3 Stoichiometry

Contents

The violent chemical reaction of bromine and phosphorus.

Chemical reactions have a profound effect on our lives. There are many examples: Food is converted to energy in the human body; nitrogen and hydrogen are combined to form ammonia, which is used as a fertilizer; fuels and plastics are produced from petroleum; the starch in plants is synthesized from carbon dioxide and water using energy from sunlight; human insulin is produced in laboratories by bacteria; cancer is induced in humans by substances from our environment; and so on, in a seemingly endless list. The central activity of chemistry is to understand chemical changes such as these, and the study of reactions occupies a central place in this book. We will examine why reactions occur, how fast they occur, and the specific pathways they follow.

In this chapter we will consider the quantities of materials consumed and produced in chemical reactions. This area of study is called **chemical stoichiometry** (pronounced stoy·kē·om′·etry). To understand chemical stoichiometry, you must first understand the concept of relative atomic masses.

3.1 Counting by Weighing

Suppose you work in a candy store that sells gourmet jelly beans by the bean. People come in and ask for 50 beans, 100 beans, 1000 beans, and so on, and you have to count them out—a tedious process at best. As a good problem solver, you try to come up with a better system. It occurs to you that it might be far more efficient to buy a scale and count the jelly beans by weighing them. How can you count jelly beans by weighing them? What information about the individual beans do you need to know?

Assume that all of the jelly beans are identical and that each has a mass of 5 g. If a customer asks for 1000 jelly beans, what mass of jelly beans would be required? Each bean has a mass of 5 g, so you would need 1000 beans $\times$ 5 g/bean, or 5000 g (5 kg). It takes just a few seconds to weigh out 5 kg of jelly beans. It would take much longer to count out 1000 of them.

In reality, jelly beans are not identical. For example, let's assume that you weigh 10 beans individually and get the following results:

Bean	Mass
1	5.1 g
2	5.2 g
3	5.0 g
4	4.8 g
5	4.9 g
6	5.0 g
7	5.0 g
8	5.1 g
9	4.9 g
10	5.0 g

Jelly beans can be counted by weighing.

Can we count these nonidentical beans by weighing? Yes. The key piece of information we need is the *average mass* of the jelly beans. Let's compute the average mass for our 10-bean sample.

$$\text{Average mass} = \frac{\text{total mass of beans}}{\text{number of beans}}$$

$$= \frac{5.1\text{ g} + 5.2\text{ g} + 5.0\text{ g} + 4.8\text{ g} + 4.9\text{ g} + 5.0\text{ g} + 5.0\text{ g} + 5.1\text{ g} + 4.9\text{ g} + 5.0\text{ g}}{10}$$

$$= \frac{50.0}{10} = 5.0\text{ g}$$

The average mass of a jelly bean is 5.0 g. Thus, to count out 1000 beans, we need to weigh out 5000 g of beans. This sample of beans, in which the beans have an average mass of 5.0 g, can be treated exactly like a sample where all of the beans ae identical. Objects do not need to have identical masses to be counted by weighing. We simply need to know the average mass of the objects. For purposes of counting, the objects *behave as though they were all identical,* as though they each actually had the average mass.

We count atoms in exactly the same way. Because atoms are so small, we deal with samples of matter that contain huge numbers of atoms. Even if we could see the atoms it would not be possible to count them directly. Thus we determine the number of atoms in a given sample by finding its mass. However, just as with jelly beans, to relate the mass to a number of atoms, we must know the average mass of the atoms.

3.2 Atomic Masses

As we saw in Chapter 2, the first quantitative information about atomic masses came from the work of Dalton, Gay-Lussac, Lavoisier, Avogadro, and Berzelius. By observing the proportions in which elements combine to form various compounds, nineteenth-century chemists calculated relative atomic masses. The modern system of atomic masses, instituted in 1961, is based on ^{12}C ("carbon twelve") as the standard. In this system, *^{12}C is assigned a mass of exactly 12 atomic mass units* (amu), and the masses of all other atoms are given relative to this standard.

The most accurate method currently available for comparing the masses of atoms involves the use of the **mass spectrometer.** In this instrument, diagramed in Fig. 3.1, atoms or molecules are passed into a beam of high-speed electrons, which knock electrons off the atoms or molecules being analyzed and change them into positive ions. An applied

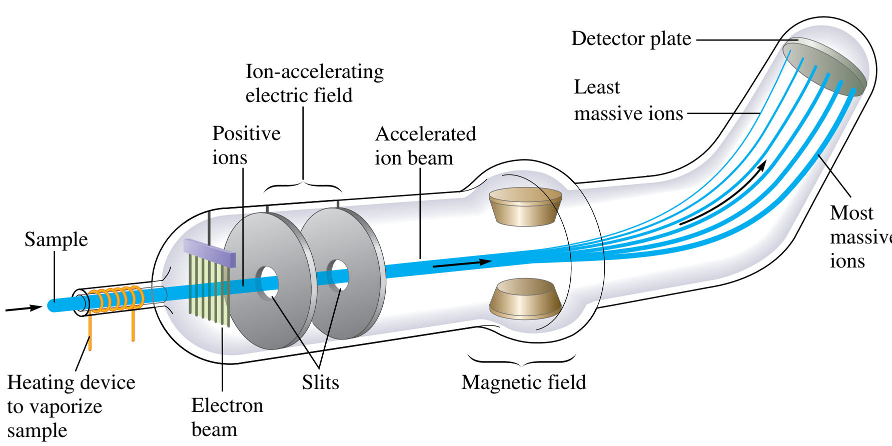


FIGURE 3.1
(left) A scientist injecting a sample into a mass spectrometer. (above) Schematic diagram of a mass spectrometer.

electric field then accelerates these ions into a magnetic field. Because an accelerating ion produces its own magnetic field, an interaction with the applied magnetic field occurs, which tends to change the path of the ion. The amount of path deflection for each ion depends on its mass—the most massive ions are deflected the smallest amount—which causes the ions to separate, as shown in Fig. 3.1. A comparison of the positions where the ions hit the detector plate gives very accurate values of their relative masses. For example, when ^{12}C and ^{13}C are analyzed in a mass spectrometer, the ratio of their masses is found to be

$$\frac{\text{Mass } ^{13}\text{C}}{\text{Mass } ^{12}\text{C}} = 1.0836129$$

Since the atomic mass unit is defined such that the mass of ^{12}C is *exactly* 12 atomic mass units, then on this same scale,

$$\text{Mass of } ^{13}\text{C} = (1.0836129)(\underset{\substack{\uparrow \\ \text{Exact number} \\ \text{by definition}}}{12} \text{ amu}) = 13.003355 \text{ amu}$$

The masses of other atoms can be determined in a similar fashion.

The mass for each element is given in the table inside the front cover of this text. This value, even though it is actually a mass, is (for historical reasons) sometimes called the *atomic weight* for each element.

Most elements occur in nature as mixtures of isotopes; thus atomic masses are usually average values.

Look at the value of the atomic mass of carbon given in this table. You might expect to see 12, since we said the system of atomic masses is based on ^{12}C. However, the number given for carbon is not 12 but 12.01. Why? The reason for this apparent discrepancy is that the carbon found on earth (natural carbon) is a mixture of the isotopes ^{12}C, ^{13}C, and ^{14}C. All three isotopes have six protons, but they have six, seven, and eight neutrons, respectively. Because natural carbon is a mixture of isotopes, the atomic mass we use for carbon is an *average value* reflecting the average of the isotopes composing it.

It is much easier to weigh out 600 hex nuts than count them one by one.

The average atomic mass for carbon is computed as follows: It is known that natural carbon is composed of 98.89% ^{12}C atoms and 1.11% ^{13}C atoms. The amount of ^{14}C is negligibly small at this level of precision. Using the masses of ^{12}C (exactly 12 amu) and ^{13}C (13.003355 amu), we can calculate the average atomic mass for natural carbon as follows:

$$98.89\% \text{ of } 12 \text{ amu} + 1.11\% \text{ of } 13.0034 \text{ amu} = (0.9889)(12 \text{ amu}) + (0.0111)(13.0034 \text{ amu}) = 12.01 \text{ amu}$$

In this text we will call the average mass for an element the **average atomic mass** or, simply, the *atomic mass* for that element.

Even though natural carbon does not contain a single atom with mass 12.01, for stoichiometric purposes, we can consider carbon to be composed of only one type of atom with a mass of 12.01. This enables us to count atoms of natural carbon by weighing a sample of carbon.

Recall from Section 3.1 that counting by weighing works if you know the *average* mass of the units being counted. Counting by weighing works just the same for atoms as for jelly beans. For natural carbon with an average mass of 12.01 atomic mass units, to obtain 1000 atoms would require weighing out 12,010 atomic mass units of natural carbon (a mixture of ^{12}C and ^{13}C).

As in the case of carbon, the mass for each element listed in the table inside the front cover of the text is an average value based on the isotopic composition of the naturally occurring element. For instance, the mass listed for hydrogen (1.008) is the average mass for natural hydrogen, which is a mixture of ^{1}H and ^{2}H (deuterium). *No* atom of hydrogen actually has the mass 1.008.

CHEMICAL IMPACT

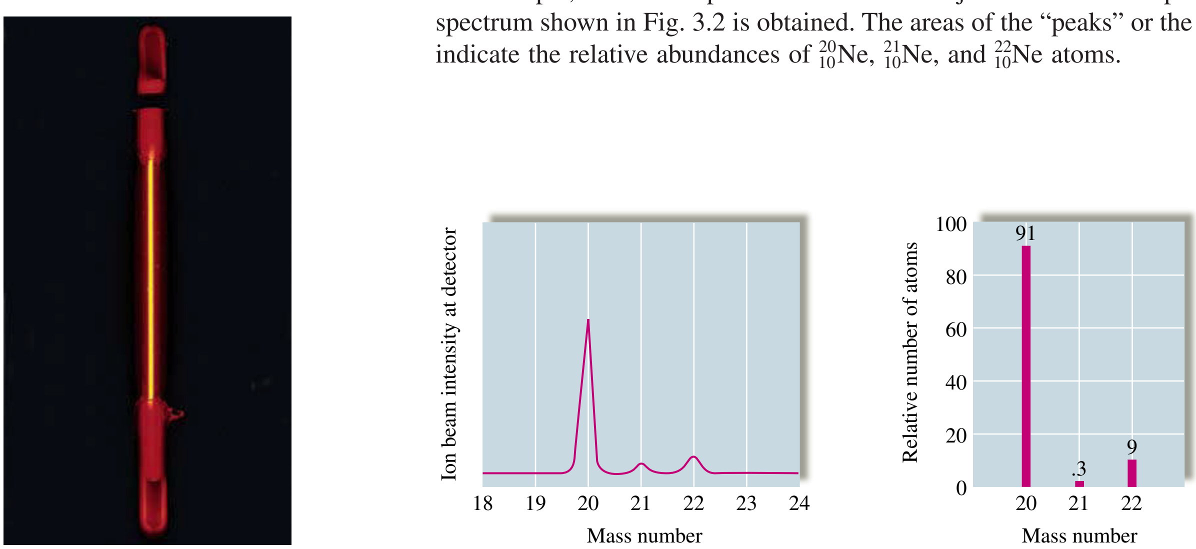
Buckyballs Teach Some History

About 250 million years ago, 90% of life on earth was destroyed in some sort of cataclysmic event. This event, which ended the Permian period and began the Triassic (the P-T boundary), is the most devastating mass extinction in the earth's history—far surpassing the catastrophe 65 million years ago that wiped out the dinosaurs (the K-T boundary). In 1979 geologist Walter Alvarez and his Nobel Prize–winning physicist father Luis Alvarez suggested that unusually high concentrations of iridium in rocks laid down at the K-T boundary meant that an asteroid had hit the earth, causing tremendous devastation. In the last 20 years much evidence has accumulated to support this hypothesis, including identification of the location of the probable crater caused by the impact in the ocean near Mexico.

Were the P-T boundary extinctions also caused by an extraterrestrial object or by some event on earth, such as a massive volcano explosion? Recent discoveries by geochemists Luann Becker of the University of Washington and Robert J. Poreda of the University of Rochester seem to strongly support the impact theory. Examining sediment from China and Japan, the team found fullerenes encapsulating argon and helium gas atoms whose isotopic composition indicates that they are extraterrestrial in origin. For example, the ratio of ${}^{3}_{2}He$ to ${}^{4}_{2}He$ found in the fullerenes is 100 times greater than the ratio for helium found in the earth's atmosphere. Likewise, the isotopic composition of the fullerene-trapped argon atoms is quite different from that found on earth.

Fullerenes include spherical C_{60} carbon molecules ("buckyballs") whose cavities can trap other atoms such as helium and argon. (See the accompanying figure.) The scientists postulate that the fullerenes originated in stars or collapsing gas clouds where the noble gas atoms were trapped as the fullerenes formed. These fullerenes were then somehow incorporated into the object that eventually hit the earth. Based on the isotopic compositions, the geochemists estimate that the impacting body must have

In addition to being useful for determining accurate mass values for individual atoms, the mass spectrometer is used to determine the isotopic composition of a natural element. For example, when a sample of natural neon is injected into a mass spectrometer, the mass spectrum shown in Fig. 3.2 is obtained. The areas of the "peaks" or the heights of the bars indicate the relative abundances of ${}^{20}_{10}Ne$, ${}^{21}_{10}Ne$, and ${}^{22}_{10}Ne$ atoms.



(a) (b) (c)

FIGURE 3.2
(a) Neon gas glowing in a discharge tube. The relative intensities of the signals recorded when natural neon is injected into a mass spectrometer, represented in terms of (b) "peaks" and (c) a bar graph. The relative areas of the peaks are 0.9092 (^{20}Ne), 0.00257 (^{21}Ne), and 0.0882 (^{22}Ne); natural neon is therefore 90.92% ^{20}Ne, 0.257% ^{21}Ne, and 8.82% ^{22}Ne.

been 10 kilometers in diameter, which is comparable in size to the asteroid that is assumed to have killed the dinosaurs.

One factor that had previously cast doubt on an asteroid collision as the cause of the P-T catastrophe was the lack of iridium found in sediments from this period. However, Becker and other scientists argue that this absence probably means the impacting object may have been a comet rather than an asteroid. It is also possible that such a blow could have intensified the volcanism already under way on earth at that time, delivering a "one-two punch" that almost obliterated life on earth, according to Becker.

It is ironic that "buckyballs," which made big news when they were recently synthesized for the first time in the laboratory, actually have been around for millions of years and have some very interesting history to teach us.

Figure from *Chemical and Engineering News,* Feb. 26, 2001, p. 9. Reprinted by permission of Joseph Wilmhoff.

Isotope ratios of the noble gas atoms inside celestial buckyballs indicate that these ancient carbon cages formed in a stellar environment, not on earth.

Sample Exercise 3.1

The Average Mass of an Element

Copper nugget.

When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in Fig. 3.3 are obtained. Use these data to compute the average mass of natural copper. (The mass values for ^{63}Cu and ^{65}Cu are 62.93 amu and 64.93 amu, respectively.)

Solution

As shown by the graph, of every 100 atoms of natural copper, 69.09 are ^{63}Cu and 30.91 are ^{65}Cu. Thus the mass of 100 atoms of natural copper is

$$(69.09\ \cancel{\text{atoms}})\left(62.93\,\frac{\text{amu}}{\cancel{\text{atom}}}\right) + (30.91\ \cancel{\text{atoms}})\left(64.93\,\frac{\text{amu}}{\cancel{\text{atom}}}\right) = 6355\ \text{amu}$$

The *average mass* of a copper atom is

$$\frac{6355\ \text{amu}}{100\ \text{atoms}} = 63.55\ \text{amu/atom}$$

This mass value is used in doing calculations involving the reactions of copper and is the value given in the table inside the front cover of this book.

Reality Check: When you finish a calculation, you should always check whether your answer makes sense. In this case our answer of 63.55 amu is between the masses of the atoms that make up natural copper. This makes sense. The answer could not be smaller than 62.93 amu or larger than 64.93 amu.

See Exercises 3.27 and 3.28.

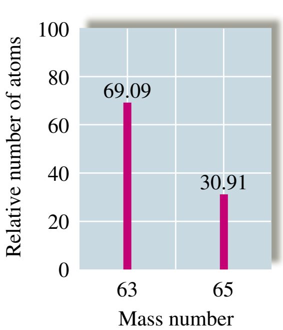


FIGURE 3.3
Mass spectrum of natural copper.

3.3 The Mole

The SI definition of the mole is the amount of a substance that contains as many entities as there are in exactly 12 g of carbon-12.

Avogadro's number is 6.022×10^{23}. One mole of anything is 6.022×10^{23} units of that substance.

Because samples of matter typically contain so many atoms, a unit of measure called the *mole* has been established for use in counting atoms. For our purposes, it is most convenient to define the **mole** (abbreviated mol) as *the number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C.* Techniques such as mass spectrometry, which count atoms very precisely, have been used to determine this number as 6.02214×10^{23} (6.022×10^{23} will be sufficient for our purposes). This number is called **Avogadro's number** to honor his contributions to chemistry. *One mole of something consists of 6.022×10^{23} units of that substance.* Just as a dozen eggs is 12 eggs, a mole of eggs is 6.022×10^{23} eggs.

The magnitude of the number 6.022×10^{23} is very difficult to imagine. To give you some idea, 1 mole of seconds represents a span of time 4 million times as long as the earth has already existed, and 1 mole of marbles is enough to cover the entire earth to a depth of 50 miles! However, since atoms are so tiny, a mole of atoms or molecules is a perfectly manageable quantity to use in a reaction (see Fig. 3.4).

How do we use the mole in chemical calculations? Recall that Avogadro's number is defined as the number of atoms in exactly 12 grams of ^{12}C. This means that 12 grams of ^{12}C contains 6.022×10^{23} atoms. It also means that a 12.01-gram sample of natural carbon contains 6.022×10^{23} atoms (a mixture of ^{12}C, ^{13}C, and ^{14}C atoms, with an average atomic mass of 12.01). Since the ratio of the masses of the samples (12 g/12.01 g) is the same as the ratio of the masses of the individual components (12 amu/12.01 amu), the two samples contain the *same number* of atoms (6.022×10^{23}).

To be sure this point is clear, think of oranges with an average mass of 0.5 pound each and grapefruit with an average mass of 1.0 pound each. Any two sacks for which the sack of grapefruit weighs twice as much as the sack of oranges will contain the same number of pieces of fruit. The same idea extends to atoms. Compare natural carbon (average mass of 12.01) and natural helium (average mass of 4.003). A sample of 12.01 grams of natural carbon contains the same number of atoms as 4.003 grams of natural helium. Both samples contain 1 mole of atoms (6.022×10^{23}). Table 3.1 gives more examples that illustrate this basic idea.

The mass of 1 mole of an element is equal to its atomic mass in grams.

Thus the mole is defined such that a sample of a natural element with a mass equal to the element's atomic mass expressed in grams contains 1 mole of atoms. This definition

FIGURE 3.4
Proceeding clockwise from the top, samples containing one mole each of copper, aluminum, iron, sulfur, iodine, and (in the center) mercury.

TABLE 3.1 Comparison of 1 Mole Samples of Various Elements

Element	Number of Atoms Present	Mass of Sample (g)
Aluminum	6.022×10^{23}	26.98
Copper	6.022×10^{23}	63.55
Iron	6.022×10^{23}	55.85
Sulfur	6.022×10^{23}	32.07
Iodine	6.022×10^{23}	126.9
Mercury	6.022×10^{23}	200.6

also fixes the relationship between the atomic mass unit and the gram. Since 6.022×10^{23} atoms of carbon (each with a mass of 12 amu) have a mass of 12 g, then

$$(6.022 \times 10^{23}\ \cancel{\text{atoms}})\left(\frac{12\ \text{amu}}{\cancel{\text{atom}}}\right) = 12\ \text{g}$$

and

$$6.022 \times 10^{23}\ \text{amu} = \underset{\substack{\uparrow \\ \text{Exact} \\ \text{number}}}{1}\ \text{g}$$

This relationship can be used to derive the unit factor needed to convert between atomic mass units and grams.

Sample Exercise 3.2 Determining the Mass of a Sample of Atoms

Americium is an element that does not occur naturally. It can be made in very small amounts in a device known as a *particle accelerator*. Compute the mass in grams of a sample of americium containing six atoms.

Solution

From the table inside the front cover of the text, we note that one americium atom has a mass of 243 amu. Thus the mass of six atoms is

$$6\ \cancel{\text{atoms}} \times 243\ \frac{\text{amu}}{\cancel{\text{atom}}} = 1.46 \times 10^{3}\ \text{amu}$$

Using the relationship

$$6.022 \times 10^{23}\ \text{amu} = 1\ \text{g}$$

we write the conversion factor for converting atomic mass units to grams:

$$\frac{1\ \text{g}}{6.022 \times 10^{23}\ \text{amu}}$$

The mass of six americium atoms in grams is

$$1.46 \times 10^{3}\ \cancel{\text{amu}} \times \frac{1\ \text{g}}{6.022 \times 10^{23}\ \cancel{\text{amu}}} = 2.42 \times 10^{-21}\ \text{g}$$

Reality Check: Since this sample contains only six atoms, the mass should be very small as the amount 2.42×10^{-21} g indicates.

See Exercise 3.33.

CHEMICAL IMPACT

Elemental Analysis Catches Elephant Poachers

In an effort to combat the poaching of elephants by controlling illegal exports of ivory, scientists are now using the isotopic composition of ivory trinkets and elephant tusks to identify the region of Africa where the elephant lived. Using a mass spectrometer, scientists analyze the ivory for the relative amounts of ^{12}C, ^{13}C, ^{14}N, ^{15}N, ^{86}Sr, and ^{87}Sr to determine the diet of the elephant and thus its place of origin. For example, because grasses use a different photosynthetic pathway to produce glucose than do trees, grasses have a slightly different $^{13}C/^{12}C$ ratio from that of trees. They have different ratios because each time a carbon atom is added in going from simpler to more complex compounds, the more massive ^{13}C is disfavored relative to ^{12}C because it reacts more slowly. Because trees use more steps to build up glucose, they end up with a smaller $^{13}C/^{12}C$ ratio in their leaves relative to grasses, and this difference is then reflected in the tissues of elephants. Thus scientists can tell whether a particular tusk came from a savanna-dwelling elephant (grass-eating) or from a tree-browsing elephant.

Similarly, because the ratios of $^{15}N/^{14}N$ and $^{87}Sr/^{86}Sr$ in elephant tusks also vary depending on the region of Africa the elephant inhabits, they can be used to trace the elephant's origin. In fact, using these techniques, scientists have reported being able to discriminate between elephants living only about 100 miles apart.

There is now international concern about the dwindling elephant populations in Africa—their numbers have decreased significantly in recent years. This concern has led to bans in the export of ivory from many countries in Africa. However, a few nations still allow ivory to be exported. Thus, to enforce the trade restrictions, the origin of a given piece of ivory must be established. It is hoped that the "isotope signature" of the ivory can be used for this purpose.

To do chemical calculations, you must understand what the mole means and how to determine the number of moles in a given mass of a substance. These procedures are illustrated in Sample Exercises 3.3 and 3.4.

Sample Exercise 3.3

Determining Moles of Atoms

Aluminum (Al) is a metal with a high strength-to-mass ratio and a high resistance to corrosion; thus it is often used for structural purposes. Compute both the number of moles of atoms and the number of atoms in a 10.0-g sample of aluminum.

(left) Pure aluminum. (right) Aluminum alloys are used for many high-quality bicycle components, such as this chain wheel.

Solution

The mass of 1 mole (6.022×10^{23} atoms) of aluminum is 26.98 g. The sample we are considering has a mass of 10.0 g. Since the mass is less than 26.98 g, this sample contains less than 1 mole of aluminum atoms. We can calculate the number of moles of aluminum atoms in 10.0 g as follows:

$$10.0\ \cancel{\text{g Al}} \times \frac{1 \text{ mol Al}}{26.98\ \cancel{\text{g Al}}} = 0.371 \text{ mol Al atoms}$$

The number of atoms in 10.0 g (0.371 mol) of aluminum is

$$0.371 \cancel{\text{mol Al}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \cancel{\text{mol Al}}} = 2.23 \times 10^{23} \text{ atoms}$$

Reality Check: One mole of Al has a mass of 26.98 g and contains 6.022×10^{23} atoms. Our sample is 10.0 g, which is roughly 1/3 of 26.98. Thus the calculated amount should be on the order of 1/3 of 6×10^{23}, which it is.

See Exercise 3.34.

Sample Exercise 3.4

Calculating Numbers of Atoms

A silicon chip used in an integrated circuit of a microcomputer has a mass of 5.68 mg. How many silicon (Si) atoms are present in the chip?

Solution

The strategy for doing this problem is to convert from milligrams of silicon to grams of silicon, then to moles of silicon, and finally to atoms of silicon:

Always check to see if your answer is sensible.

Paying careful attention to units and making sure the answer is reasonable can help you detect an inverted conversion factor or a number that was incorrectly entered in your calculator.

$$5.68 \cancel{\text{mg Si}} \times \frac{1 \text{ g Si}}{1000 \cancel{\text{mg Si}}} = 5.68 \times 10^{-3} \text{ g Si}$$

$$5.68 \times 10^{-3} \cancel{\text{g Si}} \times \frac{1 \text{ mol Si}}{28.09 \cancel{\text{g Si}}} = 2.02 \times 10^{-4} \text{ mol Si}$$

$$2.02 \times 10^{-4} \cancel{\text{mol Si}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \cancel{\text{mol Si}}} = 1.22 \times 10^{20} \text{ atoms}$$

Reality Check: Note that 5.68 mg of silicon is clearly much less than 1 mol of silicon (which has a mass of 28.09 g), so the final answer of 1.22×10^{20} atoms (compared with 6.022×10^{23} atoms) is in the right direction.

See Exercise 3.35.

Sample Exercise 3.5

Calculating the Number of Moles and Mass

Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing 5.00×10^{20} atoms and the mass of the sample.

Fragments of cobalt metal.

Solution

Note that the sample of 5.00×10^{20} atoms of cobalt is less than 1 mole (6.022×10^{23} atoms) of cobalt. What fraction of a mole it represents can be determined as follows:

$$5.00 \times 10^{20} \cancel{\text{atoms Co}} \times \frac{1 \text{ mol Co}}{6.022 \times 10^{23} \cancel{\text{atoms Co}}} = 8.30 \times 10^{-4} \text{ mol Co}$$

Since the mass of 1 mole of cobalt atoms is 58.93 g, the mass of 5.00×10^{20} atoms can be determined as follows:

$$8.30 \times 10^{-4} \cancel{\text{mol Co}} \times \frac{58.93 \text{ g Co}}{1 \cancel{\text{mol Co}}} = 4.89 \times 10^{-2} \text{ g Co}$$

Reality Check: In this case the sample contains 5×10^{20} atoms, which is approximately 1/1000 of a mole. Thus the sample should have a mass of about $(1/1000)(58.93) \cong 0.06$. Our answer of ~ 0.05 makes sense.

See Exercise 3.36.

3.4 Molar Mass

A chemical compound is, ultimately, a collection of atoms. For example, methane (the major component of natural gas) consists of molecules that each contain one carbon and four hydrogen atoms (CH_4). How can we calculate the mass of 1 mole of methane; that is, what is the mass of 6.022×10^{23} CH_4 molecules? Since each CH_4 molecule contains one carbon atom and four hydrogen atoms, 1 mole of CH_4 molecules contains 1 mole of carbon atoms and 4 moles of hydrogen atoms. The mass of 1 mole of methane can be found by summing the masses of carbon and hydrogen present:

In this case, the term 12.01 limits the number of significant figures.

$$\begin{aligned} \text{Mass of 1 mol C} &= 12.01 \text{ g} \\ \text{Mass of 4 mol H} &= \underline{4 \times 1.008 \text{ g}} \\ \text{Mass of 1 mol } CH_4 &= 16.04 \text{ g} \end{aligned}$$

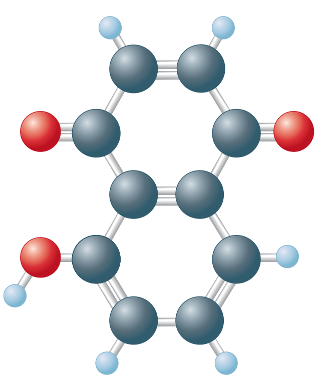
A substance's molar mass is the mass in grams of 1 mole of the substance.

Because 16.04 g represents the mass of 1 mole of methane molecules, it makes sense to call it the *molar mass* for methane. Thus the **molar mass** of a substance is *the mass in grams of one mole of the compound.* Traditionally, the term *molecular weight* has been used for this quantity. However, we will use molar mass exclusively in this text. The molar mass of a known substance is obtained by summing the masses of the component atoms as we did for methane.

Sample Exercise 3.6 Calculating Molar Mass I

Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants. The formula for juglone is $C_{10}H_6O_3$.

a. Calculate the molar mass of juglone.
b. A sample of 1.56×10^{-2} g of pure juglone was extracted from black walnut husks. How many moles of juglone does this sample represent?


Juglone

Solution

a. The molar mass is obtained by summing the masses of the component atoms. In 1 mole of juglone there are 10 moles of carbon atoms, 6 moles of hydrogen atoms, and 3 moles of oxygen atoms:

$$\begin{aligned} 10\text{ C:} \quad 10 \times 12.01 \text{ g} &= 120.1 \text{ g} \\ 6\text{ H:} \quad 6 \times 1.008 \text{ g} &= 6.048 \text{ g} \\ 3\text{ O:} \quad 3 \times 16.00 \text{ g} &= \underline{48.00 \text{ g}} \\ \text{Mass of 1 mol } C_{10}H_6O_3 &= 174.1 \text{ g} \end{aligned}$$

The mass of 1 mole of juglone is 174.1 g, which is the molar mass.

b. The mass of 1 mole of this compound is 174.1 g; thus 1.56×10^{-2} g is much less than a mole. The exact fraction of a mole can be determined as follows:

$$1.56 \times 10^{-2} \cancel{\text{g juglone}} \times \frac{1 \text{ mol juglone}}{174.1 \cancel{\text{g juglone}}} = 8.96 \times 10^{-5} \text{ mol juglone}$$

See Exercises 3.39 through 3.42.

Sample Exercise 3.7 Calculating Molar Mass II

Calcium carbonate ($CaCO_3$), also called *calcite,* is the principal mineral found in limestone, marble, chalk, pearls, and the shells of marine animals such as clams.

CHEMICAL IMPACT

Measuring the Masses of Large Molecules, or Making Elephants Fly

When a chemist produces a new molecule, one crucial property for making a positive identification is the molecule's mass. There are many ways to determine the molar mass of a compound, but one of the fastest and most accurate methods involves mass spectrometry. This method requires that the substance be put into the gas phase and ionized. The deflection that the resulting ion exhibits as it is accelerated through a magnetic field can be used to obtain a very precise value of its mass. One drawback of this method is that it is difficult to use with large molecules because they are difficult to vaporize. That is, substances that contain large molecules typically have very high boiling points, and these molecules are often damaged when they are vaporized at such high temperatures. A case in point involves proteins, an extremely important class of large biologic molecules that are quite fragile at high temperatures. Typical methods used to obtain the masses of protein molecules are slow and tedious.

Mass spectrometry has not been used previously to obtain protein masses because proteins decompose at the temperatures necessary to vaporize them. However, a new technique called *matrix-assisted laser desorption* has been developed that allows mass spectrometric determination of protein molar masses. In this technique, the large "target" molecule is embedded in a matrix of smaller molecules. The matrix is then placed in a mass spectrometer and blasted with a laser beam, which causes its disintegration. Disintegration of the matrix frees the large target molecule, which is then swept into the mass spectrometer. One researcher involved in this project likened this method to an elephant on top of a tall building: "The elephant must fly if the building is suddenly turned into fine grains of sand."

This technique allows scientists to determine the mass of huge molecules. So far researchers have measured proteins with masses up to 350,000 daltons (1 dalton is the mass of a hydrogen atom). This method, which makes mass spectrometry a routine tool for the determination of protein masses, probably will be extended to even larger molecules such as DNA and could be a revolutionary development in the characterization of biomolecules.

Calcite crystals.

a. Calculate the molar mass of calcium carbonate.

b. A certain sample of calcium carbonate contains 4.86 moles. What is the mass in grams of this sample? What is the mass of the CO_3^{2-} ions present?

Solution

a. Calcium carbonate is an ionic compound composed of Ca^{2+} and CO_3^{2-} ions. In 1 mole of calcium carbonate there are 1 mole of Ca^{2+} ions and 1 mole of CO_3^{2-} ions. The molar mass is calculated by summing the masses of the components:

$$\begin{array}{lll} 1\ Ca^{2+}: & 1 \times 40.08\ g = & 40.08\ g \\ 1\ CO_3^{2-}: & & \\ 1\ C: & 1 \times 12.01\ g = & 12.01\ g \\ 3\ O: & 3 \times 16.00\ g = & \underline{48.00\ g} \\ \multicolumn{2}{l}{\text{Mass of 1 mol } CaCO_3 =} & 100.09\ g \end{array}$$

Thus the mass of 1 mole of $CaCO_3$ (1 mol Ca^{2+} plus 1 mol CO_3^{2-}) is 100.09 g. This is the molar mass.

b. The mass of 1 mole of $CaCO_3$ is 100.09 g. The sample contains nearly 5 moles, or close to 500 g. The exact amount is determined as follows:

$$4.86\ \cancel{\text{mol } CaCO_3} \times \frac{100.09\ g\ CaCO_3}{1\ \cancel{\text{mol } CaCO_3}} = 486\ g\ CaCO_3$$

To find the mass of carbonate ions (CO_3^{2-}) present in this sample, we must realize that 4.86 moles of $CaCO_3$ contains 4.86 moles of Ca^{2+} ions and 4.86 moles of CO_3^{2-} ions. The mass of 1 mole of CO_3^{2-} ions is

$$\begin{aligned} 1\text{ C:}\quad 1 \times 12.01 &= 12.01\text{ g} \\ 3\text{ O:}\quad 3 \times 16.00 &= \underline{48.00\text{ g}} \\ \text{Mass of 1 mol } CO_3^{2-} &= 60.01\text{ g} \end{aligned}$$

Thus the mass of 4.86 moles of CO_3^{2-} ions is

$$4.86\ \cancel{\text{mol } CO_3^{2-}} \times \frac{60.01\text{ g } CO_3^{2-}}{1\ \cancel{\text{mol } CO_3^{2-}}} = 292\text{ g } CO_3^{2-}$$

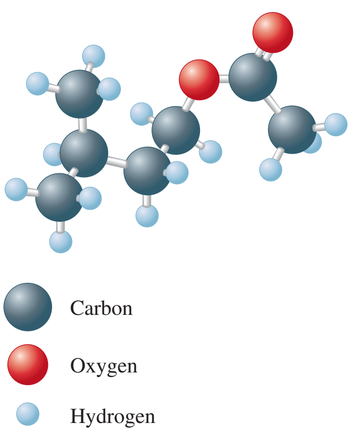
See Exercises 3.43 through 3.46.

Sample Exercise 3.8

Molar Mass and Numbers of Molecules

Isopentyl acetate ($C_7H_{14}O_2$) is the compound responsible for the scent of bananas. A molecular model of isopentyl acetate is shown in the margin below. Interestingly, bees release about 1 μg (1×10^{-6} g) of this compound when they sting. The resulting scent attracts other bees to join the attack. How many molecules of isopentyl acetate are released in a typical bee sting? How many atoms of carbon are present?

Isopentyl acetate is released when a bee stings.



Isopentyl acetate

Solution

Since we are given a mass of isopentyl acetate and want to find the number of molecules, we must first compute the molar mass:

$$7\ \cancel{\text{mol}}\text{ C} \times 12.01\frac{\text{g}}{\cancel{\text{mol}}} = 84.07\text{ g C}$$

$$14\ \cancel{\text{mol}}\text{ H} \times 1.008\frac{\text{g}}{\cancel{\text{mol}}} = 14.11\text{ g H}$$

$$2\ \cancel{\text{mol}}\text{ O} \times 16.00\frac{\text{g}}{\cancel{\text{mol}}} = \frac{32.00\text{ g O}}{130.18\text{ g}}$$

This means that 1 mole of isopentyl acetate (6.022×10^{23} molecules) has a mass of 130.18 g.

To find the number of molecules released in a sting, we must first determine the number of moles of isopentyl acetate in 1×10^{-6} g:

$$1 \times 10^{-6}\ \cancel{\text{g } C_7H_{14}O_2} \times \frac{1\text{ mol } C_7H_{14}O_2}{130.18\ \cancel{\text{g } C_7H_{14}O_2}} = 8 \times 10^{-9}\text{ mol } C_7H_{14}O_2$$

Since 1 mole is 6.022×10^{23} units, we can determine the number of molecules:

$$8 \times 10^{-9}\ \cancel{\text{mol } C_7H_{14}O_2} \times \frac{6.022 \times 10^{23}\text{ molecules}}{1\ \cancel{\text{mol } C_7H_{14}O_2}} = 5 \times 10^{15}\text{ molecules}$$

To determine the number of carbon atoms present, we must multiply the number of molecules by 7, since each molecule of isopentyl acetate contains seven carbon atoms:

$$5 \times 10^{15}\ \cancel{\text{molecules}} \times \frac{7\text{ carbon atoms}}{\cancel{\text{molecule}}} = 4 \times 10^{16}\text{ carbon atoms}$$

To show the correct number of significant figures in each calculation, we round after each step. In your calculations, always carry extra significant figures through to the end; then round.

Note: In keeping with our practice of always showing the correct number of significant figures, we have rounded after each step. However, if extra digits are carried throughout this problem, the final answer rounds to 3×10^{16}.

See Exercises 3.47 through 3.52.

3.5 Percent Composition of Compounds

There are two common ways of describing the composition of a compound: in terms of the numbers of its constituent atoms and in terms of the percentages (by mass) of its elements. We can obtain the mass percents of the elements from the formula of the compound by comparing the mass of each element present in 1 mole of the compound to the total mass of 1 mole of the compound.

For example, for ethanol, which has the formula C_2H_5OH, the mass of each element present and the molar mass are obtained as follows:

$$\text{Mass of C} = 2\ \cancel{\text{mol}} \times 12.01\frac{\text{g}}{\cancel{\text{mol}}} = 24.02\ \text{g}$$

$$\text{Mass of H} = 6\ \cancel{\text{mol}} \times 1.008\frac{\text{g}}{\cancel{\text{mol}}} = 6.048\ \text{g}$$

$$\text{Mass of O} = 1\ \cancel{\text{mol}} \times 16.00\frac{\text{g}}{\cancel{\text{mol}}} = \underline{16.00\ \text{g}}$$

$$\text{Mass of 1 mol } C_2H_5OH = 46.07\ \text{g}$$

The **mass percent** (often called the *weight percent*) of carbon in ethanol can be computed by comparing the mass of carbon in 1 mole of ethanol to the total mass of 1 mole of ethanol and multiplying the result by 100:

$$\text{Mass percent of C} = \frac{\text{mass of C in 1 mol } C_2H_5OH}{\text{mass of 1 mol } C_2H_5OH} \times 100\%$$

$$= \frac{24.02\ \text{g}}{46.07\ \text{g}} \times 100\% = 52.14\%$$

The mass percents of hydrogen and oxygen in ethanol are obtained in a similar manner:

$$\text{Mass percent of H} = \frac{\text{mass of H in 1 mol } C_2H_5OH}{\text{mass of 1 mol } C_2H_5OH} \times 100\%$$

$$= \frac{6.048\ \text{g}}{46.07\ \text{g}} \times 100\% = 13.13\%$$

$$\text{Mass percent of O} = \frac{\text{mass of O in 1 mol } C_2H_5OH}{\text{mass of 1 mol } C_2H_5OH} \times 100\%$$

$$= \frac{16.00\ \text{g}}{46.07\ \text{g}} \times 100\% = 34.73\%$$

Reality Check: Notice that the percentages add up to 100.00%; this provides a check that the calculations are correct.

Sample Exercise 3.9

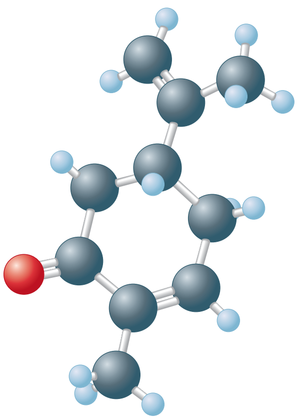
Calculating Mass Percent I

Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula ($C_{10}H_{14}O$) and mass. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone.

Solution

The masses of the elements in 1 mole of carvone are

$$\text{Mass of C in 1 mol} = 10\ \cancel{\text{mol}} \times 12.01\frac{\text{g}}{\cancel{\text{mol}}} = 120.1\ \text{g}$$



Carvone

$$\text{Mass of H in 1 mol} = 14\ \cancel{\text{mol}} \times 1.008\frac{\text{g}}{\cancel{\text{mol}}} = 14.11\ \text{g}$$

$$\text{Mass of O in 1 mol} = 1\ \cancel{\text{mol}} \times 16.00\frac{\text{g}}{\cancel{\text{mol}}} = \underline{16.00\ \text{g}}$$

$$\text{Mass of 1 mol } C_{10}H_{14}O = 150.2\ \text{g}$$

Next we find the fraction of the total mass contributed by each element and convert it to a percentage:

$$\text{Mass percent of C} = \frac{120.1\ \text{g C}}{150.2\ \text{g } C_{10}H_{14}O} \times 100\% = 79.96\%$$

$$\text{Mass percent of H} = \frac{14.11\ \text{g H}}{150.2\ \text{g } C_{10}H_{14}O} \times 100\% = 9.394\%$$

$$\text{Mass percent of O} = \frac{16.00\ \text{g O}}{150.2\ \text{g } C_{10}H_{14}O} \times 100\% = 10.65\%$$

Reality Check: Sum the individual mass percent values—they should total to 100% within round-off errors. In this case, the percentages add up to 100.00%.

See Exercises 3.59 and 3.60.

Sample Exercise 3.10

Calculating Mass Percent II

Although Fleming is commonly given credit for the discovery of penicillin, there is good evidence that penicillium mold extracts were used in the nineteenth century by Lord Joseph Lister to cure infections.

Penicillin, the first of a now large number of antibiotics (antibacterial agents), was discovered accidentally by the Scottish bacteriologist Alexander Fleming in 1928, but he was never able to isolate it as a pure compound. This and similar antibiotics have saved millions of lives that might have been lost to infections. Penicillin F has the formula $C_{14}H_{20}N_2SO_4$. Compute the mass percent of each element.

Solution

The molar mass of penicillin F is computed as follows:

$$\text{C:}\quad 14\ \cancel{\text{mol}} \times 12.01\frac{\text{g}}{\cancel{\text{mol}}} = 168.1\ \text{g}$$

$$\text{H:}\quad 20\ \cancel{\text{mol}} \times 1.008\frac{\text{g}}{\cancel{\text{mol}}} = 120.16\ \text{g}$$

$$\text{N:}\quad 2\ \cancel{\text{mol}} \times 14.01\frac{\text{g}}{\cancel{\text{mol}}} = 28.02\ \text{g}$$

$$\text{S:}\quad 1\ \cancel{\text{mol}} \times 32.07\frac{\text{g}}{\cancel{\text{mol}}} = 32.07\ \text{g}$$

$$\text{O:}\quad 4\ \cancel{\text{mol}} \times 16.00\frac{\text{g}}{\cancel{\text{mol}}} = \underline{64.00\ \text{g}}$$

$$\text{Mass of 1 mol } C_{14}H_{20}N_2SO_4 = 312.4\ g$$

$$\text{Mass percent of C} = \frac{168.1\ \text{g C}}{312.4\ \text{g } C_{14}H_{20}N_2SO_4} \times 100\% = 53.81\%$$

$$\text{Mass percent of H} = \frac{20.16\ \text{g H}}{312.4\ \text{g } C_{14}H_{20}N_2SO_4} \times 100\% = 6.453\%$$

$$\text{Mass percent of N} = \frac{28.02\ \text{g N}}{312.4\ \text{g } C_{14}H_{20}N_2SO_4} \times 100\% = 8.969\%$$

$$\text{Mass percent of S} = \frac{32.07\ \text{g S}}{312.4\ \text{g } C_{14}H_{20}N_2SO_4} \times 100\% = 10.27\%$$

Penicillin is isolated from a mold that can be grown in large quantities in fermentation tanks.

$$\text{Mass percent of O} = \frac{64.00 \text{ g O}}{312.4 \text{ g } C_{14}H_{20}N_2SO_4} \times 100\% = 20.49\%$$

Reality Check: The percentages add up to 99.99%.

See Exercises 3.61 and 3.62.

3.6 Determining the Formula of a Compound

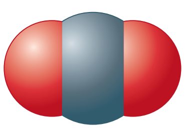
When a new compound is prepared, one of the first items of interest is the formula of the compound. This is most often determined by taking a weighed sample of the compound and either decomposing it into its component elements or reacting it with oxygen to produce substances such as CO_2, H_2O, and N_2, which are then collected and weighed. A device for doing this type of analysis is shown in Fig. 3.5. The results of such analyses provide the mass of each type of element in the compound, which can be used to determine the mass percent of each element.

We will see how information of this type can be used to compute the formula of a compound. Suppose a substance has been prepared that is composed of carbon, hydrogen, and nitrogen. When 0.1156 gram of this compound is reacted with oxygen, 0.1638 gram of carbon dioxide (CO_2) and 0.1676 gram of water (H_2O) are collected. Assuming that all the carbon in the compound is converted to CO_2, we can determine the mass of carbon originally present in the 0.1156-gram sample. To do this, we must use the fraction (by mass) of carbon in CO_2. The molar mass of CO_2 is

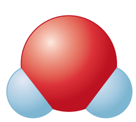
$$\text{C:} \quad 1 \cancel{\text{mol}} \times 12.01 \frac{\text{g}}{\cancel{\text{mol}}} = 12.01 \text{ g}$$

$$\text{O:} \quad 2 \cancel{\text{mol}} \times 16.00 \frac{\text{g}}{\cancel{\text{mol}}} = \underline{32.00 \text{ g}}$$

$$\text{Molar mass of } CO_2 = 44.01 \text{ g/mol}$$



CO_2



H_2O

The fraction of carbon present by mass is

$$\frac{\text{Mass of C}}{\text{Total mass of } CO_2} = \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2}$$

This factor can now be used to determine the mass of carbon in 0.1638 gram of CO_2:

$$0.1638 \cancel{\text{g } CO_2} \times \frac{12.01 \text{ g C}}{44.01 \cancel{\text{g } CO_2}} = 0.04470 \text{ g C}$$

Remember that this carbon originally came from the 0.1156-gram sample of unknown compound. Thus the mass percent of carbon in this compound is

$$\frac{0.04470 \text{ g C}}{0.1156 \text{ g compound}} \times 100\% = 38.67\% \text{ C}$$

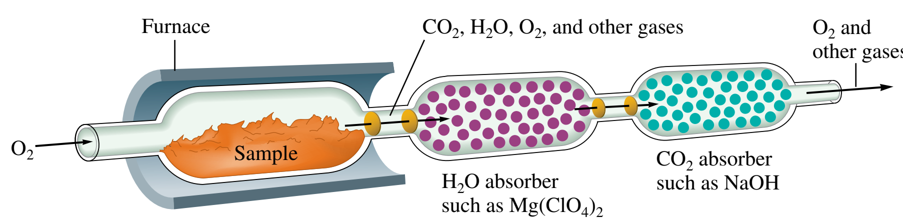


FIGURE 3.5
A schematic diagram of the combustion device used to analyze substances for carbon and hydrogen. The sample is burned in the presence of excess oxygen, which converts all its carbon to carbon dioxide and all its hydrogen to water. These products are collected by absorption using appropriate materials, and their amounts are determined by measuring the increase in masses of the absorbents.

The same procedure can be used to find the mass percent of hydrogen in the unknown compound. We assume that all the hydrogen present in the original 0.1156 gram of compound was converted to H_2O. The molar mass of H_2O is 18.02 grams, and the fraction of hydrogen by mass in H_2O is

$$\frac{\text{Mass of H}}{\text{Mass of H}_2\text{O}} = \frac{2.016 \text{ g H}}{18.02 \text{ g H}_2\text{O}}$$

Therefore, the mass of hydrogen in 0.1676 gram of H_2O is

$$0.1676 \cancel{\text{g H}_2\text{O}} \times \frac{2.016 \text{ g H}}{18.02 \cancel{\text{g H}_2\text{O}}} = 0.01875 \text{ g H}$$

The mass percent of hydrogen in the compound is

$$\frac{0.01875 \text{ g H}}{0.1156 \text{ g compound}} \times 100\% = 16.22\% \text{ H}$$

The unknown compound contains only carbon, hydrogen, and nitrogen. So far we have determined that it is 38.67% carbon and 16.22% hydrogen. The remainder must be nitrogen:

$$100.00\% - (\underset{\%\text{ C}}{\underset{\uparrow}{38.67\%}} + \underset{\%\text{ H}}{\underset{\uparrow}{16.22\%}}) = 45.11\% \text{ N}$$

We have determined that the compound contains 38.67% carbon, 16.22% hydrogen, and 45.11% nitrogen. Next we use these data to obtain the formula.

Since the formula of a compound indicates the *numbers* of atoms in the compound, we must convert the masses of the elements to numbers of atoms. The easiest way to do this is to work with 100.00 grams of the compound. In the present case, 38.67% carbon by mass means 38.67 grams of carbon per 100.00 grams of compound; 16.22% hydrogen means 16.22 grams of hydrogen per 100.00 grams of compound; and so on. To determine the formula, we must calculate the number of carbon atoms in 38.67 grams of carbon, the number of hydrogen atoms in 16.22 grams of hydrogen, and the number of nitrogen atoms in 45.11 grams of nitrogen. We can do this as follows:

$$38.67 \cancel{\text{g C}} \times \frac{1 \text{ mol C}}{12.01 \cancel{\text{g C}}} = 3.220 \text{ mol C}$$

$$16.22 \cancel{\text{g H}} \times \frac{1 \text{ mol H}}{1.008 \cancel{\text{g H}}} = 16.09 \text{ mol H}$$

$$45.11 \cancel{\text{g N}} \times \frac{1 \text{ mol N}}{14.01 \cancel{\text{g N}}} = 3.219 \text{ mol N}$$

Thus 100.00 grams of this compound contains 3.220 moles of carbon atoms, 16.09 moles of hydrogen atoms, and 3.219 moles of nitrogen atoms.

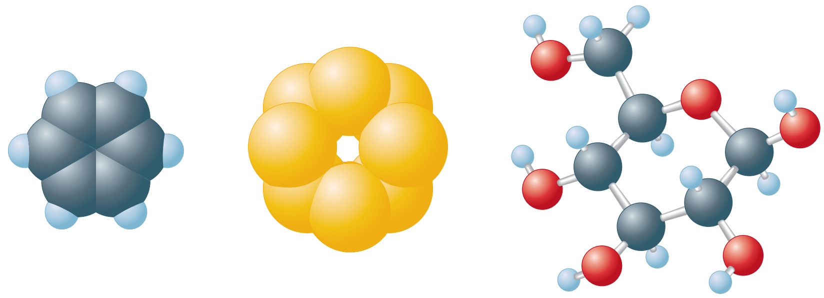
We can find the smallest *whole-number ratio* of atoms in this compound by dividing each of the mole values above by the smallest of the three:

$$\text{C:} \quad \frac{3.220}{3.220} = 1.000 = 1$$

$$\text{H:} \quad \frac{16.09}{3.220} = 4.997 = 5$$

$$\text{N:} \quad \frac{3.219}{3.220} = 1.000 = 1$$

Thus the formula might well be CH_5N. However, it also could be $C_2H_{10}N_2$ or $C_3H_{15}N_3$, and so on—that is, some multiple of the smallest whole-number ratio. Each of these alternatives also has the correct relative numbers of atoms. That is, any molecule that can



$C_6H_6 = (CH)_6$ $S_8 = (S)_8$ $C_6H_{12}O_6 = (CH_2O)_6$

FIGURE 3.6
Examples of substances whose empirical and molecular formulas differ. Notice that molecular formula = (empirical formula)$_n$, where n is an integer.

Molecular formula = (empirical formula)$_n$, where n is an integer.

be represented as $(CH_5N)_n$, where n is an integer, has the **empirical formula** CH_5N. To be able to specify the exact formula of the molecule involved, the **molecular formula,** we must know the molar mass.

Suppose we know that this compound with empirical formula CH_5N has a molar mass of 31.06 g/mol. How do we determine which of the possible choices represents the molecular formula? Since the molecular formula is always a whole number multiple of the empirical formula, we must first find the empirical formula mass for CH_5N:

$$\begin{aligned}
1\ \text{C:}\quad & 1 \times 12.01\ \text{g} = 12.01\ \text{g} \\
5\ \text{H:}\quad & 5 \times 1.008\ \text{g} = 5.040\ \text{g} \\
1\ \text{N:}\quad & 1 \times 14.01\ \text{g} = \underline{14.01\ \text{g}} \\
\text{Formula mass of } CH_5N & = 31.06\ \text{g/mol}
\end{aligned}$$

This is the same as the known molar mass of the compound. Thus in this case the empirical formula and the molecular formula are the same; this substance consists of molecules with the formula CH_5N. It is quite common for the empirical and molecular formulas to be different; some examples where this is the case are shown in Fig. 3.6.

Sample Exercise 3.11

Determining Empirical and Molecular Formulas I

Determine the empirical and molecular formulas for a compound that gives the following percentages upon analysis (in mass percents):

71.65% Cl 24.27% C 4.07% H

The molar mass is known to be 98.96 g/mol.

Solution

First, we convert the mass percents to masses in grams. In 100.00 g of this compound there are 71.65 g of chlorine, 24.27 g of carbon, and 4.07 g of hydrogen. We use these masses to compute the moles of atoms present:

$$71.65\ \cancel{\text{g Cl}} \times \frac{1\ \text{mol Cl}}{35.45\ \cancel{\text{g Cl}}} = 2.021\ \text{mol Cl}$$

$$24.27\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 2.021\ \text{mol C}$$

$$4.07\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 4.04\ \text{mol H}$$

Dividing each mole value by 2.021 (the smallest number of moles present), we obtain the empirical formula $ClCH_2$.

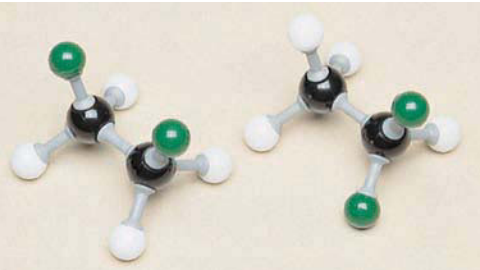


FIGURE 3.7
The two forms of dichloroethane.

To determine the molecular formula, we must compare the empirical formula mass with the molar mass. The empirical formula mass is 49.48 g/mol (confirm this). The molar mass is known to be 98.96 g/mol.

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96 \text{ g/mol}}{49.48 \text{ g/mol}} = 2$$

$$\text{Molecular formula} = (ClCH_2)_2 = Cl_2C_2H_4$$

This substance is composed of molecules with the formula $Cl_2C_2H_4$.

Notice that the method we employ here allows us to determine the molecular formula of a compound but not its structural formula. The compound $Cl_2C_2H_4$ is called *dichloroethane*. There are two forms of this compound, shown in Fig. 3.7. The form on the right was formerly used as an additive in leaded gasoline.

See Exercises 3.57 and 3.58.

Sample Exercise 3.12 Determining Empirical and Molecular Formulas II

A white powder is analyzed and found to contain 43.64% phosphorus and 56.36% oxygen by mass. The compound has a molar mass of 283.88 g/mol. What are the compound's empirical and molecular formulas?

Solution

In 100.00 g of this compound there are 43.64 g of phosphorus and 56.36 g of oxygen. In terms of moles, in 100.00 g of the compound we have

$$43.64 \cancel{\text{g P}} \times \frac{1 \text{ mol P}}{30.97 \cancel{\text{g P}}} = 1.409 \text{ mol P}$$

$$56.36 \cancel{\text{g O}} \times \frac{1 \text{ mol O}}{16.00 \cancel{\text{g O}}} = 3.523 \text{ mol O}$$

Dividing both mole values by the smaller one gives

$$\frac{1.409}{1.409} = 1 \text{ P} \quad \text{and} \quad \frac{3.523}{1.409} = 2.5 \text{ O}$$

This yields the formula $PO_{2.5}$. Since compounds must contain whole numbers of atoms, the empirical formula should contain only whole numbers. To obtain the simplest set of whole numbers, we multiply both numbers by 2 to give the empirical formula P_2O_5.

To obtain the molecular formula, we must compare the empirical formula mass to the molar mass. The empirical formula mass for P_2O_5 is 141.94.

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{283.88}{141.94} = 2$$

The molecular formula is $(P_2O_5)_2$, or P_4O_{10}.

The structural formula of this interesting compound is given in Fig. 3.8.

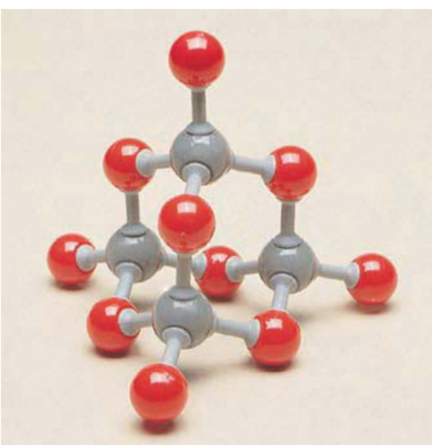


FIGURE 3.8
The structure of P_4O_{10}. Note that some of the oxygen atoms act as "bridges" between the phosphorus atoms. This compound has a great affinity for water and is often used as a desiccant, or drying agent.

See Exercise 3.59.

In Sample Exercises 3.11 and 3.12 we found the molecular formula by comparing the empirical formula mass with the molar mass. There is an alternate way to obtain the molecular formula. For example, in Sample Exercise 3.11 we know the molar mass of the compound is 98.96 g/mol. This means that 1 mole of the compound weighs 98.96 grams.

Since we also know the mass percentages of each element, we can compute the mass of each element present in 1 mole of compound:

$$\text{Chlorine: } \frac{71.65 \text{ g Cl}}{100.0 \cancel{\text{g}} \text{ compound}} \times \frac{98.96 \cancel{\text{g}}}{\text{mol}} = \frac{70.90 \text{ g Cl}}{\text{mol compound}}$$

$$\text{Carbon: } \frac{24.27 \text{ g C}}{100.0 \cancel{\text{g}} \text{ compound}} \times \frac{98.96 \cancel{\text{g}}}{\text{mol}} = \frac{24.02 \text{ g C}}{\text{mol compound}}$$

$$\text{Hydrogen: } \frac{4.07 \text{ g H}}{100.0 \cancel{\text{g}} \text{ compound}} \times \frac{98.96 \cancel{\text{g}}}{\text{mol}} = \frac{4.03 \text{ g H}}{\text{mol compound}}$$

Now we can compute moles of atoms present per mole of compound:

$$\text{Chlorine: } \frac{70.90 \cancel{\text{g Cl}}}{\text{mol compound}} \times \frac{1 \text{ mol Cl}}{35.45 \cancel{\text{g Cl}}} = \frac{2.000 \text{ mol Cl}}{\text{mol compound}}$$

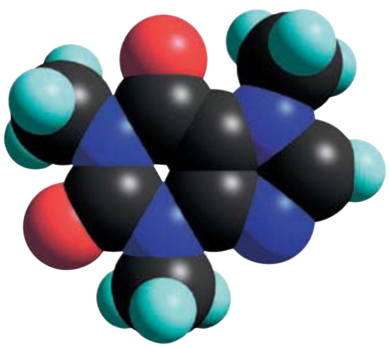
$$\text{Carbon: } \frac{24.02 \cancel{\text{g C}}}{\text{mol compound}} \times \frac{1 \text{ mol C}}{12.01 \cancel{\text{g C}}} = \frac{2.000 \text{ mol C}}{\text{mol compound}}$$

$$\text{Hydrogen: } \frac{4.03 \cancel{\text{g H}}}{\text{mol compound}} \times \frac{1 \text{ mol H}}{1.008 \cancel{\text{g H}}} = \frac{4.00 \text{ mol H}}{\text{mol compound}}$$

Thus 1 mole of the compound contains 2 mol Cl atoms, 2 mol C atoms, and 4 mol H atoms, and the molecular formula is $Cl_2C_2H_4$, as obtained in Sample Exercise 3.11.

Sample Exercise 3.13 Determining a Molecular Formula

Caffeine, a stimulant found in coffee, tea, and chocolate, contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen, and 16.49% oxygen by mass and has a molar mass of 194.2 g/mol. Determine the molecular formula of caffeine.



Computer-generated molecule of caffeine.

Solution

We will first determine the mass of each element in 1 mole (194.2 g) of caffeine:

$$\frac{49.48 \text{ g C}}{100.0 \cancel{\text{g}} \text{ caffeine}} \times \frac{194.2 \cancel{\text{g}}}{\text{mol}} = \frac{96.09 \text{ g C}}{\text{mol caffeine}}$$

$$\frac{5.15 \text{ g H}}{100.0 \cancel{\text{g}} \text{ caffeine}} \times \frac{194.2 \cancel{\text{g}}}{\text{mol}} = \frac{10.0 \text{ g H}}{\text{mol caffeine}}$$

$$\frac{28.87 \text{ g N}}{100.0 \cancel{\text{g}} \text{ caffeine}} \times \frac{194.2 \cancel{\text{g}}}{\text{mol}} = \frac{56.07 \text{ g N}}{\text{mol caffeine}}$$

$$\frac{16.49 \text{ g O}}{100.0 \cancel{\text{g}} \text{ caffeine}} \times \frac{194.2 \cancel{\text{g}}}{\text{mol}} = \frac{32.02 \text{ g O}}{\text{mol caffeine}}$$

Now we will convert to moles:

$$\text{Carbon: } \frac{96.09 \cancel{\text{g C}}}{\text{mol caffeine}} \times \frac{1 \text{ mol C}}{12.01 \cancel{\text{g C}}} = \frac{8.001 \text{ mol C}}{\text{mol caffeine}}$$

$$\text{Hydrogen: } \frac{10.0 \cancel{\text{g H}}}{\text{mol caffeine}} \times \frac{1 \text{ mol H}}{1.008 \cancel{\text{g H}}} = \frac{9.92 \text{ mol H}}{\text{mol caffeine}}$$

$$\text{Nitrogen: } \frac{56.07 \cancel{\text{g N}}}{\text{mol caffeine}} \times \frac{1 \text{ mol N}}{14.01 \cancel{\text{g N}}} = \frac{4.002 \text{ mol N}}{\text{mol caffeine}}$$

$$\text{Oxygen: } \frac{32.02 \cancel{\text{g O}}}{\text{mol caffeine}} \times \frac{1 \text{ mol O}}{16.00 \cancel{\text{g O}}} = \frac{2.001 \text{ mol O}}{\text{mol caffeine}}$$

Rounding the numbers to integers gives the molecular formula for caffeine: $C_8H_{10}N_4O_2$.

See Exercise 3.76.

The methods for obtaining empirical and molecular formulas are summarized as follows:

Empirical Formula Determination

- Since mass percentage gives the number of grams of a particular element per 100 grams of compound, base the calculation on 100 grams of compound. Each percent will then represent the mass in grams of that element.
- Determine the number of moles of each element present in 100 grams of compound using the atomic masses of the elements present.
- Divide each value of the number of moles by the smallest of the values. If each resulting number is a whole number (after appropriate rounding), these numbers represent the subscripts of the elements in the empirical formula.
- If the numbers obtained in the previous step are not whole numbers, multiply each number by an integer so that the results are all whole numbers.

Numbers very close to whole numbers, such as 9.92 and 1.08, should be rounded to whole numbers. Numbers such as 2.25, 4.33, and 2.72 should not be rounded to whole numbers.

Molecular Formula Determination

Method One

- Obtain the empirical formula.
- Compute the mass corresponding to the empirical formula.
- Calculate the ratio

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}}$$

- The integer from the previous step represents the number of empirical formula units in one molecule. When the empirical formula subscripts are multiplied by this integer, the molecular formula results. This procedure is summarized by the equation:

$$\text{Molecular formula} = (\text{empirical formula}) \times \frac{\text{molar mass}}{\text{empirical formula mass}}$$

Method Two

- Using the mass percentages and the molar mass, determine the mass of each element present in one mole of compound.
- Determine the number of moles of each element present in one mole of compound.
- The integers from the previous step represent the subscripts in the molecular formula.

Note that method two assumes that the molar mass of the compound is known accurately.

3.7 Chemical Equations

Chemical Reactions

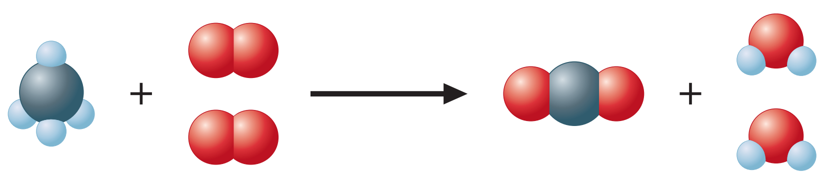
A chemical change involves a reorganization of the atoms in one or more substances. For example, when the methane (CH_4) in natural gas combines with oxygen (O_2) in the air and burns, carbon dioxide (CO_2) and water (H_2O) are formed. This process is represented by a **chemical equation** with the **reactants** (here methane and oxygen) on the left side of an arrow and the **products** (carbon dioxide and water) on the right side:

$$\underbrace{CH_4 + O_2}_{\text{Reactants}} \longrightarrow \underbrace{CO_2 + H_2O}_{\text{Products}}$$

Notice that the atoms have been reorganized. *Bonds have been broken, and new ones have been formed.* It is important to recognize that *in a chemical reaction, atoms are neither created nor destroyed. All atoms present in the reactants must be accounted for among the products.* In other words, there must be the same number of each type of atom on the product side and on the reactant side of the arrow. Making sure that this rule is obeyed is called **balancing a chemical equation** for a reaction.

The equation (shown above) for the reaction between CH_4 and O_2 is not balanced. We can see this from the following representation of the reaction:

Notice that the number of oxygen atoms (in O_2) on the left of the arrow is two, while on the right there are three O atoms (in CO_2 and H_2O). Also, there are four hydrogen atoms (in CH_4) on the left and only two (in H_2O) on the right. Remember that a chemical reaction is simply a rearrangement of the atoms (a change in the way they are organized). Atoms are not created or destroyed in a chemical reaction. Thus the reactants and products must occur in numbers that give the same number of each type of atom among both the reactants and products. Simple trial and error will allow us to figure this out for the reaction of methane with oxygen. The needed numbers of molecules are



Methane reacts with oxygen to produce the flame in a Bunsen burner.

Notice that now we have the same number of each type of atom represented among the reactants and the products.

We can represent the preceding situation in a shorthand manner by the following chemical equation:

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

We can check that the equation is balanced by comparing the number of each type of atom on both sides:

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

1 C   4 H   4 O   1 C   2 O   4 H   2 O

To summarize, we have

Reactants	Products
1 C	1 C
4 H	4 H
4 O	4 O

TABLE 3.2 Information Conveyed by the Balanced Equation for the Combustion of Methane

Reactants		Products
$CH_4(g) + 2O_2(g)$	$\longrightarrow$	$CO_2(g) + 2H_2O(g)$
1 molecule + 2 molecules	$\longrightarrow$	1 molecule + 2 molecules
1 mole + 2 moles	$\longrightarrow$	1 mole + 2 moles
6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)	$\longrightarrow$	6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)
16 g + 2 (32 g)		44 g + 2 (18 g)
80 g reactants	$\longrightarrow$	80 g products

The Meaning of a Chemical Equation

Visualization: Oxygen, Hydrogen, Soap Bubbles, and Balloons

The chemical equation for a reaction gives two important types of information: the nature of the reactants and products and the relative numbers of each.

The reactants and products in a specific reaction must be identified by experiment. Besides specifying the compounds involved in the reaction, the equation often gives the *physical states* of the reactants and products:

State	Symbol
Solid	(s)
Liquid	(l)
Gas	(g)
Dissolved in water (in aqueous solution)	(aq)

Hydrochloric acid reacts with solid sodium hydrogen carbonate to produce gaseous carbon dioxide.

For example, when hydrochloric acid in aqueous solution is added to solid sodium hydrogen carbonate, the products carbon dioxide gas, liquid water, and sodium chloride (which dissolves in the water) are formed:

$$HCl(aq) + NaHCO_3(s) \longrightarrow CO_2(g) + H_2O(l) + NaCl(aq)$$

The relative numbers of reactants and products in a reaction are indicated by the *coefficients* in the balanced equation. (The coefficients can be determined because we know that the same number of each type of atom must occur on both sides of the equation.) For example, the balanced equation

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

can be interpreted in several equivalent ways, as shown in Table 3.2. Note that the total mass is 80 grams for both reactants and products. We expect the mass to remain constant, since chemical reactions involve only a rearrangement of atoms. Atoms, and therefore mass, are conserved in a chemical reaction.

From this discussion you can see that a balanced chemical equation gives you a great deal of information.

3.8 Balancing Chemical Equations

Visualization: Conservation of Mass and Balancing Equations

An unbalanced chemical equation is of limited use. Whenever you see an equation, you should ask yourself whether it is balanced. The principle that lies at the heart of the balancing process is that atoms are conserved in a chemical reaction. The same number of each type of atom must be found among the reactants and products. It is also important to recognize that the identities of the reactants and products of a reaction are determined by experimental observation. For example, when liquid ethanol is burned in the presence of sufficient oxygen gas, the products are always carbon dioxide and water. When the equation for this reaction is balanced, the *identities* of the reactants and products must not be changed. *The formulas of the compounds must never be changed in balancing a chemical equation.* That is, the subscripts in a formula cannot be changed, nor can atoms be added or subtracted from a formula.

In balancing equations, start with the most complicated molecule.

Most chemical equations can be balanced by inspection, that is, by trial and error. It is always best to start with the most complicated molecules (those containing the greatest number of atoms). For example, consider the reaction of ethanol with oxygen, given by the unbalanced equation

$$C_2H_5OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

which can be represented by the following molecular models:

Notice that the carbon and hydrogen atoms are not balanced. There are two carbon atoms on the left and one on the right, and there are six hydrogens on the left and two on the right. We need to find the correct numbers of reactants and products so that we have the same number of all types of atoms among the reactants and products. We will balance the equation "by inspection" (a systematic trial-and-error procedure).

The most complicated molecule here is C_2H_5OH. We will begin by balancing the products that contain the atoms in C_2H_5OH. Since C_2H_5OH contains two carbon atoms, we place the coefficient 2 before the CO_2 to balance the carbon atoms:

$$\underset{\text{2 C atoms}}{C_2H_5OH(l)} + O_2(g) \rightarrow \underset{\text{2 C atoms}}{2CO_2(g)} + H_2O(g)$$

Since C_2H_5OH contains six hydrogen atoms, the hydrogen atoms can be balanced by placing a 3 before the H_2O:

$$\underset{(5+1)\text{ H}}{C_2H_5OH(l)} + O_2(g) \rightarrow 2CO_2(g) + \underset{(3\times 2)\text{ H}}{3H_2O(g)}$$

Last, we balance the oxygen atoms. Note that the right side of the preceding equation contains seven oxygen atoms, whereas the left side has only three. We can correct this by putting a 3 before the O_2 to produce the balanced equation:

$$\underbrace{\underset{1\text{ O}}{C_2H_5OH(l)} + \underset{6\text{ O}}{3O_2(g)}}_{7\text{ O}} \rightarrow \underbrace{\underset{(2\times 2)\text{ O}}{2CO_2(g)} + \underset{3\text{ O}}{3H_2O(g)}}_{7\text{ O}}$$

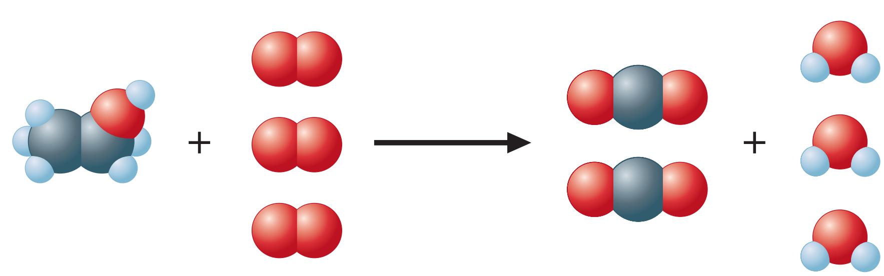
Now we check:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

2 C atoms	2 C atoms
6 H atoms	6 H atoms
7 O atoms	7 O atoms

The equation is balanced.

The balanced equation can be represented as follows:



You can see that all the elements balance.

Writing and Balancing the Equation for a Chemical Reaction

1 Determine what reaction is occurring. What are the reactants, the products, and the physical states involved?

2 Write the *unbalanced* equation that summarizes the reaction described in step 1.

3 Balance the equation by inspection, starting with the most complicated molecule(s). Determine what coefficients are necessary so that the same number of each type of atom appears on both reactant and product sides. Do not change the identities (formulas) of any of the reactants or products.

Sample Exercise 3.14 Balancing a Chemical Equation I

Chromate and dichromate compounds are carcinogens (cancer-inducing agents) and should be handled very carefully.

Chromium compounds exhibit a variety of bright colors. When solid ammonium dichromate, $(NH_4)_2Cr_2O_7$, a vivid orange compound, is ignited, a spectacular reaction occurs, as shown in the two photographs on the next page. Although the reaction is actually somewhat more complex, let's assume here that the products are solid chromium(III) oxide, nitrogen gas (consisting of N_2 molecules), and water vapor. Balance the equation for this reaction.

Solution

1 From the description given, the reactant is solid ammonium dichromate, $(NH_4)_2Cr_2O_7(s)$, and the products are nitrogen gas, $N_2(g)$, water vapor, $H_2O(g)$, and solid chromium(III) oxide, $Cr_2O_3(s)$. The formula for chromium(III) oxide can be determined by recognizing that the Roman numeral III means that Cr^{3+} ions are present. For a neutral compound, the formula must then be Cr_2O_3, since each oxide ion is O^{2-}.

2 The unbalanced equation is

$$(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + H_2O(g)$$

3 Note that nitrogen and chromium are balanced (two nitrogen atoms and two chromium atoms on each side), but hydrogen and oxygen are not. A coefficient of 4 for H_2O balances the hydrogen atoms:

$$\underbrace{(NH_4)_2}_{(4 \times 2)\ H}Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + \underbrace{4H_2O(g)}_{(4 \times 2)\ H}$$

Note that in balancing the hydrogen we also have balanced the oxygen, since there are seven oxygen atoms in the reactants and in the products.

Reality Check:

$$\underbrace{2\ N,\ 8\ H,\ 2\ Cr,\ 7\ O}_{\text{Reactant atoms}} \rightarrow \underbrace{2\ N,\ 8\ H,\ 2\ Cr,\ 7\ O}_{\text{Product atoms}}$$

The equation is balanced.

See Exercises 3.81 and 3.82.


Decomposition of ammonium dichromate.

Sample Exercise 3.15

Balancing a Chemical Equation II

The Ostwald process is described in Section 20.2.

At 1000°C, ammonia gas, $NH_3(g)$, reacts with oxygen gas to form gaseous nitric oxide, $NO(g)$, and water vapor. This reaction is the first step in the commercial production of nitric acid by the Ostwald process. Balance the equation for this reaction.

Solution

The unbalanced equation for the reaction is

$$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$$

Because all the molecules in this equation are of about equal complexity, where we start in balancing it is rather arbitrary. Let's begin by balancing the hydrogen. A coefficient of 2 for NH_3 and a coefficient of 3 for H_2O give six atoms of hydrogen on both sides:

$$2NH_3(g) + O_2(g) \rightarrow NO(g) + 3H_2O(g)$$

The nitrogen can be balanced with a coefficient of 2 for NO:

$$2NH_3(g) + O_2(g) \rightarrow 2NO(g) + 3H_2O(g)$$

Finally, note that there are two atoms of oxygen on the left and five on the right. The oxygen can be balanced with a coefficient of $\frac{5}{2}$ for O_2:

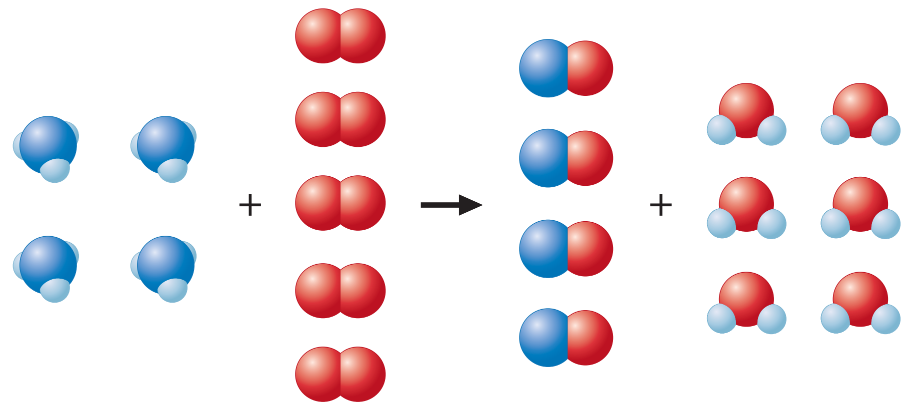
$$2NH_3(g) + \frac{5}{2}O_2(g) \rightarrow 2NO(g) + 3H_2O(g)$$

However, the usual custom is to have whole-number coefficients. We simply multiply the entire equation by 2.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

Reality Check: There are 4 N, 12 H, and 10 O on both sides, so the equation is balanced.

We can represent this balanced equation visually as



See Exercises 3.83 through 3.88.

3.9 Stoichiometric Calculations: Amounts of Reactants and Products

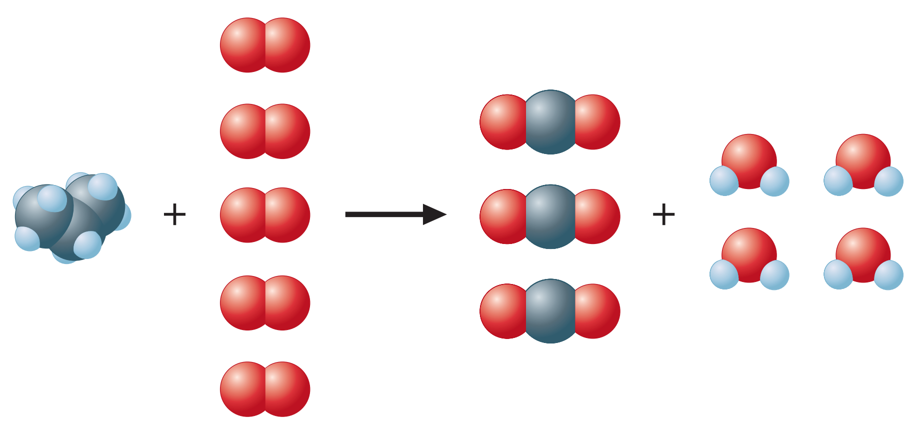
As we have seen in previous sections of this chapter, the coefficients in chemical equations represent *numbers* of molecules, not masses of molecules. However, when a reaction is to be run in a laboratory or chemical plant, the amounts of substances needed cannot be determined by counting molecules directly. Counting is always done by weighing. In this section we will see how chemical equations can be used to determine the *masses* of reacting chemicals.

To develop the principles for dealing with the stoichiometry of reactions, we will consider the reaction of propane with oxygen to produce carbon dioxide and water. We will consider the question: *"What mass of oxygen will react with 96.1 grams of propane?"* In doing stoichiometry, the first thing we must do is *write the balanced chemical equation* for the reaction. In this case the balanced equation is

Before doing any calculations involving a chemical reaction, be sure the equation for the reaction is balanced.

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

which can be visualized as



This equation means that 1 mole of C_3H_8 reacts with 5 moles of O_2 to produce 3 moles of CO_2 and 4 moles of H_2O. To use this equation to find the masses of reactants and products, we must be able to convert between masses and moles of substances. Thus we must first ask: *"How many moles of propane are present in 96.1 grams of propane?"* The molar

CHEMICAL IMPACT

High Mountains—Low Octane

The next time that you visit a gas station, take a moment to note the octane rating that accompanies the grade of gasoline that you are purchasing. The gasoline is priced according to its octane rating—a measure of the fuel's antiknock properties. In a conventional internal combustion engine, gasoline vapors and air are drawn into the combustion cylinder on the downward stroke of the piston. This air–fuel mixture is compressed on the upward piston stroke (compression stroke), and a spark from the sparkplug ignites the mix. The rhythmic combustion of the air–fuel mix occurring sequentially in several cylinders furnishes the power to propel the vehicle down the road. Excessive heat and pressure (or poor-quality fuel) within the cylinder may cause the premature combustion of the mixture—commonly known as engine "knock" or "ping." Over time, this engine knock can damage the engine, resulting in inefficient performance and costly repairs.

A consumer typically is faced with three choices of gasoline, with octane ratings of 87 (regular), 89 (midgrade), and 93 (premium). But if you happen to travel or live in the higher elevations of the Rocky Mountain states, you might be surprised to find different octane ratings at the gasoline pumps. The reason for this provides a lesson in stoichiometry. At higher elevations the air is less dense—the volume of oxygen per unit volume of air is smaller. Most engines are designed to achieve a 14:1 oxygen-to-fuel ratio in the cylinder prior to combustion. If less oxygen is available, then less fuel is required to achieve this optimal ratio. In turn, the lower volumes of oxygen and fuel result in a lower pressure in the cylinder. Because high pressure tends to promote knocking, the lower pressure within engine cylinders at higher elevations promotes a more controlled combustion of the air–fuel mixture, and therefore, octane requirements are lower. While consumers in the Rocky Mountain states can purchase three grades of gasoline, the octane ratings of these fuel blends are different from those in the rest of the United States. In Denver, Colorado, regular gasoline is 85 octane, midgrade is 87 octane, and premium is 91 octane—2 points lower than gasoline sold in most of the rest of the country.

mass of propane to three significant figures is 44.1 (that is, $3 \times 12.01 + 8 \times 1.008$). The moles of propane can be calculated as follows:

$$96.1\ \cancel{\text{g } C_3H_8} \times \frac{1 \text{ mol } C_3H_8}{44.1\ \cancel{\text{g } C_3H_8}} = 2.18 \text{ mol } C_3H_8$$

Next we must take into account the fact that each mole of propane reacts with 5 moles of oxygen. The best way to do this is to use the balanced equation to construct a **mole ratio.** In this case we want to convert from moles of propane to moles of oxygen. From the balanced equation we see that 5 moles of O_2 is required for each mole of C_3H_8, so the appropriate ratio is

$$\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8}$$

Multiplying the number of moles of C_3H_8 by this factor gives the number of moles of O_2 required:

$$2.18\ \cancel{\text{mol } C_3H_8} \times \frac{5 \text{ mol } O_2}{1\ \cancel{\text{mol } C_3H_8}} = 10.9 \text{ mol } O_2$$

Notice that the mole ratio is set up so that the moles of C_3H_8 cancel out, and the units that result are moles of O_2.

Since the original question asked for the mass of oxygen needed to react with 96.1 grams of propane, the 10.9 moles of O_2 must be converted to *grams*. Since the molar mass of O_2 is 32.0 g/mol,

$$10.9\ \cancel{\text{mol } O_2} \times \frac{32.0 \text{ g } O_2}{1\ \cancel{\text{mol } O_2}} = 349 \text{ g } O_2$$

Therefore, 349 grams of oxygen is required to burn 96.1 grams of propane.

This example can be extended by asking: *"What mass of carbon dioxide is produced when 96.1 grams of propane is combusted with oxygen?"* In this case we must convert between moles of propane and moles of carbon dioxide. This can be accomplished by looking at the balanced equation, which shows that 3 moles of CO_2 is produced for each mole of C_3H_8 reacted. The mole ratio needed to convert from moles of propane to moles of carbon dioxide is

$$\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}$$

The conversion is

$$2.18 \cancel{\text{mol } C_3H_8} \times \frac{3 \text{ mol } CO_2}{1 \cancel{\text{mol } C_3H_8}} = 6.54 \text{ mol } CO_2$$

Then, using the molar mass of CO_2 (44.0 g/mol), we calculate the mass of CO_2 produced:

$$6.54 \cancel{\text{mol } CO_2} \times \frac{44.0 \text{ g } CO_2}{1 \cancel{\text{mol } CO_2}} = 288 \text{ g } CO_2$$

We will now summarize the sequence of steps needed to carry out stoichiometric calculations.

$$96.1 \text{ g } C_3H_8 \;\boxed{\frac{1 \text{ mol } C_3H_8}{44.1 \text{ g } C_3H_8}} \Rightarrow 2.18 \text{ mol } C_3H_8 \;\boxed{\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}} \Rightarrow 6.54 \text{ mol } CO_2$$

$$\boxed{\frac{44.0 \text{ g } CO_2}{1 \text{ mol } CO_2}} \Rightarrow 288 \text{ g } CO_2$$

Calculating Masses of Reactants and Products in Chemical Reactions

1 Balance the equation for the reaction.

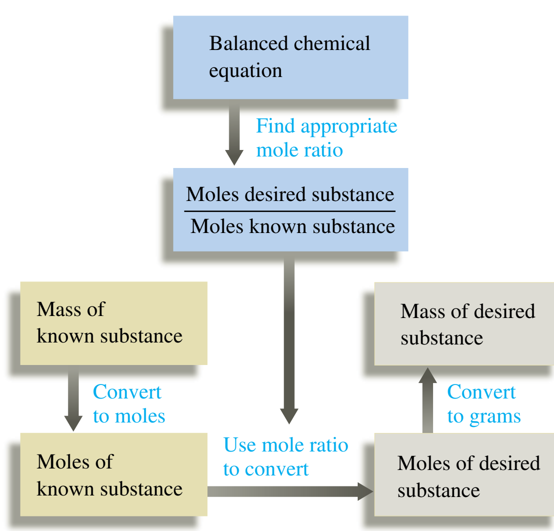
2 Convert the known mass of the reactant or product to moles of that substance.

3 Use the balanced equation to set up the appropriate mole ratios.

4 Use the appropriate mole ratios to calculate the number of moles of the desired reactant or product.

5 Convert from moles back to grams if required by the problem.

These steps are summarized by the following diagram:



Sample Exercise 3.16

Chemical Stoichiometry I

Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00 kg of lithium hydroxide?

Astronaut Sidney M. Gutierrez changes the lithium hydroxide cannisters on space shuttle Columbia. The lithium hydroxide is used to purge carbon dioxide from the air in the shuttle's cabin.

Solution

➥ **1** Using the description of the reaction, we can write the unbalanced equation:

$$\mathrm{LiOH}(s) + \mathrm{CO_2}(g) \longrightarrow \mathrm{Li_2CO_3}(s) + \mathrm{H_2O}(l)$$

The balanced equation is

$$2\mathrm{LiOH}(s) + \mathrm{CO_2}(g) \longrightarrow \mathrm{Li_2CO_3}(s) + \mathrm{H_2O}(l)$$

➥ **2** We convert the given mass of LiOH to moles, using the molar mass of LiOH ($6.941 + 16.00 + 1.008 = 23.95$ g/mol):

$$1.00\ \cancel{\mathrm{kg\ LiOH}} \times \frac{1000\ \cancel{\mathrm{g\ LiOH}}}{1\ \cancel{\mathrm{kg\ LiOH}}} \times \frac{1\ \mathrm{mol\ LiOH}}{23.95\ \cancel{\mathrm{g\ LiOH}}} = 41.8\ \mathrm{mol\ LiOH}$$

➥ **3** Since we want to determine the amount of CO_2 that reacts with the given amount of LiOH, the appropriate mole ratio is

$$\frac{1\ \mathrm{mol\ CO_2}}{2\ \mathrm{mol\ LiOH}}$$

➥ **4** We calculate the moles of CO_2 needed to react with the given mass of LiOH using this mole ratio:

$$41.8\ \cancel{\mathrm{mol\ LiOH}} \times \frac{1\ \mathrm{mol\ CO_2}}{2\ \cancel{\mathrm{mol\ LiOH}}} = 20.9\ \mathrm{mol\ CO_2}$$

➥ **5** Next we calculate the mass of CO_2, using its molar mass (44.0 g/mol):

$$20.9\ \cancel{\mathrm{mol\ CO_2}} \times \frac{44.0\ \mathrm{g\ CO_2}}{1\ \cancel{\mathrm{mol\ CO_2}}} = 9.20 \times 10^2\ \mathrm{g\ CO_2}$$

Thus 920. g of $\mathrm{CO_2}(g)$ will be absorbed by 1.00 kg of $\mathrm{LiOH}(s)$.

See Exercises 3.89 and 3.90.

Sample Exercise 3.17

Chemical Stoichiometry II

Baking soda ($NaHCO_3$) is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach:

$$\mathrm{NaHCO_3}(s) + \mathrm{HCl}(aq) \longrightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l) + \mathrm{CO_2}(aq)$$

Milk of magnesia, which is an aqueous suspension of magnesium hydroxide, is also used as an antacid:

$$\mathrm{Mg(OH)_2}(s) + 2\mathrm{HCl}(aq) \longrightarrow 2\mathrm{H_2O}(l) + \mathrm{MgCl_2}(aq)$$

Which is the more effective antacid per gram, $NaHCO_3$ or $Mg(OH)_2$?

Solution

To answer the question, we must determine the amount of HCl neutralized per gram of $NaHCO_3$ and per gram of $Mg(OH)_2$. Using the molar mass of $NaHCO_3$ (84.01 g/mol), we can determine the moles of $NaHCO_3$ in 1.00 g of $NaHCO_3$:

$$1.00\ \cancel{\mathrm{g\ NaHCO_3}} \times \frac{1\ \mathrm{mol\ NaHCO_3}}{84.01\ \cancel{\mathrm{g\ NaHCO_3}}} = 1.19 \times 10^{-2}\ \mathrm{mol\ NaHCO_3}$$

Milk of magnesia contains a suspension of $Mg(OH)_2(s)$.

Next we determine the moles of HCl using the mole ratio 1 mol HCl/1 mol $NaHCO_3$:

$$1.19 \times 10^{-2}\ \cancel{\text{mol NaHCO}_3} \times \frac{1\ \text{mol HCl}}{1\ \cancel{\text{mol NaHCO}_3}} = 1.19 \times 10^{-2}\ \text{mol HCl}$$

Thus 1.00 g of $NaHCO_3$ will neutralize 1.19×10^{-2} mol HCl.

Using the molar mass of $Mg(OH)_2$ (58.32 g/mol), we determine the moles of $Mg(OH)_2$ in 1.00 g:

$$1.00\ \cancel{\text{g Mg(OH)}_2} \times \frac{1\ \text{mol Mg(OH)}_2}{58.32\ \cancel{\text{g Mg(OH)}_2}} = 1.71 \times 10^{-2}\ \text{mol Mg(OH)}_2$$

To determine the moles of HCl that will react with this amount of $Mg(OH)_2$, we use the mole ratio 2 mol HCl/1 mol $Mg(OH)_2$:

$$1.71 \times 10^{-2}\ \cancel{\text{mol Mg(OH)}_2} \times \frac{2\ \text{mol HCl}}{1\ \cancel{\text{mol Mg(OH)}_2}} = 3.42 \times 10^{-2}\ \text{mol HCl}$$

Thus 1.00 g of $Mg(OH)_2$ will neutralize 3.42×10^{-2} mol HCl. It is a better antacid per gram than $NaHCO_3$.

See Exercises 3.91 and 3.92.

3.10 Calculations Involving a Limiting Reactant

When chemicals are mixed together to undergo a reaction, they are often mixed in **stoichiometric quantities,** that is, in exactly the correct amounts so that all reactants "run out" (are used up) at the same time. To clarify this concept, let's consider the production of hydrogen for use in the manufacture of ammonia by the **Haber process.** Ammonia, a very important fertilizer itself and a starting material for other fertilizers, is made by combining nitrogen (from the air) with hydrogen according to the equation

The details of the Haber process are discussed in Section 19.2.

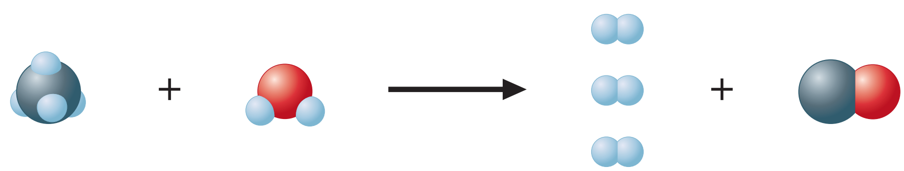
$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Hydrogen can be obtained from the reaction of methane with water vapor:

$$CH_4(g) + H_2O(g) \longrightarrow 3H_2(g) + CO(g)$$

WWW Visualization: Limiting Reactant

We can illustrate what we mean by stoichiometric quantities by first visualizing the balanced equation as follows:



Since this reaction involves one molecule of methane reacting with one molecule of water, to have stoichiometric amounts of methane and water we must have equal numbers of them, as shown in Fig. 3.9, where several stoichiometric mixtures are shown.

Suppose we want to calculate the mass of water required to react *exactly* with 2.50×10^3 kilograms of methane. That is, how much water will just consume all the 2.50×10^3 kilograms of methane, leaving no methane or water remaining?

To do this calculation, we need to recognize that we need equal numbers of methane and water molecules. Therefore, we first need to find the number of moles of methane molecules in 2.50×10^3 kg (2.50×10^6 g) of methane:

$$2.50 \times 10^6\ \cancel{\text{g CH}_4} \times \frac{1\ \text{mol CH}_4}{\underset{\uparrow\ \text{molar mass of CH}_4}{16.04\ \cancel{\text{g CH}_4}}} = 1.56 \times 10^5\ \text{mol CH}_4\ \text{molecules}$$

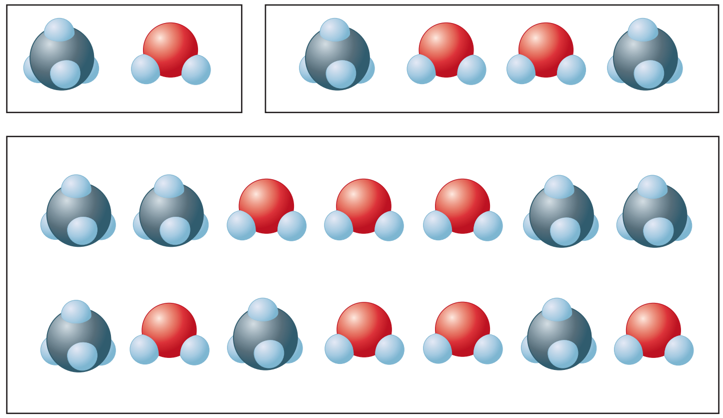


FIGURE 3.9
Three different stoichiometric mixtures of methane and water, which react one-to-one.

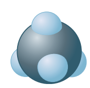
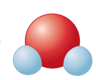
This same number of water molecules has a mass determined as follows:

$$1.56 \times 10^5 \cancel{\text{mol H}_2\text{O}} \times \frac{18.02 \text{ g}}{\cancel{\text{mol H}_2\text{O}}} = 2.81 \times 10^6 \text{ g H}_2\text{O} = 2.81 \times 10^3 \text{ kg H}_2\text{O}$$

Thus, if 2.50×10^3 kilograms of methane is mixed with 2.81×10^3 kilograms of water, both reactants will "run out" at the same time. The reactants have been mixed in stoichiometric quantities.

If, on the other hand, 2.50×10^3 kilograms of methane is mixed with 3.00×10^3 kilograms of water, the methane will be consumed before the water runs out. The water will be in *excess*; that is, there will be more water molecules than methane molecules in the reaction mixture. What is the implication of this with respect to the number of product molecules that can form?

To answer this question, consider the situation on a smaller scale. Assume we mix 10 CH_4 molecules and 17 H_2O molecules and let them react. How many H_2 and CO molecules can form?

First picture the mixture of CH_4  and H_2O  molecules as shown in Fig. 3.10.

Then imagine that groups consisting of one CH_4 molecule and one H_2O molecule (Fig. 3.10) will react to form three H_2 and one CO molecules (Fig. 3.11).

Notice that products can form only when both CH_4 and H_2O are available to react. Once the 10 CH_4 molecules are used up by reacting with 10 H_2O molecules, the remaining water

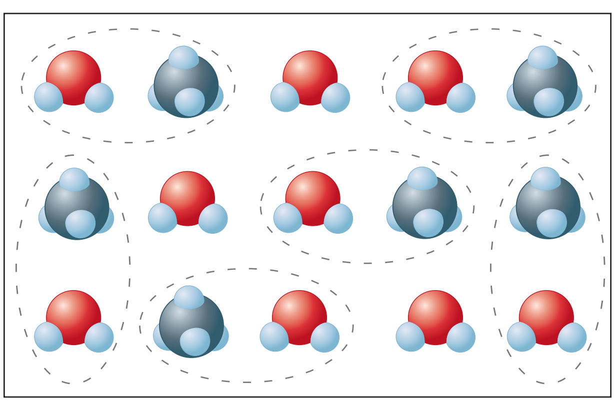


FIGURE 3.10
A mixture of CH_4 and H_2O molecules.

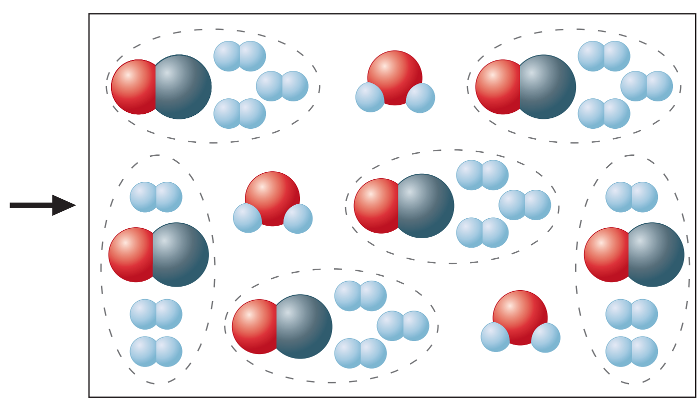


FIGURE 3.11
Methane and water have reacted to form products according to the equation $CH_4 + H_2O \longrightarrow 3H_2 + CO$.

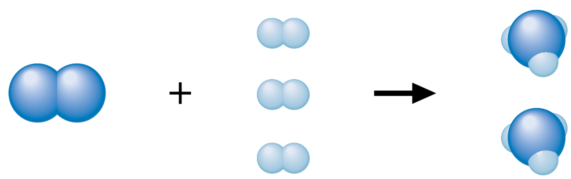
molecules cannot react. They are in excess. Thus the number of products that can form is *limited* by the methane. Once the methane is consumed, no more products can be formed, even though some water still remains. In this situation the amount of methane *limits* the amount of products that can be formed. This brings us to the concept of the **limiting reactant** (or **limiting reagent**), which is the reactant that is consumed first and that therefore limits the amounts of products that can be formed. In any stoichiometry calculation involving a chemical reaction, it is essential to determine which reactant is limiting so as to calculate correctly the amounts of products that will be formed.

To further explore the idea of a limiting reactant, consider the ammonia synthesis reaction:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Ammonia is dissolved in irrigation water to provide fertilizer for a field of corn.

Assume that 5 N_2 molecules and 9 H_2 molecules are placed in a flask. Is this a stoichiometric mixture of reactants, or will one of them be consumed before the other runs out? From the balanced equation we know that each N_2 molecule requires 3 H_2 molecules for the reaction to occur:



Thus the required H_2/N_2 ratio is $3H_2/1N_2$. In our experiment we have 9 H_2 and 5 N_2, or a ratio of $9H_2/5N_2 = 1.8H_2/1N_2$.

Since the actual ratio (1.8:1) of H_2/N_2 is less than the ratio required by the balanced equation (3:1), there is not enough hydrogen to react with all the nitrogen. That is, the hydrogen will run out first, leaving some unreacted N_2 molecules. We can visualize this as shown in Fig. 3.12.

Figure 3.12 shows that 3 of the N_2 molecules react with the 9 H_2 molecules to produce 6 NH_3 molecules:

$$3N_2 + 9H_2 \longrightarrow 6NH_3$$

This leaves 2 N_2 molecules unreacted—the nitrogen is in excess.

What we have shown here is that in this experiment the hydrogen is the limiting reactant. The amount of H_2 initially present determines the amount of NH_3 that can form. The reaction was not able to use up all the N_2 molecules because the H_2 molecules were all consumed by the first 3 N_2 molecules to react.

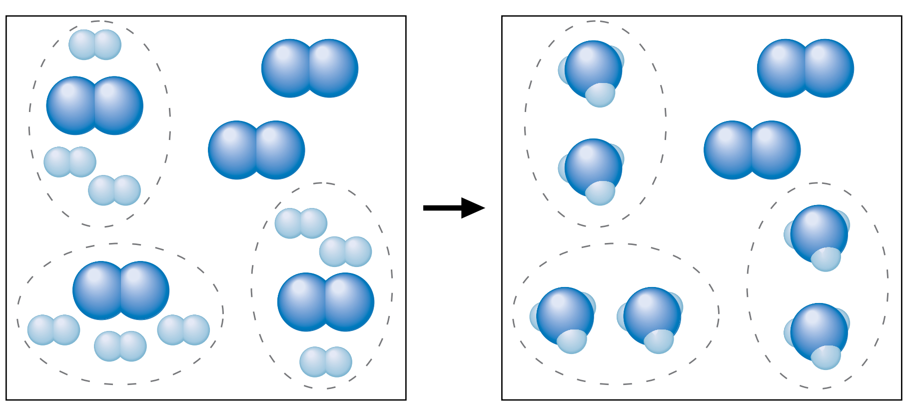


FIGURE 3.12
Hydrogen and nitrogen react to form ammonia according to the equation $N_2 + 3H_2 \longrightarrow 2NH_3$.

Another way to look at this is to determine how much H_2 would be required by 5 N_2 molecules. Multiplying the balanced equation

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

by 5 gives

$$5N_2(g) + 15H_2(g) \longrightarrow 10NH_3(g)$$

Thus 5 N_2 molecules would require 15 H_2 molecules and we have only 9. This tells us the same thing we learned earlier—the hydrogen is limiting.

The most important point here is this: *The limiting reactant limits the amount of product that can form.* The reaction that actually occurred was

$$3N_2(g) + 9H_2(g) \longrightarrow 6NH_3(g)$$

not

$$5N_2(g) + 15H_2(g) \longrightarrow 10NH_3(g)$$

Thus 6 NH_3 were formed, not 10 NH_3, because the H_2, not the N_2, was limiting.

In the laboratory or chemical plant we work with much larger quantities than the few molecules of the preceding example. Therefore, we must learn to deal with limiting reactants using moles. The ideas are exactly the same, except that we are using moles of molecules instead of individual molecules. For example, suppose 25.0 kilograms of nitrogen and 5.00 kilograms of hydrogen are mixed and reacted to form ammonia. How do we calculate the mass of ammonia produced when this reaction is run to completion (until one of the reactants is completely consumed)?

As in the preceding example, we must use the balanced equation

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

to determine whether nitrogen or hydrogen is the limiting reactant and then to determine the amount of ammonia that is formed. We first calculate the moles of reactants present:

$$25.0\ \cancel{\text{kg N}_2} \times \frac{1000\ \cancel{\text{g N}_2}}{1\ \cancel{\text{kg N}_2}} \times \frac{1\ \text{mol N}_2}{28.0\ \cancel{\text{g N}_2}} = 8.93 \times 10^2\ \text{mol N}_2$$

$$5.00\ \cancel{\text{kg H}_2} \times \frac{1000\ \cancel{\text{g H}_2}}{1\ \cancel{\text{kg H}_2}} \times \frac{1\ \text{mol H}_2}{2.016\ \cancel{\text{g H}_2}} = 2.48 \times 10^3\ \text{mol H}_2$$

Since 1 mol N_2 reacts with 3 mol H_2, the number of moles of H_2 that will react exactly with 8.93×10^2 mol N_2 is

$$8.93 \times 10^2\ \cancel{\text{mol N}_2} \times \frac{3\ \text{mol H}_2}{1\ \cancel{\text{mol N}_2}} = 2.68 \times 10^3\ \text{mol H}_2$$

Always determine which reactant is limiting.

Thus 8.93×10^2 mol N_2 requires 2.68×10^3 mol H_2 to react completely. However, in this case, only 2.48×10^3 mol H_2 is present. This means that the hydrogen will be consumed before the nitrogen. Thus hydrogen is the *limiting reactant* in this particular situation, and we must use the amount of hydrogen to compute the quantity of ammonia formed:

$$2.48 \times 10^3\ \cancel{\text{mol H}_2} \times \frac{2\ \text{mol NH}_3}{3\ \cancel{\text{mol H}_2}} = 1.65 \times 10^3\ \text{mol NH}_3$$

Converting moles to kilograms gives

$$1.65 \times 10^3\ \cancel{\text{mol NH}_3} \times \frac{17.0\ \text{g NH}_3}{1\ \cancel{\text{mol NH}_3}} = 2.80 \times 10^4\ \text{g NH}_3 = 28.0\ \text{kg NH}_3$$

Note that to determine the limiting reactant, we could have started instead with the given amount of hydrogen and calculated the moles of nitrogen required:

$$2.48 \times 10^3 \cancel{\text{mol H}_2} \times \frac{1 \text{ mol N}_2}{3 \cancel{\text{mol H}_2}} = 8.27 \times 10^2 \text{ mol N}_2$$

Thus 2.48×10^3 mol H_2 requires 8.27×10^2 mol N_2. Since 8.93×10^2 mol N_2 is actually present, the nitrogen is in excess. The hydrogen will run out first, and thus again we find that hydrogen limits the amount of ammonia formed.

A related but simpler way to determine which reactant is limiting is to compare the mole ratio of the substances required by the balanced equation with the mole ratio of reactants actually present. For example, in this case the mole ratio of H_2 to N_2 required by the balanced equation is

$$\frac{3 \text{ mol H}_2}{1 \text{ mol N}_2}$$

That is,

$$\frac{\text{mol H}_2}{\text{mol N}_2} \text{(required)} = \frac{3}{1} = 3$$

In this experiment we have 2.48×10^3 mol H_2 and 8.93×10^2 mol N_2. Thus the ratio

$$\frac{\text{mol H}_2}{\text{mol N}_2} \text{(actual)} = \frac{2.48 \times 10^3}{8.93 \times 10^2} = 2.78$$

Since 2.78 is less than 3, the actual mole ratio of H_2 to N_2 is too small, and H_2 must be limiting. If the actual H_2 to N_2 mole ratio had been greater than 3, then the H_2 would have been in excess and the N_2 would be limiting.

Sample Exercise 3.18 Stoichiometry: Limiting Reactant

Nitrogen gas can be prepared by passing gaseous ammonia over solid copper(II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1 g of NH_3 is reacted with 90.4 g of CuO, which is the limiting reactant? How many grams of N_2 will be formed?

Solution

From the description of the problem, we can obtain the following balanced equation:

$$2\text{NH}_3(g) + 3\text{CuO}(s) \longrightarrow \text{N}_2(g) + 3\text{Cu}(s) + 3\text{H}_2\text{O}(g)$$

Next we must compute the moles of NH_3 (molar mass = 17.03 g/mol) and of CuO (molar mass = 79.55 g/mol):

$$18.1 \cancel{\text{g NH}_3} \times \frac{1 \text{ mol NH}_3}{17.03 \cancel{\text{g NH}_3}} = 1.06 \text{ mol NH}_3$$

$$90.4 \cancel{\text{g CuO}} \times \frac{1 \text{ mol CuO}}{79.55 \cancel{\text{g CuO}}} = 1.14 \text{ mol CuO}$$

To determine the limiting reactant, we use the mole ratio for CuO and NH_3:

$$1.06 \cancel{\text{mol NH}_3} \times \frac{3 \text{ mol CuO}}{2 \cancel{\text{mol NH}_3}} = 1.59 \text{ mol CuO}$$

Thus 1.59 mol CuO is required to react with 1.06 mol NH_3. Since only 1.14 mol CuO is actually present, the amount of CuO is limiting; CuO will run out before NH_3 does. We can verify this conclusion by comparing the mole ratio of CuO and NH_3 required by the balanced equation

$$\frac{\text{mol CuO}}{\text{mol NH}_3}\text{(required)} = \frac{3}{2} = 1.5$$

with the mole ratio actually present

$$\frac{\text{mol CuO}}{\text{mol NH}_3}\text{(actual)} = \frac{1.14}{1.06} = 1.08$$

Since the actual ratio is too small (smaller than 1.5), CuO is the limiting reactant.

Because CuO is the limiting reactant, we must use the amount of CuO to calculate the amount of N_2 formed. From the balanced equation, the mole ratio between CuO and N_2 is

$$\frac{1 \text{ mol N}_2}{3 \text{ mol CuO}}$$

$$1.14 \cancel{\text{mol CuO}} \times \frac{1 \text{ mol N}_2}{3 \cancel{\text{mol CuO}}} = 0.380 \text{ mol N}_2$$

Using the molar mass of N_2 (28.0 g/mol), we can calculate the mass of N_2 produced:

$$0.380 \cancel{\text{mol N}_2} \times \frac{28.0 \text{ g N}_2}{1 \cancel{\text{mol N}_2}} = 10.6 \text{ g N}_2$$

See Exercises 3.99 through 3.101.

The amount of a product formed when the limiting reactant is completely consumed is called the **theoretical yield** of that product. In Sample Exercise 3.18, 10.6 grams of nitrogen represents the theoretical yield. This is the *maximum amount* of nitrogen that can be produced from the quantities of reactants used. Actually, the amount of product predicted by the theoretical yield is seldom obtained because of side reactions (other reactions that involve one or more of the reactants or products) and other complications. The *actual yield* of product is often given as a percentage of the theoretical yield. This is called the **percent yield:**

Percent yield is important as an indicator of the efficiency of a particular laboratory or industrial reaction.

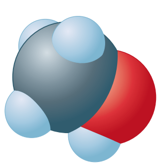
$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \text{percent yield}$$

For example, if the reaction considered in Sample Exercise 3.18 actually gave 6.63 grams of nitrogen instead of the predicted 10.6 grams, the percent yield of nitrogen would be

$$\frac{6.63 \cancel{\text{g N}_2}}{10.6 \cancel{\text{g N}_2}} \times 100\% = 62.5\%$$

Sample Exercise 3.19

Calculating Percent Yield



Methanol

Methanol (CH_3OH), also called *methyl alcohol,* is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combination of gaseous carbon monoxide and hydrogen. Suppose 68.5 kg $CO(g)$ is reacted with 8.60 kg $H_2(g)$. Calculate the theoretical yield of methanol. If 3.57×10^4 g CH_3OH is actually produced, what is the percent yield of methanol?

Solution

First, we must find out which reactant is limiting. The balanced equation is

$$2H_2(g) + CO(g) \longrightarrow CH_3OH(l)$$

Next we must calculate the moles of reactants:

$$68.5\ \cancel{\text{kg CO}} \times \frac{1000\ \cancel{\text{g CO}}}{1\ \cancel{\text{kg CO}}} \times \frac{1\ \text{mol CO}}{28.02\ \cancel{\text{g CO}}} = 2.44 \times 10^3\ \text{mol CO}$$

$$8.60\ \cancel{\text{kg H}_2} \times \frac{1000\ \cancel{\text{g H}_2}}{1\ \cancel{\text{kg H}_2}} \times \frac{1\ \text{mol H}_2}{2.016\ \cancel{\text{g H}_2}} = 4.27 \times 10^3\ \text{mol H}_2$$

To determine which reactant is limiting, we compare the mole ratio of H_2 and CO required by the balanced equation

$$\frac{\text{mol H}_2}{\text{mol CO}}\text{(required)} = \frac{2}{1} = 2$$

with the actual mole ratio

$$\frac{\text{mol H}_2}{\text{mol CO}}\text{(actual)} = \frac{4.27 \times 10^3}{2.44 \times 10^3} = 1.75$$

Since the actual mole ratio of H_2 to CO is smaller than the required ratio, *H_2 is limiting.* We therefore must use the amount of H_2 and the mole ratio between H_2 and CH_3OH to determine the maximum amount of methanol that can be produced:

$$4.27 \times 10^3\ \cancel{\text{mol H}_2} \times \frac{1\ \text{mol CH}_3\text{OH}}{2\ \cancel{\text{mol H}_2}} = 2.14 \times 10^3\ \text{mol CH}_3\text{OH}$$

Using the molar mass of CH_3OH (32.04 g/mol), we can calculate the theoretical yield in grams:

$$2.14 \times 10^3\ \cancel{\text{mol CH}_3\text{OH}} \times \frac{32.04\ \text{g CH}_3\text{OH}}{1\ \cancel{\text{mol CH}_3\text{OH}}} = 6.86 \times 10^4\ \text{g CH}_3\text{OH}$$

Thus, from the amount of reactants given, the maximum amount of CH_3OH that can be formed is 6.86×10^4 g. This is the *theoretical yield.*

The percent yield is

$$\frac{\text{Actual yield (grams)}}{\text{Theoretical yield (grams)}} \times 100 = \frac{3.57 \times 10^4\ \cancel{\text{g CH}_3\text{OH}}}{6.86 \times 10^4\ \cancel{\text{g CH}_3\text{OH}}} \times 100\% = 52.0\%$$

See Exercises 3.103 and 3.104.

Methanol is used as a fuel in Indianapolis-type racing cars.

Solving a Stoichiometry Problem Involving Masses of Reactants and Products

➥ **1 Write and balance the equation for the reaction.**

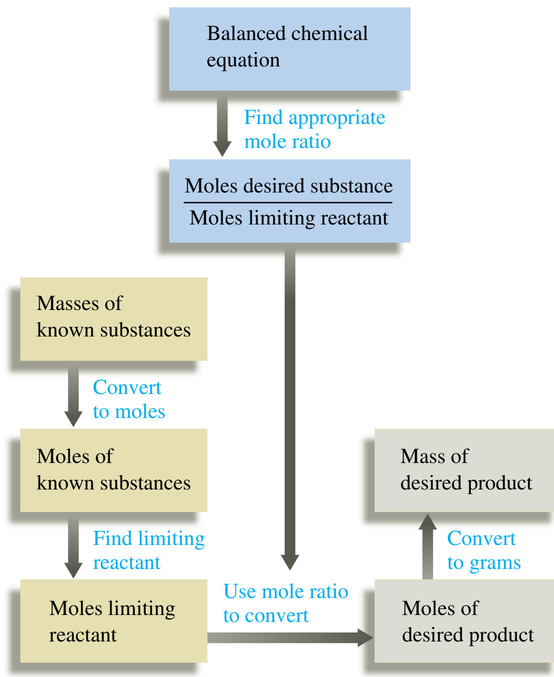
➥ **2 Convert the known masses of substances to moles.**

➥ **3 Determine which reactant is limiting.**

➥ **4 Using the amount of the limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product.**

➥ **5 Convert from moles to grams, using the molar mass.**

This process is summarized in the diagram below:



Key Terms

chemical stoichiometry

Section 3.2

mass spectrometer

average atomic mass

Section 3.3

mole

Avogadro's number

Section 3.4

molar mass

Section 3.5

mass percent

Section 3.6

empirical formula

molecular formula

For Review

Stoichiometry

- Deals with the amounts of substances consumed and/or produced in a chemical reaction.
- We count atoms by measuring the mass of the sample.
- To relate mass and the number of atoms, the average atomic mass is required.

Mole

- The amount of carbon atoms in exactly 12 g of pure ^{12}C
- 6.022×10^{23} units of a substance
- The mass of one mole of an element = the atomic mass in grams

Molar mass

- Mass (g) of one mole of a compound or element
- Obtained for a compound by finding the sum of the average masses of its constituent atoms

Section 3.7
chemical equation
reactants
products
balancing a chemical equation

Section 3.9
mole ratio

Section 3.10
stoichiometric quantities
Haber process
limiting reactant (reagent)
theoretical yield
percent yield

Percent composition
- The mass percent of each element in a compound
- Mass percent $= \dfrac{\text{mass of element in 1 mole of substance}}{\text{mass of 1 mole of substance}} \times 100\%$

Empirical formula
- The simplest whole-number ratio of the various types of atoms in a compound
- Can be obtained from the mass percent of elements in a compound

Molecular formula
- For molecular substances:
  - The formula of the constituent molecules
  - Always an integer multiple of the empirical formula
- For ionic substances:
  - The same as the empirical formula

Chemical reactions
- Reactants are turned into products.
- Atoms are neither created nor destroyed.
- All of the atoms present in the reactants must also be present in the products.

Characteristics of a chemical equation
- Represents a chemical reaction
- Reactants on the left side of the arrow, products on the right side
- When balanced, gives the relative numbers of reactant and product molecules or ions

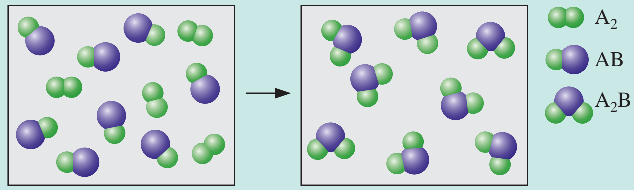
Stoichiometry calculations
- Amounts of reactants consumed and products formed can be determined from the balanced chemical equation.
- The limiting reactant is the one consumed first, thus limiting the amount of product that can form.

Yield
- The theoretical yield is the maximum amount that can be produced from a given amount of the limiting reactant.
- The actual yield, the amount of product actually obtained, is always less than the theoretical yield.
- Percent yield $= \dfrac{\text{actual yield (g)}}{\text{theoretical yield (g)}} \times 100\%$

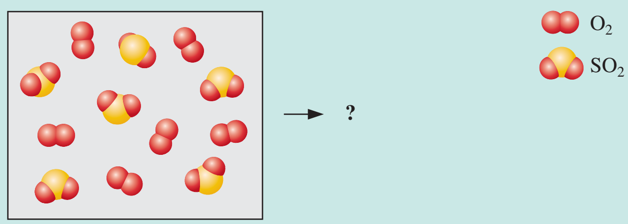
REVIEW QUESTIONS

1. Explain the concept of "counting by weighing" using marbles as your example.
2. Atomic masses are relative masses. What does this mean?
3. The atomic mass of boron (B) is given in the periodic table as 10.81, yet no single atom of boron has a mass of 10.81 amu. Explain.
4. What three conversion factors and in what order would you use them to convert the mass of a compound into atoms of a particular element in that compound—for example, from 1.00 g aspirin ($C_9H_8O_4$) to number of hydrogen atoms in the 1.00-g sample?
5. Figure 3.5 illustrates a schematic diagram of a combustion device used to analyze organic compounds. Given that a certain amount of a compound containing carbon, hydrogen, and oxygen is combusted in this device, explain how the data relating to the mass of CO_2 produced and the mass of H_2O produced can be manipulated to determine the empirical formula.

6. What is the difference between the empirical and molecular formulas of a compound? Can they ever be the same? Explain.
7. Consider the hypothetical reaction between A_2 and AB pictured below.



What is the balanced equation? If 2.50 mol A_2 is reacted with excess AB, what amount (moles) of product will form? If the mass of AB is 30.0 amu and the mass of A_2 is 40.0 amu, what is the mass of the product? If 15.0 g of AB is reacted, what mass of A_2 is required to react with all of the AB, and what mass of product is formed?
8. What is a limiting reactant problem? Explain two different strategies that can be used to solve limiting reactant problems.
9. Consider the following mixture of $SO_2(g)$ and $O_2(g)$.



If $SO_2(g)$ and $O_2(g)$ react to form $SO_3(g)$, draw a representation of the product mixture assuming the reaction goes to completion. What is the limiting reactant in the reaction? If 96.0 g of SO_2 reacts with 32.0 g O_2, what mass of product will form?
10. Why is the actual yield of a reaction often less than the theoretical yield?

Active Learning Questions

These questions are designed to be used by groups of students in class. The questions allow students to explore their understanding of concepts through discussion and peer teaching. The real value of these questions is the learning that occurs while students talk to each other about chemical concepts.

1. The following are actual student responses to the question: Why is it necessary to balance chemical equations?
   a. The chemicals will not react until you have added the correct mole ratios.
   b. The correct products will not be formed unless the right amount of reactants have been added.
   c. A certain number of products cannot be formed without a certain number of reactants.
   d. The balanced equation tells you how much reactant you need and allows you to predict how much product you'll make.
   e. A mole-to-mole ratio must be established for the reaction to occur as written.

   Justify the best choice, and for choices you did not pick, explain what is wrong with them.

2. What information do we get from a formula? From an equation?

3. You are making cookies and are missing a key ingredient—eggs. You have most of the other ingredients needed to make the cookies, except you have only 1.33 cups of butter and no eggs. You note that the recipe calls for 2 cups of butter and 3 eggs (plus the other ingredients) to make 6 dozen cookies. You call a friend and have him bring you some eggs.

a. What number of eggs do you need?

b. If you use all the butter (and get enough eggs), what number of cookies will you make?

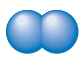
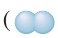
Unfortunately, your friend hangs up before you tell him how many eggs you need. When he arrives, he has a surprise for you—to save time, he has broken them all in a bowl for you. You ask him how many he brought, and he replies, "I can't remember." You weigh the eggs and find that they weigh 62.1 g. Assuming that an average egg weighs 34.21 g,

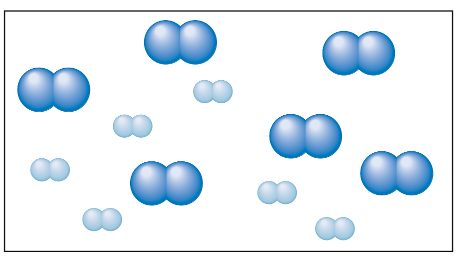
a. What quantity of butter is needed to react with all the eggs?

b. What number of cookies can you make?

c. Which will you have left over, eggs or butter?

d. What quantity is left over?

4. Nitrogen (N_2) and hydrogen (H_2) react to form ammonia (NH_3). Consider the mixture of N_2 () and H_2 () in a closed container as illustrated below:



Assuming the reaction goes to completion, draw a representation of the product mixture. Explain how you arrived at this representation.

5. For the preceding question, which of the following equations best represents the reaction?

a. $6N_2 + 6H_2 \longrightarrow 4NH_3 + 4N_2$

b. $N_2 + H_2 \longrightarrow NH_3$

c. $N + 3H \longrightarrow NH_3$

d. $N_2 + 3H_2 \longrightarrow 2NH_3$

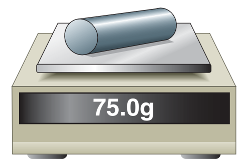
e. $2N_2 + 6H_2 \longrightarrow 4NH_3$

Justify your choice, and for choices you did not pick, explain what is wrong with them.

6. You know that chemical *A* reacts with chemical *B*. You react 10.0 g *A* with 10.0 g *B*. What information do you need to determine the amount of product that will be produced? Explain.

7. A new grill has a mass of 30.0 kg. You put 3.0 kg of charcoal in the grill. You burn all the charcoal and the grill has a mass of 30.0 kg. What is the mass of the gases given off? Explain.

8. Consider an iron bar on a balance as shown.



As the iron bar rusts, which of the following is true? Explain your answer.

a. The balance will read less than 75.0 g.

b. The balance will read 75.0 g.

c. The balance will read greater than 75.0 g.

d. The balance will read greater than 75.0 g, but if the bar is removed, the rust is scraped off, and the bar replaced, the balance will read 75.0 g.

9. You may have noticed that water sometimes drips from the exhaust of a car as it is running. Is this evidence that there is at least a small amount of water originally present in the gasoline? Explain.

Questions 10 and 11 deal with the following situation: You react chemical *A* with chemical *B* to make one product. It takes 100 g of *A* to react completely with 20 g *B*.

10. What is the mass of the product?

a. less than 10 g

b. between 20 and 100 g

c. between 100 and 120 g

d. exactly 120 g

e. more than 120 g

11. What is true about the chemical properties of the product?

a. The properties are more like chemical *A*.

b. The properties are more like chemical *B*.

c. The properties are an average of those of chemical *A* and chemical *B*.

d. The properties are not necessarily like either chemical *A* or chemical *B*.

e. The properties are more like chemical *A* or more like chemical *B*, but more information is needed.

Justify your choice, and for choices you did not pick, explain what is wrong with them.

12. Is there a difference between a homogeneous mixture of hydrogen and oxygen in a 2:1 mole ratio and a sample of water vapor? Explain.

13. Chlorine exists mainly as two isotopes, ^{37}Cl and ^{35}Cl. Which is more abundant? How do you know?

14. The average mass of a carbon atom is 12.011. Assuming you could pick up one carbon atom, estimate the chance that you would randomly get one with a mass of 12.011. Support your answer.

15. Can the subscripts in a chemical formula be fractions? Explain. Can the coefficients in a balanced chemical equation be fractions? Explain. Changing the subscripts of chemicals can balance the equations mathematically. Why is this unacceptable?

16. Consider the equation $2A + B \longrightarrow A_2B$. If you mix 1.0 mol of *A* with 1.0 mol of *B*, what amount (moles) of A_2B can be produced?

17. According to the law of conservation of mass, mass cannot be gained or destroyed in a chemical reaction. Why can't you simply add the masses of two reactants to determine the total mass of product?

18. Which of the following pairs of compounds have the same *empirical* formula?

a. acetylene, C_2H_2, and benzene, C_6H_6

b. ethane, C_2H_6, and butane, C_4H_{10}

c. nitrogen dioxide, NO_2, and dinitrogen tetroxide, N_2O_4

d. diphenyl ether, $C_{12}H_{10}O$, and phenol, C_6H_5OH

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of the book and a solution appears in the *Solutions Guide*.

Questions

19. Reference section 3.2 to find the atomic masses of ^{12}C and ^{13}C, the relative abundance of ^{12}C and ^{13}C in natural carbon, and the average mass (in amu) of a carbon atom. If you had a sample of natural carbon containing exactly 10,000 atoms, determine the number of ^{12}C and ^{13}C atoms present. What would be the average mass (in amu) and the total mass (in amu) of the carbon atoms in this 10,000-atom sample? If you had a sample of natural carbon containing 6.0221×10^{23} atoms, determine the number of ^{12}C and ^{13}C atoms present. What would be the average mass (in amu) and the total mass (in amu) of this 6.0221×10^{23} atom sample? Given that 1 g = 6.0221×10^{23} amu, what is the total mass of 1 mol of natural carbon in units of grams?

20. Avogadro's number, molar mass, and the chemical formula of a compound are three useful conversion factors. What unit conversions can be accomplished using these conversion factors?

21. If you had a mol of U.S. dollar bills and equally distributed the money to all of the people of the world, how rich would every person be? Assume a world population of 6 billion.

22. What is the difference between the molar mass and the empirical formula mass of a compound? When are these masses the same and when are they different? When different, how is the molar mass related to the empirical formula mass?

23. How is the mass percent of elements in a compound different for a 1.0-g sample versus a 100.-g sample versus a 1-mol sample of the compound?

24. A balanced chemical equation contains a large amount of information. What information is given in a balanced equation?

25. Consider the following generic reaction:

$$A_2B_2 + 2C \longrightarrow 2CB \text{ and } 2A$$

What steps and information are necessary to perform the following determinations assuming that 1.00×10^4 molecules of A_2B_2 are reacted with excess C?

a. mass of CB produced

b. atoms of A produced

c. mol of C reacted

d. percent yield of CB

26. Consider the following generic reaction:

$$Y_2 + 2XY \longrightarrow 2XY_2$$

In a limiting reactant problem, a certain quantity of each reactant is given and you are usually asked to calculate the mass of product formed. If 10.0 g of Y_2 is reacted with 10.0 g of XY, outline two methods you could use the determine which reactant is limiting (runs out first) and thus determines the mass of product formed. A method sometimes used to solve limiting reactant problems is to assume each reactant is limiting and then calculate the mass of product formed from each given quantity of reactant. How does this method work in determining which reactant is limiting?

Exercises

In this section similar exercises are paired.

Atomic Masses and the Mass Spectrometer

27. An element consists of 1.40% of an isotope with mass 203.973 amu, 24.10% of an isotope with mass 205.9745 amu, 22.10% of an isotope with mass 206.9759 amu, and 52.40% of an isotope with mass 207.9766 amu. Calculate the average atomic mass and identify the element.

28. An element "X" has five major isotopes, which are listed below along with their abundances. What is the element?

Isotope	Percent Natural Abundance	Atomic Mass
^{46}X	8.00%	45.95269
^{47}X	7.30%	46.951764
^{48}X	73.80%	47.947947
^{49}X	5.50%	48.947841
^{50}X	5.40%	49.944792

29. The element rhenium (Re) has two naturally occurring isotopes, ^{185}Re and ^{187}Re, with an average atomic mass of 186.207 amu. Rhenium is 62.60% ^{187}Re, and the atomic mass of ^{187}Re is 186.956 amu. Calculate the mass of ^{185}Re.

30. Assume silicon has three major isotopes in nature as shown in the table below. Fill in the missing information.

Isotope	Mass (amu)	Abundance
^{28}Si	27.98	______
^{29}Si	______	4.70%
^{32}Si	29.97	3.09%

31. The mass spectrum of bromine (Br_2) consists of three peaks with the following characteristics:

Mass (amu)	Relative Size
157.84	0.2534
159.84	0.5000
161.84	0.2466

How do you interpret these data?

32. Gallium arsenide, GaAs, has gained widespread use in semiconductor devices that convert light and electrical signals in fiber-optic communications systems. Gallium consists of 60.% ^{69}Ga and 40.% ^{71}Ga. Arsenic has only one naturally occurring isotope, ^{75}As. Gallium arsenide is a polymeric material, but its mass spectrum shows fragments with the formulas GaAs and Ga_2As_2. What would the distribution of peaks look like for these two fragments?

Moles and Molar Masses

33. Calculate the mass of 500. atoms of iron (Fe).

34. What number of Fe atoms and what amount (moles) of Fe atoms are in 500.0 g of iron?

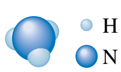
35. Diamond is a natural form of pure carbon. What number of atoms of carbon are in a 1.00-carat diamond (1.00 carat = 0.200 g)?

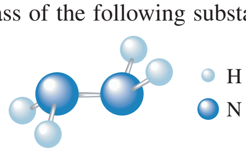
36. A diamond contains 5.0×10^{21} atoms of carbon. What amount (moles) of carbon and what mass (grams) of carbon are in this diamond?

37. Aluminum metal is produced by passing an electric current through a solution of aluminum oxide (Al_2O_3) dissolved in molten cryolite (Na_3AlF_6). Calculate the molar masses of Al_2O_3 and Na_3AlF_6.

38. The Freons are a class of compounds containing carbon, chlorine, and fluorine. While they have many valuable uses, they have been shown to be responsible for depletion of the ozone in the upper atmosphere. In 1991, two replacement compounds for Freons went into production: HFC-134a (CH_2FCF_3) and HCFC-124 ($CHClFCF_3$). Calculate the molar masses of these two compounds.

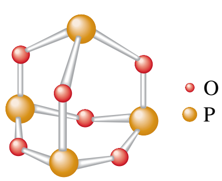
39. Calculate the molar mass of the following substances.

a. 

b. 

c. $(NH_4)_2Cr_2O_7$

40. Calculate the molar mass of the following substances.

a. 

b. $Ca_3(PO_4)_2$ **c.** Na_2HPO_4

41. What amount (moles) of compound is present in 1.00 g of each of the compounds in Exercise 39?

42. What amount (moles) of compound is present in 1.00 g of each of the compounds in Exercise 40?

43. What mass of compound is present in 5.00 mol of each of the compounds in Exercise 39?

44. What mass of compound is present in 5.00 mol of each of the compounds in Exercise 40?

45. What mass of nitrogen is present in 5.00 mol of each of the compounds in Exercise 39?

46. What mass of phosphorus is present in 5.00 mol of each of the compounds in Exercise 40?

47. What number of molecules (or formula units) are present in 1.00 g of each of the compounds in Exercise 39?

48. What number of molecules (or formula units) are present in 1.00 g of each of the compounds in Exercise 40?

49. What number of atoms of nitrogen are present in 1.00 g of each of the compounds in Exercise 39?

50. What number of atoms of phosphorus are present in 1.00 g of each of the compounds in Exercise 40?

51. Ascorbic acid, or vitamin C ($C_6H_8O_6$), is an essential vitamin. It cannot be stored by the body and must be present in the diet. What is the molar mass of ascorbic acid? Vitamin C tablets are taken as a dietary supplement. If a typical tablet contains 500.0 mg of vitamin C, what amount (moles) and what number of molecules of vitamin C does it contain?

52. The molecular formula of acetylsalicylic acid (aspirin), one of the most commonly used pain relievers, is $C_9H_8O_4$.

a. Calculate the molar mass of aspirin.

b. A typical aspirin tablet contains 500. mg of $C_9H_8O_4$. What amount (moles) of $C_9H_8O_4$ molecules and what number of molecules of acetylsalicylic acid are in a 500.-mg tablet?

53. What amount (moles) are represented by each of these samples?

a. 150.0 g Fe_2O_3 **c.** 1.5×10^{16} molecules of BF_3

b. 10.0 mg NO_2

54. What amount (moles) is represented by each of these samples?

a. 20.0 mg caffeine, $C_8H_{10}N_4O_2$

b. 2.72×10^{21} molecules of ethanol, C_2H_5OH

c. 1.50 g of dry ice, CO_2

55. What number of atoms of nitrogen are present in 5.00 g of each of the following?

a. glycine, $C_2H_5O_2N$ **c.** calcium nitrate

b. magnesium nitride **d.** dinitrogen tetroxide

56. Complete the following table.

Mass of Sample	Moles of Sample	Molecules in Sample	Total Atoms in Sample
4.24 g C_6H_6	________	________	________
________	0.224 mol H_2O	________	________
________	________	2.71×10^{22} molecules CO_2	________
________	________	________	3.35×10^{22} total atoms in CH_3OH sample

57. Aspartame is an artificial sweetener that is 160 times sweeter than sucrose (table sugar) when dissolved in water. It is marketed as Nutra-Sweet. The molecular formula of aspartame is $C_{14}H_{18}N_2O_5$.

a. Calculate the molar mass of aspartame.
b. What amount (moles) of molecules are present in 10.0 g aspartame?
c. Calculate the mass in grams of 1.56 mol aspartame.
d. What number of molecules are in 5.0 mg aspartame?
e. What number of atoms of nitrogen are in 1.2 g aspartame?
f. What is the mass in grams of 1.0×10^9 molecules of aspartame?
g. What is the mass in grams of one molecule of aspartame?

58. Chloral hydrate ($C_2H_3Cl_3O_2$) is a drug formerly used as a sedative and hypnotic. It is the compound used to make "Mickey Finns" in detective stories.
a. Calculate the molar mass of chloral hydrate.
b. What amount (moles) of $C_2H_3Cl_3O_2$ molecules are in 500.0 g chloral hydrate?
c. What is the mass in grams of 2.0×10^{-2} mol chloral hydrate?
d. What number of chlorine atoms are in 5.0 g chloral hydrate?
e. What mass of chloral hydrate would contain 1.0 g Cl?
f. What is the mass of exactly 500 molecules of chloral hydrate?

Percent Composition

59. Calculate the percent composition by mass of the following compounds that are important starting materials for synthetic polymers:
a. $C_3H_4O_2$ (acrylic acid, from which acrylic plastics are made)
b. $C_4H_6O_2$ (methyl acrylate, from which Plexiglas is made)
c. C_3H_3N (acrylonitrile, from which Orlon is made)

60. Anabolic steroids are performance enhancement drugs whose use has been banned from most major sporting activities. One anabolic steroid is fluoxymesterone ($C_{20}H_{29}FO_3$). Calculate the percent composition by mass of fluoxymesterone.

61. Several important compounds contain only nitrogen and oxygen. Place the following compounds in order of increasing mass percent of nitrogen.
a. NO, a gas formed by the reaction of N_2 with O_2 in internal combustion engines
b. NO_2, a brown gas mainly responsible for the brownish color of photochemical smog
c. N_2O_4, a colorless liquid used as fuel in space shuttles
d. N_2O, a colorless gas sometimes used as an anesthetic by dentists (known as laughing gas)

62. Arrange the following substances in order of increasing mass percent of carbon.
a. caffeine, $C_8H_{10}N_4O_2$
b. sucrose, $C_{12}H_{22}O_{11}$
c. ethanol, C_2H_5OH

63. Vitamin B_{12}, cyanocobalamin, is essential for human nutrition. It is concentrated in animal tissue but not in higher plants. Although nutritional requirements for the vitamin are quite low, people who abstain completely from animal products may develop a deficiency anemia. Cyanocobalamin is the form used in vitamin supplements. It contains 4.34% cobalt by mass. Calculate the molar mass of cyanocobalamin, assuming that there is one atom of cobalt in every molecule of cyanocobalamin.

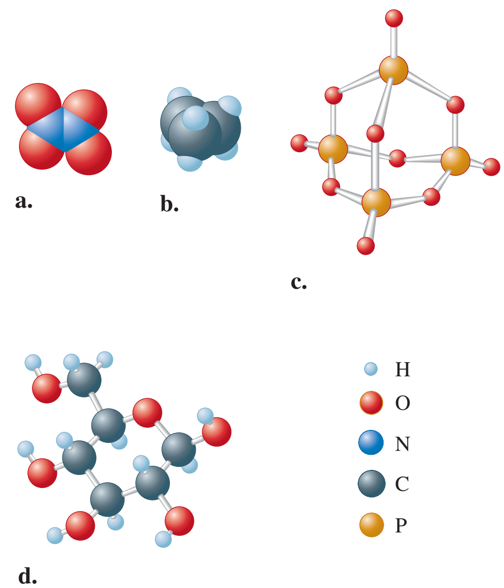
64. Fungal laccase, a blue protein found in wood-rotting fungi, is 0.390% Cu by mass. If a fungal laccase molecule contains 4 copper atoms, what is the molar mass of fungal laccase?

Empirical and Molecular Formulas

65. Express the composition of each of the following compounds as the mass percents of its elements.
a. formaldehyde, CH_2O
b. glucose, $C_6H_{12}O_6$
c. acetic acid, $HC_2H_3O_2$

66. Considering your answer to Exercise 65, which type of formula, empirical or molecular, can be obtained from elemental analysis that gives percent composition?

67. Give the empirical formula for each of the compounds represented below.



68. Determine the molecular formulas to which the following empirical formulas and molar masses pertain.
a. SNH (188.35 g/mol)
b. $NPCl_2$ (347.64 g/mol)
c. CoC_4O_4 (341.94 g/mol)
d. SN (184.32 g/mol)

69. The compound adrenaline contains 56.79% C, 6.56% H, 28.37% O, and 8.28% N by mass. What is the empirical formula for adrenaline?

70. The most common form of nylon (nylon-6) is 63.68% carbon, 12.38% nitrogen, 9.80% hydrogen, and 14.14% oxygen. Calculate the empirical formula for nylon-6.

71. There are two binary compounds of mercury and oxygen. Heating either of them results in the decomposition of the compound, with oxygen gas escaping into the atmosphere while leaving a residue of pure mercury. Heating 0.6498 g of one of the compounds leaves a residue of 0.6018 g. Heating 0.4172 g of the other compound results in a mass loss of 0.016 g. Determine the empirical formula of each compound.

72. A sample of urea contains 1.121 g N, 0.161 g H, 0.480 g C, and 0.640 g O. What is the empirical formula of urea?

73. A compound containing only sulfur and nitrogen is 69.6% S by mass; the molar mass is 184 g/mol. What are the empirical and molecular formulas of the compound?

74. Determine the molecular formula of a compound that contains 26.7% P, 12.1% N, and 61.2% Cl, and has a molar mass of 580 g/mol.

75. Adipic acid is an organic compound composed of 49.31% C, 43.79% O, and the rest hydrogen. If the molar mass of adipic acid is 146.1 g/mol, what are the empirical and molecular formulas for adipic acid?

76. Maleic acid is an organic compound composed of 41.39% C, 3.47% H, and the rest oxygen. If 0.129 mol of maleic acid has a mass of 15.0 g, what are the empirical and molecular formulas of maleic acid?

77. Many homes in rural America are heated by propane gas, a compound that contains only carbon and hydrogen. Complete combustion of a sample of propane produced 2.641 g of carbon dioxide and 1.442 g of water as the only products. Find the empirical formula of propane.

78. A compound contains only C, H, and N. Combustion of 35.0 mg of the compound produces 33.5 mg CO_2 and 41.1 mg H_2O. What is the empirical formula of the compound?

79. Cumene is a compound containing only carbon and hydrogen that is used in the production of acetone and phenol in the chemical industry. Combustion of 47.6 mg cumene produces some CO_2 and 42.8 mg water. The molar mass of cumene is between 115 and 125 g/mol. Determine the empirical and molecular formulas.

80. A compound contains only carbon, hydrogen, and oxygen. Combustion of 10.68 mg of the compound yields 16.01 mg CO_2 and 4.37 mg H_2O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of the compound?

Balancing Chemical Equations

81. Give the balanced equation for each of the following chemical reactions:

a. Glucose ($C_6H_{12}O_6$) reacts with oxygen gas to produce gaseous carbon dioxide and water vapor.

b. Solid iron(III) sulfide reacts with gaseous hydrogen chloride to form solid iron(III) chloride and hydrogen sulfide gas.

c. Carbon disulfide liquid reacts with ammonia gas to produce hydrogen sulfide gas and solid ammonium thiocyanate (NH_4SCN).

82. Give the balanced equation for each of the following.

a. The combustion of ethanol (C_2H_5OH) forms carbon dioxide and water vapor. A combustion reaction refers to a reaction of a substance with oxygen gas.

b. Aqueous solutions of lead(II) nitrate and sodium phosphate are mixed, resulting in the precipitate formation of lead(II) phosphate with aqueous sodium nitrate as the other product.

c. Solid zinc reacts with aqueous HCl to form aqueous zinc chloride and hydrogen gas.

d. Aqueous strontium hydroxide reacts with aqueous hydrobromic acid to produce water and aqueous strontium bromide.

83. Balance the following equations:

a. $Ca(OH)_2(aq) + H_3PO_4(aq) \rightarrow H_2O(l) + Ca_3(PO_4)_2(s)$

b. $Al(OH)_3(s) + HCl(aq) \rightarrow AlCl_3(aq) + H_2O(l)$

c. $AgNO_3(aq) + H_2SO_4(aq) \rightarrow Ag_2SO_4(s) + HNO_3(aq)$

84. Balance each of the following chemical equations.

a. $KO_2(s) + H_2O(l) \rightarrow KOH(aq) + O_2(g) + H_2O_2(aq)$

b. $Fe_2O_3(s) + HNO_3(aq) \rightarrow Fe(NO_3)_3(aq) + H_2O(l)$

c. $NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$

d. $PCl_5(l) + H_2O(l) \rightarrow H_3PO_4(aq) + HCl(g)$

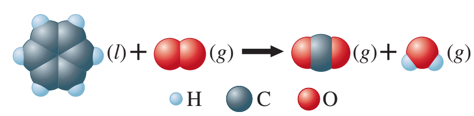
e. $CaO(s) + C(s) \rightarrow CaC_2(s) + CO_2(g)$

f. $MoS_2(s) + O_2(g) \rightarrow MoO_3(s) + SO_2(g)$

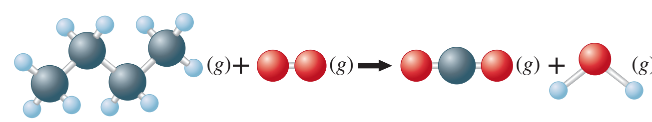
g. $FeCO_3(s) + H_2CO_3(aq) \rightarrow Fe(HCO_3)_2(aq)$

85. Balance the following equations representing combustion reactions:

a.



b.



c. $C_{12}H_{22}O_{11}(s) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

d. $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$

e. $FeO(s) + O_2(g) \rightarrow Fe_2O_3(s)$

86. Balance the following equations:

a. $Cr(s) + S_8(s) \rightarrow Cr_2S_3(s)$

b. $NaHCO_3(s) \xrightarrow{\text{Heat}} Na_2CO_3(s) + CO_2(g) + H_2O(g)$

c. $KClO_3(s) \xrightarrow{\text{Heat}} KCl(s) + O_2(g)$

d. $Eu(s) + HF(g) \rightarrow EuF_3(s) + H_2(g)$

87. Silicon is produced for the chemical and electronics industries by the following reactions. Give the balanced equation for each reaction.

a. $SiO_2(s) + C(s) \xrightarrow[\text{arc furnace}]{\text{Electric}} Si(s) + CO(g)$

b. Silicon tetrachloride is reacted with very pure magnesium, producing silicon and magnesium chloride.

c. $Na_2SiF_6(s) + Na(s) \rightarrow Si(s) + NaF(s)$

88. Glass is a mixture of several compounds, but a major constituent of most glass is calcium silicate, $CaSiO_3$. Glass can be etched by treatment with hydrofluoric acid; HF attacks the calcium silicate of the glass, producing gaseous and water-soluble products (which can be removed by washing the glass). For example, the volumetric glassware in chemistry laboratories is often graduated by using this process. Balance the following equation for the reaction of hydrofluoric acid with calcium silicate.

$$CaSiO_3(s) + HF(aq) \rightarrow CaF_2(aq) + SiF_4(g) + H_2O(l)$$

Reaction Stoichiometry

89. Over the years, the thermite reaction has been used for welding railroad rails, in incendiary bombs, and to ignite solid-fuel rocket motors. The reaction is

$$Fe_2O_3(s) + 2Al(s) \longrightarrow 2Fe(l) + Al_2O_3(s)$$

What masses of iron(III) oxide and aluminum must be used to produce 15.0 g iron? What is the maximum mass of aluminum oxide that could be produced?

90. The reaction between potassium chlorate and red phosphorus takes place when you strike a match on a matchbox. If you were to react 52.9 g of potassium chlorate ($KClO_3$) with excess red phosphorus, what mass of tetraphosphorus decaoxide (P_4O_{10}) would be produced?

$$KClO_3(s) + P_4(s) \longrightarrow P_4O_{10}(s) + KCl(s) \quad \text{(unbalanced)}$$

91. The reusable booster rockets of the U.S. space shuttle employ a mixture of aluminum and ammonium perchlorate for fuel. A possible equation for this reaction is

$$3Al(s) + 3NH_4ClO_4(s) \longrightarrow Al_2O_3(s) + AlCl_3(s) + 3NO(g) + 6H_2O(g)$$

What mass of NH_4ClO_4 should be used in the fuel mixture for every kilogram of Al?

92. One of relatively few reactions that takes place directly between two solids at room temperature is

$$Ba(OH)_2 \cdot 8H_2O(s) + NH_4SCN(s) \longrightarrow Ba(SCN)_2(s) + H_2O(l) + NH_3(g)$$

In this equation, the $\cdot\, 8H_2O$ in $Ba(OH)_2 \cdot 8H_2O$ indicates the presence of eight water molecules. This compound is called barium hydroxide octahydrate.

a. Balance the equation.

b. What mass of ammonium thiocyanate (NH_4SCN) must be used if it is to react completely with 6.5 g barium hydroxide octahydrate?

93. Bacterial digestion is an economical method of sewage treatment. The reaction

$$5CO_2(g) + 55NH_4^+(aq) + 76O_2(g) \xrightarrow{\text{bacteria}} \underbrace{C_5H_7O_2N(s)}_{\text{bacterial tissue}} + 54NO_2^-(aq) + 52H_2O(l) + 109H^+(aq)$$

is an intermediate step in the conversion of the nitrogen in organic compounds into nitrate ions. What mass of bacterial tissue is produced in a treatment plant for every 1.0×10^4 kg of wastewater containing 3.0% NH_4^+ ions by mass? Assume that 95% of the ammonium ions are consumed by the bacteria.

94. Phosphorus can be prepared from calcium phosphate by the following reaction:

$$2Ca_3(PO_4)_2(s) + 6SiO_2(s) + 10C(s) \longrightarrow 6CaSiO_3(s) + P_4(s) + 10CO(g)$$

Phosphorite is a mineral that contains $Ca_3(PO_4)_2$ plus other non-phosphorus-containing compounds. What is the maximum amount of P_4 that can be produced from 1.0 kg of phosphorite if the phorphorite sample is 75% $Ca_3(PO_4)_2$ by mass? Assume an excess of the other reactants.

95. Aspirin ($C_9H_8O_4$) is synthesized by reacting salicylic acid ($C_7H_6O_3$) with acetic anhydride ($C_4H_6O_3$). The balanced equation is

$$C_7H_6O_3 + C_4H_6O_3 \longrightarrow C_9H_8O_4 + HC_2H_3O_2$$

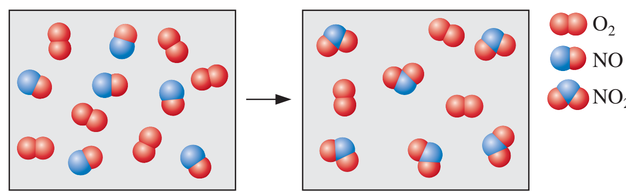
a. What mass of acetic anhydride is needed to completely consume 1.00×10^2 g salicylic acid?

b. What is the maximum mass of aspirin (the theoretical yield) that could be produced in this reaction?

96. The space shuttle environmental control system handles excess CO_2 (which the astronauts breathe out; it is 4.0% by mass of exhaled air) by reacting it with lithium hydroxide, LiOH, pellets to form lithium carbonate, Li_2CO_3, and water. If there are 7 astronauts on board the shuttle, and each exhales 20. L of air per minute, how long could clean air be generated if there were 25,000 g of LiOH pellets available for each shuttle mission? Assume the density of air is 0.0010 g/mL.

Limiting Reactants and Percent Yield

97. Consider the reaction between NO(g) and $O_2(g)$ represented below.



What is the balanced equation for this reaction and what is the limiting reactant?

98. Consider the following reaction:

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$

If a container were to have 10 molecules of O_2 and 10 molecules of NH_3 initially, how many total molecules (reactants plus products) would be present in the container after this reaction goes to completion?

99. Hydrogen peroxide is used as a cleaning agent in the treatment of cuts and abrasions for several reasons. It is an oxidizing agent that can directly kill many microorganisms; it decomposes upon contact with blood, releasing elemental oxygen gas (which inhibits the growth of anaerobic microorganisms); and it foams upon contact with blood, which provides a cleansing action. In the laboratory, small quantities of hydrogen peroxide can be prepared by the action of an acid on an alkaline earth metal peroxide, such as barium peroxide:

$$BaO_2(s) + 2HCl(aq) \longrightarrow H_2O_2(aq) + BaCl_2(aq)$$

What mass of hydrogen peroxide should result when 1.50 g of barium peroxide is treated with 25.0 mL of hydrochloric acid solution containing 0.0272 g of HCl per mL? What mass of which reagent is left unreacted?

100. Consider the following unbalanced equation:

$$Ca_3(PO_4)_2(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + H_3PO_4(aq)$$

What masses of calcium sulfate and phosphoric acid can be produced from the reaction of 1.0 kg calcium phosphate with 1.0 kg concentrated sulfuric acid (98% H_2SO_4 by mass)?

101. Hydrogen cyanide is produced industrially from the reaction of gaseous ammonia, oxygen, and methane:

$$2NH_3(g) + 3O_2(g) + 2CH_4(g) \longrightarrow 2HCN(g) + 6H_2O(g)$$

If 5.00×10^3 kg each of NH_3, O_2, and CH_4 are reacted, what mass of HCN and of H_2O will be produced, assuming 100% yield?

102. Acrylonitrile (C_3H_3N) is the starting material for many synthetic carpets and fabrics. It is produced by the following reaction.

$$2C_3H_6(g) + 2NH_3(g) + 3O_2(g) \longrightarrow 2C_3H_3N(g) + 6H_2O(g)$$

If 15.0 g C_3H_6, 10.0 g O_2, and 5.00 g NH_3 are reacted, what mass of acrylonitrile can be produced, assuming 100% yield?

103. A student prepared aspirin in a laboratory experiment using the reaction in Exercise 95. The student reacted 1.50 g salicylic acid with 2.00 g acetic anhydride. The yield was 1.50 g aspirin. Calculate the theoretical yield and the percent yield for this experiment.

104. DDT, an insecticide harmful to fish, birds, and humans, is produced by the following reaction:

$$\underset{\text{chlorobenzene}}{2C_6H_5Cl} + \underset{\text{chloral}}{C_2HOCl_3} \longrightarrow \underset{\text{DDT}}{C_{14}H_9Cl_5} + H_2O$$

In a government lab, 1142 g of chlorobenzene is reacted with 485 g of chloral.

a. What mass of DDT is formed?
b. Which reactant is limiting? Which is in excess?
c. What mass of the excess reactant is left over?
d. If the actual yield of DDT is 200.0 g, what is the percent yield?

105. Bornite (Cu_3FeS_3) is a copper ore used in the production of copper. When heated, the following reaction occurs:

$$2Cu_3FeS_3(s) + 7O_2(g) \longrightarrow 6Cu(s) + 2FeO(s) + 6SO_2(g)$$

If 2.50 metric tons of bornite is reacted with excess O_2 and the process has an 86.3% yield of copper, what mass of copper is produced?

106. Consider the following unbalanced reaction:

$$P_4(s) + F_2(g) \longrightarrow PF_3(g)$$

What mass of F_2 is needed to produce 120. g of PF_3 if the reaction has a 78.1% yield?

Additional Exercises

107. A given sample of a xenon fluoride compound contains molecules of the type XeF_n, where n is some whole number. Given that 9.03×10^{20} molecules of XeF_n weighs 0.368 g, determine the value for n in the formula.

108. Many cereals are made with high moisture content so that the cereal can be formed into various shapes before it is dried. A cereal product containing 58% H_2O by mass is produced at the rate of 1000. kg/h. What mass of water must be evaporated per hour if the final product contains only 20.% water?

109. Consider the reaction

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

Identify the limiting reagent in each of the reaction mixtures given below:

a. 50 molecules of H_2 and 25 molecules of O_2
b. 100 molecules of H_2 and 40 molecules of O_2
c. 100 molecules of H_2 and 100 molecules of O_2
d. 0.50 mol H_2 and 0.75 mol O_2
e. 0.80 mol H_2 and 0.75 mol O_2
f. 1.0 g H_2 and 0.25 mol O_2
g. 5.00 g H_2 and 56.00 g O_2

110. Some bismuth tablets, a medication used to treat upset stomachs, contain 262 mg of bismuth subsalicylate, $C_7H_5BiO_4$, per tablet. Assuming two tablets are digested, calculate the mass of bismuth consumed.

111. The empirical formula of styrene is CH; the molar mass of styrene is 104.14 g/mol. What number of H atoms are present in a 2.00-g sample of styrene?

112. Terephthalic acid is an important chemical used in the manufacture of polyesters and plasticizers. It contains only C, H, and O. Combustion of 19.81 mg terephthalic acid produces 41.98 mg CO_2 and 6.45 mg H_2O. If 0.250 mol of terephthalic acid has a mass of 41.5 g, determine the molecular formula for terephthalic acid.

113. A sample of a hydrocarbon (a compound consisting of only carbon and hydrogen) contains 2.59×10^{23} atoms of hydrogen and is 17.3% hydrogen by mass. If the molar mass of the hydrocarbon is between 55 and 65 g/mol, what amount (moles) of compound are present, and what is the mass of the sample?

114. A binary compound between an unknown element E and hydrogen contains 91.27% E and 8.73% H by mass. If the formula of the compound is E_3H_8, calculate the atomic mass of E.

115. A 0.755-g sample of hydrated copper(II) sulfate

$$CuSO_4 \cdot xH_2O$$

was heated carefully until it had changed completely to anhydrous copper(II) sulfate ($CuSO_4$) with a mass of 0.483 g. Determine the value of x. [This number is called the *number of waters of hydration* of copper(II) sulfate. It specifies the number of water molecules per formula unit of $CuSO_4$ in the hydrated crystal.]

116. ABS plastic is a tough, hard plastic used in applications requiring shock resistance. The polymer consists of three monomer units: acrylonitrile (C_3H_3N), butadiene (C_4H_6), and styrene (C_8H_8).

a. A sample of ABS plastic contains 8.80% N by mass. It took 0.605 g of Br_2 to react completely with a 1.20-g sample of ABS plastic. Bromine reacts 1:1 (by moles) with the butadiene molecules in the polymer and nothing else. What is the percent by mass of acrylonitrile and butadiene in this polymer?
b. What are the relative numbers of each of the monomer units in this polymer?

117. A sample of LSD (D-lysergic acid diethylamide, $C_{24}H_{30}N_3O$) is added to some table salt (sodium chloride) to form a mixture. Given that a 1.00-g sample of the mixture undergoes combustion to produce 1.20 g of CO_2, what is the mass percentage of LSD in the mixture?

118. Methane (CH_4) is the main component of marsh gas. Heating methane in the presence of sulfur produces carbon disulfide and hydrogen sulfide as the only products.

a. Write the balanced chemical equation for the reaction of methane and sulfur.

b. Calculate the theoretical yield of carbon disulfide when 120. g of methane is reacted with an equal mass of sulfur.

119. A potential fuel for rockets is a combination of B_5H_9 and O_2. The two react according to the following balanced equation:

$$2B_5H_9(l) + 12O_2(g) \longrightarrow 5B_2O_3(s) + 9H_2O(g)$$

If one tank in a rocket holds 126 g of B_5H_9 and another tank holds 192 g of O_2, what mass of water can be produced when the entire contents of each tank react together?

120. Silver sulfadiazine burn-treating cream creates a barrier against bacterial invasion and releases antimicrobial agents directly into the wound. If 25.0 g of Ag_2O is reacted with 50.0 g of $C_{10}H_{10}N_4SO_2$, what mass of silver sulfadiazine, $AgC_{10}H_9N_4SO_2$, can be produced, assuming 100% yield?

$$Ag_2O(s) + 2C_{10}H_{10}N_4SO_2(s) \longrightarrow 2AgC_{10}H_9N_4SO_2(s) + H_2O(l)$$

121. An iron ore sample contains Fe_2O_3 plus other impurities. A 752-g sample of impure iron ore is heated with excess carbon, producing 453 g of pure iron by the following reaction:

$$Fe_2O_3(s) + 3C(s) \longrightarrow 2Fe(s) + 3CO(g)$$

What is the mass percent of Fe_2O_3 in the impure iron ore sample? Assume that Fe_2O_3 is the only source of iron and that the reaction is 100% efficient.

122. Commercial brass, an alloy of Zn and Cu, reacts with hydrochloric acid as follows:

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

(Cu does not react with HCl.) When 0.5065 g of a certain brass alloy is reacted with excess HCl, 0.0985 g $ZnCl_2$ is eventually isolated.

a. What is the composition of the brass by mass?

b. How could this result be checked without changing the above procedure?

123. Vitamin A has a molar mass of 286.4 g/mol and a general molecular formula of C_xH_yE, where E is an unknown element. If vitamin A is 83.86% C and 10.56% H by mass, what is the molecular formula of vitamin A?

Challenge Problems

124. Natural rubidium has the average mass of 85.4678 and is composed of isotopes ^{85}Rb (mass = 84.9117) and ^{87}Rb. The ratio of atoms $^{85}Rb/^{87}Rb$ in natural rubidium is 2.591. Calculate the mass of ^{87}Rb.

125. A compound contains only carbon, hydrogen, nitrogen, and oxygen. Combustion of 0.157 g of the compound produced 0.213 g CO_2 and 0.0310 g H_2O. In another experiment, it is found that 0.103 g of the compound produces 0.0230 g NH_3. What is the empirical formula of the compound? *Hint:* Combustion involves reacting with excess O_2. Assume that all the carbon ends up in CO_2 and all the hydrogen ends up in H_2O. Also assume that all the nitrogen ends up in the NH_3 in the second experiment.

126. Nitric acid is produced commercially by the Ostwald process, represented by the following equations:

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$
$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$
$$3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(aq) + NO(g)$$

What mass of NH_3 must be used to produce 1.0×10^6 kg HNO_3 by the Ostwald process? Assume 100% yield in each reaction and assume that the NO produced in the third step is not recycled.

127. Consider a 5.430-g mixture of FeO and Fe_3O_4. You react this mixture with an excess of oxygen to form 5.779 g Fe_2O_3. Calculate the percent by mass of FeO in the original mixture.

128. A 9.780-g gaseous mixture contains ethane (C_2H_6) and propane (C_3H_8). Complete combustion to form carbon dioxide and water requires 1.120 mol of oxygen. Calculate the mass percent of ethane in the original mixture.

129. Zinc and magnesium metal each react with hydrochloric acid to make chloride salts of the respective metals, and hydrogen gas. A 10.00-g mixture of zinc and magnesium produces 0.5171 g of hydrogen gas upon being mixed with an excess of hydrochloric acid. Determine the percent magnesium by mass in the original mixture.

130. A 2.077-g sample of an element, which has an atomic mass between 40 and 55, reacts with oxygen to form 3.708 g of an oxide. Determine the formula of the oxide (and identify the element).

131. Consider a gaseous binary compound with a molar mass of 62.09 g/mol. When 1.39 g of this compound is completely burned in excess oxygen, 1.21 g of water is formed. Determine the formula of the compound. Assume water is the only product that contains hydrogen.

132. A 2.25-g sample of scandium metal is reacted with excess hydrochloric acid to produce 0.1502 g hydrogen gas. What is the formula of the scandium chloride produced in the reaction?

133. In the production of printed circuit boards for the electronics industry, a 0.60-mm layer of copper is laminated onto an insulating plastic board. Next, a circuit pattern made of a chemically resistant polymer is printed on the board. The unwanted copper is removed by chemical etching, and the protective polymer is finally removed by solvents. One etching reaction is

$$Cu(NH_3)_4Cl_2(aq) + 4NH_3(aq) + Cu(s) \longrightarrow 2Cu(NH_3)_4Cl(aq)$$

A plant needs to manufacture 10,000 printed circuit boards, each 8.0×16.0 cm in area. An average of 80.% of the copper is removed from each board (density of copper = 8.96 g/cm^3). What masses of $Cu(NH_3)_4Cl_2$ and NH_3 are needed to do this? Assume 100% yield.

134. The aspirin substitute, acetaminophen ($C_8H_9O_2N$), is produced by the following three-step synthesis:

I. $C_6H_5O_3N(s) + 3H_2(g) + HCl(aq) \longrightarrow C_6H_8ONCl(s) + 2H_2O(l)$

II. $C_6H_8ONCl(s) + NaOH(aq) \longrightarrow C_6H_7ON(s) + H_2O(l) + NaCl(aq)$

III. $C_6H_7ON(s) + C_4H_6O_3(l) \longrightarrow C_8H_9O_2N(s) + HC_2H_3O_2(l)$

The first two reactions have percent yields of 87% and 98% by mass, respectively. The overall reaction yields 3 mol of acetaminophen product for every 4 mol of $C_6H_5O_3N$ reacted.

a. What is the percent yield by mass for the overall process?

b. What is the percent yield by mass of step III?

135. An element X forms both a dichloride (XCl_2) and a tetrachloride (XCl_4). Treatment of 10.00 g XCl_2 with excess chlorine forms 12.55 g XCl_4. Calculate the atomic mass of X, and identify X.

136. When $M_2S_3(s)$ is heated in air, it is converted to $MO_2(s)$. A 4.000-g sample of $M_2S_3(s)$ shows a decrease in mass of 0.277 g when it is heated in air. What is the average atomic mass of M?

137. When aluminum metal is heated with an element from Group 6A of the periodic table, an ionic compound forms. When the experiment is performed with an unknown Group 6A element, the product is 18.56% Al by mass. What is the formula of the compound?

138. A sample of a mixture containing only sodium chloride and potassium chloride has a mass of 4.000 g. When this sample is dissolved in water and excess silver nitrate is added, a white solid (silver chloride) forms. After filtration and drying, the solid silver chloride has the mass 8.5904 g. Calculate the mass percent of each mixture component.

139. Ammonia reacts with O_2 to form either $NO(g)$ or $NO_2(g)$ according to these unbalanced equations:

$$NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$$
$$NH_3(g) + O_2(g) \longrightarrow NO_2(g) + H_2O(g)$$

In a certain experiment 2.00 mol of $NH_3(g)$ and 10.00 mol of $O_2(g)$ are contained in a closed flask. After the reaction is complete, 6.75 mol of $O_2(g)$ remains. Calculate the number of moles of $NO(g)$ in the product mixture: (*Hint:* You cannot do this problem by adding the balanced equations, because you cannot assume that the two reactions will occur with equal probability.)

140. You take 1.00 g of an aspirin tablet (a compound consisting solely of carbon, hydrogen, and oxygen), burn it in air, and collect 2.20 g CO_2 and 0.400 g H_2O. You know that the molar mass of aspirin is between 170 and 190 g/mol. Reacting 1 mole of salicylic acid with 1 mole of acetic anhydride ($C_4H_6O_3$) gives you 1 mole of aspirin and 1 mole of acetic acid ($C_2H_4O_2$). Use this information to determine the molecular formula of salicylic acid.

Integrative Problems

These problems require the integration of multiple concepts to find the solutions.

141. With the advent of techniques such as scanning tunneling microscopy, it is now possible to "write" with individual atoms by manipulating and arranging atoms on an atomic surface.

a. If an image is prepared by manipulating iron atoms and their total mass is 1.05×10^{-20} g, what number of iron atoms were used?

b. If the image is prepared on a platinum surface that is exactly 20 platinum atoms high and 14 platinum atoms wide, what is the mass (grams) of the atomic surface?

c. If the atomic surface were changed to ruthenium atoms and the same surface mass as determined in part b is used, what number of ruthenium atoms is needed to construct the surface?

142. Tetrodotoxin is a toxic chemical found in fugu pufferfish, a popular but rare delicacy in Japan. This compound has a LD_{50} (the amount of substance that is lethal to 50.% of a population sample) of 10. μg per kg of body mass. Tetrodotoxin is 41.38% carbon by mass, 13.16% nitrogen by mass, and 5.37% hydrogen by mass, with the remaining amount consisting of oxygen. What is the empirical formula of tetrodotoxin? If three molecules of tetrodotoxin has a mass of 1.59×10^{-21} g, what is the molecular formula of tetrodotoxin? What number of molecules of tetrodotoxin would be the LD_{50} dosage for a person weighing 165 lb?

143. An ionic compound MX_3 is prepared according to the following unbalanced chemical equation.

$$M + X_2 \longrightarrow MX_3$$

A 0.105-g sample of X_2 contains 8.92×10^{20} molecules. The compound MX_3 consists of 54.47% X by mass. What are the identities of M and X, and what is the correct name for MX_3? Starting with 1.00 g each of M and X_2, what mass of MX_3 can be prepared?

144. The compound As_2I_4 is synthesized by reaction of arsenic metal with arsenic triiodide. If a solid cubic block of arsenic (d = 5.72 g/cm^3) that is 3.00 cm on edge is allowed to react with 1.01×10^{24} molecules of arsenic triiodide, how much As_2I_4 can be prepared? If the percent yield of As_2I_4 was 75.6%, what mass of As_2I_4 was actually isolated?

Marathon Problems

These problems are designed to incorporate several concepts and techniques into one situation. Marathon Problems can be used in class by groups of students to help facilitate problem-solving skills.

***145.** From the information below, determine the mass of substance *C* that will be formed if 45.0 grams of substance *A* reacts with 23.0 grams of substance *B*. (Assume that the reaction between *A* and *B* goes to completion.)

a. Substance *A* is a gray solid that consists of an alkaline earth metal and carbon (37.5% by mass). It reacts with substance *B* to produce substances *C* and *D*. Forty million trillion formula units of *A* have a mass of 4.26 milligrams.

b. 47.9 grams of substance *B* contains 5.36 grams of hydrogen and 42.5 grams of oxygen.

c. When 10.0 grams of *C* is burned in excess oxygen, 33.8 grams of carbon dioxide and 6.92 grams of water are produced. A mass spectrum of substance *C* shows a parent molecular ion with a mass-to-charge ratio of 26.

d. Substance *D* is the hydroxide of the metal in substance *A*.

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146. Consider the following balanced chemical equation:

$$A + 5B \longrightarrow 3C + 4D$$

a. Equal masses of A and B are reacted. Complete each of the following with either "A is the limiting reactant because ______"; "B is the limiting reactant because ______"; or "we cannot determine the limiting reactant because ______".

i. If the molar mass of A is greater than the molar mass of B, then

ii. If the molar mass of B is greater than the molar mass of A, then

b. The products of the reaction are carbon dioxide (C) and water (D). Compound A has the same molar mass as carbon dioxide. Compound B is a diatomic molecule. Identify compound B and support your answer.

c. Compound A is a hydrocarbon that is 81.71% carbon by mass. Determine its empirical and molecular formulas.

Get help understanding core concepts and visualizing molecular-level interactions, and practice problem solving, by visiting the Online Study Center at **college.hmco.com/PIC/zumdahl7e.**