CHAPTER 5

Early Atomic Theory and Structure

Multiple lightning strikes over the desert in Arizona. Lightning occurs when electrons move to neutralize a charge difference between the clouds and the Earth.

Chapter Outline

Pure substances are classified as elements or compounds, but just what makes a substance possess its unique properties? How small a piece of salt will still taste salty? Carbon dioxide puts out fires, is used by plants to produce oxygen, and forms dry ice when solidified. But how small a mass of this material still behaves like carbon dioxide? Substances are in their simplest identifiable form at the atomic, ionic, or molecular level. Further division produces a loss of characteristic properties.

What particles lie within an atom or ion? How are these tiny particles alike? How do they differ? How far can we continue to divide them? Alchemists began the quest, early chemists laid the foundation, and modern chemists continue to build and expand on models of the atom.

5.1 Early Thoughts

The structure of matter has long intrigued and engaged us. The earliest models of the atom were developed by the ancient Greek philosophers. About 440 B.C. Empedocles stated that all matter was composed of four "elements"—earth, air, water, and fire. Democritus (about 470–370 B.C.) thought that all forms of matter were composed of tiny indivisible particles, which he called atoms, derived from the Greek word *atomos*, meaning "indivisible." He held that atoms were in constant motion and that they combined with one another in various ways. This hypothesis was not based on scientific observations. Shortly thereafter, Aristotle (384–322 B.C.) opposed the theory of Democritus and instead endorsed and advanced the Empedoclean theory. So strong was the influence of Aristotle that his theory dominated the thinking of scientists and philosophers until the beginning of the 17th century.

5.2 Dalton's Model of the Atom

More than 2000 years after Democritus, the English schoolmaster John Dalton (1766–1844) revived the concept of atoms and proposed an atomic model based on facts and experimental evidence (Figure 5.1). His theory, described in a series of papers published from 1803 to 1810, rested on the idea of a different kind of atom for each element. The essence of **Dalton's atomic model** may be summed up as follows:

Dalton's atomic model

1. Elements are composed of minute, indivisible particles called atoms.
2. Atoms of the same element are alike in mass and size.
3. Atoms of different elements have different masses and sizes.
4. Chemical compounds are formed by the union of two or more atoms of different elements.
5. Atoms combine to form compounds in simple numerical ratios, such as one to one, one to two, two to three, and so on.
6. Atoms of two elements may combine in different ratios to form more than one compound.

Dalton's atomic model stands as a landmark in the development of chemistry. The major premises of his model are still valid, but some of his statements must be modified or qualified because later investigations have shown that (1) atoms are composed of subatomic particles; (2) not all the atoms of a specific element have the same mass; and (3) atoms, under special circumstances, can be decomposed.

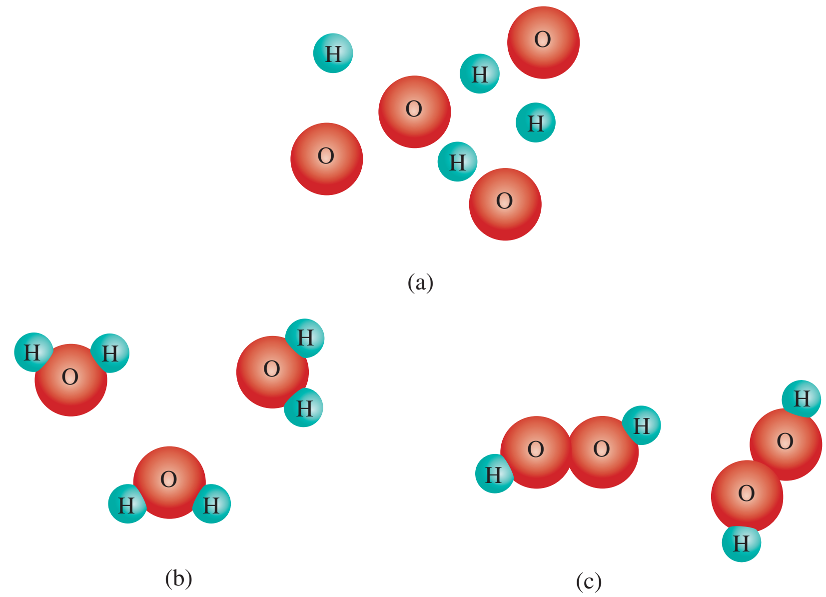


Figure 5.1
(a) Dalton's atoms were individual particles, the atoms of each element being alike in mass and size but different in mass and size from other elements. (b) and (c) Dalton's atoms combine in specific ratios to form compounds.

In chemistry we use models (theories) such as Dalton's atomic model to explain the behavior of atoms, molecules, and compounds. Models are modified to explain new information. We frequently learn the most about a system when our models (theories) fail. That is the time when we must rethink our explanation and determine whether we need to modify our model or propose a new or different model to explain the behavior.

5.3 Composition of Compounds

A large number of experiments extending over a long period have established the fact that a particular compound always contains the same elements in the same proportions by mass. For example, water always contains 11.2% hydrogen and 88.8% oxygen by mass (see Figure 5.1b). The fact that water contains hydrogen and oxygen in this particular ratio does not mean that hydrogen and oxygen cannot combine in some other ratio but rather that a compound with a different ratio would not be water. In fact, hydrogen peroxide is made up of two atoms of hydrogen and two atoms of oxygen per molecule and contains 5.9% hydrogen and 94.1% oxygen by mass; its properties are markedly different from those of water (see Figure 5.1c).

	Water	Hydrogen peroxide
	11.2% H 88.8% O	5.9% H 94.1% O
Atomic composition	2 H + 1 O	2 H + 2 O

natural law
law of definite composition

We often summarize our general observations regarding nature into a statement called a **natural law**. In the case of the composition of a compound, we use the **law of definite composition**, which states that a compound always contains two or more elements chemically combined in a definite proportion by mass.

Let's consider two elements, oxygen and hydrogen, that form more than one compound. In water, 8.0 g of oxygen are present for each gram of hydrogen. In hydrogen peroxide, 16.0 g of oxygen are present for each gram of hydrogen. The masses of oxygen are in the ratio of small whole numbers, 16:8 or 2:1. Hydrogen peroxide has twice as much oxygen (by mass) as does water. Using Dalton's atomic model, we deduce that hydrogen peroxide has twice as many oxygen atoms per hydrogen atom as water. In fact, we now write the formulas for water as H_2O and for hydrogen peroxide as H_2O_2. See Figure 5.1b and c.

The **law of multiple proportions** states atoms of two or more elements may combine in different ratios to produce more than one compound.

law of multiple proportions

Some examples of the law of multiple proportions are given in Table 5.1. The reliability of this law and the law of definite composition is the cornerstone of the science of chemistry. In essence, these laws state that (1) the composition of a particular substance will always be the same no matter what its origin or how it is formed, and (2) the composition of different compounds formed from the same elements will always be unique.

Table 5.1 Selected Compounds Showing Elements That Combine to Give More Than One Compound

Compound	Formula	Percent composition
Copper(I) chloride	$CuCl$	64.2% Cu, 35.8% Cl
Copper(II) chloride	$CuCl_2$	47.3% Cu, 52.7% Cl
Methane	CH_4	74.9% C, 25.1% H
Octane	C_8H_{18}	85.6% C, 14.4% H
Methyl alcohol	CH_4O	37.5% C, 12.6% H, 49.9% O
Ethyl alcohol	C_2H_6O	52.1% C, 13.1% H, 34.7% O
Glucose	$C_6H_{12}O_6$	40.0% C, 6.7% H, 53.3% O

You need to recognize the difference between a *law* and a *model (theory)*. A law is a summary of observed behavior. A model (theory) is an attempt to explain the observed behavior. This means that laws remain constant—that is, they do not undergo modification—while theories (models) sometimes fail and are modified or discarded over time.

5.4 The Nature of Electric Charge

You've probably received a shock after walking across a carpeted area on a dry day. You may have also experienced the static associated with combing your hair and have had your clothing cling to you. These phenomena result from an accumulation of *electric charge*. This charge may be transferred from one object to another. The properties of electric charge are as follows:

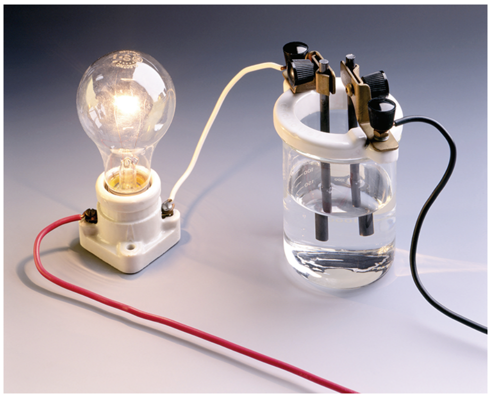
1. Charge may be of two types, positive and negative.
2. Unlike charges attract (positive attracts negative), and like charges repel (negative repels negative and positive repels positive).

3. Charge may be transferred from one object to another, by contact or induction.
4. The less the distance between two charges, the greater the force of attraction between unlike charges (or repulsion between identical charges). The force of attraction (F) can be expressed using the following equation:

$$F = \frac{kq_1q_2}{r^2}$$

where q_1 and q_2 are the charges, r is the distance between the charges, and k is a constant.

5.5 Discovery of Ions



When ions are present in a solution of salt water and an electric current is passed through the solution, the light bulb glows.

English scientist Michael Faraday (1791–1867) made the discovery that certain substances when dissolved in water conduct an electric current. He also noticed that certain compounds decompose into their elements when an electric current is passed through the compound. Atoms of some elements are attracted to the positive electrode, while atoms of other elements are attracted to the negative electrode. Faraday concluded that these atoms are electrically charged. He called them *ions* after the Greek word meaning "wanderer."

Any moving charge is an electric current. The electrical charge must travel through a substance known as a conducting medium. The most familiar conducting media are metals formed into wires.

The Swedish scientist Svante Arrhenius (1859–1927) extended Faraday's work. Arrhenius reasoned that an ion is an atom (or a group of atoms) carrying a positive or negative charge. When a compound such as sodium chloride (NaCl) is melted, it conducts electricity. Water is unnecessary. Arrhenius's explanation of this conductivity was that upon melting, the sodium chloride dissociates, or breaks up, into charged ions, Na^+ and Cl^-. The Na^+ ions move toward the negative electrode (cathode), whereas the Cl^- ions migrate toward the positive electrode (anode). Thus positive ions are called *cations*, and negative ions are called *anions*.

From Faraday's and Arrhenius's work with ions, Irish physicist G. J. Stoney (1826–1911) realized there must be some fundamental unit of electricity associated with atoms. He named this unit the electron in 1891. Unfortunately, he had no means of supporting his idea with experimental proof. Evidence remained elusive until 1897, when English physicist J. J. Thomson (1856–1940) was able to show experimentally the existence of the electron.

5.6 Subatomic Parts of the Atom

The concept of the atom—a particle so small that until recently it could not be seen even with the most powerful microscope—and the subsequent determination of its structure stand among the greatest creative intellectual human achievements.

Any visible quantity of an element contains a vast number of identical atoms. But when we refer to an atom of an element, we isolate a single atom from the multitude in order to present the element in its simplest form. Figure 5.2 shows individual atoms as we can see them today.

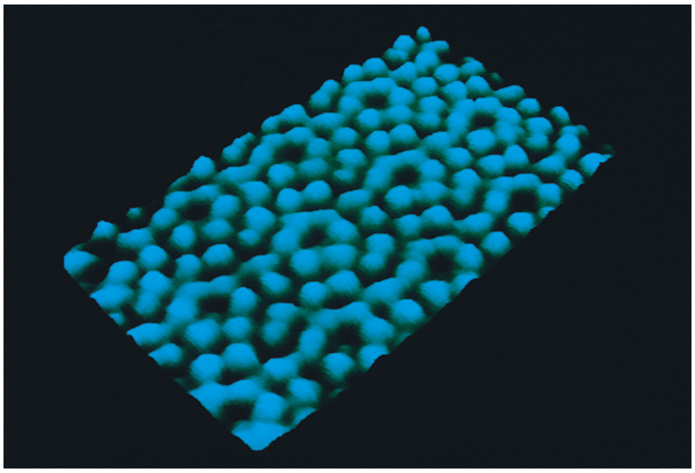


Figure 5.2
A scanning tunneling microscope shows an array of copper atoms.

What is this tiny particle we call the atom? The diameter of a single atom ranges from 0.1 to 0.5 nanometer ($1\,nm = 1 \times 10^{-9}\,m$). Hydrogen, the smallest atom, has a diameter of about 0.1 nm. To arrive at some idea of how small an atom is, consider this dot (•), which has a diameter of about 1 mm, or 1×10^{6} nm. It would take 10 million hydrogen atoms to form a line of atoms across this dot. As inconceivably small as atoms are, they contain even smaller particles, the **subatomic particles**, including electrons, protons, and neutrons.

Subatomic particles

The development of atomic theory was helped in large part by the invention of new instruments. For example, the Crookes tube, developed by Sir William Crookes (1832–1919) in 1875, opened the door to the subatomic structure of the atom (Figure 5.3). The emissions generated in a Crookes tube are called *cathode rays*. J. J. Thomson demonstrated in 1897 that cathode rays (1) travel in straight lines, (2) are negative in charge, (3) are deflected by electric and magnetic fields, (4) produce sharp shadows, and (5) are capable of moving a small paddle wheel. This was the experimental discovery of the fundamental unit of charge—the electron.

(a) (b) (c)

Figure 5.3
Crookes tube. Emissions generated in Crookes tube (a) travel in straight lines and are negative in charge, (b) are deflected by a magnetic field, and (c) provide a sharp shadow of the cross in the center of the tube.

electron

The **electron** (e^-) is a particle with a negative electrical charge and a mass of 9.110×10^{-28} g. This mass is 1/1837 the mass of a hydrogen atom. Although the actual charge of an electron is known, its value is too cumbersome for practical use and has therefore been assigned a relative electrical charge of -1. The size of an electron has not been determined exactly, but its diameter is believed to be less than 10^{-12} cm.

Protons were first observed by German physicist Eugen Goldstein (1850–1930) in 1886. However, it was Thomson who discovered the nature of the proton. He showed that the proton is a particle, and he calculated its mass to be about 1837 times that of an electron. The **proton** (p) is a particle with actual mass of 1.673×10^{-24} g. Its relative charge ($+1$) is equal in magnitude, but opposite in sign, to the charge on the electron. The mass of a proton is only very slightly less than that of a hydrogen atom.

proton

Thomson had shown that atoms contain both negatively and positively charged particles. Clearly, the Dalton model of the atom was no longer acceptable. Atoms are not indivisible but are instead composed of smaller parts. Thomson proposed a new model of the atom.

Thomson model of the atom

In the **Thomson model of the atom**, the electrons are negatively charged particles embedded in the atomic sphere. Since atoms are electrically neutral, the sphere also contains an equal number of protons, or positive charges. A neutral atom could become an ion by gaining or losing electrons.

Positive ions were explained by assuming that the neutral atom loses electrons. An atom with a net charge of $+1$ (for example, Na^+ or Li^+) has lost one electron. An atom with a net charge of $+3$ (for example, Al^{3+}) has lost three electrons (Figure 5.4a).

Negative ions were explained by assuming that additional electrons can be added to atoms. A net charge of -1 (for example, Cl^- or F^-) is produced by the addition of one electron. A net charge of -2 (for example, O^{2-} or S^{2-}) requires the addition of two electrons (Figure 5.4b).

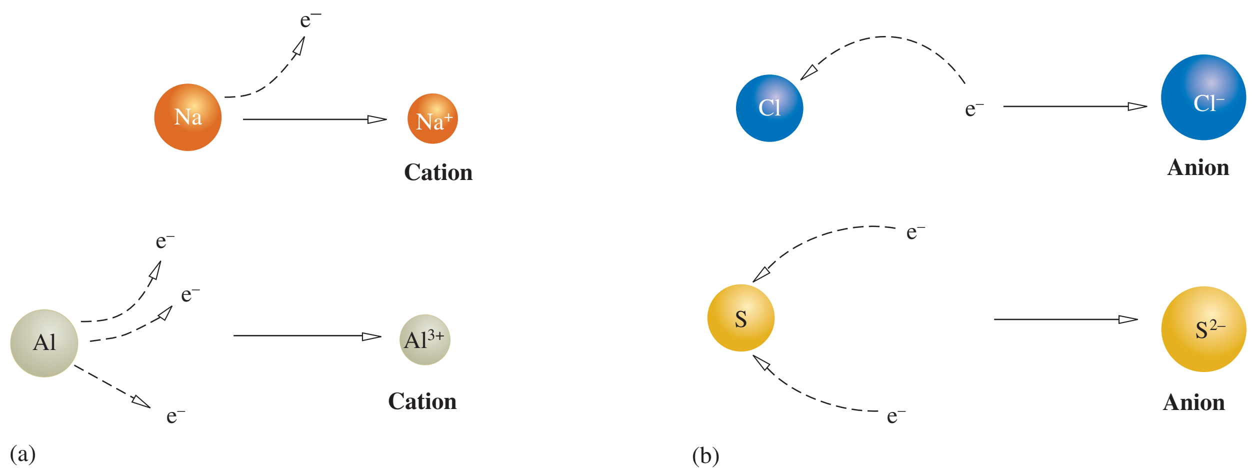


Figure 5.4
(a) When one or more electrons are lost from an atom, a cation is formed. (b) When one or more electrons are added to a neutral atom, an anion is formed.

The third major subatomic particle was discovered in 1932 by James Chadwick (1891–1974). This particle, the **neutron** (n), has neither a positive nor a negative charge and has an actual mass (1.675×10^{-24} g) which is only very slightly greater than that of a proton. The properties of these three subatomic particles are summarized in Table 5.2.

neutron

Nearly all the ordinary chemical properties of matter can be explained in terms of atoms consisting of electrons, protons, and neutrons. The discussion of atomic structure that follows is based on the assumption that atoms contain only these principal subatomic particles. Many other subatomic particles, such as mesons, positrons, neutrinos, and antiprotons, have been discovered, but it is not yet clear whether all these particles are actually present in the atom or whether they are produced by reactions occurring within the nucleus. The fields of atomic and high-energy physics have produced a long list of subatomic particles. Descriptions of the properties of many of these particles are to be found in physics textbooks.

Table 5.2 Electrical Charge and Relative Mass of Electrons, Protons, and Neutrons

Particle	Symbol	Relative electrical charge	Actual mass (g)
Electron	e^-	−1	9.110×10^{-28}
Proton	p	+1	1.673×10^{-24}
Neutron	n	0	1.675×10^{-24}

Example 5.1

The mass of a helium atom is 6.65×10^{-24} g. How many atoms are in a 4.0-g sample of helium?

SOLUTION

$$(4.0\ \text{g})\left(\frac{1\ \text{atom He}}{6.65 \times 10^{-24}\ \text{g}}\right) = 6.0 \times 10^{23}\ \text{atoms He}$$

Practice 5.1

The mass of an atom of hydrogen is 1.673×10^{-24} g. How many atoms are in a 10.0-g sample of hydrogen?

5.7 The Nuclear Atom

The discovery that positively charged particles are present in atoms came soon after the discovery of radioactivity by Henri Becquerel (1852–1908) in 1896. Radioactive elements spontaneously emit alpha particles, beta particles, and gamma rays from their nuclei (see Chapter 18).

By 1907 Ernest Rutherford (1871–1937) had established that the positively charged alpha particles emitted by certain radioactive elements are ions of the element helium. Rutherford used these alpha particles to establish the

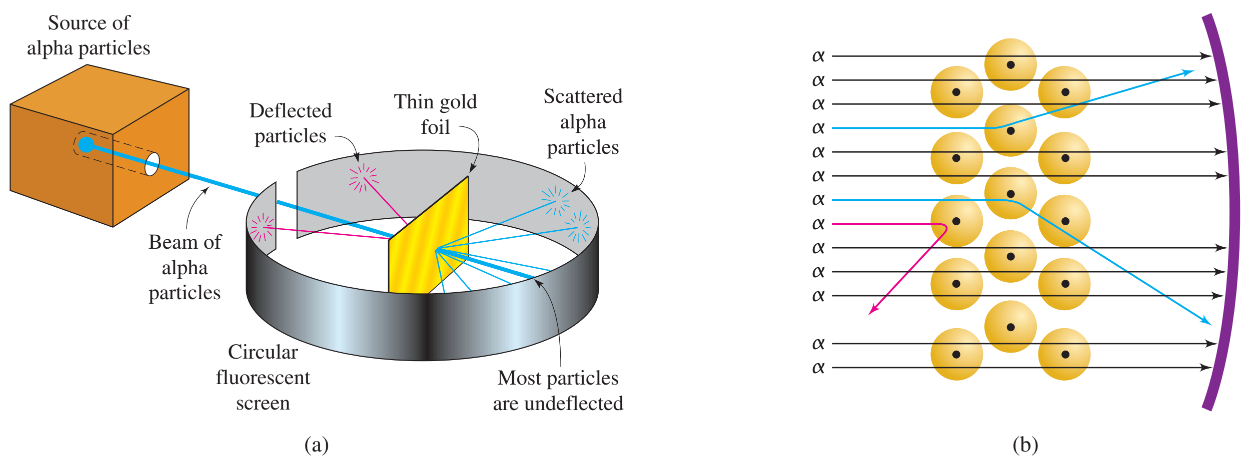


Figure 5.5
(a) Rutherford's experiment on alpha-particle scattering, where positive alpha particles (α), emanating from a radioactive source, were directed at a thin gold foil. (b) Deflection (red) and scattering (blue) of the positive alpha particles by the positive nuclei of the gold atoms.

nuclear nature of atoms. In experiments performed in 1911, he directed a stream of positively charged helium ions (alpha particles) at a very thin sheet of gold foil (about 1000 atoms thick). See Figure 5.5a. He observed that most of the alpha particles passed through the foil with little or no deflection; but a few of the particles were deflected at large angles, and occasionally one even bounced back from the foil (Figure 5.5b). It was known that like charges repel each other and that an electron with a mass of 1/1837 amu could not possibly have an appreciable effect on the path of a 4-amu alpha particle, which is about 7350 times more massive than an electron. Rutherford therefore reasoned that each gold atom must contain a positively charged mass occupying a relatively tiny volume and that, when an alpha particle approaches close enough to this positive mass, it is deflected. Rutherford spoke of this positively charged mass as the *nucleus* of the atom. Because alpha particles have relatively high masses, the extent of the deflections (some actually bounced back) indicated to Rutherford that the nucleus is very heavy and dense. (The density of the nucleus of a hydrogen atom is about 10^{12} g/cm^3—about 1 trillion times the density of water.) Because most of the alpha particles passed through the thousand or so gold atoms without any apparent deflection, he further concluded that most of an atom consists of empty space.

When we speak of the mass of an atom, we are referring primarily to the mass of the nucleus. The nucleus contains all the protons and neutrons, which represent more than 99.9% of the total mass of any atom (see Table 5.1). By way of illustration, the largest number of electrons known to exist in an atom is 111. The mass of even 111 electrons is only about 1/17 of the mass of a single proton or neutron. The mass of an atom therefore is primarily determined by the combined masses of its protons and neutrons.

General Arrangement of Subatomic Particles

The alpha-particle scattering experiments of Rutherford established that the atom contains a dense, positively charged nucleus. The later work of Chadwick demonstrated that the atom contains neutrons, which are particles with mass,

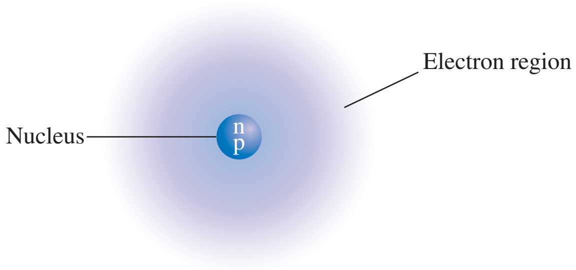


Figure 5.6
In the nuclear model of the atom, protons (p) and neutrons (n) are located in the nucleus. The electrons are found in the remainder of the atom (which is mostly empty space because electrons are very tiny).

but no charge. Rutherford also noted that light, negatively charged electrons are present and offset the positive charges in the nucleus. Based on this experimental evidence, a model of the atom and the location of its subatomic particles was devised in which each atom consists of a **nucleus** surrounded by electrons (see Figure 5.6). The nucleus contains protons and neutrons but does not contain electrons. In a neutral atom the positive charge of the nucleus (due to protons) is exactly offset by the negative electrons. Because the charge of an electron is equal to, but of opposite sign than, the charge of a proton, a neutral atom must contain exactly the same number of electrons as protons. However, this model of atomic structure provides no information on the arrangement of electrons within the atom.

nucleus

A neutral atom contains the same number of protons and electrons.

Atomic Numbers of the Elements

The **atomic number** of an element is the number of protons in the nucleus of an atom of that element. The atomic number determines the identity of an atom. For example, every atom with an atomic number of 1 is a hydrogen atom; it contains one proton in its nucleus. Every atom with an atomic number of 6 is a carbon atom; it contains 6 protons in its nucleus. Every atom with an atomic number of 92 is a uranium atom; it contains 92 protons in its nucleus. The atomic number tells us not only the number of positive charges in the nucleus but also the number of electrons in the neutral atom, since a neutral atom contains the same number of electrons and protons.

atomic number

atomic number = number of protons in the nucleus

You don't need to memorize the atomic numbers of the elements because a periodic table is usually provided in texts, in laboratories, and on examinations. The atomic numbers of all elements are shown in the periodic table on the inside front cover of this book and are also listed in the table of atomic masses on the inside front endpapers.

Carbon, shown here as a beautiful diamond, has six protons and six electrons in each atom.

5.8 Isotopes of the Elements

Shortly after Rutherford's conception of the nuclear atom, experiments were performed to determine the masses of individual atoms. These experiments showed that the masses of nearly all atoms were greater than could be accounted for by simply adding up the masses of all the protons and electrons that were known to be present in an atom. This fact led to the concept of the neutron, a particle with no charge but with a mass about the same as that of a proton. Because this particle has no charge, it was very difficult to detect, and the existence of the neutron was not proven experimentally until 1932. All atomic nuclei except that of the simplest hydrogen atom contain neutrons.

All atoms of a given element have the same number of protons. Experimental evidence has shown that, in most cases, all atoms of a given element do not have identical masses. This is because atoms of the same element may have different numbers of neutrons in their nuclei.

isotopes

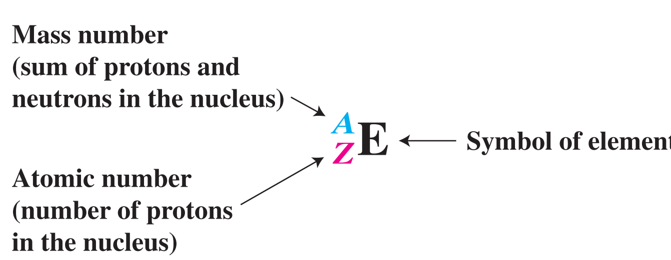
Atoms of an element having the same atomic number but different atomic masses are called **isotopes** of that element. Atoms of the various isotopes of an element therefore have the same number of protons and electrons but different numbers of neutrons.

Three isotopes of hydrogen (atomic number 1) are known. Each has one proton in the nucleus and one electron. The first isotope (protium), without a neutron, has a mass number of 1; the second isotope (deuterium), with one neutron in the nucleus, has a mass number of 2; the third isotope (tritium), with two neutrons, has a mass number of 3 (see Figure 5.7).

mass number

The three isotopes of hydrogen may be represented by the symbols ^1_1H, ^2_1H, ^3_1H, indicating an atomic number of 1 and mass numbers of 1, 2, and 3, respectively. This method of representing atoms is called *isotopic notation*. The subscript (Z) is the atomic number; the superscript (A) is the **mass number**, which is the sum of the number of protons and the number of neutrons in the nucleus. The hydrogen isotopes may also be referred to as hydrogen-1, hydrogen-2, and hydrogen-3.

The mass number of an element is the sum of the protons and neutrons in the nucleus.



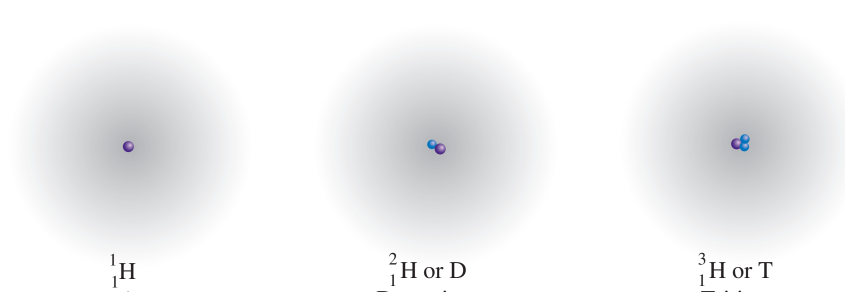
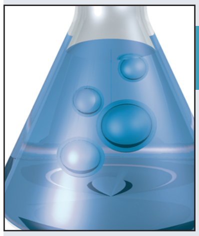


Figure 5.7
The isotopes of hydrogen. The number of protons (purple) and neutrons (blue) are shown within the nucleus. The electron (e^-) exists outside the nucleus.

CHEMISTRY IN ACTION • Isotope Detectives



Scientists are learning to use isotopes to determine the origin of drugs and gems. It turns out that isotope ratios similar to those used in carbon dating can also identify the source of cocaine or the birthplace of emeralds.

Researchers with the Drug Enforcement Agency (DEA) have created a database of the origin of coca leaves that pinpoints the origin of the leaves with a 90% accuracy. Cocaine keeps a chemical signature of the environment where it grew. Isotopes of carbon and nitrogen are found in a particular ratio based on climatic conditions in the growing region. These ratios correctly identified the source of 90% of the samples tested, according to James Ehleringer of the University of Utah, Salt Lake City. This new method can trace drugs a step further back than current techniques, which mainly look at chemicals introduced by processing practices in different locations. This could aid in tracking the original exporters and stopping production at the source.

It turns out that a similar isotopic analysis of oxygen has led researchers in France to be able to track the birthplace of emeralds. Very high quality emeralds have few inclusions (microscopic cavities). Gemologists use these inclusions and the material trapped in them to identify the source of the gem. High-quality gems can now also be identified by using an oxygen isotope ratio. These tests use an ion microscope that blasts a few atoms from the gems' surface (with virtually undetectable damage). The tiny sample is analyzed for its oxygen isotope ratio and then compared to a database from emerald mines around the world. Using the information, gemologists can determine the mine from which the emerald was born. Since emeralds from Colombian mines are valued much more highly than those from other countries, this technique can be used to help collectors know just what they are paying for, as well as to identify the history of treasured emeralds.

An uncut (left) and cut (right) emerald from Brazil.

Most of the elements occur in nature as mixtures of isotopes. However, not all isotopes are stable; some are radioactive and are continuously decomposing to form other elements. For example, of the seven known isotopes of carbon, only two, carbon-12 and carbon-13, are stable. Of the seven known isotopes of oxygen, only three—$^{16}_{8}O$, $^{17}_{8}O$, and $^{18}_{8}O$—are stable. Of the fifteen known isotopes of arsenic, $^{75}_{33}As$ is the only one that is stable.

Practice 5.2

How many protons, neutrons, and electrons are in each of the three stable isotopes of oxygen?

5.9 Atomic Mass

The mass of a single atom is far too small to measure on a balance, but fairly precise determinations of the masses of individual atoms can be made with an instrument called a *mass spectrometer* (see Figure 5.8). The mass of a single

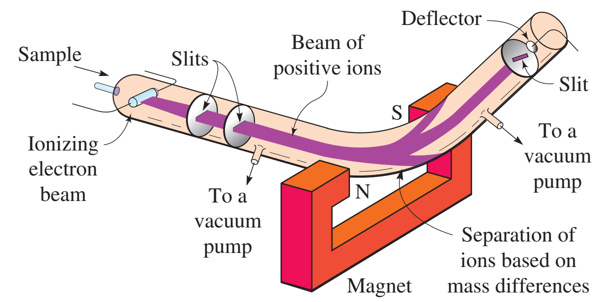


Figure 5.8
A modern mass spectrometer. A beam of positive ions is produced from the sample as it enters the chamber. The positive ions are then accelerated as they pass through slits in an electric field. When the ions enter the magnetic field, they are deflected differently, depending on mass and charge. The ions are then detected at the end of the tube. From intensity and position of the lines on the mass spectrogram, the different isotopes of the elements and their relative amounts can be determined.

hydrogen atom is 1.673×10^{-24} g. However, it is neither convenient nor practical to compare the actual masses of atoms expressed in grams; therefore, a table of relative atomic masses using *atomic mass units* was devised. (The term *atomic weight* is sometimes used instead of *atomic mass*.) The carbon isotope having six protons and six neutrons and designated carbon-12, or $^{12}_{6}C$, was chosen as the standard for atomic masses. This reference isotope was assigned a value of exactly 12 atomic mass units (amu). Thus, 1 **atomic mass unit** is defined as equal to exactly 1/12 of the mass of a carbon-12 atom. The actual mass of a carbon-12 atom is 1.9927×10^{-23} g, and that of one atomic mass unit is 1.6606×10^{-24} g. In the table of atomic masses, all elements then have values that are relative to the mass assigned to the reference isotope, carbon-12.

atomic mass unit

$1\ \text{amu} = 1.6606 \times 10^{-24}\ \text{g}$

A table of atomic masses is given on the inside front cover of this book. Hydrogen atoms, with a mass of about 1/12 that of a carbon atom, have an average atomic mass of 1.00797 amu on this relative scale. Magnesium atoms, which are about twice as heavy as carbon, have an average mass of 24.305 amu. The average atomic mass of oxygen is 15.9994 amu.

Since most elements occur as mixtures of isotopes with different masses, the atomic mass determined for an element represents the average relative mass of all the naturally occurring isotopes of that element. The atomic masses of the individual isotopes are approximately whole numbers, because the relative masses of the protons and neutrons are approximately 1.0 amu each. Yet we find that the atomic masses given for many of the elements deviate considerably from whole numbers.

The copper used in casting the Liberty bell contains a mixture of the isotopes of copper.

For example, the atomic mass of rubidium is 85.4678 amu, that of copper is 63.546 amu, and that of magnesium is 24.305 amu. The deviation of an atomic mass from a whole number is due mainly to the unequal occurrence of the various isotopes of an element. The two principal isotopes of copper are $^{63}_{29}Cu$ and $^{65}_{29}Cu$. It is apparent that copper-63 atoms are the more abundant isotope, since the atomic mass of copper, 63.546 amu, is closer to 63 than to 65 amu (see Figure 5.9). The actual values of the copper isotopes observed by mass spectra determination are shown in the following table:

Isotope	Isotopic mass (amu)	Abundance (%)	Average atomic mass (amu)
$^{63}_{29}Cu$	62.9298	69.09	63.55
$^{65}_{29}Cu$	64.9278	30.91	

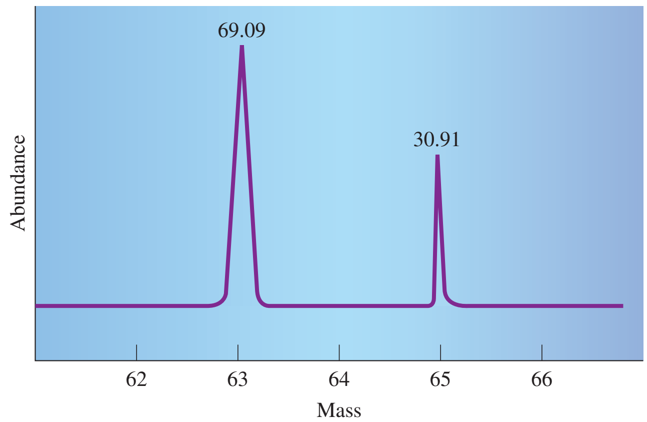


Figure 5.9
A typical reading from a mass spectrometer. The two principal isotopes of copper are shown with the abundance (%) given.

The average atomic mass can be calculated by multiplying the atomic mass of each isotope by the fraction of each isotope present and adding the results. The calculation for copper is

$$\begin{aligned}(62.9298\text{ amu})(0.6909) &= 43.48\text{ amu}\\(64.9278\text{ amu})(0.3091) &= \underline{20.07\text{ amu}}\\ &\ 63.55\text{ amu}\end{aligned}$$

The average atomic mass is a weighted average of the masses of all the isotopes present in the sample.

The **atomic mass** of an element is the average relative mass of the isotopes of that element compared to the atomic mass of carbon-12 (exactly 12.0000 . . . amu).

atomic mass

The relationship between mass number and atomic number is such that if we subtract the atomic number from the mass number of a given isotope, we obtain the number of neutrons in the nucleus of an atom of that isotope. Table 5.3 shows this method of determining the number of neutrons. For example, the fluorine atom ($^{19}_{9}F$), atomic number 9, having a mass of 19 amu, contains 10 neutrons:

$$\begin{array}{ccccc}\text{mass number} & - & \text{atomic number} & = & \text{number of neutrons}\\ 19 & - & 9 & = & 10\end{array}$$

The atomic masses given in the table on the front endpapers of this book are values accepted by international agreement. You need not memorize atomic masses. In the calculations in this book, the use of atomic masses to four significant figures will give results of sufficient accuracy. (See periodic table.)

Use four significant figures for atomic masses in this text.

Table 5.3 Determination of the Number of Neutrons in an Atom by Subtracting Atomic Number from Mass Number

	Hydrogen ($^{1}_{1}H$)	Oxygen ($^{16}_{8}O$)	Sulfur ($^{32}_{16}S$)	Fluorine ($^{19}_{9}F$)	Iron ($^{56}_{26}Fe$)
Mass number	1	16	32	19	56
Atomic number	(−)1	(−)8	(−)16	(−)9	(−)26
Number of neutrons	0	8	16	10	30

Example 5.2 How many protons, neutrons, and electrons are found in an atom of $^{14}_{6}C$?

SOLUTION The element is carbon, atomic number 6. The number of protons or electrons equals the atomic number and is 6. The number of neutrons is determined by subtracting the atomic number from the mass number: $14 - 6 = 8$.

Practice 5.3

How many protons, neutrons, and electrons are in each of these isotopes?
(a) $^{16}_{8}O$, (b) $^{80}_{35}Br$, (c) $^{235}_{92}U$, (d) $^{64}_{29}Cu$

Practice 5.4

What is the atomic number and the mass number of the elements that contain
(a) 9 electrons (b) 24 protons and 28 neutrons (c) $^{197}_{79}X$

What are the names of these elements?

Example 5.3 Chlorine is found in nature as two isotopes, $^{37}_{17}Cl$ (24.47%) and $^{35}_{17}Cl$ (75.53%). The atomic masses are 36.96590 and 34.96885 amu, respectively. Determine the average atomic mass of chlorine.

SOLUTION Multiply each mass by its percentage and add the results to find the average:

$$(0.2447)(36.96590\text{ amu}) + (0.7553)(34.96885\text{ amu})$$
$$= 35.4575\text{ amu}$$
$$= 35.46\text{ amu (4 significant figures)}$$

Practice 5.5

Silver occurs as two isotopes with atomic masses 106.9041 and 108.9047 amu, respectively. The first isotope represents 51.82% and the second 48.18%. Determine the average atomic mass of silver.

Chapter 5 Review

5.1 Early Thoughts

- Greek model of matter:
  - Four elements—earth, air, water, fire
  - Democritus—atoms (indivisible particles) make up matter
  - Aristotle—opposed atomic ideas

5.2 Dalton's Model of the Atom

KEY TERM

Dalton's atomic model

- Summary of Dalton's model of the atom:
  - Elements are composed of atoms.
  - Atoms of the same element are alike (mass and size).
  - Atoms of different elements are different in mass and size.
  - Compounds form by the union of two or more atoms of different elements.
  - Atoms form compounds in simple numerical ratios.
  - Atoms of 2 elements may combine in different ratios to form different compounds.

5.3 Composition of Compounds

KEY TERMS

Natural law
Law of definite composition
Law of multiple proportions

- The law of definite composition states that a compound always contains two or more elements combined in a definite proportion by mass.
- The law of multiple proportions states that atoms of two or more elements may combine in different ratios to form more than one compound.

5.4 The Nature of Electric Charge

- The properties of electric charge:
  - Charges are one of two types—positive or negative.
  - Unlike charges attract and like charges repel.
  - Charge is transferred from one object to another by contact or by induction.
  - The force of attraction between charges is expressed by

$$F = \frac{kq_1q_2}{r^2}.$$

5.5 Discovery of Ions

- Michael Faraday discovered electrically charged ions.
- Svante Arrhenius explained that conductivity results from the dissociation of compounds into ions:
  - Cation—positive charge—attracted to negative electrode (cathode)
  - Anion—negative charge—attracted to positive electrode (anode)

5.6 Subatomic Parts of the Atom

KEY TERMS

Subatomic particles
Electron
Proton
Thomson model of the atom
Neutron

- Atoms contain smaller subatomic particles:
  - Electron—negative charge, 1/1837 mass of proton
  - Proton—positive charge, 1.673×10^{-24} g
  - Neutron—no charge, 1.675×10^{-24} g
- Thomson model of the atom:
  - Negative electrons are embedded in a positive atomic sphere.
  - Number of protons = number of electrons in a neutral atom.
  - Ions are formed by losing or gaining electrons.

5.7 The Nuclear Atom

KEY TERMS

Nucleus
Atomic number

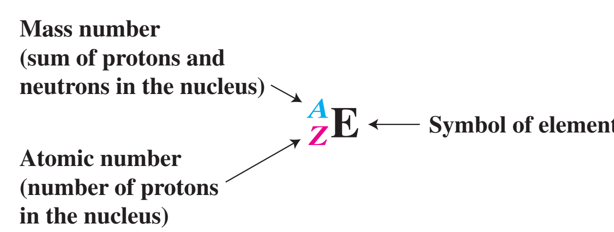
- Rutherford gold foil experiment modified the Thomson model to a nuclear model of the atom:
  - Atoms are composed of a nucleus containing protons and/or neutrons surrounded by electrons, which occupy mostly empty space.
  - Neutral atoms contain equal numbers of protons and electrons.

5.8 Isotopes of the Elements

KEY TERMS

Isotopes
Mass number

- The mass number of an element is the sum of the protons and neutrons in the nucleus.



5.9 Atomic Mass

KEY TERMS

Atomic mass unit
Atomic mass

- The average atomic mass is a weighted average of the masses of all the isotopes present in the sample.
- The number of neutrons in an atom is determined by

mass number − atomic number = number of neutrons
$A - Z =$ neutrons

Review Questions

All questions with blue numbers have answers in the appendix of the text.

1. What are the atomic numbers of (a) copper, (b) nitrogen, (c) phosphorus, (d) radium, and (e) zinc?
2. A neutron is approximately how many times heavier than an electron?
3. From the chemist's point of view, what are the essential differences among a proton, a neutron, and an electron?
4. Distinguish between an atom and an ion.
5. What letters are used to designate atomic number and mass number in isotopic notation of atoms?
6. In what ways are isotopes alike? In what ways are they different?

Paired Exercises

All exercises with blue numbers have answers in the appendix of the text.

1. In Section 5.3, there is a statement about the composition of water. It says that water (H_2O) contains 8 grams of oxygen for every 1 gram of hydrogen. Show why this statement is true.

2. In Section 5.3, there is a statement about the composition of hydrogen peroxide. It says that hydrogen peroxide (H_2O_2) contains 16 grams of oxygen for every 1 gram of hydrogen. Show why this statement is true.

3. Explain why, in Rutherford's experiments, some alpha particles were scattered at large angles by the gold foil or even bounced back.

4. What experimental evidence led Rutherford to conclude the following?
(a) The nucleus of the atom contains most of the atomic mass.
(b) The nucleus of the atom is positively charged.
(c) The atom consists of mostly empty space.

5. Describe the general arrangement of subatomic particles in the atom.

6. What part of the atom contains practically all its mass?

7. What contribution did these scientists make to atomic models of the atom?
(a) Dalton
(b) Thomson
(c) Rutherford

8. Consider the following models of the atom: (a) Dalton, (b) Thomson, (c) Rutherford. How does the location of the electrons in an atom vary? How does the location of the atom's positive matter compare?

9. Explain why the atomic masses of elements are not whole numbers.

10. Is the isotopic mass of a given isotope ever an exact whole number? Is it always? (Consider the masses of ${}^{12}_{6}C$ and ${}^{63}_{29}Cu$.)

11. What special names are given to the isotopes of hydrogen?

12. List the similarities and differences in the three isotopes of hydrogen.

13. What is the nuclear composition of the five naturally occurring isotopes of germanium having mass numbers 70, 72, 73, 74, and 76?

14. What is the nuclear composition of the five naturally occurring isotopes of zinc having mass numbers 64, 66, 67, 68, and 70?

15. Write isotopic notation symbols for each of the following:
(a) $Z = 29$, $A = 65$
(b) $Z = 20$, $A = 45$
(c) $Z = 36$, $A = 84$

16. Write isotopic notation symbols for each of the following:
(a) $Z = 47$, $A = 109$
(b) $Z = 8$, $A = 18$
(c) $Z = 26$, $A = 57$

17. Give the nuclear composition and isotopic notation for
(a) an atom containing 27 protons, 32 neutrons, and 27 electrons
(b) an atom containing 15 protons, 16 neutrons, and 15 electrons
(c) an atom containing 110 neutrons, 74 electrons, and 74 protons
(d) an atom containing 92 electrons, 143 neutrons, and 92 protons

18. Give the nuclear composition and isotopic notation for
(a) an atom containing 12 protons, 13 neutrons, and 12 electrons
(b) an atom containing 51 neutrons, 40 electrons, and 40 protons
(c) an atom containing 50 protons, 50 electrons, and 72 neutrons
(d) an atom containing 122 neutrons, 80 protons, and 80 electrons

19. An unknown element contains 24 protons, 21 electrons, and has mass number 54. Answer the following questions:
(a) What is the atomic number of this element?
(b) What is the symbol of the element?
(c) How many neutrons does it contain?

20. An unknown element contains 35 protons, 36 electrons, and has mass number 80. Answer the following questions:
(a) What is the atomic number of this element?
(b) What is the symbol of the element?
(c) How many neutrons does it contain?

21. Naturally occurring lead exists as four stable isotopes: ^{204}Pb with a mass of 203.973 amu (1.480%); ^{206}Pb, 205.974 amu (23.60%); ^{207}Pb, 206.9759 amu (22.60%); and ^{208}Pb, 207.9766 amu (52.30%). Calculate the average atomic mass of lead.

22. Naturally occurring magnesium consists of three stable isotopes: ^{24}Mg, 23.985 amu (78.99%); ^{25}Mg, 24.986 amu (10.00%); and ^{26}Mg, 25.983 amu (11.01%). Calculate the average atomic mass of magnesium.

23. 68.9257 amu is the mass of 60.4% of the atoms of an element with only two naturally occurring isotopes. The atomic mass of the other isotope is 70.9249 amu. Determine the average atomic mass of the element. Identify the element.

24. A sample of enriched lithium contains 30.00% ^{6}Li (6.015 amu) and 70.00% ^{7}Li (7.016 amu). What is the average atomic mass of the sample?

25. An average dimension for the radius of an atom is 1.0×10^{-8} cm, and the average radius of the nucleus is 1.0×10^{-13} cm. Determine the ratio of atomic volume to nuclear volume. Assume that the atom is spherical [$V = (4/3)\pi r^3$ for a sphere].

***26.** An aluminum atom has an average diameter of about 3.0×10^{-8} cm. The nucleus has a diameter of about 2.0×10^{-13} cm. Calculate the ratio of the atom's diameter to its nucleus.

Additional Exercises

All exercises with blue numbers have answers in the appendix of the text.

27. What experimental evidence supports these statements?
(a) The nucleus of an atom is small.
(b) The atom consists of both positive and negative charges.
(c) The nucleus of the atom is positive.

28. What is the relationship between the following two atoms:
(a) one atom with 10 protons, 10 neutrons, and 10 electrons; and another atom with 10 protons, 11 neutrons, and 10 electrons
(b) one atom with 10 protons, 11 neutrons, and 10 electrons; and another atom with 11 protons, 10 neutrons, and 11 electrons

29. The radius of a carbon atom in many compounds is 0.77×10^{-8} cm. If the radius of a Styrofoam ball used to represent the carbon atom in a molecular model is 1.5 cm, how much of an enlargement is this?

30. How is it possible for there to be more than one kind of atom of the same element?

31. Which element contains the largest number of neutrons per atom: ^{210}Bi, ^{210}Po, ^{210}At, or ^{211}At?

***32.** An unknown element Q has two known isotopes: ^{60}Q and ^{63}Q. If the average atomic mass is 61.5 amu, what are the relative percentages of the isotopes?

***33.** The actual mass of one atom of an unknown isotope is 2.18×10^{-22} g. Calculate the atomic mass of this isotope.

34. The mass of an atom of argon is 6.63×10^{-24} g. How many atoms are in a 40.0-g sample of argon?

35. Using the periodic table inside the front cover of the book, determine which of the first 20 elements have isotopes that you would expect to have the same number of protons, neutrons, and electrons.

36. Complete the following table with the appropriate data for each isotope given (all are neutral atoms):

Element	Symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
	^{36}Cl					
Gold			197			
		56			79	
					20	18
			58	28		

37. Complete the following table with the appropriate data for each isotope given (all are neutral atoms):

Element	Symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
	^{134}Xe					
Silver			107			
				9		
		92			143	92
			41	19		

38. Draw diagrams similar to those shown in Figure 5.4 for the following ions:
(a) F^- (b) Pb^{2+} (c) S^{2-} (d) Al^{3+}

39. Draw pictures similar to those shown in Figure 5.7 for the following isotopes of oxygen:
(a) $^{16}_{8}O$ (b) $^{17}_{8}O$ (c) $^{18}_{8}O$

40. What percent of the total mass of one atom of each of the following elements comes from electrons?
(a) iron (mass of one atom $= 9.274 \times 10^{-23}$ g)
(b) nitrogen (mass of one atom $= 2.326 \times 10^{-23}$ g)
(c) carbon (mass of one atom $= 1.994 \times 10^{-23}$ g)
(d) potassium (mass of one atom $= 6.493 \times 10^{-23}$ g)

41. What percent of the total mass of one atom of each of the following elements comes from protons?
(a) sodium (mass of one atom $= 3.818 \times 10^{-23}$ g)
(b) oxygen (mass of one atom $= 2.657 \times 10^{-23}$ g)
(c) mercury (mass of one atom $= 3.331 \times 10^{-22}$ g)
(d) fluorine (mass of one atom $= 3.155 \times 10^{-23}$ g)

42. Figure 5.6 is a representation of the nuclear model of the atom. The area surrounding the nucleus is labeled as the electron region. What is the electron region?

Challenge Exercise

All exercises with blue numbers have answers in the appendix of the text.

***43.** You have discovered a new element and are trying to determine where on the periodic table it would fit. You decide to do a mass spectrometer analysis of the sample and discover that it contains three isotopes with masses of 270.51 amu, 271.23 amu, and 269.14 amu and relative abundances of 34.07%, 55.12%, and 10.81%, respectively. Sketch the mass spectrometer reading, determine the average atomic mass of the element, estimate its atomic number, and determine its approximate location on the periodic table.

Answers to Practice Exercises

5.1 5.98×10^{24} atoms

5.2
$^{16}_{8}O$ 8p, 8e$^-$, 8n
$^{17}_{8}O$ 8p, 8e$^-$, 9n
$^{18}_{8}O$ 8p, 8e$^-$, 10n

5.3

	protons	neutrons	electrons
(a)	8	8	8
(b)	35	45	35
(c)	92	143	92
(d)	29	35	29

5.4

(a)	atomic number	9
	mass number	19
	name	potassium
(b)	atomic number	24
	mass number	52
	name	chromium
(c)	atomic number	79
	mass number	179
	name	gold

5.5 107.9 amu