Chapter 9
Chemical Calculations and Chemical Formulas

In order to explore and make use of the seemingly limitless changes that matter can undergo, chemists and chemistry students often need to answer questions that begin with, "*How much*…?" The research chemist who is developing a new cancer drug wants to know, "*How much* radioactive boron-10 do I need to make 5 g of the drug?" At a plant where the fat substitute Olestra is manufactured from white sugar and vegetable oil, a business manager asks a chemist, "*How much* sucrose and cottonseed oil should I order if we need to produce 500 Mg of Olestra per day?" In an experiment for a chemistry course you are taking, you might be asked, "How much magnesium oxide can be formed from the reaction of your measured mass of magnesium with the oxygen in the air?" This chapter and the chapters that follow provide you with the tools necessary to answer these questions and many others like them.

All of these people ask questions that begin, "How much...?"

Review Skills

The presentation of information in this chapter assumes that you can already perform the tasks listed below. You can test your readiness to proceed by answering the Review Questions at the end of the chapter. This might also be a good time to read the Chapter Objectives, which precede the Review Questions.

- Write or recognize the definition of isotope. (Section 2.4)
- Describe the general structure of molecular and ionic compounds. (Sections 3.3 and 3.5)
- Convert between the names of compounds and their chemical formulas. (Section 5.3)
- Report the answers to calculations to the correct number of significant figures. (Section 8.2)
- Use percentages as conversion factors. (Section 8.4)
- Make unit conversions. (Section 8.5)

9.1 A Typical Problem

Imagine you are a chemist at a company that makes phosphoric acid, H_3PO_4, for use in the production of fertilizers, detergents, and pharmaceuticals. The goal for your department is to produce 84.0% H_3PO_4 (and 16.0% water) in the last stage of a three-step process known as the furnace method. (We will be using the furnace method as a source of practical examples throughout this chapter and in Chapter 10.)

The first step in the furnace method is the extraction of phosphorus from phosphate rock by heating the rock with sand and coke to 2000 °C. The phosphate rock contains calcium phosphate, $Ca_3(PO_4)_2$; the sand contains silicon dioxide, SiO_2; and coke is a carbon-rich substance that can be produced by heating coal to a high temperature. At 2000 °C, these three substances react as follows:[1]

$$2Ca_3(PO_4)_2 + 6SiO_2 + 10C \rightarrow 4P + 10CO + 6CaSiO_3$$

In the second step, the step on which we will focus in this chapter, the phosphorus, P, is reacted with oxygen in air to form tetraphosphorus decoxide, P_4O_{10}:

$$4P(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$$

The third and final step is the reaction of P_4O_{10} with water to form phosphoric acid:

$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$$

Your colleagues who are responsible for the first step, the production of pure phosphorus, estimate that they can supply you with 1.09×10^4 kg of phosphorus per day. The production manager asks you to figure out the maximum mass of P_4O_{10} that can be made from this amount of phosphorus. The tools described in this chapter, in combination with the unit analysis method, enable you to satisfy this request.

Let's begin by thinking about how unit analysis might be used to solve your problem. Your production manager is asking you to convert from amount of phosphorus, P, to amount of tetraphosphorus decoxide, P_4O_{10}. To do this, you need a conversion factor relating amount of P to amount of P_4O_{10}. The chemical formula, P_4O_{10}, provides such a conversion factor. It shows that there are four atoms of phosphorus for each molecule of P_4O_{10}.

$$\frac{4 \text{ atoms P}}{1 \text{ molecule}} \quad \text{or} \quad \frac{1 \text{ molecule}}{4 \text{ atoms P}}$$

[1] To avoid potential confusion, the states—(*s*), (*l*), (*g*), and (*aq*)—are not mentioned in equations describing industrial reactions in Chapters 9 and 10. Many industrial reactions are run at temperatures and pressures under which the states of substances differ from what would be expected at room temperatures and pressures. For example, the $Ca_3(PO_4)_2$ in the first equation on this page is a solid under normal conditions but a liquid at 2000 °C; SiO_2 is also solid under normal conditions but a glass (or semisolid) at 2000 °C.

But how many atoms of phosphorus are in our 1.09×10^4 kg P? Until we know that, how can we tell how much P_4O_{10} we'll be able to produce? Unfortunately, atoms and molecules are so small and numerous that they cannot be counted directly. *We need a conversion factor that converts back and forth between any mass of the element and the number of atoms contained in that mass.*

If we can determine the number of phosphorus atoms in 1.09×10^4 kg P, we can use the second conversion factor at the bottom opf the last page to determine the number of molecules of P_4O_{10}. But your production manager doesn't want to know the number of P_4O_{10} *molecules* that you can make. She wants to know the *mass* of P_4O_{10} that can be made. That means *we also need a conversion factor that converts back and forth between any mass of the compound and the number of molecules contained in that mass*. The main goal of the first half of this chapter is to develop conversion factors that convert between mass and number of particles.

An industrial chemist must be able to calculate the amounts of reactants necessary to run an efficient reaction and predict the amount of product that will form.

9.2 Relating Mass to Number of Particles

Before we develop conversion factors to convert back and forth between any mass of an element and the number of atoms contained in that mass, let's consider a similar task that may be easier to visualize. We will calculate the number of carpenter's nails in a hardware store bin.

Imagine that you have decided to make a little money for schoolbooks by taking a temporary job doing inventory at the local hardware store. Your first task is to determine how many nails are in a bin filled with nails. There are a lot of nails in the bin, and you do not want to take the time to count them one by one. It would be a lot easier to weigh the nails and calculate the number of nails from the mass. To do this you need a conversion factor that allows you to convert from the mass of the nails in the bin to the number of nails in the bin.

Let's assume that you individually weigh 100 of the nails from the bin and find that 82 of them have a mass of 3.80 g each, 14 of them have a mass of 3.70 g each, and the last four have a mass of 3.60 g each. From this information, you can calculate the average mass of the nails in this sample, taking into consideration that 82% of the nails have a mass of 3.80 g, 14% have a mass of 3.70 g, and 4% have a mass of 3.60 g. Such an average is known as a **weighted average**. You can calculate the weighted average of the nails' masses by multiplying the decimal fraction of each subgroup of nails times the mass of one of its members and adding the results of these multiplications.

$$0.82(3.80\text{ g}) + 0.14(3.70\text{ g}) + 0.04(3.60\text{ g}) = 3.78\text{ g}$$

Thus the weighted average mass of the nails in this sample is 3.78 g. It is possible that none of the nails in our bin has this mass, but this is a good description of what we can expect the average mass of each nail in a large number of nails to be. It can be used as a conversion factor to convert between mass of nails and number of nails.

$$\left(\frac{3.78 \text{ g nails}}{1 \text{ nail}}\right)$$

We can measure the total mass of the nails in the bin and then use the conversion factor above to determine their number. For example, if the nails in the bin are found to weigh 218 pounds, the number of nails in the bin is:

$$? \text{ nails} = 218 \cancel{\text{lb}} \cancel{\text{nails}} \left(\frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}}\right)\left(\frac{1 \text{ nail}}{3.78 \cancel{\text{g}} \cancel{\text{nails}}}\right) = 2.62 \times 10^4 \text{ nails}$$

There is some uncertainty in this result due to our reliance on measuring rather than counting, but this procedure is a lot faster than the alternative of counting over 26,000 nails.

The key point to remember is that this procedure allows us to determine the number of objects in a sample of a large number of those objects without actually counting them. Our procedure allows us to count by weighing.

It is often convenient to describe numbers of objects in terms of a collective unit such as a dozen (12) or a gross (144). The number 2.62×10^4 is large and inconvenient to use. We might therefore prefer to describe the number of nails in another way. For example, we could describe them in terms of dozens of nails, but 218 pounds of our nails is 2.18×10^3 dozen nails, which is still an awkward number. We could use gross instead of dozen. A gross of objects is 144 objects. The following calculation shows how to create a conversion factor that converts between mass of nails and gross of nails:

$$\frac{? \text{ g nails}}{1 \text{ gross nails}} = \left(\frac{3.78 \text{ g nails}}{1 \cancel{\text{nail}}}\right)\left(\frac{144 \cancel{\text{nails}}}{1 \text{ gross nails}}\right) = \frac{544 \text{ g nails}}{1 \text{ gross nails}}$$

We can use this conversion factor to determine the number of gross of nails in 218 pounds of nails.

$$? \text{ gross nails} = 218 \cancel{\text{lb}} \cancel{\text{nails}} \left(\frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}}\right)\left(\frac{1 \text{ gross nails}}{544 \cancel{\text{g}} \cancel{\text{nails}}}\right) = 182 \text{ gross nails}$$

Atomic Mass and Counting Atoms by Weighing

Now let's take similar steps to "count" atoms of the element carbon. Because of the size and number of carbon atoms in any normal sample of carbon, it is impossible to count the atoms directly. Therefore, we want to develop a way of converting from mass of carbon, which we can measure, to the number of carbon atoms. To do this, we will follow steps that are similar to those we followed to "count" nails by weighing.

First, we need to know the masses of individual atoms of carbon. To describe the mass of something as small as an atom of carbon, we need a unit whose magnitude (or lack of it) is correspondingly small. The unit most often used to describe the mass of atoms is the atomic mass unit, whose symbol is **u** or **amu**. An **atomic mass unit** is defined as exactly one-twelfth the mass of an atom of carbon-12. Carbon-12 is the isotope of carbon that contains six protons, six neutrons, and six electrons. (You might want to review Section 2.5, which describes isotopes.) One atomic mass unit is equivalent to 1.660540×10^{-24} grams.

$$\begin{aligned} 1 \text{ atomic mass unit (u)} &= \frac{1}{12} \text{ mass of one carbon-12 atom} \\ &= 1.660540 \times 10^{-24} \text{ g} \end{aligned}$$

To generate a relationship between mass of carbon and number of carbon atoms, we need to know the weighted average mass of the carbon atoms found in nature. Experiments show that 98.90% of the carbon atoms in natural carbon are carbon-12, and 1.10% are carbon-13, with six protons, seven neutrons, and six electrons. Related experiments show that each carbon-13 atom has a mass of 13.003355 u. From the definition of atomic mass unit, we know that the mass of each carbon-12 atom is 12 u. The following setup shows how the weighted average mass of carbon atoms is calculated.

$$0.9890 \text{ (12 u)} + 0.0110 \text{ (13.003355 u)} = 12.011 \text{ u}$$

This value is carbon's atomic mass. Because an element's atomic mass is often described without units, carbon's atomic mass[1] is usually described as 12.011 instead of 12.011 u. The **atomic mass** of any element is the weighted average of the masses of the naturally occurring isotopes of the element. (It is very common to call this property atomic weight, but because it describes the masses of the atoms, not their weights, this text will use the term atomic mass.) Scientists have calculated the atomic masses of all elements that have stable isotopes, and they can be found on any standard periodic table, including the table in this book.

Note that no carbon atom has a mass of 12.011 u. This value is the weighted average mass of the carbon atoms found in nature. It leads to the following conversion factor for natural carbon.

$$\left(\frac{12.011 \text{ u C}}{1 \text{ C atom}}\right)$$

Although we can use the conversion factor shown above to convert between mass of carbon in atomic mass units and number of carbon atoms, let's wait to do this type of calculation until we take the next step of describing the number of atoms with a convenient collective unit, analogous to a dozen or a gross.

Just one gram of carbon has over 10^{22} carbon atoms. A dozen and a gross are both too small to be useful for conveniently describing this number of atoms. Thus chemists have created a special collective unit, called the mole, which is similar to but much greater than a dozen or a gross. A **mole** (which is abbreviated mol) is an amount of substance that contains the same number of particles as there are atoms in 12 g of carbon-12. To four significant figures, there are 6.022×10^{23} atoms in 12 g of carbon-12. Thus a mole of natural carbon is the amount of carbon that contains 6.022×10^{23} carbon atoms. The number 6.022×10^{23} is often called **Avogadro's number**.

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[1] The atomic mass without units is perhaps more correctly called the relative atomic mass (relative to the mass of carbon-12 as 12 u). We will take the common approach of calling both 12.011 u and 12.011 the atomic mass of carbon.

Objective 2

The mole is used in very much the same way as we use the collective units of trio and dozen. There are 3 items in 1 trio, as in 3 musicians in a jazz trio.

$$\left(\frac{3 \text{ musicians}}{1 \text{ jazz trio}}\right) \quad \text{or} \quad \left(\frac{3 \text{ anything}}{1 \text{ trio of anything}}\right)$$

There are 12 items in 1 dozen, as in 12 eggs in a dozen eggs.

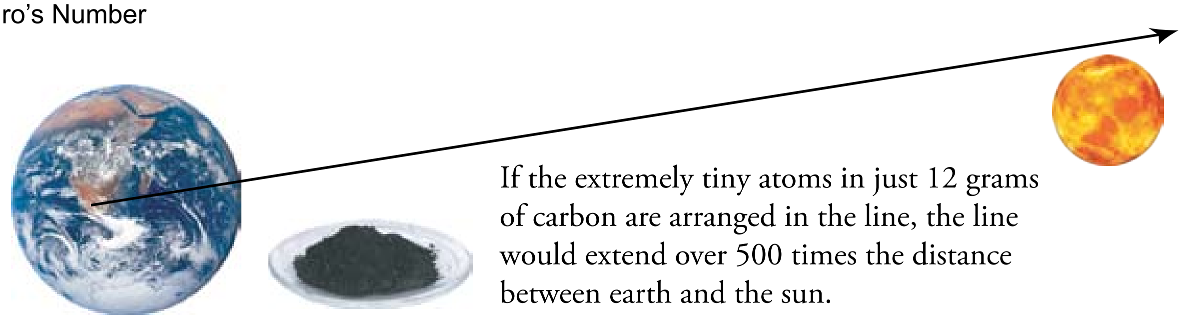
$$\left(\frac{12 \text{ eggs}}{1 \text{ dozen eggs}}\right) \quad \text{or} \quad \left(\frac{12 \text{ anything}}{1 \text{ dozen anything}}\right)$$

There are 6.022×10^{23} items in 1 mole, as in 6.022×10^{23} carbon-12 atoms in a mole of carbon-12.

$$\left(\frac{6.022 \times 10^{23}\ {}^{12}_{6}\text{C atoms}}{1 \text{ mol } {}^{12}_{6}\text{C}}\right) \quad \text{or} \quad \left(\frac{6.022 \times 10^{23} \text{ anything}}{1 \text{ mol anything}}\right)$$

Avogadro's number is unimaginably huge. For example, even though a carbon atom is extremely small, if you were to arrange the atoms contained in 12 grams (1 mole or 6.022×10^{23} atoms) of carbon in a straight line, the string of atoms would stretch over 500 times the average distance from Earth to the sun (Figure 9.1).

Figure 9.1
Avogadro's Number



According to the definition of mole, one mole of carbon-12 has a mass of 12 g, so the following conversion factor could be used to convert between mass of carbon-12 and moles of carbon-12.

$$\left(\frac{12 \text{ g C-12}}{1 \text{ mol C-12}}\right)$$

Unfortunately, natural carbon always contains carbon-13 atoms as well as carbon-12 atoms, so the conversion factor shown above is not very useful. We need a conversion factor that relates mass and moles of natural carbon instead. Because the average mass of the atoms in natural carbon (12.011 u) is slightly greater than the mass of each carbon-12 atom (12 u), the mass of a mole of natural carbon atoms has a mass slightly greater than the mass of a mole of carbon-12 atoms (12.011 g compared to 12 g). The following conversion factor can be used to convert between mass of natural carbon and number of moles of carbon atoms.

$$\left(\frac{12.011 \text{ g C}}{1 \text{ mol C}}\right)$$

Example 9.1 - Converting to Moles

Objective 4

The masses of diamonds and other gemstones are measured in carats. There are exactly 5 carats per gram. How many moles of carbon atoms are in a 0.55 carat diamond? (Assume that the diamond is pure carbon.)

A 0.55 carat diamond contains over 10^{23} carbon atoms.

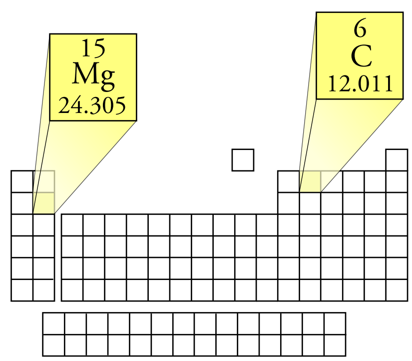
Solution

$$? \text{ mol C} = 0.55 \cancel{\text{carat C}} \left(\frac{1 \cancel{\text{g}}}{5 \cancel{\text{carat}}}\right)\left(\frac{1 \text{ mol C}}{12.011 \cancel{\text{g C}}}\right) = \mathbf{9.2 \times 10^{-3} \text{ mol C}}$$

Note that, just as we counted the nails by weighing them, we have also developed a method of counting carbon atoms by weighing. For the nails, the technique was merely convenient. If we count one nail per second, it would take over seven hours to count 26,000 nails, but we could do it if we wanted to take the time. For the carbon atoms, we have accomplished what would otherwise have been an impossible task. Even if we had the manual dexterity to pick up one carbon atom at a time, it would take us about 10^{14} centuries to count the atoms in the diamond described in Example 9.1.

Molar Mass

The mass in grams of one mole of substance is called **molar mass**. Each element has its own unique molar mass. For example, carbon's molar mass is 12.011 g/mol, and magnesium's molar mass is 24.3050 g/mol. To see why these elements have different molar masses, we need to remember that the atoms of different elements contain different numbers of protons, neutrons, and electrons, so they have different masses. The atomic masses given in the periodic table represent the different weighted average masses of the naturally occurring atoms of each element. Different atomic masses lead to different molar masses.



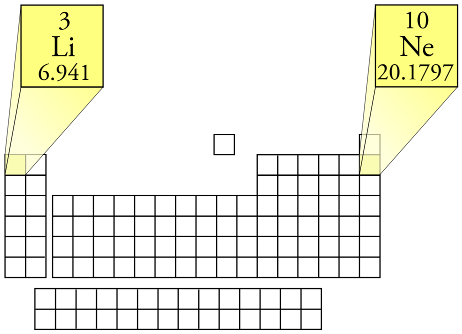
For example, the atomic mass of magnesium (24.3050) shows us that the average mass of magnesium atoms is about twice the average mass of carbon atoms (12.011), so the mass of 6.022×10^{23} magnesium atoms (the number of atoms in 1 mole of magnesium) is about twice the mass of 6.022×10^{23} carbon atoms (the number of atoms in 1 mole of carbon). Thus the molar mass of magnesium is 24.3050 g/mol, compared to carbon's molar mass of 12.011 g/mol.

The number of grams in the molar mass of an element is the same as the atomic mass. Translating atomic masses into molar masses, you can construct conversion factors that convert between the mass of an element and the number of moles of the element.

Objective 3

$$\text{Molar mass of an element} = \left(\frac{\text{(atomic mass from periodic table) g element}}{1 \text{ mol element}}\right)$$

For example, the atomic mass of the element neon listed in the periodic table is 20.1797, giving a molar mass of 20.1797 g/mol. This measurement provides the following conversion factors for converting between grams and moles of neon.



$$\left(\frac{20.1797\text{ g Ne}}{1\text{ mol Ne}}\right) \text{ or } \left(\frac{1\text{ mol Ne}}{20.1797\text{ g Ne}}\right)$$

Lithium's atomic mass is 6.941, so the conversion factors for converting between mass and moles of lithium are

$$\left(\frac{6.941\text{ g Li}}{1\text{ mol Li}}\right) \text{ or } \left(\frac{1\text{ mol Li}}{6.941\text{ g Li}}\right)$$

Example 9.2 shows how an atomic mass translates into a molar mass that allows us to convert between mass of an element and number of moles of that element.

Example 9.2 - Atomic Mass Calculations

Objective 3
Objective 4

The element boron is used as a neutron absorber in nuclear reactors. It is also used to make semiconductors and rocket propellants.

a. Write the molar mass of boron as a conversion factor that can be used to convert between grams of boron and moles of boron.
b. Calculate the mass in kilograms of 219.9 moles of boron.
c. Calculate how many moles of boron are in 0.1532 lb B.

Solution

a. The molar mass of an element comes from its atomic mass. The atomic mass of boron can be found on the periodic table inside the front cover of this text. It is **10.811**. The atomic mass of any element tells you the number of grams of that element per mole.

$$\left(\frac{\mathbf{10.811\text{ g B}}}{\mathbf{1\text{ mol B}}}\right)$$

b. $$?\text{ kg B} = 219.9\ \cancel{\text{mol B}}\left(\frac{10.811\ \cancel{\text{g}}\text{ B}}{1\ \cancel{\text{mol B}}}\right)\left(\frac{1\text{ kg}}{10^3\ \cancel{\text{g}}}\right) = \mathbf{2.377\text{ kg B}}$$

c. $$?\text{ mol B} = 0.1532\ \cancel{\text{lb B}}\left(\frac{453.6\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}}\right)\left(\frac{1\text{ mol B}}{10.811\ \cancel{\text{g B}}}\right) = \mathbf{6.428\text{ mol B}}$$

Exercise 9.1 - Atomic Mass Calculations

Objective 3
Objective 4

Gold is often sold in units of troy ounces. There are 31.10 grams per troy ounce.

a. What is the atomic mass of gold?
b. What is the mass in grams of 6.022×10^{23} gold atoms?
c. Write the molar mass of gold as a conversion factor that can be used to convert between grams of gold and moles of gold.
d. What is the mass in grams of 0.20443 moles of gold?
e. What is the mass in milligrams of 7.046×10^{-3} moles of gold?
f. How many moles of gold are in 1.00 troy ounce of pure gold?

9.3 Molar Mass and Chemical Compounds

In Section 9.2, you learned how to calculate the number of atoms—expressed in moles of atoms—in a sample of an element. Molecular substances are composed of molecules, and in this section you will learn how to calculate the number of molecules, expressed in moles of molecules, in a sample of a molecular substance. Remember that, like dozen, the collective unit mole can be used to describe the number of anything. There are 6.022×10^{23} atoms in a mole of carbon atoms, there are 6.022×10^{23} electrons in a mole of electrons, and there are 6.022×10^{23} H_2O molecules in a mole of water.

Molecular Mass and Molar Mass of Molecular Compounds

Because counting individual molecules is as impossible as counting individual atoms, we need to develop a way of converting back and forth between the number of moles of molecules in a sample of a molecular compound and the mass of the sample. To develop the tools necessary for the conversion between numbers of atoms and mass for elements, we had to determine the atomic mass of each element, which is the weighted average mass of the element's naturally occurring *atoms*. Likewise, for molecular compounds, our first step is to determine the **molecular mass** of the compound, which is the weighted average mass of the compound's naturally occurring *molecules*. This is found by adding the atomic masses of the atoms in each molecule.

Molecular mass = the sum of the atomic masses of each atom in the molecule

OBJECTIVE 5

Therefore, the molecular mass of water, H_2O, is equal to the sum of the atomic masses of two hydrogen atoms and one oxygen atom, which can be found on the periodic table.

$$\text{Molecular mass } H_2O = 2(1.00794) + 15.9994 = 18.0153$$

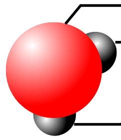
Note that the atomic mass of each element is multiplied by the number of atoms of that element in a molecule of the compound.

The number of grams in the molar mass (grams per mole) of a molecular compound is the same as its molecular mass.

$$\text{Molar mass of a molecular compound} = \left(\frac{\text{(molecular mass) g compound}}{\text{1 mol compound}}\right)$$

OBJECTIVE 6

or, for water, $\left(\dfrac{18.0153 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)$



Molar mass O: 15.9994 g/mol
Molar mass H: 1.00794 g/mol
1.00794 g/mol

Molar mass H_2O: 18.0153 g/mol

Example 9.3 shows how a molecular mass translates into a molar mass that allows us to convert between mass of a molecular compound and number of moles of that compound.

Example 9.3 - Molecular Mass Calculations

Objective 5
Objective 6
Objective 7

We have seen that the molecular compound tetraphosphorus decoxide, P_4O_{10}, is one of the substances needed for the production of phosphoric acid.

a. Write a conversion factor to convert between grams of P_4O_{10} and moles of P_4O_{10} molecules.
b. What is the mass in kilograms of 8.80 x 10^4 moles of tetraphosphorus decoxide, P_4O_{10}?

Solution

a. The molar mass of a molecular compound, such as P_4O_{10}, provides a conversion factor that converts back and forth between grams and moles of compound. This molar mass comes from the compound's molecular mass, which is the sum of the atomic masses of the atoms in a molecule. The atomic masses of phosphorus and oxygen are found on the periodic table. The atomic mass of each element is multiplied by the number of atoms of that element in a molecule of the compound.

$$\begin{aligned}\text{Molecular mass of } P_4O_{10} &= 4(\text{atomic mass P}) + 10(\text{atomic mass O}) \\ &= 4(30.9738) + 10(15.9994) \\ &= 123.895 + 159.994 = 283.889 \text{ or } 283.889 \text{ u}\end{aligned}$$

Thus there are 283.889 g P_4O_{10} in one mole of P_4O_{10}.

$$\left(\frac{283.889 \text{ g } P_4O_{10}}{1 \text{ mol } P_4O_{10}}\right)$$

b. $$? \text{ kg } P_4O_{10} = 8.80 \times 10^4 \cancel{\text{mol } P_4O_{10}} \left(\frac{283.889 \cancel{\text{g}}\ P_4O_{10}}{1 \cancel{\text{mol } P_4O_{10}}}\right)\left(\frac{1 \text{ kg}}{10^3 \cancel{\text{g}}}\right) = \mathbf{2.50 \times 10^4 \text{ kg } P_4O_{10}}$$

Exercise 9.2 - Molecular Mass Calculations

Objective 5
Objective 6
Objective 7

A typical glass of wine contains about 16 g of ethanol, C_2H_5OH.

a. What is the molecular mass of C_2H_5OH?
b. What is the mass of one mole of C_2H_5OH?
c. Write a conversion factor that will convert between mass and moles of C_2H_5OH.
d. How many moles of ethanol are in 16 grams of C_2H_5OH?
e. What is the volume in milliliters of 1.0 mole of pure C_2H_5OH? (The density of ethanol is 0.7893 g/mL.)

Ionic Compounds, Formula Units, and Formula Mass

The chemist also needs to be able to convert between mass and moles for ionic compounds. The calculations are the same as for molecular compounds, but some of the terminology is different. Remember that solid ionic compounds and molecular compounds differ in the way their particles are organized and held together (Figure 9.2). Water, a molecular substance, is composed of discrete H_2O molecules, each of which contains two hydrogen atoms and one oxygen atom. The ionic compound sodium chloride, NaCl, does not contain separate molecules. Each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. There are no sodium-chlorine atom pairs that belong together, separate from the other parts of the crystal. For this reason, we avoid using the term *molecule* when referring to an ionic compound. (You might want to review Sections 3.3 and 3.5, which describe the structure of molecular and ionic compounds.)

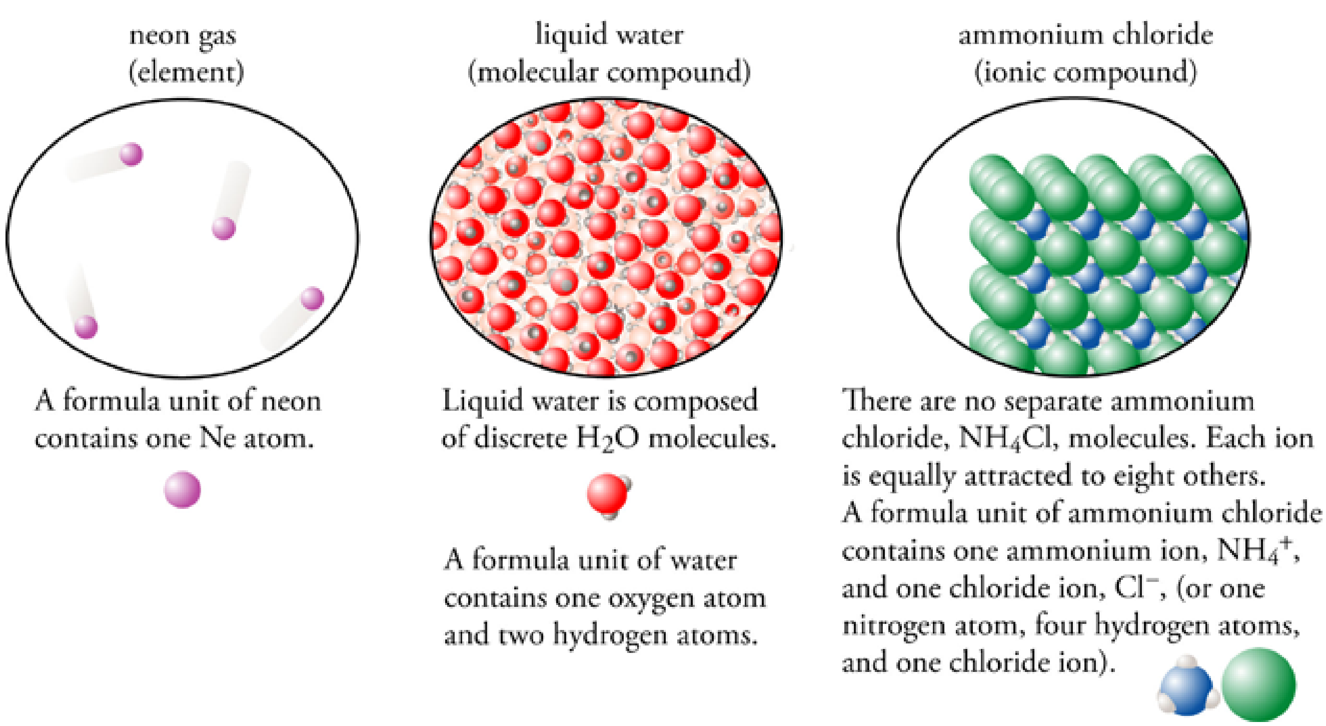
Objective 8

It is still useful to have a term *like* molecule to use when describing the composition of ionic compounds. In this text, the term *formula unit* will be used to describe ionic compounds in situations where *molecule* is used to describe molecular substances. A **formula unit** of a substance is the group represented by the substance's chemical formula, that is, a group containing the kinds and numbers of atoms or ions listed in the chemical formula. *Formula unit* is a general term that can be used in reference to elements, molecular compounds, or ionic compounds. One formula unit of the noble gas neon, Ne, contains one neon atom. In this case, the formula unit is an atom. One formula unit of water, H_2O, contains two hydrogen atoms and one oxygen atom. In this case, the formula unit is a molecule. One formula unit of ammonium chloride, NH_4Cl, contains one ammonium ion and one chloride ion—or one nitrogen atom, four hydrogen atoms, and one chloride ion (Figure 9.2).

Objective 9

Objective 9

Figure 9.2
Atoms, Molecules, and Formula Units



A sample of an element is often described in terms of the number of atoms it contains, a sample of a molecular substance can be described in terms of the number of molecules it contains, and a sample of an ionic compound is often described in terms of the number of formula units it contains. One mole of carbon contains 6.022×10^{23} carbon atoms, one mole of water contains 6.022×10^{23} molecules, and one mole of sodium chloride contains 6.022×10^{23} NaCl formula units.

As we saw with mass and moles of elements and molecular compounds, it is important to be able to convert between mass and moles of ionic substances. The development of the tools for this conversion starts with the determination of the **formula mass**, which is the weighted average of the masses of the naturally occurring formula units of the substance. (It is analogous to the atomic mass for an element and the molecular mass for a molecular substance.)

OBJECTIVE 10

Formula mass = the sum of the atomic masses of each atom in a formula unit

The formula mass of sodium chloride is equal to the sum of the atomic masses of sodium and chlorine, which can be found on the periodic table.

$$\text{Formula mass NaCl} = 22.9898 + 35.4527 = 58.4425$$

Formula mass, like *formula unit*, is a general term. The atomic mass of carbon, C, could also be called its formula mass. The molecular mass of water, H_2O, could also be called water's formula mass.

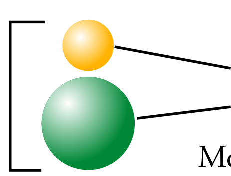
The number of grams in the molar mass (grams per mole) of any ionic compound is the same as its formula mass.

OBJECTIVE 11

$$\text{Molar mass of an ionic compound} = \left(\frac{\text{(formula mass) g compound}}{1 \text{ mol compound}}\right)$$

$$\text{or, for sodium chloride, } \left(\frac{58.4425 \text{ g NaCl}}{1 \text{ mol NaCl}}\right)$$

Formula unit NaCl



Molar mass Na: 22.9898 g/mol
Molar mass Cl: 35.4527 g/mol
Molar mass NaCl: 58.4425 g/mol

Example 9.4 - Formula Mass Calculations

Objective 10
Objective 11
Objective 12

Water from coal mines is contaminated with sulfuric acid that forms when water reacts with iron(III) sulfate, $Fe_2(SO_4)_3$.

a. Write a conversion factor to convert between grams of the ionic compound iron(III) sulfate and moles of $Fe_2(SO_4)_3$ formula units.

b. Calculate how many moles of $Fe_2(SO_4)_3$ are in 2.672 lb of $Fe_2(SO_4)_3$.

Solution

a. The molar mass, which provides a conversion factor for converting back and forth between grams and moles of an ionic compound, comes from the compound's formula mass.

Formula mass of $Fe_2(SO_4)_3$

$$= 2(\text{atomic mass Fe}) + 3(\text{atomic mass S}) + 12(\text{atomic mass O})$$

$$= 2(55.845) + 3(32.066) + 12(15.9994)$$

$$= 111.69 + 96.198 + 191.993 = 399.88 \text{ or } 399.88 \text{ u}$$

Thus the conversion factor that converts back and forth between grams moles of iron(III) sulfate is

$$\left(\frac{\mathbf{399.88\ g\ Fe_2(SO_4)_3}}{\mathbf{1\ mol\ Fe_2(SO_4)_3}}\right)$$

b. $$?\ \text{mol } Fe_2(SO_4)_3 = 2.672\ \cancel{\text{lb } Fe_2(SO_4)_3} \left(\frac{453.6\ \cancel{g}}{1\ \cancel{\text{lb}}}\right)\left(\frac{1\ \text{mol } Fe_2(SO_4)_3}{399.88\ \cancel{g\ Fe_2(SO_4)_3}}\right)$$

$$= \mathbf{3.031\ mol\ Fe_2(SO_4)_3}$$

Exercise 9.3 - Formula Mass Calculations

Objective 10
Objective 11
Objective 12

A quarter teaspoon of a typical baking powder contains about 0.4 g of sodium hydrogen carbonate, $NaHCO_3$.

a. Calculate the formula mass of sodium hydrogen carbonate.

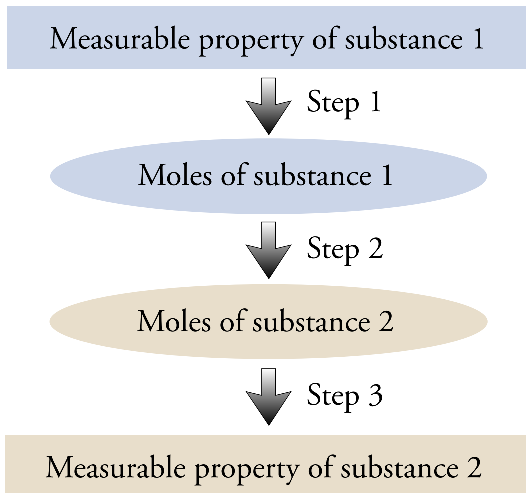
b. What is the mass in grams of one mole of $NaHCO_3$?

c. Write a conversion factor to convert between mass and moles of $NaHCO_3$.

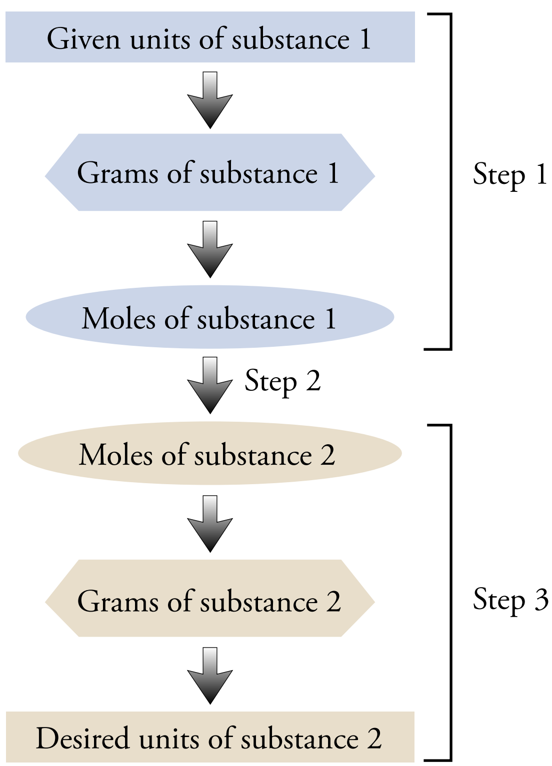
d. How many moles of $NaHCO_3$ are in 0.4 g of $NaHCO_3$?

9.4 Relationships Between Masses of Elements and Compounds

In general, the conversions you will be doing in chemistry require you to convert amount of one substance (substance 1) to amount of another substance (substance 2). Such conversions can be done in three basic steps:



Mass is often the most easily measured property, and we now know how to convert between mass in grams and moles of a substance using the substance's molar mass. The solutions to many problems will therefore follow these steps:



To complete these steps, we need one additional kind of conversion factor that converts between moles of an element and moles of a compound containing that element. We obtain this conversion factor from the compound's chemical formula. For example, the formula for hexane, C_6H_{14}, tells us that each hexane molecule contains six

Objective 13

atoms of carbon and fourteen atoms of hydrogen. A dozen C_6H_{14} molecules contain six dozen atoms of carbon and fourteen dozen atoms of hydrogen, and one mole of C_6H_{14} contains six moles of carbon atoms and fourteen moles of hydrogen atoms. These relationships lead to the following conversion factors:

$$\left(\frac{6 \text{ mol C}}{1 \text{ mol } C_6H_{14}}\right) \quad \text{and} \quad \left(\frac{14 \text{ mol H}}{1 \text{ mol } C_6H_{14}}\right)$$

One mole of the oxygen found in air, O_2, contains 6.022×10^{23} molecules and 1.204×10^{24} (two times 6.022×10^{23}) oxygen atoms. There are two moles of oxygen atoms in one mole of oxygen molecules.

$$\left(\frac{2 \text{ mol O}}{1 \text{ mol } O_2}\right)$$

Similarly, we can use ionic formulas to generate conversion factors that convert between moles of atoms of each element in an ionic compound and moles of compound. For example, the formula for calcium nitrate, $Ca(NO_3)_2$, yields the following conversion factors.

$$\left(\frac{1 \text{ mol Ca}}{1 \text{ mol } Ca(NO_3)_2}\right) \quad \text{and} \quad \left(\frac{2 \text{ mol N}}{1 \text{ mol } Ca(NO_3)_2}\right) \quad \text{and} \quad \left(\frac{6 \text{ mol O}}{1 \text{ mol } Ca(NO_3)_2}\right)$$

The collective unit of mole can also be used to describe ions. Thus the following conversion factors also come from the formula of calcium nitrate, $Ca(NO_3)_2$.

$$\left(\frac{1 \text{ mol } Ca^{2+}}{1 \text{ mol } Ca(NO_3)_2}\right) \quad \text{and} \quad \left(\frac{2 \text{ mol } NO_3^{-}}{1 \text{ mol } Ca(NO_3)_2}\right)$$

Example 9.5 - Molar Ratios of Element to Compound

Objective 13
Objective 14

Consider the molecular compound tetraphosphorus decoxide, P_4O_{10}, and the ionic compound iron(III) sulfate, $Fe_2(SO_4)_3$.

a. Write a conversion factor that converts between moles of phosphorus and moles of P_4O_{10}.
b. Write a conversion factor that converts between moles of iron and moles of iron(III) sulfate.
c. How many moles of sulfate are in one mole of iron(III) sulfate?

Solution

a. The formula P_4O_{10} shows that there are 4 atoms of phosphorus per molecule of P_4O_{10} or 4 moles of phosphorus per mole of P_4O_{10}.

$$\left(\frac{\mathbf{4 \text{ mol P}}}{\mathbf{1 \text{ mol } P_4O_{10}}}\right)$$

b. $$\left(\frac{\mathbf{2 \text{ mol Fe}}}{\mathbf{1 \text{ mol } Fe_2(SO_4)_3}}\right)$$

c. There are **three moles of SO_4^{2-}** in one mole of $Fe_2(SO_4)_3$.

Exercise 9.4 - Molar Ratios of Element to Compound

Objective 13
Objective 14

Find the requested conversion factors.

a. Write a conversion factor that converts between moles of hydrogen and moles of C_2H_5OH.
b. Write a conversion factor that converts between moles of oxygen and moles of $NaHCO_3$.
c. How many moles of hydrogen carbonate ions, HCO_3^-, are in one mole of $NaHCO_3$?

We are now ready to work the "typical problem" presented in Section 9.1.

Example 9.6 - Molecular Mass Calculations

Objective 15

What is the maximum mass of tetraphosphorus decoxide, P_4O_{10}, that could be produced from 1.09×10^4 kg of phosphorus, P?

Solution

The steps for this conversion are

$$1.09 \times 10^4 \text{ kg P} \rightarrow \text{g P} \rightarrow \text{mol P} \rightarrow \text{mol } P_4O_{10} \rightarrow \text{g } P_4O_{10} \rightarrow \text{kg } P_4O_{10}$$

Note that these follow the general steps we have been discussing.

To find a conversion factor that converts from moles of phosphorus to moles of P_4O_{10}, we look at the formula for tetraphosphorus decoxide, P_4O_{10}. It shows that each molecule of tetraphosphorus decoxide contains four atoms of phosphorus. By extension, one dozen P_4O_{10} molecules contains four dozen P atoms, and one mole of P_4O_{10} (6.022×10^{23} P_4O_{10} molecules) contains four moles of phosphorus (4 times 6.022×10^{23} P atoms). Thus the formula P_4O_{10} provides us with the following conversion factor:

$$\left(\frac{1 \text{ mol } P_4O_{10}}{4 \text{ mol P}}\right)$$

The molar mass of phosphorus can be used to convert grams of phosphorus to moles of phosphorus, and the molar mass of P_4O_{10} can be used to convert moles of P_4O_{10} to grams. Conversions of kilograms to grams and of grams to kilograms complete our setup.

Converts given mass unit into grams.
Converts moles of element into moles of compound.
Converts grams into desired mass unit.

$$? \text{ kg } P_4O_{10} = 1.09 \times 10^4 \cancel{\text{kg P}} \left(\frac{10^3 \cancel{\text{g}}}{1 \cancel{\text{kg}}}\right)\left(\frac{1 \cancel{\text{mol P}}}{30.9738 \cancel{\text{g P}}}\right)\left(\frac{1 \cancel{\text{mol } P_4O_{10}}}{4 \cancel{\text{mol P}}}\right)\left(\frac{283.889 \cancel{\text{g}}\ P_4O_{10}}{1 \cancel{\text{mol } P_4O_{10}}}\right)\left(\frac{1 \text{ kg}}{10^3 \cancel{\text{g}}}\right)$$

$$= \mathbf{2.50 \times 10^4 \text{ kg } P_4O_{10}}$$

Converts grams of element into moles.
Converts moles of compound into grams.

Sample Study Sheet 9.1 Converting Between Mass of Element and Mass of Compound Containing the Element

OBJECTIVE 15

TIP-OFF When you analyze the type of unit you have and the type of unit you want, you recognize that you are converting between a unit associated with an element and a unit associated with a compound containing that element.

GENERAL STEPS The following general procedure is summarized in Figure 9.3.

- **Convert the given unit to moles of the first substance.**

  This step often requires converting the given unit into grams, after which the grams can be converted into moles using the molar mass of the substance.

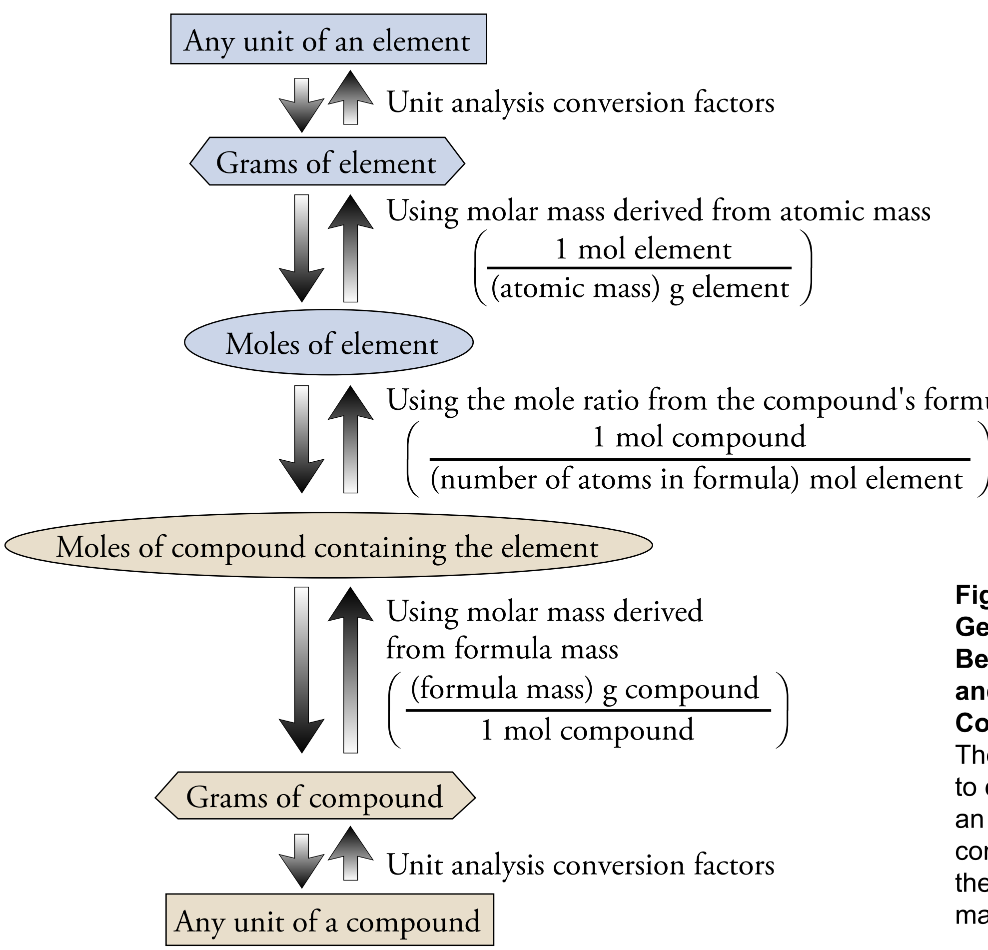
- **Convert moles of the first substance into moles of the second substance using the molar ratio derived from the formula for the compound.**

  You either convert from moles of element to moles of compound or moles of compound to moles of element.

- **Convert moles of the second substance to the desired units of the second substance.**

  This step requires converting moles of the second substance into grams of the second substance using the molar mass of the second substance, after which the grams can be converted to the specific units that you want.

EXAMPLE See Example 9.6.



OBJECTIVE 15

Figure 9.3
General Steps for Converting Between the Mass of an Element and the Mass of a Compound Containing the Element
The calculation can be set up to convert from the mass of an element to the mass of a compound (top to bottom) or from the mass of a compound to the mass of an element (bottom to top).

Exercise 9.5 - Molar Mass Calculations

Objective 15

Disulfur dichloride, S_2Cl_2, is used in vulcanizing rubber and hardening soft woods. It can be made from the reaction of pure sulfur with chlorine gas. What is the mass of S_2Cl_2 that contains 123.8 g S?

Exercise 9.6 - Molar Mass Calculations

Objective 15

Vanadium metal, used as a component of steel and to catalyze various industrial reactions, is produced from the reaction of vanadium(V) oxide, V_2O_5, and calcium metal. What is the mass in kilograms of vanadium in 2.3 kilograms of V_2O_5?

The skills developed in this chapter can be used to calculate the percentage of each element in a compound. See how this is done at the textbook's Web site.

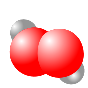
9.5 Determination of Empirical and Molecular Formulas

By this point in your study of chemistry, you have seen hundreds of chemical formulas. Have you wondered where they come from or how we know the relative numbers of atoms of each element in a compound? This section describes some of the ways chemists determine chemical formulas from experimental data.

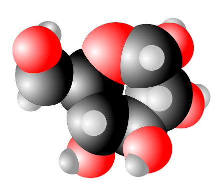
Before beginning, we need to understand the distinction between two types of chemical formulas, empirical formulas and molecular formulas. When the subscripts in a chemical formula represent the simplest ratio of the kinds of atoms in the compound, the formula is called an **empirical formula**. Most ionic compounds are described with empirical formulas. For example, chromium(III) oxide's formula, Cr_2O_3, is an empirical formula. The compound contains two chromium atoms for every three oxide atoms, and there is no lower ratio representing these relative amounts.

Molecular compounds are described with molecular formulas. A **molecular formula** describes the actual numbers of atoms of each element in a molecule. Some molecular formulas are also empirical formulas. For example, water molecules are composed of two hydrogen atoms and one oxygen atom, so water's molecular formula is H_2O. Because this formula represents the simplest ratio of hydrogen atoms to oxygen atoms in water, it is also an empirical formula.

Many molecular formulas are not empirical formulas. Hydrogen peroxide molecules contain two hydrogen atoms and two oxygen atoms, so hydrogen peroxide's molecular formula is H_2O_2. The empirical formula of hydrogen peroxide is HO. The molecular formula for glucose is $C_6H_{12}O_6$, and its empirical formula is CH_2O.



hydrogen peroxide
molecular formula, H_2O_2,
empirical formula, HO



glucose
molecular formula, $C_6H_{12}O_6$,
empirical formula, CH_2O

Determination of Empirical Formulas

If we know the relative mass or the mass percentage of each element in a compound, we can determine the compound's empirical formula. The general procedure is summarized in Study Sheet 9.2, but before we look at it, let's reason it out using a substance that is sometimes called photophor, an ingredient in signal fires, torpedoes, fireworks, and rodent poison.

This "signal pistol" or flare gun is one of 20,000 made by Remington's Bridgeport Connecticut plant, along with the green, white and red flares for it.
Photo Courtesy of John Spangler armscollectors.com

The subscripts in an empirical formula are positive integers representing the simplest ratio of the atoms of each element in the formula. For now, we can describe the empirical formula for photophor in the following way:

$$Ca_aP_b \qquad \text{a and b} = \text{positive integers}$$

(Remember that in binary compounds containing a metal and a nonmetal, the symbol for the metal is listed first.) The *a* and *b* in the formula above represent the subscripts in the empirical formula. For every *a* atoms of calcium, there are *b* atoms of phosphorus, or for every *a* moles of Ca, there are *b* moles of P. Thus the ratio of *a* to *b* describes the molar ratio of these elements in the compound.

If we can determine the number of moles of each element in any amount of a substance, we can calculate the molar ratio of these elements in the compound, which can then be simplified to the positive integers that represent the simplest molar ratio. We found in Sections 9.2 and 9.3 that we can calculate moles of a substance from grams. Thus one path to determining the empirical formula for a compound is:

Objective 16

Grams of each element in a specific amount of compound

↓

Moles of each element in that amount of compound

↓

Molar ratio of the elements

↓

Simplest molar ratio of the elements

Imagine that a sample of photophor has been analyzed and found to contain 12.368 g calcium and 6.358 g phosphorus. These masses can be converted to moles using the molar masses of the elements.

$$? \text{ mol Ca} = 12.368 \cancel{\text{g Ca}} \left(\frac{1 \text{ mol Ca}}{40.078 \cancel{\text{g Ca}}}\right) = 0.30860 \text{ mol Ca}$$

$$? \text{ mol P} = 6.358 \cancel{\text{g P}} \left(\frac{1 \text{ mol P}}{30.9738 \cancel{\text{g P}}}\right) = 0.2053 \text{ mol P}$$

The molar ratio of Ca:P is therefore 0.30860 moles of calcium to 0.2053 moles of phosphorus. We simplify it to the integers that represent the simplest molar ratio using the following steps.

- Divide each mole value by the smallest mole value, and round the answer to the nearest positive integer or common mixed fraction. (A mixed fraction contains an integer and a fraction. For example, 2½ is a mixed fraction.)

- If one of the values is a fraction, multiply all the values by the denominator of the fraction.

In this case, we divide the mole values for Ca and P by 0.2053. This step always leads to 1 mole for the smallest value. We then round the other mole values to positive integers or common mixed fractions.

$$\frac{0.30860 \text{ mol Ca}}{0.2053} = 1.503 \text{ mol Ca} \approx 1\tfrac{1}{2} \text{ mol Ca}$$

$$\frac{0.2053 \text{ mol P}}{0.2053} = 1 \text{ mol P}$$

(The mole value for calcium has been rounded off from 1.503 to 1½.)

We will restrict the examples in this text to compounds with relatively simple formulas. Thus, when you work the end-of-chapter problems, the values you will obtain at this stage will always be within 0.02 of a positive integer or a common mixed fraction, and the denominators for your fractions will be 4 or smaller.

To complete the determination of the empirical formula of photophor, we multiply each mole value by 2 to get rid of the fraction:

$$1\tfrac{1}{2} \text{ mol Ca} \times 2 = 3 \text{ mol Ca}$$
$$1 \text{ mol P} \times 2 = 2 \text{ mol P}$$

Our empirical formula is Ca_3P_2, which represents calcium phosphide. Sample Study Sheet 9.2 summarizes these steps.

Sample Study Sheet 9.2 Calculating Empirical Formulas

Objective 16

Objective 17

Tip-off You wish to calculate an empirical formula.

General Steps The following procedure is summarized in Figure 9.4.

1. If you are not given the mass of each element in grams, convert the data you *are* given to mass of each element in grams.
   - In some cases, this can be done with simple unit conversions. For example, you may be given pounds or milligrams, which can be converted to grams using unit analysis.
   - Sometimes you are given the percentage of each element in the compound. If so, assume that you have 100 g of compound, and change the percentages to grams (see Example 9.7).
2. Convert grams of each element to moles by dividing by the atomic mass of the element.
3. Divide each mole value by the smallest mole value, and round your answers to positive integers or common mixed fractions.
4. If you have a fraction after the last step, multiply all the mole values by the denominator of the fraction.
5. The resulting mole values represent the subscripts in the empirical formula.

Example See Example 9.7.

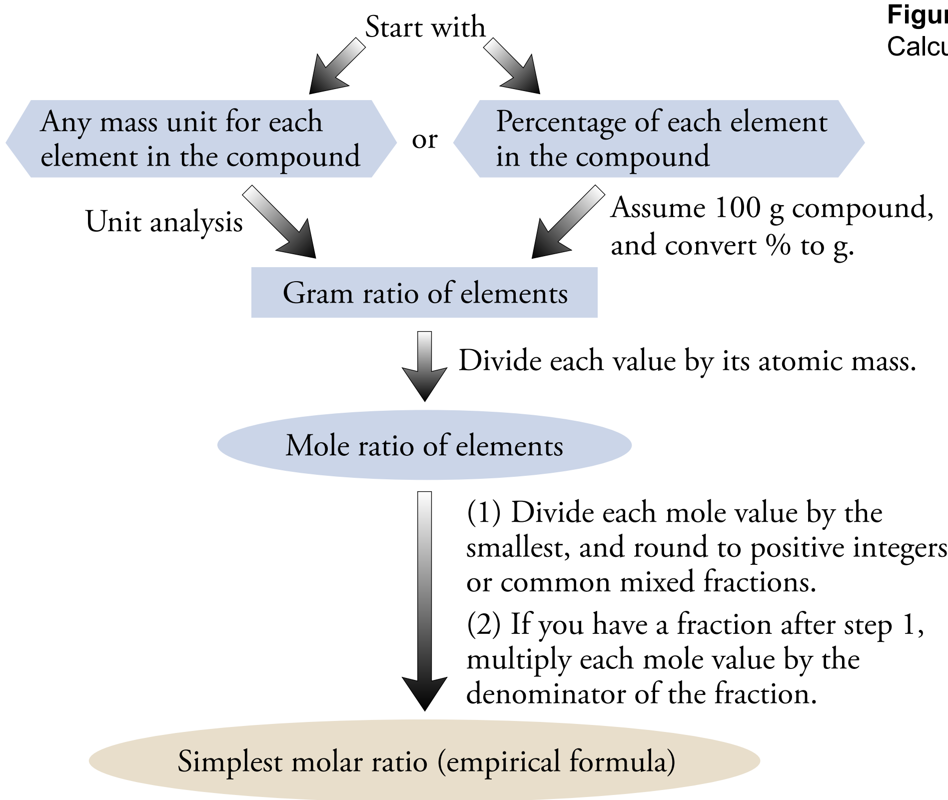


Figure 9.4
Calculating Empirical Formulas

Objective 16
Objective 17

Example 9.7 - Calculating an Empirical Formula from the Percentage of Each Element

Objective 17

An ionic compound sometimes called pearl ash is used to make special glass for color TV tubes. A sample of this compound is analyzed and found to contain 56.50% potassium, 8.75% carbon, and 34.75% oxygen. What is the empirical formula for this compound? What is its chemical name?

Solution

If we assume that we have a 100-g sample, the conversion from percentages to a gram ratio becomes very simple. We just change each "%" to "g". Thus 100 g of pearl ash would contain 56.50 g K, 8.75 g C, and 34.75 g O.

Now we can proceed with the steps described in Sample Study Sheet 9.3 to convert from grams of each element to the empirical formula.

$$?\ \text{mol K} = 56.50\ \cancel{\text{g K}} \left(\frac{1\ \text{mol K}}{39.0983\ \cancel{\text{g K}}}\right) = 1.445\ \text{mol K} \div 0.728 = 1.985\ \text{mol K} \approx 2\ \text{mol K}$$

$$?\ \text{mol C} = 8.75\ \cancel{\text{g C}} \left(\frac{1\ \text{mol C}}{12.011\ \cancel{\text{g C}}}\right) = 0.728\ \text{mol C} \div 0.728 = 1\text{mol C}$$

$$?\ \text{mol O} = 34.75\ \cancel{\text{g O}} \left(\frac{1\ \text{mol O}}{15.9994\ \cancel{\text{g O}}}\right) = 2.172\ \text{mol O} \div 0.728 = 2.984\ \text{mol O} \approx 3\ \text{mol O}$$

The empirical formula is $\mathbf{K_2CO_3}$, which is potassium carbonate.

Exercise 9.7 - Calculating an Empirical Formula

Objective 16

Bismuth ore, often called bismuth glance, contains an ionic compound consisting of the elements bismuth and sulfur. A sample of the pure compound is found to contain 32.516 g Bi and 7.484 g S. What is the empirical formula for this compound? What is its name?

Exercise 9.8 - Calculating an Empirical Formula

Objective 17

An ionic compound used in the brewing industry to clean casks and vats and in the wine industry to kill undesirable yeasts and bacteria is composed of 35.172% potassium, 28.846% sulfur, and 35.982% oxygen. What is the empirical formula for this compound?

Converting Empirical Formulas to Molecular Formulas

To see how empirical formulas can be converted to molecular formulas, let's consider an analysis of adipic acid, which is one of the substances used to make Nylon-66. (Nylon-66, originally used as a replacement for silk in women's stockings, was invented in 1935 and was produced commercially starting in 1940.) Experiment shows that adipic acid is 49.31% carbon, 6.90% hydrogen, and 43.79% oxygen, which converts to an empirical formula of $C_3H_5O_2$. From a separate experiment, the molecular mass of adipic acid is found to be 146.144. We can determine the molecular formula for adipic acid from this data.

The subscripts in a molecular formula are always whole-number multiples of the subscripts in the empirical formula. The molecular formula for adipic acid can therefore be described in the following way.

$$C_{3n}H_{5n}O_{2n} \qquad n = \text{some positive integer such as } 1, 2, 3,\ldots$$

The n can be calculated from adipic acid's molecular mass and its empirical formula mass. A substance's empirical formula mass can be calculated from the subscripts in its empirical formula and the atomic masses of the elements. The empirical formula mass of adipic acid is

$$\text{Empirical formula mass} = 3(12.011) + 5(1.00794) + 2(15.9994) = 73.072$$

Because the subscripts in the molecular formula are always a positive integer multiple of the subscripts in the empirical formula, the molecular mass is always equal to a positive integer multiple of the empirical formula mass.

$$\text{Molecular mass} = n\ (\text{empirical formula mass})$$

$$n = \text{some positive integer, such as } 1, 2, 3,\ldots$$

Therefore, we can calculate n by dividing the molecular mass by the empirical formula mass.

$$n = \frac{\text{molecular mass}}{\text{empirical formula mass}}$$

For our example, the given molecular mass of $C_3H_5O_2$ divided by the empirical formula mass is:

$$n = \frac{\text{molecular mass}}{\text{empirical formula mass}} = \frac{146.144}{73.072} = 2$$

Once you have calculated n, you can determine the molecular formula by multiplying each of the subscripts in the empirical formula by n. This gives a molecular formula for adipic acid of $C_{3(2)}H_{5(2)}O_{2(2)}$ or $C_6H_{10}O_4$ (see Special Topic 9.1 *Green Chemistry–Making Chemicals from Safer Reactants*).

Special Topic 9.1 Green Chemistry—Making Chemicals from Safer Reactants

Because benzene, C_6H_6, is a readily available and inexpensive substance that can be easily converted into many different substances, it is a common industrial starting material for making a wide range of important chemicals. A problem with using benzene, however, is that it is known to cause cancer and to be toxic in other ways as well. Thus there is strong incentive to find alternative methods for making chemicals that in the past have been made from benzene.

You discovered in this section that adipic acid is used to make nylon, but it is also used to make paint, synthetic lubricants, and plasticizers (substances that make plastics more flexible). Adipic acid is one of the important industrial chemicals conventionally produced from benzene, but recently a new process has been developed that forms adipic acid from the sugar glucose, which is much safer, instead of benzene. If you were a worker in a chemical plant making adipic acid, you would certainly prefer working with sugar instead of benzene.

Can nylon be made from sugar?

Sample Study Sheet 9.3 Calculating Molecular Formulas

OBJECTIVE 18

TIP-OFF You want to calculate a molecular formula and have been given the molecular mass of the substance and either the empirical formula or enough data to calculate the empirical formula.

GENERAL STEPS The following procedure is summarized in Figure 9.5.

- If necessary, calculate the empirical formula of the compound from the data given. (See Sample Study Sheet 9.2: Calculating Empirical Formulas.)
- Divide the molecular mass by the empirical formula mass.

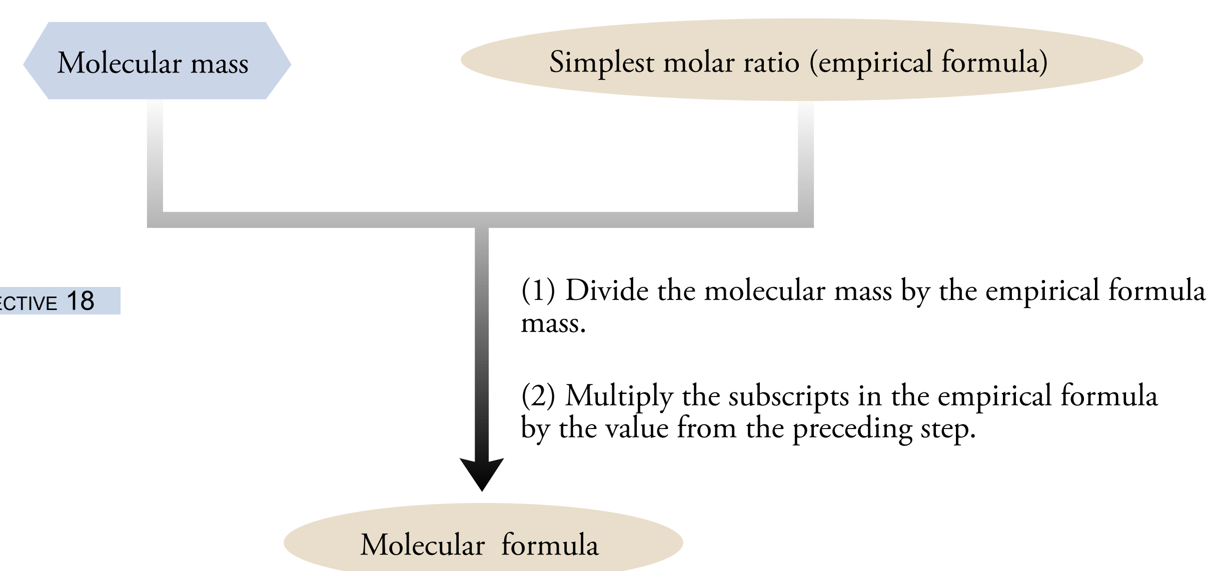
$$n = \frac{\text{molecular mass}}{\text{empirical formula mass}}$$

- Multiply each of the subscripts in the empirical formula by n to get the molecular formula.

EXAMPLE See Example 9.8.

Figure 9.5
Calculating Molecular Formulas

OBJECTIVE 18



Example 9.8 - Calculating a Molecular Formula Using the Percentage of Each Element in a Compound

Objective 17
Objective 18

A chemical called BD (or sometimes BDO), which is used in the synthesis of Spandex, has controversial uses as well. In 1999, it was added to products that claimed to stimulate the body's immune system, reduce tension, heighten sexual experience, repair muscle tissue, and cure insomnia (see Special Topic 9.2: *Safe and Effective?*). The FDA seized these products because of suspicions that BD caused at least three deaths and many severe adverse reactions. BD is composed of 53.31% carbon, 11.18% hydrogen, and 35.51% oxygen and has a molecular mass of 90.122. What is its molecular formula?

Solution

$$? \text{ mol C} = 53.31 \cancel{\text{g C}}\left(\frac{1 \text{ mol C}}{12.011 \cancel{\text{g C}}}\right) = 4.438 \text{ mol C} \div 2.219 = 2 \text{ mol C}$$

$$? \text{ mol H} = 11.18 \cancel{\text{g H}}\left(\frac{1 \text{ mol H}}{1.00794 \cancel{\text{g H}}}\right) = 11.09 \text{ mol H} \div 2.219 \approx 5 \text{ mol H}$$

$$? \text{ mol O} = 35.51 \cancel{\text{g O}}\left(\frac{1 \text{ mol O}}{15.9994 \cancel{\text{g O}}}\right) = 2.219 \text{ mol O} \div 2.219 = 1 \text{ mol O}$$

Empirical formula: C_2H_5O $\quad n = \dfrac{\text{molecular mass}}{\text{empirical formula mass}} = \dfrac{90.122}{45.061} = 2$

Molecular formula: $\mathbf{C_4H_{10}O_2}$

Exercise 9.9 - Calculating a Molecular Formula Using the Percentage of Each Element in a Compound

Objective 17
Objective 18

Compounds called polychlorinated biphenyls (PCBs) have structures similar to chlorinated insecticides such as DDT. They have been used in the past for a variety of purposes, but because they have been identified as serious pollutants, their only legal use today is as insulating fluids in electrical transformers. One PCB is composed of 39.94% carbon, 1.12% hydrogen, and 58.94% chlorine and has a molecular mass of 360.88. What is its molecular formula?

Empirical and molecular formulas can be derived from a process called combustion analysis. You can learn about this process at the textbook's Web site.

Special Topic 9.2 Safe and Effective?

The public is continually bombarded with claims for new products containing "miracle" ingredients "guaranteed" to improve our health, strength, happiness, and sex life and to give us a good night's sleep. In 1999, such claims were made on behalf of products with names such as Revitalize Plus, Serenity, Enliven, and Thunder Nectar, all of which contained a substance called 1,4-butanediol (BD). They were marketed on the Internet, sold in health food stores, and advertised in muscle-building magazines.

The claims were enticing, but according to the U.S. Food and Drug Administration (FDA), they were also unfounded. Perhaps more importantly, the FDA decided that this unapproved new drug could cause dangerously low respiratory rates, unconsciousness, seizures, and even death. Because the FDA connected BD with at least three deaths and many severe adverse reactions, they designated it a Class I health hazard, which means "its use could pose potentially life-threatening risks." The FDA seized products containing BD to prevent their causing further illness and death. Advocates of the drug submitted anecdotal evidence to show that many people have taken the substance without ill effects. As of this writing, no one knows with certainty which side is correct.

Because humans are very complex chemical factories, the positive and negative effects of a drug can be difficult to determine. To some extent, they depend on each person's unique biochemistry, as well as on interactions with other chemicals that may be present in the body. For example, the effects of BD are thought to be enhanced by alcohol and other depressants, so if an individual takes one of the "party drugs" containing BD (drugs with names like Cherry fX Bombs) and drinks too much beer, the combined depressant effect can lead to loss of consciousness, coma, and perhaps death. To minimize the uncertainties associated with individual reactions to a drug, scientists run carefully controlled tests, first on animals and only much later on humans. Until these tests are done for BD, its true effects (both positive and negative) cannot be known with confidence.

So, how do you decide whether to consume a product containing a chemical like BD? Let's consider a student named Fred who is surfing the Internet for ideas about how to relax before his final exams. One site he finds describes a product guaranteed to calm his nerves and recharge his immune system. The product seems to contain a lot of good ingredients, such as vitamins and minerals, but the most important component is tetramethylene glycol. The description of how the substance works is written in unfamiliar terminology, but Fred thinks he gets the gist, and it seems to make sense. Fortunately, he passes up the opportunity to buy this product and takes a walk in the woods to calm his nerves instead. He's never heard of the product's distributor, he knows that false and unproven advertising claims are often cloaked in pseudoscientific explanations, and he remembers reading in his chemistry book that tetramethylene glycol is another name for 1,4-butanediol (BD), a potentially dangerous substance.

The important points in this story are that it is best to stick to products from known manufacturers who have a reputation for carefully screening their ingredients; to be skeptical of claims made in advertising and on the Internet; and to keep yourself educated about substances that are suspected of being harmful. When in doubt, ask your doctor. It's part of his or her job to know about the safety and effectiveness of health products.

Chapter Glossary

Weighted average mass A mass calculated by multiplying the decimal fraction of each component in a sample by its mass and adding the results of each multiplication together.

Atomic mass unit One-twelfth the mass of a carbon-12 atom. It is sometimes called a *unified mass unit*. Its accepted abbreviation is *u*, but *amu* is sometimes used.

Atomic mass The weighted average of the masses of the naturally occurring isotopes of an element.

Mole The amount of substance that contains the same number of particles as there are atoms in 12 g of carbon-12.

Avogadro's number The number of atoms in 12 g of carbon-12. To four significant figures, it is $\mathbf{6.022 \times 10^{23}}$.

Molar mass The mass in grams of one mole of substance. (The number of grams in the molar mass of an element is the same as its atomic mass. The number of grams in the molar mass of a molecular compound is the same as its molecular mass. The number of grams in the molar mass of an ionic compound is the same as its formula mass.)

Molecular mass The weighted average of the masses of the naturally occurring molecules of a molecular substance. It is the sum of the atomic masses of the atoms in a molecule.

Formula unit A group represented by a substance's chemical formula—that is, a group containing the kinds and numbers of atoms or ions listed in the chemical formula. It is a general term that can be used in reference to elements, molecular compounds, or ionic compounds.

Formula mass The weighted average of the masses of the naturally occurring formula units of the substance. It is the sum of the atomic masses of the atoms in a formula unit.

Empirical formula A chemical formula that includes positive integers that describe the simplest ratio of the atoms of each element in a compound.

Molecular formula The chemical formula that describes the actual numbers of atoms of each element in a molecule of a compound.

You can test yourself on the glossary terms at the textbook's Web site.

Chapter Objectives

The goal of this chapter is to teach you to do the following.

1. Define all of the terms in the Chapter Glossary.

Section 9.2 Relating Mass to Number of Particles

2. Describe how a mole is similar to a dozen.
3. Given an atomic mass for an element, write a conversion factor that converts between mass and moles of that element.
4. Given a periodic table that shows atomic masses of the elements, convert between mass of an element and moles of that element.

Section 9.3 Molar Mass and Chemical Compounds

5. Given a formula for a molecular substance and a periodic table that includes atomic masses for the elements, calculate the substance's molecular mass.

6. Given enough information to calculate a molecular substance's molecular mass, write a conversion factor that converts between mass and moles of the substance.
7. Given enough information to calculate a molecular substance's molecular mass, convert between mass and moles of the substance.
8. Explain why the term molecule is better applied to molecular substances, such as water, than to ionic compounds, such as NaCl.
9. Given the name or chemical formula of a compound, identify whether it would be better to use the term *molecule* or the term *formula unit* to describe the group that contains the number of atoms or ions of each element equal to the subscript for the element in its chemical formula.
10. Given a formula for an ionic compound and a periodic table that includes atomic masses for the elements, calculate the compound's formula mass.
11. Given enough information to calculate an ionic compound's formula mass, write a conversion factor that converts between mass and moles of the compound.
12. Given enough information to calculate an ionic compound's formula mass, convert between mass and moles of the compound.

Section 9.4 Relationships Between Masses of Elements and Compounds

13. Given a formula for a compound, write conversion factors that convert between moles of atoms of each element in the compound and moles of compound.
14. Given the name or formula for an ionic compound, determine the number of moles of each ion in one mole of the compound.
15. Make conversions between the mass of a compound and the mass of an element in the compound.

Section 9.5 Determination of Empirical and Molecular Formulas

16. Given the masses of each element in a sample of a compound, calculate the compound's empirical formula.
17. Given the mass percentage of each element in a compound, calculate the compound's empirical formula.
18. Given an empirical formula for a molecular compound (or enough information to calculate the empirical formula) and given the molecular mass of the compound, determine its molecular formula.

Review Questions

1. Complete each of the following conversion factors by filling in the blank on the top of the ratio.

a. $\left(\dfrac{\_\_\_\_\ \text{g}}{1\ \text{kg}}\right)$ b. $\left(\dfrac{\_\_\_\_\ \text{mg}}{1\ \text{g}}\right)$ c. $\left(\dfrac{\_\_\_\_\ \text{kg}}{1\ \text{metric ton}}\right)$ d. $\left(\dfrac{\_\_\_\_\ \text{mg}}{1\ \text{g}}\right)$

2. Convert 3.45×10^4 kg into grams.
3. Convert 184.570 g into kilograms.
4. Convert 4.5000×10^6 g into megagrams.
5. Convert 871 Mg into grams.
6. Surinam bauxite is an ore that is 54-57% aluminum oxide, Al_2O_3. What is the mass (in kilograms) of Al_2O_3 in 1256 kg of Surinam bauxite that is 55.3% Al_2O_3?

Key Ideas

Complete the following statements by writing one of these words or phrases in each blank.

atomic mass	molecules
atoms	mole
atoms of each element	one-twelfth
formula	naturally
formula unit	numbers
formula units	simplest
impossible	weighted
kinds	whole-number

7. Because of the size and number of carbon atoms in any normal sample of carbon, it is ______________ to count the atoms directly.
8. The unit most often used to describe the mass of atoms is the atomic mass unit, whose symbol is u or amu. An atomic mass unit is defined as exactly ______________ the mass of an atom of carbon-12.
9. The atomic mass of any element is the ______________ average of the masses of the ______________ occurring isotopes of the element.
10. A mole (which is abbreviated mol) is an amount of substance that contains the same number of particles as there are ______________ in 12 g of carbon-12.
11. The number of grams in the molar mass of an element is the same as the element's ______________.
12. The molecular mass of the compound is the weighted average mass of the compound's naturally occurring ______________.
13. In this text, the term ______________ is used to describe ionic compounds in situations where molecule is used to describe molecular substances. It is the group represented by the substance's chemical formula, that is, a group containing the ______________ and ______________ of atoms or ions listed in the chemical formula.
14. One ______________ of carbon contains 6.022×10^{23} carbon atoms.
15. Formula mass is the weighted average of the masses of the naturally occurring ______________ of the substance.
16. We can convert between moles of element and moles of a compound containing that element by using the molar ratio derived from the ______________ for the compound.
17. When the subscripts in a chemical formula represent the ______________ ratio of the kinds of atoms in the compound, the formula is called an empirical formula.
18. A molecular formula describes the actual numbers of ______________ in a molecule.
19. The subscripts in a molecular formula are always ______________ multiples of the subscripts in the empirical formula.

Chapter Problems

Section 9.2 Relating Mass to Number of Particles

20. Describe how a mole is similar to a dozen. OBJECTIVE 2
21. What is the weighted average mass in atomic mass units (u) of each atom of the elements (a) sodium and (b) oxygen?
22. What is the weighted average mass in atomic mass units (u) of each atom of the elements (a) calcium and (b) neon?
23. What is the weighted average mass in grams of 6.022×10^{23} atoms of the elements (a) sulfur and (b) fluorine?
24. What is the weighted average mass in grams of 6.022×10^{23} atoms of the elements (a) bromine and (b) nickel?
25. What is the molar mass of the elements (a) zinc and (b) aluminum?
26. What is the molar mass of the elements (a) chlorine and (b) silver?

OBJECTIVE 3

27. For each of the elements (a) iron and (b) krypton, write a conversion factor that converts between mass in grams and moles of the substance.

OBJECTIVE 3

28. For each of the elements (a) manganese and (b) silicon, write a conversion factor that converts between mass in grams and moles of the substance.

OBJECTIVE 4

29. A vitamin supplement contains 50 micrograms of the element selenium in each tablet. How many moles of selenium does each tablet contain?

OBJECTIVE 4

30. A multivitamin tablet contains 40 milligrams of potassium. How many moles of potassium does each tablet contain?

OBJECTIVE 4

31. A multivitamin tablet contains 1.6×10^{-4} mole of iron per tablet. How many milligrams of iron does each tablet contain?

OBJECTIVE 4

32. A multivitamin tablet contains 1.93×10^{-6} mole of chromium. How many micrograms of chromium does each tablet contain?

Section 9.3 Molar Masses and Chemical Compounds

OBJECTIVE 5
OBJECTIVE 6

33. For each of the molecular substances (a) H_3PO_2 and (b) $C_6H_5NH_2$, calculate its molecular mass and write a conversion factor that converts between mass in grams and moles of the substance.

OBJECTIVE 5
OBJECTIVE 6

34. For each of the molecular substances (a) CF_3CHCl_2 and (b) SO_2Cl_2, calculate its molecular mass and write a conversion factor that converts between mass in grams and moles of the substance.

OBJECTIVE 7

35. Each dose of a nighttime cold medicine contains 1000 mg of the analgesic acetaminophen. Acetaminophen, or N-acetyl-p-aminophenol, has the general formula C_8H_9NO.
    a. How many moles of acetaminophen are in each dose?
    b. What is the mass in grams of 15.0 moles of acetaminophen?

OBJECTIVE 7

36. A throat lozenge contains 5.0 mg of menthol, which has the formula $C_{10}H_{20}O$.
    a. How many moles of menthol are in 5.0 mg of menthol?
    b. What is the mass in grams of 1.56 moles of menthol?

OBJECTIVE 8

37. A group of atoms that contains one atom of nitrogen and three atoms of hydrogen is called an ammonia *molecule*. A group of ions that contains one potassium ion and one fluoride ion is called a potassium fluoride *formula unit* instead of a *molecule*. Why?

38. For each of the following examples, decide whether it would be better to use the term *molecule* or *formula unit*. OBJECTIVE 9

a. Cl_2O b. Na_2O c. $(NH_4)_2SO_4$ d. $HC_2H_3O_2$

39. For each of the following examples, decide whether it would be better to use the term *molecule* or *formula unit*. OBJECTIVE 9

a. K_2SO_3 b. H_2SO_3 c. CCl_4 d. NH_4Cl

40. For each of the ionic substances (a) $BiBr_3$ and (b) $Al_2(SO_4)_3$, calculate its formula mass and write a conversion factor that converts between mass in grams and moles of the substance. OBJECTIVE 10 OBJECTIVE 11

41. For each of the following ionic substances (a) Co_2O_3 and (b) $Fe_2(C_2O_4)_3$, calculate its formula mass and write a conversion factor that converts between mass in grams and moles of the substance. OBJECTIVE 10 OBJECTIVE 11

42. A common antacid tablet contains 500 mg of calcium carbonate, $CaCO_3$. OBJECTIVE 12

a. How many moles of $CaCO_3$ does each tablet contain?

b. What is the mass in kilograms of 100.0 moles of calcium carbonate?

43. An antacid contains 200 mg of aluminum hydroxide and 200 mg of magnesium hydroxide per capsule. OBJECTIVE 12

a. How many moles of $Al(OH)_3$ does each capsule contain?

b. What is the mass in milligrams of 0.0457 mole of magnesium hydroxide?

44. Rubies and other minerals in the durable corundum family are primarily composed of aluminum oxide, Al_2O_3, with trace impurities that lead to their different colors. For example, the red color in rubies comes from a small amount of chromium replacing some of the aluminum. If a 0.78-carat ruby were pure aluminum oxide, how many moles of Al_2O_3 would be in the stone? (There are exactly 5 carats per gram.) OBJECTIVE 12

45. Many famous "rubies" are in fact spinels, which look like rubies but are far less valuable. Spinels consist primarily of $MgAl_2O_4$, whereas rubies are primarily Al_2O_3. If the Timur Ruby, a 361-carat spinel, were pure $MgAl_2O_4$, how many moles of $MgAl_2O_4$ would it contain? (There are exactly 5 carats per gram.) OBJECTIVE 12

Section 9.4 Relationship Between Masses of Elements and Compounds

46. Write a conversion factor that converts between moles of nitrogen in nitrogen pentoxide, N_2O_5, and moles of N_2O_5. OBJECTIVE 13

47. Write a conversion factor that converts between moles of oxygen in phosphoric acid, H_3PO_4, and moles of H_3PO_4. OBJECTIVE 13

48. The green granules on older asphalt roofing are chromium(III) oxide. Write a conversion factor that converts between moles of chromium ions in chromium(III) oxide, Cr_2O_3, and moles of Cr_2O_3. OBJECTIVE 13 OBJECTIVE 14

49. Calcium phosphide is used to make fireworks. Write a conversion factor that converts between moles of calcium ions in calcium phosphide, Ca_3P_2, and moles of Ca_3P_2. OBJECTIVE 13 OBJECTIVE 14

50. Ammonium oxalate is used for stain and rust removal. How many moles of ammonium ions are in 1 mole of ammonium oxalate, $(NH_4)_2C_2O_4$? OBJECTIVE 14

51. Magnesium phosphate is used as a dental polishing agent. How many moles of ions (cations and anions together) are in 1 mole of magnesium phosphate, $Mg_3(PO_4)_2$? OBJECTIVE 14

Objective 13
Objective 15

52. A nutritional supplement contains 0.405 g of $CaCO_3$. The recommended daily value of calcium is 1.000 g Ca.

a. Write a conversion factor that relates moles of calcium to moles of calcium carbonate.

b. Calculate the mass in grams of calcium in 0.405 g of $CaCO_3$.

c. What percentage of the daily value of calcium comes from this tablet?

Objective 13
Objective 15

53. A multivitamin tablet contains 0.479 g of $CaHPO_4$ as a source of phosphorus. The recommended daily value of phosphorus is 1.000 g of P.

a. Write a conversion factor that relates moles of phosphorus to moles of calcium hydrogen phosphate.

b. Calculate the mass in grams of phosphorus in 0.479 g of $CaHPO_4$.

c. What percentage of the daily value of phosphorus comes from this tablet?

Objective 15

54. A multivitamin tablet contains 10 μg of vanadium in the form of sodium metavanadate, $NaVO_3$. How many micrograms of $NaVO_3$ does each tablet contain?

Objective 15

55. A multivitamin tablet contains 5 μg of nickel in the form of nickel(II) sulfate. How many micrograms of $NiSO_4$ does each tablet contain?

Objective 15

56. There are several natural sources of the element titanium. One is the ore called rutile, which contains oxides of iron and titanium, FeO and TiO_2. Titanium metal can be made by first converting the TiO_2 in rutile to $TiCl_4$ by heating the ore to high temperature in the presence of carbon and chlorine. The titanium in $TiCl_4$ is then reduced from its +4 oxidation state to its zero oxidation state by reaction with a good reducing agent such as magnesium or sodium. What is the mass of titanium in kilograms in 0.401 Mg of $TiCl_4$?

Objective 15

57. Manganese metal is produced from the manganese(III) oxide, Mn_2O_3, which is found in manganite, a manganese ore. The manganese is reduced from its +3 oxidation state in Mn_2O_3 to the zero oxidation state of the uncharged metal by reacting the Mn_2O_3 with a reducing agent such as aluminum or carbon. How many pounds of manganese are in 1.261 tons of Mn_2O_3? (1 ton = 2000 pounds)

Section 9.5 Determination of Empirical and Molecular Formulas

58. Explain the difference between molecular formulas and empirical formulas. Give an example of a substance whose empirical formula is different from its molecular formula. Give an example of a substance whose empirical formula and molecular formula are the same.

Objective 16

59. An extremely explosive ionic compound is made from the reaction of silver compounds with ammonia. A sample of this compound is found to contain 17.261 g of silver and 0.743 g of nitrogen. What is the empirical formula for this compound? What is its chemical name?

60. A sample of an ionic compound that is often used as a dough conditioner is analyzed and found to contain 7.591 g of potassium, 15.514 g of bromine, and 9.319 g of oxygen. What is the empirical formula for this compound? What is its chemical name? Objective 16

Objective 16

61. A sample of a compound used to polish dentures and as a nutrient and dietary supplement is analyzed and found to contain 9.2402 g of calcium, 7.2183 g of phosphorus, and 13.0512 g of oxygen. What is the empirical formula for this compound?

62. A sample of an ionic compound that is used in the semiconductor industry is analyzed and found to contain 53.625 g of indium and 89.375 g of tellurium. What is the empirical formula for this compound? OBJECTIVE 16

63. An ionic compound that is 38.791% nickel, 33.011% arsenic, and 28.198% oxygen is employed as a catalyst for hardening fats used to make soap. What is the empirical formula for this compound? OBJECTIVE 17

64. An ionic compound sometimes called TKPP is used as a soap and detergent builder. It is 47.343% potassium, 18.753% phosphorus, and 33.904% oxygen. What is the empirical formula for this compound? OBJECTIVE 17

65. An ionic compound that contains 10.279% calcium, 65.099% iodine, and 24.622% oxygen is used in deodorants and in mouthwashes. What is the empirical formula for this compound? What do you think its chemical name is? (Consider the possibility that this compound contains more than one polyatomic ion.) OBJECTIVE 17

66. An ionic compound that is 62.56% lead, 8.46% nitrogen, and 28.98% oxygen is used as a mordant in the dyeing industry. A mordant helps to bind a dye to the fabric. What is the empirical formula for this compound? What do you think its name is? (Consider the possibility that this compound contains more than one polyatomic ion.) OBJECTIVE 17

67. In 1989 a controversy arose concerning the chemical daminozide, or Alar®, which was sprayed on apple trees to yield redder, firmer, and more shapely apples. Concerns about Alar's safety stemmed from the suspicion that one of its breakdown products, unsymmetrical dimethylhydrazine (UDMH), was carcinogenic. Alar is no longer sold for food uses. UDMH has the empirical formula of CNH_4 and has a molecular mass of 60.099. What is the molecular formula for UDMH? OBJECTIVE 18

68. It would be advisable for the smokers of the world to consider that nicotine was once used as an insecticide but is no longer used for that purpose because of safety concerns. Nicotine has an empirical formula of C_5H_7N and a molecular mass of 162.23. What is the molecular formula for nicotine? OBJECTIVE 18

69. Lindane is one of the chlorinated pesticides the use of which is now restricted in the United States. It is 24.78% carbon, 2.08% hydrogen, and 73.14% chlorine and has a molecular mass of 290.830. What is lindane's molecular formula? OBJECTIVE 17 OBJECTIVE 18

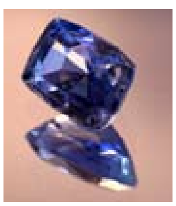
70. Hydralazine is a drug used to treat heart disease. It is 59.99% carbon, 5.03% hydrogen, and 34.98% nitrogen and has a molecular mass of 160.178. What is the molecular formula for hydralazine?

71. Melamine is a compound used to make the melamine-formaldehyde resins in very hard surface materials such as Formica®. It is 28.57% carbon, 4.80% hydrogen, and 66.63% nitrogen and has a molecular mass of 126.121. What is melamine's molecular formula? OBJECTIVE 17 OBJECTIVE 18

72. About 40 different substances called organophosphorus compounds are registered in the United States as insecticides. They are considered less damaging to the environment than some other insecticides because they breakdown relatively rapidly in the environment. The first of these organophosphorus insecticides to be produced was tetraethyl pyrophosphate, TEPP, which is 33.11% carbon, 6.95% hydrogen, 38.59% oxygen, and 21.35% phosphorus. It has a molecular mass of 290.190. What is the molecular formula for TEPP? OBJECTIVE 17 OBJECTIVE 18

Additional Problems

73. Your boss at the hardware store points you to a bin of screws and asks you to find out the approximate number of screws it contains. You weigh the screws and find that their total mass is 68 pounds. You take out 100 screws and weigh them individually, and you find that 7 screws weigh 2.65 g, 4 screws weigh 2.75 g, and 89 screws weigh 2.90 g. Calculate the weighted average mass of each screw. How many screws are in the bin? How many gross of screws are in the bin?
74. Atomic masses are derived from calculations using experimental data. As the experiments that provide this data get more precise, the data get more accurate, and the atomic masses values reported on the periodic table are revised. One source of data from 1964 reports that the element potassium is 93.10% potassium-39, which has atoms with a mass of 38.963714 u (atomic mass units), 0.0118% potassium-40, which has atoms with a mass of 39.964008 u, and 6.88% potassium-41, which has atoms with a mass of 40.961835 u. Using this data, calculate the weighted average mass of potassium atoms, in atomic mass units. Report your answer to the fourth decimal position. The weighted average mass of potassium atoms is potassium's atomic mass. How does your calculated value compare to potassium's reported atomic mass on the periodic table in this text?

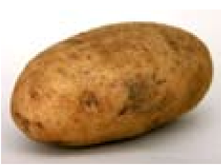


75. As a member of the corundum family of minerals, sapphire (the September birthstone) consists primarily of aluminum oxide, Al_2O_3. Small amounts of iron and titanium give it its rich dark blue color. Gem cutter Norman Maness carved a giant sapphire into the likeness of Abraham Lincoln. If this 2302-carat sapphire were pure aluminum oxide, how many moles of Al_2O_3 would it contain? (There are exactly 5 carats per gram.)

76. Emeralds are members of the beryl family, which are silicates of beryllium and aluminum with a general formula of $Be_3Al_2(SiO_3)_6$. The emerald's green color comes from small amounts of chromium in the crystal. The Viennese treasury has a cut emerald that weighs 2205 carats. If this stone were pure $Be_3Al_2(SiO_3)_6$, how many moles of $Be_3Al_2(SiO_3)_6$ would it contain? (There are exactly 5 carats per gram.)
77. Aquamarine (the March birthstone) is a light blue member of the beryl family, which is made up of natural silicates of beryllium and aluminum that have the general formula $Be_3Al_2(SiO_3)_6$. Aquamarine's bluish color is caused by trace amounts of iron(II) ions. A 43-pound aquamarine mined in Brazil in 1910 remains the largest gem-quality crystal ever found. If this stone were pure $Be_3Al_2(SiO_3)_6$, how many moles of beryllium would it contain?
78. In 1985 benitoite became the California "state gemstone." Found only in a tiny mine near Coalinga, California, it is a silicate of barium and titanium with trace impurities that cause a range of hues from colorless to blue to pink. Its general formula is $BaTi(SiO_3)_3$. If a 15-carat stone were pure $BaTi(SiO_3)_3$, how many moles of silicon would it contain?
79. November's birthstone is citrine, a yellow member of the quartz family. It is primarily silicon dioxide, but small amounts of iron(III) ions give it its yellow color. A high-quality citrine containing about 0.040 moles of SiO_2 costs around \$225. If this stone were pure SiO_2, how many carats would it weigh? (There are exactly 5 carats per gram.)

80. The gemstone tanzanite was first discovered in Tanzania in 1967. Like other gemstones, it contains impurities that give it distinct characteristics, but it is primarily $Ca_2Al_3Si_3O_{12}(OH)$. The largest tanzanite stone ever found contains about 0.0555 mole of $Ca_2Al_3Si_3O_{12}(OH)$. What is the mass of this stone in kilograms?
81. A common throat lozenge contains 29 mg of phenol, C_6H_5OH.
    a. How many moles of C_6H_5OH are there in 5.0 mg of phenol?
    b. What is the mass in kilograms of 0.9265 mole of phenol?
82. Some forms of hematite, a mineral composed of iron(III) oxide, can be used to make jewelry. Because of its iron content, hematite jewelry has a unique problem among stone jewelry. It shows signs of rusting. How many moles of iron are there in a necklace that contains 78.435 g of Fe_2O_3?
83. Beryl, $Be_3Al_2(SiO_3)_6$, is a natural source of beryllium, a known carcinogen. What is the mass in kilograms of beryllium in 1.006 Mg of $Be_3Al_2(SiO_3)_6$?
84. The element antimony is used to harden lead for use in lead-acid storage batteries. One of the principal antimony ores is stibnite, which contains antimony in the form of Sb_2S_3. Antimony is obtained through the reduction that occurs (from the +3 oxidation state to the zero oxidation state of pure antimony) when Sb_2S_3 reacts with the iron in iron scrap. What is the mass of antimony in 14.78 lb of Sb_2S_3?
85. Cermets (for *cer*amic plus *met*al) are synthetic substances with both ceramic and metallic components. They combine the strength and toughness of metal with the resistance to heat and oxidation that ceramics offer. One cermet containing molybdenum and silicon was used to coat molybdenum engine parts on space vehicles. A sample of this compound is analyzed and found to contain 14.212 g of molybdenum and 8.321 g of silicon. What is the empirical formula for this compound?
86. Blue vitriol is a common name for an ionic compound that has many purposes in industry, including the production of germicides, pigments, pharmaceuticals, and wood preservatives. A sample contains 20.238 g of copper, 10.213 g of sulfur, and 20.383 g of oxygen. What is its empirical formula? What is its name?
87. A compound that is sometimes called sorrel salt can be used to remove ink stains or to clean wood. It is 30.52% potassium, 0.787% hydrogen, 18.75% carbon, and 49.95% oxygen. What is the empirical formula for this compound?
88. The ionic compound sometimes called uranium yellow is used to produce colored glazes for ceramics. It is 7.252% sodium, 75.084% uranium, and 17.664% oxygen. What is the empirical formula for this compound?
89. An ionic compound that is 24.186% sodium, 33.734% sulfur, and 42.080% oxygen is used as a food preservative. What is its empirical formula?
90. A defoliant is an herbicide that removes leaves from trees and growing plants. One ionic compound used for this purpose is 12.711% magnesium, 37.083% chlorine, and 50.206% oxygen. What is the empirical formula for this compound? What do you think its chemical name is? (Consider the possibility that this compound contains more than one polyatomic ion.)

91. An ionic compound that is 22.071% manganese, 1.620% hydrogen, 24.887% phosphorus, and 51.422% oxygen is used as a food additive and dietary supplement. What is the empirical formula for this compound? What do you think its chemical name is? (Consider the possibility that this compound contains more than one polyatomic ion.)
92. Agent Orange, which was used as a defoliant in the Vietnam War, contained a mixture of two herbicides called 2,4,5-T and 2,4-D. (Agent Orange got its name from the orange barrels in which it was stored.) Some of the controversy that surrounds the use of Agent Orange is related to the discovery in 2,4,5-T of a trace impurity called TCDD. Although recent studies suggest that it is much more harmful to animals than to humans, TCDD has been called the most toxic small molecule known, and all uses of 2,4,5-T were banned in 1985. TCDD is 44.77% carbon, 1.25% hydrogen, 9.94% oxygen, and 44.04% chlorine and has a molecular mass of 321.97. What is the molecular formula for TCDD?
93. Thalidomide was used as a tranquilizer and flu medicine for pregnant women in Europe until it was found to cause birth defects. (The horrible effects of this drug played a significant role in the passage of the Kefauver-Harris Amendment to the Food and Drug Act, requiring that drugs be proved safe before they are put on the market.) Thalidomide is 60.47% carbon, 3.90% hydrogen, 24.78% oxygen, and 10.85% nitrogen and has a molecular mass of 258.23. What is the molecular formula for thalidomide?



94. Nabam is a fungicide used on potato plants. It is 17.94% sodium, 18.74% carbon, 2.36% hydrogen, 10.93% nitrogen, and 50.03% sulfur and has a formula mass of 256.35. Nabam is an ionic compound with a formula that is not an empirical formula. What is the formula for nabam?

Challenge Problems

95. Calamine is a naturally occurring zinc silicate that contains the equivalent of 67.5% zinc oxide, ZnO. (The term *calamine* also refers to a substance used to make calamine lotion.) What is the mass, in kilograms, of zinc in 1.347×10^4 kg of natural calamine that is 67.5% ZnO?
96. Zirconium metal, which is used to coat nuclear fuel rods, can be made from the zirconium(IV) oxide, ZrO_2, in the zirconium ore called baddeleyite (or zirconia). What maximum mass in kilograms of zirconium metal can be extracted from 1.2×10^3 kg of baddeleyite that is 53% ZrO_2?
97. Flue dust from the smelting of copper and lead contains As_2O_3. (Smelting is the heating of a metal ore until it melts, so that its metallic components can be separated.) When this flue dust is collected, it contains 90% to 95% As_2O_3. The arsenic in As_2O_3 can be reduced to the element arsenic by reaction with charcoal. What is the maximum mass, in kilograms, of arsenic that can be formed from 67.3 kg of flue dust that is 93% As_2O_3?
98. Thortveitite is a natural ore that contains from 37% to 42% scandium oxide, Sc_2O_3. Scandium metal is made by first reacting the Sc_2O_3 with ammonium hydrogen fluoride, NH_4HF_2, to form scandium fluoride, ScF_3. The scandium in ScF_3 is reduced to metallic scandium in a reaction with calcium. What is the maximum mass, in kilograms, of scandium metal that can be made from 1230.2 kilograms of thortveitite that is 39% Sc_2O_3?

99. Magnesium metal, which is used to make die-cast auto parts, missiles, and space vehicles, is obtained by the electrolysis of magnesium chloride. Magnesium hydroxide forms magnesium chloride when it reacts with hydrochloric acid. There are two common sources of magnesium hydroxide.

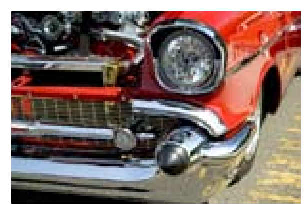
a. Magnesium ions can be precipitated from seawater as magnesium hydroxide, $Mg(OH)_2$. Each kiloliter of seawater yields about 3.0 kg of the compound. How many metric tons of magnesium metal can be made from the magnesium hydroxide derived from 1.0×10^5 kL of seawater?

b. Brucite is a natural form of magnesium hydroxide. A typical crude ore containing brucite is 29% $Mg(OH)_2$. What minimum mass, in metric tons, of this crude ore is necessary to make 34.78 metric tons of magnesium metal?

100. Spodumene is a lithium aluminum silicate containing the equivalent of 6.5% to 7.5% lithium oxide, Li_2O. Crude ore mined in North Carolina contains 15% to 20% spodumene. What maximum mass, in kilograms, of lithium could be formed from 2.538 megagrams of spodumene containing the equivalent of 7.0% Li_2O?

101. The element fluorine can be obtained by the electrolysis of combinations of hydrofluoric acid and potassium fluoride. These compounds can be made from the calcium fluoride, CaF_2, found in nature as the mineral fluorite. Fluorite's commercial name is fluorspar. Crude ores containing fluorite have 15% to 90% CaF_2. What minimum mass, in metric tons, of crude ore is necessary to make 2.4 metric tons of fluorine if the ore is 72% CaF_2?

102. Chromium metal is used in metal alloys and as a surface plating on other metals to minimize corrosion. It can be obtained by reducing the chromium(III) in chromium(III) oxide, Cr_2O_3, to the uncharged metal with finely divided aluminum. The Cr_2O_3 is found in an ore called chromite. What is the maximum mass, in kilograms, of chromium that can be made from 143.0 metric tons of Cuban chromite ore that is 38% Cr_2O_3?



103. What mass of baking powder that is 36% $NaHCO_3$ contains 1.0 mole of sodium hydrogen carbonate?

104. Roscoelite is a vanadium-containing form of mica used to make vanadium metal. Although the fraction of vanadium in the ore is variable, roscoelite can be described as $K_2V_4Al_2Si_6O_{20}(OH)_4$. Another way to describe the vanadium content of this mineral is to say that it has the equivalent of up to 28% V_2O_3.

a. What is the mass, in grams, of vanadium in 123.64 g of roscoelite?

b. What is the mass, in kilograms, of vanadium in 6.71 metric tons of roscoelite that contains the equivalent of 28% V_2O_3?

105. Hafnium metal is used to make control rods in water-cooled nuclear reactors and to make filaments in light bulbs. The hafnium is found with zirconium in zircon sand, which is about 1% hafnium(IV) oxide, HfO_2. What minimum mass, in metric tons, of zircon sand is necessary to make 120.5 kg of hafnium metal if the sand is 1.3% HfO_2?