ask if you can purchase Lego pieces in bulk. The salesperson tells you that they only sell Lego pieces by the barrel. Each barrel contains over 6 million pieces. In fact, all barrels come with exactly 6,022,042 Lego pieces. It is a strange number, but that's the way they are sold.

Then the salesperson tells you some good news. In addition to selling barrels of individual Lego pieces, the warehouse also sells barrels filled with ready-made structures. They have a large variety of ready-made structures, and it turns out that they sell barrels filled with the structure that you want to build—(Black)<sub>2</sub>(White)<sub>4</sub>. When you ask how many structures are in each barrel, you get the same answer you got before: 6,022,042. But this time, we are not talking about individual pieces. This time, we are talking about over 6 million *preassembled structures*. If we were to dump out the contents of one of these barrels, and then take apart all of the structures, we would have over 36 million individual Lego pieces from one barrel alone.

We said before, that our analogy uses individual Lego pieces to represent atoms, and Lego structures to represent molecules. Now we have added a new term: a barrel, which represents the term “mole”. A mole is just a large number (a barrel) of molecules. In our Lego analogy, each barrel had 6,022,042 (or  $6.022042 \times 10^6$ ) Lego structures. Well, it turns out that a mole of molecules is a lot larger than a barrel of Lego structures. The number is actually 17 orders of magnitude larger (100,000,000,000,000,000 times larger). So, there are  $6.022042 \times 10^{23}$  molecules in a mole. Since most of our calculations will involve other numbers that have fewer significant figures, let's make things simpler, and let's use fewer significant figures. From now on, we will use the number  $6.022 \times 10^{23}$ .

The “mole” is a very important term. It is the SI unit that is used to measure the number of molecules in a substance. For example, when we measure the number of molecules, we report our answer as 3.72 mol, or 5.3 mol, etc. As we have already discovered, one mole is defined as  $6.022042 \times 10^{23}$  molecules. This number is also called *Avogadro's number*. It is clearly a *VERY* large number. The term “mole” is the term that we use to connect our macroscopic world to the world of very small atoms and molecules.

EXERCISE 2.10. Count the number of oxygen atoms in 3.45 moles of  $\text{H}_2\text{SO}_4$ .

Answer: This might seem like simple math, and you might be tempted to do this in your head. But you need to get into the habit of always using the factor-label method to solve the problem. You will see many more problems that are not so easy to do in your head, so you need to get practice using the factor-label method now. In addition, the factor-label method serves as a way to double-check your answer. If you divide numbers instead of multiplying them (or vice versa), then the factor-label method will let you know you did it wrong, because the units won't cross off properly in the end.

In order to use the factor-label method, use the following information:

$$1 \text{ mole of } \text{H}_2\text{SO}_4 = 6.022 \times 10^{23} \text{ molecules of } \text{H}_2\text{SO}_4$$

and

$$1 \text{ molecule of } \text{H}_2\text{SO}_4 = 4 \text{ oxygen atoms}$$

Using this information, you can make the following conversion:

$$3.45 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \times \frac{4 \text{ oxygen atoms}}{1 \text{ molecule}}$$

Notice that two conversion factors were used here. The first one converts moles into a number of molecules, and the second one converts a number of molecules into a number of oxygen atoms.

When you use a calculator to multiply the numbers above (except for  $10^{23}$ ), your calculator says 83.1036. So, you have  $83.1036 \times 10^{23}$ . Before you write down your final answer, you need to make sure that your numbers and units are expressed properly. There are only three significant figures in 3.45 moles, so your answer must only have three significant figures. That gives you  $83.1 \times 10^{23}$ . Next, you need to modify the power of 10, so that there will only be one digit in front of the decimal place. By moving the decimal over one place to the left, you are making the number smaller, so you must compensate by making the power of 10 larger. This gives  $8.31 \times 10^{24}$ . Finally, you need to make sure your answer expresses the units properly:

$$\mathbf{8.31 \times 10^{24} \text{ oxygen atoms}}$$

PROBLEM 2.11. Count the number of hydrogen atoms in 0.034 moles of propane,  $\text{C}_3\text{H}_8$ .

Answer: \_\_\_\_\_

PROBLEM 2.12. Count the number of oxygen atoms in 0.872 moles of  $\text{SO}_3$ .

Answer: \_\_\_\_\_

PROBLEM 2.13. Count the number of boron atoms in 2.7 moles of  $\text{B}_2\text{H}_6$ .

Answer: \_\_\_\_\_

## 2.3 ATOMIC MASS

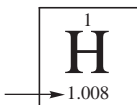
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In this section, we will focus on how we measure the mass of a single atom or a mole of atoms.

Let's start with a single atom. If you try to measure the mass of a single hydrogen atom in terms of *grams*, you will get a *very* small number (with an order of magnitude of  $10^{-24}$  g). So, rather than using very small numbers all of the time, chemists created a new term to express the mass of individual atoms, and even individual molecules. This term is called an **atomic mass unit**, and it is defined as 1 gram divided by Avogadro's number ( $6.02 \times 10^{23}$ ). So, one atomic mass unit equals  $1.66053873 \times 10^{-24}$  g.\*

You might wonder why the chemists didn't just define the new unit as being exactly  $1.0 \times 10^{-24}$  g, and there is a good answer. By defining an atomic mass unit as 1 gram divided by Avogadro's number, chemists have actually made things much easier for us. To see why, we need to take a short tour of the periodic table.

If you look at the periodic table in your textbook (or any other periodic table), you will find a number in the bottom of every box:



For example, the number at the bottom of the box for hydrogen is 1.008. That number tells us that one hydrogen atom has a mass of 1.008 atomic mass units, AND it tells us that one mole of hydrogen atoms is 1.008 grams. For purposes of calculations, everything is simplified because this one number (1.008) refers to the mass of an individual atom (measured in atomic mass units), **and** the mass of a mole (measured in grams). This only works because of the way chemists chose to define the atomic mass unit. If they had defined an atomic mass unit as  $1.0 \times 10^{-24}$  grams, then the mass of one hydrogen atom would have a different numerical value than the mass of a mole of hydrogen atoms.

The accepted symbol for atomic mass units is “u”. So, a hydrogen atom has a mass of 1.008 u. A carbon atom has a mass 12.01 u.

In this way, we can calculate the mass of a single *molecule* by adding up the mass of each atom in the molecule. For example, a water molecule is composed of two hydrogen atoms and one oxygen atom ( $\text{H}_2\text{O}$ ). The mass of a hydrogen atom is 1.008 u, and the mass of an oxygen atom is 16.00 u. Therefore, the mass of a water molecule is just  $2(1.008 \text{ u}) + (16.00 \text{ u}) = 18.02 \text{ u}$ .

\*To be more accurate, chemists actually defined the atomic mass unit based on the mass of a carbon-12 atom (the most abundant isotope of carbon) divided by 12. Another way of saying this is to simply state that an atomic mass unit is just 1 gram divided by Avogadro's number.

It is unlikely that your instructor will expect you to memorize the mass of every element on the periodic table, but you should absolutely know how to use the periodic table to find the information you need. In general, there are several elements that you will find yourself using again and again, so you may as well memorize these now. Hydrogen is 1.008 (or 1.0079, depending on how many significant figures your periodic table shows), carbon is 12.01, and oxygen is 16.00. It is OK if you don't memorize these now, but you might find that memorizing them could save you some time when you are doing problems.

EXERCISE 2.14. What is the mass of one molecule of carbonic acid ( $\text{H}_2\text{CO}_3$ )?

Answer: This molecule is composed of two hydrogen atoms, one carbon atom, and three oxygen atoms. Using a periodic table, we see that the mass of one molecule of carbonic acid is just:

$$2(1.008 \text{ u}) + (12.01 \text{ u}) + 3(16.00 \text{ u}) = 62.03 \text{ u}$$

PROBLEM 2.15. What is the mass of one molecule of barium chromate ( $\text{BaCrO}_4$ )?

Answer (consult the periodic table in your textbook): \_\_\_\_\_

PROBLEM 2.16. What is the mass of one molecule of zinc oxalate ( $\text{ZnC}_2\text{O}_4$ )?

Answer: \_\_\_\_\_

PROBLEM 2.17. What is the mass of one molecule of silver bromide ( $\text{AgBr}$ )?

Answer: \_\_\_\_\_

## 2.4 MOLAR MASS

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In the previous section we saw how to calculate the mass of a single molecule by adding up the mass of each of the atoms in the molecule. Our answer was expressed in atomic mass units. If we wanted to measure the mass of a *mole*, we would use the same exact numbers, but our units would change from atomic mass units to *grams*. If every molecule of water has a mass of 18.02 u, then every mole of water will have a mass of 18.02 grams. Notice how elegant that is. Once again, that's the beauty of the way chemists chose to define an atomic mass unit.

For any compound, one mole means  $6.022 \times 10^{23}$  molecules (by definition). But the mass of those molecules will be different from one compound to the next. For example, a mole of water molecules is  $6.022 \times 10^{23}$  molecules, and it weighs 18.02 g. A mole of liquid hydrogen is  $6.022 \times 10^{23}$   $\text{H}_2$  molecules, and it weighs only 2.016 g. In each case, we are dealing with the same number of molecules. But in the case of water, each molecule is made up of two hydrogen atoms *and* one oxy-

gen atom; whereas in the case of hydrogen, each molecule is made up of only two hydrogen atoms.

So, we see that a mole of one substance will not have the same mass as a mole of another substance. For any substance, we can determine how much mass will be in one mole of that substance. And we can say it like this: one mole of water molecules will have a mass of 18.02 grams. But there is a simpler way to say this. We can say that the *molar mass* is 18.02 *g/mol*. Notice the units that we are using now. These units (*g/mol*) indicate that we are talking about the amount of mass *per mole* of the substance. Before moving on, let's make sure that we are clear on the difference between the two terms that appear in these units (*g* and *mol*). To see the difference between these measurements, imagine that you have a room filled with people. You can add up their collective mass, or you can count the number of people. Those are clearly two different measurements.

To calculate the molar mass of a compound, all you need is the molecular formula. Let's see an example of how this works.

EXERCISE 2.18. Calculate the molar mass of ethylene, which has the following molecular formula:  $C_2H_4$ .

Answer: You are given the molecular formula:  $C_2H_4$ . That means that a *molecule* of ethylene consists of two carbon atoms and four hydrogen atoms. In the previous section you saw that one *molecule* of ethylene should have the same mass as (2 carbon atoms) + (4 hydrogen atoms). Similarly, one *mole* of ethylene should have the same mass as (2 *moles* of carbon atoms) + (4 *moles* of hydrogen atoms). All you need to know is the mass of a mole of carbon atoms and the mass of a mole of hydrogen atoms. And you can find that information on the periodic table, which tells you that carbon atoms have a mass of 12.01 g/mol, and hydrogen atoms have a mass of 1.008 g/mol.

With this information, you can calculate the molar mass of ethylene.  $C_2H_4$  should equal  $2(12.01 \text{ g/mol}) + 4(1.008 \text{ g/mol}) = 28.05 \text{ g/mol}$  (don't forget about significant figures—when you do the math above, your calculator will say 28.052, but your answer can only have *two* decimal places because 12.01 only had two decimal places). So, the molar mass of ethylene is 28.05 g/mol (if you had used a periodic table that gave you more significant figures, then your answer would have had more significant figures in it). Notice the units of your answer: g/mol. This should make sense because your answer describes the amount of grams in one mole of ethylene.

In just a few moments, you will see how to use the molar mass as a special conversion factor to convert moles into mass. But first, let's get some practice calculating molar masses using the periodic table. This might seem trivial, but this skill will be *absolutely essential* in the rest of the problems of this chapter, so you must make sure that you have mastered this skill right now. You will rarely be given the molar mass of a compound—instead, you will have to calculate it yourself using the molecular formula.

Using the periodic table in your textbook, calculate the molar mass of the following compounds:

- |   |   |
|---|---|
| 2.19. C <sub>2</sub> H <sub>6</sub> _____                 | 2.20. C <sub>3</sub> H <sub>8</sub> _____   |
| 2.21. CO <sub>2</sub> _____                               | 2.22. H <sub>2</sub> SO <sub>4</sub> _____  |
| 2.23. C <sub>6</sub> H <sub>12</sub> O <sub>6</sub> _____ | 2.24. Hg <sub>2</sub> I <sub>2</sub> _____  |
| 2.25. Al <sub>2</sub> O <sub>3</sub> _____                | 2.26. C <sub>5</sub> H <sub>5</sub> N _____ |

In this section, we have seen how to calculate the molar mass of a compound, simply by using the periodic table. It is important to realize that chemists have measured the atomic mass of each atom, and these *known values* contain MANY significant figures. The periodic tables shown in most textbooks will give atomic masses measured with only four or five significant digits. For example, hydrogen will be listed as 1.008 amu, or 1.0079 amu. We can use these values to calculate the molar mass of any compound, and our answer will have either four or five significant figures. In most problems, you will never need to use a molar mass with more than four or five significant figures. After all, most problems give you data that are reported with either three or four significant figures. But, you should realize, that atomic mass values are known with more significant figures. Therefore, if it is necessary, you should be able to calculate the molar mass of a compound using more than five significant figures (assuming you are given atomic mass information that is reported with more than five significant figures).

For example, some textbooks will report atomic masses with *six* significant figures. Hydrogen will be reported as 1.00794 amu and oxygen will be reported as 15.9994 amu. If we use these numbers to calculate the molar mass of H<sub>2</sub>O, we will get:

$$2(1.00794 \text{ g/mol}) + (15.9994 \text{ g/mol}) = \mathbf{18.0153 \text{ g/mol}}$$

So, we see that the molar mass of any compound is a value that is known to many significant figures. In general, we will only use four or five significant figures when solving problems.

## 2.5 MOLAR MASS AS A CONVERSION FACTOR: INTERCONVERTING MOLES AND MASS

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In Chapter 1, we learned how to convert units using the factor-label method, and we saw an example of a special conversion factor that helped us convert one type of measurement into another. Specifically, we saw how to use the *density* of a substance to convert from units of mass *into* units of volume, and vice versa. Although mass and volume are measured with completely different units, we were able to convert them using density as a conversion factor.

In this section, we will use a similar technique; a similar type of special conversion factor. We will convert mass into moles, and vice versa, using molar mass

as a conversion factor. But this time, we don't need to be given the conversion factor—we can calculate it ourselves. When we talked about density as a special conversion factor, we saw problems where we were given the density at a certain temperature, and then we made our calculations. If we didn't have the density, then we wouldn't be able to use it as a conversion factor. But here, when we are talking about converting mass and moles, we can calculate the conversion ourselves if we are given the molecular formula. All we need is access to a periodic table, and we will be able to calculate the molar mass. And that is why the molecular formula always provides us with an important piece of information when solving problems.

Now let's see how we can use the molar mass of a compound as a special conversion factor. Let's revisit the example of ethylene again. Previously, we saw how to use the molecular formula of ethylene ( $\text{C}_2\text{H}_4$ ) to calculate its molar mass. We calculated that the molar mass of ethylene is 28.05 g/mol. This means that:

$$1 \text{ mole of ethylene} = 28.05 \text{ gram of ethylene}$$

We can use this to convert moles into grams, and vice versa, using the following conversion factors:

$$\frac{28.05 \text{ g}}{1 \text{ mol}} = 1 \quad \text{and} \quad \frac{1 \text{ mol}}{28.05 \text{ g}} = 1$$

This is the same kind of conversion we saw in Chapter 1 (the factor-label method for converting units). But be careful: the relationships shown above are for ethylene. Different compounds will have different molar masses.

Let's see an example of how we can use these conversion factors:

**EXERCISE 2.27.** Calculate the mass (in grams) of 2.72 moles of  $\text{CO}_2$ .

**Answer:** This problem is asking you to convert moles into mass. To do this, you will need the molar mass, which the problem has not directly given. But you do have the molecular formula, and that is all you need to calculate the molar mass.

The molecular formula is  $\text{CO}_2$ . So, one *mole* should have the same mass as (1 *mole* of carbon atoms) + (2 *moles* of oxygen atoms). All you need to know is the mass of a mole of carbon atoms and the mass of a mole of oxygen atoms. By using the periodic table you'll get: 1 mole of  $\text{CO}_2$  should equal (12.01 g) + 2(16.00 g) = 44.01 g (don't forget about significant figures). So, the molar mass of  $\text{CO}_2$  is 44.01 g/mol.

Now you have the formula you need to solve the problem:

$$1 \text{ mol of } \text{CO}_2 = 44.01 \text{ g}$$

So, set up the following two conversion factors:

$$\frac{44.01 \text{ g}}{1 \text{ mol}} = 1 \quad \text{and} \quad \frac{1 \text{ mol}}{44.01 \text{ g}} = 1$$

You just have to decide which one of these to use. In this problem, you are converting moles into grams, so use the first conversion factor:

$$\begin{aligned} 2.72 \text{ mol CO}_2 \times \frac{44.01 \text{ g}}{1 \text{ mol CO}_2} &= 120 \text{ g} \\ &= 1.20 \times 10^2 \text{ g} \end{aligned}$$

Notice that your answer is expressed with three significant figures.

**PROBLEM 2.28.** Calculate the mass (in grams) of 0.89 moles of  $\text{C}_6\text{H}_{12}\text{O}_6$ . Express your answer using scientific notation.

**PROBLEM 2.29.** You have a water bottle filled with 67.2 grams of water ( $\text{H}_2\text{O}$ ). Calculate how many moles of water in are in the water bottle. (Hint: after calculating the molar mass of  $\text{H}_2\text{O}$ , write out the two conversion factors, and carefully consider which one you would use. In this problem you are converting from grams to moles instead of moles to grams).

**PROBLEM 2.30.** You have a barrel filled with 25.6 kg of water ( $\text{H}_2\text{O}$ ). Calculate how many moles of water in are in the barrel. (This problem is the same as the previous one, only you are starting with kg instead of g, so you will need to use an additional conversion factor in your calculation—to convert kg into g).

PROBLEM 2.31. A piece of copper has a mass of 2.01 kg. Calculate the number of moles of copper atoms. [In this problem, there is no molecular formula, because you are dealing with individual copper atoms, which simplifies the problem because you don't need to first *calculate* the molar mass. You can get the molar mass directly from the periodic table by looking at copper (Cu) on the periodic table].

PROBLEM 2.32. Calculate the mass (in grams) of 7.04 moles of  $\text{CaCO}_3$ . Express your answer using scientific notation.

## 2.6 USING MOLAR MASS AS A CONVERSION FACTOR TWO TIMES IN ONE PROBLEM

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So far, we have seen how to use a molecular formula to calculate molar mass, and then how to use this molar mass as a conversion factor (between moles and mass). The previous problems focused on those calculations. But there are some problems that you might see that get a bit more difficult, even though they just require the same exact calculations. To illustrate the point, consider the following example:

How many grams of silver (Ag) could be obtained if we removed all of the silver from 47.2 g of silver nitrate ( $\text{AgNO}_3$ )?

When you first read this problem, it seems like a problem that requires the conversion of mass into mass (grams of one substance into grams of another substance). So on the surface, you might be tempted to say that you have not yet learned how to do this. But in fact, you have learned all the skills you need. In order to

“see” how the problem is not so hard, you will have to train your eye to read problems in a certain way. Let’s take a closer look.

We have said that if you are given the molecular formula, then you can always convert mass into moles, and moles into mass, by first calculating the molar mass and then using it as a conversion factor. Now that you have these skills, you must keep these skills in mind when reading a problem. When a problem says: “3.07 g of  $\text{MgCl}_2$  . . .”, you must immediately think to yourself like this: they gave me the molecular formula and the mass, so I can easily calculate how many moles I have. That is what you must think to yourself while you are reading any problem. Even if the problem isn’t about that at all, you should still be thinking it. When it comes to solving problems, most of the battle is to realize what information you are given and how you can manipulate that information. Similarly, if you read a problem that says: “2.5 moles of  $\text{NH}_3$  . . .”, you should immediately think to yourself: they gave me the molecular formula and the moles, so I can easily calculate how many grams of  $\text{NH}_3$  I have.

You have to train yourself to think in this way if you want to be able to solve the more difficult problems. And when you do train yourself, you will find that the more difficult problems are really not so difficult in the first place. Let’s now revisit our “difficult” problem, and let’s apply this new type of thinking strategy to try and solve this problem:

How many grams of silver (Ag) could be obtained if we removed all of the silver from 47.2 g of silver nitrate ( $\text{AgNO}_3$ )?

Let’s analyze each half of the problem separately.

1. “How many grams of silver (Ag) could be obtained . . .”

*As you read this, you should think to yourself like this:* Whenever I have the molecular formula, I can easily interconvert grams and moles. (In this case, the molecular formula is provided—we are just talking about individual silver atoms, so the molecular formula is just Ag). You can look at the periodic table and see that the molar mass of silver is 107.9 g/mol. Therefore, *if* you can find out how many moles of silver are obtained, then you could easily convert your answer into grams.

Now let’s look at the second part of the problem:

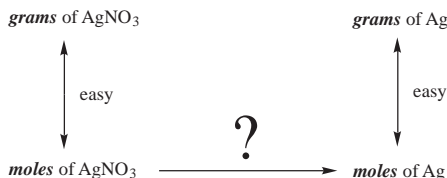
2. “. . . if we removed all of the silver from 47.2 g of silver nitrate ( $\text{AgNO}_3$ )?”

*As you read this, you should think to yourself like this:* Whenever I have the molecular formula, I can easily interconvert grams and moles. Since I know how many grams I started with, I can easily calculate how many moles I started with.

Now we can think about this problem in a new light. With the information that is given to us in the problem, we can easily calculate how many moles of silver nitrate we started with. And we have said that if we can calculate the number of moles of silver we produce, then we can convert that answer into grams of sil-

ver. So, this whole problem boils down to one simple question: how many moles of Ag are produced from one mole of  $\text{AgNO}_3$ ?

If you have trouble seeing the logic we just used, try to stare at this diagram, which represents the logic visually:



By using the concept of molar mass as a conversion factor, we can rephrase our question that makes the problem sound a lot simpler: how many moles of Ag are produced for every mole of  $\text{AgNO}_3$ ? If we look at the molecular formulas, we can easily see that there is a 1:1 ratio here. Ag appears only once in  $\text{AgNO}_3$ , so 1 mole of  $\text{AgNO}_3$  will produce 1 mole of Ag. Now we can solve our problem by using the molecular formula of  $\text{AgNO}_3$  to calculate the molar mass of  $\text{AgNO}_3$ . Once we know the molar mass, we can use it to convert the *mass* of silver nitrate (47.2 g) into *moles* of silver nitrate. If we know how many moles of  $\text{AgNO}_3$  we have, then we also know how many moles of Ag can be obtained (one mole of  $\text{AgNO}_3$  will produce one mole of Ag atoms). Finally, if we know how many moles of silver we will form, then we can use the molar mass of silver to convert moles of silver into grams of silver.

Now think about the plan we just outlined for solving this problem. We convert grams into moles (for  $\text{AgNO}_3$ ), and then we convert moles back into grams (for Ag). Essentially, we will be using molar mass as a conversion factor *twice*. Now let's use our plan to solve the problem.

**EXERCISE 2.33.** How many grams of silver (Ag) could be obtained if we removed all of the silver from 47.2 g of silver nitrate ( $\text{AgNO}_3$ )?

**Answer:** Start by calculating the molar mass of  $\text{AgNO}_3$ . Use the periodic table to do this (looking at the numbers for Ag, N, and O). You will see that the molar mass of  $\text{AgNO}_3 = (107.9 \text{ g/mol}) + (14.01 \text{ g/mol}) + 3(16.00 \text{ g/mol}) = 169.9 \text{ g/mol}$ .

Therefore, 1 mole of  $\text{AgNO}_3 = 169.9 \text{ g}$  of  $\text{AgNO}_3$ . Use this as one of your conversion factors in calculating your answer. In total, you need three conversion factors (look at the previous diagram). When you string these three conversion factors together, you get:

$$47.2 \text{ g AgNO}_3 \times \frac{1 \text{ mol AgNO}_3}{169.9 \text{ g AgNO}_3} \times \frac{1 \text{ mol Ag}}{1 \text{ mol AgNO}_3} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}}$$

If you stare at the solution above, you will see all three conversion factors. The first one converts *grams* of silver nitrate into *moles* of silver nitrate. The sec-

ond conversion factor is used to convert moles of *silver nitrate* into moles of *silver*, and the third conversion factor is used to convert *moles* of silver into *grams* of silver. When you calculate the above solution, you get: **30.0 g of Ag**

Now let's get some practice using the exact same procedure that we used above. You will need to use three conversion factors to get the answer. Try to set up the solution to this problem without looking back at the previous problem. If you get stuck, then look back to the previous problem and model your answer after that problem. Then, move on to the next problem, trying not to look back to formulate your solution. Keep going through the following problems until you can set up the solution without looking back.

PROBLEM 2.34. How many grams of silver could be obtained if we removed all of the silver from 12.07 g of silver chloride (AgCl)?

Molar mass of AgCl = \_\_\_\_\_

Solution = \_\_\_\_\_

PROBLEM 2.35. How many grams of calcium could be obtained if we removed all of the calcium from 2.20 g of calcium carbonate (CaCO<sub>3</sub>)?

Molar mass of CaCO<sub>3</sub> = \_\_\_\_\_

Solution = \_\_\_\_\_

PROBLEM 2.36. How many grams of copper could be obtained if we removed all of the copper from 3.7 g of copper iodide (CuI)?

Answer: \_\_\_\_\_

PROBLEM 2.37. How many grams of silver could be obtained if we removed all of the silver from 6.75 g of silver oxide (Ag<sub>2</sub>O)? (*Hint*: 1 mole of Ag<sub>2</sub>O will yield 2 moles of silver atoms.)

Answer: \_\_\_\_\_

We have now observed that even apparently complex problems can be greatly simplified if you can train your eye to see when you can use conversion factors. If you want to do well on your exams, it is important that you train yourself to think in this way. And we are not just talking about molar mass, but we are talking about ANY conversion factor that you learned. So far, we have seen two special conversion factors that can be used to convert one type of unit into another: first we saw density in Chapter 1, and now we have seen molar mass. To make sure that we train

our eyes to read problems in this way, let's just work through the following simple "fill-in-the-blank" problems:

PROBLEM 2.38. You are working on a problem that starts with the following words: "2.3 grams of silver sulfate ( $\text{Ag}_2\text{SO}_4$ ) . . .". As you read this problem, you should think to yourself the following: "I have been given the molecular formula, which I can use to calculate the \_\_\_\_\_ With that, I can convert grams of silver sulfate into \_\_\_\_\_ of silver sulfate.

PROBLEM 2.39. You are working on a problem that starts with the following words: "9.7 moles of magnesium oxalate ( $\text{MgC}_2\text{O}_4$ ) . . .". As you read this problem you should think to yourself the following: "I have been given the molecular formula, which I can use to calculate the \_\_\_\_\_ With that, I can convert moles of magnesium oxalate into \_\_\_\_\_ of magnesium oxalate.

## 2.7 COMBINING CONVERSION FACTORS

As you progress through this course, you will see problems where you need to use multiple conversion factors. To illustrate what we mean by this, consider the two special conversion factors that we have seen so far: *density* and *molar mass*. Let's quickly review what we have seen:

<i>Density</i>	allows us to interconvert	<i>mass and volume</i>
<i>Molar mass</i>	allows us to interconvert	<i>mass and moles</i>

Notice that mass can be interchanged in each of these conversions. So, if we put them together, we can actually convert moles into volume, or volume into moles, like this:



Let's see an example:

EXERCISE 2.40. At  $25^\circ\text{C}$ , the density of water is  $1.00\text{ g/cm}^3$ . At this temperature, what volume will be occupied by 3.89 moles of water ( $\text{H}_2\text{O}$ )?

Answer: This problem is asking you to convert *moles* of water (3.89 moles) into a measurement of *volume*. To do this, you will need to do *two* conversions (as shown in the previous diagram). You are given the molecular formula of water, so you can easily calculate the molar mass (using the values on the periodic table, you get  $18.02\text{ g/mol}$ ), which you can use to convert moles into grams. Then you are given the den-

sity, which you can use to convert from grams into volume ( $\text{g}/\text{cm}^3$ ). So, your solution will have two conversions:

$$3.89 \text{ mol} \times \frac{18.02 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ cm}^3}{1.00 \text{ g}}$$

For each of these conversions, you just need to consider which fraction to use (which units go into the numerator and which go in the denominator) in order that your units should all cross off in the end (you practiced this Chapter 1). After doing the calculation above, your answer should be **70.1 cm<sup>3</sup>**.

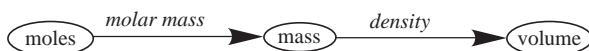
PROBLEM 2.41. At 20° C, the density of ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) is  $0.789 \text{ g}/\text{cm}^3$ . At this temperature, what volume will be occupied by 1.50 moles of ethanol?

Answer: \_\_\_\_\_

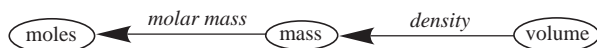
PROBLEM 2.42. At 20° C, the density of gold (Au) is  $19.3 \text{ g}/\text{cm}^3$ . At this temperature, what volume will be occupied by a bar of gold made of 9.5 moles of gold atoms?

Answer: \_\_\_\_\_

In the previous problems, we saw how to use two special conversion factors (density and molar mass) to convert from *moles* into *volume*. In these problems, we used the molar mass to convert *moles* into *grams*, and then we use density to convert *grams* into *volume*:



Similarly, we can do the exact same process in the other direction. We can convert *volume* into *moles* using the same two conversion factors. We would start by converting *volume* into *grams* (using density), and then by converting *grams* into *moles* (using molar mass):



Let's see an example:

EXERCISE 2.43. The density of ethylene glycol ( $\text{C}_2\text{H}_6\text{O}_2$ ) at 25° C is  $1.09 \text{ g}/\text{cm}^3$ . If you have a  $100.0 \text{ cm}^3$  container filled with ethylene glycol, how many moles of ethylene glycol are in the container?

Answer: In this problem, you must convert volume into moles. So two conversion factors are needed: density and molar mass. You were given the density, and

you will have to calculate the molar mass from the molecular formula. The molar mass of ethylene is:

$$2(12.01 \text{ g/mol}) + 6(1.008 \text{ g/mol}) + 2(16.00 \text{ g/mol}) = 62.07 \text{ g/mol}$$

Now you have both conversion factors that you need. In order to convert volume into moles, you will need to first convert volume into grams, and then grams into moles. So set up the following solution:

$$100.0 \text{ cm}^3 \times \frac{1.09 \text{ g C}_2\text{H}_6\text{O}_2}{1 \text{ cm}^3} \times \frac{1 \text{ mol C}_2\text{H}_6\text{O}_2}{62.07 \text{ g C}_2\text{H}_6\text{O}_2}$$

And you get: **1.76 moles of C<sub>2</sub>H<sub>6</sub>O<sub>2</sub>.**

PROBLEM 2.44. The density of ethanol (C<sub>2</sub>H<sub>5</sub>OH) at 25° C is 0.79 g/cm<sup>3</sup>. If you have a 10.5 cm<sup>3</sup> container filled with ethanol, how many moles of ethanol are in the container?

Answer: \_\_\_\_\_

PROBLEM 2.45. The density of gold (Au) at 25° C is 19.32 g/cm<sup>3</sup>. If you have a bar of gold that occupies a volume of 3.4 cm<sup>3</sup>, how many moles of gold do you have?

Answer: \_\_\_\_\_

PROBLEM 2.46. The density of iron (Fe) at 25° C is 7.9 g/cm<sup>3</sup>. If you have a bar of iron that occupies a volume of 5.2 cm<sup>3</sup>, how many moles of iron do you have?

Answer: \_\_\_\_\_

## 2.8 MASS PERCENT AND ELEMENTAL ANALYSIS

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In the beginning of this chapter, we saw the difference between an empirical formula and a molecular formula. At the time, we saw that a molecular formula is more useful because it tells us the exact number of each atom in the molecule. An empirical formula only tells us the ratio of atoms in the molecule. You might wonder why we use empirical formulas at all if they are less useful? It turns out that we can easily determine the empirical formula of an unknown compound by running one simple experiment called *elemental analysis*. In this section, we will see how to use the data from an elemental analysis to determine an empirical formula.

An elemental analysis tells us the amount of grams of each type of atom present in the molecule. We will soon see how to use this information, but first we need to focus on a term called *mass percent* (or mass %). To understand what mass %

is, let's go back to our Lego analogy. Imagine that we went to the Lego warehouse and we bought one barrel containing the following structures:



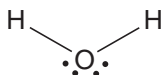
On our way out of the warehouse, we weigh our barrel and we find that the barrel weighs 100 lb. Suppose we wanted to know how much of this weight comes from the black pieces? It might be tempting to say this: there are three Lego pieces in every structure, and only one of them is black, so black pieces should be one third of the total weight (or 33 lb). The problem with that argument is that black pieces and white pieces have different weights. And we have to take that into account.

So we do it like this. We take our barrel (which has 6,022,042 structures in it), and we dump the contents onto the floor. Then we break apart all of the structures into individual Lego pieces, and we get over 12 million white pieces and over 6 million black pieces. We then sort the pieces into two piles; one pile of white pieces and one pile of black pieces. When we weigh all of the white pieces together, we get 20 lb. When we weigh all of the black pieces together, we get 80 lb. Now we have our answer. The *mass %* of the black pieces is therefore:

$$\frac{\text{mass of black pieces}}{\text{total mass}} \times 100\% = \frac{80 \text{ lb}}{100 \text{ lb}} \times 100\% = 80\%$$

We can say that the black pieces contributed 80% of the weight of the original barrel, and similarly, the white pieces contributed 20% of the weight of the original barrel. That's what we mean by *mass %*.

Now let's do the same thing with water molecules instead of Lego pieces:



If we take 1 mole of water, and we measure its mass, we see that it is 18.02 grams. If we want to know how much of this mass comes from oxygen atoms, we need to look at the mass of 1 mole of oxygen atoms. One mole of oxygen atoms is 16.00 grams. So, the *mass %* of oxygen atoms in 1 mole of  $\text{H}_2\text{O}$  is:

$$\frac{\text{mass of oxygen atoms}}{\text{total mass}} \times 100\% = \frac{16.00 \text{ g}}{18.02 \text{ g}} \times 100\% = 88.79\%$$

If we did the same kind of calculation for hydrogen, we would find that the *mass %* of hydrogen in 1 mole of  $\text{H}_2\text{O}$  is 11.21%. This should make sense because  $88.79\% + 11.21\% = 100\%$ , which is what we would expect. Sometimes rounding the individual mass percentages can cause the sum total to be 99.99%. Don't worry

about that—it is just an artifact of rounding the numbers, which is what we have to do when we report numbers with the appropriate significant figures.

The concept of mass % is a straightforward concept, but it only seems that way when you have the molecular formula in front of you. Suppose you have an unknown compound and you are given the percentage composition (that is just a list of the mass % of each element in the compound). Will you be able to reconstruct an empirical formula? What about a molecular formula? We will focus on these topics in just a few moments. For now, let's just do some simple exercises to make sure that you are comfortable with the concept of mass %.

**EXERCISE 2.47.** Consider the compound called hydrazine ( $\text{N}_2\text{H}_4$ ). What is the mass % of nitrogen in this compound?

**Answer:** First you must determine the molar mass, which can easily be derived from the molecular formula:  $2(14.01 \text{ g/mol}) + 4(1.008 \text{ g/mol}) = 32.05 \text{ g/mol}$ . So, 1 mole of hydrazine has a mass of 32.05 g (or we can say that 1 *molecule* of hydrazine has a mass of 32.05 u).

Next, we look at how much of this mass comes from nitrogen:

$$\frac{\text{mass of nitrogen atoms}}{\text{total mass}} \times 100\% = \frac{2(14.01) \text{ g}}{32.05 \text{ g}} \times 100\% = 87.43\%$$

Notice that we could have done this calculation in atomic mass units instead of grams (essentially looking at 1 molecule instead of 1 mole), and we would get the exact same answer.

**PROBLEM 2.48.** Calculate the mass % of calcium in calcium chloride ( $\text{CaCl}_2$ ):

**PROBLEM 2.49.** Calculate the mass % of chlorine in calcium chloride ( $\text{CaCl}_2$ ). Make sure that your answer to this problem plus your answer to the previous problem = 100%.

PROBLEM 2.50. Calculate the mass % of zinc in zinc sulfide (ZnS):

PROBLEM 2.51. Calculate the mass % of carbon in butanoic acid (C<sub>4</sub>H<sub>8</sub>O<sub>2</sub>):

PROBLEM 2.52. Calculate the mass % of oxygen in oxalic acid (C<sub>2</sub>H<sub>2</sub>O<sub>4</sub>):

In each of these problems, we knew the molecular formula, and we were able to use that information to calculate mass percentages. More often, we deal with *unknown* compounds, where we don't know what the molecular formula is. This is where elemental analysis becomes important. An elemental analysis is an experiment that tells us how many grams of each element were present in the original compound. With this information, it is actually easier to calculate the percentage composition. For example, suppose we have an unknown sample with a mass of 33.55 g. An elemental analysis produces 13.42 g of carbon, so we can calculate the mass % of carbon like this:

$$\frac{13.42 \text{ g}}{33.55 \text{ g}} \times 100\% = 40.00\%$$

This is the easiest way to calculate mass %. We don't have to bother with the molecular formula or with calculating molar mass, like we did earlier. We just use the information from an elemental analysis and we're done. Since it's so easy, you probably won't see too many problems like this, but there is a common type of problem that you will definitely see. When given the information from an elemental analysis, you will see problems where you must determine an empirical formula.

The procedure for doing this is not difficult, because it just involves a procedure that we have already seen and mastered: converting grams into moles. We have already seen how to convert grams into moles using molar mass as a special conversion factor. Now, we don't know the molar mass of the unknown compound, because we don't know its molecular formula. But, we do know the molar mass of individual elements. All we have to do is look at the periodic table for that. For example, we have already seen that a single carbon atom will have a mass of 12.01 u, and a mole of carbon atoms will have a mass of 12.01 grams. So, the molar mass of carbon atoms is just 12.01 g/mol. We can use that molar mass to convert *grams* of carbon into *moles* of carbon. And that is the whole technique behind solving these problems. Let's see an example of how we do it.

EXERCISE 2.53. A sample was found to contain 13.42 g of carbon, 2.25 g of hydrogen, and 17.88 g of oxygen. What is the empirical formula for this compound?

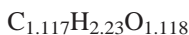
Answer: Whenever you are given the information from an elemental analysis (amounts of elements listed in grams), you will have to convert each of these amounts into moles. To do this, you will need the molar mass for carbon atoms, for hydrogen atoms, and for oxygen atoms. You get each of these values directly from the periodic table. The molar mass of carbon is 12.01 g/mol, the molar mass of hydrogen is 1.008 g/mol, and the molar mass of oxygen is 16.00 g/mol. Now, you can convert all of the data you were given into moles, using the factor-label method:

$$13.42 \text{ g carbon} \times \frac{1 \text{ mol carbon}}{12.01 \text{ g carbon}} = 1.117 \text{ mol carbon}$$

$$2.25 \text{ g hydrogen} \times \frac{1 \text{ mol hydrogen}}{1.008 \text{ g hydrogen}} = 2.23 \text{ mol hydrogen}$$

$$17.88 \text{ g oxygen} \times \frac{1 \text{ mol oxygen}}{16.00 \text{ g oxygen}} = 1.118 \text{ mol oxygen}$$

Now you know how many moles of each element were produced in the elemental analysis. You can summarize the ratio like this:



But you cannot report your answer like this, because as you'll recall, an empirical formula uses the smallest whole digits possible as subscripts. The subscripts above are not whole numbers. You might notice right away that the ratio above is just

CH<sub>2</sub>O, but sometimes you might not notice the ratio so readily. So, let's see the procedure you can use whenever you get stuck.

The trick goes like this: just choose the smallest number, and divide all other numbers by that number. In this case, the smallest number is 1.117. If you divide all of these numbers by 1.117, you get:

$$\frac{1.117}{1.117} = 1.000 \quad \frac{2.23}{1.117} = 2.00 \quad \frac{1.118}{1.117} = 1.001$$

So, with the ratio is 1:2:1 that is how you get:



Notice that, in the example above, you were only able to calculate the *empirical* formula. We still don't know what the molecular formula is. If the empirical formula is CH<sub>2</sub>O, then the molecular formula could be CH<sub>2</sub>O or C<sub>2</sub>H<sub>4</sub>O<sub>2</sub> or C<sub>3</sub>H<sub>6</sub>O<sub>3</sub>. . . . In order to determine which molecular formula you have, you will need one more piece of information. Before we do this, though, let's get practice doing problems like the example above.

Remember that the technique is to convert all the data from grams into moles. To do this you will use the molar mass of each element, which you can get from the periodic table. Once you convert everything into moles, divide all numbers by the smallest number, and you will have the ratio you need for your empirical formula.

PROBLEM 2.54. A sample was found to contain: 50.44 g of carbon, 12.70 g of hydrogen, and 33.60 g of oxygen. What is the empirical formula for this compound?

Answer: \_\_\_\_\_

PROBLEM 2.55. A sample was found to contain: 43.29 g of sulfur and 64.8 g of oxygen. What is the empirical formula for this compound?

Answer: \_\_\_\_\_

PROBLEM 2.56. A sample was found to contain: 8.982 g of phosphorous and 0.8770 g of hydrogen. What is the empirical formula for this compound?

Answer: \_\_\_\_\_

Sometimes, you will see a problem just like the problems above, BUT there will be one minor difference. Instead of giving you the data from elemental analysis *in grams* (X grams of carbon, Y grams of oxygen, etc.), the problem will give you a *percentage composition* (mass % of carbon is X%, mass % of oxygen is Y%, etc.). When you see a problem like this, there is a simple trick to put this problem into the form of all the problems above. All you do is pretend that you started with 100 g of sample. In this case, 62.5% carbon is the same as 62.5 grams of carbon.

By pretending to start with a sample size of 100 grams, you are essentially converting percentage composition into mass composition. Let's see an example:

EXERCISE 2.57. A sample has the following percentage composition: 40.00% carbon, 6.71% hydrogen, and 53.29% oxygen. What is the empirical formula for this compound?

Answer: In this problem, you are not given the actual grams of carbon, hydrogen, and oxygen. So, it seems like a different type of problem than what we are used to. Here is the trick that we just mentioned. The percentages above are true for any amount of substance, whether it is 10 grams, or 35 grams, or 100 grams of sample. If we imagine a sample size of 100 grams, then the problem becomes much simpler. 40.00% carbon becomes 40.00 grams of carbon. 6.71% hydrogen becomes 6.71 grams of hydrogen. And 53.29% oxygen becomes 53.29 grams of oxygen. And now the problem is in exactly the same format as all of the problems we just did.

Now that each of the elements is measured in *grams*, you will have to convert each of these amounts into *moles*. To do this, you will need the molar mass for carbon atoms, hydrogen atoms, and oxygen atoms. You get each of these values directly from the periodic table. The molar mass of carbon is 12.01 g/mol, the molar mass of hydrogen is 1.008 g/mol, and the molar mass of oxygen is 16.00 g/mol. You can now convert all of the data we were given into moles, using the factor-label method:

$$40.00 \text{ g carbon} \times \frac{1 \text{ mol carbon}}{12.01 \text{ g carbon}} = 3.331 \text{ mol carbon}$$

$$6.71 \text{ g hydrogen} \times \frac{1 \text{ mol hydrogen}}{1.008 \text{ g hydrogen}} = 6.66 \text{ mol hydrogen}$$

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

Now you know how many moles of each element were produced in the elemental analysis. You can summarize the ratio like this:



But you cannot report your answer like this because, as you'll remember, an empirical formula uses the smallest whole digits possible as subscripts. The ratio is 1:2:1, so your empirical formula is:



To make sure you understand the trick, try to do a problem that gives you elemental analysis data in the format of percentage composition.

PROBLEM 2.58. A sample has the following percentage composition: 50.05% sulfur and 49.95% oxygen. What is the empirical formula for this compound?

Answer: \_\_\_\_\_

In all of the problems we have just seen, we have been able to reconstruct an empirical formula from experimental data (elemental analysis) of an unknown compound. But what if we wanted to reconstruct a molecular formula? We have already seen that a molecular formula is much more useful. It turns out that we only need one more piece of information to determine the molecular formula. We need to be given the molar mass. If a problem gives us elemental analysis data AND a molar mass of the unknown compound, then we can determine the molecular formula. To see how this works, consider one of the problems that we recently did.

In exercise 2.51 (just a few pages ago), we were given the data from an elemental analysis, and we used those data to determine that the empirical formula was  $\text{CH}_2\text{O}$ . And that was where we left off. At the time, there was no way for us to tell if we were talking about  $\text{CH}_2\text{O}$  or  $\text{C}_2\text{H}_4\text{O}_2$  or  $\text{C}_3\text{H}_6\text{O}_3$  . . . etc. All of these formulas have the same ratio of carbon to hydrogen. Since we don't know which one it is, we write the empirical formula as the simplest whole numbers that show this ratio. In this case, that is  $\text{CH}_2\text{O}$ . But if we had been given the molar mass, then we would have been able to determine which one is the correct molecular formula. Let's see how this works:

EXERCISE 2.59. In Exercise 2.51, you determined that the empirical formula was  $\text{CH}_2\text{O}$ . The molar mass of this compound is 60.05 g/mol. What is the molecular formula?

Answer: If you calculate the molar mass of the compound  $\text{CH}_2\text{O}$ , you find that it is 30.03 g/mol. The compound you want is 60.05 g/mol, so you have to multiply all of your subscripts by a factor of two. When you do this, you get:



That's all there is to it.

Now let's combine it all together in the following three final problems for this chapter.

PROBLEM 2.60. A sample was found to contain: 2.270 g of carbon and 0.19105 g of hydrogen. The molar mass of the compound is 78.11 g/mol. What is the molecular formula for this compound?

PROBLEM 2.61. A sample was found to contain: 1.084 g of hydrogen, 15.06 g of nitrogen, and 34.40 g of oxygen. The molar mass of the compound is 47.02 g/mol. What is the molecular formula for this compound?

PROBLEM 2.62. A sample was found to contain: 1.062 g of hydrogen, 133.757 g of iodine, and 50.589 g of oxygen. The molar mass of the compound is 175.91 g/mol. What is the molecular formula for this compound?

QU1

Wording as wanted?