3

Mass Relations in Chemistry; Stoichiometry

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Balanced chemical equations make it possible to relate masses of reactants and products. Scales are devices commonly used to measure mass.

Ex umbris et imaginibus in veritatem! (From shadows and symbols into the truth!)

—JOHN HENRY CARDINAL NEWMAN

Chapter Outline

To this point, our study of chemistry has been largely qualitative, involving few calculations. However, chemistry is a quantitative science. From Chapter 2 we see that atoms differ from one another not only in composition (number of protons, electrons, neutrons) but also in mass. Chemical formulas, which we learned to write and name, tell us not only the atom ratios in which elements are present but also mass ratios.

The general topic of this chapter is **stoichiometry** (stoy-key-OM-e-tree), the study of mass relations in chemistry. Whether dealing with molar mass (Section 3-1), chemical formulas (Section 3-2), or chemical reactions (Section 3-3), you will be answering questions that ask "how much?" or "how many?". For example,

- how many grams of acid are needed to prepare a solution required for an experiment? (Section 3-1)
- how much iron can be obtained from a ton of iron ore? (Section 3-2)
- how much nitrogen gas is required to form a kilogram of ammonia? (Section 3-3)

3-1 The Mole

People in different professions often use special counting units. You and I eat eggs one at a time, but farmers sell them by the dozen. We spend dollar bills one at a time, but Congress distributes them by the billion. Chemists have their own counting unit, Avogadro's number (Section 2-3).

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Figure 3.1 **One-mole quantities of sugar ($C_{12}H_{22}O_{11}$), baking soda ($NaHCO_3$), and copper nails (Cu).**

The quantity represented by Avogadro's number is so important that it is given a special name, the **mole.** It is mentioned briefly on page 31. A mole represents 6.022×10^{23} items, whatever they may be (Figure 3.1).

$$1 \text{ mol H atoms} = 6.022 \times 10^{23} \text{ H atoms}$$
$$1 \text{ mol O atoms} = 6.022 \times 10^{23} \text{ O atoms}$$
$$1 \text{ mol } H_2 \text{ molecules} = 6.022 \times 10^{23}\ H_2 \text{ molecules}$$
$$1 \text{ mol } H_2O \text{ molecules} = 6.022 \times 10^{23}\ H_2O \text{ molecules}$$
$$1 \text{ mol electrons} = 6.022 \times 10^{23} \text{ electrons}$$
$$1 \text{ mol pennies} = 6.022 \times 10^{23} \text{ pennies}$$

(One mole of pennies is a lot of money. It's enough to pay all the expenses of the United States for the next billion years or so without accounting for inflation.)

A mole represents not only a specific number of particles but also a definite mass of a substance as represented by its formula (O, O_2, H_2O, NaCl, . . .). ***The* molar mass**, MM, ***in grams per mole, is numerically equal to the sum of the masses (in amu) of the atoms in the formula.***

Formula	Sum of Atomic Masses	Molar Mass, MM
O	16.00 amu	16.00 g/mol
O_2	2(16.00 amu) = 32.00 amu	32.00 g/mol
H_2O	2(1.008 amu) + 16.00 amu = 18.02 amu	18.02 g/mol
NaCl	22.99 amu + 35.45 amu = 58.44 amu	58.44 g/mol

Notice that the formula of a substance must be known to find its molar mass. It would be ambiguous, to say the least, to refer to the "molar mass of hydrogen." One mole of hydrogen atoms, represented by the symbol H, weighs 1.008 g; the molar mass of H is 1.008 g/mol. One mole of hydrogen molecules, represented by the formula H_2, weighs 2.016 g; the molar mass of H_2 is 2.016 g/mol.

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Figure 3.2 **Aspirin (acetylsalicylic acid), $C_9H_8O_4$.**

Mole-Gram Conversions

As you will see later in this chapter, it is often necessary to convert from moles of a substance to mass in grams or vice versa. Such conversions are readily made by using the general relation

$$\textit{mass} = \text{MM} \times n \tag{3.1}$$

The mass is in grams, MM is the molar mass (g/mol), and n is the amount in moles. In effect, the molar mass, MM, is a conversion factor that allows you to calculate *moles* from *mass* or vice versa.

EXAMPLE 3.1

Acetylsalicylic acid (ASA), $C_9H_8O_4$, is the active ingredient of aspirin (Figure 3.2).

a What is the mass in grams of 0.509 moles of acetylsalicylic acid (ASA)?

ANALYSIS

Information given:	moles of acetylsalicylic acid (0.509) formula for acetylsalicylic acid ($C_9H_8O_4$)
Information implied:	molar mass (MM) of acetylsalicylic acid
Asked for:	mass of acetylsalicylic acid (ASA)

continued

STRATEGY

Substitute into Equation 3.1

mass = MM × n

SOLUTION

MM of $C_9H_8O_4$	$9(12.01) + 8(1.008) + 4(16.00) = 180.15$ g/mol
mass	$\text{mass} = \text{MM} \times n = 0.509 \text{ mol} \times \dfrac{180.15 \text{ g}}{1 \text{ mol}} = 91.7 \text{ g}$

b How many moles are present in 28.00 g of $C_9H_8O_4$?

ANALYSIS

Information given:	mass of ASA (28.00 g) formula for acetylsalicylic acid ($C_9H_8O_4$)
Information implied:	molar mass (MM) of acetylsalicylic acid
Asked for:	moles of ASA in the sample

STRATEGY

Substitute into Equation 3.1.

mass = MM × n

SOLUTION

moles ASA	$n = \dfrac{\text{mass}}{\text{MM}} = \dfrac{28.00 \text{ g}}{180.15 \text{ g/mole}} = 0.1554 \text{ mol ASA}$

c How many molecules of $C_9H_8O_4$ are there in 12.00 g of acetylsalicylic acid? How many carbon atoms?

ANALYSIS

Information given:	mass of ASA (12.00 g) formula for acetylsalicylic acid ($C_9H_8O_4$)
Information implied:	molar mass (MM) of acetylsalicylic acid Avogadro's number (N_A)
Asked for:	number of molecules of ASA number of carbon atoms

STRATEGY

Follow the plan outlined in Figure 3.3 (p. 54).

$$\text{mass} \xrightarrow{\text{MM}} \text{moles} \xrightarrow{N_A} \text{molecules} \xrightarrow{9 \text{ C atoms}} \text{C atoms}$$

SOLUTION

number of molecules	$12.00 \text{ g ASA} \times \dfrac{6.022 \times 10^{23} \text{ molecules ASA}}{180.15 \text{ g ASA}} = 4.011 \times 10^{22}$
number of C atoms	$4.011 \times 10^{22} \text{ molecules ASA} \times \dfrac{9 \text{ C atoms}}{1 \text{ molecule}} = 3.610 \times 10^{23}$

END POINT

The mass of the sample in grams is always larger than the number of moles because every known substance has a molar mass greater than 1 g/mol.

We have now considered several types of conversions involving numbers of particles, moles, and grams. They are summarized in Figure 3.3. Conversions of the type we have just carried out come up over and over again in chemistry. They will be required in nearly every chapter of this text. Clearly, you must know what is meant by a mole. Remember, a mole always represents a certain number

Almost every calculation that you will make in this chapter requires that you know precisely what is meant by a mole.

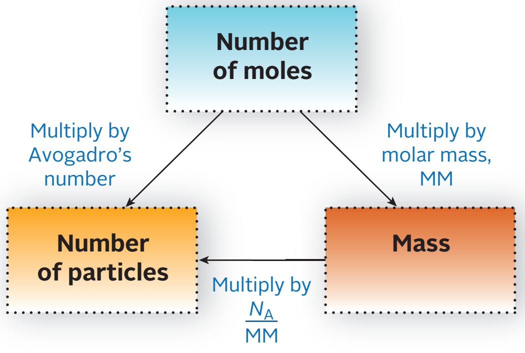

Figure 3.3 Flowchart for the interconversion of particles, mass, and moles.

of items, 6.022×10^{23}. Its mass, however, differs with the substance involved: A mole of H_2O, 18.02 g, weighs considerably more than a mole of H_2, 2.016 g, even though they both contain the same number of molecules. In the same way, a dozen bowling balls weigh a lot more than a dozen eggs, even though the number of items is the same for both.

Moles in Solution; Molarity

To obtain a given amount of a pure solid in the laboratory, you would weigh it out on a balance. Suppose, however, the solid is present as a solute dissolved in a solvent such as water. In this case, you ordinarily measure out a given volume of the water solution, perhaps using a graduated cylinder. The amount of solute you obtain in this way depends not only on the volume of solution but also on the **concentration** of solute, i.e., the amount of solute in a given amount of solution. The concentration of a solute in solution can be expressed in terms of its **molarity:**

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}} \tag{3.2}$$

The symbol [ ] is commonly used to represent the molarity of a species in solution. For a solution containing 1.20 mol of substance A in 2.50 L of solution,

$[NH_3]$ means "concentration of ammonia in moles per liter."

$$[\text{A}] = \frac{1.20 \text{ mol}}{2.50 \text{ L}} = 0.480 \text{ mol/L} = 0.480\ M$$

One liter of such a solution would contain 0.480 mol of A; 100 mL of solution would contain 0.0480 mol of A, and so on.

The molarity of a soluble solute can vary over a wide range. With sodium hydroxide, for example, we can prepare a 6 *M* solution, a 1 *M* solution, a 0.1 *M* solution, and so on. The words "concentrated" and "dilute" are often used in a qualitative way to describe these solutions. We would describe a 6 *M* solution of NaOH as concentrated; it contains a relatively large amount of solute per liter. A 0.1 *M* NaOH solution is dilute, at least in comparison to 1 *M* or 6 *M*.

To prepare a solution to a desired molarity, you first calculate the amount of solute required. This is then dissolved in enough solvent to form the required volume of solution. Suppose, for example, you want to make one liter of 0.100 *M* K_2CrO_4 solution (Figure 3.4). You first weigh out 19.4 g (0.100 mol) of K_2CrO_4 (MM = 194.20 g/mol). Then stir with enough water to form one liter (1000 mL) of solution.

The molarity of a solution can be used to calculate

- the number of moles of solute in a given volume of solution.
- the volume of solution containing a given number of moles of solute.

Here, as in so many other cases, a conversion factor approach is used (Example 3.2).

EXAMPLE 3.2

Nitric acid, HNO_3, is extensively used in the manufacture of fertilizer. A bottle containing 75.0 mL of nitric acid solution is labeled 6.0 *M* HNO_3.

a How many moles of HNO_3 are in the bottle?

ANALYSIS	
Information given:	*V* (75.0 mL) and *M* (6.0 *M*) of HNO_3 in the bottle
Information implied:	molar mass (MM) of HNO_3
Asked for:	moles of HNO_3 in the bottle

continued

STRATEGY

1. Do not forget to change the volume unit given (mL) to L.
2. Use the molarity of HNO_3 as a conversion factor:

$$\frac{6.0 \text{ mol } HNO_3}{1 \text{ L solution}}$$

SOLUTION

moles HNO_3	$75.0 \text{ mL solution} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{6.0 \text{ mol } HNO_3}{1 \text{ L solution}} = 0.45 \text{ mol}$

b A reaction needs 5.00 g of HNO_3. How many milliliters of solution are required?

ANALYSIS

Information given:	V (75.0 mL) and M (6.0 M) of HNO_3 in the bottle mass of HNO_3 required (5.00 g)
Information implied:	molar mass (MM) of HNO_3
Asked for:	volume of HNO_3 required

STRATEGY

1. Use the molarity and the molar mass of HNO_3 as conversion factors.

$$\frac{6.0 \text{ mol } HNO_3}{1 \text{ L solution}} \qquad \frac{1 \text{ mol } HNO_3}{63.02 \text{ g } HNO_3}$$

2. Follow the plan:

$$\text{mass} \xrightarrow{\text{MM}} \text{moles} \xrightarrow{M} \text{volume}$$

SOLUTION

volume HNO_3	$5.00 \text{ g } HNO_3 \times \frac{1 \text{ mol } HNO_3}{63.02 \text{ g } HNO_3} \times \frac{1 \text{ L solution}}{6.0 \text{ mol } HNO_3} = 0.013 \text{ L} = 13 \text{ mL}$

c Ten mL of water are added to the solution. What is the molarity of the resulting solution? (Assume volumes are additive.)

ANALYSIS

Information given:	V (75.0 mL) and M (6.0 M) of HNO_3 in the bottle volume of water added (10.00 mL) Assume volumes are additive.
Information implied:	molar mass (MM) of HNO_3
Asked for:	molarity (M) of diluted solution

STRATEGY

1. Adding water does not change the number of moles of solute. It does change the volume of solution.
2. Substitute into Equation 3.2. Use (75.0 mL + 10.0 mL) as the total volume.

SOLUTION

M	$\frac{0.45 \text{ mol}}{0.0750 \text{ L} + 0.0100 \text{ L}} = 5.3 \text{ mol/L} = 5.3\ M$

END POINT

The molarity of a solution decreases when water is added to the solution, but the moles of solute in solution remain the same.

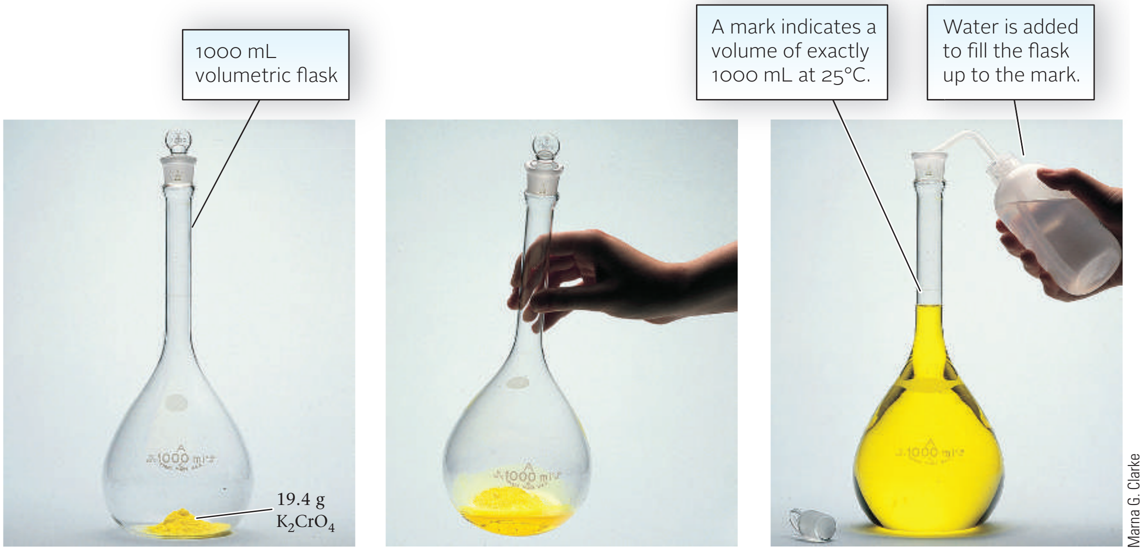

Figure 3.4 Preparing one liter of 0.100 *M* potassium chromate. A 0.100 *M* potassium chromate solution is made by adding enough water to 19.4 g of K_2CrO_4 to make one liter of solution. The weighed K_2CrO_4 (19.4 g) is transferred to a 1000-mL volumetric flask. Enough water is added to fully dissolve all of the solid by swirling. The final step is to shake the flask repeatedly until a homogeneous solution is formed.

Knowing the molarity of a solution, you can readily obtain a specified amount of solute. All you have to do is to calculate the required volume, as in Example 3.2b. Upon measuring out that volume, you should obtain the desired number of moles or grams of solute. Concentrations of reagents in the general chemistry laboratory are most often expressed in molarities. We will have more to say about molarity and other concentration units in Chapter 10.

As pointed out in Chapter 2, when an ionic solid dissolves in water, the cations and anions separate from each other. This process can be represented by a chemical equation in which the reactant is the solid and the products are the positive and negative ions in water (aqueous) solution. For the dissolving of $MgCl_2$, the equation is

$$MgCl_2(s) \longrightarrow Mg^{2+}(aq) + 2Cl^-(aq)$$

This equation tells us that one mole of $MgCl_2$ yields *one* mole of Mg^{2+} ions and *two* moles of Cl^- ions in solution. It follows that in any solution of magnesium chloride:

$$\text{molarity } Mg^{2+} = \text{molarity } MgCl_2 \qquad \text{molarity } Cl^- = 2 \times \text{molarity } MgCl_2$$

Similar relationships hold for ionic solids containing polyatomic ions (Table 2.2) or transition metal cations (Figure 3.5).

[ion] = ion subscript × [parent compound]

Figure 3.5 Charges of transition metal cations commonly found in aqueous solution.

3	4	5	6	7	8	9	10	11	12
			Cr^{3+}	Mn^{2+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^{2+}	Zn^{2+}
								Ag^+	Cd^{2+}
									Hg^{2+}

$$(NH_4)_3PO_4(s) \longrightarrow 3NH_4^+(aq) + PO_4^{3-}(aq)$$

$$\text{molarity } NH_4^+ = 3 \times \text{molarity } (NH_4)_3PO_4$$
$$\text{molarity } PO_4^{3-} = \text{molarity } (NH_4)_3PO_4$$

$$Cr_2(SO_4)_3(s) \longrightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq)$$

$$\text{molarity } Cr^{3+} = 2 \times \text{molarity } Cr_2(SO_4)_3$$
$$\text{molarity } SO_4^{2-} = 3 \times \text{molarity } Cr_2(SO_4)_3$$

For aluminum sulfate:
$[Al^{3+}] = 2[Al_2(SO_4)_3]$,
$[SO_4^{2-}] = 3[Al_2(SO_4)_3]$.

EXAMPLE 3.3

Potassium dichromate, $K_2Cr_2O_7$, is used in the tanning of leather. A flask containing 125 mL of solution is labeled 0.145 M $K_2Cr_2O_7$.

a What is the molarity of each ion in solution?

ANALYSIS

Information given:	volume, V (125 mL), and molarity, M (0.145 M), of solution
Information implied:	number of each ion in the parent compound
Asked for:	molarity of each ion in solution

STRATEGY

1. Distinguish between the parent compound and the ions.
2. Count the number of ions in the parent compound (recall that it is the subscript of the ion), and use it as a conversion factor.

$$\frac{2 \text{ mol } K^+}{1 \text{ mol } K_2Cr_2O_7} \qquad \frac{1 \text{ mol } Cr_2O_7^{2-}}{1 \text{ mol } K_2Cr_2O_7}$$

SOLUTION

$[K^+]$

$$\frac{0.145 \text{ mol } K_2Cr_2O_7}{1 \text{ L}} \times \frac{2 \text{ mol } K^+}{1 \text{ mol } K_2Cr_2O_7} = 0.290 \text{ mol/L} = \boxed{0.290\ M}$$

$[Cr_2O_7^{2-}]$

$$\frac{0.145 \text{ mol } K_2Cr_2O_7}{1 \text{ L}} \times \frac{1 \text{ mol } Cr_2O_7^{2-}}{1 \text{ mol } K_2Cr_2O_7} = 0.145 \text{ mol/L} = \boxed{0.145\ M}$$

b A sample containing 0.200 moles of K^+ is added to the $K_2Cr_2O_7$ solution. Assuming no volume change, what is the molarity of the new solution?

ANALYSIS

Information given:	volume, V (125 mL), and molarity, M (0.145 M), of solution moles of K^+ added (0.200)
Information implied:	number of each ion in the parent compound
Asked for:	molarity of $K_2Cr_2O_7$ after the addition of K^+ ions

STRATEGY

1. Use the molarity of K^+ found in (a) and substitute into Equation 3.2 to find the moles of K^+ initially.
2. Find the moles of K^+ after the addition.
3. Find $[K^+]$ again by substituting into Equation 3.2.
4. Find $[K_2Cr_2O_7]$ by relating $[K^+]$ to $K_2Cr_2O_7$.
5. The overall plan is:

$$[K^+_{\text{initial}}] \xrightarrow{V} (\text{mol } K^+)_{\text{initial}} \xrightarrow{+\ 0.200 \text{ mol } K^+} (\text{mol } K^+)_{\text{final}} \xrightarrow{V} [K^+]_{\text{final}} \xrightarrow{2\ K^+/K_2Cr_2O_7} [K_2Cr_2O_7]_{\text{final}}$$

continued

SOLUTION	
1. $(\text{mol K}^+)_{\text{initial}}$	$0.125 \text{ L} \times \dfrac{0.290 \text{ mol K}^+}{1 \text{ L}} = 0.03625$
2. $(\text{mol K}^+)_{\text{final}}$	$0.03625 + 0.200 = 0.236$
3. $[\text{K}^+]_{\text{final}}$	$\dfrac{0.236 \text{ mol K}^+}{0.125 \text{ L}} = 1.89 \text{ mol/L} = 1.89\ M$
4. $[\text{K}_2\text{Cr}_2\text{O}_7]_{\text{final}}$	$\dfrac{1.89 \text{ mol K}^+}{1 \text{ L}} \times \dfrac{1 \text{ mol K}_2\text{Cr}_2\text{O}_7}{2 \text{ mol K}^+} = 0.945 \text{ mol/L} = 0.945\ M$

END POINT

The concentration of K^+ should be twice that of $Cr_2O_7^{2-}$ in either solution. It is!

3-2 Mass Relations in Chemical Formulas

As you will see shortly, the formula of a compound can be used to determine the mass percents of the elements present. Conversely, if the percentages of the elements are known, the simplest formula can be determined. Knowing the molar mass of a molecular compound, it is possible to go one step further and find the molecular formula. In this section we will consider how these three types of calculations are carried out.

Jo Edkins, 2007. Used with permission.

Figure 3.6 Hematite ore.

Percent Composition from Formula

The **percent composition** of a compound is specified by citing the mass percents of the elements present. For example, in a 100-g sample of water there are 11.19 g of hydrogen and 88.81 g of oxygen. Hence the percentages of the two elements are

$$\frac{11.19 \text{ g H}}{100.00 \text{ g}} \times 100\% = 11.19\% \text{ H} \qquad \frac{88.81 \text{ g O}}{100.00 \text{ g}} \times 100\% = 88.81\% \text{ O}$$

We would say that the percent composition of water is 11.19% H, 88.81% O.

Knowing the formula of a compound, Fe_2O_3, you can readily calculate the mass percents of its constituent elements. It is convenient to start with one mole of compound (Example 3.4a). The formula of a compound can also be used in a straightforward way to find the mass of an element in a known mass of the compound (Example 3.4b).

EXAMPLE 3.4 GRADED

Metallic iron is most often extracted from hematite ore (Figure 3.6), which consists of iron(III) oxide mixed with impurities such as silicon dioxide, SiO_2.

a What are the mass percents of iron and oxygen in iron(III) oxide?

ANALYSIS	
Information given:	formula of the iron oxide (Fe_2O_3)
Information implied:	molar mass (MM) of Fe_2O_3
Asked for:	mass % of Fe and O in Fe_2O_3

continued

STRATEGY

1. To find the mass percent of Fe follow the plan outlined below. Start with one mole of Fe_2O_3.

$$n_{Fe_2O_3} \xrightarrow{\text{subscript}} n_{Fe} \xrightarrow{\text{MM Fe}} \text{mass Fe} \xrightarrow{\text{MM } Fe_2O_3} \%\ \text{element}$$

2. Find the mass percent of oxygen by difference. Note that the compound is made up of only two elements, Fe and O.

SOLUTION

mass % Fe	$1 \text{ mol } Fe_2O_3 \times \dfrac{2 \text{ mol Fe}}{1 \text{ mol } Fe_2O_3} \times \dfrac{55.85 \text{ g}}{1 \text{ mol Fe}} = 111.7 \text{ g Fe}$ $\dfrac{111.7 \text{ g Fe}}{159.7 \text{ g } Fe_2O_3} = 69.94\%$
mass % O	mass % O = 100% − mass % Fe = 100.00% − 69.94% = 30.06%

b How many grams of iron can be extracted from 1.000 kg of Fe_2O_3?

ANALYSIS

Information given:	mass of Fe_2O_3 (1.000 kg = 1.000×10^3 g)
Information implied:	from (a): mass % of Fe in Fe_2O_3 (69.94% = 69.94 g Fe/100.00 g Fe_2O_3)
Asked for:	mass of Fe obtained from 1.000 kg of Fe_2O_3

STRATEGY

mass of Fe = (mass Fe_2O_3)(massFe/100 g Fe_2O_3)

SOLUTION

mass Fe	$1.000 \times 10^3 \text{ g } Fe_2O_3 \times \dfrac{69.94 \text{ g Fe}}{100 \text{ g } Fe_2O_3} = 699.4 \text{ g Fe}$

c How many metric tons of hematite ore, 66.4% Fe_2O_3, must be processed to produce 1.000 kg of iron?

ANALYSIS

Information given:	mass % of Fe_2O_3 in the ore (66.4% = 66.4 g Fe_2O_3/100.0 g ore) mass of Fe needed (1.000 kg)
Information implied:	from (a): mass % of Fe in Fe_2O_3 (69.94% = 69.94 g Fe/100.00 g Fe_2O_3) factor for converting metric tons to grams.
Asked for:	mass of hematite needed to produce 1.000 kg of Fe

STRATEGY

1. Distingush between the mass % of Fe_2O_3 in the ore (66.4%) and the mass % of Fe in Fe_2O_3 (69.94%).
2. Start with 1000 g of Fe and use the mass percents as conversion factors to go from mass of Fe to mass of the ore.

$$\frac{69.94 \text{ g Fe}}{100.00 \text{ g } Fe_2O_3} \qquad \frac{66.4 \text{ g } Fe_2O_3}{100.00 \text{ g ore}}$$

3. Convert grams to metric tons.

   1 metric ton = 1×10^6 g

SOLUTION

mass of hematite needed	$1000 \text{ g Fe} \times \dfrac{100 \text{ g } Fe_2O_3}{69.94 \text{ g Fe}} \times \dfrac{100 \text{ g ore}}{66.4 \text{ g } Fe_2O_3} \times \dfrac{1 \text{ metric ton}}{1 \times 10^6 \text{ g}} = 2.15 \times 10^{-3} \text{ metric tons}$

The calculations in Example 3.4 illustrate an important characteristic of formulas. In one mole of Fe_2O_3 there are *two* moles of Fe (111.7 g) and *three* moles of O (48.00 g). This is the same as the atom ratio in Fe_2O_3, 2 atoms Fe : 3 atoms O. In general, ***the subscripts in a formula represent not only the atom ratio in which the different elements are combined but also the mole ratio.***

Can you explain why the mole ratio must equal the atom ratio?

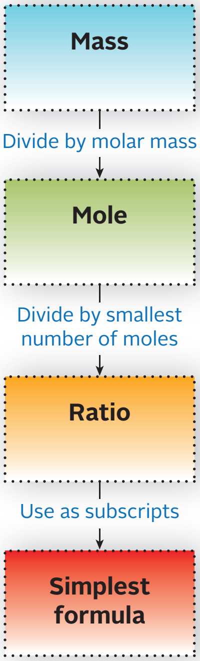

Figure 3.7 Flowchart for determining the simplest formula of a compound.

Simplest Formula from Chemical Analysis

A major task of chemical analysis is to determine the formulas of compounds. The formula found by the approach described here is the **simplest formula,** which gives the simplest whole-number ratio of the atoms present. For an ionic compound, the simplest formula is ordinarily the only one that can be written (e.g., $CaCl_2$, Cr_2O_3). For a molecular compound, the molecular formula is a whole-number multiple of the simplest formula, where that number may be 1, 2, ...

Compound	Simplest Formula	Molecular Formula	Multiple
Water	H_2O	H_2O	1
Hydrogen peroxide	HO	H_2O_2	2
Propylene	CH_2	C_3H_6	3

The analytical data leading to the simplest formula may be expressed in various ways. You may know

- the masses of the elements in a weighed sample of the compound.
- the mass percents of the elements in the compound.
- the masses of products obtained by reaction of a weighed sample of the compound.

The strategy used to calculate the simplest formula depends to some extent on which of these types of information is given. The basic objective in each case is to find the number of moles of each element, then the simplest mole ratio, and finally the simplest formula.

Figure 3.7 shows a schematic pathway for this process and Example 3.5 illustrates how a simple formula is obtained when the masses of the elements in the compound are given. The masses of the elements can also be calculated when the results of a combustion experiment (discussed on page 61) are given.

EXAMPLE 3.5

A 30.00-g sample of a white crystal used in water purification contains 4.731 g of aluminium, 8.436 g of sulfur, and 16.83 g of oxygen. Find the simplest formula of the compound.

ANALYSIS

Information given:	mass of sample (30.00 g) mass of each element in the compound: Al (4.731 g); S (8.436 g); O (16.83 g)
Information implied:	atomic masses of the elements
Asked for:	simplest formula for the compound

STRATEGY

Follow the flowchart for determining the simplest formula (Figure 3.7).

SOLUTION

moles of each element

$$\text{Al: } \frac{4.731 \text{ g Al}}{26.98 \text{ g/mol}} = 0.1753 \text{ mol}$$

$$\text{S: } \frac{8.436 \text{ g S}}{32.07 \text{ g/mol}} = 0.2630 \text{ mol}$$

$$\text{O: } \frac{16.83 \text{ g O}}{16.00 \text{ g/mol}} = 1.052 \text{ mol}$$

continued

ratios	Al: $\frac{0.1753 \text{ mol}}{0.1753 \text{ mol}} = 1$; S: $\frac{0.2630 \text{ mol}}{0.1753 \text{ mol}} = 1.5$ O: $\frac{1.052 \text{ mol}}{0.1753 \text{ mol}} = 6.0$ Since we need smallest *whole number* ratios, we multiply all ratios by 2. 2 Al : 3 S : 12 O
simplest formula	$Al_2S_3O_{12}$

END POINTS

1. A mole ratio of 1.00 A : 1.00 B : 3.33 C would imply a formula $A_3B_3C_{10}$. If the mole ratio were 1.00 A: 2.50 B : 5.50 C, the formula would be $A_2B_5C_{11}$. In general, multiply through by the smallest whole number that will give integers for all the subscripts.
2. Note that in this case, the mass of the sample was not used to find the simplest formula.

Sometimes you will be given the mass percents of the elements in a compound. If that is the case, one extra step is involved. ***Assume a 100-g sample and calculate the mass of each element in that sample.***

Suppose, for example, you are told that the percentages of K, Cr, and O in a compound are 26.6%, 35.4%, and 38.0%, respectively. It follows that in a 100.0-g sample there are

26.6 g K, 35.4 g Cr, and 38.0 g O

Working with these masses, you can go through the same procedure followed in Example 3.5 to arrive at the same answer. (Try it!)

The most complex problem of this type requires you to determine the simplest formula of a compound given only the raw data obtained from its analysis. Here, an additional step is involved; you have to determine the masses of the elements from the masses of the new compounds that are obtained from the experiment.

Simple organic compounds such as hexane (containing C and H only) or ethyl alcohol (containing C, H, and O) can be analyzed using the apparatus shown in Figure 3.8. A weighed sample of the compound is burned in oxygen in a process called *combustion*. In combustion, the C in the compound is converted to CO_2 while H is converted to H_2O. The masses of CO_2 and H_2O can be determined by measuring the increase in mass of the absorbers. If oxygen is originally present, its mass is determined by difference.

$$\text{mass O} = \text{mass of sample} - (\text{mass of C} + \text{mass of H})$$

To determine the mass of C (in CO_2) and H (in H_2O), recall that there is one atom of C (12.01 g/mol) in a molecule of CO_2 (44.01 g/mol) and two atoms of H (2×1.008 g/mol) in a molecule of H_2O (18.02 g/mol). Thus the following conversion factors apply:

$$\text{For C: } \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} \qquad \text{For H: } \frac{2(1.008) \text{ g H}}{18.02 \text{ g } H_2O}$$

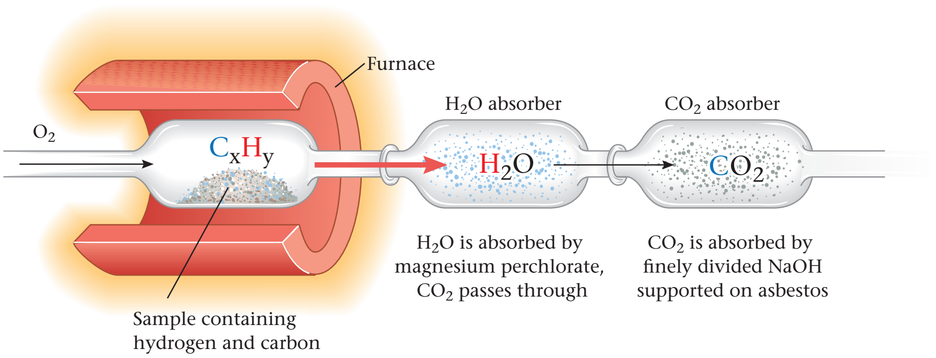

Figure 3.8 Combustion train used for carbon-hydrogen analysis. The absorbent for water is $Mg(ClO_4)_2$. Carbon dioxide is absorbed by finely divided NaOH. Alternatively, H_2O and CO_2 can be determined by gas phase chromatography.

EXAMPLE 3.6

The compound that gives vinegar its sour taste is acetic acid, which contains the elements carbon, hydrogen, and oxygen. When 5.00 g of acetic acid are burned in air, 7.33 g of CO_2 and 3.00 g of water are obtained. What is the simplest formula of acetic acid?

ANALYSIS

Information given:	elements in acetic acid (C, H, and O) mass of acetic acid (5.00 g) result of combustion analysis (7.33 g CO_2, 3.00 g H_2O)
Information implied:	molar masses (MM) of CO_2 and H_2O
Asked for:	simplest formula of acetic acid

STRATEGY

1. Find the mass of C and H by using the conversion factors:

$$\frac{12.01\text{ g C}}{44.01\text{ g CO}_2} \qquad \frac{2(1.008)\text{ g H}}{18.02\text{ g H}_2\text{O}}$$

2. Find the mass of O by difference:

mass of sample = mass of C + mass of H + mass of O

3. Follow the schematic pathway shown in Figure 3.7.

SOLUTION

1. mass of C	$7.33\text{ g CO}_2 \times \frac{12.01\text{ g C}}{44.01\text{ g CO}_2} = 2.00\text{ g C}$
mass of H	$3.00\text{ g H}_2\text{O} \times \frac{2(1.008)\text{ g H}}{18.02\text{ g H}_2\text{O}} = 0.336\text{ g H}$
2. mass of O	mass of O = mass of sample − (mass of C + mass of H) = 5.00 g − (2.00 g + 0.336 g) = 2.66 g
3. mol of each element	C: $\frac{2.00\text{ g C}}{12.01\text{ g/mol}} = 0.167$; H: $\frac{0.336\text{ g H}}{1.008\text{ g/mol}} = 0.333$; O: $\frac{2.66\text{ g O}}{16.00\text{ g/mol}} = 0.166$
ratios	C: $\frac{0.167\text{ mol}}{0.166\text{ mol}} = 1$; H: $\frac{0.333\text{ mol}}{0.167\text{ mol}} = 2$; O: $\frac{0.166\text{ mol}}{0.166\text{ mol}} = 1$
simplest formula	Using the ratios as subscripts, the simplest formula is CH_2O.

Molecular Formula from Simplest Formula

Chemical analysis always leads to the simplest formula of a compound because it gives only the simplest atom ratio of the elements. As pointed out earlier, the molecular formula is a whole-number multiple of the simplest formula. That multiple may be 1 as in H_2O, 2 as in H_2O_2, 3 as in C_3H_6, or some other integer. To find the multiple, one more piece of data is needed: the molar mass.

Sometimes it's not so easy to convert the atom ratio to simplest formula (see Problem 42 at the end of the chapter).

EXAMPLE 3.7

Vitamin C is found in many fruits and vegetables, particularly in citrus fruit. Its simplest formula is $C_3H_4O_3$ and its molar mass (determined by a mass spectrometer) is 176 g/mol. What is the molecular formula of citric acid?

continued

ANALYSIS	
Information given:	simplest formula ($C_3H_4O_3$) MM_a—actual molar mass (176 g/mol)
Asked for:	molecular formula

STRATEGY

1. Determine the molar mass of the simplest formula, MM_s.
2. Find the ratio

$$\frac{\text{actual molar mass}}{\text{simplest formula molar mass}} = \frac{MM_a}{MM_s}$$

3. To get the molecular formula, multiply all subscripts in the simplest formula by the ratio.

SOLUTION	
1. MM_s	$3(12.01)$ g C + $4(1.008)$ g H + $3(16.00)$ g O = 88.06 g $C_3H_4O_3$ g/mol
2. ratio	$\frac{MM_a}{MM_s} = \frac{176}{88.06} = 2$
3. molecular formula	$C_{3\times2}H_{4\times2}O_{3\times2}$ The molecular formula of Vitamin C is $C_6H_8O_6$.

3-3 Mass Relations in Reactions

A chemist who carries out a reaction in the laboratory needs to know how much **product** can be obtained from a given amount of starting material (**reactant**). To do this, he or she starts by writing a balanced **chemical equation**.

Writing and Balancing Chemical Equations

Chemical reactions are represented by chemical equations, which identify reactants and products. Formulas of reactants appear on the left side of the equation; those of products are written on the right. In a balanced chemical equation, there are the same number of atoms of a given element on both sides. The same situation holds for a chemical reaction that you carry out in the laboratory; atoms are conserved. For that reason, ***any calculation involving a reaction must be based on the balanced equation for that reaction.***

Beginning students are sometimes led to believe that writing a chemical equation is a simple, mechanical process. Nothing could be further from the truth. One point that seems obvious is often overlooked. ***You cannot write an equation unless you know what happens in the reaction that it represents.*** All the reactants and all the products must be identified. Moreover, you must know their formulas and physical states.

To illustrate how a relatively simple equation can be written, consider liquid hydrazine (N_2H_4). It was used to fuel the rocket engines for the Mars rover Curiosity, which landed on Mars in August 2012. It is still there (Figure 3.9). Hydrazine can react with liquid dinitrogen tetroxide to produce gaseous nitrogen and water vapor. To write a balanced equation, proceed as follows:

1. ***Write a "skeleton" equation in which the formulas of the reactants appear on the left and those of the products on the right.*** In this case,

$$N_2H_4 + N_2O_4 \longrightarrow N_2 + H_2O$$

2. ***Indicate the physical state of each reactant and product, after the formula, by writing***

(*g*) for a gaseous substance
(*l*) for a pure liquid
(*s*) for a solid
(*aq*) for an ion or molecule in water (aqueous) solution

NASA/JPL-Caltech/MSSS

Figure 3.9 Mars rover Curiosity.

Using the letters (*s*), (*l*), (*g*), (*aq*) lends a sense of physical reality to the equation.

In this case

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

3. ***Balance the equation.*** To accomplish this, start by writing a coefficient of 4 for H_2O, thus obtaining 4 oxygen atoms on both sides:

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

Now consider the hydrogen atoms. There are $4 \times 2 = 8$ H atoms on the right. To obtain 8 H atoms on the left, write a coefficient of 2 for N_2H_4:

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

Finally, consider nitrogen. There are a total of $(2 \times 2) + 2 = 6$ nitrogen atoms on the left. To balance nitrogen, write a coefficient of 3 for N_2:

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

This is the final balanced equation for the reaction of hydrazine with dinitrogen tetraoxide.

Three points concerning the balancing process are worth noting.

1. Equations are balanced by adjusting coefficients in front of formulas, never by changing subscripts within formulas. On paper, the equation discussed above could have been balanced by writing N_6 on the right, but that would have been absurd. Elemental nitrogen exists as diatomic molecules, N_2; there is no such thing as an N_6 molecule.

6N is no good either.

2. In balancing an equation, it is best to start with an element that appears in only one species on each side of the equation. In this case, either oxygen or hydrogen is a good starting point. Nitrogen would have been a poor choice, however, because there are nitrogen atoms in both reactant molecules, N_2H_4 and N_2O_4.

3. In principle, an infinite number of balanced equations can be written for any reaction. The equations

$$4N_2H_4(l) + 2N_2O_4(l) \longrightarrow 6N_2(g) + 8H_2O(g)$$
$$N_2H_4(l) + \tfrac{1}{2}N_2O_4(l) \longrightarrow \tfrac{3}{2}N_2(g) + 2H_2O(g)$$

are balanced in that there are the same number of atoms of each element on both sides. Ordinarily, the equation with the simplest whole-number coefficients

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

is preferred.

Frequently, when you are asked to balance an equation, the formulas of products and reactants are given. Sometimes, though, you will have to derive the formulas, given only the names (Example 3.8).

It's time to review Table 2.2.

EXAMPLE 3.8

Crystals of sodium hydroxide (lye) (Figure 3.10) react with carbon dioxide from air to form a colorless liquid, water, and a white powder, sodium carbonate, which is commonly added to detergents as a softening agent. Write a balanced equation for this chemical reaction.

STRATEGY

1. Translate names to formulas and write a "skeleton" equation. Recall the discussion in Sections 2-6 and 2-7.
2. Write the physical states.
3. Balance by starting with an element that appears only once on each side of the equation.
4. Check that you have the same number of atoms of each element on both sides of the equation.

continued

SOLUTION	
1. Skeleton equation	$NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$
2. Physical states	$NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$
3. Balance	$2NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$
4. Check H	2 on left 2 on right
Check C	1 on left 1 on right
Check Na	2 on left 2 on right
Check O	2 + 2 = 4 on left 3 + 1 = 4 on right
	The balanced equation is
	$2NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$

Mass Relations from Equations

The principal reason for writing balanced equations is to make it possible to relate the masses of reactants and products. Calculations of this sort are based on a very important principle:

> ***The coefficients of a balanced equation represent numbers of moles of reactants and products.***

To show that this statement is valid, recall the equation

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

The coefficients in this equation represent numbers of molecules, that is,

$$2 \text{ molecules } N_2H_4 + 1 \text{ molecule } N_2O_4 \longrightarrow 3 \text{ molecules } N_2 + 4 \text{ molecules } H_2O$$

A balanced equation remains valid if each coefficient is multiplied by the same number, including Avogadro's number, N_A:

$$2N_A \text{ molecules } N_2H_4 + 1N_A \text{ molecule } N_2O_4 \longrightarrow 3N_A \text{ molecules } N_2 + 4N_A \text{ molecules } H_2O$$

Recall from Section 3-1 that a mole represents Avogadro's number of items, N_A. Thus the equation above can be written

$$2 \text{ mol } N_2H_4 + 1 \text{ mol } N_2O_4 \longrightarrow 3 \text{ mol } N_2 + 4 \text{ mol } H_2O$$

The quantities 2 mol of N_2H_4, 1 mol of N_2O_4, 3 mol of N_2, and 4 mol of H_2O are chemically equivalent to each other in this reaction. Hence they can be used in conversion factors such as

$$\frac{2 \text{ mol } N_2H_4}{3 \text{ mol } N_2}, \frac{3 \text{ mol } N_2}{1 \text{ mol } N_2O_4}, \text{ or a variety of other combinations}$$

These are also called *stoichiometric ratios*. Use the coefficients of the balanced equations for the ratios.

If you wanted to know how many moles of hydrazine were required to form 1.80 mol of elemental nitrogen, the conversion would be

$$n_{N_2H_4} = 1.80 \text{ mol } N_2 \times \frac{2 \text{ mol } N_2H_4}{3 \text{ mol } N_2} = 1.20 \text{ mol } N_2H_4$$

To find the number of moles of nitrogen produced from 2.60 mol of N_2O_4,

$$n_{N_2} = 2.60 \text{ mol } N_2O_4 \times \frac{3 \text{ mol } N_2}{1 \text{ mol } N_2O_4} = 7.80 \text{ mol } N_2$$


Martin Walker

Figure 3.10 Sodium hydroxide pellets.

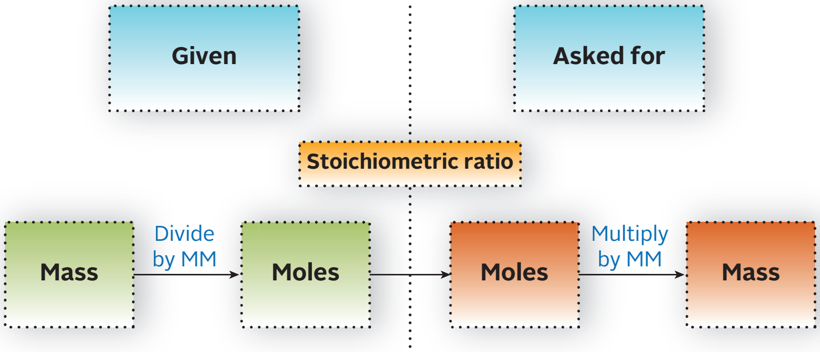

Figure 3.11 **Flowchart for mole-mass conversions.**

Simple mole relationships of this type are readily extended to relate moles of one substance to grams of another or grams of one substance to moles or molecules of another (Example 3.9). A schematic pathway that you can follow is shown in Figure 3.11.

EXAMPLE 3.9

Ammonia is used to make fertilizers (Figure 3.12) for lawns and gardens by reacting nitrogen gas with hydrogen gas. The balanced equation for the reaction is

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

STRATEGY

For all parts of this example, follow the schematic pathway shown in Figure 3.11.

a How many moles of ammonia are formed when 1.34 mol of nitrogen react?

ANALYSIS

Information given:	balanced equation: $[N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)]$ moles N_2 (1.34)
Asked for:	mol NH_3 formed

SOLUTION

mol NH_3	mol $H_2 \rightarrow$ mol NH_3 $1.34 \text{ mol } H_2 \times \dfrac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = \boxed{2.68 \text{ mol } NH_3}$

b How many grams of hydrogen are required to produce 2.75×10^3 g of ammonia?

ANALYSIS

Information given:	mass of ammonia (2.75×10^3 g) balanced equation: $[N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)]$
Information implied:	molar masses of NH_3 and H_2
Asked for:	mass of H_2 needed

SOLUTION

mass H_2	mass $NH_3 \rightarrow$ mol $NH_3 \rightarrow$ mol $H_2 \rightarrow$ mass H_2 $2.75 \times 10^3 \text{ g } NH_3 \times \dfrac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \dfrac{3 \text{ mol } H_2}{2 \text{ mol } NH_3} \times \dfrac{2.016 \text{ g } H_2}{1 \text{ mol } H_2} = \boxed{488 \text{ g } H_2}$

continued

c How many molecules of ammonia are formed when 2.92 g of hydrogen react?

ANALYSIS	
Information given:	mass of H_2 (2.92 g) balanced equation: $[N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)]$
Information implied:	molar mass of H_2 Avogadro's number (N_A)
Asked for:	molecules of NH_3 produced
SOLUTION	
molecules NH_3	mass $H_2 \rightarrow$ mol $H_2 \rightarrow$ mol $NH_3 \xrightarrow{N_A}$ molecules NH_3 $2.92\text{ g } H_2 \times \frac{1\text{ mol } H_2}{2.016\text{ g } H_2} \times \frac{2\text{ mol } NH_3}{3\text{ mol } H_2} \times \frac{6.022 \times 10^{23}\text{ molecules}}{1\text{ mol } NH_3} = \boxed{5.81 \times 10^{23}}$

d How many grams of ammonia are produced when 15.0 L of air (79% by volume nitrogen) react with an excess of hydrogen? The density of nitrogen at the conditions of the reaction is 1.25 g/L.

ANALYSIS	
Information given:	balanced equation: $[N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)]$ V_{air} (15.0 L); % by volume of N_2 in air (79%); density, d, of N_2 (1.25 g/L)
Information implied:	molar masses of N_2 and NH_3
Asked for:	mass of NH_3 produced
SOLUTION	
mass NH_3	$V_{air} \xrightarrow{\%\ N_2} V_{N_2} \xrightarrow{\text{density}}$ mass $N_2 \rightarrow$ mol $N_2 \rightarrow$ mol $NH_3 \rightarrow$ mass NH_3 $15.0\text{ L air} \times \frac{79\text{ L } N_2}{100\text{ L air}} \times \frac{1.25\text{ g } N_2}{1\text{ L } N_2} \times \frac{1\text{ mol } N_2}{28.02\text{ g } N_2} \times \frac{2\text{ mol } NH_3}{1\text{ mol } N_2}$ $\times \frac{17.03\text{ g } NH_3}{1\text{ mol } NH_3} = \boxed{18\text{ g } NH_3}$

Limiting Reactant and Theoretical Yield

When the two elements antimony and iodine are heated in contact with one another (Figure 3.13, p. 68), they react to form antimony(III) iodide.

$$2Sb(s) + 3I_2(s) \longrightarrow 2SbI_3(s)$$

The coefficients in this equation show that two moles of Sb (243.6 g) react with exactly three moles of I_2 (761.4 g) to form two moles of SbI_3 (1005.0 g). Put another way, the maximum quantity of SbI_3 that can be obtained under these conditions, assuming the reaction goes to completion and no product is lost, is 1005.0 g. This quantity is referred to as the **theoretical yield** of SbI_3.

Ordinarily, in the laboratory, reactants are not mixed in exactly the ratio required for reaction. Instead, an excess of one reactant, usually the cheaper one, is used. For example, 3.00 mol of Sb could be mixed with 3.00 mol of I_2. In that case, after the reaction is over, 1.00 mol of Sb remains unreacted.

$$\begin{aligned}\text{excess Sb} &= 3.00\text{ mol Sb originally} - 2.00\text{ mol Sb consumed}\\ &= 1.00\text{ mol Sb}\end{aligned}$$

The 3.00 mol of I_2 should be completely consumed in forming the 2.00 mol of SbI_3:

$$n_{SbI_3}\text{ formed} = 3.00\text{ mol } I_2 \times \frac{2\text{ mol } SbI_3}{3\text{ mol } I_2} = 2.00\text{ mol } SbI_3$$

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Figure 3.12 Fertilizer. The label 5-10-5 on the bag means that the fertilizer has 5% nitrogen, 10% phosphorus, and 5% potassium.

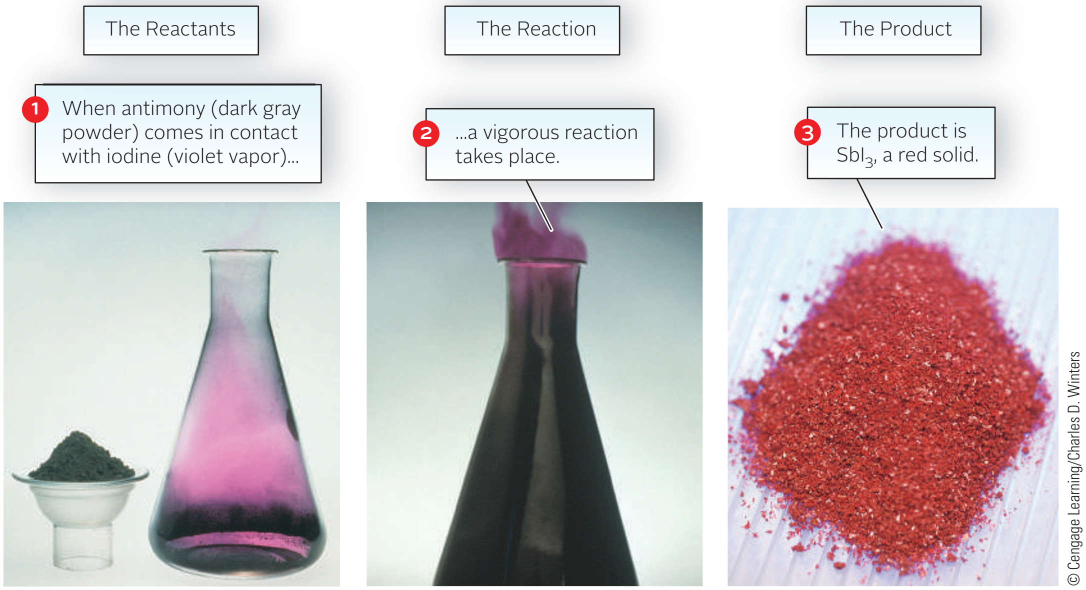

Figure 3.13 Reaction of antimony with iodine.

After the reaction is over, the solid obtained would be a mixture of product, 2.00 mol of SbI_3, with 1.00 mol of unreacted Sb.

In situations such as this, a distinction is made between the *excess reactant* (Sb) and the **limiting reactant,** I_2. The amount of product formed is determined (limited) by the amount of limiting reactant. With 3.00 mol of I_2, only 2.00 mol of SbI_3 is obtained, regardless of how large an excess of Sb is used.

Under these conditions, the theoretical yield of product is the amount produced if the limiting reactant is completely consumed. In the case just cited, the theoretical yield of SbI_3 is 2.00 mol, the amount formed from the limiting reactant, I_2.

Often you will be given the amounts of two different reactants and asked to determine which is the limiting reactant, to calculate the theoretical yield of the product and to find how much of the excess reactant is unused. To do so, it helps to follow a systematic, four-step procedure.

1. ***Calculate the amount of product that would be formed if the first reactant were completely consumed.***
2. ***Repeat this calculation for the second reactant; that is, calculate how much product would be formed if all of that reactant were consumed.***
3. ***Choose the smaller of the two amounts calculated in (1) and (2). This is the theoretical yield of product; the reactant that produces the smaller amount is the limiting reactant. The other reactant is in excess; only part of it is consumed.***
4. ***Take the theoretical yield of the product and determine how much of the reactant in excess is used up in the reaction. Subtract that from the starting amount to find the amount left.***

To illustrate how this procedure works, suppose you want to make grilled cheese sandwiches from 6 slices of cheese and 18 pieces of bread. The available cheese is enough for 6 grilled cheese sandwiches; the bread is enough for 9. Clearly, the cheese is the limiting reactant; there is an excess of bread. The theoretical yield is 6 sandwiches. Six grilled cheese sandwiches use up 12 slices of bread. Since there are 18 pieces available, 6 pieces of bread are left over.

Grilled cheese sandwiches are the favorite comfort food of WLM (William L. Masterton) and CNH.

EXAMPLE 3.10

Consider the reaction

$$2Sb(s) + 3I_2(s) \longrightarrow 2SbI_3(s)$$

a What is the limiting reactant and theoretical yield when 1.20 mol of Sb and 2.40 mol of I_2 are mixed?

ANALYSIS

Information given:	moles of each reactant: Sb (1.20), I_2 (2.40) balanced equation: $[2Sb(s) + 3I_2(s) \rightarrow 2SbI_3(s)]$
Asked for:	limiting reactant theoretical yield

STRATEGY

1. Find mol SbI_3 produced by first assuming Sb is limiting, and then assuming I_2 is limiting.

   mol Sb $\rightarrow$ mol SbI_3;   mol $I_2 \rightarrow$ mol SbI_3

2. The reactant that gives the smaller amount of SbI_3 is limiting, and the smaller amount of SbI_3 is the theoretical yield.

SOLUTION

1. mol SbI_3	$1.20 \text{ mol Sb} \times \frac{2 \text{ mol SbI}_3}{2 \text{ mol Sb}} = 1.20 \text{ mol}$ $2.40 \text{ mol I}_2 \times \frac{2 \text{ mol SbI}_3}{3 \text{ mol I}_2} = 1.60 \text{ mol}$
2. limiting reactant	1.20 mol (Sb limiting) < 1.60 mol (I_2 limiting) The limiting reactant is Sb.
theoretical yield	1.20 mol < 1.60 mol The theoretical yield is 1.20 mol SbI_3

b What is the limiting reactant and theoretical yield when 1.20 g of Sb and 2.40 g of I_2 are mixed? What mass of excess reactant is left when the reaction is complete?

ANALYSIS

Information given:	mass of each reactant: Sb (1.20 g), I_2 (2.40 g) balanced equation: $[2Sb(s) + 3I_2(s) \rightarrow 2SbI_3(s)]$
Information implied:	molar masses (MM) of SbI_3 and I_2
Asked for:	limiting reactant theoretical yield mass of excess reactant not used up

STRATEGY

1. Follow the plan outlined in Figure 3.11 and convert mass of Sb and mass of I_2 to mol SbI_3.
2. The smaller number of moles SbI_3 obtained is the theoretical yield. The reactant that yields the smaller amount is the limiting reactant.
3. Convert moles limiting reactant to mass of excess reactant. That is the mass of excess reactant consumed in the reaction.
4. Mass of excess reactant not used up = mass of excess reactant initally − mass excess reactant consumed

SOLUTION

1. mol SbI_3	$1.20 \text{ g Sb} \times \frac{1 \text{ mol Sb}}{121.8 \text{ g Sb}} \times \frac{2 \text{ mol SbI}_3}{2 \text{ mol Sb}} = 0.00985 \text{ mol SbI}_3$ $2.40 \text{ g I}_2 \times \frac{1 \text{ mol I}_2}{253.8 \text{ g I}_2} \times \frac{2 \text{ mol SbI}_3}{3 \text{ mol I}_2} = 0.006304 \text{ mol SbI}_3$

continued

2. limiting reactant theoretical yield	0.006304 mol (I_2 limiting) < 0.00985 mol (Sb limiting); thus I_2 is the limiting reactant. 0.006304 mol < 0.00985 mol The theoretical yield is 0.006304 mol (3.17 g) SbI_3. The reactant in excess is Sb.
3. mass Sb used up	$2.40 \text{ g } I_2 \times \frac{1 \text{ mol } I_2}{253.8 \text{ g } I_2} \times \frac{2 \text{ mol Sb}}{3 \text{ mol } I_2} \times \frac{121.8 \text{ g Sb}}{1 \text{ mol Sb}} = 0.768 \text{ g Sb}$
4. mass unreacted	mass unreacted = mass present initially − mass used up = 1.20 g − 0.768 g = 0.43 g

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Figure 3.14 Percent yield. If you started with 20 popcorn kernels, but only 16 of them popped, the percent yield of popcorn from this "reaction" would be 16/20 × 100% = 80%.

Remember that in deciding on the theoretical yield of product, you ***choose the smaller of the two calculated amounts.*** To see why this must be the case, refer back to Example 3.10b. There, 1.20 g of Sb was mixed with 2.40 g of I_2. Calculations show that the theoretical yield of SbI_3 is 3.17 g, and 0.43 g of Sb is left over. Thus

$$1.20 \text{ g Sb} + 2.40 \text{ g } I_2 \longrightarrow 3.17 \text{ g } SbI_3 + 0.43 \text{ g Sb}$$

This makes sense: 3.60 g of reactants yield a total of 3.60 g of products, including the unreacted antimony. Suppose, however, that 4.95 g of SbI_3 were chosen as the theoretical yield. The following nonsensical situation would arise.

$$1.20 \text{ g Sb} + 2.40 \text{ g } I_2 \longrightarrow 4.95 \text{ g } SbI_3$$

This violates the law of conservation of mass; 3.60 g of reactants cannot form 4.95 g of product.

Experimental Yield; Percent Yield

The theoretical yield is the maximum amount of product that can be obtained (Figure 3.14). In calculating the theoretical yield, it is assumed that the limiting reactant is 100% converted to product. In the real world, that is unlikely to happen. Some of the limiting reactant may be consumed in competing reactions. Some of the product may be lost in separating it from the reaction mixture. For these and other reasons, the **experimental yield** is ordinarily less than the theoretical yield. Put another way, the **percent yield** is expected to be less than 100%:

You never do quite as well as you hoped, so you never get 100%.

Remember the 6 grilled cheese sandwiches on page 68? If your dog ate one of them while you weren't looking, your yield would be 83%.

$$\text{percent yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\% \qquad (3.3)$$

EXAMPLE 3.11

Consider the reaction between magnesium and iodine to produce magnesium iodide. A student determines by calculation that with the amount of reagents he has on hand, he should produce 1.71 mol of MgI_2. After the reaction is complete, he tells his TA that he got a yield of 84.5%. How many grams of MgI_2 did he make?

ANALYSIS

Information given:	theoretical yield (1.71 mol) percent yield (84.5%)
Asked for:	mass MgI_2 actually obtained

continued

STRATEGY

1. Substitute into Equation 3.3.

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

2. Your answer will be the actual yield in moles. Convert to grams.

SOLUTION

1. actual yield	$84.5\% = \dfrac{\text{actual yield}}{1.71\ \text{mol}} \times 100\%$; actual yield $= 1.44$ mol MgI_2
2. mass MgI_2	$1.44\ \text{mol}\ MgI_2 \times \dfrac{278.1\ \text{g}\ MgI_2}{1\ \text{mol}\ MgI_2} = \boxed{402\ \text{g}}$

END POINT

If your actual yield is larger than your theoretical yield, something's wrong!

CHEMISTRY **BEYOND** THE CLASSROOM

Hydrates

Ionic compounds often separate from water solution with molecules of water incorporated into the solid. Such compounds are referred to as **hydrates**. An example is hydrated copper sulfate, which contains five moles of H_2O for every mole of $CuSO_4$. Its formula is $CuSO_4 \cdot 5H_2O$; a dot is used to separate the formulas of the two compounds $CuSO_4$ and H_2O. A Greek prefix is used to show the number of moles of water; the systematic name of $CuSO_4 \cdot 5H_2O$ is copper(II) sulfate pentahydrate.

Certain hydrates, notably $Na_2CO_3 \cdot 10H_2O$ and $FeSO_4 \cdot 7H_2O$, lose all or part of their water of hydration when exposed to dry air. This process is referred to as **efflorescence**; the glassy (vitreous) hydrate crystals crumble to a powder. Frequently, dehydration is accompanied by a color change (Figure A). When $CoCl_2 \cdot 6H_2O$ is exposed to dry air or is heated, it loses water and changes color from red to purple or blue. Crystals of this compound are used as humidity indicators and as an ingredient of invisible ink. Writing becomes visible only when the paper is heated, driving off water and leaving a blue residue.

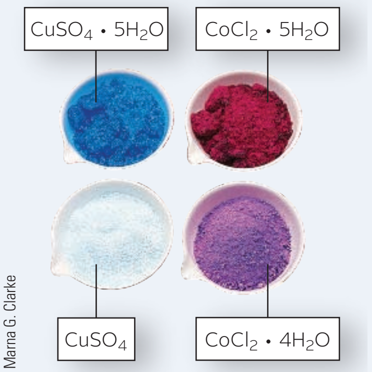


Figure A Hydrates of copper(II) sulfate and cobalt(II) chloride. $CuSO_4 \cdot 5H_2O$ (*upper left*) is blue; the anhydrous solid—that is, copper sulfate without the water—is white (*lower left*). $CoCl_2 \cdot 5H_2O$ (*upper right*) is reddish pink. Lower hydrates of cobalt such as $CoCl_2 \cdot 4H_2O$ are purple (*lower right*) or blue.

Molecular as well as ionic substances can form hydrates, but of an entirely different nature. In these crystals, sometimes referred to as *clathrates,* a molecule (such as CH_4, $CHCl_3$) is quite literally trapped in an ice-like cage of water molecules. Perhaps the best-known molecular hydrate is that of chlorine, which has the approximate composition $Cl_2 \cdot 7.3H_2O$. This compound was discovered by the great English physicist and electrochemist Michael Faraday in 1823. You can make it by bubbling chlorine gas through calcium chloride solution at 0°C; the hydrate comes down as feathery white crystals. In the winter of 1914, the German army used chlorine in chemical warfare on the Russian front against the soldiers of the Tsar. They were puzzled by its ineffectiveness; not until spring was deadly chlorine gas liberated from the hydrate, which is stable at cold temperatures.

In the 1930s when high-pressure natural gas (95% methane) pipelines were being built in the United States, it was found that the lines often became plugged in cold weather by a white,

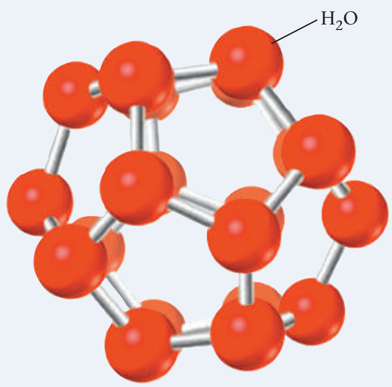


Figure B In one of the cages within which gas molecules are trapped in methane hydrate, water molecules form a pentagonal dodecahedron, a three-dimensional figure in which each of the 12 sides is a regular pentagon.

waxy solid that contained both water and methane (CH_4) molecules. Twenty years later, Walter Claussen at the University of Illinois deduced the structure of that solid, a hydrate of methane. Notice (Figure B) that CH_4 molecules are trapped within a three-dimensional cage of H_2O molecules.

In 1970 a huge deposit of methane hydrate was discovered at the bottom of the Atlantic Ocean, 330 km off the coast of North Carolina. The white solid was stable at the high pressure and low temperature (slightly above 0°C) that prevail at the ocean floor. When raised to the surface, the solid decomposed to give off copious amounts of methane gas (Figure C).

We now know that methane hydrate is widely distributed through the earth's oceans and the permafrost of Alaska and Siberia. The methane in these deposits could, upon combustion, produce twice as much energy as all the world's known resources of petroleum, natural gas, and coal. However, extracting the methane from undersea deposits has proved to be an engineering nightmare. Separating the hydrate from the mud, silt, and rocks with which it is mixed is almost as difficult as controlling its decomposition. Then there is the problem of transporting the methane to the shore, which may be 100 miles or more away.

John Pinkston and Laura Stern/Science News/U.S. Geological Survey

Figure C Methane hydrate is sometimes referred to as "the ice that burns."

Key Concepts

1. Use molar mass to relate
   - moles to mass of a substance
   (Example 3.1; Problems 1–14)
   - moles in solution; molarity
   (Examples 3.2, 3.3; Problems 15–24)
   - molecular formula to simplest formula
   (Example 3.7; Problems 43, 44)
2. Use the formula of a compound to find percent composition or its equivalent.
   (Example 3.4; Problems 25–34)
3. Find the simplest formula of a compound by chemical analysis.
   (Examples 3.5, 3.6; Problems 35–46)
4. Balance chemical equations by inspection.
   (Example 3.8; Problems 47–52)
5. Use a balanced equation to
   - relate masses of products and reactants
   (Example 3.9; Problems 53–64)
   - find the limiting reactant, theoretical yield, and percent yield
   (Examples 3.10, 3.11; Problems 65–72)

Key Equations

Molar mass $\quad \text{mass} = \text{MM} \times n$

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

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

Key Terms

limiting reactant
molar mass
molarity
mole
percent composition
percent yield
simplest formula
theoretical yield

Summary Problem

Consider oxygen gas, which is essential to our survival.

(a) How many moles are there in 19.00 g of oxygen gas?

(b) What is the mass in grams of 0.138 moles of oxygen gas?

(c) When iron is burned in oxygen (a process called "roasting"), iron(III) is formed. Write a balanced equation for this reaction.

(d) What is the molarity of a 485-mL solution made up of 0.384 g of iron(III) oxide and water?

(e) How many grams of oxygen will react with 17.28 g of iron?

(f) If 15.00 g of iron react with an excess of oxygen, how many grams of product are formed?

(g) How many grams of oxygen are required to react with an excess of iron to produce enough iron(III) oxide to make a 300.0 mL solution 0.0744 *M*?

(h) If 28.00 g of iron react with 28.00 g of oxygen, how many grams of product are formed? How much (in grams) of the reactant in excess is left?

(i) When 10.00 g of oxygen react with iron, the amount of iron(III) oxide actually obtained is 28.14 g. What is the percent yield?

(j) When oxygen is bubbled through a solution of sodium polysulfide, photographic hypo is produced. If "hypo" is made up of sodium, sulfur, and oxygen and contains 29.08% Na and 40.57% S, what is its simplest formula?

Answers

(a) **0.5938 mol**
(b) **4.42 g**
(c) **4 Fe(*s*) + 3O$_2$(*g*) $\longrightarrow$ 2Fe$_2$O$_3$(*s*)**
(d) **0.00496 *M***
(e) **7.426 g**
(f) **21.45 g**
(g) **1.07 g**
(h) **40.04 g Fe$_2$O$_3$; 15.97 g O$_2$ left**
(i) **84.58%**
(j) **Na$_2$S$_2$O$_3$**

Questions and Problems

Even-numbered questions and Challenge Problems have answers in Appendix 5 and fully worked solutions in the *Student Solutions Manual*.

3-1 The Mole

The Mole Molar Mass, and Mole-Gram Conversions

1. One chocolate chip used in making chocolate chip cookies has a mass of 0.324 g.
(a) How many chocolate chips are there in one mole of chocolate chips?
(b) If a cookie needs 15 chocolate chips, how many cookies can one make with a billionth (1×10^{-9}) of a mole of chocolate chips? (A billionth of a mole is scientifically known as a *nanomole*.)

2. The meat from one hazelnut has a mass of 0.985 g.
(a) What is the mass of a millionth of a mole (10^{-6}) of hazelnut meats? (A millionth of a mole is also called a *micromole*.)
(b) How many moles are in a pound of hazelnut meats?

3. Determine
(a) the mass of 0.429 mol of gold.
(b) the number of atoms in 0.715 g of gold.
(c) the number of moles of electrons in 0.336 g of gold.

4. How many electrons are in
(a) an ion of Sc^{3+}?
(b) a mol of Sc^{3+}?
(c) a gram of Sc^{3+}?

5. A cube of sodium has length 1.25 in. How many atoms are in that cube? (Note: $d_{Na} = 0.968$ g/cm^3.)

6. A solid circular cone made of pure platinum (d = 21.45 g/cm^3) has a diameter of 2.75 cm and a height of 3.00 inches.
(a) How many moles of platinum are there in the cone?
(b) How many electrons are there in the cone?

7. Calculate the molar masses (in grams per mole) of
(a) cane sugar, $C_{12}H_{22}O_{11}$.
(b) laughing gas, N_2O.
(c) vitamin A, $C_{20}H_{30}O$.

8. Calculate the molar mass (in grams/mol) of
(a) osmium metal, the densest naturally occurring element.
(b) baking soda, $NaHCO_3$.
(c) vitamin D, $C_{28}H_{44}O$, required for healthy bones and teeth.

9. Convert the following to moles.
(a) 4.00×10^3 g of hydrazine, a rocket propellant
(b) 12.5 g of tin(II) fluoride, the active ingredient in fluoride toothpaste
(c) 13 g of caffeine, $C_4H_5N_2O$

10. Convert to moles.
(a) 128.3 g of sucralose, $C_{12}H_{19}O_8Cl_3$, the active ingredient of the artificial sweetener Splenda™
(b) 0.3066 g of uric acid, $C_5H_4N_4O_3$, the compound that can cause gout and arthritis
(c) 2.664 g of cadmium(II) telluride used to coat solar panels

11. Calculate the mass in grams of 3.839 moles of
(a) hydrazine, a rocket propellant.
(b) caffeine, $C_4H_5N_2O$.
(c) theobromine, $C_7H_8N_4O_2$, the alkaloid present in chocolate and cocoa.

12. Calculate the mass in grams of 1.35 mol of
(a) titanium white, TiO_2, used as a paint pigment.
(b) sucralose, $C_{12}H_{19}O_8Cl_3$, the active ingredient in the artificial sweetener, Splenda™.
(c) strychnine, $C_{21}H_{22}N_2O_2$, present in rat poison.

13. Complete the following table for TNT (trinitrotoluene), $C_7H_5(NO_2)_3$.

	Number of Grams	Number of Moles	Number of Molecules	Number of N Atoms
(a)	127.2	______	______	______
(b)	______	0.9254	______	______
(c)	______	______	1.24×10^{28}	______
(d)	______	______	______	7.5×10^{22}

14. Complete the following table for citric acid, $C_6H_8O_7$, the acid found in many citrus fruits.

	Number of Grams	Number of Moles	Number of Molecules	Number of O Atoms
(a)	0.1364	______	______	______
(b)	______	1.248	______	______
(c)	______	______	4.32×10^{22}	______
(d)	______	______	______	5.55×10^{19}

Moles in Solution

15. Household ammonia used for cleaning contains about 10 g (two significant figures) of NH_3 in 100 mL (two significant figures) of solution. What is the molarity of the NH_3 in solution?

16. Household bleach contains about 5.2 g of sodium hypochlorite in 100 g (two significant figures) of water. What is the molarity of sodium hypochlorite in bleach? (Assume no volume change.)

17. What is the molarity of each ion present in aqueous solutions of the following compounds prepared by dissolving 20.00 g of each compound in water to make 4.50 L of solution?
(a) cobalt(III) chloride
(b) nickel(III) sulfate
(c) sodium permanganate
(d) iron(II) bromide

18. What is the molarity of each ion present in aqueous solutions of the following compounds prepared by dissolving 28.0 g of each compound in water to make 785 mL of solution?
(a) potassium oxide
(b) sodium hydrogen carbonate
(c) scandium(III) iodite
(d) magnesium phosphate

19. How would you prepare from the solid and pure water
(a) 0.400 L of 0.155 M $Sr(OH)_2$?
(b) 1.75 L of 0.333 M $(NH_4)_2CO_3$?

20. Starting with the solid and adding water, how would you prepare 2.00 L of 0.685 M
(a) $Ni(NO_3)_2$? (b) $CuCl_2$? (c) $C_6H_8O_6$ (vitamin C)?

21. You are asked to prepare a 0.8500 M solution of aluminum nitrate. You find that you have only 50.00 g of the solid.
(a) What is the maximum volume of solution that you can prepare?
(b) How many milliliters of this prepared solution are required to furnish 0.5000 mol of aluminum nitrate to a reaction?
(c) If 2.500 L of the prepared solution are required, how much more aluminum nitrate would you need?
(d) Fifty milliliters of a 0.450 M solution of aluminum nitrate are needed. How would you prepare the required solution from the solution prepared in (a)?

22. An experiment calls for a 0.4500 M solution of aluminum carbonate. The storeroom has only 32.00 g of the solid.
(a) What is the maximum volume of solution that can be prepared?
(b) The reaction in the experiment requires 0.1450 mol of carbonate ion. How many milliliters of this solution are required to provide 0.1450 mol?

23. A student combines two solutions of KOH and determines the molarity of the resulting solution. He records the following data:

Solution I:	30.00 mL of 0.125 M KOH
Solution II:	40.00 mL of KOH
Solution I + Solution II:	70.00 mL of 0.203 M KOH

What is the molarity of KOH in Solution II?

24. Twenty-five mL of a 0.388 M solution of Na_2SO_4 is mixed with 35.3 mL of 0.229 M Na_2SO_4. What is the molarity of the resulting solution? Assume that the volumes are additive.

3-2 Mass Relations in Chemical Formulas

25. Turquoise has the following chemical formula: $CuAl_6(PO_4)_4(OH)_8 \cdot 4H_2O$. Calculate the mass percent of each element in turquoise.

26. Diazepam is the addictive tranquilizer also known as Valium®. Its simplest formula is $C_{16}H_{13}N_2OCl$. Calculate the mass percent of each element in this compound.

27. Small amounts of tungsten (W) are usually added to steel to strengthen and harden the steel. Two ores of tungsten are ferberite ($FeWO_4$) and scheelite ($CaWO_4$). How many grams of $FeWO_4$ would contain the same mass of tungsten that is present in 725 g of $CaWO_4$?

28. Allicin is responsible for the distinctive taste and odor of garlic. Its simple formula is $C_6H_{10}O_2S$. How many grams of sulfur can be obtained from 25.0 g of allicin?

29. The active ingredient in Pepto-Bismol® (an over-the-counter remedy for an upset stomach) is bismuth subsalicylate, $C_7H_5BiO_4$. Analysis of a 1.7500-g sample of Pepto-Bismol yields 346 mg of bismuth. What percent by mass is bismuth subsalicylate in the sample? (Assume that there are no other bismuth-containing compounds in Pepto-Bismol.)

30. The active ingredient in some antiperspirants is aluminum chlorohydrate, $Al_2(OH)_5Cl$. Analysis of a 2.000-g sample of antiperspirant yields 0.334 g of aluminum. What percent (by mass) of aluminum chlorohydrate is present in the antiperspirant? (Assume that there are no other compounds containing aluminum in the antiperspirant.)

31. Combustion analysis of 1.00 g of the male sex hormone, testosterone, yields 2.90 g of CO_2 and 0.875 g H_2O. What are the mass percents of carbon, hydrogen, and oxygen in testosterone?

32. Hexachlorophene, a compound made up of atoms of carbon, hydrogen, chlorine, and oxygen, is an ingredient in germicidal soaps. Combustion of a 1.000-g sample yields 1.407 g of carbon dioxide, 0.134 g of water, and 0.523 g of chlorine gas. What are the mass percents of carbon, hydrogen, oxygen, and chlorine in hexachlorophene?

33. A compound NiX_3 is 19.67% (by mass) nickel. What is the molar mass of the compound? What is the name of the compound?

34. A compound R_2O_3 is 32.0% oxygen. What is the molar mass of R_2O_3? What is the element represented by R?

35. Manganese reacts with fluorine to form a fluoride. If 1.873 g of manganese reacts with enough fluorine to form 3.813 g of manganese fluoride, what is the simplest formula of the fluoride? Name the fluoride.

36. Nickel reacts with sulfur to form a sulfide. If 2.986 g of nickel reacts with enough sulfur to form 5.433 g of nickel sulfide, what is the simplest formula of the sulfide? Name the sulfide.

37. Determine the simplest formulas of the following compounds:

(**a**) the food enhancer monosodium glutamate (MSG), which has the composition 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na

(**b**) zircon, a diamond-like mineral, which has the composition 34.91% O, 15.32% Si, and 49.76% Zr

(**c**) nicotine, which has the composition 74.0% C, 8.65% H, and 17.4% N

38. Determine the simplest formulas of the following compounds:

(**a**) tetraethyl lead, the banned gasoline anti-knock additive, which is composed of 29.71% C, 6.234% H, and 64.07% Pb

(**b**) citric acid, present in most sour fruit, which is composed of 37.51% C, 4.20% H, and 58.29% O

(**c**) cisplatin, a drug used in chemotherapy, which is composed of 9.34% N, 2.02% H, 23.36% Cl, and 65.50% Pt

39. Nicotine is found in tobacco leaf and is mainly responsible for the addictive property of cigarette smoking. Nicotine is made up of carbon, hydrogen, and nitrogen atoms. When a 2.500-g sample of nicotine is burned, 6.782 g of CO_2 and 1.944 g of H_2O are obtained. What is the simplest formula for nicotine?

40. Methyl salicylate is a common "active ingredient" in liniments such as Ben-Gay™. It is also known as oil of wintergreen. It is made up of carbon, hydrogen, and oxygen atoms. When a sample of methyl salicylate weighing 5.287 g is burned in excess oxygen, 12.24 g of carbon dioxide and 2.505 g of water are formed. What is the simplest formula for oil of wintergreen?

41. Beta-blockers are a class of drug widely used to manage hypertension. Atenolol, a beta-blocker, is made up of carbon, hydrogen, oxygen, and nitrogen atoms. When a 5.000-g sample is burned in oxygen, 11.57 g of CO_2 and 3.721 g of H_2O are obtained. A separate experiment using the same mass of sample (5.000 g) shows that atenolol has 10.52% nitrogen. What is the simplest formula of atenolol?

42. Saccharin is the active ingredient in many sweeteners used today. It is made up of carbon, hydrogen, oxygen, sulfur, and nitrogen. When 7.500 g of saccharin are burned in oxygen, 12.6 g CO_2, 1.84 g H_2O, and 2.62 g SO_2 are obtained. Another experiment using the same mass of sample (7.500 g) shows that saccharin has 7.65% N. What is the simplest formula for saccharin?

43. Hexamethylenediamine (MM = 116.2 g/mol), a compound made up of carbon, hydrogen, and nitrogen atoms, is used in the production of nylon. When 6.315 g of hexamethylenediamine is burned in oxygen, 14.36 g of carbon dioxide and 7.832 g of water are obtained. What are the simplest and molecular formulas of this compound?

44. Dimethylhydrazine, the fuel used in the Apollo lunar descent module, has a molar mass of 60.10 g/mol. It is made up of carbon, hydrogen, and nitrogen atoms. The combustion of 2.859 g of the fuel in excess oxygen yields 4.190 g of carbon dioxide and 3.428 g of water. What are the simplest and molecular formulas for dimethylhydrazine?

45. Epsom salts are hydrated crystals of magnesium sulfate, $MgSO_4 \cdot xH_2O$. When 6.499 g of Epsom salts are dehydrated by heating, 3.173 g remain. What is the mass percent of water in the hydrated sample? What is x?

46. Sodium borate decahydrate, $Na_2B_4O_7 \cdot 10H_2O$ is commonly known as borax. It is used as a deodorizer and mold inhibitor. A sample weighing 15.86 g is heated until a constant mass is obtained indicating that all the water has been evaporated off.

(**a**) What percent, by mass of $Na_2B_4O_7 \cdot 10\ H_2O$ is water?

(**b**) What is the mass of the anhydrous sodium borate, $Na_2B_4O_7$?

3-3 Mass Relations in Reactions

Balancing Equations

47. Balance the following equations:

(**a**) $CaC_2(s) + H_2O(l) \longrightarrow Ca(OH)_2(s) + C_2H_2(g)$

(**b**) $(NH_4)_2Cr_2O_7(s) \longrightarrow Cr_2O_3(s) + N_2(g) + H_2O(g)$

(**c**) $CH_3NH_2(g) + O_2(g) \longrightarrow CO_2(g) + N_2(g) + H_2O(g)$

48. Balance the following equations:

(**a**) $C_6H_{12}O_6(s) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$

(**b**) $XeF_4(g) + H_2O(l) \longrightarrow Xe(g) + O_2(g) + HF(g)$

(**c**) $NaCl(s) + H_2O(g) + SO_2(g) + O_2(g) \longrightarrow$
$Na_2SO_4(s) + HCl(g)$

49. Write balanced equations for the reaction of sulfur with the following metals to form solids that you can take to be ionic when the anion is S^{2-}.

(**a**) potassium (**b**) magnesium (**c**) aluminum
(**d**) calcium (**e**) iron (forming Fe^{2+} ions)

50. Write balanced equations for the reaction of scandium metal to produce the scandium(III) salt with the following nonmetals:

(a) sulfur (b) chlorine
(c) nitrogen (d) oxygen (forming the oxide)

51. Write a balanced equation for

(a) the combustion (reaction with oxygen gas) of glucose, $C_6H_{12}O_6$, to give carbon dioxide and water.
(b) the reaction between xenon tetrafluoride gas and water to give xenon, oxygen, and hydrogen fluoride gases.
(c) the reaction between aluminum and iron(III) oxide to give aluminum oxide and iron.
(d) the formation of ammonia gas from its elements.
(e) the reaction between sodium chloride, sulfur dioxide gas, steam, and oxygen to give sodium sulfate and hydrogen chloride gas.

52. Write a balanced equation for the reaction between

(a) dihydrogen sulfide and sulfur dioxide gases to form sulfur solid and steam.
(b) methane, ammonia, and oxygen gases to form hydrogen cyanide gas and steam.
(c) iron(III) oxide and hydrogen gas to form molten iron and steam.
(d) uranium(IV) oxide and hydrogen fluoride gas to form uranium(IV) fluoride and steam.
(e) the combustion of ethyl alcohol (C_2H_5OH) to give carbon dioxide and water.

Mole-Mass Relations in Reactions

53. Cyanogen gas, C_2N_2, has been found in the gases of outer space. It can react with fluorine to form carbon tetrafluoride and nitrogen trifluoride.

$$C_2N_2(g) + 7F_2(g) \longrightarrow 2CF_4(g) + 2NF_3(g)$$

(a) How many moles of fluorine react with 1.37 mol of cyanogen?
(b) How many moles of CF_4 are obtained from 13.75 mol of fluorine?
(c) How many moles of cyanogen are required to produce 0.8974 mol of NF_3?
(d) How many moles of fluorine will yield 4.981 mol of nitrogen trifluoride?

54. Ammonia reacts with a limited amount of oxygen according to the equation

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

(a) How many moles of NO are obtained when 3.914 moles of oxygen are used?
(b) How many moles of oxygen are required to react with 2.611 moles of ammonia?
(c) How many moles of water are obtained when 0.8144 moles of ammonia are used?
(d) How many moles of oxygen are required to produce 0.2179 mol of water?

55. One way to remove nitrogen oxide (NO) from smokestack emissions is to react it with ammonia.

$$4NH_3(g) + 6NO(g) \longrightarrow 5N_2(g) + 6H_2O(l)$$

Calculate

(a) the mass of water produced from 0.839 mol of ammonia.
(b) the mass of NO required to react with 3.402 mol of ammonia.
(c) the mass of ammonia required to produce 12.0 g of nitrogen gas.
(d) the mass of ammonia required to react with 115 g of NO.

56. Phosphine gas reacts with oxygen according to the following equation:

$$4PH_3(g) + 8\,O_2(g) \longrightarrow P_4O_{10}(s) + 6H_2O(g)$$

Calculate

(a) the mass of tetraphosphorus decaoxide produced from 12.43 mol of phosphine.
(b) the mass of PH_3 required to form 0.739 mol of steam.
(c) the mass of oxygen gas that yields 1.000 g of steam.
(d) the mass of oxygen required to react with 20.50 g of phosphine.

57. The combustion of liquid chloroethylene, C_2H_3Cl, yields carbon dioxide, steam, and hydrogen chloride gas.

(a) Write a balanced equation for the reaction.
(b) How many moles of oxygen are required to react with 35.00 g of chloroethylene?
(c) If 25.00 g of chloroethylene react with an excess of oxygen, how many grams of each product are formed?

58. Sand is mainly silicon dioxide. When sand is heated with an excess of coke (carbon), pure silicon and carbon monoxide are produced.

(a) Write a balanced equation for the reaction.
(b) How many moles of silicon dioxide are required to form 20.00 g of silicon?
(c) How many grams of carbon monoxide are formed when 98.76 g of silicon are produced?

59. When copper(II) oxide is heated in hydrogen gas, the following reaction takes place.

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$$

A copper rod coated with copper(II) oxide has a mass of 38.72 g. The rod is heated and made to react with 5.67 L of hydrogen gas, whose density at the conditions of the experiment is 0.0519 g/L.

(a) How many grams of CuO were converted to Cu?
(b) What is the mass of the copper after all the hydrogen is consumed? (Assume that CuO is converted only to Cu.)

60. When tin comes in contact with the oxygen in the air, tin(IV) oxide, SnO_2, is formed.

$$Sn(s) + O_2(g) \longrightarrow SnO_2(s)$$

A piece of tin foil, 8.25 cm × 21.5 cm × 0.600 mm ($d = 7.28\ g/cm^3$), is exposed to oxygen.

(a) Assuming that all the tin has reacted, what is the mass of the oxidized tin foil?
(b) Air is about 21% oxygen by volume ($d = 1.309$ g/L at 25°C, 1 atm). How many liters of air are required to completely react with the tin foil?

61. A crude oil burned in electrical generating plants contains about 1.2% sulfur by mass. When the oil burns, the sulfur forms sulfur dioxide gas:

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

How many liters of SO_2 ($d = 2.60$ g/L) are produced when 1.00×10^4 kg of oil burns at the same temperature and pressure?

62. When corn is allowed to ferment, the fructose in the corn is converted to ethyl alcohol according to the following reaction

$$C_6H_{12}O_6(aq) \longrightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

(**a**) What volume of ethyl alcohol (d = 0.789 g/mL) is produced from one pound of fructose?

(**b**) Gasohol can be a mixture of 10 mL ethyl alcohol and 90 mL of gasoline. How many grams of fructose are required to produce the ethyl alcohol in one gallon of gasohol?

63. Consider the hypothetical reaction

$$8A_2B_3(s) + 3X_4(g) \longrightarrow 4A_4X_3(s) + 12B_2(g)$$

When 10.0 g of A_2B_3 (MM = 255 g/mol) react with an excess of X_4, 4.00 g of A_4X_3 are produced.

(**a**) How many moles of A_4X_3 are produced?

(**b**) What is the molar mass of A_4X_3?

64. Consider the hypothetical reaction

$$2R_2X_5 + 2\ Z_8 \longrightarrow 5X_2 + 4RZ_4$$

When 25.00 g of Z_8 (MM = 197.4 g/mol) react with an excess of R_2X_5, 21.72 g of X_2 are produced.

(**a**) How many moles of X_2 are produced?

(**b**) What is the molar mass of X_2?

65. When solid phosphorus (P_4) reacts with oxygen gas, diphosphorus pentoxide is formed. Initially 2.87 mol of phosphorus and 3.86 mol of oxygen are combined. (Assume 100% yield.)

(**a**) Write a balanced equation for the reaction.

(**b**) What is the limiting reactant?

(**c**) How many moles of excess reactant remain after the reaction is complete?

66. Chlorine and fluorine react to form gaseous chlorine trifluoride. Initially, 1.75 mol of chlorine and 3.68 mol of fluorine are combined. (Assume 100% yield for the reaction.)

(**a**) Write a balanced equation for the reaction.

(**b**) What is the limiting reactant?

(**c**) What is the theoretical yield of chlorine trifluoride in moles?

(**d**) How many moles of excess reactant remain after reaction is complete?

67. When potassium chlorate is subjected to high temperatures, it decomposes into potassium chloride and oxygen.

(**a**) Write a balanced equation for the decomposition.

(**b**) In this decomposition, the actual yield is 83.2%. If 198.5 g of oxygen are produced, how much potassium chlorate decomposed?

68. When iron and steam react at high temperatures, the following reaction takes place.

$$3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$$

How much iron must react with excess steam to form 897 g of Fe_3O_4 if the reaction yield is 69%?

69. When solid silicon tetrachloride reacts with water, solid silicon dioxide and hydrogen chloride gas are formed.

(**a**) Write a balanced equation for the reaction.

(**b**) In an experiment, 45.00 g of silicon tetrachloride are treated with 45.00 mL of water (d = 1.00 g/mL). What is the theoretical yield of HCl (in grams)?

(**c**) When the reaction is complete, 17.8 L of HCl gas (d = 1.49 g/L at the conditions of the experiment) are obtained. What is the percent yield?

(**d**) How much of the reactant in excess is unused?

70. The first step in the manufacture of nitric acid by the Ostwald process is the reaction of ammonia gas with oxygen, producing nitrogen oxide and steam. The reaction mixture contains 7.60 g of ammonia and 10.00 g of oxygen. After the reaction is complete, 6.22 g of nitrogen oxide are obtained.

(**a**) Write a balanced equation for the reaction.

(**b**) How many grams of nitrogen oxide can be theoretically obtained?

(**c**) How many grams of excess reactant are theoretically unused?

(**d**) What is the percent yield of the reaction?

71. Tin(IV) chloride is used as an external coating on glass to toughen glass containers. It is prepared by reacting tin with chlorine gas.

$$Sn(s) + 2Cl_2(g) \longrightarrow SnCl_4(l)$$

The process requires a 25.0% excess of tin and has a yield of 73.7%. Calculate the mass of tin and the volume of chlorine gas (d = 2.898 g/L at 25°C, 1 atm) needed to produce 0.500 L of $SnCl_4$ (d = 2.226 g/mL).

72. A student prepares phosphorous acid, H_3PO_3, by reacting solid phosphorus triiodide with water.

$$PI_3(s) + 3H_2O(l) \longrightarrow H_3PO_3(s) + 3HI(g)$$

The student needs to obtain 0.250 L of H_3PO_3 (d = 1.651 g/cm^3). The procedure calls for a 45.0% excess of water and a yield of 75.0%. How much phosphorus triiodide should be weighed out? What volume of water (d = 1.00 g/cm^3) should be used?

Unclassified

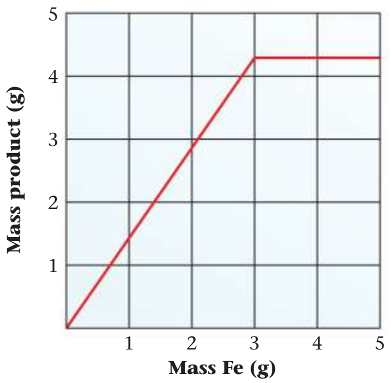
73. Cisplatin, $Pt(NH_3)_2Cl_2$, is a chemotherapeutic agent that disrupts the growth of DNA. If the current cost of Pt is \$1118.0/troy ounce (1 troy oz = 31.10 g), how many grams of cisplatin can you make with three thousand dollars worth of platinum? How many pounds?

74. Magnesium ribbon reacts with acid to produce hydrogen gas and magnesium ions. Different masses of magnesium ribbon are added to 10 mL of the acid. The volume of the hydrogen gas obtained is a measure of the number of moles of hydrogen produced by the reaction. Various measurements are given in the table below.

Experiment	Mass of Mg Ribbon (g)	Volume of Acid Used (mL)	Volume of H_2 Gas (mL)
1	0.020	10.0	21
2	0.040	10.0	42
3	0.080	10.0	82
4	0.120	10.0	122
5	0.160	10.0	122
6	0.200	10.0	122

(a) Draw a graph of the results by plotting the mass of Mg versus the volume of the hydrogen gas.
(b) What is the limiting reactant in experiment 1?
(c) What is the limiting reactant in experiment 3?
(d) What is the limiting reactant in experiment 6?
(e) Which experiment uses stoichiometric amounts of each reactant?
(f) What volume of gas would be obtained if 0.300 g of Mg ribbon were used? If 0.010 g were used?

75. Iron reacts with oxygen. Different masses of iron are burned in a constant amount of oxygen. The product, an oxide of iron, is weighed. The graph below is obtained when the mass of product obtained is plotted against the mass of iron used.



(a) How many grams of product are obtained when 0.50 g of iron are used?
(b) What is the limiting reactant when 2.00 g of iron are used?
(c) What is the limiting reactant when 5.00 g of iron are used?
(d) How many grams of iron react exactly with the amount of oxygen supplied?
(e) What is the simplest formula of the product?

76. Acetic acid ($HC_2H_3O_2$) can be prepared by the action of the *acetobacter* organism on dilute solutions of ethanol (C_2H_5OH). The equation for the reaction is

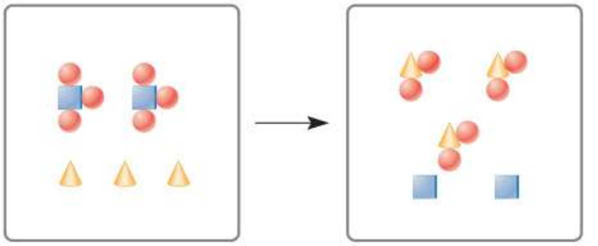
$$C_2H_5OH(aq) + O_2(g) \longrightarrow HC_2H_3O_2\,(aq) + H_2O$$

How many milliliters of a 12.5% (by volume) solution of ethanol are required to produce 175 mL of 0.664 *M* acetic acid? (Density of pure ethanol = 0.789 g/mL.)

Conceptual Problems

77. Given a pair of elements and their mass relation, answer the following questions.
(a) The mass of 4 atoms of A = the mass of 6 atoms of B. Which element has the smaller molar mass?
(b) The mass of 6 atoms of C is less than the mass of 3 atoms of the element D. Which element has more atoms/gram?
(c) Six atoms of E have a larger mass than six atoms of F. Which has more atoms/gram?
(d) Six atoms of F have the same mass as 8 atoms of G. Which has more atoms/mole?

78. The reaction between compounds made up of A (squares), B (circles), and C (triangles) is shown pictorially. Using smallest whole-number coefficients, write a balanced equation to represent the picture shown.



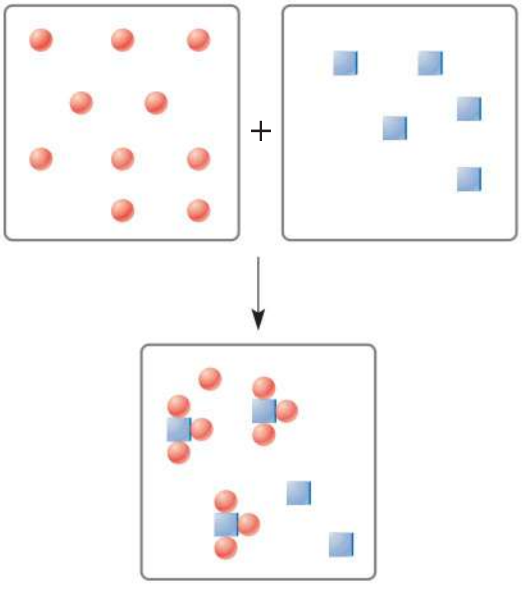
79. Represent the following equation pictorially (see Problem 78), using squares to represent A, circles to represent B, and triangles to represent C.

$$A_2B_3 + C_2 \longrightarrow C_2B_3 + A_2$$

After you have "drawn" the equation, use your drawing as a guide to balance it.

80. Nitrogen reacts with hydrogen to form ammonia. Represent each nitrogen atom by a square and each hydrogen atom with a circle. Starting with five molecules of both hydrogen and nitrogen, show pictorially what you have after the reaction is complete.

81. Consider the following diagram, where atom X is represented by a square and atom Y is represented by a circle.



(a) Write the equation for the reaction represented by the diagram.
(b) If each circle stands for a mole of Y and each square a mole of X, how many moles of X did one start with? How many moles of Y?
(c) Using the same representation described in part (b), how many moles of product are formed? How many moles of X and Y are left unreacted?

82. Box A contains 36 atoms of arsenic (As) and 27 molecules of O_2. Box B contains 18 molecules of As_2O_3. Without using your calculator, compare Box A to Box B with respect to
(a) the number of atoms of arsenic and oxygen.
(b) the number of discrete particles.
(c) mass.

83. One mol of ammonia reacts with 1.00 mol of oxygen to form nitrogen oxide and water according to the reaction

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

State which statements are true about the reaction and make the false statements true.

(**a**) All the oxygen is consumed.

(**b**) 4.00 mol NO are produced.

(**c**) 1.50 mol H_2O are produced.

(**d**) The description of the experiment does not provide enough information to determine percent yield.

(**e**) Three moles of water are produced for every two moles of NO obtained.

84. Suppose that the atomic mass of C-12 is taken to be 5.000 amu and that a mole is defined as the number of atoms in 5.000 kg of carbon-12. How many atoms would there be in one mole under these conditions? (*Hint:* There are 6.022×10^{23} C atoms in 12.00 g of C-12.)

85. Suppose that Si-28 ($^{28}_{14}Si$) is taken as the standard for expressing atomic masses and assigned an atomic mass of 10.00 amu. Estimate the molar mass of sodium nitride.

86. Answer the questions below, using **LT** (for *is less than*), **GT** (for *is greater than*), **EQ** (for *is equal to*), or **MI** (for *more information required*) in the blanks provided.

(**a**) The mass (to three significant figures) of 6.022×10^{23} atoms of Na _______ 23.0 g.

(**b**) Boron has two isotopes, B-10 (10.01 amu) and B-11 (11.01 amu). The abundance of B-10 _______ the abundance of B-11.

(**c**) If S-32 were assigned as the standard for expressing relative atomic masses and assigned an atomic mass of 10.00 amu, the atomic mass for H would be _______ 1.00 amu.

(**d**) When phosphine gas, PH_3, is burned in oxygen, tetraphosphorus decaoxide and steam are formed. In the balanced equation (using smallest whole-number coefficients) for the reaction, the sum of the coefficients on the reactant side is _______ 7.

(**e**) The mass (in grams) of one mole of bromine molecules is _______ 79.90.

87. Determine whether the statements given below are true or false.

(**a**) The mass of an atom can have the unit mole.

(**b**) In N_2O_4, the mass of the oxygen is twice that of the nitrogen.

(**c**) One mole of chlorine atoms has a mass of 35.45 g.

(**d**) Boron has an average atomic mass of 10.81 amu. It has two isotopes, B-10 (10.01 amu) and B-11 (11.01 amu). There is more naturally occurring B-10 than B-11.

(**e**) The compound $C_6H_{12}O_2N$ has for its simplest formula $C_3H_6ON_{1/2}$.

(**f**) A 558.5-g sample of iron contains ten times as many atoms as 0.5200 g of chromium.

(**g**) If 1.00 mol of ammonia is mixed with 1.00 mol of oxygen the following reaction occurs,

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

All the oxygen is consumed.

(**h**) When balancing an equation, the total number of moles of reactant molecules must equal the total number of moles of product molecules.

Challenge Problems

88. Chlorophyll, the substance responsible for the green color of leaves, has one magnesium atom per chlorophyll molecule and contains 2.72% magnesium by mass. What is the molar mass of chlorophyll?

89. By x-ray diffraction it is possible to determine the geometric pattern in which atoms are arranged in a crystal and the distances between atoms. In a crystal of silver, four atoms effectively occupy the volume of a cube 0.409 nm on an edge. Taking the density of silver to be 10.5 g/cm^3, calculate the number of atoms in one mole of silver.

90. A 5.025-g sample of calcium is burned in air to produce a mixture of two ionic compounds, calcium oxide and calcium nitride. Water is added to this mixture. It reacts with calcium oxide to form 4.832 g of calcium hydroxide. How many grams of calcium oxide are formed? How many grams of calcium nitride?

91. A mixture of potassium chloride and potassium bromide weighing 3.595 g is heated with chlorine, which converts the mixture completely to potassium chloride. The total mass of potassium chloride after the reaction is 3.129 g. What percentage of the original mixture was potassium bromide?

92. A sample of an oxide of vanadium weighing 4.589 g was heated with hydrogen gas to form water and another oxide of vanadium weighing 3.782 g. The second oxide was treated further with hydrogen until only 2.573 g of vanadium metal remained.

(**a**) What are the simplest formulas of the two oxides?

(**b**) What is the total mass of water formed in the successive reactions?

93. A sample of cocaine, $C_{17}H_{21}O_4N$, is diluted with sugar, $C_{12}H_{22}O_{11}$. When a 1.00-mg sample of this mixture is burned, 1.00 mL of carbon dioxide (d = 1.80 g/L) is formed. What is the percentage of cocaine in this mixture?

94. A 100.0-g mixture made up of $NaClO_3$, Na_2CO_3, NaCl, and $NaHCO_3$ is heated, producing 5.95 g of oxygen, 1.67 g of water, and 14.5 g of carbon dioxide. NaCl does not react under the conditions of the experiment. The equations for the reactions that take place are:

$$2\,NaClO_3(s) \longrightarrow 2\,NaCl(s) + 3O_2(g)$$

$$Na_2CO_3(s) \longrightarrow Na_2O(s) + CO_2(g)$$

$$2NaHCO_3(s) \longrightarrow Na_2O(s) + 2CO_2(g) + H_2O$$

Assuming 100% decomposition of $NaClO_3$, Na_2CO_3, and $NaHCO_3$, what is the composition of the mixture in grams?

95. An alloy made up of iron (52.6%), nickel (38.0%), cobalt (8.06%), and molybdenum (1.34%) has a density of 7.68 g/cm^3. How many molybdenum atoms are there in a block of the alloy measuring 13.0 cm × 22.0 cm × 17.5 cm?