SUPPLEMENT 1

Alternative Coverage of Moles, Molarity, and Chemical Calculations

(Clockwise from far left) Molar quantities of sulfur (32.1 grams), sucrose (table sugar) (342.3 grams), copper sulfate pentahydrate (249.7 grams), sodium chloride (58.4 grams), copper (63.6 grams), and mercury(II) oxide (216.6 grams). One mole of a substance is that quantity containing the number of grams numerically equal to its formula mass.

CONTENTS

This Supplement is designed to facilitate the teaching of courses where the introduction of the mole concept, calculations of concentration, and basic stoichiometry are necessary prior to the introduction of quantum mechanics and bonding. It is designed to follow Chapter 3 of the main textbook and replaces Sections 11-1, 11-2, 11-7, 11-8, 12-1, 12-2, and 12-4, which may then be omitted.

S1-1. The Quantity of a Substance That is Equal to Its Formula Mass in Grams Is Called a Mole

We learned in Chapter 2 that the atomic mass of an element is a relative quantity; it is the mass of one atom of the element relative to the mass of one atom of carbon-12, which by convention has a mass of exactly 12 atomic mass units, or 12 u. Consider the following table of four elements:

Element	Atomic mass/u
helium, He	4
carbon, C	12
titanium, Ti	48
molybdenum, Mo	96

This table shows that

- One carbon atom has a mass three times that of one helium atom.
- One titanium atom has a mass four times that of one carbon atom and 12 times that of one helium atom.
- One molybdenum atom has a mass twice that of one titanium atom, eight times that of one carbon atom, and 24 times that of one helium atom.

It is important to realize that we have not deduced the absolute mass of any one atom; at this point we can determine only relative masses based on our arbitrarily defined scale of carbon-12 equals 12 atomic mass units (12 u).

Consider 12 grams of carbon, 48 grams of titanium, and 96 grams of molybdenum (Figure S1.1). One titanium atom has a mass four times that of one carbon atom; therefore, 48 grams of titanium atoms must contain the same number of atoms as 12 grams of carbon. Similarly, one molybdenum atom has twice the mass of one titanium atom, so 96 grams of molybdenum must contain

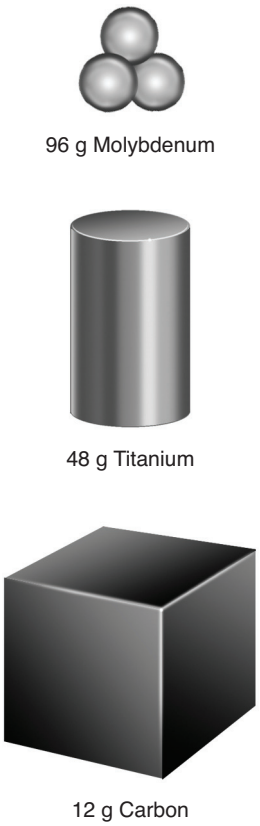


Figure S1.1 12 grams of carbon contains the same number of atoms as 48 grams of titanium and 96 grams of molybdenum.

the same number of atoms as 48 grams of titanium. We conclude that 12 grams of carbon, 48 grams of titanium, and 96 grams of molybdenum all contain the same number of atoms. If we continue this line of reasoning, we find that exactly the same number of atoms is contained in that quantity of an atomic element whose mass in grams is numerically equal to the element's atomic mass. Thus, we find from the atomic masses given on the periodic table on the inside front cover that 10.8 grams of boron, 23.0 grams of sodium, 63.6 grams of copper, and 200.6 grams of mercury all contain the same number of atoms.

All the substances we have considered here so far are **atomic substances,** that is, substances composed of only one type of atom. Now consider the following **molecular substances:**

Substance	Molecular mass/u
methane, CH_4	$12 + (4 \times 1) = 16$
oxygen, O_2	$2 \times 16 = 32$
ozone, O_3	$3 \times 16 = 48$

Like atomic masses, molecular masses are relative masses. A molecule of oxygen, O_2, has a mass of 32 u, twice that of a molecule of methane, 16 u. A molecule of ozone has a mass of 48 u, three times that of a molecule of methane. Using the same reasoning we used for atomic substances, we conclude that 16 grams of methane, 32 grams of oxygen, and 48 grams of ozone must all contain the same number of *molecules*. In addition, because the atomic mass of titanium is equal to the molecular mass of ozone, the number of *atoms* in 48 grams of titanium must equal the number of *molecules* in 48 grams of ozone.

We can eliminate the necessity of using the two separate terms atomic mass and molecular mass by using the single term **formula mass** to cover both. Likewise, a **formula unit** can refer to an atom, a molecule, or an ion. We now can extend our previous statement to say that identical numbers of formula units are contained in those quantities of different substances whose masses in grams are numerically equal to their respective formula masses. Thus, 4 grams of helium, 12 grams of carbon, 16 grams of methane, and 32 grams of oxygen all contain the same number of formula units. The formula units are atoms in the cases of helium and carbon, and molecules in the cases of methane and oxygen.

To aid in chemical calculations, chemists use a unit called a **mole**. *The quantity of a substance whose mass in grams is numerically equal to the formula mass of the substance is called a mole.* The abbreviation for mole is **mol** and is often used whenever a specific quantity of moles is written, for example, 3.12 mol $CH_4(g)$. The **molar mass** of a compound is the number of grams needed to make up one mole of the compound. For example, methane, $CH_4(g)$, has a formula mass of 16.04, and so its molar mass is 16.04 $g \cdot mol^{-1}$. Notice that formula mass is a unitless quantity, whereas molar mass has units of grams per mole. The frontispiece shows one mole of each of six common substances.

As you perform chemical calculations, you will often need to convert between mass and moles. For example, the formula mass of $CH_4(g)$ is 16.04, so there are 16.04 grams per mole of methane. This fact yields two unit conversion factors:

$$\left(\frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4}\right) = 1 \qquad \text{or} \qquad \left(\frac{16.04 \text{ g } CH_4}{1 \text{ mol } CH_4}\right) = 1$$

The factor with mass in grams in the denominator may be used to convert from grams to moles of $CH_4(g)$, and its inverse with moles in the denominator to convert from moles of $CH_4(g)$ to mass in grams.

For example, suppose you needed to know the number of moles in 50.0 grams of methane. Using the unit conversion factor with the mass in grams in the denominator, we find

$$\text{moles of } CH_4 = \underbrace{(50.0 \text{ g } CH_4)}_{\substack{\text{3 significant}\\ \text{figures}}}\underbrace{\left(\frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4}\right)}_{\substack{\text{4 significant}\\ \text{figures (1 is exact)}}} = \underbrace{3.12 \text{ mol } CH_4}_{\substack{\text{3 significant}\\ \text{figures}}}$$

Notice that the result is expressed to three significant figures and assigned the units moles of $CH_4(g)$, as required.

You can also calculate the mass of a certain number of moles of a substance. For example, let's calculate the mass of 2.16 moles of sodium chloride, $NaCl(s)$. The formula mass of $NaCl(s)$ is 58.44. The mass of $NaCl(s)$ in 2.16 moles is

$$\text{mass of NaCl} = \underbrace{(2.16 \text{ mol NaCl})}_{\substack{\text{3 significant}\\ \text{figures}}}\underbrace{\left(\frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}}\right)}_{\substack{\text{4 significant}\\ \text{figures (1 is exact)}}} = \underbrace{126 \text{ g NaCl}}_{\substack{\text{3 significant}\\ \text{figures}}}$$

Notice that we express the final result to three significant figures and that the units assigned to the result are grams of $NaCl(s)$.

When using a conversion factor, it is generally best if the numbers used in the conversion factor (if not exact) have *at least one more significant figure* than the data being converted; otherwise the conversion factor, not the actual data, will limit the number of significant figures in the result. It is important to remember that defined conversion factors such as 1 m ≡ 100 cm are always exact numbers and never limit the number of significant figures in the result.

EXAMPLE S1-1: Calculate the number of moles in (a) 28.0 grams of water (about 1 oz) and (b) 324 mg of aspirin, $C_9H_8O_4(s)$ (324 mg is the mass of aspirin in one 5-grain aspirin tablet).

Solution: (a) The formula mass of $H_2O(l)$ is 18.02. Consequently, the number of moles of water in 28.0 grams is

$$\text{moles of } H_2O = (28.0 \text{ g } H_2O)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 1.55 \text{ mol } H_2O$$

(b) The chemical formula of aspirin is $C_9H_8O_4(s)$, so its formula mass is $(9 \times 12.01) + (8 \times 1.008) + (4 \times 16.00) = 180.2$. (Recall that the atomic mass of naturally occurring carbon is 12.01.) The number of moles of aspirin in 324 milligrams is

$$\text{moles of } C_9H_8O_4 = (324 \text{ mg } C_9H_8O_4)\left(\frac{1 \text{ g}}{1000 \text{ mg}}\right)\left(\frac{1 \text{ mol } C_9H_8O_4}{180.2 \text{ g } C_9H_8O_4}\right)$$

$$= 1.80 \times 10^{-3} \text{ mol } C_9H_8O_4$$

Notice that we have to convert milligrams to grams before dividing by 180.2 grams.

This Practice Problem corresponds to 11-1 in the online Practice Problem solutions at: McQuarrieGeneralChemistry.com

PRACTICE PROBLEM S1-1: Some data for annual world chemical production are given in Appendix H. (a) Calculate the number of moles that are produced annually in the United States for each of the top five chemicals listed by mass produced: sulfuric acid, $H_2SO_4(l)$, nitrogen, $N_2(g)$, ethylene, $C_2H_4(g)$, oxygen, $O_2(g)$, and hydrogen, $H_2(g)$. (b) Which one has the largest annual production on a molar basis? (one metric ton = 1000 kg)

Answer: (a) 3.83×10^{11} mol $H_2SO_4(l)$, 9.52×10^{11} mol $N_2(g)$, 9.16×10^{11} mol $C_2H_4(g)$, 6.11×10^{11} mol $O_2(g)$, 8.78×10^{12} mol $H_2(g)$; (b) $H_2(g)$

Example S1-1 illustrates an important point. In order to calculate the number of moles in a given mass of a chemical compound, it is necessary to know the chemical formula of the compound. A mole of any compound is defined only in terms of its chemical formula. If a substance (such as coal or wood) cannot be represented by a single chemical formula, then we can give only the mass of the substance.

S1-2. One Mole of Any Substance Contains Avogadro's Number of Formula Units

It has been determined experimentally that one mole of any substance contains 6.022×10^{23} formula units (to four significant figures). This number is called **Avogadro's number** after the Italian scientist Amedeo Avogadro, who was one of the earliest scientists to distinguish between atoms and molecules (Chapter 13 Frontispiece). We say not only that one mole of any substance contains Avogadro's number of formula units but also that one mole is that mass of a substance containing Avogadro's number of formula units, or "elementary entities." For example, the atomic mass of the pure isotope carbon-12 is taken to be exactly 12, so 12.00 grams of carbon-12 contains 6.022×10^{23} atoms. Likewise, the molecular mass of water is 18.02, so 18.02 grams of water contains 6.022×10^{23} molecules.

A mole is simply a designation for Avogadro's number of "things" such as atoms and molecules, just as a dozen of eggs is a designation for twelve eggs. It is often helpful to think of one mole as a "counting unit" representing Avogadro's number of things, just as one dozen is a counting unit representing twelve things. But instead of the number 12 implied by the term dozen, the number of things in a mole is 6.022×10^{23}. A mole may be more intimidating because of the huge magnitude of Avogadro's number, but it is really the same concept as a dozen. A mole of eggs would be 6.022×10^{23} eggs, but it is not a practical measure of the number of eggs. A mole of atoms or molecules, on the other hand, is a practical measure of the number of atoms or molecules in a substance because of their small size. A few examples of things that we count in chemistry using moles are given in Table S1.1.

TABLE S1.1 Some things for which we use the counting unit "mole"

Counting unit	Number of things	Examples of things counted in moles	Mass of one mole
1 mole	6.022×10^{23}	atoms, such as aluminum, Al	26.98 g
		molecules, such as water, H_2O	18.02 g
		ions, such as Na^+	22.99 g
		elementary particles, such as electrons, e^-	0.5486 mg

We now have an alternative definition for a mole: *One mole is the mass of a substance containing Avogadro's number of formula units.* For example, referring to Table S1.1, one mole of aluminum atoms may be expressed as either 6.022×10^{23} aluminum atoms or 26.98 grams of aluminum. The formula mass of a substance is the mass in grams of one mole or 6.022×10^{23} formula units. That is,

$$1 \text{ mol Al} = 6.022 \times 10^{23} \text{ Al atoms} = 26.98 \text{ g Al}$$

Avogadro's number is an enormous number. If we were to express Avogadro's number without using scientific notation, we would have 602 200 000 000 000 000 000 000. In order to appreciate the magnitude of Avogadro's number another way, let's compute how many years it would take to spend Avogadro's number of dollars at a rate of one million dollars per second. Because there are 3.15×10^7 seconds in one year, the number of years required to spend 6.022×10^{23} dollars is

$$\text{number of years} = (6.022 \times 10^{23} \text{ dollars})\left(\frac{1 \text{ s}}{10^6 \text{ dollars}}\right)\left(\frac{1 \text{ year}}{3.15 \times 10^7 \text{ s}}\right)$$

$$= 1.91 \times 10^{10} \text{ years}$$

or 19.1 billion years (1 billion = 10^9). This interval is over four times longer than the estimated age of the earth (4.6 billion years) and is somewhat larger than the estimated age of the universe (14 billion years). This calculation illustrates just how large Avogadro's number is and, consequently, how small atoms and molecules are. Look again at the samples in shown in the frontispiece. Each of these contains 6.022×10^{23} formula units of the indicated substance.

We can use Avogadro's number to calculate the mass of a single atom or molecule, as illustrated by the following Example.

EXAMPLE S1-2: Using Avogadro's number, calculate the mass of one nitrogen molecule.

Solution: Recall that nitrogen occurs as a diatomic molecule. The formula of molecular nitrogen is N_2, so its formula mass, or molecular mass, is 28.02. Thus, there are 28.02 grams of nitrogen in one mole. Using the fact

that one mole of any substance contains 6.022×10^{23} formula units, the mass of one nitrogen molecule is

$$\begin{pmatrix}\text{mass of one}\\ \text{nitrogen molecule}\end{pmatrix} = \left(\frac{28.02\ \text{g N}_2}{1\ \text{mol N}_2}\right)\left(\frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{molecules}}\right)$$

$$= 4.653 \times 10^{-23}\ \text{g} \cdot \text{molecule}^{-1}$$

See solution to Practice Problem 11-2 online at: McQuarrieGeneralChemistry.com

PRACTICE PROBLEM S1-2: In Chapter 13 when we study gases, we shall use the masses of molecules in kilograms. Calculate the mass of a carbon dioxide, $CO_2(g)$, molecule and of a sulfur hexafluoride, $SF_6(g)$, molecule in kilograms.

Answer: CO_2, 7.308×10^{-26} kg; SF_6, 2.425×10^{-25} kg

Avogadro's number also can be used to calculate the number of atoms or molecules in a given mass of a substance. The next Example illustrates this type of calculation.

EXAMPLE S1-3: Calculate how many methane molecules and how many hydrogen and carbon atoms there are in a picogram of methane, $CH_4(g)$.

Solution: The formula mass of methane, $CH_4(g)$, is $12.01 + (4 \times 1.008) = 16.04$, so one picogram of $CH_4(g)$ consists of

molecules of CH_4

$$= (1.00\ \text{pg CH}_4)\left(\frac{1 \times 10^{-12}\ \text{g}}{1\ \text{pg}}\right)\left(\frac{1\ \text{mol CH}_4}{16.04\ \text{g CH}_4}\right)\left(\frac{6.022 \times 10^{23}\ \text{molecules CH}_4}{1\ \text{mol CH}_4}\right)$$

$$= 3.75 \times 10^{10}\ \text{CH}_4\ \text{molecules}$$

Each molecule of methane contains one carbon atom and four hydrogen atoms, so

$$\text{number of C atoms} = (3.75 \times 10^{10}\ \text{CH}_4\ \text{molecules})\left(\frac{1\ \text{C atom}}{1\ \text{CH}_4\ \text{molecule}}\right)$$

$$= 3.75 \times 10^{10}\ \text{C atoms}$$

$$\text{number of H atoms} = (3.75 \times 10^{10}\ \text{CH}_4\ \text{molecules})\left(\frac{4\ \text{H atoms}}{1\ \text{CH}_4\ \text{molecule}}\right)$$

$$= 1.50 \times 10^{11}\ \text{H atoms}$$

Realize that a picogram is a millionth of a millionth of a gram, a quantity so small that were all the molecules condensed into a liquid you would still need a high-quality microscope to see it. Nevertheless, there are well over a thousand million methane molecules in the sample. This illustrates just how small atoms and molecules are.

TABLE S1.2 Representative relationships between molar quantities

Substance	Formula	Formula mass	Molar mass/ $g \cdot mol^{-1}$	Number of particles in one mole	Number of moles
atomic chlorine	Cl	35.45	35.45	6.022×10^{23} chlorine atoms	1 mole of Cl atoms
chlorine gas	Cl_2	70.90	70.90	6.022×10^{23} chlorine molecules	1 mole of Cl_2 molecules
				12.044×10^{23} chlorine atoms	2 moles of Cl atoms
water	H_2O	18.02	18.02	6.022×10^{23} water molecules	1 mole of H_2O molecules
				12.044×10^{23} hydrogen atoms	2 moles of H atoms
				6.022×10^{23} oxygen atoms	1 mole of O atoms
sodium chloride	NaCl	58.44	58.44	6.022×10^{23} NaCl formula units	1 mole of NaCl formula units
				6.022×10^{23} sodium ions	1 mole of Na^+ ions
				6.022×10^{23} chloride ions	1 mole of Cl^- ions
barium fluoride	BaF_2	175.3	175.3	6.022×10^{23} BaF_2 formula units	1 mole of BaF_2 formula units
				6.022×10^{23} barium ions	1 mole of Ba^{2+} ions
				12.044×10^{23} fluoride ions	2 moles of F^- ions
nitrate ion	NO_3^-	62.01	62.01	6.022×10^{23} nitrate ions	1 mole of NO_3^- ions
				6.022×10^{23} nitrogen atoms	1 mole of N atoms
				18.066×10^{23} oxygen atoms	3 moles of O atoms

PRACTICE PROBLEM S1-3: Some inkjet printers produce picoliter-sized drops. How many water molecules are there in one picoliter of water? How many hydrogen and oxygen atoms does this correspond to? Take the density of water to be $1.00\ g \cdot mL^{-1}$.

Answer: 3.34×10^{13} water molecules, 6.68×10^{13} hydrogen atoms, 3.34×10^{13} oxygen atoms.

See solution to Practice Problem 11-3 online at: McQuarrieGeneralChemistry.com

Table S1.2 summarizes the relationships between molar quantities.

We conclude this section with the official SI definition of a mole: "The mole is the amount of substance of a system which contains as many elementary entities as there are atoms in exactly 0.012 kilograms of carbon-12. When the mole is used, the elementary entities must be specified and may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles." Note that because the atomic mass of carbon-12 is exactly 12 by definition, a mole of carbon-12 contains exactly 12 grams (= 0.012 kg) of carbon. This SI definition of a mole is equivalent to the other definitions given in this section.

S1-3. The Coefficients in Chemical Equations Can Be Interpreted as Numbers of Moles

A subject of great practical importance in chemistry is the determination of what quantity of product can be obtained from a given quantity of reactants. For example, the reaction between hydrogen and nitrogen to produce ammonia, $NH_3(g)$, can be described by the equation

$$3\,H_2(g) + N_2(g) \rightarrow 2\,NH_3(g) \qquad \text{(S1.1)}$$

where the coefficients in the equation are called **balancing coefficients** or **stoichiometric coefficients**. We might wish to know how much $NH_3(g)$ is produced when 10.0 grams of $N_2(g)$ reacts with excess $H_2(g)$. Recall from Section 3-2 that we can interpret the balancing coefficients in a chemical equation in a number of ways. To determine how much of one substance can be obtained from another, we interpret the equation in terms of moles. Thus, we interpret Equation S1.1 as

$$3 \text{ mol } H_2(g) + 1 \text{ mol } N_2(g) \rightarrow 2 \text{ mol } NH_3(g)$$

This result is important. It tells us that the stoichiometric or balancing coefficients are the relative numbers of moles of each substance in a balanced chemical equation.

We can also interpret the hydrogen-nitrogen reaction in terms of masses. If we convert moles to masses by multiplying by the appropriate molar masses, then we get

$$6.05 \text{ g } H_2(g) + 28.02 \text{ g } N_2(g) \rightarrow 34.07 \text{ g } NH_3(g)$$

Note that the total mass is the same on the two sides of the equation, in accord with the law of conservation of mass. Table S1.3 summarizes the various interpretations of the equation for the reaction of hydrogen with nitrogen to pro-

TABLE S1.3 The various interpretations of two chemical equations

Interpretation	**$3H_2$**	**+**	**N_2**	**→**	**$2NH_3$**
molecular:	3 molecules	+	1 molecule	→	2 molecules
molar:	3 moles	+	1 mole	→	2 moles
mass:	6.05 grams	+	28.02 grams	→	34.07 grams
Interpretation	**2Na**	**+**	**Cl_2**	**→**	**2NaCl**
molecular:	2 atoms	+	1 molecule	→	2 ion pairs or 2 formula units
molar:	2 moles	+	1 mole	→	2 moles
mass:	45.98 grams	+	70.90 grams	→	116.88 grams

duce ammonia as well as that of the reaction of sodium with chlorine to produce sodium chloride.

We are now ready to calculate how much ammonia is produced when a given quantity of nitrogen or hydrogen is used. Let's calculate how many moles of $NH_3(g)$ can be produced from 10.0 moles of $N_2(g)$, assuming that an excess amount of $H_2(g)$ is available. According to Equation S1.1, two moles of $NH_3(g)$ are produced from each mole of $N_2(g)$. We can express this relationship as a **stoichiometric unit conversion factor**,

$$1 \text{ mol } N_2 \Leftrightarrow 2 \text{ mol } NH_3$$

which can also be expressed as

$$\frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 1 \qquad \text{or} \qquad \frac{1 \text{ mol } N_2}{2 \text{ mol } NH_3} = 1$$

When working with stoichiometric unit conversion factors, we choose the form suitable to the conversion of the given units. In this case, because we are converting from moles of nitrogen to moles of ammonia, we use the conversion factor with moles of ammonia in the numerator and moles of nitrogen in the denominator. Thus, we find that 10.0 moles of $N_2(g)$ yields

$$\text{moles of } NH_3 = (10.0 \text{ mol } N_2)\left(\frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2}\right) = 20.0 \text{ mol } NH_3$$

We can also calculate the number of grams of $NH_3(g)$ produced by using the fact that one mole of $NH_3(g)$ corresponds to 17.03 grams of $NH_3(g)$:

$$\begin{aligned}\text{mass of } NH_3 \text{ produced} &= (20.0 \text{ mol } NH_3)\left(\frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3}\right)\\ &= 341 \text{ g } NH_3\end{aligned}$$

Thus, we see that ratios of the balancing coefficients in a chemical equation are also unit conversion factors that allow us to convert from the number of moles of one substance into the number of moles of any other species consumed or produced in that reaction. The following Examples and Practice Problems illustrate the use of stoichiometric coefficients as conversion factors in chemical equations.

EXAMPLE S1-4: How many grams of $NH_3(g)$ can be produced from 8.50 grams of $H_2(g)$, assuming that an excess amount of $N_2(g)$ is available? What is the minimum mass of $N_2(g)$ required?

Solution: The equation for the reaction is $3\,H_2(g) + N_2(g) \rightarrow 2\,NH_3(g)$. The number of moles of $H_2(g)$ corresponding to 8.50 grams is

$$\text{moles of } H_2 = (8.50 \text{ g } H_2)\left(\frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2}\right) = 4.22 \text{ mol } H_2$$

The number of moles of $NH_3(g)$ is obtained by using the stoichiometric unit conversion factor between $NH_3(g)$ and $H_2(g)$,

$$\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} = 1$$

which we obtain directly from the balanced chemical equation. Therefore,

$$\text{moles of } NH_3 = (4.22 \text{ mol } H_2)\left(\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}\right) = 2.81 \text{ mol } NH_3$$

The mass of $NH_3(g)$ is given by

$$\text{mass of } NH_3 = (2.81 \text{ mol } NH_3)\left(\frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3}\right) = 47.9 \text{ g } NH_3$$

To calculate the minimum mass of $N_2(g)$ required to react completely with 8.50 grams, or 4.22 moles of $H_2(g)$, we first calculate the number of moles of $N_2(g)$ required:

$$\text{moles of } N_2 = (4.22 \text{ mol } H_2)\left(\frac{1 \text{ mol } N_2}{3 \text{ mol } H_2}\right) = 1.41 \text{ mol } N_2$$

As always, the stoichiometric unit conversion factor is obtained from the balanced chemical equation. The minimum mass of $N_2(g)$ required is given by

$$\text{mass of } N_2 = (1.41 \text{ mol } N_2)\left(\frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2}\right) = 39.5 \text{ g } N_2$$

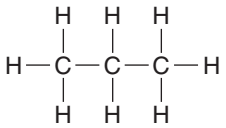
See solution to Practice Problem 11-9 online at: McQuarrieGeneralChemistry.com

PRACTICE PROBLEM S1-4: A frequently used method for preparing oxygen in the laboratory is by the thermal decomposition of potassium chlorate, $KClO_3(s)$ (Figure S1.2). This reaction is described by the unbalanced equation

$$KClO_3(s) \rightarrow KCl(s) + O_2(g) \quad (\textit{unbalanced})$$

(a) Balance this equation. (b) How many moles of $O_2(g)$ can be prepared from 0.50 moles of $KClO_3(s)$? (c) How many grams of $O_2(g)$ can be prepared from 30.6 grams of $KClO_3(s)$?

Answer: (a) $2\,KClO_3(s) \rightarrow 2\,KCl(s) + 3\,O_2(g)$; (b) 0.75 moles; (c) 12.0 grams



The formula of propane, $C_3H_8(g)$. Its formula is sometimes written as $CH_3CH_2CH_3(g)$ to emphasize the structure. Propane is stored in metal tanks and often used as a fuel in areas not serviced by natural gas pipelines.

EXAMPLE S1-5: Propane, $C_3H_8(g)$, a common fuel, burns in oxygen according to the equation

$$C_3H_8(g) + 5\,O_2(g) \rightarrow 3\,CO_2(g) + 4\,H_2O(l)$$

(a) How many grams of $O_2(g)$ are required to burn 75.0 grams of $C_3H_8(g)$? (b) How many grams of $H_2O(l)$ and $CO_2(g)$ are then produced?

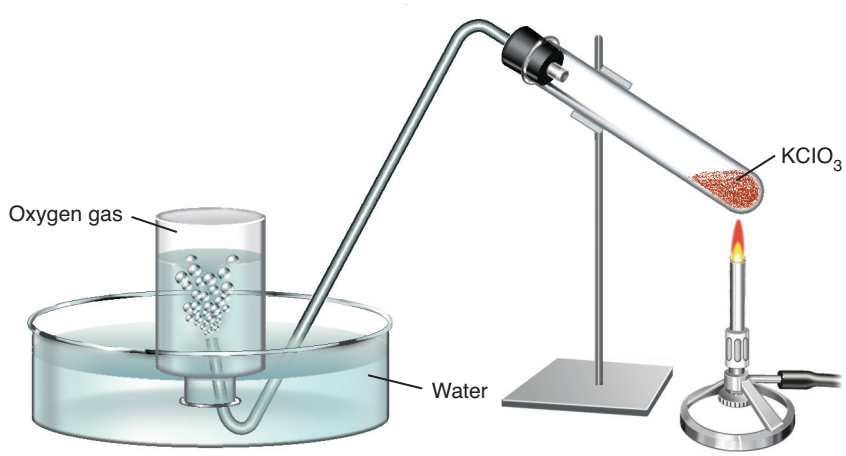


Figure S1.2 A typical experimental setup for the production of small amounts of oxygen by gently heating potassium chlorate, $KClO_3(s)$. Because it is only slightly soluble in water, the oxygen is collected by the displacement of water from an inverted bottle.

Solution: (a) The chemical equation states that five moles of $O_2(g)$ are required to burn one mole of $C_3H_8(g)$. The molecular mass of $C_3H_8(g)$ is 44.10, so 75.0 grams of $C_3H_8(g)$ corresponds to

$$\text{moles of } C_3H_8 = (75.0 \text{ g } C_3H_8)\left(\frac{1 \text{ mol } C_3H_8}{44.09 \text{ g } C_3H_8}\right) = 1.70 \text{ mol } C_3H_8$$

The number of moles of $O_2(g)$ required is

$$\text{moles of } O_2 = (1.70 \text{ mol } C_3H_8)\left(\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8}\right) = 8.50 \text{ mol } O_2$$

To find out how many grams of $O_2(g)$ this is, we multiply the number of moles by the molar mass of $O_2(g)$:

$$\text{mass of } O_2 = (8.50 \text{ mol } O_2)\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}\right) = 272 \text{ g } O_2$$

Thus, we see that 272 grams of $O_2(g)$ are required to burn 75.0 grams of $C_3H_8(g)$. (Remember whenever possible to use at least one more significant figure in your constants than are present in your data.)

(b) According to the chemical equation, three moles of $CO_2(g)$ and four moles of $H_2O(l)$ are produced for each mole of $C_3H_8(g)$ burned. Therefore, the number of moles of $CO_2(g)$ produced is

$$\text{moles of } CO_2 = (1.70 \text{ mol } C_3H_8)\left(\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}\right) = 5.10 \text{ mol } CO_2$$

The number of moles of $H_2O(l)$ produced is

$$\text{moles of } H_2O = (1.70 \text{ mol } C_3H_8)\left(\frac{4 \text{ mol } H_2O}{1 \text{ mol } C_3H_8}\right) = 6.80 \text{ mol } H_2O$$

The moles of $CO_2(g)$ and $H_2O(l)$ are converted to the mass in grams by multiplying by the respective molar masses:

$$\text{mass of } CO_2 = (5.10 \text{ mol } CO_2)\left(\frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\right) = 224 \text{ g } CO_2$$

$$\text{mass of } H_2O = (6.80 \text{ mol } H_2O)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right) = 123 \text{ g } H_2O$$

These are the quantities of $CO_2(g)$ and $H_2O(l)$ produced when 75.0 grams of propane are burned.

We can summarize the results of this example by

$$\underset{75.0\text{ g}}{C_3H_8(g)} + \underset{272\text{ g}}{5\,O_2(g)} \rightarrow \underset{224\text{ g}}{3\,CO_2(g)} + \underset{123\text{ g}}{4\,H_2O(l)}$$

Notice that the total mass on each side of the chemical reaction is the same, as it must be according to the principle of conservation of mass.

See solution to Practice Problem 11-10 online at: McQuarrieGeneralChemistry.com

PRACTICE PROBLEM S1-5: Deuterated ammonia, $ND_3(g)$, can be prepared by reacting lithium nitride, $Li_3N(s)$, with heavy water, $D_2O(l)$, according to

$$Li_3N(s) + 3\,D_2O(l) \rightarrow 3\,LiOD(s) + ND_3(g)$$

How many milligrams of heavy water are required to produce 7.15 milligrams of $ND_3(g)$? Given that the density of heavy water is 1.106 g·mL^{-1} at 25°C, how many milliliters of heavy water are required? Take the atomic mass of deuterium, D, to be 2.014.

Answer: 21.4 mg; 0.0194 mL

In Example S1-4, we represented the reaction between $H_2(g)$ and $N_2(g)$ to form $NH_3(g)$ by Equation S1.1. We could also have represented it by

$$\tfrac{3}{2}H_2(g) + \tfrac{1}{2}N_2(g) \rightarrow NH_3(g) \qquad \text{(S1.2)}$$

Before leaving this section, we shall show that the result we obtained in Example S1-4 does not depend upon which equation we use to represent the reaction. Starting again with 8.50 grams of $H_2(g)$, or 4.22 moles of $H_2(g)$, the number of moles of $NH_3(g)$ produced may be determined using the stoichiometric unit conversion factors from Equation S1.2, namely,

$$\frac{1 \text{ mol } NH_3}{\frac{3}{2} \text{ mol } H_2} = 1$$

Therefore,

$$\text{moles of } NH_3(g) = (4.22 \text{ mol } H_2)\left(\frac{1 \text{ mol } NH_3}{\frac{3}{2} \text{ mol } H_2}\right) = 2.81 \text{ mol } NH_3$$

which is exactly the same result as we obtained in Example S1-4 using Equation S1.2. You can see that this is so because

$$\frac{1 \text{ mol } NH_3}{\frac{3}{2} \text{ mol } H_2} = \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}$$

In other words, the stoichiometric unit conversion factors are identical because they consist of *ratios* of stoichiometric coefficients. Of course, this must be so for physical reasons; it would be ridiculous to obtain different amounts of products depending upon how we (arbitrarily) choose to represent the reaction by a chemical equation.

S1-4. Calculations Involving Chemical Reactions Are Carried Out in Terms of Moles

For calculations involving chemical reactions and mass, the procedure is first to convert mass to moles, then convert moles of one substance to moles of another by using the balancing coefficients in the chemical equation, and then convert moles into mass. An understanding of the flowchart in Figure S1.3 will allow you to do most calculations involving chemical equations and mass. The following calculation further illustrates the use of Figure S1.3.

Small quantities of bromine can be prepared in the laboratory by heating manganese(IV) oxide, concentrated sulfuric acid, and potassium bromide under a fume hood. The equation describing this reaction is

$$MnO_2(s) + 2\,H_2SO_4(aq) + 2\,KBr(aq) \rightarrow MnSO_4(aq) + K_2SO_4(aq) + 2\,H_2O(l) + Br_2(l)$$

Let's calculate how many grams of $MnO_2(s)$ and $KBr(aq)$ are required to produce 225 grams of $Br_2(l)$ in an excess of $H_2SO_4(aq)$. From the balanced chemical equation, we see that

$$1 \text{ mol } Br_2 \Leftrightarrow 1 \text{ mol } MnO_2 \quad \text{and} \quad 1 \text{ mol } Br_2 \Leftrightarrow 2 \text{ mol KBr}$$

The number of moles of $Br_2(l)$ corresponding to 225 grams is

$$\text{moles of } Br_2 = (225 \text{ g } Br_2)\left(\frac{1 \text{ mol } Br_2}{159.8 \text{ g } Br_2}\right) = 1.41 \text{ mol}$$

To convert from moles of $Br_2(l)$ to moles of $MnO_2(s)$, we use the stoichiometric coefficients from the balanced equation as outlined in Figure S1.3:

$$\text{moles of } MnO_2 \text{ required} = (1.41 \text{ mol } Br_2)\left(\frac{1 \text{ mol } MnO_2}{1 \text{ mol } Br_2}\right) = 1.41 \text{ mol}$$

and

$$\text{moles of KBr required} = (1.41 \text{ mol } Br_2)\left(\frac{2 \text{ mol KBr}}{1 \text{ mol } Br_2}\right) = 2.82 \text{ mol}$$

Therefore, the number of grams of each reactant required is

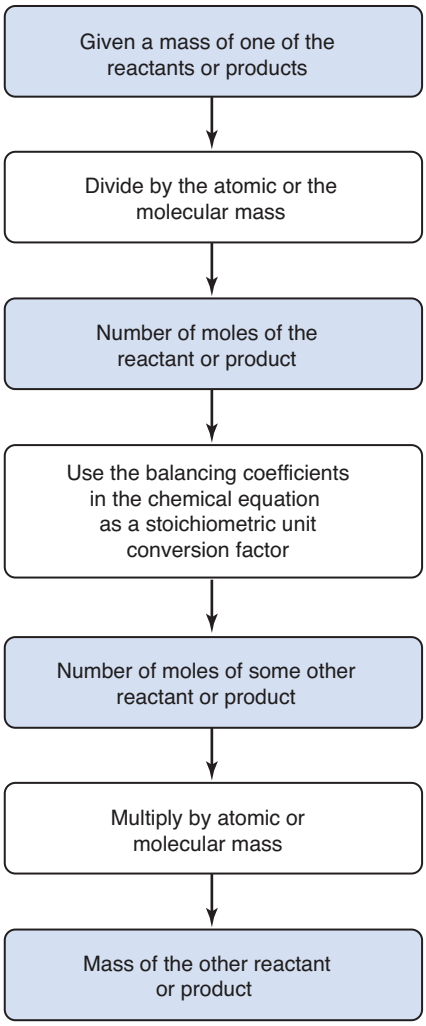


Figure S1.3 Flow diagram of the procedure for calculating the mass or the number of moles from chemical equations. The essence of the method is to realize that we convert from moles of one substance to moles of another substance in a chemical equation by using ratios of the stoichiometric coefficients.

$$\text{mass of } MnO_2 \text{ required} = (1.41 \text{ mol } MnO_2)\left(\frac{86.94 \text{ g } MnO_2}{1 \text{ mol } MnO_2}\right) = 123 \text{ g}$$

and

$$\text{mass of KBr required} = (2.82 \text{ mol KBr})\left(\frac{119.0 \text{ g KBr}}{1 \text{ mol KBr}}\right) = 336 \text{ g}$$

The next two Examples further illustrate the calculation of quantities involved in chemical reactions.

EXAMPLE S1-6: Diiodine pentoxide, $I_2O_5(s)$, is a reagent used for the quantitative determination of carbon monoxide. The equation for the reaction is

$$I_2O_5(s) + 5\,CO(g) \rightarrow I_2(s) + 5\,CO_2(g)$$

A gas sample containing carbon monoxide is collected from the exhaust of an engine. If 0.098 grams of $I_2(s)$ are produced from the reaction of the $CO(g)$ in the gas sample with excess $I_2O_5(s)$, then how many grams of $CO(g)$ are present in the sample?

Solution: We see from the chemical equation that five moles of $CO(g)$ produce one mole of $I_2(s)$, or that

$$1 \text{ mol } I_2 \Leftrightarrow 5 \text{ mol CO}$$

We carry the calculations through to one more significant figure than the final answer to avoid rounding errors.

The number of moles of $I_2(s)$ corresponding to 0.098 grams is

$$\text{moles of } I_2 = (0.098 \text{ g } I_2)\left(\frac{1 \text{ mol } I_2}{253.8 \text{ } I_2}\right) = 3.86 \times 10^{-4} \text{ mol } I_2$$

The number of moles of $CO(g)$ is

$$\text{moles of } CO(g) = (3.86 \times 10^{-4} \text{ mol } I_2)\left(\frac{5 \text{ mol CO}}{1 \text{ mol } I_2}\right) = 0.00193 \text{ mol CO}$$

and the mass of $CO(g)$ is given by

$$\text{mass of CO} = (0.00193 \text{ mol CO})\left(\frac{28.01 \text{ g CO}}{1 \text{ mol CO}}\right) = 0.054 \text{ g CO}$$

PRACTICE PROBLEM S1-6: Finely divided sulfur ignites spontaneously in fluorine gas to produce sulfur hexafluoride, as described by the chemical equation

$$S(s) + 3\,F_2(g) \rightarrow SF_6(g)$$

See solution to Practice Problem 11-11 online at: McQuarrieGeneralChemistry.com

How many grams of sulfur hexafluoride, $SF_6(g)$, can be produced from 5.00 grams of sulfur? How many grams of fluorine gas are required to react with the 5.00 grams of sulfur?

Answer: 22.8 grams of $SF_6(g)$; 17.8 grams of $F_2(g)$

EXAMPLE S1-7: Phosphorus reacts directly with sodium metal to produce sodium phosphide, as described by the chemical equation

$$3\,Na(s) + P(s) \rightarrow Na_3P(s)$$

How many grams of sodium phosphide, $Na_3P(s)$, can be produced from 10.0 grams of sodium metal?

Solution: We see from the chemical equation that one mole of $Na_3P(s)$ is produced from three moles of $Na(s)$, or that

$$1 \text{ mol } Na_3P \Leftrightarrow 3 \text{ mol Na}$$

The number of moles of $Na(s)$ is given by

$$\text{moles of Na} = (10.0 \text{ g Na})\left(\frac{1 \text{ mol Na}}{22.99 \text{ g Na}}\right) = 0.435 \text{ mol}$$

The number of moles of $Na_3P(g)$ produced is given by

$$\text{moles of } Na_3P = (0.435 \text{ mol Na})\left(\frac{1 \text{ mol } Na_3P}{3 \text{ mol Na}}\right) = 0.145 \text{ mol}$$

and the number of grams of Na_3P produced is given by

$$\text{mass of } Na_3P = (0.145 \text{ mol } Na_3P)\left(\frac{99.94 \text{ g } Na_3P}{1 \text{ mol } Na_3P}\right) = 14.5 \text{ g } Na_3P$$

This calculation can also be done in one operation:

$$\text{mass of } Na_3P = (10.0 \text{ g Na})\left(\frac{1 \text{ mol Na}}{22.99 \text{ g Na}}\right)\left(\frac{1 \text{ mol } Na_3P}{3 \text{ mol Na}}\right)\left(\frac{99.94 \text{ g } Na_3P}{1 \text{ mol } Na_3P}\right)$$
$$= 14.5 \text{ g } Na_3P$$

You should try to become comfortable doing a calculation like this in one step.

PRACTICE PROBLEM S1-7: Phosphorus triiodide can be prepared by the direct combination of phosphorus and iodine:

$$2\,P(s) + 3\,I_2(s) \rightarrow 2\,PI_3(s)$$

(a) How many grams of $PI_3(s)$ can be prepared from 1.25 grams of $P(s)$?
(b) How many grams of $I_2(s)$ are required to react with the 1.25 grams of $P(s)$?

Answer: (a) 16.6 grams $PI_3(s)$; (b) 15.4 grams $I_2(s)$

See solution to Practice Problem 11-12 online at: McQuarrieGeneralChemistry.com

S1-5. Molarity Is the Most Common Unit of Concentration

Most chemical and biological processes take place in solution, particularly in aqueous solution. As we have seen in Section 2-3, a solution is a mixture that is homogeneous at the molecular level. From a molecular point of view, the species in a solution are uniformly dispersed among one another (Figure S1.4).

The most common type of solution is a solid dissolved in a liquid. The solid that is dissolved is called the **solute**, and the liquid in which it is dissolved is called the **solvent**. The terms solvent and solute are merely terms of convenience because all the components of a solution are uniformly dispersed throughout the solution. When NaCl(*s*) is dissolved in water, we say that NaCl(*s*) is the solute and $H_2O(l)$ is the solvent. The process of dissolving NaCl(*s*) in water is represented by the equation

$$NaCl(s) \xrightarrow[H_2O(l)]{} Na^+(aq) + Cl^-(aq)$$

where $H_2O(l)$ under the arrow tells us that water is the solvent. The species $Na^+(aq)$ and $Cl^-(aq)$ represent a sodium ion and a chloride ion in an aqueous solution. As Figure S1.5 illustrates, these ions are solvated by water molecules; that is, they are surrounded by a loosely bound shell of water molecules.

When a small quantity of sodium chloride is added to a beaker of water, the sodium chloride dissolves completely, leaving no crystals at the bottom of the beaker. As more and more sodium chloride is added, we reach a point where no more sodium chloride can dissolve, so that any further sodium chloride crystals that we add simply remain at the bottom of the beaker. Such a solution is called a **saturated solution**, and the maximum quantity of solute dissolved is called the **solubility** of that solute. Solubility can be expressed in a variety of units, but quite commonly it is expressed as grams of solute per 100 grams of solvent. For example, we can say that the solubility of NaCl(*s*) in water at 20°C is about 36 grams per 100 grams of $H_2O(l)$.

It is important to realize that the solubility of a substance is the maximum quantity that can be dissolved in a saturated solution at a particular temperature. The solubility of NaCl(*s*) at 20°C is about 36 grams per 100 grams of

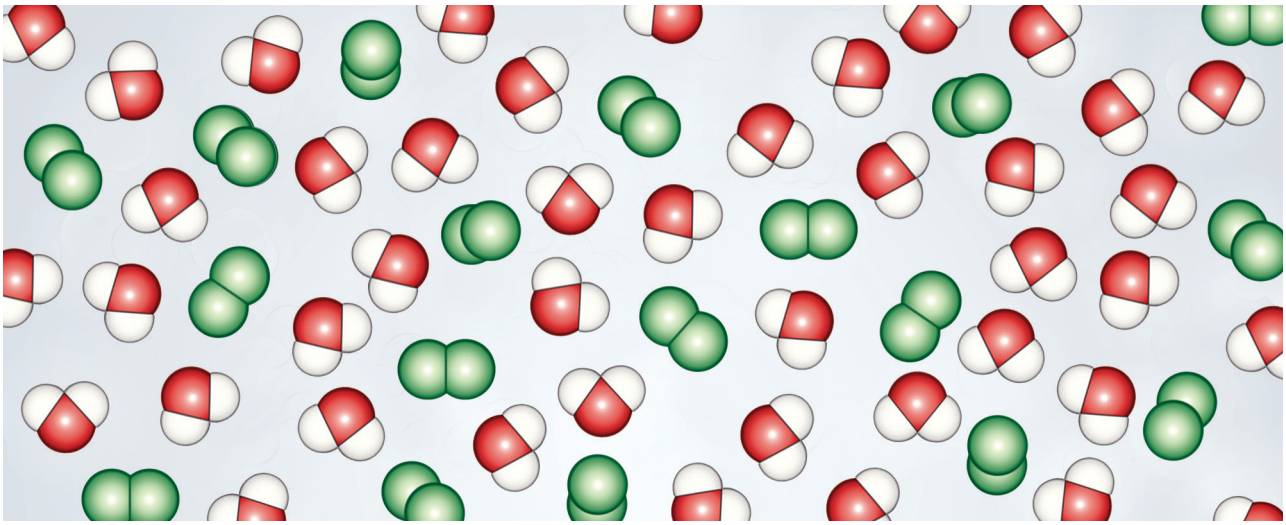


Figure S1.4 A solution is a homogeneous mixture at the molecular level. The species in a solution are uniformly dispersed among one another.

$H_2O(l)$. If we add 50 grams of $NaCl(s)$ to 100 grams of $H_2O(l)$ at 20°C, then 36 grams dissolve and 14 grams are left as undissolved $NaCl(s)$. The solution is saturated. If we add 25 grams of $NaCl(s)$ to 100 grams of $H_2O(l)$, then all the $NaCl(s)$ dissolves to form what is called an **unsaturated solution**; that is, a solution in which we can dissolve more of the solute.

In most cases the solubility of a substance depends on temperature. The effect of temperature on the solubility of several salts in water is shown in Figure S1.6. Almost all substances become more soluble in water as the temperature increases. For example, potassium nitrate is about five times more soluble in water at 40°C than it is at 0°C.

The **concentration** of solute in a solution describes the quantity of solute dissolved in a given quantity of solvent or a given quantity of solution. A common method of expressing the concentration of a solute is **molarity,** which is denoted by the symbol *M*. Molarity is defined as the number of moles of solute per liter of solution:

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \tag{S1.3}$$

Equation S1.3 can be expressed symbolically as

$$M = \frac{n}{V} \tag{S1.4}$$

where *M* represents the molarity of the solution, *n* is the number of moles of solute dissolved in the solution, and *V* is the total volume of the solution in liters. To see how to use Equation S1.4, let's calculate the molarity of a solution prepared by dissolving 62.3 grams of sucrose, $C_{12}H_{22}O_{11}(s)$, in enough water to form 0.500 liters of solution. The formula mass of sucrose is 342.3, so 62.3 grams corresponds to

$$\text{moles of sucrose} = (62.3 \text{ g sucrose})\left(\frac{1 \text{ mol sucrose}}{342.3 \text{ g sucrose}}\right) = 0.182 \text{ mol}$$

The molarity of the solution is given by

$$M = \frac{n}{V} = \frac{0.182 \text{ mol}}{0.500 \text{ L}} = 0.364 \text{ mol}\cdot\text{L}^{-1} = 0.364 \text{ M}$$

We say that the concentration of sucrose in the solution is 0.364 **molar**, which we write as 0.364 M. The unit of molarity is written as M.

The definition of molarity involves the total volume of the solution, not just the volume of the solvent. Suppose we wish to prepare one liter of a 0.100-M aqueous solution of potassium dichromate, $K_2Cr_2O_7(aq)$. We would prepare the solution by weighing out 0.100 moles (29.4 grams) of $K_2Cr_2O_7(s)$, dissolving it in less than one liter of water, say, about 500 mL, and then adding water while stirring until the final volume of the solution is precisely one liter. As shown in Figure S1.7, we use a **volumetric flask**, which is a piece of glassware used to prepare precise volumes. It would be incorrect to add 0.100 moles of $K_2Cr_2O_7(s)$

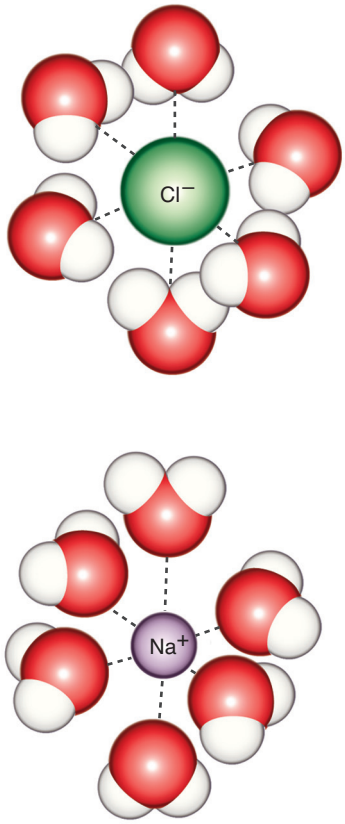


Figure S1.5 Ions in aqueous solutions are surrounded by a loosely bound shell of water molecules. Such ions are said to be solvated.

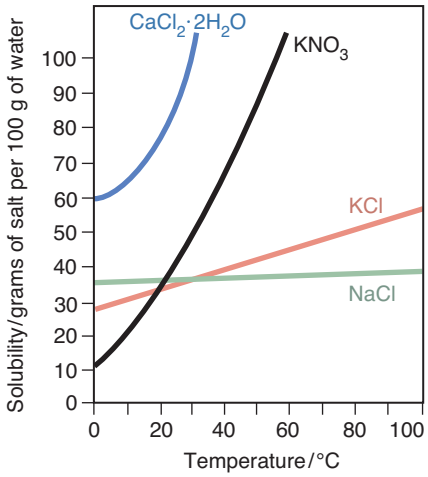


Figure S1.6 The solubility of most salts increases with increasing temperature.

(a) (b) (c)

Figure S1.7 The procedure used to prepare one liter of a solution of a specific molarity, such as 0.100 M $K_2Cr_2O_7(aq)$. (a) The 0.100 moles of $K_2Cr_2O_7(s)$ (29.4 grams) are weighed out and (b) added to a one-liter volumetric flask that is only partially filled with water. (c) The $K_2Cr_2O_7(s)$ is dissolved, and then more water is added to bring the final volume up to the one-liter mark on the flask. The solution is swirled to ensure uniform mixing.

to one liter of water; the final volume of such a solution is not precisely one liter because the added $K_2Cr_2O_7(s)$ changes the volume from 1.00 liter to 1.02 liters. The following example illustrates the procedure for making up a solution of a specified molarity.

THE PRECISION OF VOLUMETRIC GLASSWARE. The exact precision of volumetric glassware varies with the size and quality of the glassware used. The precision is rated using the International Organization for Standardization (ISO) system where "Class A" volumetric glassware (the sort found in most general chemistry laboratories) has the lowest precision, with subsequent letters designating glassware of increasing precision. Class A volumetric glassware has a precision that is at least 0.1% of the volume being measured. Thus, a Class A 250-mL volumetric flask has a precision of at least ± 0.25 mL, or about four significant figures at the stated temperature. For good analytical work, it is critical that the precision of the glassware be greater than that of the desired solution. Throughout this text we do not specify significant figures for volumetric measurements; rather, we assume that the glassware used has a precision greater than that of the desired solutions.

EXAMPLE S1-8: Potassium bromide, $KBr(s)$, is used by veterinarians to treat epilepsy in dogs. Explain how you would prepare 250 mL of a 0.600-M aqueous $KBr(aq)$ solution.

Solution: From Equation S1.4 and the specified volume and concentration, we can calculate the number of moles of KBr(s) required. Equation S1.4 can be written as

$$n = MV \tag{S1.5}$$

so

$$\text{moles of KBr} = (0.600\ \text{M})(250\ \text{mL})\left(\frac{1\ \text{L}}{1000\ \text{mL}}\right) = 0.150\ \text{mol}$$

We can convert moles to grams by multiplying by the formula mass of KBr(s):

$$\text{mass of KBr} = (0.150\ \text{mol KBr})\left(\frac{119.0\ \text{g KBr}}{1\ \text{mol KBr}}\right) = 17.9\ \text{g}$$

To prepare the solution, we add 17.9 grams of KBr(s) to a 250-mL volumetric flask that is partially filled with distilled water. We swirl the flask until the salt is dissolved and then dilute the solution to the 250-mL mark on the flask and swirl it again to assure uniformity. We would not add the KBr(s) to 250 mL of water, because the volume of the resulting solution would not necessarily be 250 mL.

PRACTICE PROBLEM S1-8: Ammonium selenate, $(NH_4)_2SeO_4(s)$, is used as a mothproofing agent. Describe how you would prepare 0.500 L of a 0.155-M aqueous solution of ammonium selenate.

See solution to Practice Problem 12-1 online at: McQuarrieGeneralChemistry.com

Answer: Dissolve 13.9 grams of ammonium selenate in less than 500 mL of water and then dilute to 0.500 L using a volumetric flask.

Occasionally, the concentration of a solution is given as the mass percentage of the solute. For example, commercial sulfuric acid is sold as a solution that is 96.7% H_2SO_4 and 3.3% water by mass. If you know the density of such a solution, you can calculate its molarity. The density of the sulfuric acid solution is 1.84 g·mL^{-1} at 20°C. The mass of H_2SO_4 in one liter of solution is given by

$$\begin{pmatrix}\text{mass of } H_2SO_4 \\ \text{per liter of solution}\end{pmatrix} = \left(\frac{1000\ \text{mL}}{1\ \text{L}}\right)\left(\frac{1.84\ \text{g solution}}{1\ \text{mL}}\right)\left(\frac{96.7\ \text{g } H_2SO_4}{100\ \text{g solution}}\right)$$

$$= 1780\ \text{g } H_2SO_4 \text{ per liter of solution}$$

and the number of moles of $H_2SO_4(aq)$ per liter of solution—or, in other words, the molarity—is given by

$$\text{molarity of } H_2SO_4(aq) = \left(\frac{1780\ \text{g } H_2SO_4}{1\ \text{L solution}}\right)\left(\frac{1\ \text{mol } H_2SO_4}{98.08\ \text{g } H_2SO_4}\right) = 18.1\ \text{M}$$

EXAMPLE S1-9: Ammonia is sold as an aqueous solution that is 28% NH_3 by mass and has a density of 0.90 g·mL^{-1} at 20°C. Calculate the molarity of this solution.

Solution: The mass of NH_3 in one liter of solution is

$$\left(\begin{array}{c}\text{mass of } NH_3\\ \text{per liter of solution}\end{array}\right) = \left(\frac{1000\text{ mL}}{1\text{ L}}\right)\left(\frac{0.90\text{ g solution}}{1\text{ mL solution}}\right)\left(\frac{28\text{ g } NH_3}{100\text{ g solution}}\right)$$

$$= 250\text{ g } NH_3 \text{ per liter of solution}$$

The molarity is given by

$$\text{molarity of } NH_3 = \left(\frac{250\text{ g } NH_3}{1\text{ L solution}}\right)\left(\frac{1\text{ mol } NH_3}{17.03\text{ g } NH_3}\right) = 15\text{ M}$$

See solution to Practice Problem 12-2 online at: McQuarrieGeneralChemistry.com

PRACTICE PROBLEM S1-9: A concentrated sodium hydroxide solution is 50.0% NaOH by mass and has a density of 1.525 g·mL^{-1} at 20°C. Calculate the molarity of the solution.

Answer: 19.1 M

It is often necessary in laboratory work to prepare a more dilute solution from a more concentrated stock solution, as we might do with the concentrated sodium hydroxide solution above. In such cases, a certain volume of a solution of known molarity is diluted with a certain volume of pure solvent to produce the final solution with the desired molarity. The key point to recognize in carrying out such **dilution** calculations is that the number of moles of solute does not change on dilution with solvent (Figure S1.8). Thus, from Equation S1.5, we have

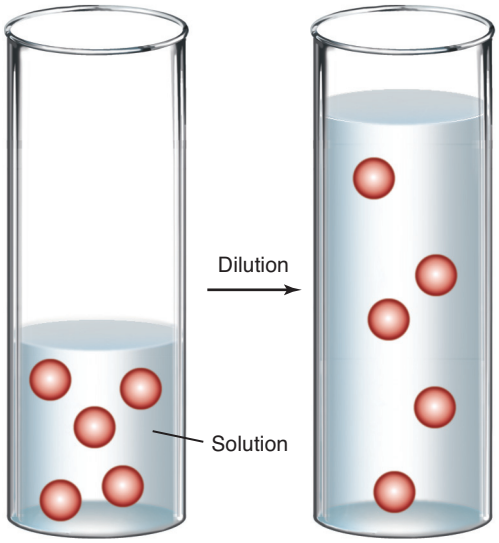


Figure S1.8 When we dilute a solution, the volume of solvent is increased but the number of moles of solute (spheres) stays the same.

$$\text{number of moles of solute before dilution} = n_1 = M_1V_1$$

$$\text{number of moles of solute after dilution} = n_2 = M_2V_2$$

But $n_1 = n_2$, and so we have

$$M_1V_1 = M_2V_2 \qquad \text{(dilution)} \qquad \text{(S1.6)}$$

The following Example illustrates a dilution calculation.

EXAMPLE S1-10: Compute the volume of the 19.1-M concentrated NaOH(aq) solution in Practice Problem S1-9 required to produce 500 mL of 3.0-M NaOH(aq).

Solution: From Equation S1.6, we have

$$M_1V_1 = M_2V_2$$

$$(19.1\ \text{mol}\cdot\text{L}^{-1})\ (V_1) = (3.0\ \text{mol}\cdot\text{L}^{-1})\ (0.500\ \text{L})$$

Thus,

$$V_1 = \frac{(0.500\ \text{L})(3.0\ \text{mol}\cdot\text{L}^{-1})}{19.1\ \text{mol}\cdot\text{L}^{-1}} = 0.079\ \text{L}$$

or V_1 = 79 mL. To make the 3.0 M NaOH(aq) solution, we add 79 mL of 19.1 M NaOH(aq) to a 500-mL volumetric flask that is about half-filled with water, swirl the solution, and dilute with water to the 500-mL mark on the flask. Finally, we swirl again to make the new solution homogeneous.

PRACTICE PROBLEM S1-10: Commercial nitric acid, $HNO_3(aq)$, is a 15.9-M aqueous solution. How would you prepare one liter of 6.00 M $HNO_3(aq)$ solution from commercial nitric acid?

Answer: Dilute 377 mL of the 15.9 M $HNO_3(aq)$ to one liter using a volumetric flask.

See solution to Practice Problem 12-3 online at: McQuarrieGeneralChemistry.com

S1-6. Molarity Is Used in Stoichiometric Calculations for Reactions That Take Place in Solution

The concept of molarity allows us to extend the types of calculations we discussed in Section S1-3 to reactions that take place in solution. For example, a standard laboratory preparation of small quantities of bromine involves the reaction described by the chemical equation

$$MnO_2(s) + 4\,HBr(aq) \rightarrow MnBr_2(aq) + Br_2(l) + 2\,H_2O(l)$$

What volume of 8.84 M HBr(aq) solution would be required to react completely with 3.62 grams of $MnO_2(s)$?

From the chemical equation, we see that one mole of $MnO_2(s)$ requires four moles of HBr(aq) to react completely, or that

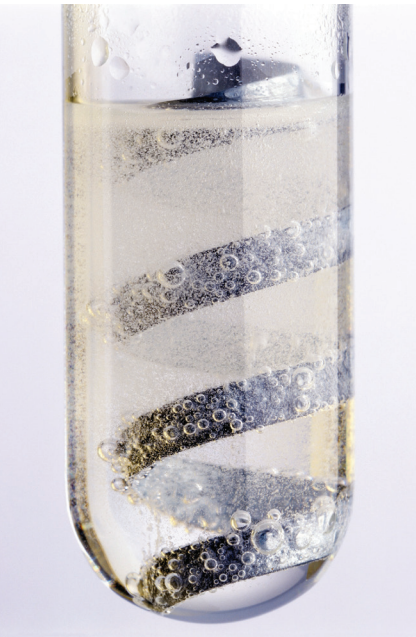


Figure S1.9 Zinc metal reacts with an aqueous solution of hydrochloric acid. The bubbles are hydrogen gas escaping from the solution.

$$1 \text{ mol } MnO_2 \Leftrightarrow 4 \text{ mol HBr}$$

The number of moles of $MnO_2(s)$ is given by

$$\text{moles of } MnO_2 = (3.62 \text{ g } MnO_2)\left(\frac{1 \text{ mol } MnO_2}{86.94 \text{ g } MnO_2}\right) = 0.0416 \text{ mol}$$

and the corresponding number of moles of $HBr(aq)$ required is

$$\text{moles of HBr} = (0.0416 \text{ mol } MnO_2)\left(\frac{4 \text{ mol HBr}}{1 \text{ mol } MnO_2}\right) = 0.166 \text{ mol}$$

When working with molarity in chemical calculations it is often helpful to rewrite the unit M as a unit conversion factor between liters and moles. For example, the concentration of the 8.84 M $HBr(aq)$ solution may be expressed as either

$$\frac{8.84 \text{ mol HBr}}{1 \text{ L}} = 1 \quad \text{or} \quad \frac{1 \text{ L}}{8.84 \text{ mol HBr}} = 1$$

Making use of the second expression above, the volume of the solution is given by

$$\text{volume of solution} = (0.166 \text{ mol HBr})\left(\frac{1 \text{ L}}{8.84 \text{ mol HBr}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 18.8 \text{ mL}$$

We could have also determined the volume required by solving Equation S1.5 for V to obtain

$$V = \frac{n}{M} = \frac{0.166 \text{ mol}}{8.84 \text{ mol}\cdot\text{L}^{-1}} = 0.0188 \text{ L} = 18.8 \text{ mL}$$

The following Example illustrates a calculation involving a reaction between a solution and a solid.

Figure S1.10 Drāno® consists of a mixture of pieces of aluminum and $NaOH(s)$. When Drāno® is added to water, the aluminum reacts with the $NaOH(aq)$ to produce hydrogen gas.

EXAMPLE S1-11: Zinc reacts with hydrochloric acid, $HCl(aq)$ (Figure S1.9), according to the equation

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

Calculate how many grams of zinc react with 50.0 mL of 6.00 M $HCl(aq)$.

Solution: The equation for the reaction indicates that one mole of $Zn(s)$ reacts with two moles of $HCl(aq)$. First we determine the number of moles of $HCl(aq)$ in 50.0 mL of a 6.00 M $HCl(aq)$ solution:

$$\text{moles of HCl} = (50.0 \text{ mL})\left(\frac{1 \text{ L}}{1000 \text{ mL}}\right)\left(\frac{6.00 \text{ mol}}{1 \text{ L}}\right) = 0.300 \text{ mol}$$

The number of grams of zinc that react is given by

$$\text{mass of Zn} = (0.300 \text{ mol HCl})\left(\frac{1 \text{ mol Zn}}{2 \text{ mol HCl}}\right)\left(\frac{65.41 \text{ g Zn}}{1 \text{ mol Zn}}\right) = 9.81 \text{ g Zn}$$

In essence the molarity of the solution and the stoichiometric coefficients from the balanced equation are conversion factors that allow us to convert the volume of HCl(*aq*) given to the mass of zinc reacted.

PRACTICE PROBLEM S1-11: Aluminum reacts with a moderately concentrated sodium hydroxide solution (Figure S1.10) according to:

$$2\,Al(s) + 2\,NaOH(aq) + 6\,H_2O(l) \rightarrow 2\,NaAl(OH)_4(aq) + 3\,H_2(g)$$

How many grams of aluminum will react with 30.0 mL of 6.00 M NaOH(*aq*) solution?

Answer: 4.86 grams

See solution to Practice Problem 12-6 online at: McQuarrieGeneralChemistry.com

TERMS YOU SHOULD KNOW

atomic substances *S1-2*
molecular substances *S1-2*
formula mass *S1-2*
formula unit *S1-2*
mole (mol) *S1-2*
molar mass *S1-2*
Avogadro's number *S1-4*
balancing coefficient *S1-8*
stoichiometric coefficient *S1-8*
stoichiometric unit conversion factor *S1-9*
solute *S1-16*
solvent *S1-16*
saturated solution *S1-16*
solubility, grams of solute per 100 g of solvent *S1-16*
unsaturated solution *S1-17*
concentration *S1-17*
molarity, moles of solute per 1 L of solution, *M* *S1-17*
molar, M *S1-17*
volumetric flask *S1-17*
dilution *S1-20*

EQUATIONS YOU SHOULD KNOW HOW TO USE

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \quad \text{(S1.3)}$$

$$M = \frac{n}{V} \quad \text{(S1.4) definition of molarity}$$

$$n = MV \quad \text{(S1.5) number of moles of solute in a volume of solution (volume expressed in liters)}$$

$$M_1V_1 = M_2V_2 \quad \text{(S1.6) used for dilution calculations}$$

PROBLEMS

The number in parentheses indicates the corresponding problem number from the main text and the solutions manual.

MOLES

S1-1 (11-1). Calculate the number of moles in

(a) 52.0 grams of calcium carbonate, $CaCO_3(s)$

(b) 250.0 milliliters of ethanol, $CH_3CH_2OH(l)$, with a density of 0.76 g·mL^{-1}

(c) 28.1 grams of carbon dioxide gas, $CO_2(g)$

(d) 5.55×10^{22} molecules of sulfur hexafluoride gas, $SF_6(g)$

S1-2 (11-2). Calculate the mass in grams of

(a) 3.00 moles of Hg(*l*)

(b) 1.872×10^{24} molecules of iron(III) hydroxide, $Fe(OH)_3(s)$

(c) 1.0 mole of ^{18}O atoms

(d) 2.0 moles of nitrogen gas, $N_2(g)$

CALCULATIONS INVOLVING CHEMICAL REACTIONS

S1-3 (11-31). The combustion of propane may be described by the chemical equation

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

How many grams of oxygen are required to burn completely 10.0 grams of propane?

S1-4 (11-32). Iodine is prepared both in the laboratory and commercially by adding $Cl_2(g)$ to an aqueous solution containing sodium iodide according to

$$2\,NaI(aq) + Cl_2(g) \rightarrow I_2(s) + 2\,NaCl(aq)$$

How many grams of sodium iodide must be used to produce 50.0 grams of iodine?

S1-5 (11-33). Small quantities of chlorine can be prepared in the laboratory by the reaction described by the equation

$$MnO_2(s) + 4\,HCl(aq) \rightarrow MnCl_2(aq) + Cl_2(g) + 2\,H_2O(l)$$

How many grams of chlorine can be prepared from 100.0 grams of manganese(II) oxide?

S1-6 (11-34). Small quantities of oxygen can be prepared in the laboratory by heating potassium chlorate, $KClO_3(s)$. The equation for the reaction is

$$2\,KClO_3(s) \rightarrow 2\,KCl(s) + 3\,O_2(g)$$

Calculate how many grams of $O_2(g)$ can be produced from heating 10.0 grams of $KClO_3(s)$.

MOLARITY AND CONCENTRATION

S1-7 (12-1). Calculate the molarity of a saturated solution of sodium hydrogen carbonate (baking soda), $NaHCO_3(aq)$ that contains 69.0 grams in 1.00 liter of solution.

S1-8 (12-2). Sodium hydroxide is extremely soluble in water. A saturated solution contains 572 grams of $NaOH(s)$ per liter of solution. Calculate the molarity of a saturated $NaOH(aq)$ solution.

S1-9 (12-3). A saturated solution of calcium hydroxide, $Ca(OH)_2(aq)$, contains 0.185 grams per 100 milliliters of solution. Calculate the molarity of a saturated calcium hydroxide solution.

S1-10 (12-4). A cup of coffee may contain as much as 300 milligrams of caffeine, $C_8H_{10}N_4O_2(s)$. Calculate the molarity of caffeine in one cup of coffee (4 cups = 0.946 liters).

S1-11 (12-5). Calculate the number of moles of solute in

(a) 25.46 milliliters of a 0.1255 M $K_2Cr_2O_7(aq)$ solution

(b) 50 μL of a 0.020 M $C_6H_{12}O_6(aq)$ solution

S1-12 (12-6). Calculate the number of moles of solute in

(a) 50.0 μL of a 0.200 M $NaCl(aq)$ solution

(b) 2.00 milliliters of a 2.00-mM $H_2SO_4(aq)$ solution

PREPARATIONS OF SOLUTIONS

S1-13 (12-7). How many milliliters of 18.0 M $H_2SO_4(aq)$ are required to prepare 500 milliliters of 0.30 M $H_2SO_4(aq)$?

S1-14 (12-8). How many milliliters of 12.0 M $HCl(aq)$ are required to prepare 250 milliliters of 1.0 M $HCl(aq)$?

REACTIONS IN SOLUTION

S1-15 (12-17). When silicon is heated with an aqueous solution of sodium hydroxide, sodium silicate and hydrogen gas are formed according to

$$Si(s) + 2\,NaOH(aq) + H_2O(l) \rightarrow Na_2SiO_3(aq) + 2\,H_2(g)$$

How many milliliters of 6.00 M $NaOH(aq)$ are required to react with 12.5 grams of silicon? How many grams of hydrogen gas will be produced?

S1-16 (12-18). Hydrogen peroxide can be prepared by the reaction of barium peroxide with sulfuric acid according to the equation

$$BaO_2(s) + H_2SO_4(aq) \rightarrow BaSO_4(s) + H_2O_2(aq)$$

How many milliliters of 3.75 M $H_2SO_4(aq)$ are required to react completely with 17.6 grams of $BaO_2(s)$?

S1-17 (12-23). Zinc reacts with hydrochloric acid according to the reaction equation

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

How many milliliters of 2.00 M $HCl(aq)$ are required to react with 2.55 grams of $Zn(s)$?

S1-18 (12-24). Bromine is obtained commercially from natural brines from wells in Michigan and Arkansas by the reaction described by the equation

$$Cl_2(g) + 2\,NaBr(aq) \rightarrow 2\,NaCl(aq) + Br_2(l)$$

If the concentration of $NaBr(aq)$ is 4.00×10^{-3} M, how many grams of bromine can be obtained per cubic meter of brine? How many grams of chlorine are required?

ADDITIONAL PROBLEMS

S1-19 (11-70). Determine the mass of ammonium nitrate, $NH_4NO_3(s)$, that has the same number of nitrogen atoms as 2.0 liters of liquid nitrogen, $N_2(l)$. Take the density of liquid nitrogen to be $0.808\ g{\cdot}mL^{-1}$.

S1-20 (11-74). Lithium is the only Group 1 metal that yields the normal oxide, $Li_2O(s)$, when it is burned in excess oxygen. The other alkali metals react with excess oxygen as shown below.

(a) $2\,Na(s) + O_2(g) \rightarrow \underset{\text{sodium peroxide}}{Na_2O_2(s)}$

(b) $K(s) + O_2(g) \rightarrow \underset{\text{potassium superoxide}}{KO_2(s)}$

(c) $Rb(s) + O_2(g) \rightarrow \underset{\text{rubidium superoxide}}{RbO_2(s)}$

(d) $Cs(s) + O_2(g) \rightarrow \underset{\text{cesium superoxide}}{CsO_2(s)}$

Calculate how much product is formed when 0.600 grams of each alkali metal are burned in excess oxygen.

S1-21 (11-76). Glucose is used as an energy source by the human body. The overall reaction in the body is described by the equation

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

Calculate the number of grams of oxygen required to convert 28.0 grams of glucose to $CO_2(g)$ and $H_2O(l)$. Also compute the number of grams of $CO_2(g)$ produced.

S1-22 (11-77). A hydrated form of copper(II) sulfate, $CuSO_4{\cdot}nH_2O(s)$, is heated to drive off all the waters of hydration. If we start with 9.40 grams of hydrated salt and have 5.25 grams of anhydrous $CuSO_4(s)$ after heating, find the number of water molecules, n, associated with each $CuSO_4$ formula unit.

S1-23 (11-86). Nitric acid, $HNO_3(aq)$, is made commercially from ammonia by the Ostwald process, which was developed by the German chemist Wilhelm Ostwald. The process consists of three steps:

(1) $4\,NH_3(g) + 5\,O_2(g) \rightarrow 4\,NO(g) + 6\,H_2O(g)$

(2) $2\,NO(g) + O_2(g) \rightarrow 2\,NO_2(g)$

(3) $3\,NO_2(g) + H_2O(l) \rightarrow 2\,HNO_3(aq) + NO(g)$

How many kilograms of nitric acid can be produced from 6.40×10^4 kilograms of ammonia?

S1-24 (12-9). Explain how you would prepare 500 milliliters of a 0.250-M aqueous solution of sucrose, $C_{12}H_{22}O_{11}(aq)$. This solution is used frequently in biological experiments.

S1-25 (12-45). What molarity of a $CaCl_2(aq)$ solution should you use if you want the molarity of the $Cl^-(aq)$ ions in the solution to be 0.100 M?

S1-26 (12-52). Describe how you would prepare 250 milliliters of 0.12-M sulfuric acid from a stock solution labeled 8.0-M sulfuric acid.

S1-27* (11-94). In 1773 Ben Franklin wrote in a letter about calming the waves on Clapham pond using a small quantity of oil,

> "*...the oil, though not more than a teaspoonful, produced an instant calm over a space several yards square which spread amazingly and extended itself gradually till it reached the lee side, making all that quarter of the pond, perhaps half an acre, as smooth as a looking glass.*"

Assuming Franklin used castor oil, which has a formula mass of about 180 and a density of about $0.96\ g{\cdot}mL^{-1}$, and that the oil forms a one-molecule-thick monolayer on the pond, *estimate* Avogadro's number using Franklin's data. Why must we assume the oil forms a monolayer? What other assumptions must we make?