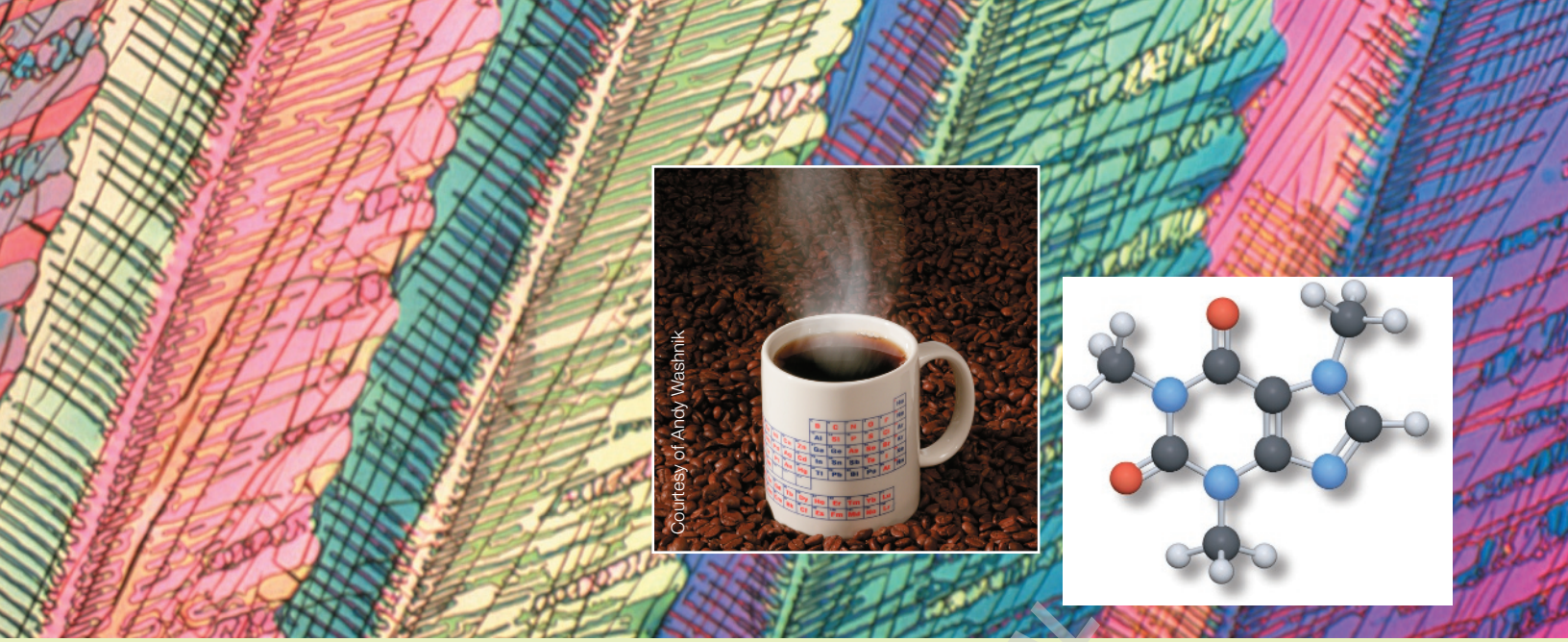


Michael W. Davidson at Florida State University

1 FUNDAMENTAL CONCEPTS OF CHEMISTRY

CHAPTER CONTENTS

1.1 Atoms, Molecules, and Compounds
1.2 Measurements in Chemistry
1.3 Chemical Problem Solving
1.4 Counting Atoms: The Mole
1.5 Amounts of Compounds
1.6 Aqueous Solutions
1.7 Writing Chemical Equations
1.8 The Stoichiometry of Chemical Reactions
1.9 Yields of Chemical Reactions
1.10 The Limiting Reactant

LEARNING OBJECTIVES

Upon completion of this chapter you should be able to

- recognize elemental symbols and names of the elements, and name compounds from molecular pictures
- recognize the SI units commonly used in chemistry, and perform some common unit conversions
- analyze and solve problems in a consistent, organized fashion
- solve mass–number–molar mass-type problems
- perform mole–mass–number conversions
- calculate concentrations of solutions and of diluted solutions
- balance chemical reactions
- calculate the amount of a product from the amounts of the reactants and a balanced chemical reaction
- calculate yields of chemical reactions
- solve limiting-reagent-type problems

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OUR WORLD IS MOLECULAR

How would you describe the differences between a cup of coffee and a cup of hot water? What probably come to mind are the aroma, the dark colour, and the taste of a good cup of coffee. Coffee's action as a stimulant is another obvious difference. These properties come from the chemical compounds that hot water dissolves from ground coffee beans. These compounds are molecules constructed from different atoms bound together in very specific arrangements. The molecule that makes coffee a stimulant is caffeine. Our background photo is a magnification of crystals of pure caffeine, and the inset is a ball-and-stick model of this molecule.

The properties of a particular molecule are due to the types and number of atoms it contains and how those atoms are arranged in space. Caffeine is a stimulant because it has the same shape as one part of cyclic adenosine monophosphate (cyclic AMP), a molecule that helps to regulate the supply of energy in the brain. Once caffeine has been absorbed into the blood and carried to the brain, it binds to an enzyme that normally binds with and controls the supply of cyclic AMP. As a result, the enzyme can no longer bind with cyclic AMP, the brain's supply of this energy-regulating molecule is increased, and we feel stimulated.

Molecular sizes and shapes play key roles in determining chemical and physical properties. The immense variety of chemical and physical properties displayed by substances in the natural world mirrors an equally immense variety of different types of molecules. However, variety need not come from a large number of different elements. The molecules that make up a cup of coffee are made up almost entirely of atoms of just five elements: hydrogen, carbon, oxygen, nitrogen, and sulphur. Carbon, in particular, is capable of combining in many different ways, generating molecules with elaborate structures.

Even simple molecules often have strikingly different properties. For example, carbon and oxygen are the base elements of two different simple compounds. Whereas a molecule of carbon monoxide contains one oxygen atom and one carbon atom, carbon dioxide contains two atoms of oxygen and one atom of carbon. Although these molecules have some common properties (both are colourless, odourless gases), the difference in chemistry caused by a change of one atom is profound. We produce and exhale carbon dioxide as a natural by-product of metabolism. This compound is relatively harmless to humans. In contrast, carbon monoxide is a deadly poison, even at very low concentrations.

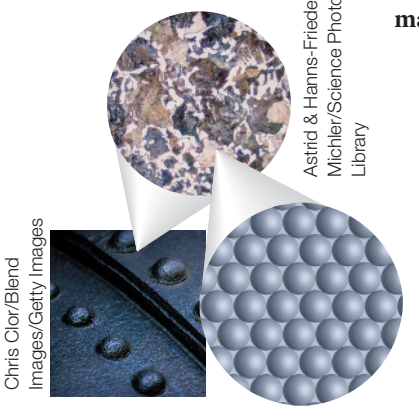
This chapter introduces the ideas of atoms and molecules, how to quantify them, and the basics of chemical reactions.

KEY CONCEPT

Properties of molecules (the microscopic level) translate into properties of materials (the macroscopic level).

1.1 ATOMS, MOLECULES, AND COMPOUNDS

Every substance has physical properties that we can measure and describe, including shape, colour, and texture. For example, the iron girder shown in Figure 1-1 has a lustrous silvery colour and a smooth texture. Substances also have chemical properties, such as their flammability. **Macroscopic** physical and chemical properties can be observed with the eye. The underlying or **microscopic** structure of a chemical substance requires magnification to be observed. The magnified view of iron shown in Figure 1-1 reveals fissures and pits that are not visible in the macroscopic view. Still further magnification eventually reveals the building blocks of **matter**. This, the **molecular** or **atomic view**, is an essential part of every chemist's thinking.



Astrid & Hanns-Frieder Michler/Science Photo Library

Chris Clor/Blend Images/Getty Images

Figure 1-1

Iron appears different at the macroscopic, microscopic, and atomic levels.

Atoms

The fundamental unit of a chemical substance is called an **atom**. The word is derived from the Greek word *atomos,* meaning "uncuttable." Atoms are extremely small, with diameters of the order of one ten billionth of a metre. To give you some idea of just how small that is, a sample of carbon the size of the period at the end of this sentence contains more atoms than the number of stars in the Milky Way. Any sample of matter large enough for us to see or feel contains an unfathomable number of atoms. Many experiments confirm the existence of atoms. Focus 1-1 details some of the more recent evidence.

Molecules

Atoms combine to make all the substances in the world around us, but they do so in very orderly ways. Most substances that we encounter in day-to-day life are made up of small units called molecules.

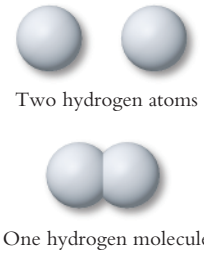


Figure 1-2

A hydrogen molecule can be represented by connecting two spheres together, with each sphere representing one hydrogen atom.

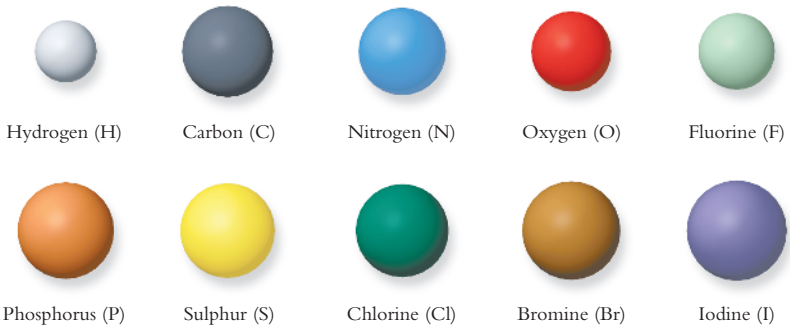


Figure 1-3

Colour-coded scale models of 10 types of atoms that appear frequently in this book.

The simplest molecules contain just two atoms and are termed **diatomic**. For example, a molecule of hydrogen is made up of two hydrogen atoms. Figure 1-2 represents a diatomic hydrogen molecule as two spheres connected together.

Because chemistry deals mostly with the behaviour of molecules, this book emphasizes chemistry's molecular foundation. Throughout this book, you will see many figures that represent molecules, with each atom represented by a sphere. Although there are more than 100 different types of atoms, only about 20 are encountered frequently in our world. Many of the molecules described in this book are made up of just 10 different types of atoms: hydrogen, carbon, nitrogen, oxygen, fluorine, phosphorus, sulphur, fluorine, chlorine, bromine, and iodine. Figure 1-3 shows the colour scheme that we use to represent these atoms.

With practice, you can recognize simple molecules, such as carbon dioxide and water, just by looking at their models. Figure 1-4 shows scale models of a few molecules whose names may be familiar to you. The structures of larger molecules, however, are too complex to be recognized at a glance. Consequently, chemists have created a shorthand language of symbols, formulas, and equations that convey information about atoms and molecules in a simple manner. A specific letter or pair of letters designates each type of atom. These symbols, in turn, are combined into formulas that describe the compositions of more complex chemical substances. Formulas can then be used to write chemical equations that describe how molecules change during a chemical reaction.

The name *hydrogen* refers to both atoms and molecules. To minimize confusion, we refer to *atomic hydrogen* when we mean hydrogen atoms and *molecular hydrogen* when we mean hydrogen molecules.

Chemical Formulas

A chemical **compound** is a substance that contains more than one **element**. The relative amounts of the elements in a particular compound do not change: every molecule of a particular chemical substance contains a characteristic number of atoms of its constituent elements.

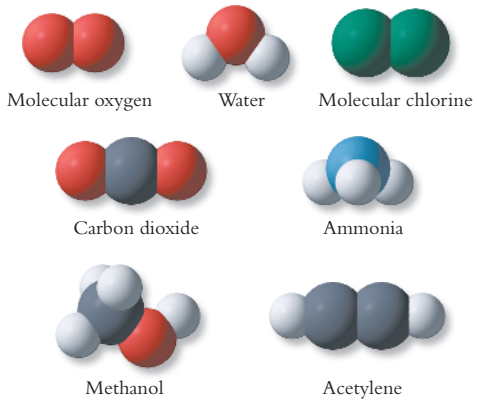
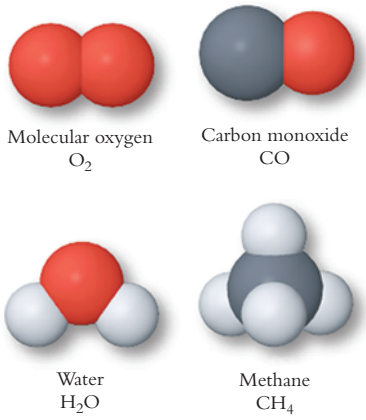


Figure 1-4

Models of seven relatively simple molecules.

For example, every water molecule contains two hydrogen atoms and one oxygen atom. To describe this atomic composition, chemists write the **chemical formula** for water as H_2O.

The chemical formula for water shows how formulas are constructed. The formula lists the **elemental symbols** found in the compound, in this case H (hydrogen) and O (oxygen). A subscript number after an element's symbol denotes how many atoms of that element are present in the molecule. The subscript 2 in the formula for water indicates that each molecule contains two hydrogen atoms. No subscript is used when only one atom of an element is present, as is the case for the oxygen atom in a water molecule. Molecules always contain whole numbers of atoms, and so the subscripts in chemical formulas of molecular substances are always integers.

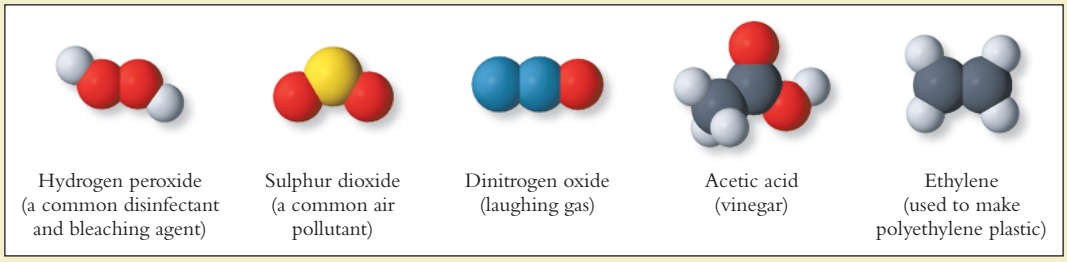


Molecular oxygen O_2 | Carbon monoxide CO | Water H_2O | Methane CH_4

Molecules vary considerably in complexity. Molecular oxygen is made up of two oxygen atoms, so its chemical formula is O_2. A carbon monoxide molecule contains one atom of carbon and one atom of oxygen, so its chemical formula is CO. Each molecule of methane, the major constituent of natural gas, contains one carbon atom and four hydrogen atoms, so its formula is CH_4. You will encounter still more complex structures, such as methanol (made up of three different elements, CH_4O) and pantothenic acid (made up of four different elements, $C_9H_{17}NO_5$) as you progress through this book.

SECTION EXERCISES

- **1.1.1** What are the elemental symbols for cerium, cesium, copper, calcium, and carbon?
- **1.1.2** What are the names of the elements represented by the symbols Zr, Ni, Sn, W, Se, Be, and Au?
- **1.1.3** Molecular pictures of some common molecules are shown here. What are their chemical formulas?



Hydrogen peroxide (a common disinfectant and bleaching agent) | Sulphur dioxide (a common air pollutant) | Dinitrogen oxide (laughing gas) | Acetic acid (vinegar) | Ethylene (used to make polyethylene plastic)

NEED MORE PRACTICE? Try Problems 1.1–1.8

1.2 MEASUREMENTS IN CHEMISTRY

The knowledge that allows chemists to describe, interpret, and predict the behaviour of chemical substances is gained by making careful experimental measurements. The properties of a sample can be divided into **physical properties**, which can be measured without observing a chemical reaction, and **chemical properties**, which are displayed only during a **chemical transformation**. Physical properties include familiar attributes such as size, colour, and mass. Some chemical properties also are familiar to us. As examples, bleach reacts chemically with many coloured substances to destroy their colours, molecular oxygen reacts chemically with many fuels to generate heat, and salt is used to lower the freezing point of water on Canadian roads in the winter.

Physical Properties

Length, area, and **volume** measure the size of an object. In addition to its volume, every object possesses a certain quantity of matter, called its **mass** (*m*). Mass measurements are particularly important in chemistry. Consequently, highly accurate balances are essential instruments in chemistry laboratories. A modern balance, whether in a delicatessen or in a chemistry laboratory (Figure 1-5), provides a quick and accurate digital readout of an object's mass.



Figure 1-5

Masses are determined using balances that compare the gravitational force acting on the object being weighed to that on a standard mass. (*a*) Delicatessen balance. (*b*) A high precision balance typical of those in chemistry laboratories.

Chemists measure time (t) because they want to know how long it takes for chemical transformations to occur. Some chemical reactions, such as the conversion of green plants into petroleum, may take millions of years. Other chemical processes, such as an explosion of dynamite, are incredibly fast. Whereas wristwatches typically measure time only to the nearest second, scientists have developed instruments that make it possible to study processes that occur in less than 10^{-18} (0.000 000 000 000) second.

Most of us associate **temperature** (T) with the concepts of hot and cold. More accurately, however, temperature is the property of matter that determines the direction of heat flow. Heat flows naturally from a warm object to a cool object, from higher temperature to lower temperature. Heat is a form of energy, and because energy changes in chemical systems have important consequences, chemists are interested in temperature changes that occur during chemical transformations.

All experimental sciences rely on quantitative measurements of properties. Every measurement gives a numerical result that has three aspects: a numerical **magnitude**; an indicator of scale, called a **unit**; and a **precision**. Each aspect is essential, and all three must be reported to make a measurement scientifically valuable.

FOCUS 1-1 Tools for Discovery: Atomic-Level Microscopy

Atoms and molecules are much too tiny to see, even with the most powerful light microscopes. In recent years, however, scientists have developed a set of immensely powerful magnifying techniques that make it possible to visualize how individual atoms are arranged in solids. These techniques are scanning tunnelling microscopy (STM) and atomic force microscopy (AFM).

Scanning tunnelling microscopy and AFM generate images of atoms and molecules using highly miniaturized lever arms with atomically sharp tips. These probes respond to the contours of individual atoms. In AFM, the tip can be moved with extreme precision across a surface. The tip responds to individual atoms on that surface, moving up and down by tiny amounts as it passes over each atom. A laser beam focused on the edge of the lever arm detects this tiny motion, and an optical detector creates an image of the surface.

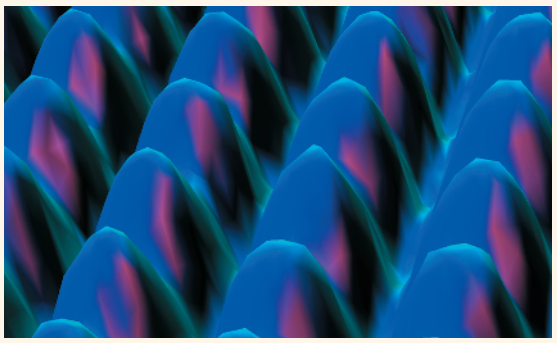
At their highest sensitivities, STM and AFM generate images that show how atoms are arranged on the surfaces they probe. Figure (*a*) shows individual atoms on the surface of nickel metal.

Increasingly, scientists are using STM and AFM to explore chemical reactions. For example, molecules often undergo chemical reactions when they adhere to a metal surface. Scanning tunnelling microscopy can be used to monitor the atomic changes that take place during such chemical reactions. It is even possible to use STM to manipulate individual atoms. Figures (*b*) and (*c*) show how iron atoms can be arranged on the surface of copper metal to make an "atomic corral."

Whereas STM is best suited for imaging atoms, AFM is more appropriate for larger structures. Deoxyribonucleic acid (DNA) is visible as a strand (Figure (*d*)), surrounded by coloured regions that are protein molecules bound to the DNA.

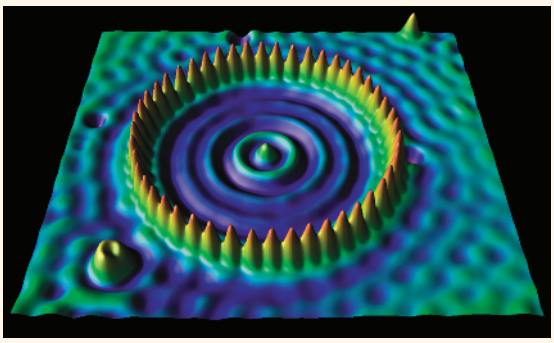
(*continued*)

The next generation of molecular probes combine magnetic resonance imaging (MRI) with the atomic force microscope. Physicians use MRI to perform brain scans to pinpoint tumours or blood clots. By coupling MRI with an atomic force scanner and cooling the sample to extremely low temperatures, it is now possible to detect an individual spinning electron in an atom.



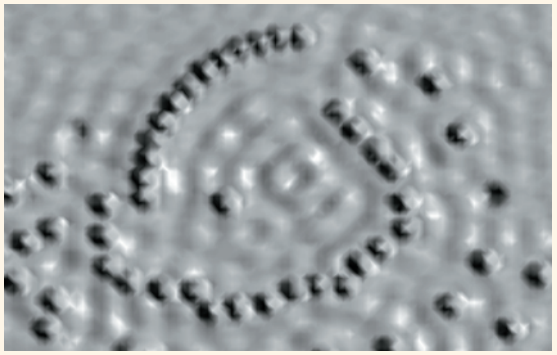
Courtesy of IBM Corporate Archives, Almaden

(*a*) Nickel metal surface



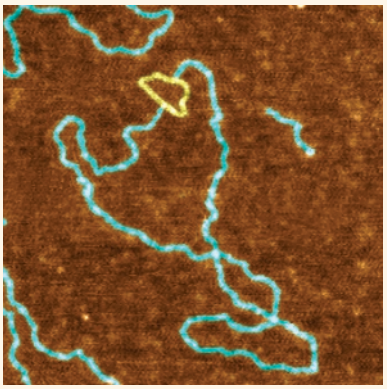
Courtesy of IBM Corporate Archives, Almaden

(*c*) A complete "corral" of iron atoms



Courtesy of IBM Corporate Archives, Almaden

(*b*) Iron atoms forming a ring



Torunn Berge/Science Photo Library

(*d*) A strand of DNA

Magnitude

The magnitudes of experimental values in science range from vanishingly small to astronomically large, as summarized in Figure 1-6. To simplify manipulating very large and very small numbers, we use **scientific notation** (see Appendix A for a review). For example, the diameter of a carbon atom is 0.000 000 0003 metres. This number is more conveniently expressed as 3×10^{-10} m.

Scientists routinely study objects whose sizes extend far beyond the narrow range encountered in daily life. Physicists, for example, study atomic nuclei measuring 10^{-15} m across, and astronomers study our universe, which spans about 10^{23} m. Chemists are most often interested in matter on the smaller side of this range. Length measurements in the laboratory vary from metres to subatomic sizes, 10^{-12} m. To further simplify the use of very large and very small numbers, scientists use prefixes that change the unit sizes by multiples of 10. For example, the prefix pico means "10^{-12}." The symbol for picometre is pm. The diameter of a carbon atom is 3×10^{-10} m, which is 300 pm:

$$\begin{aligned} \text{Carbon atom diameter} &= 3 \times 10^{-10}\ \text{m} = 300 \times 10^{-12}\ \text{m} \\ &= 300\ \text{pm} \qquad\qquad \uparrow \text{pico} \end{aligned}$$

The most common magnitude prefixes are listed in Table 1-1.

TABLE 1-1 Frequently Used Prefixes for Magnitudes

Prefix	Symbol	Number	Exponential Notation
atto	a	0.000 000 000 000 000 001	10^{-18}
femto	f	0.000 000 000 000 001	10^{-15}
pico	p	0.000 000 000 001	10^{-12}
nano	n	0.000 000 001	10^{-9}
micro	μ	0.000 001	10^{-6}
milli	m	0.001	10^{-3}
centi	c	0.01	10^{-2}
kilo	k	1000	10^{3}
mega	M	1 000 000	10^{6}
giga	G	1 000 000 000	10^{9}
tera	T	1 000 000 000 000	10^{12}

Units

The units associated with a numerical value are just as important as the value itself. It is *essential* to include a unit with every experimental value. A unit of measurement is an agreed-upon standard with which other values are compared. Scientists use the metre as the standard unit of length. The metre was originally chosen to be 10^{-7} times the length of a line from the North Pole to the equator. It is now defined as the distance travelled by light through a vacuum in $\frac{1}{299\,792\,458}$ of a second. Volume is measured in **litres** or cubic metres. Temperature is measured in degrees **Celsius (°C)** or **kelvins (K)**.

The international scientific community prefers to work exclusively with a single set of units, the **Système International (SI)**, which expresses each fundamental physical quantity in decimally (power of 10) related units. The six base units of the SI used in this book are listed in Table 1-2. Other units are based on some combination of these base units. The unit for volume, for example, is obtained from the base unit for length: a cube that measures 1 metre on a side has a volume of 1 cubic metre.

Unit Conversions

It is often necessary to convert measurements from one set of units to another. Chemists frequently need to convert volumes from one unit to another. One enormous benefit of using the SI system is that conversions from one unit to another are simple, often requiring only multiplying by some power of 10. For example, the SI unit of volume is the cubic

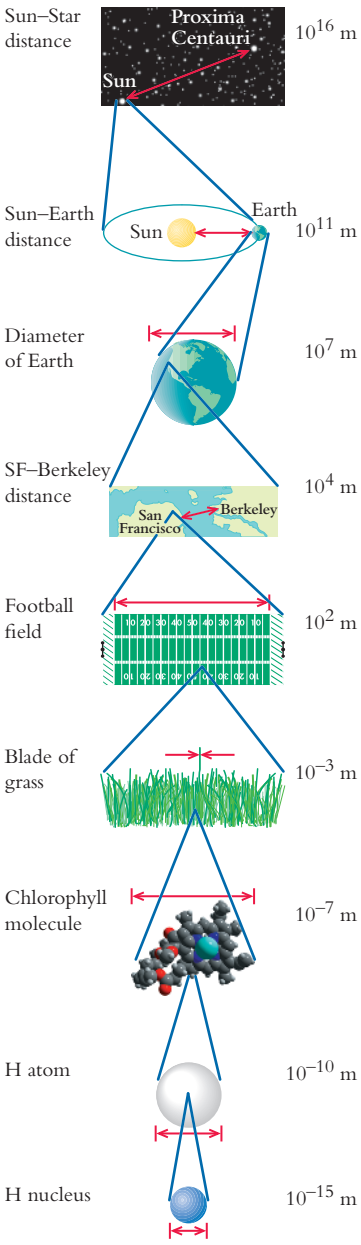


Figure 1-6

Dimensions of objects known to humans span more than 30 orders of magnitude, from interstellar distances (>10^{16} m) to the diameter of a hydrogen nucleus (~10^{-15} m).

TABLE 1-2 Base SI Units

Quantity	Unit	Symbol
Mass	kilogram	kg
Length	metre	m
Time	second	s
Temperature	kelvin	K
Amount	mole	mol
Electric current	ampere	A

metre, but chemists usually work with much smaller volumes. Hence, chemists often express volume using the litre (L), which is defined to be exactly 10^{-3} m^3. Another volume unit in common use is the millilitre (mL), or 10^{-3} L. The millilitre is the same as the cubic centimetre (cm^3).

$$1\ \text{L} = 10^{-3}\ \text{m}^3,\ 1\ \text{mL} = 10^{-3}\ \text{L} = 1\ \text{cm}^3$$

Some commonly used units are not in the SI system, however, and so it is useful to be able to perform unit conversions. A unit conversion starts with an equality between different units. For example, the equivalence between two commonly used pressure units, the bar and the Torr, is

$$1\ \text{bar} = 750.06\ \text{Torr}$$

KEY CONCEPT

The correct conversion ratio leads to cancellation of unwanted units.

Each equality gives two **conversion ratios.** The equality between bar and Torr can be rearranged to give two ratios:

$$1 = \frac{1\ \text{bar}}{750.06\ \text{Torr}} \quad \text{and} \quad 1 = \frac{750.06\ \text{Torr}}{1\ \text{bar}}$$

To convert from one unit to another, we multiply by the ratio that leads to an appropriate cancellation of units. For example, a pressure of exactly 2 bar is expressed in Torr as follows:

$$2\ \text{bar}\left(\frac{750.06\ \text{Torr}}{1\ \text{bar}}\right) = 1500.12\ \text{Torr}$$

A unit equality may link SI units and non-SI units (1 bar = 750.06 Torr), decimally related units (10^6 $cm^3 = 1$ m^3), or base units and derived units (1 L = 10^{-3} m^3). Some of the more common unit equalities are given on the inside back cover of this text. Scientists use two units for temperature, the Celsius (°C) scale and the Kelvin (K) scale. These scales are shown schematically in Figure 1-7. Unlike other scientific units, the unit size of the Celsius and Kelvin scales is the same, but their zero points differ. For both scales, the difference in temperature between the freezing and boiling points of water is defined to be 100 units. However, the temperature at which ice melts to liquid water is 0 °C and 273.15 K.

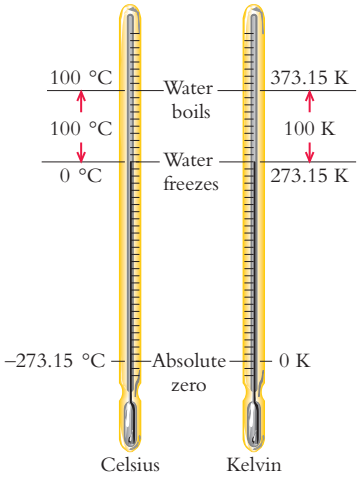


Figure 1-7

Celsius and Kelvin temperature scales. Kelvins and degrees Celsius have the same unit size but different zero points. "Room temperature" is typically about 295 K (22 °C).

The conversion between kelvins and degrees Celsius is therefore straightforward. A temperature change of 1 °C is the same as a temperature change of 1 K. To convert from one scale to the other, add or subtract 273.15:

$$T(\text{K}) = T(^\circ\text{C}) + 273.15 \quad \text{or} \quad T(^\circ\text{C}) = T(\text{K}) - 273.15$$

Precision and Accuracy

The exactness of a measurement is expressed by its **precision**. This concept can be explained with an example. Suppose three swimmers are discussing the temperature of a swimming pool. The first dips a finger in the water and says that the temperature is "about 24 °C." The second examines an immersed pool thermometer and reports the temperature to be 26 °C. The third swimmer, who has been monitoring daily variations in the pool's temperature, uses a portable precision digital thermometer and reports, "According to my precision thermometer, the pool temperature is 25.8 °C."

The swimmers have measured the water temperature using different measuring devices with different levels of precision. The first swimmer's "finger test" is precise to about 3 °C: $T = 24 \pm 3$ °C (read "twenty-four plus or minus three degrees"). The pool thermometer gives a reading that is precise to the nearest degree: $T = 26 \pm 1$ °C. The precision thermometer measures to the nearest tenth of a degree: $T = 25.8 \pm 0.1$ °C. An actual temperature between 25.7 and 25.9 °C falls within the precision ranges of all three measurements, so all three are correct to within their stated limits of precision.

Whereas precision describes the exactness of a measurement, **accuracy** describes how close a measurement is to the true value. Figure 1-8 illustrates the difference between precision and accuracy. An expert archer shoots arrows with high precision, with all the arrows

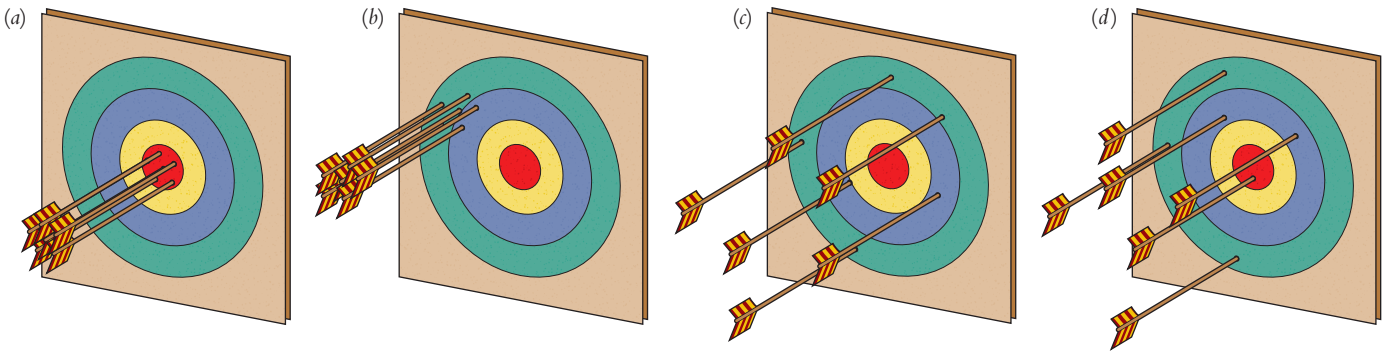


Figure 1-8

Patterns of arrows striking a target illustrate the notions of precision and accuracy. (*a*) is precise and accurate. (*b*) is precise but inaccurate. (*c*) is neither precise nor accurate. (*d*) is an imprecise pattern with one accurate shot.

hitting the target near the same spot (*a* and *b*). Under calm conditions, the expert is also quite accurate, shooting all the arrows into the bull's-eye (*a*). In a strong wind, however, the archer may be precise but not accurate, regularly missing the bull's-eye (*b*). In contrast, a novice shooter is not very precise (*c* and *d*): the arrows scatter over a large area. The novice usually is not accurate (*c*), but an occasional arrow may accurately find the bull's-eye (*d*).

The goal of any measurement in science is to be as accurate as possible. However, determining the accuracy of a measurement is much harder than determining precision. An archer can examine the target to see if the arrows found their mark, but a scientific researcher studying a new phenomenon does not know the correct value. A scientist can assess precision by making additional measurements to find out how closely repeated measurements agree with one another. In contrast, assessing accuracy requires careful attention to the design of an experiment and the instruments used in that experiment. The instruments must be used to measure some known value in order to assess their ability to accurately measure unknown quantities.

Significant Figures

KEY CONCEPT

Proper use of significant figures is important to tell the reader what the accuracy and precision of the measurement are.

Scientific measurements should always include both magnitude and precision. Indicating the precision limits with a plus/minus (±) statement is cumbersome, particularly when many numerical values are reported at one time. Scientists have agreed to simplify the reporting of precision. Experimental measurements are written so that there is an uncertainty of up to one unit in the last reported digit. For example, a temperature reported as 26 °C is greater than 25 °C but less than 27 °C, or 26 ± 1 °C. A temperature reported as 25.8 °C is 25.8 ± 0.1 °C. The number of digits expressed in a numerical value is called the number of **significant figures**. The value 26 has two significant figures, whereas 25.8 has three significant figures.

Zeros can present a problem when determining the precision of a numerical value. This is because zeros are needed both to locate the decimal point and to express precision. From the statement that the Sun is 150 000 000 km from the Earth, we cannot tell whether the measurement is precise to nine significant figures or whether the zeros are there only to put the decimal point in the right place. Scientific notation eliminates this ambiguity because a power of 10 locates the decimal point, leaving us free to indicate the precision by the number of digits. For example, a distance of 1.5×10^8 km means that the number is precise to $\pm 0.1 \times 10^8$ km; the number has two significant figures. If this distance is known to a precision of $\pm 0.01 \times 10^8$ km, then it is written as 1.50×10^8 km, with three significant figures. Zeros at the end of a number with a decimal point always indicate increased precision. Writing "0.010" in scientific notation clarifies that this number has two significant figures rather than three or four: $0.010 = 1.0 \times 10^{-2}$.

To determine how many significant figures there are in a particular numerical value, read the number from left to right and count all the digits, starting with the first digit that is not zero. To avoid ambiguities about trailing zeros, we place a decimal point after a value when its trailing zeros are significant. For example, "110" has only two significant figures (110 ± 10), whereas "110." has three significant figures (110 ± 1). Here are some examples of significant figures:

Number	Significant Figures	Number	Significant Figures
500	1	505	3
0.05	1	5.00×10^3	3
55	2	5.000	4
50.	2	505.0	4

SECTION EXERCISES

- **1.2.1** Convert the following measurements to scientific notation and express in base SI units: 0.000 463 L, 17 935 km, and 260 000 h (precise to three significant figures).
- **1.2.2** One light-year is the distance light travels in exactly one year. The speed of light is 1.08×10^9 km/h. Express the speed of light and the length of one light-year in SI units.
- **1.2.3** Convert each of these measurements to SI units: 155 pounds (mass of a typical person, 1 kg = 2.2 pounds), 110.0 yards (full length of a Canadian football field, 1 yard = 91.44 cm), 39 °C (body temperature of someone with a slight fever), and 365.2422 days (length of one year).

NEED MORE PRACTICE? Try Problems 1.9–1.14

1.3 CHEMICAL PROBLEM SOLVING

Many students have difficulty solving chemistry problems, but problem solving is a skill that you can learn and master. To help you, we devote extra attention to setting up and solving problems. Throughout this book, the Examples include information about methods for finding solutions.

Each Example includes a strategy that describes how to approach the problem. After describing the solution, we offer a Practice Exercise designed to help you master the particular type of problem while the example is fresh in your mind. The answers to these Practice Exercises appear at the end of each chapter, along with the answers to the Section Exercises.

Chemical problems that call for calculations often involve several tasks. It can be hard to recognize where to start and how to negotiate these tasks. A stepwise procedure can help you master these types of calculations. We recommend the sequence of steps shown in the Solving Quantitative Problems box. When you see a problem that appears complex, apply the stepwise procedure.

The first step in any problem is to identify what you are asked to find. Next, try to visualize what is going on. Think about the chemistry and draw pictures to help you see what is taking place. Once you have a good idea about the problem, determine which equations apply to the calculation and organize the data that are provided. We identify the most important equations by giving them numbers and screening them in yellow, and we recommend that you memorize these equations. Once you have organized the equations and data, you can perform appropriate manipulations and substitutions to carry out the calculation correctly. After you have finished the problem, check to see that your answer makes sense. Checking whether your answer is reasonable will keep you from making careless errors and train you to reject answers that make no physical sense.

Example 1-1 provides a first illustration of this stepwise procedure.

Solving Quantitative Problems

Step 1 **Determine what is asked for.**
What type of problem is it? What are you asked to find?

Step 2 **Visualize the problem.**
Draw pictures that illustrate what takes place.
If chemical changes occur, draw molecular pictures.

Step 3 **Organize the data.**
What data are available?
How are the data related to what is asked for?

Step 4 **Identify a process to solve the problem.**
What concepts are required? What equations apply?

Step 5 **Manipulate the equations.**
If necessary, do calculations in steps. Solve for what is asked for.

Step 6 **Substitute and calculate.**
Keep track of units.
Use the correct number of significant figures.

Step 7 **Does the result make sense?**
Are the units consistent? Is the result sensible?

EXAMPLE 1-1 PROBLEM-SOLVING STRATEGY



Courtesy of International Silver Plating

Spoons are electroplated with silver to give them an attractive finish.

Electroplating is a process by which a metal such as copper is coated with another metal such as silver or chromium. The transfer of metal atoms is driven by an electrical current. In an electroplating process, a spoon is coated with silver from a silver rod. In the process, 1.0×10^{21} atoms are transferred from the rod to the spoon, and the rod loses 0.179 g of its mass. Use this information and the density of silver (10.50 g/cm^3) to estimate the volume occupied by one silver atom.

Strategy: Apply the seven-step approach. The first two steps are part of the strategy.

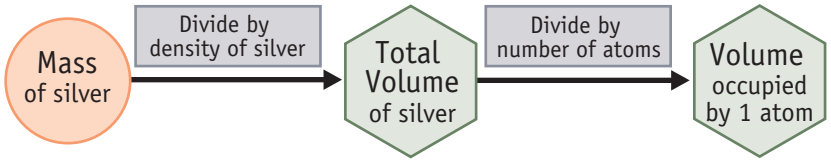
1. Determine what is asked for. The problem asks for the volume of a silver atom.
2. Visualize the problem. Draw a picture that shows the setup and the information provided in the problem. A picture often helps to interpret and summarize a problem. In this case, the figure should show the spoon, the silver wire, and some indication that atoms are transferred from the wire to the spoon.

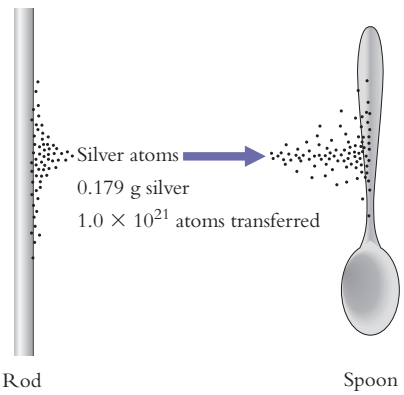
Solution:

3. Organize the data. First, what data are given in the problem?

$$\text{Mass of silver} = 0.179 \text{ g}$$
$$\text{Number of atoms transferred} = 1.0 \times 10^{21} \text{ atoms}$$
$$\text{Density of silver metal} = 10.50 \text{ g/cm}^3$$

4. Identify a process to solve the problem. The question asks about the volume of one silver atom. Mass and volume are related through **density**: $\rho = \mu\omega$ From this equation, we can calculate the total volume of the silver atoms. The problem also gives the total number of silver atoms transferred from the wire to the spoon. The volume of a single atom is the total volume divided by the number of atoms. Oftentimes, a flowchart helps to summarize the process:





(*continued*)

5. Manipulate the equations as needed. The flowchart shows two division steps.

   To begin the calculations, we use the mass and density to calculate the total volume:

$$\rho = \frac{m}{V} \quad \text{so} \quad V = \frac{m}{\rho}$$

   This is V_{total} which we divide by the number of atoms to find the volume of a single atom:

$$V_{atom} = \frac{V_{total}}{\text{Number of atoms}}$$

6. Substitute and calculate.

$$V_{total} = \frac{m}{\rho} = \frac{0.179\ \text{g}}{10.50\ \text{g/cm}^3} = 1.70 \times 10^{-2}\ \text{cm}^3$$

$$V_{atom} = \frac{V_{total}}{\text{Number of atoms}} = \frac{1.70 \times 10^{-2}\ \text{cm}^3}{1.0 \times 10^{21}\ \text{atoms}} = 1.7 \times 10^{-23}\ \text{cm}^3/\text{atom}$$

7. ✓ **Does the Result Make Sense?** Is this a "reasonable" value for the volume of an atom? At this stage of your study of chemistry, perhaps you cannot easily answer this question. Atoms are unimaginably small, however, and a volume of 10^{-23} cm^3 certainly is unimaginably small. Also, the units are correct: cm^3 is a measure of volume, and the calculation yields units of volume per atom, as the question asked.

Common Pitfalls: For this, and for any problem, the most common pitfall is not using a structured approach to solving it. We strongly suggest you follow the seven-step procedure as often as possible when solving problems.

Practice Exercise 1.1

A cylindrical jar has an inside diameter of 3.00 cm and a height of 8.00 cm. The empty jar weighs 185.65 g. Filled with gasoline, the same jar weighs 225.40 g. Find the density of gasoline.

NEED MORE PRACTICE? Try Problems 1.15 and 1.16

SECTION EXERCISES

- **1.3.1** The standard unit of length in the land of Ferdovia is the frud (2.000 frud = 10^{-2} m). Fast-food restaurants in Ferdovia sell fruit drinks in rectangular cartons measuring 50.00 by 40.00 by 80.0 fruds. The Frod family left a freshly opened carton on their kitchen table, and their pet ferret (Fred) drank some of the fruit drink. Mrs. Frod found the height of the remaining liquid to be 61 fruds. How many cubic metres of fruit drink did Fred drink?
- **1.3.2** A technician used a section of glass tubing to measure the density of a liquid that the laboratory needed to identify. The inside diameter of the tubing was known to be 0.87 mm. An empty piece of the tubing weighed 0.785 g. When liquid was drawn into the tubing to a height of 4.0 cm, the tubing and liquid weighed 0.816 g. Find the density of the liquid.
- **1.3.3** Canada imports an average of 1.229×10^6 barrels of oil per day. One barrel contains 158.99 litres. If the average density of oil is 900. g/L, calculate the mass (in tonnes) of oil imported by Canada in one year. 1 tonne = 10^6 g.

NEED MORE PRACTICE? Try Problems 1.17–1.20

1.4 COUNTING ATOMS: THE MOLE

Chemistry is a quantitative science, and chemists frequently measure amounts of matter, which means measuring the numbers of atoms. Counting atoms is difficult, but we can easily measure the mass of a sample of matter. To convert a mass measurement into a statement about the number of atoms in a sample, we must know the mass of an individual atom.

There is an enormous difference between the mass of one atom and the masses of samples measured in the laboratory. For example, a good laboratory balance measures mass values from about 10^{-6} to 10^3 g. An atom, on the other hand, has a mass between 10^{-24} and 10^{-21} g. Even a tiny mass of 10^{-6} g contains approximately 10^{16} atoms.

The Mole and Avogadro's Number

KEY CONCEPT

Chemical calculations are built around the mole.

To avoid having to work with unimaginably large numbers, chemists use a convenient unit called the **mole (mol):** *one mole is the number of atoms in exactly 12 g of the pure isotope carbon-12.*

KEY CONCEPT

One mole is the number of atoms in exactly 12 g of the pure isotope carbon-12.

Using mass spectrometers, scientists have determined that the mass of a ^{12}C atom is $1.992\,646 \times 10^{-23}$ g. This experimental mass combined with the definition of the mole gives the number of atoms in 1 mole:

$$\left(\frac{12\text{ g}}{1\text{ mol}}\right)\left(\frac{1\text{ atom}}{1.992\,646 \times 10^{-23}\text{ g}}\right) = 6.022\,142 \times 10^{23}\text{ atoms/mol}$$

The number of items in 1 mole is important enough to be given its own name and symbol. It is known as **Avogadro's number**, commonly symbolized N_A. Although Avogadro's number is known with seven-figure accuracy, four figures are enough for most calculations: $N_A = 6.022 \times 10^{23}$ items/mol.

The mole is a convenient unit for chemical amounts because the mass of most typical laboratory samples contains between 0.01 and 10 moles of atoms or molecules. For example, 1.20 g of ^{12}C contains 6.02×10^{22} atoms of carbon, which is 0.100 mole. The convenience of using a larger unit to describe many small items is not limited to atoms and molecules. Eggs are sold by the dozen, and paper is sold by the ream; chemists "count" atoms and molecules by the mole. There are 12 eggs in a dozen eggs, 500 sheets of paper in a ream of paper, and 6.022×10^{23} atoms in a mole of copper (Figure 1-9).

How large is a mole? If a computer were to count items at one million per second, it would take about 19 billion years to reach Avogadro's number. If every person on the planet (population = 7 billion) were to spend 1 million dollars per second, it would take about 19 years to spend 1 mole of dollars. One mole of Canadian one-dollar coins would cover the surface of the Earth to a depth of over 1,000 km!



Courtesy of Andy Washnik

Figure 1-9

Different units are useful for measuring different-sized items. Eggs are sold by the dozen, paper is sold by the ream, and atoms are measured by the mole. Twenty-four pre-1982 Canadian pennies contain about 1 mole of copper atoms.

Molar Mass

The **molar mass (*M*)** of any substance is the mass of 1 mole of that substance. Each isotope of a particular element has a different mass. Therefore, the mass of 1 mole of any **isotope** has a unique value, its *isotopic molar mass.* This characteristic molar mass can be found by multiplying the mass of one atom of that isotope by Avogadro's number. For example, mass spectrometry experiments reveal that one atom of carbon-13 has a mass of 2.15928×10^{-23}, from which we can calculate the isotopic molar mass of ^{13}C atoms

$$(2.15\,928 \times 10^{-23}\text{ g})\left(\frac{6.022\,142 \times 10^{23}}{1\text{ mol}}\right) = 13.0035\text{ g/mol}$$

A sample of a chemical element is usually a mixture of all of its stable isotopes. The carbon in a sample of graphite is 99.892% ^{12}C and 1.108% ^{13}C; elemental titanium is a mixture of five stable isotopes, and tin has 10 isotopes, each with its own particular mass. Every element has a characteristic distribution of isotopes, and each isotope of the element contributes proportionally to the molar mass of the element. The mass contributed by each isotope is proportional to its fractional abundance, as given by Equation 1-1:

$$\text{Elemental molar mass} = \sum(\text{Fractional abundance})(\text{Isotopic molar mass}) \qquad (1\text{-}1)$$

The symbol Σ is the Greek letter sigma. It means "take the sum of the following quantities."

Equation 1-1 combines multiplication and addition steps. The table below shows the fractional abundance and molar mass of each of the most abundant titanium isotopes.

KEY CONCEPT

The molar mass of any naturally occurring element is the sum of the contributions from its isotopes.

Isotope	Isotopic Molar Mass (g/mol)	Fractional Abundance
^{46}Ti	45.952 63	0.0825
^{47}Ti	46.951 76	0.0744
^{48}Ti	47.947 95	0.7322
^{49}Ti	48.947 87	0.0541
^{50}Ti	49.944 79	0.0518

Using Equation 1-1, the elemental molar mass of titanium is

$$M_{\text{Ti}} = (0.0825)(45.952\,63) + (0.0744)(46.951\,76) + (0.7322)(47.947\,95) + (0.0541)(48.947\,87) + (0.0518)(49.944\,79) = 47.63 \text{ g/mol}$$

The standard molar mass of titanium is slightly higher than this value because of the existence of several other more massive, but less abundant, isotopes.

For purposes of chemical bookkeeping, it is unnecessary to know the isotopic molar masses and isotopic distributions of the elements. All we need to know is the mass of 1 mole of an element containing its natural composition of isotopes. These molar masses usually are included in the periodic table, and they appear on the inside front and back covers of this textbook.

Mass–Mole–Atom Conversions

Molar mass can be thought of as a conversion factor between the mass in grams and the number of moles. These conversions are essential in chemistry, because chemists count amounts of substances in moles but routinely measure masses in grams. To determine the amount (number of moles) in a sample, we measure the mass of the sample and then divide by the molar mass of the substance:

$$n = \frac{m}{M} \tag{1-2}$$

If we know the moles and wish to convert to mass, we rearrange Equation 1-2 and multiply that number of moles by the molar mass: $m = nM$.

Suppose, for example, that a bracelet contains 168 g of silver (Ag). To determine the number of moles of silver in the bracelet, divide the mass by the molar mass of Ag, which is 107.87 g/mol:

$$n = \frac{m}{M} \quad \text{so} \quad \frac{168 \text{ g Ag}}{107.87 \text{ g/mol}} = 1.56 \text{ mol Ag}$$

If, for example, a chemical reaction calls for 0.250 mole of silver, we multiply moles by molar mass to determine the mass of Ag that should be used:

$$m = nM \quad \text{so} \quad m = (0.250 \text{ mol Ag})(107.87 \text{ g/mol}) = 27.0 \text{ g Ag}$$

Avogadro's number is the conversion factor that links the number of moles with the number of individual particles. To determine the number of atoms in a sample of an element, we multiply the number of moles by Avogadro's number: $\# = nN_{\text{A}}$. Likewise, if a sample contains a certain number of atoms, the number of moles in the sample can be calculated by dividing by Avogadro's number:

$$n = \frac{\#}{N_{\text{A}}} \tag{1-3}$$

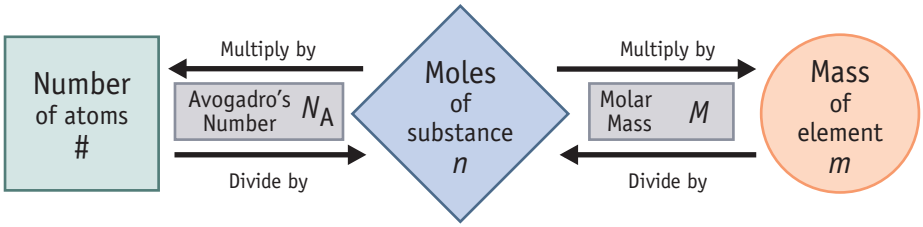


Figure 1-10

This flowchart shows the conversion processes among the number of moles, the mass, and the number of atoms for a sample of an element.

The number of silver atoms in the bracelet that contains 1.56 moles of Ag is found by multiplying the amount in moles by Avogadro's number:

$$\# = nN_A \quad \text{so}$$

$$\# = (1.56 \text{ mol Ag})(6.022 \times 10^{23} \text{ atoms/mol}) = 9.39 \times 10^{23} \text{ atoms Ag}$$

Radioactivity detectors, which "count" the number of nuclei that decay, provide an example of a conversion from number of particles to moles. If a radioactivity measurement on a blood sample from a patient receiving radiation treatment indicates that 1.75×10^4 nuclei of ^{99M}Tc decay in one hour, dividing by Avogadro's number converts to moles:

$$n = \frac{\#}{N_A} \quad \text{so} \quad n = \frac{1.75 \times 10^4 \text{ nuclei } ^{99m}\text{Tc}}{6.022 \times 10^{23} \text{ nuclei/mol}} = 2.91 \times 10^{-20} \text{ mol } ^{99m}\text{Tc}$$

Avogadro's number and molar mass make it possible to convert readily among the mass of a pure element, the number of moles, and the number of atoms in the sample. These conversions are represented schematically in the flowchart shown in Figure 1-10.

Moles occupy the central position of this flowchart because the mole is the unit that chemists use in almost all chemical calculations. When you set out to solve a chemical problem, first interpret the question on the atomic/molecular level. The second part of chemical problem solving often involves quantitative calculations, which usually require working with moles.

SECTION EXERCISES

1.4.1 Use the data in the following table to calculate the molar mass of naturally occurring sulphur.

Isotope	Isotopic Molar Mass (g/mol)	Abundance (%)
^{32}S	31.972 07	95.02
^{33}S	32.971 46	0.75
^{34}S	33.967 86	4.21
^{36}S	35.967 09	0.02

1.4.2 Calculate the number of atoms present in 2.55 g of each of these elements: (a) lithium; (b) titanium; and (c) gold.

1.4.3 Elemental radon (M = 222 g/mol) is a radioactive gas found at low concentrations in some homes in Canada. The Canadian Nuclear Safety Commission specifies that radon concentrations higher than 2.7×10^{-17} g/L are unhealthy. At this concentration, how many moles of radon would there be in a living room of volume 120 m^3? How many atoms is this?

NEED MORE PRACTICE? Try Problems 1.21–1.26

1.5 AMOUNTS OF COMPOUNDS

KEY CONCEPT

A molecule is a combination of two or more atoms held together in a specific shape by attractive forces.

As described in the section above, chemists use the mole to describe amounts of substances. The procedures introduced in that section work equally well for compounds as for elements. To apply these procedures, we need to be able to determine the molar masses of chemical compounds. We begin by detailing a procedure for quickly drawing organic molecules. Knowing how to do this saves time when discussing larger molecules.

Line Structures

LOOKING FORWARD In Chapters 10 and 11 on organic chemistry, we make extensive use of line structures.

The chemistry of living processes is complex, and many carbon-based molecules found in living organisms have complex structures. Because of this complexity, chemists have developed **line structures**, which are compact representations of the **structural formulas** of carbon compounds. Line structures are constructed according to the following guidelines:

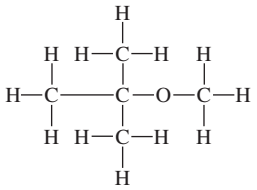
Guidelines for Line Structures

1. All bonds except C—H bonds are shown as lines.
2. C—H bonds are not shown in the line structure.
3. Single bonds are shown as single lines; double bonds are shown as two lines; triple bonds are shown as three lines.
4. Carbon atoms are not labelled.
5. All atoms except carbon and hydrogen are labelled with their elemental symbols.
6. Hydrogen atoms are labelled when they are attached to any atom other than carbon.

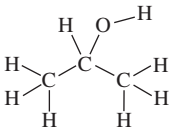
Example 1-2 illustrates these guidelines for three relatively simple molecules.

EXAMPLE 1-2 DRAWING LINE STRUCTURES

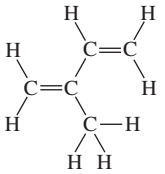
Construct line structures for the compounds with the following structural formulas:



$C_5H_{12}O$
Methyl tert-butyl ether
(a key antiknock ingredient in gasoline)



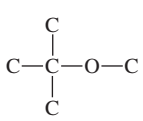
C_3H_8
2-Propanol
(rubbing alcohol)

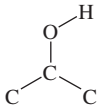


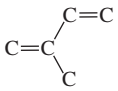
C_5H_8
Isoprene
(the building block of natural rubber)

Strategy: A line structure is built by using information about how atoms are connected in the molecule. The structural formula is the basis for the line structure. Apply the guidelines to convert the structural formula representation into a line-structure representation.

Solution: The first two guidelines state that bonds are represented by lines and that all C—H bonds are ignored. Thus, we remove the C—H bonds from the structural formula, leaving a bond framework. According to the third guideline, the double bonds remain as two lines:

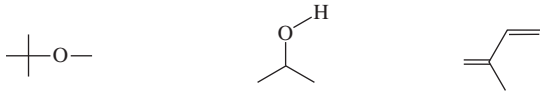






According to the fourth guideline, carbon atoms are not labelled, so we remove the C's:

These are the line structures for the three molecules:

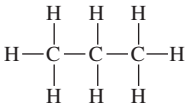


✓ Does the Result Make Sense? To check that the line structures represent the correct substances, count the number of intersections and line ends, which should equal the number of C atoms in the compound. The first line structure has five, the second has three, and the third has five, matching the chemical formulas.

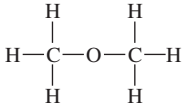
Common Pitfalls: When learning line structures, it is common to miss the fact that there is a carbon atom at the end of each line (that does not have an atom explicitly shown) and at the intersection of two or more lines.

Practice Exercise 1.2

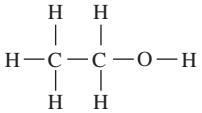
Draw the line structures for the following compounds:



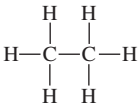
Propane
C_3H_8



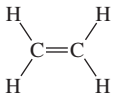
Dimethyl ether
$(CH_3)_2O$



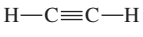
Ethanol
C_2H_5OH



Ethane
C_2H_6



Ethylene
C_2H_4

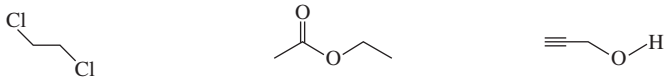


Acetylene
C_2H_2

Although line structures are convenient, a chemist must be able to convert them back into chemical formulas. The reconstruction of a complete formula from a line structure relies on the most important general feature of carbon chemistry: in all neutral molecules containing carbon, each carbon atom has *four* chemical bonds. Keep in mind that carbon atoms are not shown in a line structure, so the first step in constructing a structural formula from a line structure is to place a C at every line intersection and at the end of every line. Then add singly bonded hydrogen atoms (—H) until every carbon atom has four bonds. Example 1-3 illustrates this procedure.

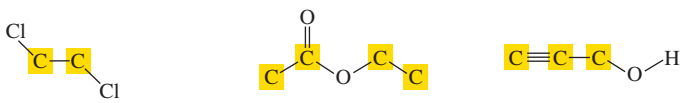
EXAMPLE 1-3 CONVERTING LINE STRUCTURES

Construct the structural formulas and determine the chemical formulas from the following line drawings:



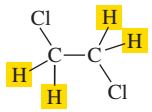
Strategy: Line drawings show all structural features except carbon atoms and C—H bonds. Convert a line drawing into a structural formula in two steps. First, place a C at any unlabelled line end and at each line intersection. Second, add C—H bonds until each carbon atom has four bonds. The chemical formula is then obtained by counting the atoms of each element.

Solution: Place a C at each intersection and line end:

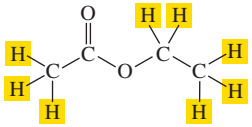


(*continued*)

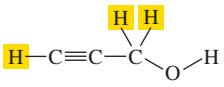
Now add C—H bonds until each carbon atom has a total of four bonds. Each carbon atom in the first structure has two bonds, so each carbon needs two C—H bonds to complete the structural formula:



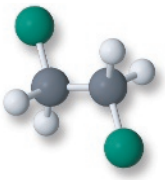
The second structure has four carbon atoms. The two end carbons have just one bond, so each needs three C—H bonds. The carbon with the double bond to oxygen already has its complete set of four bonds. The other inner carbon atom has a bond to C and a bond to O, so it needs two C—H bonds:

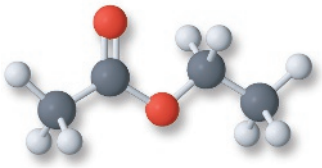


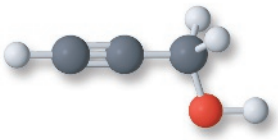
The third structure contains a triple bond. The end carbon needs one C—H bond, but the other carbon atom of the triple bond already has four bonds. The next carbon has two bonds, one to its carbon neighbour and one to the oxygen. Two C—H bonds are needed to give this atom its usual set of four bonds:



These are the correct structural formulas. Now it is a simple matter to count the number of atoms of each element and write the chemical formulas of the compounds: $C_2H_4Cl_2$, $C_4H_8O_2$, and C_3H_4O.



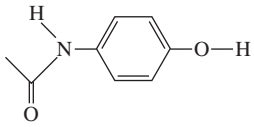




✓ Does the Result Make Sense? Check the consistencies of the structural formulas by counting the number of bonds associated with each carbon atom. If you have converted the line structure correctly, each carbon atom will have four bonds.

Common Pitfalls: It is easy at first to forget that each carbon makes a total of four bonds.

Practice Exercise 1.3
Convert the following line structure into structural and chemical formulas:



The structural formulas of the compounds in these examples are not very complicated, yet their line structures are clearer than other representations. In ever-larger carbon-containing compounds, the simplification provided by line drawings is a great help to chemists. Line structures are not cluttered by the C—H bonds in the molecule, making them easier both to draw and to interpret.

KEY CONCEPT

The molar mass of a compound is found by adding together the molar masses of all of its elements, taking into account the number of moles of each element present.

Molar Masses of Chemical Compounds

Just as each element has a characteristic molar mass, so does every chemical compound. Chemical compounds are composed of atoms bound together into molecules, or **ions** clustered together in electrically neutral aggregates. In either case, a chemical formula describes the atomic composition of a compound.

One mole of any chemical compound is 1 mole of its chemical formula unit. Here are some examples:

- One mole of diatomic oxygen is 1 mole of O_2 molecules. Each molecule contains two oxygen atoms, so 1 mole of O_2 molecules contains 2 moles of oxygen atoms.
- One mole of methane (CH_4) is 1 mole of CH_4 molecules, so it contains 1 mole of carbon atoms and 4 moles of hydrogen atoms.
- One mole of sodium chloride (NaCl) contains 1 mole each of Na^+ cations and Cl^- anions.
- One mole of magnesium chloride ($MgCl_2$) contains 1 mole of Mg^{2+} and 2 moles of Cl^-.

When atoms combine to form molecules, the atoms retain their atomic identities and characteristic molar masses. Thus, we can add elemental molar masses to obtain the molar mass of any compound.

One mole of O_2 molecules contains 2 moles of O atoms, so the molar mass of molecular O_2 is calculated as follows:

$$\left(\frac{2 \text{ mol O}}{1 \text{ mol } O_2}\right)\left(\frac{15.999 \text{ g}}{1 \text{ mol O}}\right) = 31.998 \text{ g/mol } O_2$$

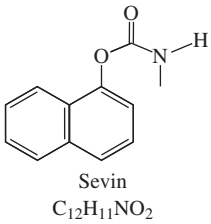
$$M_{O_2} = 31.998 \text{ g/mol}$$

You can verify that the molar masses of CH_4, $NaNO_3$, and $MgCl_2$ are 16.043 g/mol, 84.994 g/mol, and 95.211 g/mol, respectively.

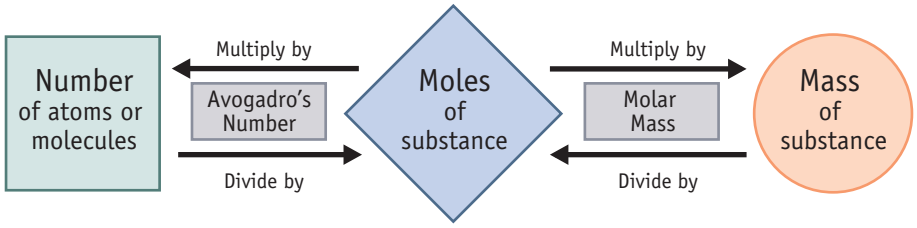
As chemical formulas become more complicated, the calculation of molar mass requires more steps, as shown in Examples 1-4 and 1-5.

EXAMPLE 1-4 MOLAR MASS OF A MOLECULE

The line drawing and chemical formula of Sevin, a common insecticide, appear in the margin. Determine the molar mass of Sevin.



Sevin
$C_{12}H_{11}NO_2$



Strategy: To determine the molar mass of a substance, we need its chemical formula and elemental molar masses. From the chemical formula, determine the number of moles of each element contained in 1 mole of the substance. Multiply each elemental molar mass by the number of moles of that element and add the results for each element.

Solution: The line structure demonstrates how the atoms are connected in the molecule, but we do not need this information because the chemical formula is provided. We use the chemical formula to calculate the mass of each element in 1 mole of the compound:

$$\left(\frac{12 \text{ mol C}}{1 \text{ mol } C_{12}H_{11}NO_2}\right)\left(\frac{12.011 \text{ g C}}{\text{mol C}}\right) = \left(\frac{144.13 \text{ g C}}{1 \text{ mol } C_{12}H_{11}NO_2}\right)$$

$$\left(\frac{11 \text{ mol H}}{1 \text{ mol } C_{12}H_{11}NO_2}\right)\left(\frac{1.0079 \text{ g H}}{\text{mol H}}\right) = \left(\frac{11.087 \text{ g H}}{1 \text{ mol } C_{12}H_{11}NO_2}\right)$$

(*continued*)

$$\left(\frac{1\text{ mol N}}{1\text{ mol } C_{12}H_{11}NO_2}\right)\left(\frac{14.007\text{ g C}}{\text{mol N}}\right) = \left(\frac{14.007\text{ g N}}{1\text{ mol } C_{12}H_{11}NO_2}\right)$$

$$\left(\frac{2\text{ mol O}}{1\text{ mol } C_{12}H_{11}NO_2}\right)\left(\frac{15.999\text{ g O}}{\text{mol O}}\right) = \left(\frac{31.998\text{ g O}}{1\text{ mol } C_{12}H_{11}NO_2}\right)$$

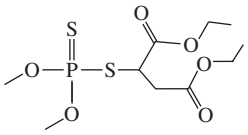
Now add the individual components to find the molar mass of the compound:

$$M = \left(\frac{144.13\text{ g C}}{1\text{ mol } C_{12}H_{11}NO_2}\right) + \left(\frac{11.087\text{ g H}}{1\text{ mol } C_{12}H_{11}NO_2}\right) + \left(\frac{14.007\text{ g N}}{1\text{ mol } C_{12}H_{11}NO_2}\right) + \left(\frac{31.998\text{ g O}}{1\text{ mol } C_{12}H_{11}NO_2}\right)$$
$$= 201.22\text{ g/mol}$$

✓ **Does the Result Make Sense?** We can do a quick check to see that the magnitude of the molar mass is about right by limiting the use of significant figures. Sevin contains 12 C and 11 H atoms, which contribute (12 × 12) + (11 × 1) = 155 g/mol to its molar mass. Add the contributions of two O atoms and one N atom: (2 × 16) + (1 × 14) = 46, giving 201 g/mol.

Common Pitfall: Not multiplying the atomic molar masses by the numbers of atoms in the formula.

Practice Exercise 1.4
Determine the molar mass of Malathion, an insecticide whose line structure and chemical formula appear below.



Malathion
$C_{10}H_{19}O_6PS_2$

NEED MORE PRACTICE? Try Problems 1.27–1.30

EXAMPLE 1-5 MOLAR MASS OF AN IONIC COMPOUND

Metal salts often exist as **hydrates**. One example is iron(II) nitrate hexahydrate. Determine the molar mass of this compound.

Strategy: Before we can calculate a molar mass, we need a chemical formula. Then we can calculate the masses of each of the elements in 1 mole of the compound.

Solution: Iron(II) has a charge of +2, and the nitrate anion carries a –1 charge, so there must be two nitrates for every one iron ion, and *hexa-* stands for 6, so the chemical formula is $Fe(NO_3)_2 \cdot 6H_2O$. It is convenient to deconstruct the formula into a list of elements, removing the parentheses and including the atoms in the waters of hydration.

One mole of $Fe(NO_3)_2 \cdot 6H_2O$ contains 1 mol Fe, 12 mol H, 2 mol N, and 12 mol O:

$$\begin{aligned}
(1\text{ mol Fe})\,(55.845\text{ g/mol}) &= 55.845\text{ g Fe}\\
(12\text{ mol H})\,(1.0079\text{ g/mol}) &= 12.095\text{ g H}\\
(2\text{ mol N})\,(14.007\text{ g/mol}) &= 28.014\text{ g N}\\
(12\text{ mol O})\,(15.999\text{ g/mol}) &= 191.99\text{ g O}\\
\hline
M_{Fe(NO_3)_2\,\cdot\,6H_2O} &= 287.94\text{ g/mol}
\end{aligned}$$

✓ **Does the Result Make Sense?** Another way to calculate this molar mass is by working with the individual components of the formula. Each formula unit of the compound contains one iron(II) cation, two nitrate anions, and six waters of hydration: $M_{Fe(NO_3)_2\cdot\,6H_2O} = M_{Fe^{2+}} + 2\,M_{NO_3^-} + 6\,M_{H_2O}$. Verify that this procedure gives the same result.

Common Pitfalls: It is easy to miss the parentheses around a group of atoms, but it is important not to do so.

Practice Exercise 1.5
Determine the molar mass of $Al(NO_3)_3 \cdot 9H_2O$.

Mass–Mole–Number Conversions for Compounds

As emphasized in Section 1.4, many of the calculations in chemistry involve converting back and forth among the mass of a substance, the number of moles, and the number of atoms and/or molecules. These calculations are all centred on the mole. The connections shown in Figure 1-10 apply to chemical compounds as well as to atoms of pure elements. Molar mass and Avogadro's number provide links between the mass of a sample, the number of moles, and the number of molecules, according to Equations 1-2 and 1-3.

For example, the following calculation shows how many moles and molecules of Sevin are in a 5.0 g sample (the molar mass of the insecticide Sevin was calculated in Example 1-4):

$$n = \frac{m}{M} = \frac{5.0\text{ g}}{201.22\text{ g/mol}} = 0.0248\text{ mol Sevin}$$

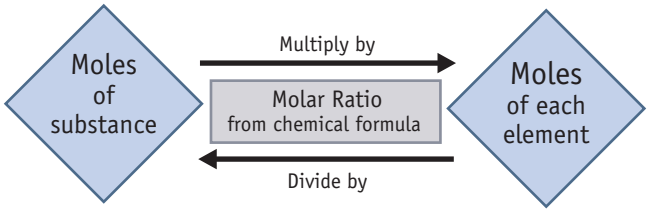
$$\# = nN_A = (0.0248\text{ mol Sevin})(6.022\times10^{23}\text{ molecules/mol}) = 1.5\times10^{22}\text{ molecules of Sevin}$$

Notice that, once again, moles are at the centre of the scheme. To repeat, *calculations of chemical amounts centre on the mole.*

KEY CONCEPT

Molar masses are used to convert between moles of substance and mass.

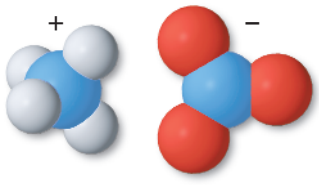
Mass–mole–number calculations often involve atoms *within* a compound as well as the compound itself. The chemical formula provides the link between the moles of a compound and the number of moles of the compound's individual elements:



Example 1-6 illustrates mass–mole–number conversions involving both elements and compounds.

EXAMPLE 1-6 ELEMENTAL CONTENT

Ammonium nitrate (NH_4NO_3) is used as a fertilizer because it is a good source of nitrogen atoms. It is such a good source, in fact, that it ranks among the top industrial chemicals produced yearly in the world. How many moles of nitrogen atoms are present in a 454 g bag of NH_4NO_3 fertilizer? How many atoms is this?



NH_4^+ Ammonium NO_3^- Nitrate

Strategy: First think about the chemistry of the problem, and then construct a mathematical solution. Use the seven-step procedure.

1. The problem asks for the number of moles and atoms of nitrogen in a 454 g sample of NH_4NO_3.
2. Visualize NH_4NO_3. (Always think atoms and molecules.) Ammonium and nitrate are common **polyatomic** ions whose chemical formulas you should remember.

Solution:

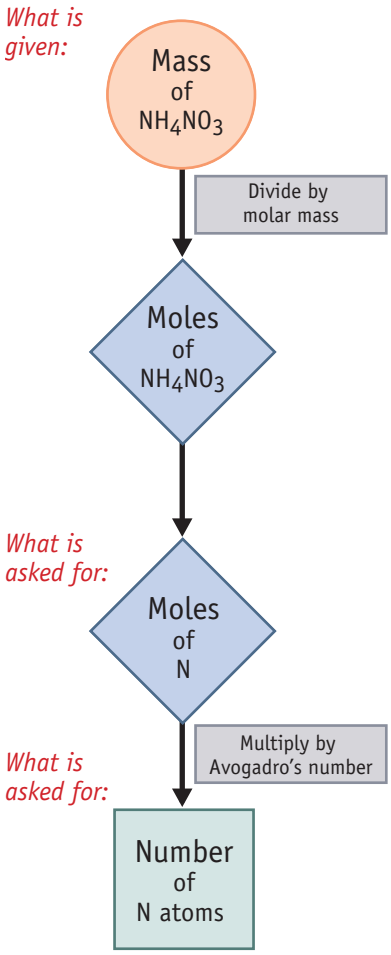
3. The problem gives the mass of the sample:

$$\text{Mass of sample} = 454\text{ g } NH_4NO_3$$

4. Equations 1-2 and 1-3 link mass, moles, and atoms:

$$n = \frac{m}{M} \quad \text{and} \quad n = \frac{\#}{N_A}$$

(*continued*)



5. Often, it helps to draw a flowchart that organizes the steps necessary to analyze and solve a problem. A flowchart for this problem appears in the margin.
6. Now we are ready to work through the calculations. Determine the molar mass of NH_4NO_3:

$$\begin{aligned} M &= (2 \text{ mol N})(14.007 \text{ g/mol}) + (4 \text{ mol H})(1.0079 \text{ g/mol}) \\ &\quad + (3 \text{ mol O})(15.999 \text{ g/mol}) \\ &= 80.043 \text{ g/mol} \end{aligned}$$

Divide the sample mass by the molar mass to determine moles of NH_4NO_3:

$$n = \frac{m}{M} = \left(\frac{454 \text{ g } NH_4NO_3}{80.043 \text{ g/mol}}\right) = 5.672 \text{ mol } NH_4NO_3$$

The chemical formula reveals that every mole of NH_4NO_3 contains 2 moles of nitrogen atoms:

$$(5.672 \text{ mol } NH_4NO_3)\left(\frac{2 \text{ mol N atoms}}{1 \text{ mol } NH_4NO_3}\right) = 11.34 \text{ mol N atoms}$$

As a final answer, this value should be rounded to three significant figures to match the data (454 g): 11.3 mol N atoms. For the next step of the overall calculation, however, the value should retain a fourth digit to avoid possible rounding errors. Finish the problem by rearranging Equation 1-3, and use Avogadro's number to convert from the number of moles to the number of atoms:

$$\begin{aligned} \# &= nN_A = (11.34 \text{ mol N})(6.022 \times 10^{23} \text{ atoms/mol}) \\ &= 6.83 \times 10^{24} \text{ N atoms} \end{aligned}$$

The final result is rounded to three significant figures to match the three significant figures given for the mass of the bag of fertilizer.

7. ✓ **Does the Result Make Sense?** Although 10^{24} is an immense number of atoms, we know that 1 mol contains 6×10^{23} atoms. Numbers of the order of 10^{24} are reasonable when calculating the number of atoms in a sample of everyday size.

Common Pitfalls: It is tempting to simply use the number of significant digits shown on your calculator, but make sure you round off the answer appropriately.

Practice Exercise 1.6
How many moles and atoms of hydrogen are present in 125 g of water (H_2O)?

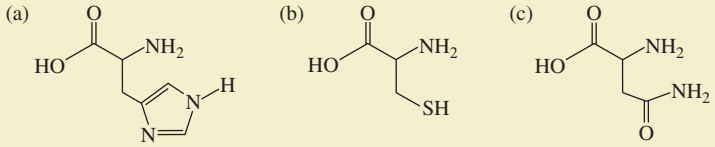
NEED MORE PRACTICE? Try Problems 1.31–1.36

SECTION EXERCISES

1.5.1 Calculate the molar masses of the following compounds:

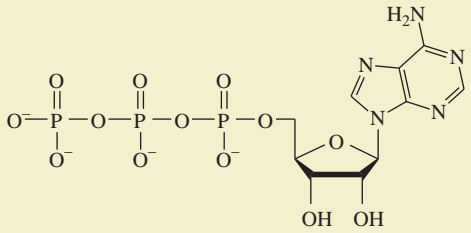
Compound	Name	Uses
(a) $NaAsO_2$	Sodium arsenite	Insecticide, herbicide, fungicide
(b) Na_4SiO_4	Sodium silicate	Soaps, adhesives, flame retardant
(c) $Ca(ClO)_2$	Calcium hypochlorite	Bactericide, bleach, pool cleaner
(d) $Ba(NO_3)_2$	Barium nitrate	Green fireworks
(e) Tl_2SO_4	Thallium sulphate	Rat and ant poisons

1.5.2 Twenty different amino acids are the essential building blocks of proteins. Calculate the molar masses of these three:



Histidine $C_6H_9N_3O_2$ | Cysteine $C_3H_7NO_2S$ | Asparagine $C_4H_8N_2O_3$

1.5.3 Adenosine triphosphate (ATP, shown below) is the body's principal energy storage molecule. The formula of ATP is $C_{10}H_{16}N_5O_{13}P_3$. (a) How many moles of ATP are in a 50.0 mg sample? (b) How many molecules of ATP is this? (c) How many atoms of phosphorus are in the sample? (d) How many atoms of nitrogen are there? (e) What is the mass of oxygen in 50.0 mg of ATP? (f) How many total atoms are there in 50.0 mg of ATP?



1.6 AQUEOUS SOLUTIONS

We developed the concept of the mole in terms of pure chemical substances, but many chemical reactions take place in solution. To treat solution reactions quantitatively, we need ways to apply the mole concept to solutions. A substance used to dissolve solutes is a **solvent**, and a pure substance dissolved in solution is a **solute**. Most of the time, the solvent is a liquid and is present in much larger quantities than any solutes.

Chemists use many different liquid solvents, but we focus most of our attention on water. When water is the solvent, the solution is said to be **aqueous**. A rich array of chemistry occurs in aqueous solution, including many geological and biochemical processes. Aqueous solutions dominate the chemistry of the Earth and the biosphere. The oceans, for example, are aqueous solutions of various cations and anions, with sodium and chloride being the most abundant. The oceans can be thought of as huge aqueous solvent vessels for the remarkably complex chemistry of our world. Blood is an aqueous system that contains many ionic species, most notably carbonate, sodium, and potassium. In fact, the human body is mostly water, and much of the biochemistry of life takes place in aqueous solution.

Molarity

Any solution contains at least two chemical species: the solvent and one or more solutes. The mass of a solution is the sum of the masses of the solvent and all dissolved solutes. To answer questions such as "How much is there?" about solutions, we need to know the amount of each solute present in a specified volume of solution. The amount of a solute in a solution is given by the **concentration (*c*)**, which is the ratio of the amount of solute to the amount of solution. In chemistry the most common measure of concentration is **molarity**, which is the number of *moles of solute (n)* divided by the total *volume of the solution (V)* in litres. In this case,

KEY CONCEPT

Concentration is the amount of substance per unit volume of solution.

$$c = \frac{\text{Moles of solute}}{\text{Total volume of solution}} \quad \text{or} \quad c = \frac{n}{V} \tag{1-4}$$

Equation 1-4 defines concentration in terms of molarity, but note that many other units of concentration are used in chemistry. (Several of these are detailed in Chapter 9.) If a solution has a concentration of, for example, 2.00 mol/L, we would write $c = 2.00$ M. **Be careful not to confuse this unit of concentration (M = mol/L) with the symbol for molar mass (*M* = g/mol).** Confusion also arises because many chemists will refer to the "molarity of the solution" instead of the "concentration of the solution, expressed as molarity."

LOOKING FORWARD Other concentration units are introduced in Chapter 9.

An aqueous solution is prepared by dissolving a measured quantity of solid or liquid in enough water to give a desired final volume. For example, we might dissolve 2.05 g of table sugar (sucrose) in just enough water to make the final volume 25.0 mL. To determine the concentration of the sugar solution, we must convert the mass of sucrose into moles and convert the solution volume into litres. To begin, use the formula of sucrose, which is $C_{12}H_{22}O_{11}$, to calculate the molar mass, 342 g/mol. Next, convert mass to moles in the usual manner:

$$n = \frac{m}{M} = \frac{2.05 \text{ g sugar}}{342 \text{ g/mol}} = 5.994 \times 10^{-3} \text{ mol sugar}$$

Then, convert the volume of the solution from millilitres to litres:

$$(25.0 \text{ mL})\,(10^{-3} \text{ L/mL}) = 2.50 \times 10^{-2} \text{ L}$$

Finally, divide moles of sugar by volume in litres to obtain concentration in mol/L:

$$c = \frac{n}{V} = \frac{5.994 \times 10^{-3} \text{ mol}}{2.50 \times 10^{-2} \text{ L}} = 0.240 \text{ M}$$

Concentrations are used so frequently in chemistry that shorthand notations are essential. For example, the concentration of a species is expressed by enclosing its formula in square brackets: $[C_{12}H_{22}O_{11}] = 0.240$ M for the sugar solution example above.

EXAMPLE 1-7 PREPARING A SOLUTION

What mass of nickel(II) chloride hexahydrate is required to prepare 250 mL of aqueous solution with a concentration of 0.255 M?

Strategy: As with all calculations of chemical amounts, we must work with moles. Because grams are asked for, we must do a mole–mass conversion; this requires the molar mass of the substance, which in turn requires that we know the chemical formula.

Solution: Begin by finding the chemical formula from the name, nickel(II) chloride hexahydrate. Chloride ion carries a –1 charge, and the (II) indicates that nickel is a +2 cation. Electrical neutrality requires two chloride ions for every nickel ion. The name also tells us that the salt contains six water molecules for each unit of nickel(II) chloride. Thus, each formula unit of the salt contains one Ni^{2+} cation, two Cl^- anions, and six water molecules: $NiCl_2 \cdot 6H_2O$.

To determine the mass of $NiCl_2 \cdot 6H_2O$ required to prepare the solution, first calculate the number of moles of the salt required, and then use the molar mass to determine the mass in grams:

$$n = cV = \left(\frac{0.255 \text{ mol}}{\text{L}}\right)\left(\frac{10^{-3} \text{ L}}{1 \text{ L}}\right)(250. \text{ mL}) = 6.375 \times 10^{-2} \text{ mol } NiCl_2 \cdot 6H_2O$$

We need the molar mass of $NiCl_2 \cdot 6H_2O$ to convert moles to mass. The molar mass of $NiCl_2 \cdot 6H_2O$ is 237.69 g/mol:

$$m = nM = (6.375 \times 10^{-2} \text{ mol})\,(237.69 \text{ g/mol}) = 15.15 \text{ g}$$

To prepare the desired solution, we would first weigh 15.2 g of solid using a balance. (Rounding to three significant figures is consistent with the target molarity stated in the problem.) We would transfer this solid into a 250 mL volumetric flask, add enough water to dissolve the solid, and then continue adding water and mixing until the solution level matched the 250 mL mark on the flask. Note that molarity is moles per litre of *solution*, and making this solution will require slightly less than 250 mL of water.

✓ **Does the Result Make Sense?** The mass is substantially less than the molar mass of the compound. Given that the volume is considerably less than 1 L and the target concentration is considerably less than 1 M, this is a reasonable result.

Practice Exercise 1.7
Calculate the mass of ammonium nitrate needed to prepare 100 mL of aqueous solution with a concentration of 0.400 M.

Species in Solution

To understand the chemical behaviour of solutions, we must "think molecules." Before working any problem involving aqueous solutions, begin with the question, "What chemical species are present in the solution?" There will always be an abundance of water molecules in an aqueous solution. In addition, there will be solute species, which may be molecules or ions.

An aqueous solution of a molecular substance such as sugar ($C_{12}H_{22}O_{11}$) or ethanol (C_2H_5OH) contains individual molecules in a sea of water molecules (Figure 1-11). We know from measurements of electrical conductivity that these solutes dissolve as neutral molecules. Figure 1-12 shows that pure water does not conduct electricity, and neither does a solution of sugar in water. This result shows that these solutions contain no mobile charged particles. Sugar and ethanol dissolve as neutral molecules.

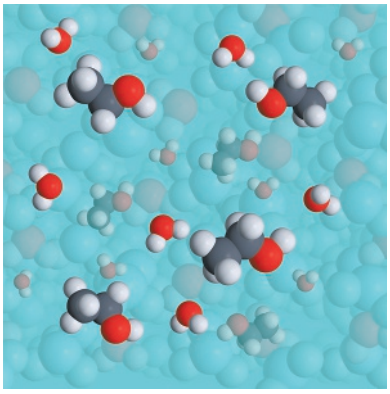


Figure 1-11
Molecular picture of an aqueous solution of ethanol, which contains ethanol molecules and water molecules.

In contrast to sugar, solid NaCl dissolves in water to give a liquid that conducts electricity. Figure 1-12 shows that a solution of NaCl is a good conductor. When an ionic compound dissolves in water, its component cations and anions are free to move about in the solution. Mixing leads to a uniform distribution of Na^+ and Cl^- ions throughout the entire solution, with each ion surrounded by a sheath of water molecules, as shown in Figure 1-13.

The presence of ions in solution is what gives a NaCl solution the ability to conduct electricity. If positively and negatively charged wires are dipped into the solution, the ions in the solution respond to the charges on the wires. Chloride anions move toward the positive wire, and sodium cations move toward the negative wire. This directed movement of ions in solution is a flow of electrical current. Pure water, which has virtually no dissolved ions, does not conduct electricity. Any solution formed by dissolving an ionic solid in water conducts electricity. Ordinary tap water, for example, contains ionic impurities that make it an electrical conductor.



Ken Karp

Figure 1-12
Pure water (*left*) and a solution of sugar (*right*) do not conduct electricity because they contain virtually no ions. A solution of salt (*centre*) conducts electricity well because it contains mobile cations and anions.

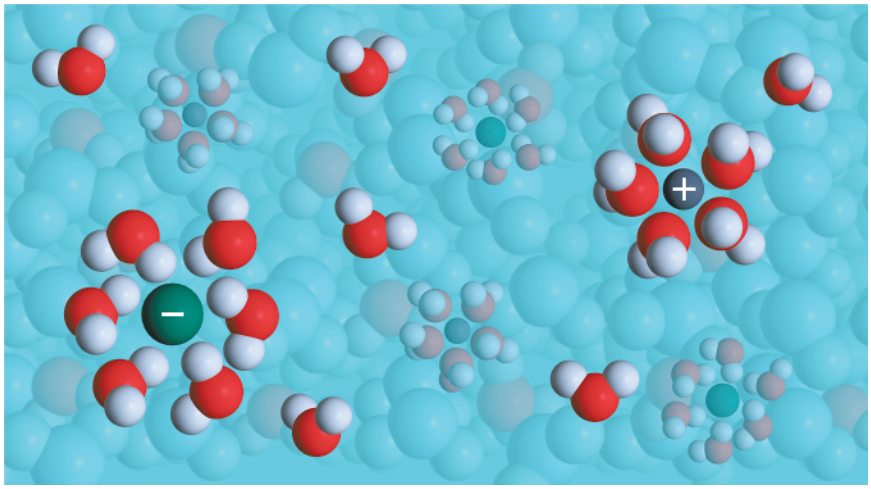


Figure 1-13
Molecular picture of a solution of sodium chloride in water. All the molecules and ions move freely about, but overall electrical neutrality is maintained because the total amount of anionic charge equals the total amount of cationic charge.

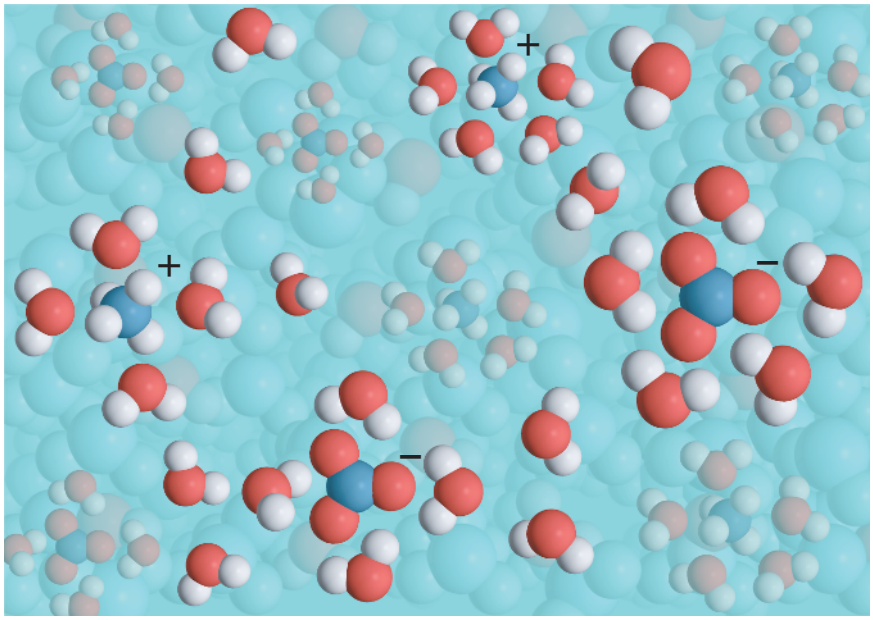


Figure 1-14

Molecular view of an aqueous solution of ammonium nitrate. Ammonium ions separate from nitrate ions, but each of these species remains intact as polyatomic clusters of atoms.

Notice that the molecular view in Figure 1-13 includes ions as well as neutral molecules. "Molecular view" means a view of how matter looks at the atomic/molecular level. Thus, such illustrations may contain ions and individual atoms as well as molecules.

KEY CONCEPT

Polyatomic ions remain intact when a salt dissolves in water.

When a salt containing polyatomic ions dissolves in water, the cations separate from the anions, but each polyatomic ion remains intact. An example is ammonium nitrate, composed of NH_4^+ polyatomic cations and NH_3^- polyatomic anions. (Appendix G contains a list of common polyatomic ions.) Ammonium nitrate dissolves in water to give a solution containing NH_4^+ cations and NH_3^- anions, as Figure 1-14 illustrates.

Because aqueous solutions containing ions play a central role in the world around us, the chemistry of such solutions is discussed in depth in several chapters of this book. To understand the chemistry of aqueous solutions containing ions, *it is essential that you learn to recognize the common ions at a glance.* This is especially true of the polyatomic ions.

Concentrations of Ionic Solutions

The ratio of cations and anions in a solution of a salt is determined by its chemical formula. For example, a solution of NaCl contains Na^+ cations and Cl^- anions in 1:1 ratio: a 1.0 M solution of NaCl has 1.0 M each of Na^+ and Cl^-. In contrast, $NiCl_2$ contains 2 moles of Cl^- anions for every 1 mole of Na^{2+} cations. This ratio is maintained when nickel(II) chloride dissolves in water; a 1.0 M solution of $NiCl_2$ has 1.0 M Ni^{2+} and 2.0 M Cl^-:

$$\underset{1\text{ mol}\Uparrow}{NiCl_2(s)} \xrightarrow{H_2O} \underset{1\text{ mol}\Uparrow}{Ni^{2+}(aq)} + \underset{2\text{ mol}\Uparrow}{Cl^-(aq) + Cl^-(aq)}$$

The (*s*) and (*aq*) in the equation indicate phases: (*s*) designates solid and (*aq*) designates aqueous solution.

Notice that preserving the cation–anion ratio ensures that the solution is electrically neutral overall. The total positive charge carried by 1 mole of Ni^{2+} equals the total negative charge carried by 2 moles of Cl^-.

Example 1-8 treats concentrations of polyatomic ions.

EXAMPLE 1-8 MOLARITY OF IONS IN SOLUTION

Find the concentrations (in mol/L) of the ionic species present in 250 mL of an aqueous solution containing 1.75 g of ammonium sulphate, $(NH_4)_2SO_4$.

Strategy: The chemical formula identifies the ions that are present in the final solution. The formula also tells us how many moles of each ion are present in 1 mole of the salt. Use mass, molar mass, and volume to calculate molarity.

Solution: Looking at the formula, we recognize the ammonium cation$^+$, and the sulphate anion$^-$, When ammonium sulphate dissolves in water, it dissociates into its component polyatomic ions. Each mole of salt will produce 2 mol NH_4^+ and 1 mol SO_4^{2-}.

After identifying what species are present in solution, we can calculate molarities. Begin with the molar mass of the salt:

$$\begin{aligned}(2\text{ mol N})(14.01\text{ g/mol}) &= 28.02\text{ g N}\\ (8\text{ mol H})(1.008\text{ g/mol}) &= 8.064\text{ g H}\\ (1\text{ mol S})(32.06\text{ g/mol}) &= 32.06\text{ g S}\\ (4\text{ mol O})(16.00\text{ g/mol}) &= 64.00\text{ g O}\\ \hline 1\text{ mol }(NH_4)_2SO_4 &= 132.14\text{ g}\end{aligned}$$

Use the molar mass to determine the number of moles of salt and, in turn, the number of moles of each ion. Notice that even though we are aiming for molarities of ionic species, we determine the moles of the salt. This is because the mass that is provided is the *mass of the salt.* To find the moles of salt (n), this mass must be divided by the *molar mass of the salt*:

$$n = \frac{m}{M} = \frac{1.75\text{ g}}{132.14\text{ g/mol}} = 1.324 \times 10^{-2}\text{ mol }(NH_4)_2SO_4$$

From the number of moles of salt, calculate the number of moles of each ion:

$$(1.324 \times 10^{-2}\text{ mol }(NH_4)_2SO_4)\left(\frac{2\text{ mol }NH_4^+}{1\text{ mol }(NH_4)_2SO_4}\right) = 2.648 \times 10^{-2}\text{ mol }NH_4^+$$

$$(1.324 \times 10^{-2}\text{ mol }(NH_4)_2SO_4)\left(\frac{1\text{ mol }SO_4^{2-}}{1\text{ mol }(NH_4)_2SO_4}\right) = 1.324 \times 10^{-2}\text{ mol }SO_4^{2-}$$

Now find the molarities of the individual ions:

$$(250.\text{ mL})(10^{-3}\text{ L/mL}) = 0.250\text{ L}$$

$$[NH_4^+] = \frac{2.648 \times 10^{-2}\text{ mol}}{0.250\text{ L}} = 1.06 \times 10^{-1}\text{ M}$$

$$[SO_4^{2-}] = \frac{1.324 \times 10^{-2}\text{ mol}}{0.250\text{ L}} = 5.30 \times 10^{-2}\text{ M}$$

✓ Does the Result Make Sense? The concentration of ammonium ions is twice that of sulphate ions. This is consistent with the 2:1 ratio of ammonium to sulphate in the chemical formula. You should understand that when we describe this solution as "0.0530 M $(NH_4)_2SO_4$" stating moles of salt per litre of solution, we are describing a solution that actually contains a *mixture of ions, each with its own molarity.*

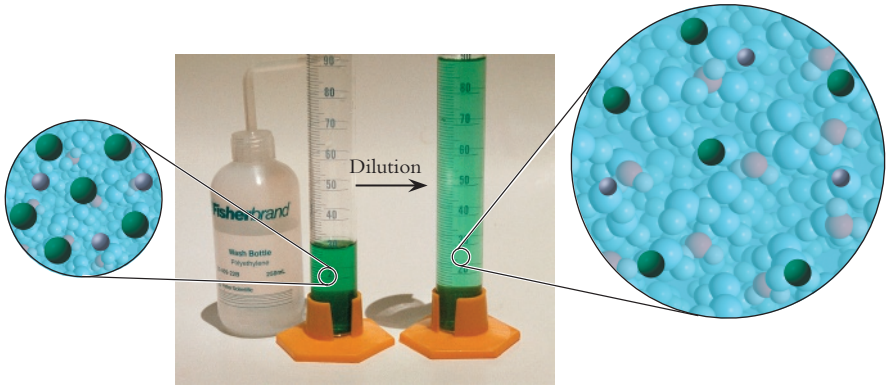
Common Pitfalls: Do not forget that salts dissociate in solution. When we say that a 1.0 M solution of a salt is prepared, it means we dissolve 1.0 moles of salt in 1 L of water, but there will be at least 2.0 mol or more of ions in this solution.

Remember to carry an extra significant figure until the final calculation.

Use dimensional analysis to obtain correct conversion factors. Rearrange 1 mL = 10^{-3} L to give 1 = (10^{-3} L/mL) with mL in the denominator so that it cancels mL in the numerator.

Practice Exercise 1.8
Calculate the molarities of the ions present in the solution of nickel(II) chloride hexahydrate prepared in Example 1-7.

NEED MORE PRACTICE? Try Problems 1.37 and 1.38



Courtesy of John Olmsted

Figure 1-15

When a concentrated solution of $NiCl_2$ (*left*) is diluted by adding more solvent, the resulting solution (*right*) contains the same numbers of Ni^{2+} and Cl^- ions. The solution is more dilute, however, because the ions are spread around in a larger volume of solvent.

Dilutions

One of the common methods for preparing a solution of known concentration is to dilute an existing solution of higher concentration by adding more solvent. In a **dilution**, the *amount of solute* remains the same, but the *volume of the solution* increases. Thus, a dilution results in a solution of lower molarity.

The most important feature of a dilution is that *the number of moles of solute does not change during the dilution.* Consider the dilution of an aqueous solution of nickel(II) chloride, as shown in Figure 1-15. The photograph shows that the green colour, which is due to the Ni^{2+} ions, fades when the solution is diluted. This fading is visible evidence that the concentration has decreased. At the molecular level, there are the same *numbers* of Ni^{2+} and Cl^- ions in the solution before the dilution and after the dilution. However, because the volume of the solution increases, the concentration of the new solution is lower than the concentration of the original solution.

The quantitative aspects of dilutions can be found from the fact that the number of moles of solute does not change during a dilution from the initial (i) solution to the final (f) solution:

$$n_i = n_f$$

KEY CONCEPT

Diluting a solution decreases the concentrations of the solutes.

Because we can arrange Equation 1-4 to give $n = cV$, this leads to a simple equation:

$$c_i V_i = c_{ff} V \qquad (1\text{-}5)$$

Equation 1-5 is convenient for dilution calculations. If any three of the quantities are known, we can calculate the fourth, as Example 1-9 shows.

EXAMPLE 1-9 DILUTION OF A SOLUTION

Aqueous hydrochloric acid, HCl, is usually sold as a 12.0 M solution, commonly referred to as *concentrated HCl.* A chemist needs to prepare 2.00 L of 1.50 M HCl for a number of different applications. What volume of concentrated HCl solution should the chemist use in the dilution?

Strategy: The seven-step problem-solving approach is appropriate.

1. The question asks for the volume of concentrated HCl required to prepare a dilute solution.
2. Visualization: A concentrated solution is diluted to a larger volume. The chemist will remove a sample from the bottle of concentrated HCl solution and mix it with water.

Solution:

3. The information available is as follows:

$$\text{Concentrated HCl} = 12.0\ \text{M} \quad \text{Diluted HCl} = 1.50\ \text{M}$$
$$\text{Final volume} = 2.00\ \text{L}$$

4. This is a dilution process, so we use Equation 1-5.
5. Rearrange Equation 1-5 to solve for V_i, the volume required: $V_i = \dfrac{c_f V_f}{c_i}$.
6. Do the calculation. In Step 3, we identified the appropriate values:

$$c_f = 1.50\ \text{M}, \quad V_f = 2.00\ \text{L}, \quad c_i = 12.0\ \text{M}$$

$$V_i = \frac{(1.50\ \text{M})(2.00\ \text{L})}{12.0\ \text{M}} = 0.250\ \text{L}$$



© Tony Freeman/PhotoEdit

7. ✓ **Does the Result Make Sense?** The volume has the correct units, and the amount required is smaller than the total volume. This is reasonable.

Common Pitfalls: Make sure that if the solution is diluted, its concentration decreases. If your calculation says otherwise, you have incorrectly rearranged Equation 1-5.

Practice Exercise 1.9

What volume (in mL) of concentrated ammonia (14.8 M) should be used to prepare 0.500 L of a 0.350 M solution?

NEED MORE PRACTICE? Try Problems 1.41 and 1.42

When preparing acid solutions, water should never be added directly to concentrated acid. Addition of water often causes rapid local heating that may lead to a dangerous acid splash. Instead, concentrated acid is always added to the water. Additional water can then be added safely to the less concentrated solution that has resulted. The solution in Example 1-9 is prepared by using a volume-measuring device, such as a graduated cylinder, to measure 0.250 L (250 mL) of concentrated HCl. About 0.5 L of water is added to a 2.00 L **volumetric flask**, and then the concentrated acid from the graduated cylinder is poured slowly into the water. Filling the volumetric flask to the mark with additional water completes the dilution.

Chemists often need to work with solutions with a variety of concentrations. Often, a relatively concentrated "stock" solution is prepared by weighing. Then, a portion of this solution is diluted volumetrically. In procedures of this kind, it is important to keep track of the volumes, as illustrated in Example 1-10 and shown in the flowchart in Figure 1-16.

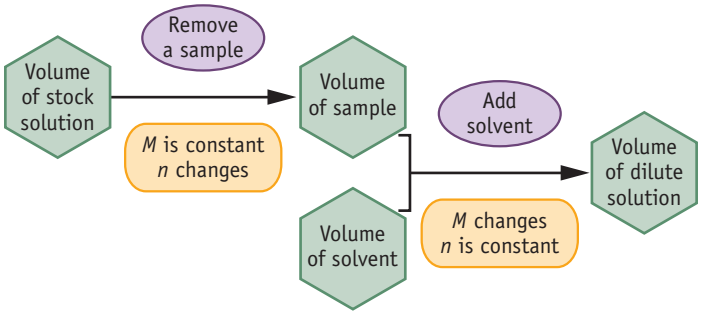


Figure 1-16

Flowchart showing the connections among measured quantities in dilution procedures.

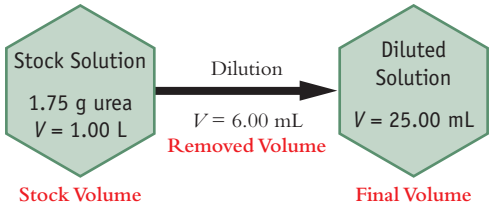
EXAMPLE 1-10 PREPARING SOLUTIONS

An agricultural chemist wished to study the effect of varying fertilizer applications on the growth of tomato plants. The chemist prepared a stock aqueous solution of urea, $(NH_2)_2CO$, by dissolving 1.75 g of this compound in water to make 1.00 L of solution. Then, she prepared a series of more dilute solutions to apply to her tomato plants. One of these solutions contained 5.00 mL of stock solution diluted to give a final volume of 25.00 mL. What was the concentration of urea in this diluted solution?

Strategy: Use the seven-step method.

1. We are asked to determine the concentration of the dilute solution. This requires us first to determine the concentration of the stock solution, so we need to do two separate calculations.

2 and 3. A flowchart, shown below, helps us to visualize the problem and organizes the data. The flowchart shows that the calculation has two parts. First, we treat the stock solution. Then, we use the result to treat the diluted solution.



Solution:

4. We have mass and volume information for the stock solution, so we use Equations 1-2 and 1-4:

$$n = \frac{m}{M} \quad \text{and} \quad c = \frac{n}{V}$$

For the diluted solution, we use Equation 1-5:

$$c_i V_i = c_f V_f$$

5. Manipulate the equations to obtain equations for the molarity of the stock and diluted solutions:

$$c_{stock} = \frac{m}{(M)\,(V_{stock})} \quad \text{and} \quad c_{final} = \frac{(c_{stock})\,(V_{removed})}{V_{final}}$$

We need the molar mass of urea:

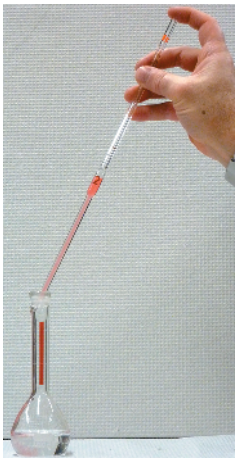
$$\begin{aligned} M &= (4 \text{ mol H})\,(1.0079 \text{ g/mol}) + (2 \text{ mol N})\,(14.007 \text{ g/mol}) \\ &\quad + (1 \text{ mol C})\,(12.011 \text{ g/mol}) + (1 \text{ mol O})\,(15.999 \text{ g/mol}) \\ &= 60.056 \text{ g/mol} \end{aligned}$$

6. Substitute and calculate:

$$c_{stock} = \frac{m}{(M)\,(V_{stock})} = \frac{1.75 \text{ g}}{(60.056 \text{ g/mol})\,(1.00 \text{ L})} = 2.914 \times 10^{-2} \text{ M}$$

$$c_{final} = \frac{(c_{stock})\,(V_{removed})}{V_{final}} = \frac{(2.914 \times 10^{-2} \text{ M})\,(5.00 \text{ mL})}{25.00 \text{ mL}} = 5.82 \times 10^{-3} \text{ M}$$

7. ✓ **Does the Result Make Sense?** To check for reasonableness, make sure you have used the correct volumes. The *stock solution* was prepared in a volume of 1.00 L. The *dilution* was from 5.00 mL to 25.00 mL, a factor of 5. We can see by inspection that the final concentration is lower than the stock concentration by a factor of about 5.



Courtesy of Jenna E. Davies

Practice Exercise 1.10

Calculate the concentration of ammonium nitrate in a solution prepared by dissolving 3.20 g of the salt in enough water to make 100 mL of solution, and then diluting 5.00 mL of this solution to a volume of 25.00 mL.

NEED MORE PRACTICE? Try Problems 1.39 and 1.40

SECTION EXERCISES

- **1.6.1** An urban gardener buys a 1.0 kg bag of ammonium nitrate to fertilize house plants. Calculate the concentrations of the ionic species present in a stock solution prepared by dissolving 7.5 g of the salt in 0.300 L of water.
- **1.6.2** Plants can be badly damaged by solutions that are highly concentrated. A friend of the urban gardener recommends against using fertilizer solutions that are greater than 1.0 mM (1.0×10^{-3} M) in total nitrogen. What volume (in mL) of the stock solution from Section Exercise 1.6.1 will be needed to prepare 750 mL of 1.0 mM fertilizer solution?
- **1.6.3** Draw a molecular picture of a portion of the solution described in Section Exercise 1.6.2. Make sure the solution is electrically neutral (omit water molecules to simplify your drawing).

NEED MORE PRACTICE? Try Problems 1.43 and 1.44

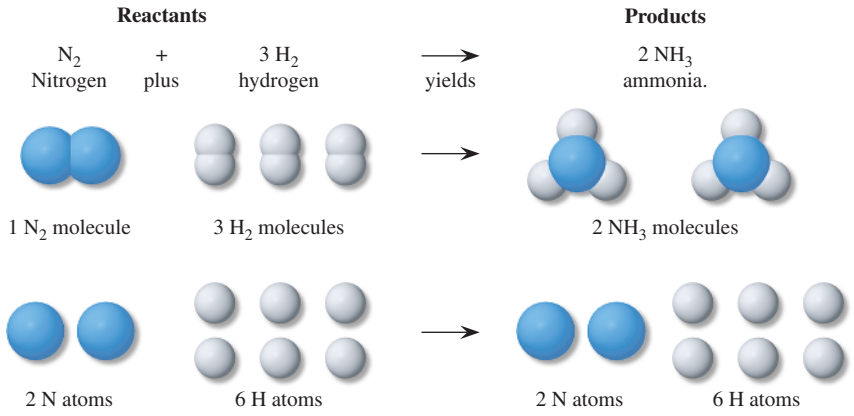
1.7 WRITING CHEMICAL EQUATIONS

Balanced Equations

In chemical reactions, *the amount of each element is always conserved.* This is consistent with the statements of Dalton's atomic theory. In addition, *the total amount of electrical charge is always conserved.* This is the law of conservation of charge. A **balanced chemical equation** describes a chemical reaction in which the amounts of all elements and of electrical charge are conserved. We refer to the species that take part in a chemical reaction as **reagents. Reactants** are reagents that are consumed in a reaction, whereas **products** are reagents that are produced in a reaction. In addition to providing a list of reactants and products of the reaction, a balanced chemical equation tells us the relative amounts of all the reagents.

As an example, consider the industrial synthesis of ammonia (NH_3). Ammonia is made by the Haber process, a single chemical reaction between molecules of hydrogen (H_2) and nitrogen (N_2). Although it is simple, this synthesis has immense industrial importance. Canada produces some 4 billion kilograms of ammonia annually.

In this synthesis, the reactants are N_2 and H_2, which react to produce NH_3. Here is the balanced chemical equation:



KEY CONCEPT

Molecules of reagents react in whole number ratios to form products.

All balanced chemical equations have the following features:

1. The reactants appear on the left and the products appear on the right. The arrow joining them indicates the direction of reaction.
2. An integer precedes the formula of each substance. These numbers are the **stoichiometric coefficients.** When no number appears (as for N_2 in this equation), the stoichiometric coefficient is 1.

3. The stoichiometric coefficients in a chemical equation are the smallest integers that give a balanced equation.
4. Charge is conserved. In this equation, all participants are neutral species, so charge is conserved regardless of the stoichiometric coefficients.

Stoichiometric coefficients describe the relative numbers of molecules involved in the reaction. In any actual reaction, immense numbers of molecules are involved, but the relative numbers are always related through the stoichiometric coefficients. Furthermore, these coefficients describe both the relative numbers of *molecules* and the relative numbers of *moles* involved in the reaction. For example, the Haber reaction always involves immense numbers of molecules, but the equation describing the synthesis of ammonia tells us the following:

1. Each molecule of nitrogen reacts with three molecules of hydrogen to give two molecules of ammonia.
2. Each mole of molecular nitrogen reacts with 3 moles of molecular hydrogen to give 2 moles of ammonia.

KEY CONCEPT

The number of atoms of each element is conserved in a chemical reaction.

Keep in mind that the key feature of balanced chemical equations is the **conservation law**:

The number of atoms of each element is conserved in any chemical reaction.

Balancing Equations

The stoichiometric coefficients in a balanced chemical equation must be chosen so that the atoms of each element are conserved. Many chemical equations can be balanced by inspection. Balancing by inspection means changing stoichiometric coefficients until the number of atoms of each element is the same on each side of the arrow. Usually, we can tell what changes need to be made by looking closely at the reaction and matching the numbers of atoms of each element on both sides of the equation. Consider the following example.

Propane, which is used as a fuel for gas barbecues, reacts with molecular oxygen to form carbon dioxide and water:

$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

To determine if this equation is balanced, make a list of the elements and numbers of atoms on each side:



$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

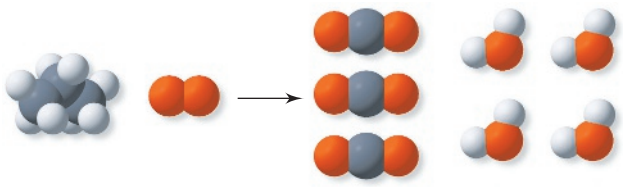
$$3\,C + 8\,H + 2\,O \longrightarrow 1\,C + 2\,H + 3\,O$$

The equation is not balanced, because there are too many carbon and hydrogen atoms on the left and too many oxygen atoms on the right. We need to change the numbers of molecules by changing the stoichiometric coefficients until the numbers of atoms of each element are equal.

It is easiest to balance a chemical equation one element at a time, starting with the elements that appear in only one substance on each side. Notice that all of the carbon atoms in propane end up in carbon dioxide molecules, and all of propane's hydrogen atoms appear in water molecules. This feature allows us to balance carbon and hydrogen easily.

To take care of the three carbon atoms per propane molecule, we need three molecules of CO_2. Thus, the carbon atoms are balanced by changing the stoichiometric coefficient of CO_2 from 1 to 3. In this reaction the ratio of CO_2 to propane is 3:1. Similarly, we need four

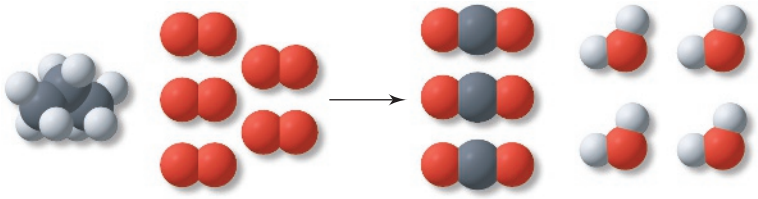
molecules of water for the eight hydrogen atoms in one molecule of propane. Using this information, we modify the equation as follows:



$$C_3H_8 + O_2 \longrightarrow 3\ CO_2 + 4\ H_2O$$

$$3\ C + 8\ H + 2\ O \longrightarrow 3\ C + 8\ H + 10\ O$$

The situation is looking better because atoms of carbon and hydrogen are now conserved. However, the equation is still not balanced because there are too few oxygen atoms on the left side. To balance oxygen, we must give O_2 a stoichiometric coefficient of 5:



$$C_3H_8 + 5\ O_2 \longrightarrow 3\ CO_2 + 4\ H_2O$$

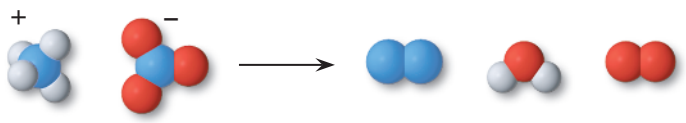
$$3\ C + 8\ H + 10\ O \longrightarrow 3\ C + 8\ H + 10\ O$$

Now the equation is balanced. Combustion of one molecule of propane gas produces three molecules of CO_2 and four molecules of H_2O and consumes five molecules of O_2. Also, combustion of 1 mole of propane produces 3 moles of CO_2 and 4 moles of H_2O and consumes 5 moles of O_2. Example 1-11 shows how to balance another reaction.

EXAMPLE 1-11 BALANCING CHEMICAL REACTIONS

Ammonium nitrate (NH_4NO_3) explodes when it is heated above 300 °C. The products are three gases: molecular nitrogen, molecular oxygen, and steam (water vapour). Write a balanced equation for the explosion of ammonium nitrate.

Strategy: The description in the problem tells us what happens to the starting material: NH_4NO_3 breaks apart into molecules of N_2, O_2, and H_2O. An unbalanced form of the equation can be written from this description. Then, we must balance each element in turn by inspection.



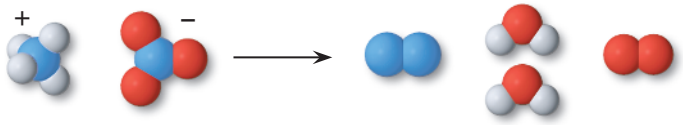
$$NH_4NO_3 \longrightarrow N_2 + H_2O + O_2$$

Solution: Count atoms of each element to see if the elements are in balance:

$$2\ N + 4\ H + 3\ O \longrightarrow 2\ N + 2\ H + 3\ O$$

(*continued*)

Focus first on the elements that appear in only one reactant and one product, nitrogen and hydrogen in this case. Nitrogen is already balanced. To balance the H atoms, change the stoichiometric coefficient of water from 1 to 2:



$$NH_4NO_3 \longrightarrow N_2 + 2\ H_2O + O_2$$

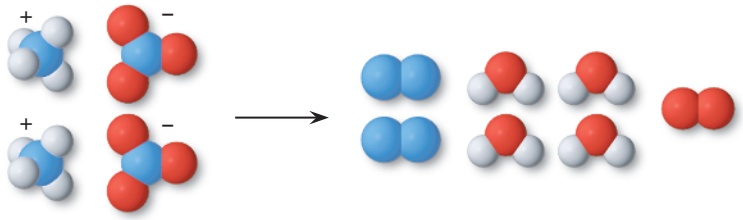
$$2\ N + 4\ H + 3\ O \longrightarrow 2\ N + 4\ H + 4\ O$$

Now, atoms of nitrogen and hydrogen are conserved, but the equation is not balanced with respect to oxygen. We cannot change the 1:1 ratio of NH_4NO_3 to N_2 and the 1:2 ratio of NH_4NO_3 to H_2O, because these ratios must be retained to keep the hydrogen and nitrogen atoms conserved. Thus, we must adjust the coefficient of O_2 to balance with respect to oxygen. Numerically, we can take care of oxygen by changing the coefficient of O_2 to ½:

$$NH_4NO_3 \longrightarrow N_2 + 2\ H_2O + \tfrac{1}{2}\ O_2$$

$$2\ N + 4\ H + 3\ O \longrightarrow 2\ N + 4\ H + 3\ O$$

This balances the equation, but it is unrealistic from a molecular perspective because there is no such thing as half a molecule. To get rid of the ½ without unbalancing the reaction, we multiply all the coefficients by 2:



$$2\ NH_4NO_3 \longrightarrow 2\ N_2 + 4\ H_2O + O_2$$

$$4\ N + 8\ H + 6\ O \longrightarrow 4\ N + 8\ H + 6\ O$$

✓ **Does the Result Make Sense?** Count the atoms of each element to be sure that the reaction is balanced and reasonable from a molecular perspective.

Common Pitfalls: Missing the subscripts or parentheses in a chemical formula will make it impossible to balance a reaction.

Practice Exercise 1.11
Some racing cars use methanol, CH_3OH, as their fuel. Balance the combustion reaction of methanol.

NEED MORE PRACTICE? Try Problems 1.45–1.54

Later, we show that it can be convenient to use fractional coefficients such as $^1/_2$. When this is the case, the coefficient refers to $^1/_2$ mol, not $^1/_2$ molecule.

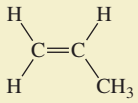
SECTION EXERCISES

- **1.7.1** Although gasoline is a complex mixture of molecules, the chemical reaction that takes place in an automobile engine can be represented by combustion of one of its components, octane (C_8H_{18}). Such burning of fossil fuels releases millions of tonnes of carbon dioxide into the Earth's atmosphere each year. Write a balanced equation for the combustion of octane.
- **1.7.2** Additional reactions accompany the combustion of octane. For example, molecular nitrogen reacts with molecular oxygen in an automobile cylinder to give nitrogen oxide. After leaving the engine, nitrogen oxide reacts with atmospheric oxygen to give nitrogen dioxide, the red-brown gas seen in the air over many urban areas. Write balanced equations for these two reactions.

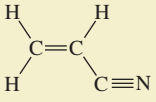
1.7.3 Acrylonitrile is used to make synthetic fibres such as Orlon. About 1.5 billion kg of acrylonitrile is produced each year. Balance the following chemical equation, which shows how acrylonitrile is made from propene, ammonia, and oxygen:

$$C_3H_6 + NH_3 + O_2 \longrightarrow C_3H_3N + H_2O$$

Propene



Acrylonitrile



1.8 THE STOICHIOMETRY OF CHEMICAL REACTIONS

Sections 1.4 and 1.5 describe how to use the relationships among atoms, moles, and masses to answer "how much?" questions about *individual substances.* Combining these ideas with the concept of a balanced chemical equation lets us answer "how much?" questions about *chemical reactions.* The study of the amounts of materials consumed and produced in chemical reactions is called **stoichiometry.**

Stoichiometry is pronounced "stoy-key-om-etry." It combines two Greek words, *stoicheion* ("element") and *metron* ("measure").

A chemical synthesis requires the proper amounts of starting materials for a successful outcome. Just as a cake recipe provides the amounts of ingredients needed for successful baking, a balanced chemical equation is a chemical recipe for successful synthesis. For example, how many grams of hydrogen do we need to produce 68 g of ammonia? To find out, we begin with the balanced chemical equation:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

According to the balanced equation, the synthesis of 2.0 moles of ammonia requires 3.0 moles of hydrogen and 1.0 moles of nitrogen. However, the balanced equation does not give us any direct information about the *masses* involved in the synthesis. One gram of N_2 plus 3 g of H_2 *does not* make 2 g of NH_3. Remember that calculations in chemistry are based on the concept of the mole.

In other words, a chemical recipe for making ammonia requires amounts in moles, not masses. Thus, to prepare 68 g of ammonia, we need to know the number of moles of ammonia in 68 g. To do that, we must divide the mass of ammonia by the molar mass of ammonia, 17.0 g/mol, to convert grams to moles:

$$n = \frac{m}{M} = \frac{68\text{ g NH}_3}{17.0\text{ g/mol}} = 4.0\text{ mol NH}_3$$

How many moles of H_2 are needed to make 4.0 moles of NH_3? According to the balanced equation, 2 moles of NH_3 requires 3 moles of H_2. The ratio of stoichiometric coefficients from a balanced equation is called the **stoichiometric ratio.** We now apply this ratio to the desired amount of NH_3.

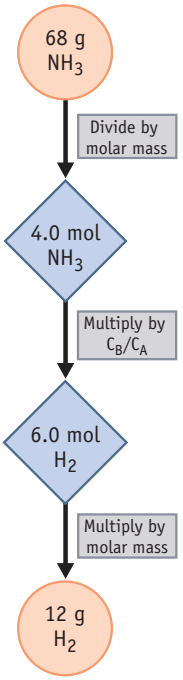
$$4.0\text{ mol NH}_3\left(\frac{3\text{ mol H}_2}{2\text{ mol NH}_3}\right) = 6.0\text{ mol H}_2$$

Note that the units "mol NH_3" cancel, leaving only "mol H_2." All that remains is to convert from moles of H_2 to mass: $m = nM = (6.0\text{ mol H}_2)(2.02\text{ g/mol}) = 12\text{ g H}_2$.

The same procedure can be used to show that the synthesis of 68 g of NH_3 requires 56 g of N_2.

To summarize, the amounts of different reagents that participate in a chemical reaction are related through the stoichiometric coefficients in the balanced chemical equation. To convert from moles of one reagent to moles of any other reagent, multiply by the stoichiometric ratio that leads to proper cancellation of units:

$$n_B = \left(\frac{\text{Coefficient}_B}{\text{Coefficient}_A}\right) n_A \qquad (1\text{-}6)$$



A flowchart illustrates the logic of the stoichiometric ratio.

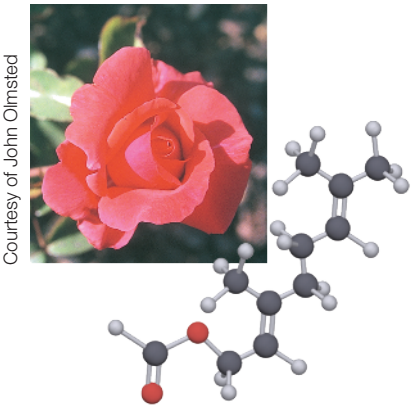
Examples 1-12 and 1-13 illustrate this procedure.

EXAMPLE 1-12 HOW MUCH PRODUCT CAN BE MADE?

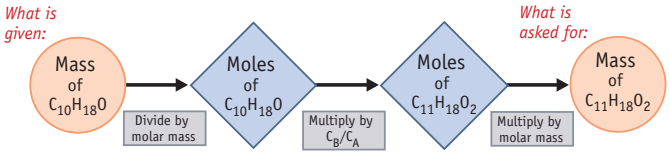
Geranyl formate is used as a synthetic rose essence in cosmetics. The compound is prepared from formic acid and geraniol:

$$HCO_2H + C_{10}H_{18}O \longrightarrow C_{11}H_{18}O_2 + H_2O$$

A perfumery needs some geranyl formate for a batch of perfume. How many grams of geranyl formate can a chemist make from 375 g of geraniol?


Courtesy of John Olmsted

Strategy: The stoichiometric coefficients state that 1 mole of each reactant will produce 1 mole of each product. A flowchart summarizes the steps used to convert the mass of geraniol into the mass of geranyl formate.



Solution: First, convert the mass of geraniol into moles by dividing by the molar mass:

$$n = \frac{m}{M} = \frac{375 \text{ g geraniol}}{154.2 \text{ g/mol}} = 2.432 \text{ mol geraniol}$$

A 1:1 stoichiometric ratio links moles of geraniol to moles of geranyl formate:

$$\text{Moles of geranyl formate} = \left(\frac{1 \text{ mol geranyl formate}}{1 \text{ mol geraniol}}\right)(2.432 \text{ mol geraniol})$$

Thus, 2.432 moles of geraniol can produce 2.432 moles of geranyl formate. To finish the problem, multiply by the molar mass of geranyl formate to determine the mass of geranyl formate produced in the reaction. The mole units are eliminated, leaving gram units for the answer:

$$(2.432 \text{ mol geranyl formate})\left(\frac{182.3 \text{ g geranyl formate}}{1 \text{ mol geranyl formate}}\right) = 443 \text{ g geranyl formate}$$

✓ **Does the Result Make Sense?** The result has three significant figures, just like the initial mass of geraniol. Geranyl formate has a slightly larger molar mass than geraniol, and we can see that the mass of the product is slightly more than the mass of the starting material.

Practice Exercise 1.12
What mass of formic acid is required to produce 443 g of geranyl formate?

NEED MORE PRACTICE? Try Problems 1.59 and 1.60

EXAMPLE 1-13 AMOUNTS OF REACTANTS AND PRODUCTS

Poisonous hydrogen cyanide (HCN) is an important industrial chemical. It is produced from methane (CH_4), ammonia, and molecular oxygen. The reaction also produces water. An industrial manufacturer wants to convert 175 kg of methane into HCN. How much molecular oxygen will be required for this synthesis?

Strategy:

1. This is a stoichiometry problem, in which we are asked to find the mass of a reactant.
2. To visualize this kind of problem, we need a balanced chemical equation, as follows:

Write the unbalanced form of the equation from the information given; namely methane, ammonia, and molecular oxygen react to form hydrogen cyanide and water:

$$CH_4 + NH_3 + O_2 \longrightarrow HCN + H_2O$$

Balance the equation by inspection, one element at a time. Carbon and nitrogen are in balance when there are equal numbers of molecules of CH_4, NH_3, and HCN. Hydrogen and oxygen, however, are not in balance. There are seven hydrogen atoms on the left but only three on the right. The coefficient for HCN cannot be changed without unbalancing C and N. Thus, to balance H, we must multiply the coefficient for H_2O by 3:

$$CH_4 + NH_3 + O_2 \longrightarrow HCN + 3\ H_2O$$

This gives three oxygen atoms on the right, which corresponds to $\frac{3}{2}$ O_2 molecules among the starting materials:

$$CH_4 + NH_3 + \tfrac{3}{2}\ O_2 \longrightarrow HCN + 3\ H_2O$$

To avoid the impossible $\frac{1}{2}$ molecule, multiply all of the stoichiometric coefficients by 2. This gives the balanced equation:

$$2\ CH_4 + 2\ NH_3 + 3\ O_2 \longrightarrow 2\ HCN + 6\ H_2O$$

Solution:

3. The problem gives only one piece of data: 175 kg of CH_4 will be converted into HCN. The problem asks about another starting material, so we will need two molar masses:

$$M_{CH_4} = 16.04\ \text{g/mol},\ M_{O_2} = 32.00\ \text{g/mol}$$

4. Knowing the mass of one reagent, we are asked to find the mass of another. We need to determine molar amounts and then convert to kilograms. The calculations require two equations for interconversion of moles and mass:

$$n = \frac{m}{M} \quad \text{and} \quad n_B = \left(\frac{\text{Coefficient}_B}{\text{Coefficient}_A}\right) n_A$$

5. After converting mass to moles and n_A to n_B, we need to convert moles back to mass: $m = nM$
6. Begin by converting the mass of methane to moles:

$$(175\ \text{kg}\ CH_4)\left(\frac{10^3\ \text{g}}{1\ \text{kg}\ CH^4}\right) = 1.75 \times 10^5\ \text{g}\ CH_4$$

$$\frac{1.75 \times 10^5\ \text{g}\ CH_4}{16.04\ \text{g/mol}} = 1.091 \times 10^4\ \text{mol}\ CH_4$$

Next, use the ratio of coefficients to find the number of moles of O_2 that are required for the synthesis:

$$(1.091 \times 10^4\ \text{mol}\ CH_4)\left(\frac{3\ \text{mol}\ O_2}{2\ \text{mol}\ CH_4}\right) = 1.637 \times 10^4\ \text{mol}\ O_2$$

Notice that including the formula in the unit, "mol O_2" instead of "mol," for example, clarifies the cancellation of units. Finally, use the molar mass of O_2 to convert from moles back to mass:

$$(1.637 \times 10^4\ \text{mol}\ O_2)\left(\frac{32.00\ \text{g}\ O_2}{1\ \text{mol}\ O_2}\right)\left(\frac{1\ \text{kg}}{10^3\ \text{g}}\right) = 524\ \text{kg}\ O_2$$

7. ✓ **Does the Result Make Sense?** The units of the result, kilograms, indicate that the calculations are in order, because the question asked for the mass of oxygen. The mass of O_2 required for the synthesis, 524 kg, is a large quantity, but it is reasonable because the reaction also involves a large quantity of methane (175 kg).

Common Pitfalls: Inadvertently inverting the ratio of coefficients will give an incorrect result. Inspect the balanced reaction and predict whether there will be a greater or smaller number of moles of product than reactant.

Practice Exercise 1.13

In this synthesis, what mass of ammonia will be required, and what mass of HCN could be prepared?

NEED MORE PRACTICE? Try Problems 1.55–1.58

SECTION EXERCISES

1.8.1 Phosphorus trichloride is produced from the reaction of solid phosphorus and chlorine gas:

$$P_4 + 6\ Cl_2 \longrightarrow 4\ PCl_3$$

(a) What mass of phosphorus trichloride can be prepared from 75.0 g of phosphorus?

(b) What mass of chlorine will be consumed in the reaction?

1.8.2 Tooth enamel consists, in part, of $Ca_5(PO_4)_3(OH)$. Tin(II) fluoride (also called *stannous fluoride*) is added to some toothpastes because it exchanges F^- for OH^- to give the more decay-resistant compound fluoroapatite, $Ca_5(PO_4)_3F$. In addition, this reaction produces tin(II) oxide and water. What mass of $Ca_5(PO_4)_3(OH)$ can be converted to $Ca_5(PO_4)_3F$ by reaction with 0.115 g of tin(II) fluoride?

1.8.3 Combustion reactions require molecular oxygen. In an automobile the fuel-injection system must be adjusted to provide the right "mix" of gasoline and air. Compute the number of grams of oxygen required to react completely with 1.00 L of octane (C_8H_{18}, which has a density $\rho = 0.80$ g/mL). What masses of water and carbon dioxide are produced in this reaction?

NEED MORE PRACTICE? Try Problems 1.61 and 1.62

1.9 YIELDS OF CHEMICAL REACTIONS

Under practical conditions, chemical reactions almost always produce smaller amounts of products than the amounts predicted by stoichiometric analysis. There are three major reasons for this.

1. Many reactions stop before reaching completion. For example, solid sodium metal reacts with oxygen gas to give sodium peroxide (Na_2O_2). As the reaction proceeds, a solid crust of the product builds up on the surface of the metal, as Figure 1-17 shows. The crust prevents oxygen from reaching the sodium metal, and the reaction stops, even though both starting materials are still present.

   Other reactions do not go to completion because they reach dynamic equilibrium. While reactant molecules continue to form product molecules, product molecules also interact to re-form reactant molecules. The Haber reaction and many precipitation reactions, described later in this chapter, are examples of reactions that reach dynamic equilibrium rather than going to completion. We treat chemical equilibria in detail in Chapters 14–16.

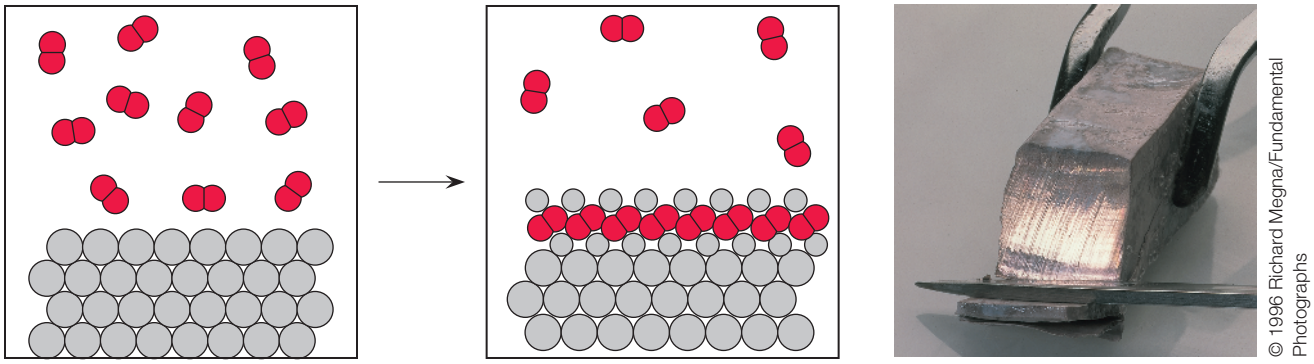


Figure 1-17

The reaction of sodium metal with oxygen gas produces a crust of sodium peroxide, which blocks additional reactant molecules from reaching the unreacted metal.

2. Competing reactions often consume some of the starting materials. For example, sodium metal reacts with water to produce sodium hydroxide. If a sample of oxygen is contaminated with water vapour, both O_2 and H_2O will react with the sodium metal. The more water present in the gas mixture, the less Na_2O_2 will be formed.
3. When the product of a reaction is purified and isolated, some of it is inevitably lost during the collection process. Gases may escape while being pumped out of a reactor. Liquids adhere to glass surfaces, making it impossible to transfer every drop of a liquid product. Likewise, it is impossible to scrape every trace of a solid material from a reaction vessel.

We can compare the incomplete formation of a chemical product with the process of harvesting. When a farmer harvests a fruit crop, some of the fruit is left on the trees. Some may spill off the truck on the way to market. Some of the fruit may be unsuitable for sale. The result is that less than 100% of the potential crop eventually finds its way to market. Likewise, a synthetic chemist loses some of the potential product of a reaction at each step in the preparation.

The amount of a product obtained from a reaction is often reported as a **yield**. The amount of product predicted by stoichiometry is the **theoretical yield**, whereas the amount actually obtained is the **actual yield**. The **percent yield** is the percentage of the theoretical amount that is actually obtained:

$$\text{Percent yield} = 100\%\left(\frac{\text{Actual amount}}{\text{Theoretical amount}}\right) \tag{1-7}$$

When we calculate a percent yield, the amounts can be expressed in either moles or mass, provided that both the actual and theoretical amounts are in the same units. Example 1-14 shows how to use Equation 1-7.

EXAMPLE 1-14 CALCULATING PERCENT YIELD

According to Example 1-12, it is possible to make 443 g of geranyl formate from 375 g of geraniol. A chemist making geranyl formate for the preparation of a perfume uses 375 g of starting material and collects 417 g of purified product. What is the percent yield of this synthesis?



Courtesy of John Olmsted

Strategy: To calculate a percent yield, we need to compare the actual amount obtained in the synthesis with the theoretical amount that could be produced, using Equation 1-7.

Solution: The theoretical yield of geranyl formate is 443 g. This is the amount of product that would result from complete conversion of geraniol into geranyl formate. The actual yield, 417 g, is the quantity of the desired product that the chemist collects. The percent yield is their ratio multiplied by 100%:

$$\text{Percent yield} = (100\%)\left(\frac{417\text{ g}}{443\text{ g}}\right) = 94.1\%$$

✓ **Does the Result Make Sense?** Yields can never exceed the theoretical amount, so a percent yield must always be less than 100%. A percent yield between 90% and 100% is quite reasonable

Practice Exercise 1.14

If the reaction described in Example 1-13 produces 195 kg of HCN from 175 kg of methane reactant, what is the percent yield?

NEED MORE PRACTICE? Try Problems 1.63 and 1.64

If the percent yield of a reaction is already known, we can calculate how much of a product to expect from a synthesis that uses a known amount of starting material. For example, the Haber synthesis of ammonia stops when 13% of the starting materials have formed products. Knowing this, how much ammonia could an industrial producer expect to make from 2.0 tonnes of molecular hydrogen? First, calculate the theoretical yield:

$$\text{Theoretical yield} = \frac{(2.0 \text{ tonne } H_2)(10^6 \text{ g/tonne})}{(2.016 \text{ g/mol } H_2)}\left(\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}\right)$$
$$= 6.61 \times 10^5 \text{ mol } NH_3$$

Rearrange Equation 1-7 and substitute known values:

$$\text{Actual yield} = (6.61 \times 10^5 \text{ mol})\left(\frac{13\%}{100\%}\right) = 8.60 \times 10^4 \text{ mol } NH_3$$

Finally, convert back to mass:

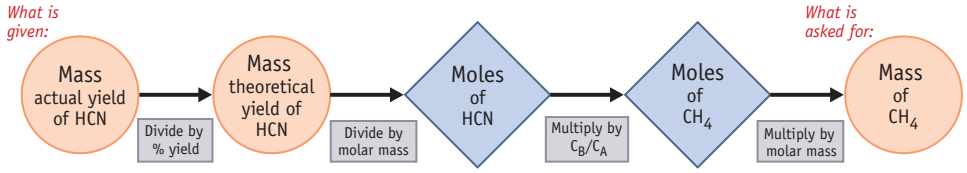
$$\text{Actual mass} = (8.60 \times 10^4 \text{ mol})(17.03 \text{ g/mol})(10^{-6} \text{ tonne/g}) = 1.5 \text{ tonne}$$

Chemists often perform yield calculations in reverse to determine the masses of reactants needed to obtain a desired amount of product. If the percent yield is known from previous experiments, we can calculate the amount of starting material needed to make a specific amount of product. Example 1-15 shows how to do this.

EXAMPLE 1-15 CALCULATING REACTANT MASS FROM YIELD

The industrial production of hydrogen cyanide is described in Example 1-13. If the yield of this synthesis is 97.5%, how many kilograms of methane should be used to produce 1.50×10^5 kg of HCN?

Strategy: This is a two-stage problem that requires a yield calculation and a conversion among molar amounts of products and reactants. This is typical of yield calculations. Use the percent yield to determine the theoretical yield, and then use the stoichiometric ratios to calculate masses of starting materials. A flowchart summarizes the stepwise strategy.



Solution: First, rearrange the equation for percent yield to find the theoretical yield that will give an actual yield of 1.50×10^5 kg:

$$\text{Theoretical yield} = (100\%)\left(\frac{\text{Actual yield}}{\text{Percent yield}}\right)$$

$$\text{Theoretical yield} = (100\%)\left(\frac{1.50 \times 10^5 \text{ kg}}{97.5\%}\right) = 1.538 \times 10^5 \text{ kg HCN}$$

$$(1.538 \times 10^5 \text{ kg HCN})(10^3 \text{ g/kg}) = 1.538 \times 10^8 \text{ g HCN}$$

$$\frac{1.538 \times 10^8 \text{ g HCN}}{27.03 \text{ g/mol}} = 5.690 \times 10^6 \text{ mol HCN}$$

Next, to convert this mass to moles, apply the stoichiometric ratio and multiply by the molar mass of methane:

$$(5.690 \times 10^6 \text{ mol HCN})\left(\frac{2 \text{ mol } CH_4}{2 \text{ mol HCN}}\right) = 5.690 \times 10^6 \text{ mol } CH_4$$

$$(5.690 \times 10^6 \text{ mol } CH_4)(16.04 \text{ g/mol})(10^{-3} \text{ kg/g}) = 9.13 \times 10^4 \text{ kg } CH_4$$

✓ **Does the Result Make Sense?** Compare the mass of the reactant with the mass of the product. Notice that both masses are the same order of magnitude, 10^5 kg. This makes sense when the yield is close to 100%, the stoichiometric ratio is 1:1, and the molar masses are comparable.

Common Pitfalls: Doing too many of the steps at once may lead to errors. Follow the procedure outlined above for this type of problem.

Practice Exercise 1.15
What mass of NH_3 is required for this synthesis?

NEED MORE PRACTICE? Try Problems 1.65–1.68

Suppose a chemical reaction gives a 50% yield. Is that a good yield or a poor yield? The answer depends on the circumstances. Many valuable chemicals are manufactured in processes that require several steps. The overall yield of a multistep synthesis is the product of the yields of the individual steps. For example, suppose that a reaction requires two steps. If the first reaction has a yield of 85% and the second a yield of 65%, then the overall yield is (100%)(0.85)(0.65) = 55%. Notice that the overall yield is lower than the yield of the least efficient step. For this reason, the more steps there are in a synthesis, the more important the yields of the individual reactions become. For example, the synthesis of a growth hormone might require a sequence of 16 different chemical reactions. At an average yield per step of 95%, the overall yield of this synthesis is 44%, but at an average yield of 75% the overall yield of the synthesis falls to 1.0%!

Multistep reactions that have only moderate yields at each step are wasteful and expensive. For this reason, chemists devote much time, effort, and ingenuity to devising reaction sequences and conditions that improve the yields of chemical syntheses. Our most important industrial chemicals are produced in billion-kilogram quantities on an annual basis. Here, improving the synthesis yield by even a few tenths of a percent can save a company millions of dollars each year. A good example is fertilizer production, discussed in Focus 1-2.

FOCUS 1-2 Chemistry and Life: Feeding the World

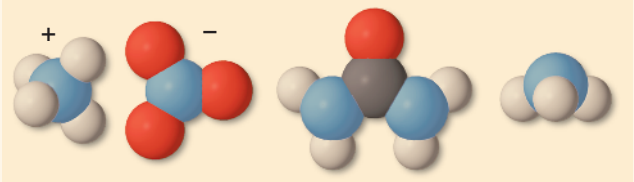
Fertilizers are immensely important to humanity. Agriculture requires fertilizers because growing plants remove various chemical elements from the soil. In a fully contained ecosystem, decaying organic matter replenishes the soil, but the elements contained in crops that are harvested and shipped elsewhere are not replenished. Thus, intensive agriculture inevitably depletes the soil of essential elements, which must be replaced by fertilization.

Lacking means of fertilization, primitive people farmed a plot of land until the nutrients in the soil were exhausted. Then, they moved to new ground, where they burned the natural vegetation and began farming again. This slash-and-burn method of growing food is still used extensively in South America, where this human activity is destroying vast tracts of rain forest.

Fertilization using animal products has been practised since ancient times. Animal manure returns nutrients to the soil, replenishing elements that are depleted as crops are grown and harvested. It is likely that the use of animal fertilizers quickly followed the domestication of goats, sheep, and cattle.

Growing plants require a variety of chemical elements. Nitrogen, phosphorus, and potassium are required in the greatest amounts, but plants also need trace amounts of calcium, copper, iron, zinc, and other elements. By far, the most substantial need is for nitrogen. The Earth's atmosphere is 80% molecular nitrogen, but plants cannot use N_2. Instead, most plants absorb nitrogen from the soil in the form of nitrate ions.

The production of nitrogen fertilizers is a major activity of the chemical industry. Every year, the top 15 chemicals in industrial production include several nitrogen-containing compounds whose major use is in fertilizers. Molecular nitrogen serves as the primary source of nitrogen for chemical production. Gaseous ammonia (NH_3), which is synthesized from N_2 and H_2, can be injected directly into the ground, where it dissolves in moisture in the soil and serves as a fertilizer. Ammonia is more widely used in reactions with acids to produce other fertilizers: ammonia and nitric acid produce ammonium nitrate (NH_4NO_3), whereas ammonia and sulphuric acid produce ammonium sulphate. These chemicals and urea $(NH_2)_2CO$ are among the most important fertilizers.



NH_4NO_3 $(NH_2)_2CO$ NH_3

Crop yields can rise dramatically with the use of commercial fertilizers. For example, in 1910, a hectare of land in Ontario produced about 3140 kg of corn. In 2007, that same hectare of land produced 7600 kg of corn. This large increase was partly due to the use of fertilizers. Worldwide, approximately 10 billion hectares of land are used to grow food crops. This would probably be enough land to feed the world's population if the entire area could be fertilized commercially. It has been estimated that world crop production would increase by about 50% if about $100 per hectare were spent to apply modern chemical fertilizers. However, it would cost about $160 trillion to produce this additional food. Furthermore, the use of chemical fertilizers can lead to the contamination of streams, lakes, and bays with phosphates and nitrates.

Contamination of the world's oceans by fertilizer runoff is an emerging environmental problem. Ironically, the very fertilizers that make the land bloom can cause the oceans to wither. When water containing excess nitrogen drains into the seas, the nitrogen triggers rapid growth of phytoplankton. These microscopic algae eventually die, and as they rot, they consume the oxygen dissolved in seawater. Without oxygen, animal life of all types, including shellfish, crustaceans, and fish, can no longer survive. Oxygen-starved areas have existed in the Gulf of Mexico and Chesapeake Bay for some time, but in 2003 the United Nations Environmental Program reported that the known dead zones have doubled in the last decade of the twentieth century, to nearly 150 dead zones around the world. Among these zones are portions of the Baltic and Adriatic seas.

In addition, the production of fertilizers from molecular nitrogen is very energy-intensive. Finding ways to produce non-contaminating and inexpensive fertilizers without huge energy investment is a major challenge to today's chemical industry.

As world supplies of petroleum are depleted, and as the Earth's population steadily increases, society will be forced to develop more efficient ways to make fertilizer. Genetic engineering offers a promising solution. There is a remarkable class of bacteria called *Rhizobium* that live in the roots of leguminous plants such as soybeans, peas, and peanuts. This organism can convert molecular nitrogen into ammonia. The plant and the bacterium have a symbiotic relationship. Ammonia produced by the bacterium nourishes the plant, and the plant provides other nutrients to the bacterium. Exactly how the bacterium converts nitrogen to ammonia is the subject of vigorous research. Scientists hope eventually to transfer the bacterial gene responsible for the conversion of nitrogen to ammonia into the cells of non-leguminous plants.



Bloomberg/Getty Images, Inc.

SECTION EXERCISES

1.9.1 When heated, potassium chlorate decomposes to potassium chloride and gaseous molecular oxygen:

$$2\ KClO_3 \longrightarrow 2\ KCl + 3\ O_2$$

What is the theoretical yield of oxygen when 5.00 g of potassium chlorate decomposes? Calculate the percent yield if a 5.00 g sample gives 1.84 g O_2 on decomposition. Give possible reasons why the actual yield is less than the theoretical yield.

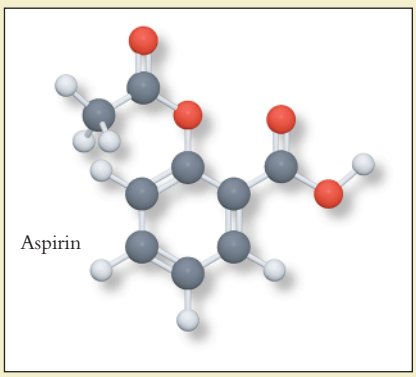
1.9.2 Sulphuric acid is made from elemental sulphur in a three-step industrial process. How many tonnes of sulphur are required to produce 15 tonnes of H_2SO_4 if the three steps have yields of 94%, 92.5%, and 97%?

1.9.3 Aspirin is produced by treating salicylic acid with acetic anhydride:

$$2\ C_7H_6O_3 + C_4H_6O_3 \longrightarrow 2\ C_9H_8O_4 + H_2O$$

Salicylic acid — Acetic anhydride — Aspirin

If the synthesis has an 87% yield, what mass of salicylic acid should be used to produce 75 g of aspirin?



Aspirin

1.10 THE LIMITING REACTANT

The reactions encountered so far in this chapter were set up so that all the starting materials were used up at the same time. In other words, there were no leftovers at the end of the reaction. Often, chemical reactions are run with an excess of one or more starting materials. This means that one reactant will "run out" before the others. The reactant that runs out is called the **limiting reactant** because it limits the amount of product that can be made. The other starting materials are said to be in *excess*.

We illustrate the limiting reactant concept with the bicycle analogy shown in Figure 1-18. It takes two wheels, a frame, and a chain to build a bicycle. Imagine that a bicycle shop has the following parts on hand: eight wheels, five frames, and five chains. How many bicycles can be built with these parts? There are enough frames and chains to make five bicycles. However, because each bicycle needs two wheels, eight wheels are enough to make only four bicycles. Even though there are more wheels around the shop than any of the other parts, the shop will run out of wheels first. After four bicycles are built, the shop inventory will be four bicycles, no wheels, one frame, and one chain. There are not enough parts left to make a fifth bicycle because there is a shortage of two wheels. In this bicycle shop, wheels are the limiting reactant.

Chemical reactions must be analyzed in this same way. Instead of comparing numbers of bicycle parts, we compare the number of moles of each starting material on hand with the number of moles required to make the desired product. Consider the Haber synthesis again. Figure 1-19 shows a limiting reactant situation for this synthesis. If we start with six molecules of H_2 and four molecules of N_2, the six molecules of H_2 will combine with two molecules of N_2 to make four molecules of NH_3. When all the H_2 molecules are consumed, two molecules of N_2 will be left over. Here, H_2 is the limiting reactant. This molecular analysis applies equally well to numbers of moles. In molar terms, if 6 moles of H_2 and 4 moles of N_2 react completely, all the H_2 is consumed, 4 moles of NH_3 are produced, and 2 moles of N_2 are left over.

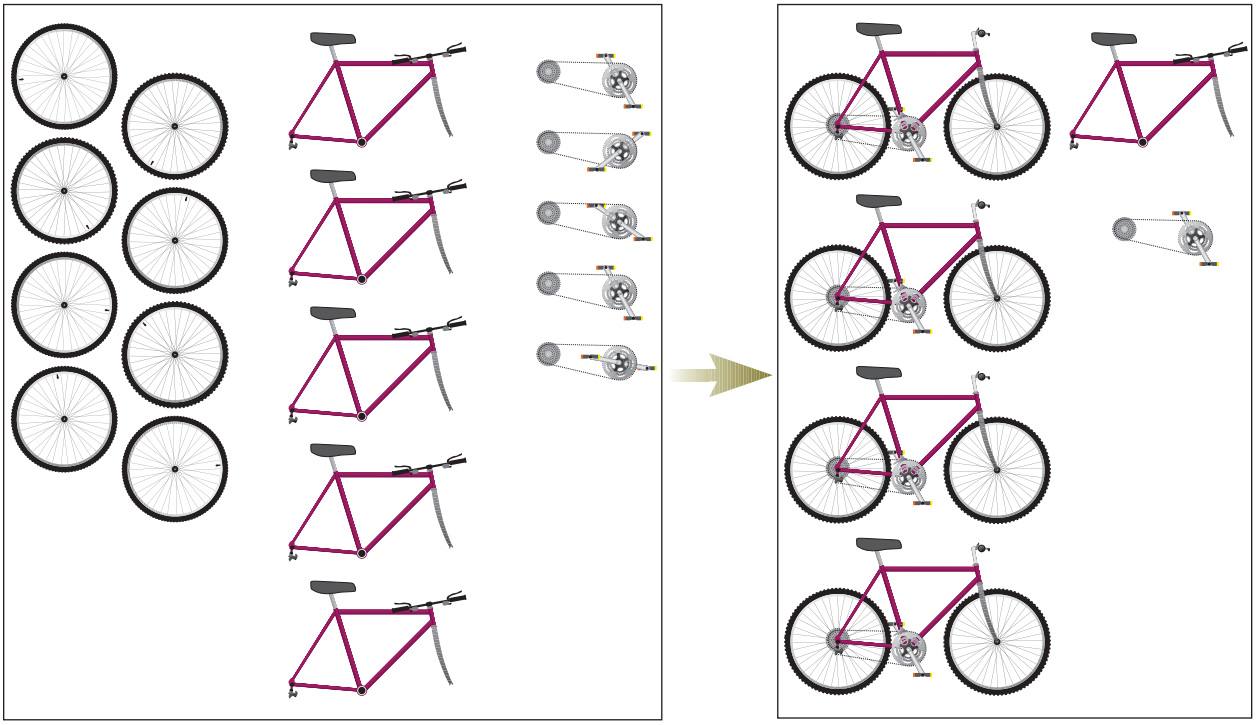


Figure 1-18

The number of bicycles that can be assembled is limited by whichever part runs out first. In the inventory shown in this figure, wheels are that part.

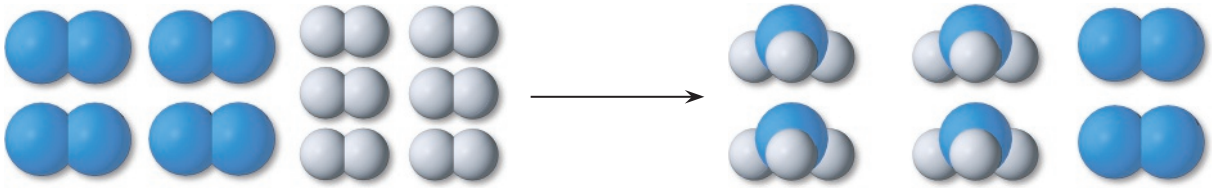


Figure 1-19

A molecular view of a limiting reactant situation for the ammonia synthesis. To make four molecules of NH_3 requires two molecules of N_2 and six molecules of H_2. If we start with four molecules of N_2 and six molecules of H_2, then H_2 is the limiting reactant.

To solve a quantitative limiting reactant problem, we identify the limiting reactant by working with amounts in moles and the stoichiometric coefficients from the balanced equation. For the ammonia synthesis, if we start with 84.0 g of molecular nitrogen and 24.2 g of molecular hydrogen, what mass of ammonia can be prepared? First, convert from masses to moles:

$$\frac{84.0 \text{ g } N_2}{28.01 \text{ g/mol}} = 3.00 \text{ mol } N_2 \qquad \frac{24.2 \text{ g } H_2}{2.016 \text{ g/mol}} = 12.0 \text{ mol } H_2$$

According to the balanced equation, the stoichiometric ratio of nitrogen to ammonia is 1:2.

$$N_2 + 3\ H_2 \longrightarrow 2\ NH_3$$

There are 3.00 moles of N_2, which is enough to make 6.00 moles of NH_3.

$$(3.00 \text{ mol } N_2)\left(\frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2}\right) = 6.00 \text{ mol } NH_3$$

The stoichiometric ratio of hydrogen to ammonia is 3:2. There are 12.0 moles of H_2, which is enough to make 8.00 moles of NH_3.

$$(12.0 \text{ mol } H_2)\left(\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}\right) = 8.00 \text{ mol } NH_3$$

There is enough H_2 on hand to make 8.00 moles of NH_3, but only enough N_2 to make 6.00 moles of NH_3. Nitrogen is the limiting reactant, and hydrogen is present in excess. Once 6.00 moles of ammonia have been produced, there will still be 3.00 moles of hydrogen left over. However, no more ammonia can be produced because there is no nitrogen left for the reaction.

KEY CONCEPT

The limiting reactant is the one whose number of moles divided by its stoichiometric coefficient has the smallest value.

We can identify the limiting reactant by dividing each amount in moles by the stoichiometric coefficient for that reactant. To see how this works, rearrange the ratios for H_2 and N_2.

$$\left(\frac{12.0 \text{ mol } H_2}{3 \text{ mol } H_2}\right)(2 \text{ mol } NH_3) = 4.00\ (2 \text{ mol } NH_3) = 8.00 \text{ mol } NH_3$$

$$\left(\frac{6.00 \text{ mol } N_2}{2 \text{ mol } N_2}\right)(2 \text{ mol } NH_3) = 3.00\ (2 \text{ mol } NH_3) = 6.00 \text{ mol } NH_3$$

The reactant with the smaller ratio of amount to coefficient can produce a smaller amount of product. That reactant (N_2 in this case) is limiting because it runs out first. This leads to the following generalization:

The limiting reactant is the one whose number of moles divided by its stoichiometric coefficient has the smallest value.

Tables of Amounts

LOOKING FORWARD Making tables of amounts is a necessary skill for doing equilibrium calculations (Chapters 14, 15, and 16).

A table of amounts is a convenient way to organize the data and summarize the calculations of a stoichiometry problem. Such a table helps to identify the limiting reactant, shows how much product will form during the reaction, and indicates how much of the excess reactant will be left over. A table of amounts has the balanced chemical equation at the top. The table has one column for each substance involved in the reaction and three rows listing amounts. The first row lists the starting amounts for all the substances. The second row shows the changes that occur during the reaction, and the last row lists the amounts present at the end of the reaction. Here is a table of amounts for the ammonia example:

Reaction	**N_2**	+	**3 H_2**	⟶	**2 NH_3**
Starting amount (mol)	3.00		12.0		0
Change in amount (mol)	–3.00		–9.00		+6.00
Final amount (mol)	0		3.0		6.00

A table of amounts has three important features, all involving *changes in amounts:*

1. ***Changes are negative for reactants and positive for products***. This is because the amounts of reactants decrease during the reaction and the amounts of products increase during the reaction.
2. ***All changes are related by stoichiometry***. Each ratio of changes in amount equals the ratio of stoichiometric coefficients in the balanced equation. In the example above, the changes in amounts for H_2 and N_2 are in the ratio 3:1, the same as the ratio for the coefficients of H_2 and N_2 in the balanced equation.
3. In each column, ***the starting amount plus the change in amount gives the final amount***. Each entry in the final amount row of the table is the sum of the previous entries in the same column.

KEY CONCEPT

The quantity of the limiting reactant determines how much products are formed.

The table of amounts for the ammonia example illustrates these features:

1. The changes are negative for N_2 and H_2, which are consumed, and positive for NH_3, which is produced.
2. The changes in amount for H_2 and N_2 are in the ratio of 3:1, the same as the stoichiometric ratio for the coefficients of H_2 and N_2 in the balanced equation.
3. The change of –9.00 moles for H_2 leaves (12.0 – 9.00) = 3.0 moles as the final amount.

You can use these three key features to construct a table of amounts for any reaction.

When a reaction goes to completion, the limiting reactant is consumed completely, so its final amount must be zero. Two facts help identify the limiting reactant. First, final amounts can never be negative. If a negative amount appears in the bottom row of the table, a mistake has been made in identifying the limiting reactant. Second, the limiting reactant is always the one whose ratio of moles to stoichiometric coefficient is smallest. The smallest ratio indicates the limiting reactant, in this case N_2. Thus, dividing the starting amounts by coefficients is a quick way to identify the limiting reactant. Example 1-16 applies the features of limiting reactants and amounts tables.

EXAMPLE 1-16 LIMITING REACTANT CALCULATION

Nitric acid, a leading industrial chemical, is used in the production of fertilizers and explosives. One step in the industrial production of nitric acid is the reaction of ammonia with molecular oxygen to form nitrogen oxide:

$$4\ NH_3 + 5\ O_2 \longrightarrow 4\ NO + 6\ H_2O$$

In a study of this reaction, a chemist mixed 125 g of ammonia with 256 g of oxygen and allowed them to react to completion. What masses of NO and H_2O were produced, and which reactant was left over and what was its final mass?

Strategy: We apply the seven-step problem-solving method, but without numbering the steps. We are asked about masses of reagents in a chemical reaction. From the fact that starting amounts of both reactants are given, we identify this as a limiting reactant problem (one reactant will be consumed before the other is consumed). A table of amounts will help organize the information, and we need the balanced chemical equation, which is given in the statement of the problem.

Solution: The data available are starting masses of the reactants: 125 g of NH_3 and 256 g of O_2. We also need the molar masses, which are easily calculated:

$$M_{NH_3} = 17.03\ \text{g/mol},\ M_{O_2} = 32.00\ \text{g/mol},\ M_{NO} = 30.01\ \text{g/mol},\ M_{H_2O} = 18.02\ \text{g/mol}$$

As always in stoichiometric calculations, we need to work with moles:

$$n_{NH_3} = \frac{125\ \text{g}}{17.03\ \text{g/mol}} = 7.34\ \text{mol},\ n_{O_2} = \frac{256\ \text{g}}{32.00\ \text{g/mol}} = 8.00\ \text{mol}$$

Determine the limiting reactant by dividing each starting amount by the stoichiometric coefficient for that reactant:

$$\frac{7.34\ \text{mol}\ NH_3}{4\ \text{mol}\ NH_3} = 1.84, \qquad \frac{8.00\ \text{mol}\ O_2}{5\ \text{mol}\ O_2} = 1.60$$

The ratio for O_2 is smaller, so oxygen is the limiting reactant.

Now set up and complete a table of amounts. Because O_2 is limiting, it will be completely consumed, so its change is 8.00 moles. Knowing this, we calculate all other changes in amounts using ratios of stoichiometric coefficients:

$$\text{Change in } NH_3 = (8.00\ \text{mol}\ O_2)\left(\frac{4\ \text{mol}\ NH_3}{5\ \text{mol}\ O_2}\right) = 6.40\ \text{mol}\ NH_3\ \text{used}$$

$$\text{Change in NO} = (8.00\ \text{mol}\ O_2)\left(\frac{4\ \text{mol NO}}{5\ \text{mol}\ O_2}\right) = 6.40\ \text{mol NO produced}$$

$$\text{Change in } H_2O = (8.00\ \text{mol}\ O_2)\left(\frac{6\ \text{mol}\ H_2O}{5\ \text{mol}\ O_2}\right) = 9.60\ \text{mol}\ H_2O\ \text{produced}$$

Remember that changes in amounts are negative for reactants and positive for products. Here is the complete amounts table:

Reaction	**4 NH_3**	+	**5 O_2**	⟶	**4 NO**	+	**6 H_2O**
Starting amount (mol)	7.34		8.00		0		0
Change in amount (mol)	–6.40		–8.00		+6.40		+9.60
Final amount (mol)	0.94		0.00		6.40		9.60

All that remains is to convert back to masses using $m = nM$. We leave it to you to verify that the masses present at the end of the reaction are 16 g of NH_3, 192 g of NO, and 173 g of H_2O.

✓ **Does the Result Make Sense?** Are these masses reasonable? First, it is reasonable that none of the final amounts is negative. Second, we can check to see if total mass is conserved. We started with 125 g of ammonia and 256 g of oxygen, a total of 381 g. The masses present at the end also total 381 g, indicating that the calculations are consistent and the results are reasonable.

Common Pitfalls: Make sure the "change in amount" numbers are based on the stoichiometry of the balanced reaction.

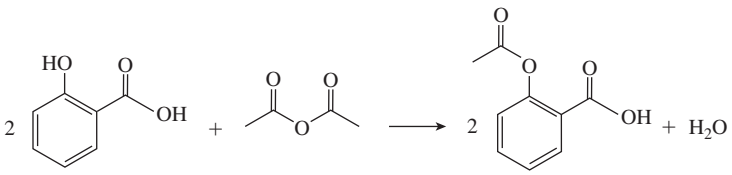
Practice Exercise 1.16

Construct an amounts table for this reaction if the starting amounts are 85.15 g of NH_3 and 224 g of O_2.

A reaction that is carried out under limiting reactant conditions nevertheless has a yield that generally will be less than 100%. The reasons why reactions yield less than the theoretical amounts, given in Section 1.9, apply to all reactions. When a reaction operates under limiting reactant conditions, we calculate the theoretical yield assuming that the limiting reactant will be completely consumed. We then determine the percent yield as described in Section 1.9. Example 1-17 shows how to do this.

EXAMPLE 1-17 YIELD UNDER LIMITING REACTANT CONDITIONS

The synthesis of aspirin appears in Section Exercise 1.9.3:



	$C_7H_6O_3$	$C_4H_6O_3$	$C_9H_8O_4$
Chemical formula:			
Name:	Salicylic acid	Acetic anhydride	Aspirin
M:	138.1 g/mol	102.1 g/mol	180.2 g/mol

Suppose a chemist started with 152 g of salicylic acid and 86.8 g of acetic anhydride and produced 133 g of aspirin. What is the yield of this reaction?

Strategy: The problem asks for a yield, so we identify this as a yield problem. In addition, we recognize this as a limiting reactant situation because we are given the masses of both starting materials. First, identify the limiting reactant by working with moles and stoichiometric coefficients. Then, carry out standard stoichiometry calculations to determine the theoretical amount that could form. A table of amounts helps organize these calculations. Calculate the percent yield from the theoretical amount and the actual amount formed.

Solution: Begin by converting mass to moles. The problem gives the molar masses.

$$\frac{152\text{ g } C_7H_6O_3}{138.1\text{ g/mol}} = 1.10\text{ mol } C_7H_6O_3 \qquad \frac{86.8\text{ g } C_4H_6O_3}{102.1\text{ g/mol}} = 0.850\text{ mol } C_4H_6O_3$$

Now identify the limiting reactant:

$$\frac{1.10\text{ mol } C_7H_6O_3}{2\text{ mol } C_7H_6O_3} = 0.550 \qquad \frac{0.850\text{ mol } C_4H_6O_3}{1\text{ mol } C_4H_6O_3} = 0.850$$

Because it has the smaller ratio, $C_7H_6O_3$ is the limiting reactant.
We determine the entries in the change row using the amount of the limiting reactant and stoichiometric ratios.

$$(1.10\text{ mol } C_7H_6O_3)\left(\frac{1\text{ mol } C_4H_6O_3}{2\text{ mol } C_7H_6O_3}\right) = 0.550\text{ mol } C_4H_6O_3 \text{ consumed}$$

$$(1.10\text{ mol } C_7H_6O_3)\left(\frac{2\text{ mol } C_9H_8O_4}{2\text{ mol } C_7H_6O_3}\right) = 1.10\text{ mol } C_9H_8O_4 \text{ produced}$$

$$(1.10\text{ mol } C_7H_6O_3)\left(\frac{1\text{ mol } H_2O}{2\text{ mol } C_7H_6O_3}\right) = 0.550\text{ mol } H_2O \text{ produced}$$

(*continued*)

We can now complete the table of amounts:

Reaction	2 $C_7H_6O_3$	+	$C_4H_6O_3$	⟶	2 $C_9H_8O_4$	+	H_2O
Starting amount (mol)	1.10		0.850		0		0
Change in amount (mol)	–1.10		–0.550		+1.10		+0.550
Final amount (mol)	0		0.300		1.10		0.550

The final amount in this table represents the theoretical yield. To determine the percent yield, we need to have actual and theoretical amounts in the same units. In this example, we work with moles:

$$\text{Actual yield} = \frac{133 \text{ g } C_9H_8O_4}{180.2 \text{ g } C_9H_8O_4/\text{mol}} = 0.738 \text{ mol}$$

$$\text{Percent yield} = (100\%)\left(\frac{0.738 \text{ mol}}{1.10 \text{ mol}}\right) = 67.1\%$$

✓ **Does the Result Make Sense?** The entries in the amounts table are all reasonable, and the yield turns out to be less than 100%, as we know must be the case.

Practice Exercise 1.17
If 152 g of salicylic acid and 41.2 g of acetic anhydride react to form 114 g of aspirin, what is the percent yield?

Masses and number ratios cannot be used to determine the "limiting reactant" in bicycle manufacture either. Five chains are fewer in number than eight wheels and have less mass than eight wheels. Nevertheless, the wheels are used up before the chains are.

Notice that neither the masses of the reactants nor the number of moles indicate directly which reactant is limiting. In this example, salicylic acid ($C_7H_6O_3$) is the limiting reactant even though the mass of salicylic acid and the number of moles of salicylic acid are both larger than the corresponding amounts of acetic anhydride ($C_4H_6O_3$).

It is often desirable to run practical syntheses under limiting reactant conditions, even though this may appear to be wasteful. For example, the aspirin synthesis typically is run with salicylic acid as the limiting reactant. Making salicylic acid the limiting reactant means that the product mixture will contain acetic anhydride and aspirin but a negligible amount of salicylic acid. This is advantageous for two reasons. First, salicylic acid and aspirin have similar molecular structures, making it difficult to separate these two substances. In contrast, aspirin is easy to separate from acetic anhydride. Second, salicylic acid is more expensive than acetic anhydride, and it is more economical to consume the more expensive reactant completely than to leave some of it unreacted.

SECTION EXERCISES

- **1.10.1** Suppose that the industrial synthesis of HCN described in Example 1-13 is carried out using 5.00×10^2 kg each of ammonia and methane in excess oxygen. What is the maximum mass of HCN that could be produced? Which reactant would be left over and what would be its final mass?
- **1.10.2** Lithium metal is one of the few substances that reacts directly with molecular nitrogen:

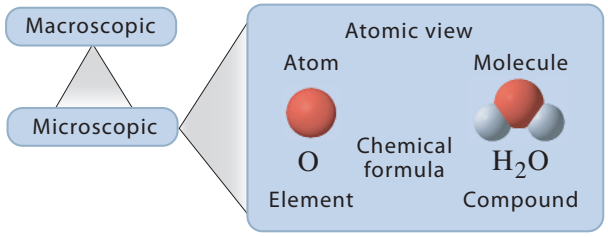
$$6 \text{ Li} + N_2 \longrightarrow 2 \text{ } Li_3N$$

  What mass of the product, lithium nitride, can be prepared from 4.5 g of lithium metal and 9.5 g of molecular nitrogen?
- **1.10.3** Solid carbon reacts with chlorine gas to give carbon tetrachloride. If 150 g of carbon reacts with 250 g of Cl_2 with an 85% yield, what mass of CCl_4 is produced?

NEED MORE PRACTICE? Try Problems 1.69–1.76

CHAPTER 1 VISUAL SUMMARY

1.1 Atoms, Molecules, and Compounds



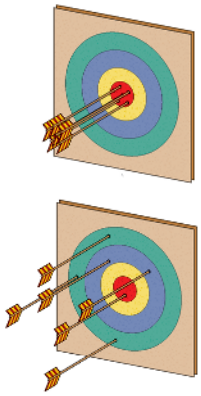
LEARNING OBJECTIVE: Recognize elemental symbols and names of the elements, and name compounds from molecular pictures.

SKILLS TO MASTER: Recognizing elemental symbols and names of the elements

Naming compounds from molecular pictures

KEY CONCEPTS: Properties of molecules (the microscopic level) translate into properties of materials (the macroscopic level).

1.2 Measurements in Chemistry



Precision	Accuracy
Significant figures	
Time (t)	Volume (V)
Temperature (T)	Length
Scientific notation	Mass (m)

$$T\,(\text{K}) = T\,(^\circ\text{C}) + 273.15$$
$$T\,(^\circ\text{C}) = T\,(\text{K}) - 273.15$$

LEARNING OBJECTIVE: Recognize the SI units commonly used in chemistry, and perform some common unit conversions.

SKILLS TO MASTER: Working with unit conversions

Determining the correct number of significant figures in the result of a calculation

KEY CONCEPTS: The correct conversion ratio leads to cancellation of unwanted units.

Proper use of significant figures is important to tell the reader what the accuracy and precision of the measurement are.

1.3 Chemical Problem Solving

Solving Quantitative Problems

Step 1. **Determine what is asked for.**
Step 2. **Visualize the problem.**
Step 3. **Organize the data.**
Step 4. **Identify a process to solve the problem.**
Step 5. **Manipulate the equations.**
Step 6. **Substitute and calculate.**
Step 7. **Does the result make sense?**

LEARNING OBJECTIVE: Analyze and solve problems in a consistent, organized fashion.

1.4 Counting Atoms: The Mole

One mole is the number of atoms in exactly 12 g of pure carbon-12.

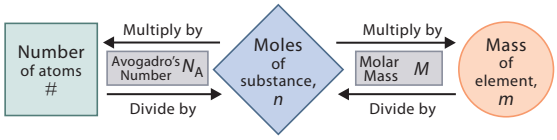
Molar mass (M)
Avogadro's number (N_A)

Elemental molar mass is the sum of the contribution of its isotopes.

$M = \Sigma(\text{Fractional abundance})(\text{Isotopic } M)$

Chemical calculations are built around the mole.

$$n = \frac{\#}{N_A} \qquad n = \frac{m}{M}$$



LEARNING OBJECTIVE: Solve mass–number–molar mass-type problems.

SKILLS TO MASTER: Determining elemental molar masses

KEY CONCEPTS: Chemical calculations are built around the mole.

One mole is the number of atoms in exactly 12 g of the pure isotope carbon-12.

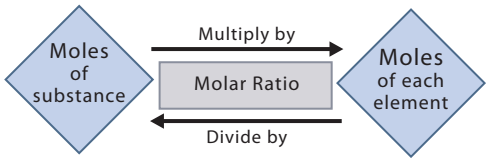
The molar mass of any naturally occurring element is the sum of the contributions from its isotopes.

1.5 Amounts of Compounds

$M_{Compound} = \Sigma(\#\ Atoms)(M_{Element})$

Element	#	M (g/mol)	Sum (g/mol)
N	2	14.007	28.014
H	4	1.008	4.032
O	3	15.999	47.997
NH_4NO_3			M = 80.043 g/mol





LEARNING OBJECTIVE: Perform mole–mass–number conversions.

SKILLS TO MASTER: Drawing line structures of compounds

Calculating molar masses of compounds

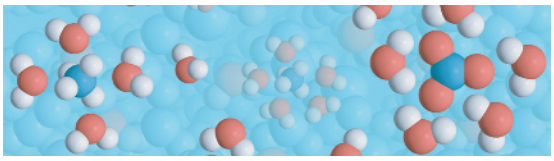
KEY CONCEPTS: A molecule is a combination of two or more atoms held together in a specific shape by attractive forces.

The molar mass of a compound is found by adding together the molar masses of all of its elements, taking into account the number of moles of each element present.

Molar masses are used to convert between moles of substance and mass.

1.6 Aqueous Solutions

Solvent
Solute
Aqueous



Polyatomic ions remain intact.

$c = \frac{n}{V}$ $\quad c_i V_i = c_f V_f$

Concentration (c)
Molarity Dilution

LEARNING OBJECTIVE: Calculate concentrations of solutions and of diluted solutions.

KEY CONCEPTS: Concentration is the amount of substance per unit volume of solution.

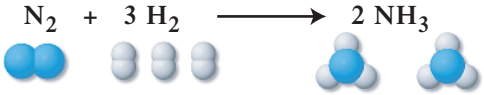
Polyatomic ions remain intact when a salt dissolves in water.

Diluting a solution decreases the concentrations of the solutes.

1.7 Writing Chemical Equations

Stoichiometric coefficients
Balanced chemical equation

Reagents
Reactants Products



Atoms are conserved.

LEARNING OBJECTIVE: Balance chemical reactions.

KEY CONCEPTS: Molecules of reagents react in whole number ratios to form products.

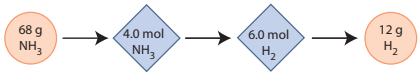
The number of atoms of each element is conserved in any chemical reaction.

1.8 The Stoichiometry of Chemical Reactions

$n_B = \left(\frac{Coeff_B}{Coeff_A}\right) n_A$

Stoichiometric ratio

Calculations centre on the mole.



Stoichiometry

LEARNING OBJECTIVE: Calculate the amount of a product from the amounts of the reactants and a balanced chemical equation.

SKILLS TO MASTER: Using stoichiometric ratios

1.9 Yields of Chemical Reactions

Yield
Theoretical yield
Actual yield
Percent yield

$\text{Percent yield} = 100\% \left(\frac{\text{Actual amount}}{\text{Theoretical amount}}\right)$

LEARNING OBJECTIVE: Calculate yields of chemical reactions.

1.10 The Limiting Reactant

Limiting reactant
Smallest value of moles divided by coefficient

Reaction:	N_2	+ $3\ H_2$	→ $2\ NH_3$
Starting amount (mol)	3.00	12.0	0
Change in amount (mol)	−3.00	−9.00	+6.00
Final amount (mol)	0	3.0	6.00

LEARNING OBJECTIVE: Solve limiting-reagent-type problems.

SKILLS TO MASTER: Constructing tables of amounts

KEY CONCEPTS: The limiting reactant is the one whose number of moles divided by its stoichiometric coefficient has the smallest value.

The quantity of the limiting reactant determines how much products are formed.

LEARNING EXERCISES

ilw = interactive learning ware problem. These can be found in *WileyPLUS*.

1.1 Make a list of all terms new to you in Chapter 1. Give a one-sentence definition for each. Consult the glossary if you need help.

1.2 Diagram the process for converting from the mass of a compound of a known chemical formula to the number of atoms of one of its constituent elements. Include all necessary equations and conversion factors.

1.3 Describe what a solution of magnesium nitrate looks like to a molecular-sized observer.

1.4 Update your list of memory bank equations. For each equation, add a phrase that describes the kinds of chemical problems for which the equation is used.

PROBLEMS

Atoms, Molecules, and Compounds

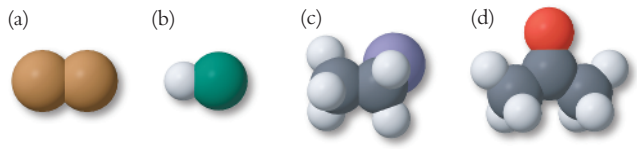
1.1 What is the elemental symbol for each of the following elements? (a) hydrogen; (b) helium; (c) hafnium; (d) nitrogen; (e) neon; and (f) niobium.

1.2 What is the elemental symbol for each of the following elements? (a) potassium; (b) platinum; (c) plutonium; (d) lead; (e) palladium; and (f) phosphorus.

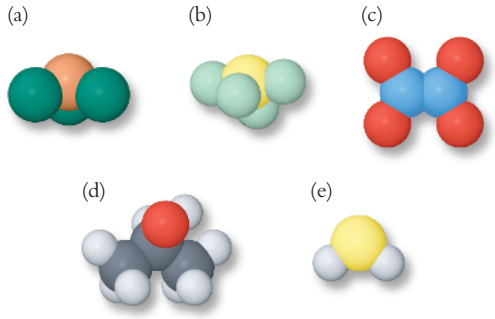
1.3 What is the name of each of the following elements? (a) As; (b) Ar; (c) Al; (d) Am; (e) Ag; (f) Au; (g) At; and (h) Ac.

1.4 What is the name of each of the following elements? (a) Br; (b) Be; (c) B; (d) Bk; (e) Ba; and (f) Bi.

1.5 Examine the following molecular pictures and determine the corresponding molecular formulas.



1.6 Examine the following molecular pictures and determine the corresponding molecular formulas.



1.7 Write the chemical formula of the compound whose molecules contain (a) four atoms of chlorine and one atom of carbon; (b) two atoms of oxygen and two atoms of hydrogen; (c) four atoms of phosphorus and ten atoms of oxygen; (d) two atoms of iron and three atoms of sulphur.

1.8 Write the chemical formula of the compound whose molecules contain (a) five carbon atoms and twelve hydrogen atoms; (b) four atoms of fluorine and one atom of silicon; (c) two atoms of nitrogen and five atoms of oxygen; (d) one iron atom and three chlorine atoms.

Measurements in Chemistry

1.9 Convert to scientific notation: (a) 100 000, precise to ±1; (b) ten thousand, precise to ±1000; (c) 0.000 400; (d) 0.0003; and (e) 275.3.

1.10 Convert to scientific notation: (a) 175 906; (b) 0.000 0605; (c) two and a half million, precise to ±100; and (d) two and a half billion, precise to ± one million.

1.11 Express each of the following in SI base units using scientific notation (e.g., 1.45 mm = 1.45×10^{-3} m): (a) 432 kg; (b) 624 ps; (c) 1024 ng; (d) 93 000 km, precise to ±10; (e) 1 day; and (f) 0.0426 in.

1.12 Express each of the following in SI base units using scientific notation: (a) 1 week; (b) 1.35 mm; (c) 15 km; (d) 4.567 μs; (e) 6.45 mL; and (f) 47 kg.

1.13 The mass unit most commonly used for precious stones is the carat. 1 carat = 3.168 grains, and 1 g = 15.4 grains. Find the total mass in kilograms (kg) of a ring that contains a 5.0×10^{-1} carat diamond and 7.00 g of gold.

1.14 What is the total mass in grams, expressed in scientific notation with the correct number of significant figures, of a solution containing 2.000 kg of water, 6.5 g of sodium chloride, and 47.546 g of sugar?

Chemical Problem Solving

1.15 A chemist who wished to verify the density of water constructed a cylindrical container of aluminum 11.43 cm high, with an inside radius measuring 2.22 cm. The empty cylinder had a mass of 93.054 g. When filled with water, its mass was 270.064 g. Find the density of water from these data, expressing your result in SI units with the correct precision (cylinder volume, $V = \pi r^2 h$).

1.16 A chemist who prepared a new organic liquid wanted to determine its density. Having only a small sample to work with, the chemist had to use a small container. A tube with a volume of 8.00×10^{-3} cm^3 weighed 0.4763 g when empty and 0.4827 g when filled with the liquid. Compute the density of the liquid in g/cm^3.

1.17 ilw Determine which possesses more mass: a sphere of gold with a diameter of 2.00 cm or a cube of silver measuring 2.00 cm on each side (sphere volume, $V = 4\pi r^3/3$).

1.18 Determine which possesses more mass: a sphere of zinc with a diameter of 2.00 cm or a cube of aluminum measuring 2.00 cm on each side (sphere volume, $V = 4\pi r^3/3$).

1.19 Bromine is one of two elements that are liquids at room temperature (mercury is the other). The density of bromine at room temperature is 3.12 g/mL. What volume of bromine is required if a chemist needs 36.5 g for an experiment?

1.20 The density of gasoline at room temperature is 0.70 g/mL. If the gas tank of a car holds 12.0 gallons (45.4 L), what is the mass of a tankful of gasoline?

Counting Atoms: The Mole

1.21 ilw Use the data in the following table to calculate the molar mass of naturally occurring argon:

Isotope	Isotopic Molar Mass (g/mol)	Abundance (%)
^{36}Ar	35.967 55	0.337
^{38}Ar	37.962 72	0.063
^{40}Ar	39.962 38	99.600

1.22 Use the data in the following table to calculate the molar mass of naturally occurring silicon:

Isotope	Isotopic Molar Mass (g/mol)	Abundance (%)
^{28}Si	27.976 93	92.23
^{29}Si	28.976 49	3.10
^{30}Si	29.973 76	4.67

1.23 Calculate the number of moles in the following: (a) 7.85 g of Fe; (b) 65.5 μg of carbon; (c) 4.68 mg of Si; and (d) 1.46 tonnes of Al (1 tonne = 10^3 kg).

1.24 Calculate the number of moles in the following: (a) 3.67 kg of titanium; (b) 7.9 mg of calcium; (c) 1.56 g of ruthenium; and (d) 9.63 pg of technetium.

1.25 Determine the number of atoms present in 5.86 mg of each of the following elements: (a) beryllium; (b) phosphorus; (c) zirconium; and (d) uranium.

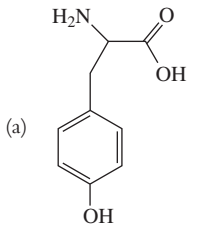
1.26 Determine the number of atoms present in 2.44 kg of each of the following elements: (a) argon; (b) silver; (c) iodine; and (d) potassium.

Amounts of Compounds

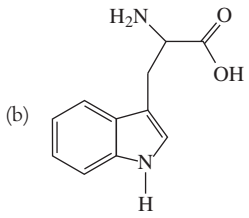
1.27 Calculate the molar mass of each of the following substances: (a) carbon tetrachloride; (b) potassium sulphide; (c) O_3 (ozone); (d) lithium bromide; (e) GaAs (a semiconductor); and (f) silver nitrate.

1.28 Calculate the molar mass of each of the following compounds: (a) ammonium carbonate; (b) N_2O (laughing gas); (c) calcium carbonate; (d) NH_3 (ammonia); (e) sodium sulphate; and (f) C_4H_{10} (butane).

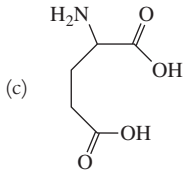
1.29 Determine the molecular formula and calculate the molar mass of each of the following essential amino acids:



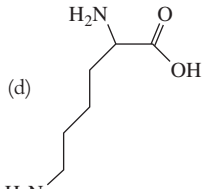
Tyrosine



Tryptophan

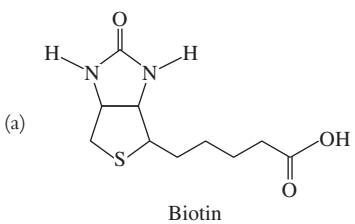


Glutamic acid

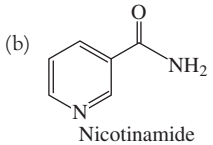


Lysine

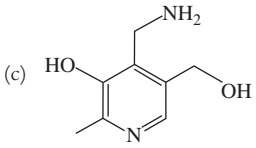
1.30 Determine the molecular formula and calculate the molar mass of each of the following B vitamins:



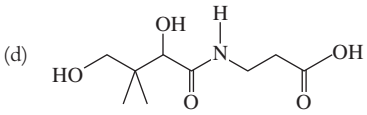
Biotin



Nicotinamide



Pyridoxamine



Pantothenic acid

1.31 Calculate the number of molecules present in 5.86 mg of each of the following compounds: (a) methane (CH_4); (b) phosphorus trichloride; (c) dimethyl ether (C_2H_6O); and (d) uranium hexafluoride.

1.32 The Canadian Recommended Dietary Allowances (RDA) of several vitamins for an adult female follow. In each case, calculate how many molecules are in the Canadian RDA. (a) 75. mg vitamin C, $C_6H_8O_6$; (b) 400. mg folic acid, $C_{19}H_{19}N_7O_6$; (c) 0.7 mg vitamin A, $C_{20}H_{30}O$; and (d) 15 mg vitamin E, $C_{29}H_{50}O_2$.

1.33 Calculate the mass of the following: (a) 3.75×10^5 molecules of methane; (b) 2.5×10^9 molecules of adrenaline, $C_9H_{13}NO_3$; and (c) one molecule of chlorophyll, $C_{55}H_{72}MgN_4O_5$.

1.34 Calculate the mass of the following: (a) 1.0×10^{15} molecules of ozone; (b) 5.000×10^3 molecules of cholesterol, $C_{27}H_{46}O$; and (c) one molecule of vitamin B_{12}, $C_{63}H_{88}CoN_{14}O_{14}P$.

1.35 ilw One carrot may contain 0.75 mg of vitamin A ($C_{20}H_{30}O$). How many moles of vitamin A is this? How many molecules? How many hydrogen atoms are in 0.75 mg of vitamin A? What mass of hydrogen is this?

1.36 A particular oral contraceptive contains 0.035 mg of ethynyl estradiol in each pill. The formula of this compound is $C_{20}H_{24}O_2$. How many moles of ethynyl estradiol are there in one pill? How many molecules is this? How many carbon atoms are in a 0.035 mg sample of ethynyl estradiol? What mass of carbon is this?

Aqueous Solutions

1.37 (a) Calculate the molarities of the ionic species in 1.50×10^2 mL of aqueous solution that contains 4.68 g of magnesium chloride. (b) Draw a molecular picture that shows a portion of this solution, showing the relative proportions of the ions present.

1.38 A chemist places 3.25 g of sodium carbonate in a 250 mL volumetric flask and fills it to the mark with water. (a) Calculate the molarities of the major ionic species. (b) Draw a molecular picture that shows a portion of this solution, making sure the portion is electrically neutral.

1.39 ilw A student prepares a solution by dissolving 4.75 g of solid KOH in enough water to make 275 mL of solution. (a) Calculate the molarities of the major ionic species present. (b) Calculate the molarities of the major ionic species present if 25.00 mL of this solution is added to a 100 mL volumetric flask, and water is added to the mark. (c) Draw molecular pictures of portions of the solutions in (a) and (b), showing how they differ.

1.40 A mildly antiseptic mouthwash can be prepared by dissolving sodium chloride in water. (a) Calculate the molarities of the ionic species present in 0.150 L of solution containing 27.0 g of sodium chloride. (b) Calculate the new molarities if 50.0 mL of this solution is diluted with water to give 450. mL of a new solution. (c) Draw

molecular pictures of portions of the solutions described in (a) and (b), showing how they differ.

1.41 Concentrated hydrochloric acid is 12.1 M. What volume of this solution should be used to prepare 0.500 L of 0.125 M HCl?

1.42 Concentrated ammonia is 14.8 M. What volume of this solution should be used to prepare 1.25 L of 0.500 M NH_3?

1.43 Calculate the concentrations of the major ionic species present in each of the following solutions: (a) 4.55 g of Na_2CO_3 in 245 mL of solution; (b) 27.45 mg of NH_4Cl in 1.55×10^{-2} L of solution; and (c) 1.85 kg of potassium sulphate in 5.75×10^3 L of solution.

1.44 Calculate the concentrations of the major ionic species present in each of the following solutions: (a) 1.54 g of sodium hydrogen carbonate in 75.0 mL of solution; (b) 1.44 mg of $FeCl_3$ in 2.75 mL of solution; and (c) 8.75 kg of KNO_3 in 235 L of solution.

Writing Chemical Equations

1.45 Balance the following chemical equations:

(a) $NH_4NO_3 \longrightarrow N_2O + H_2O$

(b) $P_4O_{10} + H_2O \longrightarrow H_3PO_4$

(c) $HIO_3 \longrightarrow I_2O_5 + H_2O$

(d) $As + Cl_2 \longrightarrow AsCl_5$

1.46 Balance the following chemical equations:

(a) $N_2O_5 + H_2O \longrightarrow HNO_3$

(b) $KClO_3 \longrightarrow KCl + O_2$

(c) $Fe + O_2 + H_2O \longrightarrow Fe(OH)_2$

(d) $P_4 + Cl_2 \longrightarrow PCl_3$

1.47 Draw a molecular picture that illustrates reaction 1.45 (d).

1.48 Draw a molecular picture that illustrates reaction 1.46 (d).

1.49 Balance the chemical equations for the following important industrial processes:

(a) Molecular hydrogen and carbon monoxide react to form methanol (CH_3OH).

(b) $CaO + C \longrightarrow CO + CaC_2$

(c) $C_2H_4 + O_2 + HCl \longrightarrow C_2H_4Cl_2 + H_2O$

1.50 The following reactions play roles in the manufacture of nitric acid. Write a balanced equation for each of them.

(a) $NH_3 + O_2 \longrightarrow NO + H_2O$

(b) Nitrogen oxide and molecular oxygen combine to give nitrogen dioxide.

(c) Nitrogen dioxide and water react to form nitric acid and nitrogen oxide.

(d) $NH_3 + O_2 \longrightarrow N_2 + H_2O$

(e) $NH_3 + NO \longrightarrow N_2 + H_2O$

1.51 Draw a molecular picture that illustrates reaction 1.49 (a).

1.52 Draw a molecular picture that illustrates reaction 1.50 (b).

1.53 Here are some reactions involving water. Balance the equations:

(a) $Ca(OH)_2 + H_3PO_4 \longrightarrow H_2O + Ca_3(PO_4)_2$

(b) $Na_2O_2 + H_2O \longrightarrow NaOH + H_2O_2$

(c) $BF_3 + H_2O \longrightarrow HF + H_3BO_3$

(d) $NH_3 + CuO \longrightarrow Cu + N_2 + H_2O$

1.54 Here are some reactions of phosphorus and its compounds. Balance the equations:

(a) $P_4 + Na \longrightarrow Na_3P$

(b) $Na_3P + H_2O \longrightarrow PH_3 + NaOH$

(c) $PH_3 + O_2 \longrightarrow P_4O_{10} + H_2O$

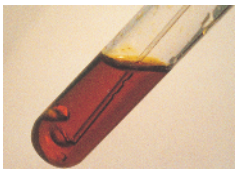
(d) $P_4O_{10} + H_2O \longrightarrow H_3PO_4$

The Stoichiometry of Chemical Reactions

1.55 For each reaction in Problem 1.49, calculate the mass of the second reactant that is required to react completely with 5.00 g of the first reactant.

1.56 For each reaction in Problem 1.50, calculate the mass (in kg) of the first reactant that is required to react completely with 875 kg of the second reactant.

1.57 Iodine can be prepared by bubbling chlorine gas through an aqueous solution of sodium iodide:



Courtesy of John Olmsted

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

What mass (in g) of NaI is required to produce 1.50 kg of I_2?

1.58 The fertilizer ammonium sulphate is prepared by the reaction between ammonia and sulphuric acid:

$$2\ NH_3(g) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq)$$

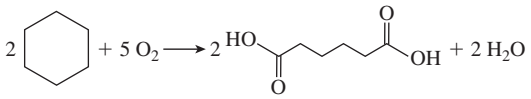
What mass (in kg) of NH_3 is needed to make 3.50 tonnes (1 tonne = 1000 kg) of ammonium sulphate?

1.59 The fermentation of sugar to produce ethyl alcohol occurs by the following reaction:

$$C_6H_{12}O_6(s) \xrightarrow{\text{yeast}} 2\ C_2H_5OH(l) + 2\ CO_2(g)$$

What mass of ethyl alcohol can be made from 1.00 kg of sugar?

1.60 One starting material for the preparation of nylon is adipic acid. Adipic acid is produced from the oxidation of cyclohexane:



Cyclohexane C_6H_{12} Adipic acid $C_6H_{10}O_4$

If 375 kg of cyclohexane reacts with an unlimited supply of oxygen, how much adipic acid (in kg) can be formed?

1.61 ilw The use of Freons is being phased out because of the damage these compounds do to the stratospheric ozone layer. One of the Freons, CCl_2F_2, is manufactured from carbon tetrachloride and hydrogen fluoride. The other reaction product is hydrogen chloride. If a manufacturer wishes to convert 175 kg of carbon tetrachloride into CCl_2F_2, what is the minimum mass of hydrogen fluoride required? What would be the masses of the two products obtained if this reaction is 100% efficient?



CCl_2F_2

1.62 One way to clean up emissions of sulphur dioxide formed during the burning of coal is to pass the exhaust gases through a water slurry of powdered limestone. In the reaction, sulphur dioxide reacts with calcium carbonate and oxygen gas to produce calcium sulphate and carbon dioxide. A coal-burning plant that uses coal with a typical sulphur content produces 2.00×10^4 tonnes (1 tonne = 1000 kg) of sulphur dioxide in a month. How many tonnes of calcium carbonate will be consumed during the scrubbing process, assuming 100% efficiency? How many tonnes of calcium sulphate waste will the plant have to dispose of? (The process is actually about 90% efficient.)

Yields of Chemical Reactions

1.63 ilw Approximately 12 billion kilograms of phosphoric acid are produced annually for fertilizers, detergents, and agents for water treatment. Phosphoric acid can be prepared by heating the mineral fluoroapatite with sulphuric acid in the presence of water:

$$\underset{\text{Fluoroapati}}{Ca_5(PO_4)_3F} + 5\,H_2SO_4 + 10\,H_2O \longrightarrow \underset{\text{Phosphoric acid}}{3\,H_3PO_4} + 5\,CaSO_4 \cdot 2\,H_2O + HF$$

If every kilogram of fluoroapatite yields 4.00×10^2 g of phosphoric acid, what is the percent yield?

1.64 When HgO is heated, it decomposes into elemental mercury and molecular oxygen gas. If 60.0 g of Hg is obtained from 80.0 g of the oxide, what is the percent yield of the reaction?

1.65 Phenobarbital is a sleep-inducing drug whose chemical formula is $C_{12}H_{12}N_2O_3$. It is manufactured in an eight-step process starting from toluene, C_7H_8. Theoretically, each molecule of toluene yields one molecule of phenobarbital. If each of the steps has a yield of 88%, what mass of toluene is needed to manufacture 25 kg of phenobarbital?

1.66 Most of the ammonia produced by the Haber process is used as fertilizer. A second important use of NH_3 is in the production of nitric acid, a top-15 industrial chemical. Nitric acid is produced by a three-step synthesis called the *Ostwald process:*

$$4\,NH_3 + 5\,O_2 \longrightarrow 4\,NO + 6\,H_2O$$
$$2\,NO + O_2 \longrightarrow 2\,NO_2$$
$$3\,NO_2 + H_2O \longrightarrow 2\,HNO_3 + NO$$

The NO is recycled so that every mole of ammonia theoretically yields 1 mole of nitric acid. Starting with 7.50×10^2 kg of ammonia, what mass of nitric acid can be produced if the average yield of the three steps is 94.5% efficient?

1.67 What is the yield of the reaction described in Problem 1.61 if the manufacturer obtains 105 kg of CCl_2F_2 from 175 kg of CCl_4? Given this reaction yield, what masses of CCl_4 and hydrogen fluoride should be used in order to make 155 kg of CCl_2F_2?

1.68 If the oxidation reaction described in Problem 1.60 is 76.5% efficient, what mass of cyclohexane is required to produce 3.50 kg of adipic acid?

The Limiting Reactant

1.69 A fast-food restaurant makes double cheeseburgers using one hamburger roll, two quarter-pound patties of beef, one slice of cheese, a fourth of a tomato, and 15 g of shredded lettuce. At the start of the day, the restaurant manager has 12 dozen rolls, 40 lb of beef patties, 2 packages of sliced cheese containing 65 slices each, 40 tomatoes, and 1 kg of lettuce. Which ingredient will run out first, and how many cheeseburgers can be made?

1.70 A hardware store sells a "handyman assortment" of wood screws containing the following:

Type	Number of Screws	Mass per Screw (g)
#8, $^1/_2$″	10	12.0
#8, 1″	8	21.5
#10, $^1/_2$″	8	14.5
#10, $1^1/_2$″	5	31.0
#12, 2″	4	45.0

The store buys these screws in 1 lb boxes (1 lb = 454 g). From two boxes of each type of screws, how many handyman assortments can the store prepare?

1.71 ilw Ammonia is produced industrially using the Haber process:

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$$

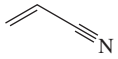
Suppose that an industrial reactor is charged with 75.0 kg each of N_2 and H_2. Use a table of amounts to determine what mass of ammonia could be produced if the reaction went to completion.

1.72 Acrylonitrile is an important building block for synthetic fibres and plastics. Over 1.4 billion kilograms of acrylonitrile are produced in the United States each year. The compound is synthesized from propene in the following reaction:

$$2\,C_3H_6(g) + 2\,NH_3(g) + 3\,O_2(g) \longrightarrow 2\,C_3H_3N(l) + 6\,H_2O(g)$$



Propene



Acrylonitrile

Use a table of amounts to determine how many kilograms of acrylonitrile can be prepared from 1.50×10^3 kg of propene, 6.80×10^2 kg of ammonia, and 1.92×10^3 kg of oxygen.

1.73 For each of the reactions given in Problem 1.49, consider an industrial process that starts with 1.00 tonne (1000 kg) of each reactant. Construct a table of amounts, identify the limiting reactant, and determine the maximum mass of each product that could be produced.

1.74 For each of the reactions given in Problem 1.50, consider an industrial process that starts with 7.50×10^3 kg of each reactant. Construct a table of amounts, identify the limiting reactant, and determine the maximum mass of each product that could be produced.

1.75 Elemental phosphorus, P_4, reacts vigorously with oxygen to give P_4O_{10}. Use a table of amounts to determine how much P_4O_{10} can be prepared from 3.75 g of P_4 and 6.55 g of O_2, and how much of the excess reactant will remain at the end of the reaction.

1.76 The following unbalanced reaction is called the *thermite reaction*. It releases tremendous amounts of energy and is sometimes used to generate heat for welding:

$$Al + Fe_3O_4 \longrightarrow Fe + Al_2O_3$$



Courtesy of Patrick Watson

Use a table of amounts to determine the masses of all substances present after the reaction if 2.00×10^2 g of Al and 7.00×10^2 g of Fe_3O_4 react to completion.

ADDITIONAL PAIRED PROBLEMS

1.77 Draw a molecular picture that represents a homogeneous mixture of molecular fluorine and molecular chlorine in a 3:1 ratio of fluorine to chlorine. Both are gases, and both are diatomic molecules. The drawing should include at least 20 molecules. Use Figure 1-4 as a guide.

1.78 Gasoline does not dissolve in water. Instead, molecules of gasoline float on the surface of water. Draw a molecular picture that shows a heterogeneous mixture of gasoline and water. Draw enough molecules of each to show the structure of the mixture clearly. Use circles for water molecules and ovals for gasoline molecules.

1.79 Ethylene glycol is used as antifreeze in car radiators. The freezing temperature of ethylene glycol is –11.5 °C. Convert this freezing temperature to kelvins (K).

1.80 The melting point of silver metal is 1235 K. Convert this melting point to degrees Celsius (°C).

1.81 Perform the following calculations and report your answers with the correct number of significant figures:

(a) $$\frac{(6.531\times10^{13})(6.02\times10^{23})}{(435)(2.000)}$$

(b) $$\frac{4.476+(3.44)(5.6223)+5.666}{(4.3)(7\times10^{4})}$$

1.82 Perform the following calculations, and report your answers with the correct number of significant figures:

(a) $$\frac{(3.141\,59)(4.599\times10^{6})-(1.12\times10^{7})}{(4.756\times10^{8})+(3.67\times10^{4})}$$

(b) $$\frac{(6.577\times10^{-6})+0.003\,69+(8.234\times10^{-4})}{(0.000\,256\,7)+(6.9377\times10^{-5})}$$

1.83 Draw appropriately scaled and coloured molecular pictures of each of the following molecules: HCN (linear); H_2O (bent, O in the middle); CO; and NNO (linear). Use Figure 1-4 as a guideline.

1.84 Draw appropriately scaled and coloured molecular pictures of each of the following molecules: CO_2 (linear, C in the middle); ClO_2 (bent, Cl in the middle); ClBr; and HOCl (linear, O in the middle). Use Figure 1-4 as a guide.

1.85 Calculate the mass of lithium that has the same number of atoms as 5.75 g of platinum.

1.86 Calculate the mass of neon that has the same number of atoms as 37.5 mg of xenon.

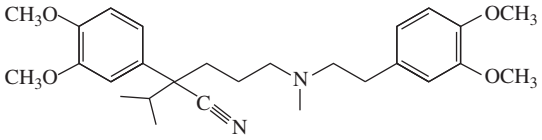
1.87 The chemical formula for aluminum sulphate is $Al_2(SO_4)_3$. (a) Compute its molar mass. (b) Compute the number of moles contained in 25.0 g of this compound. (c) Determine its percent composition. (d) Determine the mass of this compound that contains 1.00 mole of O.

1.88 The chemical formula for nickel sulphate hexahydrate is $NiSO_4 \cdot 6H_2O$. (a) Compute its molar mass. (b) Compute the number of moles contained in 25.0 g of this compound. (c) Determine its percent composition. (d) Determine the mass of this compound that contains 1.00 mole of O.

1.89 Seawater at 25 °C contains 8.3 mg/L of oxygen. What molarity is this?

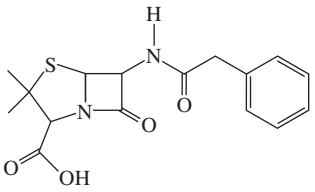
1.90 The oceans contain 1.3 mg of fluoride ions per litre. What molarity is this?

1.91 Heart disease causes 31% of the deaths in Canada. However, the death rate from heart disease has dropped significantly in recent years partly because of the development of new drugs for heart therapy by chemists working in the pharmaceutical industry. One of these drugs is verapamil, used for the treatment of arrhythmia, angina, and hypertension. A tablet contains 120.0 mg of verapamil. Determine the following quantities: (a) the molar mass of verapamil; (b) the number of moles of verapamil in one tablet; and (c) the number of nitrogen atoms in one tablet.



Verapamil
$C_{27}H_{38}O_4N_2$

1.92 Penicillin was discovered in 1928 by Alexander Fleming, who was a bacteriologist at the University of London. This molecule was originally isolated from a mould that contaminated some of Fleming's experiments. Penicillin destroys bacterial cells without harming animal cells, so it has been used as an antibiotic and has saved countless lives. The molecular formula of penicillin G is $C_{16}H_{18}N_2O_4S$. (a) Calculate the molar mass of penicillin G. (b) How many moles of penicillin are in a 50.0 mg sample? (c) How many carbon atoms are in a 50.0 mg sample? (d) What is the mass of sulphur in a 75.0 mg sample?

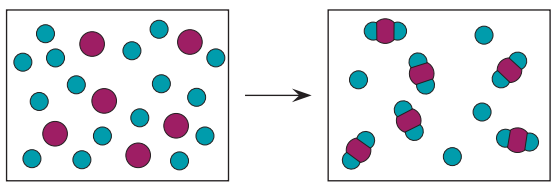


Penicillin G
$C_{16}H_{18}N_2O_4S$

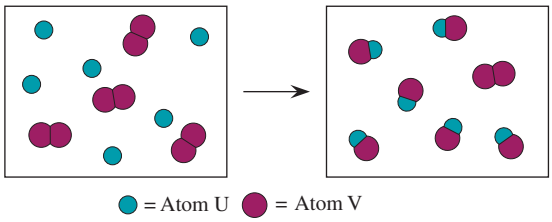
1.93 Although xenon is a "noble gas," it reacts with fluorine and oxygen. When xenon gas reacts with F_2 gas, XeF_4 is one of the products. If 5.00 g of xenon reacts with excess fluorine and generates 4.00 g of XeF_4, what is the percent yield and how much xenon remains unreacted, assuming no other product forms?

1.94 Hydrogen fluoride is produced industrially by the action of sulphuric acid on CaF_2. If 365 kg of CaF_2 is treated with excess sulphuric acid and 155 kg of HF is produced, what is the percent yield and how much CaF_2 remains unreacted, assuming no other fluorine-containing product forms?

1.95 The reaction of atom X with atom Y is represented in the following diagram. Write the balanced equation for this reaction.



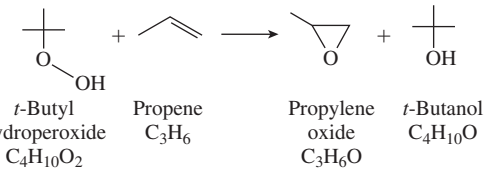
1.96 The reaction of element U with element V is represented in the diagram shown below. Write the balanced equation for this reaction.



1.97 A 4.6 g sample of ethanol (C_2H_5OH) is burned completely in air. (a) How many moles of H_2O are formed? (b) How many molecules? (c) How many grams?

1.98 An automobile engine completely burns 3.5 g of octane (C_8H_{18}). (a) How many moles of CO_2 are formed? (b) How many molecules? (c) How many grams?

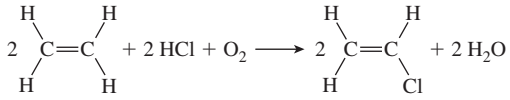
1.99 Propylene oxide is used primarily in the synthesis of poly(propylene glycol), a polymer used in the manufacture of automobile seats, bedding, and carpets. Around 2 billion kilograms of this compound are produced annually in the United States. Propylene oxide is produced by the following reaction:



t-Butyl hydroperoxide $C_4H_{10}O_2$ | Propene C_3H_6 | Propylene oxide C_3H_6O | *t*-Butanol $C_4H_{10}O$

(a) How many kilograms of propylene oxide can be prepared from 75 kg of *t*-butyl hydroperoxide? (b) What mass of propene will be required for the synthesis?

1.100 Vinyl chloride, one of the top 20 industrial compounds, is used primarily to make the polymer poly(vinyl chloride), better known as PVC. This versatile polymer is used to make piping, siding, gutters, floor tile, clothing, and toys. Vinyl chloride is made by oxychlorination of ethylene (C_2H_4). The overall balanced equation follows:



Ethylene | Vinyl chloride

Annual production of vinyl chloride is about 6.8 billion kilograms. What minimum masses (in billions of kilograms) of C_2H_4 and HCl are required to produce this much vinyl chloride?

ADDITIONAL PROBLEMS

1.101 The diameter of a chlorine atom is 2.00×10^2 pm. How many chlorine atoms lined up end to end would form a line 2.54 cm long?

1.102 The day is defined to contain exactly 24 hours of 60 minutes each, and each minute contains exactly 60 seconds. The year is 365.24 days long. How many seconds are there in a century? Express your answer in scientific notation with the correct number of significant figures.

1.103 How many atoms of chromium-53 are there in a 0.123 g sample of chromium metal? (The isotopic abundance of ^{53}Cr is 9.5%.)

1.104 Solution A is prepared by dissolving 90.0 g of Na_3PO_4 in enough water to make 1.5 L of solution. Solution B is 2.5 L of 0.705 M Na_2SO_4. (a) What is the molar concentration of Na_3PO_4 in Solution A? (b) How many millilitres of Solution A will give 2.50 g of Na_3PO_4? (c) A 50.0 mL sample of Solution B is mixed with a 75.00 mL sample of Solution A. Calculate the concentration of Na ions in the final solution.

1.105 The seventh-ranked industrial chemical in U.S. production is phosphoric acid, whose chemical formula is H_3PO_4. One method of manufacture starts with elemental phosphorus, which is burned in air; the product of this reaction then reacts with water to give the final product. In 1995, the United States manufactured 2.619×10^8 pounds of phosphoric acid. How many moles is this? If 15% of this material was made by burning elemental phosphorus, how many moles and how many kilograms of phosphorus were consumed?

1.106 Police officers confiscate a packet of white powder that they believe contains heroin. Purification by a forensic chemist yields a 38.70 mg sample for combustion analysis. This sample gives 97.46 mg of CO_2 and 20.81 mg of H_2O. A second sample is analyzed for its nitrogen content, which is 3.8%. Show by calculations whether these data are consistent with the formula for heroin, $C_{21}H_{22}NO_5$.

1.107 In the Haber synthesis of ammonia, N_2 and H_2 react at high temperature, but they never react completely. In a typical reaction, 24.0 kg of H_2 and 84.0 kg of N_2 react to produce 68 kg of NH_3. Using a table of amounts, find the theoretical yield, the percent yield, and the masses of H_2 and N_2 that remain unreacted, assuming that no other products form.

1.108 One set of reactants for rocket fuel is hydrazine and hydrogen peroxide, which react vigorously when mixed:

$$N_2H_4(l) + H_2O_2(l) \longrightarrow N_2(g) + H_2O(g) \text{ (unbalanced)}$$

The density of liquid hydrazine is 1.44 g/mL, and that of hydrogen peroxide is 1.01 g/mL. What volume ratio of these two liquids should be used if both fuels are to be used up at the same time?

1.109 The largest single use of sulphuric acid is for the production of phosphate fertilizers. The acid reacts with calcium phosphate in a 2:1 mole ratio to give calcium sulphate and calcium dihydrogen phosphate. The mixture is crushed and spread on fields, where the salts dissolve in rain water. (Calcium phosphate, commonly found in phosphate rock, is too insoluble to be a direct source of phosphate for plants.) (a) Write a balanced equation for the reaction of sulphuric acid with calcium phosphate. (b) How many kilograms each of sulphuric acid and calcium phosphate are required to produce 50.0 kg of the calcium sulphate–dihydrogen phosphate mixture? (c) How many moles of phosphate ion will this mixture provide?

1.110 The CO_2 exhaled by astronauts must be removed from the spacecraft atmosphere. One way to do this is with solid LiOH:

$$CO_2(g) + 2\ LiOH(s) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

The CO_2 output of an astronaut is about 1.0 kg/day. What is the minimum mass of LiOH required for a six-day space shuttle flight involving five astronauts?

INTEGRATIVE AND ADVANCED PROBLEMS

1.111 On average, a sample of human blood with dimensions of 0.1 mm × 0.1 mm × 0.1 mm contains 6.0×10^3 red blood cells. The volume of blood in a typical adult is about 5 L. How many red blood cells are there in an adult?

1.112 An object will float on water, even though it is made of a material that is denser than water, if its net density is less than 1. Thus, a hollow sphere with a large enough interior volume will float, because air has a very low density (1.189 g/L at room temperature and pressure). Determine the volume percentage that the interior of a sphere must have to float if it is made of aluminum. If the sphere has an interior volume of 1.50 L, how thin must its metal shell be? (*Hint:* for a sphere, $V = \frac{4}{3}\pi r^3$ and $A = 4\pi r^2$.) Repeat your calculations for copper.

1.113 Consult the *CRC Handbook of Chemistry and Physics* and compile two lists, one of the Z values of the elements that have just one stable isotope and the other of the Z values of the elements that have four or more stable isotopes. Determine if there are any patterns in the atomic number values of the elements occurring in each of these lists.

1.114 Aluminum sulphate is used in the manufacture of paper and in the water purification industry. In the solid state, aluminum sulphate is a hydrate. The formula is $Al_2(SO_4)_3 \cdot 18H_2O$. (a) How many grams of sulphur are there in 0.570 moles of solid aluminum sulphate? (b) How many water molecules are there in a 5.1 g sample of solid aluminum sulphate? (c) How many moles of sulphate ions are there in a sample of solid aluminum sulphate that contains 12.5 moles of oxygen atoms? (d) An aqueous solution of aluminum sulphate contains 1.25% by mass aluminum and has a density of 1.05 g/mL. What is the molarity of aluminum ions in the solution?

1.115 CF_3CH_2F (HFC-134a) has replaced chlorofluorocarbon compounds for use as refrigerants. HFC-134a is produced from trichloroethylene by the following reactions:

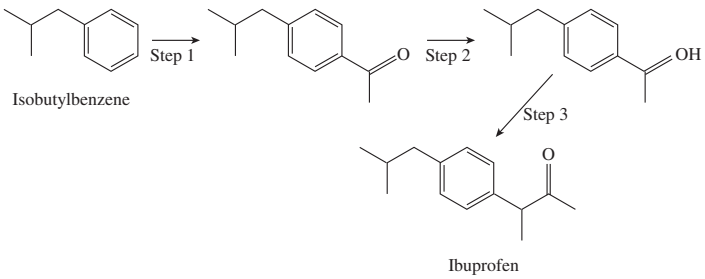
$$CCl_2{=}CHCl + 3\ HF \xrightarrow{\text{catalyst}} CF_3CH_2Cl + 2\ HCl$$

$$CF_3CH_2Cl + HF \xrightarrow{\text{catalyst}} CF_3CH_2F + HCl$$

In 1999, over 1.3 million tonnes of HFC-134a were produced, and production is increasing at a rate of nearly 20% annually. Manufacturers sell HFC-134a for $5.50/kg. Assume that you are a chemist working for a company that sold 65 million kilograms of HFC-134a last year. (a) If the synthesis had a reaction yield of 47%, how much trichloroethylene was used to produce the HFC-134a? (b) You have developed a new catalyst that improves the yield of the synthesis of HFC-134a from 47% to 53%. If the company uses the same mass of trichloroethylene for next year's production, how many kilograms of HFC-134a can the company produce if they use your new process? (c) Assuming that the selling price of HFC-134a does not change, how much more money will the company make next year because of your work?

1.116 Uranium tetrafluoride (UF_4) is an intermediate in the production of uranium hexafluoride (UF_6). One method of manufacturing UF_4 consists of reacting UO_2 with HF according to the (unbalanced) reaction $UO_2 + HF \longrightarrow UF_4 + H_2O$. Assuming a yield of 99%, calculate the volume of 31.8 M HF required (per day) to manufacture 25.0 tonnes of UF4 (per day).

1.117 Ibuprofen, a popular analgesic, can be synthesized by a three-step process:



If the yields of Steps 1, 2, and 3 are 91%, 90%, and 85%, respectively, calculate the mass (in kg) of ibuprofen that can be synthesized from 100 kg of the starting material, isobutylbenzene, and unlimited amounts of the reagents (which are not shown above).

ANSWERS TO SECTION EXERCISES

1.1.1 Ce, Cs, Cu, Ca, and C

1.1.2 Zirconium, nickel, tin, tungsten, selenium, beryllium, and gold

1.1.3 (a) H_2O_2; (b) SO_2; (c) N_2O; (d) $C_2H_4O_2$; and (e) C_2H_4

1.2.1 4.63×10^{-7} m^3, 1.7935×10^7 m, and 9.36×10^8 s

1.2.2 3.0×10^8 m/s and 9.5×10^{15} m

1.2.3 70.3 kg, 110. m, 312 K, and $3.155\,693 \times 10^7$ s

1.3.1 4.8×10^{-3} m^3

1.3.2 1.3 g/cm^3

1.3.3 6.42×10^7 tonne/y

1.4.1 32.06 g/mol

1.4.2 (a) 2.21×10^{23} atoms; (b) 3.21×10^{22} atoms; and (c) 7.80×10^{21} atoms

1.4.3 1.46×10^{-14} mol and 8.79×10^9 atoms

1.5.1 (a) 129.91 g/mol; (b) 184.04 g/mol; (c) 142.98 g/mol; (d) 261.34 g/mol; and (e) 504.82 g/mol

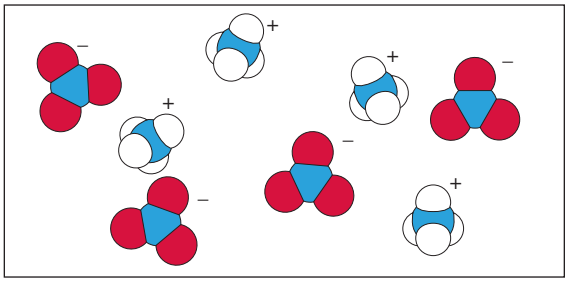
1.5.2 (a) 155.16 g/mol; (b) 121.16 g/mol; (c) 132.12 g/mol

1.5.3 (a) 9.86×10^{-5} mol; (b) 5.94×10^{19} molecules; (c) 1.78×10^{20} atoms; (d) 2.97×10^{20} atoms; (e) 20.5 mg; and (f) 2.79×10^{21} atoms

1.6.1 $[NH_4^+] = 0.31$ M, $[NO_3^-] = 0.31$ M

1.6.2 1.2 mL

1.6.3



1.7.1 $2\ C_8H_{18} + 25\ O_2 \longrightarrow 16\ CO_2 + 18\ H_2O$

1.7.2 $N_2 + O_2 \longrightarrow 2\ NO$; $2\ NO + O_2 \longrightarrow 2\ NO_2$

1.7.3 $2\ C_3H_6 + 2\ NH_3 + 3\ O_2 \longrightarrow 2\ C_3H_3N + 6\ H_2O$

1.8.1 (a) 333 g PCl_3 can be prepared; (b) 258 g Cl_2 is consumed.

1.8.2 0.737 g

1.8.3 2.80×10^3 g O_2 is required; 1.14×10^3 g H_2O and 2.47×10^3 g CO_2 are produced.

1.9.1 Theoretical yield is 1.96 g, and percent yield is 94.0%. Possible reasons are that not all the $KClO_3$ decomposes or that some O_2 escapes and is not collected.

1.9.2 Start with 5.8 tonnes of sulphur.

1.9.3 66 g salicylic acid is required.

1.10.1 The maximum mass is 793 kg HCN, and 29 kg CH_4 is left.

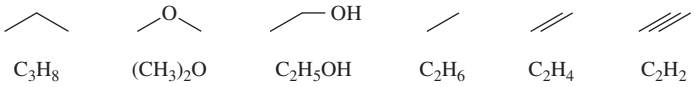
1.10.2 7.5 g

1.10.3 2.3×10^2 g

ANSWERS TO PRACTICE EXERCISES

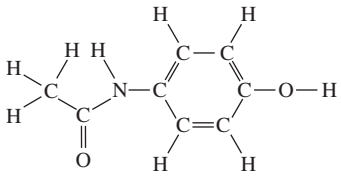
1.1 0.703 g/cm^3

1.2



C_3H_8 $(CH_3)_2O$ C_2H_5OH C_2H_6 C_2H_4 C_2H_2

1.3



Chemical formula: $C_8H_9NO_2$

1.4 330.36 g/mol

1.5 375.13 g/mol

1.6 13.9 moles and 8.36×10^{24} atoms

1.7 3.20 g

1.8 $[N_i^{2+}] = 0.255$ M; $[Cl^-] = 0.510$ M

1.9 11.8 mL

1.10 8.00×10^{-2} M

1.11 $2\ CH_3OH + 3\ O_2 \longrightarrow 2\ CO_2 + 4\ H_2O$

1.12 112 g formic acid is required.

1.13 186 kg of NH_3 is required as a starting material, and the reaction produces 295 kg of HCN.

1.14 66.1% yield

1.15 9.69×10^4 kg of NH_3 is required.

1.16

Reaction	**4 NH_3**	+	**5 O_2**	⟶	**4 NO**	+	**6 H_2O**
Start (mol)	5.00		7.00		0		0
Change (mol)	–5.00		–6.25		+5.00		+7.50
Final (mol)	0		0.75		5.00		7.50

1.17 Yield = 78.4%