TOPIC 1

Introduction: MATTER, ENERGY, AND MEASUREMENT

You probably learned most of this material in first-year chemistry. When you finish reviewing this topic, be sure you are able to

- Distinguish among elements, compounds, and mixtures
- Explain that separation of mixtures is based on their physical properties
- Use SI units for measurement
- Apply dimensional analysis and significant figures to calculations

Classifications of Matter

Section 1.2

Matter may be classified as either pure substances or mixtures.

Pure substances are either elements or compounds.

An **element** is a substance all of whose atoms contain the same number of protons.

A **compound** is a relatively stable combination of two or more chemically bonded elements in a specific ratio.

Mixtures consist of two or more substances.

Homogeneous mixtures are uniform throughout. Air, seawater, and a nickel coin (a mixture of copper and nickel metals called an alloy) are examples.

Heterogeneous mixtures vary in texture and appearance throughout the sample. Rocks, wood, polluted air, and muddy water are examples.

Properties of Matter

Section 1.3

Identify evidence of chemical and physical changes in matter. **TRA-1.A.**

Physical properties can be measured without changing the identity or composition of the substance. Physical properties include color, density, melting point, and hardness.

Chemical properties describe the way a substance changes (reacts) to form other substances. The flammability of gasoline is a chemical property because the gasoline reacts with oxygen to form carbon dioxide and water.

Intensive properties do not depend on the amount of substance in a sample. Temperature, density, and boiling point are intensive properties.

Extensive properties depend on the quantity of the sample. Energy content, mass, and volume are examples of extensive properties.

A **physical change** changes the appearance of a substance but does not change its composition. Changes of physical state, from solid to liquid or from liquid to gas, are examples.

A **chemical change** (also called a chemical reaction) transforms a substance into a different substance or substances. When the chief component of natural gas, methane, burns in air, the methane and the oxygen from the air are transformed into carbon dioxide and water.

Differences in physical properties, as determined by intermolecular forces (see Topic 11), are used to separate the components of mixtures.

Filtration separates a solid from a liquid.

Distillation separates substances based on their differences in boiling point.

Chromatography is a technique that separates substances based on their differences in intermolecular forces and their abilities to dissolve in various solvents. Chromatography is discussed in more detail in Topic 13.

Section 1.4 The Nature of Energy

Energy is the capacity to do work or to transfer heat.

Heat is the energy transferred from one object to another because of a difference in temperature.

Kinetic energy is the energy of motion. The magnitude of the kinetic energy (KE) of a particle depends on its mass (m) and velocity (v): $KE = \frac{1}{2} mv^2$.

Potential energy is "stored energy." One form of potential energy is chemical energy, which arises from chemical bonds, the attractions and repulsions between atoms. More details about chemical energy are found in Topics, 5, 8, and 19.

Section 1.5 Units of Measurement

Chemists often use preferred units called SI units after the French *Système International d' Unités*. Table 1.1 lists the base SI units and their symbols.

Table 1.1 SI base units.

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s or sec
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A or amp
Luminous intensity	Candela	cd

Table 1.2 lists metric prefixes that indicate decimal fractions or multiples of various units.

Table 1.2 Selected prefixes used in the metric system.

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) $= 1 \times 10^9$ m
Mega	M	10^6	1 megameter (Mm) $= 1 \times 10^6$ m
Kilo	k	10^3	1 kilometer (km) $= 1 \times 10^3$ m
Deci	d	10^{-1}	1 decimeter (dm) $= 0.1$ m
Centi	c	10^{-2}	1 centimeter (cm) $= 0.01$ m
Milli	m	10^{-3}	1 millimeter (mm) $= 0.001$ m
Micro	μ*	10^{-6}	1 micrometer (μm) $= 1 \times 10^{-6}$ m
Nano	n	10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
Pico	p	10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m
Femto	f	10^{-15}	1 femtometer (fm) $= 1 \times 10^{-15}$ m

*This is the Greek letter mu (pronounced "mew").

Temperature is commonly measured using either the Celsius scale or the Kelvin scale.

$$K = °C + 273$$

Derived units are units derived from SI base units.

Volume, the space occupied by a substance, is commonly measured in cubic meters, m^3, or cubic centimeters, cm^3. A non-SI unit commonly used by chemists is the liter, L. One liter is the volume of a cube measuring exactly 10 cm on a side.

$$1\ L = 1000\ cm^3 = 1000\ mL$$

$$1\ cm^3 = 1\ mL$$

Density, the amount of matter packed into a given space, is often measured in g/cm^3 for liquids and solids and g/L for gases.

$$\text{Density} = \text{mass}/\text{volume}$$

Units of Energy

Heat is the energy transferred from one object to another because of a difference in temperature.

A **calorie** is an informal but still used unit for heat energy. One calorie is the amount of energy required to raise the temperature of one gram of water by one degree Celsius. The large Calorie (spelled with a capital C) is used to measure food energy. 1 Cal = 1000 cal.

A **joule** is the SI unit of energy. One calorie is equal to 4.184 joules. 1 cal = 4.184 J.

Section 1.6 Uncertainty in Measurement and Significant Figures

Exact numbers are known exactly and are usually defined or counted. "One liter equals 1000 cm^3" describes a defined number. "There are 32 students in this class" describes a counted number.

Inexact numbers have some degree of error or uncertainty associated with them. All measured numbers are inexact.

Measured numbers are generally reported in such a way that only the last digit is uncertain. **Significant figures** are all digits of a measured number, including the uncertain one.

Zeros in measured numbers are either significant or merely there to locate the decimal place. The following guidelines describe when zeros are significant:

1. Zeros between nonzero digits are always significant.
2. Zeros at the beginning of a number are never significant.
3. Zeros at the end of a number are significant only when the number contains a decimal point.

In calculations involving measured quantities, the least certain measurement limits the certainty of the calculated quantity and determines the number of significant figures in the final answer. Exact numbers do not limit the certainty.

For multiplication and division, the number of significant figures in the answer is determined by the measurement with the least number of significant figures.

For addition and subtraction, the result has the same number of decimal places as the measurement with the least number of decimal places.

Common misconception: The guidelines for determining the number of significant figures in a result obtained by carrying measured quantities through calculations do not always give the correct number of significant figures. This is principally why the AP test usually allows full credit for answers reported to plus or minus one significant figure. In this book, numerical answers are usually rounded to three significant figures.

Dimensional Analysis Section 1.7

Dimensional analysis is a way of converting a written question into an algebraic equation, followed by manipulating factors until the unit of the known quantity is converted into the unit of the unknown quantity.

Example:

How many microseconds are there in one year?

Solution:

The algebraic equivalent to the given question is:

$x\ \mu\text{s} = 1\ \text{yr}.$

Now multiply the right side of the equation by what is known about a year in such a way that the unit of years cancels, giving another unit. Continue to do this until the result has the unit of microseconds, μs.

$x\ \mu\text{s} = 1\ \text{yr}\ (365\ \text{days/yr})\ (24\ \text{h/day})\ (60\ \text{min/h})\ (60\ \text{s/min})$
$(10^6\ \mu\text{s/s}) = 3.15 \times 10^{13} \mu\text{s}$

TOPIC

2 ATOMS, MOLECULES, AND IONS

Except for the information on mass spectrometry, you may have learned much of this material in first-year chemistry. Although nomenclature is not specifically tested on the AP exam, it is helpful to review the names and formulas of ionic and molecular compounds. When you finish reviewing this topic, be sure you are able to:

- Describe the basic structure of the atom using the terms protons, neutrons, electrons, nucleus, electron cloud, atomic number, mass number, isotope, and atomic mass
- Describe the key experimental evidence that led scientists to understand the modern atom
- Cite specific experimental evidence that supports various atomic models and evidence that is contradictory
- Identify elements and individual isotopes using data from mass spectra
- Justify that, in a pure sample of a compound, the ratio of the masses of its constituent elements is always the same
- Distinguish between a molecular and an ionic compound and a molecular and an empirical formula
- Name common ions and ionic and molecular compounds and write their formulas
- Recognize the structures and names of alkanes, alcohols, and carboxylic acids

Section 2.1 The Atomic Theory of Matter

Scientists formulate models based on experimental observations. They then use these scientific models to make predictions and test the predictions with experiments. When new data are inconsistent with a model's predictions, that model must be revised or replaced. The development and refinement of the atomic theory of matter illustrate this fundamental process of science.

Dalton's Atomic Theory

Nineteenth-century English chemist John Dalton proposed that matter is composed of tiny indivisible particles called atoms. Figure 2.1 summarizes Dalton's basic ideas.

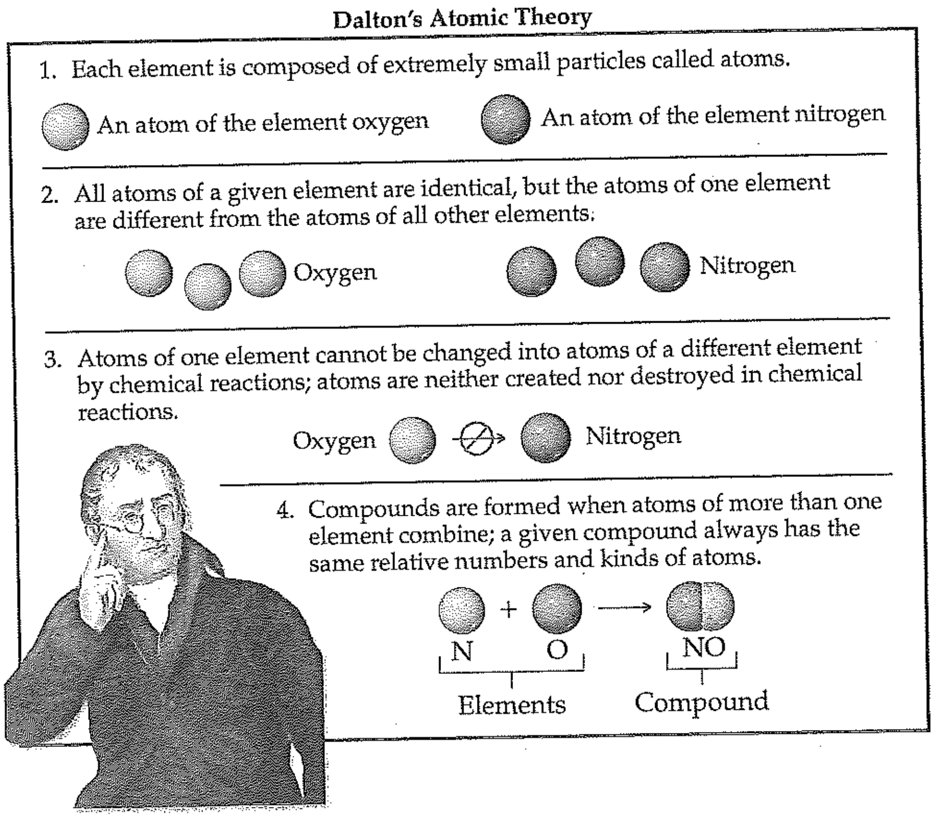


Figure 2.1 Summary of Dalton's atomic theory.

Dalton's theory explains three fundamental laws:

The **law of constant composition** states that in a given compound, the relative numbers and kinds of atoms are constant.

The **law of conservation of mass** states that the mass of materials does not change in a chemical reaction.

The **law of multiple proportions** states that when two elements combine to form a compound, their masses always exist in a ratio of small whole numbers.

Today we accept three of Dalton's four ideas expressed in Figure 2.1. Only the second idea is incorrect. We now know that atoms of a given element are not identical. Evidence from mass spectra (described later in this topic) clearly demonstrates that atoms of the same element can be composed of different isotopes, each having different masses and different numbers of neutrons.

The Thomson Model

J. J. Thomson showed that cathode rays are streams of negative particles, and he is credited with discovering the electron. Thomson postulated that all atoms contain electrons, contradicting Dalton's assumption that atoms are indivisible. Thomson proposed that an atom consisted of a uniform positive sphere in which electrons were embedded like plums in a pudding (Figure 2.2).

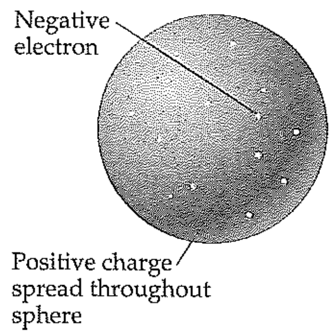


Figure 2.2 J. J. Thomson's plum-pudding model of the atom. Ernest Rutherford proved this model wrong.

The Rutherford Experiment

Ernest Rutherford's gold-foil experiment, illustrated in Figure 2.3, showed that atoms are mostly empty space having a tiny dense nucleus. Rutherford's evidence is inconsistent with Thomson's assumption that atoms are solid particles. Rutherford proposed

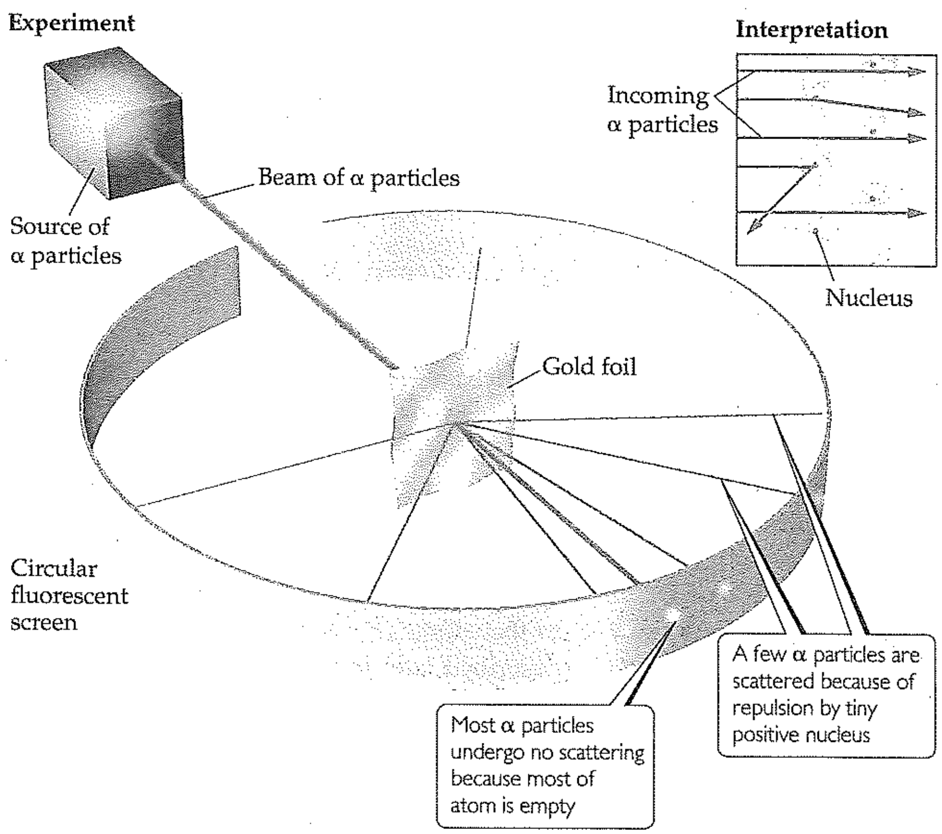


Figure 2.3 Rutherford's α-scattering experiment. When α particles pass through a gold foil, most pass through undeflected but some are scattered, a few at very large angles. According to the plum-pudding model of the atom, the particles should experience only very minor deflections. The nuclear model of the atom explains why a few α particles are deflected at large angles. For clarity, the nuclear atom is shown here as a sphere, but remember that most of the space around the nucleus is empty except for the tiny electrons moving around.

that electrons circle the nucleus much as planets orbit the sun. However, he offered no explanation why an atom is stable. Classical physics predicts that orbiting electrons would lose energy and fall into the nucleus but this does not happen.

Niels Bohr explained why atoms are stable. He postulated that the lines of the atomic emission spectrum of hydrogen (described in Topic 6) represent transitions of electrons from one allowed energy state to another. Atoms are stable because their electrons occupy fixed energy states preventing orbital decay.

Your Turn 2.1

For each atomic model proposed by Dalton, Thomson, and Rutherford, cite at least one piece of experimental evidence that is inconsistent with that model.

The Modern View of Atomic Structure

Section 2.3

SPQ-1.B.

Explain the quantitative relationship between the mass spectrum of an element and the masses of the element's isotopes. See also Section 2.4.

In studying chemistry, it is often convenient to think of atoms as fundamental, indivisible units of matter. However, atoms are composed of three basic subatomic particles: **protons, electrons,** and **neutrons**.

Atoms consist of a tiny, dense, positively charged nucleus surrounded by a cloud of negative electrons.

The **nucleus** contains positively charged protons and neutral neutrons.

Atoms are electrically neutral because each atom contains the same number of protons as electrons.

Atoms can gain electrons to form negatively charged ions called **anions** or they can lose electrons to become positively charged ions called **cations**.

The **atomic number** is the number of protons in the nucleus.

An **element** is a substance containing atoms that have the same number of protons. Each element is defined by its atomic number.

The **mass number** is the number of protons and neutrons in the nucleus of an atom.

Isotopes are atoms containing the same number of protons but different mass numbers.

Symbols are often used to denote various elements and to distinguish isotopes. For example, the isotope of carbon containing six protons and six neutrons is designated like this:

12 = mass number = number of protons plus neutrons

$^{12}_{6}C$

6 = atomic number = number of protons

Because carbon is the element that always contains six protons, often the atomic number designation is omitted and the following symbols all designate the same isotope of carbon:

$$^{12}_{6}C = {}^{12}C = \text{carbon-12} = \text{C-12} \text{ (Read all these symbols as "carbon twelve.")}$$

Carbon has several isotopes, each having six protons and a different mass number. The respective number of neutrons of an isotope is calculated by subtracting the atomic number from the mass number. Here are symbols for various isotopes of carbon showing the number of neutrons present in each isotope. Isotopes are distinguished by their mass numbers.

$^{11}_{6}C$	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
5 neutrons	6 neutrons	7 neutrons	8 neutrons
C-11	C-12	C-13	C-14

Section 2.4 Atomic Weights

The **atomic mass unit, amu**, is a convenient way to express the relative masses of tiny atoms. One amu equals 1.66054×10^{-24} g. However, it is more useful to compare the masses of atoms to the mass of one carbon-12 isotope. One carbon-12 atom has a defined mass of exactly 12 amu.

Atomic mass, sometimes called atomic weight, is the weighted average mass of all the isotopes of an element based on the abundance of each isotope found on Earth. Atomic masses are expressed in amu. All atomic masses reported on the periodic table are based on the carbon-12 standard. For example, the atomic mass of magnesium is 24.3050 amu. This means that the average mass of all the magnesium isotopes is a little more than twice the mass of a carbon-12 atom.

Mass Spectrometry

Mass spectrometry is a method that measures precise masses of atoms and molecules.

A **mass spectrometer** is an instrument that bombards a sample with high-energy electrons. It converts the sample to charged particles, which are accelerated and deflected in a magnetic field. The extent of deflection depends on the mass of the particle, thereby separating different particles according to their masses. The spectrometer detects the masses and relative abundances of the charged particles.

A **mass spectrum** is a graph of intensity of the detector signal versus particle atomic mass.

Figure 2.4a shows a schematic diagram of a mass spectrometer. Figure 2.4b shows a typical mass spectrum.

Mass spectrometers dramatically demonstrate the existence of isotopes, and they accurately measure the individual masses and relative abundances of these isotopes.

Mass spectrometry is the most accurate way to determine atomic masses.

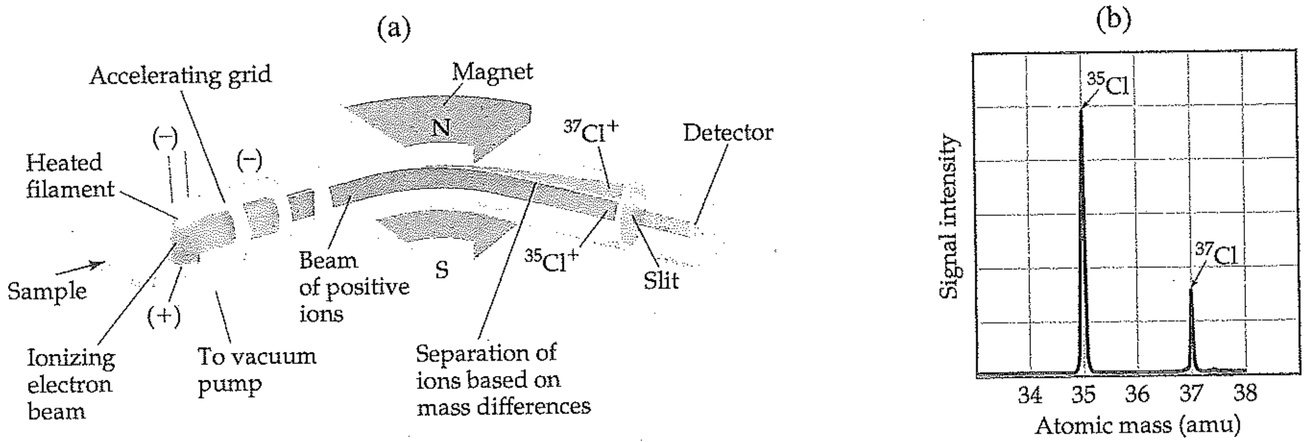


Figure 2.4 (a) A mass spectrometer. A sample of Cl atoms are introduced at the left and are ionized to Cl^{+} ions, which are then directed through a magnetic field. The paths of the ions of the two Cl isotopes diverge as they pass through the field. (b) Mass spectrum of atomic chlorine. The fractional abundances of the isotopes ^{35}Cl and ^{37}Cl are indicated by the relative signal intensities of the beams reaching the detector of the mass spectrometer.

Example:

Detailed analysis of data from the mass spectrum of chlorine shown in Figure 2.4(b) reveals that there are two different isotopes of chlorine. Cl-35 has a mass of 34.969 amu and a relative abundance of 75.77%. Cl-37 has a mass of 36.966 amu and a relative abundance of 24.23%. Calculate the atomic mass of chlorine.

Solution:

The atomic mass of an element is the weighted average of the masses of the isotopes. Multiply each individual mass by its relative abundance.

$$\textit{Average atomic mass} =$$
$$(\textit{mass of C-35})(\textit{abundance of Cl-35}) + (\textit{mass of Cl-37})(\textit{abundance of Cl-37})$$
$$= (34.969\ amu)(0.7577) + (36.966\ amu)(0.2423)$$
$$= 26.496 + 8.957 = 35.453 = 35.45\ amu$$

(This result compares favorably, within significant figures, with 35.453, the atomic mass of chlorine found on the periodic table.)

The method of radiocarbon dating of ancient artifacts uses a technique called accelerator mass spectrometry, AMS. The mass spectrometer measures the ratio of C-12 to C-14. Because C-14 is radioactive, it decays to N-14 at a known rate. The less carbon-14 an object contains, the older it is. The age of a sample is calculated from the measured C-12:C-14 ratio.

Besides measuring the masses of isotopes, mass spectrometry accurately measures the masses of molecules and provides a powerful method to identify them. The high-energy beam of electrons striking a molecule produces a "parent ion," which breaks into a collection of smaller pieces that are characteristic of the molecule. The resulting mass spectrum shows the molecular mass (mass of the parent ion) of the sample and a pattern of fragmented masses that is a characteristic "fingerprint" of the molecule.

Your Turn 2.2

Figure 2.5 shows the mass spectrum of lead. How many different isotopes of lead are represented in the figure? Justify your answer. Identify each isotope. Tell how many electrons, protons, and neutrons are contained in the atoms of each isotope of lead.

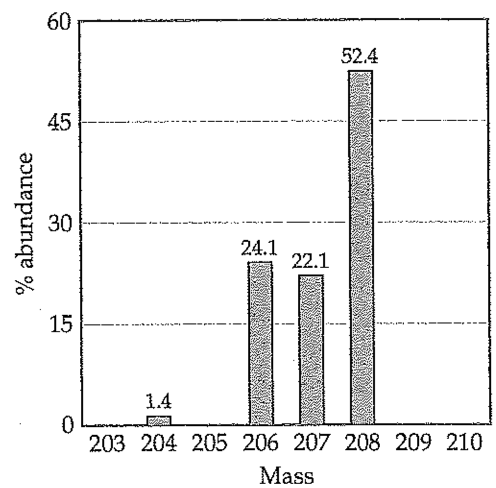


Figure 2.5 The mass spectrum of lead.

Section 2.5

The Periodic Table

The **periodic table** is an arrangement of elements in order of increasing atomic number, with elements having similar properties placed in vertical columns. The vertical columns are called **groups** or families and the horizontal rows are called **periods**. Figure 2.6 shows the periodic table with the symbol, atomic number, and atomic mass of each element. It also shows two commonly used numbering systems for the groups. Elements are classified as metals, nonmetals, and metalloids.

Table 2.1 shows the special names given to four element groups.

Table 2.1 Names given to four groups of elements on the periodic table.

Group Number	Name of Group
1 or 1A	alkali metals
2 or 2A	alkaline earth metals
17 or 7A	halogens
18 or 8A	noble gases

Main Group Representative Elements | Transition metals | Main Group Representative Elements

☐ Metals ☐ Metalloids ☐ Nonmetals

	1A[a] 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H** 1.00794																	2 **He** 4.002602
2	3 **Li** 6.941	4 **Be** 9.0121831											5 **B** 10.811	6 **C** 12.0107	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.99840316	10 **Ne** 20.1797
3	11 **Na** 22.989770	12 **Mg** 24.3050											13 **Al** 26.981538	14 **Si** 28.0855	15 **P** 30.973762	16 **S** 32.065	17 **Cl** 35.453	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.955908	22 **Ti** 47.867	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.938044	26 **Fe** 55.845	27 **Co** 58.933194	28 **Ni** 58.6934	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.64	33 **As** 74.92160	34 **Se** 78.971	35 **Br** 79.904	36 **Kr** 83.80
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.90584	40 **Zr** 91.224	41 **Nb** 92.90637	42 **Mo** 95.95	43 **Tc** [98]	44 **Ru** 101.07	45 **Rh** 102.90550	46 **Pd** 106.42	47 **Ag** 107.8682	48 **Cd** 112.414	49 **In** 114.818	50 **Sn** 118.710	51 **Sb** 121.760	52 **Te** 127.60	53 **I** 126.90447	54 **Xe** 131.293
6	55 **Cs** 132.905453	56 **Ba** 137.327	71 **Lu** 174.967	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.84	75 **Re** 186.207	76 **Os** 190.23	77 **Ir** 192.217	78 **Pt** 195.078	79 **Au** 196.966569	80 **Hg** 200.59	81 **Tl** 204.3833	82 **Pb** 207.2	83 **Bi** 208.98038	84 **Po** [208.98]	85 **At** [209.99]	86 **Rn** [222.02]
7	87 **Fr** [223.02]	88 **Ra** [226.03]	103 **Lr** [262.11]	104 **Rf** [267.1]	105 **Db** [268.1]	106 **Sg** [269.1]	107 **Bh** [270.1]	108 **Hs** [269.1]	109 **Mt** [278.2]	110 **Ds** [281.2]	111 **Rg** [282.2]	112 **Cn** [285.2]	113 **Nh** [286.2]	114 **Fl** [289.2]	115 **Mc** [289.2]	116 **Lv** [293.2]	117 **Ts** [293.2]	118 **Og** [294.2]

Lanthanide series	57 **La** 138.9055	58 **Ce** 140.116	59 **Pr** 140.90766	60 **Nd** 144.24	61 **Pm** [145]	62 **Sm** 150.36	63 **Eu** 151.964	64 **Gd** 157.25	65 **Tb** 158.92534	66 **Dy** 162.50	67 **Ho** 164.93033	68 **Er** 167.259	69 **Tm** 168.93422	70 **Yb** 173.04
Actinide series	89 **Ac** [227.03]	90 **Th** 232.0377	91 **Pa** 231.03588	92 **U** 238.02891	93 **Np** [237.05]	94 **Pu** [244.06]	95 **Am** [243.06]	96 **Cm** [247.07]	97 **Bk** [247.07]	98 **Cf** [251.08]	99 **Es** [252.08]	100 **Fm** [257.10]	101 **Md** [258.10]	102 **No** [259.10]

[a]The labels on top (1A, 2A, etc.) are common American usage. The labels below these (1, 2, etc.) are those recommended by the International Union of Pure and Applied Chemistry (IUPAC).

Atomic weights in brackets are for the longest-lived or most important isotope of radioactive elements.
Further information is available at http://www.webelements.com

Figure 2.6 The periodic table of the elements.

Molecules and Molecular Compounds Section 2.6

Explain the quantitative relationship between the elemental composition by mass and the empirical formula of a pure substance. See also Sections 3.4 and 3.5. **SPQ-2.A.**

Although the atom is the smallest representative particle of an element, most matter is composed of molecules or ions, which are combinations of atoms.

A **molecule** is an assembly of two or more atoms tightly bonded together. For example, a molecule that is made up of two atoms is called a diatomic molecule. Figure 2.7 shows the names, formulas, and pictorial representations of some simple molecules.

Seven elements normally occur as **diatomic molecules**. They are hydrogen, oxygen, nitrogen, fluorine, chlorine, bromine, and iodine. The formulas for these diatomic molecules are written like this:

$$H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2$$

The subscript displayed in each formula indicates that two atoms are present in each molecule.

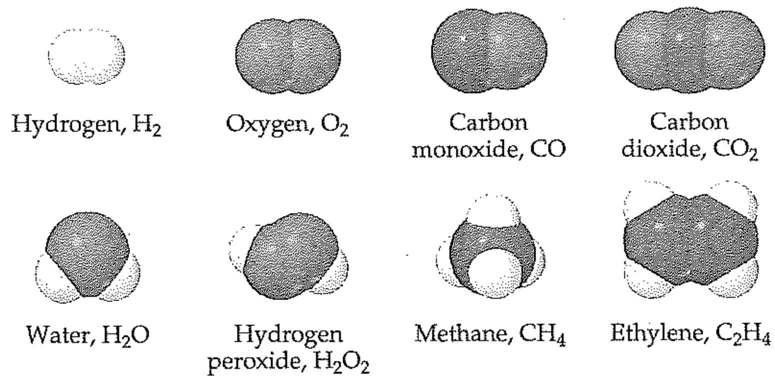


Figure 2.7 Molecular models. Notice how the chemical formulas of these simple molecules correspond to their compositions.

Common misconception: Chemists often use the name *oxygen* to refer to both O and O_2, even though the latter's official name is *dioxygen* to distinguish it from monatomic oxygen. For chemists, the correct species can easily be inferred by the context of the sentence. For example, the oxygen we breathe is O_2, and the oxygen in the water molecule is O. Most texts use the monatomic names for the diatomic elements. Pay close attention to the context in which these names are used to determine the exact meaning.

Your Turn 2.3

Tell which chemical form of chlorine is implied in these two sentences: (a) Chlorine is a toxic gas. (b) Common table salt contains the element chlorine.

Compounds are substances consisting of two or more different tightly bonded elements. Generally, there are two types of compounds: molecular compounds and ionic compounds.

Molecular compounds are composed of molecules and usually contain only nonmetals. A **molecular formula** indicates the actual number and type of atoms in the molecule and is the most often used formula for molecular compounds.

Figure 2.8 shows various ways chemists represent molecular compounds.

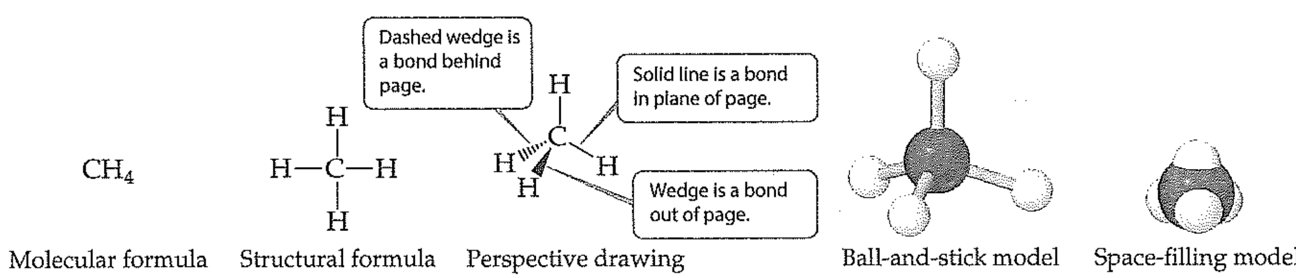


Figure 2.8 Different representations of the methane (CH_4) molecule. Structural formulas, perspective drawings, ball-and-stick models, and space-filling models correspond to the molecular formula, and each helps us visualize the three-dimensional arrangement of atoms.

Ions and Ionic Compounds

Section 2.7

Ions are charged particles composed of single atoms (in which case they are called **monatomic ions**) or aggregates of atoms (in which case they are called **polyatomic ions**). A **cation** is a positive ion, and an **anion** is a negative ion.

Ionic compounds are composed of ions and usually contain both metals and nonmetals.

An **empirical formula** gives only the relative number of atoms of each type in the compound. Empirical formulas are usually used for ionic compounds and sometimes used for molecular compounds.

Your Turn 2.4

a. Which formulas in Figure 2.7 represent molecular formulas? Which represent empirical formulas? Explain your reasoning. b. Why not use H, O, HO, and CH_2 for the formulas of hydrogen, oxygen, hydrogen peroxide, and ethylene, respectively?

Unlike molecular compounds, ionic compounds do not consist of discrete molecules. Ionic compounds are a collection of many ions arranged in a regular pattern, as shown in Figure 2.9. Rather than draw the ionic arrangement of sodium chloride, chemists use the much simpler empirical formula, NaCl, to represent the compound.

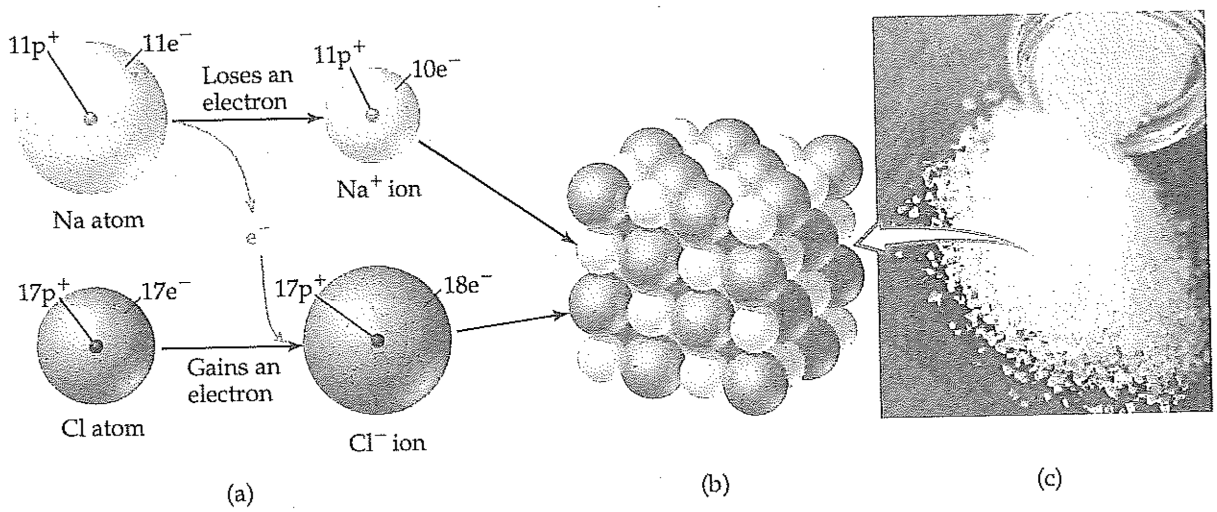


Figure 2.9 Formation of an ionic compound. (a) The transfer of an electron from a Na atom to a Cl atom leads to the formation of a Na^+ ion and a Cl^- ion. (b) Arrangement of these ions in solid sodium chloride, NaCl. (c) A sample of NaCl crystals.

Naming Inorganic Compounds

Section 2.8

Metal atoms can lose electrons to become monatomic cations.

Nonmetal atoms gain electrons to become monatomic anions.

The periodic table is useful in remembering the names and charges of monatomic cations and anions. Figure 2.10 shows the common names and charges of ions derived from elements on the left and right sides of the periodic table. Notice that the cations of the A groups carry a positive charge equal to the group number, and the anions carry a negative charge equal to the group number minus 8. Transition metals tend to form cations of more than one charge value.

Common misconception: The names and formulas of monatomic anions need not be memorized. Simply locate the atom on the periodic table, assign the charge based on the group number, and change the ending to -ide.

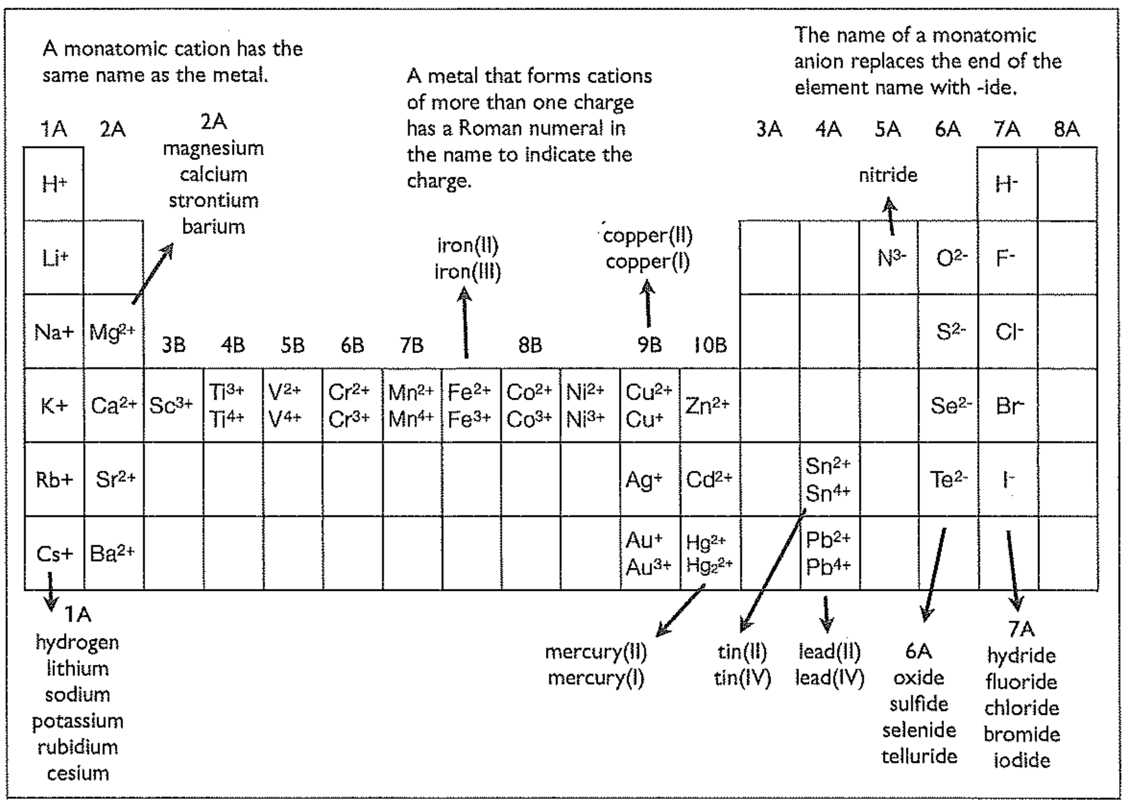


Figure 2.10 Charges of some monatomic cations and anions are consistent within groups on the periodic table. Notice that some transition metals form ions having more than one charge.

A few common diatomic anions also end in -ide.

OH^- hydroxide CN^- cyanide O_2^{2-} peroxide

Names of Polyatomic Cations

Polyatomic (containing more than one atom) cations formed from nonmetals end in -ium.

NH_4^+ = ammonium $CH_3NH_3^+$ = methylammonium
$(CH_3)_2NH_2^+$ = dimethylammonium

Polyatomic Oxyanions (Anions Containing Oxygen)

Some polyatomic anions end in -ate or -ite. These are called oxyanions because they contain oxygen. The ending -ate denotes the most common oxyanion of an element. The ending -ite refers to an oxyanion having the same charge but one fewer oxygen. Because there is no logical way to predict their formulas or charges, these must be memorized. Table 2.2 lists the names and formulas of some common oxyanions.

Some polyatomic oxyanions have a hydrogen ion attached. The word *hydrogen* or *dihydrogen* is used to indicate anions derived by adding H^+ to an oxyanion. An added hydrogen ion changes the charge by 1+.

Table 2.2 Names and formulas of some common oxyanions listed by charge.

Charge	**3–**	**2–**	**1–**
Name and formula	phosphate PO_4^{3-}	hydrogen phosphate HPO_4^{2-}	dihydrogen phosphate $H_2PO_4^-$
		sulfate SO_4^{2-}	hydrogen sulfate HSO_4^-
		carbonate CO_3^{2-}	hydrogen carbonate HCO_3^-
			nitrate NO_3^-
			acetate (also ethanoate) CH_3COO^-

Chlorine forms a series of four oxyanions. Prefixes and suffixes are used to distinguish them, as shown in Table 2.3.

Table 2.3 Formulas and names of halogen oxyanions.

ClO^-	hypochlorite	hypo- denotes one fewer oxygen than -ite
ClO_2^-	chlorite	-ite denotes one fewer oxygen than -ate
ClO_3^-	chlorate	-ate denotes one more oxygen than -ite
ClO_4^-	perchlorate	per- denotes one more oxygen than -ate

Names of Ionic Compounds

Names of ionic compounds consist of the cation name followed by the anion name, as shown in Table 2.4.

Table 2.4 How to write formulas and names of ionic compounds.

Write a formula: Combine the cation and anion in a ratio such that the charge of the compound is zero. Use a subscript to indicate the number of each anion and cation, and use parentheses when necessary for polyatomic ions.		**Name a compound:** Name the cation, then name the anion.	
Cation	Anion	Formula	Name
Li^+	F^-	LiF	lithium fluoride
Mg^{2+}	Br^-	$MgBr_2$	magnesium bromide
Fe^{3+}	O^{2-}	Fe_2O_3	iron(III) oxide *
Fe^{2+}	O^{2-}	FeO	iron(II) oxide *
Ca^{2+}	N^{3-}	Ca_3N_2	calcium nitride
Ba^{2+}	PO_4^{3-}	$Ba_3(PO_4)_2$	barium phosphate
NH_4^+	SO_4^{2-}	$(NH_4)_2SO_4$	ammonium sulfate

* The Roman numerals III and II refer to the 3+ and 2+ charges of the iron cations, respectively. Notice that the charge of each iron cation is inferred by the ratio in which iron combines with the 2− charge of the oxide ion.

Your Turn 2.5

Name the following ionic compounds: KI, $MgSO_4$, FeS, Al_2O_3, $Pb_3(PO_4)_2$.

Names and Formulas of Acids

The names and formulas of some common acids, an important class of hydrogen-containing compounds, are listed in Table 2.5. In all cases, the formulas of acids are composed of one or more hydrogen ions added to a common monatomic anion or oxyanion. Acids of monatomic ions are called binary acids, and acids of oxyanions are called oxyacids.

Table 2.5 Names and formulas of some common acids.

Binary Acids		**Oxyacids**	
HF	hydrofluoric acid	HNO_3	nitric acid
HCl	hydrochloric acid	H_2SO_4	sulfuric acid
HBr	hydrobromic acid	H_3PO_4	phosphoric acid
HI	hydroiodic acid	CH_3COOH	acetic acid
H_2S	hydrosulfuric acid	H_2CO_3	carbonic acid
H_2Se	hydroselenic acid	HNO_2	nitrous acid
		H_2SO_3	sulfurous acid
		$HClO_4$	perchloric acid
		$HClO_3$	chloric acid
		$HClO_2$	chlorous acid
		HClO	hypochlorous acid

To name binary acids, replace the -ide ending of the anion with -ic acid, and add the prefix hydro-.

_____ide becomes hydro______ic acid.

Bromide, Br^-, becomes hydrobromic acid, HBr.

To name oxyacids, replace the -ate ending of the oxyanion with -ic acid or the -ite ending of the oxyanion with -ous acid.

_____ate becomes _____ic acid.

Nitrate, NO_3^-, becomes nitric acid, HNO_3.

_____ite becomes _____ous acid.

Hypochlorite, ClO^-, becomes hypochlorous acid, HClO.

Some common exceptions:

Phosphate, PO_4^{3-}, becomes phosPHORic acid, H_3PO_4.

Sulfate, SO_4^{2-}, becomes sulfURic acid, H_2SO_4.

Names of Binary Molecular Compounds

A binary molecular compound contains two nonmetals. The rules for naming binary molecular compounds are the following:

1. Name the first element.
2. Name the second element and change the ending to -ide.
3. Use prefixes to denote how many of each element are present in the formula. The prefixes are shown in Table 2.6. (Notice that a prefix ending in a- drops the a- when it precedes -oxide.)

Table 2.6 Prefixes and example names for binary molecular compounds formed between nonmetals.

Prefix	Meaning	Prefix	Meaning	Formula	Name
mono-	1	hexa-	6	CO_2	carbon dioxide
di-	2	hepta-	7	SCl_3	sulfur trichloride
tri-	3	octa-	8	N_2O_5	dinitrogen pentoxide
tetra-	4	nona-	9	P_4O_{10}	tetraphosphorus decoxide
penta-	5	deca-	10	CO	carbon monoxide

Common misconception: Although there are similarities, the naming system for binary molecular compounds should not be applied to naming ionic compounds. Ionic compounds, usually those that contain at least one metal atom, do not use prefixes unless the prefix is already part of an ion name, as is true of the ion dihydrogen phosphate, $H_2PO_4^-$.

Your Turn 2.6

Name the following compounds: SCl_2, CF_4, BrI_3, PBr_5, SF_6.

Section 2.9 Some Simple Organic Compounds

Organic chemistry is the study of carbon compounds.

Hydrocarbons are compounds containing only carbon and hydrogen.

Alkanes are hydrocarbons containing only C—C single bonds.

Table 2.7 shows the names and formulas of some common alkanes. Notice that in each name a prefix indicates the number of carbon atoms in the formula. The suffix -ane indicates that the formula is an alkane.

Table 2.7 Names and formulas of some simple alkanes. The prefix of the name tells how many carbon atoms are in the formula.

Number of Carbon Atoms	Prefix	Name	Formula
1	meth-	methane	CH_4
2	eth-	ethane	CH_3CH_3
3	prop-	propane	$CH_3CH_2CH_3$
4	but-	butane	$CH_3CH_2CH_2CH_3$
5	pent-	pentane	$CH_3CH_2CH_2CH_2CH_3$
6	hex	hexane	$CH_3CH_2CH_2CH_2CH_2CH_3$
7	hept-	heptane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$
8	oct-	octane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
9	non-	nonane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
10	dec-	decane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

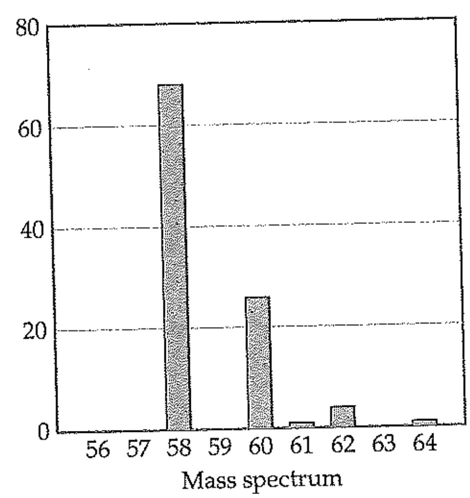
The same prefixes are used to name organic compounds having **functional groups**, groups of atoms that relate to the structure and properties of an organic compound. For example, —OH is the functional group of a class of organic compounds known as **alcohols**. The functional group of a **carboxylic acid** is —COOH. Table 2.8 lists the names and formulas of simple alcohols and carboxylic acids. In Chapter 24 of *Chemistry: The Central Science*, you will study organic compounds in more detail.

Table 2.8 Names and formulas of some simple alcohols and carboxylic acids. The prefix tells how many carbon atoms are in the formula.

Alcohols		**Carboxylic Acids**	
Name	**Formula**	**Name**	**Formula**
methanol	CH_3OH	methanoic acid	H_2COOH
ethanol	CH_3CH_2OH	ethanoic acid	CH_3COOH
1-propanol	$CH_3CH_2CH_2OH$	propanoic acid	CH_3CH_2COOH
2-propanol	$CH_3CHOHCH_3$		
1-butanol	$CH_3CH_2CH_2CH_2OH$	butanoic acid	$CH_3CH_2CH_2COOH$
2-butanol	$CH_3CH_2CHOHCH_3$		

Your Turn 2.7

Predict which element or mixture of elements the mass spectrum below represents. Justify your answer with evidence.



Multiple Choice Questions

1. *Fluorine gas is bubbled through an aqueous solution of calcium bromate. Besides water, this statement refers to which chemical formulas?*

   A) F_2 *and* $CaBr_2$

   B) F *and* $CaBr_2$

   C) F_2 *and* $Ca(BrO_3)_2$

   D) F_2 *and* $CaBrO_3$

2. *Magnesium nitride reacts with water to form ammonia and magnesium hydroxide. Besides water, this statement refers to which chemical formulas?*

   A) Mg_3N_2, NH_3, *and* $Mg(OH)_2$

   B) $Mg(NO_3)_2$, NH_4^+, *and* MgH_2

   C) $Mg(NO_3)_2$, NH_3, *and* MgH_2

   D) Mg_3N_2, NH_3, *and* $MgOH$

3. *The mineral spinel is an ionic compound containing only the elements magnesium, aluminum, and oxygen. Its simplest formula is probably*

   A) $MgAlO_3$

   B) Mg_2AlO_4

   C) $MgAl_2O_4$

   D) $Mg_2Al_2O_3$

4. *The mineral chromite,* $FeCr_2O_4$, *consists of a mixture of iron(II) oxide and chromium(III) oxide. What is the most likely ratio of iron(II) oxide to chromium(III) oxide in chromite?*

   A) 1:1

   B) 1:2

   C) 2:3

   D) 3:2

5. *Which formula represents a peroxide?*

   A) K_2O

   B) K_2O_2

   C) KO_2

   D) CaO

6. *What is the general formula for an alkaline earth metal hydride?*

A) MH

B) M_2H

C) MH_2

D) $M(OH)_2$

7. *Which of the following compounds contains only four carbon atoms?*

A) *propane*

B) *butanoic acid*

C) *ethylmethyl ether*

D) *2-pentanol*

8. *Bromine has just two major isotopes giving it an atomic mass of 79.904 amu. Based on this information, which of the following statements can explain the atomic mass value?*

A) *The isotope Br-81 is more common than Br-79.*

B) *Br-79 and Br-81 exist in about equal proportions.*

C) *Br-78 is about twice as abundant as Br-81.*

D) *The two major isotopes of Br have 45 and 46 neutrons.*

9. *Which is a collection of only molecular compounds?*

A) NO, CS_2, PCl_3, HBr

B) $NaNO_3$, CCl_4, CuS

C) Ar, NH_3, SF_4, PCl_5

D) Cl_2, CCl_4, NO_2, SF_6

10. *Which is true of the* $^{243}Am^{3+}$ *ion?*

	Protons	Electrons	Neutrons
A)	148	148	243
B)	95	98	243
C)	95	95	148
D)	95	92	148

Refer to the following information and Figure 2.4 to answer Questions 11–16.

The mass spectrum of a natural abundance of chlorine atoms is shown in Figure 2.4. Detailed analysis shows that the two stable isotopes of chlorine have masses of 34.969 amu and 36.966 amu.

11. *Which are the mass numbers of the two isotopes of chlorine?*

    A) *34.969 and 36.966*

    B) *34 and 36*

    C) *35 and 37*

    D) *17 and 17*

12. *Estimate the approximate % abundance of the lighter isotope.*

    A) *20*

    B) *25*

    C) *50*

    D) *75*

13. *How many types of molecules with different masses exist in a sample of chlorine gas if the sample exists entirely as diatomic molecules?*

    A) *1*

    B) *2*

    C) *3*

    D) *4*

14. *Estimate the approximate mass of the most abundant naturally occurring Cl_2 molecule.*

    A) *70*

    B) *71*

    C) *72*

    D) *74*

15. *How many neutrons does the less abundant chlorine atom have?*

    A) *17*

    B) *18*

    C) *19*

    D) *20*

16. *Why are the individual masses of the two isotopes not integers?*

A) *Atomic mass of an element is the average mass of all isotopes.*

B) *The masses of individual protons and neutrons are not integers.*

C) *Mass number is the total number of protons and neutrons in an isotope.*

D) *Mass number of an element is the average mass of all isotopes.*

17. *A compound whose empirical formula is C_2H_4O has a molar mass that lies between 100 and 150 g/mol. What is the molecular formula of the compound?*

A) C_2H_4O

B) $C_4H_8O_2$

C) $C_6H_{12}O_3$

D) $C_6H_{12}O_2$

18. *Find the empirical formula for a compound only one element of which is a metal. The compound's percentage composition by mass is 40.0% metal, 12.0% C, and 48.0% O.*

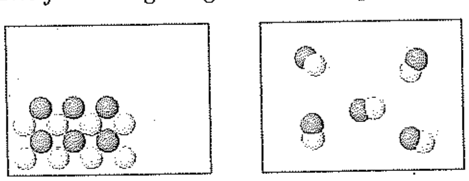
A) $CaCO_3$

B) Na_2CO_3

C) $NaHCO_3$

D) $Al_2(CO_3)_3$

19. *The following diagrams best represent which pair of substances?*



(i) (ii)

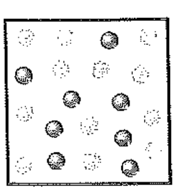
A) NaCl Cl_2

B) KF NO

C) Ag CO

D) H_2O H_2

Answer Questions 20–22 using the following diagram, which represents a sample of a fictional element called chemistrium, Ch, having two different isotopes.



20. *Estimate the natural abundance of the more common isotope.*

    A) 8

    B) 12

    C) 0.40

    D) 0.60

21. *If Ch forms diatomic molecules, how many peaks will the mass spectrum of this element display?*

    A) 1

    B) 2

    C) 3

    D) 4

22. *If Ch has 10 protons, and the two isotopes have 10 neutrons and 15 neutrons, respectively, which best represents the approximate atomic mass of chemistrium?*

    A) 21

    B) 22

    C) 23

    D) 25

Free Response Questions

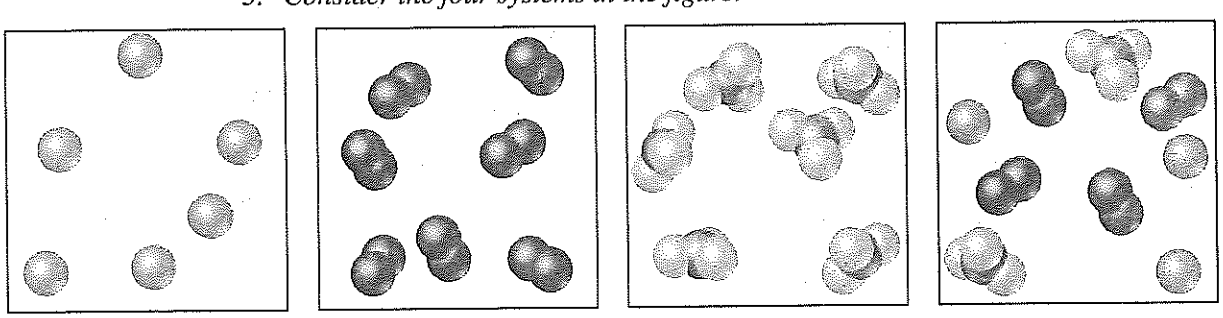
1. *Use the data provided for Multiple Choice Questions 11–16 to perform the following mathematical operations:*

   a. *Calculate the mass of the chlorine molecule having the largest molecular mass.*

   b. *Calculate the % abundance of the more abundant chlorine isotope.*

2. *Iodine, like chlorine, is a halogen and forms similar compounds. Write the names and formulas of the four oxyanions and the four oxyacids of iodine.*

3. *Consider the four systems in the figure:*



I II III IV

   a. *Identify each system as a pure substance or a mixture. Explain your answer.*

   b. *Identify each pure substance as an element or a compound. Justify your answer.*

c. When physically divided, which system, III or IV, will always contain particles that have identical ratios of masses? Explain using atomic molecular theory.

4. *Examine Figure 2.1 and tell which of the four postulates of Dalton's atomic theory is(are) incorrect according to the modern view of the atom. Justify your answer with evidence that demonstrates why each postulate is incorrect.*

5. *Examine Figure 2.4b and use it to estimate the relative abundances of chlorine-35 and chlorine-37. Is your estimate consistent with the atomic mass of chlorine? Justify your answer.*

6. *Examine the different representations of methane in Figure 2.8. Draw five analogous representations of the molecule ethane. Tell what kind of information each representation yields.*

7. *a. Why is an empirical formula used to represent sodium chloride in Figure 2.9?*
   b. A magnifying glass reveals that common table salt is composed of many tiny cubic crystals. How does the arrangement of ions in solid sodium chloride explain this observation?

TOPIC

3

CHEMICAL REACTIONS AND REACTION STOICHIOMETRY

You may have learned some of this material in first-year chemistry. When you finish reviewing this topic, be sure you are able to:

- Apply the law of conservation of mass to balance a chemical equation using symbols for atoms and molecules and particle drawings
- Use the mole as a quantitative model for chemical composition
- Calculate moles, mass, number of particles, and volume of a gas and interconvert those quantities
- Determine the identity or purity of various substances using calculations of mass data
- Calculate the percentage composition of a compound
- Calculate the empirical and molecular formulas of a compound and of a hydrate from combustion and decomposition data
- Calculate the masses and moles of reactants and products using stoichiometry
- Determine limiting reactants and percent yields from experimental data

Chemical Equations

Section 3.1

Represent changes in matter with a balanced chemical or net ionic equation:

a. *For physical changes.*

b. *For given information about the identity of the reactants and/or product.*

c. *For ions in a given chemical reaction.* See also Sections 4.2 and 4.3 for net ionic equations.

TRA-1.B.

Represent a given chemical reaction or physical process with a consistent particulate model. See also Section 16.2.

TRA-1.C.

Stoichiometry is the area of study that examines the quantities of substances involved in chemical reactions.

A **chemical reaction** also called a chemical change, is a process by which one or more substances are converted to other substances.

Chemical equations use chemical formulas to symbolically represent chemical reactions. For example, the chemical sentence "Hydrogen burns in air to produce water" can be expressed as chemical symbols or as a molecular picture of the particles, as shown in Figure 3.1.

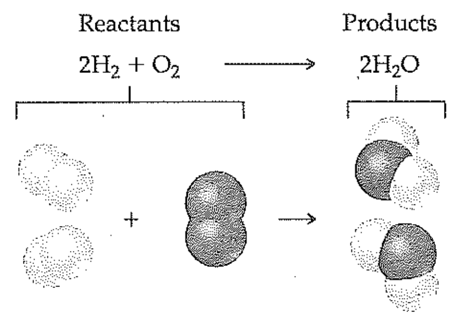


Figure 3.1 A balanced chemical equation.

When the reaction is more complex, chemists prefer to write the equation using only chemical symbols to represent the particle pictures. For example, the following chemical equation describes how butane, C_4H_{10}, burns in air:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

$$\text{reactants} \rightarrow \text{products}$$

The equation is a "chemical sentence" written in the symbolic "words" of chemical formulas and symbols. The formulas on the left are reactants, and the formulas on the right are products. The equation reads:

"Two molecules of gaseous butane react with thirteen molecules of oxygen gas to produce eight molecules of carbon dioxide gas and ten molecules of liquid water."

A **balanced chemical equation** has the same number of atoms of each element on each side of the arrow. The coefficients preceding each formula "balance" the equation. Notice that, because of the coefficients, on each side of the arrow there are eight carbon atoms, twenty hydrogen atoms, and twenty-six oxygen atoms.

The symbols (g), (l), (s), and (aq) are used to designate the physical state of each reactant and product: gas, liquid, solid, and aqueous (dissolved in water).

Your Turn 3.1

Write a chemical sentence to illustrate how to "read" the following chemical equation:

$$CO_2(g) + 2NH_3(g) + H_2O(l) \rightarrow 2NH_4^+(aq) + CO_3^{2-}(aq)$$

Convert each formula of the equation to a particle representation.

To balance simple equations, start by balancing atoms other than hydrogen or oxygen. Balance hydrogen atoms next to last, and balance oxygen atoms last.

Example:

Balance the following equation:

$H_2SO_4(aq) + NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

Solution:

1. *S and C are already balanced, so start with Na.*

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

2. *Now balance carbon.*

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

3. *Next balance H.*

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2CO_2(g) + 2H_2O(l) + Na_2SO_4(aq)$

4. *Check to see that oxygen is balanced and double-check all other atoms. (As in algebraic equations, the numeral 1 is usually not written.)*

Simple Patterns of Chemical Reactivity

Section 3.2

Predicting the products of chemical reactions is an essential skill to acquire in the study of chemistry. Sometimes reactions fall into simple patterns, and recognizing these patterns can be helpful in predicting which products will be produced from given reactants. Topic 4 addresses predicting products of chemical reactions in more depth. For now, here are a few patterns to learn to recognize.

1. In a combination reaction, two elements combine to form one compound. (Other combinations are possible, and Topic 4 describes better ways to predict what will happen.)

Example:

A metal reacts with a nonmetal to produce an ionic compound. Write an equation to describe what happens when solid sodium is exposed to chlorine gas.

Solution:

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Your Turn 3.2

Write an equation to describe what happens when solid magnesium metal reacts at high temperature with nitrogen gas.

2. In a decomposition reaction, one reactant changes to two or more products.

Example:

Upon heating, a metal carbonate decomposes to yield a metal oxide and carbon dioxide. Write an equation to describe what happens when solid magnesium carbonate is heated.

Solution:

$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

Your Turn 3.3

Write an equation to describe what happens when liquid water is decomposed to its elements by an electric current.

3. A combustion reaction usually involves oxygen, often from air, reacting with hydrocarbons or other organic molecules containing carbon, hydrogen, and oxygen to produce carbon dioxide and water.

Example:

Write an equation to describe what happens when liquid ethanol burns in air.

Solution:

$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

(Recall that encoded in the name "ethanol" is a two-carbon alcohol having the –OH functional group.)

Your Turn 3.4

Write an equation to describe what happens when liquid hexane is burned in air.

Recall that a physical change is one in which the appearance of the substance changes but its chemical composition does not change. Examples are changes of physical state from solid to liquid or from liquid to gas. Like chemical reactions, physical changes are represented using chemical symbols. For example,

The melting of ice: $H_2O(s) \rightarrow H_2O(l)$

The evaporation of water: $H_2O(l) \rightarrow H_2O(g)$

TRA-1.D.

Explain the relationship between macroscopic characteristics and bond interactions for:

a. Chemical processes.

b. Physical processes.

Chemical changes usually involve the breaking and/or formation of chemical bonds (see Topic 8). Processes that involve only changes in intermolecular forces (see Section 11.2), such as phase changes, are typical of physical changes.

The distinction between chemical and physical changes is sometimes arbitrary. For example, plausible arguments can be made for the dissolution of a salt in water as either a physical or chemical change. The process involves breaking ionic bonds and formation of ion-dipole interactions between ions and solvent (see Section 13.1).

Avogadro's Number and the Mole

Section 3.4

SPQ-1.A.

Calculate quantities of a substance or its relative number of particles using dimensional analysis and the mole concept. See also Sections 1.7 and 2.4.

Avogadro's number ($N_A = 6.02 \times 10^{23}\ mol^{-1}$) provides a way to count the number of atoms, ions, molecules, or formula units in a measured mass of any pure substance.

Avogadro's number is 6.02×10^{23}. It represents the number of atoms in exactly 12 g of isotopically pure ^{12}C. Because all atomic masses are based on ^{12}C, the atomic mass of any element, expressed in grams, represents Avogadro's number (6.02×10^{23}) of atoms of that element.

A **mole** is the amount of matter that contains 6.02×10^{23} atoms, ions, molecules, or formula units.

The **molar mass** of a substance is the mass in grams of one mole of that substance. To calculate a molar mass of any substance, add the atomic masses of all the atoms in its formula. (For convenience, atomic masses are often rounded to three significant figures.) Table 3.1 shows the molar masses of various substances.

Common misconception: "Molar mass" is a universal term that is often used to replace the terms atomic mass, molecular mass, and formula mass. Molar mass is used to express the mass of one mole of any substance, whether it is an atom (atomic mass), a molecule (molecular mass), an ion (formula mass), or an ionic compound (formula mass).

Table 3.1 The molar mass of any substance is the mass in grams of 6.02×10^{23} particles, or one mole, of that substance.

Formula	Number of Particles	Representative Particles	Molar Mass	Alternative Term
Ar	6.02×10^{23}	Atoms	39.9 g/mol	Atomic mass
CO_2	6.02×10^{23}	Molecules	44.0 g/mol	Molecular mass
NaBr	6.02×10^{23}	Formula units	103 g/mol	Formula mass
CO_3^{2-}	6.02×10^{23}	Ions	60.0 g/mol	Formula mass

Grams, moles, and representative particles (atoms, molecules, ions, or formula units) are converted from one to another using the "mole road" described in Figure 3.2.

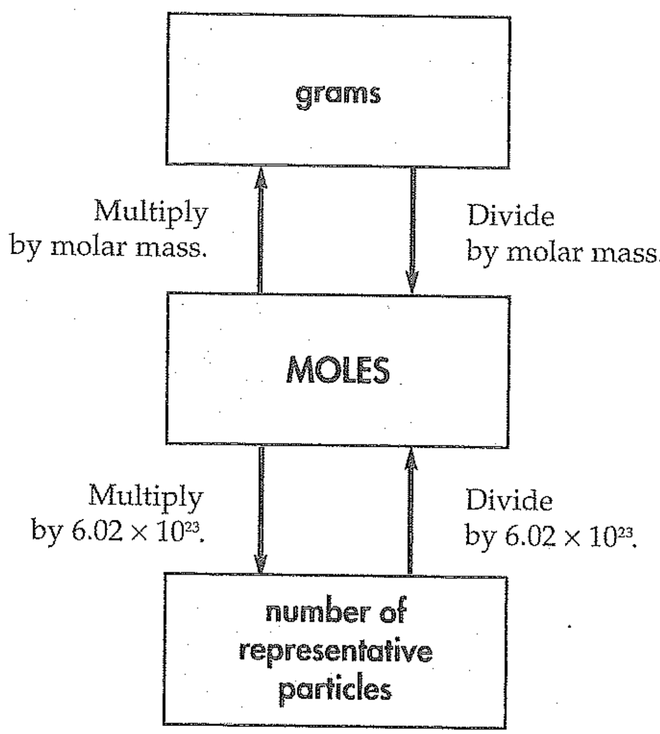


Figure 3.2 The "mole road." Divide to convert to moles. Multiply to convert from moles.

Calculating Percentage Composition of a Compound

The percentage composition of a compound is the percentage by mass contributed by each element in the compound. To calculate the percentage composition of an element in any formula, divide the molar mass of the element multiplied by the number of times it appears in the formula, by the molar mass of the formula, and then multiply by 100.

% composition = 100 × (molar mass of element × subscript for element)/(molar mass of substance)

Example:

What is the percentage composition of Na_2CO_3?

Solution:

$$\% Na = 100 \times 2(23.0)\ g\ Na/[2(23.0) + 12.0 + 3(16.0)\ g] = 43.4\%\ Na$$

$$\% C = 100 \times 12.0\ g\ C/[2(23.0) + 12.0 + 3(16.0)\ g] = 11.3\%\ C$$

$$\% O = 100 \times 3(16.0)\ g/[2(23.0) + 12.0 + 3(16.0)\ g] = 45.3\%\ O$$

Your Turn 3.5

Determine the percentage composition of $Ca_3(PO_4)_2$.

Empirical Formulas from Analyses

Section 3.5

The **empirical formula** for a compound expresses the simplest ratio of atoms in the formula. The percentage composition of a compound can be determined experimentally by chemical analysis, and the empirical formula can be calculated from the percentage composition.

Example:

What is the empirical formula of a compound containing 68.4% chromium and 31.6% oxygen?

Solution:

In chemistry, percentage always means mass percentage, unless specified otherwise. The data mean that for every 100 g of compound, there are 68.4 g of Cr and 31.6 g of O.

1. *Write the formula using number of grams.*

$Cr_{68.4}\ O_{31.6}$

2. *Convert grams to moles by dividing by the molar mass of each element.*

$Cr_{68.4/52.0}\ O_{31.6/16.0} = Cr_{1.315}\ O_{1.975}$

3. *Convert to small numbers by dividing each mole quantity by the smaller mole quantity.*

$Cr_{1.315/1.315}\ O_{1.975/1.315} = Cr_1\ O_{1.5}$

4. *If necessary, multiply each mole quantity by a small whole number that converts all quantities to whole numbers.*

$Cr_{1\times2}O_{1.5\times2} = Cr_2O_3$

Your Turn 3.6

Analysis of a poison victim's blood revealed a compound containing 15.8% Na, 51.4% As, and 32.8% O. Determine the simplest formula for the poison.

Molecular Formulas from Empirical Formulas

A **molecular formula** tells exactly how many atoms are in one molecule of the compound. The subscripts in a molecular formula are always whole-number multiples of the subscripts in the empirical formula. Molecular formulas can be determined from empirical formulas if the molar mass of the compound is known.

Example:

A compound containing only carbon, hydrogen, and oxygen is 63.16% C and 8.77% H. Mass spectrometry shows that the compound has a molar mass of 114 g/mol. What are its empirical formula and its molecular formula?

Solution:

1. *Write the formula using number of grams.*

$C_{63.16}\,H_{8.77}\,O_{28.07}$

(Calculate the grams of oxygen by subtracting the grams of carbon and the grams of hydrogen from 100 *g.*

$x\,g\,O = 100\,g - 63.16\,g\,C - 8.77\,g\,H = 28.07\,g\,O.)$

2. *Convert grams to moles by dividing by the molar mass of each element.*

$C_{63.16/12.0}\,H_{8.77/1.00}\,O_{28.07/16.0} = C_{5.263}\,H_{8.77}\,O_{1.754}$

3. *Convert to small numbers by dividing each mole quantity by the smallest mole quantity.*

$C_{5.263/1.754}\,H_{8.77/1.754}\,O_{1.754/1.754} = C_3\,H_5\,O_1$

empirical formula

4. *All quantities are whole numbers.*

5. *Divide the known molar mass by the mass of one mole of the empirical formula. The result produces the integer by which you multiply the empirical formula to obtain the molecular formula.*

$(114\,g/mol)/(57.0\,g/mol) = 2$

$C_{3\times2}\,H_{5\times2}\,O_{1\times2} = C_6\,H_{10}\,O_2$

molecular formula

Empirical Formulas from Combustion Analysis

When a compound containing carbon, hydrogen, and oxygen is completely combusted, all the carbon is converted to carbon dioxide, and all the hydrogen becomes water. The empirical formula of the compound can be calculated from the measured masses of the products.

Example:

A 3.489-g sample of a compound containing C, H, and O yields 7.832 g of CO_2 and 1.922 g of water upon combustion. What is the simplest formula of the compound?

The unbalanced chemical equation is

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

X is the number of moles of carbon in the compound because all the carbon in the compound is converted to CO_2. X equals the number of moles of CO_2 because there is one mole of carbon in one mole of CO_2.

$$X = mol\ C = 7.832\ g\ CO_2/44.0\ g/mol = 0.178\ mol\ C$$

Y is the number of moles of hydrogen in the compound because all the hydrogen becomes water. Y equals twice the number of moles of water because there are two moles of hydrogen in one mole of water.

$$Y = mol\ H = (1.922\ g\ H_2O/18.0\ g/mol) \times 2 = 0.2136\ mol\ H$$

Z is the number of moles of O. To obtain the number of grams of O in the compound, subtract the number of grams of C ($X\ mol \times 12.0\ g/mol$) and the number of grams of H ($Y\ mol \times 1.00\ g/mol$) from the total grams of the compound. Convert the result to moles of O by dividing by 16.0 g/mol.

$$\begin{aligned} Z = mol\ O &= [3.489 - (12.0)(0.178) - (1.00)(0.2136)]/16.0 \\ &= 0.0712\ mol\ O \end{aligned}$$

Convert to small numbers by dividing each mole quantity by the smallest mole quantity.

$$C_{0.178}\,H_{0.2136}\,O_{0.0712} =$$

$$C_{0.178/0.0712}\,H_{0.2136/0.0712}\,O_{0.0712/0.0712} = C_{2.5}H_3O_1$$

Finally, multiply each mole quantity by a small whole number that converts all quantities to whole numbers.

$$C_{2.5\times 2}\,H_{3\times 2}\,O_{1\times 2} = C_5H_6O_2$$

Your Turn 3.7

A 3.489-g sample of an unknown compound containing only C, H, and O yields 7.832 g of carbon dioxide and 1.922 g of water upon combustion. a. Determine the empirical formula of the compound. b. The mass spectrum shows that the compound has a molar mass of 196 g/mol. Determine the molecular formula of the compound. c. Write a balanced equation for the combustion of the compound.

Formulas of Hydrates

Ionic compounds often form crystal structures called **hydrates** by acquiring one or more water molecules per formula unit. For example, solid sodium thiosulfate decahydrate, $Na_2S_2O_3 \cdot 10H_2O$, has ten water molecules per formula unit of sodium thiosulfate. Heating a sample of hydrate causes it to lose water. The number of water molecules per formula unit can be calculated from the mass difference before and after heating.

Example:

When 21.91 g of a hydrate of copper(II) sulfate is heated to drive off the water, 14.00 g of anhydrous copper(II) sulfate remains. What is the formula of the hydrate?

Solution:

This is a variation of an empirical formula problem. The solution lies in calculating the ratio of H_2O moles to the moles of $CuSO_4$. The chemical equation is

$$\underset{21.91\ g}{CuSO_4 \cdot X\,H_2O(s)} \rightarrow \underset{14.00\ g}{CuSO_4(s)} + X\,H_2O(g)$$

Calculate the grams of water by subtracting the grams of copper(II) sulfate from the grams of the hydrate.

$x\ g\ H_2O = 21.91\ g - 14.00\ g = 7.91\ g\ H_2O$

$mol\ H_2O = 7.91\ g\ H_2O/18.0\ g/mol = 0.439\ mol\ H_2O$

$mol\ CuSO_4 = 14.00\ g/159.5\ g/mol = 0.08777\ mol\ CuSO_4$

$mol\ H_2O/mol\ CuSO_4 = 0.439\ mol/0.08777\ mol = 5.00$

The formula has five moles of water per mole of copper(II) sulfate: $CuSO_4 \cdot 5H_2O$.

Your Turn 3.8

Upon heating, 39.208 g of sodium thiosulfate hydrate yields 18.328 g of sodium thiosulfate. Determine the formula of the hydrate.

Quantitative Information from Balanced Equations

Section 3.6

Explain changes in the amounts of reactants and products based on the balanced reaction equation for a chemical process. See also Section 4.6.

SPQ-4.A.

Stoichiometry explores the quantities of substances involved in chemical reactions. The coefficients in a balanced chemical equation indicate both the relative number of molecules (or formula units) involved in the reaction, and the relative number of moles. For example, the equation for the combustion of octane, C_8H_{18}, a component of gasoline, is

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

The four coefficients that balance the equation are proportional to one another and can be used to relate mole quantities of reactants and/or products.

Example:

How many moles of octane will burn in the presence of 37.0 moles of oxygen gas?

Solution:

The balanced equation tells us that 2 moles of octane will burn in 25 moles of oxygen gas, so the answer to the question involves the ratio 2 mol C_8H_{18} per 25 mol O_2.

$$x\ mol\ C_8H_{18} = 37.0\ mol\ O_2\ (2\ mol\ C_8H_{18}/25\ mol\ O_2) = 37.0 \times 2/25$$
$$= 2.96\ mol\ C_8H_{18}.$$

Alternatively, we can use the mole road:

To convert mol O_2 to mol C_8H_{18}, multiply by 2/25.

$$x\ mol\ C_8H_{18} = 37.0\ mol(2/25) = 2.96\ mol\ C_8H_{18}$$

(Always multiply by the coefficient at the head of the arrow and divide by the coefficient at the tail of the arrow.)

$$mol\ C_8H_{18} \underset{\times 2/25}{\overset{\times 25/2}{\rightleftarrows}} mol\ O_2$$

To convert mol C_8H_{18} to mol O_2, multiply by 25/2.

Figure 3.3 shows an expanded mole road to include the relationship between the moles of any reactant or product in a balanced chemical equation.

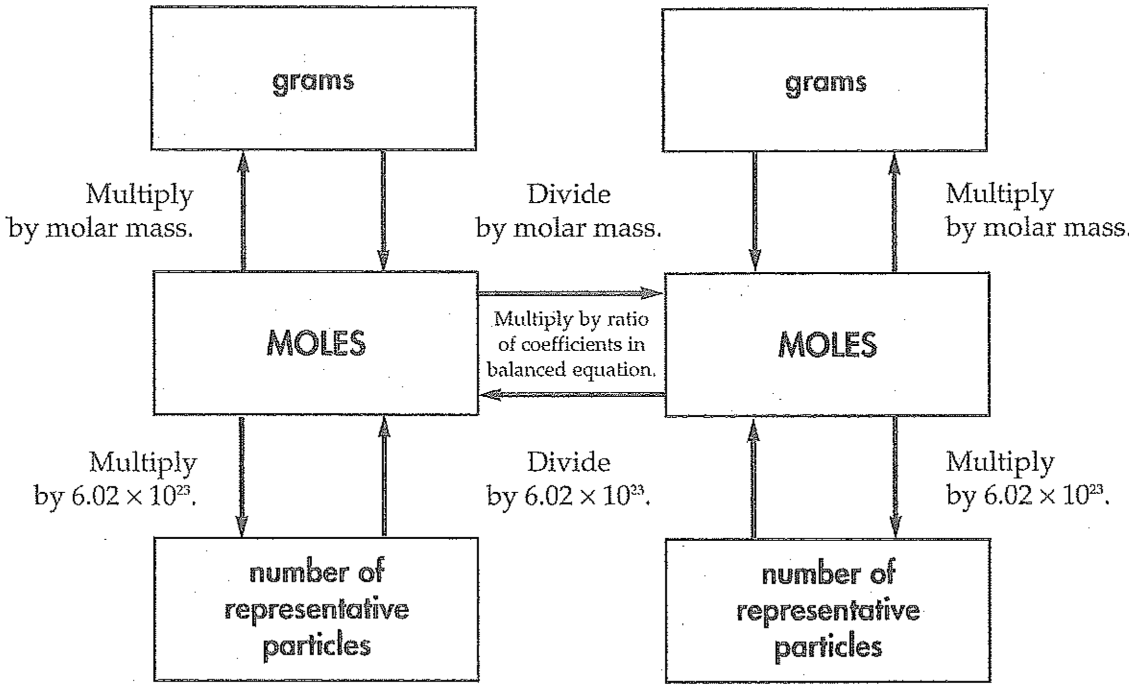


Figure 3.3 The stoichiometry "mole road." Divide to convert to moles. Multiply to convert from moles. Multiply by the ratio of coefficients in a balanced equation to convert moles of one substance to moles of another substance.

This stoichiometry mole road allows us to relate any quantity in a balanced equation to any other quantity in the same equation.

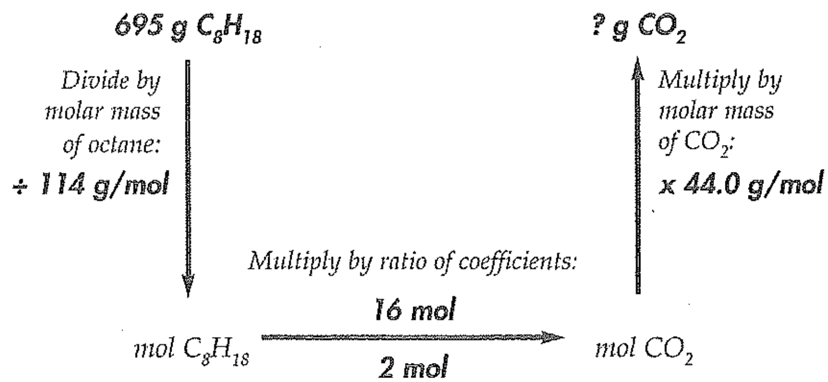
Example:

How many grams of carbon dioxide are obtained when 695 g (about 1 gallon) of octane are burned in oxygen?

Solution:

Follow the road map:

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



$$x \text{ g } CO_2 = (695 \text{ g})(1 \text{ mol}/114 \text{ g})(16 \text{ mol}/2 \text{ mol})(44.0 \text{ g/mol}) = 2150 \text{ g } CO_2$$

Your Turn 3.9

Calculate the number of grams of oxygen gas needed to produce 152 g of carbon dioxide in the combustion of gasoline.

Explain the quantitative relationship between the elemental composition by mass and the composition of substances in a mixture. See also Section 1.2.

SPQ-2.B.

A pure substance contains, atoms, molecules or formula units of a single element or a single compound. Mixtures contain atoms, molecules, or formula units of two or more substances, whose relative proportions can vary.

Chemists use stoichiometric calculations to determine the purity of a substance as a percentage.

Example:

Chalk is composed of a mixture of calcium carbonate and calcium sulfate. To determine the percentage of calcium carbonate in the chalk, a student reacts 100.0 g of chalk with excess hydrochloric acid and finds that the reaction produces 33.0 g of carbon dioxide. What is the percentage of calcium carbonate in the chalk? Assume that calcium sulfate does not react with acid and that calcium carbonate reacts according to the following balanced equation:

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

Solution:

The amount of carbon dioxide generated by reaction with the calcium carbonate in the chalk is directly proportional to the amount of calcium carbonate in the mixture.

$$x\ g\ CaCO_3 = 33.0\ g\ CO_2(1\ mol\ CO_2/44.0\ g\ CO_2)$$
$$(1\ mol\ CaCO_3/1\ mol\ CO_2)(100.0\ g\ CaCO_3)/1\ mol\ CaCO_3)$$
$$= 75.0\ g\ CaCO_3$$
$$\%\ CaCO_3 = (g\ CaCO_3/g\ chalk)(100) = (75.0\ g/100.0\ g)(100) = 75.0\%$$

Your Turn 3.10

The purity of a mixture of sodium sulfate with sodium chloride is determined by gravimetric analysis. A solid 2.172-g sample of the mixture is dissolved in water. Excess solution of barium chloride is added to precipitate all the sulfate as barium sulfate. The barium sulfate precipitate is filtered, dried, and weighed, indicating a mass of 1.234 g. a. Write a chemical equation for the precipitation reaction. b. Calculate the percentage of sodium sulfate in the original sample.

Section 3.7 Limiting Reactants

The **limiting reactant** is the reactant that is completely consumed in a chemical reaction. The limiting reactant limits the amount of products formed.

The **excess reactant** is usually the other reactant. Some of the excess reactant is left unreacted when the limiting reactant is completely consumed.

In a **stoichiometric mixture**, both reactants are limiting and both are completely consumed by the reaction. There is an excess of neither. Because of their subtle nature, quantitative limiting-reactant problems are among the most difficult. The mole road is a useful tool in solving limiting-reactant problems.

Common misconception: Stoichiometry calculations can calculate only how much reactant reacts or how much product is formed. Stoichiometry cannot calculate how much excess reactant is left unreacted. To calculate how much excess reactant remains after the reaction is complete, first calculate how much is consumed and then subtract that amount from how much total reactant was initially present.

Example:

255 g of octane and 1510 g of oxygen gas are present at the beginning of a reaction that goes to completion and forms carbon dioxide and water according to the following equation:

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

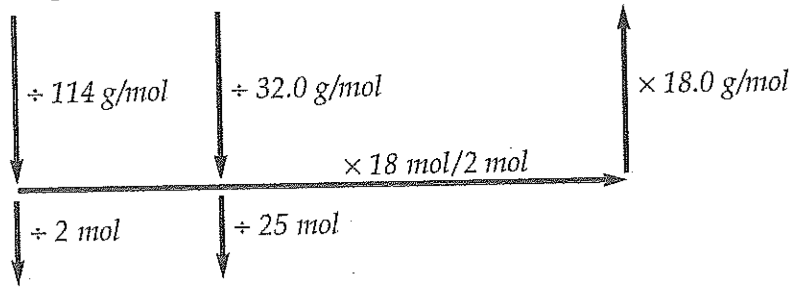
a. *What is the limiting reactant?*
b. *How many grams of water are formed when the limiting reactant is completely consumed?*
c. *How many grams of the reactant that is in excess are consumed?*
d. *How many grams of excess reactant are left unreacted?*

Solution:

a. *To find the limiting reactant, first compare the number of moles of reactants relative to the ratio in which they react. To do this, divide the number of moles of each reactant by its corresponding coefficient that balances the equation. The resulting lower number always identifies the limiting reactant. Then use the mole road to solve the problem based on the identified limiting reactant.*

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

255 g *1510 g* *? g*



1.12 < 1.89

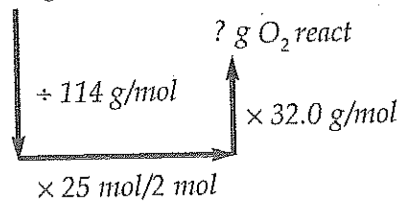
limiting reactant

a. *C_8H_{18} is the limiting reactant because* 1.12 < 1.89.

b. $x\ g\ H_2O = (255\ g/114\ g/mol)(18\ mol/2\ mol)(18.0\ g/mol)$
$= 362\ g\ H_2O$

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

255 g *1510 g* *? g*



$x\ g\ O_2 = (255\ g/114\ g/mol)(25\ mol/2\ mol)(32.0\ g/mol)$
$= 895\ g\ O_2\ react$

d. $g\ O_2$ *left unreacted* $= 1510\ g - 895\ g = 615\ g\ O_2$ *unreacted*

Theoretical, Actual, and Percent Yields

The **theoretical yield** of a reaction is the quantity of product that is calculated to form.

The **actual yield** is the amount of product actually obtained and is usually less than the theoretical yield.

The **percent theoretical yield** relates the actual yield to the theoretical yield:

Percent theoretical yield = (actual yield)/(theoretical yield) × 100

Example:

In the previous example, the theoretical yield of water is calculated to be 362 g. What is the percent yield if the actual yield of water is only 312 g?

Solution:

Percent theoretical yield = 312 g/362 g × 100 = 86.2%

Your Turn 3.11

Nitric acid is used to etch silver in making jewelry. The process produces the toxic gas nitrogen monoxide gas:

$$3Ag(s) + 4HNO_3(l) \longrightarrow 3AgNO_3(aq) + NO(g) + 2H_2O(l)$$

a. *If 84.0 g of silver is mixed with 75.0 g of nitric acid, what is the limiting reactant?*
b. *How many grams of water can be produced?*
c. *How many grams of excess reactant remain?*
d. *If 7.12 g of NO is produced by the reaction, calculate the percent yield.*

Multiple Choice Questions

1. *Ammonia forms when hydrogen gas reacts with nitrogen gas. If equal numbers of moles of nitrogen and hydrogen are combined, the maximum number of moles of ammonia that could be formed will be equal to*

   A) *the number of moles of hydrogen*

   B) *the number of moles of nitrogen*

   C) *two-thirds the number of moles of hydrogen*

   D) *twice the number of moles of nitrogen*

2. *If $C_4H_{10}O$ undergoes complete combustion, what is the sum of the coefficients when the equation is completed and balanced using smallest whole numbers?*

   A) 8

   B) 16

   C) 22

   D) 25

3. *What are the products when lithium carbonate is heated?*

   A) $LiOH + CO_2$

   B) $Li_2O + CO_2$

   C) $LiO + CO_2$

   D) $LiC + O_2$

4. *Beginning with 48 moles of H_2, how many moles of $Cu(NH_3)_4Cl_2(aq)$ can be obtained if the synthesis of $Cu(NH_3)_4Cl_2(aq)$ is carried out through the following sequential reactions? Assume that a 50% yield of product(s) is (are) obtained in each reaction.*

   1. $3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$
   2. $4NH_3(g) + CuSO_4(aq) \rightarrow Cu(NH_3)_4SO_4(aq)$
   3. $Cu(NH_3)_4SO_4 + 2NaCl \rightarrow Cu(NH_3)_4Cl_2(aq) + Na_2SO_4(aq)$

   A) 1

   B) 2

   C) 4

   D) 8

5. *What mass of water can be obtained from 4.0 g of H_2 and 16 g of O_2?*

   $2H_2(g) + O_2(g) \rightarrow 2H_2O$

   A) 9.0 g

   B) 18 g

   C) 36 g

   D) 54 g

6. *The empirical formula of pyrogallol is C_2H_2O and its molar mass is 126. Its molecular formula is*

   A) C_2H_2O

   B) $C_4H_4O_2$

   C) $C_2H_6O_3$

   D) $C_6H_6O_3$

7. *What is the maximum amount of water that can be prepared from the reaction of 20.0 g of HBr with 20.0 g of $Ca(OH)_2$?*

   $2HBr + Ca(OH)_2 \rightarrow CaBr_2 + 2H_2O$

   A) (20/81) (2/2) (18) g

   B) (20/74) (2/2) (18) g

   C) (20/81) (2/1) (18) g

   D) (20/74) (2/1) (18) g

8. *How many moles of ozone, O_3, could be formed from 96.0 g of oxygen gas, O_2?*

   A) 0.500

   B) 1.00

   C) 2.00

   D) 3.00

9. *The percentage of oxygen in $C_8H_{12}O_2$ is*

   A) (16/140) (100)

   B) (32/140) (100)

   C) (16/124) (100)

   D) (140/32) (100)

10. *A compound contains 48.0% O and 40.0% Ca, and the remainder is C. What is its empirical formula?*

A) $O_3C_2Ca_2$

B) O_3CCa_2

C) O_3CCa

D) O_3CCa_2

11. *Calculate the maximum number of grams of $MgCl_2$ that can be prepared from the reaction of 20.0 g of HCl with 20.0 g of $Mg(OH)_2$.*

A) $(20.0/36.5)(1/2)(58.3)$

B) $(20.0/36.5)(95.3)$

C) $(20.0/36.5)(1/2)(95.3)$

D) $(20.0/58.3)(1/2)(95.3)$

12. *Calculate the maximum number of grams of CO_2 that can be produced from 50.0 g each of sulfuric acid and sodium hydrogen carbonate. The unbalanced equation is*
$NaHCO_3 + H_2SO_4 \rightarrow CO_2 + H_2O + Na_2SO_4$

A) $(50.0/98.0)(1/2)(44.0)$

B) $(50.0/98.0)(44.0)$

C) $(50.0/84.0)(2)(44.0)$

D) $(50.0/84.0)(44.0)$

13. *Heating Br_2O_7 causes it to decompose to its gaseous elements. What is the ratio of bromine to oxygen molecules in the product?*

A) 1 to 7

B) 2 to 7

C) 7 to 2

D) 7 to 1

14. *A sample of hydrated copper(II) sulfate, $CuSO_4 \cdot XH_2O$, weighs 24.95 g. When the water is driven off, the anhydrous form weighs 15.95 g. What is the value of x in the formula of the hydrated salt?*

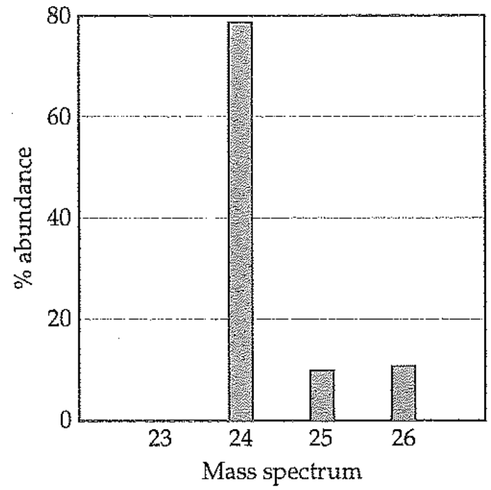
A) 2

B) 3

C) 4

D) 5

15. The mass spectrum shown is most likely that of



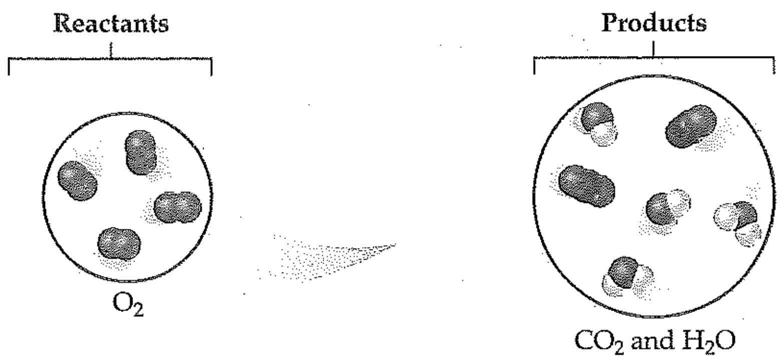
A) magnesium

B) sodium

C) a mixture of magnesium and sodium

D) a mixture of Cr, Mn, and Fe

16. The figure shows oxygen molecules as reactants and carbon dioxide and water as products. How many molecules of what other reactant would you draw to make this a balanced equation?



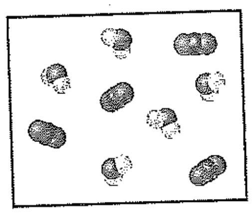
A) 1 C_2H_8

B) 4 CH_2

C) 1 C_2H_6

D) 2 CH_4

17. *The molecules shown represent a collection of carbon dioxide and water molecules formed by the complete combustion of a hydrocarbon. What is the empirical formula of the hydrocarbon?*



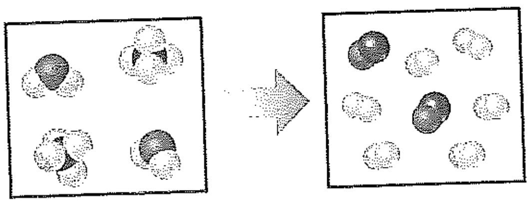
A) $C_4H_{10}O_{13}$

B) C_4H_{10}

C) C_2H_5

D) $C_4H_{10}O_2$

18. *The following depiction represents the reaction between water and methane. How many moles of what products will be produced from 3 moles of methane and 4 moles of water?*



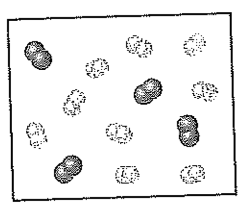
A) 3 mol CO + 6 mol H_2

B) 3 mol CO + 9 mol H_2

C) 4 mol CO + 6 mol H_2

D) 4 mol CO + 12 mol H_2

19. *In the following diagram the dark spheres represent N and the light spheres represent H. If the reaction of nitrogen with hydrogen to form ammonia goes to completion, how many molecules of each species would you draw to represent all the molecules in the final mixture?*



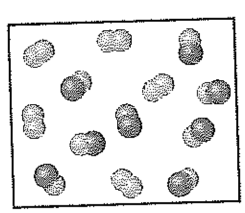
A) 6 NH_3 only

B) 6 NH_3 + 1 N_2

C) 4 NH_3 + 2N_2 + 3 H_2

D) 4 NH_3 + 1 H_2

20. *Nitrogen monoxide reacts with oxygen to form nitrogen dioxide. If the picture represents the reactant molecules, how many nitrogen dioxide molecules would you draw as products if the reaction had a yield of 50%?*



A) 4

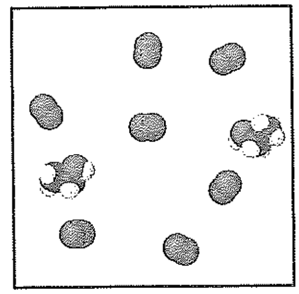
B) 6

C) 8

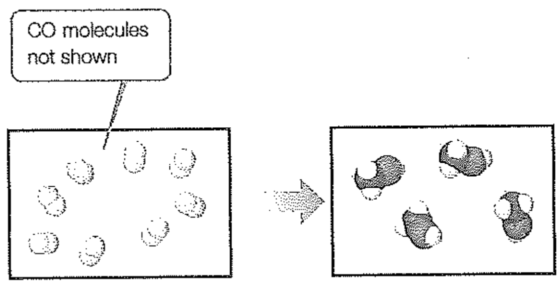
D) 10

21. *If complete combustion of the following mixture of ethanol and oxygen goes to completion, what will be the correct number of moles of each species in the final mixture?*

	CH_3CH_2OH	O_2	CO_2	H_2O
A)	0	0	4	6
B)	0	1	4	6
C)	1	3	2	3
D)	1	0	2	3



22. *Hydrogen reacts with carbon monoxide to form methanol. The following diagram depicts the reactants and products with the carbon monoxide molecules not shown. How many molecules of carbon monoxide would you draw to complete the diagram?*



A) 2

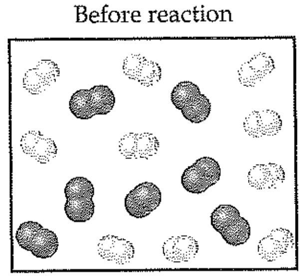
B) 4

C) 6

D) 8

23. *The diagram shows a reaction mixture of hydrogen molecules (light spheres) and oxygen molecules (dark spheres). What is the limiting reactant and how many molecules of the excess reactant will be left after complete reaction?*

Before reaction



A) Hydrogen, 2

B) Oxygen, 2

C) Hydrogen, 3

D) Oxygen, 3

Free Response Questions

1. *Combustion of 8.652 g of a compound containing C, H, O, and N yields 11.088 g of CO_2, 3.780 g of H_2O, and 3.864 g of NO_2.*

   a. *How many moles of C, H, and N are contained in the sample?*

   b. *How many grams of oxygen are contained in the sample?*

   c. *What is the simplest formula of the compound?*

   d. *If the molar mass of the compound lies between 200 and 300, what is its molecular formula?*

   e. *Write and balance a chemical equation for the combustion of the compound.*

2. *A student finds that 344.0 g of a pure sample of the mineral gypsum contains 80.00 g of calcium.*

   a. *Show a mathematical calculation to demonstrate that the pure mineral sample does not contain pure calcium sulfate.*

   b. *If gypsum is hydrated calcium sulfate, use the data from the experiment to derive the chemical formula for gypsum. (How many waters of hydration does the formula contain?)*

3. *Decomposition of 36.54 g of a pure solid compound produces 4.06 g of nitrogen gas, 10.44 g of water, and a solid metal oxide. The metal oxide is found to contain 68.42% chromium.*

   a. *What is the simplest formula for the metal oxide?*

   b. *What is the oxidation number of chromium in the oxide?*

   c. *How many moles of each element are present in the unknown compound?*

   d. *What is the simplest formula for the unknown compound?*

   e. *Express the formula of the compound with a common cation–anion pair, and name the compound.*

   f. *Write and balance a chemical equation for the decomposition reaction.*