TOPIC 1

Introduction: MATTER 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

- Know the differences between elements, compounds, and mixtures
- Understand that separation of mixtures is based on their physical properties
- Use SI units for measurement
- Apply dimensional analysis and significant figures to calculations

Classification of Matter — Section 1.2

Matter is classified as 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

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 properties are used to separate the components of mixtures.

Filtration separates a solid from a liquid.

Distillation separates substances based on their differences in boiling points.

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 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.

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

Uncertainty in Measurement

Section 1.5

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.

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

For addition and subtraction, the result has the same number of decimal places as the measurement with the fewest 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, all numerical answers are usually rounded to three significant figures.

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 sentence is:

$x\ \mu s = 1\ yr.$

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.

$x\ \mu s = 1\ yr\ (365\ days/yr)\ (24\ h/day)\ (60\ min/h)\ (60\ s/min)$

$(10^6\ \mu s/s) = 3.15 \times 10^{13} \mu s$

TOPIC

2

ATOMS, MOLECULES, AND IONS

The content in this topic is the basis for mastering Learning Objectives 1.1, 1.13, and 1.14 as found in the Curriculum Framework.

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:

- Know the basic structure of the atom and the key experimental evidence that led scientists to understand the modern atom
- Cite specific experimental evidence that contradicts various atomic models
- Use data from mass spectra to identify elements and individual isotopes
- Justify that, in a pure sample of a compound, the ratio of the masses of its constituent elements is always the same
- Know the difference 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
- Know the structures and names of alkanes, alcohols, and carboxylic acids

The Atomic Theory of Matter

Section 2.1

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 illustrates 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 summarize Dalton's basic ideas.

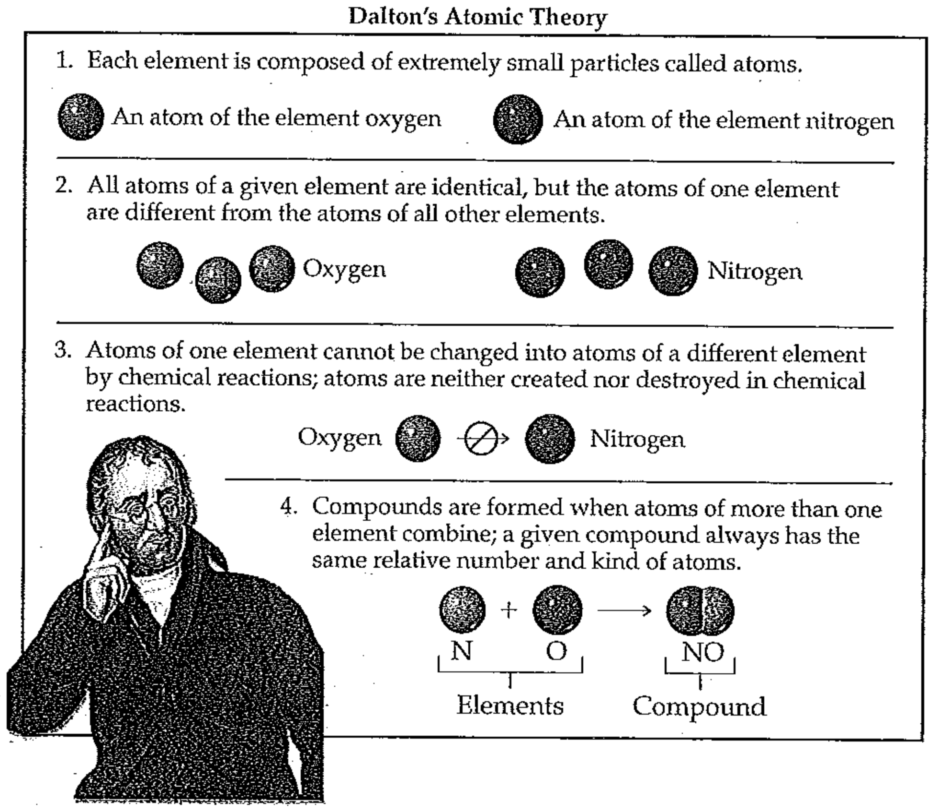

Figure 2.1 Summary of Dalton's atomic theory.

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).

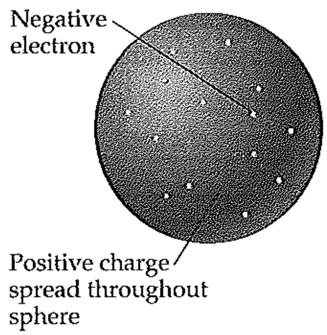

Figure 2.2 J. J. Thomson's plum-pudding model of the atom. Ernest Rutherford proved this model wrong.

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

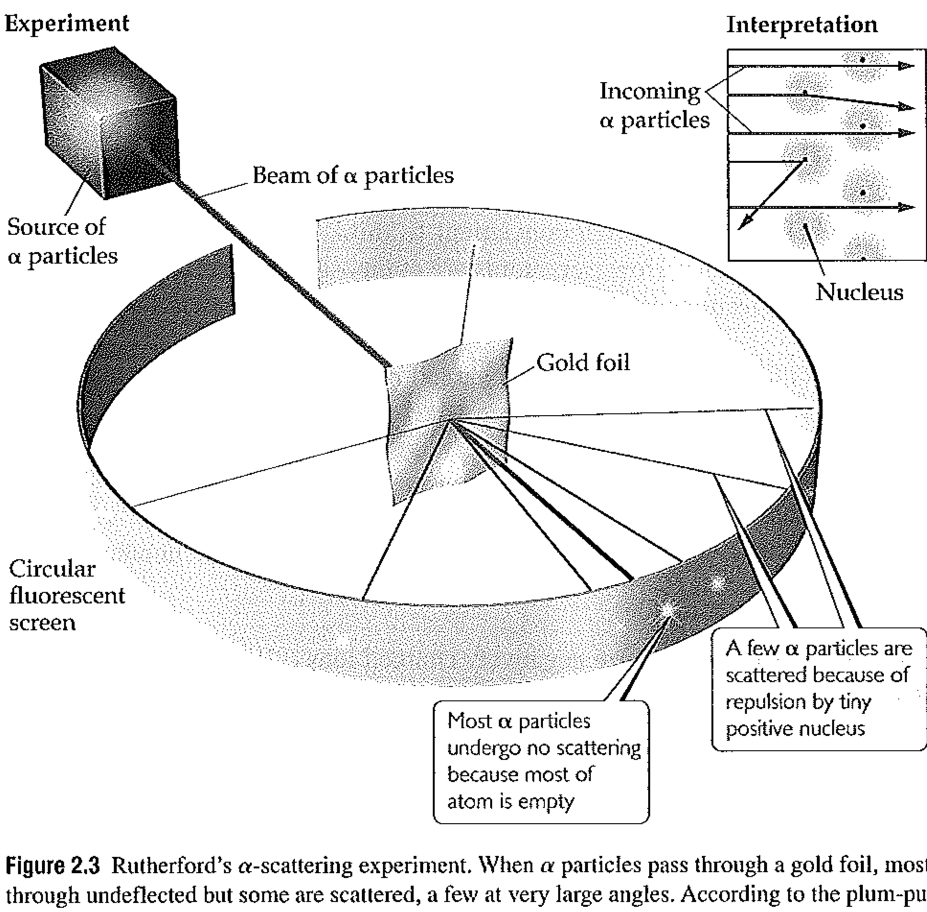

Figure 2.3 Rutherford's α-scattering experiment. When α particles pass through a gold foil, most pass through undeflected but some are scattered, a few at very large angles. According to the plum-pudding model of the atom, the particles should experience only very minor deflections. The nuclear model 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.

proposed that electrons circle the nucleus much like 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 and 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. Write your answer in the space provided.

Section 2.3 The Modern View of Atomic Structure

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 all of whose atoms contain 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:

$$\begin{array}{l} ^{12} = \text{mass number} = \text{number of protons plus neutrons} \\ \mathrm{C} \\ _{6} = \text{atomic number} = \text{number of protons} \end{array}$$

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:

$$^{12}_{6}C = {}^{12}C = \text{carbon-12} = \text{C-12}$$

Carbon has several isotopes and each is distinguished by its mass number. Their respective number of neutrons is calculated by subtracting the atomic number from the mass number. Here are symbols for various isotopes of carbon with the number of neutrons each isotope possesses.

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

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 is the weighted average mass of all the isotopes of an element based on the abundance of each isotope found on the 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 24.3040 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.

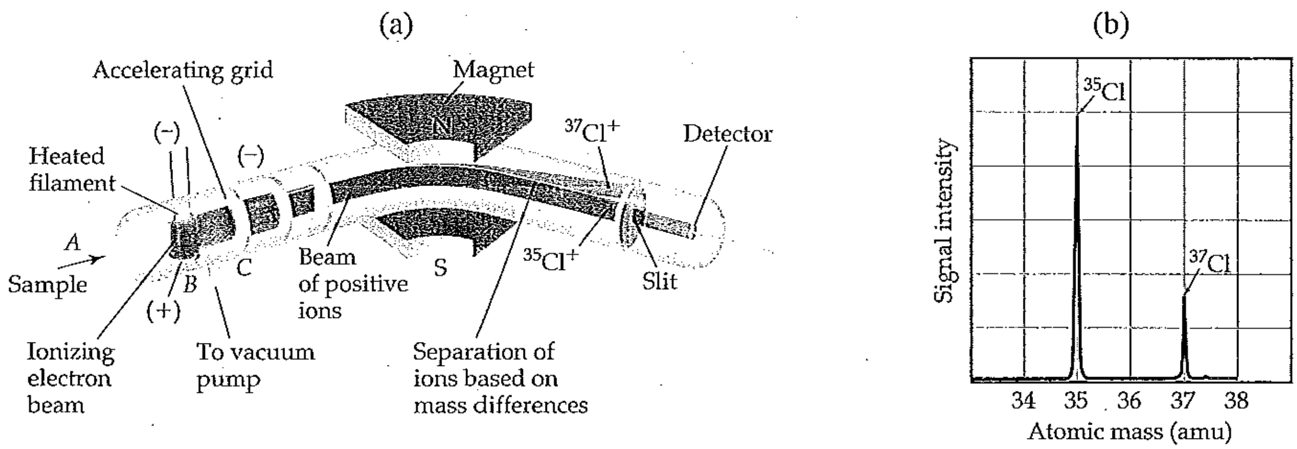

Figure 2.4 (a) A mass spectrometer. Cl atoms are introduced at A and are ionized to form Cl^+ ions, which are then directed through a magnetic field. The paths of the ions of the two Cl isotopes diverge as they pass through the field. (b) Mass spectrum of atomic chlorine. The fractional abundances of the isotopes ^{36}Cl and ^{37}Cl are indicated by the relative signal intensities of the beams reaching the detector of the mass spectrometer.

Mass spectrometers dramatically demonstrate the existence of isotopes and they accurately measure the individual masses of each isotope and their relative abundances.

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

Example:

Detailed analysis of data from the mass spectrum of chlorine shown in Figure 2.5 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.

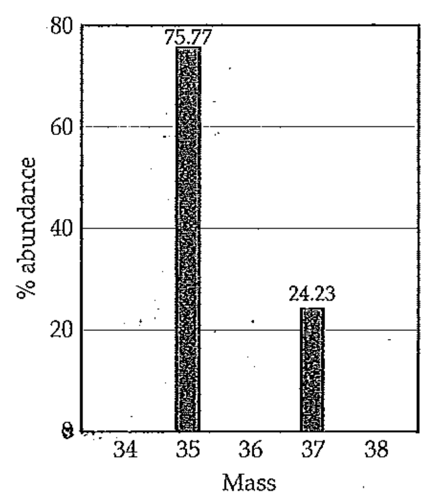

Figure 2.5 Mass spectrum of atomic chlorine.

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.

Average atomic mass =
(*mass of Cl*-35)(*abundance of Cl*-35) + (*mass of Cl*-37)
(*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 radio carbon 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 2.6 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. Write your answer in the space provided.

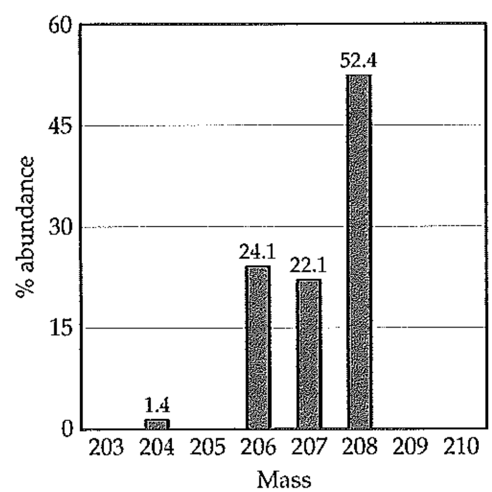

Figure 2.6 The mass spectrum of lead.

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 2.7 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.

Main groups | Transition metals | Main groups

	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 9.012182											5 B 10.811	6 C 12.0107	7 N 14.0067	8 O 15.9994	9 F 18.998403	10 Ne 20.1797
3	11 Na 22.989770	12 Mg 24.3050											13 Al 26.981538	14 Si 28.0855	15 P 30.973761	16 S 32.065	17 Cl 35.453	18 Ar 39.948
4	19 K 39.0933	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	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 78.96	35 Br 79.904	36 Kr 83.80
5	37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc [98]	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	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 132.90545	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 196.96655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 203.98038	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 [261.11]	105 Db [262.11]	106 Sg [266.12]	107 Bh [264.12]	108 Hs [269.13]	109 Mt [268.14]	110 Ds [281.15]	111 Rg [272.15]	112 Cn [285]	113 [284]	114 Fl [289]	115 [288]	116 Lv [292]	117 [294]	118 [294]

*Lanthanide series	57 *La 138.9055	58 Ce 140.116	59 Pr 140.90765	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 164.93032	68 Er 167.259	69 Tm 168.93421	70 Yb 173.04
†Actinide series	89 †Ac [227.03]	90 Th 232.0381	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.
The names and symbols for elements 110 and above have not yet been decided.
Atomic weights in brackets are the masses of the longest-lived or most important isotope of radioactive elements.
Further information is available at http://www.webelements.com
The production of element 116 was reported in May 1999 by scientists at Lawrence Berkeley National Laboratory.

Figure 2.7 The periodic table of the elements.

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

Molecules and Molecular Compounds

Section 2.6

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 2.8 shows the names, formulas, and pictorial representations of some simple molecules.

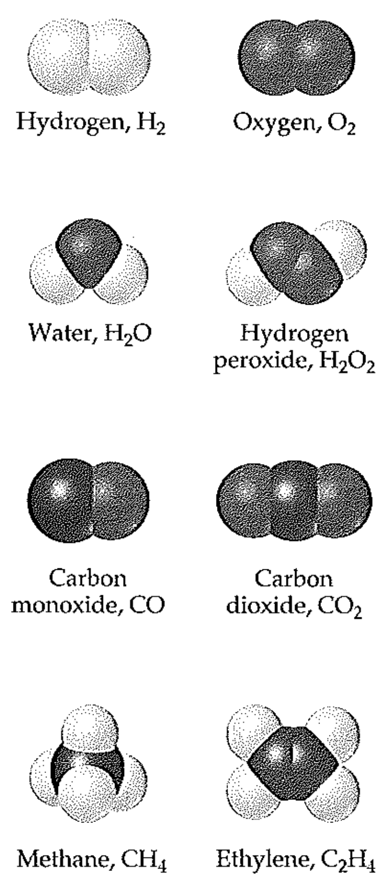

Figure 2.8 Molecular models. Notice how the chemical formulas of these simple molecules correspond to their compositions.

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.

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. Write your answer in the space provided.

Compounds are substances consisting of two or more different 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 2.9 shows various ways chemists represent molecular compounds.

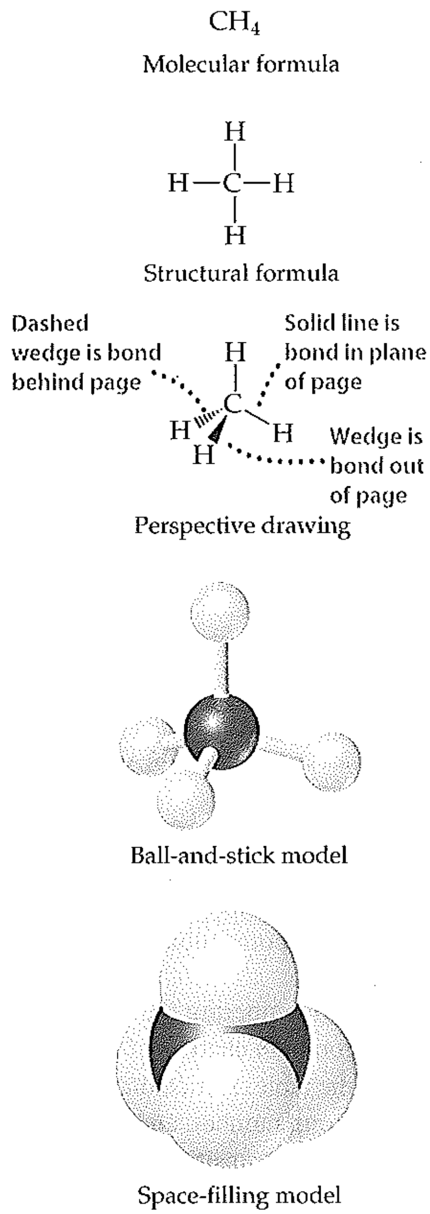

Figure 2.9 Different representations of the methane (CH_4) molecule. Structural formulas, perspective drawings, ball-and-stick models, and space-filling models correspond to the molecular formula, and each helps us visualize the three-dimensional arrangement of atoms.

Ions and Ionic Compounds — Section 2.7

Ions are charged particles composed of single atoms (called **monatomic ions**) or aggregates of atoms (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.

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 2.10. Rather than draw the ionic arrangement of sodium chloride, chemists use the much simpler empirical formula, NaCl, to represent the compound.

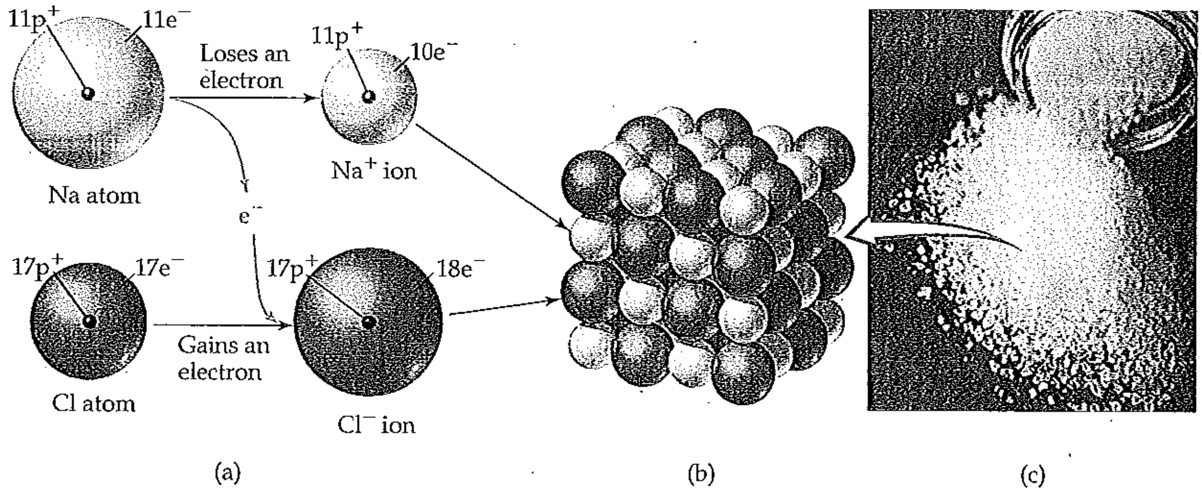

Figure 2.10 Formation of an ionic compound. (a) The transfer of an electron from a Na atom to a Cl atom leads to the formation of a Na^+ ion and a Cl^- ion. (b) Arrangement of these ions in solid sodium chloride, NaCl. (c) A sample of NaCl crystals.

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 charges of monatomic cations and anions. Figure 2.11 shows the common 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.

1A	2A	3B	4B	5B	6B	7B	8B			9B	10B	3A	4A	5A	6A	7A	8A
H^+																H^-	
Li^+														N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}			S^{2-}	Cl^-	
K^+	Ca^{2+}	Sc^{3+}	Ti^{2+} Ti^{4+}	V^{2+} V^{4+}	Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{4+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^{2+} Cu^+	Zn^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}									Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}		Te^{2-}	I^-	
Cs^+	Ba^{2+}									Au^+ Au^{3+}	Hg^{2+} Hg_2^{2+}		Pb^{2+} Pb^{4+}				

Figure 2.11 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.

Naming Inorganic Compounds — Section 2.8

Names of Monatomic Cations

Monatomic cations have the same name as the metal. If the metal forms cations of different charges, a Roman numeral in the name indicates the charge.

K^+ = potassium ion Ca^{2+} = calcium ion
Co^{2+} = cobalt(II) ion Co^{3+} = cobalt(III) ion

Notice that most transition metals form cations with more than one charge. However there are four common exceptions.

Sc^{3+} = scandium ion Ag^+ = silver ion
Zn^{2+} = zinc ion Cd^{2+} = cadmium ion

Most non-transition metals form cations having only a single charge. Note the two common exceptions.

Sn^{2+} = tin(II) ion Sn^{4+} = tin(IV) ion
Pb^{2+} = lead(II) ion Pb^{4+} = lead(IV) ion

Names of Monatomic Anions

Monatomic anions replace the end of the element name with -ide as shown in Table 2.2.

Table 2.2. Names and formulas of common monatomic ions.

Group 1A	Group 5A	Group 6A	Group 7A
H^- hydride			
	N^{3-} nitride	O^{2-} oxide	F^- fluoride
		S^{2-} sulfide	Cl^- chloride
		Se^{2-} selenide	Br^- bromide
		Te^{2-} telluride	I^- iodide

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.

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. Because there is no logical way to predict their formulas or charges these must be memorized.

Table 2.3 lists the names and formulas of some common oxyanions (polyatomic anions containing oxygen).

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 2.3 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^-

Oxyanions are polyatomic anions that usually end in -ate or -ite. 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.

The halogens form a series of four oxyanions and use prefixes to distinguish them.

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

Names of Ionic Compounds

Names of ionic compounds consist of the cation name followed by the anion name.

$MgBr_2$ = magnesium bromide
$Ca_3(PO_4)_2$ = calcium phosphate
Fe_2O_3 = iron(III) oxide
FeO = iron(II) oxide

(Remember that the Roman numerals III and II refer to the 3+ and 2+ charges of the iron cations, respectively. Notice that the charge on the iron atom in each case is inferred by the ratio in which iron combines with the 2− charged oxide ion.)

Your Turn 2.4

Name the following ionic compounds: KI, $MgSO_4$, FeS, Al_2O_3, $Pb_3(PO_4)_2$. Write your answers in the space provided.

Names and Formulas of Acids

The names and formulas of some common acids, an important class of hydrogen-containing compounds, are listed in Table 2.4. 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 2.4 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 giving it an -ide ending.
3. Use prefixes to denote how many of each element are present in the formula. The prefixes are shown in Table 2.5.

Table 2.5 Prefixes used in naming binary compounds formed between nonmetals.

Prefix	Meaning
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

CO_2 = carbon dioxide

SO_3 = sulfur trioxide

N_2O_5 = dinitrogen pentoxide

P_4O_{10} = tetraphosphorus decoxide

Your Turn 2.5

Name the following compounds: SCl_2, CF_4, BrI_3, PBr_5, SF_6. Write your answers in the space provided.

Some Simple Organic Compounds

Section 2.9

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 2.6 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 prefix is followed by the suffix -ane to indicate that the formula is an alkane.

Table 2.6 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 give rise to the structure and properties of an organic compound. For example, the functional group of a class of organic compounds classified as **alcohols** is —OH. The functional group of a **carboxylic acid** is —COOH. Table 2.7 lists the names and formulas of simple alcohols and carboxylic acids. In Chapter 25 of *Chemistry: The Central Science*, you will study organic compounds in more detail.

Table 2.7 Names and formulas of some simple alcohols and carboxylic acids. The prefix tells how many carbon atoms are in the formulas.

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	$CH_3CH_2CHOHCH_3$		