Stoichiometry of Formulas and Equations

3

Measuring Natural Substances Taxol, a substance with strong anticancer activity, was discovered in the bark of the Pacific yew using principles you'll learn in this chapter. Here we discuss how to determine a formula and find the amount of substance involved in a chemical change.

Outline

Key Principles
to focus on while studying this chapter

- The *mole* (mol) is the SI unit for *amount of substance* and contains *Avogadro's number* (6.022×10^{23}) of chemical entities (atoms, molecules, or ions). It has the same numerical value in grams as a single entity of the substance has in atomic mass units; for example, 1 molecule of H_2O weighs 18.02 amu and 1 mol of H_2O molecules weighs 18.02 g. Therefore, if the amount of a substance is expressed in moles, we know the number of entities in a given mass of it, which means that amount, mass, and number are interconvertible. ***(Section 3.1)***
- The subscripts in the chemical formula for a compound provide quantitative information about the amounts of each element in one mole of the compound. In an *empirical formula*, the subscripts show the *relative* numbers of moles of each element in the compound; in a *molecular formula*, they show the *actual* numbers. Isomers are different compounds with the same molecular formula. ***(Section 3.2)***
- In a *balanced equation*, chemical formulas preceded by integer (whole-number) coefficients give the same number of each kind of atom on the left *(reactants)* as on the right *(products)* but with atoms in different combinations. ***(Section 3.3)***
- Using molar ratios from the balanced equation, we calculate the amount of one substance from the amount of any other involved in the reaction. During a typical reaction, one substance (the *limiting reactant*) is used up, so it limits the amount of product that can form; the other reactant(s) is (are) in excess. The *theoretical yield*, the amount indicated by the balanced equation, is never obtained in the lab because of competing *side reactions*, incompleteness of the main reaction, and inability to collect all of the product. ***(Section 3.4)***
- For reactions in solution, we determine amounts of substances from their *concentration (molarity)* and the solution volume. To dilute a solution, we add *solvent,* which lowers the amount of *solute* dissolved in each unit volume. ***(Section 3.5)***

CONCEPTS & SKILLS TO REVIEW before studying 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)
- mass laws in chemical reactions (Section 2.2)

Chemistry is, above all, a practical science. Imagine that you're a biochemist who has extracted a substance with medicinal activity from a natural source: what is its formula, and what quantity of metabolic products will establish a safe dosage level? 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.

3.1 • THE MOLE

All the ideas in this chapter rely on understanding a key concept related to a unit called the *mole*. In daily life, we often measure things by counting or by weighing: we weigh beans or rice, but we count eggs or pencils. And we use counting units (a dozen pencils) or mass units (a kilogram of beans) 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 them.*

Defining the Mole

The **mole** (abbreviated **mol**) is the SI unit for *amount of substance*. It is defined as *the amount of a substance that contains the same number of entities as 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:

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

Thus,

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

THINK OF IT THIS WAY
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 United States 70 miles deep. However, atoms and molecules are not ordinary objects: a mole of water molecules (about 18 mL) can be swallowed in one gulp!

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 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, thus, contains 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,

1 atom of S	has a mass of	32.07 amu	and	1 mol (6.022×10^{23} atoms) of S	has a mass of	32.07 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

16
S
32.07

Note, also, that since atomic masses are relative, 1 Fe atom weighs 55.85/32.07 as much as 1 S atom, and 1 mol of Fe weighs 55.85/32.07 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 there are 1 dozen eggs from their weight or that there is 1 kilogram of beans from their count, because eggs and 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 2.13). 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:

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

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:

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

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

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

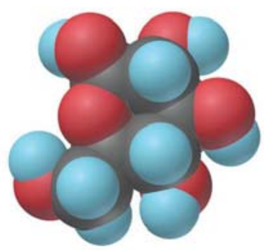
*The mass value in the periodic table has no units because it is a *relative* atomic mass, given by the atomic mass (in amu) divided by 1 amu ($\frac{1}{12}$ mass of one ^{12}C atom in amu):

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

Therefore, you use the same number for the atomic mass and for the molar mass.

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



Glucose

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.

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

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

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

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

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.
   - *From amount (mol) to number of entities,* multiply by Avogadro's number:

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

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

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

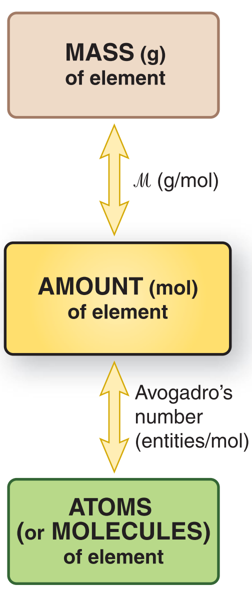
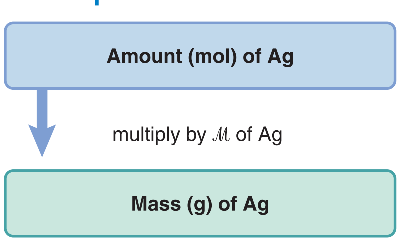


Figure 3.2 Mass-mole-number relationships for elements.

Amount-Mass-Number Conversions Involving Elements We begin with amount-mass-number relationships of elements. As Figure 3.2 shows, *convert mass (in grams) 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.

Road Map



Sample Problem 3.1 Calculating the Mass of a Given 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).

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 (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.

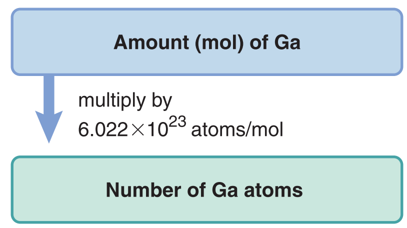
FOLLOW-UP PROBLEM 3.1 Graphite is the crystalline form of carbon used in "lead" pencils. How many moles of carbon are in 315 mg of graphite? Include a road map that shows how you planned the solution. (See Brief Solutions at the end of the chapter for the solution to this and all other follow-up problems.)

Sample Problem 3.2 Calculating the Number of Entities in a Given 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



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 PROBLEM 3.2 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.

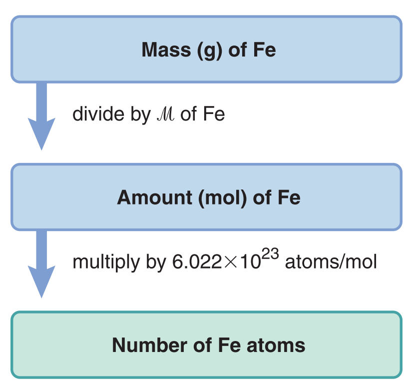
For the next sample problem, note that mass and number of entities relate directly to amount (mol), but *not* to each other. Therefore, *to convert between mass and number, first convert to amount.*

Sample Problem 3.3 Calculating the Number of Entities in a Given 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



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

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

Converting from 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 ~100 g/(~60 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 PROBLEM 3.3 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.

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.

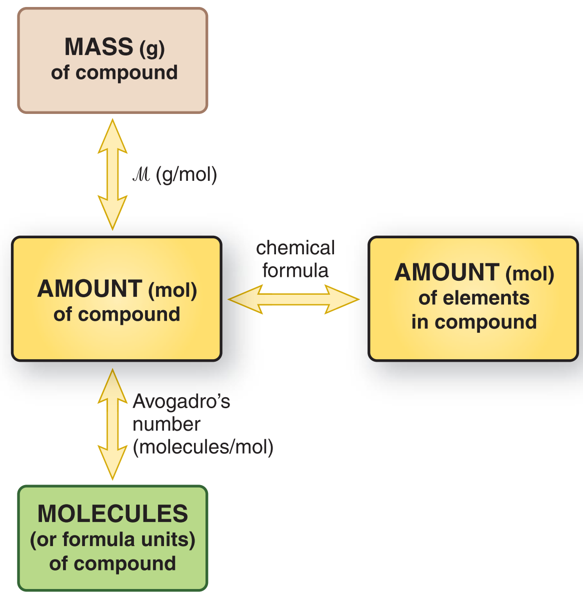


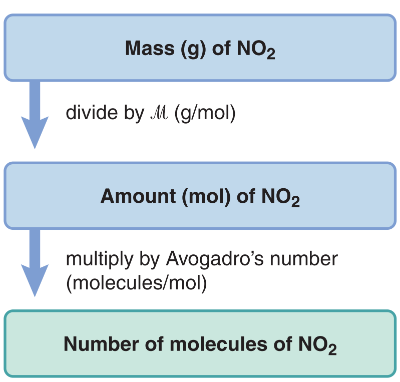
Figure 3.3 Amount-mass-number relationships for compounds. Use the chemical formula to find the amount (mol) of each element in a compound.

Sample Problem 3.4 Calculating the Number of Chemical Entities in a Given 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 2.15). 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:

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

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

$$\text{Amount (mol) of NO}_2 = 8.92 \cancel{\text{ g NO}_2} \times \frac{1 \text{ mol NO}_2}{46.01 \cancel{\text{ g NO}_2}} = 0.194 \text{ mol NO}_2$$

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

$$\text{No. of molecules} = 0.194 \cancel{\text{ mol NO}_2} \times \frac{6.022\times10^{23} \text{ NO}_2 \text{ molecules}}{1 \cancel{\text{ mol NO}_2}} = \boxed{1.17\times10^{23} \text{ NO}_2 \text{ molecules}}$$

Check Rounding, we get $(\sim0.2 \text{ mol})(6\times10^{23}) = 1.2\times10^{23}$, so the answer seems correct.

FOLLOW-UP PROBLEM 3.4 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.

In Sample Problem 3.5, we go a step further and find the number of atoms of an element in the sample of a compound.

Sample Problem 3.5 Calculating the Number of Chemical Entities in a Given 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 need the formula to find the amount (mol) and then multiply by Avogadro's number to find the number of formula units. [The road map steps are for part (a).] **(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.

Road Map

Mass (g) of $(NH_4)_2CO_3$

↓ divide by $\mathcal{M}$ (g/mol)

Amount (mol) of $(NH_4)_2CO_3$

↓ multiply by 6.022×10^{23} formula units/mol

Number of formula units of $(NH_4)_2CO_3$

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

$$\begin{aligned}\mathcal{M} &= (2\times\mathcal{M}\text{ of N}) + (8\times\mathcal{M}\text{ of H}) + (1\times\mathcal{M}\text{ of C}) + (3\times\mathcal{M}\text{ of O})\\ &= (2\times14.01\text{ g/mol N}) + (8\times1.008\text{ g/mol H}) + 12.01\text{ g/mol C}\\ &\quad + (3\times16.00\text{ g/mol O})\\ &= 96.09\text{ g/mol }(NH_4)_2CO_3\end{aligned}$$

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

$$\begin{aligned}\text{Amount (mol) of }(NH_4)_2CO_3 &= 41.6\ \cancel{\text{g }(NH_4)_2CO_3}\times\frac{1\text{ mol }(NH_4)_2CO_3}{96.09\ \cancel{\text{g }(NH_4)_2CO_3}}\\ &= 0.433\text{ mol }(NH_4)_2CO_3\end{aligned}$$

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:

$$\begin{aligned}\text{No. of O atoms} &= 2.61\times10^{23}\ \cancel{\text{formula units }(NH_4)_2CO_3}\times\frac{3\text{ O atoms}}{1\ \cancel{\text{formula unit }(NH_4)_2CO_3}}\\ &= \boxed{7.83\times10^{23}\text{ O atoms}}\end{aligned}$$

Check In (a), the units are correct. Since the mass is less than half the molar mass ($\sim42/96 < 0.5$), the number of formula units should be less than half Avogadro's number ($\sim2.6\times10^{23}/6.0\times10^{23} < 0.5$).

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

FOLLOW-UP PROBLEM 3.5 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?

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. 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:

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

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.

Sample Problem 3.6 Calculating the Mass Percent of Each Element in a Compound from the Formula

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

Plan We know the relative amounts (mol) of the elements from the formula (6 C, 12 H, 6 O), and we have to find the mass % of each element. We multiply the amount of each by its molar mass to find its mass. Dividing each mass by the mass of 1 mol of glucose gives the mass fraction of each element, and multiplying by 100 gives each mass %. The calculation steps for any element (X) are shown in the road map.

Road Map

Amount (mol) of element X in 1 mol of glucose

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

Mass (g) of X in 1 mol of glucose

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

Mass fraction of X in glucose

↓ multiply by 100

Mass % of X in glucose

Solution Converting amount (mol) of C to mass (g):
We have 6 mol of C in 1 mol of glucose, so

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

Calculating the mass of 1 mol of glucose ($C_6H_{12}O_6$):

$$\begin{aligned}\mathcal{M} &= (6 \times \mathcal{M}\text{ of C}) + (12 \times \mathcal{M}\text{ of H}) + (6 \times \mathcal{M}\text{ of O})\\ &= (6 \times 12.01\ \text{g/mol C}) + (12 \times 1.008\ \text{g/mol H}) + (6 \times 16.00\ \text{g/mol O})\\ &= 180.16\ \text{g/mol } C_6H_{12}O_6\end{aligned}$$

Finding the mass fraction of C in glucose:

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

Changing to mass %:

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

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

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

$$= \boxed{6.714\ \text{mass \% H}}$$

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

$$= \boxed{53.29\ \text{mass \% O}}$$

Check The answers make sense: even though there are equal numbers of moles of O and of C in the compound, the mass % of O is greater than the mass % of C because the molar mass of O is greater than the molar mass of C. The mass % of H is small because the molar mass of H is small. The 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 PROBLEM 3.6 Agronomists base the effectiveness of fertilizers on their nitrogen content. Ammonium nitrate is a common fertilizer. Calculate the mass percent of N in ammonium nitrate.

Determining the Mass of an Element from Its Mass Percent 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:

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

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

$$\text{Mass (g) of O} = 15.5\ \text{g } NO_2 \times \frac{2\ \text{mol} \times \mathcal{M}\ \text{of O (g/mol)}}{\text{mass (g) of 1 mol } NO_2}$$

$$= 15.5\ \text{g } \cancel{NO_2} \times \frac{32.00\ \text{g O}}{46.01\ \text{g } \cancel{NO_2}} = 10.8\ \text{g O}$$

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 carbon in 16.55 g of glucose.

Plan To find the mass of C in the sample of glucose, we multiply the mass of the sample by the mass of 6 mol of C divided by the mass of 1 mol of glucose.

Solution Finding the mass of C in a given mass of glucose:

$$\text{Mass (g) of C} = \text{mass (g) of glucose} \times \frac{6\ \text{mol C} \times \mathcal{M}\ \text{of C (g/mol)}}{\text{mass (g) of 1 mol glucose}}$$

$$= 16.55\ \cancel{\text{g glucose}} \times \frac{72.06\ \text{g C}}{180.16\ \cancel{\text{g glucose}}} = \boxed{6.620\ \text{g C}}$$

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

FOLLOW-UP PROBLEM 3.7 Use the information in Follow-up Problem 3.6 to find the mass (g) of N in 35.8 kg of ammonium nitrate.

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, molecules, or formula units).
- The mass (in grams) of a mole of the entity has the same numerical value as the mass (in amu) of the individual 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 an 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 numbers of moles, and thus the *relative* numbers 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 that

- The **molecular formula** shows the *actual* number of atoms of each element in a molecule: the molecular formula 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 see 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.

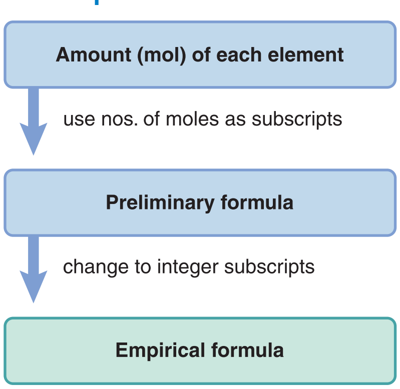
Sample Problem 3.8 demonstrates these math steps.

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

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 is 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 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 PROBLEM 3.9 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.]

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:

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

From Mass Percents to Formula Instead of giving masses of each element, analytical laboratories provide mass percents. We use these steps to find a formula:

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. Proceed as above 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 contains 40.0 mass % C, 6.71 mass % H, and 53.3 mass % O.
(a) Determine the empirical formula of lactic acid.
(b) Determine the molecular formula.

(a) Determining the empirical formula

Plan We know the mass % of each element and must convert each to an integer subscript. 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:

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

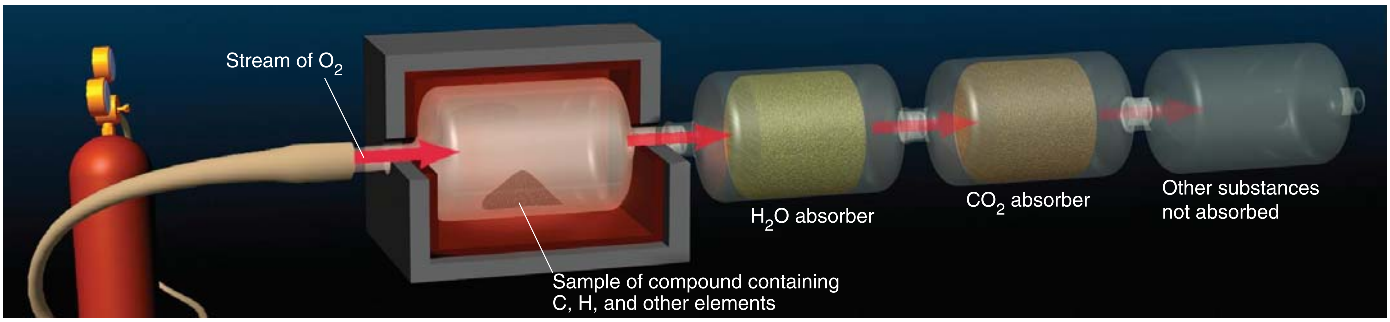
Check The calculated molecular formula has the same ratio of moles of elements (3/6/3) as the empirical formula (1/2/1) and corresponds to the given molar mass:

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

FOLLOW-UP PROBLEM 3.10 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?

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, and the H_2O and CO_2 that form 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:

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

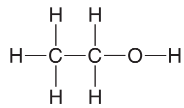
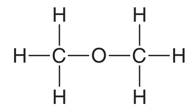
FOLLOW-UP PROBLEM 3.11 A dry-cleaning solvent ($\mathcal{M}$ = 146.99 g/mol) that contains C, H, and Cl is suspected to be a cancer-causing agent. When a 0.250-g sample was studied by combustion analysis, 0.451 g of CO_2 and 0.0617 g of H_2O formed. Find the molecular formula.

Isomers

A molecular formula tells the *actual* number of each type of atom, providing as much information as possible from mass analysis. Yet *different compounds can have the* ***same*** *molecular formula* because the atoms can bond to each other in different arrangements to give more than one *structural formula*. **Isomers** are two or more compounds with the same molecular formula but different properties. The simplest type of isomerism, called *constitutional,* or *structural, isomerism,* occurs when the atoms link together in different arrangements. The pair of constitutional isomers shown in Table 3.2 share the molecular formula C_2H_6O but have very different properties because they are different compounds. In this case, they are even different *types* of compounds—one is an alcohol, and the other an ether.

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

Table 3.2 Two Constitutional Isomers of C_2H_6O

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

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.
- Combustion analysis provides data on the masses of C and H 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 involving a small group of 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 3.5). This information is expressed by a **chemical equation,** which shows the identities and quantities of substances in a chemical or physical change.

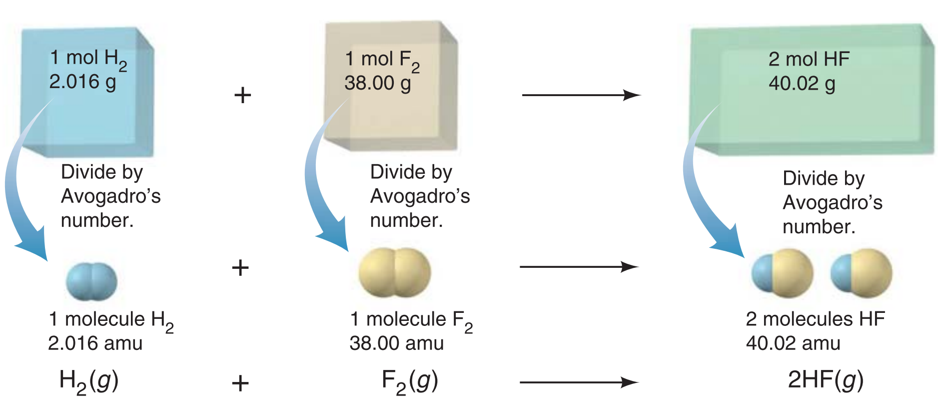


Figure 3.5 The formation of HF on the macroscopic and molecular levels.

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:**

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

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

   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:

$$\_\_Mg + \_\_O_2 \longrightarrow \underline{1}\,MgO$$

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

$$\underline{1}\,Mg + \_\_O_2 \longrightarrow \underline{1}\,MgO$$

   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:

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

   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:

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

   - 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:

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

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

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

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

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

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

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

As you saw in Figure 3.5, *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 3.6 depicts this reaction on three levels:

- *Macroscopic level (photos),* as it appears in the laboratory
- *Molecular 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

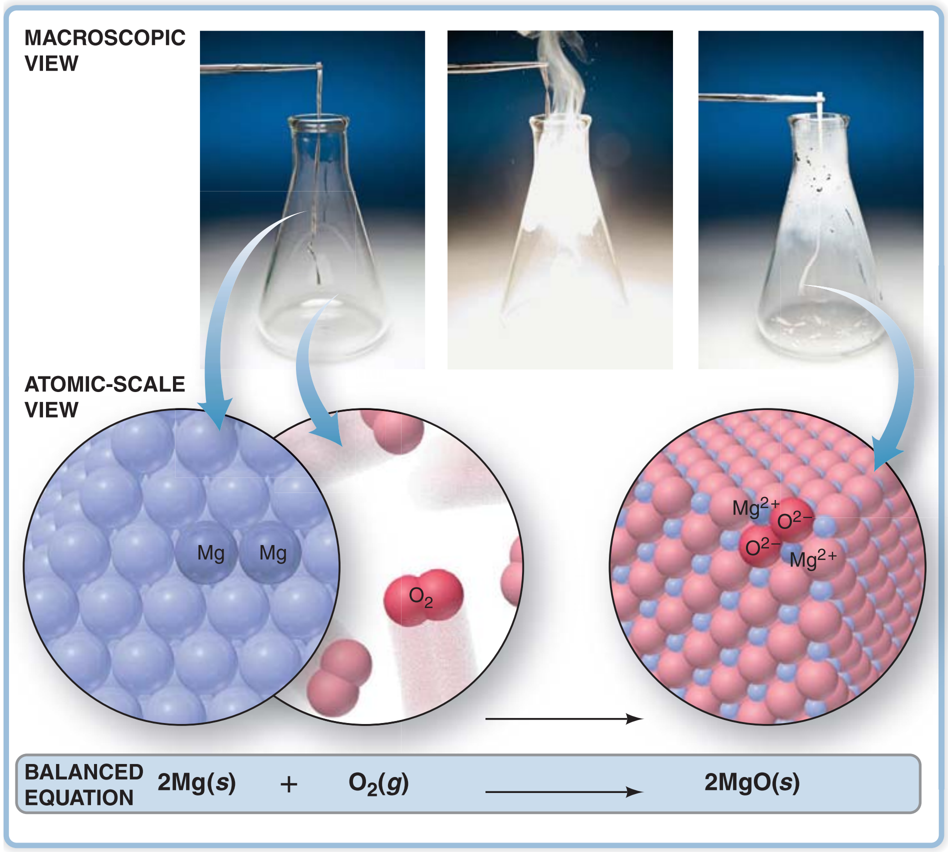


Figure 3.6 A three-level view of the reaction between magnesium and oxygen.

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. 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.

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:

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

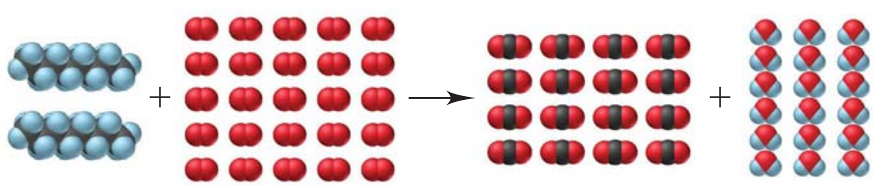
5. *Specify* states of matter. C_8H_{18} is liquid; O_2, CO_2, and H_2O vapor are gases:

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

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

FOLLOW-UP PROBLEM 3.12 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.
(d) Explosive decomposition of dynamite: liquid nitroglycerine ($C_3H_5N_3O_9$) explodes to produce a mixture of gases—carbon dioxide, water vapor, nitrogen, and oxygen.

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:

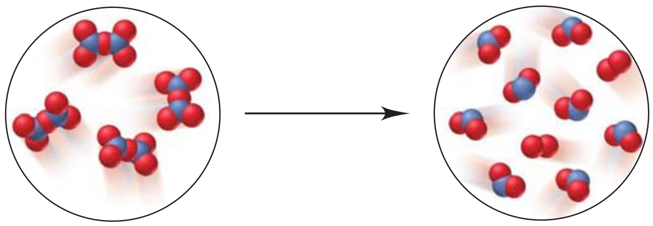


$$2C_8H_{18}(l) \quad + \quad 25O_2(g) \quad \longrightarrow \quad 16CO_2(g) \quad + \quad 18H_2O(g)$$

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 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:

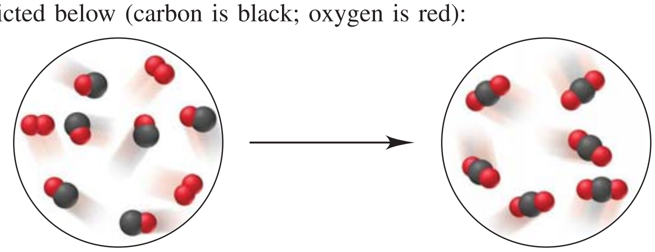
$$4N_2O_5 \longrightarrow 8NO_2 + 2O_2$$

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



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 can 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; Table 3.3 presents the quantitative information contained in this equation.

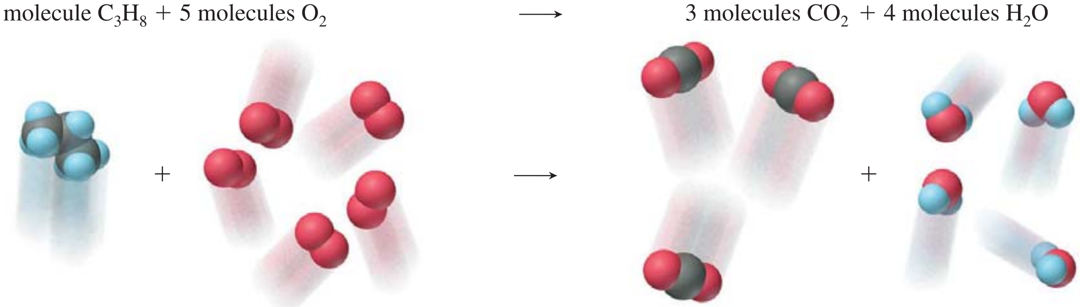
Here's a typical problem that shows how stoichiometric equivalence is used to create conversion factors: in the combustion of propane, how many moles of O_2 are consumed when 10.0 mol of H_2O is produced? To solve this problem, we have to find the molar ratio between O_2 and H_2O. From the balanced equation,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

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

Table 3.3 Information Contained in a Balanced Equation

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

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" *(the factor at left)* to cancel "mol H_2O":

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

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

*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 3.7 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.

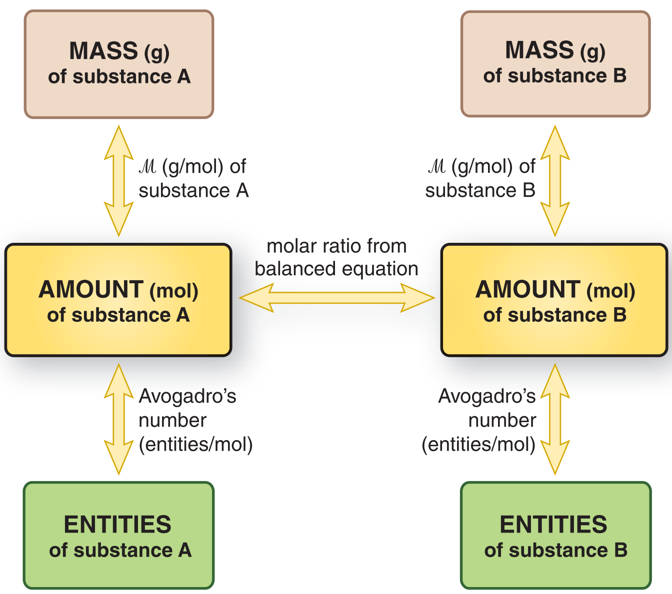


Figure 3.7 Summary of amount-mass-number relationships in a chemical equation. Start at any box (known) and move to any other (unknown) by using the conversion factor on the arrow. As always, convert to amount (mol) first.

Sample Problem 3.14 Calculating Quantities of Reactants and Products: Amount (mol) to Amount (mol)

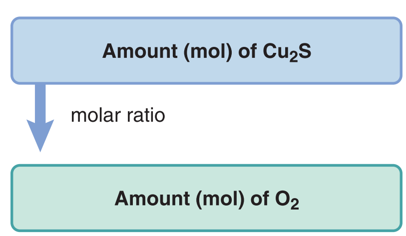
Problem In a lifetime, the average American uses more than half a ton (> 500 kg) of copper in coins, plumbing, and wiring. Copper is obtained from sulfide ores, such as chalcocite [copper(I) sulfide] by a multistep process. After initial grinding, the ore is "roasted" (heated strongly with oxygen gas) 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).

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).

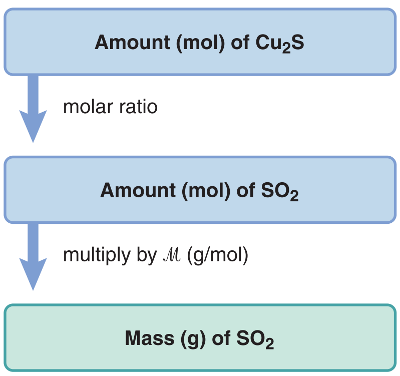
Comment A *common mistake* is to invert the conversion factor; that calculation would be

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

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 PROBLEM 3.14 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.

Road Map 3.15



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 top road map).

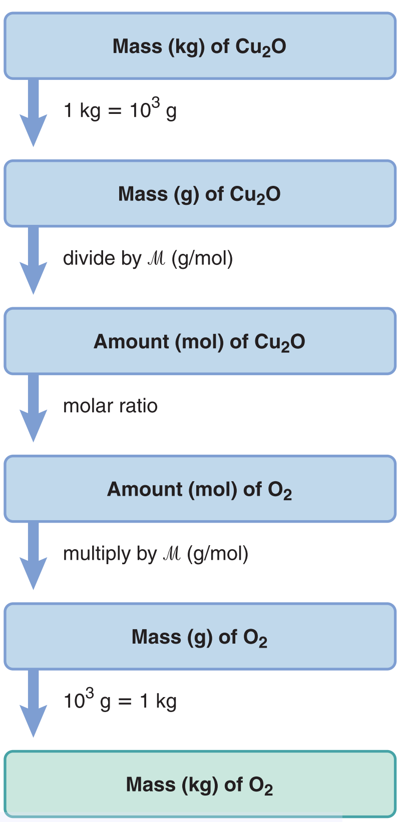
Solution Combining the two conversion steps into one calculation, we have

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

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 PROBLEM 3.15 In the thermite reaction, what amount (mol) of iron forms when 1.85×10^{25} formula units of aluminum oxide reacts? Write a road map to show how to plan the solution.

Road Map 3.16



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 amount of O_2 to g and then kg (see bottom road map).

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

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

Converting from moles of Cu_2O to moles of O_2:

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

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

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

Check The units are correct. Rounding to check the math, for example, in the final step, $\sim$30 mol $\times$ 30 g/mol $\times$ 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 PROBLEM 3.16 During the thermite reaction, 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.

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 reactant(s) 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:

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

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

1 car body 4 tires 1 car

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

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

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

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

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 car assembly "reaction," the reaction table would be

Quantity	1 car body	+	4 tires	$\longrightarrow$	1 car
Initial	1500		4000		0
Change	−1000		−4000		+1000
Final	500		0		1000

The body of the table shows the following important points:

- In the Initial line, "product" has not yet formed, so the entry is "0 cars."
- In the Change line, since the reactants (car bodies and tires) 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 (cars) increases, so the change in its quantity has a *positive* sign.
- For the Final line, we *add* the Change and Initial lines. Notice that one reactant (car bodies) is in excess, while the limiting reactant (tires) is used up.

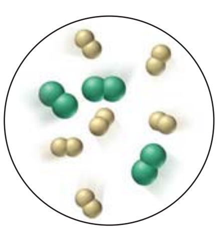
THINK OF IT THIS WAY
Limiting Reactants in Everyday Life

In addition to the industrial setting of a car assembly plant, limiting-"reactant" situations arise in daily life all the time. 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 making cheeseburgers for a picnic, and you have 10 buns, 12 meat patties, and 8 slices of cheese. Here, the cheese limits the number of cheeseburgers you can make. Or, there are 16 students in a cell biology lab but only 13 microscopes. The number of times limiting-reactant situations arise is almost limitless.

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

The following problems examine these ideas from several aspects. In Sample Problem 3.17, we solve the problem by looking at a molecular scene; in Sample Problem 3.18, we start with the amounts (mol) of two reactants, and in Sample Problem 3.19, we start with masses of two reactants.

Sample Problem 3.17 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

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

$$\text{For } Cl_2\text{: Molecules of } ClF_3 = 3 \cancel{\text{molecules of } Cl_2} \times \frac{2 \text{ molecules of } ClF_3}{1 \cancel{\text{molecule of } Cl_2}}$$

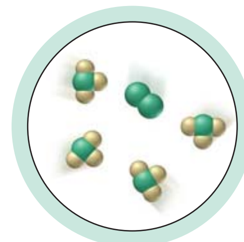
$$= 6 \text{ molecules of } ClF_3$$

$$\text{For } F_2\text{: Molecules of } ClF_3 = 6 \cancel{\text{molecules of } F_2} \times \frac{2 \text{ molecules of } ClF_3}{3 \cancel{\text{molecules of } F_2}}$$

$$= \tfrac{12}{3} \text{ molecules of } ClF_3 = 4 \text{ molecules of } ClF_3$$

Because it forms less product, F_2 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:

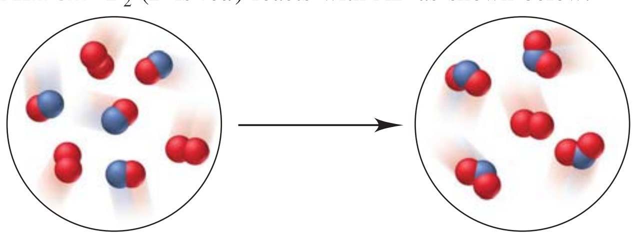
Molecules	$Cl_2(g)$	+	$3F_2(g)$	$\longrightarrow$	$2ClF_3(g)$
Initial	3		6		0
Change	−2		−6		+4
Final	1		0		4



(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 (2Cl, 6F) $\longrightarrow$ products (2Cl, 6F). 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 PROBLEM 3.17 B_2 (B is *red*) reacts with AB as shown below:



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

Sample Problem 3.18 Calculating Quantities in a Limiting-Reactant Problem: Amount to Amount

Problem In another preparation of ClF_3 (see Sample Problem 3.17), 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 3.17 and 3.18 is that the relative quantities of reactants *do not* determine which is limiting, but rather the quantity 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,

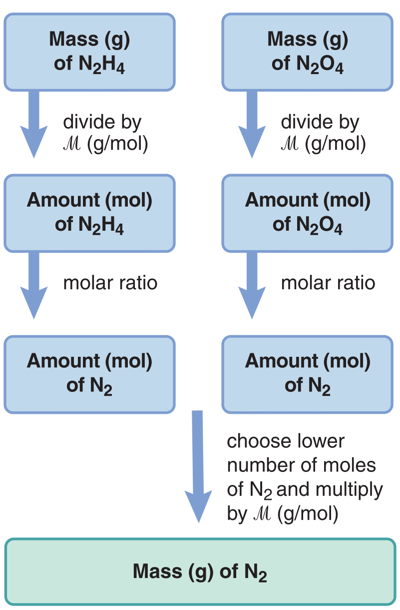
- Sample Problem 3.17 has an F_2/Cl_2 ratio of 6/3, or 2/1, which is less than the required molar ratio of 3/1, so F_2 is limiting and Cl_2 is in excess.
- Sample Problem 3.18 has an F_2/Cl_2 ratio of 3.00/0.750, which is greater than the required molar ratio of 3/1, so Cl_2 is limiting and F_2 is in excess.

FOLLOW-UP PROBLEM 3.18 For the reaction in Follow-up Problem 3.17, how many moles of product form from 1.5 mol of each reactant?

Sample Problem 3.19 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 tetraoxide (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) Write a reaction table for this process.

Road Map



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)** We use the values based on the limiting reactant for the reaction table.

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

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

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

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

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

Amount (mol)	$2N_2H_4(g)$	+	$N_2O_4(g)$	$\longrightarrow$	$3N_2(g)$	+	$4H_2O(g)$
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 The number of grams of N_2O_4 is more than that of 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 PROBLEM 3.19 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?

Theoretical, Actual, and Percent Reaction Yields

Up until now, we've assumed that 100% of the limiting reactant becomes product and that we use perfect lab technique to 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 3.8). In the rocket fuel reaction in Sample Problem 3.19, 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 one example in Chapter 4 and many more later in the book.)
   - Isolating a product often requires several steps, and some product is lost during each one. With careful lab technique, we 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).

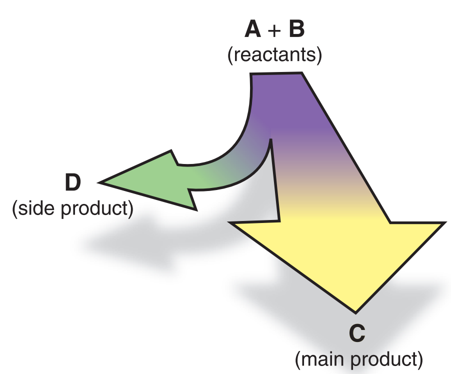


Figure 3.8 The effect of a side reaction on the yield of the main product.

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%.

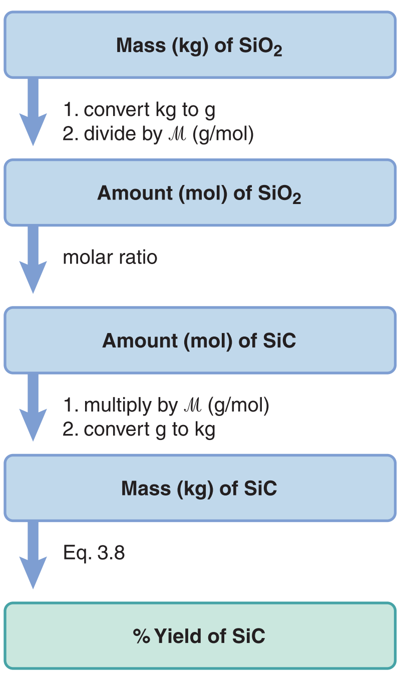
In the multistep synthesis of a complex compound, the overall yield can be surprisingly low, even if the yield of each step is high. For example, suppose a six-step synthesis has a 90.0% yield for each step. To find the overall percent yield, *express the yield of each step as a decimal, multiply them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

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

Sample Problem 3.20 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?

Road Map



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

FOLLOW-UP PROBLEM 3.20 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.

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.

3.5 • FUNDAMENTALS OF SOLUTION STOICHIOMETRY

In popular media, you may see a chemist in a lab coat, surrounded by glassware, mixing colored solutions that froth and give off billowing fumes. Most reactions in solution are not this dramatic, and good technique requires safer mixing procedures; however, it is true that aqueous solution chemistry is central to laboratory chemistry. And many environmental reactions and almost all biochemical reactions occur in solution. Liquid solutions are easier to store than gases and easier to mix than solids, and the amounts of substances in solution can be measured precisely.

We know the amounts of pure substances by converting their masses into number of moles. But for dissolved substances, we need the *concentration*—the number of moles per volume of solution—to find the volume that contains a given number of moles. In this section, we first discuss *molarity,* the most common way to express concentration (Chapter 13 covers others). Then, we see how to dilute a concentrated solution and how to use stoichiometric calculations for reactions in solution.

Expressing Concentration in Terms of Molarity

A solution consists of a smaller quantity of one substance, the **solute,** dissolved in a larger quantity of another, the **solvent.** When it dissolves, the solute's chemical entities become evenly dispersed throughout the solvent. The **concentration** of a solution is often expressed as *the quantity of solute dissolved in a given quantity of solution.*

Concentration is an *intensive* property (like density or temperature; Section 1.4) and thus is independent of the solution volume: a 50-L tank of a solution has the *same concentration* (solute quantity/solution quantity) as a 50-mL beaker of the solution. **Molarity (*M*)** expresses the concentration in units of *moles of solute per liter of solution:*

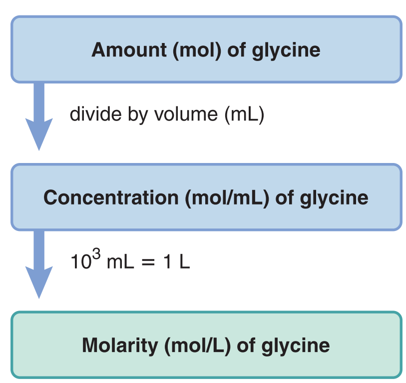
$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \quad \text{or} \quad M = \frac{\text{mol solute}}{\text{L soln}} \tag{3.9}$$

Sample Problem 3.21 Calculating the Molarity of a Solution

Problem Glycine has the simplest structure of the 20 amino acids that make up proteins. What is the molarity of a solution that contains 0.715 mol of glycine in 495 mL?

Plan The molarity is the number of moles of solute in each liter of solution. We divide the number of moles (0.715 mol) by the volume (495 mL) and convert the volume to liters to find the molarity (see the road map).

Road Map



Solution

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

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

FOLLOW-UP PROBLEM 3.21 How many moles of KI are in 84 mL of 0.50 *M* KI? Include a road map that shows how you planned the solution.

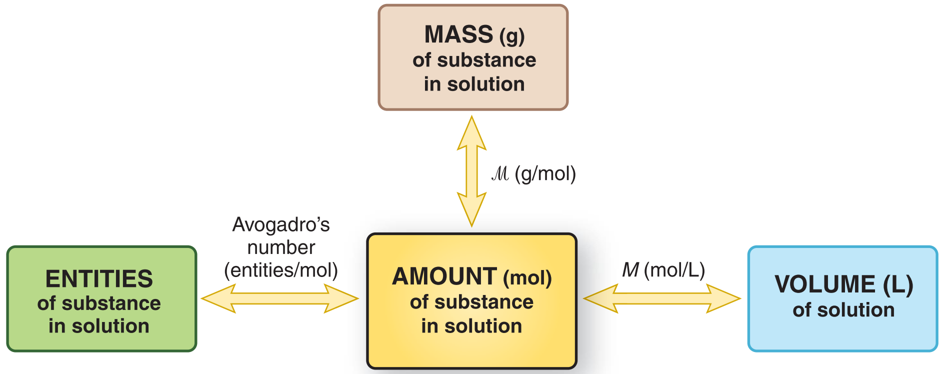


Figure 3.9 Summary of amount-mass-number relationships in solution. The amount (mol) of a substance in solution is related to the volume (L) of solution through the molarity (M; mol/L). As always, convert the given quantity to amount (mol) first.

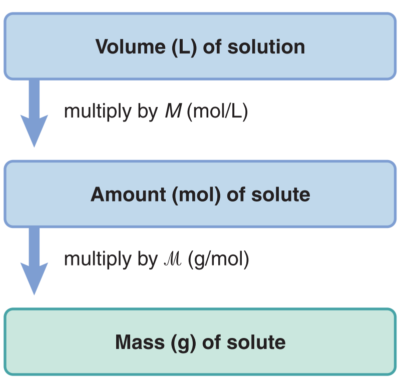
Amount-Mass-Number Conversions Involving Solutions

Like many intensive properties, molarity can be used as a *conversion factor* between volume (L) of solution and amount (mol) of solute, from which we can find the mass or the number of entities of solute (Figure 3.9), as applied in Sample Problem 3.22.

Sample Problem 3.22 Calculating Mass of Solute in a Given Volume of Solution

Problem Biochemists often study reactions in solutions containing phosphate ion, commonly found in cells. How many grams of solute are in 1.75 L of 0.460 M sodium hydrogen phosphate?

Road Map



Plan We need the mass (g) of solute, so we multiply the known solution volume (1.75 L) by the known molarity (0.460 M) to find the amount (mol) of solute and convert it to mass (g) using the solute's molar mass (see the road map).

Solution Calculating amount (mol) of solute in solution:

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

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

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

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

FOLLOW-UP PROBLEM 3.22 In biochemistry laboratories, solutions of sucrose (table sugar, $C_{12}H_{22}O_{11}$) are used in high-speed centrifuges to separate the parts of a biological cell. How many liters of 3.30 M sucrose contain 135 g of solute? Include a road map that shows how you planned the solution.

Diluting a Solution

A concentrated solution (higher molarity) is converted to a dilute solution (lower molarity) by adding solvent, which means the solution volume increases but the amount (mol) of solute stays the same. As a result, the dilute solution contains *fewer solute particles per unit volume* and, thus, has a lower concentration than the concentrated solution (Figure 3.10). If you need several different dilute solutions, prepare a concentrated solution *(stock solution)* and store it and dilute it as needed.

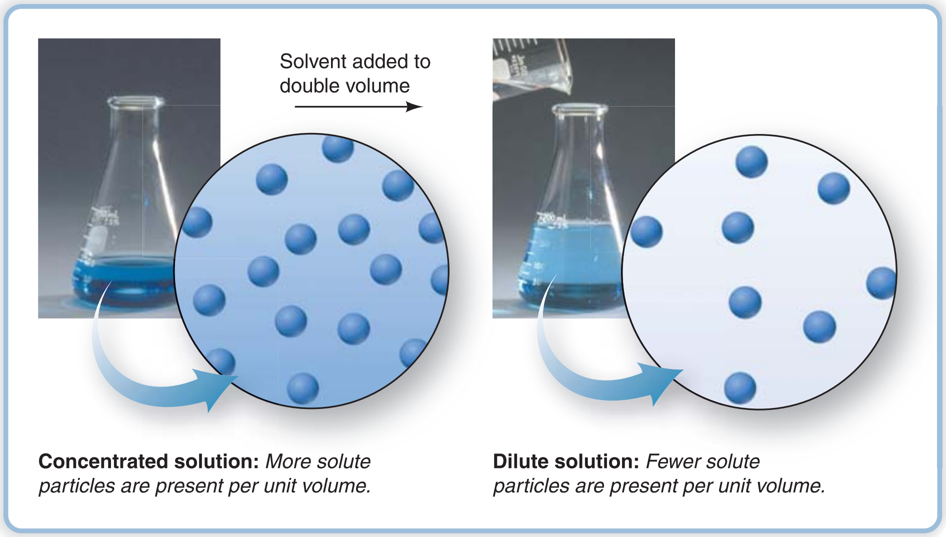


Figure 3.10 Converting a concentrated solution to a dilute solution.

Sample Problem 3.23 Preparing a Dilute Solution from a Concentrated Solution

Problem Isotonic saline is 0.15 M aqueous NaCl. It simulates the total concentration of ions in many cellular fluids, and its uses range from cleaning contact lenses to washing red blood cells. How would you prepare 0.80 L of isotonic saline from a 6.0 M stock solution?

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

Road Map

Volume (L) of dilute solution

↓ multiply by M (mol/L) of dilute solution

Amount (mol) of NaCl in dilute solution = Amount (mol) of NaCl in concentrated solution

↓ divide by M (mol/L) of concentrated solution

Volume (L) of concentrated solution

Solution Finding amount (mol) of solute in dilute solution:

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

Finding amount (mol) of solute in concentrated solution:

Because we add only solvent to dilute the solution,

$$\text{Amount (mol) of NaCl in dil soln} = \text{amount (mol) of NaCl in conc soln} = 0.12\ \text{mol NaCl}$$

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

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

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

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

FOLLOW-UP PROBLEM 3.23 A chemical engineer dilutes a stock solution of sulfuric acid by adding 25.0 m^3 of 7.50 M acid to enough water to make 500. m^3. What is the concentration of sulfuric acid in the diluted solution in g/mL?

To solve dilution problems and others involving a change in concentration, apply the following relationship:

$$M_{\text{dil}} \times V_{\text{dil}} = \text{amount (mol)} = M_{\text{conc}} \times V_{\text{conc}} \qquad (3.10)$$

where M and V are the molarity and volume of the *dil*ute (subscript "dil") and *conc*entrated (subscript "conc") solutions. Using the values in Sample Problem 3.23, for example, and solving Equation 3.10 for V_{conc} gives

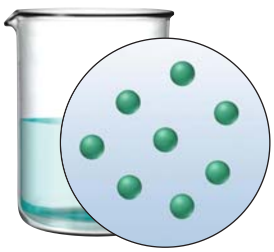
$$V_{conc} = \frac{M_{dil} \times V_{dil}}{M_{conc}} = \frac{0.15\ \cancel{M} \times 0.80\ \mathrm{L}}{6.0\ \cancel{M}}$$
$$= 0.020\ \mathrm{L}$$

Notice that Sample Problem 3.23 had the same calculation broken into two parts:

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

In the next sample problem, we use a variation of this relationship, with molecular scenes showing numbers of particles, to visualize changes in concentration.

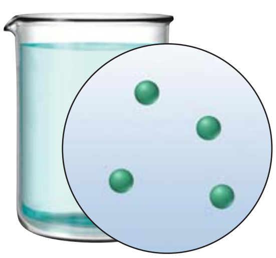
Sample Problem 3.24 Visualizing Changes in Concentration



Problem The top circle at left represents a unit volume of a solution. Draw a circle representing a unit volume of the solution after each of these changes:
(a) For every 1 mL of solution, 1 mL of solvent is added.
(b) One-third of the solvent is boiled off.

Plan Given the starting solution, we have to find the number of solute particles in a unit volume after each change. The number of particles per unit volume, N, is directly related to the number of moles per unit volume, M, so we can use a relationship similar to Equation 3.10 to find the number of particles. **(a)** The volume increases, so the final solution is more dilute—fewer particles per unit volume. **(b)** Some solvent is lost, so the final solution is more concentrated—more particles per unit volume.

(a)

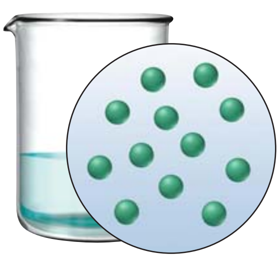


Solution **(a)** Finding the number of particles in the dilute solution, N_{dil}:

$$N_{dil} \times V_{dil} = N_{conc} \times V_{conc}$$

thus, $$N_{dil} = N_{conc} \times \frac{V_{conc}}{V_{dil}} = 8\ \text{particles} \times \frac{1\ \cancel{\mathrm{mL}}}{2\ \cancel{\mathrm{mL}}} = 4\ \text{particles}$$

(b)



(b) Finding the number of particles in the concentrated solution, N_{conc}:

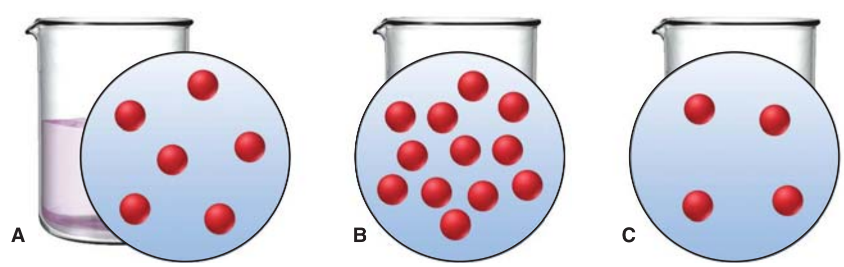
$$N_{dil} \times V_{dil} = N_{conc} \times V_{conc}$$

thus, $$N_{conc} = N_{dil} \times \frac{V_{dil}}{V_{conc}} = 8\ \text{particles} \times \frac{1\ \cancel{\mathrm{mL}}}{\frac{2}{3}\ \cancel{\mathrm{mL}}} = 12\ \text{particles}$$

Check In (a), the volume is doubled (from 1 mL to 2 mL), so the number of particles should be halved; $\frac{1}{2}$ of 8 is 4. In (b), the volume is $\frac{2}{3}$ of the original, so the number of particles should be $\frac{3}{2}$ of the original; $\frac{3}{2}$ of 8 is 12.

Comment In (b), we assumed that only solvent boils off. This is true with nonvolatile solutes, such as ionic compounds, but in Chapter 13, we'll encounter solutions in which both solvent *and* solute are volatile.

FOLLOW-UP PROBLEM 3.24 The circle labeled A represents a unit volume of a solution. Explain the changes that must be made to A to obtain the unit volumes in B and C.



Stoichiometry of Reactions in Solution

Solving stoichiometry problems for reactions in solution requires the additional step of converting the volume of reactant or product in solution to amount (mol):

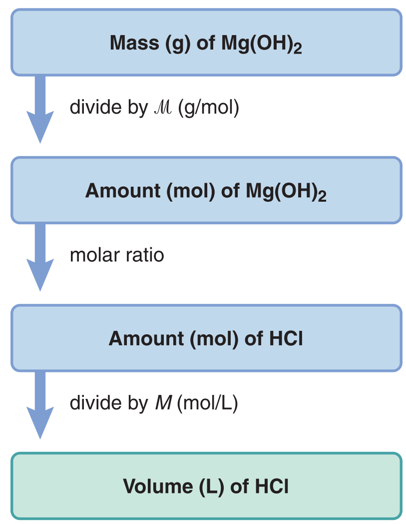
1. Balance the equation.
2. Find the amount (mol) of one substance from the volume and molarity.
3. Relate it to the stoichiometrically equivalent amount of another substance.
4. Convert to the desired units.

Sample Problem 3.25 Calculating Quantities of Reactants and Products for a Reaction in Solution

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

Plan We are given the mass (0.10 g) of magnesium hydroxide, $Mg(OH)_2$, that reacts with the acid. We also know the acid concentration (0.10 *M*) and must find the acid volume. After writing the balanced equation, we convert the mass (g) of $Mg(OH)_2$ to amount (mol) and use the molar ratio to find the amount (mol) of HCl that reacts with it. Then, we use the molarity of HCl to find the volume (L) that contains this amount (see the road map).

Road Map



Solution Writing the balanced equation:

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

Converting from mass (g) of $Mg(OH)_2$ to amount (mol):

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

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

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

Converting from amount (mol) of HCl to volume (L):

$$\text{Volume (L) of HCl} = 3.4\times10^{-3}\ \cancel{\text{mol HCl}} \times \frac{1 \text{ L}}{0.10\ \cancel{\text{mol HCl}}} = \boxed{3.4\times10^{-2} \text{ L}}$$

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

Comment In Chapter 4, you'll see that this equation is an oversimplification, because HCl and $MgCl_2$ exist in solution as separated ions.

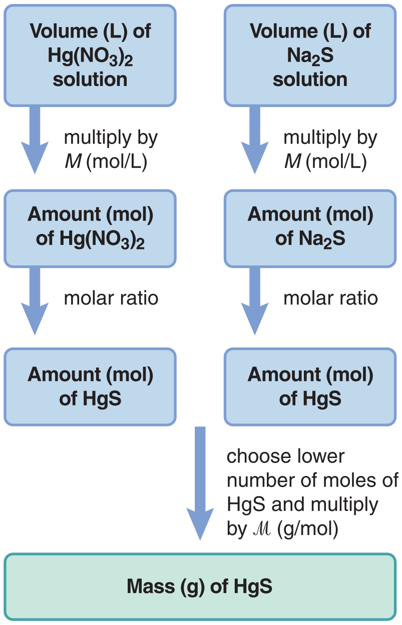
FOLLOW-UP PROBLEM 3.25 Another active ingredient in some antacids is aluminum hydroxide. Which is more effective at neutralizing stomach acid, magnesium hydroxide or aluminum hydroxide? [*Hint:* "Effectiveness" refers to the amount of acid that reacts with a given mass of antacid. You already know the effectiveness of 0.10 g of $Mg(OH)_2$.]

Except for the additional step of finding amounts (mol) in solution, limiting-reactant problems for reactions in solution are handled just like other such problems.

Sample Problem 3.26 Solving Limiting-Reactant Problems for Reactions in Solution

Problem Mercury and its compounds have uses from fillings for teeth (as a mixture with silver, copper, and tin) to the production of chlorine. Because of their toxicity, however, soluble mercury compounds, such as mercury(II) nitrate, must be removed from industrial wastewater. One removal method reacts the wastewater with sodium sulfide solution to produce solid mercury(II) sulfide and sodium nitrate solution. In a laboratory simulation, 0.050 L of 0.010 *M* mercury(II) nitrate reacts with 0.020 L of 0.10 *M* sodium sulfide.
(a) How many grams of mercury(II) sulfide form? **(b)** Write a reaction table for this process.

Road Map



Plan This is a limiting-reactant problem because *the quantities of two reactants are given.* After balancing the equation, we determine the limiting reactant. From the molarity and volume of each solution, we calculate the amount (mol) of each reactant. Then, we use the molar ratio to find the amount of product (HgS) that each reactant forms. The limiting reactant forms fewer moles of HgS, which we convert to mass (g) of HgS using its molar mass (see the road map). We use the amount of HgS formed from the limiting reactant in the reaction table.

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

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

Finding the amount (mol) of HgS formed from $Hg(NO_3)_2$: Combining the steps gives

$$\text{Amount (mol) of HgS} = 0.050\ \cancel{\text{L soln}} \times \frac{0.010\ \cancel{\text{mol Hg(NO}_3)_2}}{1\ \cancel{\text{L soln}}} \times \frac{1\ \text{mol HgS}}{1\ \cancel{\text{mol Hg(NO}_3)_2}}$$
$$= \mathbf{5.0{\times}10^{-4}\ mol\ HgS}$$

Finding the amount (mol) of HgS from Na_2S: Combining the steps gives

$$\text{Amount (mol) of HgS} = 0.020\ \cancel{\text{L soln}} \times \frac{0.10\ \cancel{\text{mol Na}_2\text{S}}}{1\ \cancel{\text{L soln}}} \times \frac{1\ \text{mol HgS}}{1\ \cancel{\text{mol Na}_2\text{S}}}$$
$$= \mathbf{2.0{\times}10^{-3}\ mol\ HgS}$$

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

$$\text{Mass (g) of HgS} = 5.0{\times}10^{-4}\ \cancel{\text{mol HgS}} \times \frac{232.7\ \text{g HgS}}{1\ \cancel{\text{mol HgS}}} = \boxed{0.12\ \text{g HgS}}$$

(b) With $Hg(NO_3)_2$ as the limiting reactant, the reaction table is

Amount (mol)	$Hg(NO_3)_2(aq)$	+	$Na_2S(aq)$	$\longrightarrow$	$HgS(s)$	+	$2NaNO_3(aq)$
Initial	$5.0{\times}10^{-4}$		$2.0{\times}10^{-3}$		0		0
Change	$-5.0{\times}10^{-4}$		$-5.0{\times}10^{-4}$		$+5.0{\times}10^{-4}$		$+1.0{\times}10^{-3}$
Final	0		$1.5{\times}10^{-3}$		$5.0{\times}10^{-4}$		$1.0{\times}10^{-3}$

A large excess of Na_2S remains after the reaction. Note that the amount of $NaNO_3$ formed is twice the amount of $Hg(NO_3)_2$ consumed, as the balanced equation shows.

Check As a check on our choice of the limiting reactant, let's use the alternative method in Sample Problem 3.19 (see Comment, p. 97).

Finding amount (mol) of reactants given:

$$\text{Amount (mol) of Hg(NO}_3)_2 = 0.050\ \cancel{\text{L soln}} \times \frac{0.010\ \text{mol Hg(NO}_3)_2}{1\ \cancel{\text{L soln}}} = 5.0{\times}10^{-4}\ \text{mol Hg(NO}_3)_2$$

$$\text{Amount (mol) of Na}_2\text{S} = 0.020\ \cancel{\text{L soln}} \times \frac{0.10\ \text{mol Na}_2\text{S}}{1\ \cancel{\text{L soln}}} = 2.0{\times}10^{-3}\ \text{mol Na}_2\text{S}$$

The molar ratio of the reactants is 1 $Hg(NO_3)_2$/1 Na_2S. Therefore, $Hg(NO_3)_2$ is limiting because there is less of it than we would need to react with all of the available Na_2S.

FOLLOW-UP PROBLEM 3.26 Despite the toxicity of lead, many of its compounds are still used to make pigments. **(a)** When 268 mL of 1.50 *M* lead(II) acetate reacts with 130. mL of 3.40 *M* sodium chloride, how many grams of solid lead(II) chloride can form? (Sodium acetate solution also forms.) **(b)** Using the abbreviation "Ac" for the acetate ion, write a reaction table for the process.

Summary of Section 3.5

- When reactions occur in solution, amounts of reactants and products are given in terms of concentration and volume.
- Molarity is the number of moles of solute dissolved in one liter of solution. A concentrated solution (higher molarity) is converted to a dilute solution (lower molarity) by adding solvent.
- By using molarity as a conversion factor, we can apply the principles of stoichiometry to reactions in solution.

CHAPTER REVIEW GUIDE

The following sections provide many aids to help you study this chapter. (Numbers in parentheses refer to pages, unless noted otherwise.)

Learning Objectives

These are concepts and skills to review after studying this chapter.

Related section (§), sample problem (SP), and upcoming end-of-chapter problem (EP) numbers are listed in parentheses.

1. Realize the usefulness of the mole concept, and use the relation between molecular (or formula) mass and molar mass to calculate the molar mass of any substance (§3.1) (EPs 3.1–3.5, 3.7–3.10)
2. Understand the relationships among amount of substance (in moles), mass (in grams), and number of chemical entities and convert from one to any other (§3.1) (SPs 3.1–3.5) (EPs 3.6, 3.11–3.16, 3.19)
3. Find a mass percent and use it to find the mass of element in a given mass of compound (§3.1) (SPs 3.6, 3.7) (EPs 3.17, 3.18, 3.20–3.23)
4. Determine the empirical and molecular formulas of a compound from mass analysis of its elements (§3.2) (SPs 3.8–3.11) (EPs 3.24–3.34)
5. Balance an equation given formulas or names, and use molar ratios to calculate amounts of reactants and products for reactions of pure or dissolved substances (§3.3, 3.5) (SPs 3.12–3.16, 3.25) (EPs 3.35–3.40, 3.63, 3.76, 3.77)
6. Understand why one reactant limits the yield of product, and solve limiting-reactant problems for reactions of pure or dissolved substances (§3.4, 3.5) (SPs 3.17–3.19, 3.26) (EPs 3.41–3.55, 3.62, 3.64, 3.74, 3.75)
7. Explain the reasons for lower-than-expected yields and the distinction between theoretical and actual yields, and calculate percent yield (§3.4) (SP 3.20) (EPs 3.56–3.61, 3.65)
8. Understand the meaning of concentration and the effect of dilution, and calculate molarity or mass of dissolved solute (§3.5) (SPs 3.21–3.24) (EPs 3.66–3.73, 3.78)

Key Terms

These important terms appear in boldface in the chapter and are defined again in the Glossary.

stoichiometry (72)

Section 3.1
mole (mol) (72)
Avogadro's number (72)
molar mass ($\mathcal{M}$) (73)

Section 3.2
empirical formula (80)
molecular formula (80)
structural formula (80)
combustion analysis (83)
isomer (84)

Section 3.3
chemical equation (85)
reactant (86)
product (86)
balancing (stoichiometric) coefficient (86)

Section 3.4
limiting reactant (93)
theoretical yield (97)
side reaction (97)
actual yield (97)
percent yield (% yield) (98)

Section 3.5
solute (99)
solvent (99)
concentration (99)
molarity (M) (99)

Key Equations and Relationships

Numbered and screened concepts are listed for you to refer to or memorize.

3.1 Number of entities in one mole (72):

$$1 \text{ mol contains } 6.022\times10^{23} \text{ entities (to 4 sf)}$$

3.2 Converting amount (mol) to mass (g) using $\mathcal{M}$ (74):

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

3.3 Converting mass (g) to amount (mol) using $1/\mathcal{M}$ (74):

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

3.4 Converting amount (mol) to number of entities (74):

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

3.5 Converting number of entities to amount (mol) (74):

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

3.6 Calculating mass % (78):

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 (79):

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

3.8 Calculating percent yield (98):

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

3.9 Defining molarity (99):

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

3.10 Diluting a concentrated solution (101):

$$M_{\text{dil}} \times V_{\text{dil}} = \text{amount (mol)} = M_{\text{conc}} \times V_{\text{conc}}$$

BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS

Compare your own solutions to these calculation steps and answers.

3.1

Mass (mg) of C

↓ 10^3 mg = 1 g

Mass (g) of C

↓ divide by $\mathcal{M}$ (g/mol)

Amount (mol) of C

$$\text{Amount (mol) of C} = 315\ \cancel{\text{mg C}} \times \frac{1\ \cancel{\text{g}}}{10^3\ \cancel{\text{mg}}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 2.62\times10^{-2}\ \text{mol C}$$

3.2

No. of N_2 molecules

↓ divide by Avogadro's number (molecules/mol)

Amount (mol) of N_2

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

Amount (mol) of N

$$\text{Amount (mol) of N} = 9.72\times10^{21}\ \cancel{\text{N}_2\ \text{molecules}} \times \frac{1\ \cancel{\text{mol N}_2}}{6.022\times10^{23}\ \cancel{\text{N}_2\ \text{molecules}}} \times \frac{2\ \text{mol N}}{1\ \cancel{\text{mol N}_2}} = 3.23\times10^{-2}\ \text{mol N}$$

3.3

No. of Mn atoms

↓ divide by Avogadro's number (molecules/mol)

Amount (mol) of Mn

↓ multiply by $\mathcal{M}$ (g/mol)

Mass (g) of Mn

$$\text{Mass (g) of Mn} = 3.22\times10^{20}\ \cancel{\text{Mn atoms}} \times \frac{1\ \cancel{\text{mol Mn}}}{6.022\times10^{23}\ \cancel{\text{Mn atoms}}} \times \frac{54.94\ \text{g Mn}}{1\ \cancel{\text{mol Mn}}} = 2.94\times10^{-2}\ \text{g Mn}$$

3.4

No. of formula units of NaF

↓ divide by Avogadro's number (formula units/mol)

Amount (mol) of NaF

↓ multiply by $\mathcal{M}$ (g/mol)

Mass (g) of NaF

$$\mathcal{M} = (1 \times \mathcal{M}\ \text{of Na}) + (1 \times \mathcal{M}\ \text{of F}) = 22.99\ \text{g/mol} + 19.00\ \text{g/mol} = 41.99\ \text{g/mol}$$

$$\text{Mass (g) of NaF} = 1.19\times10^{19}\ \cancel{\text{NaF formula units}} \times \frac{1\ \cancel{\text{mol NaF}}}{6.022\times10^{23}\ \cancel{\text{NaF formula units}}} \times \frac{41.99\ \text{g NaF}}{1\ \cancel{\text{mol NaF}}} = 8.30\times10^{-4}\ \text{NaF}$$

3.5 (a)
$$\text{Mass (g) of P}_4\text{O}_{10} = 4.65\times10^{22}\ \cancel{\text{molecules P}_4\text{O}_{10}} \times \frac{1\ \cancel{\text{mol P}_4\text{O}_{10}}}{6.022\times10^{23}\ \cancel{\text{molecules P}_4\text{O}_{10}}} \times \frac{283.88\ \text{g P}_4\text{O}_{10}}{1\ \cancel{\text{mol P}_4\text{O}_{10}}} = 21.9\ \text{g P}_4\text{O}_{10}$$

(b)
$$\text{No. of P atoms} = 4.65\times10^{22}\ \cancel{\text{molecules P}_4\text{O}_{10}} \times \frac{4\ \text{P atoms}}{1\ \cancel{\text{molecule P}_4\text{O}_{10}}} = 1.86\times10^{23}\ \text{P atoms}$$

3.6
$$\text{Mass \% of N} = \frac{2\ \cancel{\text{mol N}} \times \dfrac{14.01\ \cancel{\text{g N}}}{1\ \cancel{\text{mol N}}}}{80.05\ \cancel{\text{g NH}_4\text{NO}_3}} \times 100 = 35.00\ \text{mass \% N}$$

3.7
$$\text{Mass (g) of N} = 35.8\ \cancel{\text{kg NH}_4\text{NO}_3} \times \frac{10^3\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{0.3500\ \text{g N}}{1\ \cancel{\text{g NH}_4\text{NO}_3}} = 1.25\times10^4\ \text{g N}$$

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

3.9
$$\text{Amount (mol) of S} = 2.88\ \cancel{\text{g S}} \times \frac{1\ \text{mol S}}{32.07\ \cancel{\text{g S}}} = 0.0898\ \text{mol S}$$

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

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

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

3.10 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: $C_{20}H_{12}$

3.11
$$\text{Mass (g) of C} = 0.451\ \cancel{\text{g CO}_2} \times \frac{12.01\ \text{g C}}{44.01\ \cancel{\text{g CO}_2}} = 0.123\ \text{g C}$$

Similarly, there is 0.00690 g H.

Mass (g) of Cl $= 0.250\text{ g} - (0.123\text{ g} + 0.00690\text{ g}) = 0.120\text{ g Cl}$
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: $C_6H_4Cl_2$

3.12 (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) $PCl_3(g) + 3HF(g) \longrightarrow PF_3(g) + 3HCl(g)$
(d) $4C_3H_5N_3O_9(l) \longrightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)$

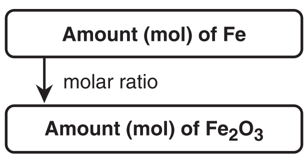
3.13 From the depiction, we have

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

Or,

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

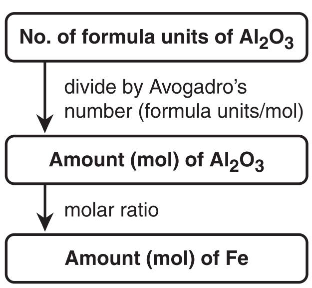
3.14



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

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

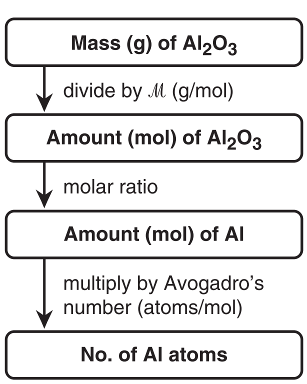
3.15



Amount (mol) of Fe

$$= 1.85\times10^{25}\ \cancel{\text{formula units } Al_2O_3} \times \frac{1\ \cancel{\text{mol } Al_2O_3}}{6.022\times10^{23}\ \cancel{\text{formula units } Al_2O_3}} \times \frac{2\text{ mol Fe}}{1\ \cancel{\text{mol } Al_2O_3}}$$
$$= 61.4\text{ mol Fe}$$

3.16



$$\text{No. of Al atoms} = 1.00\ \cancel{\text{g } Al_2O_3} \times \frac{1\ \cancel{\text{mol } Al_2O_3}}{101.96\ \cancel{\text{g } Al_2O_3}} \times \frac{2\ \cancel{\text{mol Al}}}{1\ \cancel{\text{mol } Al_2O_3}} \times \frac{6.022\times10^{23}\text{ Al atoms}}{1\ \cancel{\text{mol Al}}}$$
$$= 1.18\times10^{22}\text{ Al atoms}$$

3.17 $4AB + 2B_2 \longrightarrow 4AB_2$, 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_2 is the limiting reactant; one B_2 molecule is in excess.

3.18

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

$$\text{Amount (mol) of } AB_2 = 1.5\ \cancel{\text{mol } B_2} \times \frac{2\text{ mol } AB_2}{1\ \cancel{\text{mol } B_2}} = 3.0\text{ mol } AB_2$$

Therefore, 1.5 mol of AB_2 can form.

3.19 $2Al(s) + 3S(s) \longrightarrow Al_2S_3(s)$
Mass (g) of Al_2S_3 formed from 10.0 g of Al

$$= 10.0\ \cancel{\text{g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{26.98\ \cancel{\text{g Al}}} \times \frac{1\ \cancel{\text{mol } Al_2S_3}}{2\ \cancel{\text{mol Al}}} \times \frac{150.17\text{ g } Al_2S_3}{1\ \cancel{\text{mol } Al_2S_3}}$$
$$= 27.8\text{ g } Al_2S_3$$

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

$$= \text{total mass of Al} - \text{mass of Al used}$$
$$= 10.0\text{ g Al} - \left(15.0\ \cancel{\text{g S}} \times \frac{1\ \cancel{\text{mol S}}}{32.07\ \cancel{\text{g S}}} \times \frac{2\ \cancel{\text{mol Al}}}{3\ \cancel{\text{mol S}}} \times \frac{26.98\text{ g Al}}{1\ \cancel{\text{mol Al}}}\right)$$
$$= 1.6\text{ g Al}$$

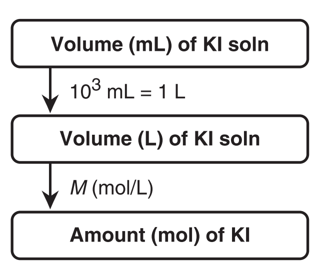
(We would obtain the same answer if sulfur were shown more correctly as S_8.)

3.20 $CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

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

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

3.21



$$\text{Amount (mol) of KI} = 84\ \cancel{\text{mL soln}} \times \frac{1\ \cancel{\text{L}}}{10^3\ \cancel{\text{mL}}} \times \frac{0.50\text{ mol KI}}{1\ \cancel{\text{L soln}}}$$
$$= 0.042\text{ mol KI}$$

3.22

Mass (g) of sucrose

↓ divide by $\mathcal{M}$ (g/mol)

Amount (mol) of sucrose

↓ divide by M (mol/L)

Volume (L) of sucrose soln

Volume (L) of sucrose soln

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

$$= 0.120\ \text{L soln}$$

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

Mass (g) of H_2SO_4/mL soln

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

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

3.24 To obtain B, the total volume of sol A was reduced by half:

$$V_{\text{conc}} = V_{\text{dil}} \times \frac{N_{\text{dil}}}{N_{\text{conc}}} = 1.0\ \text{mL} \times \frac{6\ \cancel{\text{particles}}}{12\ \cancel{\text{particles}}} = 0.50\ \text{mL}$$

To obtain C, $\frac{1}{2}$ of a volume of solvent was added for every volume of A:

$$V_{\text{dil}} = V_{\text{conc}} \times \frac{N_{\text{conc}}}{N_{\text{dil}}} = 1.0\ \text{mL} \times \frac{6\ \cancel{\text{particles}}}{4\ \cancel{\text{particles}}} = 1.5\ \text{mL}$$

3.25 $Al(OH)_3(s) + 3HCl(aq) \longrightarrow AlCl_3(aq) + 3H_2O(l)$

Volume (L) of HCl neutralized

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

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

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

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

3.26 (a) For $Pb(C_2H_3O_2)_2$:

$$\text{Amount (mol) of } Pb(C_2H_3O_2)_2 = 0.268\ \cancel{\text{L}} \times \frac{1.50\ \text{mol } Pb(C_2H_3O_2)_2}{1\ \cancel{\text{L}}}$$

$$= 0.402\ \text{mol } Pb(C_2H_3O_2)_2$$

$$\text{Amount (mol) of } PbCl_2 = 0.402\ \cancel{\text{mol } Pb(C_2H_3O_2)_2}$$

$$\times \frac{1\ \text{mol } PbCl_2}{1\ \cancel{\text{mol } Pb(C_2H_3O_2)_2}} = 0.402\ \text{mol } PbCl_2$$

$$\text{For NaCl: Amount (mol) of NaCl} = \frac{3.40\ \text{mol NaCl}}{1\ \cancel{\text{L}}} \times 0.130\ \cancel{\text{L}}$$

$$= 0.442\ \text{mol NaCl}$$

$$\text{Amount (mol) of } PbCl_2 = 0.442\ \cancel{\text{mol NaCl}} \times \frac{1\ \text{mol } PbCl_2}{2\ \cancel{\text{mol NaCl}}}$$

$$= 0.221\ \text{mol } PbCl_2$$

Thus, NaCl is the limiting reactant.

$$\text{Mass (g) of } PbCl_2 = 0.22\ \cancel{\text{mol } PbCl_2} \times \frac{278.1\ \text{g } PbCl_2}{1\ \cancel{\text{mol } PbCl_2}}$$

$$= 61.5\ \text{g } PbCl_2$$

(b) Using "Ac" for acetate ion and with NaCl limiting:

Amount (mol)	$Pb(Ac)_2$	+	2NaCl	$\longrightarrow$	$PbCl_2$	+	2NaAc
Initial	0.402		0.442		0		0
Change	−0.221		− 0.442		+0.221		+0.442
Final	0.181		0		0.221		0.442

PROBLEMS

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

The Mole

(Sample Problems 3.1 to 3.7)

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?

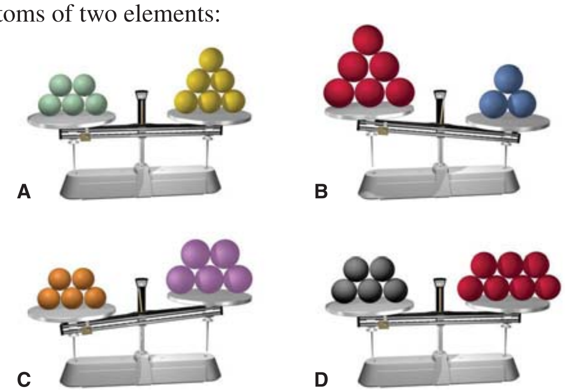
3.2 (a) How many moles of C atoms are in 1 mol of sucrose ($C_{12}H_{22}O_{11}$)?
(b) How many C atoms are in 2 mol of sucrose?

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 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?

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

3.8 Calculate the molar mass of each of the following:
(a) $(NH_4)_3PO_4$ (b) CH_2Cl_2 (c) $CuSO_4 \cdot 5H_2O$ (d) BrF_3

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

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

3.11 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 $CaSO_4 \cdot 2H_2O$

3.12 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

3.13 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

3.14 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 $Bi(NO_3)_3 \cdot 5H_2O$

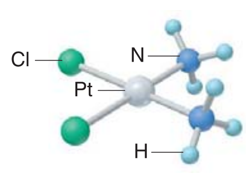
3.15 Calculate each of the following quantities:
(a) Mass (g) of 8.35 mol of copper(I) carbonate
(b) Mass (g) of 4.04×10^{20} molecules of dinitrogen pentoxide
(c) Amount (mol) and number of formula units in 78.9 g of sodium perchlorate
(d) Number of sodium ions, perchlorate ions, chlorine atoms, and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

3.16 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)

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

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

3.19 Cisplatin *(right)*, or Platinol, is 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.



3.20 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.

3.21 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, $CO(NH_2)_2$.

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

3.23 Hemoglobin, a protein in red blood cells, carries O_2 from the lungs to the body's cells. Iron (as ferrous ion, Fe^{2+}) makes up 0.33 mass % of hemoglobin. If the molar mass of hemoglobin is 6.8×10^4 g/mol, how many Fe^{2+} ions are in one molecule?

Determining the Formula of an Unknown Compound

(Sample Problems 3.8 to 3.11)

3.24 Which of the following sets of information allows you to obtain the molecular formula of a covalent compound? In each case that allows it, explain how you would proceed (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

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

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

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

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

3.29 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

3.30 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

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

3.32 A 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?

3.33 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?

3.34 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

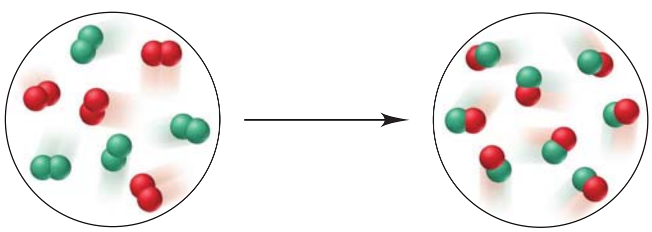
(Sample Problems 3.12 and 3.13)

3.35 In the process of balancing the equation

$$Al + Cl_2 \longrightarrow AlCl_3$$

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

3.36 The 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) $4A_2 + 4B_2 \longrightarrow 8AB$

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

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

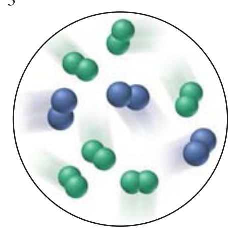
3.39 Convert the following into balanced equations:
(a) When gallium metal is heated in oxygen gas, it melts and forms solid gallium(III) oxide.
(b) Liquid hexane burns in oxygen gas to form carbon dioxide gas and water vapor.
(c) When solutions of calcium chloride and sodium phosphate are mixed, solid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

3.40 Convert the following into balanced equations:
(a) When lead(II) nitrate solution is added to potassium iodide solution, solid lead(II) iodide forms and potassium nitrate solution remains.
(b) Liquid disilicon hexachloride reacts with water to form solid silicon dioxide, hydrogen chloride gas, and hydrogen gas.
(c) When nitrogen dioxide is bubbled into water, a solution of nitric acid forms and gaseous nitrogen monoxide is released.

Calculating Quantities of Reactant and Product

(Sample Problems 3.14 to 3.20)

3.41 The circle 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?

3.42 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?

3.43 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?

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

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

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

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

3.46 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?

3.47 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?

3.48 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?

3.49 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?

3.50 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:

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

How many grams of the excess reactant remains?

3.51 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:

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

How many grams of the excess reactant remain?

3.52 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?

3.53 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?

3.54 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?

3.55 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?

3.56 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?

3.57 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?

3.58 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)?

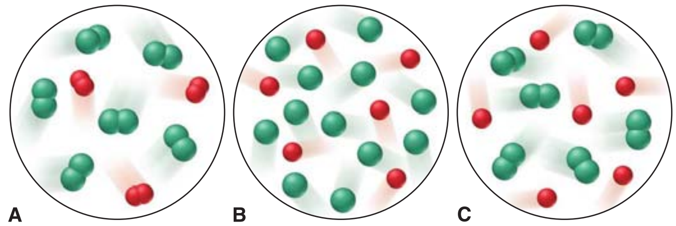
3.59 What is the percent yield of a reaction in which 200. g of phosphorus trichloride reacts with excess water to form 128 g of HCl and aqueous phosphorous acid (H_3PO_3)?

3.60 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.

3.61 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?

3.62 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?

3.63 Gaseous dichlorine monoxide decomposes readily to chlorine and oxygen 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?

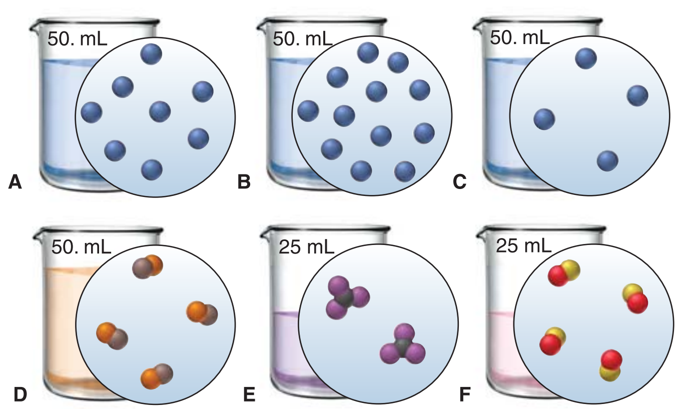
3.64 Butane gas is compressed and used as a liquid fuel in disposable cigarette lighters and lightweight camping stoves. Suppose a lighter contains 5.50 mL of butane ($d = 0.579$ g/mL).
(a) How many grams of oxygen are needed to burn the butane completely?
(b) How many moles of H_2O form when all the butane burns?
(c) How many total molecules of gas form when the butane burns completely?

3.65 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?

Fundamentals of Solution Stoichiometry

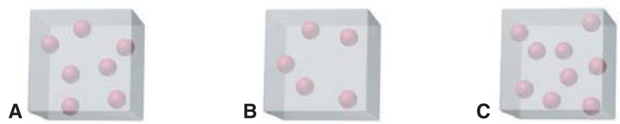
(Sample Problems 3.21 to 3.26)

3.66 Six different aqueous solutions (with solvent molecules omitted for clarity) are represented in the beakers below, and their total volumes are noted.



(a) Which solution has the highest molarity? (b) Which solutions have the same molarity? (c) If you mix solutions A and C, does the resulting solution have a higher, a lower, or the same molarity as solution B? (d) After 50. mL of water is added to solution D, is its molarity higher, lower, or the same as the molarity of solution F after 75 mL is added to it? (e) How much solvent must be evaporated from solution E for it to have the same molarity as solution A?

3.67 Boxes A, B, and C represent a unit volume of three solutions of the same solute. Which box, B or C, represents the solution that has (a) more solute added; (b) more solvent added; (c) higher molarity; (d) lower concentration?



3.68 Calculate each of the following quantities:
(a) Mass (g) of solute in 185.8 mL of 0.267 M calcium acetate
(b) Molarity of 500. mL of solution containing 21.1 g of potassium iodide
(c) Amount (mol) of solute in 145.6 L of 0.850 M sodium cyanide

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

3.70 Calculate each of the following quantities:
(a) Molarity of a solution prepared by diluting 37.00 mL of 0.250 M potassium chloride to 150.00 mL
(b) Molarity of a solution prepared by diluting 25.71 mL of 0.0706 M ammonium sulfate to 500.00 mL
(c) Molarity of sodium ion in a solution made by mixing 3.58 mL of 0.348 M sodium chloride with 500. mL of 6.81×10^{-2} M sodium sulfate (assume volumes are additive)

3.71 Calculate each of the following quantities:
(a) Volume (L) of 2.050 M copper(II) nitrate that must be diluted with water to prepare 750.0 mL of a 0.8543 M solution
(b) Volume (L) of 1.63 M calcium chloride that must be diluted with water to prepare 350. mL of a 2.86×10^{-2} M chloride ion solution
(c) Final volume (L) of a 0.0700 M solution prepared by diluting 18.0 mL of 0.155 M lithium carbonate with water

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

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

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

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

3.75 How many grams of NaH_2PO_4 are needed to react with 43.74 mL of 0.285 M NaOH?

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

3.76 How many grams of solid barium sulfate form when 35.0 mL of 0.160 M barium chloride reacts with 58.0 mL of 0.065 M sodium sulfate? Aqueous sodium chloride forms also.

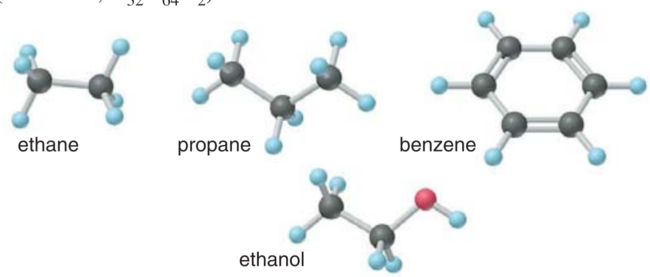
3.77 How many moles of excess reactant are present when 350. mL of 0.210 M sulfuric acid reacts with 0.500 L of 0.196 M sodium hydroxide to form water and aqueous sodium sulfate?

3.78 Muriatic acid, an industrial grade of concentrated HCl, is used to clean masonry and cement. Its concentration is 11.7 M.
(a) Write instructions for diluting the concentrated acid to make 3.0 gallons of 3.5 M acid for routine use (1 gal = 4 qt; 1 qt = 0.946 L). (b) How many milliliters of the muriatic acid solution contain 9.66 g of HCl?

Comprehensive Problems

3.79 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$\cdot xH_2O$.

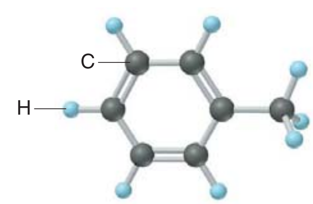
3.80 Hydrogen-containing fuels have a "fuel value" based on their mass % H. Rank the following compounds from highest fuel value to lowest: ethane, propane, benzene, ethanol, cetyl palmitate (whale oil, $C_{32}H_{64}O_2$).



3.81 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.

3.82 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?

3.83 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?



3.84 During studies of the reaction in Sample Problem 3.19,

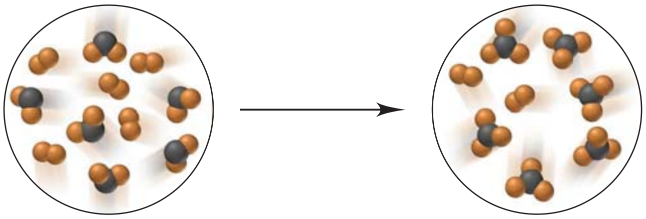
$$2N_2H_4(l) + N_2O_4(l) \longrightarrow 3N_2(g) + 4H_2O(g)$$

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?

3.85 The following circles represent a chemical reaction between AB_2 and B_2:

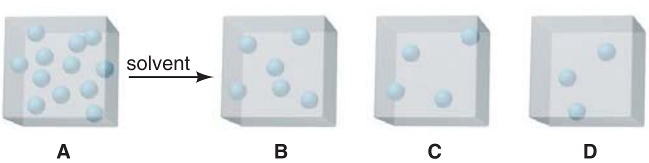


(a) Write a balanced equation for the reaction. (b) What is the limiting reactant? (c) How many moles of product can be made from 3.0 mol of B_2 and 5.0 mol of AB_2? (d) How many moles of excess reactant remain after the reaction in part (c)?

3.86 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. (c) Find the molarity of NaCl in seawater at 15°C (d of seawater at 15°C = 1.025 g/mL).

3.87 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 concentration of a solution is an intensive property, but the amount of solute in a solution is an extensive property.

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



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

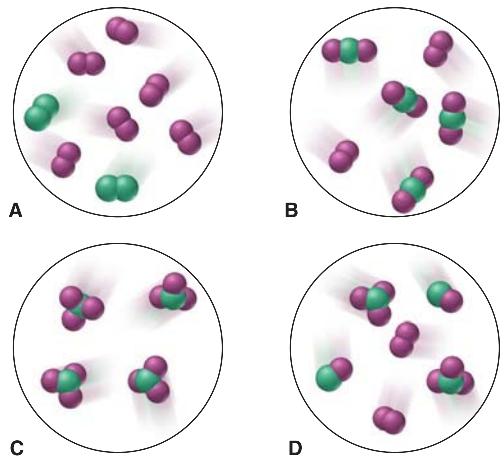
3.90 For the reaction between solid tetraphosphorus trisulfide and oxygen gas to form solid tetraphosphorus decoxide and sulfur dioxide gas, write a balanced equation. Show the equation (see Table 3.3) in terms of (a) molecules, (b) moles, and (c) grams.

3.91 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?

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

3.93 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

3.94 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?



3.95 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?

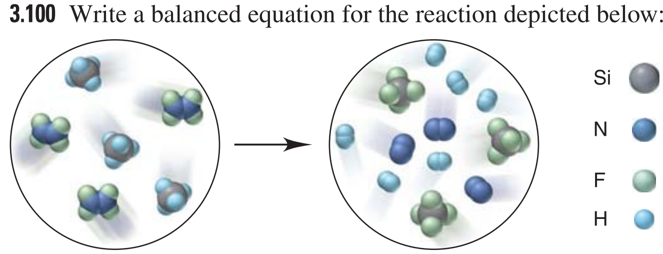
3.96 Nitrogen (N), phosphorus (P), and potassium (K) are the main nutrients in plant fertilizers. By industry convention, the numbers on a label refer to the mass percents of N, P_2O_5, and K_2O, in that order. Calculate the N/P/K ratio of a 30/10/10 fertilizer in terms of moles of each element, and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

3.97 What mass percents of ammonium sulfate, ammonium hydrogen phosphate, and potassium chloride would you use to prepare 10/10/10 plant fertilizer (see Problem 3.96)?

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

3.99 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?

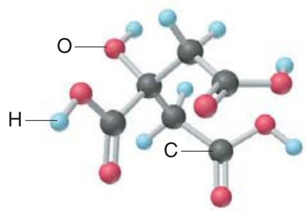
3.100 Write a balanced equation for the reaction depicted below:



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

3.101 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?

3.102 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?



3.103 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?

3.104 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?

3.105 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?

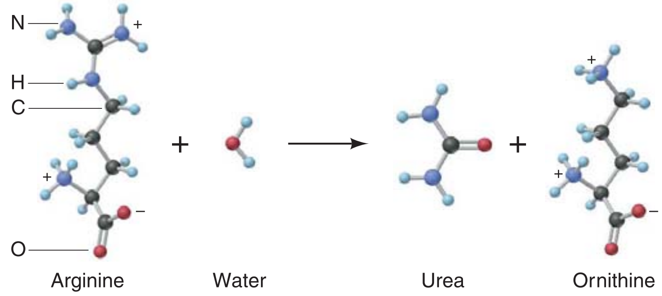
3.106 Hydroxyapatite, $Ca_5(PO_4)_3(OH)$, is the main mineral component of dental enamel, dentin, and bone, and thus has many medical uses. Coating it on metallic implants (such as titanium alloys and stainless steels) helps the body accept the implant. In the form of powder and beads, it is used to fill bone voids, which 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 of hydroxyapatite could form from 100. g of 85% phosphoric acid and 100. g of calcium hydroxide?

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

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

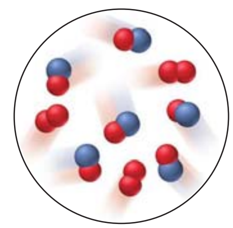
In one preparation, 3.077 g of salicylic acid and 5.50 mL of acetic anhydride react to form 3.281 g of aspirin. (a) Which is the limiting reactant (d of acetic anhydride = 1.080 g/mL)? (b) What is the percent yield of this reaction?

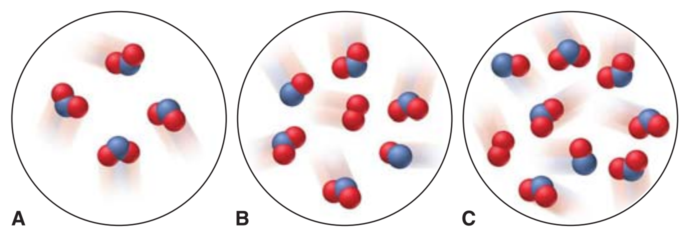
3.108 The human body excretes nitrogen in the form of urea, NH_2CONH_2. The key step in its biochemical formation is the reaction of water with arginine to produce urea and ornithine:



(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?

3.109 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?





3.110 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?

3.111 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?