2

ATOMS, MOLECULES, AND IONS

WHAT'S AHEAD

2.1 | The Atomic Theory of Matter

Take a moment to appreciate the great variety of colors, textures, and other properties in the materials that surround you—the array of colors in a flower, the texture of the fabric in your clothes, the solubility of sugar in a cup of coffee or the transparency and beauty of a diamond. How can we explain the striking and seemingly infinite variety of properties of the materials that make up our world? What makes diamond transparent and hard whereas table salt is brittle and dissolves in water? Aluminum conducts electricity, but aluminum oxide does not. Paper burns in the presence of oxygen gas but not in the presence of nitrogen gas. What accounts for these differences? The answers to all such questions lie in the structures of atoms, which determine the physical and chemical properties of matter.

Although materials vary greatly in their properties, everything is formed from only about 100 different kinds of atoms. In a sense, these different atoms are like the 26 letters of the English alphabet that join in different combinations to form the immense number of words in our language. But what rules govern the ways in which atoms combine? How do the properties of a substance relate to the kinds of atoms it contains? Indeed, what is an atom like, and what makes the atoms of one element different from those of another?

A helicopter engine is composed of many smaller parts, just as any substance on Earth is composed of countless atoms and molecules to give it its unique characteristics.

In this chapter, we introduce the basic structure of atoms and discuss the formation of molecules and ions. This knowledge provides you with the foundation you need to understand the chapters that follow.

By the end of this section, you should be able to

- Understand Dalton's postulates.

Philosophers from the earliest times speculated about the nature of the fundamental "stuff" from which the world is made. Democritus (460–370 BCE) and other early Greek philosophers described the material world as made up of tiny indivisible particles that they called *atomos*, meaning "indivisible" or "uncuttable." Later, however, Plato and Aristotle formulated the notion that there can be no ultimately indivisible particles, and the "atomic" view of matter faded for many centuries during which Aristotelean philosophy dominated Western culture.

The notion of **atoms** reemerged in Europe during the seventeenth century. As chemists learned to measure the amounts of elements that reacted with one another to form new substances, the ground was laid for an atomic theory that linked the idea of elements with the idea of atoms. That theory came from the work of John Dalton during the period from 1803 to 1807. Dalton's atomic theory was based on four postulates (see **Figure 2.1**).

A good theory explains known facts; Dalton's theory explained several laws of chemical combination known at the time.

- The *law of constant composition*, based on postulate 4:

  In a given compound, the relative numbers and kinds of atoms are constant.

- The **law of conservation of mass,** based on postulate 3:

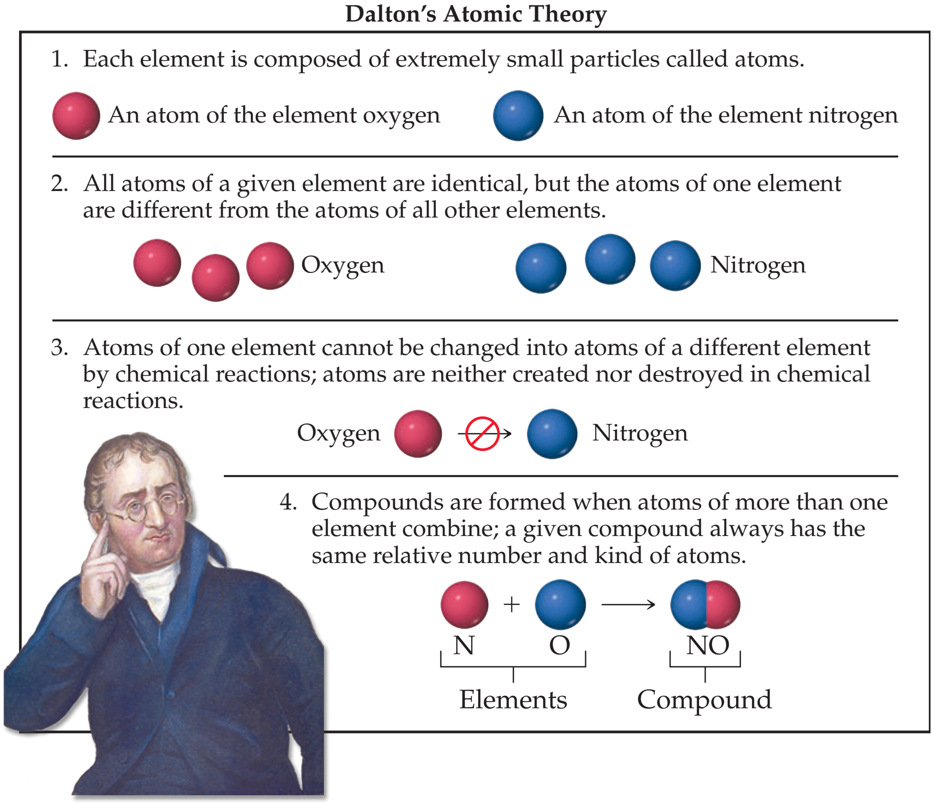
  The total mass of materials present after a chemical reaction is the same as the total mass present before the reaction.

A good theory also predicts new facts; Dalton used his theory to deduce

- The **law of multiple proportions:**

  If two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in the ratio of small whole numbers.

We can illustrate this law by considering water and hydrogen peroxide, both of which consist of the elements hydrogen and oxygen. In forming water, 8.0 g of oxygen combines with 1.0 g of hydrogen. In forming hydrogen peroxide, 16.0 g of oxygen combines with 1.0 g of hydrogen. Thus, the ratio of the masses of oxygen per gram of hydrogen in the two compounds is 2:1. Using Dalton's atomic theory, we conclude that hydrogen peroxide contains twice as many atoms of oxygen per hydrogen atom than does water.



◀ **Figure 2.1 Dalton's atomic theory.*** John Dalton (1766–1844), the son of a poor English weaver, began teaching at age 12. He spent most of his years in Manchester, where he taught both grammar school and college. His lifelong interest in meteorology led him to study gases, then chemistry, and eventually atomic theory. Despite his humble beginnings, Dalton gained a strong scientific reputation during his lifetime.

Self-Assessment Exercise

2.1 In an experiment, 7.0 g of nitrogen reacted with exactly 16.0 g of oxygen to form a single compound. What would be the total mass of the compound?

(a) 7.0 g

(b) 16 g

(c) 23 g

Exercises

2.2 A 1.0-g sample of carbon dioxide (CO_2) is fully decomposed into its elements, yielding 0.273 g of carbon and 0.727 g of oxygen. (**a**) What is the ratio of the mass of O to C? (**b**) If a sample of a different compound decomposes into 0.429 g of carbon and 0.571 g of oxygen, what is its ratio of the mass of O to C? (**c**) According to Dalton's atomic theory, what is the empirical formula of the second compound?

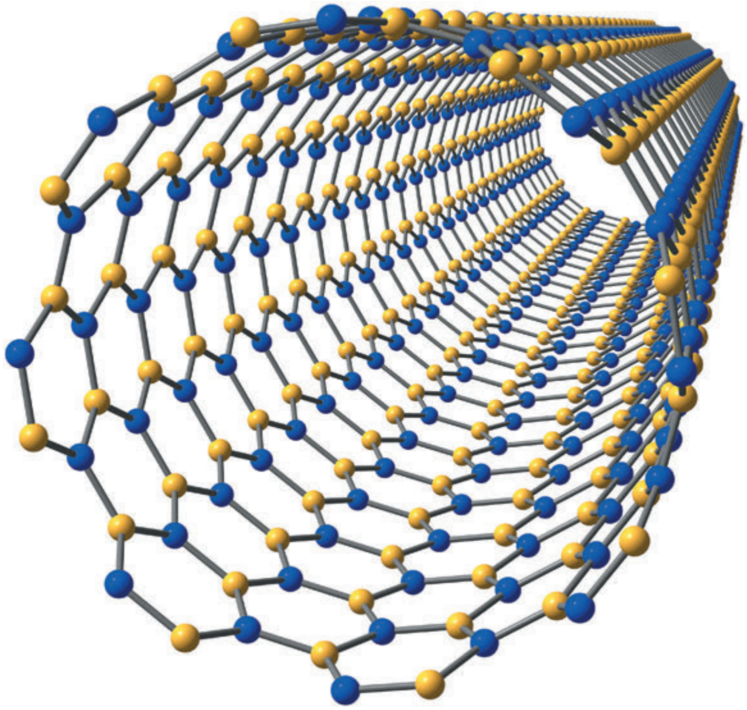
2.3 A chemist finds that 30.82 g of nitrogen will react with 17.60, 35.20, 70.40, or 88.00 g of oxygen to form four different compounds. (**a**) Calculate the mass of oxygen per gram of nitrogen in each compound. (**b**) How do the numbers in part (a) support Dalton's atomic theory?

Answers to Self-Assessment Exercises

2.1 (c)

*Dalton, "Atomic Theory" 1844.

2.2 | The Discovery of Atomic Structure



Dalton based his conclusions about atoms on chemical observations made in the laboratory. By assuming the existence of atoms, he was able to account for the laws of constant composition and of multiple proportions. But neither Dalton nor those who followed him during the century after his work was published had any direct evidence for the existence of atoms. Today, however, we can measure the properties of individual atoms and even provide images of them. The picture was obtained by a technique called scanning tunneling microscopy. The color was added to the image by computer to help distinguish its features. Each gold sphere is a silicon atom.

By the end of this section, you should be able to

- Describe the main experiments that led to the discovery of the electron and to the nuclear model of the atom

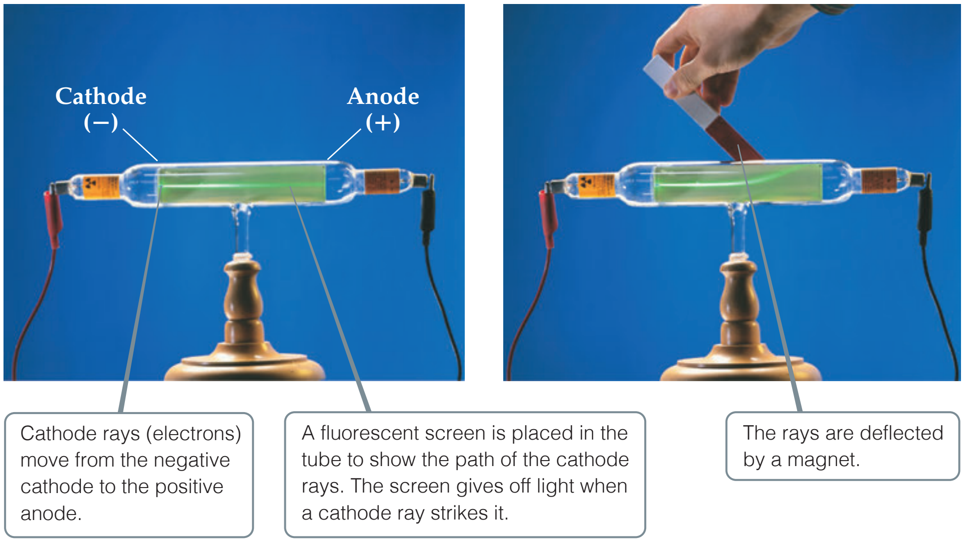
As scientists developed methods for probing the nature of matter, the supposedly indivisible atom began to show signs of a more complex structure, and today we know that the atom is composed of **subatomic particles**. Before we summarize the current model, we briefly consider a few of the landmark discoveries that led to that model. We will see that the atom is composed in part of electrically charged particles, some with a positive charge and some with a negative charge. As we discuss the development of our current model of the atom, keep in mind this fact: *Particles with the same charge repel one another, whereas particles with opposite charges attract one another*.

Cathode Rays and Electrons

During the mid-1800s, scientists began to study electrical discharge through a glass tube pumped almost empty of air (**Figure 2.2**). When a high voltage was applied to the electrodes in the tube, radiation was produced between the electrodes. This radiation, called **cathode rays**, originated at the negative electrode and traveled to the positive electrode. Although the rays could not be seen, their presence was detected because they cause certain materials to *fluoresce,* or to give off light.

Experiments showed that cathode rays are deflected by electric or magnetic fields in a way consistent with there being a stream of negative electrical charge. The British scientist J. J. Thomson (1856–1940) observed that cathode rays are the same regardless of the identity of the cathode material. In a paper published in 1897, Thomson described cathode rays as streams of negatively charged particles that we now call **electrons**.

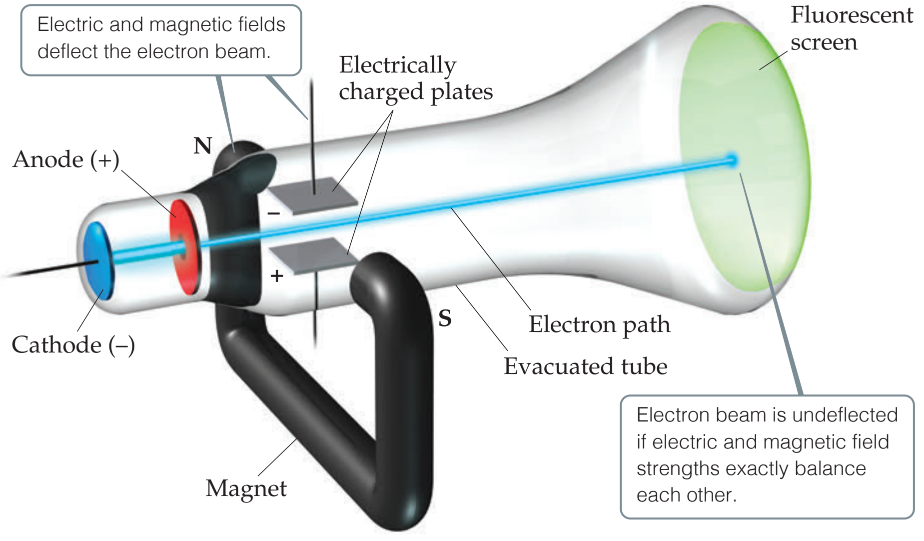
Go Figure If the fluorescent screen were removed from the tube, would cathode rays still be generated? Would you be able to see them?



▲ **Figure 2.2 Cathode-ray tube.**

Thomson constructed a cathode-ray tube having a hole in the anode through which the cathode rays could pass. Electrically charged plates and a magnet were positioned perpendicular to the beam, and a fluorescent screen that would give off light when struck with a cathode ray was located at one end (**Figure 2.3**). Because the electron is a negatively charged particle, the electric field deflected the rays in one direction, and the magnetic field deflected them in the opposite direction. Thomson adjusted the strengths of the fields so that the effects balanced each other, allowing the electrons to travel in a straight path to the screen. Knowing the strengths that resulted in the straight path made it possible to calculate a value of 1.76×10^8 coulombs* per gram for the ratio of the electron's electrical charge to its mass.

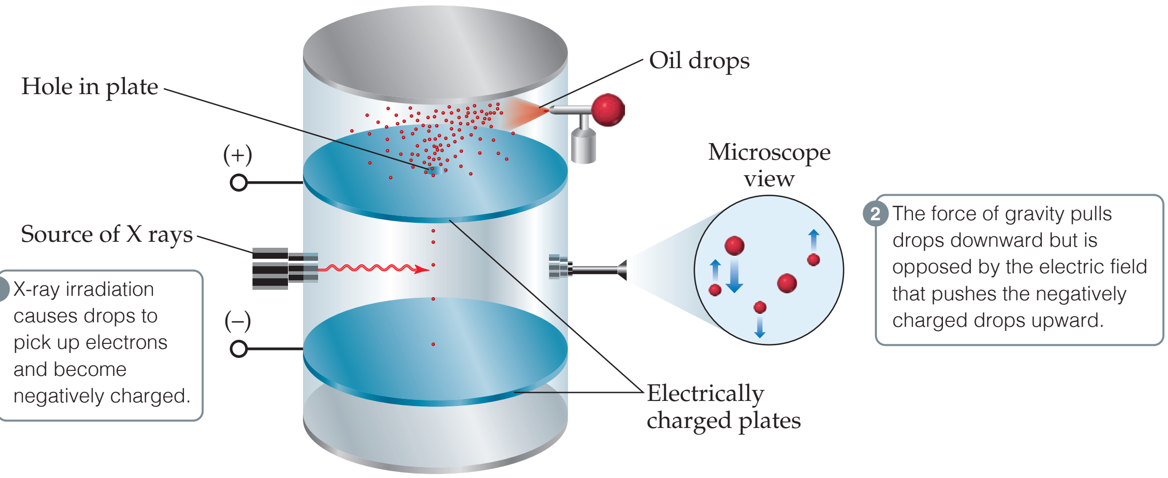
Go Figure If no magnetic field were applied, would you expect the electron beam to be deflected upward or downward by the electric field?



▲ **Figure 2.3 Cathode-ray tube with perpendicular magnetic and electric fields.** The cathode rays (electrons) originate at the cathode and are accelerated toward the anode, which has a hole in its center. A narrow beam of electrons passes through the hole and travels to the fluorescent screen that glows when struck by a cathode ray.

*The coulomb (C) is the SI unit for electrical charge.

Go Figure Are the masses of the oil drops changed significantly when electrons accumulate on them?



▲ **Figure 2.4 Millikan's oil-drop experiment to measure the charge of the electron.** Small drops of oil are allowed to fall between electrically charged plates. Millikan measured how varying the voltage between the plates affected the rate of fall. From these data he calculated the negative charge on the drops. Because the charge on any drop was always some integral multiple of 1.602×10^{-19} C, Millikan deduced this value to be the charge of a single electron.

Once the charge-to-mass ratio of the electron was known, measuring either quantity allowed scientists to calculate the other. In 1909, Robert Millikan (1868–1953) of the University of Chicago succeeded in measuring the charge of an electron by performing the experiment described in **Figure 2.4**. He then calculated the mass of the electron by using his experimental value for the charge, 1.602×10^{-19} C, and Thomson's charge-to-mass ratio, 1.76×10^{8} C/g:

$$\text{Electron mass} = \frac{1.602 \times 10^{-19}\,\cancel{C}}{1.76 \times 10^{8}\,\cancel{C}/g} = 9.10 \times 10^{-28}\,g$$

This result agrees well with the currently accepted value for the electron mass, 9.10938×10^{-28} g. This mass is about 2000 times smaller than that of hydrogen, the lightest atom.

▲ **Figure 2.5 Marie Sklodowska Curie (1867–1934).** In 1903, Henri Becquerel, Marie Curie, and her husband, Pierre, were jointly awarded the Nobel Prize in Physics for their pioneering work on radioactivity (a term she introduced). In 1911, Marie Curie won a second Nobel Prize, this time in chemistry for her discovery of the elements polonium and radium.

Radioactivity

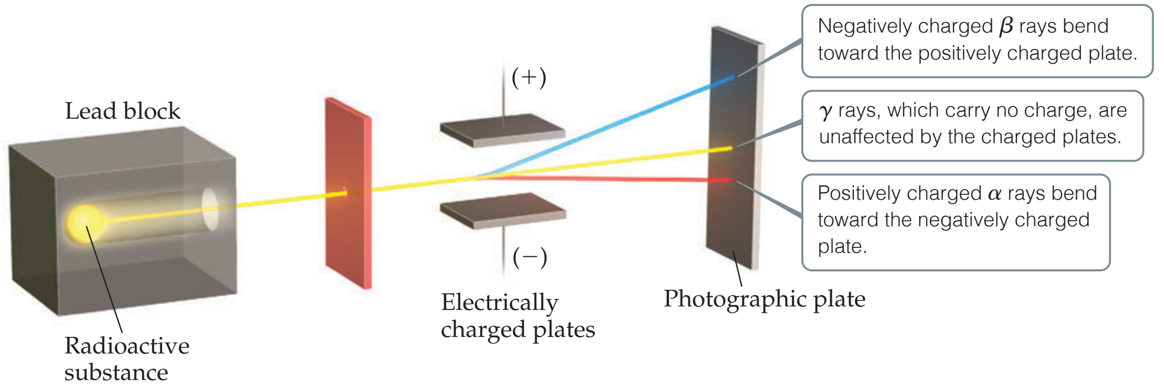
In 1896 the French scientist Henri Becquerel (1852–1908) discovered that a compound of uranium spontaneously emits high-energy radiation. This spontaneous emission of radiation is called **radioactivity**. At Becquerel's suggestion, Marie Curie (**Figure 2.5**) and her husband, Pierre, began experiments to identify and isolate the source of radioactivity in the compound. They concluded that it was the uranium atoms.

Further study of radioactivity, principally by the British scientist Ernest Rutherford, revealed three types of radiation: alpha (α), beta (β), and gamma (γ). Rutherford (1871–1937) was a very important figure in this period of atomic science. After working at Cambridge University with J. J. Thomson, he moved to McGill University in Montreal, where he did research on radioactivity that led to his 1908 Nobel Prize in Chemistry. In 1907 he returned to England as a faculty member at Manchester University, where he did his famous α-particle scattering experiments, described further in this chapter.

Rutherford showed that the paths of α and β radiation are bent by an electric field, although in opposite directions; while γ radiation is unaffected by the field (**Figure 2.6**). From this finding he concluded that α and β rays consist of fast-moving electrically

Go Figure

Which subatomic particle—proton, neutron, or electron—is equivalent to a β ray?
β rays are deflected to a greater extent than α rays because (**a**) they are lighter, or (**b**) they are more highly charged.



▲ **Figure 2.6 Behavior of alpha (α), beta (β), and gamma (γ) rays in an electric field.**

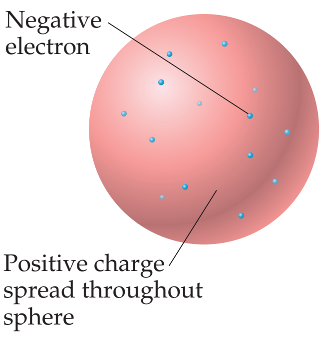
charged particles. In fact, β particles are nothing more than high-speed electrons that can be considered the radioactive equivalent of cathode rays. Because of their negative charge, they are attracted to a positively charged plate. The α particles have a positive charge and are attracted to a negative plate. In units of the charge of the electron, β particles have a charge of 1− and α particles a charge of 2+. Each α particle has a mass about 7400 times that of an electron. Gamma radiation is high-energy electromagnetic radiation similar to X rays; it does not consist of particles and it carries no charge.

The Nuclear Model of the Atom

With growing evidence that the atom is composed of smaller particles, scientists gave attention to how the particles fit together. During the early 1900s, Thomson reasoned that because electrons contribute only a very small fraction of an atom's mass, they probably are responsible for an equally small fraction of the atom's size. He proposed that the atom consists of a uniform positive sphere of matter in which the mass is evenly distributed and in which the electrons are embedded like raisins in a pudding or seeds in a watermelon (**Figure 2.7**). This *plum-pudding model,* named after a traditional English dessert, was very short-lived.

In 1910, Rutherford was studying the angles at which α particles were deflected, or *scattered*, as they passed through a thin sheet of gold foil (**Figure 2.8**). He discovered that almost all the particles passed directly through the foil without deflection, with a few particles deflected about 1°, consistent with Thomson's plum-pudding model. For the sake of completeness, Rutherford suggested that Ernest Marsden (1889–1970), an undergraduate student working in the laboratory, look for scattering at large angles. To everyone's surprise, a small amount of scattering was observed at large angles, with some particles scattered back in the direction from which they had come. The explanation for these results was not immediately obvious, but they were clearly inconsistent with Thomson's plum-pudding model.

Rutherford explained the results by postulating the **nuclear model** of the atom, in which most of the mass of each gold atom and all of its positive charge reside in a very small, extremely dense region that he called the **nucleus**. He postulated further that most of the volume of an atom is empty space in which electrons move around the nucleus. In the α-scattering experiment, most of the particles passed through the foil unscattered because they did not encounter the minute nucleus of any gold atom. Occasionally, however, an α particle came close to a gold nucleus. In such

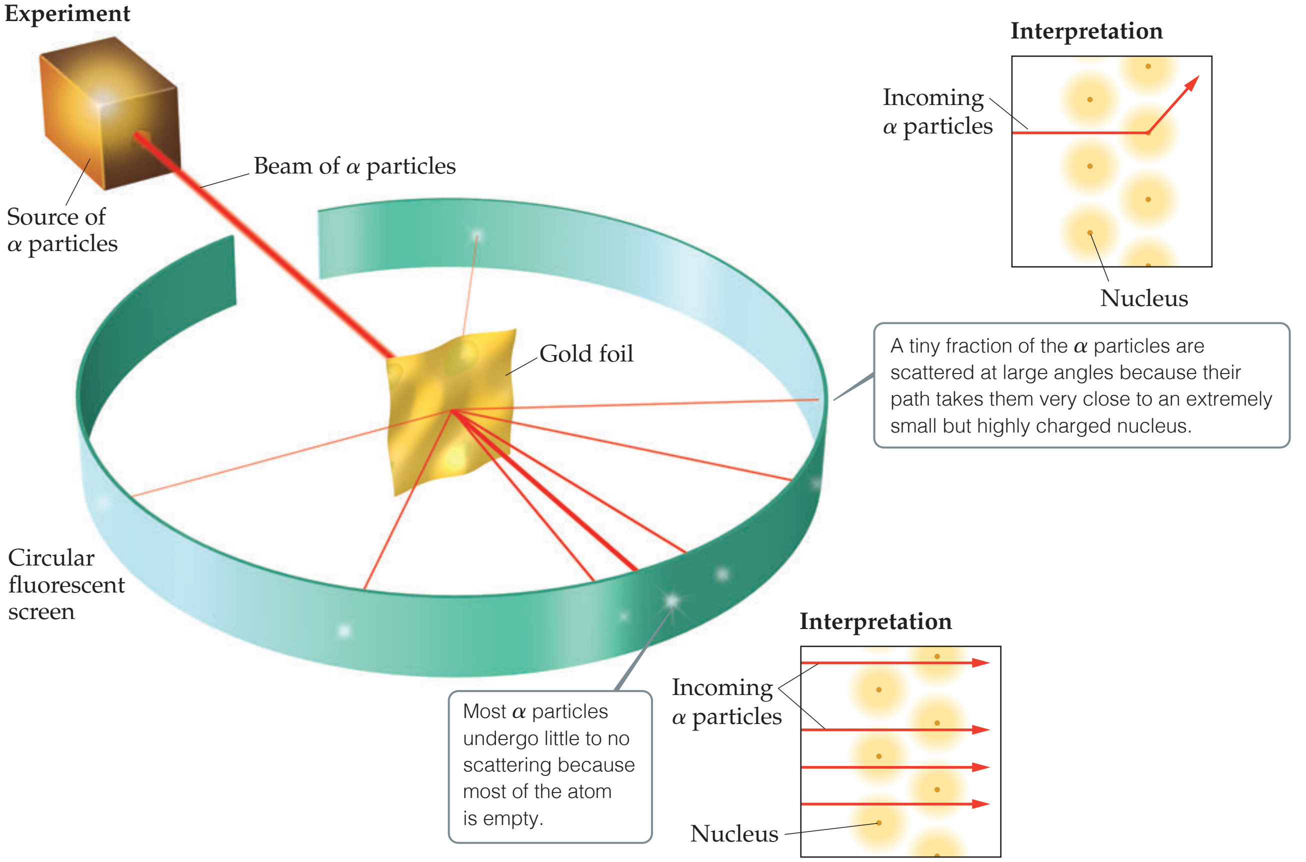


▲ **Figure 2.7 J. J. Thomson's plum-pudding model of the atom.** Ernest Rutherford and Ernest Marsden proved this model wrong.

encounters, the repulsion between the highly positive charge of the gold nucleus and the positive charge of the α particle was strong enough to deflect the particle, as shown in Figure 2.8.

Subsequent experiments led to the discovery of positive particles (**protons**) and neutral particles (**neutrons**) in the nucleus. Protons were discovered in 1919 by Rutherford and neutrons in 1932 by British scientist James Chadwick (1891–1972). Thus, the atom is composed of electrons, protons, and neutrons.

Go Figure What is the charge on the particles that form the beam? Will they be attracted to or repelled from the positively charged gold nuclei?



▲ **Figure 2.8 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. Although the nuclear atom has been depicted here as a yellow sphere, it is important to realize that most of the space around the nucleus contains only the low-mass electrons.

Self-Assessment Exercise

2.4 Which experiment enabled the mass of an electron to be calculated?

(a) J J Thompson's cathode ray tube experiment

(b) R Millikan's oil drop experiment

(c) E Rutherford's gold foil experiment

Exercises

2.5 An unknown particle is caused to move between two electrically charged plates, as illustrated in Figure 2.6. You hypothesize that the particle is a proton. (**a**) If your hypothesis is correct, would the particle be deflected in the same or opposite direction as the β rays? (**b**) Would it be deflected by a smaller or larger amount than the β rays?

2.6 What fraction of the α particles in Rutherford's gold foil experiment are scattered at large angles? Assume the gold foil is two layers thick, as shown in Figure 2.8, and that the approximate diameters of a gold atom and its nucleus are 270 pm and 1.0×10^{-2} pm, respectively. *Hint:* Calculate the cross sectional area occupied by the nucleus as a fraction of that occupied by the atom. Assume that the gold nuclei in each layer are offset from each other.

Answers to Self-Assessment Exercises

2.4 (b)

2.3 | The Modern View of Atomic Structure

Ernest Rutherford (1871–1937), whom Einstein called 'the second Newton', was born and educated in New Zealand. In 1895, he was the first overseas student ever to be awarded a position at the Cavendish Laboratory at Cambridge University in England, where he worked with J J Thompson. In 1898, he joined the faculty of McGill University in Canada, where he did the research on radioactivity that led to his being awarded the 1908 Nobel Prize in Chemistry. In 1907, he moved back to England to join Manchester University where, in 1910, he performed his famous α-particle scattering experiments that led to the nuclear model of the atom. In 1992, his native New Zealand honored Rutherford by putting his likeness, along with his Nobel Prize medal, on its $100 currency note. Element 104 is named in his honor.

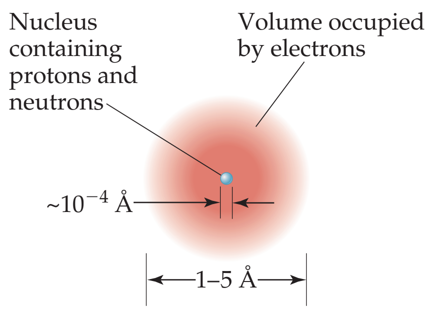
By the end of this section, you should be able to

- Describe the structure of the atom in terms of its subatomic particles
- Use symbols to indicate the composition of an isotope

Since Rutherford's time, as physicists have learned more and more about atomic nuclei, the list of particles that make up nuclei has grown and continues to increase. As chemists, however, we can take a simple view of the atom because only three subatomic particles—the proton, neutron, and electron—have a bearing on chemical behavior.

Go Figure

What is the approximate diameter of the nucleus in units of pm?



▲ **Figure 2.9 The structure of the atom.** A cloud of rapidly moving electrons occupies most of the volume of the atom. The nucleus occupies a tiny region at the center of the atom and is composed of the protons and neutrons. The nucleus contains virtually all the mass of the atom.

As noted earlier, the charge of an electron is -1.602×10^{-19} C. The charge of a proton is opposite in sign but equal in magnitude to that of an electron: $+1.602 \times 10^{-19}$ C. The quantity 1.602×10^{-19} C is called the **electronic charge**. For convenience, the charges of atomic and subatomic particles are usually expressed as multiples of this charge rather than as coulombs. Thus, the charge of an electron is 1– and that of a proton is 1+. Neutrons are electrically neutral (which is how they received their name). *Every atom has an equal number of electrons and protons, so atoms have no net electrical charge.*

Protons and neutrons reside in the tiny nucleus of the atom. The vast majority of an atom's volume is the space in which the electrons reside (Figure 2.9). Most atoms have diameters between 30 pm and 300 pm. The diameter of a chlorine atom, for example, is 175 pm.

Electrons are attracted to the protons in the nucleus by the electrostatic force that exists between particles of opposite electrical charge. In later chapters, we will see that the strength of the attractive forces between electrons and nuclei can be used to explain many of the differences among different elements.

Atoms have extremely small masses. The mass of the heaviest known atom, for example, is approximately 4×10^{-22} g. Because it would be cumbersome to express such small masses in grams, we use the **atomic mass unit** (u), where $1 \text{ u} = 1.66054 \times 10^{-24}$ g. A proton has a mass of 1.0073 u, a neutron 1.0087 u, and an electron 5.486×10^{-4} u (Table 2.1). Because it takes 1836 electrons to equal the mass of one proton, and 1839 electrons to equal the mass of a single neutron, the nucleus accounts for nearly the entire mass of an atom.

TABLE 2.1 Comparison of the Proton, Neutron, and Electron

Particle	Charge	Mass (u)
Proton	Positive (1+)	1.0073
Neutron	None (neutral)	1.0087
Electron	Negative (1−)	5.486×10^{-4}

The diameter of an atomic nucleus is approximately 1^{-10} fm, only a small fraction of the diameter of the atom as a whole. You can appreciate the relative sizes of the atom and its nucleus by imagining that if the hydrogen atom were as large as a football stadium, the nucleus would be the size of a small marble. Because the tiny nucleus carries most of the mass of the atom in such a small volume, it has an incredibly high density—on the order of 10^{13}–10^{14} g/cm^3. A matchbox full of material of such density would weigh over 2.5 billion tons!

Figure 2.9 incorporates the features we have just discussed. Electrons play the major role in chemical reactions. The significance of representing the region containing electrons as an indistinct cloud will become clear in later chapters when we consider the energies and spatial arrangements of the electrons. For now, however, we have all the information we need to discuss many topics that form the basis of everyday uses of chemistry.

Atomic Numbers, Mass Numbers, and Isotopes

What makes an atom of one element different from an atom of another element? The atoms of each element have a *characteristic number of protons*. The number of protons in an atom of any particular element is called that element's **atomic number**. Because an atom has no net electrical charge, the number of electrons it contains must equal the number of protons. All atoms of carbon, for example, have six protons and six electrons, whereas all atoms of oxygen have eight protons and eight electrons. Thus, carbon has atomic number 6, and oxygen has atomic number 8. The atomic number of each element is listed with the name and symbol of the element on the front inside cover of the text.

Atoms of a given element can differ in the number of neutrons they contain and, consequently, in mass. For example, while most atoms of carbon have six neutrons, some have more and some have less. The symbol $^{12}_{6}\text{C}$ (read "carbon twelve," carbon-12) represents the carbon atom containing six protons and six neutrons, whereas carbon atoms

Sample Exercise 2.1
Atomic Size

The diameter of a small coin is 17.9 mm, and the diameter of a silver atom is 288 pm. How many silver atoms could be arranged side by side across the diameter of the coin?

SOLUTION

The unknown is the number of silver (Ag) atoms. Using the relationship 1 Ag atom = 288 pm as a conversion factor relating number of atoms and distance, we start with the diameter of the coin, first converting this distance into picometers and then using the diameter of the Ag atom to convert distance to number of Ag atoms:

$$\text{Ag atoms} = (17.9\ \cancel{\text{mm}})\left(\frac{10^{-3}\ \cancel{\text{m}}}{1\ \cancel{\text{mm}}}\right)\left(\frac{1\ \text{pm}}{10^{-12}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{Ag atom}}{288\ \text{pm}}\right)$$

$$= 6.22 \times 10^{7}\ \text{Ag atoms}$$

That is, 62.2 million silver atoms could sit side by side across the coin!

Practice Exercise

The diameter of a carbon atom is 1.54 Å. (**a**) Express this diameter in picometers. (**b**) How many carbon atoms could be aligned side by side across the width of a pencil line that is 0.20 mm wide?

A CLOSER LOOK Basic Forces

Four basic forces are known in nature: (1) gravitational, (2) electromagnetic, (3) strong nuclear, and (4) weak nuclear. *Gravitational forces* are attractive forces that act between all objects in proportion to their masses. Gravitational forces between atoms or between subatomic particles are so small that they are of no chemical significance.

Electromagnetic forces are attractive or repulsive forces that act between either electrically charged or magnetic objects. The magnitude of the electric force between two charged particles is given by *Coulomb's law*: $F = kQ_1Q_2/d^2$, where Q_1 and Q_2 are the magnitudes of the charges on the two particles, d is the distance between their centers, and k is a constant determined by the units for Q and d. A negative value for the force indicates attraction, whereas a positive value indicates repulsion. Electric forces are of primary importance in determining the chemical properties of elements.

All nuclei except those of hydrogen atoms contain two or more protons. Because like charges repel, electrical repulsion would cause the protons to fly apart if the *strong nuclear force* did not keep them together. As the name implies, this force can be quite strong but only when particles are extremely close together, as are the protons and neutrons in a nucleus. At this distance, the attractive strong nuclear force is stronger than the positive–positive repulsive electric force and holds the nucleus together.

The *weak nuclear force* is weaker than the electric force and the strong nuclear force but stronger than the gravitational force. We are aware of its existence only because it shows itself in certain types of radioactivity.

Related Exercise: 2.124

that contain six protons and eight neutrons have mass number 14, are represented as ${}^{14}_{6}C$ and are referred to as carbon-14.

The atomic number is indicated by the subscript; the superscript, called the **mass number**, is the number of protons plus neutrons in the atom:

Mass number (number of protons plus neutrons)

${}^{12}_{6}C$ ← Symbol of element

Atomic number (number of protons or electrons)

Because all atoms of a given element have the same atomic number, the subscript is redundant and is often omitted. Thus, the symbol for carbon-12 can be represented simply as ${}^{12}C$.

Atoms with identical atomic numbers but different mass numbers (that is, the same number of protons but different numbers of neutrons) are called **isotopes** of one another. Several isotopes of carbon are listed in Table 2.2. We will generally use the notation with superscripts only when referring to a particular isotope of an element. It is important to keep in mind that the isotopes of any given element are all alike chemically. A carbon dioxide molecule that contains a ${}^{13}C$ atom behaves for all practical purposes identically to one that contains a ${}^{12}C$ atom.

TABLE 2.2 Some Isotopes of Carbon[a]

Symbol	Number of Protons	Number of Electrons	Number of Neutrons
^{11}C	6	6	5
^{12}C	6	6	6
^{13}C	6	6	7
^{14}C	6	6	8

[a] Almost 99% of the carbon found in nature is ^{12}C.

Sample Exercise 2.2
Determining the Number of Subatomic Particles in Atoms

How many protons, neutrons, and electrons are in an atom of (**a**) ^{197}Au, (**b**) strontium-90?

SOLUTION

(**a**) The superscript 197 is the mass number (protons + neutrons). According to the list of elements given on the front inside cover, gold has atomic number 79. Consequently, an atom of ^{197}Au has 79 protons, 79 electrons, and 197 − 79 = 118 neutrons. (**b**) The atomic number of strontium is 38. Thus, all atoms of this element have 38 protons and 38 electrons. The strontium-90 isotope has 90 − 38 = 52 neutrons.

Practice Exercise

Which of these atoms has the largest number of neutrons? (**a**) ^{148}Eu (**b**) ^{157}Dy (**c**) ^{149}Nd (**d**) ^{162}Ho (**e**) ^{159}Gd

Sample Exercise 2.3
Writing Symbols for Atoms

Magnesium has three isotopes with mass numbers 24, 25, and 26. (**a**) Write the complete chemical symbol (superscript and subscript) for each. (**b**) How many neutrons are in an atom of each isotope?

SOLUTION

(**a**) Magnesium has atomic number 12, so all atoms of magnesium contain 12 protons and 12 electrons. The three isotopes are therefore represented by $^{24}_{12}Mg$, $^{25}_{12}Mg$, and $^{26}_{12}Mg$. (**b**) The number of neutrons in each isotope is the mass number minus the number of protons. The numbers of neutrons in an atom of each isotope are therefore 12, 13, and 14, respectively.

Practice Exercise

Give the complete chemical symbol for the atom that contains 82 protons, 82 electrons, and 126 neutrons.

Self-Assessment Exercises

2.7 What subatomic particle(s) account for the vast majority of the mass of an atom of helium?

(a) Electrons
(b) Protons
(c) Neutrons
(d) Electrons and protons
(e) Protons and neutrons

2.8 What is the symbol for the isotope of oxygen-18?

(a) $^{10}_{8}O$
(b) $^{18}_{8}O$
(c) $^{10}_{18}O$

Exercises

2.9 The radius of an atom of tungsten (W) is about 2.10 units. (**a**) Express this distance in nanometers (nm) and in picometers (pm). (**b**) How many tungsten atoms would have to be lined up to create a wire of 2.0 mm? (**c**) If the atom is assumed to be a sphere, what is the volume in m^3 of a single W atom?

2.10 Answer the following questions without referring to Table 2.1: (**a**) What are the main subatomic particles that make up the atom? (**b**) What is the relative charge (in multiples of the electronic charge) of each of the particles? (**c**) Which of the particles is the most massive? (**d**) Which is the least massive?

2.11 Consider an atom of ^{32}P. (**a**) How many protons, neutrons, and electrons does this atom contain? (**b**) What is the symbol of the atom obtained by adding one proton to ^{32}P? (**c**) What is the symbol of the atom obtained by adding one neutron to ^{32}P? (**d**) Are either of the atoms obtained in parts (b) and (c) isotopes of ^{32}P? If so which one?

2.12 (**a**) Define atomic number and mass number. (**b**) Which of these can vary without changing the identity of the element?

2.13 How many protons, neutrons, and electrons are in the following atoms? (**a**) ^{84}Kr, (**b**) ^{200}Hg, (**c**) ^{59}Co, (**d**) ^{55}Mn, (**e**) ^{239}U, (**f**) ^{181}Ta.

2.14 Fill in the gaps in the following table, assuming each column represents a neutral atom.

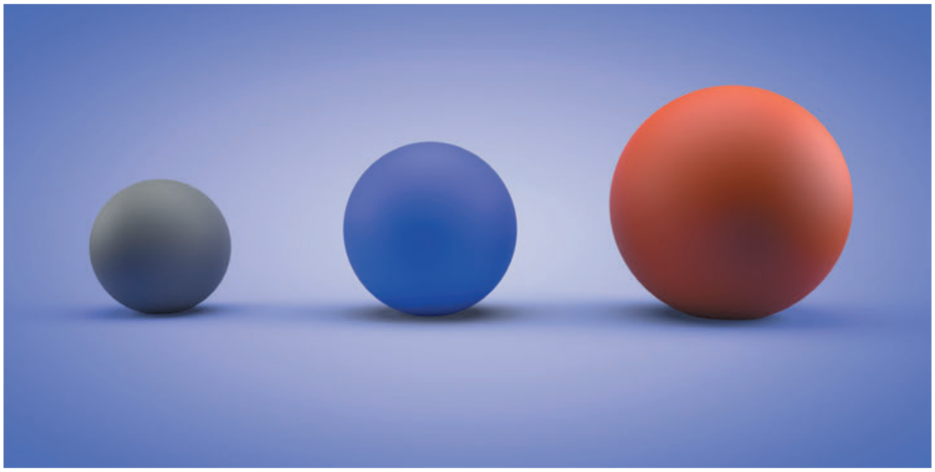
Symbol	^{159}Tb				
Protons		29			37
Neutrons		34	53		
Electrons			42	34	
Mass no.				79	85

2.15 Write the correct symbol, with both superscript and subscript, for each of the following. Use the list of elements in the front inside cover as needed: (**a**) the isotope of hafnium that contains 106 neutrons, (**b**) the isotope of mercury with mass number 201, (**c**) the isotope of rhenium with mass number 187, (**d**) the isotope of calcium that has an equal number of protons and neutrons.

Answers to Self-Assessment Exercises

2.7 (e) **2.8** (b)

2.4 Atomic Weights



We are quite used to objects of the same shape but of different sizes—a golf ball, a tennis ball, and a soccer ball for example. Atoms of different elements also have a different size and weight. By the end of this section, you should be able to

- Calculate the atomic weight of an element from its isotopic composition.

Atoms are small pieces of matter, so they have mass. In this section, we discuss the mass scale used for atoms and introduce the concept of *atomic weights*.

The Atomic Mass Scale

Scientists of the nineteenth century were aware that atoms of different elements have different masses. They found, for example, that each 100.0 g of water contains 11.1 g of hydrogen and 88.9 g of oxygen. Thus, water contains $88.9/11.1 = 8$ times as much oxygen, by mass, as hydrogen. Once scientists understood that water contains two hydrogen atoms for each oxygen atom, they concluded that an oxygen atom must have $2 \times 8 = 16$ times as much mass as a hydrogen atom. Hydrogen, the lightest atom, was arbitrarily assigned a relative mass of 1 (no units). Atomic masses of other elements were at first determined relative to this value. Thus, oxygen was assigned an atomic mass of 16.

Today we can determine the masses of individual atoms with a high degree of accuracy. For example, we know that the ^{1}H atom has a mass of 1.6735×10^{-24} g and the ^{16}O atom has a mass of 2.6560×10^{-23} g. As we noted in Section 2.3, it is convenient to use the **atomic mass unit** when dealing with these extremely small masses:

$$1\ \text{u} = 1.66054 \times 10^{-24}\ \text{g} \quad \text{and} \quad 1\ \text{g} = 6.02214 \times 10^{23}\ \text{u}$$

The atomic mass unit is presently defined by assigning a mass of exactly 12 u to a chemically unbound atom of the ^{12}C isotope of carbon. In these units, an ^{1}H atom has a mass of 1.0078 u and an ^{16}O atom has a mass of 15.9949 u.

Atomic Weight

Most elements occur in nature as mixtures of isotopes. We can determine the *average atomic mass* of an element, usually called the element's **atomic weight**, by summing (indicated by the Greek sigma, Σ) over the masses of its isotopes multiplied by their relative abundances:

$$\text{Atomic weight} = \sum_{\substack{\text{over all}\\ \text{isotopes of}\\ \text{the element}}} [(\text{isotope mass}) \times (\text{fractional isotope abundance})] \qquad [2.1]$$

Naturally occurring carbon, for example, is composed of 98.93% ^{12}C and 1.07% ^{13}C. The masses of these isotopes are 12 u (exactly) and 13.00335 u, respectively, making the atomic weight of carbon

$$(0.9893)(12\ \text{u}) + (0.0107)(13.00335\ \text{u}) = 12.01\ \text{u}$$

The atomic weights of the elements are listed in both the periodic table and the table of elements on the front inside cover of this text.

Sample Exercise 2.4
Calculating the Atomic Weight of an Element from Isotopic Abundances

Naturally occurring chlorine is 75.78% ^{35}Cl (atomic mass 34.969 u) and 24.22% ^{37}Cl (atomic mass 36.966 u). Calculate the atomic weight of chlorine.

SOLUTION

We can calculate the atomic weight by multiplying the abundance of each isotope by its mass and summing these products. Because 75.78% = 0.7578 and 24.22% = 0.2422, we have

$$\begin{aligned}\text{Atomic weight} &= (0.7578)(34.969\ \text{u}) + (0.2422)(36.966\ \text{u})\\ &= 26.50\ \text{u} + 8.953\ \text{u}\\ &= 35.45\ \text{u}\end{aligned}$$

This answer makes sense: The atomic weight, which is actually the average atomic mass, is between the masses of the two isotopes and is closer to the value of ^{35}Cl, the more abundant isotope.

Practice Exercise

Three isotopes of silicon occur in nature: ^{28}Si (92.23%), atomic mass 27.97693 u; ^{29}Si (4.68%), atomic mass 28.97649 u; and ^{30}Si (3.09%), atomic mass 29.97377 u. Calculate the atomic weight of silicon.

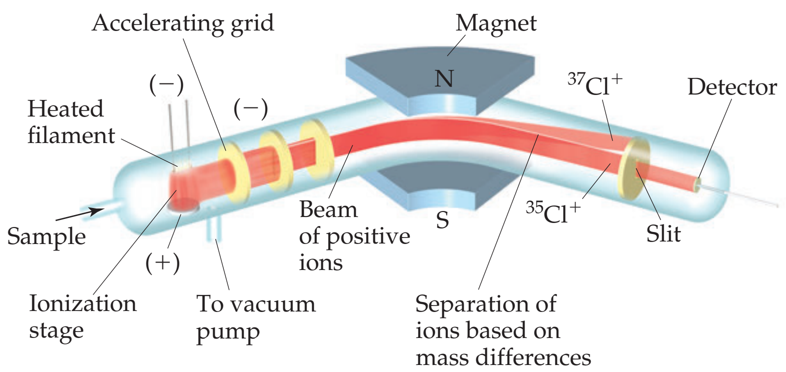
A CLOSER LOOK The Mass Spectrometer

The most accurate means for determining atomic weights is provided by the **mass spectrometer** (**Figure 2.10**). There are various designs of mass spectrometers, but they all operate on similar principles. The first step is to get atoms or molecules into the gas phase. Sometimes the sample to be analyzed is already a gas, whereas in other cases heating, application of an electric field, or a pulse of laser light may be needed to create gas-phase atoms or molecules. Next, the gas-phase species must be converted to positively charged particles called *ions*. There are many approaches to creating ions, including bombardment with beams of high-energy electrons or chemical reactions with other gas-phase molecules. Once gas-phase ions have been produced, they are accelerated toward a negatively charged grid. After the ions pass through the grid, they encounter two slits that allow only a narrow beam of ions to pass. This beam then passes between the poles of a magnet, which deflects the ions into a curved path. For ions with the same charge, the extent of deflection depends on mass—the more massive the ion, the less the deflection. The ions are thereby separated according to their masses. By changing the strength of the magnetic field or the accelerating voltage on the grid, ions of various masses can be selected to enter the detector.

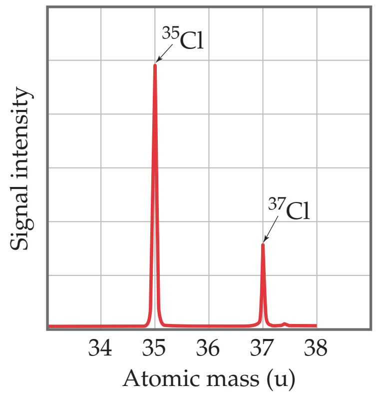
A graph of the intensity of the detector signal versus ion atomic mass is called a *mass spectrum* (**Figure 2.11**). Analysis of a mass spectrum gives both the masses of the ions reaching the detector and their relative abundances, which are obtained from the signal intensities. Knowing the atomic mass and the abundance of each isotope allows us to calculate the atomic weight of an element, as shown in Sample Exercise 2.4.

Mass spectrometers are used extensively today to identify chemical compounds and analyze mixtures of substances. Any molecule that loses electrons can fall apart, forming an array of positively charged fragments. The mass spectrometer measures the masses of these fragments, producing a chemical "fingerprint" of the molecule and providing clues about how the atoms were connected in the original molecule. Thus, a chemist might use this technique to determine the molecular structure of a newly synthesized compound, to analyze proteins in the human genome, or to identify a pollutant in the environment.

Related Exercises: 2.19, 2.73, 2.74, 2.99, 2.109, 2.110



▲ **Figure 2.10 A mass spectrometer.** Cl atoms are first ionized to form Cl^+ ions, accelerated with an electric field, and finally their path is directed by a magnetic field. The paths of the ions of the two Cl isotopes diverge as they pass through the field.



▲ **Figure 2.11 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.

Self-Assessment Exercise

2.16 Bromine has two main isotopes of similar abundance: ^{79}Br and ^{81}Br. Estimate (without using a calculator) the atomic weight of bromine.

(a) ~79 u
(b) ~80 u
(c) ~81 u

Exercises

2.17 (**a**) What isotope is used as the standard in establishing the atomic mass scale? (**b**) The atomic weight of boron is reported as 10.81, yet no atom of boron has the mass of 10.81 u. Explain.

2.18 Iron has three major isotopes: ^{54}Fe (atomic mass = 53.9396 u; abundance 5.85%), ^{56}Fe (atomic mass = 55.9349 u; abundance 91.75%), and ^{57}Fe (atomic mass = 56.9354 u; abundance 2.12 %). Calculate the atomic weight of iron.

2.19 (**a**) Thomson's cathode-ray tube (Figure 2.3) and the mass spectrometer (Figure 2.10) both involve the use of electric or magnetic fields to deflect charged particles. What are the charged particles involved in each of these experiments? (**b**) What are the labels on the axes of a mass spectrum? (**c**) To measure the mass spectrum of an atom, the atom must first lose one or more electrons. Which would you expect to be deflected more by the same setting of the electric and magnetic fields, a Cl^+ or a Cl^{2+} ion?

2.20 Naturally occurring lead has the following isotopic abundances:

Isotope	Abundance (%)	Atomic mass (u)
^{204}Pb	1.4	203.9730
^{206}Pb	24.1	205.9744
^{207}Pb	22.1	206.9759
^{208}Pb	52.4	207.9766

(**a**) What is the average atomic mass of Pb? (**b**) Sketch the mass spectrum of Pb.

Answers to Self-Assessment Exercises

2.16 (b)

2.5 The Periodic Table

The shininess of metals is an obvious similarity that they share, but many elements show pattens in their reactivity that are not so evident from their physical appearance. One triumph of classification was the ordering of the elements into the Periodic Table. It enables us to understand trends in the properties of the elements, make predictions about their chemical reactivity and even suggest the characteristics of elements that are yet to be discovered.

By the end of this section, you should be able to

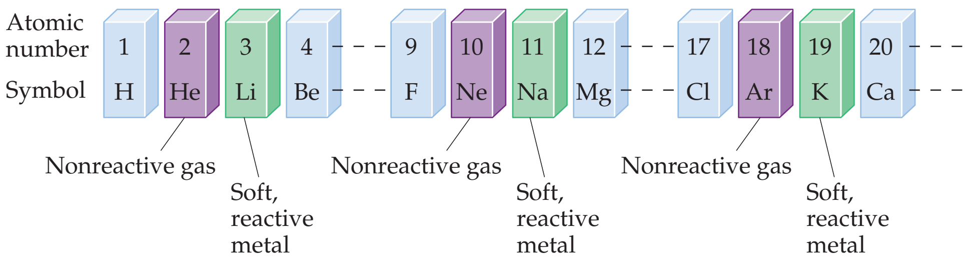
- Describe the general features of the periodic table

As the list of known elements expanded during the early 1800s, attempts were made to find patterns in chemical behavior. These efforts culminated in the development of the periodic table in 1869. We will have much to say about the periodic table in later chapters, but it is so important and useful that you should become acquainted with it now. You will quickly learn that *the periodic table is the most significant tool that chemists use for organizing and remembering chemical facts*.

Many elements show strong similarities to one another. The elements lithium (Li), sodium (Na), and potassium (K) are all soft, very reactive metals, for example. The elements helium (He), neon (Ne), and argon (Ar) are all nonreactive gases. If the elements are arranged in order of increasing atomic number, their chemical and physical properties show a repeating, or *periodic*, pattern. For example, each of the soft, reactive metals—lithium, sodium, and potassium—comes immediately after one of the nonreactive gases—helium, neon, and argon, respectively—as shown in **Figure 2.12**.

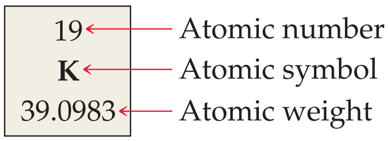
Go Figure

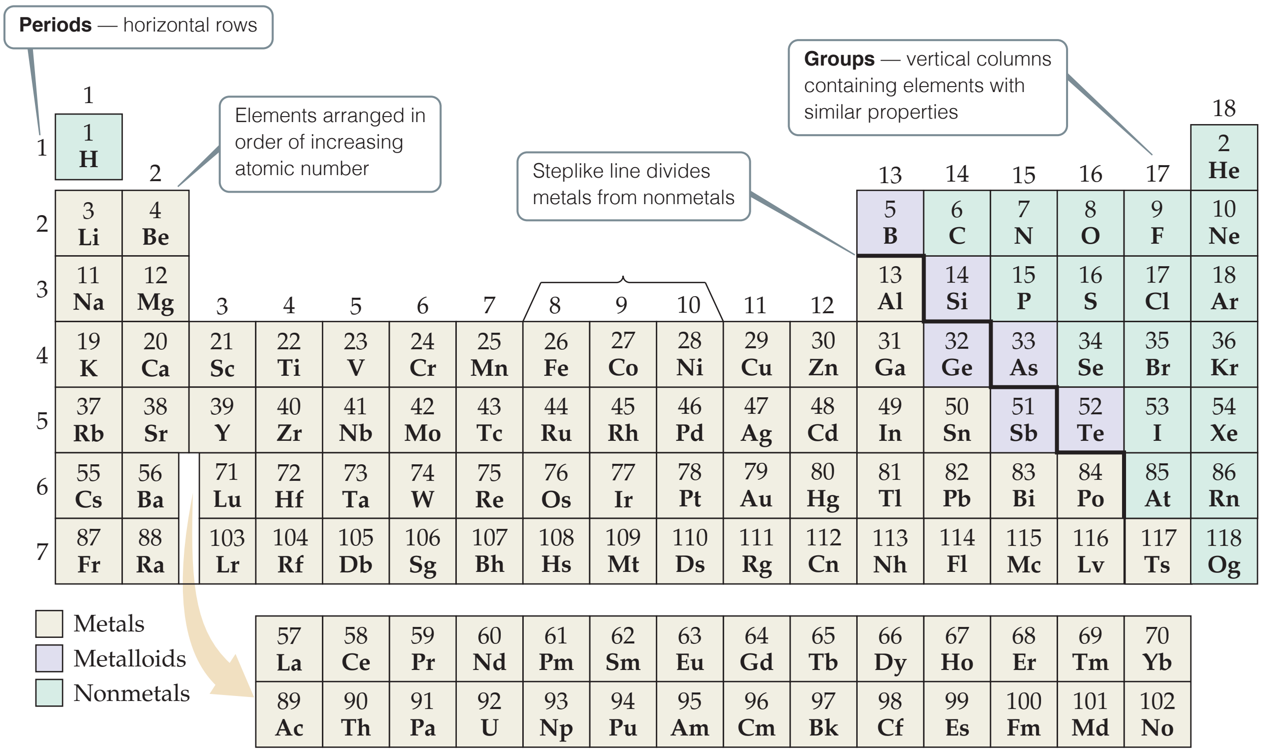
If F is a reactive nonmetal, which other element or elements shown here are likely to be reactive nonmetals?



▲ **Figure 2.12 Arranging elements by atomic number reveals a periodic pattern of properties.** This pattern is the basis of the periodic table.

The arrangement of elements in order of increasing atomic number, with elements having similar properties placed in vertical columns, is known as the **periodic table** (Figure 2.13). The table shows the atomic number and atomic symbol for each element, and the atomic weight is often given as well, as in this typical entry for potassium:





▲ **Figure 2.13 Periodic table of elements.**

You might notice slight variations in periodic tables from one book to another or between those in the lecture hall and in the text. These are simply matters of style, or they might concern the particular information included. There are no fundamental differences.

The horizontal rows of the periodic table are called **periods**. The first period consists of only two elements, hydrogen (H) and helium (He). The second and third periods consist of eight elements each. The fourth and fifth periods contain 18 elements. The sixth and seventh periods have 32 elements each, but, in order to fit on a page, 14 of the elements from each period (atomic numbers 57–70 and 89–102) appear at the bottom of the table.

The vertical columns are **groups**. These are numbered from 1 to 18, as shown in Figure 2.13.

We will use the IUPAC's proposed convention and number the groups from 1 through 18.

Elements in a group often exhibit similarities in physical and chemical properties. For example, the "coinage metals"—copper (Cu), silver (Ag), and gold (Au)—belong to Group 11. These elements are less reactive than most metals, which is why they have been traditionally used throughout the world to make coins. Many other groups in the periodic table also have names, listed in Table 2.3.

We will learn in Chapters 6 and 7 that elements in a group have similar properties because they have the same arrangement of electrons at the periphery of their atoms. However, we need not wait until then to make good use of the periodic table; after all, the chemists who developed the table knew nothing about electrons! We can use the table, as they intended, to correlate behaviors of elements and to help us remember many facts.

The color code of Figure 2.13 shows that, except for hydrogen, all the elements on the left and in the middle of the table are **metallic elements**, or **metals**. All the metallic elements share characteristic properties, such as luster and high electrical and heat conductivity, and all of them except mercury (Hg) are solid at room temperature.* The metals are separated from the **nonmetallic elements**, or **nonmetals**, by a stepped line that runs from boron (B) to astatine (At). (Note that hydrogen, although on the left side of the table, is a nonmetal.) At room temperature and pressure, some of the nonmetals are gaseous, some are solid, and one is liquid. Nonmetals generally differ from metals in appearance (Figure 2.14) and in other physical properties. Many of the elements that lie along the line that separates metals from nonmetals have properties that fall between those of metals and nonmetals. These elements are often referred to as **metalloids**.

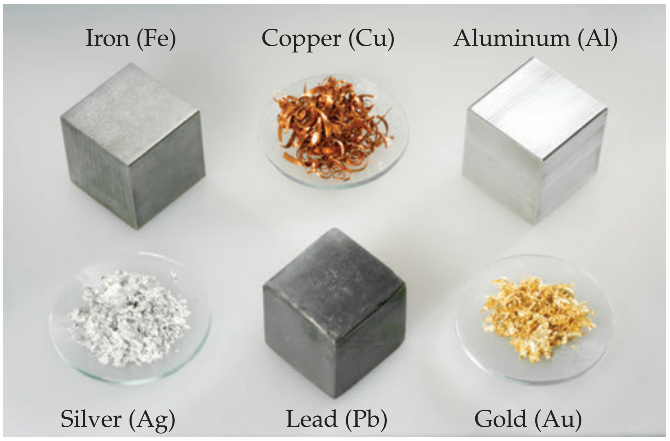
TABLE 2.3 Names of Some Groups in the Periodic Table

Group	Name	Elements
1	Alkali metals	Li, Na, K, Rb, Cs, Fr
2	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
16	Chalcogens	O, S, Se, Te, Po
17	Halogens	F, Cl, Br, I, At
18	Noble gases	He, Ne, Ar, Kr, Xe, Rn

*All metals become liquids if heated sufficiently. Hg simply has the lowest melting point of any metallic element. Although sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and gallium (Ga) are solids at room temperature, they all melt at temperatures below 100 °C.

Go Figure Name two ways in which the metals shown here differ in general appearance from the nonmetals.

Metals



Nonmetals

▲ **Figure 2.14** Examples of metals and nonmetals.

Sample Exercise 2.5
Using the Periodic Table

Which two of these elements would you expect to show the greatest similarity in chemical and physical properties: B, Ca, F, He, Mg, P?

SOLUTION

Elements in the same group of the periodic table are most likely to exhibit similar properties. We therefore expect Ca and Mg to be most alike because they are in the same group (2, the alkaline earth metals).

Practice Exercise

Locate Na (sodium) and Br (bromine) in the periodic table. Give the atomic number of each and classify each as metal, metalloid, or nonmetal.

Self-Assessment Exercise

2.21 The halogens are a group of elements, two of which are gases, one of which is a liquid, and two solid at room temperature. In spite of these obvious physical differences, there are many chemical similarities between the elements. What are the elements that are classed as halogens?

(a) Li, Be, B, C, N, O, F, Ne
(b) Be, Mg, Ca, Sr, Ba, Ra
(c) F, Cl, Br, I, At

Exercises

2.22 For each of the following elements, write its chemical symbol, locate it in the periodic table, give its atomic number, and indicate whether it is a metal, metalloid, or nonmetal: (**a**) radon, (**b**) tellurium, (**c**) cadmium, (**d**) chromium, (**e**) barium, (**f**) selenium, (**g**) sulphur.

2.23 The elements of Group 14 show an interesting change in properties moving down the group. Give the name and chemical symbol of each element in the group and label it as a nonmetal, metalloid, or metal.

Answers to Self-Assessment Exercises

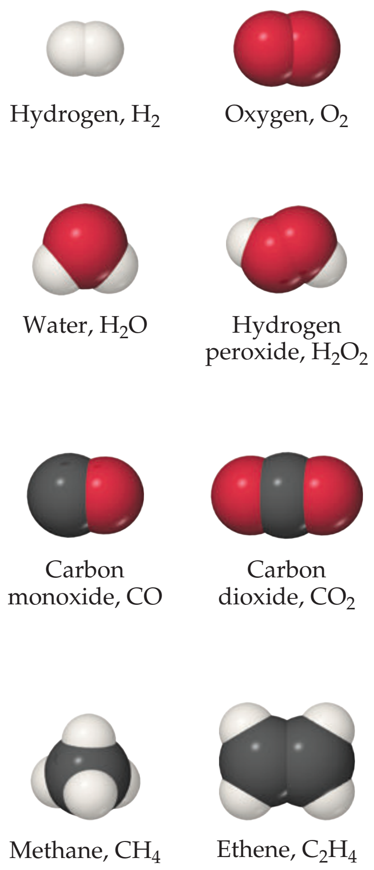
2.21 (c)

2.6 Molecules and Molecular Compounds

DNA, responsible for carrying the genetic instructions of all organisms, is one of the most complex molecules, yet it is composed of just five elements: C, H, O, N, and P. Knowing how the elements join together is fundamental to our understanding of chemistry. By the end of this section, you should be able to

- Recognize the different types of formula used to describe molecules

Even though the atom is the smallest representative sample of an element, only the noble-gas elements are normally found in nature as isolated atoms. Most matter is composed of molecules or ions. We examine molecules here and ions in Section 2.7.



▲ **Figure 2.15 Molecular models.** Notice how the chemical formulas of these simple molecules correspond to their compositions.

Molecules and Chemical Formulas

Several elements are found in nature in molecular form—two or more of the same type of atom bound together. For example, most of the oxygen in air consists of molecules that contain two oxygen atoms. As we saw in Section 1.2, we represent this molecular oxygen by the **chemical formula** O_2 (read "oh two"). The subscript tells us that two oxygen atoms are present in each molecule. A molecule made up of two atoms is called a **diatomic molecule**.

Oxygen also exists in another molecular form known as *ozone*. Molecules of ozone consist of three oxygen atoms, making the chemical formula O_3. Even though "normal" oxygen (O_2) and ozone (O_3) are both composed only of oxygen atoms, they exhibit very different chemical and physical properties. For example, O_2 is essential for life, but O_3 is toxic; O_2 is odorless, whereas O_3 has a sharp, pungent smell.

The elements that normally occur as diatomic molecules are hydrogen, oxygen, nitrogen, and the halogens (H_2, O_2, N_2, F_2, Cl_2, Br_2, and I_2). Except for hydrogen, these diatomic elements are clustered on the right side of the periodic table.

Compounds composed of molecules contain more than one type of atom and are called **molecular compounds**. A molecule of the compound methane, for example, consists of one carbon atom and four hydrogen atoms and is therefore represented by the chemical formula CH_4. Lack of a subscript on the C indicates one atom of C per methane molecule. Several common molecules of both elements and compounds are shown in **Figure 2.15.** Notice how the composition of each substance is given by its chemical formula. Notice also that these substances are composed only of nonmetallic elements.

Most of the molecular substances we will encounter contain only nonmetals.

Molecular and Empirical Formulas

Chemical formulas that indicate the actual numbers of atoms in a molecule are called **molecular formulas.** (The formulas in Figure 2.17 are molecular formulas.) Chemical formulas that give only the relative number of atoms of each type in a molecule are called **empirical formulas**. The subscripts in an empirical formula are always the smallest possible whole-number ratios. The molecular formula for hydrogen peroxide is H_2O_2, for example, whereas its empirical formula is HO. The molecular formula for ethene is C_2H_4, and its empirical formula is CH_2. For many substances, the molecular formula and the empirical formula are identical, as in the case of water, H_2O.

Whenever we know the molecular formula of a compound, we can determine its empirical formula. The converse is not true, however. If we know the empirical formula of a substance, we cannot determine its molecular formula unless we have more information. So why do chemists bother with empirical formulas? As we will see in Chapter 3, certain common methods of analyzing substances lead to the empirical formula only. For example, if you decomposed hydrogen peroxide H_2O_2 into its elements and weighed them, you could determine that there were equal numbers of hydrogen and oxygen atoms, but you would not know if the molecular formula was HO, H_2O_2, H_3O_3, or the like. Once the empirical formula is known, additional experiments can give the information needed to convert the empirical formula to the molecular one. In addition, there are many substances that do not exist as isolated molecules, one example being ionic compounds that are discussed later in this chapter. For these substances, we must rely on empirical formulas.

Sample Exercise 2.6
Relating Empirical and Molecular Formulas

Write the empirical formulas for (**a**) glucose, a substance also known as either blood sugar or dextrose—molecular formula $C_6H_{12}O_6$; (**b**) nitrous oxide, a substance used as an anesthetic and commonly called laughing gas—molecular formula N_2O.

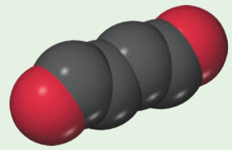
SOLUTION

(**a**) The subscripts of an empirical formula are the smallest whole-number ratios. The smallest ratios are obtained by dividing each subscript by the largest common factor, in this case 6. The resultant empirical formula for glucose is CH_2O.

(**b**) Because the subscripts in N_2O are already the lowest integral numbers, the empirical formula for nitrous oxide is the same as its molecular formula, N_2O.

Practice Exercise

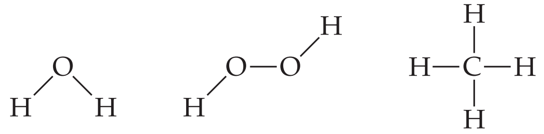
Tetracarbon dioxide is an unstable oxide of carbon with the following molecular structure:



What are the molecular and empirical formulas of this substance? (**a**) C_2O_2, CO_2 (**b**) C_4O, CO (**c**) CO_2, CO_2 (**d**) C_4O_2, C_2O (**e**) C_2O, CO_2

Picturing Molecules

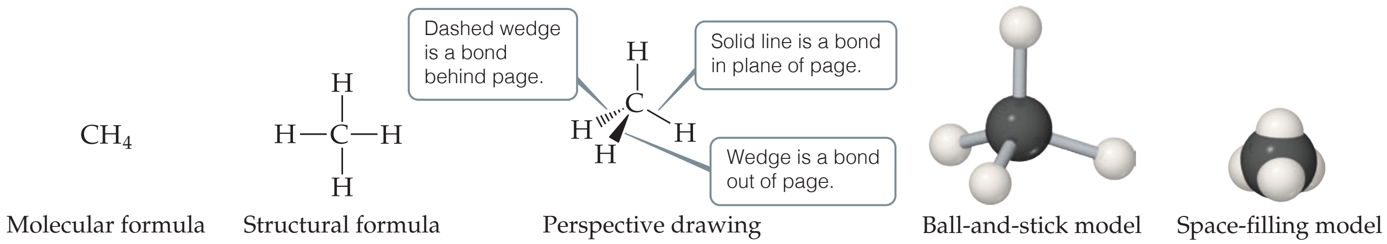
The molecular formula of a substance does not show how its atoms are joined together. A **structural formula** is needed to convey that information, as in the following examples:



Water　Hydrogen peroxide　Methane

The atoms are represented by their chemical symbols, and lines are used to represent the bonds that hold the atoms together.

Go Figure Which model, the ball-and-stick or the space-filling, more effectively shows the angles between bonds around a central atom?



▲ **Figure 2.16 Different representations of the methane (CH_4) molecule.** Structural formulas, perspective drawings, ball-and-stick models, and space-filling models.

A structural formula does not typically depict the actual geometry of the molecule, that is, the actual angles at which atoms are joined; for that, more sophisticated representations are needed (**Figure 2.16**).

- **Perspective drawings** use wedges and dashed lines to depict bonds that are not in the plane of the paper. This gives a crude sense of the three-dimensional shape of a molecule.
- **Ball-and-stick models** show atoms as spheres and bonds as sticks. This type of model has the advantage of accurately representing the angles at which the atoms are attached to one another in a molecule (Figure 2.16). Sometimes the chemical symbols of the elements are superimposed on the balls, but often the atoms are identified simply by color.
- **Space-filling models** depict what a molecule would look like if the atoms were scaled up in size (Figure 2.16). These models show the relative sizes of the atoms, but the angles between atoms, which help define their molecular geometry, are often more difficult to see than in ball-and-stick models. Because space-filling models give a good representation of the true size of a molecule, they are useful for picturing how two molecules might fit together or pack in the solid state. As with ball-and-stick models, the identities of the atoms are typically indicated by color.

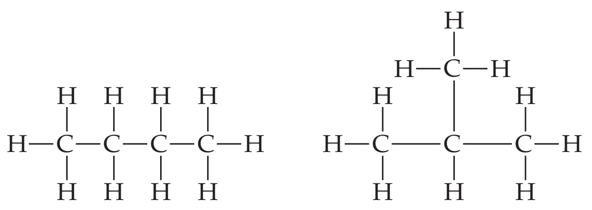
Self-Assessment Exercise

2.24 Lactic acid can build up in muscle tissue during strenuous exercise, sometimes leading to painful 'cramps'. The molecular formula of lactic acid is $C_3H_6O_3$. What is its empirical formula?

(a) CHO
(b) CH_2O
(c) $C_3H_6O_3$

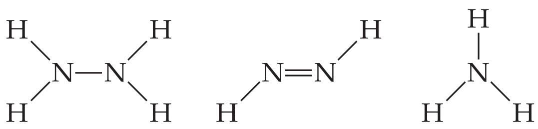
Exercises

2.25 The structural formulas of the compounds *n*-butane and isobutane are shown here. (**a**) Determine the molecular formula of each. (**b**) Determine the empirical formula of each. (**c**) Which formulas—empirical, molecular, or structural—allow you determine these are different compounds?



n-butane isobutane

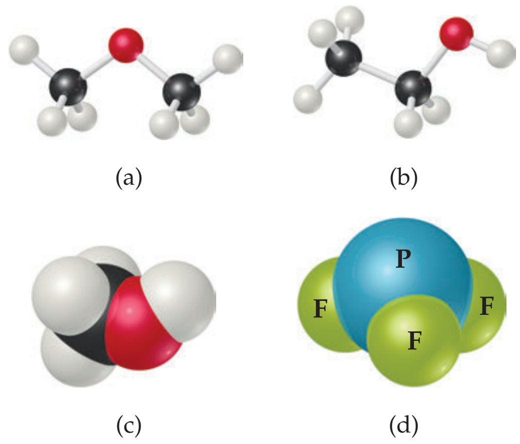
2.26 What are the molecular and empirical formulas for each of the following compounds?



2.27 Write the empirical formula corresponding to each of the following molecular formulas: (**a**) Al_2Br_6, (**b**) C_8H_{10}, (**c**) $C_4H_8O_2$, (**d**) P_4O_{10}, (**e**) $C_6H_4Cl_2$, (**f**) $B_3N_3H_6$.

2.28 How many hydrogen atoms are in each of the following: (**a**) C_2H_5OH, (**b**) $Ca(C_2H_5COO)_2$, (**c**) $(NH_4)_3PO_4$?

2.29 Write the molecular and structural formulas for the compounds represented by the following molecular models:



Answers to Self-Assessment Exercises

2.24 (b)

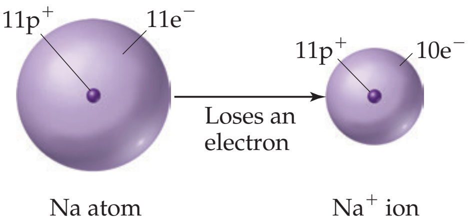
2.7 Ions and Ionic Compounds

Salt has always been an essential part of our diet. In regions where it is scarce, it commands a high price. During the Roman Empire, soldiers were paid, in part, with an allocation of salt (*sal* being the Latin word for salt) and this is the origin of our word 'salary'. By the end of this section, you should be able to

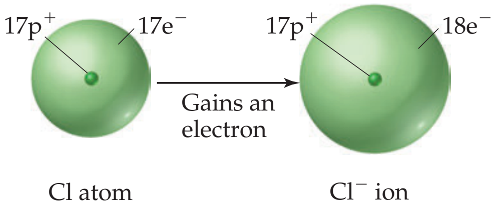
- Predict the charge on the ion formed by elements of Group 1, 2, 15, 16, and 17.
- Write the formula of an ionic compound given the charges of the ions

If electrons are removed from or added to an atom, a charged particle called an **ion** is formed. An ion with a positive charge is a **cation** (pronounced CAT-ion); a negatively charged ion is an **anion** (AN-ion).

To see how ions form, consider the sodium atom, which has 11 protons and 11 electrons. If this atom loses one electron, the resulting cation has 11 protons and 10 electrons, which means it has a net charge of 1+.



The net charge on an ion is represented by a superscript. The superscripts +, 2+, and 3+, for instance, mean a net charge resulting from the *loss* of one, two, and three electrons, respectively. The superscripts −, 2−, and 3− represent net charges resulting from the *gain* of one, two, and three electrons, respectively. Chlorine, with 17 protons and 17 electrons, for example, can gain an electron in chemical reactions, producing the Cl^- ion:



In general, metal atoms tend to lose electrons to form cations and nonmetal atoms tend to gain electrons to form anions. Thus, ionic compounds tend to be composed of both metal cations and nonmetal anions, as in NaCl.

Sample Exercise 2.7
Writing Chemical Symbols for Ions

Give the chemical symbol, including superscript indicating mass number, for (**a**) the ion with 22 protons, 26 neutrons, and 19 electrons; and (**b**) the ion of sulfur that has 16 neutrons and 18 electrons.

SOLUTION

(**a**) The number of protons is the atomic number of the element. A periodic table or list of elements tells us that the element with atomic number 22 is titanium (Ti). The mass number (protons plus neutrons) of this isotope of titanium is 22 + 26 = 48. Because the ion has three more protons than electrons, it has a net charge of 3+ and is designated $^{48}Ti^{3+}$.

(**b**) The periodic table tells us that sulfur (S) has an atomic number of 16. Thus, each atom or ion of sulfur contains 16 protons. We are told that the ion also has 16 neutrons, meaning the mass number is 16 + 16 = 32. Because the ion has 16 protons and 18 electrons, its net charge is 2− and the ion symbol is $^{32}S^{2-}$.

In general, we will focus on the net charges of ions and ignore their mass numbers unless the circumstances dictate that we specify a certain isotope.

Practice Exercise

In which of the following species is the difference between the number of protons and the number of electrons largest?
(**a**) Ti^{2+} (**b**) P^{3-} (**c**) Mn (**d**) Se^{2-} (**e**) Ce^{4+}

In addition to simple ions such as Na^+ and Cl^-, there are **polyatomic ions**, such as NH_4^+ (ammonium ion) and SO_4^{2-} (sulfate ion), which consist of atoms joined as in a molecule, but carrying a net positive or negative charge. Polyatomic ions will be discussed in Section 2.8.

It is important to realize that the chemical properties of ions are very different from the chemical properties of the atoms from which the ions are derived. The addition or removal of one or more electrons produces a charged species with behavior very different from that of its associated atom or group of atoms. For example, sodium metal reacts violently with water, but an ionic compound containing sodium ions, such as NaCl, does not.

Predicting Ionic Charges

As noted in Table 2.3, the elements of Group 18 are called the noble-gas elements. The noble gases are chemically nonreactive elements that form very few compounds. Many atoms gain or lose electrons to end up with the same number of electrons as the noble gas closest to them in the periodic table. We might deduce that atoms tend to acquire the electron arrangements of the noble gases because these electron arrangements are very stable. Nearby elements can obtain these same stable arrangements by losing or gaining electrons. For example, the loss of one electron from an atom of sodium leaves it with the same number of electrons as in a neon atom (10). Similarly, when chlorine gains an electron, it ends up with 18, the same number of electrons as in argon. This simple observation will be helpful for now to account for the formation of ions. A deeper explanation awaits us in Chapter 8, where we discuss chemical bonding.

Sample Exercise 2.8

Predicting Ionic Charge

Predict the charge expected for the most stable ion of barium and the most stable ion of oxygen.

SOLUTION

We will assume that barium and oxygen form ions that have the same number of electrons as the nearest noble-gas atom. From the periodic table, we see that barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can attain a stable arrangement of 54 electrons by losing two electrons, forming the Ba^{2+} cation.

Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can attain this stable electron arrangement by gaining two electrons, forming the O^{2-} anion.

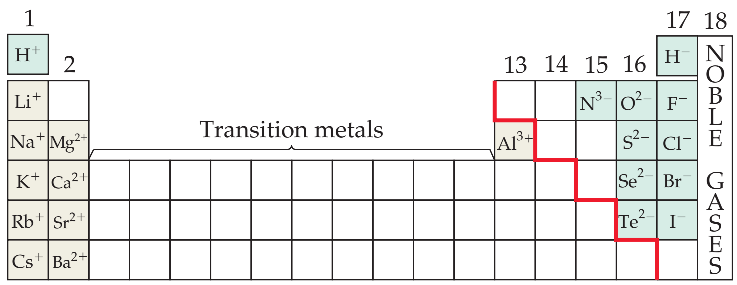
Practice Exercise

Although it is helpful to know that many ions have the electron arrangement of a noble gas, many elements, especially among the metals, form ions that do not have a noble-gas electron arrangement. Use the periodic table, Figure 2.13, to determine which of the following ions has a noble-gas electron arrangement, and which do not. For those that do, indicate the noble-gas arrangement they match: **(a)** Ti^{4+}, **(b)** Mn^{2+}, **(c)** Pb^{2+}, **(d)** Te^{2-}, **(e)** Zn^{2+}.

The periodic table is very useful for remembering ionic charges, especially those of elements on the left and right sides of the table. As **Figure 2.17** shows, the charges of these ions relate in a simple way to their positions in the table: The Group 1 elements (alkali metals) form 1+ ions, the Group 2 elements (alkaline earth metals) form 2+ ions, the Group 17 elements (halogens) form 1− ions, and the Group 16 elements form 2− ions. As noted in the Practice Exercise of Sample Exercise 2.8, many of the other groups do not lend themselves to such simple rules.

Go Figure

The most common ions for silver, zinc, and scandium are Ag^{+}, Zn^{2+}, and Sc^{3+}. Locate the boxes in which you would place these ions in this table. Which of these ions has the same number of electrons as a noble-gas element?

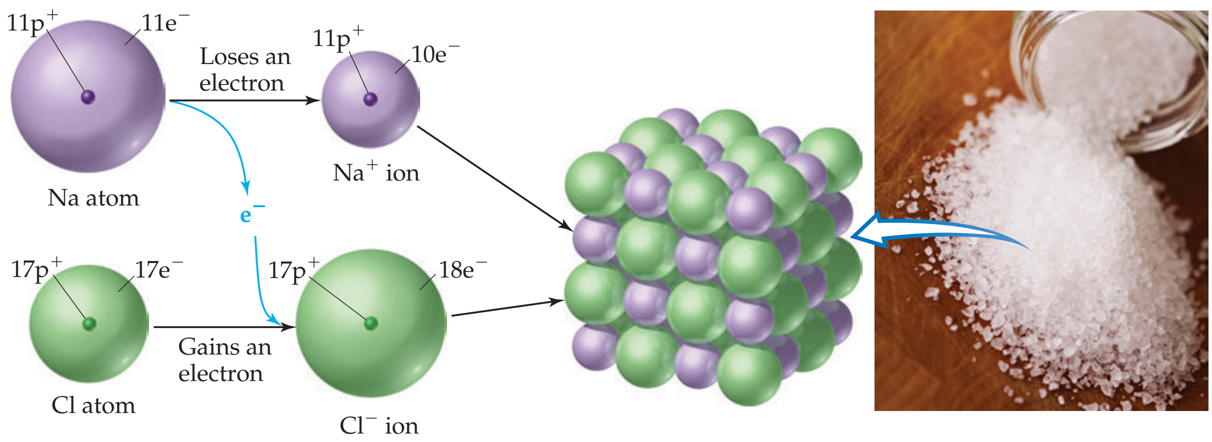


▲ **Figure 2.17 Predictable charges of some common ions.** Notice that the red stepped line that divides metals from nonmetals also separates cations from anions. Hydrogen forms both 1+ and 1− ions.

Ionic Compounds

A great deal of chemical activity involves the transfer of electrons from one substance to another. **Figure 2.18** shows that when elemental sodium is allowed to react with elemental chlorine, an electron transfers from a sodium atom to a chlorine atom, forming a Na^{+} ion and a Cl^{-} ion. Because objects of opposite charges attract, the Na^{+} and the Cl^{-} ions bind together to form the compound sodium chloride (NaCl). Sodium chloride, which we know better as common table salt, is an example of an **ionic compound**, a compound made up of cations and anions.

Go Figure Is it possible to pick out a unique cluster of atoms that can be thought of as a sodium chloride molecule?



▲ **Figure 2.18 Formation of an ionic compound.** The transfer of an electron from a sodium atom to a chlorine atom leads to the formation of a Na^+ ion and a Cl^- ion. These ions are arranged in a lattice in solid sodium chloride, NaCl.

We can often tell whether a compound is ionic (consisting of ions) or molecular (consisting of molecules) from its composition. As a general rule, cations are metal ions and anions are nonmetal ions. Consequently,

Ionic compounds are generally combinations of metals and nonmetals, as in NaCl.

Sample Exercise 2.9
Identifying Ionic and Molecular Compounds

Which of these compounds would you expect to be ionic: N_2O, Na_2O, $CaCl_2$, SF_4?

SOLUTION

We predict that Na_2O and $CaCl_2$ are ionic compounds because they are composed of a metal combined with a nonmetal. We predict (correctly) that N_2O and SF_4 are molecular compounds because they are composed entirely of nonmetals.

▶ Practice Exercise

Give a reason why each of the following statements is a safe prediction:

(a) Every compound of Rb with a nonmetal is ionic in character.
(b) Every compound of nitrogen with a halogen element is a molecular compound.
(c) The compound $MgKr_2$ does not exist.
(d) Na and K are very similar in the compounds they form with nonmetals.
(e) If contained in an ionic compound, calcium (Ca) will be in the form of the doubly charged ion, Ca^{2+}.

The ions in ionic compounds are arranged in three-dimensional structures, as Figure 2.18 shows for NaCl. Because there is no discrete "molecule" of NaCl, we are able to write only an empirical formula for this substance. This is true for most other ionic compounds.

We can write the empirical formula for an ionic compound if we know the charges of the ions. Because chemical compounds are always electrically neutral, the ions in an ionic compound always occur in such a ratio that the total positive charge equals the total negative charge. Thus, there is one Na^+ to one Cl^- in NaCl, one Ba^{2+} to two Cl^- in $BaCl_2$, and so forth.

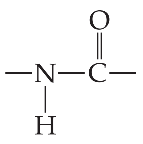
As you consider these and other examples, you will see that if the charges on the cation and anion are equal, the subscript on each ion is 1. If the charges are not equal, the charge on one ion (without its sign) will become the subscript on the other ion. For example, the ionic compound formed from Mg (which forms Mg^{2+} ions) and N (which forms N^{3-} ions) is Mg_3N_2:

$$Mg^{2+} \quad N^{3-} \longrightarrow Mg_3N_2$$

There is one caveat to using this approach. Remember that the empirical formula should be the smallest possible whole-number ratio of the two elements. So the empirical formula for the ionic compound formed between Ti^{4+} and O^{2-} is TiO_2 rather than Ti_2O_4.

CHEMISTRY AND LIFE Elements Required by Living Organisms

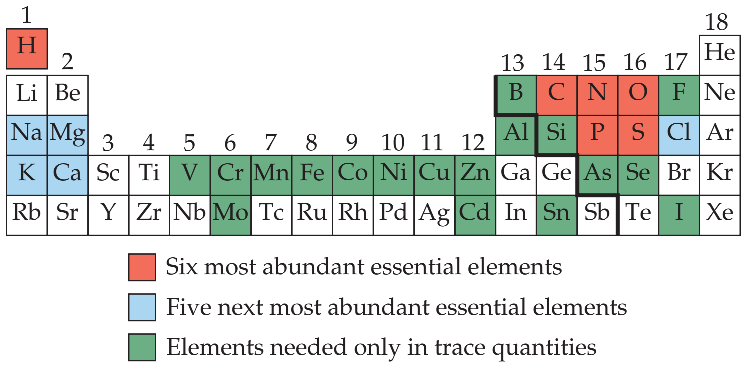
The elements essential to life are highlighted in color in **Figure 2.19.** More than 97% of the mass of most organisms is made up of just six of these elements—oxygen, carbon, hydrogen, nitrogen, phosphorus, and sulfur. Water is the most common compound in living organisms, accounting for at least 70% of the mass of most cells. In the solid components of cells, carbon is the most prevalent element by mass. Carbon atoms are found in a vast variety of organic molecules, bonded either to other carbon atoms or to atoms of other elements. Nearly all proteins, for example, contain the carbon-based group



which occurs repeatedly in the molecules.

In addition, 23 other elements have been found in various living organisms. Five are ions required by all organisms: Ca^{2+}, Cl^-, Mg^{2+}, K^+, and Na^+. Calcium ions, for example, are necessary for the formation of bone and transmission of nervous system signals. Many other elements are needed in only very small quantities and consequently are called *trace* elements. For example, trace quantities of copper are required in the diet of humans to aid in the synthesis of hemoglobin.

Related Exercise: 2.113



▲ **Figure 2.19 Elements essential to life.**

Sample Exercise 2.10

Using Ionic Charge to Write Empirical Formulas for Ionic Compounds

Write the empirical formula of the compound formed by (**a**) Al^{3+} and Cl^- ions, (**b**) Al^{3+} and O^{2-} ions, (**c**) Mg^{2+} and NO_3^- ions.

SOLUTION

(**a**) Three Cl^- ions are required to balance the charge of one Al^{3+} ion, making the empirical formula $AlCl_3$.

(**b**) Two Al^{3+} ions are required to balance the charge of three O^{2-} ions. A 2:3 ratio is needed to balance the total positive charge of 6+ and the total negative charge of 6–. The empirical formula is Al_2O_3.

(**c**) Two NO_3^- ions are needed to balance the charge of one Mg^{2+}, yielding $Mg(NO_3)_2$. Note that the formula for the polyatomic ion, NO_3^-, must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion.

▶ Practice Exercise

Which of the following nonmetals will form an ionic compound with Sc^{3+} that has a 1:1 ratio of cations to anions?
(**a**) Ne (**b**) F (**c**) O (**d**) N

Self-Assessment Exercises

2.30 Which group in the periodic table contains elements that form a 2- anion?

(a) Group 2

(b) Group 12

(c) Group 16

2.31 Which is the formula of the ionic compound formed from sodium cations and sulfur anion?

(a) Na_2S

(b) NaS

(c) NaS_2

Exercises

2.32 Predict whether each of the following compounds is molecular or ionic: (**a**) $HClO_4$ (**b**) CH_3OCH_3 (**c**) $Mg(NO_3)_2$ (**d**) H_2S (**e**) $TiCl_4$ (**f**) K_2O_2 (**g**) PCl_5 (**h**) $P(OH)_3$.

2.33 Fill in the gaps in the following table:

Symbol	$^{58}Fe^{2+}$			
Protons		50		40
Neutrons		68	78	50
Electrons			54	38
Net charge		4+	2−	

2.34 Each of the following elements is capable of forming an ion in chemical reactions. By referring to the periodic table, predict the charge of the most stable ion of each: (**a**) Be, (**b**) Rb, (**c**) As, (**d**) In, (**e**) At.

2.35 Using the periodic table to guide you, predict the chemical formula of the compound formed by the following elements: (**a**) Ga and F, (**b**) Li and H, (**c**) Al and I, (**d**) K and S.

2.36 Predict the chemical formulas of the ionic compound formed by (**a**) Fe^{3+} and OH^-, (**b**) Cs^+ and NO_3^-, (**c**) V^{2+} and CH_3COO^-, (**d**) Li^+ and PO_4^{3-}, (**e**) In^{3+} and O^{2-}.

2.37 Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

Ion	K^+	NH_4^+	Mg^{2+}	Fe^{3+}
Cl^-	KCl			
OH^-				
CO_3^{2-}				
PO_4^{3-}				

Answers to Self-Assessment Exercises

2.30 (c) **2.31** (a)

2.8 | Naming Inorganic Compounds

The names and chemical formulas of compounds are essential vocabulary in chemistry. The system used in naming substances is called **chemical nomenclature**, from the Latin words *nomen* (name) and *calare* (to call). By the end of this section, you should be able to

- Convert between the name and formula of an inorganic compound

There are more than 50 million known chemical substances. Naming them all would be a hopelessly complicated task if each had a name independent of all others. Many important substances that have been known for a long time, such as water (H_2O) and ammonia (NH_3), do have traditional names (called *common names*). For most substances, however, we rely on a set of rules that leads to an informative and unique name that conveys the composition of the substance.

The rules for chemical nomenclature are based on the division of substances into categories. The major division is between organic and inorganic compounds. *Organic compounds* contain carbon and hydrogen, often in combination with oxygen, nitrogen, or other elements. All others are *inorganic compounds*. Early chemists associated organic compounds with plants and animals and inorganic compounds with the nonliving portion of our world, for example minerals and water. Although this distinction is no longer pertinent, the classification between organic and inorganic compounds continues to be useful. In this section, we consider the basic rules for naming three categories of inorganic compounds: ionic compounds, acids, and molecular compounds.

Names and Formulas of Ionic Compounds

Recall from Section 2.7 that ionic compounds usually consist of metal ions combined with nonmetal ions. The metals form the cations, and the nonmetals form the anions.

1. **Cations**

a. *Cations formed from metal atoms have the same name as the metal:*

Na^+	sodium ion	Zn^{2+}	zinc ion	Al^{3+}	aluminum ion

b. *If a metal can form cations with different charges, the positive charge is indicated by a Roman numeral in parentheses following the name of the metal:*

Fe^{2+}	iron(II) ion	Cu^+	copper(I) ion
Fe^{3+}	iron(III) ion	Cu^{2+}	copper(II) ion

Ions of the same element that have different charges have different chemical and physical properties, such as color (**Figure 2.20**).

Most metals that form cations with different charges are *transition metals*, elements that occur in the middle of the periodic table, from Group 3 to Group 12. The metals that form only one cation (only one possible charge) are those of Group 1 and Group 2, as well as Al^{3+} (Group 13) and two transition-metal ions: Ag^+ (Group 12) and Zn^{2+} (Group 11). Charges are not expressed when naming these ions. However, if there is any doubt in your mind whether a metal forms more than one cation, use a Roman numeral to indicate the charge. It is never wrong to do so, even though it may be unnecessary.

An older method still widely used for distinguishing between differently charged ions of a metal uses the endings *-ous* and *-ic* added to the root of the element's Latin name:

Fe^{2+}	ferrous ion	Cu^+	cuprous ion
Fe^{3+}	ferric ion	Cu^{2+}	cupric ion

Although we will only rarely use these older names in this text, you might encounter them elsewhere.

▲ **Figure 2.20 Different ions of the same element have different properties.** Both substances shown are compounds of iron. The grey substance on the left is Fe_3O_4, which contains Fe^{2+} and Fe^{3+} ions. The red substance on the right is Fe_2O_3, which contains Fe^{3+} ions.

c. *Cations formed from molecules composed of nonmetal atoms have names that end in* -ium:

NH_4^+	ammonium ion	H_3O^+	hydronium ion

These two ions are the only ions of this kind that we will encounter frequently in the text.

The names and formulas of some common cations are shown in **Table 2.4** and on the back inside cover of the text. The ions on the left side in Table 2.4 are the monatomic ions that do not have more than one possible charge. Those on the right side are either polyatomic cations or cations with more than one possible charge. The Hg_2^{2+} ion is unusual because, even though it is a metal ion, it is not monatomic. It is called the mercury(I) ion because it can be thought of as two Hg^+ ions bound together. The cations that you will encounter most frequently in this text are shown in boldface. You should learn these cations first.

TABLE 2.4 Common Cations[a]

Charge	Formula	Name	Formula	Name
1+	**H^+**	**hydrogen ion**	**NH_4^+**	**ammonium ion**
	Li^+	lithium ion	Cu^+	copper(I) or cuprous ion
	Na^+	**sodium ion**		
	K^+	**potassium ion**		
	Cs^+	cesium ion		
	Ag^+	**silver ion**		
2+	**Mg^{2+}**	**magnesium ion**	Co^{2+}	cobalt(II) or cobaltous ion
	Ca^{2+}	**calcium ion**	**Cu^{2+}**	**copper(II)** or cupric ion
	Sr^{2+}	strontium ion	**Fe^{2+}**	**iron(II)** or ferrous ion
	Ba^{2+}	barium ion	Mn^{2+}	manganese(II) or manganous ion
	Zn^{2+}	**zinc ion**	Hg_2^{2+}	mercury(I) or mercurous ion
	Cd^{2+}	cadmium ion	Hg^{2+}	mercury(II) or mercuric ion
			Ni^{2+}	nickel(II) or nickelous ion
			Pb^{2+}	**lead(II)** or plumbous ion
			Sn^{2+}	tin(II) or stannous ion
3+	**Al^{3+}**	**aluminum ion**	Cr^{3+}	chromium(III) or chromic ion
			Fe^{3+}	**iron(III)** or ferric ion

[a]The ions we use most often in this course are in boldface. Learn them first.

2. Anions

a. *The names of monatomic anions are formed by replacing the ending of the name of the element with* -ide*:*

H^-	hydride ion	O^{2-}	oxide ion	N^{3-}	nitride ion

A few polyatomic anions also have names ending in *-ide*:

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

b. *Polyatomic anions containing oxygen have names ending in either* -ate *or* -ite *and are called* **oxyanions**. The *-ate* is used for the most common or representative oxyanion of an element, and *-ite* is used for an oxyanion that has the same charge but one O atom fewer:

NO_3^-	nitrate ion	SO_4^{2-}	sulfate ion
NO_2^-	nitrite ion	SO_3^{2-}	sulfite ion

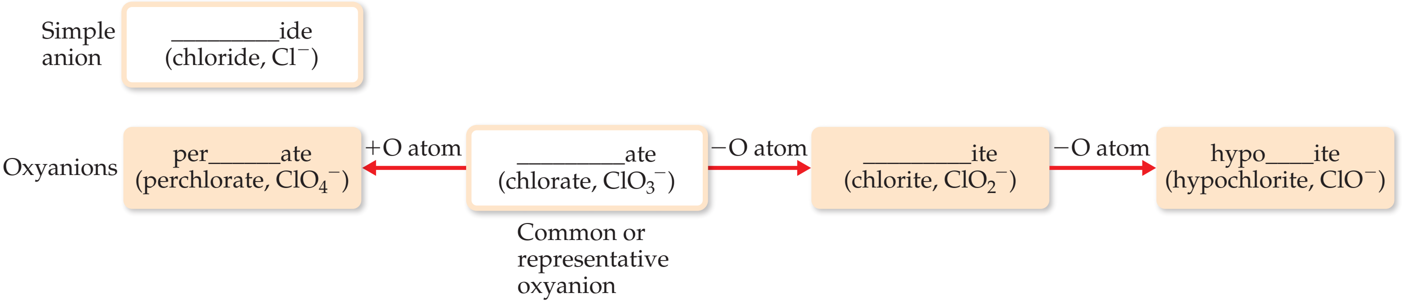
Prefixes are used when the series of oxyanions of an element extends to four members, as with the halogens. The prefix *per-* indicates one more O atom than the oxyanion ending in *-ate; hypo-* indicates one O atom fewer than the oxyanion ending in *-ite*:

ClO_4^-	perchlorate ion (one more O atom than chlorate)
ClO_3^-	chlorate ion
ClO_2^-	chlorite ion (one O atom fewer than chlorate)
ClO^-	hypochlorite ion (one O atom fewer than chlorite)

These rules are summarized in **Figure 2.21**.

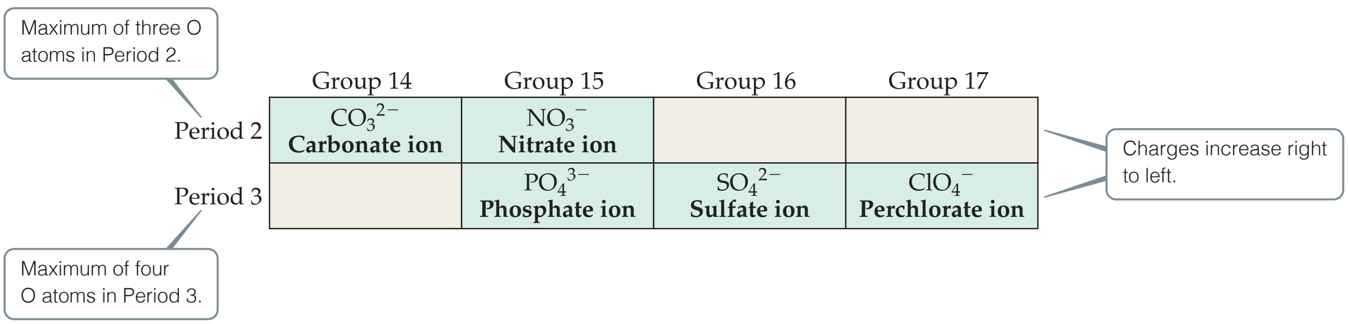
Go Figure

Name the anion obtained by removing one oxygen atom from the perbromate ion, BrO_4^-.



▲ **Figure 2.21 Procedure for naming anions.** Here the simple anion and polyatomic oxyanions derived from chlorine are used as an example.

Figure 2.22 can help you remember the charge and number of oxygen atoms in the various oxyanions. Notice that C and N, both Period 2 elements, have only three O atoms each, whereas the Period 3 elements P, S, and Cl have four O atoms each. Beginning at the lower right in Figure 2.22, note that ionic charge increases from right to left, from 1− for ClO_4^- to 3− for PO_4^{3-}. In the second period the charges also increase from right to left, from 1− for NO_3^- to 2− for CO_3^{2-}. Notice also that although each of the anions in Figure 2.22 ends in *-ate*, the ClO_4^- ion also has a *per-* prefix.



▲ **Figure 2.22 Common oxyanions.** The composition and charges of common oxyanions are related to their location in the periodic table.

Sample Exercise 2.11

Determining the Formula of an Oxyanion from Its Name

Based on the formula for the sulfate ion, predict the formula for (**a**) the selenate ion and (**b**) the selenite ion. (Sulfur and selenium are both in Group 16 and form analogous oxyanions.)

SOLUTION

(**a**) The sulfate ion is SO_4^{2-}. The analogous selenate ion is therefore SeO_4^{2-}.

(**b**) The ending *-ite* indicates an oxyanion with the same charge but one O atom fewer than the corresponding oxyanion that ends in *-ate*. Thus, the formula for the selenite ion is SeO_3^{2-}.

Practice Exercise

Which of the following oxyanions is incorrectly named? (**a**) ClO_2^-, chlorate (**b**) IO_4^-, periodate (**c**) SO_3^{2-}, sulfite (**d**) IO_3^-, iodate (**e**) NO_2^-, nitrite

c. *Anions derived by adding* H^+ *to an oxyanion are named by adding as a prefix the word* hydrogen *or* dihydrogen, *as appropriate*:

CO_3^{2-}	carbonate ion	PO_4^{3-}	phosphate ion
HCO_3^-	hydrogen carbonate ion	$H_2PO_4^-$	dihydrogen phosphate ion

Notice that each H^+ added reduces the negative charge of the parent anion by one. An older method for naming some of these ions uses the prefix *bi-*. Thus, the HCO_3^- ion is commonly called the bicarbonate ion, and HSO_4^- is sometimes called the bisulfate ion.

The names and formulas of the common anions are listed in **Table 2.5** and on the back inside cover of the text. Those anions whose names end in *-ide* are listed on the left portion of Table 2.5, and those whose names end in *-ate* are listed on the right. The most common of these ions are shown in boldface. You should learn the names and formulas of these anions first. The formulas of the ions whose names end with *-ite* can be derived from those ending in *-ate* by removing an O atom. Notice the location of the monatomic ions in the periodic table. Those of Group 17 always have a 1− charge (F^-, Cl^-, Br^-, and I^-), and those of Group 16 have a 2− charge (O^{2-} and S^{2-}).

TABLE 2.5 Common Anions[a]

Charge	Formula	Name	Formula	Name
1−	H^-	hydride ion	**CH_3COO^-** (or $C_2H_3O_2^-$)	**acetate ion**
	F^-	**fluoride ion**	ClO_3^-	chlorate ion
	Cl^-	**chloride ion**	**ClO_4^-**	**perchlorate ion**
	Br^-	**bromide ion**	**NO_3^-**	**nitrate ion**
	I^-	**iodide ion**	MnO_4^-	permanganate ion
	CN^-	cyanide ion		
	OH^-	**hydroxide ion**		
2−	**O^{2-}**	**oxide ion**	**CO_3^{2-}**	**carbonate ion**
	O_2^{2-}	peroxide ion	CrO_4^{2-}	chromate ion
	S^{2-}	**sulfide ion**	$Cr_2O_7^{2-}$	dichromate ion
			SO_4^{2-}	**sulfate ion**
3−	N^{3-}	nitride ion	**PO_4^{3-}**	**phosphate ion**

[a]The ions we use most often are in boldface. Learn them first.

3. Ionic Compounds

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

$CaCl_2$	calcium chloride
$Al(NO_3)_3$	aluminum nitrate
$Cu(ClO_4)_2$	copper(II) perchlorate (or cupric perchlorate)

In the chemical formulas for aluminum nitrate and copper(II) perchlorate, parentheses followed by the appropriate subscript are used because the compounds contain two or more polyatomic ions.

Sample Exercise 2.12

Determining the Names of Ionic Compounds from Their Formulas

Name the ionic compounds (**a**) K_2SO_4, (**b**) $Ba(OH)_2$, (**c**) $FeCl_3$.

SOLUTION

In naming ionic compounds, it is important to recognize polyatomic ions and to determine the charge of cations with variable charge.

(**a**) The cation is K^+, the potassium ion, and the anion is SO_4^{2-}, the sulfate ion, making the name potassium sulfate. (If you thought the compound contained S^{2-} and O^{2-} ions, you failed to recognize the polyatomic sulfate ion.)

(**b**) The cation is Ba^{2+}, the barium ion, and the anion is OH^-, the hydroxide ion: barium hydroxide.

(**c**) You must determine the charge of Fe in this compound because an iron atom can form more than one cation. Because the compound contains three chloride ions, Cl^-, the cation must be Fe^{3+}, the iron(III), or ferric, ion. Thus, the compound is iron(III) chloride or ferric chloride.

Practice Exercise

Which of the following ionic compounds is incorrectly named? (**a**) $Zn(NO_3)_2$, zinc nitrate (**b**) $TeCl_4$, tellurium(IV) chloride (**c**) Fe_2O_3, diiron oxide (**d**) BaO, barium oxide (**e**) $Mn_3(PO_4)_2$, manganese(II) phosphate

Names and Formulas of Acids

Acids are an important class of hydrogen-containing compounds, and they are named in a special way. For our present purposes, an *acid* is a substance whose molecules yield hydrogen ions (H^+) when dissolved in water. When we encounter the chemical formula for an acid at this stage of the course, it will be written with H as the first element, as in HCl and H_2SO_4.

An acid is composed of an anion connected to enough H^+ ions to neutralize, or balance, the anion's charge. Thus, the SO_4^{2-} ion requires two H^+ ions, forming H_2SO_4. The name of an acid is related to the name of its anion, as summarized in **Figure 2.23**.

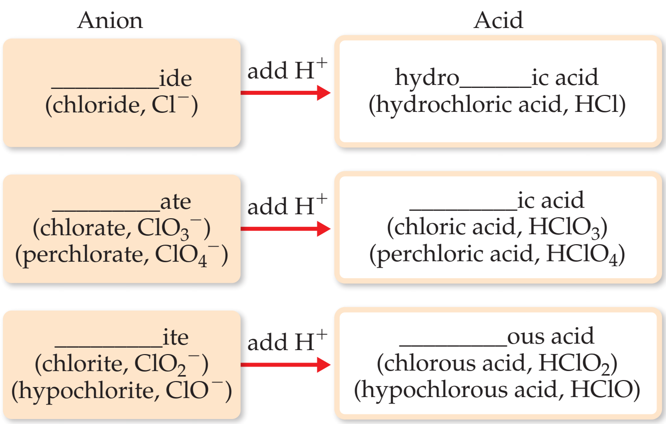
How to Name Acids

1. *Acids containing anions whose names end in* -ide *are named by changing the* -ide *ending to* -ic, *adding the prefix* hydro- *to this anion name, and then following with the word* acid:

Anion	Corresponding Acid
Cl^- (chloride)	HCl (hydrochloric acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)

2. *Acids containing anions whose names end in* -ate *or* -ite *are named by changing* -ate *to* -ic *and* -ite *to* -ous *and then adding the word* acid. Prefixes in the anion name are retained in the name of the acid:

Anion		Corresponding Acid	
ClO_4^-	(perchlorate)	$HClO_4$	(perchloric acid)
ClO_3^-	(chlorate)	$HClO_3$	(chloric acid)
ClO_2^-	(chlorite)	$HClO_2$	(chlorous acid)
ClO^-	(hypochlorite)	HClO	(hypochlorous acid)



▲ **Figure 2.23 Procedure for naming acids.** Here acids containing chlorine are used as an example. Prefixes used for oxyanions, such as, *per-* and *hypo-*, are retained in the acids derived from those anions.

Sample Exercise 2.13
Relating the Names and Formulas of Acids

Name the acids (**a**) HCN, (**b**) HNO_3, (**c**) H_2SO_4, (**d**) H_2SO_3.

SOLUTION

(**a**) The anion from which this acid is derived is CN^-, the cyanide ion. Because this ion has an *-ide* ending, the acid is given a *hydro-* prefix and an *-ic* ending: hydrocyanic acid. Only water solutions of HCN are referred to as hydrocyanic acid. The pure compound, which is a gas under normal conditions, is called hydrogen cyanide. Both hydrocyanic acid and hydrogen cyanide are *extremely* toxic.

(**b**) Because NO_3^- is the nitrate ion, HNO_3 is called nitric acid (the *-ate* ending of the anion is replaced with an *-ic* ending in naming the acid).

(**c**) Because SO_4^{2-} is the sulfate ion, H_2SO_4 is called sulfuric acid.

(**d**) Because SO_3^{2-} is the sulfite ion, H_2SO_3 is sulfurous acid (the *-ite* ending of the anion is replaced with an *-ous* ending).

Practice Exercise

Which of the following acids are incorrectly named? For those that are, provide a correct name or formula.
(**a**) hydrofluoric acid, HF (**b**) nitrous acid, HNO_3 (**c**) perbromic acid, $HBrO_4$ (**d**) iodic acid, HI (**e**) selenic acid, H_2SeO_4

TABLE 2.6 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*

Names and Formulas of Binary Molecular Compounds

The procedures used for naming *binary* (two-element) molecular compounds are similar to those used for naming ionic compounds:

How to Name Binary Molecular Compounds

1. *The name of the element farther to the left in the periodic table (closest to the metals) is usually written first.* An exception occurs when the compound contains oxygen and chlorine, bromine, or iodine (any halogen except fluorine), in which case oxygen is written last.
2. *If both elements are in the same group, the one closer to the bottom of the table is named first.*
3. *The name of the second element is given an* -ide *ending.*
4. *Greek prefixes* (Table 2.6) indicate the number of atoms of each element. (Exception: The prefix *mono-* is never used with the first element.) When the prefix ends in *a* or *o* and the name of the second element begins with *a* vowel, the *a* or *o* of the prefix is often dropped.

The following examples illustrate these rules:

Cl_2O	dichlorine monoxide	NF_3	nitrogen trifluoride
N_2O_4	dinitrogen tetroxide	P_4S_{10}	tetraphosphorus decasulfide

Rule 4 is necessary because we cannot predict formulas for most molecular substances the way we can for ionic compounds. Molecular compounds that contain hydrogen and one other element are an important exception, however. These compounds can be treated as if they were neutral substances containing H^+ ions and anions. Thus, you can predict that the substance named hydrogen chloride has the formula HCl, containing one H^+ to balance the charge of one Cl^-. (The name *hydrogen chloride* is used only for the pure compound; water solutions of HCl are called hydrochloric acid. The distinction, which is important, will be explained in Section 4.1.) Similarly, the formula for hydrogen sulfide is H_2S because two H^+ ions are needed to balance the charge on S^{2-}.

Sample Exercise 2.14

Relating the Names and Formulas of Binary Molecular Compounds

Name the compounds (**a**) SO_2, (**b**) PCl_5, (**c**) Cl_2O_3.

SOLUTION

The compounds consist entirely of nonmetals, so they are molecular rather than ionic. Using the prefixes in Table 2.6 , we have (**a**) sulfur dioxide, (**b**) phosphorus pentachloride, (**c**) dichlorine trioxide.

Practice Exercise

Give the name for each of the following binary compounds of carbon: (**a**) CS_2, (**b**) CO, (**c**) C_3O_2, (**d**) CBr_4, (**e**) CF.

Self-Assessment Exercise

2.38 What is the name of $NaClO_2$?

(**a**) Sodium chlorine oxide

(**b**) Sodium chlorite

(**c**) Sodium chlorate

Exercises

2.39 Give the chemical formula for (**a**) chromate ion (**b**) bromide ion (**c**) nitrite ion (**d**) sulphite ion (**e**) permanganate ion.

2.40 Give the names and charges of the cation and anion in each of the following compounds: (**a**) CaO, (**b**) Na_2SO_4, (**c**) $KClO_4$, (**d**) $Fe(NO_3)_2$, (**e**) $Cr(OH)_3$.

2.41 Name the following ionic compounds: (**a**) Li_2O, (**b**) $FeCl_3$, (**c**) NaClO, (**d**) $CaSO_3$, (**e**) $Cu(OH)_2$, (**f**) $Fe(NO_3)_2$, (**g**) $Ca(CH_3COO)_2$, (**h**) $Cr_2(CO_3)_3$, (**i**) K_2CrO_4, (**j**) $(NH_4)_2SO_4$.

2.42 Write the chemical formulas for the following compounds: (**a**) aluminum hydroxide, (**b**) potassium sulfate, (**c**) copper(I) oxide, (**d**) zinc nitrate, (**e**) mercury(II) bromide, (**f**) iron(III) carbonate, (**g**) sodium hypobromite.

2.43 Give the name or chemical formula, as appropriate, for each of the following acids: (**a**) $HBrO_3$, (**b**) HBr, (**c**) H_3PO_4, (**d**) hypochlorous acid, (**e**) iodic acid, (**f**) sulfurous acid.

2.44 Give the name or chemical formula, as appropriate, for each of the following binary molecular substances: (**a**) SF_6, (**b**) IF_5, (**c**) XeO_3, (**d**) dinitrogen tetroxide, (**e**) hydrogen cyanide, (**f**) tetraphosphorus hexasulfide.

2.45 Write the chemical formula for each substance mentioned in the following word descriptions (use the front inside cover to find the symbols for the elements you do not know). (**a**) Zinc carbonate can be heated to form zinc oxide and carbon dioxide. (**b**) On treatment with hydrofluoric acid, silicon dioxide forms silicon tetrafluoride and water. (**c**) Sulfur dioxide reacts with water to form sulfurous acid. (**d**) The substance phosphorus trihydride, commonly called phosphine, is a toxic gas. (**e**) Perchloric acid reacts with cadmium to form cadmium(II) perchlorate. (**f**) Vanadium(III) bromide is a colored solid.

Answers to Self-Assessment Exercises

2.37 (b)

2.9 | Some Simple Organic Compounds

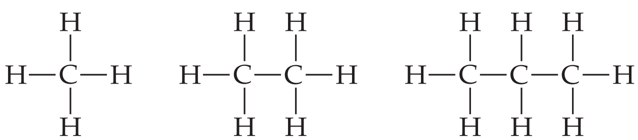
Organic compounds were believed to require living cells for their synthesis, either plant or animal in origin. Now there are many ways in which we may synthesize organic compound from inorganic materials. They represent over 99% of all known compounds and each one has carbon as a central element. By the end of this section, you should be able to

- Recognize and name simple alkanes and alcohols

The study of compounds of carbon is called **organic chemistry**, and as noted earlier, compounds that contain carbon and hydrogen, often in combination with oxygen, nitrogen, or other elements, are called *organic compounds*. Organic compounds are a very important part of chemistry, far outnumbering all other types of chemical substances. We will examine organic compounds in a systematic way later on, but you will encounter many examples of them throughout the text. Here we present a brief introduction to some of the simplest organic compounds and the ways in which they are named.

Alkanes

Compounds that contain only carbon and hydrogen are called **hydrocarbons**. In the simplest class of hydrocarbons, **alkanes**, each carbon is bonded to four other atoms. The three smallest alkanes are methane (CH_4), ethane (C_2H_6), and propane (C_3H_8). The structural formulas of these three alkanes are as follows:

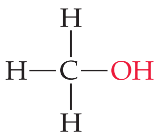


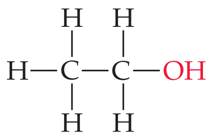
Methane Ethane Propane

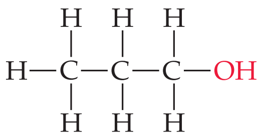
Although hydrocarbons are binary molecular compounds, they are not named like the binary inorganic compounds discussed in Section 2.8. Instead, each alkane has a name that ends in *-ane*. The alkane with four carbons is called *butane*. For alkanes with five or more carbons, the names are derived from prefixes like those in Table 2.6. An alkane with eight carbon atoms, for example, is *octane* (C_8H_{18}), where the *octa-* prefix for eight is combined with the *-ane* ending for an alkane.

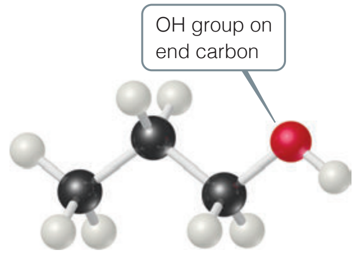
Some Derivatives of Alkanes

Other classes of organic compounds are obtained when one or more hydrogen atoms in an alkane are replaced with *functional groups*, which are specific groups of atoms. An **alcohol**, for example, is obtained by replacing an H atom of an alkane with an —OH group. The name of the alcohol is derived from that of the alkane by adding an *-ol* ending:

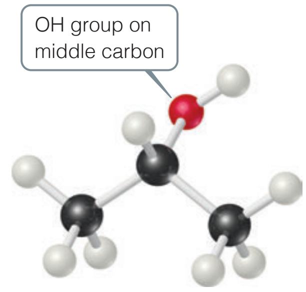

Methanol


Ethanol


1-Propanol



1-Propanol



2-Propanol

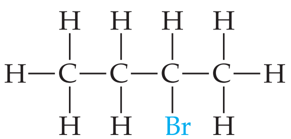
▲ **Figure 2.24 The two forms (isomers) of propanol.**

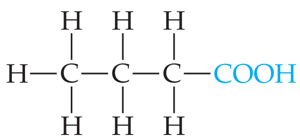
Alcohols have properties that are very different from those of the alkanes from which the alcohols are obtained. For example, methane, ethane, and propane are all colorless gases under normal conditions, whereas methanol, ethanol, and propanol are colorless liquids. We will discuss the reasons for these differences in Chapter 11.

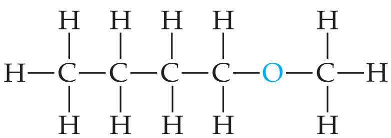
The prefix “1” in the name 1-propanol indicates that the replacement of H with OH has occurred at one of the “outer” carbon atoms rather than the “middle” carbon atom. A different compound, called 2-propanol is obtained when the OH functional group is attached to the middle carbon atom (**Figure 2.24**).

Compounds with the same molecular formula but different arrangements of atoms are called **isomers**. For example, 1-propanol and 2-propanol are *structural isomers*, compounds that have the same molecular formula but different structural formulas. There are many different kinds of isomers, as we will discover later in this book.

As already noted, many different functional groups can replace one or more of the hydrogens on an alkane—for example, one or more of the halogens, or a special grouping of carbon and oxygen atoms, such as the carboxylic acid group, —COOH. Here are a few examples of functional groups you will be encountering in the chapters that lie ahead (the functional group is highlighted in blue):


2-Bromobutane


Butyric acid


Methyl butyl ether

Much of the richness of organic chemistry is possible because organic compounds can form long chains of carbon–carbon bonds. The series of alkanes that begins with methane, ethane, and propane and the series of alcohols that begins with methanol, ethanol, and propanol can both be extended for as long as we desire, in principle. The properties of alkanes and alcohols change as the chains get longer. Octanes, which are alkanes with eight carbon atoms, are liquids under normal conditions. If the alkane series is extended to tens of thousands of carbon atoms, we obtain *polyethylene*, a solid substance that is used to make thousands of plastic products, such as plastic bags, food containers, and laboratory equipment.

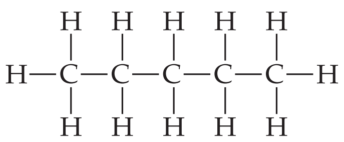
Sample Exercise 2.15

Writing Structural and Molecular Formulas for Hydrocarbons

Assuming the carbon atoms in *pentane* are in a linear chain, write (**a**) the structural formula and (**b**) the molecular formula for this alkane.

SOLUTION

(**a**) Alkanes contain only carbon and hydrogen, and each carbon is attached to four other atoms. The name *pentane* contains the prefix *penta-* for five (Table 2.6), and we are told that the carbons are in a linear chain. If we then add enough hydrogen atoms to make four bonds to each carbon, we obtain the structural formula



This form of pentane is often called *n*-pentane, where the *n*- stands for "normal" because all five carbon atoms are in one line in the structural formula.

(**b**) Once the structural formula is written, we determine the molecular formula by counting the atoms present. Thus, *n*-pentane has the molecular formula C_5H_{12}.

Practice Exercise

(**a**) What is the molecular formula of hexane, the alkane with six carbons? (**b**) What are the name and molecular formula of an alcohol derived from hexane?

STRATEGIES FOR SUCCESS How to Take a Test

At about this time in your study of chemistry, you are likely to face your first examination. The best way to prepare is to study, do homework diligently, and get help from the instructor on any material that is unclear or confusing. (See the advice for learning and studying chemistry presented in the preface of the book.) We present here some general guidelines for taking tests.

Depending on the nature of your course, the exam could consist of a variety of different types of questions.

1. **Multiple-choice questions** In large-enrollment courses, the most common kind of test question is the multiple-choice question. Many of the practice exercise problems in this book are written in this format to give you practice at this style of question. When faced with this type of problem, the first thing to realize is that the instructor has written the question so that at first glance all the answers appear to be correct. Thus, you should not jump to the conclusion that because one of the choices looks correct, it must be correct.

   If a multiple-choice question involves a calculation, do the calculation, check your work, and *only then* compare your answer with the choices. Keep in mind, though, that your instructor has anticipated the most common errors you might make in solving a given problem and has probably listed the incorrect answers resulting from those errors. Always double-check your reasoning and use dimensional analysis to arrive at the correct numeric answer and the correct units.

   In multiple-choice questions that do not involve calculations, if you are not sure of the correct choice, eliminate all the choices you know for sure to be incorrect. The reasoning you use in eliminating incorrect choices may offer insight into which of the remaining choices is correct.

2. **Calculations in which you must show your work** In questions of this kind, you may receive partial credit even if you do not arrive at the correct answer, depending on whether the instructor can follow your line of reasoning. It is important, therefore, to be neat and organized in your calculations. Pay particular attention to what information is given and to what your unknown is. Think about how you can get from the given information to your unknown.

   You may want to write a few words or a diagram on the test paper to indicate your approach. Then write out your calculations as neatly as you can. Show the units for every number you write down, and use dimensional analysis as much as you can, showing how units cancel.

3. **Questions requiring drawings** Questions of this kind will come later in the course, but it is useful to talk about them here. (You should review this box before each exam to remind yourself of good exam-taking practices.) Be sure to label your drawing as completely as possible.

Self-Assessment Exercise

2.46 Do isomers have the same molecular formula?

(a) Yes

(b) No

Exercises

2.47 (**a**) What is a hydrocarbon? (**b**) Pentane is the alkane with a chain of five carbon atoms. Write a structural formula for this compound and determine its molecular and empirical formulas.

2.48 (**a**) What is a functional group? (**b**) What functional group characterizes an alcohol? (**c**) Write a structural formula for 1-pentanol, the alcohol derived from pentane by making a substitution on one of the carbon atoms.

2.49 Chloropropane is derived from propane by substituting Cl for H on one of the carbon atoms. (**a**) Draw the structural formulas for the two isomers of chloropropane. (**b**) Suggest names for these two compounds.

Answers to Self-Assessment Exercises

2.45 (a)

Chapter Summary and Key Terms

THE ATOMIC THEORY OF MATTER; THE DISCOVERY OF ATOMIC STRUCTURE (SECTION 2.1 AND 2.2) **Atoms** are the basic building blocks of matter. They are the smallest units of an element that can combine with other elements. Atoms are composed of even smaller particles, called **subatomic particles**. Some of these subatomic particles are charged and follow the usual behavior of charged particles: Particles with the same charge repel one another, whereas particles with opposite charges are attracted to one another.

We considered some of the important experiments that led to the discovery and characterization of subatomic particles. Thomson's experiments on the behavior of **cathode rays** in magnetic and electric fields led to the discovery of the electron and allowed its charge-to-mass ratio to be measured. Millikan's oil-drop experiment determined the charge of the electron. Becquerel's discovery of **radioactivity**, the spontaneous emission of radiation by atoms, gave further evidence that the atom has a substructure. Rutherford's studies of how α particles scatter when passing through thin metal foils led to the **nuclear model** of the atom, showing that the atom has a dense, positively charged **nucleus**.

THE MODERN VIEW OF ATOMIC STRUCTURE (SECTION 2.3) Atoms have a nucleus that contains **protons** and **neutrons**; **electrons** move in the space around the nucleus. The magnitude of the charge of the electron, 1.602×10^{-19} C, is called the **electronic charge**. The charges of particles are usually represented as multiples of this charge—an electron has a 1− charge, and a proton has a 1+ charge. The masses of atoms are usually expressed in terms of **atomic mass units** ($1\ u = 1.66054 \times 10^{-24}$ g).

Elements can be classified by **atomic number**, the number of protons in the nucleus of an atom. All atoms of a given element have the same atomic number. The **mass number** of an atom is the sum of the numbers of protons and neutrons. Atoms of the same element that differ in mass number are known as **isotopes**.

ATOMIC WEIGHTS (SECTION 2.4) The atomic mass scale is defined by assigning a mass of exactly 12 u to a ^{12}C atom. The **atomic weight** (average atomic mass) of an element can be calculated from the relative abundances and masses of that element's isotopes. The **mass spectrometer** provides the most direct and accurate means of experimentally measuring atomic (and molecular) weights.

THE PERIODIC TABLE (SECTION 2.5) The **periodic table** is an arrangement of the elements in order of increasing atomic number. Elements with similar properties are placed in vertical columns. The elements in a column are known as a **group**. The elements in a horizontal row are known as a **period**. The **metallic elements** (**metals**), which comprise the majority of the elements, dominate the left side and the middle of the table; the **nonmetallic elements (nonmetals)** are located on the upper right side. Many of the elements that lie along the line that separates metals from nonmetals are **metalloids**.

MOLECULES AND MOLECULAR COMPOUNDS (SECTION 2.6) Atoms can combine to form **molecules**. Compounds composed of molecules (**molecular compounds**) usually contain only nonmetallic elements. A molecule that contains two atoms is called a **diatomic molecule**. The composition of a substance is given by its **chemical formula**. A molecular substance can be represented by its **empirical formula**, which gives the relative numbers of atoms of each kind, but is usually represented by its **molecular formula**, which gives the actual numbers of each type of atom in a molecule. **Structural formulas** show the order in which the atoms in a molecule are connected. **Ball-and-stick models** and **space-filling models** convey additional information about the shapes of molecules.

IONS AND IONIC COMPOUNDS (SECTION 2.7) Atoms can either gain or lose electrons, forming charged particles called **ions**. Metals tend to lose electrons, becoming positively charged ions (**cations**). Nonmetals tend to gain electrons, forming negatively charged ions (**anions**). Because **ionic compounds** are electrically neutral, containing both cations and anions, they usually contain both metallic and nonmetallic elements. Atoms that are joined together, as in a molecule, but carry a net charge are called **polyatomic ions**. The chemical formulas used for ionic compounds are empirical formulas, which can be written readily if the charges of the ions are known. The total positive charge of the cations in an ionic compound must equal the total negative charge of the anions.

NAMING INORGANIC COMPOUNDS (SECTION 2.8) The set of rules for naming chemical compounds is called **chemical nomenclature**. We studied the systematic rules used for naming three classes of inorganic substances: ionic compounds, acids, and binary molecular compounds.

In naming an ionic compound, the cation is named first and then the anion. Cations formed from metal atoms have the same name as the metal. If the metal can form cations of differing charges, the charge is given using Roman numerals. Monatomic anions have names ending in *-ide*. Polyatomic anions containing oxygen and another element (**oxyanions**) have names ending in *-ate* or *-ite*. In naming binary molecular compounds, Greek prefixes are used to denote the number of each element in the molecular formula, and the element farthest to the left in the periodic table (closest to metallic elements) is generally written first.

SOME SIMPLE ORGANIC COMPOUNDS (SECTION 2.9) **Organic chemistry** is the study of compounds that contain carbon. The simplest class of organic molecules is the **hydrocarbons**, which contain only carbon and hydrogen. Hydrocarbons in which each carbon atom is attached to four other atoms are called **alkanes**. Alkanes have names that end in *-ane*, such as methane and ethane. Other organic compounds are formed when an H atom of a hydrocarbon is replaced with a functional group. An **alcohol**, for example, is a compound in which an H atom of a hydrocarbon is replaced by an OH functional group. Alcohols have names that end in *-ol*, such as methanol and ethanol. Compounds with the same molecular formula but different bonding arrangements of their constituent atoms are called **isomers**.

Learning Outcomes After studying this chapter, you should be able to:

- List the basic postulates of Dalton's atomic theory. (Section 2.1) *Related Exercises: 2.2, 2.3, 2.60, 2.61*
- Describe the key experiments that led to the discovery of electrons and to the nuclear model of the atom. (Section 2.2) *Related Exercises: 2.5, 2.6, 2.62, 2.63*
- Describe the structure of the atom in terms of protons, neutrons, and electrons and express the relative electrical charges and masses of these subatomic particles. (Section 2.3) *Related Exercises: 2.10, 2.65*
- Use chemical symbols together with atomic number and mass number to express the subatomic composition of isotopes. (Section 2.3) *Related Exercises: 2.11, 2.12, 2.14, 2.33*
- Calculate the atomic weight of an element from the masses of individual atoms and a knowledge of natural abundances. (Section 2.4) *Related Exercises: 2.18, 2.20, 2.71, 2.72*
- Describe how elements are organized in the periodic table by atomic number and by similarities in chemical behavior, giving rise to periods and groups. (Section 2.5) *Related Exercises: 2.23, 2.76*
- Identify the locations of metals and nonmetals in the periodic table. (Section 2.5) *Related Exercises: 2.22, 2.75*
- Distinguish between molecular substances and ionic substances in terms of their composition. (Section 2.6 and 2.7) *Related Exercises: 2.32, 2.36, 2.37, 2.87*
- Distinguish between empirical formulas and molecular formulas. (Section 2.6) *Related Exercises: 2.25, 2.26, 2.27, 2.77*
- Describe how molecular formulas and structural formulas are used to represent the compositions of molecules. (Section 2.6) *Related Exercises: 2.29, 2.81*
- Explain how ions are formed by the gain or loss of electrons and use the periodic table to predict the charges of common ions. (Section 2.7) *Related Exercises: 2.34, 2.82, 2.83*
- Write the empirical formulas of ionic compounds, given the charges of their component ions. (Section 2.7) *Related Exercises: 2.35, 2.36, 2.37, 2.85*
- Write the name of an ionic compound given its chemical formula or write the chemical formula given its name. (Section 2.8) *Related Exercises: 2.41, 2.42, 2.90, 2.91*
- Name or write chemical formulas for binary inorganic compounds and for acids. (Section 2.8) *Related Exercises: 2.43, 2.44, 2.92, 2.93*
- Identify organic compounds and name simple alkanes and alcohols. (Section 2.9) *Related Exercises: 2.47, 2.48, 2.49, 2.96*

Key Equations

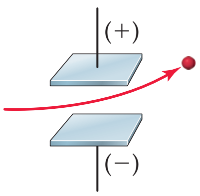
- $$\text{Atomic weight} = \sum_{\substack{\text{over}\\ \text{all isotopes}}} [(\text{isotope mass}) \times (\text{fractional isotope abundance})] \qquad [2.1]$$ Calculating atomic weight as a fractionally weighted average of isotopic masses.

Exercises

Visualizing Concepts

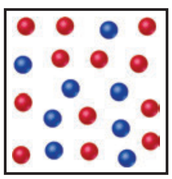
These exercises are intended to probe your understanding of key concepts rather than your ability to utilize formulas and perform calculations. Exercises with red numbers have answers in the back of the book.

2.50 A charged particle moves between two electrically charged plates, as shown here.

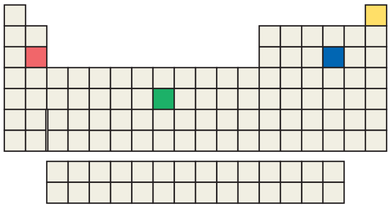


(**a**) What is the sign of the electrical charge on the particle? (**b**) As the charge on the plates is increased, would you expect the bending to increase, decrease, or stay the same? (**c**) As the mass of the particle is increased while the speed of the particles remains the same, would you expect the bending to increase, decrease, or stay the same? [Section 2.2]

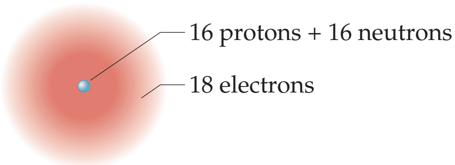
2.51 The following diagram is a representation of 20 atoms of a fictitious element, which we will call nevadium (Nv). The red spheres are ^{293}Nv, and the blue spheres are ^{295}Nv. (**a**) Assuming that this sample is a statistically representative sample of the element, calculate the percent abundance of each element. (**b**) If the mass of ^{293}Nv is 293.15 u and that of ^{295}Nv is 295.15 u, what is the atomic weight of Nv? [Section 2.4]



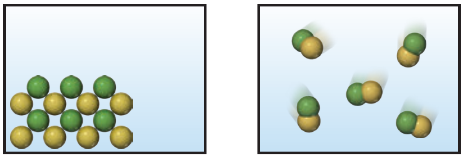
2.52 Four of the boxes in the following periodic table are colored. Which of these are metals and which are nonmetals? Which one is an alkaline earth metal? Which one is a noble gas? [Section 2.5]



2.53 Does the following drawing represent a neutral atom or an ion? Write its complete chemical symbol, including mass number, atomic number, and net charge (if any). [Sections 2.3 and 2.7]

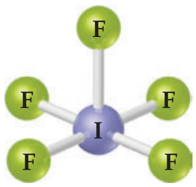


2.54 Which of the following diagrams most likely represents an ionic compound, and which represents a molecular one? Explain your choice. [Sections 2.6 and 2.7]

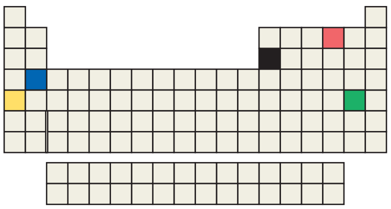


(i) (ii)

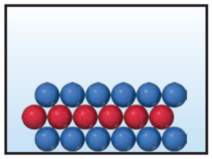
2.55 Write the chemical formula for the following compound. Is the compound ionic or molecular? Name the compound. [Sections 2.6 and 2.8]



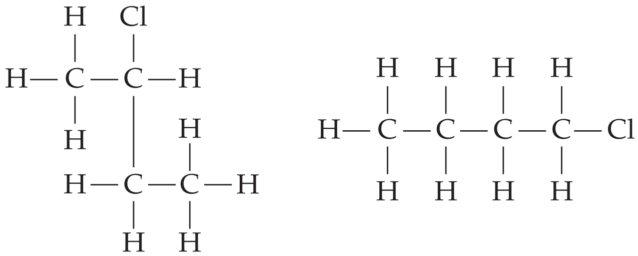
2.56 Five of the boxes in the following periodic table are colored. Predict the charge on the ion associated with each of these elements. [Section 2.7]



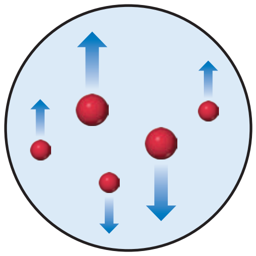
2.57 The following diagram represents an ionic compound in which the red spheres represent cations and the blue spheres represent anions. Which of the following formulas is consistent with the drawing? KBr, K_2SO_4, $Ca(NO_3)_2$, $Fe_2(SO_4)_3$. Name the compound. [Sections 2.7 and 2.8]



2.58 Are these two compounds isomers? Explain. [Section 2.9]



2.59 In the Millikan oil-drop experiment (see Figure 2.4), the tiny oil drops are observed through the viewing lens as rising, stationary, or falling, as shown here. (**a**) What causes their rate of fall to vary from their rate in the absence of an electric field? (**b**) Why do some drops move upward? [Section 2.2]



The following exercises are divided into sections that deal with specific topics in the chapter. The exercises are grouped in pairs, with the answers given in the back of the book to the odd-numbered exercises, as indicated by the red exercise numbers. Those exercises whose numbers appear in brackets are more challenging than the nonbracketed exercises.

The Atomic Theory of Matter and the Discovery of Atomic Structure (Sections 2.1 and 2.2)

2.60 Sodium reacts with oxygen in air to form two compounds: sodium oxide and sodium peroxide. In forming sodium oxide, 23.0 g of sodium combines with 8.0 g of hydrogen. In forming sodium peroxide, 23.0 g of sodium combines with 16.0 g of oxygen. (**a**) What are the mass ratios of oxygen in the two compounds? (**b**) What fundamental law does this experiment demonstrate?

2.61 In a series of experiments, a chemist prepared three different compounds that contain only iodine and fluorine and determined the mass of each element in each compound:

Compound	Mass of Iodine (g)	Mass of Fluorine (g)
1	4.75	3.56
2	7.64	3.43
3	9.41	9.86

(**a**) Calculate the mass of fluorine per gram of iodine in each compound. (**b**) How do the numbers in part (a) support the atomic theory?

2.62 Discovering which of the three subatomic particles proved to be the most difficult—the proton, neutron, or electron? Why?

2.63 Millikan determined the charge on the electron by studying the static charges on oil drops falling in an electric field (Figure 2.4). A student carried out this experiment using several oil drops for her measurements and calculated the charges on the drops. She obtained the following data:

Droplet	Calculated Charge (C)
A	1.60×10^{-19}
B	3.15×10^{-19}
C	4.81×10^{-19}
D	6.31×10^{-19}

(**a**) What is the significance of the fact that the droplets carried different charges? (**b**) What conclusion can the student draw from these data regarding the charge of the electron? (**c**) What value (and to how many significant figures) should she report for the electronic charge?

The Modern View of Atomic Structure; Atomic Weights (Sections 2.3 and 2.4)

2.64 The radius of an atom of copper (Cu) is about 140 pm. (**a**) Express this distance in millimeters (mm). (**b**) How many Cu atoms would have to be placed side by side to span a distance of 5.0 mm? (**c**) If you assume that the Cu atom is a sphere, what is the volume in cm^3 of a single atom?

2.65 Determine whether each of the following statements is true or false. If false, correct the statement to make it true: (**a**) The nucleus has most of the mass and comprises most of the volume of an atom. (**b**) Every atom of a given element has the same number of protons. (**c**) The number of electrons in an atom equals the number of neutrons in the atom. (**d**) The protons in the nucleus of the helium atom are held together by a force called the strong nuclear force.

2.66 Consider an atom of ^{58}Ni. (**a**) How many protons, neutrons, and electrons does this atom contain? (**b**) What is the symbol of the ion obtained by removing two electrons from ^{58}Ni? (**c**) What is the symbol for the isotope of ^{58}Ni that possesses 33 neutrons?

2.67 (**a**) Which two of the following are isotopes of the same element: $^{106}_{46}X$, $^{107}_{46}X$, $^{107}_{47}X$? (**b**) What is the identity of the element whose isotopes you have selected?

2.68 Each of the following isotopes is used in medicine. Indicate the number of protons and neutrons in each isotope: (**a**) samarium-153, (**b**) lutetium-177, (**c**) bismuth-213, (**d**) molybdenum-99, (**e**) lead-212, (**f**) caesium-131.

2.69 Fill in the gaps in the following table, assuming each column represents a neutral atom.

Symbol	^{89}Y				
Protons		78			89
Neutrons			123		
Electrons			81	50	
Mass no.		195		119	227

2.70 One way in which Earth's evolution as a planet can be understood is by measuring the amounts of certain isotopes in rocks. One quantity recently measured is the ratio of ^{129}Xe to ^{130}Xe in some minerals. In what way do these two isotopes differ from one another? In what respects are they the same?

2.71 (**a**) What is the mass in u of a carbon-12 atom? (**b**) Why is the atomic weight of carbon reported as 12.011 in the table of elements and the periodic table in the front inside cover of this text?

2.72 Bromine has two naturally occurring isotopes, bromine-79 (atomic mass = 78.9183 u; abundance = 50.69%) and bromine-81 (atomic mass = 80.9163 u; abundance = 49.31%). Calculate the atomic weight of bromine.

2.73 Consider the mass spectrometer shown in Figure 2.10. Determine whether each of the following statements is true or false. If false, correct the statement to make it true: (**a**) The paths of neutral (uncharged) atoms are not affected by the magnet. (**b**) The height of each peak in the mass spectrum is inversely proportional to the mass of that isotope. (**c**) For a given element, the number of peaks in the spectrum is equal to the number of naturally occurring isotopes of that element.

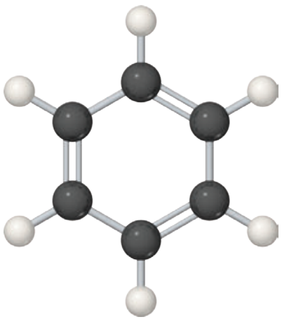
2.74 Mass spectrometry is more often applied to molecules than to atoms. We will see in Chapter 3 that the *molecular weight* of a molecule is the sum of the atomic weights of the atoms in the molecule. The mass spectrum of H_2 is taken under conditions that prevent decomposition into H atoms. The two naturally occurring isotopes of hydrogen are ^{1}H (atomic mass = 1.00783 u; abundance 99.9885%) and ^{2}H (atomic mass = 2.01410 u; abundance 0.0115%). (**a**) How many peaks will the mass spectrum have? (**b**) Give the relative atomic masses of each of these peaks. (**c**) Which peak will be the largest, and which the smallest?

The Periodic Table, Molecules, Molecular Compounds, Ions, and Ionic Compounds (Sections 2.5, 2.6, and 2.7)

2.75 Locate each of the following elements in the periodic table; give its name and atomic number, and indicate whether it is a metal, metalloid, or nonmetal: (**a**) Hg, (**b**) At, (**c**) Mo, (**d**) W, (**e**) Sn, (**f**) V, (**g**) K.

2.76 For each of the following elements, write its chemical symbol, determine the name of the group to which it belongs (Table 2.3), and indicate whether it is a metal, metalloid, or nonmetal: (**a**) polonium, (**b**) strontium, (**c**) neon, (**d**) rubidium, (**e**) bromine.

2.77 Ball-and-stick representations of benzene, a colorless liquid often used in organic chemistry reactions, and acetylene, a gas used as a fuel for high-temperature welding, are shown here. (**a**) Determine the molecular formula of each. (**b**) Determine the empirical formula of each.


benzene

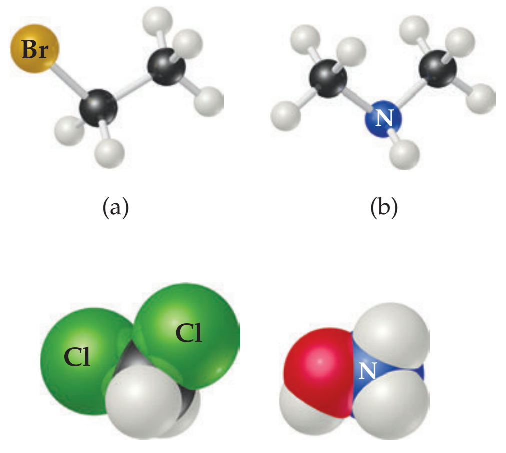

acetylene

2.78 Two substances have the same molecular and empirical formulas. Does this mean that they must be the same compound?

2.79 Determine the molecular and empirical formulas of the following: (**a**) the organic solvent *benzene*, which has six carbon atoms and six hydrogen atoms; (**b**) the compound *silicon tetrachloride*, which has a silicon atom and four chlorine atoms and is used in the manufacture of computer chips; (**c**) the reactive substance *diborane*, which has two boron atoms and six hydrogen atoms; (**d**) the sugar called *glucose,* which has six carbon atoms, twelve hydrogen atoms, and six oxygen atoms.

2.80 How many of the indicated atoms are represented by each chemical formula: (**a**) carbon atoms in $C_4H_9COOCH_3$, (**b**) oxygen atoms in $Ca(ClO_3)_2$, (**c**) hydrogen atoms in $(NH_4)_2HPO_4$?

2.81 Write the molecular and structural formulas for the compounds represented by the following models:



(a) (b) (c) (d)

2.82 Fill in the gaps in the following table:

Symbol	$^{133}Cs^+$			
Protons		35	15	
Neutrons		46	16	30
Electrons			18	20
Net charge		1−		5+

2.83 Using the periodic table, predict the charge of the most stable ion of the following elements: (**a**) Li, (**b**) Ba, (**c**) Po, (**d**) I, (**e**) Sb.

2.84 The most common charge associated with selenium is 2−. Indicate the chemical formulas you would expect for compounds formed between selenium and (**a**) barium, (**b**) lithium, (**c**) aluminum.

2.85 Predict the chemical formulas of the compounds formed by the following pairs of ions: (**a**) Cr^{3+} and CN^-, (**b**) Mn^{2+} and ClO_4^-, (**c**) Na^+ and $Cr_2O_7^{2-}$, (**d**) Cd^{2+} and CO_3^{2-}, (**e**) Ti^{4+} and O^{2-}.

2.86 Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

Ion	Na^+	Ca^{2+}	Fe^{2+}	Al^{3+}
O^{2-}	Na_2O			
NO_3^-				
SO_4^{2-}				
AsO_4^{3-}				

2.87 Predict whether each of the following compounds is molecular or ionic: (**a**) BI_3 (**b**) $N(CH_3)_3$ (**c**) $Zr(NO_3)_2$ (**d**) N_2H_4 (**e**) $OsCO_3$ (**f**) H_2SO_4 (**g**) HgS (**h**) IOH.

Naming Inorganic Compounds; Some Simple Organic Compounds (Sections 2.8 and 2.9)

2.88 Selenium, an element required nutritionally in trace quantities, forms compounds analogous to sulfur. Name the following ions: (**a**) SeO_4^{2-}, (**b**) Se^{2-}, (**c**) HSe^-, (**d**) $HSeO_3^-$.

2.89 Give the names and charges of the cation and anion in each of the following compounds: (**a**) CuS, (**b**) Ag_2SO_4, (**c**) $Al(ClO_3)_3$, (**d**) $Co(OH)_2$, (**e**) $PbCO_3$.

2.90 Name the following ionic compounds: (**a**) KCN, (**b**) $NaBrO_2$, (**c**) $Sr(OH)_2$, (**d**) CoTe, (**e**) $Fe_2(CO_3)_3$, (**f**) $Cr(NO_3)_3$, (**g**) $(NH_4)_2SO_3$, (**h**) NaH_2PO_4, (**i**) $KMnO_4$, (**j**) $Ag_2Cr_2O_7$.

2.91 Give the chemical formula for each of the following ionic compounds: (**a**) sodium phosphate, (**b**) zinc nitrate, (**c**) barium bromate, (**d**) iron(II) perchlorate, (**e**) cobalt(II) hydrogen carbonate, (**f**) chromium(III) acetate, (**g**) potassium dichromate.

2.92 Provide the name or chemical formula, as appropriate, for each of the following acids: (**a**) hydroiodic acid, (**b**) chloric acid, (**c**) nitrous acid, (**d**) H_2CO_3, (**e**) $HClO_4$, (**f**) CH_3COOH.

2.93 The oxides of nitrogen are very important components in urban air pollution. Name each of the following compounds: (**a**) N_2O, (**b**) NO, (**c**) NO_2, (**d**) N_2O_5, (**e**) N_2O_4.

2.94 Assume that you encounter the following sentences in your reading. What is the chemical formula for each substance mentioned? (**a**) Sodium hydrogen carbonate is used as a deodorant. (**b**) Calcium hypochlorite is used in some bleaching solutions. (**c**) Hydrogen cyanide is a very poisonous gas. (**d**) Magnesium hydroxide is used as a cathartic. (**e**) Tin(II) fluoride has been used as a fluoride additive in toothpastes. (**f**) When cadmium sulfide is treated with sulfuric acid, fumes of hydrogen sulfide are given off.

2.95 (**a**) What is meant by the term *isomer*? (**b**) Among the four alkanes, ethane, propane, butane, and pentane, which is capable of existing in isomeric forms?

2.96 Consider the following organic substances: ethylethanoate, ethylmethylether, hexanol, and propanone. (**a**) Which of these molecules contains three carbons? (**b**) Which of these molecules contain a C = O group?

2.97 Draw the structural formulas for four structural isomers of C_4H_9Br.

Additional Exercises

These exercises are not divided by category, although they are roughly in the order of the topics in the chapter.

2.98 Suppose a scientist repeats the Millikan oil-drop experiment but reports the charges on the drops using an unusual (and imaginary) unit called the *warmomb* (wa). The scientist obtains the following data for four of the drops:

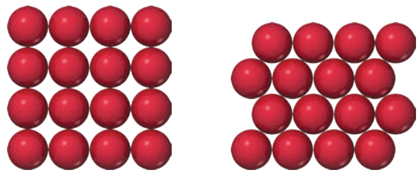
Droplet	Calculated Charge (wa)
A	3.84×10^{-8}
B	4.80×10^{-8}
C	2.88×10^{-8}
D	8.64×10^{-8}

(**a**) If all the droplets were the same size, which would fall most slowly through the apparatus? (**b**) From these data, what is the best choice for the charge of the electron in warmombs? (**c**) Based on your answer to part (b), how many electrons are there on each of the droplets? (**d**) What is the conversion factor between warmombs and coulombs?

2.99 The natural abundance of ^{3}He is 0.000137%. (**a**) How many protons, neutrons, and electrons are in an atom of ^{3}He? (**b**) Based on the sum of the masses of their subatomic particles, which is expected to be more massive, an atom of ^{3}He or an atom of ^{3}H (which is also called *tritium*)? (**c**) Based on your answer to part (b), what would need to be the precision of a mass spectrometer that is able to differentiate between peaks that are due to $^{3}He^{+}$ and $^{3}H^{+}$?

2.100 A cube of gold that is 1.00 cm on a side has a mass of 19.3 g. A single gold atom has a mass of 197.0 u. (**a**) How many gold atoms are in the cube? (**b**) From the information given, estimate the diameter in Å of a single gold atom. (**c**) What assumptions did you make in arriving at your answer for part (**b**)?

2.101 The diameter of a rubidium atom is 495 pm We will consider two different ways of placing the atoms on a surface. In arrangement A, all the atoms are lined up with one another to form a square grid. Arrangement B is called a *close-packed* arrangement because the atoms sit in the "depressions" formed by the previous row of atoms:



A B

(**a**) Using arrangement A, how many Rb atoms could be placed on a square surface that is 1.0 cm on a side? (**b**) How many Rb atoms could be placed on a square surface that is 1.0 cm on a side, using arrangement B? (**c**) By what factor has the number of atoms on the surface increased in going to arrangement B from arrangement A? If extended to three dimensions, which arrangement would lead to a greater density for Rb metal?

2.102 (**a**) Assuming the dimensions of the nucleus and atom shown in Figure 2.9, what fraction of the *volume* of the atom is taken up by the nucleus? (**b**) Using the mass of the proton from Table 2.1 and assuming its diameter is 1.0×10^{-15} m, calculate the density of a proton in g/cm^3.

2.103 Identify the element represented by each of the following symbols and give the number of protons and neutrons in each: (**a**) $^{11}_{5}X$ (**b**) $^{75}_{33}X$ (**c**) $^{86}_{36}X$ (**d**) $^{67}_{30}X$.

2.104 The nucleus of ^{6}Li is a powerful absorber of neutrons. It exists in the naturally occurring metal to the extent of 7.5%. In the era of nuclear deterrence, large quantities of lithium were processed to remove ^{6}Li for use in hydrogen bomb production. The lithium metal remaining after removal of ^{6}Li was sold on the market. (**a**) What are the compositions of the nuclei of ^{6}Li and ^{7}Li? (**b**) The atomic masses of ^{6}Li and ^{7}Li are 6.015122 and 7.016004 u, respectively. A sample of lithium depleted in the lighter isotope was found on analysis to contain 1.442% ^{6}Li. What is the average atomic weight of this sample of the metal?

2.105 The element argon has three naturally occurring isotopes, with 18, 20, and 22 neutrons in the nucleus, respectively. (**a**) Write the full chemical symbols for these three isotopes. (**b**) Describe the similarities and differences between the three kinds of atoms of argon.

2.106 The element chromium (Cr) consists of four naturally occurring isotopes with atomic masses 49.9460, 51.9405, 52.9407, and 53.9389 u. The relative abundances of these four isotopes are 4.3, 83.8, 9.5, and 2.4%, respectively. From these data, calculate the atomic weight of chromium.

2.107 Copper (Cu) consists of two naturally occurring isotopes with masses of 62.9296 and 64.9278 u. (**a**) How many protons and neutrons are in the nucleus of each isotope? Write the complete atomic symbol for each, showing the atomic number and mass number. (**b**) The average atomic mass of Cu is 63.55 u. Calculate the abundance of each isotope.

2.108 Using a suitable reference such as the *CRC Handbook of Chemistry and Physics* or http://www.webelements.com, look up the following information for nickel: (**a**) the number of known isotopes, (**b**) the atomic masses (in u), (**c**) the natural abundances of the five most abundant isotopes.

2.109 There are two different isotopes of bromine atoms. Under normal conditions, elemental bromine consists of Br_2 molecules, and the mass of a Br_2 molecule is the sum of the masses of the two atoms in the molecule. The mass spectrum of Br_2 consists of three peaks:

Mass (u)	Relative Size
157.836	0.2569
159.834	0.4999
161.832	0.2431

(**a**) What is the origin of each peak (of what isotopes does each consist)? (**b**) What is the mass of each isotope? (**c**) Determine the average molecular mass of a Br_2 molecule. (**d**) Determine the average atomic mass of a bromine atom. (**e**) Calculate the abundances of the two isotopes.

2.110 It is common in mass spectrometry to assume that the mass of a cation is the same as that of its parent atom. (**a**) Using data in Table 2.1, determine the number of significant figures that must be reported before the difference in masses of ^{1}H and $^{1}H^{+}$ is significant. (**b**) What percentage of the mass of an ^{1}H atom does the electron represent?

2.111 From the following list of elements—Mg, Li, Tl, Pb, Se, Cl, Xe, Si, C—pick the one that best fits each description. Use each element only once: (**a**) an alkali metal, (**b**) an alkaline earth metal, (**c**) a noble gas, (**d**) a halogen, (**e**) a metalloid in Group 14, (**f**) a nonmetal listed in Group 14, (**g**) a metal that forms a 3+ ion, (**h**) a nonmetal that forms a 2− ion, (**i**) an element that is used as radiation shielding.

2.112 The first atoms of seaborgium (Sg) were identified in 1974. The longest-lived isotope of Sg has a mass number of 266. (**a**) How many protons, electrons, and neutrons are in an ^{266}Sg atom? (**b**) Atoms of Sg are very unstable, and it is therefore difficult to study this element's properties. Based on the position of Sg in the periodic table, what element should it most closely resemble in its chemical properties?

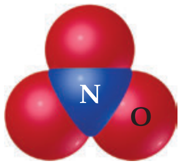
2.113 The explosion of an atomic bomb releases many radioactive isotopes, including strontium-90. Considering the location of strontium in the periodic table, suggest a reason for the fact that this isotope is particularly dangerous for human health.

2.114 A U.S. 1-cent coin (a penny) has a diameter of 19 mm and a thickness of 1.5 mm. Assume the coin is made of pure copper, whose density and approximate market price are 8.9 g/cm^3 and \$2.40 per pound, respectively. Calculate the value of the copper in the coin, assuming its thickness is uniform.

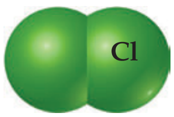
2.115 The U.S. Mint produces a dollar coin called the American Silver Eagle that is made of nearly pure silver. This coin has a diameter of 41 mm and a thickness of 2.5 mm. The density and approximate market price of silver are 10.5 g/cm^3 and

\$0.51 per gram, respectively. Calculate the value of the silver in the coin, assuming its thickness is uniform.

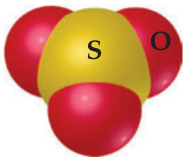
2.116 From the molecular structures shown here, identify the one that corresponds to each of the following species: (**a**) chlorine gas; (**b**) propane; (**c**) nitrate ion; (**d**) sulfur trioxide; (**e**) methyl chloride, CH_3Cl.



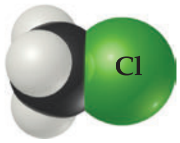
(i)



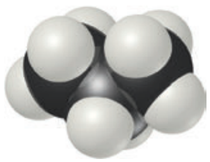
(ii)



(iii)



(iv)

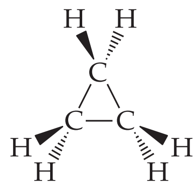

(v)

2.117 Name each of the following chlorides. Assuming that the compounds are ionic, what charge is associated with the metallic element in each case? (**a**) AgCl, (**b**) $TiCl_4$, (**c**) $IrCl_3$, (**d**) LiCl.

2.118 Fill in the blanks in the following table:

Cation	Anion	Formula	Name
			Sodium carbonate
Ni^{2+}	CH_3COO^-		
		$Cu(ClO_4)_2$	
Ca^{2+}	F^-		
		$NaMnO_4$	
			Zinc sulfide
		Mg_3N_2	

2.119 Cyclopropane is an interesting hydrocarbon. Instead of having three carbons in a row, the three carbons form a ring, as shown in this perspective drawing (see Figure 2.16 for a prior example of this kind of drawing):



Cyclopropane was at one time used as an anesthetic, but its use was discontinued, in part because it is highly flammable. (**a**) What is the empirical formula of cyclopropane? How does it differ from that of propane? (**b**) The three carbon atoms are necessarily in a plane. What do the different wedges mean? (**c**) What change would you make to the structure shown to illustrate chlorocyclopropane? Are there isomers of chlorocyclopropane?

2.120 Elements in the same group of the periodic table often form oxyanions with the same general formula. The anions are also named in a similar fashion. Based on these observations, suggest a chemical formula or name, as appropriate, for each of the following ions: (**a**) BrO_4^-, (**b**) SeO_3^{2-}, (**c**) arsenate ion, (**d**) hydrogen tellurate ion.

2.121 Carbonic acid occurs in carbonated beverages. When allowed to react with lithium hydroxide, it produces lithium carbonate. Lithium carbonate is used to treat depression and bipolar disorder. Write chemical formulas for carbonic acid, lithium hydroxide, and lithium carbonate.

2.122 Give the chemical names of each of the following familiar compounds: (**a**) NaCl (table salt), (**b**) $NaHCO_3$ (baking soda), (**c**) NaOCl (in many bleaches), (**d**) NaOH (caustic soda), (**e**) $(NH_4)_2CO_3$ (smelling salts), (**f**) $CaSO_4$ (plaster of Paris).

2.123 Many familiar substances have common, unsystematic names. For each of the following, give the correct systematic name: (**a**) saltpeter, KNO_3; (**b**) soda ash, Na_2CO_3; (**c**) lime, CaO; (**d**) muriatic acid, HCl; (**e**) Epsom salts, $MgSO_4$; (**f**) milk of magnesia, $Mg(OH)_2$.

2.124 Because many ions and compounds have very similar names, there is great potential for confusing them. Write the correct chemical formulas to distinguish between (**a**) sodium carbonate and sodium bicarbonate, (**b**) potassium peroxide and potassium oxide, (**c**) calcium sulfide and calcium sulfate, (**d**) manganese (II) oxide and manganese (III) oxide, (**e**) hydride ion and hydroxide ion, (**f**) magnesium nitride and magnesium nitrite, (**g**) silver nitrate and silver nitrite, (**h**) cuprous oxide and cupric oxide.

2.125 In what part of the atom does the strong nuclear force operate?