CHAPTER THREE

Stoichiometry of Formulas and Equations

Key Principles

- The *mole* (*mol*) is the standard unit for *amount of substance* and consists of *Avogadro's number* (6.022×10^{23}) of atoms, molecules, or ions. It has the same numerical value in grams as a single entity of the substance has in atomic mass units; for example, 1 molecule of H_2O weighs 18.02 amu and 1 mol of H_2O weighs 18.02 g. Therefore, if the amount of a substance is expressed in moles, we know the number of entities in a given mass of it.
- The subscripts in a *chemical formula* provide quantitative information about the amounts of each element in a mole of compound. In an *empirical formula,* the subscripts show the *relative* number of moles of each element; in a *molecular formula,* they show the *actual* number. Isomers are different compounds with the same molecular formula.
- In a *balanced equation,* formulas preceded by integer *balancing coefficients* are used to show the same numbers of each kind of atom on the left (*reactants*) as on the right (*products*) but in different combinations; we can therefore use the amount of one substance to calculate the amount of any other.
- During a typical reaction, one substance (the *limiting reactant*) is used up, so it limits the amount of product that can form; the other reactant(s) are in excess. The *theoretical yield,* the amount based on the balanced equation, is never obtained in the lab because of competing *side reactions* and physical losses.
- For reactions in solution, we determine amounts of substances from their *concentration* (*molarity*) and volume. To dilute a solution, we add *solvent,* which lowers the amount of *solute* dissolved in each unit volume.

Outline

Concepts & Skills to Review Before You Study This Chapter

- isotopes and atomic mass (Section 2.5)
- names and formulas of compounds (Section 2.8)
- molecular mass of a compound (Section 2.8)
- empirical and molecular formulas (Section 2.8)
- mass laws in chemical reactions (Section 2.2)

Chemistry is a practical science. Just imagine how useful it could be to determine the formula of a compound from the masses of its elements or to predict the amounts of substances consumed and produced in a reaction. Suppose you are a polymer chemist preparing a new plastic: how much of this new material will a given polymerization reaction yield? Or suppose you're a chemical engineer studying rocket engine thrust: what amount of exhaust gases will a test of this fuel mixture produce? Perhaps you are on a team of environmental chemists examining coal samples: what quantity of air pollutants will this sample release when burned? Or, maybe you're a biomedical researcher who has extracted a new cancer-preventing substance from a tropical plant: what is its formula, and what quantity of metabolic products will establish a safe dosage level? You can answer such questions and countless others like them with a knowledge of **stoichiometry** (pronounced "stoy-key-AHM-uh-tree"; from the Greek *stoicheion,* "element or part," and *metron,* "measure"), the study of the quantitative aspects of chemical formulas and reactions.

3.1 THE MOLE

All the ideas and skills discussed in this chapter depend on an understanding of the *mole* concept, so let's begin there. In daily life, we typically measure things out by counting or by weighing, with the choice based on convenience. It is more convenient to weigh beans or rice than to count individual pieces, and it is more convenient to count eggs or pencils than to weigh them. To measure such things, we use mass units (a kilogram of rice) or counting units (a dozen pencils). Similarly, daily life in the laboratory involves measuring substances. However, an obvious problem arises when we try to do this. The atoms, ions, molecules, or formula units are the entities that react with one another, so we would like to know the numbers of them that we mix together. But, how can we possibly count entities that are so small? To do this, chemists have devised a unit called the mole to *count chemical entities by weighing them.*

Defining the Mole

The **mole** (abbreviated **mol**) is the SI unit for amount of substance. It is defined as *the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12.* This number is called **Avogadro's number,** in honor of the 19th-century Italian physicist Amedeo Avogadro, and as you can tell from the definition, it is enormous:*

One mole (1 mol) contains 6.022×10^{23} entities (to four significant figures) **(3.1)**

Thus,

1 mol of carbon-12	contains	6.022×10^{23} carbon-12 atoms
1 mol of H_2O	contains	6.022×10^{23} H_2O molecules
1 mol of NaCl	contains	6.022×10^{23} NaCl formula units

However, the mole is not just a counting unit like the dozen, which specifies only the *number* of objects. The definition of the mole specifies the *number* of objects in a fixed *mass* of substance. Therefore, *1 mole of a substance represents a fixed number of chemical entities* ***and*** *has a fixed mass.* To see why this is important, consider the marbles in Figure 3.1A, which we'll use as

*A mole of any ordinary object is a staggering amount: a mole of periods (.) lined up side by side would equal the radius of our galaxy; a mole of marbles stacked tightly together would cover the United States 70 miles deep. However, atoms and molecules are not ordinary objects: a mole of water molecules (about 18 mL) can be swallowed in one gulp!

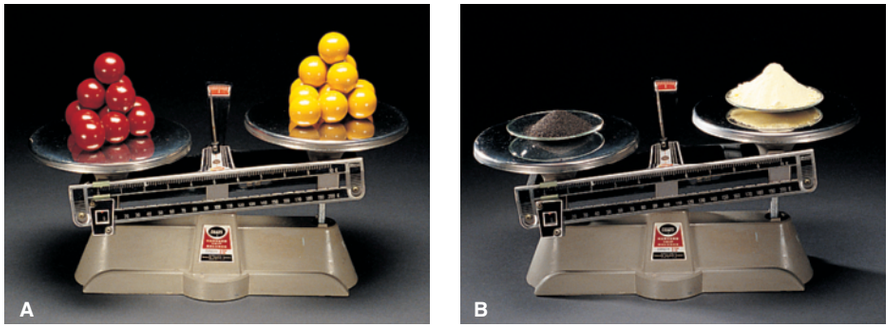


Figure 3.1 Counting objects of fixed relative mass. A, If marbles had a fixed mass, we could count them by weighing them. Each red marble weighs 7 g, and each yellow marble weighs 4 g, so 84 g of red marbles and 48 g of yellow marbles each contains 12 marbles. Equal numbers of the two types of marbles always have a 7:4 mass ratio of red:yellow marbles. **B,** Because atoms of a substance have a fixed mass, we can weigh the substance to count the atoms; 55.85 g of Fe (left pan) and 32.07 g of S (right pan) each contains 6.022×10^{23} atoms (1 mol of atoms). Any two samples of Fe and S that contain equal numbers of atoms have a 55.85:32.07 mass ratio of Fe:S.

an analogy for atoms. Suppose you have large groups of red marbles and yellow marbles; each red marble weighs 7 g and each yellow marble weighs 4 g. Right away you know that there are 12 marbles in 84 g of red marbles or in 48 g of yellow marbles. Moreover, because one red marble weighs $\frac{7}{4}$ as much as one yellow marble, any given *number* of red and of yellow marbles always has this 7:4 *mass* ratio. By the same token, any given *mass* of red and of yellow marbles always has a 4:7 *number* ratio. For example, 280 g of red marbles contains 40 marbles, and 280 g of yellow marbles contains 70 marbles. As you can see, the fixed masses of the marbles allow you to count marbles by weighing them.

Atoms have fixed masses, too. Let's recall a key point from Chapter 2: the atomic mass of an element (which appears on the periodic table) is the weighted average of the masses of its naturally occurring isotopes. That is, all iron (Fe) atoms have an atomic mass of 55.85 amu, all sulfur (S) atoms have an atomic mass of 32.07 amu, and so forth.

The central relationship between the mass of one atom and the mass of 1 mole of those atoms is that *the atomic mass of an element expressed in* ***amu*** *is numerically the same as the mass of 1 mole of atoms of the element expressed in* ***grams.*** You can see this from the definition of the mole, which referred to the number of atoms in "**12** g of carbon-**12.**" Thus,

1 Fe atom	has a mass of	55.85 amu	and	1 mol of Fe atoms	has a mass of	55.85 g
1 S atom	has a mass of	32.07 amu	and	1 mol of S atoms	has a mass of	32.07 g
1 O atom	has a mass of	16.00 amu	and	1 mol of O atoms	has a mass of	16.00 g
1 O_2 molecule	has a mass of	32.00 amu	and	1 mol of O_2 molecules	has a mass of	32.00 g

Moreover, because of their fixed atomic masses, we know that 55.85 g of Fe atoms and 32.07 g of S atoms each contains 6.022×10^{23} atoms. As with marbles of fixed mass, one Fe atom weighs $\frac{55.85}{32.07}$ as much as one S atom, and 1 mol of Fe atoms weighs $\frac{55.85}{32.07}$ as much as 1 mol of S atoms (Figure 3.1B).

A similar relationship holds for compounds: *the molecular mass (or formula mass) of a compound expressed in* ***amu*** *is numerically the same as the mass of 1 mole of the compound expressed in* ***grams.*** Thus, for example,

1 molecule of H_2O	has a mass of	18.02 amu	and	1 mol of H_2O (6.022×10^{23} molecules)	has a mass of	18.02 g
1 formula unit of NaCl	has a mass of	58.44 amu	and	1 mol of NaCl (6.022×10^{23} formula units)	has a mass of	58.44 g

To summarize the two key points about the usefulness of the mole concept:

- The mole maintains the *same mass relationship* between macroscopic samples as exists between individual chemical entities.
- The mole relates the *number* of chemical entities to the *mass* of a sample of those entities.

A grocer cannot obtain 1 dozen eggs by weighing them because eggs vary in mass. But a chemist *can* obtain 1 mol of copper atoms (6.022×10^{23} atoms) simply

Figure 3.2 One mole of some familiar substances. One mole of a substance is the amount that contains 6.022×10^{23} atoms, molecules, or formula units. From left to right: 1 mol (172.19 g) of writing chalk (calcium sulfate dihydrate), 1 mol (32.00 g) of gaseous O_2, 1 mol (63.55 g) of copper, and 1 mol (18.02 g) of liquid H_2O.

by weighing 63.55 g of copper. Figure 3.2 shows 1 mol of some familiar elements and compounds.

Molar Mass

The **molar mass ($\mathcal{M}$)** of a substance is the mass per mole of its entities (atoms, molecules, or formula units). Thus, molar mass has units of grams per mole (g/mol). The periodic table is indispensable for calculating the molar mass of a substance. Here's how the calculations are done:

1. **Elements.** You find the molar mass of an element simply by looking up its atomic mass in the periodic table and then noting whether the element occurs naturally as individual atoms or as molecules.

- *Monatomic elements.* For elements that occur as individual atoms, the molar mass is the numerical value from the periodic table expressed in units of grams per mole.* Thus, the molar mass of neon is 20.18 g/mol, the molar mass of iron is 55.85 g/mol, and the molar mass of gold is 197.0 g/mol.
- *Molecular elements.* For elements that occur as molecules, you must know the molecular formula to determine the molar mass. For example, oxygen exists normally in air as diatomic molecules, so the molar mass of O_2 molecules is twice that of O atoms:

$$\text{Molar mass } (\mathcal{M}) \text{ of } O_2 = 2 \times \mathcal{M} \text{ of O} = 2 \times 16.00 \text{ g/mol} = 32.00 \text{ g/mol}$$

The most common form of sulfur exists as octatomic molecules, S_8:

$$\mathcal{M} \text{ of } S_8 = 8 \times \mathcal{M} \text{ of S} = 8 \times 32.07 \text{ g/mol} = 256.6 \text{ g/mol}$$

2. **Compounds.** *The molar mass of a compound is the sum of the molar masses of the atoms of the elements in the formula.* For example, the formula of sulfur dioxide (SO_2) tells us that 1 mol of SO_2 molecules contains 1 mol of S atoms and 2 mol of O atoms:

$$\begin{aligned}\mathcal{M} \text{ of } SO_2 &= \mathcal{M} \text{ of S} + (2 \times \mathcal{M} \text{ of O}) = 32.07 \text{ g/mol} + (2 \times 16.00 \text{ g/mol})\\ &= 64.07 \text{ g/mol}\end{aligned}$$

Similarly, for ionic compounds, such as potassium sulfide (K_2S), we have

$$\begin{aligned}\mathcal{M} \text{ of } K_2S &= (2 \times \mathcal{M} \text{ of K}) + \mathcal{M} \text{ of S} = (2 \times 39.10 \text{ g/mol}) + 32.07 \text{ g/mol}\\ &= 110.27 \text{ g/mol}\end{aligned}$$

A key point to note is that *the subscripts in a formula refer to individual atoms (or ions), as well as to moles of atoms (or ions).* Table 3.1 presents this idea for glucose ($C_6H_{12}O_6$), the essential sugar in energy metabolism.

*The mass value in the periodic table has no units because it is a *relative* atomic mass, given by the atomic mass (in amu) divided by 1 amu ($\frac{1}{12}$ mass of one ^{12}C atom in amu):

$$\text{Relative atomic mass} = \frac{\text{atomic mass } (\cancel{\text{amu}})}{\frac{1}{12} \text{ mass of } ^{12}\text{C } (\cancel{\text{amu}})}$$

Therefore, you use the same number for the atomic mass (weighted average mass of one atom in amu) and the molar mass (mass of 1 mole of atoms in grams).

Table 3.1 Information Contained in the Chemical Formula of Glucose, $C_6H_{12}O_6$ ($\mathcal{M}$ = 180.16 g/mol)

	Carbon (C)	Hydrogen (H)	Oxygen (O)
Atoms/molecule of compound	6 atoms	12 atoms	6 atoms
Moles of atoms/mole of compound	6 mol of atoms	12 mol of atoms	6 mol of atoms
Atoms/mole of compound	$6(6.022\times10^{23})$ atoms	$12(6.022\times10^{23})$ atoms	$6(6.022\times10^{23})$ atoms
Mass/molecule of compound	6(12.01 amu) = 72.06 amu	12(1.008 amu) = 12.10 amu	6(16.00 amu) = 96.00 amu
Mass/mole of compound	72.06 g	12.10 g	96.00 g

Interconverting Moles, Mass, and Number of Chemical Entities

One of the reasons the mole is such a convenient unit for laboratory work is that it allows you to calculate the mass or number of entities of a substance in a sample if you know the amount (number of moles) of the substance. Conversely, if you know the mass or number of entities of a substance, you can calculate the number of moles.

The molar mass, which expresses the equivalent relationship between 1 mole of a substance and its mass in grams, can be used as a conversion factor. We multiply by the molar mass of an element or compound ($\mathcal{M}$, in g/mol) to convert a given amount (in moles) to mass (in grams):

$$\text{Mass (g)} = \text{no. of moles} \times \frac{\text{no. of grams}}{1\ \text{mol}} \tag{3.2}$$

Or, we divide by the molar mass (multiply by $1/\mathcal{M}$) to convert a given mass (in grams) to amount (in moles):

$$\text{No. of moles} = \text{mass (g)} \times \frac{1\ \text{mol}}{\text{no. of grams}} \tag{3.3}$$

In a similar way, we use Avogadro's number, which expresses the equivalent relationship between 1 mole of a substance and the number of entities it contains, as a conversion factor. We multiply by Avogadro's number to convert amount of substance (in moles) to the number of entities (atoms, molecules, or formula units):

$$\text{No. of entities} = \text{no. of moles} \times \frac{6.022\times10^{23}\ \text{entities}}{1\ \text{mol}} \tag{3.4}$$

Or, we divide by Avogadro's number to do the reverse:

$$\text{No. of moles} = \text{no. of entities} \times \frac{1\ \text{mol}}{6.022\times10^{23}\ \text{entities}} \tag{3.5}$$

Converting Moles of Elements For problems involving mass-mole-number relationships of elements, keep these points in mind:

- To convert between amount (mol) and mass (g), use the molar mass ($\mathcal{M}$ in g/mol).
- To convert between amount (mol) and number of entities, use Avogadro's number (6.022×10^{23} entities/mol). For elements that occur as molecules, use the molecular formula to find atoms/mol.
- Mass and number of entities relate directly to number of moles, *not* to each other. Therefore, to convert between number of entities and mass, *first convert to number of moles.* For example, to find the number of atoms in a given mass,

$$\text{No. of atoms} = \text{mass } \cancel{(g)} \times \frac{1\ \cancel{\text{mol}}}{\text{no. } \cancel{\text{of grams}}} \times \frac{6.022\times10^{23}\ \text{atoms}}{1\ \cancel{\text{mol}}}$$

These relationships are summarized in Figure 3.3 and demonstrated in Sample Problem 3.1.

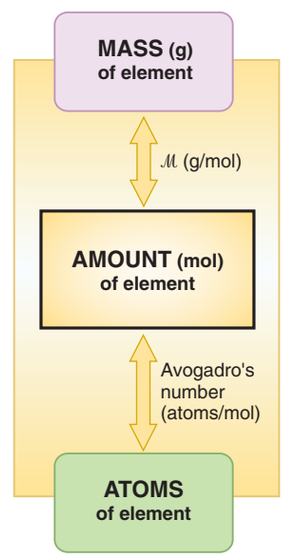
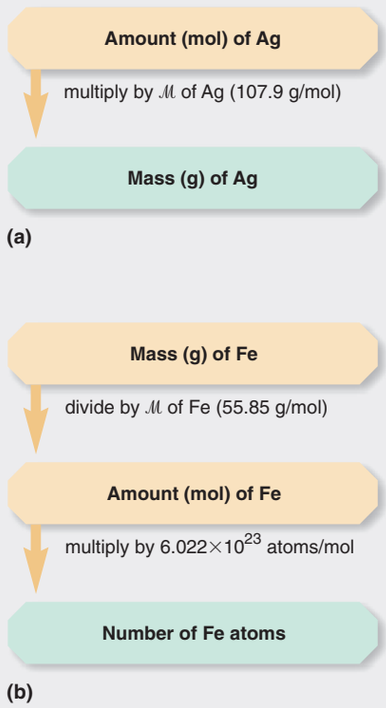


Figure 3.3 Summary of the mass-mole-number relationships for elements. The amount (mol) of an element is related to its mass (g) through the molar mass ($\mathcal{M}$ in g/mol) and to its number of atoms through Avogadro's number (6.022×10^{23} atoms/mol). For elements that occur as molecules, Avogadro's number gives *molecules* per mole.

SAMPLE PROBLEM 3.1 Calculating the Mass and Number of Atoms in a Given Number of Moles of an Element

Problem **(a)** Silver (Ag) is used in jewelry and tableware but no longer in U.S. coins. How many grams of Ag are in 0.0342 mol of Ag?
(b) Iron (Fe), the main component of steel, is the most important metal in industrial society. How many Fe atoms are in 95.8 g of Fe?



(a) Determining the mass (g) of Ag

Plan We know the number of moles of Ag (0.0342 mol) and have to find the mass (in g). To convert *moles* of Ag to *grams* of Ag, we multiply by the *molar mass* of Ag, which we find in the periodic table (see roadmap a).

Solution Converting from moles of Ag to grams:

$$\text{Mass (g) of Ag} = 0.0342\ \cancel{\text{mol Ag}} \times \frac{107.9\ \text{g Ag}}{1\ \cancel{\text{mol Ag}}} = 3.69\ \text{g Ag}$$

Check We rounded the mass to three significant figures because the number of moles has three. The units are correct. About 0.03 mol × 100 g/mol gives 3 g; the small mass makes sense because 0.0342 is a small fraction of a mole.

(b) Determining the number of Fe atoms

Plan We know the grams of Fe (95.8 g) and need the number of Fe atoms. We cannot convert directly from grams to atoms, so we first convert to moles by dividing grams of Fe by its molar mass. [This is the reverse of the step in part (a).] Then, we multiply number of moles by Avogadro's number to find number of atoms (see roadmap b).

Solution Converting from grams of Fe to moles:

$$\text{Moles of Fe} = 95.8\ \cancel{\text{g Fe}} \times \frac{1\ \text{mol Fe}}{55.85\ \cancel{\text{g Fe}}} = 1.72\ \text{mol Fe}$$

Converting from moles of Fe to number of atoms:

$$\text{No. of Fe atoms} = 1.72\ \cancel{\text{mol Fe}} \times \frac{6.022\times10^{23}\ \text{atoms Fe}}{1\ \cancel{\text{mol Fe}}}$$

$$= 10.4\times10^{23}\ \text{atoms Fe} = 1.04\times10^{24}\ \text{atoms Fe}$$

Check When we approximate the mass of Fe and the molar mass of Fe, we have ~100 g/(~50 g/mol) = 2 mol. Therefore, the number of atoms should be about twice Avogadro's number: $2(6\times10^{23}) = 1.2\times10^{24}$.

FOLLOW-UP PROBLEM 3.1 **(a)** Graphite is the crystalline form of carbon used in "lead" pencils. How many moles of carbon are in 315 mg of graphite?
(b) Manganese (Mn) is a transition element essential for the growth of bones. What is the mass in grams of 3.22×10^{20} Mn atoms, the number found in 1 kg of bone?

Converting Moles of Compounds Solving mass-mole-number problems involving compounds requires a very similar approach to the one for elements. We need the chemical formula to find the molar mass and to determine the moles of a given element in the compound. These relationships are shown in Figure 3.4, and an example is worked through in Sample Problem 3.2.

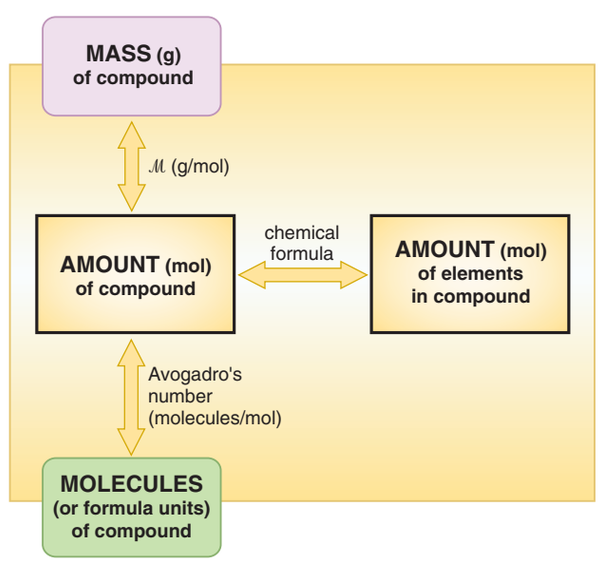
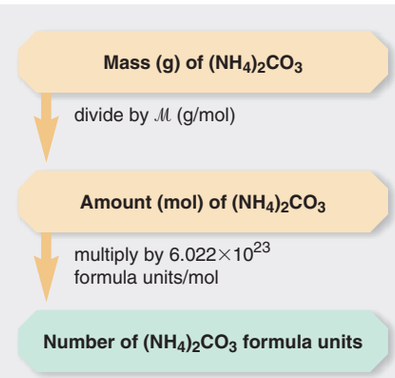


Figure 3.4 Summary of the mass-mole-number relationships for compounds. Moles of a compound are related to grams of the compound through the molar mass (ℳ in g/mol) and to the number of molecules (or formula units) through Avogadro's number (6.022×10^{23} molecules/mol). To find the number of molecules (or formula units) in a given mass, or vice versa, convert the information to moles first. With the chemical formula, you can calculate mass-mole-number information about each component element.

SAMPLE PROBLEM 3.2 Calculating the Moles and Number of Formula Units in a Given Mass of a Compound

Problem Ammonium carbonate is a white solid that decomposes with warming. Among its many uses, it is a component of baking powder, fire extinguishers, and smelling salts. How many formula units are in 41.6 g of ammonium carbonate?

Plan We know the mass of compound (41.6 g) and need to find the number of formula units. As we saw in Sample Problem 3.1(b), to convert grams to number of entities, we have to find number of moles first, so we must divide the grams by the molar mass ($\mathcal{M}$). For this, we need $\mathcal{M}$, so we determine the formula (see Table 2.5) and take the sum of the elements' molar masses. Once we have the number of moles, we multiply by Avogadro's number to find the number of formula units.



Solution The formula is $(NH_4)_2CO_3$. Calculating molar mass:

$$\begin{aligned}\mathcal{M} &= (2 \times \mathcal{M} \text{ of N}) + (8 \times \mathcal{M} \text{ of H}) + (1 \times \mathcal{M} \text{ of C}) + (3 \times \mathcal{M} \text{ of O})\\ &= (2 \times 14.01 \text{ g/mol}) + (8 \times 1.008 \text{ g/mol}) + 12.01 \text{ g/mol} + (3 \times 16.00 \text{ g/mol})\\ &= 96.09 \text{ g/mol}\end{aligned}$$

Converting from grams to moles:

$$\text{Moles of } (NH_4)_2CO_3 = 41.6 \cancel{\text{ g } (NH_4)_2CO_3} \times \frac{1 \text{ mol } (NH_4)_2CO_3}{96.09 \cancel{\text{ g } (NH_4)_2CO_3}} = 0.433 \text{ mol } (NH_4)_2CO_3$$

Converting from moles to formula units:

$$\begin{aligned}\text{Formula units of } (NH_4)_2CO_3 &= 0.433 \cancel{\text{ mol } (NH_4)_2CO_3}\\ &\quad \times \frac{6.022\times10^{23} \text{ formula units } (NH_4)_2CO_3}{1 \cancel{\text{ mol } (NH_4)_2CO_3}}\\ &= 2.61\times10^{23} \text{ formula units } (NH_4)_2CO_3\end{aligned}$$

Check The units are correct. The mass is less than half the molar mass ($\sim 42/96 < 0.5$), so the number of formula units should be less than half Avogadro's number ($\sim 2.6\times10^{23}/6.0\times10^{23} < 0.5$).

Comment A *common mistake* is to forget the subscript 2 outside the parentheses in $(NH_4)_2CO_3$, which would give a much lower molar mass.

FOLLOW-UP PROBLEM 3.2 Tetraphosphorus decaoxide reacts with water to form phosphoric acid, a major industrial acid. In the laboratory, the oxide is used as a drying agent.

(a) What is the mass (in g) of 4.65×10^{22} molecules of tetraphosphorus decaoxide?

(b) How many P atoms are present in this sample?

Mass Percent from the Chemical Formula

Each element in a compound constitutes its own particular portion of the compound's mass. For an individual molecule (or formula unit), we use the molecular (or formula) mass and chemical formula to find the mass percent of any element X in the compound:

$$\text{Mass \% of element X} = \frac{\text{atoms of X in formula} \times \text{atomic mass of X (amu)}}{\text{molecular (or formula) mass of compound (amu)}} \times 100$$

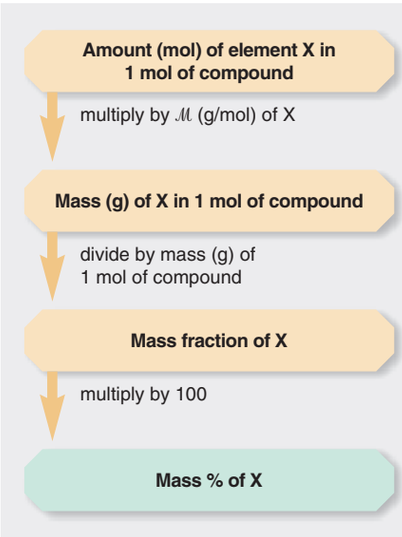
The formula also tells the number of *moles* of each element in the compound, so we can use the molar mass to find the mass percent of each element on a mole basis:

$$\text{Mass \% of element X} = \frac{\text{moles of X in formula} \times \text{molar mass of X (g/mol)}}{\text{mass (g) of 1 mol of compound}} \times 100 \quad (3.6)$$

As always, the individual mass percents of the elements in the compound must add up to 100% (within rounding). As Sample Problem 3.3 demonstrates, an important practical use of mass percent is to determine the amount of an element in any size sample of a compound.

SAMPLE PROBLEM 3.3 Calculating Mass Percents and Masses of Elements in a Sample of a Compound

Problem In mammals, lactose (milk sugar) is metabolized to glucose ($C_6H_{12}O_6$), the key nutrient for generating chemical potential energy.
(a) What is the mass percent of each element in glucose?
(b) How many grams of carbon are in 16.55 g of glucose?



(a) Determining the mass percent of each element

Plan We know the relative numbers of moles of the elements in glucose from the formula (6 C, 12 H, 6 O). We multiply the number of moles of each element by its molar mass to find grams. Dividing each element's mass by the mass of 1 mol of glucose gives the mass fraction of each element, and multiplying each fraction by 100 gives the mass percent. The calculation steps for any element X are shown in the roadmap.

Solution Calculating the mass of 1 mol of $C_6H_{12}O_6$:

$$\begin{aligned}\mathcal{M} &= (6 \times \mathcal{M} \text{ of C}) + (12 \times \mathcal{M} \text{ of H}) + (6 \times \mathcal{M} \text{ of O})\\ &= (6 \times 12.01 \text{ g/mol}) + (12 \times 1.008 \text{ g/mol}) + (6 \times 16.00 \text{ g/mol})\\ &= 180.16 \text{ g/mol}\end{aligned}$$

Converting moles of C to grams: There are 6 mol of C in 1 mol of glucose, so

$$\text{Mass (g) of C} = 6\ \cancel{\text{mol C}} \times \frac{12.01 \text{ g C}}{1\ \cancel{\text{mol C}}} = 72.06 \text{ g C}$$

Finding the mass fraction of C in glucose:

$$\text{Mass fraction of C} = \frac{\text{total mass C}}{\text{mass of 1 mol glucose}} = \frac{72.06 \text{ g}}{180.16 \text{ g}} = 0.4000$$

Finding the mass percent of C:

$$\text{Mass \% of C} = \text{mass fraction of C} \times 100 = 0.4000 \times 100 = \boxed{40.00 \text{ mass \% C}}$$

Combining the steps for each of the other two elements in glucose:

$$\text{Mass \% of H} = \frac{\text{mol H} \times \mathcal{M} \text{ of H}}{\text{mass of 1 mol glucose}} \times 100 = \frac{12\ \cancel{\text{mol H}} \times \dfrac{1.008\ \cancel{\text{g}}\ \text{H}}{1\ \cancel{\text{mol H}}}}{180.16\ \cancel{\text{g}}} \times 100 = \boxed{6.714 \text{ mass \% H}}$$

$$\text{Mass \% of O} = \frac{\text{mol O} \times \mathcal{M} \text{ of O}}{\text{mass of 1 mol glucose}} \times 100 = \frac{6\ \cancel{\text{mol O}} \times \dfrac{16.00\ \cancel{\text{g}}\ \text{O}}{1\ \cancel{\text{mol O}}}}{180.16\ \cancel{\text{g}}} \times 100 = \boxed{53.29 \text{ mass \% O}}$$

Check The answers make sense: even though there are equal numbers of moles of O and C in the compound, the mass % of O is greater than the mass % of C because the molar mass of O is greater than the molar mass of C. The mass % of H is small because the molar mass of H is small. The total of the mass percents is 100.00%.

(b) Determining the mass (g) of carbon

Plan To find the mass of C in the glucose sample, we multiply the mass of the sample by the mass fraction of C from part (a).

Solution Finding the mass of C in a given mass of glucose (with units for mass fraction):

$$\text{Mass (g) of C} = \text{mass of glucose} \times \text{mass fraction of C} = 16.55\ \cancel{\text{g glucose}} \times \frac{0.4000 \text{ g C}}{1\ \cancel{\text{g glucose}}} = \boxed{6.620 \text{ g C}}$$

Check Rounding shows that the answer is "in the right ballpark": 16 g times less than 0.5 parts by mass should be less than 8 g.

Comment 1. A *more direct approach* to finding the mass of element in any mass of compound is similar to the approach we used in Sample Problem 2.1 and eliminates the need

to calculate the mass fraction. Just multiply the given mass of compound by the ratio of the total mass of element to the mass of 1 mol of compound:

$$\text{Mass (g) of C} = 16.55\ \cancel{\text{g glucose}} \times \frac{72.06\ \text{g C}}{180.16\ \cancel{\text{g glucose}}} = 6.620\ \text{g C}$$

2. From here on, you should be able to determine the molar mass of a compound, so that calculation will no longer be shown.

FOLLOW-UP PROBLEM 3.3 Ammonium nitrate is a common fertilizer. Agronomists base the effectiveness of fertilizers on their nitrogen content.
(a) Calculate the mass percent of N in ammonium nitrate.
(b) How many grams of N are in 35.8 kg of ammonium nitrate?

SECTION SUMMARY

A mole of substance is the amount that contains Avogadro's number (6.022×10^{23}) of chemical entities (atoms, molecules, or formula units). The mass (in grams) of a mole has the same numerical value as the mass (in amu) of the entity. Thus, the mole allows us to count entities by weighing them. Using the molar mass ($\mathcal{M}$, g/mol) of an element (or compound) and Avogadro's number as conversion factors, we can convert among amount (mol), mass (g), and number of entities. The mass fraction of element X in a compound is used to find the mass of X in any amount of the compound.

3.2 DETERMINING THE FORMULA OF AN UNKNOWN COMPOUND

In Sample Problem 3.3, we knew the formula and used it to find the mass percent (or mass fraction) of an element in a compound *and* the mass of the element in a given mass of the compound. In this section, we do the reverse: use the masses of elements in a compound to find its formula. We'll present the mass data in several ways and then look briefly at molecular structures.

Empirical Formulas

An analytical chemist investigating a compound decomposes it into simpler substances, finds the mass of each component element, converts these masses to numbers of moles, and then arithmetically converts the moles to whole-number (integer) subscripts. This procedure yields the empirical formula, the *simplest whole-number ratio* of moles of each element in the compound (see Section 2.8). Let's see how to obtain the subscripts from the moles of each element.

Analysis of an unknown compound shows that the sample contains 0.21 mol of zinc, 0.14 mol of phosphorus, and 0.56 mol of oxygen. Because the subscripts in a formula represent individual atoms or moles of atoms, we write a preliminary formula that contains fractional subscripts: $Zn_{0.21}P_{0.14}O_{0.56}$. Next, we convert these fractional subscripts to whole numbers using one or two simple arithmetic steps (rounding when needed):

1. Divide each subscript by the smallest subscript:

$$Zn_{\frac{0.21}{0.14}}P_{\frac{0.14}{0.14}}O_{\frac{0.56}{0.14}} \longrightarrow Zn_{1.5}P_{1.0}O_{4.0}$$

   This step alone often gives integer subscripts.
2. If any of the subscripts is still not an integer, multiply through by the *smallest integer* that will turn all subscripts into integers. Here, we multiply by 2, the smallest integer that will make 1.5 (the subscript for Zn) into an integer:

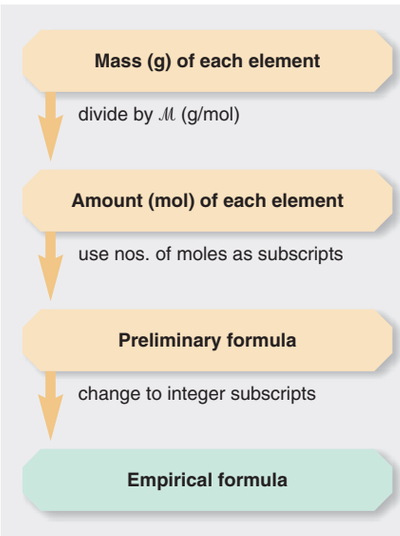
$$Zn_{(1.5\times2)}P_{(1.0\times2)}O_{(4.0\times2)} \longrightarrow Zn_{3.0}P_{2.0}O_{8.0},\ \text{or}\ Zn_3P_2O_8$$

   Notice that the *relative* number of moles has not changed because we multiplied *all* the subscripts by 2.

Always check that the subscripts are the smallest set of integers with the same ratio as the original numbers of moles; that is, 3:2:8 is *in the same ratio* as 0.21:0.14:0.56. A more conventional way to write this formula is $Zn_3(PO_4)_2$; the compound is zinc phosphate, a dental cement.

The following three sample problems (3.4, 3.5, and 3.6) demonstrate how other types of compositional data are used to determine chemical formulas. In the first problem, the empirical formula is found from data given as grams of each element rather than as moles.

SAMPLE PROBLEM 3.4 Determining an Empirical Formula from Masses of Elements



Problem Elemental analysis of a sample of an ionic compound showed 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What is the empirical formula and name of the compound?

Plan This problem is similar to the one we just discussed, except that we are given element *masses,* so we must convert the masses into integer subscripts. We first divide each mass by the element's molar mass to find *number of moles.* Then we construct a preliminary formula and convert the numbers of moles to integers.

Solution Finding moles of elements:

$$\text{Moles of Na} = 2.82\ \cancel{\text{g Na}} \times \frac{1\ \text{mol Na}}{22.99\ \cancel{\text{g Na}}} = 0.123\ \text{mol Na}$$

$$\text{Moles of Cl} = 4.35\ \cancel{\text{g Cl}} \times \frac{1\ \text{mol Cl}}{35.45\ \cancel{\text{g Cl}}} = 0.123\ \text{mol Cl}$$

$$\text{Moles of O} = 7.83\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.489\ \text{mol O}$$

Constructing a preliminary formula: $Na_{0.123}Cl_{0.123}O_{0.489}$

Converting to integer subscripts (dividing all by the smallest subscript):

$$Na_{\frac{0.123}{0.123}}Cl_{\frac{0.123}{0.123}}O_{\frac{0.489}{0.123}} \longrightarrow Na_{1.00}Cl_{1.00}O_{3.98} \approx Na_1Cl_1O_4, \text{ or } NaClO_4$$

We rounded the subscript of O from 3.98 to 4. The empirical formula is $NaClO_4$; the name is sodium perchlorate.

Check The moles seem correct because the masses of Na and Cl are slightly more than 0.1 of their molar masses. The mass of O is greatest and its molar mass is smallest, so it should have the greatest number of moles. The ratio of subscripts, 1:1:4, is the same as the ratio of moles, 0.123:0.123:0.489 (within rounding).

FOLLOW-UP PROBLEM 3.4 An unknown metal M reacts with sulfur to form a compound with the formula M_2S_3. If 3.12 g of M reacts with 2.88 g of S, what are the names of M and M_2S_3? (*Hint:* Determine the number of moles of S and use the formula to find the number of moles of M.)

Molecular Formulas

If we know the molar mass of a compound, we can use the empirical formula to obtain the molecular formula, the *actual* number of moles of each element in 1 mol of compound. In some cases, such as water (H_2O), ammonia (NH_3), and methane (CH_4), the empirical and molecular formulas are identical, but in many others the molecular formula is a *whole-number multiple* of the empirical formula. Hydrogen peroxide, for example, has the empirical formula HO and the molecular formula H_2O_2. Dividing the molar mass of H_2O_2 (34.02 g/mol) by the empirical formula mass (17.01 g/mol) gives the whole-number multiple:

$$\text{Whole-number multiple} = \frac{\text{molar mass (g/mol)}}{\text{empirical formula mass (g/mol)}} = \frac{34.02\ \cancel{\text{g/mol}}}{17.01\ \cancel{\text{g/mol}}} = 2.000 = 2$$

Instead of giving compositional data in terms of masses of each element, analytical laboratories provide it as mass percents. From this, we determine the

empirical formula by (1) assuming 100.0 g of compound, which allows us to express mass percent directly as mass, (2) converting the mass to number of moles, and (3) constructing the empirical formula. With the molar mass, we can also find the whole-number multiple and then the molecular formula.

SAMPLE PROBLEM 3.5 Determining a Molecular Formula from Elemental Analysis and Molar Mass

Problem During excessive physical activity, lactic acid ($\mathcal{M}$ = 90.08 g/mol) forms in muscle tissue and is responsible for muscle soreness. Elemental analysis shows that this compound contains 40.0 mass % C, 6.71 mass % H, and 53.3 mass % O.
(a) Determine the empirical formula of lactic acid.
(b) Determine the molecular formula.

(a) Determining the empirical formula
Plan We know the mass % of each element and must convert each to an integer subscript. Although the mass of lactic acid is not given, mass % is the same for any mass of compound, so we can assume 100.0 g of lactic acid and express each mass % directly as grams. Then, we convert grams to moles and construct the empirical formula as we did in Sample Problem 3.4.
Solution Expressing mass % as grams, assuming 100.0 g of lactic acid:

$$\text{Mass (g) of C} = \frac{\text{40.0 parts C by mass}}{\text{100 parts by mass}} \times 100.0\text{ g} = 40.0\text{ g C}$$

Similarly, we have 6.71 g of H and 53.3 g of O.
Converting from grams of each element to moles:

$$\text{Moles of C} = \text{mass of C} \times \frac{1}{\mathcal{M}\text{ of C}} = 40.0\ \cancel{\text{g C}} \times \frac{1\text{ mol C}}{12.01\ \cancel{\text{g C}}} = 3.33\text{ mol C}$$

Similarly, we have 6.66 mol of H and 3.33 mol of O.
Constructing the preliminary formula: $C_{3.33}H_{6.66}O_{3.33}$
Converting to integer subscripts:

$$C_{\frac{3.33}{3.33}}H_{\frac{6.66}{3.33}}O_{\frac{3.33}{3.33}} \longrightarrow C_{1.00}H_{2.00}O_{1.00} \approx C_1H_2O_1\text{; the empirical formula is } \boxed{CH_2O}$$

Check The numbers of moles seem correct: the masses of C and O are each slightly more than 3 times their molar masses (e.g., for C, 40 g/(12 g/mol) > 3 mol), and the mass of H is over 6 times its molar mass.

(b) Determining the molecular formula
Plan The molecular formula subscripts are whole-number multiples of the empirical formula subscripts. To find this whole number, we divide the given molar mass (90.08 g/mol) by the empirical formula mass, which we find from the sum of the elements' molar masses. Then we multiply the whole number by each subscript in the empirical formula.
Solution The empirical-formula molar mass is 30.03 g/mol. Finding the whole-number multiple:

$$\text{Whole-number multiple} = \frac{\mathcal{M}\text{ of lactic acid}}{\mathcal{M}\text{ of empirical formula}} = \frac{90.08\ \cancel{\text{g/mol}}}{30.03\ \cancel{\text{g/mol}}} = 3.000 = 3$$

Determining the molecular formula:

$$C_{(1\times3)}H_{(2\times3)}O_{(1\times3)} = \boxed{C_3H_6O_3}$$

Check The calculated molecular formula has the same ratio of moles of elements (3:6:3) as the empirical formula (1:2:1) and corresponds to the given molar mass:

$$\mathcal{M}\text{ of lactic acid} = (3 \times \mathcal{M}\text{ of C}) + (6 \times \mathcal{M}\text{ of H}) + (3 \times \mathcal{M}\text{ of O})$$
$$= (3 \times 12.01) + (6 \times 1.008) + (3 \times 16.00) = 90.08\text{ g/mol}$$

FOLLOW-UP PROBLEM 3.5 One of the most widespread environmental carcinogens (cancer-causing agents) is benzo[*a*]pyrene ($\mathcal{M}$ = 252.30 g/mol). It is found in coal dust, in cigarette smoke, and even in charcoal-grilled meat. Analysis of this hydrocarbon shows 95.21 mass % C and 4.79 mass % H. What is the molecular formula of benzo[*a*]pyrene?

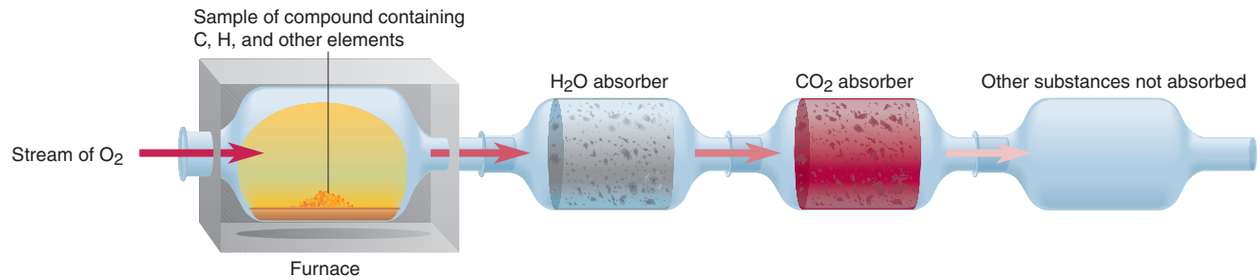


Figure 3.5 Combustion apparatus for determining formulas of organic compounds. A sample of compound that contains C and H (and perhaps other elements) is burned in a stream of O_2 gas. The CO_2 and H_2O formed are absorbed separately, while any other element oxides are carried through by the O_2 gas stream. H_2O is absorbed by $Mg(ClO_4)_2$; CO_2 is absorbed by NaOH. The increases in mass of the absorbers are used to calculate the amounts (mol) of C and H in the sample.

Combustion Analysis of Organic Compounds Still another type of compositional data is obtained through **combustion analysis,** a method used to measure the amounts of carbon and hydrogen in a combustible organic compound. The unknown compound is burned in pure O_2 in an apparatus that consists of a combustion chamber and chambers containing compounds that absorb either H_2O or CO_2 (Figure 3.5). All the H in the unknown is converted to H_2O, which is absorbed in the first chamber, and all the C is converted to CO_2, which is absorbed in the second. By weighing the contents of the chambers before and after combustion, we find the masses of CO_2 and H_2O and use them to calculate the masses of C and H in the compound, from which we find the empirical formula.

Many organic compounds also contain at least one other element, such as oxygen, nitrogen, or a halogen. As long as a third element doesn't interfere with the absorption of CO_2 and H_2O, we calculate its mass by subtracting the masses of C and H from the original mass of the compound.

SAMPLE PROBLEM 3.6 Determining a Molecular Formula from Combustion Analysis

Problem Vitamin C ($\mathcal{M}$ = 176.12 g/mol) is a compound of C, H, and O found in many natural sources, especially citrus fruits. When a 1.000-g sample of vitamin C is placed in a combustion chamber and burned, the following data are obtained:

$$\text{Mass of } CO_2 \text{ absorber after combustion} = 85.35 \text{ g}$$
$$\text{Mass of } CO_2 \text{ absorber before combustion} = 83.85 \text{ g}$$
$$\text{Mass of } H_2O \text{ absorber after combustion} = 37.96 \text{ g}$$
$$\text{Mass of } H_2O \text{ absorber before combustion} = 37.55 \text{ g}$$

What is the molecular formula of vitamin C?

Plan We find the masses of CO_2 and H_2O by subtracting the masses of the absorbers before the reaction from the masses after. From the mass of CO_2, we use the mass fraction of C in CO_2 to find the mass of C (see Comment in Sample Problem 3.3). Similarly, we find the mass of H from the mass of H_2O. The mass of vitamin C (1.000 g) minus the sum of the C and H masses gives the mass of O, the third element present. Then, we proceed as in Sample Problem 3.5: calculate numbers of moles using the elements' molar masses, construct the empirical formula, determine the whole-number multiple from the given molar mass, and construct the molecular formula.

Solution Finding the masses of combustion products:

$$\text{Mass (g) of } CO_2 = \text{mass of } CO_2 \text{ absorber after} - \text{mass before}$$
$$= 85.35 \text{ g} - 83.85 \text{ g} = 1.50 \text{ g } CO_2$$
$$\text{Mass (g) of } H_2O = \text{mass of } H_2O \text{ absorber after} - \text{mass before}$$
$$= 37.96 \text{ g} - 37.55 \text{ g} = 0.41 \text{ g } H_2O$$

Calculating masses of C and H using their mass fractions:

$$\text{Mass of element} = \text{mass of compound} \times \frac{\text{mass of element in compound}}{\text{mass of 1 mol of compound}}$$

$$\text{Mass (g) of C} = \text{mass of } CO_2 \times \frac{1 \text{ mol C} \times \mathcal{M} \text{ of C}}{\text{mass of 1 mol } CO_2} = 1.50 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2}$$

$$= 0.409 \text{ g C}$$

$$\text{Mass (g) of H} = \text{mass of } H_2O \times \frac{2 \text{ mol H} \times \mathcal{M} \text{ of H}}{\text{mass of 1 mol } H_2O} = 0.41 \text{ g } H_2O \times \frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O}$$

$$= 0.046 \text{ g H}$$

Calculating the mass of O:

$$\text{Mass (g) of O} = \text{mass of vitamin C sample} - (\text{mass of C} + \text{mass of H})$$
$$= 1.000 \text{ g} - (0.409 \text{ g} + 0.046 \text{ g}) = 0.545 \text{ g O}$$

Finding the amounts (mol) of elements: Dividing the mass in grams of each element by its molar mass gives 0.0341 mol of C, 0.046 mol of H, and 0.0341 mol of O.
Constructing the preliminary formula: $C_{0.0341}H_{0.046}O_{0.0341}$
Determining the empirical formula: Dividing through by the smallest subscript gives

$$C_{\frac{0.0341}{0.0341}}H_{\frac{0.046}{0.0341}}O_{\frac{0.0341}{0.0341}} = C_{1.00}H_{1.3}O_{1.00}$$

By trial and error, we find that 3 is the smallest integer that will make all subscripts approximately into integers:

$$C_{(1.00\times3)}H_{(1.3\times3)}O_{(1.00\times3)} = C_{3.00}H_{3.9}O_{3.00} \approx C_3H_4O_3$$

Determining the molecular formula:

$$\text{Whole-number multiple} = \frac{\mathcal{M} \text{ of vitamin C}}{\mathcal{M} \text{ of empirical formula}} = \frac{176.12 \text{ g/mol}}{88.06 \text{ g/mol}} = 2.000 = 2$$

$$C_{(3\times2)}H_{(4\times2)}O_{(3\times2)} = \boxed{C_6H_8O_6}$$

Check The element masses seem correct: carbon makes up slightly more than 0.25 of the mass of CO_2 (12 g/44 g > 0.25), as do the masses in the problem (0.409 g/1.50 g > 0.25). Hydrogen makes up slightly more than 0.10 of the mass of H_2O (2 g/18 g > 0.10), as do the masses in the problem (0.046 g/0.41 g > 0.10). The molecular formula has the same ratio of subscripts (6:8:6) as the empirical formula (3:4:3) and adds up to the given molar mass:

$$(6 \times \mathcal{M} \text{ of C}) + (8 \times \mathcal{M} \text{ of H}) + (6 \times \mathcal{M} \text{ of O}) = \mathcal{M} \text{ of vitamin C}$$
$$(6 \times 12.01) + (8 \times 1.008) + (6 \times 16.00) = 176.12 \text{ g/mol}$$

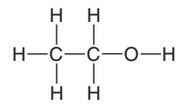
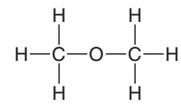
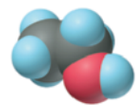
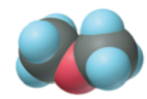
Comment In determining the subscript for H, if we string the calculation steps together, we obtain the subscript 4.0, rather than 3.9, and don't need to round:

$$\text{Subscript of H} = 0.41 \text{ g } H_2O \times \frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} \times \frac{1}{0.0341 \text{ mol}} \times 3 = 4.0$$

FOLLOW-UP PROBLEM 3.6 A dry-cleaning solvent ($\mathcal{M}$ = 146.99 g/mol) that contains C, H, and Cl is suspected to be a cancer-causing agent. When a 0.250-g sample was studied by combustion analysis, 0.451 g of CO_2 and 0.0617 g of H_2O formed. Find the molecular formula.

Isomers A molecular formula tells the *actual* number of each type of atom, providing as much information as possible from mass analysis. Yet *different compounds can have the* ***same*** *molecular formula* because the atoms can bond to each other in different arrangements to give more than one *structural formula*. **Isomers** are compounds with the same molecular formula but different properties. The simplest type of isomerism, called *constitutional,* or *structural, isomerism,* occurs when the atoms link together in different arrangements. The pair of constitutional isomers shown in Table 3.2 (on the next page) share the molecular formula C_2H_6O

Table 3.2 Constitutional Isomers of C_2H_6O

Property	Ethanol	Dimethyl Ether
$\mathcal{M}$ (g/mol)	46.07	46.07
Boiling point	78.5°C	−25°C
Density (at 20°C)	0.789 g/mL (liquid)	0.00195 g/mL (gas)
Structural formula		
Space-filling model		

but have very different properties because they are different compounds. In this case, they are even different *types* of compounds—one is an alcohol, and the other an ether.

As the number and kinds of atoms increase, the number of constitutional isomers—that is, the number of structural formulas that can be written for a given molecular formula—also increases: C_2H_6O has the two that you've seen, C_3H_8O has three, and $C_4H_{10}O$, seven. (We'll discuss this and other types of isomerism fully later in the text.)

SECTION SUMMARY

From the masses of elements in an unknown compound, the relative amounts (in moles) can be found and the empirical formula determined. If the molar mass is known, the molecular formula can also be determined. Methods such as combustion analysis provide data on the masses of elements in a compound, which can be used to obtain the formula. Because atoms can bond in different arrangements, more than one compound may have the same molecular formula (constitutional isomers).

3.3 WRITING AND BALANCING CHEMICAL EQUATIONS

Perhaps the most important reason for thinking in terms of moles is because it greatly clarifies the amounts of substances taking part in a reaction. Comparing masses doesn't tell the ratio of substances reacting but comparing numbers of moles does. It allows us to view substances as large populations of interacting particles rather than as grams of material. To clarify this idea, consider the formation of hydrogen fluoride gas from H_2 and F_2, a reaction that occurs explosively at room temperature. If we weigh the gases, we find that

2.016 g of H_2 and 38.00 g of F_2 react to form 40.02 g of HF

This information tells us little except that mass is conserved. However, if we convert these masses (in grams) to amounts (in moles), we find that

1 mol of H_2 and 1 mol of F_2 react to form 2 mol of HF

This information reveals that equal-size populations of H_2 and F_2 molecules combine to form twice as large a population of HF molecules. Dividing through by Avogadro's number shows us the chemical event that occurs between individual molecules:

1 H_2 molecule and 1 F_2 molecule react to form 2 HF molecules

Figure 3.6 shows that when we express the reaction in terms of moles, *the macroscopic (molar) change corresponds to the submicroscopic (molecular) change.* As you'll see, a balanced chemical equation shows both changes.

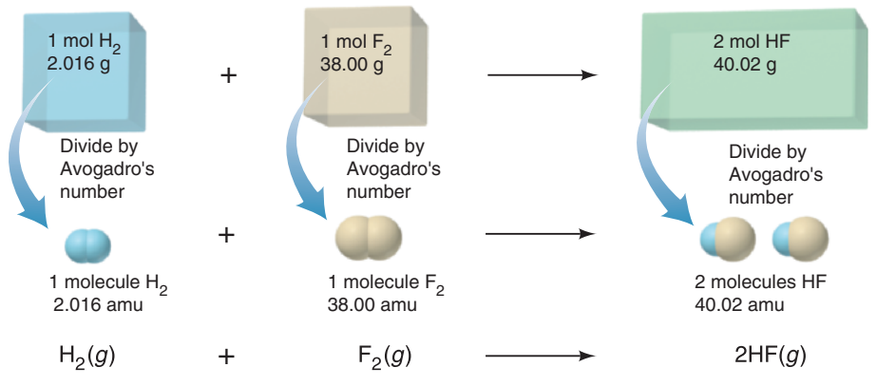


Figure 3.6 The formation of HF gas on the macroscopic and molecular levels. When 1 mol of H_2 (2.016 g) and 1 mol of F_2 (38.00 g) react, 2 mol of HF (40.02 g) forms. Dividing by Avogadro's number shows the change at the molecular level.

A **chemical equation** is a statement in formulas that expresses the identities and quantities of the substances involved in a chemical or physical change. Equations are the "sentences" of chemistry, just as chemical formulas are the "words" and atomic symbols the "letters." The left side of an equation shows the amount of each substance present before the change, and the right side shows the amounts present afterward. *For an equation to depict these amounts accurately, it must be balanced; that is, the same number of each type of atom must appear on both sides of the equation.* This requirement follows directly from the mass laws and the atomic theory:

- In a chemical process, atoms cannot be created, destroyed, or changed, only rearranged into different combinations.
- A formula represents a fixed ratio of the elements in a compound, so a different ratio represents a different compound.

Consider the chemical change that occurs in an old-fashioned photographic flashbulb: magnesium wire and oxygen gas yield powdery magnesium oxide. (Light and heat are produced as well, but here we're concerned only with the substances involved.) Let's convert this chemical statement into a balanced equation through the following steps:

1. *Translating the statement.* We first translate the chemical statement into a "skeleton" equation: chemical formulas arranged in an equation format. All the substances that react during the change, called **reactants,** are placed to the left of a "yield" arrow, which points to all the substances produced, called **products:**

$$\overbrace{\_\_Mg \quad + \quad \_\_O_2}^{\text{reactants}} \xrightarrow{\text{yield}} \underset{\text{product}}{\_\_MgO}$$

magnesium and oxygen yield magnesium oxide

At the beginning of the balancing process, we put a blank in front of each substance to remind us that we have to account for its atoms.

2. *Balancing the atoms.* The next step involves shifting our attention back and forth from right to left in order to *match the number of each type of atom on each side.* At the end of this step, each blank will contain a **balancing (stoichiometric) coefficient,** a numerical multiplier of *all the atoms* in the formula that follows it. In general, balancing is easiest when we

- Start with the most complex substance, the one with the largest number of atoms or different types of atoms.
- End with the least complex substance, such as an element by itself.

In this case, MgO is the most complex, so we place a coefficient 1 *in front of* the compound:

$$\_\_Mg + \_\_O_2 \longrightarrow \_1\_MgO$$

To balance the Mg in MgO on the right, we place a 1 in front of Mg on the left:

$$\underline{1}\,Mg + \underline{\ }\,O_2 \longrightarrow \underline{1}\,MgO$$

The O atom on the right must be balanced by one O atom on the left. One-half an O_2 molecule provides one O atom:

$$\underline{1}\,Mg + \underline{\tfrac{1}{2}}\,O_2 \longrightarrow \underline{1}\,MgO$$

In terms of number and type of atom, the equation is balanced.

3. *Adjusting the coefficients.* There are several conventions about the final form of the coefficients:

- In most cases, *the smallest whole-number coefficients are preferred.* Whole numbers allow entities such as O_2 molecules to be treated as intact particles. One-half of an O_2 molecule cannot exist, so we multiply the equation by 2:

$$2Mg + 1O_2 \longrightarrow 2MgO$$

- We used the coefficient 1 to remind us to balance each substance. In the final form, a coefficient of 1 is implied just by the presence of the formula of the substance, so we don't need to write it:

$$2Mg + O_2 \longrightarrow 2MgO$$

(This convention is similar to not writing a subscript 1 in a formula.)

4. *Checking.* After balancing and adjusting the coefficients, always check that the equation is balanced:

$$\text{Reactants } (2\ Mg,\ 2\ O) \longrightarrow \text{products } (2\ Mg,\ 2\ O)$$

5. *Specifying the states of matter.* The final equation also indicates the physical state of each substance or whether it is dissolved in water. The abbreviations that are used for these states are solid (*s*), liquid (*l*), gas (*g*), and aqueous solution (*aq*). From the original statement, we know that the Mg "wire" is solid, the O_2 is a gas, and the "powdery" MgO is also solid. The balanced equation, therefore, is

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Of course, the key point to realize is, as was pointed out in Figure 3.6, *the balancing coefficients refer to both individual chemical entities and moles of chemical entities*. Thus, 2 mol of Mg and 1 mol of O_2 yield 2 mol of MgO. Figure 3.7 shows this reaction from three points of view—as you see it on the macroscopic level, as chemists (and you!) can imagine it on the atomic level (darker colored atoms represent the stoichiometry), and on the symbolic level of the chemical equation.

Keep in mind these other key points about the balancing process:

- A coefficient operates on *all the atoms in the formula* that follows it: 2MgO means 2 × (MgO), or 2 Mg atoms and 2 O atoms; $2Ca(NO_3)_2$ means 2 × [$Ca(NO_3)_2$], or 2 Ca atoms, 4 N atoms, and 12 O atoms.
- In balancing an equation, *chemical formulas cannot be altered.* In step 2 of the example, we *cannot* balance the O atoms by changing MgO to MgO_2 because MgO_2 has a different elemental composition and thus is a different compound.
- We *cannot add other reactants or products* to balance the equation because this would represent a different reaction. For example, we *cannot* balance the O atoms by changing O_2 to O or by adding one O atom to the products, because the chemical statement does not say that the reaction involves O atoms.

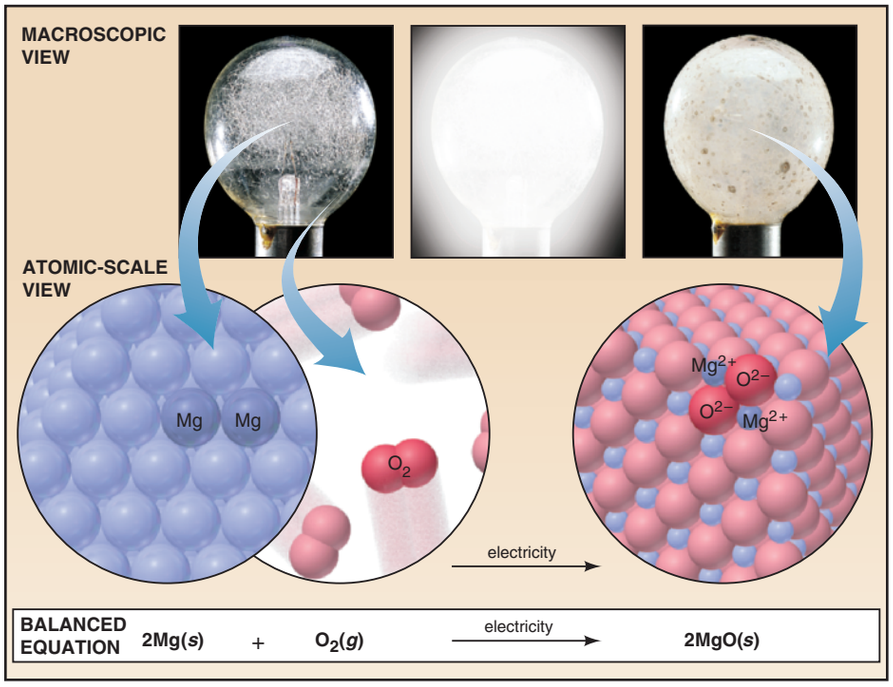


Figure 3.7 A three-level view of the chemical reaction in a flashbulb. The photos present the macroscopic view that you see. Before the reaction occurs, a fine magnesium filament is surrounded by oxygen (*left*). After the reaction, white, powdery magnesium oxide coats the bulb's inner surface (*right*). The blow-up arrows lead to an atomic-scale view, a representation of the chemist's mental picture of the reaction. The darker colored spheres show the stoichiometry. By knowing the substances before and after a reaction, we can write a balanced equation (*bottom*), the chemist's symbolic shorthand for the change.

- A balanced equation remains balanced even if you multiply all the coefficients by the same number. For example,

$$4Mg(s) + 2O_2(g) \longrightarrow 4MgO(s)$$

is also balanced: it is just the same balanced equation multiplied by 2. However, we balance an equation with the *smallest* whole-number coefficients.

SAMPLE PROBLEM 3.7 Balancing Chemical Equations

Problem Within the cylinders of a car's engine, the hydrocarbon octane (C_8H_{18}), one of many components of gasoline, mixes with oxygen from the air and burns to form carbon dioxide and water vapor. Write a balanced equation for this reaction.

Solution

1. *Translate* the statement into a skeleton equation (with coefficient blanks). Octane and oxygen are reactants; "oxygen from the air" implies molecular oxygen, O_2. Carbon dioxide and water vapor are products:

$$\_\_C_8H_{18} + \_\_O_2 \longrightarrow \_\_CO_2 + \_\_H_2O$$

2. *Balance the atoms.* We start with the most complex substance, C_8H_{18}, and balance O_2 last:

$$\underline{1}\,C_8H_{18} + \_\_O_2 \longrightarrow \_\_CO_2 + \_\_H_2O$$

The C atoms in C_8H_{18} end up in CO_2. Each CO_2 contains one C atom, so 8 molecules of CO_2 are needed to balance the 8 C atoms in each C_8H_{18}:

$$\underline{1}\,C_8H_{18} + \_\_O_2 \longrightarrow \underline{8}\,CO_2 + \_\_H_2O$$

The H atoms in C_8H_{18} end up in H_2O. The 18 H atoms in C_8H_{18} require a coefficient 9 in front of H_2O:

$$\underline{1}\,C_8H_{18} + \_\_O_2 \longrightarrow \underline{8}\,CO_2 + \underline{9}\,H_2O$$

There are 25 atoms of O on the right (16 in $8CO_2$ plus 9 in $9H_2O$), so we place the coefficient $\frac{25}{2}$ in front of O_2:

$$\underline{1}\,C_8H_{18} + \underline{\tfrac{25}{2}}\,O_2 \longrightarrow \underline{8}\,CO_2 + \underline{9}\,H_2O$$

3. *Adjust the coefficients.* Multiply through by 2 to obtain whole numbers:

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

4. *Check* that the equation is balanced:

Reactants (16 C, 36 H, 50 O) $\longrightarrow$ products (16 C, 36 H, 50 O)

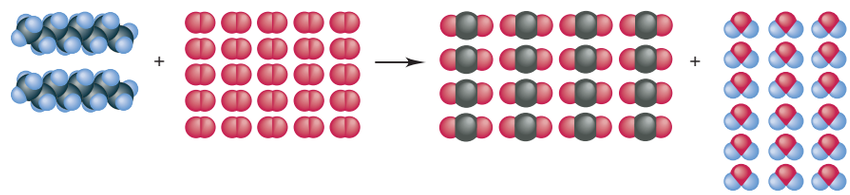
5. *Specify* states of matter. C_8H_{18} is liquid; O_2, CO_2, and H_2O vapor are gases:

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

Comment This is an example of a combustion reaction. *Any* compound containing C and H that burns in an excess of air produces CO_2 and H_2O.

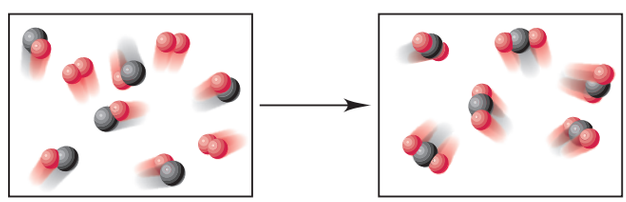
FOLLOW-UP PROBLEM 3.7 Write a balanced equation for each chemical statement:
(a) A characteristic reaction of Group 1A(1) elements: chunks of sodium react violently with water to form hydrogen gas and sodium hydroxide solution.
(b) The destruction of marble statuary by acid rain: aqueous nitric acid reacts with calcium carbonate to form carbon dioxide, water, and aqueous calcium nitrate.
(c) Halogen compounds exchanging bonding partners: phosphorus trifluoride is prepared by the reaction of phosphorus trichloride and hydrogen fluoride; hydrogen chloride is the other product. The reaction involves gases only.
(d) Explosive decomposition of dynamite: liquid nitroglycerine ($C_3H_5N_3O_9$) explodes to produce a mixture of gases—carbon dioxide, water vapor, nitrogen, and oxygen.

Viewing an equation in a schematic molecular scene is a great way to focus on the essence of the change—the rearrangement of the atoms from reactants to products. Here's a simple schematic for the combustion of octane:



$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

We can also derive a balanced equation from a molecular scene. Let's do this for the following change (black = carbon, red = oxygen):



We can see from the colors that the reactant box includes carbon monoxide (CO) and oxygen (O_2) molecules, and the product box contains carbon dioxide (CO_2). Once we identify the substances with formulas, we count the number of each kind of molecule and put the numbers and formulas in equation format. There are six CO, three O_2, and six CO_2, that is, twice as many CO as O_2 and the same number as CO_2. So, given that they are all gases, we adjust the coefficients and have

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$

SECTION SUMMARY

To conserve mass and maintain the fixed composition of compounds, a chemical equation must be balanced in terms of number and type of each atom. A balanced equation has reactant formulas on the left of a yield arrow and product formulas on the right. Balancing coefficients are integer multipliers for *all* the atoms in a formula and apply to the individual entity or to moles of entities.

3.4 CALCULATING AMOUNTS OF REACTANT AND PRODUCT

A balanced equation contains a wealth of quantitative information relating individual chemical entities, amounts of chemical entities, and masses of substances. It is essential for all calculations involving amounts of reactants and products: *if you know the number of moles of one substance, the balanced equation tells you the number of moles of all the others in the reaction.*

Stoichiometrically Equivalent Molar Ratios from the Balanced Equation

In a balanced equation, *the number of moles of one substance is stoichiometrically equivalent to the number of moles of any other substance.* The term *stoichiometrically equivalent* means that a definite amount of one substance is formed from, produces, or reacts with a definite amount of the other. These quantitative relationships are expressed as *stoichiometrically equivalent molar ratios* that we use as conversion factors to calculate these amounts. Table 3.3 presents the quantitative information contained in the equation for the combustion of propane, a hydrocarbon fuel used in cooking and water heating:

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

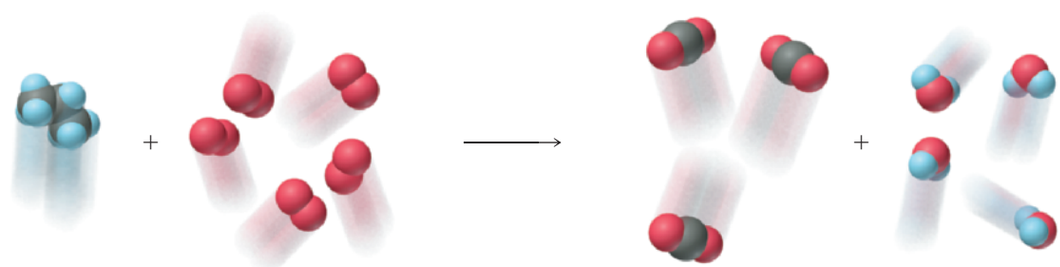
If we view the reaction quantitatively in terms of C_3H_8, we see that

1 mol of C_3H_8 reacts with 5 mol of O_2
1 mol of C_3H_8 produces 3 mol of CO_2
1 mol of C_3H_8 produces 4 mol of H_2O

Therefore, in this reaction,

1 mol of C_3H_8 is stoichiometrically equivalent to 5 mol of O_2
1 mol of C_3H_8 is stoichiometrically equivalent to 3 mol of CO_2
1 mol of C_3H_8 is stoichiometrically equivalent to 4 mol of H_2O

Table 3.3 Information Contained in a Balanced Equation

Viewed in Terms of	Reactants $C_3H_8(g) + 5O_2(g)$	$\longrightarrow$	Products $3CO_2(g) + 4H_2O(g)$
Molecules	1 molecule C_3H_8 + 5 molecules O_2	$\longrightarrow$	3 molecules CO_2 + 4 molecules H_2O
			
Amount (mol)	1 mol C_3H_8 + 5 mol O_2	$\longrightarrow$	3 mol CO_2 + 4 mol H_2O
Mass (amu)	44.09 amu C_3H_8 + 160.00 amu O_2	$\longrightarrow$	132.03 amu CO_2 + 72.06 amu H_2O
Mass (g)	44.09 g C_3H_8 + 160.00 g O_2	$\longrightarrow$	132.03 g CO_2 + 72.06 g H_2O
Total mass (g)	204.09 g	$\longrightarrow$	204.09 g

We chose to look at C_3H_8, but any two of the substances are stoichiometrically equivalent to each other. Thus,

3 mol of CO_2 is stoichiometrically equivalent to 4 mol of H_2O
5 mol of O_2 is stoichiometrically equivalent to 3 mol of CO_2

and so on.

Here's a typical problem that shows how stoichiometric equivalence is used to create conversion factors: in the combustion of propane, how many moles of O_2 are consumed when 10.0 mol of H_2O are produced? To solve this problem, we have to find the molar ratio between O_2 and H_2O. From the balanced equation, we see that for every 5 mol of O_2 consumed, 4 mol of H_2O is formed:

5 mol of O_2 is stoichiometrically equivalent to 4 mol of H_2O

We can construct two conversion factors from this equivalence, depending on the quantity we want to find:

$$\frac{5 \text{ mol } O_2}{4 \text{ mol } H_2O} \quad \text{or} \quad \frac{4 \text{ mol } H_2O}{5 \text{ mol } O_2}$$

Because we want to find moles of O_2 and we know moles of H_2O, we choose "5 mol O_2/4 mol H_2O" to cancel "mol H_2O":

$$\text{Moles of } O_2 \text{ consumed} = 10.0 \cancel{\text{ mol } H_2O} \times \frac{5 \text{ mol } O_2}{4 \cancel{\text{ mol } H_2O}} = 12.5 \text{ mol } O_2$$

$$\text{mol } H_2O \xrightarrow[\text{conversion factor}]{\text{molar ratio as}} \text{mol } O_2$$

Obviously, we could not have solved this problem without the balanced equation. Here is a general approach for solving *any* stoichiometry problem that involves a reaction:

1. Write a balanced equation for the reaction.
2. Convert the given mass (or number of entities) of the first substance to amount (mol).
3. Use the appropriate molar ratio from the balanced equation to calculate the amount (mol) of the second substance.
4. Convert the amount of the second substance to the desired mass (or number of entities).

This approach is shown in Figure 3.8 and demonstrated in the following sample problem.

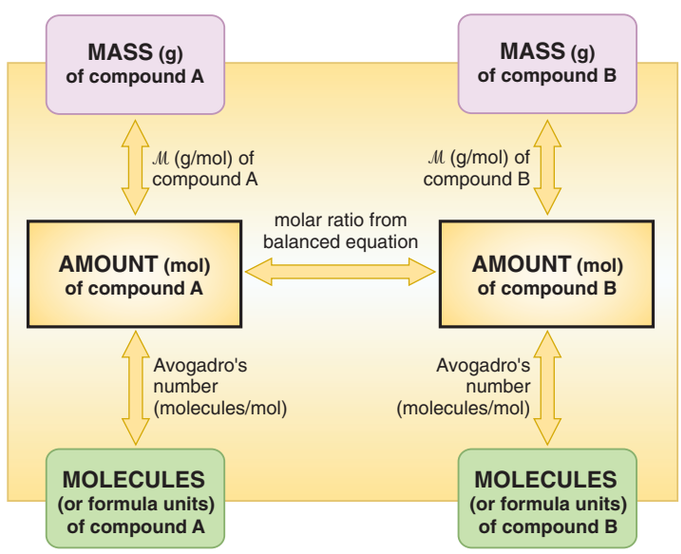


Figure 3.8 Summary of the mass-mole-number relationships in a chemical reaction. The amount of one substance in a reaction is related to that of any other. Quantities are expressed in terms of grams, moles, or number of entities (atoms, molecules, or formula units). Start at any box in the diagram (known) and move to any other box (unknown) by using the information on the arrows as conversion factors. As an example, if you know the mass (in g) of A and want to know the number of molecules of B, the path involves three calculation steps:
1. Grams of A to moles of A, using the molar mass ($\mathcal{M}$) of A
2. Moles of A to moles of B, using the molar ratio from the balanced equation
3. Moles of B to molecules of B, using Avogadro's number

Steps 1 and 3 refer to calculations discussed in Section 3.1 (see Figure 3.4).

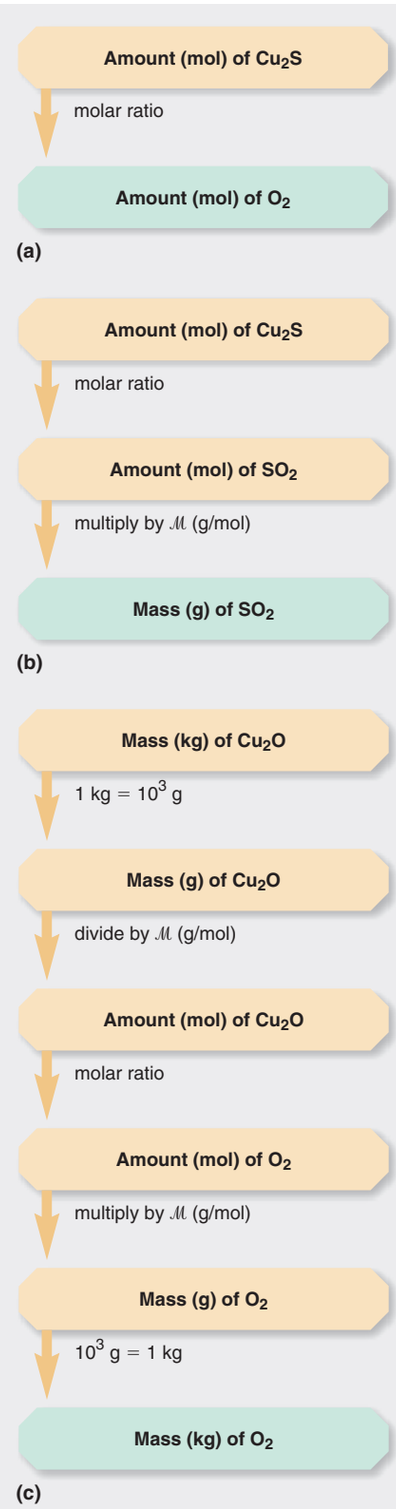
SAMPLE PROBLEM 3.8 Calculating Amounts of Reactants and Products

Problem In a lifetime, the average American uses 1750 lb (794 kg) of copper in coins, plumbing, and wiring. Copper is obtained from sulfide ores, such as chalcocite, or copper(I) sulfide, by a multistep process. After an initial grinding, the first step is to "roast" the ore (heat it strongly with oxygen gas) to form powdered copper(I) oxide and gaseous sulfur dioxide.

(a) How many moles of oxygen are required to roast 10.0 mol of copper(I) sulfide?

(b) How many grams of sulfur dioxide are formed when 10.0 mol of copper(I) sulfide is roasted?

(c) How many kilograms of oxygen are required to form 2.86 kg of copper(I) oxide?



(a) Determining the moles of O_2 needed to roast 10.0 mol of Cu_2S

Plan We *always* write the balanced equation first. The formulas of the reactants are Cu_2S and O_2, and the formulas of the products are Cu_2O and SO_2, so we have

$$2Cu_2S(s) + 3O_2(g) \longrightarrow 2Cu_2O(s) + 2SO_2(g)$$

We are given the *moles* of Cu_2S and need to find the *moles* of O_2. The balanced equation shows that 3 mol of O_2 is needed for every 2 mol of Cu_2S consumed, so the conversion factor is "3 mol O_2/2 mol Cu_2S" (see roadmap a).

Solution Calculating number of moles of O_2:

$$\text{Moles of } O_2 = 10.0\ \cancel{\text{mol } Cu_2S} \times \frac{3 \text{ mol } O_2}{2\ \cancel{\text{mol } Cu_2S}} = \boxed{15.0 \text{ mol } O_2}$$

Check The units are correct, and the answer is reasonable because this O_2/Cu_2S molar ratio (15:10) is equivalent to the ratio in the balanced equation (3:2).

Comment A *common mistake* is to use the incorrect conversion factor; the calculation would then be

$$\text{Moles of } O_2 = 10.0 \text{ mol } Cu_2S \times \frac{2 \text{ mol } Cu_2S}{3 \text{ mol } O_2} = \frac{6.67 \text{ mol}^2\ Cu_2S}{1 \text{ mol } O_2}$$

Such strange units should signal that you made an error in setting up the conversion factor. In addition, the answer, 6.67, is *less* than 10.0, whereas the balanced equation shows that *more* moles of O_2 than of Cu_2S are needed. Be sure to think through the calculation when setting up the conversion factor and canceling units.

(b) Determining the mass (g) of SO_2 formed from 10.0 mol of Cu_2S

Plan Here we need the *grams* of product (SO_2) that form from the given *moles* of reactant (Cu_2S). We first find the moles of SO_2 using the molar ratio from the balanced equation (2 mol SO_2/2 mol Cu_2S) and then multiply by its molar mass (64.07 g/mol) to find grams of SO_2. The steps appear in roadmap b.

Solution Combining the two conversion steps into one calculation, we have

$$\text{Mass (g) of } SO_2 = 10.0\ \cancel{\text{mol } Cu_2S} \times \frac{2\ \cancel{\text{mol } SO_2}}{2\ \cancel{\text{mol } Cu_2S}} \times \frac{64.07 \text{ g } SO_2}{1\ \cancel{\text{mol } SO_2}} = \boxed{641 \text{ g } SO_2}$$

Check The answer makes sense, since the molar ratio shows that 10.0 mol of SO_2 are formed and each mole weighs about 64 g. We rounded to three significant figures.

(c) Determining the mass (kg) of O_2 that yields 2.86 kg of Cu_2O

Plan Here the mass of product (Cu_2O) is known, and we need the mass of reactant (O_2) that reacts to form it. We first convert the quantity of Cu_2O from *kilograms* to *moles* (in two steps, as shown in roadmap c). Then, we use the molar ratio (3 mol O_2/2 mol Cu_2O) to find the *moles* of O_2 required. Finally, we convert *moles* of O_2 to *kilograms* (in two steps).

Solution Converting from kilograms of Cu_2O to moles of Cu_2O: Combining the mass unit conversion with the mass-to-mole conversion gives

$$\text{Moles of } Cu_2O = 2.86\ \cancel{\text{kg } Cu_2O} \times \frac{10^3\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1 \text{ mol } Cu_2O}{143.10\ \cancel{\text{g } Cu_2O}} = 20.0 \text{ mol } Cu_2O$$

Converting from moles of Cu_2O to moles of O_2:

$$\text{Moles of } O_2 = 20.0\ \cancel{\text{mol } Cu_2O} \times \frac{3 \text{ mol } O_2}{2\ \cancel{\text{mol } Cu_2O}} = 30.0 \text{ mol } O_2$$

Converting from moles of O_2 to kilograms of O_2: Combining the mole-to-mass conversion with the mass unit conversion gives

$$\text{Mass (kg) of } O_2 = 30.0\ \cancel{\text{mol } O_2} \times \frac{32.00\ \text{g } O_2}{1\ \cancel{\text{mol } O_2}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}} = 0.960\ \text{kg } O_2$$

Check The units are correct. Round off to check the math: for example, in the final step, ~30 mol × 30 g/mol × 1 kg/10^3 g = 0.90 kg. The answer seems reasonable: even though the amount (mol) of O_2 is greater than the amount (mol) of Cu_2O, the mass of O_2 is less than the mass of Cu_2O because $\mathcal{M}$ of O_2 is less than $\mathcal{M}$ of Cu_2O.

Comment This problem highlights a key point for solving stoichiometry problems: *convert the information given into moles.* Then, use the appropriate molar ratio and any other conversion factors to complete the solution.

FOLLOW-UP PROBLEM 3.8 Thermite is a mixture of iron(III) oxide and aluminum powders that was once used to weld railroad tracks. It undergoes a spectacular reaction to yield solid aluminum oxide and molten iron.

(a) How many grams of iron form when 135 g of aluminum reacts?

(b) How many atoms of aluminum react for every 1.00 g of aluminum oxide formed?

Chemical Reactions That Involve a Limiting Reactant

In the problems we've considered up to now, the amount of *one* reactant was given, and we assumed there was enough of any other reactant for the first reactant to be completely used up. For example, to find the amount of SO_2 that forms when 100 g of Cu_2S reacts, we convert the grams of Cu_2S to moles and assume that the Cu_2S reacts with as much O_2 as needed. Because all the Cu_2S is used up, its initial amount determines, or limits, how much SO_2 can form. We call Cu_2S the **limiting reactant** (or *limiting reagent*) because the product stops forming once the Cu_2S is gone, no matter how much O_2 is present.

Animation: Limiting Reagent
Online Learning Center

Suppose, however, that the amounts of both Cu_2S *and* O_2 are given in the problem, and we need to find out how much SO_2 forms. We first have to determine whether Cu_2S *or* O_2 is the limiting reactant (that is, which one is completely used up) because the amount of that reactant limits how much SO_2 can form. The other reactant is *in excess,* and whatever amount of it is not used is left over.

To clarify the idea of a limiting reactant, let's consider a situation from real life. A car assembly plant has 1500 car bodies and 4000 tires. How many cars can be made with the supplies on hand? Does the plant manager need to order more car bodies or more tires? Obviously, 4 tires are required for each car body, so the "balanced equation" is

$$1\ \text{car body} + 4\ \text{tires} \longrightarrow 1\ \text{car}$$

How much "product" (cars) can we make from the amount of each "reactant"?

$$1500\ \cancel{\text{car bodies}} \times \frac{1\ \text{car}}{1\ \cancel{\text{car body}}} = 1500\ \text{cars}$$

$$4000\ \cancel{\text{tires}} \times \frac{1\ \text{car}}{4\ \cancel{\text{tires}}} = 1000\ \text{cars}$$

The number of tires limits the number of cars because less "product" (fewer cars) can be produced from the available tires. There will be 1500 − 1000 = 500 car bodies in excess, and they cannot be turned into cars until more tires are delivered.

Now let's apply these ideas to solving chemical problems. In limiting-reactant problems, the amounts of two (or more) reactants are given, and we must first determine which is limiting. To do this, just as we did with the cars, we first note how much of each reactant *should* be present to completely use up the other, and then we compare it with the amount that is *actually* present. Simply put, the limiting reactant is the one there is not enough of; that is, *it is the reactant that limits the amount of the other reactant that can react, and thus the amount of*

product that can form. In mathematical terms, *the limiting reactant is the one that yields the* ***lower*** *amount of product*.

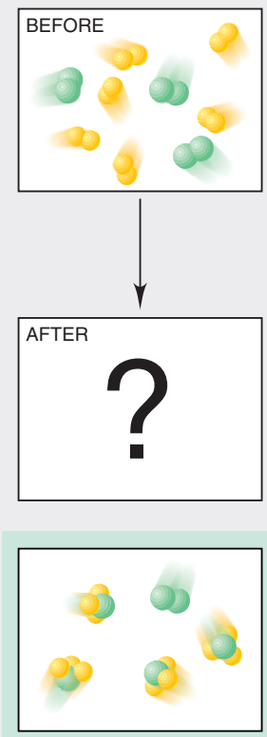
We'll examine limiting reactants in the following two sample problems. Sample Problem 3.9 has two parts, and in both we have to identify the limiting reactant. In the first part, we look at a simple molecular view of a reaction and compare the number of molecules to find the limiting reactant; in the second part, we start with the amounts (mol) of two reactants and perform two calculations, each of which assumes an excess of one of the reactants, to see which reactant forms less product. Then, in Sample Problem 3.10, we go through a similar process but start with the masses of the two reactants.

SAMPLE PROBLEM 3.9 Using Molecular Depictions to Solve a Limiting-Reactant Problem

Problem Nuclear engineers use chlorine trifluoride in the processing of uranium fuel for power plants. This extremely reactive substance is formed as a gas in special metal containers by the reaction of elemental chlorine and fluorine.

(a) Suppose the box shown at right represents a container of the reactant mixture before the reaction occurs (with chlorine colored green). Name the limiting reactant and draw the container contents after the reaction is complete.

(b) When the reaction is run again with 0.750 mol of Cl_2 and 3.00 mol of F_2, what mass of chlorine trifluoride will be prepared?



(a) Determining the limiting reactant and drawing the container contents

Plan We first write the balanced equation. From its name, we know that chlorine trifluoride consists of one Cl atom bonded to three F atoms, ClF_3. Elemental chlorine and fluorine refer to the diatomic molecules Cl_2 and F_2. All the substances are gases. To find the limiting reactant, we compare the number of molecules we have of each reactant, with the number we need for the other to react completely. The limiting reactant limits the amount of the other reactant that can react and the amount of product that will form.

Solution The balanced equation is

$$Cl_2(g) + 3F_2(g) \longrightarrow 2ClF_3(g)$$

The equation shows that two ClF_3 molecules are formed for every one Cl_2 molecule and three F_2 molecules that react. Before the reaction, there are three Cl_2 molecules (six Cl atoms). For all the Cl_2 to react, we need three times three, or nine, F_2 molecules (18 F atoms). But there are only six F_2 molecules (12 F atoms). Therefore, F_2 is the limiting reactant because it limits the amount of Cl_2 that can react, and thus the amount of ClF_3 that can form. After the reaction, as the box at right depicts, all 12 F atoms and four of the six Cl atoms make four ClF_3 molecules, and one Cl_2 molecule remains in excess.

Check The equation is balanced: reactants (2 Cl, 6 F) $\longrightarrow$ products (2 Cl, 6 F), and, in the boxes, the number of each type of atom before the reaction equals the number after the reaction. You can check the choice of limiting reactant by examining the reaction from the perspective of Cl_2: Two Cl_2 molecules are enough to react with the six F_2 molecules in the container. But there are three Cl_2 molecules, so there is not enough F_2.

(b) Calculating the mass of ClF_3 formed

Plan We first determine the limiting reactant by using the molar ratios from the balanced equation to convert the moles of each reactant to moles of ClF_3 formed, assuming an excess of the other reactant. Whichever reactant forms fewer moles of ClF_3 is limiting. Then we use the molar mass of ClF_3 to convert this lower number of moles to grams.

Solution Determining the limiting reactant:

Finding moles of ClF_3 from moles of Cl_2 (assuming F_2 is in excess):

$$\text{Moles of } ClF_3 = 0.750\ \cancel{\text{mol } Cl_2} \times \frac{2 \text{ mol } ClF_3}{1\ \cancel{\text{mol } Cl_2}} = 1.50 \text{ mol } ClF_3$$

Finding moles of ClF_3 from moles of F_2 (assuming Cl_2 is in excess):

$$\text{Moles of } ClF_3 = 3.00\ \cancel{\text{mol } F_2} \times \frac{2 \text{ mol } ClF_3}{3\ \cancel{\text{mol } F_2}} = 2.00 \text{ mol } ClF_3$$

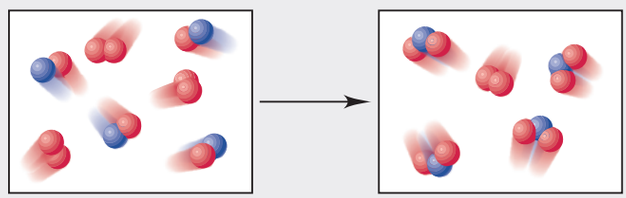
In this experiment, Cl_2 is limiting because it forms fewer moles of ClF_3. Calculating grams of ClF_3 formed:

$$\text{Mass (g) of } ClF_3 = 1.50\ \cancel{\text{mol } ClF_3} \times \frac{92.45\text{ g } ClF_3}{1\ \cancel{\text{mol } ClF_3}} = 139\text{ g } ClF_3$$

Check Let's check our reasoning that Cl_2 is the limiting reactant by assuming, for the moment, that F_2 is limiting. In that case, all 3.00 mol of F_2 would react to form 2.00 mol of ClF_3. Based on the balanced equation, however, that amount of product would require that 1.00 mol of Cl_2 reacted. But that is impossible because only 0.750 mol of Cl_2 is present.

Comment Note that a reactant can be limiting even though it is present in the greater amount. It is the *reactant molar ratio in the balanced equation* that is the determining factor. In both parts (a) and (b), F_2 is present in greater amount than Cl_2. However, in (a), the F_2/Cl_2 ratio is 6/3, or 2/1, which is less than the required molar ratio of 3/1, so F_2 is limiting; in (b), the F_2/Cl_2 ratio is 3.00/0.750, greater than the required 3/1, so F_2 is in excess.

FOLLOW-UP PROBLEM 3.9 B_2 (red spheres) reacts with AB as shown below:

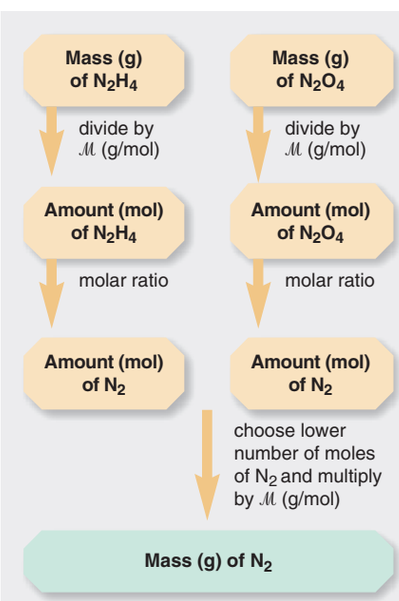


(a) Write a balanced equation for the reaction, and determine the limiting reactant.
(b) How many moles of product can form from the reaction of 1.5 mol of each reactant?

SAMPLE PROBLEM 3.10 Calculating Amounts of Reactant and Product in a Limiting-Reactant Problem

Problem A fuel mixture used in the early days of rocketry is composed of two liquids, hydrazine (N_2H_4) and dinitrogen tetraoxide (N_2O_4), which ignite on contact to form nitrogen gas and water vapor. How many grams of nitrogen gas form when 1.00×10^2 g of N_2H_4 and 2.00×10^2 g of N_2O_4 are mixed?

Plan We first write the balanced equation. *Because the amounts of two reactants are given, we know this is a limiting-reactant problem.* To determine which reactant is limiting, we calculate the mass of N_2 formed from each reactant *assuming an excess of the other.* We convert the grams of each reactant to moles and use the appropriate molar ratio to find the moles of N_2 each forms. Whichever yields *less* N_2 is the limiting reactant. Then, we convert this lower number of moles of N_2 to mass. The roadmap shows the steps.



Solution Writing the balanced equation:

$$2N_2H_4(l) + N_2O_4(l) \longrightarrow 3N_2(g) + 4H_2O(g)$$

Finding the moles of N_2 from the moles of N_2H_4 (if N_2H_4 is limiting):

$$\text{Moles of } N_2H_4 = 1.00\times10^2\ \cancel{\text{g } N_2H_4} \times \frac{1\text{ mol } N_2H_4}{32.05\ \cancel{\text{g } N_2H_4}} = 3.12\text{ mol } N_2H_4$$

$$\text{Moles of } N_2 = 3.12\ \cancel{\text{mol } N_2H_4} \times \frac{3\text{ mol } N_2}{2\ \cancel{\text{mol } N_2H_4}} = \mathbf{4.68\ mol\ N_2}$$

Finding the moles of N_2 from the moles of N_2O_4 (if N_2O_4 is limiting):

$$\text{Moles of } N_2O_4 = 2.00\times10^2\ \cancel{\text{g } N_2O_4} \times \frac{1\text{ mol } N_2O_4}{92.02\ \cancel{\text{g } N_2O_4}} = 2.17\text{ mol } N_2O_4$$

$$\text{Moles of } N_2 = 2.17\ \cancel{\text{mol } N_2O_4} \times \frac{3\text{ mol } N_2}{1\ \cancel{\text{mol } N_2O_4}} = \mathbf{6.51\ mol\ N_2}$$

Thus, N_2H_4 is the limiting reactant because it yields fewer moles of N_2. Converting from moles of N_2 to grams:

$$\text{Mass (g) of } N_2 = 4.68\ \cancel{\text{mol } N_2} \times \frac{28.02\text{ g } N_2}{1\ \cancel{\text{mol } N_2}} = 131\text{ g } N_2$$

Check The mass of N_2O_4 is greater than that of N_2H_4, but there are fewer moles of N_2O_4 because its $\mathcal{M}$ is much higher. Round off to check the math: for N_2H_4, 100 g N_2H_4 × 1 mol/32 g ≈ 3 mol; ~3 mol × $\frac{3}{2}$ ≈ 4.5 mol N_2; ~4.5 mol × 30 g/mol ≈ 135 g N_2.

Comment 1. Here are two *common mistakes* in solving limiting-reactant problems:

- The limiting reactant is not the *reactant* present in fewer moles (2.17 mol of N_2O_4 vs. 3.12 mol of N_2H_4). Rather, it is the reactant that forms fewer moles of *product.*
- Similarly, the limiting reactant is not the *reactant* present in lower mass. Rather, it is the reactant that forms the lower mass of *product.*

2. Here is an *alternative approach* to finding the limiting reactant. Find the moles of each reactant that would be needed to react with the other reactant. Then see which amount actually given in the problem is sufficient. That substance is in excess, and the other substance is limiting. For example, the balanced equation shows that 2 mol of N_2H_4 reacts with 1 mol of N_2O_4. The moles of N_2O_4 needed to react with the given moles of N_2H_4 are

$$\text{Moles of } N_2O_4 \text{ needed} = 3.12\ \cancel{\text{mol } N_2H_4} \times \frac{1 \text{ mol } N_2O_4}{2\ \cancel{\text{mol } N_2H_4}} = 1.56 \text{ mol } N_2O_4$$

The moles of N_2H_4 needed to react with the given moles of N_2O_4 are

$$\text{Moles of } N_2H_4 \text{ needed} = 2.17\ \cancel{\text{mol } N_2O_4} \times \frac{2 \text{ mol } N_2H_4}{1\ \cancel{\text{mol } N_2O_4}} = 4.34 \text{ mol } N_2H_4$$

We are given 2.17 mol of N_2O_4, which is *more* than the amount of N_2O_4 needed (1.56 mol) to react with the given amount of N_2H_4, and we are given 3.12 mol of N_2H_4, which is *less* than the amount of N_2H_4 needed (4.34 mol) to react with the given amount of N_2O_4. Therefore, N_2H_4 is limiting, and N_2O_4 is in excess. Once we determine this, we continue with the final calculation to find the amount of N_2.

FOLLOW-UP PROBLEM 3.10 How many grams of solid aluminum sulfide can be prepared by the reaction of 10.0 g of aluminum and 15.0 g of sulfur? How much of the nonlimiting reactant is in excess?

Chemical Reactions in Practice: Theoretical, Actual, and Percent Yields

Up until now, we've been optimistic about the amount of product obtained from a reaction. We have assumed that 100% of the limiting reactant becomes product, that ideal separation and purification methods exist for isolating the product, and that we use perfect lab technique to collect all the product formed. In other words, we have assumed that we obtain the **theoretical yield,** the amount indicated by the stoichiometrically equivalent molar ratio in the balanced equation.

It's time to face reality. The theoretical yield is *never* obtained, for reasons that are largely uncontrollable. For one thing, although the major reaction predominates, many reactant mixtures also proceed through one or more **side reactions** that form smaller amounts of different products (Figure 3.9). In the rocket fuel reaction in Sample Problem 3.10, for example, the reactants might form some NO in the following side reaction:

$$2N_2O_4(l) + N_2H_4(l) \longrightarrow 6NO(g) + 2H_2O(g)$$

This reaction decreases the amounts of reactants available for N_2 production (see Problem 3.81 at the end of the chapter). Even more important, as we'll discuss in later chapters, many reactions seem to stop before they are complete, which leaves some limiting reactant unused. But, even when a reaction does go completely to product, losses occur in virtually every step of the separation procedure used to isolate the product from the reaction mixture. With careful technique, you can minimize these losses but never eliminate them.

The amount of product that you actually obtain is the **actual yield.** Theoretical and actual yields are expressed in units of amount (moles) or mass (grams).

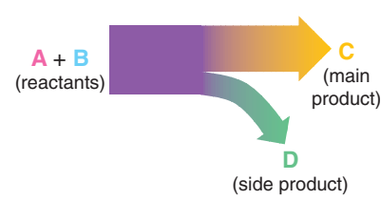


Figure 3.9 The effect of side reactions on yield. One reason the theoretical yield is never obtained is that other reactions lead some of the reactants along side paths to form undesired products.

The **percent yield (% yield)** is the actual yield expressed as a percentage of the theoretical yield:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \qquad (3.7)$$

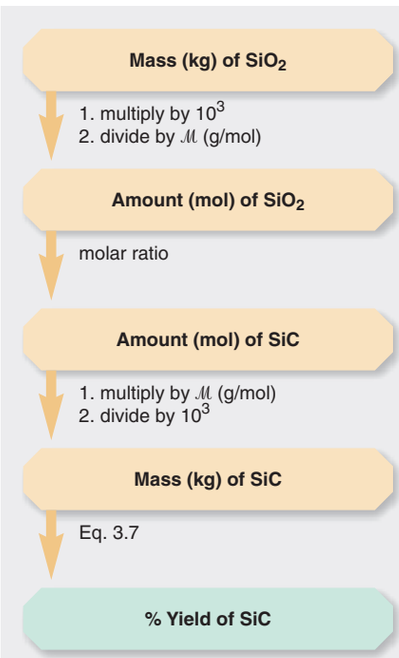
Because the actual yield *must* be less than the theoretical yield, the percent yield is *always* less than 100%. In multistep reaction sequences, the percent yield of each step is expressed as a fraction and multiplied by the others to find the overall yield. The result may sometimes be surprising. For example, suppose a six-step reaction sequence has a 90.0% yield for each step, which is quite high. Even so, the overall percent yield would be

$$\text{Overall \% yield} = (0.900 \times 0.900 \times 0.900 \times 0.900 \times 0.900 \times 0.900) \times 100 = 53.1\%$$

SAMPLE PROBLEM 3.11 Calculating Percent Yield

Problem Silicon carbide (SiC) is an important ceramic material that is made by allowing sand (silicon dioxide, SiO_2) to react with powdered carbon at high temperature. Carbon monoxide is also formed. When 100.0 kg of sand is processed, 51.4 kg of SiC is recovered. What is the percent yield of SiC from this process?

Plan We are given the actual yield of SiC (51.4 kg), so we need the theoretical yield to calculate the percent yield. After writing the balanced equation, we convert the given mass of SiO_2 (100.0 kg) to amount (mol). We use the molar ratio to find the amount of SiC formed and convert that amount to mass (kg) to obtain the theoretical yield [see Sample Problem 3.8(c)]. Then, we use Equation 3.7 to find the percent yield (see the roadmap).



Solution Writing the balanced equation:

$$SiO_2(s) + 3C(s) \longrightarrow SiC(s) + 2CO(g)$$

Converting from kilograms of SiO_2 to moles:

$$\text{Moles of } SiO_2 = 100.0 \text{ kg } SiO_2 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol } SiO_2}{60.09 \text{ g } SiO_2} = 1664 \text{ mol } SiO_2$$

Converting from moles of SiO_2 to moles of SiC: The molar ratio is 1 mol SiC/1 mol SiO_2, so

$$\text{Moles of } SiO_2 = \text{moles of SiC} = 1664 \text{ mol SiC}$$

Converting from moles of SiC to kilograms:

$$\text{Mass (kg) of SiC} = 1664 \text{ mol SiC} \times \frac{40.10 \text{ g SiC}}{1 \text{ mol SiC}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 66.73 \text{ kg SiC}$$

Calculating the percent yield:

$$\text{\% yield of SiC} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{51.4 \text{ kg}}{66.73 \text{ kg}} \times 100 = 77.0\%$$

Check Rounding shows that the mass of SiC seems correct: ~1500 mol × 40 g/mol × 1 kg/1000 g = 60 kg. The molar ratio of SiC:SiO_2 is 1:1, and the $\mathcal{M}$ of SiC is about two-thirds ($\sim\frac{40}{60}$) the $\mathcal{M}$ of SiO_2, so 100 kg of SiO_2 should form about 66 kg of SiC.

FOLLOW-UP PROBLEM 3.11 Marble (calcium carbonate) reacts with hydrochloric acid solution to form calcium chloride solution, water, and carbon dioxide. What is the percent yield of carbon dioxide if 3.65 g of the gas is collected when 10.0 g of marble reacts?

SECTION SUMMARY

The substances in a balanced equation are related to each other by stoichiometrically equivalent molar ratios, which can be used as conversion factors to find the moles of one substance given the moles of another. In limiting-reactant problems, the amounts of two (or more) reactants are given, and one of them limits the amount of product that forms. The limiting reactant is the one that forms the lower amount of product. In practice, side reactions, incomplete reactions, and physical losses result

in an actual yield of product that is less than the theoretical yield, the amount based solely on the molar ratio. The percent yield is the actual yield expressed as a percentage of the theoretical yield. In multistep reaction sequences, the overall yield is found by multiplying the percent yields for each step.

3.5 FUNDAMENTALS OF SOLUTION STOICHIOMETRY

Many environmental reactions and almost all biochemical reactions occur in solution, so an understanding of reactions in solution is extremely important in chemistry and related sciences. We'll discuss solution chemistry at many places in the text, but here we focus on solution stoichiometry. Only one aspect of the stoichiometry of dissolved substances is different from what we've seen so far. We know the amounts of pure substances by converting their masses directly into moles. For dissolved substances, we must know the *concentration*—the number of moles present in a certain volume of solution—to find the volume that contains a given number of moles. Of the various ways to express concentration, the most important is *molarity,* so we discuss it here (and wait until Chapter 13 to discuss the other ways). Then, we see how to prepare a solution of a specific molarity and how to use solutions in stoichiometric calculations.

Expressing Concentration in Terms of Molarity

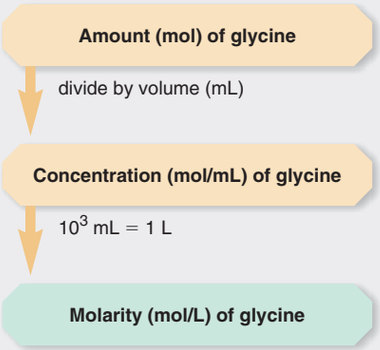
A typical solution consists of a smaller amount of one substance, the **solute,** dissolved in a larger amount of another substance, the **solvent.** When a solution forms, the solute's individual chemical entities become evenly dispersed throughout the available volume and surrounded by solvent molecules. The **concentration** of a solution is usually expressed as *the amount of solute dissolved in a given amount of solution.* Concentration is an *intensive* quantity (like density or temperature) and thus independent of the volume of solution: a 50-L tank of a given solution has the *same concentration* (solute amount/solution amount) as a 50-mL beaker of the solution. **Molarity *(M)*** expresses the concentration in units of *moles of solute per liter of solution:*

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \quad \text{or} \quad M = \frac{\text{mol solute}}{\text{L soln}} \tag{3.8}$$

SAMPLE PROBLEM 3.12 Calculating the Molarity of a Solution

Problem Glycine (H_2NCH_2COOH) is the simplest amino acid. What is the molarity of an aqueous solution that contains 0.715 mol of glycine in 495 mL?

Plan The molarity is the number of moles of solute in each liter of solution. We are given the number of moles (0.715 mol) and the volume (495 mL), so we divide moles by volume and convert the volume to liters to find the molarity (see the roadmap).



Solution

$$\text{Molarity} = \frac{0.715 \text{ mol glycine}}{495 \cancel{\text{mL soln}}} \times \frac{1000 \cancel{\text{mL}}}{1 \text{ L}} = 1.44\ M \text{ glycine}$$

Check A quick look at the math shows about 0.7 mol of glycine in about 0.5 L of solution, so the concentration should be about 1.4 mol/L, or 1.4 *M.*

FOLLOW-UP PROBLEM 3.12 How many moles of KI are in 84 mL of 0.50 *M* KI?

Mole-Mass-Number Conversions Involving Solutions

Molarity can be thought of as a conversion factor used to convert between volume of solution and amount (mol) of solute, from which we then find the mass or the number of entities of solute. Figure 3.10 (on the next page) shows this new stoichiometric relationship, and Sample Problem 3.13 applies it.

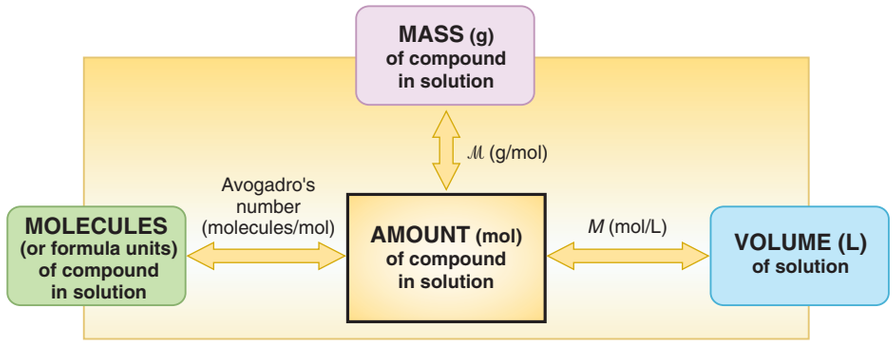
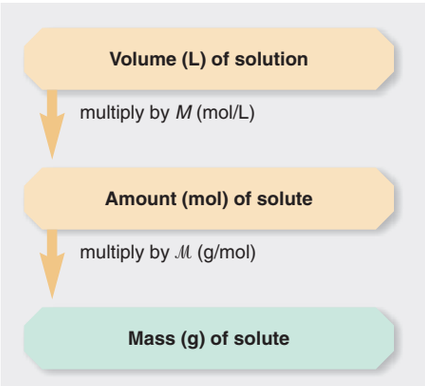


Figure 3.10 Summary of mass-mole-number-volume relationships in solution. The amount (in moles) of a compound in solution is related to the volume of solution in liters through the molarity (*M*) in moles per liter. The other relationships shown are identical to those in Figure 3.4, except that here they refer to the quantities *in solution*. As in previous cases, to find the quantity of substance expressed in one form or another, convert the given information to moles first.



SAMPLE PROBLEM 3.13 Calculating Mass of Solute in a Given Volume of Solution

Problem A *buffered* solution maintains acidity as a reaction occurs. In living cells, phosphate ions play a key buffering role, so biochemists often study reactions in such solutions. How many grams of solute are in 1.75 L of 0.460 *M* sodium monohydrogen phosphate?

Plan We know the solution volume (1.75 L) and molarity (0.460 *M*), and we need the mass of solute. We use the known quantities to find the amount (mol) of solute and then convert moles to grams with the solute molar mass, as shown in the roadmap.

Solution Calculating moles of solute in solution:

$$\text{Moles of } Na_2HPO_4 = 1.75\ \cancel{\text{L soln}} \times \frac{0.460\ \text{mol } Na_2HPO_4}{1\ \cancel{\text{L soln}}} = 0.805\ \text{mol } Na_2HPO_4$$

Converting from moles of solute to grams:

$$\text{Mass (g) } Na_2HPO_4 = 0.805\ \cancel{\text{mol } Na_2HPO_4} \times \frac{141.96\ \text{g } Na_2HPO_4}{1\ \cancel{\text{mol } Na_2HPO_4}} = 114\ \text{g } Na_2HPO_4$$

Check The answer seems to be correct: ~1.8 L of 0.5 mol/L contains 0.9 mol, and 150 g/mol × 0.9 mol = 135 g, which is close to 114 g of solute.

FOLLOW-UP PROBLEM 3.13 In biochemistry laboratories, solutions of sucrose (table sugar, $C_{12}H_{22}O_{11}$) are used in high-speed centrifuges to separate the parts of a biological cell. How many liters of 3.30 *M* sucrose contain 135 g of solute?

Animation: Making a Solution
Online Learning Center

Dilution of Molar Solutions

A concentrated solution (higher molarity) can be converted to a dilute solution (lower molarity) by *adding solvent* to it. The solution volume increases while the number of moles of solute remains the same. Thus, a given volume of the final (dilute) solution contains fewer solute particles and has a lower concentration than the original (concentrated) solution did (Figure 3.11). If various low concentrations of a solution are used frequently, it is common practice to prepare a more concentrated solution (called a *stock solution*), which is stored and diluted as needed.

SAMPLE PROBLEM 3.14 Preparing a Dilute Solution from a Concentrated Solution

Problem Isotonic saline is a 0.15 *M* aqueous solution of NaCl that simulates the total concentration of ions found in many cellular fluids. Its uses range from a cleansing rinse for contact lenses to a washing medium for red blood cells. How would you prepare 0.80 L of isotonic saline from a 6.0 *M* stock solution?

Plan To dilute a concentrated solution, we add only solvent, so the *moles of solute are the same in both solutions.* We know the volume (0.80 L) and molarity (0.15 *M*) of the dilute (dil) NaCl solution we need, so we find the moles of NaCl it contains and then find the volume of concentrated (conc; 6.0 *M*) NaCl solution that contains the same number of moles. Then, we dilute this volume with solvent *up to* the final volume (see roadmap).

Solution Finding moles of solute in dilute solution:

$$\text{Moles of NaCl in dil soln} = 0.80\ \cancel{\text{L soln}} \times \frac{0.15\ \text{mol NaCl}}{1\ \cancel{\text{L soln}}} = 0.12\ \text{mol NaCl}$$

Finding moles of solute in concentrated solution: Because we add only solvent to dilute the solution,

$$\text{Moles of NaCl in dil soln} = \text{moles of NaCl in conc soln} = 0.12\ \text{mol NaCl}$$

Finding the volume of concentrated solution that contains 0.12 mol of NaCl:

$$\text{Volume (L) of conc NaCl soln} = 0.12\ \cancel{\text{mol NaCl}} \times \frac{1\ \text{L soln}}{6.0\ \cancel{\text{mol NaCl}}} = 0.020\ \text{L soln}$$

To prepare 0.80 L of dilute solution, place 0.020 L of 6.0 *M* NaCl in a 1.0-L cylinder, add distilled water (~780 mL) to the 0.80-L mark, and stir thoroughly.

Check The answer seems reasonable because a small volume of concentrated solution is used to prepare a large volume of dilute solution. Also, the ratio of volumes (0.020 L:0.80 L) is the same as the ratio of concentrations (0.15 *M*:6.0 *M*).

Comment An *alternative approach* to solving dilution problems uses the formula

$$M_{\text{dil}} \times V_{\text{dil}} = \text{number of moles} = M_{\text{conc}} \times V_{\text{conc}} \qquad (3.9)$$

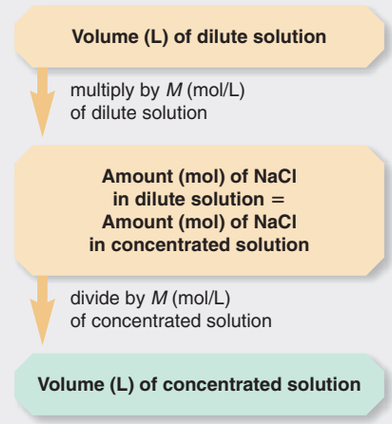
where the *M* and *V* terms are the molarity and volume of the *dil*ute and *conc*entrated solutions. In this problem, we need the volume of concentrated solution, V_{conc}:

$$V_{\text{conc}} = \frac{M_{\text{dil}} \times V_{\text{dil}}}{M_{\text{conc}}} = \frac{0.15\ \cancel{M} \times 0.80\ \text{L}}{6.0\ \cancel{M}} = 0.020\ \text{L}$$

The method worked out in the Solution (above) is actually the same calculation broken into two parts to emphasize the thinking process:

$$V_{\text{conc}} = 0.80\ \cancel{\text{L}} \times \frac{0.15\ \cancel{\text{mol NaCl}}}{1\ \cancel{\text{L}}} \times \frac{1\ \text{L}}{6.0\ \cancel{\text{mol NaCl}}} = 0.020\ \text{L}$$

FOLLOW-UP PROBLEM 3.14 To prepare a fertilizer, an engineer dilutes a stock solution of sulfuric acid by adding 25.0 m^3 of 7.50 *M* acid to enough water to make 500. m^3. What is the mass (in g) of sulfuric acid per milliliter of the diluted solution?



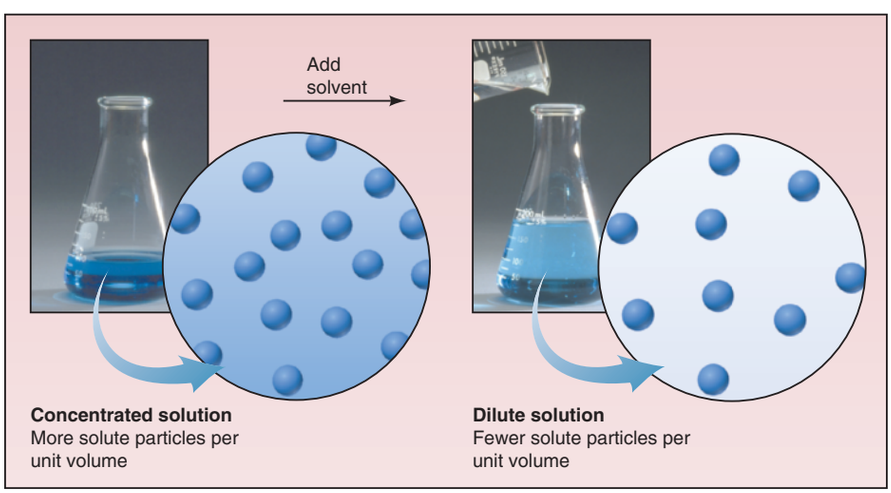


Figure 3.11 Converting a concentrated solution to a dilute solution. When a solution is diluted, only solvent is added. The solution volume increases while the total number of moles of solute remains the same. Therefore, as shown in the blow-up views, a unit volume of concentrated solution contains more solute particles than the same unit volume of dilute solution.

Stoichiometry of Chemical Reactions in Solution

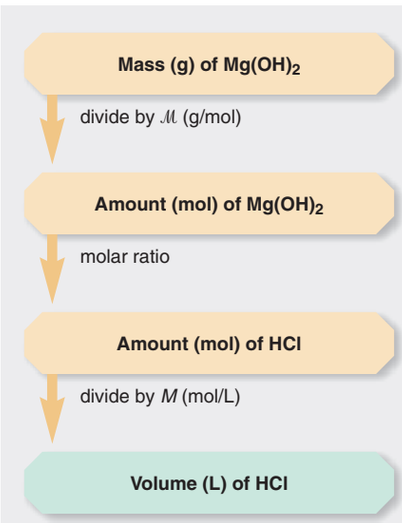
Animation: Preparing a Solution by Dilution
Online Learning Center

Solving stoichiometry problems for reactions in solution requires the same approach as before, with the additional step of converting the volume of reactant or product to moles: (1) balance the equation, (2) find the number of moles of one substance, (3) relate it to the stoichiometrically equivalent number of moles of another substance, and (4) convert to the desired units.

SAMPLE PROBLEM 3.15 Calculating Amounts of Reactants and Products for a Reaction in Solution

Problem Specialized cells in the stomach release HCl to aid digestion. If they release too much, the excess can be neutralized with an antacid to avoid discomfort. A common antacid contains magnesium hydroxide, $Mg(OH)_2$, which reacts with the acid to form water and magnesium chloride solution. As a government chemist testing commercial antacids, you use 0.10 *M* HCl to simulate the acid concentration in the stomach. How many liters of "stomach acid" react with a tablet containing 0.10 g of $Mg(OH)_2$?

Plan We know the mass of $Mg(OH)_2$ (0.10 g) that reacts and the acid concentration (0.10 *M*), and we must find the acid volume. After writing the balanced equation, we convert the grams of $Mg(OH)_2$ to moles, use the molar ratio to find the moles of HCl that react with these moles of $Mg(OH)_2$, and then use the molarity of HCl to find the volume that contains this number of moles. The steps appear in the roadmap.



Solution Writing the balanced equation:

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

Converting from grams of $Mg(OH)_2$ to moles:

$$\text{Moles of } Mg(OH)_2 = 0.10\ \cancel{\text{g } Mg(OH)_2} \times \frac{1 \text{ mol } Mg(OH)_2}{58.33\ \cancel{\text{g } Mg(OH)_2}} = 1.7\times10^{-3} \text{ mol } Mg(OH)_2$$

Converting from moles of $Mg(OH)_2$ to moles of HCl:

$$\text{Moles of HCl} = 1.7\times10^{-3}\ \cancel{\text{mol } Mg(OH)_2} \times \frac{2 \text{ mol HCl}}{1\ \cancel{\text{mol } Mg(OH)_2}} = 3.4\times10^{-3} \text{ mol HCl}$$

Converting from moles of HCl to liters:

$$\text{Volume (L) of HCl} = 3.4\times10^{-3}\ \cancel{\text{mol HCl}} \times \frac{1 \text{ L}}{0.10\ \cancel{\text{mol HCl}}}$$

$$= \boxed{3.4\times10^{-2} \text{ L}}$$

Check The size of the answer seems reasonable: a small volume of dilute acid (0.034 L of 0.10 *M*) reacts with a small amount of antacid (0.0017 mol).

Comment The reaction as written is an oversimplification; in reality, HCl and $MgCl_2$ exist as separated ions in solution. This point will be covered in great detail in Chapters 4 and 18.

FOLLOW-UP PROBLEM 3.15 Another active ingredient found in some antacids is aluminum hydroxide. Which is more effective at neutralizing stomach acid, magnesium hydroxide or aluminum hydroxide? [*Hint:* Effectiveness refers to the amount of acid that reacts with a given mass of antacid. You already know the effectiveness of 0.10 g of $Mg(OH)_2$.]

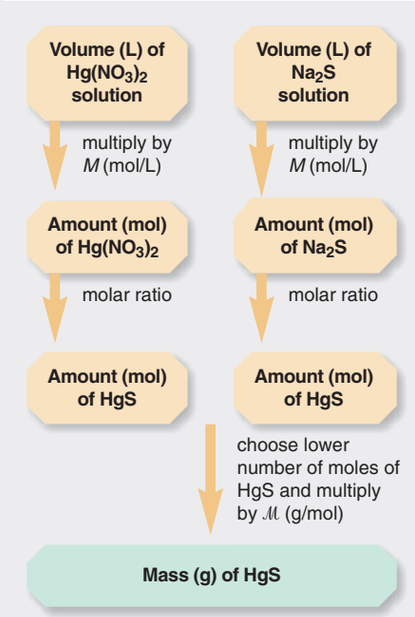
In limiting-reactant problems for reactions in solution, we first determine which reactant is limiting and then determine the yield, as demonstrated in the next sample problem.

SAMPLE PROBLEM 3.16 Solving Limiting-Reactant Problems for Reactions in Solution

Problem Mercury and its compounds have many uses, from fillings for teeth (as a mixture with silver, copper, and tin) to the industrial production of chlorine. Because of their toxicity, however, soluble mercury compounds, such as mercury(II) nitrate, must be removed from industrial wastewater. One removal method reacts the wastewater with

sodium sulfide solution to produce solid mercury(II) sulfide and sodium nitrate solution. In a laboratory simulation, 0.050 L of 0.010 *M* mercury(II) nitrate reacts with 0.020 L of 0.10 *M* sodium sulfide. How many grams of mercury(II) sulfide form?

Plan This is a limiting-reactant problem because *the amounts of two reactants are given.* After balancing the equation, we must determine the limiting reactant. The molarity (0.010 *M*) and volume (0.050 L) of the mercury(II) nitrate solution tell us the moles of one reactant, and the molarity (0.10 *M*) and volume (0.020 L) of the sodium sulfide solution tell us the moles of the other. Then, as in Sample Problem 3.10, we use the molar ratio to find the moles of HgS that form from each reactant, *assuming the other reactant is present in excess.* The limiting reactant is the one that forms fewer moles of HgS, which we convert to mass using the HgS molar mass. The roadmap shows the process.



Solution Writing the balanced equation:

$$Hg(NO_3)_2(aq) + Na_2S(aq) \longrightarrow HgS(s) + 2NaNO_3(aq)$$

Finding moles of HgS assuming $Hg(NO_3)_2$ is limiting: Combining the steps gives

$$\text{Moles of HgS} = 0.050\ \cancel{\text{L soln}} \times \frac{0.010\ \cancel{\text{mol Hg(NO}_3)_2}}{1\ \cancel{\text{L soln}}} \times \frac{1\ \text{mol HgS}}{1\ \cancel{\text{mol Hg(NO}_3)_2}}$$

$$= \mathbf{5.0\times10^{-4}\ mol\ HgS}$$

Finding moles of HgS assuming Na_2S is limiting: Combining the steps gives

$$\text{Moles of HgS} = 0.020\ \cancel{\text{L soln}} \times \frac{0.10\ \cancel{\text{mol Na}_2\text{S}}}{1\ \cancel{\text{L soln}}} \times \frac{1\ \text{mol HgS}}{1\ \cancel{\text{mol Na}_2\text{S}}}$$

$$= \mathbf{2.0\times10^{-3}\ mol\ HgS}$$

$Hg(NO_3)_2$ is the limiting reactant because it forms fewer moles of HgS. Converting the moles of HgS formed from $Hg(NO_3)_2$ to grams:

$$\text{Mass (g) of HgS} = 5.0\times10^{-4}\ \cancel{\text{mol HgS}} \times \frac{232.7\ \text{g HgS}}{1\ \cancel{\text{mol HgS}}}$$

$$= 0.12\ \text{g HgS}$$

Check As a check, let's use the alternative method for finding the limiting reactant (see Comment in Sample Problem 3.10). Finding moles of reactants available:

$$\text{Moles of Hg(NO}_3)_2 = 0.050\ \cancel{\text{L soln}} \times \frac{0.010\ \text{mol Hg(NO}_3)_2}{1\ \cancel{\text{L soln}}} = 5.0\times10^{-4}\ \text{mol Hg(NO}_3)_2$$

$$\text{Moles of Na}_2\text{S} = 0.020\ \cancel{\text{L soln}} \times \frac{0.10\ \text{mol Na}_2\text{S}}{1\ \cancel{\text{L soln}}} = 2.0\times10^{-3}\ \text{mol Na}_2\text{S}$$

The molar ratio of the reactants is 1 $Hg(NO_3)_2$/1 Na_2S. Therefore, $Hg(NO_3)_2$ is limiting because there are fewer moles of it than are needed to react with the moles of Na_2S. Finding grams of product from moles of limiting reactant and the molar ratio:

$$\text{Mass (g) of HgS} = 5.0\times10^{-4}\ \cancel{\text{mol Hg(NO}_3)_2} \times \frac{1\ \cancel{\text{mol HgS}}}{1\ \cancel{\text{mol Hg(NO}_3)_2}} \times \frac{232.7\ \text{g HgS}}{1\ \cancel{\text{mol HgS}}}$$

$$= 0.12\ \text{g HgS}$$

FOLLOW-UP PROBLEM 3.16 Even though gasoline sold in the United States no longer contains lead, this metal persists in the environment as a poison. Despite their toxicity, many compounds of lead are still used to make pigments.

(a) What volume of 1.50 *M* lead(II) acetate contains 0.400 mol of Pb^{2+} ions?

(b) When this volume reacts with 125 mL of 3.40 *M* sodium chloride, how many grams of solid lead(II) chloride can form? (Sodium acetate solution also forms.)

SECTION SUMMARY

When reactions occur in solution, reactant and product amounts are given in terms of concentration and volume. Molarity is the number of moles of solute dissolved in one liter of solution. Using molarity as a conversion factor, we apply the principles of stoichiometry to all aspects of reactions in solution.

For Review and Reference (Numbers in parentheses refer to pages, unless noted otherwise.)

Learning Objectives

To help you review these learning objectives, the numbers of related sections (§), sample problems (SP), and upcoming end-of-chapter problems (EP) are listed in parentheses.

1. Realize the usefulness of the mole concept, and use the relation between molecular (or formula) mass and molar mass to calculate the molar mass of any substance (§ 3.1) (EPs 3.1–3.5, 3.7–3.10)
2. Understand the relationships among amount of substance (in moles), mass (in grams), and number of chemical entities and convert from one to any other (§ 3.1) (SP 3.1, 3.2) (EPs 3.6, 3.11–3.16, 3.19)
3. Use mass percent to find the mass of element in a given mass of compound (§ 3.1) (SP 3.3) (EPs 3.17, 3.18, 3.20–3.23)
4. Determine the empirical and molecular formulas of a compound from mass analysis of its elements (§ 3.2) (SPs 3.4–3.6) (EPs 3.24–3.34)
5. Balance an equation given formulas or names, and use molar ratios to calculate amounts of reactants and products for reactions of pure or dissolved substances (§ 3.3 and 3.5) (SPs 3.7, 3.8, 3.15) (EPs 3.35–3.46, 3.62, 3.71, 3.72)
6. Understand why one reactant limits the yield of product, and solve limiting-reactant problems for reactions of pure or dissolved substances (§ 3.4, 3.5) (SPs 3.9, 3.10, 3.16) (EPs 3.47–3.54, 3.61, 3.73, 3.74)
7. Explain the reasons for lower-than-expected yields and the distinction between theoretical and actual yields, and calculate percent yield (§ 3.4) (SP 3.11) (EPs 3.55–3.60, 3.63)
8. Understand the meaning of concentration and the effect of dilution, and calculate molarity or mass of dissolved solute (§ 3.5) (SPs 3.12–3.14) (EPs 3.64–3.70, 3.75)

Key Terms

stoichiometry (70)

Section 3.1
mole (mol) (70)
Avogadro's number (70)
molar mass ($\mathcal{M}$) (72)

Section 3.2
combustion analysis (80)
isomer (81)

Section 3.3
chemical equation (83)
reactant (83)
product (83)
balancing (stoichiometric) coefficient (83)

Section 3.4
limiting reactant (90)
theoretical yield (93)
side reaction (93)
actual yield (93)
percent yield (% yield) (94)

Section 3.5
solute (95)
solvent (95)
concentration (95)
molarity (M) (95)

Key Equations and Relationships

3.1 Number of entities in one mole (70):

1 mole contains 6.022×10^{23} entities (to 4 sf)

3.2 Converting amount (mol) to mass using $\mathcal{M}$ (73):

$$\text{Mass (g)} = \text{no. of moles} \times \frac{\text{no. of grams}}{1\text{ mol}}$$

3.3 Converting mass to amount (mol) using $1/\mathcal{M}$ (73):

$$\text{No. of moles} = \text{mass (g)} \times \frac{1\text{ mol}}{\text{no. of grams}}$$

3.4 Converting amount (mol) to number of entities (73):

$$\text{No. of entities} = \text{no. of moles} \times \frac{6.022\times10^{23}\text{ entities}}{1\text{ mol}}$$

3.5 Converting number of entities to amount (mol) (73):

$$\text{No. of moles} = \text{no. of entities} \times \frac{1\text{ mol}}{6.022\times10^{23}\text{ entities}}$$

3.6 Calculating mass % (75):

$$\text{Mass \% of element X} = \frac{\text{moles of X in formula} \times \text{molar mass of X (g/mol)}}{\text{mass (g) of 1 mol of compound}} \times 100$$

3.7 Calculating percent yield (94):

$$\text{\% yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

3.8 Defining molarity (95):

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \quad \text{or} \quad M = \frac{\text{mol solute}}{\text{L soln}}$$

3.9 Diluting a concentrated solution (97):

$$M_{\text{dil}} \times V_{\text{dil}} = \text{number of moles} = M_{\text{conc}} \times V_{\text{conc}}$$

Brief Solutions to Follow-up Problems

3.1 (a) Moles of C $= 315\ \cancel{\text{mg C}} \times \dfrac{1\ \cancel{\text{g}}}{10^3\ \cancel{\text{mg}}} \times \dfrac{1\text{ mol C}}{12.01\ \cancel{\text{g C}}}$

$= 2.62\times10^{-2}$ mol C

(b) Mass (g) of Mn $= 3.22\times10^{20}\ \cancel{\text{Mn atoms}} \times \dfrac{1\ \cancel{\text{mol Mn}}}{6.022\times10^{23}\ \cancel{\text{Mn atoms}}} \times \dfrac{54.94\text{ g Mn}}{1\ \cancel{\text{mol Mn}}}$

$= 2.94\times10^{-2}$ g Mn

3.2 (a) Mass (g) of P_4O_{10}

$= 4.65\times10^{22}\ \cancel{\text{molecules } P_4O_{10}} \times \dfrac{1\ \cancel{\text{mol } P_4O_{10}}}{6.022\times10^{23}\ \cancel{\text{molecules } P_4O_{10}}} \times \dfrac{283.88\text{ g } P_4O_{10}}{1\ \cancel{\text{mol } P_4O_{10}}}$

$= 21.9$ g P_4O_{10}

(b) No. of P atoms $= 4.65\times10^{22}$ ~~molecules P_4O_{10}~~

$$\times \frac{4 \text{ atoms P}}{1 \text{ molecule } P_4O_{10}}$$

$= 1.86\times10^{23}$ P atoms

3.3 (a) Mass % of N $= \dfrac{2 \text{ mol N} \times \dfrac{14.01 \text{ g N}}{1 \text{ mol N}}}{80.05 \text{ g } NH_4NO_3} \times 100$

$= 35.00$ mass % N

(b) Mass (g) of N $= 35.8 \text{ kg } NH_4NO_3 \times \dfrac{10^3 \text{ g}}{1 \text{ kg}} \times \dfrac{0.3500 \text{ g N}}{1 \text{ g } NH_4NO_3}$

$= 1.25\times10^4$ g N

3.4 Moles of S $= 2.88 \text{ g S} \times \dfrac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.0898$ mol S

Moles of M $= 0.0898 \text{ mol S} \times \dfrac{2 \text{ mol M}}{3 \text{ mol S}} = 0.0599$ mol M

Molar mass of M $= \dfrac{3.12 \text{ g M}}{0.0599 \text{ mol M}} = 52.1$ g/mol

M is chromium, and M_2S_3 is chromium(III) sulfide.

3.5 Assuming 100.00 g of compound, we have 95.21 g of C and 4.79 g of H:

Moles of C $= 95.21 \text{ g C} \times \dfrac{1 \text{ mol C}}{12.01 \text{ g C}}$

$= 7.928$ mol C

Also, 4.75 mol H

Preliminary formula $= C_{7.928}H_{4.75} \approx C_{1.67}H_{1.00}$

Empirical formula $= C_5H_3$

Whole-number multiple $= \dfrac{252.30 \text{ g/mol}}{63.07 \text{ g/mol}} = 4$

Molecular formula $= C_{20}H_{12}$

3.6 Mass (g) of C $= 0.451 \text{ g } CO_2 \times \dfrac{12.01 \text{ g C}}{44.01 \text{ g } CO_2}$

$= 0.123$ g C

Also, 0.00690 g H

Mass (g) of Cl $= 0.250 \text{ g} - (0.123 \text{ g} + 0.00690 \text{ g}) = 0.120$ g Cl

Moles of elements

$= 0.0102$ mol C; 0.00685 mol H; 0.00339 mol Cl

Empirical formula $= C_3H_2Cl$; multiple $= 2$

Molecular formula $= C_6H_4Cl_2$

3.7 (a) $2Na(s) + 2H_2O(l) \longrightarrow H_2(g) + 2NaOH(aq)$

(b) $2HNO_3(aq) + CaCO_3(s) \longrightarrow$

$H_2O(l) + CO_2(g) + Ca(NO_3)_2(aq)$

(c) $PCl_3(g) + 3HF(g) \longrightarrow PF_3(g) + 3HCl(g)$

(d) $4C_3H_5N_3O_9(l) \longrightarrow$

$12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)$

3.8 $Fe_2O_3(s) + 2Al(s) \longrightarrow Al_2O_3(s) + 2Fe(l)$

(a) Mass (g) of Fe

$= 135 \text{ g Al} \times \dfrac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \dfrac{2 \text{ mol Fe}}{2 \text{ mol Al}} \times \dfrac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$

$= 279$ g Fe

(b) No. of Al atoms $= 1.00 \text{ g } Al_2O_3 \times \dfrac{1 \text{ mol } Al_2O_3}{101.96 \text{ g } Al_2O_3}$

$\times \dfrac{2 \text{ mol Al}}{1 \text{ mol } Al_2O_3} \times \dfrac{6.022\times10^{23} \text{ Al atoms}}{1 \text{ mol Al}}$

$= 1.18\times10^{22}$ Al atoms

3.9 (a) $2AB + B_2 \longrightarrow 2AB_2$

In the boxes, the AB/B_2 ratio is 4/3, which is less than the 2/1 ratio in the equation. Thus, there is not enough AB, so it is the limiting reactant; note that one B_2 is in excess.

(b) Moles of $AB_2 = 1.5 \text{ mol AB} \times \dfrac{2 \text{ mol } AB_2}{2 \text{ mol AB}} = 1.5 \text{ mol } AB_2$

Moles of $AB_2 = 1.5 \text{ mol } B_2 \times \dfrac{2 \text{ mol } AB_2}{1 \text{ mol } B_2} = 3.0 \text{ mol } AB_2$

3.10 $2Al(s) + 3S(s) \longrightarrow Al_2S_3(s)$

Mass (g) of Al_2S_3 formed from Al

$= 10.0 \text{ g Al} \times \dfrac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \dfrac{1 \text{ mol } Al_2S_3}{2 \text{ mol Al}} \times \dfrac{150.17 \text{ g } Al_2S_3}{1 \text{ mol } Al_2S_3}$

$= 27.8 \text{ g } Al_2S_3$

Similarly, mass (g) of Al_2S_3 formed from S $= 23.4$ g Al_2S_3.

Therefore, S is limiting reactant, and 23.4 g of Al_2S_3 can form.

Mass (g) of Al in excess

= total mass of Al − mass of Al used

$= 10.0$ g Al

$$- \left(15.0 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} \times \frac{2 \text{ mol Al}}{3 \text{ mol S}} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}}\right)$$

$= 1.6$ g Al

(We would obtain the same answer if sulfur were shown more correctly as S_8.)

3.11 $CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Theoretical yield (g) of CO_2

$= 10.0 \text{ g } CaCO_3 \times \dfrac{1 \text{ mol } CaCO_3}{100.09 \text{ g } CaCO_3} \times \dfrac{1 \text{ mol } CO_2}{1 \text{ mol } CaCO_3}$

$\times \dfrac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 4.40 \text{ g } CO_2$

% yield $= \dfrac{3.65 \text{ g } CO_2}{4.40 \text{ g } CO_2} \times 100 = 83.0\%$

3.12 Moles of KI $= 84 \text{ mL soln} \times \dfrac{1 \text{ L}}{10^3 \text{ mL}} \times \dfrac{0.50 \text{ mol KI}}{1 \text{ L soln}}$

$= 0.042$ mol KI

3.13 Volume (L) of soln

$= 135 \text{ g sucrose} \times \dfrac{1 \text{ mol sucrose}}{342.30 \text{ g sucrose}} \times \dfrac{1 \text{ L soln}}{3.30 \text{ mol sucrose}}$

$= 0.120$ L soln

3.14 M_{dil} of $H_2SO_4 = \dfrac{7.50\ M \times 25.0 \text{ m}^3}{500. \text{ m}^3} = 0.375\ M\ H_2SO_4$

Mass (g) of H_2SO_4/mL soln

$= \dfrac{0.375 \text{ mol } H_2SO_4}{1 \text{ L soln}} \times \dfrac{1 \text{ L}}{10^3 \text{ mL}} \times \dfrac{98.09 \text{ g } H_2SO_4}{1 \text{ mol } H_2SO_4}$

$= 3.68\times10^{-2}$ g/mL soln

3.15 $Al(OH)_3(s) + 3HCl(aq) \longrightarrow AlCl_3(aq) + 3H_2O(l)$

Volume (L) of HCl consumed

$= 0.10 \text{ g } Al(OH)_3 \times \dfrac{1 \text{ mol } Al(OH)_3}{78.00 \text{ g } Al(OH)_3}$

$\times \dfrac{3 \text{ mol HCl}}{1 \text{ mol } Al(OH)_3} \times \dfrac{1 \text{ L soln}}{0.10 \text{ mol HCl}}$

$= 3.8\times10^{-2}$ L soln

Therefore, $Al(OH)_3$ is more effective than $Mg(OH)_2$.

3.16 (a) Volume (L) of soln

$= 0.400 \text{ mol } Pb^{2+}$

$\times \dfrac{1 \text{ mol } Pb(C_2H_3O_2)_2}{1 \text{ mol } Pb^{2+}} \times \dfrac{1 \text{ L soln}}{1.50 \text{ mol } Pb(C_2H_3O_2)_2}$

$= 0.267$ L soln

(b) $Pb(C_2H_3O_2)_2(aq) + 2NaCl(aq) \longrightarrow$

$PbCl_2(s) + 2NaC_2H_3O_2(aq)$

Mass (g) of $PbCl_2$ from $Pb(C_2H_3O_2)_2$ soln $= 111$ g $PbCl_2$

Mass (g) of $PbCl_2$ from NaCl soln $= 59.1$ g $PbCl_2$

Thus, NaCl is the limiting reactant, and 59.1 g of $PbCl_2$ can form.

Problems

Problems with **colored** numbers are answered in Appendix E. Sections match the text and provide the numbers of relevant sample problems. Bracketed problems are grouped in pairs (indicated by a short rule) that cover the same concept. Comprehensive Problems are based on material from any section or previous chapter.

The Mole

(Sample Problems 3.1 to 3.3)

3.1 The atomic mass of Cl is 35.45 amu, and the atomic mass of Al is 26.98 amu. What are the masses in grams of 2 mol of Al atoms and of 3 mol of Cl atoms?

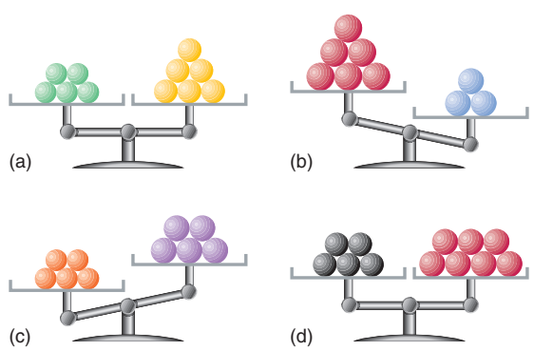
3.2 (a) How many moles of C atoms are in 1 mol of sucrose ($C_{12}H_{22}O_{11}$)?
(b) How many C atoms are in 1 mol of sucrose?

3.3 Why might the expression "1 mol of nitrogen" be confusing? What change would remove any uncertainty? For what other elements might a similar confusion exist? Why?

3.4 How is the molecular mass of a compound the same as the molar mass, and how is it different?

3.5 What advantage is there to using a counting unit (the mole) in chemistry rather than a mass unit?

3.6 Each of the following balances weighs the indicated numbers of atoms of two elements:



Which element—left, right, or neither,
(a) Has the higher molar mass?
(b) Has more atoms per gram?
(c) Has fewer atoms per gram?
(d) Has more atoms per mole?

3.7 Calculate the molar mass of each of the following:
(a) $Sr(OH)_2$ (b) N_2O (c) $NaClO_3$ (d) Cr_2O_3

3.8 Calculate the molar mass of each of the following:
(a) $(NH_4)_3PO_4$ (b) CH_2Cl_2 (c) $CuSO_4{\cdot}5H_2O$ (d) BrF_5

3.9 Calculate the molar mass of each of the following:
(a) SnO_2 (b) BaF_2 (c) $Al_2(SO_4)_3$ (d) $MnCl_2$

3.10 Calculate the molar mass of each of the following:
(a) N_2O_4 (b) C_8H_{10} (c) $MgSO_4{\cdot}7H_2O$ (d) $Ca(C_2H_3O_2)_2$

3.11 Calculate each of the following quantities:
(a) Mass in grams of 0.57 mol of $KMnO_4$
(b) Moles of O atoms in 8.18 g of $Mg(NO_3)_2$
(c) Number of O atoms in 8.1×10^{-3} g of $CuSO_4{\cdot}5H_2O$

3.12 Calculate each of the following quantities:
(a) Mass in kilograms of 3.8×10^{20} molecules of NO_2
(b) Moles of Cl atoms in 0.0425 g of $C_2H_4Cl_2$
(c) Number of H^- ions in 4.92 g of SrH_2

3.13 Calculate each of the following quantities:
(a) Mass in grams of 0.64 mol of $MnSO_4$
(b) Moles of compound in 15.8 g of $Fe(ClO_4)_3$
(c) Number of N atoms in 92.6 g of NH_4NO_2

3.14 Calculate each of the following quantities:
(a) Total number of ions in 38.1 g of CaF_2
(b) Mass in milligrams of 3.58 mol of $CuCl_2{\cdot}2H_2O$
(c) Mass in kilograms of 2.88×10^{22} formula units of $Bi(NO_3)_3{\cdot}5H_2O$

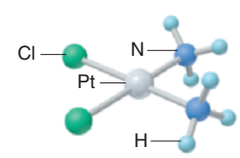
3.15 Calculate each of the following quantities:
(a) Mass in grams of 8.41 mol of copper(I) carbonate
(b) Mass in grams of 2.04×10^{21} molecules of dinitrogen pentaoxide
(c) Number of moles and formula units in 57.9 g of sodium perchlorate
(d) Number of sodium ions, perchlorate ions, Cl atoms, and O atoms in the mass of compound in part (c)

3.16 Calculate each of the following quantities:
(a) Mass in grams of 3.52 mol of chromium(III) sulfate decahydrate
(b) Mass in grams of 9.64×10^{24} molecules of dichlorine heptaoxide
(c) Number of moles and formula units in 56.2 g of lithium sulfate
(d) Number of lithium ions, sulfate ions, S atoms, and O atoms in the mass of compound in part (c)

3.17 Calculate each of the following:
(a) Mass % of H in ammonium bicarbonate
(b) Mass % of O in sodium dihydrogen phosphate heptahydrate

3.18 Calculate each of the following:
(a) Mass % of I in strontium periodate
(b) Mass % of Mn in potassium permanganate

3.19 Cisplatin *(right)*, or Platinol, is a powerful drug used in the treatment of certain cancers. Calculate (a) the moles of compound in 285.3 g of cisplatin; (b) the number of hydrogen atoms in 0.98 mol of cisplatin.



3.20 Propane is widely used in liquid form as a fuel for barbecue grills and camp stoves. For 75.3 g of propane, calculate (a) the moles of compound; (b) the grams of carbon.

3.21 The effectiveness of a nitrogen fertilizer is determined mainly by its mass % N. Rank the following fertilizers in terms of their effectiveness: potassium nitrate; ammonium nitrate; ammonium sulfate; urea, $CO(NH_2)_2$.

3.22 The mineral galena is composed of lead(II) sulfide and has an average density of 7.46 g/cm^3. (a) How many moles of lead(II) sulfide are in 1.00 ft^3 of galena? (b) How many lead atoms are in 1.00 dm^3 of galena?

3.23 Hemoglobin, a protein in red blood cells, carries O_2 from the lungs to the body's cells. Iron (as ferrous ion, Fe^{2+}) makes up

0.33 mass % of hemoglobin. If the molar mass of hemoglobin is 6.8×10^4 g/mol, how many Fe^{2+} ions are in one molecule?

Determining the Formula of an Unknown Compound

(Sample Problems 3.4 to 3.6)

3.24 Which of the following sets of information allows you to obtain the molecular formula of a covalent compound? In each case that allows it, explain how you would proceed (write a solution "Plan").
(a) Number of moles of each type of atom in a given sample of the compound
(b) Mass % of each element and the total number of atoms in a molecule of the compound
(c) Mass % of each element and the number of atoms of one element in a molecule of the compound
(d) Empirical formula and the mass % of each element in the compound
(e) Structural formula of the compound

3.25 What is the empirical formula and empirical formula mass for each of the following compounds?
(a) C_2H_4 (b) $C_2H_6O_2$ (c) N_2O_5 (d) $Ba_3(PO_4)_2$ (e) Te_4I_{16}

3.26 What is the empirical formula and empirical formula mass for each of the following compounds?
(a) C_4H_8 (b) $C_3H_6O_3$ (c) P_4O_{10} (d) $Ga_2(SO_4)_3$ (e) Al_2Br_6

3.27 What is the molecular formula of each compound?
(a) Empirical formula CH_2 ($\mathcal{M}$ = 42.08 g/mol)
(b) Empirical formula NH_2 ($\mathcal{M}$ = 32.05 g/mol)
(c) Empirical formula NO_2 ($\mathcal{M}$ = 92.02 g/mol)
(d) Empirical formula CHN ($\mathcal{M}$ = 135.14 g/mol)

3.28 What is the molecular formula of each compound?
(a) Empirical formula CH ($\mathcal{M}$ = 78.11 g/mol)
(b) Empirical formula $C_3H_6O_2$ ($\mathcal{M}$ = 74.08 g/mol)
(c) Empirical formula HgCl ($\mathcal{M}$ = 472.1 g/mol)
(d) Empirical formula $C_7H_4O_2$ ($\mathcal{M}$ = 240.20 g/mol)

3.29 Determine the empirical formula of each of the following compounds:
(a) 0.063 mol of chlorine atoms combined with 0.22 mol of oxygen atoms
(b) 2.45 g of silicon combined with 12.4 g of chlorine
(c) 27.3 mass % carbon and 72.7 mass % oxygen

3.30 Determine the empirical formula of each of the following compounds:
(a) 0.039 mol of iron atoms combined with 0.052 mol of oxygen atoms
(b) 0.903 g of phosphorus combined with 6.99 g of bromine
(c) A hydrocarbon with 79.9 mass % carbon

3.31 A sample of 0.600 mol of a metal M reacts completely with excess fluorine to form 46.8 g of MF_2.
(a) How many moles of F are in the sample of MF_2 that forms?
(b) How many grams of M are in this sample of MF_2?
(c) What element is represented by the symbol M?

3.32 A sample of 0.370 mol of a metal oxide (M_2O_3) weighs 55.4 g.
(a) How many moles of O are in the sample?
(b) How many grams of M are in the sample?
(c) What element is represented by the symbol M?

3.33 Cortisol ($\mathcal{M}$ = 362.47 g/mol), one of the major steroid hormones, is a key factor in the synthesis of protein. Its profound effect on the reduction of inflammation explains its use in the treatment of rheumatoid arthritis. Cortisol is 69.6% C, 8.34% H, and 22.1% O by mass. What is its molecular formula?

3.34 Menthol ($\mathcal{M}$ = 156.3 g/mol), a strong-smelling substance used in cough drops, is a compound of carbon, hydrogen, and oxygen. When 0.1595 g of menthol was subjected to combustion analysis, it produced 0.449 g of CO_2 and 0.184 g of H_2O. What is menthol's molecular formula?

Writing and Balancing Chemical Equations

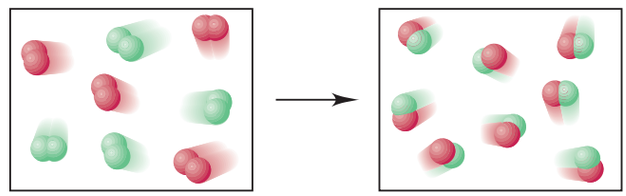
(Sample Problem 3.7)

3.35 In the process of balancing the equation

$$Al + Cl_2 \longrightarrow AlCl_3$$

Student I writes: $Al + Cl_2 \longrightarrow AlCl_2$
Student II writes: $Al + Cl_2 + Cl \longrightarrow AlCl_3$
Student III writes: $2Al + 3Cl_2 \longrightarrow 2AlCl_3$
Is the approach of Student I valid? Student II? Student III? Explain.

3.36 The boxes below represent a chemical reaction between elements A (red) and B (green):



Which of the following best represents the balanced equation for the reaction?
(a) $2A + 2B \longrightarrow A_2 + B_2$ (b) $A_2 + B_2 \longrightarrow 2AB$
(c) $B_2 + 2AB \longrightarrow 2B_2 + A_2$ (d) $4A_2 + 4B_2 \longrightarrow 8AB$

3.37 Write balanced equations for each of the following by inserting the correct coefficients in the blanks:
(a) $\_\_Cu(s) + \_\_S_8(s) \longrightarrow \_\_Cu_2S(s)$
(b) $\_\_P_4O_{10}(s) + \_\_H_2O(l) \longrightarrow \_\_H_3PO_4(l)$
(c) $\_\_B_2O_3(s) + \_\_NaOH(aq) \longrightarrow \_\_Na_3BO_3(aq) + \_\_H_2O(l)$
(d) $\_\_CH_3NH_2(g) + \_\_O_2(g) \longrightarrow \_\_CO_2(g) + \_\_H_2O(g) + \_\_N_2(g)$

3.38 Write balanced equations for each of the following by inserting the correct coefficients in the blanks:
(a) $\_\_Cu(NO_3)_2(aq) + \_\_KOH(aq) \longrightarrow \_\_Cu(OH)_2(s) + \_\_KNO_3(aq)$
(b) $\_\_BCl_3(g) + \_\_H_2O(l) \longrightarrow \_\_H_3BO_3(s) + \_\_HCl(g)$
(c) $\_\_CaSiO_3(s) + \_\_HF(g) \longrightarrow \_\_SiF_4(g) + \_\_CaF_2(s) + \_\_H_2O(l)$
(d) $\_\_(CN)_2(g) + \_\_H_2O(l) \longrightarrow \_\_H_2C_2O_4(aq) + \_\_NH_3(g)$

3.39 Convert the following into balanced equations:
(a) When gallium metal is heated in oxygen gas, it melts and forms solid gallium(III) oxide.
(b) Liquid hexane burns in oxygen gas to form carbon dioxide gas and water vapor.
(c) When solutions of calcium chloride and sodium phosphate are mixed, solid calcium phosphate forms and sodium chloride remains in solution.

3.40 Convert the following into balanced equations:
(a) When lead(II) nitrate solution is added to potassium iodide solution, solid lead(II) iodide forms and potassium nitrate solution remains.

(b) Liquid disilicon hexachloride reacts with water to form solid silicon dioxide, hydrogen chloride gas, and hydrogen gas.
(c) When nitrogen dioxide is bubbled into water, a solution of nitric acid forms and gaseous nitrogen monoxide is released.

Calculating Amounts of Reactant and Product

(Sample Problems 3.8 to 3.11)

3.41 Potassium nitrate decomposes on heating, producing potassium oxide and gaseous nitrogen and oxygen:

$$4KNO_3(s) \longrightarrow 2K_2O(s) + 2N_2(g) + 5O_2(g)$$

To produce 88.6 kg of oxygen, how many (a) moles of KNO_3 must be heated? (b) Grams of KNO_3 must be heated?

3.42 Chromium(III) oxide reacts with hydrogen sulfide (H_2S) gas to form chromium(III) sulfide and water:

$$Cr_2O_3(s) + 3H_2S(g) \longrightarrow Cr_2S_3(s) + 3H_2O(l)$$

To produce 421 g of Cr_2S_3, (a) how many moles of Cr_2O_3 are required? (b) How many grams of Cr_2O_3 are required?

3.43 Calculate the mass of each product formed when 33.61 g of diborane (B_2H_6) reacts with excess water:

$$B_2H_6(g) + H_2O(l) \longrightarrow H_3BO_3(s) + H_2(g) \text{ [unbalanced]}$$

3.44 Calculate the mass of each product formed when 174 g of silver sulfide reacts with excess hydrochloric acid:

$$Ag_2S(s) + HCl(aq) \longrightarrow AgCl(s) + H_2S(g) \text{ [unbalanced]}$$

3.45 Elemental phosphorus occurs as tetratomic molecules, P_4. What mass of chlorine gas is needed to react completely with 355 g of phosphorus to form phosphorus pentachloride?

3.46 Elemental sulfur occurs as octatomic molecules, S_8. What mass of fluorine gas is needed to react completely with 17.8 g of sulfur to form sulfur hexafluoride?

3.47 Many metals react with oxygen gas to form the metal oxide. For example, calcium reacts as follows:

$$2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$$

You wish to calculate the mass of calcium oxide that can be prepared from 4.20 g of Ca and 2.80 g of O_2.
(a) How many moles of CaO can form from the given mass of Ca?
(b) How many moles of CaO can form from the given mass of O_2?
(c) Which is the limiting reactant?
(d) How many grams of CaO can form?

3.48 Metal hydrides react with water to form hydrogen gas and the metal hydroxide. For example,

$$SrH_2(s) + 2H_2O(l) \longrightarrow Sr(OH)_2(s) + 2H_2(g)$$

You wish to calculate the mass of hydrogen gas that can be prepared from 5.63 g of SrH_2 and 4.80 g of H_2O.
(a) How many moles of H_2 can form from the given mass of SrH_2?
(b) How many moles of H_2 can form from the given mass of H_2O?
(c) Which is the limiting reactant?
(d) How many grams of H_2 can form?

3.49 Calculate the maximum numbers of moles and grams of iodic acid (HIO_3) that can form when 685 g of iodine trichloride reacts with 117.4 g of water:

$$ICl_3 + H_2O \longrightarrow ICl + HIO_3 + HCl \text{ [unbalanced]}$$

What mass of the excess reactant remains?

3.50 Calculate the maximum numbers of moles and grams of H_2S that can form when 158 g of aluminum sulfide reacts with 131 g of water:

$$Al_2S_3 + H_2O \longrightarrow Al(OH)_3 + H_2S \text{ [unbalanced]}$$

What mass of the excess reactant remains?

3.51 When 0.100 mol of carbon is burned in a closed vessel with 8.00 g of oxygen, how many grams of carbon dioxide can form? Which reactant is in excess, and how many grams of it remain after the reaction?

3.52 A mixture of 0.0359 g of hydrogen and 0.0175 mol of oxygen in a closed container is sparked to initiate a reaction. How many grams of water can form? Which reactant is in excess, and how many grams of it remain after the reaction?

3.53 Aluminum nitrite and ammonium chloride react to form aluminum chloride, nitrogen, and water. What mass of each substance is present after 62.5 g of aluminum nitrite and 54.6 g of ammonium chloride react completely?

3.54 Calcium nitrate and ammonium fluoride react to form calcium fluoride, dinitrogen monoxide, and water vapor. What mass of each substance is present after 16.8 g of calcium nitrate and 17.50 g of ammonium fluoride react completely?

3.55 Two successive reactions, A $\longrightarrow$ B and B $\longrightarrow$ C, have yields of 82% and 65%, respectively. What is the overall percent yield for conversion of A to C?

3.56 Two successive reactions, D $\longrightarrow$ E and E $\longrightarrow$ F, have yields of 48% and 73%, respectively. What is the overall percent yield for conversion of D to F?

3.57 What is the percent yield of a reaction in which 41.5 g of tungsten(VI) oxide (WO_3) reacts with excess hydrogen gas to produce metallic tungsten and 9.50 mL of water (d = 1.00 g/mL)?

3.58 What is the percent yield of a reaction in which 200. g of phosphorus trichloride reacts with excess water to form 128 g of HCl and aqueous phosphorous acid (H_3PO_3)?

3.59 When 18.5 g of methane and 43.0 g of chlorine gas undergo a reaction that has an 80.0% yield, what mass of chloromethane (CH_3Cl) forms? Hydrogen chloride also forms.

3.60 When 56.6 g of calcium and 30.5 g of nitrogen gas undergo a reaction that has a 93.0% yield, what mass of calcium nitride forms?

3.61 Cyanogen, $(CN)_2$, has been observed in the atmosphere of Titan, Saturn's largest moon, and in the gases of interstellar nebulas. On Earth, it is used as a welding gas and a fumigant. In its reaction with fluorine gas, carbon tetrafluoride and nitrogen trifluoride gases are produced. What mass of carbon tetrafluoride forms when 80.0 g of each reactant is used?

3.62 Gaseous butane is compressed and used as a liquid fuel in disposable cigarette lighters and lightweight camping stoves. Suppose a lighter contains 6.50 mL of butane (d = 0.579 g/mL).
(a) How many grams of oxygen are needed to burn the butane completely?
(b) How many moles of CO_2 form when all the butane burns?
(c) How many total molecules of gas form when the butane burns completely?

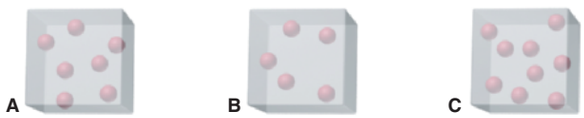
3.63 Sodium borohydride ($NaBH_4$) is used industrially in many organic syntheses. One way to prepare it is by reacting sodium hydride with gaseous diborane (B_2H_6). Assuming a 95.5% yield,

how many grams of $NaBH_4$ can be prepared by reacting 7.88 g of sodium hydride and 8.12 g of diborane?

Fundamentals of Solution Stoichiometry

(Sample Problems 3.12 to 3.16)

3.64 Box A represents a unit volume of a solution. Choose from boxes B and C the one representing the same unit volume of solution that has (a) more solute added; (b) more solvent added; (c) higher molarity; (d) lower concentration.



3.65 Calculate each of the following quantities:
(a) Grams of solute in 175.8 mL of 0.207 *M* calcium acetate
(b) Molarity of 500. mL of solution containing 21.1 g of potassium iodide
(c) Moles of solute in 145.6 L of 0.850 *M* sodium cyanide

3.66 Calculate each of the following quantities:
(a) Volume in liters of 2.26 *M* potassium hydroxide that contains 8.42 g of solute
(b) Number of Cu^{2+} ions in 52 L of 2.3 *M* copper(II) chloride
(c) Molarity of 275 mL of solution containing 135 mmol of glucose

3.67 Calculate each of the following quantities:
(a) Molarity of a solution prepared by diluting 37.00 mL of 0.250 *M* potassium chloride to 150.00 mL
(b) Molarity of a solution prepared by diluting 25.71 mL of 0.0706 *M* ammonium sulfate to 500.00 mL
(c) Molarity of sodium ion in a solution made by mixing 3.58 mL of 0.288 *M* sodium chloride with 500. mL of 6.51×10^{-3} *M* sodium sulfate (assume volumes are additive)

3.68 Calculate each of the following quantities:
(a) Volume of 2.050 *M* copper(II) nitrate that must be diluted with water to prepare 750.0 mL of a 0.8543 *M* solution
(b) Volume of 1.03 *M* calcium chloride that must be diluted with water to prepare 350. mL of a 2.66×10^{-2} *M* chloride ion solution
(c) Final volume of a 0.0700 *M* solution prepared by diluting 18.0 mL of 0.155 *M* lithium carbonate with water

3.69 A sample of concentrated nitric acid has a density of 1.41 g/mL and contains 70.0% HNO_3 by mass.
(a) What mass of HNO_3 is present per liter of solution?
(b) What is the molarity of the solution?

3.70 Concentrated sulfuric acid (18.3 *M*) has a density of 1.84 g/mL.
(a) How many moles of sulfuric acid are present per milliliter of solution?
(b) What is the mass % of H_2SO_4 in the solution?

3.71 How many milliliters of 0.383 *M* HCl are needed to react with 16.2 g of $CaCO_3$?

$$2HCl(aq) + CaCO_3(s) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

3.72 How many grams of NaH_2PO_4 are needed to react with 38.74 mL of 0.275 *M* NaOH?

$$NaH_2PO_4(s) + 2NaOH(aq) \longrightarrow Na_3PO_4(aq) + 2H_2O(l)$$

3.73 How many grams of solid barium sulfate form when 25.0 mL of 0.160 *M* barium chloride reacts with 68.0 mL of 0.055 *M* sodium sulfate? Aqueous sodium chloride is the other product.

3.74 Which reactant is in excess and by how many moles when 350.0 mL of 0.210 *M* sulfuric acid reacts with 0.500 L of 0.196 *M* sodium hydroxide to form water and aqueous sodium sulfate?

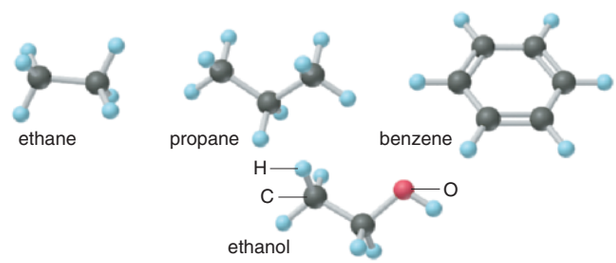
3.75 Muriatic acid, an industrial grade of concentrated HCl, is used to clean masonry and etch cement for painting. Its concentration is 11.7 *M*.
(a) Write instructions for diluting the concentrated acid to make 5.0 gallons of 3.5 *M* acid for routine use (1 gal = 4 qt; 1 qt = 0.946 L).
(b) How many milliliters of the muriatic acid solution contain 9.55 g of HCl?

Comprehensive Problems

Problems with an asterisk (*) are more challenging.

3.76 Narceine is a narcotic in opium. It crystallizes from water solution as a hydrate that contains 10.8 mass % water. If the molar mass of narceine hydrate is 499.52 g/mol, determine *x* in narceine·xH_2O.

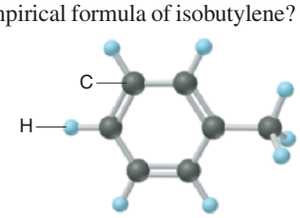
3.77 Hydrogen-containing fuels have a "fuel value" based on their mass % H. Rank the following compounds from highest mass % H to lowest: ethane, propane, benzene, ethanol, cetyl palmitate (whale oil, $C_{32}H_{64}O_2$).



3.78 Convert the following descriptions of reactions into balanced equations:
(a) In a gaseous reaction, hydrogen sulfide burns in oxygen to form sulfur dioxide and water vapor.
(b) When crystalline potassium chlorate is heated to just above its melting point, it reacts to form two different crystalline compounds, potassium chloride and potassium perchlorate.
(c) When hydrogen gas is passed over powdered iron(III) oxide, iron metal and water vapor form.
(d) The combustion of gaseous ethane in air forms carbon dioxide and water vapor.
(e) Iron(II) chloride is converted to iron(III) fluoride by treatment with chlorine trifluoride gas. Chlorine gas is also formed.

3.79 Isobutylene is a hydrocarbon used in the manufacture of synthetic rubber. When 0.847 g of isobutylene was analyzed by combustion (using an apparatus similar to that in Figure 3.5), the gain in mass of the CO_2 absorber was 2.657 g and that of the H_2O absorber was 1.089 g. What is the empirical formula of isobutylene?

3.80 One of the compounds used to increase the octane rating of gasoline is toluene *(right)*. Suppose 15.0 mL of toluene (d = 0.867 g/mL) is consumed when a sample of gasoline burns in air.



(a) How many grams of oxygen are needed for complete combustion of the toluene?

(b) How many total moles of gaseous products form?
(c) How many molecules of water vapor form?

3.81 During studies of the reaction in Sample Problem 3.10,

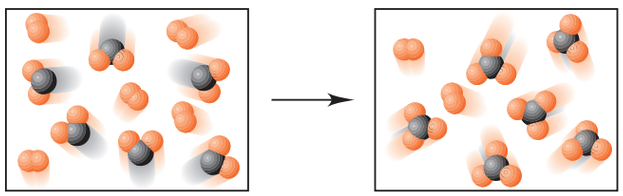
$$N_2O_4(l) + 2N_2H_4(l) \longrightarrow 3N_2(g) + 4H_2O(g)$$

a chemical engineer measured a less-than-expected yield of N_2 and discovered that the following side reaction occurs:

$$2N_2O_4(l) + N_2H_4(l) \longrightarrow 6NO(g) + 2H_2O(g)$$

In one experiment, 10.0 g of NO formed when 100.0 g of each reactant was used. What is the highest percent yield of N_2 that can be expected?

3.82 The following boxes represent a chemical reaction between AB_2 and B_2:

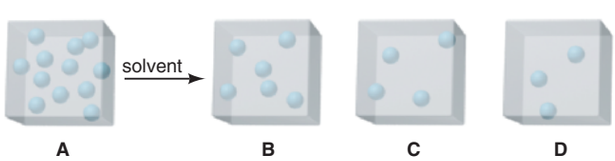


(a) Write a balanced equation for the reaction.
(b) What is the limiting reactant in this reaction?
(c) How many moles of product can be made from 3.0 mol of B_2 and 5.0 mol of AB_2?
(d) How many moles of excess reactant remain after the reaction in part (c)?

3.83 Seawater is approximately 4.0% by mass dissolved ions. About 85% of the mass of the dissolved ions is from NaCl.
(a) Calculate the mass percent of NaCl in seawater.
(b) Calculate the mass percent of Na^+ ions and of Cl^- ions in seawater.
(c) Calculate the molarity of NaCl in seawater at 15°C (d of seawater at 15°C = 1.025 g/mL).

3.84 Is each of the following statements true or false? Correct any that are false:
(a) A mole of one substance has the same number of atoms as a mole of any other substance.
(b) The theoretical yield for a reaction is based on the balanced chemical equation.
(c) A limiting-reactant problem is presented when the quantity of available material is given in moles for one of the reactants.
(d) The concentration of a solution is an intensive property, but the amount of solute in a solution is an extensive property.

3.85 Box A represents one unit volume of solution A. Which box—B, C, or D—represents one unit volume after adding enough solvent to solution A to (a) triple its volume; (b) double its volume; (c) quadruple its volume?



3.86 In each pair, choose the larger of the indicated quantities or state that the samples are equal:
(a) Entities: 0.4 mol of O_3 molecules or 0.4 mol of O atoms
(b) Grams: 0.4 mol of O_3 molecules or 0.4 mol of O atoms
(c) Moles: 4.0 g of N_2O_4 or 3.3 g of SO_2
(d) Grams: 0.6 mol of C_2H_4 or 0.6 mol of F_2
(e) Total ions: 2.3 mol of sodium chlorate or 2.2 mol of magnesium chloride
(f) Molecules: 1.0 g of H_2O or 1.0 g of H_2O_2
(g) Na^+ ions: 0.500 L of 0.500 M NaBr or 0.0146 kg of NaCl
(h) Mass: 6.02×10^{23} atoms of ^{235}U or 6.02×10^{23} atoms of ^{238}U

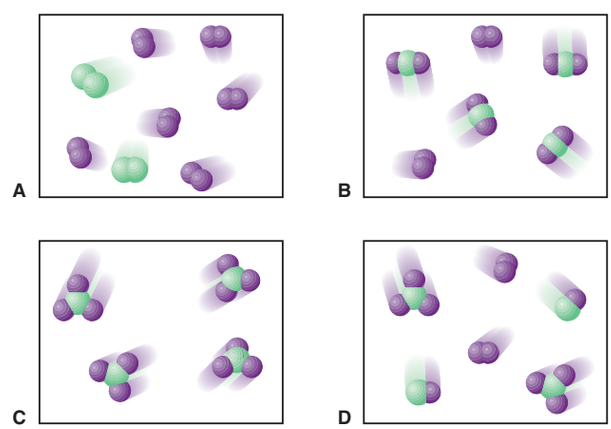
3.87 Balance the equation for the reaction between solid tetraphosphorus trisulfide and oxygen gas to form solid tetraphosphorus decaoxide and gaseous sulfur dioxide. Tabulate the equation (see Table 3.2) in terms of (a) molecules, (b) moles, and (c) grams.

3.88 Hydrogen gas has been suggested as a clean fuel because it produces only water vapor when it burns. If the reaction has a 98.8% yield, what mass of hydrogen forms 85.0 kg of water?

3.89 Assuming that the volumes are additive, what is the concentration of KBr in a solution prepared by mixing 0.200 L of 0.053 M KBr with 0.550 L of 0.078 M KBr?

3.90 Calculate each of the following quantities:
(a) Moles of compound in 0.588 g of ammonium bromide
(b) Number of potassium ions in 68.5 g of potassium nitrate
(c) Mass in grams of 5.85 mol of glycerol ($C_3H_8O_3$)
(d) Volume of 2.55 mol of chloroform ($CHCl_3$; d = 1.48 g/mL)
(e) Number of sodium ions in 2.11 mol of sodium carbonate
(f) Number of atoms in 10.0 μg of cadmium
(g) Number of atoms in 0.0015 mol of fluorine gas

3.91 Elements X (green) and Y (purple) react according to the following equation: $X_2 + 3Y_2 \longrightarrow 2XY_3$. Which molecular scene represents the product of the reaction?



3.92 Hydrocarbon mixtures are used as fuels. How many grams of $CO_2(g)$ are produced by the combustion of 200. g of a mixture that is 25.0% CH_4 and 75.0% C_3H_8 by mass?

* **3.93** Nitrogen (N), phosphorus (P), and potassium (K) are the main nutrients in plant fertilizers. According to an industry convention, the numbers on the label refer to the mass percents of N, P_2O_5, and K_2O, in that order. Calculate the N:P:K ratio of a 30:10:10 fertilizer in terms of moles of each element, and express it as x:y:1.0.

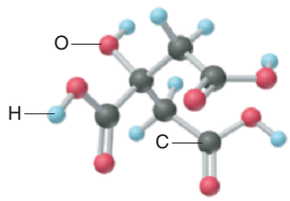
3.94 A 0.652-g sample of a pure strontium halide reacts with excess sulfuric acid, and the solid strontium sulfate formed is separated, dried, and found to weigh 0.755 g. What is the formula of the original halide?

* **3.95** When carbon-containing compounds are burned in a limited amount of air, some $CO(g)$ as well as $CO_2(g)$ is produced. A

gaseous product mixture is 35.0 mass % CO and 65.0 mass % CO_2. What is the mass % C in the mixture?

3.96 Ferrocene, first synthesized in 1951, was the first organic iron compound with Fe—C bonds. An understanding of the structure of ferrocene gave rise to new ideas about chemical bonding and led to the preparation of many useful compounds. In the combustion analysis of ferrocene, which contains only Fe, C, and H, a 0.9437-g sample produced 2.233 g of CO_2 and 0.457 g of H_2O. What is the empirical formula of ferrocene?

*** 3.97** Citric acid *(right)* is concentrated in citrus fruits and plays a central metabolic role in nearly every animal and plant cell.



(a) What are the molar mass and formula of citric acid?
(b) How many moles of citric acid are in 1.50 qt of lemon juice (d = 1.09 g/mL) that is 6.82% citric acid by mass?

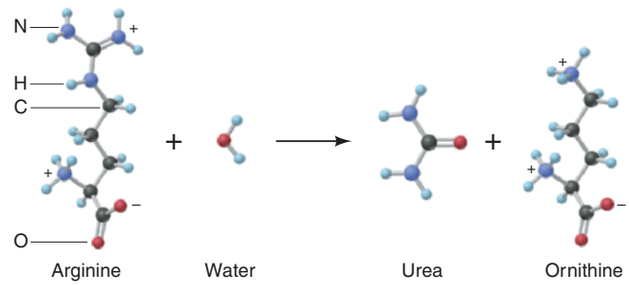
*** 3.98** Fluorine is so reactive that it forms compounds with materials inert to other treatments.
(a) When 0.327 g of platinum is heated in fluorine, 0.519 g of a dark red, volatile solid forms. What is its empirical formula?
(b) When 0.265 g of this red solid reacts with excess xenon gas, 0.378 g of an orange-yellow solid forms. What is the empirical formula of this compound, the first noble gas compound formed?
(c) Fluorides of xenon can be formed by direct reaction of the elements at high pressure and temperature. Depending on conditions, the product mixture may include the difluoride, the tetrafluoride, and the hexafluoride. Under conditions that produce only the tetra- and hexafluorides, 1.85×10^{-4} mol of xenon reacted with 5.00×10^{-4} mol of fluorine, and 9.00×10^{-6} mol of xenon was found in excess. What are the mass percents of each xenon fluoride in the product mixture?

3.99 Hemoglobin is 6.0% heme ($C_{34}H_{32}FeN_4O_4$) by mass. To remove the heme, hemoglobin is treated with acetic acid and NaCl to form hemin ($C_{34}H_{32}N_4O_4FeCl$). At a crime scene, a blood sample contains 0.45 g of hemoglobin.
(a) How many grams of heme are in the sample?
(b) How many moles of heme?
(c) How many grams of Fe?
(d) How many grams of hemin could be formed for a forensic chemist to measure?

*** 3.100** Manganese is a key component of extremely hard steel. The element occurs naturally in many oxides. A 542.3-g sample of a manganese oxide has an Mn:O ratio of 1.00:1.42 and consists of braunite (Mn_2O_3) and manganosite (MnO).
(a) What masses of braunite and manganosite are in the ore?
(b) What is the ratio Mn^{3+}:Mn^{2+} in the ore?

3.101 Sulfur dioxide is a major industrial gas used primarily for the production of sulfuric acid, but also as a bleach and food preservative. One way to produce it is by roasting iron pyrite (iron disulfide, FeS_2) in oxygen, which yields the gas and solid iron(III) oxide. What mass of each of the other three substances are involved in producing 1.00 kg of sulfur dioxide?

3.102 The human body excretes nitrogen in the form of urea, NH_2CONH_2. The key biochemical step in urea formation is the reaction of water with arginine to produce urea and ornithine:



(a) What is the mass percent of nitrogen in urea, arginine, and ornithine? (b) How many grams of nitrogen can be excreted as urea when 143.2 g of ornithine is produced?

3.103 Aspirin (acetylsalicylic acid, $C_9H_8O_4$) can be made by reacting salicylic acid ($C_7H_6O_3$) with acetic anhydride [$(CH_3CO)_2O$]:

$$C_7H_6O_3(s) + (CH_3CO)_2O(l) \longrightarrow C_9H_8O_4(s) + CH_3COOH(l)$$

In one reaction, 3.027 g of salicylic acid and 6.00 mL of acetic anhydride react to form 3.261 g of aspirin.
(a) Which is the limiting reactant (d of acetic anhydride = 1.080 g/mL)?
(b) What is the percent yield of this reaction?

*** 3.104** When powdered zinc is heated with sulfur, a violent reaction occurs, and zinc sulfide forms:

$$Zn(s) + S_8(s) \longrightarrow ZnS(s) \quad \text{[unbalanced]}$$

Some of the reactants also combine with oxygen in air to form zinc oxide and sulfur dioxide. When 85.2 g of Zn reacts with 52.4 g of S_8, 105.4 g of ZnS forms. What is the percent yield of ZnS? (b) If all the remaining reactants combine with oxygen, how many grams of each of the two oxides form?

*** 3.105** High-temperature superconducting oxides hold great promise in the utility, transportation, and computer industries. (a) One superconductor is $La_{2-x}Sr_xCuO_4$. Calculate the molar mass of this oxide when $x = 0$, $x = 1$, and $x = 0.163$ (the last characterizes the compound with optimum superconducting properties).
(b) Another common superconducting oxide is made by heating a mixture of barium carbonate, copper(II) oxide, and yttrium(III) oxide, followed by further heating in O_2:

$$4BaCO_3(s) + 6CuO(s) + Y_2O_3(s) \longrightarrow 2YBa_2Cu_3O_{6.5}(s) + 4CO_2(g)$$

$$2YBa_2Cu_3O_{6.5}(s) + \tfrac{1}{2}O_2(g) \longrightarrow 2YBa_2Cu_3O_7(s)$$

When equal masses of the three reactants are heated, which reactant is limiting?
(c) After the product in part (b) is removed, what is the mass percent of each reactant in the solid mixture remaining?

*** 3.106** The zirconium oxalate $K_2Zr(C_2O_4)_3(H_2C_2O_4)\cdot H_2O$ was synthesized by mixing 1.60 g of $ZrOCl_2\cdot 8H_2O$ with 5.20 g of $H_2C_2O_4\cdot 2H_2O$ and an excess of aqueous KOH. After 2 months, 1.20 g of crystalline product was obtained, as well as aqueous KCl and water. Calculate the percent yield.