3 Stoichiometry of Formulas and Equations

Concepts and Skills to Review Before You Study This Chapter

- atomic mass (Section 2.5)
- names and formulas of compounds (Section 2.8)
- molecular (or formula) mass (Section 2.8)
- molecular and structural formulas, and ball-and-stick and space-filling models (Section 2.8)

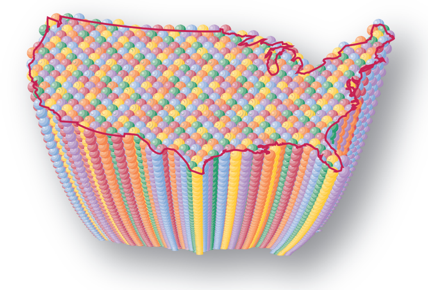
Chemistry is, above all, a practical science. Imagine that you're a biochemist who has extracted a substance with medicinal activity from a tropical plant: what is its formula, and what quantity of metabolic products will establish a safe dosage level? Thousands of biologically active compounds have been found in plants and many are used in modern medicines. For example, in 1963, the U.S. Food and Drug Administration (FDA) approved the drug *vincristine,* which is derived from the Madagascar periwinkle *(photo),* for treatment of lymphomas and childhood cancers. Or, suppose you're a chemical engineer studying rocket-fuel thrust: what amount of propulsive gases will a fuel produce? Perhaps you're on a team of environmental chemists examining coal samples: what quantity of air pollutants will a sample produce when burned? Or, maybe you're a polymer chemist preparing a plastic with unusual properties: how much of this new material will the polymerization reaction yield? You can answer countless questions like these 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 formulas and reactions.

IN THIS CHAPTER . . . *We relate the mass of a substance to the number of chemical entities comprising it (atoms, ions, molecules, or formula units) and apply this relationship to formulas and equations.*

- We discuss the *mole,* the chemist's unit for amount of a substance, and use it to convert between mass and number of entities.
- We also use it to derive a chemical formula from the results of mass analysis.
- We see whether two key types of formulas relate to molecular structures.
- We learn how to write chemical equations and how to balance them in terms of the amounts of substances reacting and produced.
- We calculate the amounts of reactants and products in a reaction and see why one of the reactants limits the amount of product that can form and, thus, the reaction yield.

3.1 THE MOLE

In daily life, we often measure things by weighing or by counting: we weigh coffee beans or rice, but we count eggs or pencils. And we use mass units (a kilogram of coffee beans) or counting units (a dozen pencils) to express the amount. Similarly, daily life in the laboratory involves measuring substances. We want to know the numbers of chemical entities—atoms, ions, molecules, or formula units—that react with each other, but how can we possibly count or weigh such minute objects? As you'll see, chemists have devised a unit, called the *mole,* to *count chemical entities by weighing a very large number of them.*



Imagine a Mole of . . .

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 continental United States 70 miles deep. However, atoms and molecules are not ordinary objects: you can swallow a mole of water molecules (about 18 mL) in one gulp!

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 the number of atoms in 12 g of carbon-12.* This number, called **Avogadro's number** (in honor of the 19th-century Italian physicist Amedeo Avogadro), is enormous:

$$\text{One mole (1 mol) contains } 6.022\times10^{23} \text{ entities (to four significant figures)} \quad \textbf{(3.1)}$$

A counting unit, like *dozen,* tells you the number of objects but not their mass; a mass unit, like *kilogram,* tells you the mass of the objects but not their number. The mole tells you both—the *number* of objects in a given *mass* of substance:

1 mol of carbon-12 contains 6.022×10^{23} carbon-12 atoms *and* has a mass of 12 g

What does it mean that the mole unit allows you to count entities by weighing the sample? Suppose you have a sample of carbon-12 and want to know the number of atoms present. You find that the sample weighs 6 g, so it is 0.5 mol of carbon-12 and contains $0.5(6.022\times10^{23})$ or 3.011×10^{23} atoms:

6 g of carbon-12 is 0.5 mol of carbon-12 and contains 3.011×10^{23} atoms

Knowing the amount (in moles), the mass (in grams), and the number of entities becomes very important when we mix different substances to run a reaction. The central relationship between masses on the atomic scale and on the macroscopic scale is the same for elements and compounds:

- *Elements.* The mass in *atomic mass units (amu)* of one atom of an element is the *same numerically* as the mass in *grams (g)* of 1 mole of atoms of the element. Recall from Chapter 2 that each atom of an element is considered to have the *atomic mass* given in the periodic table *(see margin)*. Thus,

16
S
32.06

1 atom of S	has a mass of	32.06 amu	and	1 mol (6.022×10^{23} atoms) of S	has a mass of	32.06 g
1 atom of Fe	has a mass of	55.85 amu	and	1 mol (6.022×10^{23} atoms) of Fe	has a mass of	55.85 g

Note, also, that since atomic masses are relative, 1 Fe atom weighs 55.85/32.06 as much as 1 S atom, and 1 mol of Fe weighs 55.85/32.06 as much as 1 mol of S.

- *Compounds.* The mass in *atomic mass units (amu)* of one molecule (or formula unit) of a compound is the *same numerically* as the mass in *grams (g)* of 1 mole of the compound. Thus, for example,

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

Here, too, because masses are relative, 1 H_2O molecule weighs 18.02/58.44 as much as 1 NaCl formula unit, and 1 mol of H_2O weighs 18.02/58.44 as much as 1 mol of NaCl.

The two key points to remember about the importance of the mole unit are

- The *mole* lets us relate the *number* of entities to the *mass* of a sample of those entities.
- The mole maintains the *same numerical relationship* between mass on the atomic scale (atomic mass units, amu) and mass on the macroscopic scale (grams, g).

In everyday terms, a grocer *does not* know that there are 1 dozen eggs from their weight or that there is 1 kilogram of coffee beans from their count, because eggs and coffee beans do not have fixed masses. But, by weighing out 63.55 g (1 mol) of copper, a chemist *does* know that there are 6.022×10^{23} copper atoms, because all copper atoms have an atomic mass of 63.55 amu. Figure 3.1 shows 1 mole of some familiar elements and compounds.

Figure 3.1 One mole (6.022×10^{23} entities) of some familiar substances. From left to right: 1 mol of copper (63.55 g), of liquid H_2O (18.02 g), of sodium chloride (table salt, 58.44 g), of sucrose (table sugar, 342.3 g), and of aluminum (26.98 g).

Determining Molar Mass

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

1. *Elements*. To find the molar mass, look up the atomic mass and note whether the element is monatomic or molecular.

- *Monatomic elements.* The molar mass is the periodic-table value in grams per mole.* For example, the molar mass of neon is 20.18 g/mol, and the molar mass of gold is 197.0 g/mol.
- *Molecular elements.* You must know the formula to determine the molar mass (see Figure 2.15). For example, in air, oxygen exists most commonly as diatomic molecules, so the molar mass of O_2 is twice that of O:

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

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

$$\mathcal{M} \text{ of } S_8 = 8 \times \mathcal{M} \text{ of S} = 8 \times 32.06 \text{ g/mol} = 256.5 \text{ g/mol}$$

2. *Compounds. The molar mass is the sum of the molar masses of the atoms in the formula.* Thus, from the formula of sulfur dioxide, SO_2, we know that 1 mol of SO_2 molecules contains 1 mol of S atoms and 2 mol of O atoms:

$$\mathcal{M} \text{ of } SO_2 = \mathcal{M} \text{ of S} + (2 \times \mathcal{M} \text{ of O}) = 32.06 \text{ g/mol} + (2 \times 16.00 \text{ g/mol}) = 64.06 \text{ g/mol}$$

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

$$\mathcal{M} \text{ of } K_2S = (2 \times \mathcal{M} \text{ of K}) + \mathcal{M} \text{ of S} = (2 \times 39.10 \text{ g/mol}) + 32.06 \text{ g/mol} = 110.26 \text{ g/mol}$$

Thus, *subscripts in a formula refer to individual atoms (or ions) as well as to moles of atoms (or ions).* Table 3.1 summarizes these ideas for glucose, $C_6H_{12}O_6$ *(see margin),* the essential sugar in energy metabolism.

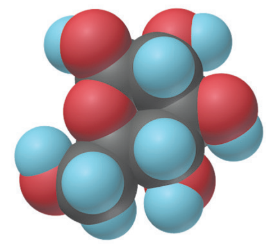

Glucose

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

Converting Between Amount, Mass, and Number of Chemical Entities

One of the most common skills in the lab—and on exams—is converting between amount (mol), mass (g), and number of entities of a substance.

1. *Converting between amount and mass.* If you know the amount of a substance, you can find its mass, and vice versa. The molar mass ($\mathcal{M}$), which expresses the equivalence between 1 mole of a substance and its mass in grams, is the conversion factor between amount and mass:

$$\frac{\text{no. of grams}}{1 \text{ mol}} \quad \text{or} \quad \frac{1 \text{ mol}}{\text{no. of grams}}$$

- *From amount (mol) to mass (g),* multiply by the molar mass to cancel the mole unit:

$$\text{Mass (g)} = \text{amount } (\cancel{\text{mol}}) \times \frac{\text{no. of grams}}{1 \cancel{\text{mol}}} \quad \textbf{(3.2)}$$

*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 (with different units) for the atomic mass and for the molar mass.

- *From mass (g) to amount (mol),* divide by the molar mass (multiply by $1/\mathcal{M}$) to cancel the mass unit:

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

2. *Converting between amount and number.* Similarly, if you know the amount (mol), you can find the number of entities, and vice versa. Avogadro's number, which expresses the equivalence between 1 mole of a substance and the number of entities it contains, is the conversion factor between amount and number of entities:

$$\frac{6.022\times10^{23}\text{ entities}}{1\text{ mol}} \quad \text{or} \quad \frac{1\text{ mol}}{6.022\times10^{23}\text{ entities}}$$

- *From amount (mol) to number of entities,* multiply by Avogadro's number to cancel the mole unit:

$$\text{No. of entities} = \text{amount }(\cancel{\text{mol}}) \times \frac{6.022\times10^{23}\text{ entities}}{1\ \cancel{\text{mol}}} \tag{3.4}$$

- *From number of entities to amount (mol),* divide by Avogadro's number to cancel the number of entities:

$$\text{Amount (mol)} = \cancel{\text{no. of entities}} \times \frac{1\text{ mol}}{6.022\times10^{23}\ \cancel{\text{entities}}} \tag{3.5}$$

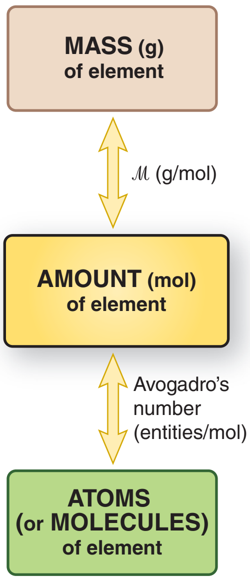


Figure 3.2 Mass-mole-number relationships for elements.

Amount-Mass-Number Conversions Involving Elements We begin with amount-mass-number relationships of elements. As Figure 3.2 shows, *convert mass or number of entities (atoms or molecules) to amount (mol) first.* For molecular elements, Avogadro's number gives *molecules* per mole.

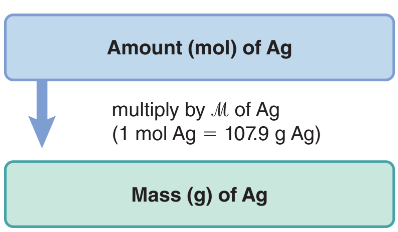
Let's work through a series of sample problems that show these conversions for both elements and compounds.

SAMPLE PROBLEM 3.1 Converting Between Mass and Amount of an Element

Problem 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?

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

Road Map



Solution Converting from amount (mol) of Ag to mass (g):

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

Check We rounded the mass to three significant figures because the amount (in mol) 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.

FOLLOW-UP PROBLEMS

*Brief Solutions for **all** Follow-up Problems appear at the end of the chapter.*

3.1A Graphite is the crystalline form of carbon used in "lead" pencils. How many moles of carbon are in 315 mg of graphite? Include a road map that shows how you planned the solution.

3.1B A soda can contains about 14 g of aluminum (Al), the most abundant element in Earth's crust. How many soda cans can be made from 52 mol of Al? Include a road map that shows how you planned the solution.

SOME SIMILAR PROBLEMS 3.12(a) and 3.13(a)

SAMPLE PROBLEM 3.2 Converting Between Number of Entities and Amount of an Element

Problem Gallium (Ga) is a key element in solar panels, calculators, and other light-sensitive electronic devices. How many Ga atoms are in 2.85×10^{-3} mol of gallium?

Plan We know the amount of gallium (2.85×10^{-3} mol) and need the number of Ga atoms. We multiply amount (mol) by Avogadro's number to find number of atoms (see the road map).

Road Map

Amount (mol) of Ga

↓ multiply by Avogadro's number (1 mol Ga = 6.022×10^{23} Ga atoms)

Number of Ga atoms

Solution Converting from amount (mol) of Ga to number of atoms:

$$\text{No. of Ga atoms} = 2.85\times10^{-3}\ \cancel{\text{mol Ga}} \times \frac{6.022\times10^{23}\ \text{Ga atoms}}{1\ \cancel{\text{mol Ga}}}$$

$$= \boxed{1.72\times10^{21}\ \text{Ga atoms}}$$

Check The number of atoms has three significant figures because the number of moles does. When we round amount (mol) of Ga and Avogadro's number, we have $\sim(3\times10^{-3}\ \text{mol})(6\times10^{23}\ \text{atoms/mol}) = 18\times10^{20}$, or 1.8×10^{21} atoms, so our answer seems correct.

FOLLOW-UP PROBLEMS

3.2A At rest, a person inhales 9.72×10^{21} nitrogen molecules in an average breath of air. How many moles of nitrogen atoms are inhaled? (*Hint:* In air, nitrogen occurs as a diatomic molecule.) Include a road map that shows how you planned the solution.

3.2B A tank contains 325 mol of compressed helium (He) gas. How many He atoms are in the tank? Include a road map that shows how you planned the solution.

SOME SIMILAR PROBLEMS 3.12(b) and 3.13(b)

For the next sample problem, note that mass and number of entities relate directly to amount (mol), but *not* to each other. Therefore, *to convert between mass and number, first convert to amount.*

SAMPLE PROBLEM 3.3 Converting Between Number of Entities and Mass of an Element

Problem Iron (Fe) is the main component of steel and, thus, the most important metal in industrial society; it is also essential in the body. How many Fe atoms are in 95.8 g of Fe?

Plan We know the mass of Fe (95.8 g) and need the number of Fe atoms. We cannot convert directly from mass to number, so we first convert to amount (mol) by dividing mass of Fe by its molar mass. Then, we multiply amount (mol) by Avogadro's number to find number of atoms (see the road map).

Road Map

Mass (g) of Fe

↓ divide by $\mathcal{M}$ of Fe (55.85 g Fe = 1 mol Fe)

Amount (mol) of Fe

↓ multiply by Avogadro's number (1 mol Fe = 6.022×10^{23} Fe atoms)

Number of Fe atoms

Solution Converting from mass (g) of Fe to amount (mol):

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

Converting from amount (mol) of Fe to number of Fe atoms:

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

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

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

Check Rounding the mass and the molar mass of Fe, we have $\sim100\ \text{g}/(\sim60\ \text{g/mol}) = 1.7$ mol. Therefore, the number of atoms should be a bit less than twice Avogadro's number: $<2(6\times10^{23}) = <1.2\times10^{24}$, so the answer seems correct.

FOLLOW-UP PROBLEMS

3.3A 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? Include a road map that shows how you planned the solution.

3.3B Pennies minted after 1982 are made of zinc plated with a thin coating of copper (Cu); the copper layer on each penny has a mass of 0.0625 g. How many Cu atoms are in a penny? Include a road map that shows how you planned the solution.

SOME SIMILAR PROBLEMS 3.12(c) and 3.13(c)

Amount-Mass-Number Conversions Involving Compounds Only one new step is needed to solve amount-mass-number problems involving compounds: we need the chemical formula to find the molar mass and the amount of each element in the compound. The relationships are shown in Figure 3.3, and Sample Problems 3.4 and 3.5 apply them to compounds with simple and more complicated formulas, respectively.

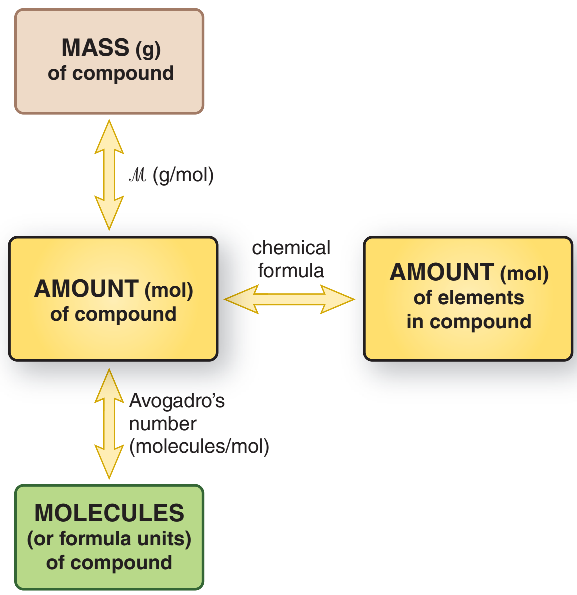


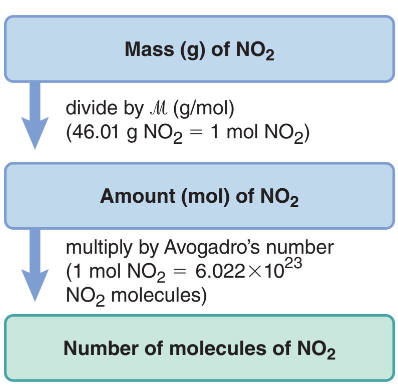
Figure 3.3 Amount-mass-number relationships for compounds. Use the chemical formula to find the amount (mol) of each element in a compound.

SAMPLE PROBLEM 3.4 Converting Between Number of Entities and Mass of a Compound I

Problem Nitrogen dioxide is a component of urban smog that forms from gases in car exhaust. How many molecules are in 8.92 g of nitrogen dioxide?

Plan We know the mass of compound (8.92 g) and need to find the number of molecules. As you just saw in Sample Problem 3.3, to convert mass to number of entities, we have to find the amount (mol). To do so, we divide the mass by the molar mass ($\mathcal{M}$), which we calculate from the molecular formula (see Sample Problem 2.16). Once we have the amount (mol), we multiply by Avogadro's number to find the number of molecules (see the road map).

Road Map



Solution The formula is NO_2. Calculating the molar mass:

$$\begin{aligned}\mathcal{M} &= (1 \times \mathcal{M} \text{ of N}) + (2 \times \mathcal{M} \text{ of O})\\ &= 14.01 \text{ g/mol} + (2 \times 16.00 \text{ g/mol})\\ &= 46.01 \text{ g/mol}\end{aligned}$$

Converting from mass (g) of NO_2 to amount (mol):

$$\begin{aligned}\text{Amount (mol) of } NO_2 &= 8.92 \cancel{\text{ g } NO_2} \times \frac{1 \text{ mol } NO_2}{46.01 \cancel{\text{ g } NO_2}}\\ &= 0.194 \text{ mol } NO_2\end{aligned}$$

Converting from amount (mol) of NO_2 to number of molecules:

$$\text{No. of molecules} = 0.194\ \cancel{\text{mol } NO_2} \times \frac{6.022\times10^{23}\ NO_2 \text{ molecules}}{1\ \cancel{\text{mol } NO_2}}$$

$$= \boxed{1.17\times10^{23}\ NO_2 \text{ molecules}}$$

Check Rounding, we get ($\sim$0.2 mol)(6×10^{23}) = 1.2×10^{23}, so the answer seems correct.

FOLLOW-UP PROBLEMS

3.4A Fluoride ion is added to drinking water to prevent tooth decay. What is the mass (g) of sodium fluoride in a liter of water that contains 1.19×10^{19} formula units of the compound? Include a road map that shows how you planned the solution.

3.4B Calcium chloride is applied to highways in winter to melt accumulated ice. A snow-plow truck applies 400 lb of $CaCl_2$ per mile of highway. How many formula units of the compound are applied per mile? Include a road map that shows how you planned the solution.

SOME SIMILAR PROBLEMS 3.14–3.19

SAMPLE PROBLEM 3.5 Converting Between Number of Entities and Mass of a Compound II

Problem Ammonium carbonate is a white solid that decomposes with warming. It has many uses, for example, as a component in baking powder, fire extinguishers, and smelling salts.

(a) How many formula units are in 41.6 g of ammonium carbonate?

(b) How many O atoms are in this sample?

Plan **(a)** We know the mass of compound (41.6 g) and need to find the number of formula units. As in Sample Problem 3.4, we find the amount (mol) and then multiply by Avogadro's number to find the number of formula units. (A road map for this step would be the same as the one in Sample Problem 3.4.) **(b)** To find the number of O atoms, we multiply the number of formula units by the number of O atoms in one formula unit (see the road map).

Road Map

Number of formula units of $(NH_4)_2CO_3$

multiply by number of O atoms in one formula unit [1 formula unit of $(NH_4)_2CO_3$ = 3 O atoms]

Number of O atoms

Solution **(a)** The formula is $(NH_4)_2CO_3$ (see Table 2.5). Calculating the molar mass:

$$\begin{aligned}\mathcal{M} &= (2 \times \mathcal{M} \text{ of N}) + (8 \times \mathcal{M} \text{ of H}) + (1 \times \mathcal{M} \text{ of C}) + (3 \times \mathcal{M} \text{ of O})\\ &= (2 \times 14.01 \text{ g/mol N}) + (8 \times 1.008 \text{ g/mol H}) + 12.01 \text{ g/mol C}\\ &\quad + (3 \times 16.00 \text{ g/mol O})\\ &= 96.09 \text{ g/mol } (NH_4)_2CO_3\end{aligned}$$

Converting from mass (g) to amount (mol):

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

$$= 0.433 \text{ mol } (NH_4)_2CO_3$$

Converting from amount (mol) to formula units:

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

(b) Finding the number of O atoms:

$$\text{No. of O atoms} = 2.61\times10^{23}\ \cancel{\text{formula units } (NH_4)_2CO_3} \times \frac{3 \text{ O atoms}}{1\ \cancel{\text{formula unit } (NH_4)_2CO_3}}$$

$$= \boxed{7.83\times10^{23} \text{ O atoms}}$$

Check In (a), the units are correct. Since the mass is less than half the molar mass ($\sim 42/96 < 0.5$), 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 PROBLEMS

3.5A Tetraphosphorus decoxide reacts with water to form phosphoric acid, a major industrial acid. In the laboratory, the oxide is a drying agent.
(a) What is the mass (g) of 4.65×10^{22} molecules of tetraphosphorus decoxide?
(b) How many P atoms are present in this sample?

3.5B Calcium phosphate is added to some foods, such as yogurt, to boost the calcium content and is also used as an anticaking agent.
(a) How many formula units are in 75.5 g of calcium phosphate?
(b) How many phosphate ions are present in this sample?

SOME SIMILAR PROBLEMS 3.14–3.19

The Importance of Mass Percent

For many purposes, it is important to know how much of an element is present in a given amount of compound. A biochemist may want the ionic composition of a mineral nutrient; an atmospheric chemist may be studying the carbon content of a fuel; a materials scientist may want the metalloid composition of a semiconductor. In this section, we find the composition of a compound in terms of mass percent and use it to find the mass of each element in the compound.

Determining Mass Percent from a Chemical Formula Each element contributes a fraction of a compound's mass, and that fraction multiplied by 100 gives the element's mass percent. Finding the mass percent is similar on the molecular and molar scales:

- *For a molecule (or formula unit) of compound,* use the molecular (or formula) mass and chemical formula to find the mass percent of any element X in the compound:

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

- *For a mole of compound,* use the molar mass and formula to find the mass percent of each element on a mole basis:

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

As always, the individual mass percents add up to 100% (within rounding). In Sample Problem 3.6, we determine the mass percent of each element in a compound.

Road Map

Amount (mol) of element X in 1 mol of ammonium nitrate

↓ multiply by $\mathcal{M}$ (g/mol) of X

Mass (g) of X in 1 mol of ammonium nitrate

↓ divide by mass (g) of 1 mol of compound

Mass fraction of X in ammonium nitrate

↓ multiply by 100

Mass % of X in ammonium nitrate

SAMPLE PROBLEM 3.6 **Calculating the Mass Percent of Each Element in a Compound from the Formula**

Problem The effectiveness of fertilizers depends on their nitrogen content. Ammonium nitrate is a common fertilizer. What is the mass percent of each element in ammonium nitrate?

Plan We know the relative amounts (mol) of the elements from the formula, and we have to find the mass % of each element. We multiply the amount of each element by its molar mass to find its mass. Dividing each element's mass by 1 mol of ammonium nitrate gives the mass fraction of that element, and multiplying the mass fraction by 100 gives the mass %. The calculation steps for any element (X) are shown in the road map.

Solution The formula is NH_4NO_3 (see Table 2.5). In 1 mol of NH_4NO_3, there are 2 mol of N, 4 mol of H, and 3 mol of O.
Converting amount (mol) of N to mass (g): We have 2 mol of N in 1 mol of NH_4NO_3, so

$$\text{Mass (g) of N} = 2\ \cancel{\text{mol N}} \times \frac{14.01\ \text{g N}}{1\ \cancel{\text{mol N}}} = 28.02\ \text{g N}$$

Calculating the mass of 1 mol of NH_4NO_3:

$$\begin{aligned}\mathcal{M} &= (2 \times \mathcal{M} \text{ of N}) + (4 \times \mathcal{M} \text{ of H}) + (3 \times \mathcal{M} \text{ of O})\\ &= (2 \times 14.01 \text{ g/mol N}) + (4 \times 1.008 \text{ g/mol H}) + (3 \times 16.00 \text{ g/mol O})\\ &= 80.05 \text{ g/mol } NH_4NO_3\end{aligned}$$

Finding the mass fraction of N in NH_4NO_3:

$$\text{Mass fraction of N} = \frac{\text{total mass of N}}{\text{mass of 1 mol } NH_4NO_3} = \frac{28.02 \text{ g N}}{80.05 \text{ g } NH_4NO_3} = 0.3500$$

Changing to mass %:

$$\begin{aligned}\text{Mass \% of N} &= \text{mass fraction of N} \times 100 = 0.3500 \times 100\\ &= \boxed{35.00 \text{ mass \% N}}\end{aligned}$$

Combining the steps for each of the other elements in NH_4NO_3:

$$\begin{aligned}\text{Mass \% of H} &= \frac{\text{mol H} \times \mathcal{M} \text{ of H}}{\text{mass of 1 mol } NH_4NO_3} \times 100 = \frac{4 \cancel{\text{ mol H}} \times \frac{1.008 \text{ g H}}{1 \cancel{\text{ mol H}}}}{80.05 \text{ g } NH_4NO_3} \times 100\\ &= \boxed{5.037 \text{ mass \% H}}\end{aligned}$$

$$\begin{aligned}\text{Mass \% of O} &= \frac{\text{mol O} \times \mathcal{M} \text{ of O}}{\text{mass of 1 mol } NH_4NO_3} \times 100 = \frac{3 \cancel{\text{ mol O}} \times \frac{16.00 \text{ g O}}{1 \cancel{\text{ mol O}}}}{80.05 \text{ g } NH_4NO_3} \times 100\\ &= \boxed{59.96 \text{ mass \% O}}\end{aligned}$$

Check The answers make sense. The mass % of O is greater than that of N because there are more moles of O in the compound and the molar mass of O is greater. The mass % of H is small because its molar mass is small. The sum of the mass percents is 100.00%.

Comment 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 PROBLEMS

3.6A In mammals, lactose (milk sugar) is metabolized to glucose ($C_6H_{12}O_6$), the key nutrient for generating chemical potential energy. Calculate the mass percent of C in glucose.

3.6B For many years, compounds known as *chlorofluorocarbons* were used as refrigerants, until it was discovered that the chlorine atoms in these compounds destroy ozone molecules in the atmosphere. The compound CCl_3F is a chlorofluorocarbon with a high chlorine content. Calculate the mass percent of Cl in CCl_3F.

SOME SIMILAR PROBLEMS 3.20–3.23

Determining the Mass of an Element from Its Mass Fraction Sample Problem 3.6 shows that *an element always constitutes the same fraction of the mass of a given compound* (see Equation 3.6). We can use that fraction to find the mass of element in any mass of a compound:

$$\text{Mass of element} = \text{mass of compound} \times \frac{\text{mass of element in 1 mol of compound}}{\text{mass of 1 mol of compound}} \quad \textbf{(3.7)}$$

For example, to find the mass of oxygen in 15.5 g of nitrogen dioxide, we have

$$\begin{aligned}\text{Mass (g) of O} &= 15.5 \text{ g } NO_2 \times \frac{2 \text{ mol} \times \mathcal{M} \text{ of O (g/mol)}}{\text{mass (g) of 1 mol } NO_2}\\ &= 15.5 \cancel{\text{ g } NO_2} \times \frac{32.00 \text{ g O}}{46.01 \cancel{\text{ g } NO_2}} = 10.8 \text{ g O}\end{aligned}$$

SAMPLE PROBLEM 3.7 Calculating the Mass of an Element in a Compound

Problem Use the information in Sample Problem 3.6 to determine the mass (g) of nitrogen in 650. g of ammonium nitrate.

Plan To find the mass of N in the sample of ammonium nitrate, we multiply the mass of the sample by the mass of 2 mol of N divided by the mass of 1 mol of ammonium nitrate.

Solution Finding the mass of N in a given mass of ammonium nitrate:

$$\text{Mass (g) of N} = \text{mass (g) of } NH_4NO_3 \times \frac{2 \text{ mol N} \times \mathcal{M} \text{ of N (g/mol)}}{\text{mass (g) of 1 mol } NH_4NO_3}$$

$$= 650.\ \cancel{\text{g } NH_4NO_3} \times \frac{28.02 \text{ g N}}{80.05\ \cancel{\text{g } NH_4NO_3}} = \boxed{228 \text{ g N}}$$

Check Rounding shows that the answer is "in the right ballpark": N accounts for about one-third of the mass of NH_4NO_3 and $\frac{1}{3}$ of 700 g is 233 g.

FOLLOW-UP PROBLEMS

3.7A Use the information in Follow-up Problem 3.6A to find the mass (g) of C in 16.55 g of glucose.

3.7B Use the information in Follow-up Problem 3.6B to find the mass (g) of Cl in 112 g of CCl_3F.

SOME SIMILAR PROBLEMS 3.27 and 3.28

Summary of Section 3.1

- A mole of substance is the amount that contains Avogadro's number (6.022×10^{23}) of chemical entities (atoms, ions, molecules, or formula units).
- The mass (in grams) of a mole of a given entity (atom, ion, molecule, or formula unit) 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 a given amount of the compound.

3.2 DETERMINING THE FORMULA OF AN UNKNOWN COMPOUND

In Sample Problems 3.6 and 3.7, we used a compound's formula to find the mass percent (or mass fraction) of each element in it *and* the mass of each element in any size sample of it. In this section, we do the reverse: we use the masses of elements in a compound to find the formula. Then, we look briefly at the relationship between molecular formula and molecular structure.

Let's compare three common types of formula, using hydrogen peroxide as an example:

- The **empirical formula** is derived from mass analysis. It shows the *lowest* whole number of moles, and thus the *relative* number of atoms, of each element in the compound. For example, in hydrogen peroxide, there is 1 part by mass of hydrogen for every 16 parts by mass of oxygen. Because the atomic mass of hydrogen is 1.008 amu and that of oxygen is 16.00 amu, there is one H atom for every O atom. Thus, the empirical formula is HO.

Recall from Section 2.8, p. 75, that

- The **molecular formula** shows the *actual* number of atoms of each element in a molecule: the molecular formula of hydrogen peroxide is H_2O_2, twice the empirical formula.
- The **structural formula** shows the relative *placement and connections of atoms* in the molecule: the structural formula is H—O—O—H.

Let's focus here on how to determine empirical and molecular formulas.

Empirical Formulas

A chemist studying an unknown compound goes through a three-step process to find the empirical formula:

1. Determine the mass (g) of each component element.
2. Convert each mass (g) to amount (mol), and write a preliminary formula.
3. Convert the amounts (mol) mathematically to whole-number (integer) subscripts. To accomplish this math conversion,
   - Divide each subscript by the smallest subscript, and
   - If necessary, multiply through by the *smallest integer* that turns all subscripts into integers.

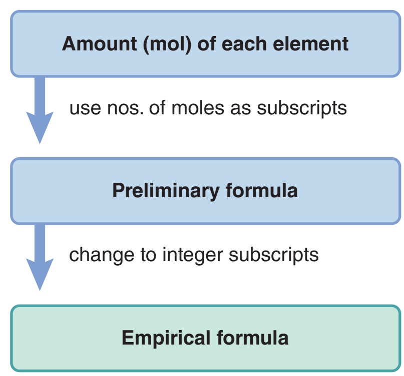
Sample Problem 3.8 demonstrates these math steps.

SAMPLE PROBLEM 3.8 Determining an Empirical Formula from Amounts of Elements

Problem A sample of an unknown compound contains 0.21 mol of zinc, 0.14 mol of phosphorus, and 0.56 mol of oxygen. What is the empirical formula?

Plan We are given the amount (mol) of each element as a fraction. We use these fractional amounts directly in a preliminary formula as subscripts of the element symbols. Then, we convert the fractions to whole numbers.

Road Map



Solution Using the fractions to write a preliminary formula, with the symbols Zn for zinc, P for phosphorus, and O for oxygen:

$$Zn_{0.21}P_{0.14}O_{0.56}$$

Converting the fractions to whole numbers:

1. Divide each subscript by the smallest one, which in this case is 0.14:

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

2. Multiply through by the *smallest integer* that turns all subscripts into integers. We multiply by 2 to make 1.5 (the subscript for Zn) into an integer:

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

Check The integer subscripts must be the smallest integers with the same ratio as the original fractional numbers of moles: 3/2/8 is *the same ratio* as 0.21/0.14/0.56.

Comment A more conventional way to write this formula is $Zn_3(PO_4)_2$; this compound is zinc phosphate, formerly used widely as a dental cement.

FOLLOW-UP PROBLEMS

3.8A A sample of a white solid contains 0.170 mol of boron and 0.255 mol of oxygen. What is the empirical formula?

3.8B A sample of an unknown compound contains 6.80 mol of carbon and 18.1 mol of hydrogen. What is the empirical formula?

SOME SIMILAR PROBLEMS 3.42(a) and 3.43(a)

Sample Problems 3.9–3.11 show how other types of compositional data are used to determine chemical formulas.

SAMPLE PROBLEM 3.9 **Determining an Empirical Formula from Masses of Elements**

Problem Analysis of a sample of an ionic compound yields 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What are the empirical formula and the name of the compound?

Plan This problem is similar to Sample Problem 3.8, except that we are given element *masses* that we must convert into integer subscripts. We first divide each mass by the element's molar mass to find the amount (mol). Then we construct a preliminary formula and convert the amounts (mol) to integers.

Solution Finding amount (mol) of each element:

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

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

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

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

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

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

The empirical formula is $NaClO_4$; the name is sodium perchlorate.

Check The numbers of 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 PROBLEMS

3.9A A sample of an unknown compound is found to contain 1.23 g of H, 12.64 g of P, and 26.12 g of O. What is the empirical formula?

3.9B 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 amount (mol) of S, and use the formula to find the amount (mol) of M.]

SOME SIMILAR PROBLEMS 3.42(b), 3.43(b), 3.46(b), and 3.47(b)

Molecular Formulas

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

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

Multiplying the empirical formula subscripts by 2 gives the molecular formula:

$$H_{(1\times2)}O_{(1\times2)} \text{ gives } H_2O_2$$

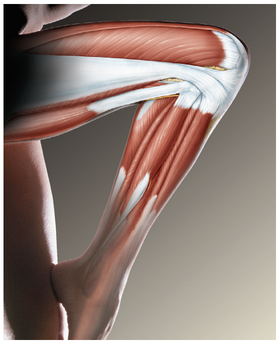
Since the molar mass of hydrogen peroxide is twice as large as the empirical formula mass, the molecular formula has twice the number of atoms as the empirical formula.

Instead of giving compositional data as masses of each element, analytical laboratories provide mass percents. We use this kind of data as follows:

1. Assume 100.0 g of compound to express each mass percent directly as mass (g).
2. Convert each mass (g) to amount (mol).

3. Derive the empirical formula.
4. Divide the molar mass of the compound by the empirical formula mass to find the whole-number multiple and the molecular formula.

SAMPLE PROBLEM 3.10 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 has 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. The mass of the sample of lactic acid is not given, but the mass percents are the same for any sample of it. Therefore, we assume there is 100.0 g of lactic acid and express each mass % as a number of grams. Then, we construct the empirical formula as in Sample Problem 3.9.

Solution Expressing mass % as mass (g) by assuming 100.0 g of lactic acid:

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

Similarly, we have 6.71 g of H and 53.3 g of O.

Converting from mass (g) of each element to amount (mol):

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

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:

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

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 of 1.

(b) Determining the molecular formula

Plan The molecular formula subscripts are whole-number multiples of the empirical formula subscripts. To find this multiple, 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 each subscript in the empirical formula by the multiple.

Solution The empirical formula 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:

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

FOLLOW-UP PROBLEMS

3.10A 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, cigarette smoke, and even 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?

3.10B Caffeine ($\mathcal{M}$ = 194.2 g/mol) is a stimulant found in coffee, tea, soft drinks, and chocolate. Elemental analysis of caffeine shows 49.47 mass % C, 5.19 mass % H, 28.86 mass % N, and 16.48 mass % O. What is the molecular formula of caffeine?

SOME SIMILAR PROBLEMS 3.44, 3.45, and 3.49

Combustion Analysis of Organic Compounds Still another type of compositional data is obtained through **combustion analysis,** used to measure the amounts of carbon and hydrogen in a combustible organic compound. The unknown compound is burned in an excess of pure O_2; during the combustion, the compound's carbon and hydrogen react with the oxygen to form CO_2 and H_2O, respectively, which are absorbed in separate containers (Figure 3.4). By weighing the absorbers before and after combustion, we find the masses of CO_2 and H_2O and use them to find the masses of C and H in the compound; from these results, we find the empirical formula. Many organic compounds also contain oxygen, nitrogen, or a halogen. As long as the third element doesn't interfere with the absorption of H_2O and CO_2, we calculate its mass by subtracting the masses of C and H from the original mass of the compound.

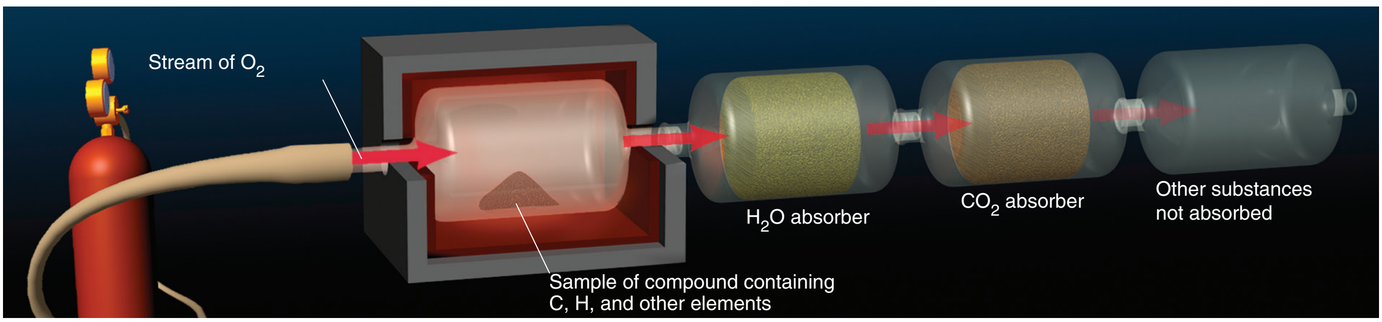


Figure 3.4 Combustion apparatus for determining formulas of organic compounds. A sample of an organic compound is burned in a stream of O_2. The resulting H_2O is absorbed by $Mg(ClO_4)_2$, and the CO_2 is absorbed by NaOH on asbestos.

SAMPLE PROBLEM 3.11 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 burned in a combustion apparatus, the following data are obtained:

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

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 and after the combustion. From the mass of CO_2, we use Equation 3.7 to find the mass of C. 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 masses of C and H gives the mass of O, the third element present. Then, we proceed as in Sample Problem 3.10: calculate amount (mol) of each element using its molar mass, 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:

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

Calculating masses (g) of C and H using Equation 3.7:

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

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

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

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

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

Calculating mass (g) 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 (g) 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

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

We find that 3 is the smallest integer that makes all subscripts 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:

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

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

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 the preliminary formula (0.0341/0.046/0.0341), and it gives the known molar mass:

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

Comment The subscript we calculated for H was 3.9, which we rounded to 4. But, if we had strung the calculation steps together, we would have obtained 4.0:

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

FOLLOW-UP PROBLEMS

3.11A 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 were formed. Find the molecular formula.

3.11B Anabolic steroids are sometimes used illegally by athletes to increase muscle strength. A forensic chemist analyzes some tablets suspected of being a popular steroid. He determines that the substance in the tablets contains only C, H, and O and has a molar mass of 300.42 g/mol. When a 1.200-g sample is studied by combustion analysis, 3.516 g of CO_2 and 1.007 g of H_2O are collected. What is the molecular formula of the substance in the tablets?

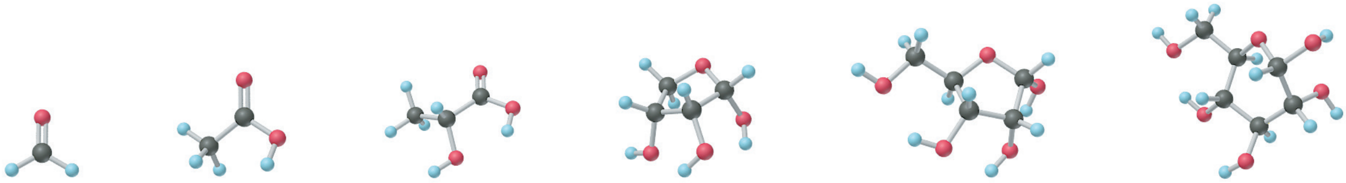
A SIMILAR PROBLEM 3.51

Chemical Formulas and Molecular Structures; Isomers

A formula represents a real, three-dimensional object. The structural formula makes this point, with its relative placement of atoms, but do empirical and molecular formulas contain structural information?

Table 3.2 Some Compounds with Empirical Formula CH_2O (Composition by Mass: 40.0% C, 6.71% H, 53.3% O)

Name	Molecular Formula	Whole-Number Multiple	$\mathcal{M}$ (g/mol)	Use or Function
Formaldehyde	CH_2O	1	30.03	Disinfectant; biological preservative
Acetic acid	$C_2H_4O_2$	2	60.05	Acetate polymers; vinegar (5% solution)
Lactic acid	$C_3H_6O_3$	3	90.08	Causes milk to sour; forms in muscles during exercise
Erythrose	$C_4H_8O_4$	4	120.10	Forms during sugar metabolism
Ribose	$C_5H_{10}O_5$	5	150.13	Component of many nucleic acids and vitamin B_2
Glucose	$C_6H_{12}O_6$	6	180.16	Major nutrient for energy in cells

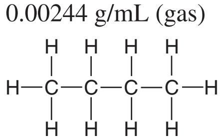
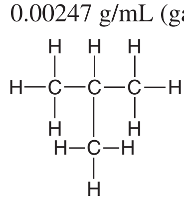
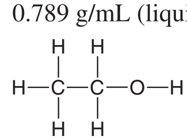
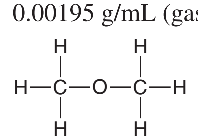
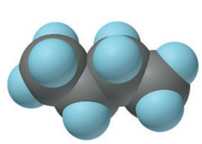
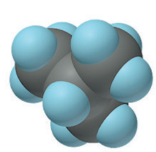
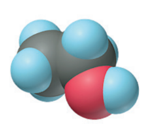
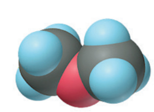


CH_2O $C_2H_4O_2$ $C_3H_6O_3$ $C_4H_8O_4$ $C_5H_{10}O_5$ $C_6H_{12}O_6$

1. *Different compounds with the same empirical formula*. The empirical formula tells nothing about molecular structure because it is based solely on mass analysis. In fact, different compounds can have the *same* empirical formula. NO_2 and N_2O_4 are inorganic cases, and there are numerous organic ones. For example, many compounds have the empirical formula CH_2 (the general formula is C_nH_{2n}, with n an integer greater than or equal to 2), such as ethylene (C_2H_4) and propylene (C_3H_6), starting materials for two common plastics. Table 3.2 shows some biological compounds with the same empirical formula, CH_2O.

2. *Isomers: Different compounds with the same molecular formula.* A molecular formula also tells nothing about structure. Different compounds can have the *same* molecular formula because their atoms can bond in different arrangements to give more than one *structural formula.* **Isomers** are compounds with the same molecular formula, and thus molar mass, but different properties. *Constitutional,* or *structural, isomers* occur when the atoms link together in different arrangements. Table 3.3 shows two pairs of examples. The left pair, butane and 2-methylpropane, share the molecular formula C_4H_{10}. One has a four-C chain and the other a one-C branch off a three-C chain. Both

Table 3.3 Two Pairs of Constitutional Isomers

	C_4H_{10}		C_2H_6O	
Property	**Butane**	**2-Methylpropane**	**Ethanol**	**Dimethyl Ether**
$\mathcal{M}$ (g/mol)	58.12	58.12	46.07	46.07
Boiling point	−0.5°C	−11.6°C	78.5°C	−25°C
Density (at 20°C)	0.00244 g/mL (gas)	0.00247 g/mL (gas)	0.789 g/mL (liquid)	0.00195 g/mL (gas)
Structural formula				
Space-filling model				

are small alkanes, so their properties are similar, but not identical. The two compounds with the molecular formula C_2H_6O have very different properties; indeed, they are different classes of organic compound—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 two structural formulas (Table 3.3), C_3H_8O has three, and $C_4H_{10}O$ seven. Imagine how many there are for $C_{16}H_{19}N_3O_4S$! Of all the possible isomers with this molecular formula, only one is the antibiotic ampicillin (Figure 3.5). We'll discuss these and other types of isomerism fully later in the text.

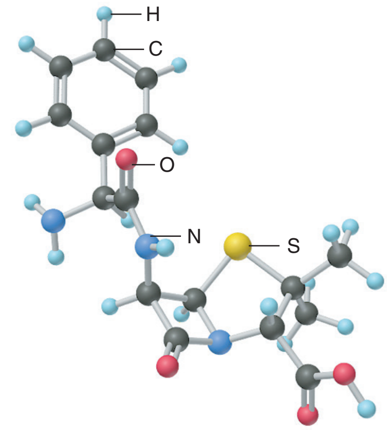


Figure 3.5 The antibiotic ampicillin.

Summary of Section 3.2

- From the masses of elements in a compound, their relative numbers of moles are found, which gives the empirical formula.
- If the molar mass of the compound is known, the molecular formula, the actual numbers of moles of each element, can also be determined, because the molecular formula is a whole-number multiple of the empirical formula.
- Combustion analysis provides data on the masses of carbon and hydrogen in an organic compound, which are used to obtain the formula.
- Atoms can bond in different arrangements (structural formulas). Two or more compounds with the same molecular formula are constitutional isomers.

3.3 WRITING AND BALANCING CHEMICAL EQUATIONS

Thinking in terms of amounts, rather than masses, allows us to view reactions as large populations of interacting particles rather than as grams of material. For example, for the formation of HF from H_2 and F_2, if we weigh the substances, we find that

Macroscopic level (grams): 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 (g) to amounts (mol), we find that

Macroscopic level (moles): 1 mol of H_2 and 1 mol of F_2 react to form 2 mol of HF

This information reveals that an enormous number of H_2 molecules react with just as many F_2 molecules to form twice as many HF molecules. Dividing by Avogadro's number gives the reaction between individual molecules:

Molecular level: 1 molecule of H_2 and 1 molecule of F_2 react to form 2 molecules of HF

Thus, *the macroscopic (molar) change corresponds to the submicroscopic (molecular) change* (Figure 3.6). This information forms the essence of a **chemical equation,** a statement that uses formulas to express the identities and quantities of substances in a chemical or physical change.

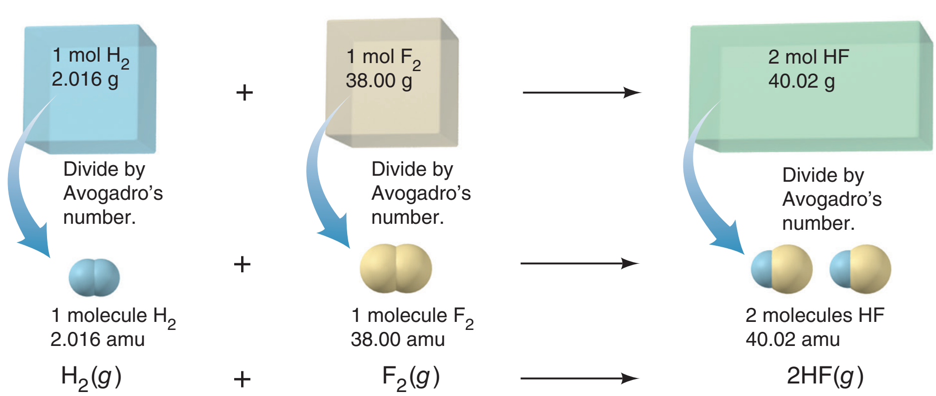


Figure 3.6 The formation of HF on the macroscopic and molecular levels and written as a balanced chemical equation.

Steps for Balancing an Equation To present a chemical change quantitatively, the equation must be *balanced: the same number of each type of atom must appear on both sides.* As an example, here is a description of a chemical change that occurs in many fireworks and in a common lecture demonstration: a magnesium strip burns in oxygen gas to yield powdery magnesium oxide. (Light and heat are also produced, but we are concerned here only with substances.) Converting this description into a balanced equation involves the following steps:

1. *Translating the statement.* We first translate the chemical statement into a "skeleton" equation: the substances present *before* the change, called **reactants,** are placed to the left of a yield arrow, which points to the substances produced *during* the change, called **products:**

$$\overbrace{\_\_\mathrm{Mg} \quad + \quad \_\_\mathrm{O_2}}^{\text{reactants}} \xrightarrow{\text{yield}} \overset{\text{product}}{\_\_\mathrm{MgO}}$$

$$\text{magnesium and oxygen} \quad \text{yield} \quad \text{magnesium oxide}$$

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

2. *Balancing the atoms.* By shifting our attention back and forth, we *match the numbers of each type of atom on the left and the right of the yield arrow.* In each blank, we place 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 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 that blank:

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

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

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

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

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

   In terms of numbers of each type of atom, the equation is balanced.

3. *Adjusting the coefficients.* There are several conventions about the final coefficients:
   - In most cases, *the smallest whole-number coefficients are preferred.* In this case, one-half of an O_2 molecule cannot exist, so we multiply the equation by 2:

$$2\mathrm{Mg} + 1\mathrm{O_2} \longrightarrow 2\mathrm{MgO}$$

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

$$2\mathrm{Mg} + \mathrm{O_2} \longrightarrow 2\mathrm{MgO}$$

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

4. *Checking.* After balancing and adjusting the coefficients, we 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 used for these states are shown in the margin. From the original statement, we know that a Mg "strip" is solid, O_2 is a gas, and "powdery" MgO is also solid. The balanced equation, therefore, is

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

g for gas
l for liquid
s for solid
aq for aqueous solution

As you saw in Figure 3.6, *balancing coefficients refer to both individual chemical entities and moles of entities*. Thus,

2 atoms of Mg and 1 molecule of O_2 yield 2 formula units of MgO
2 moles of Mg and 1 mole of O_2 yield 2 moles of MgO

Figure 3.7 depicts this reaction on three levels:

- *Macroscopic level (photos),* as it appears in the laboratory
- *Atomic level (blow-up circles),* as chemists imagine it (with darker colored atoms representing the stoichiometry)
- *Symbolic level,* in the form of the balanced chemical equation

Keep in mind several key points about the balancing process:

- A coefficient operates on *all* the atoms in the formula that follows it:

  2MgO *means* $2 \times$ (MgO), or 2 Mg atoms + 2 O atoms
  $2Ca(NO_3)_2$ *means* $2 \times [Ca(NO_3)_2]$, or 2 Ca atoms + 4 N atoms + 12 O atoms

- Chemical formulas *cannot* be altered. Thus, in step 2 of the example, we *cannot* balance the O atoms by changing MgO to MgO_2 because MgO_2 is a different compound.
- Other reactants or products *cannot* be added. Thus, we *cannot* balance the O atoms by changing the reactant from O_2 molecules to O atoms or by adding an O atom to the products. The description of the reaction mentions oxygen gas, which consists of O_2 molecules, *not* separate O atoms.
- A balanced equation remains balanced if you multiply all the coefficients by the same number. For example,

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

  is also balanced because the coefficients have just been multiplied by 2. However, *by convention,* we balance an equation with the *smallest* whole-number coefficients.

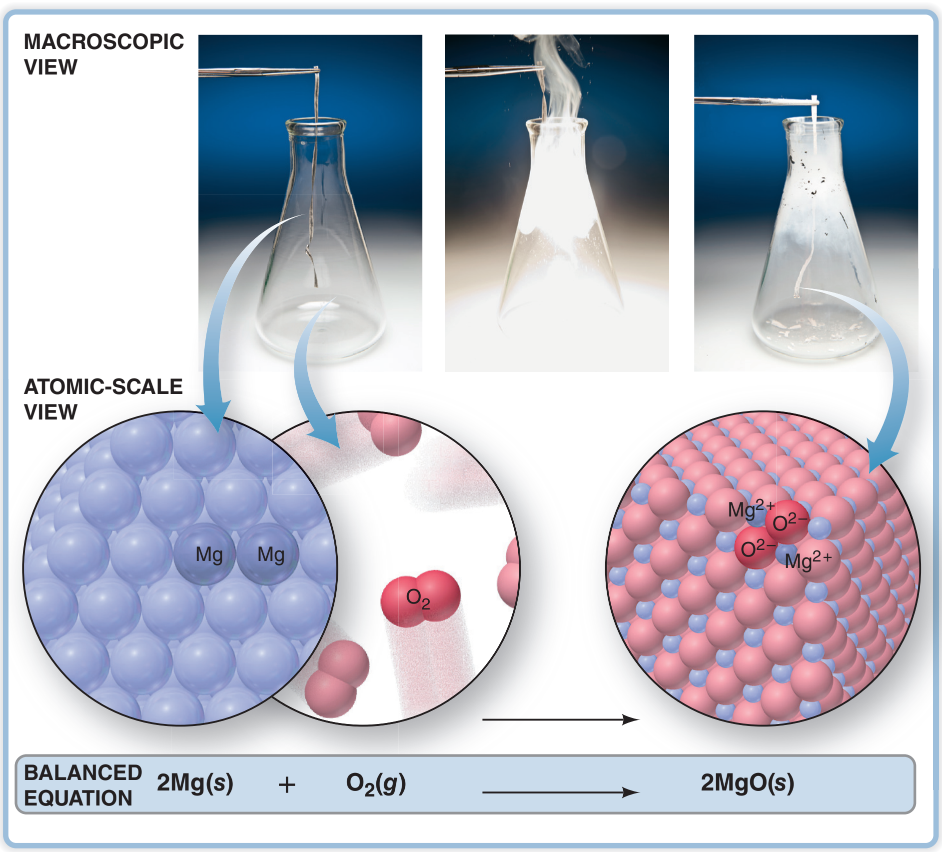


Figure 3.7 A three-level view of the reaction between magnesium and oxygen.

SAMPLE PROBLEM 3.12 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.* 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 the 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) ⟶ 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.

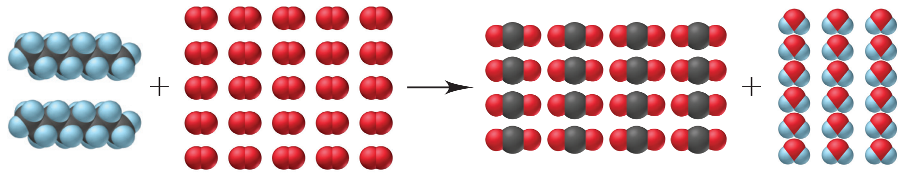
FOLLOW-UP PROBLEMS

3.12A Write a balanced equation for each of the following:
(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.

3.12B Write a balanced equation for each of the following:
(a) 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.
(b) A reaction that takes place in a self-contained breathing apparatus: solid potassium superoxide (KO_2) reacts with carbon dioxide gas to produce oxygen gas and solid potassium carbonate.
(c) The production of iron from its ore in a blast furnace: solid iron(III) oxide reacts with carbon monoxide gas to produce solid iron metal and carbon dioxide gas.

SOME SIMILAR PROBLEMS 3.56–3.61

Visualizing a Reaction with a Molecular Scene A great way to focus on the rearrangement of atoms from reactants to products is by visualizing an equation as a molecular scene. Here's a representation of the combustion of octane we just balanced:

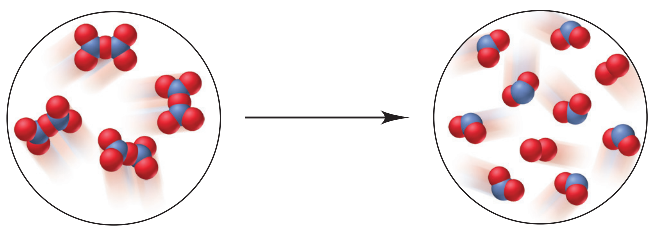


$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

Now let's work through a sample problem to do the reverse—derive a balanced equation from a molecular scene.

SAMPLE PROBLEM 3.13 Balancing an Equation from a Molecular Scene

Problem The following molecular scenes depict an important reaction in nitrogen chemistry (nitrogen is blue; oxygen is red):



Write a balanced equation for this reaction.

Plan To write a balanced equation, we first have to determine the formulas of the molecules and obtain coefficients by counting the number of each type of molecule. Then, we arrange this information in the correct equation format, using the smallest whole-number coefficients and including states of matter.

Solution The reactant circle shows only one type of molecule. It has two N and five O atoms, so the formula is N_2O_5; there are four of these molecules. The product circle shows two different molecules, one with one N and two O atoms, and the other with two O atoms; there are eight NO_2 and two O_2. Thus, we have

$$4N_2O_5 \longrightarrow 8NO_2 + 2O_2$$

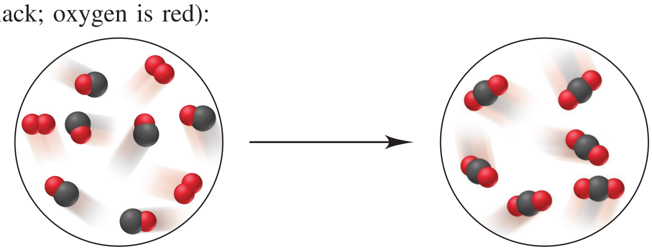
Writing the balanced equation with the smallest whole-number coefficients and all substances as gases:

$$\boxed{2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)}$$

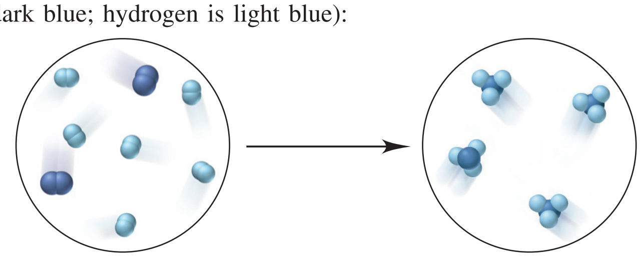
Check Reactant (4 N, 10 O) $\longrightarrow$ products (4 N, 8 + 2 = 10 O)

FOLLOW-UP PROBLEMS

3.13A Write a balanced equation for the important atmospheric reaction depicted below (carbon is black; oxygen is red):



3.13B Write a balanced equation for the important industrial reaction depicted below (nitrogen is dark blue; hydrogen is light blue):



A SIMILAR PROBLEM 3.55

Summary of Section 3.3

- A chemical equation has reactant formulas on the left of a yield arrow and product formulas on the right.
- A balanced equation has the same number of each type of atom on both sides.
- 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 QUANTITIES OF REACTANT AND PRODUCT

A balanced equation is essential for all calculations involving chemical change: *if you know the number of moles of one substance, the balanced equation tells you the number of moles of the others.*

Stoichiometrically Equivalent Molar Ratios from the Balanced Equation

In a balanced equation, *the amounts (mol) of substances are stoichiometrically equivalent to each other,* which means that a specific amount of one substance is formed from, produces, or reacts with a specific amount of the other. The quantitative relationships are expressed as *stoichiometrically equivalent molar ratios* that we use as conversion factors to calculate the amounts. For example, consider 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

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. A balanced equation contains a wealth of quantitative information relating individual chemical entities, amounts (mol) of substances, and masses of substances, and Table 3.4 presents the quantitative information contained in this equation.

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

As with any equivalent quantities, we can construct two conversion factors, depending on the quantity we want to find:

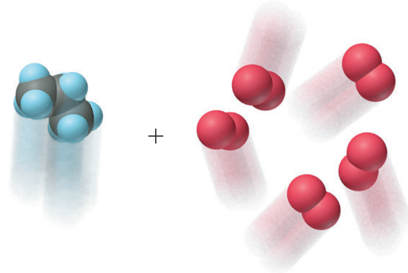
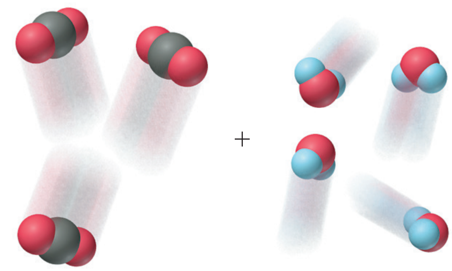
$$\frac{5\ \text{mol } O_2}{4\ \text{mol } H_2O} \quad \text{or} \quad \frac{4\ \text{mol } H_2O}{5\ \text{mol } O_2}$$

Since we want to find the amount (mol) of O_2 and we know the amount (mol) of H_2O, we choose "5 mol O_2/4 mol H_2O" *(above left)* to cancel "mol H_2O":

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

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

Table 3.4 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
			$\longrightarrow$		
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	

*You **cannot** solve this type of problem without the balanced equation.* Here is an approach for solving *any* stoichiometry problem that involves a reaction:

1. Write the balanced equation.
2. When necessary, convert the known mass (or number of entities) of one substance to amount (mol) using its molar mass (or Avogadro's number).
3. Use the molar ratio to calculate the unknown amount (mol) of the other substance.
4. When necessary, convert the amount of that other substance to the desired mass (or number of entities) using its molar mass (or Avogadro's number).

Figure 3.8 summarizes the possible relationships among quantities of substances in a reaction, and Sample Problems 3.14–3.16 apply three of them in the first chemical step of converting copper ore to copper metal.

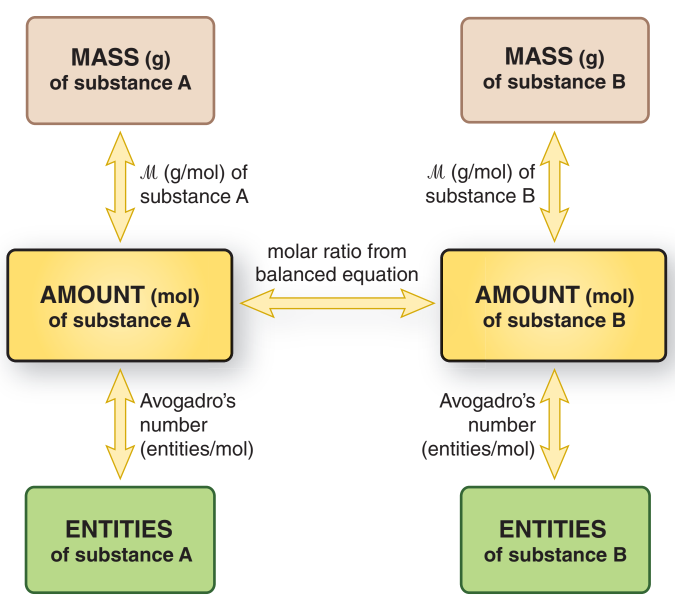
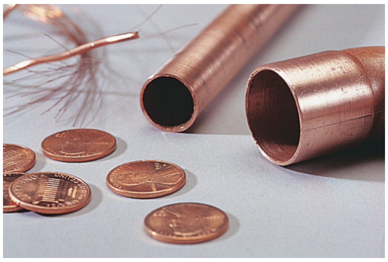


Figure 3.8 Summary of amount-mass-number relationships in a chemical equation. Start at any box (known) and move to any other (unknown) by using the conversion factor on the arrow. As always, convert to amount (mol) first.



Road Map

Amount (mol) of Cu_2S

↓ molar ratio (2 mol Cu_2S = 3 mol O_2)

Amount (mol) of O_2

SAMPLE PROBLEM 3.14 Calculating Quantities of Reactants and Products: Amount (mol) to Amount (mol)

Problem In a lifetime, the average American uses more than a half ton (>500 kg) of copper in coins, plumbing, and wiring. Copper is obtained from sulfide ores, such as chalcocite [copper(I) sulfide] by a multistep process. After initial grinding, the ore is "roasted" (heated strongly with oxygen gas) to form powdered copper(I) oxide and gaseous sulfur dioxide. How many moles of oxygen are required to roast 10.0 mol of copper(I) sulfide?

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 know the amount of Cu_2S (10.0 mol) and must find the amount (mol) of O_2 that is needed to roast it. The balanced equation shows that 3 mol of O_2 is needed for 2 mol of Cu_2S, so the conversion factor for finding amount (mol) of O_2 is "3 mol O_2/2 mol Cu_2S" (see the road map).

Solution Calculating the amount of O_2:

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

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

Comment A *common mistake* is to invert the conversion factor; that calculation would be

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

The strange units should alert you that an error was made in setting up the conversion factor. Also note that this answer, 6.67, is *less* than 10.0, whereas the equation shows that there should be *more* moles of O_2 (3 mol) than moles of Cu_2S (2 mol). Be sure to think through the calculation when setting up the conversion factor and canceling units.

FOLLOW-UP PROBLEMS

3.14A 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. How many moles of iron(III) oxide are needed to form 3.60×10^3 mol of iron? Include a road map that shows how you planned the solution.

3.14B The tarnish that forms on objects made of silver is solid silver sulfide; it can be removed by reacting it with aluminum metal to produce silver metal and solid aluminum sulfide. How many moles of aluminum are required to remove 0.253 mol of silver sulfide from a silver bowl? Include a road map that shows how you planned the solution.

A SIMILAR PROBLEM 3.68(a)

Road Map

Amount (mol) of Cu_2S

↓ molar ratio (2 mol Cu_2S = 2 mol SO_2)

Amount (mol) of SO_2

↓ multiply by $\mathcal{M}$ (g/mol) (1 mol SO_2 = 64.07 g SO_2)

Mass (g) of SO_2

SAMPLE PROBLEM 3.15 Calculating Quantities of Reactants and Products: Amount (mol) to Mass (g)

Problem During the roasting process, how many grams of sulfur dioxide form when 10.0 mol of copper(I) sulfide reacts?

Plan Referring to the balanced equation in Sample Problem 3.14, here we are given amount of reactant (10.0 mol of Cu_2S) and need the mass (g) of product (SO_2) that forms. We find the amount (mol) of SO_2 using the molar ratio (2 mol SO_2/2 mol Cu_2S) and then multiply by its molar mass (64.07 g/mol) to find the mass (g) of SO_2 (see the road map).

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 is formed and each mole weighs about 64 g. We rounded to three significant figures.

FOLLOW-UP PROBLEMS

3.15A In the thermite reaction (see Follow-up Problem 3.14A), what amount (mol) of iron forms when 1.85×10^{25} formula units of iron(III) oxide reacts? Write a road map to show how to plan the solution.

3.15B In the reaction that removes silver tarnish (see Follow-up Problem 3.14B), how many moles of silver are produced when 32.6 g of silver sulfide reacts? Write a road map to show how you planned the solution.

SOME SIMILAR PROBLEMS 3.68(b), 3.69, 3.70(a), and 3.71(a)

SAMPLE PROBLEM 3.16 Calculating Quantities of Reactants and Products: Mass to Mass

Problem During the roasting of chalcocite, how many kilograms of oxygen are required to form 2.86 kg of copper(I) oxide?

Plan In this problem, we know the mass of the product, Cu_2O (2.86 kg), and we need the mass (kg) of O_2 that reacts to form it. Therefore, we must convert from mass of product to amount of product to amount of reactant to mass of reactant. We convert the mass of Cu_2O from kg to g and then to amount (mol). Then, we use the molar ratio (3 mol O_2/2 mol Cu_2O) to find the amount (mol) of O_2 required. Finally, we convert the amount of O_2 to g and then kg (see the road map).

Road Map

Mass (kg) of Cu_2O
↓ 1 kg = 10^3 g
Mass (g) of Cu_2O
↓ divide by $\mathcal{M}$ (g/mol) (143.10 g Cu_2O = 1 mol Cu_2O)
Amount (mol) of Cu_2O
↓ molar ratio (2 mol Cu_2O = 3 mol O_2)
Amount (mol) of O_2
↓ multiply by $\mathcal{M}$ (g/mol) (1 mol O_2 = 32.00 g O_2)
Mass (g) of O_2
↓ 10^3 g = 1 kg
Mass (kg) of O_2

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

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

Converting from moles of Cu_2O to moles of O_2:

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

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

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

Check The units are correct. Rounding 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 The three related sample problems (3.14–3.16) highlight the main point for solving stoichiometry problems: *convert the information given into amount (mol).* Then, use the appropriate molar ratio and any other conversion factors to complete the solution.

FOLLOW-UP PROBLEMS

3.16A During the thermite reaction (see Follow-up Problems 3.14A and 3.15A), how many atoms of aluminum react for every 1.00 g of aluminum oxide that forms? Include a road map that shows how you planned the solution.

3.16B During the reaction that removes silver tarnish (see Follow-up Problems 3.14B and 3.15B), how many grams of aluminum react to form 12.1 g of aluminum sulfide? Include a road map that shows how you planned the solution.

SOME SIMILAR PROBLEMS 3.70(b), 3.71(b), and 3.72–3.75

Reactions That Occur in a Sequence

In many situations, a product of one reaction becomes a reactant for the next in a sequence of reactions. For stoichiometric purposes, when the same (common) substance forms in one reaction and reacts in the next, we eliminate it in an **overall (net) equation.** The steps in writing the overall equation are

1. Write the sequence of balanced equations.
2. Adjust the equations arithmetically to cancel the common substance(s).
3. Add the adjusted equations together to obtain the overall balanced equation.

Sample Problem 3.17 shows the approach by continuing with the copper recovery process that started in Sample Problem 3.14.

SAMPLE PROBLEM 3.17 Writing an Overall Equation for a Reaction Sequence

Problem Roasting is the first step in extracting copper from chalcocite. In the next step, copper(I) oxide reacts with powdered carbon to yield copper metal and carbon monoxide gas. Write a balanced overall equation for the two-step sequence.

Plan To obtain the overall equation, we write the individual equations in sequence, adjust coefficients to cancel the common substance (or substances), and add the equations together. In this case, only Cu_2O appears as a product in one equation and a reactant in the other, so it is the common substance.

Solution Writing the individual balanced equations:

$$2Cu_2S(s) + 3O_2(g) \longrightarrow 2Cu_2O(s) + 2SO_2(g) \quad [\text{equation 1; see Sample Problem 3.14}]$$
$$Cu_2O(s) + C(s) \longrightarrow 2Cu(s) + CO(g) \quad [\text{equation 2}]$$

Adjusting the coefficients: Since 2 mol of Cu_2O form in equation 1 but 1 mol of Cu_2O reacts in equation 2, we double *all* the coefficients in equation 2 to use up the Cu_2O:

$$2Cu_2S(s) + 3O_2(g) \longrightarrow 2Cu_2O(s) + 2SO_2(g) \quad [\text{equation 1}]$$
$$2Cu_2O(s) + 2C(s) \longrightarrow 4Cu(s) + 2CO(g) \quad [\text{equation 2, doubled}]$$

Adding the two equations and canceling the common substance: We keep the reactants of both equations on the left and the products of both equations on the right:

$$2Cu_2S(s) + 3O_2(g) + \cancel{2Cu_2O(s)} + 2C(s) \longrightarrow \cancel{2Cu_2O(s)} + 2SO_2(g) + 4Cu(s) + 2CO(g)$$

or,
$$2Cu_2S(s) + 3O_2(g) + 2C(s) \longrightarrow 2SO_2(g) + 4Cu(s) + 2CO(g)$$

Check Reactants (4 Cu, 2 S, 6 O, 2 C) ⟶ products (4 Cu, 2 S, 6 O, 2 C)

Comment **1.** Even though Cu_2O *does* participate in the chemical change, it is not involved in the reaction stoichiometry. An overall equation *may not* show which substances actually react; for example, C(*s*) and $Cu_2S(s)$ do not interact directly in this reaction sequence, even though both are shown as reactants.
2. The SO_2 formed in copper recovery contributes to acid rain, so chemists have devised microbial and electrochemical methods to extract metals without roasting sulfide ores. Such methods are examples of *green chemistry;* we'll discuss another on page 124.
3. These reactions were shown to explain how to obtain an overall equation. The actual extraction of copper is more complex, as you'll see in Chapter 22.

FOLLOW-UP PROBLEMS

3.17A The SO_2 formed in copper recovery reacts in air with oxygen and forms sulfur trioxide. This gas, in turn, reacts with water to form a sulfuric acid solution that falls in rain. Write a balanced overall equation for this process.

3.17B During a lightning strike, nitrogen gas can react with oxygen gas to produce nitrogen monoxide. This gas then reacts with the gas ozone, O_3, to produce nitrogen dioxide gas and oxygen gas. The nitrogen dioxide that is produced is a pollutant in smog. Write a balanced overall equation for this process.

SOME SIMILAR PROBLEMS 3.76 and 3.77

Reaction Sequences in Organisms Multistep reaction sequences called *metabolic pathways* occur throughout biological systems. (We discuss them again in Chapter 17.) For example, in most cells, the chemical energy in glucose is released through a sequence of about 30 individual reactions. The product of each reaction step is the reactant of the next, so that all the common substances cancel, and the overall equation is

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$$

We eat food that contains glucose, inhale O_2, and excrete CO_2 and H_2O. In our cells, these reactants and products are many steps apart: O_2 never reacts *directly* with glucose, and CO_2 and H_2O are formed at various, often distant, steps along the sequence of reactions. Even so, the molar ratios in the overall equation are the same as if the glucose burned in a combustion chamber filled with O_2 and formed CO_2 and H_2O directly (Figure 3.9).

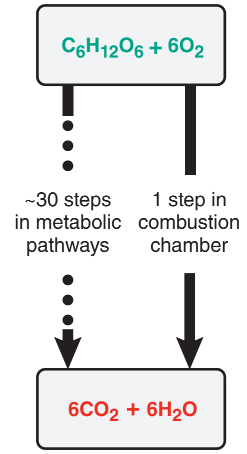


Figure 3.9 An overall equation equals the sum of the individual steps.

Reactions That Involve a Limiting Reactant

In problems up to now, the amount of *one* reactant was given, and we assumed there was enough of the other reactants to react with it completely. For example, suppose we want the amount (mol) of SO_2 that forms when 5.2 mol of Cu_2S reacts with O_2:

$$2Cu_2S(s) + 3O_2(g) \longrightarrow 2Cu_2O(s) + 2SO_2(g) \quad [\text{equation 1; see Sample Problem 3.14}]$$

We assume the 5.2 mol of Cu_2S reacts with as much O_2 as needed. Because all the Cu_2S reacts, its initial amount of 5.2 mol determines, or *limits,* the amount of SO_2 that can form, no matter how much more O_2 is present. In this situation, we call Cu_2S the **limiting reactant** (or *limiting reagent*).

Suppose, however, you know the amounts of both Cu_2S *and* O_2 and need to find out how much SO_2 forms. You first have to determine whether Cu_2S *or* O_2 is the limiting reactant—that is, which one is completely used up—because that reactant limits how much SO_2 can form. The reactant that is *not* limiting is present *in excess,* which means the amount that doesn't react is left over. ❯ To determine which is the limiting reactant, we use the molar ratios in the balanced equation to perform a series of calculations to see *which reactant forms* ***less*** *product*.

Determining the Limiting Reactant Let's clarify these ideas in a much more appetizing situation. Suppose you have a job making ice cream sundaes. Each sundae requires two scoops (12 oz) of ice cream, one cherry, and 50 mL of syrup:

$$2 \text{ scoops (12 oz)} + 1 \text{ cherry} + 50 \text{ mL syrup} \longrightarrow 1 \text{ sundae}$$

A mob of 25 ravenous school kids enters, and each one wants a sundae with vanilla ice cream and chocolate syrup. You have 300 oz of vanilla ice cream (at 6 oz per scoop), 30 cherries, and 1 L of syrup: can you feed them all? A series of calculations based on the balanced equation shows the number of sundaes you can make from each ingredient:

$$\text{Ice cream: No. of sundaes} = 300 \cancel{\text{oz}} \times \frac{1 \cancel{\text{scoop}}}{6 \cancel{\text{oz}}} \times \frac{1 \text{ sundae}}{2 \cancel{\text{scoops}}} = 25 \text{ sundaes}$$

$$\text{Cherries: No. of sundaes} = 30 \cancel{\text{cherries}} \times \frac{1 \text{ sundae}}{1 \cancel{\text{cherry}}} = 30 \text{ sundaes}$$

$$\text{Syrup: No. of sundaes} = 1000 \cancel{\text{mL syrup}} \times \frac{1 \text{ sundae}}{50 \cancel{\text{mL syrup}}} = 20 \text{ sundaes}$$

Of the reactants (ice cream, cherry, syrup), the syrup forms the *least* product (sundaes), so it is the limiting "reactant." When all the syrup has been used up, some ice cream and cherries are "unreacted" so they are in excess:

$$300 \text{ oz (50 scoops)} + 30 \text{ cherries} + 1 \text{ L syrup} \longrightarrow$$
$$\mathbf{20\ sundaes} + 60 \text{ oz (10 scoops)} + 10 \text{ cherries}$$

Figure 3.10 on the next page shows a similar example with different initial (starting) quantities.

Limiting "Reactants" in Everyday Life

Limiting-"reactant" situations arise in business all the time. A car assembly-plant manager must order more tires if there are 1500 car bodies and only 4000 tires, and a clothes manufacturer must cut more sleeves if there are 320 sleeves for 170 shirt bodies. You've probably faced such situations in daily life as well. A muffin recipe calls for 2 cups of flour and 1 cup of sugar, but you have 3 cups of flour and only $\frac{3}{4}$ cup of sugar. Clearly, the flour is in excess and the sugar limits the number of muffins you can make. Or, you're in charge of making cheeseburgers for a picnic, and you have 10 buns, 12 meat patties, and 15 slices of cheese. Here, the number of buns limits how many cheeseburgers you can make. Or, there are 26 students and only 23 microscopes in a cell biology lab. You'll find that limiting-"reactant" situations are almost limitless.

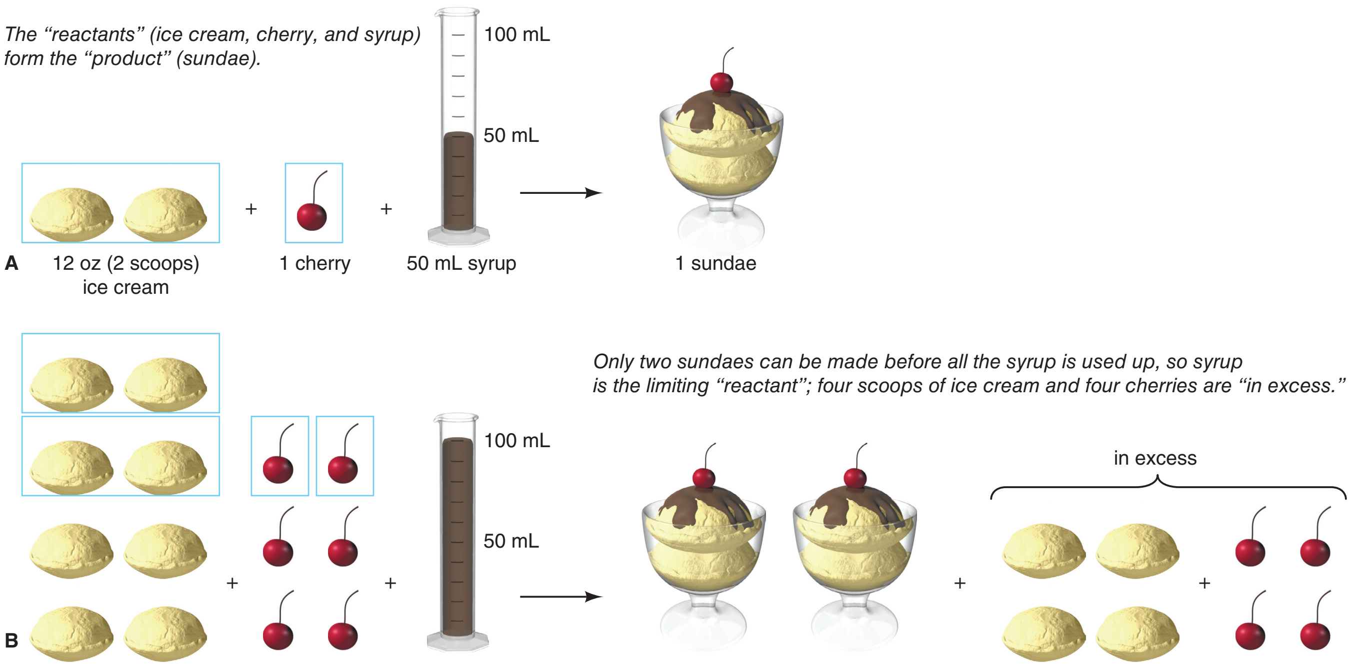


Figure 3.10 An ice cream sundae analogy for limiting reactants.

Using Reaction Tables in Limiting-Reactant Problems A good way to keep track of the quantities in a limiting-reactant problem is with a *reaction table*. The balanced equation appears at the top for the column heads. The table shows the

- *Initial* quantities of reactants and products *before* the reaction
- *Change* in the quantities of reactants and products *during* the reaction
- *Final* quantities of reactants and products remaining *after* the reaction

For example, for the ice-cream sundae "reaction," the reaction table would be

Quantity	12 oz (2 scoops)	+	1 cherry	+	50 mL syrup	⟶	1 sundae
Initial	300 oz (50 scoops)		30 cherries		1000 mL syrup		0 sundaes
Change	−240 oz (40 scoops)		−20 cherries		−1000 mL syrup		+20 sundaes
Final	60 oz (10 scoops)		10 cherries		0 mL syrup		20 sundaes

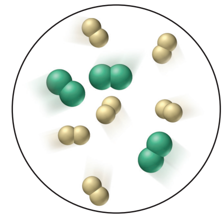
The body of the table shows the following important points:

- In the **Initial** line, "product" has not yet formed, so the entry is "0 sundaes."
- In the **Change** line, since the reactants (ice cream, cherries, and syrup) are used during the reaction, their quantities decrease, so the changes in their quantities have a *negative* sign. At the same time, the quantity of product (sundaes) increases, so the change in its quantity has a *positive* sign.
- For the **Final** line, we *add* the Change and Initial lines. Notice that some reactants (ice cream and cherries) are in excess, while the limiting reactant (syrup) is used up.

Solving Limiting-Reactant Problems In limiting-reactant problems, *the amounts of two (or more) reactants are given, and we first determine which is limiting*. To do this, just as we did with the ice cream sundaes, we use the balanced equation to solve a series of calculations to see how much product forms from the given amount of each reactant: the limiting reactant is the one that yields the *least* amount of product.

The following problems examine these ideas from several aspects. In Sample Problem 3.18, we solve the problem by looking at a molecular scene; in Sample Problem 3.19, we start with the amounts (mol) of two reactants; and in Sample Problem 3.20, we start with masses of two reactants.

SAMPLE PROBLEM 3.18 Using Molecular Depictions in a Limiting-Reactant Problem



Problem Nuclear engineers use chlorine trifluoride to prepare uranium fuel for power plants. The compound is formed as a gas by the reaction of elemental chlorine and fluorine. The circle in the margin shows a representative portion of the reaction mixture before the reaction starts (chlorine is *green;* fluorine is *yellow*).
(a) Find the limiting reactant.
(b) Write a reaction table for the process.
(c) Draw a representative portion of the mixture after the reaction is complete. (*Hint:* The ClF_3 molecule has Cl bonded to three individual F atoms.)

Plan **(a)** We have to find the limiting reactant. The first step is to write the balanced equation, so we need the formulas and states of matter. From the name, chlorine trifluoride, we know the product consists of one Cl atom bonded to three F atoms, or ClF_3. Elemental chlorine and fluorine are the diatomic molecules Cl_2 and F_2, and all three substances are gases. To find the limiting reactant, we find the number of molecules of product that would form from the numbers of molecules of each reactant: whichever forms less product is the limiting reactant. **(b)** We use these numbers of molecules to write a reaction table. **(c)** We use the numbers in the Final line of the table to draw the scene.

Solution **(a)** The balanced equation is

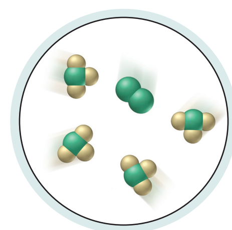
$$\mathrm{Cl_2}(g) + 3\mathrm{F_2}(g) \longrightarrow 2\mathrm{ClF_3}(g)$$

$$\text{For Cl}_2\text{: Molecules of ClF}_3 = 3\ \cancel{\text{molecules of Cl}_2} \times \frac{2\ \text{molecules of ClF}_3}{1\ \cancel{\text{molecule of Cl}_2}}$$
$$= 6\ \text{molecules of ClF}_3$$

$$\text{For F}_2\text{: Molecules of ClF}_3 = 6\ \cancel{\text{molecules of F}_2} \times \frac{2\ \text{molecules of ClF}_3}{3\ \cancel{\text{molecules of F}_2}}$$
$$= \tfrac{12}{3}\ \text{molecules of ClF}_3 = 4\ \text{molecules of ClF}_3$$

Because it forms less product, $\boxed{F_2 \text{ is the limiting reactant.}}$
(b) Since F_2 is the limiting reactant, all of it (6 molecules) is used in the Change line of the reaction table:

Molecules	$Cl_2(g)$	+	$3F_2(g)$	$\longrightarrow$	$2ClF_3(g)$
Initial	3		6		0
Change	−2		−6		+4
Final	1		0		4

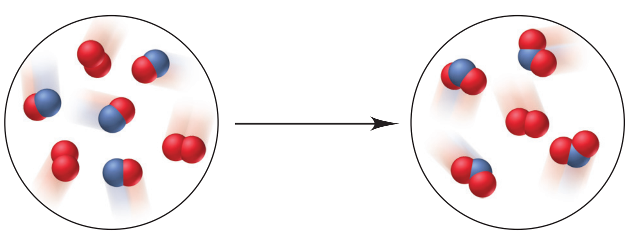


(c) The representative portion of the final reaction mixture *(see margin)* includes 1 molecule of Cl_2 (the reactant in excess) and 4 molecules of product ClF_3.

Check The equation is balanced: reactants (2 Cl, 6 F) $\longrightarrow$ products (2 Cl, 6 F). And, as shown in the circles, the numbers of each type of atom before and after the reaction are equal. Let's think through our choice of limiting reactant. From the equation, one Cl_2 needs three F_2 to form two ClF_3. Therefore, the three Cl_2 molecules in the circle depicting reactants need nine (3 × 3) F_2. But there are only six F_2, so there is not enough F_2 to react with the available Cl_2; or put the other way, there is too much Cl_2 to react with the available F_2. From either point of view, F_2 is the limiting reactant.

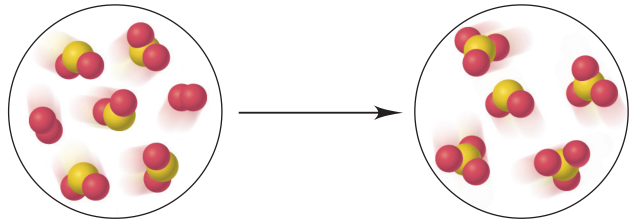
FOLLOW-UP PROBLEMS

3.18A B_2 (B is *red*) reacts with AB as shown below:



Write a balanced equation for the reaction, and determine the limiting reactant.

3.18B Sulfur dioxide gas reacts with oxygen gas to produce sulfur trioxide, as shown below (sulfur is *yellow*; oxygen is *red*):



Write a balanced equation for the reaction, and determine the limiting reactant.
A SIMILAR PROBLEM 3.64

SAMPLE PROBLEM 3.19 Calculating Quantities in a Limiting-Reactant Problem: Amount to Amount

Problem In another preparation of ClF_3 (see Sample Problem 3.18), 0.750 mol of Cl_2 reacts with 3.00 mol of F_2. **(a)** Find the limiting reactant. **(b)** Write a reaction table.

Plan **(a)** We find the limiting reactant by calculating the amount (mol) of ClF_3 formed from the amount (mol) of each reactant: the reactant that forms fewer moles of ClF_3 is limiting. **(b)** We enter those values into the reaction table.

Solution **(a)** Determining the limiting reactant:
Finding amount (mol) of ClF_3 from amount (mol) of Cl_2:

$$\text{Amount (mol) of ClF}_3 = 0.750\ \cancel{\text{mol Cl}_2} \times \frac{2\ \text{mol ClF}_3}{1\ \cancel{\text{mol Cl}_2}} = 1.50\ \text{mol ClF}_3$$

Finding amount (mol) of ClF_3 from amount (mol) of F_2:

$$\text{Amount (mol) of ClF}_3 = 3.00\ \cancel{\text{mol F}_2} \times \frac{2\ \text{mol ClF}_3}{3\ \cancel{\text{mol F}_2}} = 2.00\ \text{mol ClF}_3$$

In this case, Cl_2 is limiting because it forms fewer moles of ClF_3.
(b) Writing the reaction table, with Cl_2 limiting:

Amount (mol)	$Cl_2(g)$	+	$3F_2(g)$	$\longrightarrow$	$2ClF_3(g)$
Initial	0.750		3.00		0
Change	−0.750		−2.25		+1.50
Final	0		0.75		1.50

Check Let's check that Cl_2 is the limiting reactant by assuming, for the moment, that F_2 is limiting. If that were true, all 3.00 mol of F_2 would react to form 2.00 mol of ClF_3. However, based on the balanced equation, obtaining 2.00 mol of ClF_3 would require 1.00 mol of Cl_2, and only 0.750 mol of Cl_2 is present. Thus, Cl_2 must be the limiting reactant.

Comment A major point to note from Sample Problems 3.18 and 3.19 is that the relative amounts of reactants *do not* determine which is limiting, but rather the amount of product formed, which is based on the *molar ratio in the balanced equation*. In both problems, there is more F_2 than Cl_2. However,

- Sample Problem 3.18 has an F_2/Cl_2 ratio of 6/3, or 2/1, which is less than the required molar ratio of 3/1, so F_2 is limiting and Cl_2 is in excess.
- Sample Problem 3.19 has an F_2/Cl_2 ratio of 3.00/0.750, which is greater than the required molar ratio of 3/1, so Cl_2 is limiting and F_2 is in excess.

FOLLOW-UP PROBLEMS

3.19A In the reaction in Follow-up Problem 3.18A, how many moles of product form from 1.5 mol of each reactant?

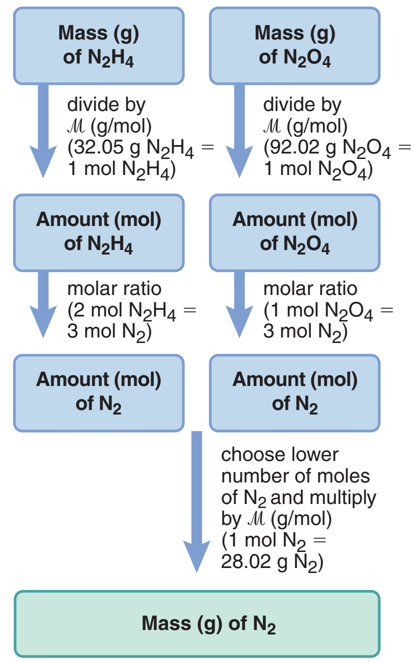
3.19B In the reaction in Follow-up Problem 3.18B, 4.2 mol of SO_2 reacts with 3.6 mol of O_2. How many moles of SO_3 are produced?
A SIMILAR PROBLEM 3.110

SAMPLE PROBLEM 3.20 Calculating Quantities in a Limiting-Reactant Problem: Mass to Mass

Problem A fuel mixture used in the early days of rocketry consisted of two liquids, hydrazine (N_2H_4) and dinitrogen tetroxide (N_2O_4), which ignite on contact to form nitrogen gas and water vapor.
(a) 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?
(b) How many grams of the excess reactant remain unreacted when the reaction is over?
(c) Write a reaction table for this process.

Plan The amounts of two reactants are given, which means this is a limiting-reactant problem. **(a)** To determine the mass of product formed, we must find the limiting reactant by calculating which of the given masses of reactant forms *less* nitrogen gas. As always, we first write the balanced equation. We convert the grams of each reactant to moles using its molar mass and then use the molar ratio from the balanced equation to find the number of moles of N_2 each reactant forms. Next, we convert the lower amount of N_2 to mass (see the road map). **(b)** To determine the mass of excess reactant, we use the molar ratio to calculate the mass of excess reactant that is required to react with the given amount of the limiting reactant. We subtract that mass from the given amount of excess reactant; this difference is the mass of unreacted excess reactant.
(c) We use the values based on the limiting reactant for the reaction table.

Road Map



Solution **(a)** Writing the balanced equation:

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

Finding the amount (mol) of N_2 from the amount (mol) of each reactant

For N_2H_4: Amount (mol) of $N_2H_4 = 1.00\times10^2\ \cancel{\text{g } N_2H_4} \times \dfrac{1 \text{ mol } N_2H_4}{32.05\ \cancel{\text{g } N_2H_4}} = 3.12 \text{ mol } N_2H_4$

$$\text{Amount (mol) of } N_2 = 3.12\ \cancel{\text{mol } N_2H_4} \times \frac{3 \text{ mol } N_2}{2\ \cancel{\text{mol } N_2H_4}} = 4.68 \text{ mol } N_2$$

For N_2O_4: Amount (mol) of $N_2O_4 = 2.00\times10^2\ \cancel{\text{g } N_2O_4} \times \dfrac{1 \text{ mol } N_2O_4}{92.02\ \cancel{\text{g } N_2O_4}} = 2.17 \text{ mol } N_2O_4$

$$\text{Amount (mol) of } N_2 = 2.17\ \cancel{\text{mol } N_2O_4} \times \frac{3 \text{ mol } N_2}{1\ \cancel{\text{mol } N_2O_4}} = 6.51 \text{ mol } N_2$$

Thus, N_2H_4 is the limiting reactant because it yields less N_2.
Converting from amount (mol) of N_2 to mass (g):

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

(b) Finding the mass (g) of N_2O_4 that reacts with 1.00×10^2 g of N_2H_4:

$$\text{Mass (g) of } N_2O_4 = 1.00\times10^2\ \cancel{\text{g } N_2H_4} \times \frac{1\ \cancel{\text{mol } N_2H_4}}{32.05\ \cancel{\text{g } N_2H_4}} \times \frac{1\ \cancel{\text{mol } N_2O_4}}{2\ \cancel{\text{mol } N_2H_4}} \times \frac{92.02 \text{ g } N_2O_4}{1\ \cancel{\text{mol } N_2O_4}}$$
$$= 144 \text{ g } N_2O_4$$

$$\text{Mass (g) of } N_2O_4 \text{ in excess} = \text{initial mass of } N_2O_4 - \text{mass of } N_2O_4 \text{ reacted}$$
$$= 2.00\times10^2 \text{ g } N_2O_4 - 144 \text{ g } N_2O_4 = \boxed{56 \text{ g } N_2O_4}$$

(c) With N_2H_4 as the limiting reactant, the reaction table is

Amount (mol)	$2N_2H_4(l)$	+	$N_2O_4(l)$	$\longrightarrow$	$3N_2(g)$	+	$4H_2O(g)$
Initial	3.12		2.17		0		0
Change	−3.12		−1.56		+4.68		+6.24
Final	0		0.61		4.68		6.24

Check There are more grams of N_2O_4 than N_2H_4, but there are fewer moles of N_2O_4 because its $\mathcal{M}$ is much higher. Rounding 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.** Recall this *common mistake* in solving limiting-reactant problems: The limiting reactant is not the *reactant* present in fewer moles (or grams). Rather, it is the reactant that forms fewer moles (or grams) of *product.*

2. An *alternative approach* to finding the limiting reactant compares "How much is needed?" with "How much is given?" That is, based on the balanced equation,

- Find the amount (mol) of each reactant needed to react with the other reactant.
- Compare that *needed* amount with the *given* amount in the problem statement. There will be *more* than enough of one reactant (excess) and *less* than enough of the other (limiting).

For example, the balanced equation for this problem shows that 2 mol of N_2H_4 reacts with 1 mol of N_2O_4. The amount (mol) of N_2O_4 needed to react with the given 3.12 mol of N_2H_4 is

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

The amount of N_2H_4 needed to react with the given 2.17 mol of N_2O_4 is

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

We are given 2.17 mol of N_2O_4, which is *more* than the 1.56 mol of N_2O_4 needed, and we are given 3.12 mol of N_2H_4, which is *less* than the 4.34 mol of N_2H_4 needed. Therefore, N_2H_4 is limiting, and N_2O_4 is in excess.

FOLLOW-UP PROBLEMS

3.20A 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 many grams of the nonlimiting reactant are in excess?

3.20B Butane gas (C_4H_{10}) is used as the fuel in disposable lighters. It burns in oxygen to form carbon dioxide gas and water vapor. What mass of carbon dioxide is produced when 4.65 g of butane is burned in 10.0 g of oxygen? How many grams of the excess reactant remain unreacted when the reaction is over?

SOME SIMILAR PROBLEMS 3.78–3.83

Figure 3.11 provides an overview of all of the stoichiometric relationships we've discussed in this chapter.

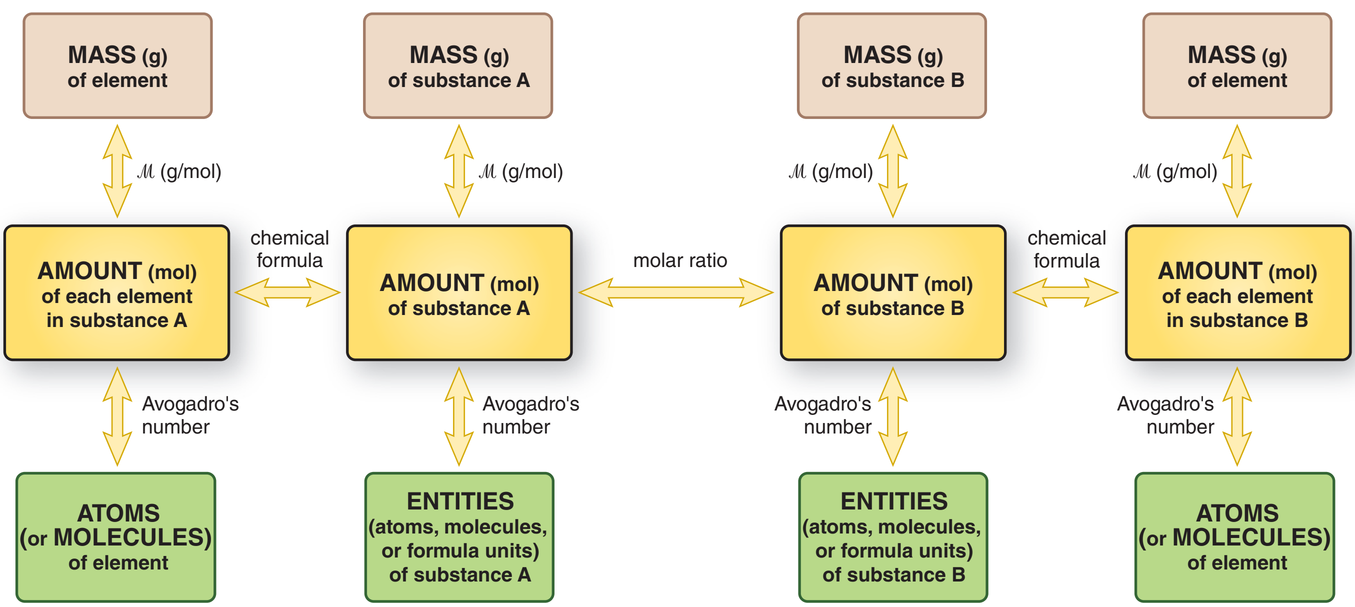


Figure 3.11 An overview of amount-mass-number stoichiometric relationships.

Theoretical, Actual, and Percent Reaction Yields

Up until now, we've assumed that 100% of the limiting reactant becomes product, that ideal methods exist for isolating the product, and that we have perfect lab technique and collect all the product. In theory, this may happen, but in reality, it doesn't, and chemists recognize three types of reaction yield:

1. *Theoretical yield.* The amount of product calculated from the molar ratio in the balanced equation is the **theoretical yield.** But, there are several reasons why the theoretical yield is *never* obtained:
   - Reactant mixtures often proceed through **side reactions** that form different products (Figure 3.12). In the rocket fuel reaction in Sample Problem 3.20, for example, the reactants might form some NO in the following side reaction:
     $$N_2H_4(l) + 2N_2O_4(l) \longrightarrow 6NO(g) + 2H_2O(g)$$
     This reaction decreases the amounts of reactants available for N_2 production.
   - Even more important, many reactions seem to stop before they are complete, so some limiting reactant is unused. (We'll see why in Chapter 4.)
   - Physical losses occur in every step of a separation (see Tools of the Laboratory, Section 2.9): some solid clings to filter paper, some distillate evaporates, and so forth. With careful technique, you can minimize, but never eliminate, such losses.
2. *Actual yield.* Given these reasons for obtaining less than the theoretical yield, the amount of product actually obtained is the **actual yield.** Theoretical and actual yields are expressed in units of amount (moles) or mass (grams).
3. *Percent yield.* The **percent yield (% yield)** is the actual yield expressed as a percentage of the theoretical yield:
   $$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \qquad \textbf{(3.8)}$$
   By definition, the actual yield is less than the theoretical yield, so the percent yield is *always* less than 100%.

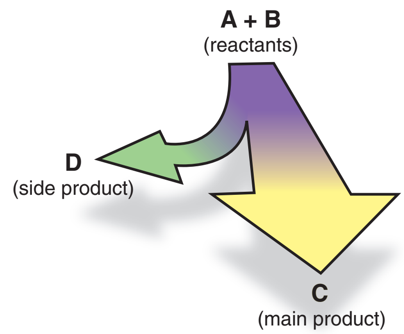
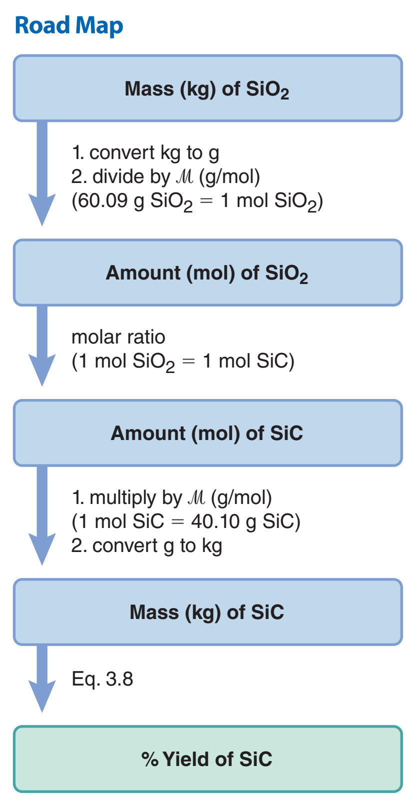


Figure 3.12 The effect of a side reaction on the yield of the main product.

SAMPLE PROBLEM 3.21 Calculating Percent Yield

Problem Silicon carbide (SiC) is an important ceramic material made by reacting sand (silicon dioxide, SiO_2) with powdered carbon at a 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 it to mass (kg) to obtain the theoretical yield. Then, we use Equation 3.8 to find the percent yield (see the road map).



Solution Writing the balanced equation:

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

Converting from mass (kg) of SiO_2 to amount (mol):

$$\text{Amount (mol) of } SiO_2 = 100.0\ \cancel{\text{kg } SiO_2} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1 \text{ mol } SiO_2}{60.09\ \cancel{\text{g } SiO_2}} = 1664 \text{ mol } SiO_2$$

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

$$\text{Amount (mol) of } SiO_2 = \text{moles of SiC} = 1664 \text{ mol SiC}$$

Converting from amount (mol) of SiC to mass (kg):

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

Calculating the percent yield:

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

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

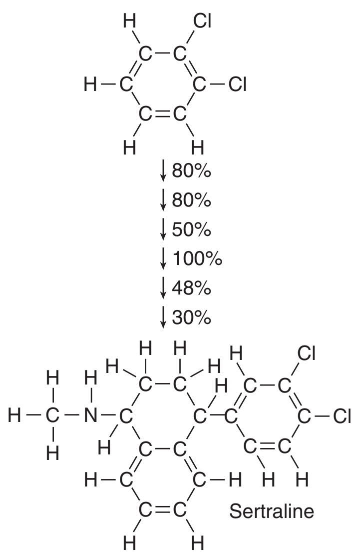
FOLLOW-UP PROBLEMS

3.21A Marble (calcium carbonate) reacts with hydrochloric acid solution to form calcium chloride solution, water, and carbon dioxide. Find the percent yield of carbon dioxide if 3.65 g is collected when 10.0 g of marble reacts.

3.21B Sodium carbonate, also known as *soda ash,* is used in glassmaking. It is obtained from a reaction between sodium chloride and calcium carbonate; calcium chloride is the other product. Calculate the percent yield of sodium carbonate if 92.6 g is collected when 112 g of sodium chloride reacts with excess calcium carbonate.

SOME SIMILAR PROBLEMS 3.88–3.91

Starting with 100 g of 1,2-dichlorobenzene ...



... the yield of Sertraline is only 4.6 g.

Figure 3.13 Low overall yield in a multistep synthesis.

Yields in Multistep Syntheses In the multistep synthesis of a complex compound, the overall yield can be surprisingly low, even if the yield of each step is high. For example, suppose a six-step synthesis has a 90.0% yield for each step. To find the overall percent yield, *express the yield of each step as a decimal, multiply all the decimal amounts together, and then convert back to a percentage*. The overall recovery is only slightly more than 50%:

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

Such multistep sequences are common in laboratory syntheses of medicines, dyes, pesticides, and many other organic compounds. For example, the antidepressant Sertraline is prepared from a simple starting compound in six steps, with yields of 80%, 80%, 50%, 100%, 48%, and 30%, respectively, and an overall percent yield of only 4.6% (Figure 3.13). Because a typical synthesis begins with large amounts of inexpensive, simple reactants and ends with small amounts of expensive, complex products, the overall yield greatly influences the commercial potential of a product.

Atom Economy: A Green Chemistry Perspective on Yield In the relatively new field of **green chemistry,** academic, industrial, and government chemists develop methods that reduce or prevent the release of harmful substances into the environment and the wasting of energy resources.

One way that green chemists evaluate a synthetic route is to focus on its *atom economy,* the proportion of reactant atoms that end up in the desired product. The efficiency of a synthesis is quantified in terms of the *percent atom economy:*

$$\% \text{ atom economy} = \frac{\text{no. of moles} \times \text{molar mass of desired product}}{\text{sum of (no. of moles} \times \text{molar mass) for all products}} \times 100$$

Consider two synthetic routes—one starting with benzene (C_6H_6), the other with butane (C_4H_{10})—for the production of maleic anhydride ($C_4H_2O_3$), a key substance in the manufacture of polymers, dyes, medicines, pesticides, and other products:

Route 1. $2C_6H_6(l) + 9O_2(g) \rightarrow \cdots \rightarrow 2C_4H_2O_3(l) + 4H_2O(l) + 4CO_2(g)$

Route 2. $2C_4H_{10}(g) + 7O_2(g) \rightarrow \cdots \rightarrow 2C_4H_2O_3(l) + 8H_2O(l)$

Let's compare the efficiency of these routes in terms of percent atom economy:

Route 1:

$$\begin{aligned}\% \text{ atom economy} &= \frac{2 \times \mathcal{M} \text{ of } C_4H_2O_3}{(2 \times \mathcal{M} \text{ of } C_4H_2O_3) + (4 \times \mathcal{M} \text{ of } H_2O) + (4 \times \mathcal{M} \text{ of } CO_2)} \times 100 \\ &= \frac{2 \times 98.06 \cancel{\text{ g}}}{(2 \times 98.06 \cancel{\text{ g}}) + (4 \times 18.02 \cancel{\text{ g}}) + (4 \times 44.01 \cancel{\text{ g}})} \times 100 \\ &= 44.15\%\end{aligned}$$

Route 2:

$$\% \text{ atom economy} = \frac{2 \times \mathcal{M} \text{ of } C_4H_2O_3}{(2 \times \mathcal{M} \text{ of } C_4H_2O_3) + (8 \times \mathcal{M} \text{ of } H_2O)} \times 100$$

$$= \frac{2 \times 98.06 \cancel{g}}{(2 \times 98.06 \cancel{g}) + (8 \times 18.02 \cancel{g})} \times 100$$

$$= 57.63\%$$

From the perspective of atom economy, route 2 is preferable because a larger percentage of reactant atoms end up in the desired product. It is also a "greener" approach than route 1 because it avoids the use of the toxic reactant benzene and does not produce CO_2, a gas that contributes to global warming.

Summary of Section 3.4

- The substances in a balanced equation are related to each other by stoichiometrically equivalent molar ratios, which are used as conversion factors to find the amount (mol) of one substance given the amount of another.
- In limiting-reactant problems, the quantities of two (or more) reactants are given, and the limiting reactant is the one that forms the lower quantity of product. Reaction tables show the initial and final quantities of all reactants and products, as well as the changes in those quantities.
- In practice, side reactions, incomplete reactions, and physical losses result in an actual yield of product that is less than the theoretical yield (the quantity based on the molar ratio from the balanced equation), giving a percent yield less than 100%. In multistep reaction sequences, the overall yield is found by multiplying the yields for each step.
- Atom economy, or the proportion of reactant atoms found in the product, is one criterion for choosing a "greener" reaction process.

CHAPTER REVIEW GUIDE

Learning Objectives

Relevant section (§) and/or sample problem (SP) numbers appear in parentheses.

Understand These Concepts

1. The definition of the mole unit (§3.1)
2. Relation between the mass of a chemical entity (in amu) and the mass of a mole of that entity (in g) (§3.1)
3. The relations among amount of substance (in mol), mass (in g), and number of chemical entities (§3.1)
4. Mole-mass-number information in a chemical formula (§3.1)
5. The difference between empirical and molecular formulas of a compound (§3.2)
6. How more than one substance can have the same empirical formula and the same molecular formula (isomers) (§3.2)
7. The importance of balancing equations for the quantitative study of chemical reactions (§3.3)
8. Mole-mass-number information in a balanced equation (§3.4)
9. The relation between amounts of reactants and amounts of products (§3.4)
10. Why one reactant limits the amount of product (§3.4)
11. The causes of lower-than-expected yields and the distinction between theoretical and actual yields (§3.4)

Master These Skills

1. Calculating the molar mass of any substance (§3.1; SPs 3.4–3.6)
2. Converting between amount of substance (in moles), mass (in grams), and number of chemical entities (SPs 3.1–3.5)
3. Using mass percent to find the mass of an element in a given mass of compound (SPs 3.6, 3.7)
4. Determining empirical and molecular formulas of a compound from mass percents and molar masses of elements (SPs 3.8–3.10)
5. Determining a molecular formula from combustion analysis (SP 3.11)
6. Converting a chemical statement or a molecular depiction into a balanced equation (SPs 3.12, 3.13)
7. Using stoichiometrically equivalent molar ratios to convert between amounts of reactants and products in reactions (SPs 3.14–3.16)
8. Writing an overall equation from a series of equations (SP 3.17)
9. Solving limiting-reactant problems for reactions (SPs 3.18–3.20)
10. Calculating percent yield (SP 3.21)

Key Terms

Page numbers appear in parentheses.

stoichiometry (91)

Section 3.1

mole (mol) (92)

Avogadro's number (92)

molar mass ($\mathcal{M}$) (92)

Section 3.2

empirical formula (100)

molecular formula (101)

structural formula (101)

combustion analysis (104)

isomer (106)

Section 3.3

chemical equation (107)

reactant (108)

product (108)

balancing (stoichiometric) coefficient (108)

Section 3.4

overall (net) equation (116)

limiting reactant (117)

theoretical yield (123)

side reaction (123)

actual yield (123)

percent yield (% yield) (123)

green chemistry (124)

Key Equations and Relationships

Page numbers appear in parentheses.

3.1 Number of entities in one mole (92):

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

3.2 Converting amount (mol) to mass (g) using $\mathcal{M}$ (93):

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

3.3 Converting mass (g) to amount (mol) using $1/\mathcal{M}$ (94):

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

3.4 Converting amount (mol) to number of entities (94):

$$\text{No. of entities} = \text{amount (mol)} \times \frac{6.022\times10^{23}\text{ entities}}{1\text{ mol}}$$

3.5 Converting number of entities to amount (mol) (94):

$$\text{Amount (mol)} = \text{no. of entities} \times \frac{1\text{ mol}}{6.022\times10^{23}\text{ entities}}$$

3.6 Calculating mass % (98):

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 Finding the mass of an element in any mass of compound (99):

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

3.8 Calculating percent yield (123):

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

BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS

3.1A $\text{Amount (mol) of C} = 315\ \cancel{\text{mg C}} \times \frac{1\ \cancel{\text{g}}}{10^3\ \cancel{\text{mg}}} \times \frac{1\text{ mol C}}{12.01\ \cancel{\text{g C}}}$

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

See Road Map 3.1A.

3.1B $\text{Number of cans} = 52\ \cancel{\text{mol Al}} \times \frac{26.98\ \cancel{\text{g Al}}}{1\ \cancel{\text{mol Al}}} \times \frac{1\text{ can}}{14\ \cancel{\text{g Al}}}$

$= 100\text{ cans}$

See Road Map 3.1B.

Road Map 3.1A

$\boxed{\text{Mass (mg) of C}}$

↓ 10^3 mg = 1 g

$\boxed{\text{Mass (g) of C}}$

↓ divide by $\mathcal{M}$ (g/mol) (12.01 g = 1 mol C)

$\boxed{\text{Amount (mol) of C}}$

Road Map 3.1B

$\boxed{\text{Amount (mol) of Al}}$

↓ multiply by $\mathcal{M}$ (g/mol) (1 mol Al = 26.98 g Al)

$\boxed{\text{Mass (g) of Al}}$

↓ 14 g Al = 1 soda can

$\boxed{\text{No. of cans}}$

3.2A $\text{Amount (mol) of N} = 9.72\times10^{21}\ \cancel{\text{N}_2\text{ molecules}} \times \frac{1\ \cancel{\text{mol N}_2}}{6.022\times10^{23}\ \cancel{\text{N}_2\text{ molecules}}} \times \frac{2\text{ mol N}}{1\ \cancel{\text{mol N}_2}}$

$= 3.23\times10^{-2}\text{ mol N}$

See Road Map 3.2A.

3.2B $\text{No. of He atoms} = 325\ \cancel{\text{mol He}} \times \frac{6.022\times10^{23}\text{ He atoms}}{1\ \cancel{\text{mol He}}}$

$= 1.96\times10^{26}\text{ He atoms}$

See Road Map 3.2B.

Road Map 3.2A

$\boxed{\text{No. of N}_2\text{ molecules}}$

↓ divide by Avogadro's number (6.022×10^{23} N_2 molecules = 1 mol N_2)

$\boxed{\text{Amount (mol) of N}_2}$

↓ use chemical formula (1 mol N_2 = 2 mol N)

$\boxed{\text{Amount (mol) of N}}$

Road Map 3.2B

$\boxed{\text{Amount (mol) of He}}$

↓ multiply by Avogadro's number (1 mol He = 6.022×10^{23} He atoms)

$\boxed{\text{No. of He atoms}}$

3.3A $\text{Mass (g) of Mn} = 3.22\times10^{20}\ \cancel{\text{Mn atoms}} \times \frac{1\ \cancel{\text{mol Mn}}}{6.022\times10^{23}\ \cancel{\text{Mn atoms}}} \times \frac{54.94\text{ g Mn}}{1\ \cancel{\text{mol Mn}}}$

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

See Road Map 3.3A.

3.3B No. of Cu atoms $= 0.0625\ \cancel{\text{g Cu}} \times \dfrac{1\ \cancel{\text{mol Cu}}}{63.55\ \cancel{\text{g Cu}}} \times \dfrac{6.022\times10^{23}\ \text{Cu atoms}}{1\ \cancel{\text{mol Cu}}}$

$= 5.92\times10^{20}$ Cu atoms

See Road Map 3.3B.

Road Map 3.3A

No. of Mn atoms

↓ divide by Avogadro's number (6.022×10^{23} Mn atoms = 1 mol Mn)

Amount (mol) of Mn

↓ multiply by $\mathcal{M}$ (g/mol) (1 mol Mn = 54.94 g Mn)

Mass (g) of Mn

Road Map 3.3B

Mass (g) of Cu

↓ divide by $\mathcal{M}$ (g/mol) (63.55 g Cu = 1 mol Cu)

Amount (mol) of Cu

↓ multiply by Avogadro's number (1 mol Cu = 6.022×10^{23} Cu atoms)

No. of Cu atoms

3.4A $\mathcal{M} = (1 \times \mathcal{M}\text{ of Na}) + (1 \times \mathcal{M}\text{ of F})$

$= 22.99\ \text{g/mol} + 19.00\ \text{g/mol} = 41.99\ \text{g/mol}$

Mass (g) of NaF

$= 1.19\times10^{19}\ \cancel{\text{NaF formula units}} \times \dfrac{1\ \cancel{\text{mol NaF}}}{6.022\times10^{23}\ \cancel{\text{NaF formula units}}} \times \dfrac{41.99\ \text{g NaF}}{1\ \cancel{\text{mol NaF}}}$

$= 8.30\times10^{-4}$ g NaF

See Road Map 3.4A.

3.4B $\mathcal{M} = (1 \times \mathcal{M}\text{ of Ca}) + (2 \times \mathcal{M}\text{ of Cl})$

$= 40.08\ \text{g/mol} + (2 \times 35.45\ \text{g/mol}) = 110.98\ \text{g/mol}$

No. of $CaCl_2$ formula units $= 400\ \cancel{\text{lb CaCl}_2} \times \dfrac{453.6\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}} \times \dfrac{1\ \text{mol CaCl}_2}{110.98\ \cancel{\text{g CaCl}_2}} \times \dfrac{6.022\times10^{23}\ \text{CaCl}_2\ \text{formula units}}{1\ \text{mol CaCl}_2}$

$= 1\times10^{27}$ $CaCl_2$ formula units

See Road Map 3.4B.

Road Map 3.4A

No. of formula units of NaF

↓ divide by Avogadro's number (6.022×10^{23} formula units = 1 mol NaF)

Amount (mol) of NaF

↓ multiply by $\mathcal{M}$ (g/mol) (1 mol NaF = 41.99 g NaF)

Mass (g) of NaF

Road Map 3.4B

Mass (lb) of $CaCl_2$

↓ 1 lb = 453.6 g

Mass (g) of $CaCl_2$

↓ divide by $\mathcal{M}$ (g/mol) (110.98 g $CaCl_2$ = 1 mol $CaCl_2$)

Amount (mol) of $CaCl_2$

↓ multiply by Avogadro's number (1 mol $CaCl_2$ = 6.022×10^{23} $CaCl_2$ formula units)

No. of formula units of $CaCl_2$

3.5A (a) Mass (g) of P_4O_{10}

$= 4.65\times10^{22}\ \cancel{\text{molecules P}_4\text{O}_{10}} \times \dfrac{1\ \cancel{\text{mol P}_4\text{O}_{10}}}{6.022\times10^{23}\ \cancel{\text{molecules P}_4\text{O}_{10}}} \times \dfrac{283.88\ \text{g P}_4\text{O}_{10}}{1\ \cancel{\text{mol P}_4\text{O}_{10}}}$

$= 21.9$ g P_4O_{10}

(b) No. of P atoms $= 4.65\times10^{22}\ \cancel{\text{molecules P}_4\text{O}_{10}} \times \dfrac{4\ \text{atoms P}}{1\ \cancel{\text{molecule P}_4\text{O}_{10}}}$

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

3.5B (a) No. of formula units of $Ca_3(PO_4)_2$

$= 75.5\ \cancel{\text{g Ca}_3(\text{PO}_4)_2} \times \dfrac{1\ \cancel{\text{mol Ca}_3(\text{PO}_4)_2}}{310.18\ \cancel{\text{g Ca}_3(\text{PO}_4)_2}} \times \dfrac{6.022\times10^{23}\ \text{formula units Ca}_3(\text{PO}_4)_2}{1\ \cancel{\text{mol Ca}_3(\text{PO}_4)_2}}$

$= 1.47\times10^{23}$ formula units $Ca_3(PO_4)_2$

(b) No. of PO_4^{3-} ions $= 1.47\times10^{23}\ \cancel{\text{formula units Ca}_3(\text{PO}_4)_2} \times \dfrac{2\ \text{PO}_4^{3-}\ \text{ions}}{1\ \cancel{\text{formula unit Ca}_3(\text{PO}_4)_2}}$

$= 2.94\times10^{23}$ PO_4^{3-} ions

3.6A Mass % of C $= \dfrac{6\ \cancel{\text{mol C}} \times \dfrac{12.01\ \text{g C}}{1\ \cancel{\text{mol C}}}}{180.16\ \text{g C}_6\text{H}_{12}\text{O}_6} \times 100$

$= 40.00$ mass % C

3.6B Mass % of Cl $= \dfrac{3\ \cancel{\text{mol Cl}} \times \dfrac{35.45\ \text{g Cl}}{1\ \cancel{\text{mol Cl}}}}{137.36\ \text{g CCl}_3\text{F}} \times 100$

$= 77.42$ mass % Cl

3.7A Mass (g) of C $= 16.55\ \cancel{\text{g C}_6\text{H}_{12}\text{O}_6} \times \dfrac{72.06\ \text{g C}}{180.16\ \cancel{\text{g C}_6\text{H}_{12}\text{O}_6}}$

$= 6.620$ g C

3.7B Mass (g) of Cl $= 112\ \cancel{\text{g CCl}_3\text{F}} \times \dfrac{106.35\ \text{g Cl}}{137.36\ \cancel{\text{g CCl}_3\text{F}}}$

$= 86.7$ g Cl

3.8A Preliminary formula: $B_{0.170}O_{0.255}$

Divide by smaller subscript: $B_{\frac{0.170}{0.170}}O_{\frac{0.255}{0.170}} = B_{1.00}O_{1.50}$

Multiply by 2: $B_{2\times1.00}O_{2\times1.50} = B_{2.00}O_{3.00} = B_2O_3$

3.8B Preliminary formula: $C_{6.80}H_{18.1}$

Divide by smallest subscript: $C_{\frac{6.80}{6.80}}O_{\frac{18.1}{6.80}} = C_{1.00}H_{2.66}$

Multiply by 3: $C_{3\times1.00}H_{3\times2.66} = C_{3.00}H_{7.98} = C_3H_8$

3.9A Amount (mol) of H $= 1.23\ \cancel{\text{g H}} \times \dfrac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}}$

$= 1.22$ mol H

Similarly, there are 0.408 mol P and 1.63 mol O.

Preliminary formula: $H_{1.22}P_{0.408}O_{1.63}$

Divide by smallest subscript:

$$H_{\frac{1.22}{0.408}}P_{\frac{0.408}{0.408}}O_{\frac{1.63}{0.408}} = H_{2.99}P_{1.00}O_{4.00} = H_3PO_4$$

3.9B Amount (mol) of S $= 2.88\ \cancel{\text{g S}} \times \dfrac{1\ \text{mol S}}{32.06\ \cancel{\text{g S}}}$

$= 0.0898$ mol S

Amount (mol) of M $= 0.0898\ \cancel{\text{mol S}} \times \dfrac{2\ \text{mol M}}{3\ \cancel{\text{mol S}}} = 0.0599$ 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.

BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS (continued)

3.10A Assuming 100.00 g of compound, we have 95.21 g of C and 4.79 g of H:

$$\text{Amount (mol) of C} = 95.21\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 7.928\ \text{mol C}$$

Similarly, there is 4.75 mol H.
Preliminary formula: $C_{7.928}H_{4.75} \approx C_{1.67}H_{1.00}$
Empirical formula: C_5H_3

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

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

3.10B Assuming 100.00 g of compound, we have 49.47 g C, 5.19 g H, 28.86 g N, and 16.48 g O:

$$\text{Amount (mol) of C} = 49.47\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 4.119\ \text{mol C}$$

Similarly, there are 5.15 mol H, 2.060 mol N, and 1.030 mol O.
Preliminary formula: $C_{4.119}H_{5.15}N_{2.060}O_{1.030} = C_{4.00}H_{5.00}N_{2.00}O_{1.00}$
Empirical formula: $C_4H_5N_2O$

$$\text{Whole-number multiple} = \frac{194.2\ \cancel{\text{g/mol}}}{97.10\ \cancel{\text{g/mol}}} = 2$$

Molecular formula: $C_8H_{10}N_4O_2$

3.11A $\text{Mass (g) of C} = 0.451\ \cancel{\text{g } CO_2} \times \dfrac{12.01\ \text{g C}}{44.01\ \cancel{\text{g } CO_2}}$

$$= 0.123\ \text{g C}$$

Similarly, there is 0.00690 g H.
Mass (g) of Cl = 0.250 g − (0.123 g + 0.00690 g) = 0.120 g Cl
Amount (mol) of elements: 0.0102 mol C; 0.00685 mol H; 0.00339 mol Cl
Empirical formula: C_3H_2Cl
Whole-number multiple = 2
Molecular formula: $C_6H_4Cl_2$

3.11B $\text{Mass (g) of C} = 3.516\ \cancel{\text{g } CO_2} \times \dfrac{12.01\ \text{g C}}{44.01\ \cancel{\text{g } CO_2}}$

$$= 0.9595\ \text{g C}$$

Similarly, there is 0.1127 g H.
Mass (g) of O = 1.200 g − (0.9595 g + 0.1127 g) = 0.128 g O
Amount (mol) of elements: 0.07989 mol C; 0.1118 mol H; 0.00800 mol O
Empirical formula: $C_{10}H_{14}O$
Whole-number multiple: 2
Molecular formula: $C_{20}H_{28}O_2$

3.12A (a) $2Na(s) + 2H_2O(l) \longrightarrow H_2(g) + 2NaOH(aq)$
(b) $2HNO_3(aq) + CaCO_3(s) \longrightarrow CO_2(g) + H_2O(l) + Ca(NO_3)_2(aq)$
(c) $PCl_3(g) + 3HF(g) \longrightarrow PF_3(g) + 3HCl(g)$

3.12B (a) $4C_3H_5N_3O_9(l) \longrightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)$
(b) $4KO_2(s) + 2CO_2(g) \longrightarrow 3O_2(g) + 2K_2CO_3(s)$
(c) $Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$

3.13A From the depiction, we have

$$6CO + 3O_2 \longrightarrow 6CO_2$$

Or,

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

3.13B From the depiction, we have

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

Or,

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

3.14A $Fe_2O_3(s) + 2Al(s) \longrightarrow Al_2O_3(s) + 2Fe(l)$

$$\text{Amount (mol) of } Fe_2O_3 = 3.60\times10^3\ \cancel{\text{mol Fe}} \times \frac{1\ \text{mol } Fe_2O_3}{2\ \cancel{\text{mol Fe}}}$$
$$= 1.80\times10^3\ \text{mol } Fe_2O_3$$

See Road Map 3.14A.

3.14B $3Ag_2S(s) + 2Al(s) \longrightarrow 6Ag(s) + Al_2S_3(s)$

$$\text{Amount (mol) of Al} = 0.253\ \cancel{\text{mol } Ag_2S} \times \frac{2\ \text{mol Al}}{3\ \cancel{\text{mol } Ag_2S}}$$
$$= 0.169\ \text{mol Al}$$

See Road Map 3.14B.

Road Map 3.14A

Amount (mol) of Fe
↓ molar ratio (2 mol Fe = 1 mol Fe_2O_3)
Amount (mol) of Fe_2O_3

Road Map 3.14B

Amount (mol) of Ag_2S
↓ molar ratio (3 mol Ag_2S = 2 mol Al)
Amount (mol) of Al

3.15A Amount (mol) of Fe

$$= 1.85\times10^{25}\ \cancel{\text{formula units } Fe_2O_3} \times \frac{1\ \cancel{\text{mol } Fe_2O_3}}{6.022\times10^{23}\ \cancel{\text{formula units } Fe_2O_3}} \times \frac{2\ \text{mol Fe}}{1\ \cancel{\text{mol } Fe_2O_3}}$$
$$= 61.4\ \text{mol Fe}$$

See Road Map 3.15A.

3.15B
$$\text{Amount (mol) of Ag} = 32.6\ \cancel{\text{g } Ag_2S} \times \frac{1\ \text{mol } Ag_2S}{247.9\ \cancel{\text{g } Ag_2S}} \times \frac{6\ \text{mol Ag}}{3\ \text{mol } Ag_2S}$$
$$= 0.263\ \text{mol Ag}$$

See Road Map 3.15B.

Road Map 3.15A

No. of formula units of Fe_2O_3
↓ divide by Avogadro's number (6.022×10^{23} Fe_2O_3 formula units = 1 mol Fe_2O_3)
Amount (mol) of Fe_2O_3
↓ molar ratio (1 mol Fe_2O_3 = 2 mol Fe)
Amount (mol) of Fe

Road Map 3.15B

Mass (g) of Ag_2S
↓ divide by $\mathcal{M}$ (g/mol) (247.9 g Ag_2S = 1 mol Ag_2S)
Amount (mol) of Ag_2S
↓ molar ratio (3 mol Ag_2S = 6 mol Ag)
Amount (mol) of Ag

3.16A

$$\text{No. of Al atoms} = 1.00\ \cancel{\text{g Al}_2\text{O}_3} \times \frac{1\ \cancel{\text{mol Al}_2\text{O}_3}}{101.96\ \cancel{\text{g Al}_2\text{O}_3}} \times \frac{2\ \cancel{\text{mol Al}}}{1\ \cancel{\text{mol Al}_2\text{O}_3}} \times \frac{6.022\times10^{23}\ \text{Al atoms}}{1\ \cancel{\text{mol Al}}}$$
$$= 1.18\times10^{22}\ \text{Al atoms}$$

Mass (g) of Al_2O_3

↓ divide by $\mathcal{M}$ (g/mol) (101.96 g Al_2O_3 = 1 mol Al_2O_3)

Amount (mol) of Al_2O_3

↓ molar ratio (1 mol Al_2O_3 = 2 mol Al)

Amount (mol) of Al

↓ multiply by Avogadro's number (1 mol Al = 6.022×10^{23} Al atoms)

No. of Al atoms

3.16B $$\text{Mass (g) of Al} = 12.1\ \cancel{\text{g Al}_2\text{S}_3} \times \frac{1\ \cancel{\text{mol Al}_2\text{S}_3}}{150.14\ \cancel{\text{g Al}_2\text{S}_3}} \times \frac{2\ \cancel{\text{mol Al}}}{1\ \cancel{\text{mol Al}_2\text{S}_3}} \times \frac{26.98\ \text{g Al}}{1\ \cancel{\text{mol Al}}}$$
$$= 4.35\ \text{g Al}$$

Mass (g) of Al_2S_3

↓ divide by $\mathcal{M}$ (g/mol) (150.14 g Al_2S_3 = 1 mol Al_2S_3)

Amount (mol) of Al_2S_3

↓ molar ratio (1 mol Al_2S_3 = 2 mol Al)

Amount (mol) of Al

↓ multiply by $\mathcal{M}$ (g/mol) (1 mol Al = 26.98 g Al)

Mass (g) of Al

3.17A

$$\begin{array}{l} 2SO_2(g) + O_2(g) \longrightarrow \cancel{2SO_3(g)} \\ \cancel{2SO_3(g)} + 2H_2O(l) \longrightarrow 2H_2SO_4(aq) \\ \hline 2SO_2(g) + O_2(g) + 2H_2O(l) \longrightarrow 2H_2SO_4(aq) \end{array}$$

3.17B

$$\begin{array}{l} N_2(g) + \cancel{O_2(g)} \longrightarrow \cancel{2NO(g)} \\ \cancel{2NO(g)} + 2O_3(g) \longrightarrow 2NO_2(g) + \cancel{2}O_2(g) \\ \hline N_2(g) + 2O_3(g) \longrightarrow 2NO_2(g) + O_2(g) \end{array}$$

3.18A $4AB(g) + 2B_2(g) \longrightarrow 4AB_2(g)$

Or $2AB(g) + B_2(g) \longrightarrow 2AB_2(g)$

$$\text{For AB: Molecules of } AB_2 = 4\cancel{AB} \times \frac{2AB_2}{2\cancel{AB}} = 4AB_2$$

$$\text{For } B_2\text{: Molecules of } AB_2 = 3\cancel{B_2} \times \frac{2AB_2}{1\cancel{B_2}} = 6AB_2$$

Thus, AB is the limiting reactant; one B_2 molecule is in excess.

3.18B $4SO_2(g) + 2O_2(g) \longrightarrow 4SO_3(g)$

Or $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$

$$\text{For } SO_2\text{: Molecules of } SO_3 = 5\cancel{SO_2} \times \frac{2SO_3}{2\cancel{SO_2}} = 5SO_3$$

$$\text{For } O_2\text{: Molecules of } SO_3 = 2\cancel{O_2} \times \frac{2SO_3}{1\cancel{O_2}} = 4SO_3$$

Thus, O_2 is the limiting reactant; one SO_2 molecule is in excess.

3.19A $$\text{Amount (mol) of } AB_2 = 1.5\ \cancel{\text{mol AB}} \times \frac{2\ \text{mol AB}_2}{2\ \cancel{\text{mol AB}}} = 1.5\ \text{mol AB}_2$$

$$\text{Amount (mol) of } AB_2 = 1.5\ \cancel{\text{mol B}_2} \times \frac{2\ \text{mol AB}_2}{1\ \cancel{\text{mol B}_2}} = 3.0\ \text{mol AB}_2$$

Therefore, 1.5 mol of AB_2 can form.

3.19B $$\text{Amount (mol) of } SO_3 = 4.2\ \cancel{\text{mol SO}_2} \times \frac{2\ \text{mol SO}_3}{2\ \cancel{\text{mol SO}_2}} = 4.2\ \text{mol SO}_3$$

$$\text{Amount (mol) of } SO_3 = 3.6\ \cancel{\text{mol O}_2} \times \frac{2\ \text{mol SO}_3}{1\ \cancel{\text{mol O}_2}} = 7.2\ \text{mol SO}_3$$

Therefore, 4.2 mol of SO_3 is produced.

3.20A $2Al(s) + 3S(s) \longrightarrow Al_2S_3(s)$

Mass (g) of Al_2S_3 formed from 10.0 g of Al

$$= 10.0\ \cancel{\text{g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{26.98\ \cancel{\text{g Al}}} \times \frac{1\ \cancel{\text{mol Al}_2\text{S}_3}}{2\ \cancel{\text{mol Al}}} \times \frac{150.14\ \text{g Al}_2\text{S}_3}{1\ \cancel{\text{mol Al}_2\text{S}_3}}$$
$$= 27.8\ \text{g Al}_2\text{S}_3$$

Similarly, mass (g) of Al_2S_3 formed from 15.0 g of S = 23.4 g Al_2S_3. Thus, S is the limiting reactant, and 23.4 g of Al_2S_3 forms.

Mass (g) of Al in excess

$$= \text{total mass of Al} - \text{mass of Al used}$$
$$= 10.0\ \text{g Al} - \left(15.0\ \cancel{\text{g S}} \times \frac{1\ \cancel{\text{mol S}}}{32.06\ \cancel{\text{g S}}} \times \frac{2\ \cancel{\text{mol Al}}}{3\ \cancel{\text{mol S}}} \times \frac{26.98\ \text{g Al}}{1\ \cancel{\text{mol Al}}}\right)$$
$$= 1.6\ \text{g Al}$$

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

3.20B $2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$

Mass (g) of CO_2 formed from 4.65 g of C_4H_{10}

$$= 4.65\ \cancel{\text{g C}_4\text{H}_{10}} \times \frac{1\ \cancel{\text{mol C}_4\text{H}_{10}}}{58.12\ \cancel{\text{g C}_4\text{H}_{10}}} \times \frac{8\ \cancel{\text{mol CO}_2}}{2\ \cancel{\text{mol C}_4\text{H}_{10}}} \times \frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol CO}_2}}$$
$$= 14.1\ \text{g CO}_2$$

Similarly, the mass (g) of CO_2 formed from 10.0 g of O_2 = 8.46 g CO_2.

Thus, O_2 is the limiting reactant, and 8.46 g of CO_2 forms.

Mass (g) of C_4H_{10} in excess

$$= \text{total mass of } C_4H_{10} - \text{mass of } C_4H_{10} \text{ used}$$
$$= 4.65\ \text{g} - \left(10.0\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g O}_2}} \times \frac{2\ \cancel{\text{mol C}_4\text{H}_{10}}}{13\ \cancel{\text{mol O}_2}} \times \frac{58.12\ \text{g C}_4\text{H}_{10}}{1\ \cancel{\text{mol C}_4\text{H}_{10}}}\right)$$
$$= 1.86\ \text{g C}_4\text{H}_{10}$$

BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS (continued)

3.21A $CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

$$\text{Theoretical yield (g) of } CO_2 = 10.0\ \cancel{g\ CaCO_3} \times \frac{1\ \cancel{mol\ CaCO_3}}{100.09\ \cancel{g\ CaCO_3}} \times \frac{1\ \cancel{mol\ CO_2}}{1\ \cancel{mol\ CaCO_3}} \times \frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}} = 4.40\ g\ CO_2$$

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

3.21B $2NaCl(aq) + CaCO_3(s) \longrightarrow Na_2CO_3(s) + CaCl_2(aq)$

Theoretical yield (g) of Na_2CO_3

$$= 112\ \cancel{g\ NaCl} \times \frac{1\ \cancel{mol\ NaCl}}{58.44\ \cancel{g\ NaCl}} \times \frac{1\ \cancel{mol\ Na_2CO_3}}{2\ \cancel{mol\ NaCl}} \times \frac{105.99\ g\ Na_2CO_3}{1\ \cancel{mol\ Na_2CO_3}} = 102\ g\ Na_2CO_3$$

$$\%\ \text{yield} = \frac{92.6\ \cancel{g\ Na_2CO_3}}{102\ \cancel{g\ Na_2CO_3}} \times 100 = 90.8\%$$

PROBLEMS

Problems with **colored** numbers are answered in Appendix E and worked in detail in the Student Solutions Manual. Problem sections match those in the text and give the numbers of relevant sample problems. Most offer Concept Review Questions, Skill-Building Exercises (grouped in pairs covering the same concept), and Problems in Context. The Comprehensive Problems are based on material from any section or previous chapter.

The Mole

(Sample Problems 3.1 to 3.7)

Concept Review Questions

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 3 mol of Al atoms and of 2 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 2 mol of sucrose?

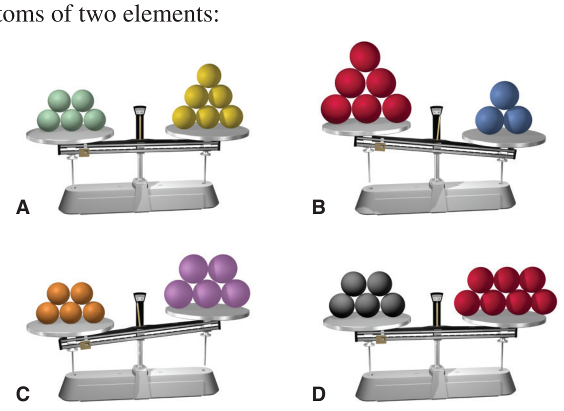
3.3 Why might the expression "1 mol of chlorine" be confusing? What change would remove any uncertainty? For what other elements might a similar confusion exist? Why?

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) for amount of substance rather than a mass unit?

3.6 You need to calculate the number of P_4 molecules that can form from 2.5 g of $Ca_3(PO_4)_2$. Draw a road map for solving this and write a Plan, without doing any calculations.

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



For each balance, 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?

Skill-Building Exercises (grouped in similar pairs)

3.8 Calculate the molar mass of each of the following:
(a) $Sr(OH)_2$ (b) N_2O_3 (c) $NaClO_3$ (d) Cr_2O_3

3.9 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_3

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

3.11 Calculate the molar mass of each of the following:
(a) N_2O_4 (b) C_4H_9OH (c) $MgSO_4{\cdot}7H_2O$ (d) $Ca(C_2H_3O_2)_2$

3.12 Calculate each of the following quantities:
(a) Mass (g) of 0.346 mol of Zn
(b) Number of F atoms in 2.62 mol of F_2
(c) Number of Ca atoms in 28.5 g of Ca

3.13 Calculate each of the following quantities:
(a) Amount (mol) of Mn atoms in 62.0 mg of Mn
(b) Amount (mol) for 1.36×10^{22} atoms of Cu
(c) Mass (g) of 8.05×10^{24} Li atoms

3.14 Calculate each of the following quantities:
(a) Mass (g) of 0.68 mol of $KMnO_4$
(b) Amount (mol) of O atoms in 8.18 g of $Ba(NO_3)_2$
(c) Number of O atoms in 7.3×10^{-3} g of $CaSO_4{\cdot}2H_2O$

3.15 Calculate each of the following quantities:
(a) Mass (kg) of 4.6×10^{21} molecules of NO_2
(b) Amount (mol) of Cl atoms in 0.0615 g of $C_2H_4Cl_2$
(c) Number of H^- ions in 5.82 g of SrH_2

3.16 Calculate each of the following quantities:
(a) Mass (g) of 6.44×10^{-2} mol of $MnSO_4$
(b) Amount (mol) of compound in 15.8 kg of $Fe(ClO_4)_3$
(c) Number of N atoms in 92.6 mg of NH_4NO_2

3.17 Calculate each of the following quantities:
(a) Total number of ions in 38.1 g of SrF_2
(b) Mass (kg) of 3.58 mol of $CuCl_2{\cdot}2H_2O$
(c) Mass (mg) of 2.88×10^{22} formula units of $Bi(NO_3)_3{\cdot}5H_2O$

3.18 Calculate each of the following quantities:
(a) Mass (g) of 8.35 mol of copper(I) carbonate
(b) Mass (g) of 4.04×10^{20} molecules of dinitrogen pentoxide
(c) Amount (mol) and number of formula units in 78.9 g of sodium perchlorate
(d) Number of sodium ions, perchlorate ions, chlorine atoms, and oxygen atoms in the mass of compound in part (c)

3.19 Calculate each of the following quantities:
(a) Mass (g) of 8.42 mol of chromium(III) sulfate decahydrate
(b) Mass (g) of 1.83×10^{24} molecules of dichlorine heptoxide
(c) Amount (mol) and number of formula units in 6.2 g of lithium sulfate
(d) Number of lithium ions, sulfate ions, sulfur atoms, and oxygen atoms in the mass of compound in part (c)

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

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

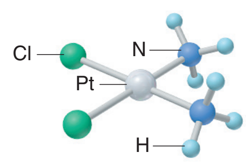
3.22 Calculate each of the following:
(a) Mass fraction of C in cesium acetate
(b) Mass fraction of O in uranyl sulfate trihydrate (the uranyl ion is UO_2^{2+})

3.23 Calculate each of the following:
(a) Mass fraction of Cl in calcium chlorate
(b) Mass fraction of N in dinitrogen trioxide

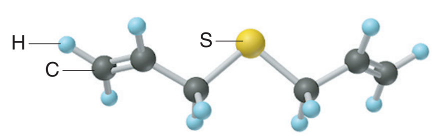
Problems in Context

3.24 Oxygen is required for the metabolic combustion of foods. Calculate the number of atoms in 38.0 g of oxygen gas, the amount absorbed from the lungs in about 15 min when a person is at rest.

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



3.26 Allyl sulfide *(below)* gives garlic its characteristic odor. Calculate (a) the mass (g) of 2.63 mol of allyl sulfide; (b) the number of carbon atoms in 35.7 g of allyl sulfide.



3.27 Iron reacts slowly with oxygen and water to form a compound commonly called rust ($Fe_2O_3 \cdot 4H_2O$). For 45.2 kg of rust, calculate (a) moles of compound; (b) moles of Fe_2O_3; (c) grams of Fe.

3.28 Propane is widely used in liquid form as a fuel for barbecue grills and camp stoves. For 85.5 g of propane, calculate (a) moles of compound; (b) grams of carbon.

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

3.30 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?

3.31 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 3.8 to 3.11)

Concept Review Questions

3.32 What is the difference between an empirical formula and a molecular formula? Can they ever be the same?

3.33 List three ways compositional data may be given in a problem that involves finding an empirical formula.

3.34 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 (draw a road map and write a Plan for a solution).
(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 mass % of each element
(e) Structural formula

3.35 Is $MgCl_2$ an empirical or a molecular formula for magnesium chloride? Explain.

Skill-Building Exercises (grouped in similar pairs)

3.36 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}

3.37 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

3.38 Give the name, empirical formula, and molar mass of the compound depicted in Figure P3.38.

3.39 Give the name, empirical formula, and molar mass of the compound depicted in Figure P3.39.

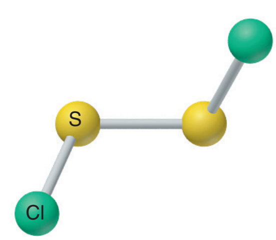


Figure P3.38

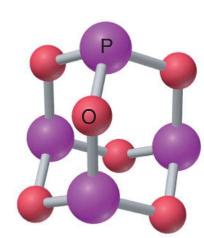


Figure P3.39

3.40 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)

3.41 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)

3.42 Find the empirical formula 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

3.43 Find the empirical formula 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

3.44 An oxide of nitrogen contains 30.45 mass % N. (a) What is the empirical formula of the oxide? (b) If the molar mass is 90±5 g/mol, what is the molecular formula?

3.45 A chloride of silicon contains 79.1 mass % Cl. (a) What is the empirical formula of the chloride? (b) If the molar mass is 269 g/mol, what is the molecular formula?

3.46 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?

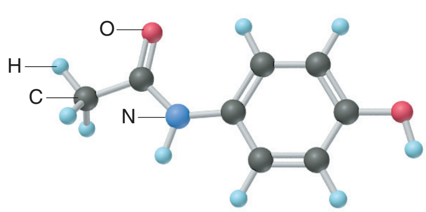
3.47 A 0.370-mol sample of a metal oxide (M_2O_3) weighs 55.4 g.
(a) How many moles of O are in the sample?
(b) How many grams of M are in the sample?
(c) What element is represented by the symbol M?

Problems in Context

3.48 Nicotine is a poisonous, addictive compound found in tobacco. A sample of nicotine contains 6.16 mmol of C, 8.56 mmol of H, and 1.23 mmol of N [1 mmol (1 millimole) = 10^{-3} mol]. What is the empirical formula of nicotine?

3.49 Cortisol ($\mathcal{M}$ = 362.47 g/mol) is a steroid hormone involved in protein synthesis. Medically, it has a major use in reducing inflammation from rheumatoid arthritis. Cortisol is 69.6% C, 8.34% H, and 22.1% O by mass. What is its molecular formula?

3.50 Acetaminophen *(below)* is a popular nonaspirin pain reliever. What is the mass % of each element in acetaminophen?



3.51 Menthol ($\mathcal{M}$ = 156.3 g/mol), the strong-smelling substance in many cough drops, is a compound of carbon, hydrogen, and oxygen. When 0.1595 g of menthol was burned in a combustion apparatus, 0.449 g of CO_2 and 0.184 g of H_2O formed. What is menthol's molecular formula?

Writing and Balancing Chemical Equations

(Sample Problems 3.12 and 3.13)

Concept Review Questions

3.52 What three types of information does a balanced chemical equation provide? How?

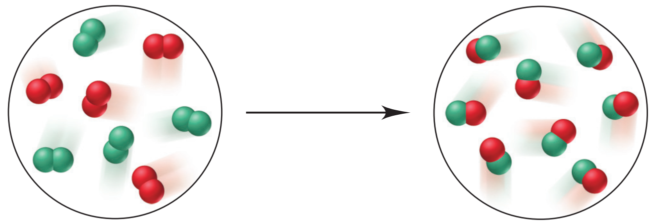
3.53 How does a balanced chemical equation apply the law of conservation of mass?

3.54 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.

3.55 The scenes below represent a chemical reaction between elements A *(red)* and B *(green)*:



Which 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) $4A_2 + 4B_2 \longrightarrow 8AB$

Skill-Building Exercises (grouped in similar pairs)

3.56 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) $\_\_CH_3NH_2(g) + \_\_O_2(g) \longrightarrow \_\_CO_2(g) + \_\_H_2O(g) + \_\_N_2(g)$

3.57 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) $\_\_(CN)_2(g) + \_\_H_2O(l) \longrightarrow \_\_H_2C_2O_4(aq) + \_\_NH_3(g)$

3.58 Write balanced equations for each of the following by inserting the correct coefficients in the blanks:
(a) $\_\_SO_2(g) + \_\_O_2(g) \longrightarrow \_\_SO_3(g)$
(b) $\_\_Sc_2O_3(s) + \_\_H_2O(l) \longrightarrow \_\_Sc(OH)_3(s)$
(c) $\_\_H_3PO_4(aq) + \_\_NaOH(aq) \longrightarrow \_\_Na_2HPO_4(aq) + \_\_H_2O(l)$
(d) $\_\_C_6H_{10}O_5(s) + \_\_O_2(g) \longrightarrow \_\_CO_2(g) + \_\_H_2O(g)$

3.59 Write balanced equations for each of the following by inserting the correct coefficients in the blanks:
(a) $\_\_As_4S_6(s) + \_\_O_2(g) \longrightarrow \_\_As_4O_6(s) + \_\_SO_2(g)$
(b) $\_\_Ca_3(PO_4)_2(s) + \_\_SiO_2(s) + \_\_C(s) \longrightarrow \_\_P_4(g) + \_\_CaSiO_3(l) + \_\_CO(g)$
(c) $\_\_Fe(s) + \_\_H_2O(g) \longrightarrow \_\_Fe_3O_4(s) + \_\_H_2(g)$
(d) $\_\_S_2Cl_2(l) + \_\_NH_3(g) \longrightarrow \_\_S_4N_4(s) + \_\_S_8(s) + \_\_NH_4Cl(s)$

3.60 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.

3.61 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.

Problem in Context

3.62 Loss of atmospheric ozone has led to an ozone "hole" over Antarctica. The loss occurs in part through three consecutive steps:
(1) Chlorine atoms react with ozone (O_3) to form chlorine monoxide and molecular oxygen.
(2) Chlorine monoxide forms ClOOCl.
(3) ClOOCl absorbs sunlight and breaks into chlorine atoms and molecular oxygen.
(a) Write a balanced equation for each step.
(b) Write an overall balanced equation for the sequence.

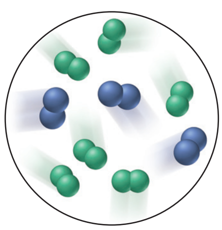
Calculating Quantities of Reactant and Product

(Sample Problems 3.14 to 3.21)

Concept Review Questions

3.63 What does the term *stoichiometrically equivalent molar ratio* mean, and how is it applied in solving problems?

3.64 The scene below represents a mixture of A_2 and B_2 before they react to form AB_3.



(a) What is the limiting reactant?
(b) How many molecules of product can form?

3.65 Percent yields are generally calculated from masses. Would the result be the same if amounts (mol) were used instead? Why?

Skill-Building Exercises (grouped in similar pairs)

3.66 Reactants A and B form product C. Draw a road map and write a Plan to find the mass (g) of C when 25 g of A reacts with excess B.

3.67 Reactants D and E form product F. Draw a road map and write a Plan to find the mass (g) of F when 27 g of D reacts with 31 g of E.

3.68 Chlorine gas can be made in the laboratory by the reaction of hydrochloric acid and manganese(IV) oxide:

$$4HCl(aq) + MnO_2(s) \longrightarrow MnCl_2(aq) + 2H_2O(g) + Cl_2(g)$$

When 1.82 mol of HCl reacts with excess MnO_2, how many (a) moles of Cl_2 and (b) grams of Cl_2 form?

3.69 Bismuth oxide reacts with carbon to form bismuth metal:

$$Bi_2O_3(s) + 3C(s) \longrightarrow 2Bi(s) + 3CO(g)$$

When 283 g of Bi_2O_3 reacts with excess carbon, how many (a) moles of Bi_2O_3 react and (b) moles of Bi form?

3.70 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 56.6 kg of oxygen, how many (a) moles of KNO_3 and (b) grams of KNO_3 must be heated?

3.71 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, how many (a) moles of Cr_2O_3 and (b) grams of Cr_2O_3 are required?

3.72 Calculate the mass (g) of each product formed when 43.82 g of diborane (B_2H_6) reacts with excess water:

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

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

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

3.74 Elemental phosphorus occurs as tetratomic molecules, P_4. What mass (g) of chlorine gas is needed to react completely with 455 g of phosphorus to form phosphorus pentachloride?

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

3.76 Solid iodine trichloride is prepared in two steps: first, a reaction between solid iodine and gaseous chlorine to form solid iodine monochloride; then, treatment with more chlorine.
(a) Write a balanced equation for each step.
(b) Write a balanced equation for the overall reaction.
(c) How many grams of iodine are needed to prepare 2.45 kg of final product?

3.77 Lead can be prepared from galena [lead(II) sulfide] by first roasting the galena in oxygen gas to form lead(II) oxide and sulfur dioxide. Heating the metal oxide with more galena forms the molten metal and more sulfur dioxide.
(a) Write a balanced equation for each step.
(b) Write an overall balanced equation for the process.
(c) How many metric tons of sulfur dioxide form for every metric ton of lead obtained?

3.78 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 (g) of calcium oxide that can be prepared from 4.20 g of Ca and 2.80 g of O_2.
(a) What amount (mol) of CaO can be produced from the given mass of Ca?
(b) What amount (mol) of CaO can be produced from the given mass of O_2?
(c) Which is the limiting reactant?
(d) How many grams of CaO can be produced?

3.79 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 (g) of hydrogen gas that can be prepared from 5.70 g of SrH_2 and 4.75 g of H_2O.

(a) What amount (mol) of H_2 can be produced from the given mass of SrH_2?
(b) What amount (mol) of H_2 can be produced from the given mass of H_2O?
(c) Which is the limiting reactant?
(d) How many grams of H_2 can be produced?

3.80 Calculate the maximum numbers of moles and grams of iodic acid (HIO_3) that can form when 635 g of iodine trichloride reacts with 118.5 g of water:

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

How many grams of the excess reactant remains?

3.81 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:

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

How many grams of the excess reactant remain?

3.82 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?

3.83 A mixture of 0.0375 g of hydrogen and 0.0185 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?

3.84 Aluminum nitrite and ammonium chloride react to form aluminum chloride, nitrogen, and water. How many grams of each substance are present after 72.5 g of aluminum nitrite and 58.6 g of ammonium chloride react completely?

3.85 Calcium nitrate and ammonium fluoride react to form calcium fluoride, dinitrogen monoxide, and water vapor. How many grams of each substance are present after 16.8 g of calcium nitrate and 17.50 g of ammonium fluoride react completely?

3.86 Two successive reactions, $A \longrightarrow B$ and $B \longrightarrow C$, have yields of 73% and 68%, respectively. What is the overall percent yield for conversion of A to C?

3.87 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?

3.88 What is the percent yield of a reaction in which 45.5 g of tungsten(VI) oxide (WO_3) reacts with excess hydrogen gas to produce metallic tungsten and 9.60 mL of water (d = 1.00 g/mL)?

3.89 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)?

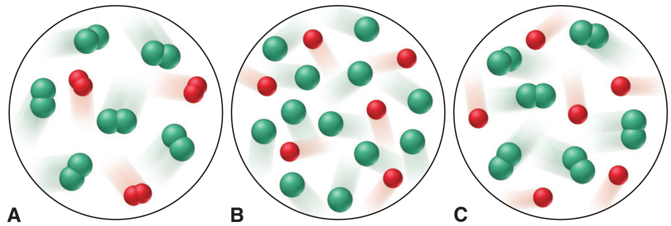
3.90 When 20.5 g of methane and 45.0 g of chlorine gas undergo a reaction that has a 75.0% yield, what mass (g) of chloromethane (CH_3Cl) forms? Hydrogen chloride also forms.

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

Problems in Context

3.92 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 (g) of carbon tetrafluoride forms when 60.0 g of each reactant is used?

3.93 Gaseous dichlorine monoxide decomposes readily to chlorine *(green)* and oxygen *(red)* gases.
(a) Which scene best depicts the product mixture after the decomposition?



(b) Write the balanced equation for the decomposition.
(c) If each oxygen atom represents 0.050 mol, how many molecules of dichlorine monoxide were present before the decomposition?

3.94 An intermediate step in the production of nitric acid involves the reaction of ammonia with oxygen gas to form nitrogen monoxide and water. How many grams of nitrogen monoxide can form in the reaction of 485 g of ammonia with 792 g of oxygen?

3.95 Butane gas is compressed and used as a liquid fuel in disposable cigarette lighters and lightweight camping stoves. Suppose a lighter contains 5.50 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 H_2O form when all the butane burns?
(c) How many total molecules of gas form when the butane burns completely?

3.96 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 an 88.5% yield, how many grams of $NaBH_4$ can be prepared by reacting 7.98 g of sodium hydride and 8.16 g of diborane?

Comprehensive Problems

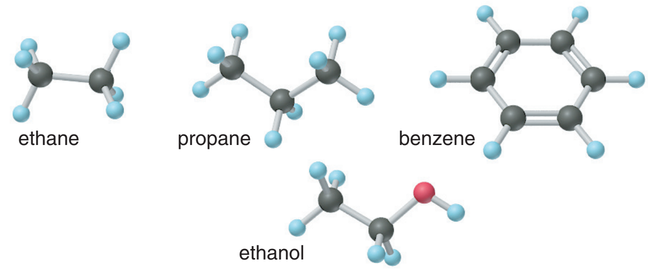
3.97 The mole is defined in terms of the carbon-12 atom. Use the definition to find (a) the mass in grams equal to 1 atomic mass unit; (b) the ratio of the gram to the atomic mass unit.

3.98 The first sulfur-nitrogen compound was prepared in 1835 and has been used to synthesize many others. In the early 1980s, researchers made another such compound that conducts electricity like a metal. Mass spectrometry of the compound shows a molar mass of 184.27 g/mol, and analysis shows it to contain 2.288 g of S for every 1.000 g of N. What is its molecular formula?

3.99 Hydroxyapatite, $Ca_5(PO_4)_3(OH)$, is the main mineral component of dental enamel, dentin, and bone. Coating the compound on metallic implants (such as titanium alloys and stainless steels) helps the body accept the implant. When placed in bone voids, the powder encourages natural bone to grow into the void. Hydroxyapatite is prepared by adding aqueous phosphoric acid to a dilute slurry of calcium hydroxide. (a) Write a balanced equation for this preparation. (b) What mass (g) of hydroxyapatite could form from 100. g of 85% phosphoric acid and 100. g of calcium hydroxide?

3.100 Narceine is a narcotic in opium that crystallizes from solution as a hydrate that contains 10.8 mass % water and has a molar mass of 499.52 g/mol. Determine x in narceine·xH_2O.

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



3.102 Serotonin ($\mathcal{M}$ = 176 g/mol) transmits nerve impulses between neurons. It contains 68.2% C, 6.86% H, 15.9% N, and 9.08% O by mass. What is its molecular formula?

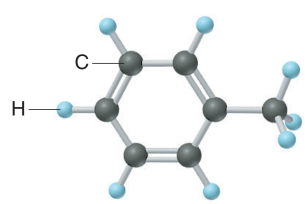
3.103 In 1961, scientists agreed that the atomic mass unit (amu) would be defined as $\frac{1}{12}$ the mass of an atom of ^{12}C. Before then, it was defined as $\frac{1}{16}$ the *average* mass of an atom of naturally occurring oxygen (a mixture of ^{16}O, ^{17}O, and ^{18}O). The current atomic mass of oxygen is 15.9994 amu. (a) Did Avogadro's number change after the definition of an amu changed and, if so, in what direction? (b) Did the definition of the mole change? (c) Did the mass of a mole of a substance change? (d) Before 1961, was Avogadro's number 6.02×10^{23} (to three significant figures), as it is today?

3.104 Convert the following descriptions 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.

3.105 Isobutylene is a hydrocarbon used in the manufacture of synthetic rubber. When 0.847 g of isobutylene was subjected to combustion analysis, 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?

3.106 The multistep smelting of ferric oxide to form elemental iron occurs at high temperatures in a blast furnace. In the first step, ferric oxide reacts with carbon monoxide to form Fe_3O_4. This substance reacts with more carbon monoxide to form iron(II) oxide, which reacts with still more carbon monoxide to form molten iron. Carbon dioxide is also produced in each step. (a) Write an overall balanced equation for the iron-smelting process. (b) How many grams of carbon monoxide are required to form 45.0 metric tons of iron from ferric oxide?

3.107 One of the compounds used to increase the octane rating of gasoline is toluene *(right)*. Suppose 20.0 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?



3.108 During studies of the reaction in Sample Problem 3.20,

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

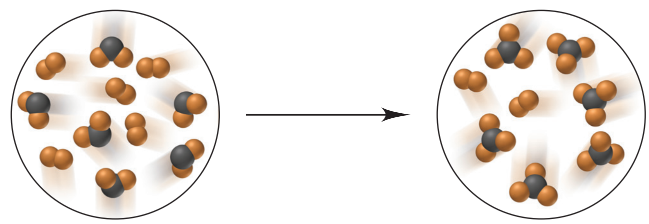
a chemical engineer measured a less-than-expected yield of N_2 and discovered that the following side reaction occurs:

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

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?

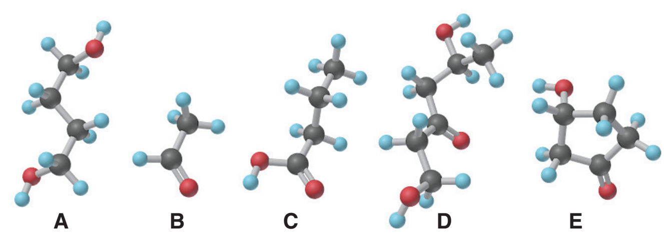
3.109 A 0.652-g sample of a pure strontium halide reacts with excess sulfuric acid. The solid strontium sulfate formed is separated, dried, and found to weigh 0.755 g. What is the formula of the original halide?

3.110 The following scenes represent a chemical reaction between AB_2 and B_2:



(a) Write a balanced equation for the reaction. (b) What is the limiting reactant? (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)?

3.111 Which of the following models represent compounds having the same empirical formula? What is the empirical formula mass of this common formula?



3.112 The zirconium oxalate $K_2Zr(C_2O_4)_3(H_2C_2O_4)\cdot H_2O$ was synthesized by mixing 1.68 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, 1.25 g of crystalline product was obtained, along with aqueous KCl and water. Calculate the percent yield.

3.113 Seawater is approximately 4.0% by mass dissolved ions, 85% of which are from NaCl. (a) Find the mass % of NaCl in seawater. (b) Find the mass % of Na^+ ions and of Cl^- ions in seawater.

3.114 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 being stated when the available quantity of one of the reactants is given in moles.
(d) The empirical and molecular formulas of a compound are always different.

3.115 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) Grams: 6.02×10^{23} atoms of ^{235}U or 6.02×10^{23} atoms of ^{238}U

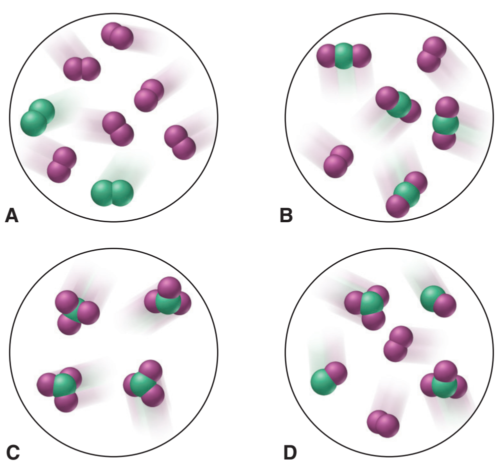
3.116 For the reaction between solid tetraphosphorus trisulfide and oxygen gas to form solid tetraphosphorus decoxide and sulfur dioxide gas, write a balanced equation. Show the equation (see Table 3.4) in terms of (a) molecules, (b) moles, and (c) grams.

3.117 Hydrogen gas is considered a clean fuel because it produces only water vapor when it burns. If the reaction has a 98.8% yield, what mass (g) of hydrogen forms 105 kg of water?

3.118 Solar winds composed of free protons, electrons, and α particles bombard Earth constantly, knocking gas molecules out of the atmosphere. In this way, Earth loses about 3.0 kg of matter per second. It is estimated that the atmosphere will be gone in about 50 billion years. Use this estimate to calculate (a) the mass (kg) of Earth's atmosphere and (b) the amount (mol) of nitrogen, which makes up 75.5 mass % of the atmosphere.

3.119 Calculate each of the following quantities:
(a) Amount (mol) of 0.588 g of ammonium bromide
(b) Number of potassium ions in 88.5 g of potassium nitrate
(c) Mass (g) of 5.85 mol of glycerol ($C_3H_8O_3$)
(d) Volume (L) of 2.85 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 25.0 μg of cadmium
(g) Number of atoms in 0.0015 mol of fluorine gas

3.120 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?



3.121 Hydrocarbon mixtures are used as fuels. (a) 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? (b) A 252-g gaseous mixture of CH_4 and C_3H_8 burns in excess O_2, and 748 g of CO_2 gas is collected. What is the mass % of CH_4 in the mixture?

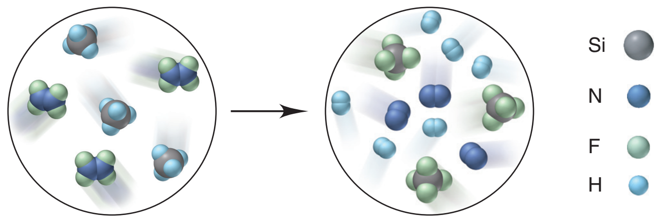
3.122 Nitrogen (N), phosphorus (P), and potassium (K) are the main nutrients in plant fertilizers. By industry convention, the numbers on a label refer to the mass percents of N, P_2O_5, and K_2O, in that order. Calculate the N/P/K ratio of a 30/10/10 fertilizer in terms of moles of each element, and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

3.123 What mass percents of ammonium sulfate, ammonium hydrogen phosphate, and potassium chloride would you use to prepare 10/10/10 plant fertilizer (see Problem 3.122)?

3.124 Ferrocene, 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?

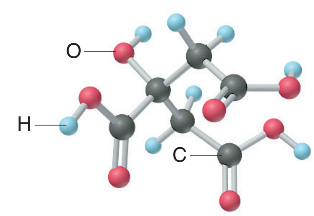
3.125 When carbon-containing compounds are burned in a limited amount of air, some $CO(g)$ as well as $CO_2(g)$ is produced. A gaseous product mixture is 35.0 mass % CO and 65.0 mass % CO_2. What is the mass % of C in the mixture?

3.126 Write a balanced equation for the reaction depicted below:



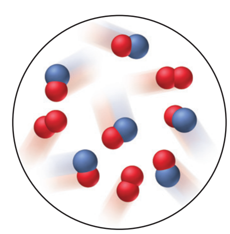
If each reactant molecule represents 1.25×10^{-2} mol and the reaction yield is 87%, how many grams of Si-containing product form?

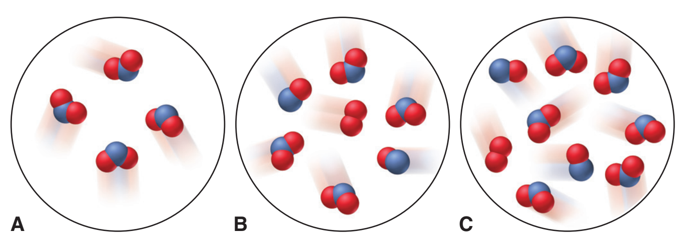
3.127 Citric acid *(right)* is concentrated in citrus fruits and plays a central metabolic role in nearly every animal and plant cell. (a) What are the molar mass and formula of citric acid? (b) How many moles of citric acid are in 1.50 qt of lemon juice ($d = 1.09$ g/mL) that is 6.82% citric acid by mass?



3.128 Various nitrogen oxides, as well as sulfur oxides, contribute to acidic rainfall through complex reaction sequences. Nitrogen and oxygen combine during the high-temperature combustion of fuels in air to form nitrogen monoxide gas, which reacts with more oxygen to form nitrogen dioxide gas. In contact with water vapor, nitrogen dioxide forms aqueous nitric acid and more nitrogen monoxide. (a) Write balanced equations for these reactions. (b) Use the equations to write one overall balanced equation that does *not* include nitrogen monoxide and nitrogen dioxide. (c) How many metric tons (t) of nitric acid form when 1350 t of atmospheric nitrogen is consumed (1 t = 1000 kg)?

3.129 Nitrogen monoxide reacts with elemental oxygen to form nitrogen dioxide. The scene at right represents an initial mixture of reactants. If the reaction has a 66% yield, which of the scenes below (A, B, or C) best represents the final product mixture?



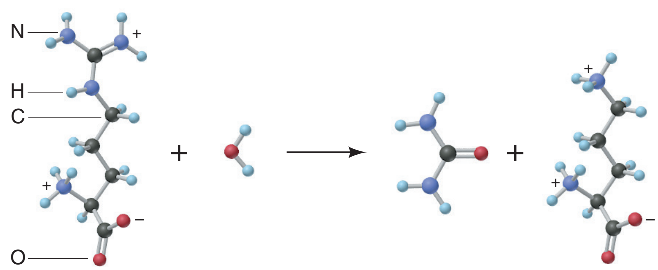


3.130 Fluorine is so reactive that it forms compounds with several of the noble gases.
(a) When 0.327 g of platinum is heated in fluorine, 0.519 g of a dark red, volatile solid forms. What is its empirical formula?
(b) When 0.265 g of this red solid reacts with excess xenon gas, 0.378 g of an orange-yellow solid forms. What is the empirical formula of this compound, the first to contain a noble gas?
(c) Fluorides of xenon can be formed by direct reaction of the elements at high pressure and temperature. 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?

3.131 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, which forms hemin ($C_{34}H_{32}N_4O_4FeCl$). A blood sample from a crime scene contains 0.65 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?

3.132 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) How many grams of braunite and of manganosite are in the ore? (b) What is the Mn^{3+}/Mn^{2+} ratio in the ore?

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



(a) What is the mass % of nitrogen in urea, in arginine, and in ornithine? (b) How many grams of nitrogen can be excreted as urea when 135.2 g of ornithine is produced?

3.134 Aspirin (acetylsalicylic acid, $C_9H_8O_4$) is 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 preparation, 3.077 g of salicylic acid and 5.50 mL of acetic anhydride react to form 3.281 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? (c) What is the percent atom economy of this reaction?

3.135 The rocket fuel hydrazine (N_2H_4) is made by the three-step Raschig process, which has the following overall equation:

$$NaOCl(aq) + 2NH_3(aq) \longrightarrow N_2H_4(aq) + NaCl(aq) + H_2O(l)$$

What is the percent atom economy of this process?

3.136 Lead(II) chromate ($PbCrO_4$) is used as the yellow pigment for marking traffic lanes but is banned from house paint because of the risk of lead poisoning. It is produced from chromite ($FeCr_2O_4$), an ore of chromium:

$$4FeCr_2O_4(s) + 8K_2CO_3(aq) + 7O_2(g) \longrightarrow 2Fe_2O_3(s) + 8K_2CrO_4(aq) + 8CO_2(g)$$

Lead(II) ion then replaces the K^+ ion. If a yellow paint is to have 0.511% $PbCrO_4$ by mass, how many grams of chromite are needed per kilogram of paint?

3.137 Ethanol (CH_3CH_2OH), the intoxicant in alcoholic beverages, is also used to make other organic compounds. In concentrated sulfuric acid, ethanol forms diethyl ether and water:

$$2CH_3CH_2OH(l) \longrightarrow CH_3CH_2OCH_2CH_3(l) + H_2O(g)$$

In a side reaction, some ethanol forms ethylene and water:

$$CH_3CH_2OH(l) \longrightarrow CH_2CH_2(g) + H_2O(g)$$

(a) If 50.0 g of ethanol yields 35.9 g of diethyl ether, what is the percent yield of diethyl ether? (b) If 45.0% of the ethanol that did not produce the ether reacts by the side reaction, what mass (g) of ethylene is produced?

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

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

Some of the reactants also combine with oxygen in air to form zinc oxide and sulfur dioxide. When 83.2 g of Zn reacts with 52.4 g of S_8, 104.4 g of ZnS forms.
(a) 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?

3.139 Cocaine ($C_{17}H_{21}O_4N$) is a natural substance found in coca leaves, which have been used for centuries as a local anesthetic and stimulant. Illegal cocaine arrives in the United States either as the pure compound or as the hydrochloride salt ($C_{17}H_{21}O_4NHCl$). At 25°C, the salt is very soluble in water (2.50 kg/L), but cocaine is much less so (1.70 g/L).
(a) What is the maximum mass (g) of the salt that can dissolve in 50.0 mL of water?
(b) If this solution is treated with NaOH, the salt is converted to cocaine. How much more water (L) is needed to dissolve it?

3.140 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 masses of this oxide when $x = 0$, $x = 1$, and $x = 0.163$.
(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 % of each reactant in the remaining solid mixture?