2 Atoms, Molecules, and Ions

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A worker in Thailand piles up salt crystals.

Where does one start in learning chemistry? Clearly we must consider some essential vocabulary and something about the origins of the science before we can proceed very far. Thus, while Chapter 1 provided background on the fundamental ideas and procedures of science in general, Chapter 2 covers the specific chemical background necessary for understanding the material in the next few chapters. The coverage of these topics is necessarily brief at this point. We will develop these ideas more fully as it becomes appropriate to do so. A major goal of this chapter is to present the systems for naming chemical compounds to provide you with the vocabulary necessary to understand this book and to pursue your laboratory studies.

Because chemistry is concerned first and foremost with chemical changes, we will proceed as quickly as possible to a study of chemical reactions (Chapters 3 and 4). However, before we can discuss reactions, we must consider some fundamental ideas about atoms and how they combine.

2.1 The Early History of Chemistry

Chemistry has been important since ancient times. The processing of natural ores to produce metals for ornaments and weapons and the use of embalming fluids are just two applications of chemical phenomena that were utilized prior to 1000 B.C.

The Greeks were the first to try to explain why chemical changes occur. By about 400 B.C. they had proposed that all matter was composed of four fundamental substances: fire, earth, water, and air. The Greeks also considered the question of whether matter is continuous, and thus infinitely divisible into smaller pieces, or composed of small, indivisible particles. Supporters of the latter position were Demokritos* of Abdera (*c.* 460–*c.* 370 B.C.) and Leucippos, who used the term *atomos* (which later became *atoms*) to describe these ultimate particles. However, because the Greeks had no experiments to test their ideas, no definitive conclusion could be reached about the divisibility of matter.

The next 2000 years of chemical history were dominated by a pseudoscience called *alchemy.* Some alchemists were mystics and fakes who were obsessed with the idea of turning cheap metals into gold. However, many alchemists were serious scientists, and this period saw important advances: The alchemists discovered several elements and learned to prepare the mineral acids.

The foundations of modern chemistry were laid in the sixteenth century with the development of systematic metallurgy (extraction of metals from ores) by a German, Georg Bauer (1494–1555), and the medicinal application of minerals by a Swiss alchemist/physician known as Paracelsus (full name: Philippus Theophrastus Bombastus von Hohenheim [1493–1541]).

The first "chemist" to perform truly quantitative experiments was Robert Boyle (1627–1691), who carefully measured the relationship between the pressure and volume of air. When Boyle published his book *The Skeptical Chymist* in 1661, the quantitative sciences of physics and chemistry were born. In addition to his results on the quantitative behavior of gases, Boyle's other major contribution to chemistry consisted of his ideas about the chemical elements. Boyle held no preconceived notion about the number of elements. In his view, a substance was an element unless it could be broken down into two or more simpler substances. As Boyle's experimental definition of an element became generally accepted, the list of known elements began to grow, and the Greek system of

*Democritus is an alternate spelling.

CHEMICAL IMPACT

There's Gold in Them There Plants!

Gold has always held a strong allure. For example, the alchemists were obsessed with finding a way to transform cheap metals into gold. Also, when gold was discovered in California in 1849, a frantic rush occurred to that area and other areas in the west. Although gold is still valuable, most of the high-grade gold ores have been exhausted. This leaves the low-grade ores—ores with low concentrations of gold that are expensive to process relative to the amount of gold finally obtained.

Now two scientists have come across a novel way to concentrate the gold from low-grade ores. Christopher Anderson and Robert Brooks of Massey University in Palmerston North, New Zealand, have found plants that accumulate gold atoms as they grow in soil containing gold ore [*Nature* 395 (1998): 553]. The plants brassica (of the mustard family) and chicory seem especially effective as botanical "gold miners." When these plants are dried and burned (after having grown in gold-rich soil), the resulting ash contains approximately 150 ppm (parts per million) of gold. (1 ppm gold represents 1 g of gold in 10^6 g of sample.)

The New Zealand scientists were able to double the amount of gold taken from the soil by the plants by treating the soil with ammonium thiocyanate (NH_4SCN). The thiocyanate, which reacts with the gold, making it more available to the plants, subsequently breaks down in the soil and therefore poses no environmental hazard.

Thus plants seem to hold great promise as gold miners. They are efficient and reliable and will never go on strike.

This plant from the mustard family is a newly discovered source of gold.

four elements finally died. Although Boyle was an excellent scientist, he was not always right. For example, he clung to the alchemists' views that metals were not true elements and that a way would eventually be found to change one metal into another.

The phenomenon of combustion evoked intense interest in the seventeenth and eighteenth centuries. The German chemist Georg Stahl (1660–1734) suggested that a substance he called "phlogiston" flowed out of the burning material. Stahl postulated that a substance burning in a closed container eventually stopped burning because the air in the container became saturated with phlogiston. Oxygen gas, discovered by Joseph Priestley (1733–1804),* an English clergyman and scientist (Fig. 2.1), was found to support vigorous combustion and was thus supposed to be low in phlogiston. In fact, oxygen was originally called "dephlogisticated air."

FIGURE 2.1
The Priestley Medal is the highest honor given by the American Chemical Society. It is named for Joseph Priestley, who was born in England on March 13, 1733. He performed many important scientific experiments, among them the discovery that a gas later identified as carbon dioxide could be dissolved in water to produce *seltzer*. Also, as a result of meeting Benjamin Franklin in London in 1766, Priestley became interested in electricity and was the first to observe that graphite was an electrical conductor. However, his greatest discovery occurred in 1774 when he isolated oxygen by heating mercuric oxide.

Because of his nonconformist political views, Priestley was forced to leave England. He died in the United States in 1804.

*Oxygen gas was actually first observed by the Swedish chemist Karl W. Scheele (1742–1786), but because his results were published after Priestley's, the latter is commonly credited with the discovery of oxygen.

2.2 Fundamental Chemical Laws

By the late eighteenth century, combustion had been studied extensively; the gases carbon dioxide, nitrogen, hydrogen, and oxygen had been discovered; and the list of elements continued to grow. However, it was Antoine Lavoisier (1743–1794), a French chemist (Fig. 2.2), who finally explained the true nature of combustion, thus clearing the way for the tremendous progress that was made near the end of the eighteenth century. Lavoisier, like Boyle, regarded measurement as the essential operation of chemistry. His experiments, in which he carefully weighed the reactants and products of various reactions, suggested that *mass is neither created nor destroyed.* Lavoisier's verification of this **law of conservation of mass** was the basis for the developments in chemistry in the nineteenth century.

Oxygen is from the French *oxygène,* meaning "generator of acid," because it was initially considered to be an integral part of all acids.

Lavoisier's quantitative experiments showed that combustion involved oxygen (which Lavoisier named), not phlogiston. He also discovered that life was supported by a process that also involved oxygen and was similar in many ways to combustion. In 1789 Lavoisier published the first modern chemistry textbook, *Elementary Treatise on Chemistry,* in which he presented a unified picture of the chemical knowledge assembled up to that time. Unfortunately, in the same year the text was published, the French Revolution broke out. Lavoisier, who had been associated with collecting taxes for the government, was executed on the guillotine as an enemy of the people in 1794.

After 1800, chemistry was dominated by scientists who, following Lavoisier's lead, performed careful weighing experiments to study the course of chemical reactions and to determine the composition of various chemical compounds. One of these chemists, a Frenchman, Joseph Proust (1754–1826), showed that *a given compound always contains exactly the same proportion of elements by mass.* For example, Proust found that the substance copper carbonate is always 5.3 parts copper to 4 parts oxygen to 1 part carbon (by mass). The principle of the constant composition of compounds, originally called "Proust's law," is now known as the **law of definite proportion.**

Proust's discovery stimulated John Dalton (1766–1844), an English schoolteacher (Fig. 2.3), to think about atoms as the particles that might compose elements. Dalton reasoned that if elements were composed of tiny individual particles, a given compound should always contain the same combination of these atoms. This concept explained why the same relative masses of elements were always found in a given compound.

But Dalton discovered another principle that convinced him even more of the existence of atoms. He noted, for example, that carbon and oxygen form two different compounds that contain different relative amounts of carbon and oxygen, as shown by the following data:

	Mass of Oxygen That Combines with 1 g of Carbon
Compound I	1.33 g
Compound II	2.66 g

Dalton noted that compound II contains twice as much oxygen per gram of carbon as compound I, a fact that could easily be explained in terms of atoms. Compound I might be CO, and compound II might be CO_2.* This principle, which was found to apply to compounds of other elements as well, became known as the **law of multiple proportions:** *When two elements form a series of compounds, the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers.*

To make sure the significance of this observation is clear, in Sample Exercise 2.1 we will consider data for a series of compounds consisting of nitrogen and oxygen.

Sample Exercise 2.1 Illustrating the Law of Multiple Proportions

The following data were collected for several compounds of nitrogen and oxygen:

	Mass of Nitrogen That Combines with 1 g of Oxygen
Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

Show how these data illustrate the law of multiple proportions.

Solution

For the law of multiple proportions to hold, the ratios of the masses of nitrogen combining with 1 gram of oxygen in each pair of compounds should be small whole numbers. We therefore compute the ratios as follows:

$$\frac{A}{B} = \frac{1.750}{0.875} = \frac{2}{1}$$

$$\frac{B}{C} = \frac{0.875}{0.4375} = \frac{2}{1}$$

$$\frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

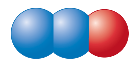
These results support the law of multiple proportions.

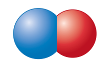
See Exercises 2.27 and 2.28.

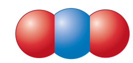
*Subscripts are used to show the numbers of atoms present. The number 1 is understood (not written). The symbols for the elements and the writing of chemical formulas will be illustrated further in Sections 2.6 and 2.7.

The significance of the data in Sample Exercise 2.1 is that compound A contains twice as much nitrogen (N) per gram of oxygen (O) as does compound B and that compound B contains twice as much nitrogen per gram of oxygen as does compound C.

These data can be explained readily if the substances are composed of molecules made up of nitrogen atoms and oxygen atoms. For example, one set of possibilities for compounds A, B, and C is

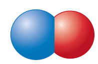
A:  $\frac{N}{O} = \frac{2}{1}$

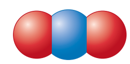
B:  $\frac{N}{O} = \frac{1}{1}$

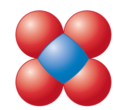
C:  $\frac{N}{O} = \frac{1}{2}$

Now we can see that compound A contains two atoms of N for every atom of O, whereas compound B contains one atom of N per atom of O. That is, compound A contains twice as much nitrogen per given amount of oxygen as does compound B. Similarly, since compound B contains one N per O and compound C contains one N per *two* O's, the nitrogen content of compound C per given amount of oxygen is half that of compound B.

Another set of compounds that fits the data in Sample Exercise 2.1 is

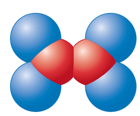
A:  $\frac{N}{O} = \frac{1}{1}$

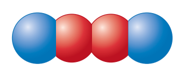
B:  $\frac{N}{O} = \frac{1}{2}$

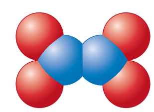
C:  $\frac{N}{O} = \frac{1}{4}$

Verify for yourself that these compounds satisfy the requirements.

Still another set that works is

A:  $\frac{N}{O} = \frac{4}{2}$

B:  $\frac{N}{O} = \frac{2}{2}$

C:  $\frac{N}{O} = \frac{2}{4}$

See if you can come up with still another set of compounds that satisfies the data in Sample Exercise 2.1. How many more possibilities are there?

In fact, an infinite number of other possibilities exists. Dalton could not deduce absolute formulas from the available data on relative masses. However, the data on the composition of compounds in terms of the relative masses of the elements supported his hypothesis that each element consisted of a certain type of atom and that compounds were formed from specific combinations of atoms.

2.3 Dalton's Atomic Theory

In 1808 Dalton published *A New System of Chemical Philosophy,* in which he presented his theory of atoms:

These statements are a modern paraphrase of Dalton's ideas.

1. Each element is made up of tiny particles called atoms.
2. The atoms of a given element are identical; the atoms of different elements are different in some fundamental way or ways.
3. Chemical compounds are formed when atoms of different elements combine with each other. A given compound always has the same relative numbers and types of atoms.
4. Chemical reactions involve reorganization of the atoms—changes in the way they are bound together. The atoms themselves are not changed in a chemical reaction.

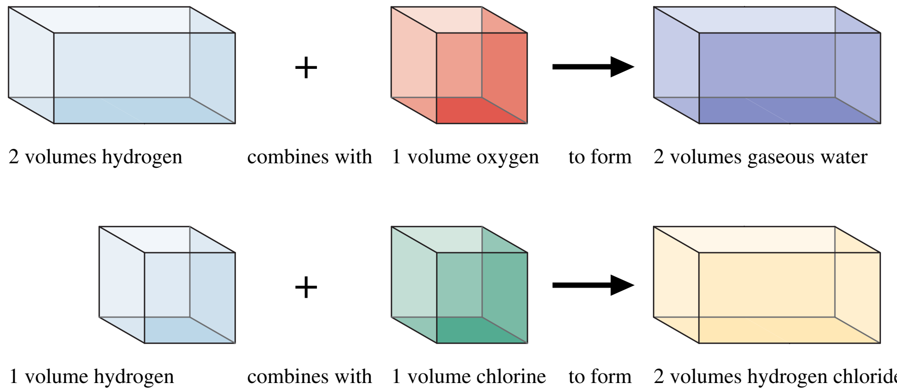


FIGURE 2.4
A representation of some of Gay-Lussac's experimental results on combining gas volumes.

Joseph Louis Gay-Lussac, a French physicist and chemist, was remarkably versatile. Although he is now known primarily for his studies on the combining of volumes of gases, Gay-Lussac was instrumental in the studies of many of the other properties of gases. Some of Gay-Lussac's motivation to learn about gases arose from his passion for ballooning. In fact, he made ascents to heights of over 4 miles to collect air samples, setting altitude records that stood for about 50 years. Gay-Lussac also was the codiscoverer of boron and the developer of a process for manufacturing sulfuric acid. As chief assayer of the French mint, Gay-Lussac developed many techniques for chemical analysis and invented many types of glassware now used routinely in labs. Gay-Lussac spent his last 20 years as a lawmaker in the French government.

It is instructive to consider Dalton's reasoning on the relative masses of the atoms of the various elements. In Dalton's time water was known to be composed of the elements hydrogen and oxygen, with 8 grams of oxygen present for every 1 gram of hydrogen. If the formula for water were OH, an oxygen atom would have to have 8 times the mass of a hydrogen atom. However, if the formula for water were H_2O (two atoms of hydrogen for every oxygen atom), this would mean that each atom of oxygen is 16 times as massive as *each* atom of hydrogen (since the ratio of the mass of one oxygen to that of *two* hydrogens is 8 to 1). Because the formula for water was not then known, Dalton could not specify the relative masses of oxygen and hydrogen unambiguously. To solve the problem, Dalton made a fundamental assumption: He decided that nature would be as simple as possible. This assumption led him to conclude that the formula for water should be OH. He thus assigned hydrogen a mass of 1 and oxygen a mass of 8.

Using similar reasoning for other compounds, Dalton prepared the first table of **atomic masses** (sometimes called **atomic weights** by chemists, since mass is often determined by comparison to a standard mass—a process called *weighing*). Many of the masses were later proved to be wrong because of Dalton's incorrect assumptions about the formulas of certain compounds, but the construction of a table of masses was an important step forward.

Although not recognized as such for many years, the keys to determining absolute formulas for compounds were provided in the experimental work of the French chemist Joseph Gay-Lussac (1778–1850) and by the hypothesis of an Italian chemist named Amadeo Avogadro (1776–1856). In 1809 Gay-Lussac performed experiments in which he measured (under the same conditions of temperature and pressure) the volumes of gases that reacted with each other. For example, Gay-Lussac found that 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of gaseous water and that 1 volume of hydrogen reacts with 1 volume of chlorine to form 2 volumes of hydrogen chloride. These results are represented schematically in Fig. 2.4.

In 1811 Avogadro interpreted these results by proposing that *at the same temperature and pressure, equal volumes of different gases contain the same number of particles.* This assumption (called **Avogadro's hypothesis**) makes sense if the distances between the particles in a gas are very great compared with the sizes of the particles. Under these conditions, the volume of a gas is determined by the number of molecules present, not by the size of the individual particles.

If Avogadro's hypothesis is correct, Gay-Lussac's result,

$$\text{2 volumes of hydrogen react with 1 volume of oxygen} \longrightarrow \text{2 volumes of water vapor}$$

can be expressed as follows:

$$\text{2 molecules* of hydrogen react with 1 molecule of oxygen} \longrightarrow \text{2 molecules of water}$$

*A *molecule* is a collection of atoms (see Section 2.6).

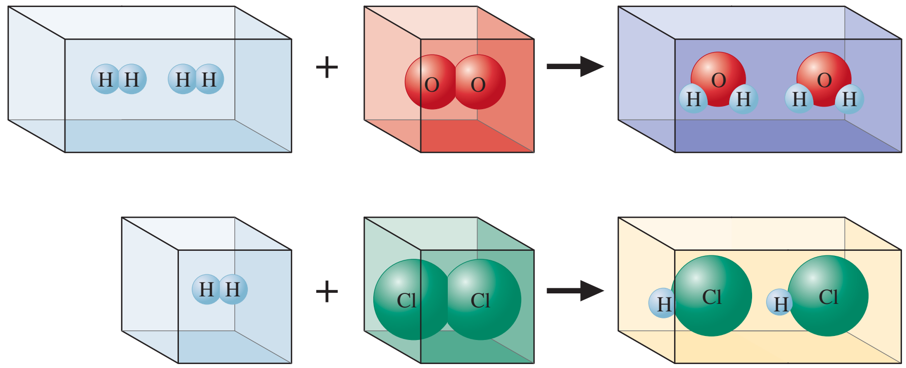


FIGURE 2.5
A representation of combining gases at the molecular level. The spheres represent atoms in the molecules.

These observations can best be explained by assuming that gaseous hydrogen, oxygen, and chlorine are all composed of diatomic (two-atom) molecules: H_2, O_2, and Cl_2, respectively. Gay-Lussac's results can then be represented as shown in Fig. 2.5. (Note that this reasoning suggests that the formula for water is H_2O, not OH as Dalton believed.)

The Italian chemist Stanislao Cannizzaro (1826–1910) cleared up the confusion in 1860 by doing a series of molar mass determinations that convinced the scientific community that the correct atomic mass of carbon is 12. For more information, see *From Caveman to Chemist* by Hugh Salzberg (American Chemical Society, 1991), p. 223.

Unfortunately, Avogadro's interpretations were not accepted by most chemists, and a half-century of confusion followed, in which many different assumptions were made about formulas and atomic masses.

During the nineteenth century, painstaking measurements were made of the masses of various elements that combined to form compounds. From these experiments a list of relative atomic masses could be determined. One of the chemists involved in contributing to this list was a Swede named Jöns Jakob Berzelius (1779–1848), who discovered the elements cerium, selenium, silicon, and thorium and developed the modern symbols for the elements used in writing the formulas of compounds.

2.4 Early Experiments to Characterize the Atom

On the basis of the work of Dalton, Gay-Lussac, Avogadro, and others, chemistry was beginning to make sense. The concept of atoms was clearly a good idea. Inevitably, scientists began to wonder about the nature of the atom. What is an atom made of, and how do the atoms of the various elements differ?

The Electron

The first important experiments that led to an understanding of the composition of the atom were done by the English physicist J. J. Thomson (Fig. 2.6), who studied electrical discharges in partially evacuated tubes called **cathode-ray tubes** (Fig. 2.7) during the period from 1898 to 1903. Thomson found that when high voltage was applied to the tube, a "ray" he called a *cathode ray* (because it emanated from the negative electrode, or cathode) was produced. Because this ray was produced at the negative electrode and was repelled by the negative pole of an applied electric field (see Fig. 2.8), Thomson postulated that the ray was a stream of negatively charged particles, now called **electrons.** From experiments in which he measured the deflection of the beam of electrons in a magnetic field, Thomson determined the *charge-to-mass ratio* of an electron:

$$\frac{e}{m} = -1.76 \times 10^8 \text{ C/g}$$

where e represents the charge on the electron in coulombs (C) and m represents the electron mass in grams.

CHEMICAL IMPACT

Berzelius, Selenium, and Silicon

Jöns Jakob Berzelius was probably the best experimental chemist of his generation and, given the crudeness of his laboratory equipment, maybe the best of all time. Unlike Lavoisier, who could afford to buy the best laboratory equipment available, Berzelius worked with minimal equipment in very plain surroundings. One of Berzelius's students described the Swedish chemist's workplace: "The laboratory consisted of two ordinary rooms with the very simplest arrangements; there were neither furnaces nor hoods, neither water system nor gas. Against the walls stood some closets with the chemicals, in the middle the mercury trough and the blast lamp table. Beside this was the sink consisting of a stone water holder with a stopcock and a pot standing under it. [Next door in the kitchen] stood a small heating furnace."

Comparison of Several of Berzelius's Atomic Masses with the Modern Values

	Atomic Mass	
Element	**Berzelius's Value**	**Current Value**
Chlorine	35.41	35.45
Copper	63.00	63.55
Hydrogen	1.00	1.01
Lead	207.12	207.2
Nitrogen	14.05	14.01
Oxygen	16.00	16.00
Potassium	39.19	39.10
Silver	108.12	107.87
Sulfur	32.18	32.07

In these simple facilities Berzelius performed more than 2000 experiments over a 10-year period to determine accurate atomic masses for the 50 elements then known. His success can be seen from the data in the table at left. These remarkably accurate values attest to his experimental skills and patience.

Besides his table of atomic masses, Berzelius made many other major contributions to chemistry. The most important of these was the invention of a simple set of symbols for the elements along with a system for writing the formulas of compounds to replace the awkward symbolic representations of the alchemists. Although some chemists, including Dalton, objected to the new system, it was gradually adopted and forms the basis of the system we use today.

In addition to these accomplishments, Berzelius discovered the elements cerium, thorium, selenium, and silicon. Of these elements, selenium and silicon are particularly important in today's world. Berzelius discovered selenium in 1817 in connection with his studies of sulfuric acid. For years selenium's toxicity has been known, but only recently have we become aware that it may have a positive effect on human

Visualization: Cathode-Ray Tube

One of Thomson's primary goals in his cathode-ray tube experiments was to gain an understanding of the structure of the atom. He reasoned that since electrons could be produced from electrodes made of various types of metals, *all* atoms must contain electrons. Since atoms were known to be electrically neutral, Thomson further assumed that atoms also must contain some positive charge. Thomson postulated that an atom consisted of a

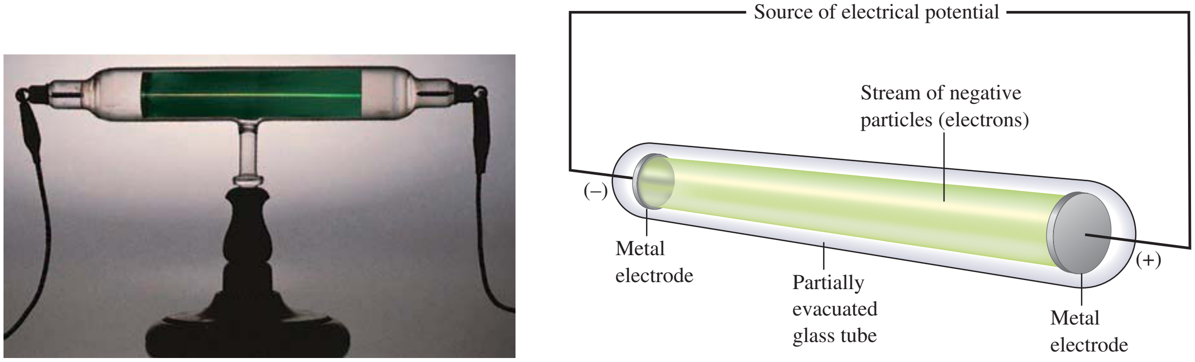


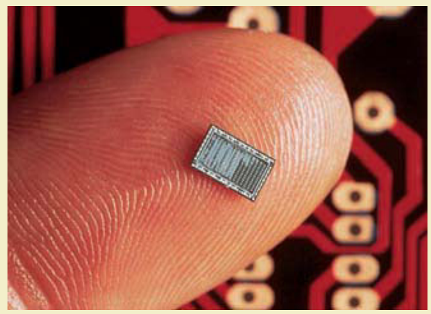
FIGURE 2.7
A cathode-ray tube. The fast-moving electrons excite the gas in the tube, causing a glow between the electrodes. The green color in the photo is due to the response of the screen (coated with zinc sulfide) to the electron beam.

The Alchemists' Symbols for Some Common Elements and Compounds

Substance	Alchemists' Symbol
Silver	☽
Lead	♄
Tin	♃
Platinum	☽☉
Sulfuric acid	
Alcohol	
Sea salt	

health. Studies have shown that trace amounts of selenium in the diet may protect people from heart disease and cancer. One study based on data from 27 countries showed an inverse relationship between the cancer death rate and the selenium content of soil in a particular region (low cancer death rate in areas with high selenium content). Another research paper reported an inverse relationship between the selenium content of the blood and the incidence of breast cancer in women. A study reported in 1998 used the toenail clippings of 33,737 men to show that selenium seems to protect against prostate cancer. Selenium is also found in the heart muscle and may play an important role in proper heart function. Because of these and other studies, selenium's reputation has improved, and many scientists are now studying its function in the human body.

Silicon is the second most abundant element in the earth's crust, exceeded only by oxygen. As we will see in Chapter 10, compounds involving silicon bonded to oxygen make up most of the earth's sand, rock, and soil. Berzelius prepared silicon in its pure form in 1824 by heating silicon tetrafluoride (SiF_4) with potassium metal. Today, silicon forms the basis for the modern microelectronics industry centered near San Francisco in a place that has come to be known as "Silicon Valley." The technology of the silicon chip (see figure) with its printed circuits has transformed computers from room-sized monsters with thousands of unreliable vacuum tubes to desktop and notebook-sized units with trouble-free "solid-state" circuitry.



A silicon chip.

See E. J. Holmyard, *Alchemy* (New York: Penguin Books, 1968).

diffuse cloud of positive charge with the negative electrons embedded randomly in it. This model, shown in Fig. 2.9, is often called the *plum pudding model* because the electrons are like raisins dispersed in a pudding (the positive charge cloud), as in plum pudding, a favorite English dessert.

In 1909 Robert Millikan (1868–1953), working at the University of Chicago, performed very clever experiments involving charged oil drops. These experiments allowed

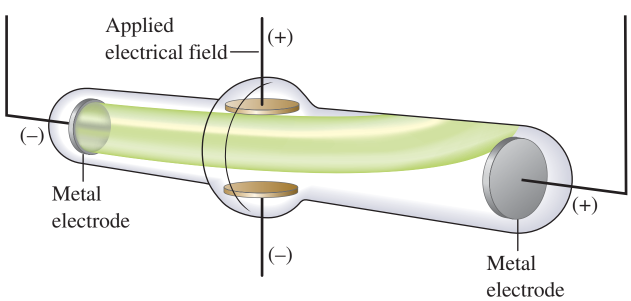


FIGURE 2.8
Deflection of cathode rays by an applied electric field.

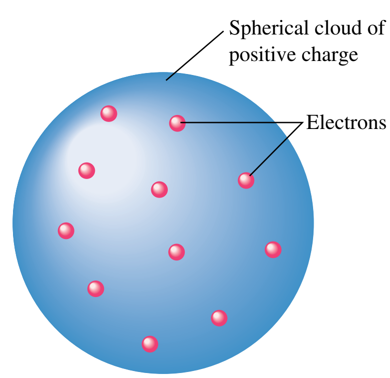


FIGURE 2.9
The plum pudding model of the atom.

Visualization: Millikan's Oil Drop Experiment

A technician using a scanner to monitor the uptake of radioactive iodine in a patient's thyroid.

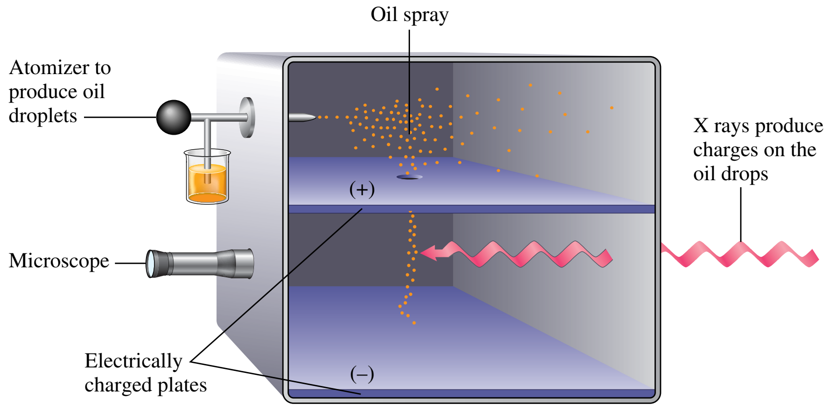


FIGURE 2.10
A schematic representation of the apparatus Millikan used to determine the charge on the electron. The fall of charged oil droplets due to gravity can be halted by adjusting the voltage across the two plates. This voltage and the mass of the oil drop can then be used to calculate the charge on the oil drop. Millikan's experiments showed that the charge on an oil drop is always a whole-number multiple of the electron charge.

him to determine the magnitude of the electron charge (see Fig. 2.10). With this value and the charge-to-mass ratio determined by Thomson, Millikan was able to calculate the mass of the electron as 9.11×10^{-31} kilogram.

Radioactivity

In the late nineteenth century scientists discovered that certain elements produce high-energy radiation. For example, in 1896 the French scientist Henri Becquerel found accidentally that a piece of a mineral containing uranium could produce its image on a photographic plate in the absence of light. He attributed this phenomenon to a spontaneous emission of radiation by the uranium, which he called **radioactivity.** Studies in the early twentieth century demonstrated three types of radioactive emission: gamma (γ) rays, beta (β) particles, and alpha (α) particles. A γ ray is high-energy "light"; a β particle is a high-speed electron; and an α particle has a 2+ charge, that is, a charge twice that of the electron and with the opposite sign. The mass of an α particle is 7300 times that of the electron. More modes of radioactivity are now known, and we will discuss them in Chapter 18. Here we will consider only α particles because they were used in some crucial early experiments.

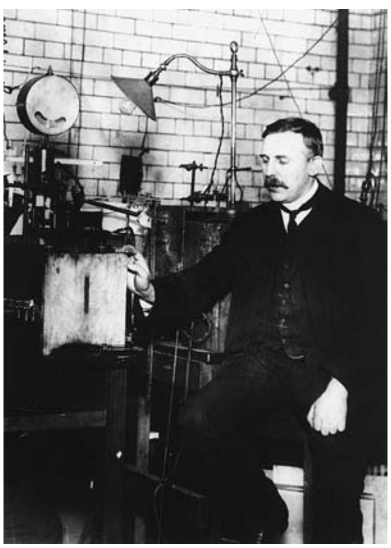


FIGURE 2.11
Ernest Rutherford (1871–1937) was born on a farm in New Zealand. In 1895 he placed second in a scholarship competition to attend Cambridge University but was awarded the scholarship when the winner decided to stay home and get married. As a scientist in England, Rutherford did much of the early work on characterizing radioactivity. He named the α and β particles and the γ ray and coined the term *half-life* to describe an important attribute of radioactive elements. His experiments on the behavior of α particles striking thin metal foils led him to postulate the nuclear atom. He also invented the name *proton* for the nucleus of the hydrogen atom. He received the Nobel Prize in chemistry in 1908.

The Nuclear Atom

In 1911 Ernest Rutherford (Fig. 2.11), who performed many of the pioneering experiments to explore radioactivity, carried out an experiment to test Thomson's plum pudding model. The experiment involved directing α particles at a thin sheet of metal foil, as illustrated in Fig. 2.12. Rutherford reasoned that if Thomson's model were accurate, the massive α particles should crash through the thin foil like cannonballs through gauze, as shown in Fig. 2.13(a). He expected the α particles to travel through the foil with, at the most, very minor deflections in their paths. The results of the experiment were very different from those Rutherford anticipated. Although most of the α particles passed straight through, many of the particles were deflected at large angles, as shown in Fig. 2.13(b), and some were reflected, never hitting the detector. This outcome was a great surprise to Rutherford. (He wrote that this result was comparable with shooting a howitzer at a piece of paper and having the shell reflected back.)

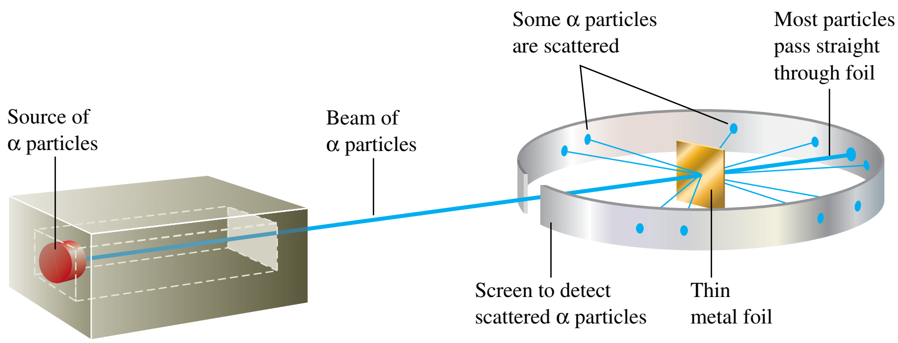


FIGURE 2.12
Rutherford's experiment on α-particle bombardment of metal foil.

Rutherford knew from these results that the plum pudding model for the atom could not be correct. The large deflections of the α particles could be caused only by a center of concentrated positive charge that contains most of the atom's mass, as illustrated in Fig. 2.13(b). Most of the α particles pass directly through the foil because the atom is mostly open space. The deflected α particles are those that had a "close encounter" with the massive positive center of the atom, and the few reflected α particles are those that made a "direct hit" on the much more massive positive center.

In Rutherford's mind these results could be explained only in terms of a **nuclear atom**—an atom with a dense center of positive charge (the **nucleus**) with electrons moving around the nucleus at a distance that is large relative to the nuclear radius.

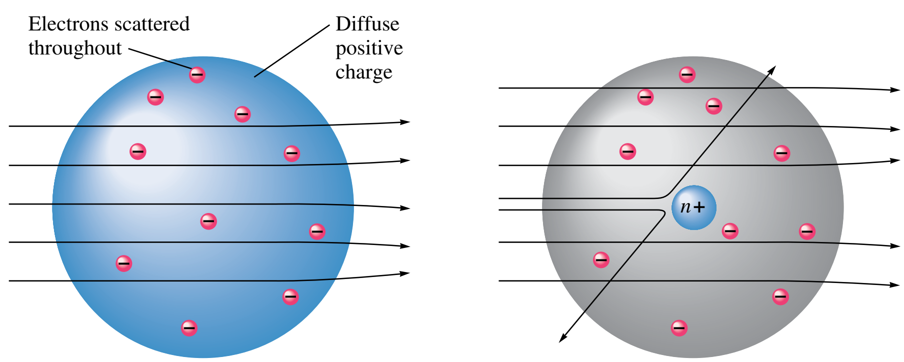
2.5 The Modern View of Atomic Structure: An Introduction

The forces that bind the positively charged protons in the nucleus will be discussed in Chapter 18.

In the years since Thomson and Rutherford, a great deal has been learned about atomic structure. Because much of this material will be covered in detail in later chapters, only an introduction will be given here. The simplest view of the atom is that it consists of a tiny nucleus (with a diameter of about 10^{-13} cm) and electrons that move about the nucleus at an average distance of about 10^{-8} cm from it (see Fig. 2.14).

As we will see later, the chemistry of an atom mainly results from its electrons. For this reason, chemists can be satisfied with a relatively crude nuclear model. The nucleus is assumed to contain **protons,** which have a positive charge equal in magnitude to the electron's negative charge, and **neutrons,** which have virtually the same mass as a proton but no charge. The masses and charges of the electron, proton, and neutron are shown in Table 2.1.

WWW Visualization: Gold Foil Experiment



(a) (b)

FIGURE 2.13
(a) The expected results of the metal foil experiment if Thomson's model were correct. (b) Actual results.

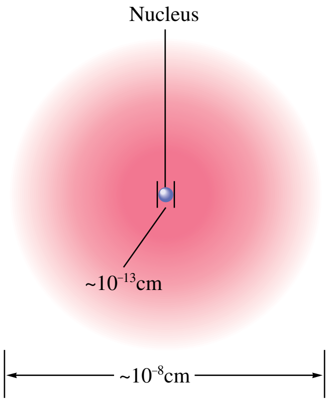


FIGURE 2.14
A nuclear atom viewed in cross section. Note that this drawing is not to scale.

The *chemistry* of an atom arises from its electrons.

If the atomic nucleus were the size of this ball bearing, a typical atom would be the size of this stadium.

TABLE 2.1 The Mass and Charge of the Electron, Proton, and Neutron

Particle	Mass	Charge*
Electron	9.11×10^{-31} kg	1−
Proton	1.67×10^{-27} kg	1+
Neutron	1.67×10^{-27} kg	None

*The magnitude of the charge of the electron and the proton is 1.60×10^{-19} C.

Two striking things about the nucleus are its small size compared with the overall size of the atom and its extremely high density. The tiny nucleus accounts for almost all the atom's mass. Its great density is dramatically demonstrated by the fact that a piece of nuclear material about the size of a pea would have a mass of 250 million tons!

An important question to consider at this point is, *"If all atoms are composed of these same components, why do different atoms have different chemical properties?"* The answer to this question lies in the number and the arrangement of the electrons. The electrons constitute most of the atomic volume and thus are the parts that "intermingle" when atoms combine to form molecules. Therefore, the number of electrons possessed by a given atom greatly affects its ability to interact with other atoms. As a result, the atoms of different elements, which have different numbers of protons and electrons, show different chemical behavior.

A sodium atom has 11 protons in its nucleus. Since atoms have no net charge, the number of electrons must equal the number of protons. Therefore, a sodium atom has 11 electrons moving around its nucleus. It is *always* true that a sodium atom has 11 protons and 11 electrons. However, each sodium atom also has neutrons in its nucleus, and different types of sodium atoms exist that have different numbers of neutrons. For example, consider the sodium atoms represented in Fig. 2.15. These two atoms are **isotopes,** or *atoms with the same number of protons but different numbers of neutrons.* Note that the symbol for one particular type of sodium atom is written

Mass number ↘
$^{23}_{11}Na$ ← Element symbol
Atomic number ↗

where the **atomic number** Z (number of protons) is written as a subscript, and the **mass number** A (the total number of protons and neutrons) is written as a superscript. (The particular atom represented here is called "sodium twenty-three." It has 11 electrons, 11 protons, and 12 neutrons.) Because the chemistry of an atom is due to its electrons, isotopes show almost identical chemical properties. In nature most elements contain mixtures of isotopes.

Mass number ↘
Atomic number → $^{A}_{Z}X$ ← Element symbol

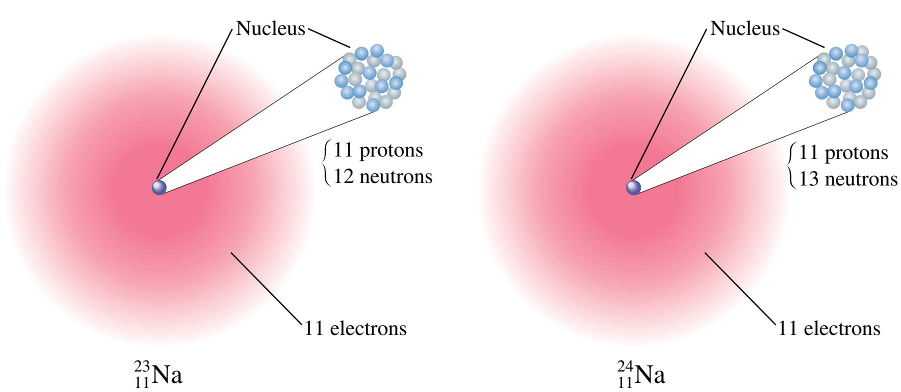


FIGURE 2.15
Two isotopes of sodium. Both have 11 protons and 11 electrons, but they differ in the number of neutrons in their nuclei.

CHEMICAL IMPACT

Reading the History of Bogs

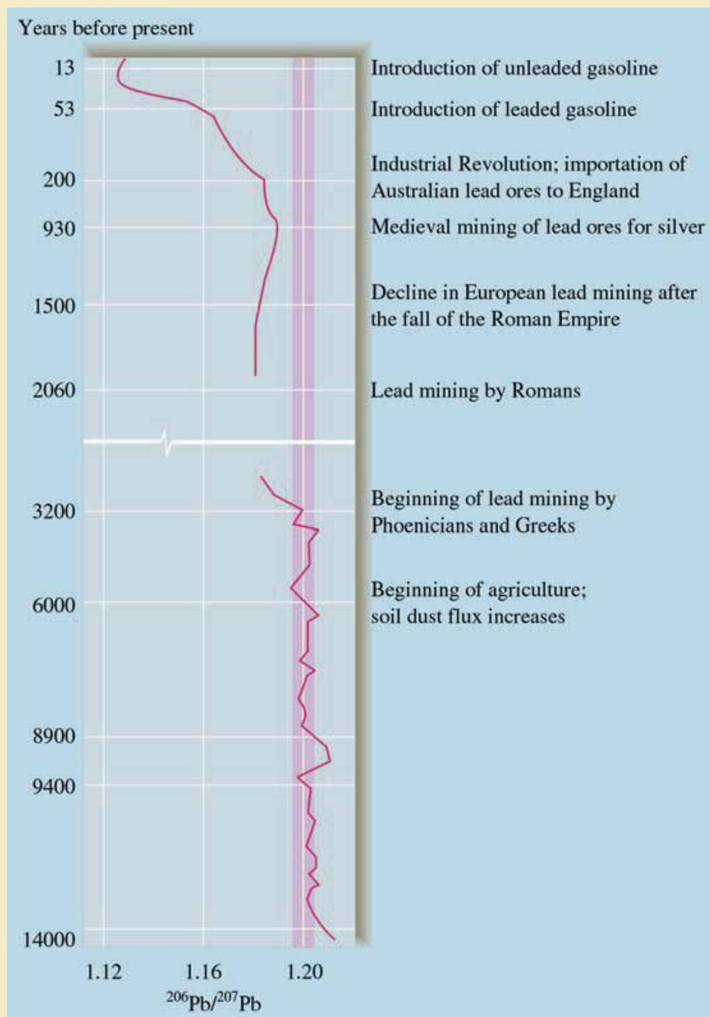
Scientists often "read" the history of the earth and its inhabitants using very different "books" than traditional historians. For example, the disappearance of the dinosaurs 65 million years ago in an "instant" of geological time was a great mystery until unusually high iridium and osmium levels were discovered at a position in the earth's crust corresponding to that time. These high levels of iridium and osmium suggested that an extraterrestrial object had struck the earth 65 million years ago with catastrophic results for the dinosaurs. Since then, the huge buried crater caused by the object has been discovered on the Yucatan Peninsula, and virtually everyone is now convinced that this is the correct explanation for the disappearing dinosaurs.

History is also being "read" by scientists studying ice cores from glaciers in Iceland. Now Swiss scientists have found that ancient peat bogs can furnish a reliable historical record. Geochemist William Shotyk of the University of Bern has found a 15,000-year window on history by analyzing the lead content of core samples from a Swiss mountainside peat bog [*Science* 281 (1998): 1635]. Various parts of the core samples were dated by ^{14}C dating techniques (see Chapter 18, Section 18.4, for more information) and analyzed for their scandium and lead contents. Also, the $^{206}Pb/^{207}Pb$ ratio was measured for each sample. These data are represented in the accompanying figure. Notice that the $^{206}Pb/^{207}Pb$ ratio remains very close to 1.20 (see the red band in the figure) from 14,000 years to 3200 years. The value of 1.20 is the same as the average $^{206}Pb/^{207}Pb$ ratio in the earth's soil.

The core also reveals that the total lead and scandium levels increased simultaneously at the 6000-year mark but that the $^{206}Pb/^{207}Pb$ ratio remained close to 1.20. This coincides with the beginning of agriculture in Europe, which caused more soil dust to enter the atmosphere.

Significantly, about 3000 years ago the $^{206}Pb/^{207}Pb$ ratio decreased markedly. This also corresponds in the core sample to an increase in total lead content out of proportion to the increase in scandium. This indicates the lead no longer resulted from soil dust but from other activities of humans—lead mining had begun. Since the 3000-year mark, the $^{206}Pb/^{207}Pb$ ratio has remained well below 1.20, indicating that human use of lead ores has become the dominant source of airborne lead. This is confirmed by the sharp decline in the ratio beginning 200 years ago that corresponds to the importation into England of Australian lead ores having low $^{206}Pb/^{207}Pb$ ratios.

So far only lead has been used to read the history in the bog. However, Shotyk's group is also measuring the changes in the levels of copper, zinc, cadmium, arsenic, mercury, and antimony. More interesting stories are sure to follow.



Geochemist William Shotyk's analysis of the lead content of ice core samples reveals a 15,000-year history of lead levels. (Note: Dates are based on calibrated radiocarbon dating. Because the core was retrieved in two segments, a break in data occurs between 2060 and 3200 years before present.)

Sample Exercise 2.2 Writing the Symbols for Atoms

Write the symbol for the atom that has an atomic number of 9 and a mass number of 19. How many electrons and how many neutrons does this atom have?

Solution

The atomic number 9 means the atom has 9 protons. This element is called *fluorine,* symbolized by F. The atom is represented as

$$^{19}_{9}F$$

and is called "fluorine nineteen." Since the atom has 9 protons, it also must have 9 electrons to achieve electrical neutrality. The mass number gives the total number of protons and neutrons, which means that this atom has 10 neutrons.

See Exercises 2.43 through 2.46.

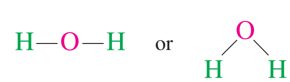
2.6 Molecules and Ions

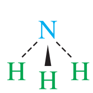
From a chemist's viewpoint, the most interesting characteristic of an atom is its ability to combine with other atoms to form compounds. It was John Dalton who first recognized that chemical compounds are collections of atoms, but he could not determine the structure of atoms or their means for binding to each other. During the twentieth century we learned that atoms have electrons and that these electrons participate in bonding one atom to another. We will discuss bonding thoroughly in Chapters 8 and 9; here we will introduce some simple bonding ideas that will be useful in the next few chapters.

The forces that hold atoms together in compounds are called **chemical bonds.** One way that atoms can form bonds is by *sharing electrons.* These bonds are called **covalent bonds,** and the resulting collection of atoms is called a **molecule.** Molecules can be represented in several different ways. The simplest method is the **chemical formula,** in which the symbols for the elements are used to indicate the types of atoms present and subscripts are used to indicate the relative numbers of atoms. For example, the formula for carbon dioxide is CO_2, meaning that each molecule contains 1 atom of carbon and 2 atoms of oxygen.

Visualization: Covalent Bonding

Examples of molecules that contain covalent bonds are hydrogen (H_2), water (H_2O), oxygen (O_2), ammonia (NH_3), and methane (CH_4). More information about a molecule is given by its **structural formula,** in which the individual bonds are shown (indicated by lines). Structural formulas may or may not indicate the actual shape of the molecule. For example, water might be represented as



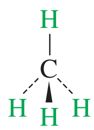


Ammonia

The structure on the right shows the actual shape of the water molecule. Scientists know from experimental evidence that the molecule looks like this. (We will study the shapes of molecules further in Chapter 8.) The structural formula for ammonia is shown in the margin at left.

Note that atoms connected to the central atom by dashed lines are behind the plane of the paper, and atoms connected to the central atom by wedges are in front of the plane of the paper.

In a compound composed of molecules, the individual molecules move around as independent units. For example, a molecule of methane gas can be represented in several ways. The structural formula for methane (CH_4) is shown in Fig. 2.16. The **space-filling**



Methane

FIGURE 2.16
The structural formula for methane.

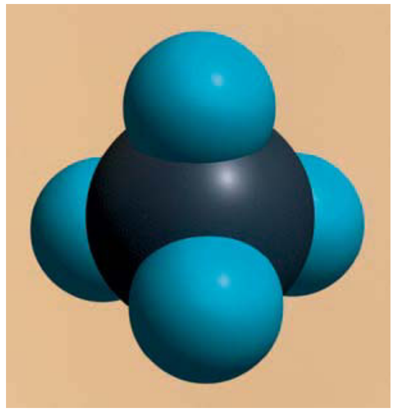


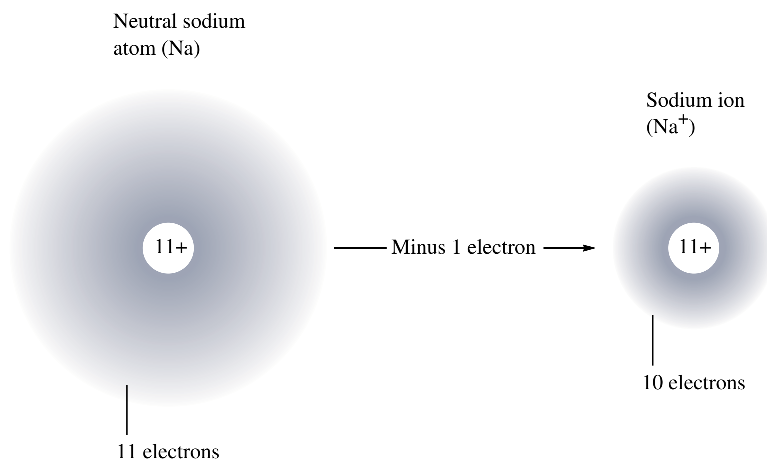
FIGURE 2.17
Space-filling model of methane. This type of model shows both the relative sizes of the atoms in the molecule and their spatial relationships.

FIGURE 2.18
Ball-and-stick model of methane.

model of methane, which shows the relative sizes of the atoms as well as their relative orientation in the molecule, is given in Fig. 2.17. **Ball-and-stick models** are also used to represent molecules. The ball-and-stick structure of methane is shown in Fig. 2.18.

A second type of chemical bond results from attractions among ions. An **ion** is an atom or group of atoms that has a net positive or negative charge. The best-known ionic compound is common table salt, or sodium chloride, which forms when neutral chlorine and sodium react.

To see how the ions are formed, consider what happens when an electron is transferred from a sodium atom to a chlorine atom (the neutrons in the nuclei will be ignored):

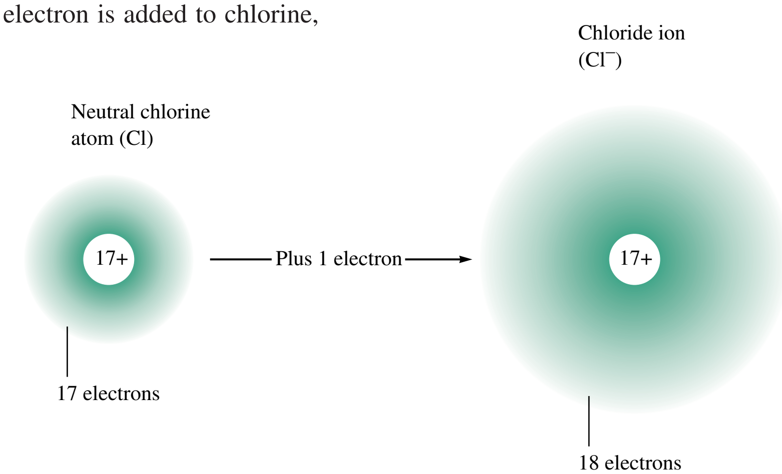


Na^+ is usually called the *sodium ion* rather than the sodium cation. Also Cl^- is called the *chloride ion* rather than the chloride anion. In general, when a specific ion is referred to, the word *ion* rather than cation or anion is used.

With one electron stripped off, the sodium, with its 11 protons and only 10 electrons, now has a net 1+ charge—it has become a *positive ion.* A positive ion is called a **cation.** The sodium ion is written as Na^+, and the process can be represented in shorthand form as

$$Na \longrightarrow Na^+ + e^-$$

If an electron is added to chlorine,



the 18 electrons produce a net 1− charge; the chlorine has become an *ion with a negative charge*—an **anion.** The chloride ion is written as Cl^-, and the process is represented as

$$Cl + e^- \longrightarrow Cl^-$$

Because anions and cations have opposite charges, they attract each other. This *force of attraction between oppositely charged ions* is called **ionic bonding.** As illustrated in Fig. 2.19, sodium metal and chlorine gas (a green gas composed of Cl_2 molecules) react

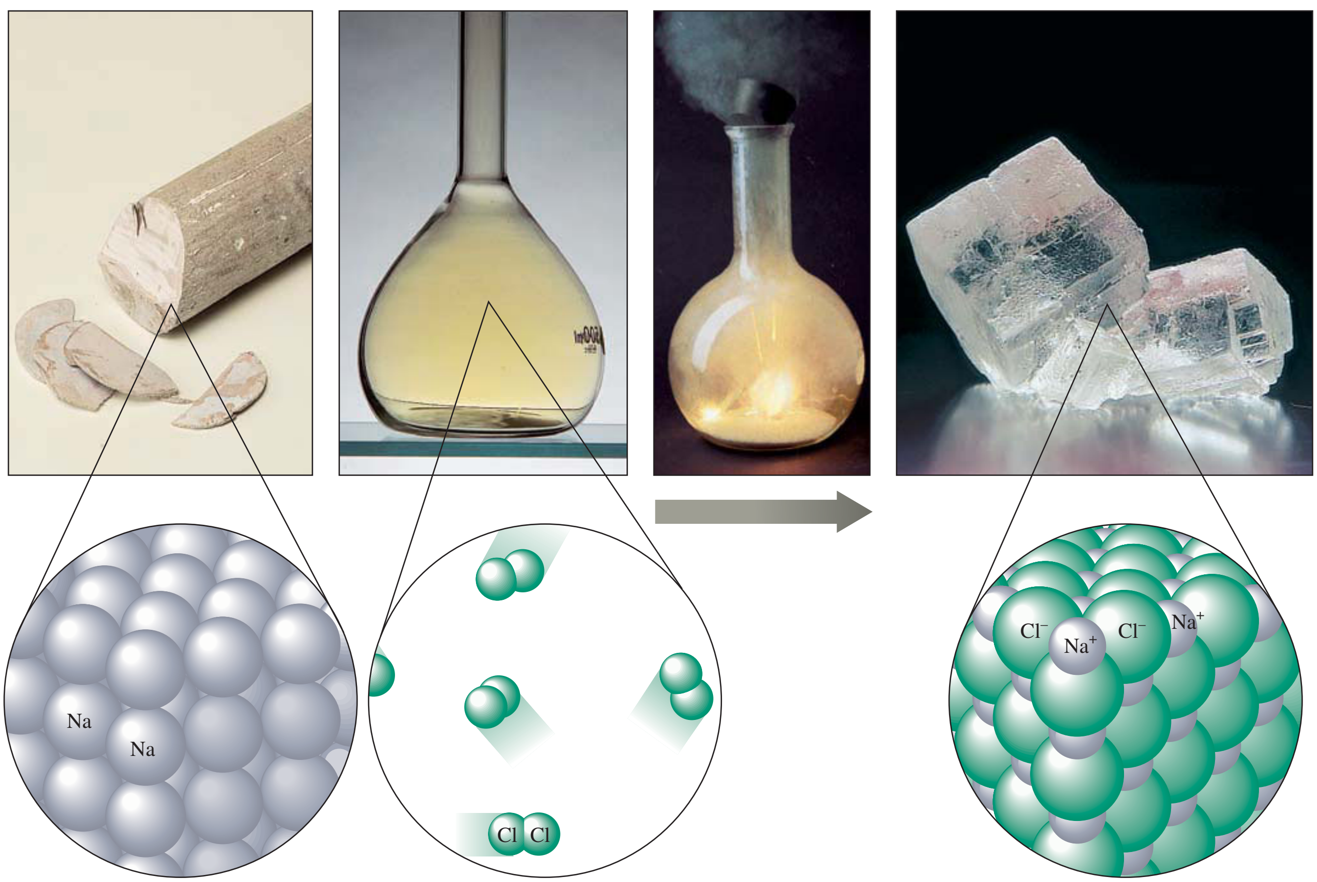


FIGURE 2.19
Sodium metal (which is so soft it can be cut with a knife and which consists of individual sodium atoms) reacts with chlorine gas (which contains Cl_2 molecules) to form solid sodium chloride (which contains Na^+ and Cl^- ions packed together).

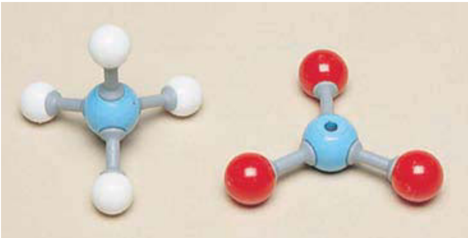


FIGURE 2.20
Ball-and-stick models of the ammonium ion (NH_4^+) and the nitrate ion (NO_3^-).

to form solid sodium chloride, which contains many Na^+ and Cl^- ions packed together and forms the beautiful colorless cubic crystals shown in Fig. 2.19.

A solid consisting of oppositely charged ions is called an **ionic solid,** or a **salt.** Ionic solids can consist of simple ions, as in sodium chloride, or of **polyatomic** (many atom) **ions,** as in ammonium nitrate (NH_4NO_3), which contains ammonium ions (NH_4^+) and nitrate ions (NO_3^-). The ball-and-stick models of these ions are shown in Fig. 2.20.

2.7 An Introduction to the Periodic Table

In a room where chemistry is taught or practiced, a chart called the **periodic table** is almost certain to be found hanging on the wall. This chart shows all the known elements and gives a good deal of information about each. As our study of chemistry progresses, the usefulness of the periodic table will become more obvious. This section will simply introduce it to you.

Visualization: Comparison of a Molecular Compound and an Ionic Compound

A simplified version of the periodic table is shown in Fig. 2.21. The letters in the boxes are the symbols for the elements; these abbreviations are based on the current element names or the original names (see Table 2.2). The number shown above each symbol is the *atomic number* (number of protons) for that element. For example, carbon (C) has atomic number 6, and lead (Pb) has atomic number 82. Most of the elements are **metals.** Metals have characteristic physical properties such as efficient conduction of heat and electricity, malleability (they can be hammered into thin sheets), ductility (they can be pulled into wires), and (often) a lustrous appearance. Chemically, metals tend to *lose* electrons to form positive ions. For example, copper is a typical metal. It is lustrous (although it tarnishes readily); it is an excellent conductor of electricity (it is widely used in electrical wires); and it is readily formed into various shapes, such as pipes for water systems. Copper is also found in many salts, such as the beautiful blue copper sulfate, in which copper is present as Cu^{2+} ions. Copper is a member of the transition metals—the metals shown in the center of the periodic table.

Metals tend to form positive ions; nonmetals tend to form negative ions.

The relatively few **nonmetals** appear in the upper-right corner of the table (to the right of the heavy line in Fig. 2.21), except hydrogen, a nonmetal that resides in the upper-left corner. The nonmetals lack the physical properties that characterize the metals. Chemically, they tend to *gain* electrons in reactions with metals to form negative ions. Nonmetals often bond to each other by forming covalent bonds. For example, chlorine is a typical nonmetal. Under normal conditions it exists as Cl_2 molecules; it reacts with metals to form salts containing Cl^- ions (NaCl, for example); and it forms covalent bonds with nonmetals (for example, hydrogen chloride gas, HCl).

Elements in the same vertical column in the periodic table form a *group* (or *family*) and generally have similar properties.

The periodic table is arranged so that elements in the same vertical columns (called **groups** or **families**) have *similar chemical properties.* For example, all of the **alkali metals,** members of Group 1A—lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr)—are very active elements that readily form ions with a 1+ charge when they react with nonmetals. The members of Group 2A—beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra)—are called the **alkaline earth metals.** They all form ions with a 2+ charge when they react with nonmetals. The **halogens,** the members of Group 7A—fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At)—all form diatomic molecules. Fluorine, chlorine, bromine, and iodine all react with metals to form salts containing ions with a 1− charge (F^-, Cl^-, Br^-, and I^-). The members of Group 8A—helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn)—are known as the **noble gases.** They all exist under normal conditions as monatomic (single-atom) gases and have little chemical reactivity.

Samples of chlorine gas, liquid bromine, and solid iodine.

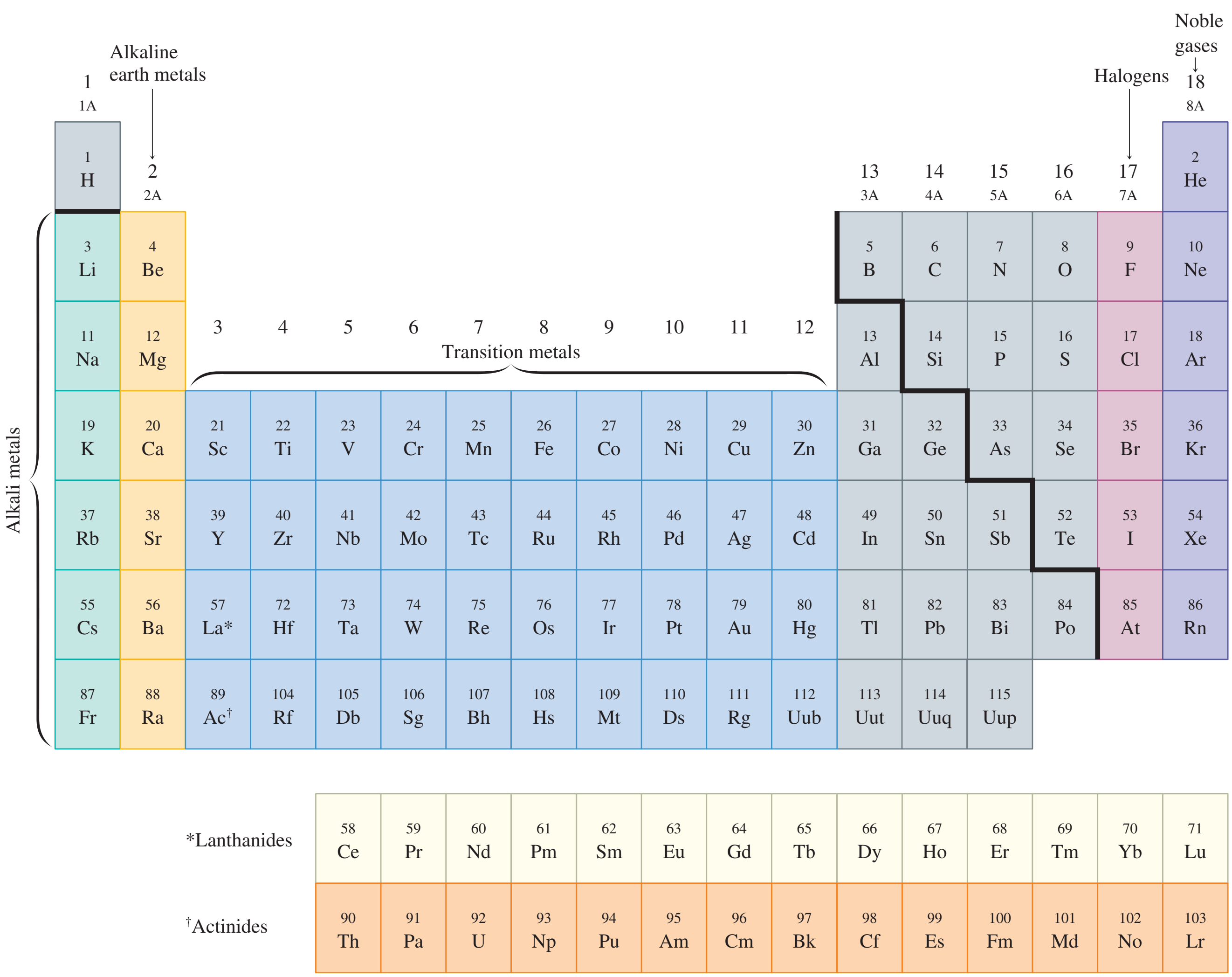


FIGURE 2.21
The periodic table.

TABLE 2.2 The Symbols for the Elements That Are Based on the Original Names

Current Name	Original Name	Symbol
Antimony	Stibium	Sb
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Mercury	Hydrargyrum	Hg
Potassium	Kalium	K
Silver	Argentum	Ag
Sodium	Natrium	Na
Tin	Stannum	Sn
Tungsten	Wolfram	W

CHEMICAL IMPACT

Hassium Fits Right in

Hassium, element 108, does not exist in nature but must be made in a particle accelerator. It was first created in 1984 and can be made by shooting magnesium-26 ($^{26}_{12}Mg$) atoms at curium-248 ($^{248}_{96}Cm$) atoms. The collisions between these atoms produce some hassium-265 ($^{265}_{108}Hs$) atoms. The position of hassium in the periodic table (see Fig. 2.21) in the vertical column containing iron, ruthenium, and osmium suggests that hassium should have chemical properties similar to these metals. However, it is not easy to test this prediction—only a few atoms of hassium can be made at a given time and they last for only about 9 seconds. Imagine having to get your next lab experiment done in 9 seconds!

Amazingly, a team of chemists from the Lawrence Berkeley National Laboratory in California, the Paul Scherrer Institute and the University of Bern in Switzerland, and the Institute of Nuclear Chemistry in Germany have done experiments to characterize the chemical behavior of hassium. For example, they have observed that hassium atoms react with oxygen to form a hassium oxide compound of the type expected from its position on the periodic table. The team has also measured other properties of hassium, including the energy released as it undergoes nuclear decay to another atom.

This work would have surely pleased Dmitri Mendeleev (see Fig. 7.23), who originally developed the periodic table and showed its power to predict chemical properties.

Note from Fig. 2.21 that alternate sets of symbols are used to denote the groups. The symbols 1A through 8A are the traditional designations, whereas the numbers 1 to 18 have been suggested recently. In this text the 1A to 8A designations will be used.

The horizontal rows of elements in the periodic table are called **periods.** Horizontal row 1 is called the *first period* (it contains H and He); row 2 is called the *second period* (elements Li through Ne); and so on.

We will learn much more about the periodic table as we continue with our study of chemistry. Meanwhile, when an element is introduced in this text, you should always note its position on the periodic table.

Another format of the periodic table will be discussed in Section 7.11.

2.8 Naming Simple Compounds

When chemistry was an infant science, there was no system for naming compounds. Names such as sugar of lead, blue vitrol, quicklime, Epsom salts, milk of magnesia, gypsum, and laughing gas were coined by early chemists. Such names are called *common names.* As chemistry grew, it became clear that using common names for compounds would lead to unacceptable chaos. Nearly 5 million chemical compounds are currently known. Memorizing common names for these compounds would be an impossible task.

The solution, of course, is to adopt a *system* for naming compounds in which the name tells something about the composition of the compound. After learning the system, a chemist given a formula should be able to name the compound or, given a name, should be able to construct the compound's formula. In this section we will specify the most important rules for naming compounds other than organic compounds (those based on chains of carbon atoms).

We will begin with the systems for naming inorganic **binary compounds**—compounds composed of two elements—which we classify into various types for easier recognition. We will consider both ionic and covalent compounds.

TABLE 2.3 Common Monatomic Cations and Anions

Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Cesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulfide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminum		
Ag^+	Silver		

Binary Ionic Compounds (Type I)

Binary ionic compounds contain a positive ion (cation) always written first in the formula and a negative ion (anion). In naming these compounds, the following rules apply:

1. The cation is always named first and the anion second.
2. A monatomic (meaning "one-atom") cation takes its name from the name of the element. For example, Na^+ is called sodium in the names of compounds containing this ion.
3. A monatomic anion is named by taking the root of the element name and adding *-ide.* Thus the Cl^- ion is called chloride.

A monatomic cation has the same name as its parent element.

Some common monatomic cations and anions and their names are given in Table 2.3.

The rules for naming binary ionic compounds are illustrated by the following examples:

In formulas of ionic compounds, simple ions are represented by the element symbol: Cl means Cl^-, Na means Na^+, and so on.

Compound	Ions Present	Name
NaCl	Na^+, Cl^-	Sodium chloride
KI	K^+, I^-	Potassium iodide
CaS	Ca^{2+}, S^{2-}	Calcium sulfide
Li_3N	Li^+, N^{3-}	Lithium nitride
CsBr	Cs^+, Br^-	Cesium bromide
MgO	Mg^{2+}, O^{2-}	Magnesium oxide

Sample Exercise 2.3

Naming Type I Binary Compounds

Name each binary compound.

a. CsF **b.** $AlCl_3$ **c.** LiH

Solution

a. CsF is cesium fluoride.
b. $AlCl_3$ is aluminum chloride.
c. LiH is lithium hydride.

Notice that, in each case, the cation is named first, and then the anion is named.

See Exercise 2.55.

Visualization: Formation of Ionic Compounds

Formulas from Names

So far we have started with the chemical formula of a compound and decided on its systematic name. The reverse process is also important. For example, given the name calcium hydroxide, we can write the formula as $Ca(OH)_2$ because we know that calcium forms only Ca^{2+} ions and that, since hydroxide is OH^-, two of these anions will be required to give a neutral compound.

TABLE 2.4 Common Type II Cations

Ion	Systematic Name
Fe^{3+}	Iron(III)
Fe^{2+}	Iron(II)
Cu^{2+}	Copper(II)
Cu^{+}	Copper(I)
Co^{3+}	Cobalt(III)
Co^{2+}	Cobalt(II)
Sn^{4+}	Tin(IV)
Sn^{2+}	Tin(II)
Pb^{4+}	Lead(IV)
Pb^{2+}	Lead(II)
Hg^{2+}	Mercury(II)
Hg_2^{2+}*	Mercury(I)
Ag^{+}	Silver†
Zn^{2+}	Zinc†
Cd^{2+}	Cadmium†

*Note that mercury(I) ions always occur bound together to form Hg_2^{2+} ions.
†Although these are transition metals, they form only one type of ion, and a Roman numeral is not used.

Binary Ionic Compounds (Type II)

In the binary ionic compounds considered earlier (Type I), the metal present forms only a single type of cation. That is, sodium forms only Na^+, calcium forms only Ca^{2+}, and so on. However, as we will see in more detail later in the text, there are many metals that form more than one type of positive ion and thus form more than one type of ionic compound with a given anion. For example, the compound $FeCl_2$ contains Fe^{2+} ions, and the compound $FeCl_3$ contains Fe^{3+} ions. In a case such as this, the *charge on the metal ion must be specified.* The systematic names for these two iron compounds are iron(II) chloride and iron(III) chloride, respectively, where the *Roman numeral indicates the charge of the cation.*

Another system for naming these ionic compounds that is seen in the older literature was used for metals that form only two ions. *The ion with the higher charge has a name ending in* -ic, *and the one with the lower charge has a name ending in* -ous. In this system, for example, Fe^{3+} is called the ferric ion, and Fe^{2+} is called the ferrous ion. The names for $FeCl_3$ and $FeCl_2$ are then ferric chloride and ferrous chloride, respectively. In this text we will use the system that employs Roman numerals. Table 2.4 lists the systematic names for many common type II cations.

Sample Exercise 2.4 Formulas from Names for Type I Binary Compounds

Given the following systematic names, write the formula for each compound:

a. potassium iodide
b. calcium oxide
c. gallium bromide

Solution

Name	Formula	Comments
a. potassium iodide	KI	Contains K^+ and I^-.
b. calcium oxide	CaO	Contains Ca^{2+} and O^{2-}.
c. gallium bromide	$GaBr_3$	Contains Ga^{3+} and Br^-. Must have $3Br^-$ to balance charge of Ga^{3+}.

See Exercise 2.55.

Sample Exercise 2.5 Naming Type II Binary Compounds

1. Give the systematic name for each of the following compounds:

a. CuCl **b.** HgO **c.** Fe_2O_3

2. Given the following systematic names, write the formula for each compound:

a. Manganese(IV) oxide
b. Lead(II) chloride

Type II binary ionic compounds contain a metal that can form more than one type of cation.

A compound must be electrically neutral.

Solution

All of these compounds include a metal that can form more than one type of cation. Thus we must first determine the charge on each cation. This can be done by recognizing that a compound must be electrically neutral; that is, the positive and negative charges must exactly balance.

1.

Formula	Name	Comments
a. CuCl	Copper(I) chloride	Because the anion is Cl^-, the cation must be Cu^+ (for charge balance), which requires a Roman numeral I.
b. HgO	Mercury(II) oxide	Because the anion is O^{2-}, the cation must be Hg^{2+} [mercury(II)].
c. Fe_2O_3	Iron(III) oxide	The three O^{2-} ions carry a total charge of 6−, so two Fe^{3+} ions [iron(III)] are needed to give a 6+ charge.

2.

Name	Formula	Comments
a. Manganese(IV) oxide	MnO_2	Two O^{2-} ions (total charge 4−) are required by the Mn^{4+} ion [manganese(IV)].
b. Lead(II) chloride	$PbCl_2$	Two Cl^- ions are required by the Pb^{2+} ion [lead(II)] for charge balance.

See Exercise 2.56.

A compound containing a transition metal usually requires a Roman numeral in its name.

Note that the use of a Roman numeral in a systematic name is required only in cases where more than one ionic compound forms between a given pair of elements. This case most commonly occurs for compounds containing transition metals, which often form more than one cation. *Elements that form only one cation do not need to be identified by a Roman numeral.* Common metals that do not require Roman numerals are the Group 1A elements, which form only 1+ ions; the Group 2A elements, which form only 2+ ions; and aluminum, which forms only Al^{3+}. The element silver deserves special mention at this point. In virtually all its compounds silver is found as the Ag^+ ion. Therefore, although silver is a transition metal (and can potentially form ions other than Ag^+), silver compounds are usually named without a Roman numeral. Thus AgCl is typically called silver chloride rather than silver(I) chloride, although the latter name is technically correct. Also, a Roman numeral is not used for zinc compounds, since zinc forms only the Zn^{2+} ion.

As shown in Sample Exercise 2.5, when a metal ion is present that forms more than one type of cation, the charge on the metal ion must be determined by balancing the positive and negative charges of the compound. To do this you must be able to recognize the common cations and anions and know their charges (see Tables 2.3 and 2.5).

Crystals of copper(II) sulfate.

Sample Exercise 2.6 Naming Binary Compounds

1. Give the systematic name for each of the following compounds:

   a. $CoBr_2$ **b.** $CaCl_2$ **c.** Al_2O_3

2. Given the following systematic names, write the formula for each compound:

   a. Chromium(III) chloride
   b. Gallium iodide

Solution

1.

Formula	**Name**	**Comments**
a. $CoBr_2$	Cobalt(II) bromide	Cobalt is a transition metal; the compound name must have a Roman numeral. The two Br^- ions must be balanced by a Co^{2+} ion.
b. $CaCl_2$	Calcium chloride	Calcium, an alkaline earth metal, forms only the Ca^{2+} ion. A Roman numeral is not necessary.
c. Al_2O_3	Aluminum oxide	Aluminum forms only the Al^{3+} ion. A Roman numeral is not necessary.

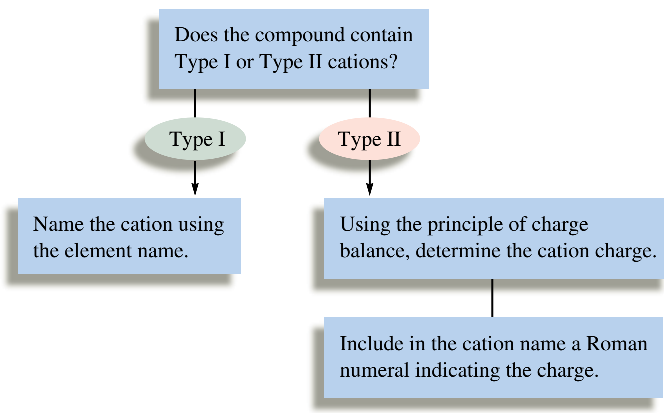
2.

Name	**Formula**	**Comments**
a. Chromium(III) chloride	$CrCl_3$	Chromium(III) indicates that Cr^{3+} is present, so 3 Cl^- ions are needed for charge balance.
b. Gallium iodide	GaI_3	Gallium always forms 3+ ions, so 3 I^- ions are required for charge balance.

See Exercises 2.57 and 2.58.

Various chromium compounds dissolved in water. From left to right: $CrCl_2$, $K_2Cr_2O_7$, $Cr(NO_3)_3$, $CrCl_3$, K_2CrO_4.

The following flowchart is useful when you are naming binary ionic compounds:



The common Type I and Type II ions are summarized in Fig. 2.22. Also shown in Fig. 2.22 are the common monatomic ions.

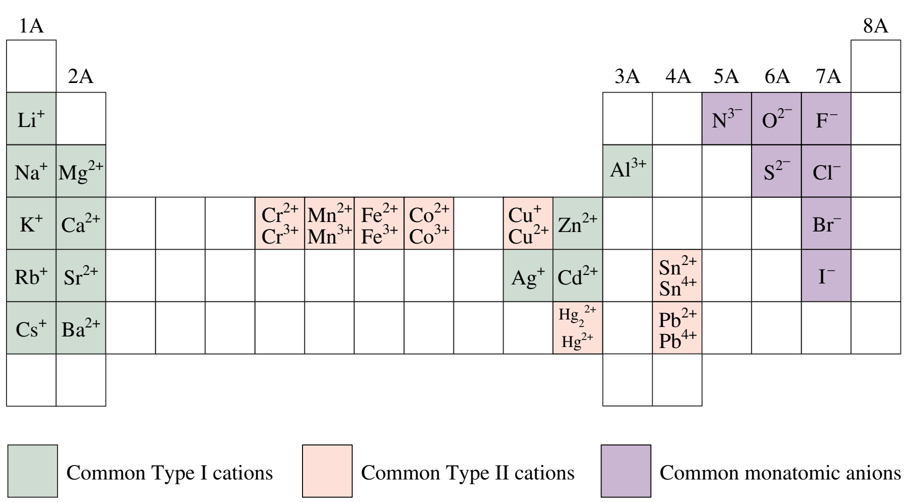


FIGURE 2.22
The common cations and anions.

TABLE 2.5 Common Polyatomic Ions

Ion	Name	Ion	Name
Hg_2^{2+}	Mercury(I)	NCS^-	Thiocyanate
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate (bicarbonate is a widely used common name)
NO_3^-	Nitrate		
SO_3^{2-}	Sulfite		
SO_4^{2-}	Sulfate	ClO^-	Hypochlorite
HSO_4^-	Hydrogen sulfate (bisulfate is a widely used common name)	ClO_2^-	Chlorite
		ClO_3^-	Chlorate
		ClO_4^-	Perchlorate
OH^-	Hydroxide	$C_2H_3O_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$H_2PO_4^-$	Dihydrogen phosphate	O_2^{2-}	Peroxide
		$C_2O_4^{2-}$	Oxalate

Ionic Compounds with Polyatomic Ions

Polyatomic ion formulas must be memorized.

We have not yet considered ionic compounds that contain polyatomic ions. For example, the compound ammonium nitrate, NH_4NO_3, contains the polyatomic ions NH_4^+ and NO_3^-. Polyatomic ions are assigned special names that *must be memorized* to name the compounds containing them. The most important polyatomic ions and their names are listed in Table 2.5.

Note in Table 2.5 that several series of anions contain an atom of a given element and different numbers of oxygen atoms. These anions are called **oxyanions.** When there are two members in such a series, the name of the one with the smaller number of oxygen atoms ends in *-ite* and the name of the one with the larger number ends in *-ate*—for example, sulfite (SO_3^{2-}) and sulfate (SO_4^{2-}). When more than two oxyanions make up a series, *hypo-* (less than) and *per-* (more than) are used as prefixes to name the members of the series with the fewest and the most oxygen atoms, respectively. The best example involves the oxyanions containing chlorine, as shown in Table 2.5.

Sample Exercise 2.7 Naming Compounds Containing Polyatomic Ions

1. Give the systematic name for each of the following compounds:

a. Na_2SO_4
b. KH_2PO_4
c. $Fe(NO_3)_3$
d. $Mn(OH)_2$
e. Na_2SO_3
f. Na_2CO_3

2. Given the following systematic names, write the formula for each compound:

a. Sodium hydrogen carbonate
b. Cesium perchlorate

c. Sodium hypochlorite
d. Sodium selenate
e. Potassium bromate

Solution

1.

Formula	Name	Comments
a. Na_2SO_4	Sodium sulfate	
b. KH_2PO_4	Potassium dihydrogen phosphate	
c. $Fe(NO_3)_3$	Iron(III) nitrate	Transition metal—name must contain a Roman numeral. The Fe^{3+} ion balances three NO_3^- ions.
d. $Mn(OH)_2$	Manganese(II) hydroxide	Transition metal—name must contain a Roman numeral. The Mn^{2+} ion balances three OH^- ions.
e. Na_2SO_3	Sodium sulfite	
f. Na_2CO_3	Sodium carbonate	

2.

Name	Formula	Comments
a. Sodium hydrogen carbonate	$NaHCO_3$	Often called sodium bicarbonate.
b. Cesium perchlorate	$CsClO_4$	
c. Sodium hypochlorite	$NaOCl$	
d. Sodium selenate	Na_2SeO_4	Atoms in the same group, like sulfur and selenium, often form similar ions that are named similarly. Thus SeO_4^{2-} is selenate, like SO_4^{2-} (sulfate).
e. Potassium bromate	$KBrO_3$	As above, BrO_3^- is bromate, like ClO_3^- (chlorate).

See Exercises 2.59 and 2.60.

Binary Covalent Compounds (Type III)

In *binary covalent compounds,* the element names follow the same rules as for binary ionic compounds.

Binary covalent compounds are formed between *two nonmetals*. Although these compounds do not contain ions, they are named very similarly to binary ionic compounds.

In the naming of binary covalent compounds, the following rules apply:

1. The first element in the formula is named first, using the full element name.
2. The second element is named as if it were an anion.
3. Prefixes are used to denote the numbers of atoms present. These prefixes are given in Table 2.6.
4. The prefix *mono-* is never used for naming the first element. For example, CO is called carbon monoxide, *not* monocarbon monoxide.

TABLE 2.6 Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

To see how these rules apply, we will now consider the names of the several covalent compounds formed by nitrogen and oxygen:

Compound	Systematic Name	Common Name
N_2O	Dinitrogen monoxide	Nitrous oxide
NO	Nitrogen monoxide	Nitric oxide
NO_2	Nitrogen dioxide	
N_2O_3	Dinitrogen trioxide	
N_2O_4	Dinitrogen tetroxide	
N_2O_5	Dinitrogen pentoxide	

Notice from the preceding examples that to avoid awkward pronunciations, we often drop the final *o* or *a* of the prefix when the element begins with a vowel. For example, N_2O_4 is called dinitrogen tetroxide, *not* dinitrogen tetr*a*oxide, and CO is called carbon monoxide, *not* carbon mon*o*oxide.

Some compounds are always referred to by their common names. The two best examples are water and ammonia. The systematic names for H_2O and NH_3 are never used.

Sample Exercise 2.8

Naming Type III Binary Compounds

1. Name each of the following compounds:

a. PCl_5
b. PCl_3
c. SO_2

2. From the following systematic names, write the formula for each compound:

a. Sulfur hexafluoride
b. Sulfur trioxide
c. Carbon dioxide

Solution

1.

Formula	Name
a. PCl_5	Phosphorus pentachloride
b. PCl_3	Phosphorus trichloride
c. SO_2	Sulfur dioxide

2.

Name	Formula
a. Sulfur hexafluoride	SF_6
b. Sulfur trioxide	SO_3
c. Carbon dioxide	CO_2

See Exercises 2.61 and 2.62.

The rules for naming binary compounds are summarized in Fig. 2.23. Prefixes to indicate the number of atoms are used only in Type III binary compounds (those containing two nonmetals). An overall strategy for naming compounds is given in Fig. 2.24.

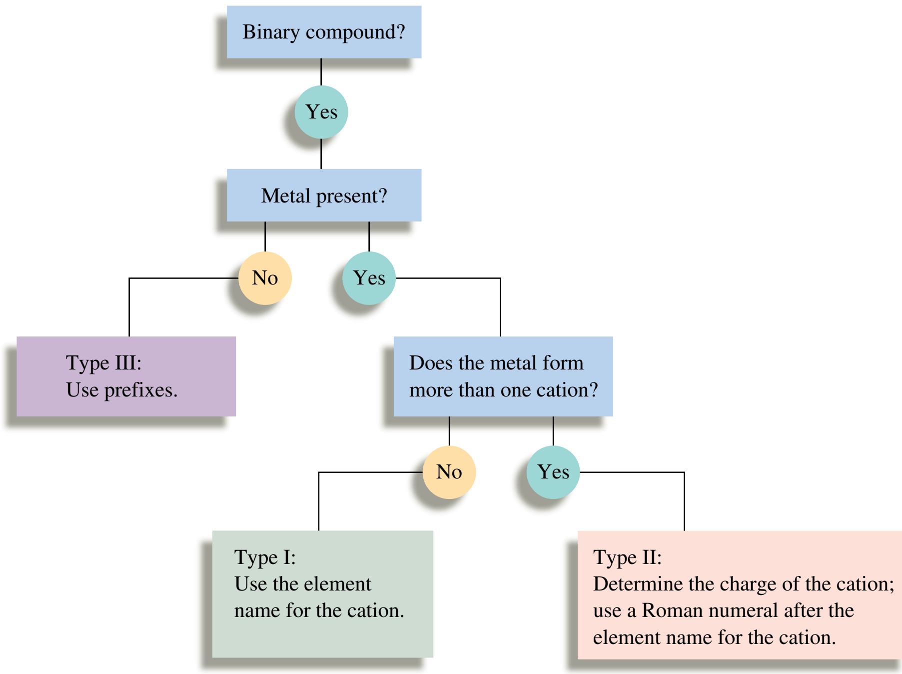


FIGURE 2.23
A flowchart for naming binary compounds.

Sample Exercise 2.9 Naming Various Types of Compounds

1. Give the systematic name for each of the following compounds:

a. P_4O_{10}
b. Nb_2O_5
c. Li_2O_2
d. $Ti(NO_3)_4$

2. Given the following systematic names, write the formula for each compound:

a. Vanadium(V) fluoride
b. Dioxygen difluoride
c. Rubidium peroxide
d. Gallium oxide

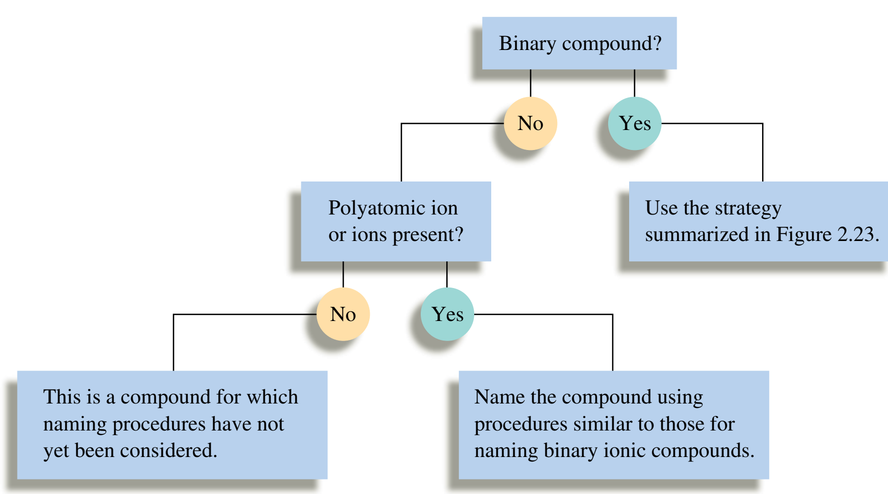


FIGURE 2.24
Overall strategy for naming chemical compounds.

Solution

1.

Compound	Name	Comment
a. P_4O_{10}	Tetraphosphorus decaoxide	Binary covalent compound (Type III), so prefixes are used. The *a* in *deca-* is sometimes dropped.
b. Nb_2O_5	Niobium(V) oxide	Type II binary compound containing Nb^{5+} and O^{2-} ions. Niobium is a transition metal and requires a Roman numeral.
c. Li_2O_2	Lithium peroxide	Type I binary compound containing the Li^+ and O_2^{2-} (peroxide) ions.
d. $Ti(NO_3)_4$	Titanium(IV) nitrate	Not a binary compound. Contains the Ti^{4+} and NO_3^- ions. Titanium is a transition metal and requires a Roman numeral.

2.

Name	Chemical Formula	Comment
a. Vanadium(V) fluoride	VF_5	The compound contains V^{5+} ions and requires five F^- ions for charge balance.
b. Dioxygen difluoride	O_2F_2	The prefix *di-* indicates two of each atom.
c. Rubidium peroxide	Rb_2O_2	Because rubidium is in Group 1A, it forms only 1+ ions. Thus two Rb^+ ions are needed to balance the 2− charge on the peroxide ion (O_2^{2-}).
d. Gallium oxide	Ga_2O_3	Because gallium is in Group 3A, like aluminum, it forms only 3+ ions. Two Ga^{3+} ions are required to balance the charge on three O^{2-} ions.

See Exercises 2.63, 2.65, and 2.66.

Acids

Acids can be recognized by the hydrogen that appears first in the formula.

When dissolved in water, certain molecules produce a solution containing free H^+ ions (protons). These substances, **acids,** will be discussed in detail in Chapters 4, 14, and 15. Here we will simply present the rules for naming acids.

An acid can be viewed as a molecule with one or more H^+ ions attached to an anion. The rules for naming acids depend on whether the anion contains oxygen. If the *anion does not contain oxygen,* the acid is named with the prefix *hydro-* and the suffix *-ic.* For example, when gaseous HCl is dissolved in water, it forms hydrochloric acid. Similarly, HCN and H_2S dissolved in water are called hydrocyanic and hydrosulfuric acids, respectively.

When the *anion contains oxygen,* the acidic name is formed from the root name of the anion with a suffix of *-ic* or *-ous,* depending on the name of the anion.

1. If the anion name ends in *-ate,* the suffix *-ic* is added to the root name. For example, H_2SO_4 contains the sulfate anion (SO_4^{2-}) and is called sulfuric acid; H_3PO_4 contains the phosphate anion (PO_4^{3-}) and is called phosphoric acid; and $HC_2H_3O_2$ contains the acetate ion ($C_2H_3O_2^-$) and is called acetic acid.
2. If the anion has an *-ite* ending, the *-ite* is replaced by *-ous.* For example, H_2SO_3, which contains sulfite (SO_3^{2-}), is named sulfurous acid; and HNO_2, which contains nitrite (NO_2^-), is named nitrous acid.

TABLE 2.7 Names of Acids* That Do Not Contain Oxygen

Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
HCN	Hydrocyanic acid
H_2S	Hydrosulfuric acid

*Note that these acids are aqueous solutions containing these substances.

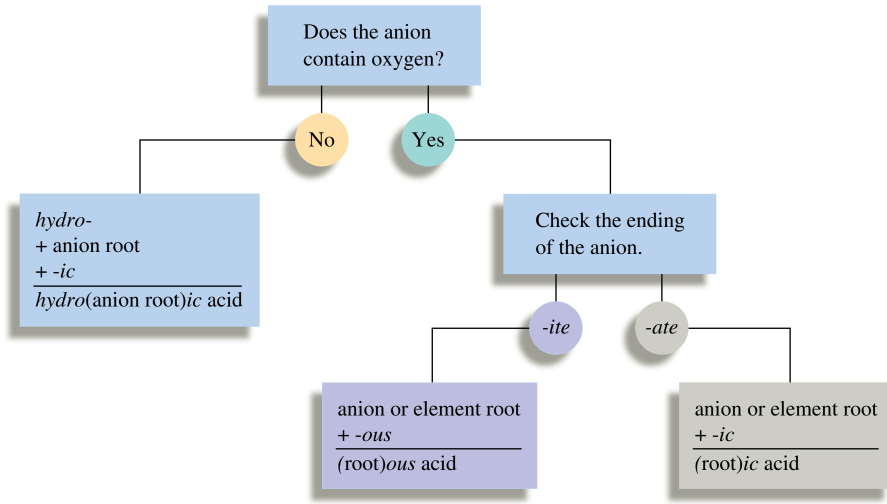


FIGURE 2.25
A flowchart for naming acids. An acid is best considered as one or more H^+ ions attached to an anion.

TABLE 2.8 Names of Some Oxygen-Containing Acids

Acid	Name
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic Acid

The application of these rules can be seen in the names of the acids of the oxyanions of chlorine:

Acid	Anion	Name
$HClO_4$	Perchlor*ate*	Perchlor*ic* acid
$HClO_3$	Chlor*ate*	Chlor*ic* acid
$HClO_2$	Chlor*ite*	Chlor*ous* acid
$HClO$	Hypochlor*ite*	Hypochlor*ous* acid

The names of the most important acids are given in Tables 2.7 and 2.8. An overall strategy for naming acids is shown in Fig. 2.25.

Key Terms

Section 2.2
law of conservation of mass
law of definite proportion
law of multiple proportions

Section 2.3
atomic masses
atomic weights
Avogadro's hypothesis

Section 2.4
cathode-ray tube
electron
radioactivity
nuclear atom
nucleus

Section 2.5
proton
neutron
isotopes
atomic number
mass number

For Review

Fundamental laws
- Conservation of mass
- Definite proportion
- Multiple proportions

Dalton's atomic theory
- All elements are composed of atoms.
- All atoms of a given element are identical.
- Chemical compounds are formed when atoms combine.
- Atoms are not changed in chemical reactions but the way they are bound together changes.

Early atomic experiments and models
- Thomson model
- Millikan experiment
- Rutherford experiment
- Nuclear model

Section 2.6

chemical bond
covalent bond
molecule
chemical formula
structural formula
space-filling model
ball-and-stick model
ion
cation
anion
ionic bond
ionic solid (salt)
polyatomic ion

Section 2.7

periodic table
metal
nonmetal
group (family)
alkali metals
alkaline earth metals
halogens
noble gases
period

Section 2.8

binary compounds
binary ionic compounds
oxyanions
binary covalent compounds
acid

Atomic structure

- Small dense nucleus contains protons and neutrons.
  - Protons—positive charge
  - Neutrons—no charge
- Electrons reside outside the nucleus in the relatively large remaining atomic volume.
  - Electrons—negative charge, small mass (1/1840 of proton)
- Isotopes have the same atomic number but different mass numbers.

Atoms combine to form molecules by sharing electrons to form covalent bonds.

- Molecules are described by chemical formulas.
- Chemical formulas show number and type of atoms.
  - Structural formula
  - Ball-and-stick model
  - Space-filling model

Formation of ions

- Cation—formed by loss of an electron, positive charge
- Anion—formed by gain of an electron, negative charge
- Ionic bonds—formed by interaction of cations and anions

The periodic table organizes elements in order of increasing atomic number.

- Elements with similar properties are in columns, or groups.
- Metals are in the majority and tend to form cations.
- Nonmetals tend to form anions.

Compounds are named using a system of rules depending on the type of compound.

- Binary compounds
  - Type I—contain a metal that always forms the same cation
  - Type II—contain a metal that can form more than one cation
  - Type III—contain two nonmetals
- Compounds containing a polyatomic ion

REVIEW QUESTIONS

1. Use Dalton's atomic theory to account for each of the following.
   a. the law of conservation of mass
   b. the law of definite proportion
   c. the law of multiple proportions
2. What evidence led to the conclusion that cathode rays had a negative charge?
3. What discoveries were made by J. J. Thomson, Henri Becquerel, and Lord Rutherford? How did Dalton's model of the atom have to be modified to account for these discoveries?
4. Consider Ernest Rutherford's alpha-particle bombardment experiment illustrated in Figure 2.12. How did the results of this experiment lead Rutherford away from the plum pudding model of the atom to propose the nuclear model of the atom?
5. Do the proton and the neutron have exactly the same mass? How do the masses of the proton and neutron compare to the mass of the electron? Which particles make the greatest contribution to the mass of an atom? Which particles make the greatest contribution to the chemical properties of an atom?
6. What is the distinction between atomic number and mass number? Between mass number and atomic mass?
7. Distinguish between the terms *family* and *period* in connection with the periodic table. For which of these terms is the term *group* also used?
8. The compounds $AlCl_3$, $CrCl_3$, and ICl_3 have similar formulas, yet each follows a different set of rules to name it. Name these compounds, and then compare and contrast the nomenclature rules used in each case.

9. When metals react with nonmetals, an ionic compound generally results. What is the predicted general formula for the compound formed between an alkali metal and sulfur? Between an alkaline earth metal and nitrogen? Between aluminum and a halogen?
10. How would you name $HBrO_4$, KIO_3, $NaBrO_2$, and HIO? Refer to Table 2.5 and the acid nomenclature discussion in the text.

Active Learning Questions

These questions are designed to be used by groups of students in class. The questions allow students to explore their understanding of concepts through discussion and peer teaching. The real value of these questions is the learning that occurs while students talk to each other about chemical concepts.

1. Which of the following is true about an individual atom? Explain.
- **a.** An individual atom should be considered to be a solid.
- **b.** An individual atom should be considered to be a liquid.
- **c.** An individual atom should be considered to be a gas.
- **d.** The state of the atom depends on which element it is.
- **e.** An individual atom cannot be considered to be a solid, liquid, or gas.

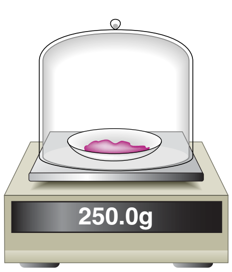
Justify your choice, and for choices you did not pick, explain what is wrong with them.

2. How would you go about finding the number of "chalk molecules" it takes to write your name on the board? Provide an explanation of all you would need to do and a sample calculation.

3. These questions concern the work of J. J. Thomson.
- **a.** From Thomson's work, which particles do you think he would feel are most important for the formation of compounds (chemical changes) and why?
- **b.** Of the remaining two subatomic particles, which do you place second in importance for forming compounds and why?
- **c.** Propose three models that explain Thomson's findings and evaluate them. To be complete you should include Thomson's findings.

4. Heat is applied to an ice cube in a closed container until only steam is present. Draw a representation of this process, assuming you can see it at an extremely high level of magnification. What happens to the size of the molecules? What happens to the total mass of the sample?

5. You have a chemical in a sealed glass container filled with air. The setup is sitting on a balance as shown below. The chemical is ignited by means of a magnifying glass focusing sunlight on the reactant. After the chemical has completely burned, which of the following is true? Explain your answer.



- **a.** The balance will read less than 250.0 g.
- **b.** The balance will read 250.0 g.
- **c.** The balance will read greater than 250.0 g.
- **d.** Cannot be determined without knowing the identity of the chemical.

6. You take three compounds consisting of two elements and decompose them. To determine the relative masses of *X*, *Y*, and *Z*, you collect and weigh the elements, obtaining the following data:

Elements in Compound	Masses of Elements
X and *Y*	$X = 0.4$ g, $Y = 4.2$ g
Y and *Z*	$Y = 1.4$ g, $Z = 1.0$ g
X and *Y*	$X = 2.0$ g, $Y = 7.0$ g

- **a.** What are the assumptions in solving this problem?
- **b.** What are the relative masses of *X*, *Y*, and *Z*?
- **c.** What are the chemical formulas of the three compounds?
- **d.** If you decompose 21 g of compound *XY*, how much of each element is present?

7. The vitamin niacin (nicotinic acid, $C_6H_5NO_2$) can be isolated from a variety of natural sources such as liver, yeast, milk, and whole grain. It also can be synthesized from commercially available materials. Which source of nicotinic acid, from a nutritional view, is best for use in a multivitamin tablet? Why?

8. One of the best indications of a useful theory is that it raises more questions for further experimentation than it originally answered. Does this apply to Dalton's atomic theory? Give examples.

9. Dalton assumed that all atoms of the same element were identical in all their properties. Explain why this assumption is not valid.

10. Evaluate each of the following as an acceptable name for water:
- **a.** dihydrogen oxide
- **b.** hydroxide hydride
- **c.** hydrogen hydroxide
- **d.** oxygen dihydride

11. Why do we call $Ba(NO_3)_2$ barium nitrate, but we call $Fe(NO_3)_2$ iron(II) nitrate?

12. Why is calcium dichloride not the correct systematic name for $CaCl_2$?

13. The common name for NH_3 is ammonia. What would be the systematic name for NH_3? Support your answer.

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the *Solutions Guide*.

Questions

14. What refinements had to be made in Dalton's atomic theory to account for Gay-Lussac's results on the combining volumes of gases?

15. When hydrogen is burned in oxygen to form water, the composition of water formed does not depend on the amount of oxygen reacted. Interpret this in terms of the law of definite proportion.
16. The two most reactive families of elements are the halogens and the alkali metals. How do they differ in their reactivities?
17. Explain the law of conservation of mass, the law of definite proportion, and the law of multiple proportions.
18. Section 2.3 describes the postulates of Dalton's atomic theory. With some modifications, these postulates hold up very well regarding how we view elements, compounds, and chemical reactions today. Answer the following questions concerning Dalton's atomic theory and the modifications made today.
    a. The atom can be broken down into smaller parts. What are the smaller parts?
    b. How are atoms of hydrogen identical to each other and how can they be different from each other?
    c. How are atoms of hydrogen different from atoms of helium? How can H atoms be similar to He atoms?
    d. How is water different from hydrogen peroxide (H_2O_2) even though both compounds are composed of only hydrogen and oxygen?
    e. What happens in a chemical reaction and why is mass conserved in a chemical reaction?
19. The contributions of J. J. Thomson and Ernest Rutherford led the way to today's understanding of the structure of the atom. What were their contributions?
20. What is the modern view of the structure of the atom?
21. The number of protons in an atom determines the identity of the atom. What does the number and arrangement of the electrons in an atom determine? What does the number of neutrons in an atom determine?
22. Distinguish between the following terms.
    a. molecule versus ion
    b. covalent bonding versus ionic bonding
    c. molecule versus compound
    d. anion versus cation
23. Which of the following statements are true? For the false statements, correct them.
    a. Most of the known elements are metals.
    b. Element 118 should be a nonmetal.
    c. Hydrogen has mostly metallic properties.
    d. A family of elements is also known as a period of elements.
    e. When an alkaline earth metal, A, reacts with a halogen, X, the formula of the covalent compound formed should be A_2X.
24. Each of the following compounds has three possible names listed for it. For each compound, what is the correct name and why aren't the other names used?
    a. N_2O: nitrogen oxide, nitrogen(I) oxide, dinitrogen monoxide
    b. Cu_2O: copper oxide, copper(I) oxide, dicopper monoxide
    c. Li_2O: lithium oxide, lithium(I) oxide, dilithium monoxide

Exercises

In this section similar exercises are paired.

Development of the Atomic Theory

25. When mixtures of gaseous H_2 and gaseous Cl_2 react, a product forms that has the same properties regardless of the relative amounts of H_2 and Cl_2 used.
    a. How is this result interpreted in terms of the law of definite proportion?
    b. When a volume of H_2 reacts with an equal volume of Cl_2 at the same temperature and pressure, what volume of product having the formula HCl is formed?
26. A reaction of 1 liter of chlorine gas (Cl_2) with 3 liters of fluorine gas (F_2) yields 2 liters of a gaseous product. All gas volumes are at the same temperature and pressure. What is the formula of the gaseous product?

27. Hydrazine, ammonia, and hydrogen azide all contain only nitrogen and hydrogen. The mass of hydrogen that combines with 1.00 g of nitrogen for each compound is 1.44×10^{-1} g, 2.16×10^{-1} g, and 2.40×10^{-2} g, respectively. Show how these data illustrate the law of multiple proportions.
28. Consider 100.0-g samples of two different compounds consisting only of carbon and oxygen. One compound contains 27.2 g of carbon and the other has 42.9 g of carbon. How can these data support the law of multiple proportions if 42.9 is not a multiple of 27.2? Show that these data support the law of multiple proportions.

29. Early tables of atomic weights (masses) were generated by measuring the mass of a substance that reacts with 1.00 g of oxygen. Given the following data and taking the atomic mass of hydrogen as 1.00, generate a table of relative atomic masses for oxygen, sodium, and magnesium.

Element	Mass That Combines with 1.00 g Oxygen	Assumed Formula
Hydrogen	0.126 g	HO
Sodium	2.875 g	NaO
Magnesium	1.500 g	MgO

How do your values compare with those in the periodic table? How do you account for any differences?
30. Indium oxide contains 4.784 g of indium for every 1.000 g of oxygen. In 1869, when Mendeleev first presented his version of the periodic table, he proposed the formula In_2O_3 for indium oxide. Before that time it was thought that the formula was InO. What values for the atomic mass of indium are obtained using these two formulas? Assume that oxygen has an atomic mass of 16.00.

The Nature of the Atom

31. From the information in this chapter on the mass of the proton, the mass of the electron, and the sizes of the nucleus and the atom, calculate the densities of a hydrogen nucleus and a hydrogen atom.
32. If you wanted to make an accurate scale model of the hydrogen atom and decided that the nucleus would have a diameter of 1 mm, what would be the diameter of the entire model?

33. In an experiment it was found that the total charge on an oil drop was 5.93×10^{-18} C. How many negative charges does the drop contain?
34. A chemist in a galaxy far, far away performed the Millikan oil drop experiment and got the following results for the charges on

various drops. Use these data to calculate the charge of the electron in zirkombs.

2.56×10^{-12} zirkombs	7.68×10^{-12} zirkombs
3.84×10^{-12} zirkombs	6.40×10^{-13} zirkombs

35. What are the symbols of the following metals: sodium, radium, iron, gold, manganese, lead.

36. What are the symbols of the following nonmetals: fluorine, chlorine, bromine, sulfur, oxygen, phosphorus?

37. Give the names of the metals that correspond to the following symbols: Sn, Pt, Hg, Mg, K, Ag.

38. Give the names of the nonmetals that correspond to the following symbols: As, I, Xe, He, C, Si.

39. **a.** Classify the following elements as metals or nonmetals:

Mg	Si	Rn
Ti	Ge	Eu
Au	B	Am
Bi	At	Br

b. The distinction between metals and nonmetals is really not a clear one. Some elements, called *metalloids,* are intermediate in their properties. Which of these elements would you reclassify as metalloids? What other elements in the periodic table would you expect to be metalloids?

40. **a.** List the noble gas elements. Which of the noble gases has only radioactive isotopes? (This situation is indicated on most periodic tables by parentheses around the mass of the element. See inside front cover.)

b. Which lanthanide element and which transition element have only radioactive isotopes?

41. In the periodic table, how many elements are found in

a. Group 2A? **c.** the nickel group?
b. the oxygen family? **d.** Group 8A?

42. In the periodic table, how many elements are found

a. in the halogen group?
b. in the alkali family?
c. in the lanthanide series?
d. classified as transition metals?

43. How many protons and neutrons are in the nucleus of each of the following atoms? In a neutral atom of each element, how many electrons are present?

a. ^{79}Br **d.** ^{133}Cs
b. ^{81}Br **e.** ^{3}H
c. ^{239}Pu **f.** ^{56}Fe

44. What number of protons and neutrons are contained in the nucleus of each of the following atoms? Assuming each atom is uncharged, what number of electrons are present?

a. $^{235}_{92}U$ **d.** $^{208}_{82}Pb$
b. $^{13}_{6}C$ **e.** $^{86}_{37}Rb$
c. $^{57}_{26}Fe$ **f.** $^{41}_{20}Ca$

45. Write the atomic symbol ($^{A}_{Z}X$) for each of the following isotopes.

a. $Z = 8$, number of neutrons = 9
b. the isotope of chlorine in which $A = 37$
c. $Z = 27$, $A = 60$
d. number of protons = 26, number of neutrons = 31
e. the isotope of I with a mass number of 131
f. $Z = 3$, number of neutrons = 4

46. Write the atomic symbol ($^{A}_{Z}X$) for each of the isotopes described below.

a. number of protons = 27, number of neutrons = 31
b. the isotope of boron with mass number 10
c. $Z = 12$, $A = 23$
d. atomic number 53, number of neutrons = 79
e. $Z = 9$, number of neutrons = 10
f. number of protons = 29, mass number 65

47. What is the symbol for an ion with 63 protons, 60 electrons, and 88 neutrons? If an ion contains 50 protons, 68 neutrons, and 48 electrons, what is its symbol?

48. What is the symbol of an ion with 16 protons, 18 neutrons, and 18 electrons? What is the symbol for an ion that has 16 protons, 16 neutrons, and 18 electrons?

49. Complete the following table:

Symbol	Number of Protons in Nucleus	Number of Neutrons in Nucleus	Number of Electrons	Net Charge
$^{238}_{92}U$				
	20	20		2+
	23	28	20	
$^{89}_{39}Y$				
	35	44	36	
	15	16		3−

50.

Symbol	Number of Protons in Nucleus	Number of Neutrons in Nucleus	Number of Electrons	Net Charge
$^{53}_{26}Fe^{2+}$				
	26	33		3+
	85	125	86	
	13	14	10	
		76	54	2−

51. For each of the following sets of elements, label each as either noble gases, halogens, alkali metals, alkaline earth metals, or transition metals.

a. Ti, Fe, Ag **d.** Ne, Kr, Xe
b. Mg, Sr, Ba **e.** F, Br, I
c. Li, K, Rb

52. Consider the elements of Group 4A (the "carbon family"): C, Si, Ge, Sn, and Pb. What is the trend in metallic character as one goes down this group? What is the trend in metallic character going from left to right across a period in the periodic table?

53. Would you expect each of the following atoms to gain or lose electrons when forming ions? What ion is the most likely in each case?

a. Ra **b.** In **c.** P **d.** Te **e.** Br **f.** Rb

54. For each of the following atomic numbers, use the periodic table to write the formula (including the charge) for the simple *ion* that the element is most likely to form in ionic compounds.

a. 13 **b.** 34 **c.** 56 **d.** 7 **e.** 87 **f.** 35

Nomenclature

55. Name the compounds in parts a–d and write the formulas for the compounds in parts e–h.

a. NaBr
b. Rb_2O
c. CaS
d. AlI_3
e. strontium fluoride
f. aluminum selenide
g. potassium nitride
h. magnesium phosphide

56. Name the compounds in parts a–d and write the formulas for the compounds in parts e–h.

a. Hg_2O
b. $FeBr_3$
c. CoS
d. $TiCl_4$
e. tin(II) nitride
f. cobalt(III) iodide
g. mercury(II) oxide
h. chromium(VI) sulfide

57. Name each of the following compounds:

a. CsF **b.** Li_3N **c.** Ag_2S **d.** MnO_2 **e.** TiO_2 **f.** Sr_3P_2

58. Write the formula for each of the following compounds:

a. zinc chloride
b. tin(IV) fluoride
c. calcium nitride
d. aluminum sulfide
e. mercury(I) selenide
f. silver iodide

59. Name each of the following compounds:

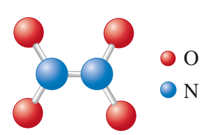
a. $BaSO_3$ **b.** $NaNO_2$ **c.** $KMnO_4$ **d.** $K_2Cr_2O_7$

60. Write the formula for each of the following compounds:

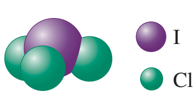
a. chromium(III) hydroxide
b. magnesium cyanide
c. lead(IV) carbonate
d. ammonium acetate

61. Name each of the following compounds:

a.



b.



c. SO_2
d. P_2S_5

62. Write the formula for each of the following compounds:

a. diboron trioxide
b. arsenic pentafluoride
c. dinitrogen monoxide
d. sulfur hexachloride

63. Name each of the following compounds:

a. CuI
b. CuI_2
c. CoI_2
d. Na_2CO_3
e. $NaHCO_3$
f. S_4N_4
g. SF_6
h. NaOCl
i. $BaCrO_4$
j. NH_4NO_3

64. Name each of the following compounds:

a. $HC_2H_3O_2$
b. NH_4NO_2
c. Co_2S_3
d. ICl
e. $Pb_3(PO_4)_2$
f. KIO_3
g. H_2SO_4
h. Sr_3N_2
i. $Al_2(SO_3)_3$
j. SnO_2
k. Na_2CrO_4
l. HClO

65. Write the formula for each of the following compounds:

a. sulfur difluoride
b. sulfur hexafluoride
c. sodium dihydrogen phosphate
d. lithium nitride
e. chromium(III) carbonate
f. tin(II) fluoride
g. ammonium acetate
h. ammonium hydrogen sulfate
i. cobalt(III) nitrate
j. mercury(I) chloride
k. potassium chlorate
l. sodium hydride

66. Write the formula for each of the following compounds:

a. chromium(VI) oxide
b. disulfur dichloride
c. nickel(II) fluoride
d. potassium hydrogen phosphate
e. aluminum nitride
f. ammonia
g. manganese(IV) sulfide
h. sodium dichromate
i. ammonium sulfite
j. carbon tetraiodide

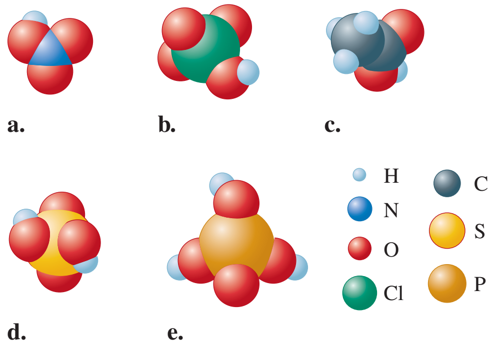
67. Write the formula for each of the following compounds:

a. sodium oxide
b. sodium peroxide
c. potassium cyanide
d. copper(II) nitrate
e. selenium tetrabromide
f. iodous acid
g. lead(IV) sulfide
h. copper(I) chloride
i. gallium arsenide
j. cadmium selenide
k. zinc sulfide
l. nitrous acid
m. diphosphorus pentoxide

68. Write the formula for each of the following compounds:

a. ammonium hydrogen phosphate
b. mercury(I) sulfide
c. silicon dioxide
d. sodium sulfite
e. aluminum hydrogen sulfate
f. nitrogen trichloride
g. hydrobromic acid
h. bromous acid
i. perbromic acid
j. potassium hydrogen sulfide
k. calcium iodide
l. cesium perchlorate

69. Name the following acids illustrated below.



70. Each of the following compounds is incorrectly named. What is wrong with each name, and what is the correct name for each compound?

a. $FeCl_3$, iron chloride
b. NO_2, nitrogen(IV) oxide
c. CaO, calcium(II) monoxide
d. Al_2S_3, dialuminum trisulfide
e. $Mg(C_2H_3O_2)_2$, manganese diacetate
f. $FePO_4$, iron(II) phosphide
g. P_2S_5, phosphorous sulfide
h. Na_2O_2, sodium oxide
i. HNO_3, nitrate acid
j. H_2S, sulfuric acid

Additional Exercises

71. Chlorine has two natural isotopes: $^{37}_{17}Cl$ and $^{35}_{17}Cl$. Hydrogen reacts with chlorine to form the compound HCl. Would a given amount of hydrogen react with different masses of the two chlorine isotopes? Does this conflict with the law of definite proportion? Why or why not?

72. Which of the following statements is(are) *true?* For the false statements, correct them.

a. All particles in the nucleus of an atom are charged.
b. The atom is best described as a uniform sphere of matter in which electrons are embedded.
c. The mass of the nucleus is only a very small fraction of the mass of the entire atom.
d. The volume of the nucleus is only a very small fraction of the total volume of the atom.
e. The number of neutrons in a neutral atom must equal the number of electrons.

73. The isotope of an unknown element, X, has a mass number of 79. The most stable ion of the isotope has 36 electrons and forms a binary compound with sodium having a formula of Na_2X. Which of the following statements is(are) *true?* For the false statements, correct them.

a. The binary compound formed between X and fluorine will be a covalent compound.
b. The isotope of X contains 38 protons.
c. The isotope of X contains 41 neutrons.
d. The identity of X is strontium, Sr.

74. For each of the following ions, indicate the total number of protons and electrons in the ion. For the positive ions in the list, predict the formula of the simplest compound formed between each positive ion and the oxide ion. For the negative ions in the list, predict the formula of the simplest compound formed between each negative ion and the aluminum ion.

a. Fe^{2+}	**e.** S^{2-}
b. Fe^{3+}	**f.** P^{3-}
c. Ba^{2+}	**g.** Br^{-}
d. Cs^{+}	**h.** N^{3-}

75. The formulas and common names for several substances are given below. Give the systematic names for these substances.

a. sugar of lead	$Pb(C_2H_3O_2)_2$
b. blue vitrol	$CuSO_4$
c. quicklime	CaO
d. Epsom salts	$MgSO_4$
e. milk of magnesia	$Mg(OH)_2$
f. gypsum	$CaSO_4$
g. laughing gas	N_2O

76. Identify each of the following elements:

a. a member of the same family as oxygen whose most stable ion contains 54 electrons
b. a member of the alkali metal family whose most stable ion contains 36 electrons
c. a noble gas with 18 protons in the nucleus
d. a halogen with 85 protons and 85 electrons

77. An element's most stable ion forms an ionic compound with bromine, having the formula XBr_2. If the ion of element X has a mass number of 230 and has 86 electrons, what is the identity of the element, and how many neutrons does it have?

78. A certain element has only two naturally occurring isotopes: one with 18 neutrons and the other with 20 neutrons. The element forms 1− charged ions when in ionic compounds. Predict the identity of the element. What number of electrons does the 1− charged ion have?

79. The designations 1A through 8A used for certain families of the periodic table are helpful for predicting the charges on ions in binary ionic compounds. In these compounds, the metals generally take on a positive charge equal to the family number, while the nonmetals take on a negative charge equal to the family number minus eight. Thus the compound between sodium and chlorine contains Na^+ ions and Cl^- ions and has the formula NaCl. Predict the formula and the name of the binary compound formed from the following pairs of elements.

a. Ca and N	**e.** Ba and I
b. K and O	**f.** Al and Se
c. Rb and F	**g.** Cs and P
d. Mg and S	**h.** In and Br

80. By analogy with phosphorous compounds, name the following: Na_3AsO_4, H_3AsO_4, $Mg_3(SbO_4)_2$.

81. A sample of H_2SO_4 contains 2.02 g of hydrogen, 32.07 g of sulfur, and 64.00 g of oxygen. How many grams of sulfur and grams of oxygen are present in a second sample of H_2SO_4 containing 7.27 g of hydrogen?

82. In a reaction, 34.0 g of chromium(III) oxide reacts with 12.1 g of aluminum to produce chromium and aluminum oxide. If 23.3 g of chromium is produced, what mass of aluminum oxide is produced?

Challenge Problems

83. The elements in one of the groups in the periodic table are often called the coinage metals. Identify the elements in this group based on your own experience.

84. Reaction of 2.0 L of hydrogen gas with 1.0 L of oxygen gas yields 2.0 L of water vapor. All gases are at the same temperature and pressure. Show how these data support the idea that oxygen gas is a diatomic molecule. Must we consider hydrogen to be a diatomic molecule to explain these results?

85. A combustion reaction involves the reaction of a substance with oxygen gas. The complete combustion of any hydrocarbon (binary compound of carbon and hydrogen) produces carbon dioxide and water as the only products. Octane is a hydrocarbon that is found in gasoline. Complete combustion of octane produces 8 liters of carbon dioxide for every 9 liters of water vapor (both measured at the same temperature and pressure). What is the ratio of carbon atoms to hydrogen atoms in a molecule of octane?

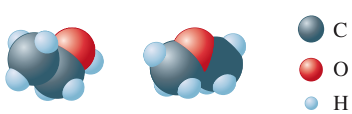
86. A chemistry instructor makes the following claim: "Consider that if the nucleus were the size of a grape, the electrons would be about 1 *mile* away on average." Is this claim reasonably accurate? Provide mathematical support.

87. Two elements, R and Q, combine to form two binary compounds. In the first compound, 14.0 g of R combines with 3.00 g of Q. In the second compound, 7.00 g of R combines with 4.50 g of Q. Show that these data are in accord with the law of multiple proportions. If the formula of the second compound is RQ, what is the formula of the first compound?

88. The early alchemists used to do an experiment in which water was boiled for several days in a sealed glass container. Eventually, some solid residue would appear in the bottom of the flask, which was interpreted to mean that some of the water in the flask had been converted into "earth." When Lavoisier repeated this experiment, he found that the water weighed the same before and after heating and the mass of the flask plus the solid residue equaled the original mass of the flask. Were the alchemists correct? Explain what really happened. (This experiment is described in the article by A. F. Scott in *Scientific American,* January 1984.)

89. Each of the following statements is true, but Dalton might have had trouble explaining some of them with his atomic theory. Give explanations for the following statements.

   a. The space-filling models for ethyl alcohol and dimethyl ether are shown below.



   These two compounds have the same composition by mass (52% carbon, 13% hydrogen, and 35% oxygen), yet the two have different melting points, boiling points, and solubilities in water.

   b. Burning wood leaves an ash that is only a small fraction of the mass of the original wood.

   c. Atoms can be broken down into smaller particles.

   d. One sample of lithium hydride is 87.4% lithium by mass, while another sample of lithium hydride is 74.9% lithium by mass. However, the two samples have the same properties.

90. You have two distinct gaseous compounds made from element X and element Y. The mass percents are as follows:

Compound I: 30.43% X, 69.57% Y
Compound II: 63.64% X, 36.36% Y

In their natural standard states, element X and element Y exist as gases. (Monatomic? Diatomic? Triatomic? That is for you to determine.) When you react "gas X" with "gas Y" to make the products, you get the following data (all at standard pressure and temperature):

1 volume "gas X" + 2 volumes "gas Y" $\longrightarrow$ 2 volumes compound I

2 volumes "gas X" + 1 volume "gas Y" $\longrightarrow$ 2 volumes compound II

Assume the simplest possible formulas for reactants and products in the chemical equations above. Then, determine the relative atomic masses of element X and element Y.

Integrative Problems

These problems require the integration of multiple concepts to find the solutions.

91. What is the systematic name of Ta_2O_5? If the charge on the metal remained constant and then sulfur was substituted for oxygen, how would the formula change? What is the difference in the total number of protons between Ta_2O_5 and its sulfur analog?

92. A binary ionic compound is known to contain a cation with 51 protons and 48 electrons. The anion contains one-third the number of protons as the cation. The number of electrons in the anion is equal to the number of protons plus 1. What is the formula of this compound? What is the name of this compound?

93. Using the information in Table 2.1, answer the following questions. In an ion with an unknown charge, the total mass of all the electrons was determined to be 2.55×10^{-26} g, while the total mass of its protons was 5.34×10^{-23} g. What is the identity and charge of this ion? What is the symbol and mass number of a neutral atom whose total mass of its electrons is 3.92×10^{-26} g, while its neutrons have a mass of 9.35×10^{-23} g?

Marathon Problem

This problem is designed to incorporate several concepts and techniques into one situation. Marathon Problems can be used in class by groups of students to help facilitate problem-solving skills.

94. You have gone back in time and are working with Dalton on a table of relative masses. Following are his data.

0.602 g gas A reacts with 0.295 g gas B
0.172 g gas B reacts with 0.401 g gas C
0.320 g gas A reacts with 0.374 g gas C

   a. Assuming simplest formulas (AB, BC, and AC), construct a table of relative masses for Dalton.

b. Knowing some history of chemistry, you tell Dalton that if he determines the volumes of the gases reacted at constant temperature and pressure, he need not assume simplest formulas. You collect the following data:

6 volumes gas A + 1 volume gas B → 4 volumes product
1 volume gas B + 4 volumes gas C → 4 volumes product
3 volumes gas A + 2 volumes gas C → 6 volumes product

Write the simplest balanced equations, and find the actual relative masses of the elements. Explain your reasoning.

Get help understanding core concepts and visualizing molecular-level interactions, and practice problem solving, by visiting the Online Study Center at **college.hmco.com/PIC/zumdahl7e.**