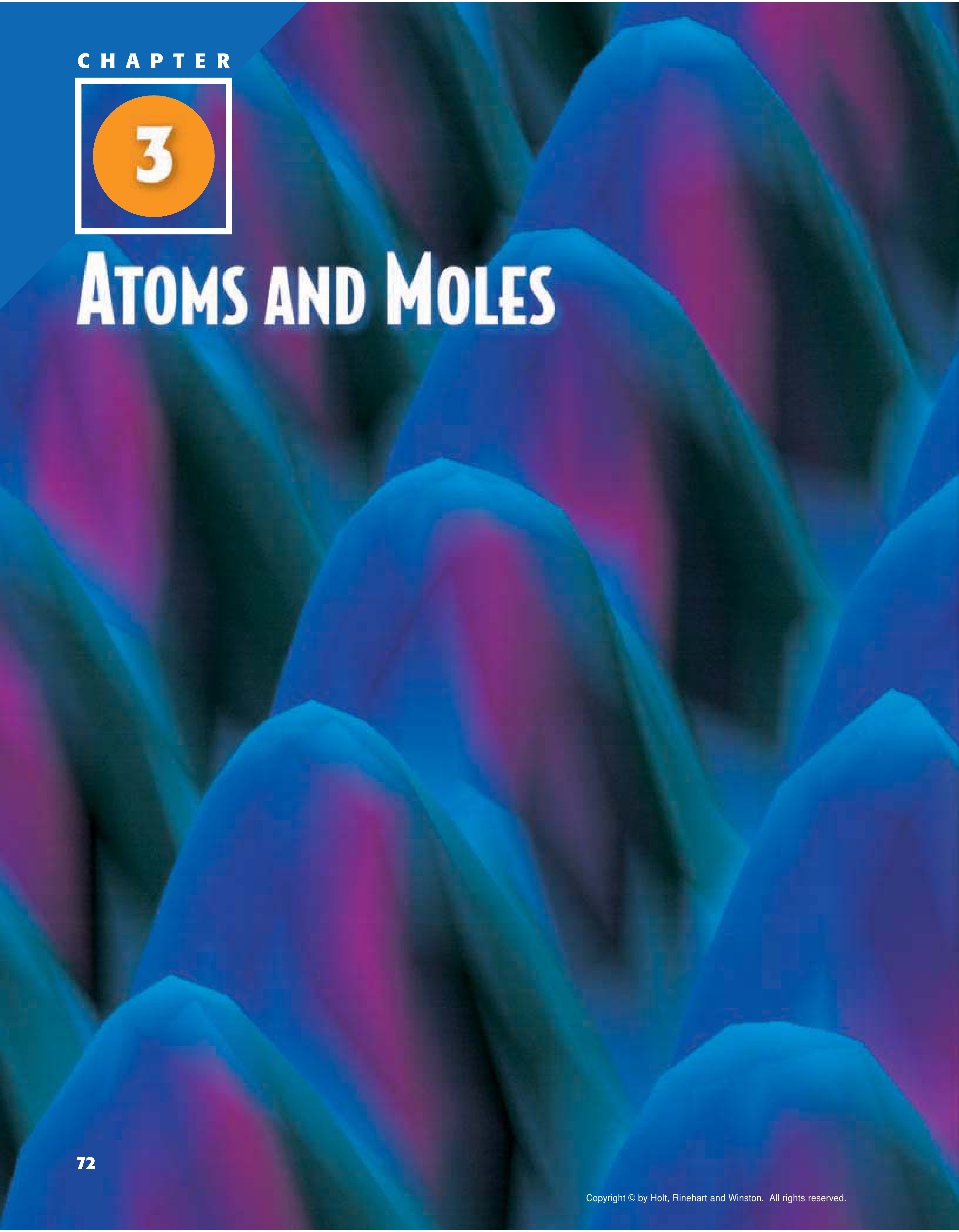
CHAPTER 3

ATOMS AND MOLES



Until recently, if you wanted to see an image of atoms, the best you could hope to see was an artists's drawing of atoms. Now, with the help of powerful microscopes, scientists are able to obtain images of atoms. One such microscope is known as the scanning tunneling microscope, which took the image of the nickel atoms shown on the opposite page. As its name implies, this microscope scans a surface, and it can come as close as a billionth of a meter to a surface to get an image. The images that these microscopes provide help scientists understand atoms.

START-UP ACTIVITY

SAFETY PRECAUTIONS

Forces of Attraction

PROCEDURE

1. Spread some **salt** and **pepper** on a piece of **paper** that lies on a flat surface. Mix the salt and pepper but make sure that the salt and pepper are not clumped together.
2. Rub a **plastic spoon** with a **wool cloth.**
3. Hold the spoon just above the salt and pepper.
4. Clean off the spoon by using a **towel.** Rub the spoon with the wool cloth and bring the spoon slowly toward the salt and pepper from a distance.

ANALYSIS

1. What happened when you held your spoon right above the salt and pepper? What happened when you brought your spoon slowly toward the salt and pepper?
2. Why did the salt and pepper jump up to the spoon?
3. When the spoon is brought toward the paper from a distance, which is the first substance to jump to the spoon? Why?

Pre-Reading Questions

1. **What is an atom?**
2. **What particles make up an atom?**
3. **Where are the particles that make up an atom located?**
4. **Name two types of electromagnetic radiation.**

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SECTION 1

Substances Are Made of Atoms

KEY TERMS

- **law of definite proportions**
- **law of conservation of mass**
- **law of multiple proportions**

OBJECTIVES

1. **State** the three laws that support the existence of atoms.
2. **List** the five principles of John Dalton's atomic theory.

Atomic Theory

As early as 400 BCE, a few people believed in an *atomic theory,* which states that atoms are the building blocks of all matter. Yet until recently, even scientists had never seen evidence of atoms. Experimental results supporting the existence of atoms did not appear until more than 2000 years after the first ideas about atoms emerged. The first of these experimental results indicated that all chemical compounds share certain characteristics.

What do you think an atom looks like? Many people think that an atom looks like the diagram in **Figure 1a.** However, after reading this chapter, you will find that the diagram in **Figure 1b** is a better model of an atom.

Recall that a compound is a pure substance composed of atoms of two or more elements that are chemically combined. These observations about compounds and the way that compounds react led to the development of the law of definite proportions, the law of conservation of mass, and the law of multiple proportions. Experimental observations show that these laws also support the current atomic theory.

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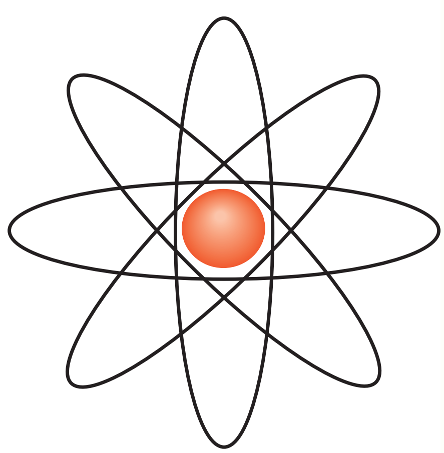


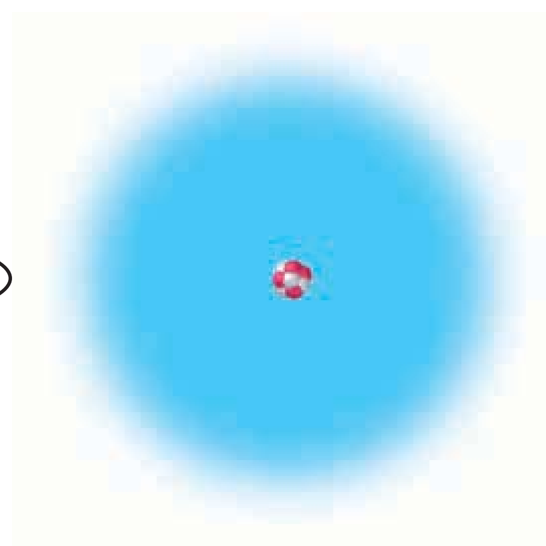


Figure 1
a Many people believe that an atom looks like this diagram.

b This diagram is a better model of the atom.

The Law of Definite Proportions

The **law of definite proportions** states that two samples of a given compound are made of the same elements in exactly the same proportions by mass regardless of the sizes or sources of the samples. Notice the composition of ethylene glycol, as shown in **Figure 2.** Every sample of ethylene glycol is composed of three elements in the following proportions by mass:

51.56% oxygen, 38.70% carbon, and 9.74% hydrogen

The law of definite proportions also states that every molecule of ethylene glycol is made of the same number and types of atoms. A molecule of ethylene glycol has the formula $C_2H_6O_2$, so the law of definite proportions tells you that all other molecules of ethylene glycol have the same formula.

Table salt (sodium chloride) is another example that shows the law of definite proportions. Any sample of table salt consists of two elements in the following proportions by mass:

60.66% chlorine and 39.34% sodium

Every sample of table salt also has the same proportions of ions. As a result, every sample of table salt has the same formula, NaCl.

As chemists of the 18th century began to gather data during their studies of matter, they first began to recognize the law of definite proportions. Their conclusions led to changes in the atomic theory.

law of definite proportions

the law that states that a chemical compound always contains the same elements in exactly the same proportions by weight or mass

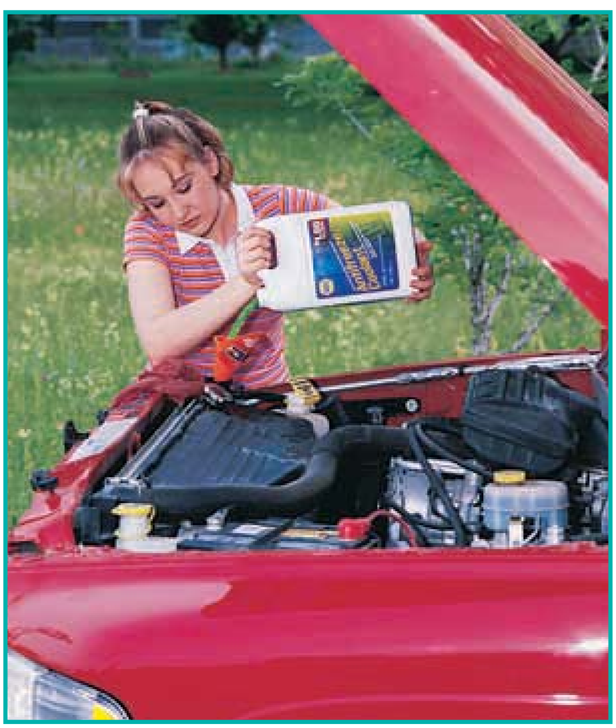
STUDY TIP

USING THE ILLUSTRATIONS

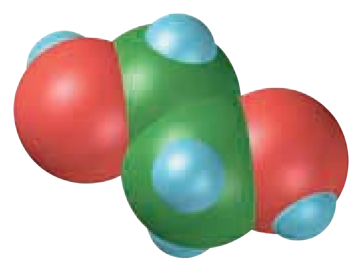
The illustrations in the text will help you make the connection between what you can see, such as a beaker of chemicals, and what you cannot see, such as the atoms that make up those chemicals. Notice that the model in **Figure 2** shows how the atoms of a molecule of ethylene glycol are arranged.

- To practice thinking at the particle level, draw pictures of water molecules and copper atoms.

Figure 2
a Ethylene glycol is the main component of automotive antifreeze.

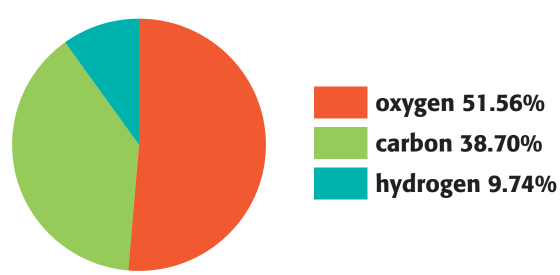


b Ethylene glycol is composed of carbon, oxygen, and hydrogen.



c Ethylene glycol is made of exact proportions of these elements regardless of the size of the sample or its source.

Ethylene Glycol Composition by Mass



The Law of Conservation of Mass

law of conservation of mass

the law that states that mass cannot be created or destroyed in ordinary chemical and physical changes

As early chemists studied more chemical reactions, they noticed another pattern. Careful measurements showed that the mass of a reacting system does not change. The **law of conservation of mass** states that the mass of the reactants in a reaction equals the mass of the products. **Figure 3** shows several reactions that show the law of conservation of mass. For example, notice the combined mass of the sulfur atom and the oxygen molecule equals the mass of the sulfur dioxide molecule.

Also notice that **Figure 3** shows that the sum of the mass of the chlorine molecule and the mass of the phosphorus trichloride molecule is slightly smaller than the mass of the phosphorus pentachloride molecule. This difference is the result of rounding off and of correctly using significant figures.

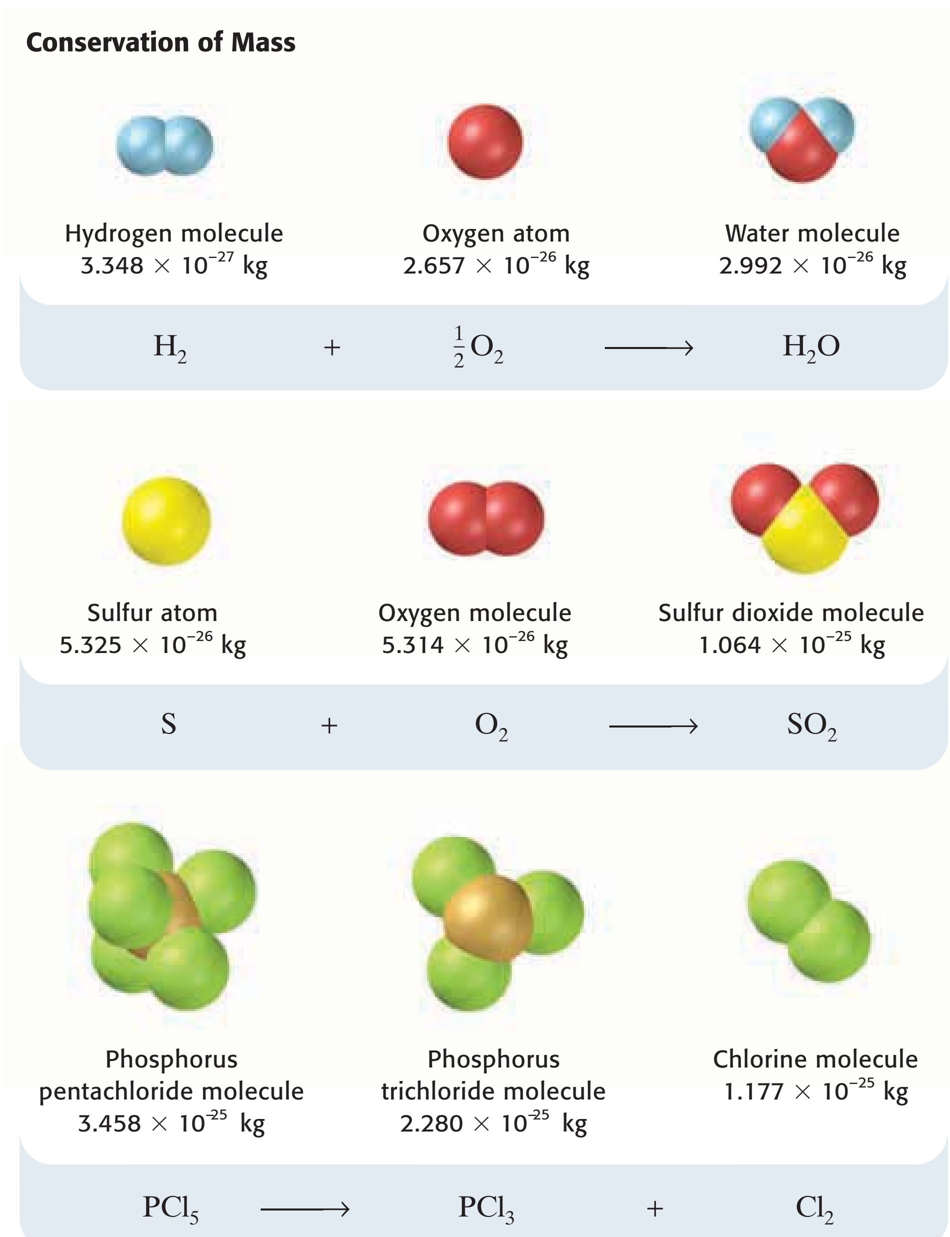
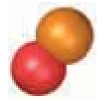


Figure 3
The total mass of a system remains the same whether atoms are combined, separated, or rearranged. Here, mass is expressed in kilograms (kg).

Table 1 Compounds of Nitrogen and Oxygen and the Law of Multiple Proportions

Name of compound	Description	As shown in figures	Formula	Mass O (g)	Mass N (g)	$\frac{\text{Mass O}(g)}{\text{Mass N}(g)}$
Nitrogen monoxide	colorless gas that reacts readily with oxygen		NO	16.00	14.01	$\frac{16.00\text{ g O}}{14.01\text{ g N}} = \frac{1.14\text{ g O}}{1\text{ g N}}$
Nitrogen dioxide	poisonous brown gas in smog		NO_2	32.00	14.01	$\frac{32.00\text{ g O}}{14.01\text{ g N}} = \frac{2.28\text{ g O}}{1\text{ g N}}$

The Law of Multiple Proportions

Table 1 lists information about the compounds nitrogen monoxide and nitrogen dioxide. For each compound, the table also lists the ratio of the mass of oxygen to the mass of nitrogen. So, 1.14 g of oxygen combine with 1 g of nitrogen when nitrogen monoxide forms. In addition, 2.28 g of oxygen combine with 1 g of nitrogen when nitrogen dioxide forms. The ratio of the masses of oxygen in these two compounds is exactly 1.14 to 2.28 or 1 to 2. This example illustrates **the law of multiple proportions:** If two or more different compounds are composed of the same two elements, the ratio of the masses of the second element (which combines with a given mass of the first element) is always a ratio of small whole numbers.

The law of multiple proportions may seem like an obvious conclusion given the molecules' diagrams and formulas shown. But remember that the early chemists did not know the formulas for compounds. In fact, chemists still have not actually seen these molecules. Scientists think that molecules have these formulas because of these mass data.

law of multiple proportions

the law that states that when two elements combine to form two or more compounds, the mass of one element that combines with a given mass of the other is in the ratio of small whole numbers

Dalton's Atomic Theory

In 1808, John Dalton, an English school teacher, used the Greek concept of the atom and the law of definite proportions, the law of conservation of mass, and the law of multiple proportions to develop an atomic theory. Dalton believed that a few kinds of atoms made up all matter.

According to Dalton, elements are composed of only one kind of atom and compounds are made from two or more kinds of atoms. For example, the element copper consists of only one kind of atom, as shown in **Figure 4.** Notice that the compound iodine monochloride consists of two kinds of atoms joined together. Dalton also reasoned that only whole numbers of atoms could combine to form compounds, such as iodine monochloride. In this way, Dalton revised the early Greek idea of atoms into a scientific theory that could be tested by experiments.

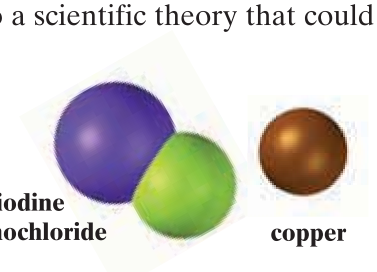


Figure 4
An element, such as copper, is made of only one kind of atom. In contrast, a compound, such as iodine monochloride, can be made of two or more kinds of atoms.

Dalton's Theory Contains Five Principles

Dalton's atomic theory can be summarized by the following statements:

1. All matter is composed of extremely small particles called *atoms,* which cannot be subdivided, created, or destroyed.
2. Atoms of a given element are identical in their physical and chemical properties.
3. Atoms of different elements differ in their physical and chemical properties.
4. Atoms of different elements combine in simple, whole-number ratios to form compounds.
5. In chemical reactions, atoms are combined, separated, or rearranged but never created, destroyed, or changed.

Dalton's theory explained most of the chemical data that existed during his time. As you will learn later in this chapter, data gathered since Dalton's time shows that the first two principles are not true in all cases.

Today, scientists can divide an atom into even smaller particles. Technology has also enabled scientists to destroy and create atoms. Another feature of atoms that Dalton could not detect is that many atoms will combine with like atoms. Oxygen, for example, is generally found as O_2, a molecule made of two oxygen atoms. Sulfur is found as S_8. Because some parts of Dalton's theory have been shown to be incorrect, his theory has been modified and expanded as scientists learn more about atoms.

1 Section Review

UNDERSTANDING KEY IDEAS

1. What is the atomic theory?
2. What is a compound?
3. State the laws of definite proportions, conservation of mass, and multiple proportions.
4. According to Dalton, what is the difference between an element and a compound?
5. What are the five principles of Dalton's atomic theory?
6. Which of Dalton's five principles still apply to the structure of an atom?

CRITICAL THINKING

7. What law is described by the fact that carbon dioxide consists of 27.3% carbon and 72.7% oxygen by mass?
8. What law is described by the fact that the ratio of the mass of oxygen in carbon dioxide to the mass of oxygen in carbon monoxide is 2:1?
9. Three compounds contain the elements sulfur, S, and fluorine, F. How do the following data support the law of multiple proportions?

   compound A: 1.188 g F for every 1.000 g S

   compound B: 2.375 g F for every 1.000 g S

   compound C: 3.563 g F for every 1.000 g S

SECTION

2 Structure of Atoms

KEY TERMS

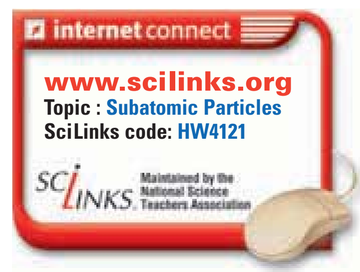
- electron
- nucleus
- proton
- neutron
- atomic number
- mass number
- isotope

OBJECTIVES

1. **Describe** the evidence for the existence of electrons, protons, and neutrons, and describe the properties of these subatomic particles.
2. **Discuss** atoms of different elements in terms of their numbers of electrons, protons, and neutrons, and define the terms *atomic number* and *mass number.*
3. **Define** *isotope,* and determine the number of particles in the nucleus of an isotope.

Subatomic Particles

Experiments by several scientists in the mid-1800s led to the first change to Dalton's atomic theory. Scientists discovered that atoms can be broken into pieces after all. These smaller parts that make up atoms are called *subatomic particles.* Many types of subatomic particles have since been discovered. The three particles that are most important for chemistry are the electron, the proton, and the neutron.



Electrons Were Discovered by Using Cathode Rays

The first evidence that atoms had smaller parts was found by researchers who were studying electricity, not atomic structure. One of these scientists was the English physicist J. J. Thomson. To study current, Thomson pumped most of the air out of a glass tube. He then applied a voltage to two metal plates, called *electrodes,* which were placed at either end of the tube. One electrode, called the *anode,* was attached to the positive terminal of the voltage source, so it had a positive charge. The other electrode, called a *cathode,* had a negative charge because it was attached to the negative terminal of the voltage source.

Thomson observed a glowing beam that came out of the cathode and struck the anode and the nearby glass walls of the tube. So, he called these rays *cathode rays.* The glass tube Thomson used is known as a *cathode-ray tube* (CRT). CRTs have become an important part of everyday life. They are used in television sets, computer monitors, and radar displays.

Figure 5
The image on a television screen or a computer monitor is produced when cathode rays strike the special coating on the inside of the screen.

An Electron Has a Negative Charge

Thomson knew the rays must have come from the atoms of the cathode because most of the atoms in the air had been pumped out of the tube. Because the cathode ray came from the negatively charged cathode, Thomson reasoned that the ray was negatively charged.

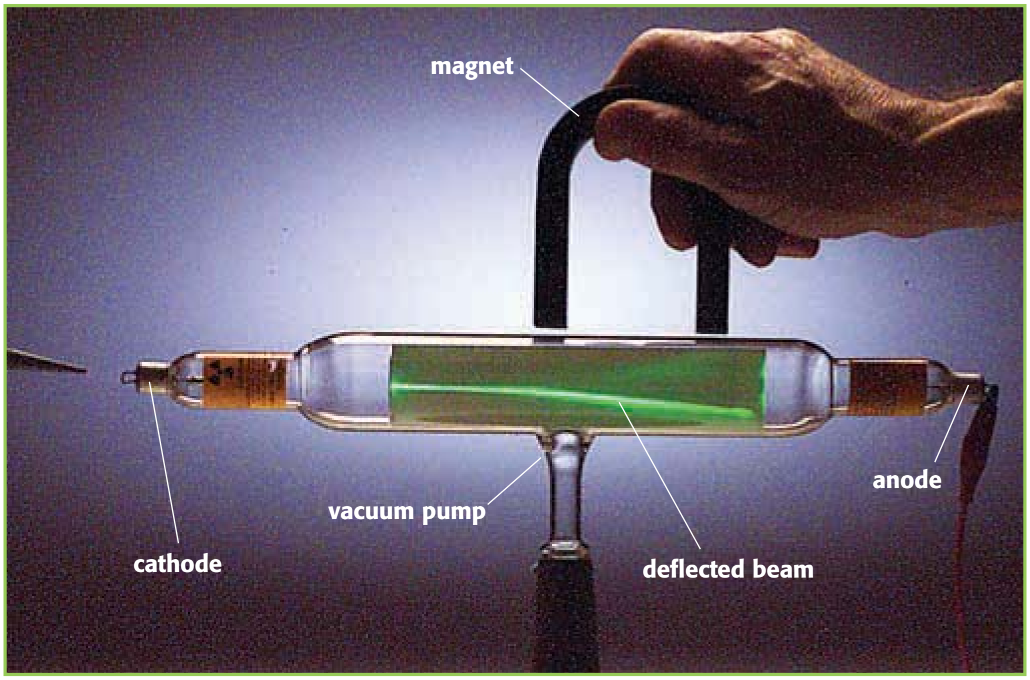


Figure 6
A magnet near the cathode-ray tube causes the beam to be deflected. The deflection indicates that the particles in the beam have a negative charge.

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He confirmed this prediction by seeing how electric and magnetic fields affected the cathode ray. **Figure 6** shows what Thomson saw when he placed a magnet near the tube. Notice that the beam is deflected by the magnet. Other researchers had shown that moving negative charges are deflected this way.

Thomson also observed that when a small paddle wheel was placed in the path of the rays, the wheel would turn. This observation suggested that the cathode rays consisted of tiny particles that were hitting the paddles of the wheel.

Thomson's experiments showed that a cathode ray consists of particles that have mass and a negative charge. These particles are called **electrons.** **Table 2** lists the properties of an electron. Later experiments, which used different metals for cathodes, confirmed that electrons are a part of atoms of all elements.

electron
a subatomic particle that has a negative electric charge

Electrons are negatively charged, but atoms have no charge. Therefore, atoms must contain some positive charges that balance the negative charges of the electrons. Scientists realized that positive charges must exist in atoms and began to look for more subatomic particles. Scientists also recognized that atoms must have other particles because an electron was found to have much less mass than an atom does.

Table 2 Properties of an Electron

Name	Symbol	As shown in figures	Charge	Common charge notation	Mass (kg)
Electron	e, e^-, or ${}^{\;\,0}_{-1}e$		-1.602×10^{-19} C	-1	9.109×10^{-31} kg

Rutherford Discovered the Nucleus

Thomson proposed that the electrons of an atom were embedded in a positively charged ball of matter. His picture of an atom, which is shown in **Figure 7,** was named the *plum-pudding model* because it resembled plum pudding, a dessert consisting of a ball of cake with pieces of fruit in it. Ernest Rutherford, one of Thomson's former students, performed experiments in 1909 that disproved the plum-pudding model of the atom.

Rutherford's team of researchers carried out the experiment shown in **Figure 8.** A beam of small, positively charged particles, called *alpha particles*, was directed at a thin gold foil. The team measured the angles at which the particles were deflected from their former straight-line paths as they came out of the foil.

Rutherford found that most of the alpha particles shot at the foil passed straight through the foil. But a very small number of particles were deflected, in some cases backward, as shown in **Figure 8.** This result greatly surprised the researchers—it was very different from what Thomson's model predicted. As Rutherford said, "It was almost as if you fired a 15-inch shell into a piece of tissue paper and it came back and hit you." After thinking about the startling result for two years, Rutherford finally came up with an explanation. He went on to reason that only a very concentrated positive charge in a tiny space within the gold atom could possibly repel the fast-moving, positively charged alpha particles enough to reverse the alpha particles' direction of travel.

Rutherford also hypothesized that the mass of this positive-charge containing region, called the *nucleus,* must be larger than the mass of the alpha particle. If not, the incoming particle would have knocked the positive charge out of the way. The reason that most of the alpha particles were undeflected, Rutherford argued, was that most parts of the atoms in the gold foil were empty space.

This part of the model of the atom is still considered true today. The **nucleus** is the dense, central portion of the atom. The nucleus has all of the positive charge, nearly all of the mass, but only a very small fraction of the volume of the atom.

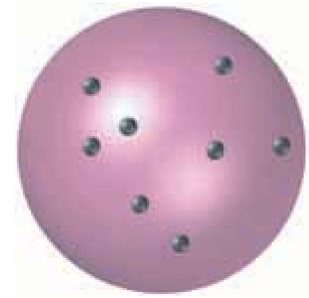
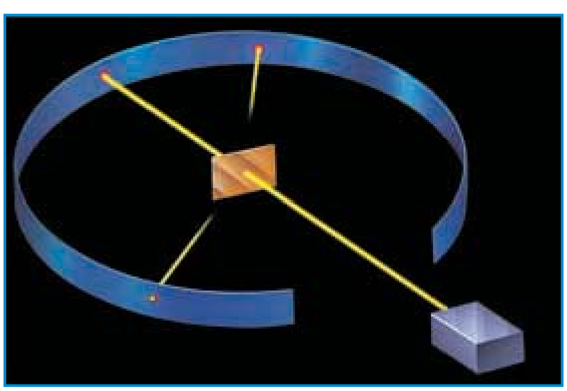


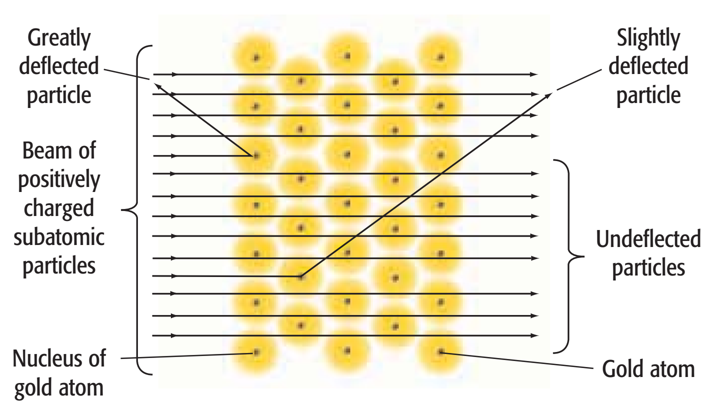
Figure 7
Thomson's model of an atom had negatively charged electrons embedded in a ball of positive charge.

nucleus
an atom's central region, which is made up of protons and neutrons

Figure 8



a In the gold foil experiment, small positively charged particles were directed at a thin foil of gold atoms.



b The pattern of deflected alpha particles supported Rutherford's hypothesis that gold atoms were mostly empty space.

Figure 9
If the nucleus of an atom were the size of a marble, then the whole atom would be about the size of a football stadium.

Protons and Neutrons Compose the Nucleus

By measuring the numbers of alpha particles that were deflected and the angles of deflection, scientists calculated the radius of the nucleus to be less than $\frac{1}{10\,000}$ of the radius of the whole atom. **Figure 9** gives you a better idea of these sizes. Even though the radius of an entire atom is more than 10 000 times larger than the radius of its nucleus, an atom is still extremely small. The unit used to express atomic radius is the picometer (pm). One picometer equals 10^{-12} m.

The positively charged particles that repelled the alpha particles in the gold foil experiments and that compose the nucleus of an atom are called **protons.** The charge of a proton was calculated to be exactly equal in magnitude but opposite in sign to the charge of an electron. Later experiments showed that the proton's mass is almost 2000 times the mass of an electron.

proton

a subatomic particle that has a positive charge and that is found in the nucleus of an atom; the number of protons of the nucleus is the atomic number, which determines the identity of an element

Because protons and electrons have equal but opposite charges, a neutral atom must contain equal numbers of protons and electrons. But solving this mystery led to another: the mass of an atom (except hydrogen atoms) is known to be greater than the combined masses of the atom's protons and electrons. What could account for the rest of the mass? Hoping to find an answer, scientists began to search for a third subatomic particle.

About 30 years after the discovery of the electron, Irene Joliot-Curie (the daughter of the famous scientists Marie and Pierre Curie) discovered that when alpha particles hit a sample of beryllium, a beam that could go through almost anything was produced.

The British scientist James Chadwick found that this beam was not deflected by electric or magnetic fields. He concluded that the particles carried no electric charge. Further investigation showed that these neutral particles, which were named **neutrons,** are part of all atomic nuclei (except the nuclei of most hydrogen atoms).

neutron

a subatomic particle that has no charge and that is found in the nucleus of an atom

Table 3 **Properties of a Proton and a Neutron**

Name	Symbol	As shown in figures	Charge	Common charge notation	Mass (kg)
Proton	p, p^+, or ${}^{1}_{+1}p$		$+1.602 \times 10^{-19}$ C	+1	1.673×10^{-27} kg
Neutron	n or ${}^{1}_{0}n$		0 C	0	1.675×10^{-27} kg

Protons and Neutrons Can Form a Stable Nucleus

Table 3 lists the properties of a neutron and a proton. Notice that the charge of a neutron is commonly assigned the value 0 while that of a proton is +1. How do protons that are positively charged come together to form a nucleus? In fact, the formation of a nucleus with protons seems impossible if you just consider *Coulomb's law.* Coulomb's law states that the closer two charges are, the greater the force between them. In fact, the force increases by a factor of 4 as the distance is halved. In addition, the larger the two charges are the greater the force between them. If the charges are opposite, they attract one another. If both charges have the same sign, they repel one another.

If you keep Coulomb's law in mind, it is easy to understand why—with the exception of some hydrogen atoms—no atoms have nuclei that are composed of only protons. All protons have a +1 charge. So, the repulsive force between two protons is large when two protons are close together, such as within a nucleus.

Protons, however, do form stable nuclei despite the repulsive force between them. A strong attractive force between these protons overcomes the repulsive force at small distances. Because neutrons also add attractive forces without being subject to repulsive charge-based forces, some neutrons can help stabilize a nucleus. Thus, all atoms that have more than one proton also have neutrons.

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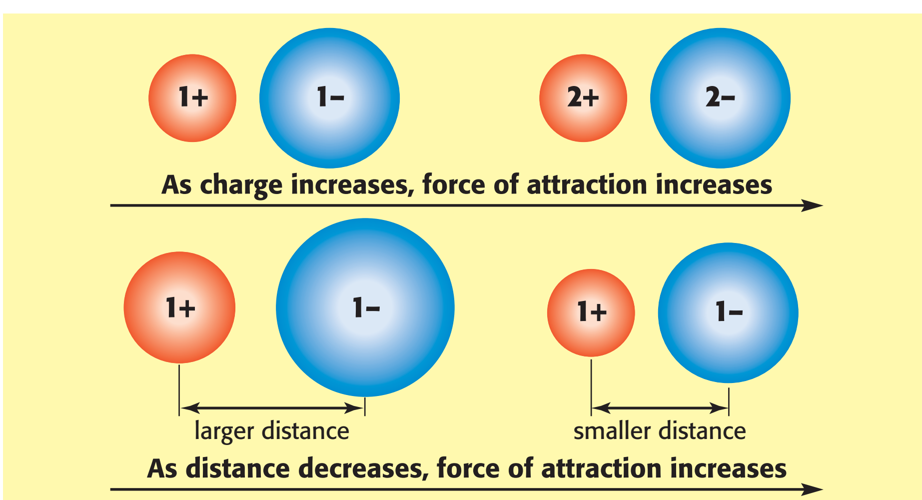


Figure 10
This figure shows that the larger two charges are, the greater the force between the charges. In addition, the figure shows the smaller the distance between two charges, the greater the force between the charges.

Atomic Number and Mass Number

All atoms consist of protons and electrons. Most atoms also have neutrons. Protons and neutrons make up the small, dense nuclei of atoms. The electrons occupy the space surrounding the nucleus. For example, an oxygen atom has protons and neutrons surrounded by electrons. But that description fits all other atoms, such as atoms of carbon, nitrogen, silver, and gold. How, then, do the atoms of one element differ from those of another element? Elements differ from each other in the number of protons their atoms contain.

Atomic Number Is the Number of Protons of the Nucleus

atomic number

the number of protons in the nucleus of an atom; the atomic number is the same for all atoms of an element

The number of protons that an atom has is known as the atom's **atomic number.** For example, the atomic number of hydrogen is 1 because the nucleus of each hydrogen atom has one proton. The atomic number of oxygen is 8 because all oxygen atoms have eight protons. Because each element has a unique number of protons in its atoms, no two elements have the same atomic number. So an atom whose atomic number is 8 must be an oxygen atom.

To date, scientists have identified 113 elements, whose atomic numbers range from 1 to 114. The element whose atomic number is 113 has yet to be discovered. Note that atomic numbers are always whole numbers. For example, an atom cannot have 2.5 protons.

The atomic number also reveals the number of electrons in an atom of an element. For atoms to be neutral, the number of negatively charged electrons must equal the number of positively charged protons. Therefore, if you know the atomic number of an atom, you immediately know the number of protons and the number of electrons found in that atom. **Figure 11** shows a model of an oxygen atom, whose atomic number is 8 and which has 8 electrons surrounding a nucleus that has 8 protons. The atomic number of gold is 79, so an atom of gold must have 79 electrons surrounding a nucleus of 79 protons. The next step in describing an atom's structure is to find out how many neutrons the atom has.

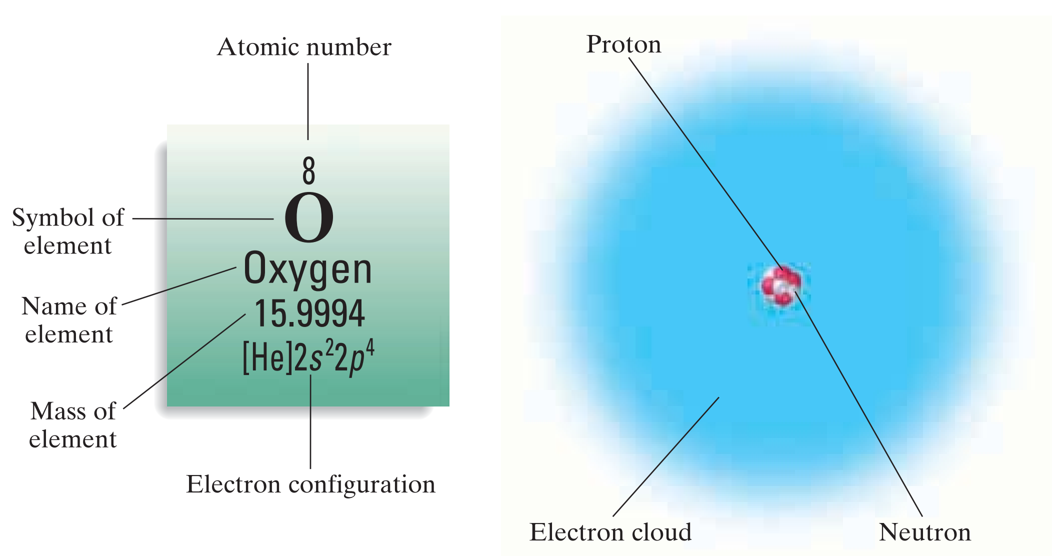


Figure 11
The atomic number for oxygen, as shown on the periodic table, tells you that the oxygen atom has 8 protons and 8 electrons.

Mass Number Is the Number of Particles of the Nucleus

Every atomic nucleus can be described not only by its atomic number but also by its mass number. The **mass number** is equal to the total number of particles of the nucleus—that is, the total number of protons and neutrons. For example, a particular atom of neon has a mass number of 20, as shown in **Figure 12.** Therefore, the nucleus of this atom has a total of 20 protons and neutrons. Because the atomic number for an atom of neon is 10, neon has 10 protons. You can calculate the number of neutrons in a neon atom by subtracting neon's atomic number (the number of protons) from neon's mass number (the number of protons and neutrons).

mass number
the sum of the numbers of protons and neutrons of the nucleus of an atom

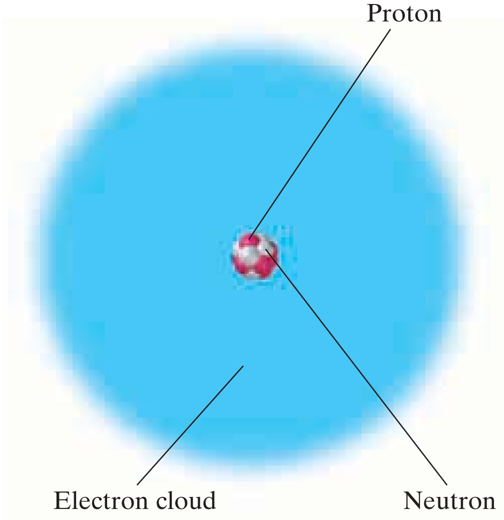
$$\textit{mass number} - \textit{atomic number} = \textit{number of neutrons}$$

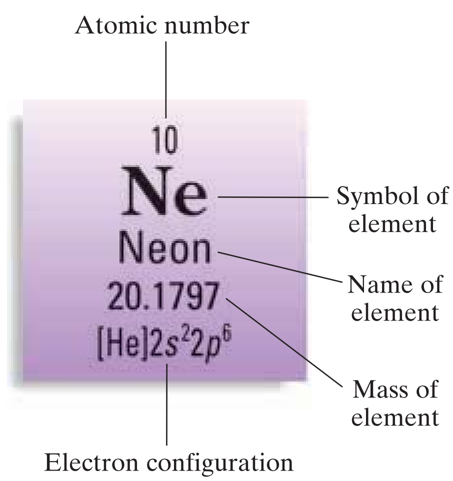
In this example, the neon atom has 10 neutrons.

$$\begin{array}{rr} \textit{number of protons and neutrons (mass number)} = & 20 \\ - \;\textit{number of protons (atomic number)} = & 10 \\ \hline \textit{number of neutrons} = & 10 \end{array}$$

Unlike the atomic number, which is the same for all atoms of an element, mass number can vary among atoms of a single element. In other words, all atoms of an element have the same number of protons, but they can have different numbers of neutrons. The atomic number of every hydrogen atom is 1, but hydrogen atoms can have mass numbers of 1, 2, or 3. These atoms differ from one another in having 0, 1, and 2 neutrons, respectively. Another example is oxygen. The atomic number of every oxygen atom is 8, but oxygen atoms can have mass numbers of 16, 17, or 18. These atoms differ from one another in having 8, 9, and 10 neutrons, respectively.

Figure 12
The neon atom has 10 protons, 10 neutrons, and 10 electrons. This atom's mass number is 20, or the sum of the numbers of protons and neutrons in the atom.





SAMPLE PROBLEM A

Determining the Number of Particles in an Atom

How many protons, electrons, and neutrons are present in an atom of copper whose atomic number is 29 and whose mass number is 64?

1 Gather information.

- The atomic number of copper is 29.
- The mass number of copper is 64.

2 Plan your work.

- The atomic number indicates the number of protons in the nucleus of a copper atom.
- A copper atom must be electrically neutral, so the number of electrons equals the number of protons.
- The mass number indicates the total number of protons and neutrons in the nucleus of a copper atom.

3 Calculate.

- *atomic number* (29) = *number of protons* = 29
- *number of protons* = *number of electrons* = 29
- *mass number* (64) − *atomic number* (29) = *number of neutrons* = 35

4 Verify your results.

- *number of protons* (29) + *number of neutrons* (35) = *mass number* (64)

PRACTICE HINT

Check that the atomic number and the number of protons are the same. Also check that adding the numbers of protons and neutrons equals the mass number.

PRACTICE

PROBLEM SOLVING SKILL

1. How many protons and electrons are in an atom of sodium whose atomic number is 11?
2. An atom has 13 protons and 14 neutrons. What is its mass number?
3. Calculate the mass number for an atom that has 45 neutrons and 35 electrons.
4. An atom of an element has 54 protons. Some of the element's atoms have 77 neutrons, while other atoms have 79 neutrons. What are the atomic numbers and mass numbers of the two types of atoms of this element?

Different Elements Can Have the Same Mass Number

The atomic number identifies an element. For example, copper has the atomic number 29. All copper atoms have nuclei that have 29 protons. Each of these atoms also has 29 electrons. Any atom that has 29 protons must be a copper atom.

In contrast, knowing just the mass number does not help you identify the element. For example, some copper atom nuclei have 36 neutrons. These copper atoms have a mass number of 65. But zinc atoms that have 30 protons and 35 neutrons also have mass numbers of 65.

Atomic Structures Can Be Represented by Symbols

Each element has a name, and the same name is given to all atoms of an element. For example, sulfur is composed of sulfur atoms. Recall that each element also has a symbol, and the same symbol is used to represent one of the element's atoms. Thus, S represents a single atom of sulfur, 2S represents two sulfur atoms, and 8S represents eight sulfur atoms. However, chemists write S_8 to indicate that the eight sulfur atoms are joined together and form a molecule of sulfur, as shown in the model in **Figure 13.**

Atomic number and mass number are sometimes written with an element's symbol. The atomic number always appears on the lower left side of the symbol. For example, the symbols for the first five elements are written with atomic numbers as follows:

$$_{1}\mathrm{H} \quad _{2}\mathrm{He} \quad _{3}\mathrm{Li} \quad _{4}\mathrm{Be} \quad _{5}\mathrm{B}$$

Note that these subscript numbers give no new information. They simply indicate the atomic number of a particular element. On the other hand, mass numbers provide information that specifies particular atoms of an element. Mass numbers are written on the upper left side of the symbol. The following are the symbols of stable atoms of the first five elements written with mass numbers:

$$^{1}\mathrm{H} \quad ^{2}\mathrm{H} \quad ^{3}\mathrm{He} \quad ^{4}\mathrm{He} \quad ^{6}\mathrm{Li} \quad ^{7}\mathrm{Li} \quad ^{9}\mathrm{Be} \quad ^{10}\mathrm{B} \quad ^{11}\mathrm{B}$$

Both numbers may be written with the symbol. For example, the most abundant kind of each of the first five elements can be represented by the following symbols:

$$^{1}_{1}\mathrm{H} \quad ^{4}_{2}\mathrm{He} \quad ^{7}_{3}\mathrm{Li} \quad ^{9}_{4}\mathrm{Be} \quad ^{11}_{5}\mathrm{B}$$

An element may be represented by more than one notation. For example, the following notations represent the different atoms of hydrogen:

$$^{1}_{1}\mathrm{H} \quad ^{2}_{1}\mathrm{H} \quad ^{3}_{1}\mathrm{H}$$

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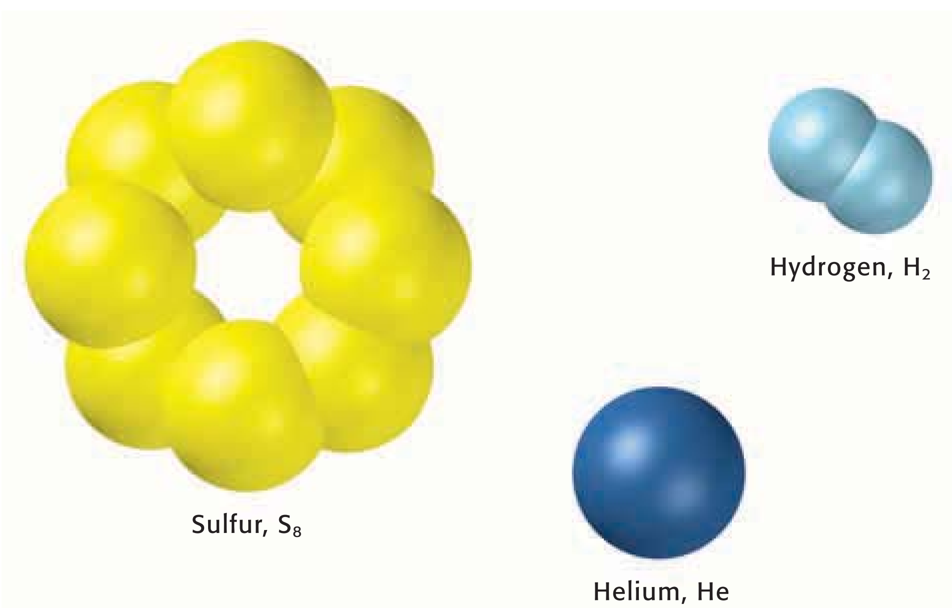
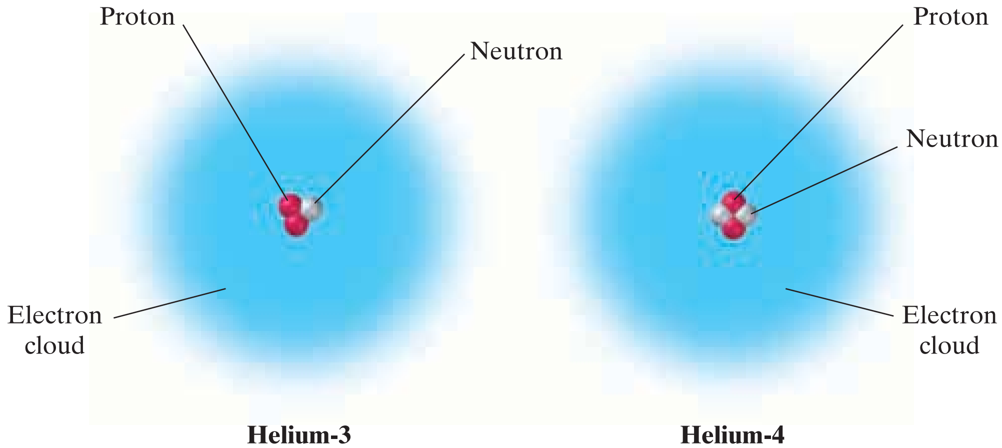


Figure 13
In nature, elemental sulfur exists as eight sulfur atoms joined in a ring, elemental hydrogen exists as a molecule of two hydrogen atoms, and elemental helium exists as single helium atoms.

Figure 14
The two stable isotopes of helium are helium-3 and helium-4. The nucleus of a helium-4 atom is known as an *alpha particle*.



isotope
an atom that has the same number of protons (atomic number) as other atoms of the same element but has a different number of neutrons (atomic mass)

Isotopes of an Element Have the Same Atomic Number

All atoms of an element have the same atomic number and the same number of protons. However, atoms do not necessarily have the same number of neutrons. Atoms of the same element that have different numbers of neutrons are called **isotopes.** The two atoms modeled in **Figure 14** are stable isotopes of helium.

There are two standard methods of identifying isotopes. One method is to write the mass number with a hyphen after the name of an element. For example, the helium isotope shown on the left in **Figure 14** is written helium-3, while the isotope shown on the right is written as helium-4. The second method shows the composition of a nucleus as the isotope's nuclear symbol. Using this method, the notations for the two helium isotopes shown in **Figure 14** are written below.

$$^{3}_{2}He \quad ^{4}_{2}He$$

Notice that all isotopes of an element have the same atomic number. However, their atomic masses are not the same because the number of neutrons of the atomic nucleus of each isotope varies. In the case of helium, both isotopes have two protons in their nuclei. However, helium-3 has one neutron, while helium-4 has two neutrons.

Table 4 lists the four stable isotopes of lead. The least abundant of these isotopes is lead-204, while the most common is lead-208. Why do all lead atoms have 82 protons and 82 electrons?

Table 4 The Stable Isotopes of Lead

Name of atom	Symbol	Number of neutrons	Mass number	Mass (kg)	Abundance (%)
Lead-204	$^{204}_{82}Pb$	122	204	203.973	1.4
Lead-206	$^{206}_{82}Pb$	124	206	205.974	24.1
Lead-207	$^{207}_{82}Pb$	125	207	206.976	22.1
Lead-208	$^{208}_{82}Pb$	126	208	207.977	52.4

SAMPLE PROBLEM B

Determining the Number of Particles in Isotopes

Calculate the numbers of protons, electrons, and neutrons in oxygen-17 and in oxygen-18.

1 Gather information.

- The mass numbers for the two isotopes are 17 and 18.

2 Plan your work.

- An oxygen atom must be electrically neutral.

3 Calculate.

- *atomic number* = *number of protons* = *number of electrons* = 8
- *mass number* − *atomic number* = *number of neutrons*
- For oxygen-17, 17 − 8 = 9 neutrons
- For oxygen-18, 18 − 8 = 10 neutrons

4 Verify your results.

- The two isotopes have the same numbers of protons and electrons and differ only in their numbers of neutrons.

PRACTICE HINT

The only difference between the isotopes of an element is the number of neutrons in the atoms of each isotope.

PRACTICE

1. Chlorine has two stable isotopes, chlorine-35 and chlorine-37. The atomic number of chlorine is 17. Calculate the numbers of protons, electrons, and neutrons each isotope has.
2. Calculate the numbers of protons, electrons, and neutrons for each of the following isotopes of calcium: $^{42}_{20}Ca$ and $^{44}_{20}Ca$.

PROBLEM SOLVING SKILL

2 Section Review

UNDERSTANDING KEY IDEAS

1. Describe the differences between electrons, protons, and neutrons.
2. How are isotopes of the same element alike?
3. What subatomic particle was discovered with the use of a cathode-ray tube?

PRACTICE PROBLEMS

4. Write the symbol for element X, which has 22 electrons and 22 neutrons.
5. Determine the numbers of electrons, protons, and neutrons for each of the following:

   a. $^{80}_{35}Br$ **b.** $^{106}_{46}Pd$ **c.** $^{133}_{55}Cs$
6. Calculate the atomic number and mass number of an isotope that has 56 electrons and 82 neutrons.

CRITICAL THINKING

7. Why must there be an attractive force to explain the existence of stable nuclei?
8. Are hydrogen-3 and helium-3 isotopes of the same element? Explain your answer.

SECTION

3 Electron Configuration

Key Terms

- orbital
- electromagnetic spectrum
- ground state
- excited state
- quantum number
- Pauli exclusion principle
- electron configuration
- aufbau principle
- Hund's rule

Objectives

1. **Compare** the Rutherford, Bohr, and quantum models of an atom.
2. **Explain** how the wavelengths of light emitted by an atom provide information about electron energy levels.
3. **List** the four quantum numbers, and describe their significance.
4. **Write** the electron configuration of an atom by using the Pauli exclusion principle and the aufbau principle.

Atomic Models

Soon after the atomic theory was widely accepted by scientists, they began constructing models of atoms. Scientists used the information that they had about atoms to build these models. They knew, for example, that an atom has a densely packed nucleus that is positively charged. This conclusion was the only way to explain the data from Rutherford's gold foil experiments.

Building a model helps scientists imagine what may be happening at the microscopic level. For this very same reason, the illustrations in this book provide pictures that are models of chemical compounds to help you understand the relationship between the macroscopic and microscopic worlds. Scientists knew that any model they make may have limitations. A model may even have to be modified or discarded as new information is found. This is exactly what happened to scientists' models of the atom.

Rutherford's Model Proposed Electron Orbits

The experiments of Rutherford's team led to the replacement of the plum-pudding model of the atom with a nuclear model of the atom. Rutherford suggested that electrons, like planets orbiting the sun, revolve around the nucleus in circular or elliptical orbits. **Figure 15** shows Rutherford's model. Because opposite charges attract, the negatively charged electrons should be pulled into the positively charged nucleus. Because Rutherford's model could not explain why electrons did not crash into the nucleus, his model had to be modified.

The Rutherford model of the atom, in turn, was replaced only two years later by a model developed by Niels Bohr, a Danish physicist. The Bohr model, which is shown in **Figure 16,** describes electrons in terms of their energy levels.

Figure 15
According to Rutherford's model of the atom, electrons orbit the nucleus just as planets orbit the sun.

Bohr's Model Confines Electrons to Energy Levels

According to Bohr's model, electrons can be only certain distances from the nucleus. Each distance corresponds to a certain quantity of energy that an electron can have. An electron that is as close to the nucleus as it can be is in its lowest energy level. The farther an electron is from the nucleus, the higher the energy level that the electron occupies. The difference in energy between two energy levels is known as a *quantum* of energy.

The energy levels in Bohr's model can be compared to the rungs of a ladder. A person can go up and down the ladder only by stepping on the rungs. When standing on the first rung, the person has the lowest potential energy. By climbing to the second rung, the person increases his or her potential energy by a fixed, definite quantity. Because the person cannot stand between the rungs on the ladder, the person's potential energy cannot have a continuous range of values. Instead, the values can be only certain, definite ones. In the same way, Bohr's model states that an electron can be in only one energy level or another, not between energy levels. Bohr also concluded that an electron did not give off energy while in a given energy level.

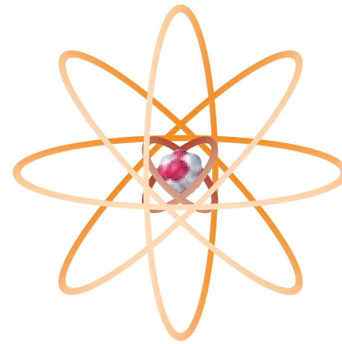


Figure 16
According to Bohr's model of the atom, electrons travel around the nucleus in specific energy levels.

Electrons Act Like Both Particles and Waves

Thomson's experiments demonstrated that electrons act like particles that have mass. Although the mass of an electron is extremely small, electrons in a cathode ray still have enough mass to turn a paddle wheel.

In 1924, Louis de Broglie pointed out that the behavior of electrons according to Bohr's model was similar to the behavior of waves. For example, scientists knew that any wave confined in space can have only certain frequencies. The frequency of a wave is the number of waves that pass through a given point in one second. De Broglie suggested that electrons could be considered waves confined to the space around a nucleus. As waves, electrons could have only certain frequencies. These frequencies could correspond to the specific energy levels in which electrons are found.

Other experiments also supported the wave nature of electrons. Like light waves, electrons can change direction through diffraction. Diffraction refers to the bending of a wave as the wave passes by the edge of an object, such as a crystal. Experiments also showed that electron beams, like waves, can interfere with each other.

Figure 17 shows the present-day model of the atom, which takes into account both the particle and wave properties of electrons. According to this model, electrons are located in **orbitals,** regions around a nucleus that correspond to specific energy levels. Orbitals are regions where electrons are likely to be found. Orbitals are sometimes called *electron clouds* because they do not have sharp boundaries. When an orbital is drawn, it shows where electrons are most likely to be. Because electrons can be in other places, the orbital has a fuzzy boundary like a cloud.

As an analogy to an electron cloud, imagine the spinning blades of a fan. You know that each blade can be found within the spinning image that you see. However, you cannot tell exactly where any one blade is at a particular moment.

orbital
a region in an atom where there is a high probability of finding electrons

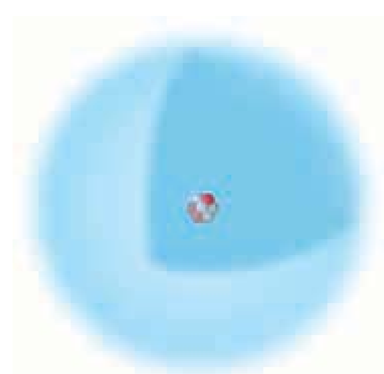


Figure 17
According to the current model of the atom, electrons are found in orbitals.

Electrons and Light

By 1900, scientists knew that light could be thought of as moving waves that have given frequencies, speeds, and wavelengths.

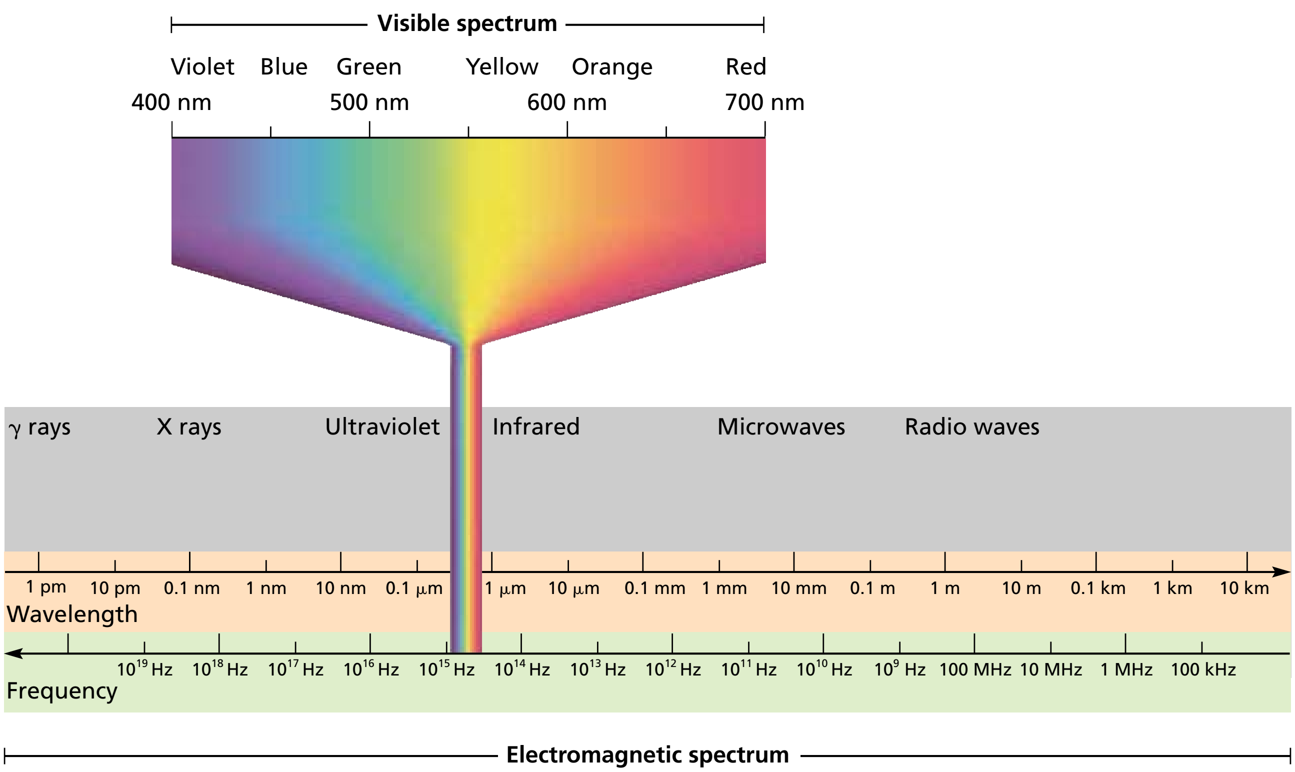
In empty space, light waves travel at 2.998×10^8 m/s. At this speed, light waves take only 500 s to travel the 150 million kilometers between the sun and Earth. The *wavelength* is the distance between two consecutive peaks or troughs of a wave. The distance of a wavelength is usually measured in meters. The wavelength of light can vary from 10^5 m to less than 10^{-13} m. This broad range of wavelengths makes up the **electromagnetic spectrum,** which is shown in **Figure 18.** Notice in **Figure 18** that our eyes are sensitive to only a small portion of the electromagnetic spectrum. This sensitivity ranges from 700 nm, which is about the value of wavelengths of red light, to 400 nm, which is about the value of wavelengths of violet light.

electromagnetic spectrum
all of the frequencies or wavelengths of electromagnetic radiation

In 1905, Albert Einstein proposed that light also has some properties of particles. His theory would explain a phenomenon known as the *photoelectric effect.* This effect happens when light strikes a metal and electrons are released. What confused scientists was the observation that for a given metal, no electrons were emitted if the light's frequency was below a certain value, no matter how long the light was on. Yet if light were just a wave, then any frequency eventually should supply enough energy to remove an electron from the metal.

Einstein proposed that light has the properties of both waves and particles. According to Einstein, light can be described as a stream of particles, the energy of which is determined by the light's frequency. To remove an electron, a particle of light has to have at least a minimum energy and therefore a minimum frequency.

Figure 18
The electromagnetic spectrum is composed of light that has a broad range of wavelengths. Our eyes can detect only the visible spectrum.



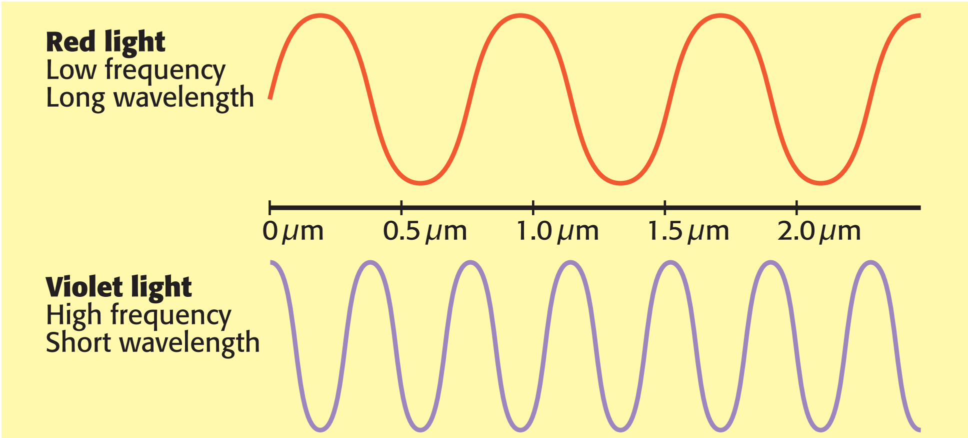


Figure 19 The frequency and wavelength of a wave are inversely related. As frequency increases, wavelength decreases.

Light Is an Electromagnetic Wave

When passed through a glass prism, sunlight produces the visible spectrum—all of the colors of light that the human eye can see. You can see from **Figure 18** on the previous page that the visible spectrum is only a tiny portion of the electromagnetic spectrum. The electromagnetic spectrum also includes X rays, ultraviolet and infrared light, microwaves, and radio waves. Each of these electromagnetic waves is referred to as *light,* although we cannot see these wavelengths.

Figure 19 shows the frequency and wavelength of two regions of the spectrum that we see: red and violet lights. If you compare red and violet lights, you will notice that red light has a low frequency and a long wavelength. But violet light has a high frequency and a short wavelength. The frequency and wavelength of a wave are inversely related.

Light Emission

When a high-voltage current is passed through a tube of hydrogen gas at low pressure, lavender-colored light is seen. When this light passes through a prism, you can see that the light is made of only a few colors. This spectrum of a few colors is called a *line-emission spectrum.* Experiments with other gaseous elements show that each element has a line-emission spectrum that is made of a different pattern of colors.

In 1913, Bohr showed that hydrogen's line-emission spectrum could be explained by assuming that the hydrogen atom's electron can be in any one of a number of distinct energy levels. The electron can move from a low energy level to a high energy level by absorbing energy. Electrons at a higher energy level are unstable and can move to a lower energy level by releasing energy. This energy is released as light that has a specific wavelength. Each different move from a particular energy level to a lower energy level will release light of a different wavelength.

Bohr developed an equation to calculate all of the possible energies of the electron in a hydrogen atom. His values agreed with those calculated from the wavelengths observed in hydrogen's line-emission spectrum. In fact, his values matched with the experimental values so well that his atomic model that is described earlier was quickly accepted.

Light Provides Information About Electrons

Normally, if an electron is in a state of lowest possible energy, it is in a **ground state.** If an electron gains energy, it moves to an **excited state.** An electron in an excited state will release a specific quantity of energy as it quickly "falls" back to its ground state. This energy is emitted as certain wavelengths of light, which give each element a unique line-emission spectrum.

ground state

the lowest energy state of a quantized system

excited state

a state in which an atom has more energy than it does at its ground state

Figure 20 shows the wavelengths of light in a line-emission spectrum for hydrogen, through which a high-voltage current was passed. The high-voltage current may supply enough energy to move an electron from its ground state, which is represented by $n = 1$ in **Figure 20,** to a higher excited state for an electron in a hydrogen atom, represented by $n > 1$. Eventually, the electron will lose energy and return to a lower energy level. For example, the electron may fall from the $n = 7$ energy level to the $n = 3$ energy level. Notice in **Figure 20** that when this drop happens, the electron emits a wavelength of infrared light. An electron in the $n = 6$ energy level can also fall to the $n = 2$ energy level. In this case, the electron emits a violet light, which has a shorter wavelength than infrared light does.

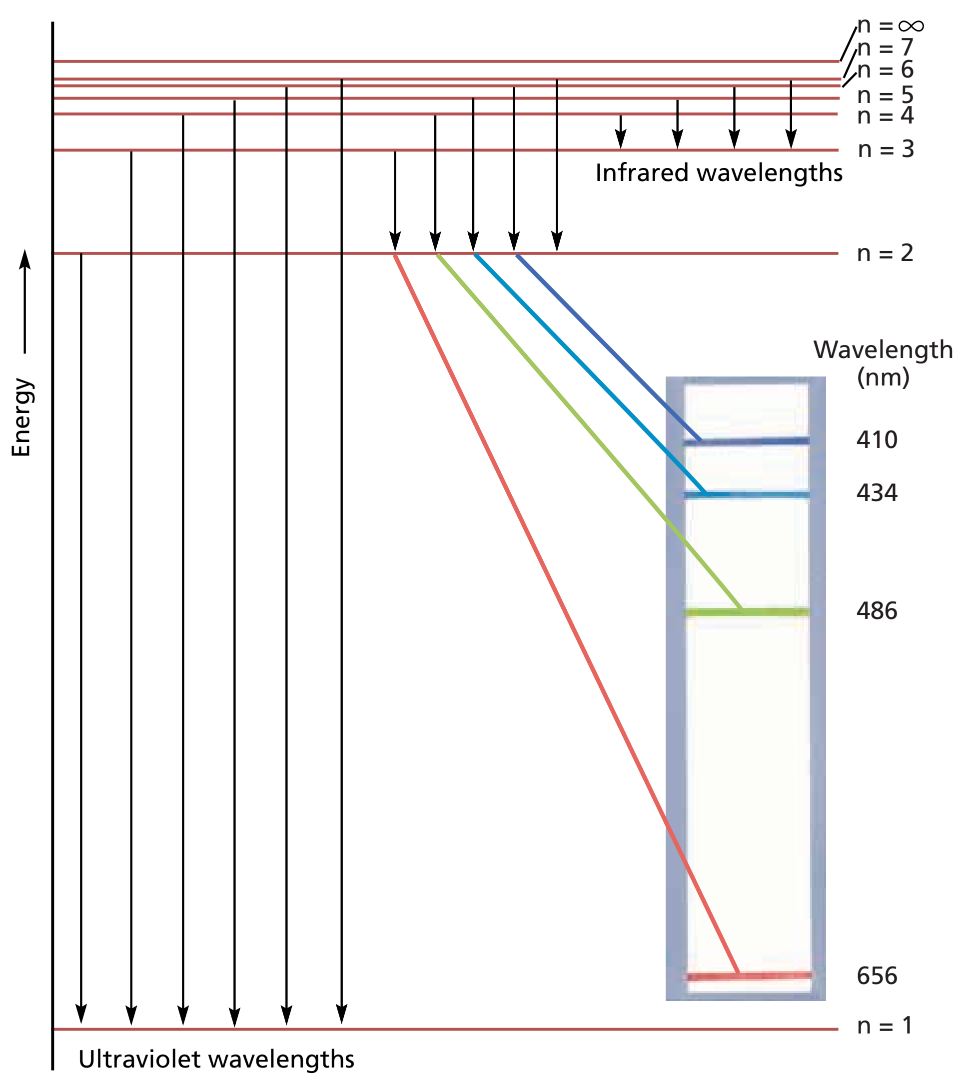


Figure 20
An electron in a hydrogen atom can move between only certain energy states, shown as $n = 1$ to $n = 7$. In dropping from a higher energy state to a lower energy state, an electron emits a characteristic wavelength of light.

Table 5 Quantum Numbers of the First 30 Atomic Orbitals

n	*l*	*m*	Orbital name	Number of orbitals
1	0	0	1*s*	1
2	0	0	2*s*	1
2	1	−1, 0, 1	2*p*	3
3	0	0	3*s*	1
3	1	−1, 0, 1	3*p*	3
3	2	−2, −1, 0, 1, 2	3*d*	5
4	0	0	4*s*	1
4	1	−1, 0, 1	4*p*	3
4	2	−2, −1, 0, 1, 2	4*d*	5
4	3	−3, −2, −1, 0, 1, 2, 3	4*f*	7

Quantum Numbers

The present-day model of the atom, in which electrons are located in orbitals, is also known as the *quantum model.* According to this model, electrons within an energy level are located in orbitals, regions of high probability for finding a particular electron. However, the model does not explain how the electrons move about the nucleus to create these regions.

To define the region in which electrons can be found, scientists have assigned four **quantum numbers** to each electron. **Table 5** lists the quantum numbers for the first 30 atomic orbitals. The *principal quantum number,* symbolized by n, indicates the main energy level occupied by the electron. Values of n are positive integers, such as 1, 2, 3, and 4. As n increases, the electron's distance from the nucleus and the electron's energy increases.

quantum number
a number that specifies the properties of electrons

The main energy levels can be divided into sublevels. These sublevels are represented by the *angular momentum quantum number,* l. This quantum number indicates the shape or type of orbital that corresponds to a particular sublevel. Chemists use a letter code for this quantum number. A quantum number $l = 0$ corresponds to an *s* orbital, $l = 1$ to a *p* orbital, $l = 2$ to a *d* orbital, and $l = 3$ to an *f* orbital. For example, an orbital with $n = 3$ and $l = 1$ is called a 3*p* orbital, and an electron occupying that orbital is called a 3*p* electron.

The *magnetic quantum number,* symbolized by m, is a subset of the l quantum number. It also indicates the numbers and orientations of orbitals around the nucleus. The value of m takes whole-number values, depending on the value of l. The number of orbitals includes one *s* orbital, three *p* orbitals, five *d* orbitals, and seven *f* orbitals.

The *spin quantum number,* symbolized by $+\frac{1}{2}$ or $-\frac{1}{2}$ (↑ or ↓), indicates the orientation of an electron's magnetic field relative to an outside magnetic field. A single orbital can hold a maximum of two electrons, which must have opposite spins.

Electron Configurations

Figure 21 shows the shapes and orientations of the *s, p,* and *d* orbitals. Each orbital that is shown can hold a maximum of two electrons. The discovery that two, but no more than two, electrons can occupy a single orbital was made in 1925 by the German chemist Wolfgang Pauli. This rule is known as the **Pauli exclusion principle.**

Pauli exclusion principle

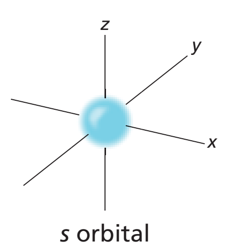
the principle that states that two particles of a certain class cannot be in the exact same energy state

Another way of stating the Pauli exclusion principle is that no two electrons in the same atom can have the same four quantum numbers. The two electrons can have the same value of *n* by being in the same main energy level. These two electrons can also have the same value of *l* by being in orbitals that have the same shape. And, these two electrons may have the same value of *m* by being in the same orbital. But these two electrons cannot have the same spin quantum number. If one electron has the value of $+\frac{1}{2}$, then the other electron must have the value of $-\frac{1}{2}$.

The arrangement of electrons in an atom is usually shown by writing an **electron configuration.** Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. An electron configuration of an atom shows the lowest-energy arrangement of the electrons for the element.

electron configuration

the arrangement of electrons in an atom



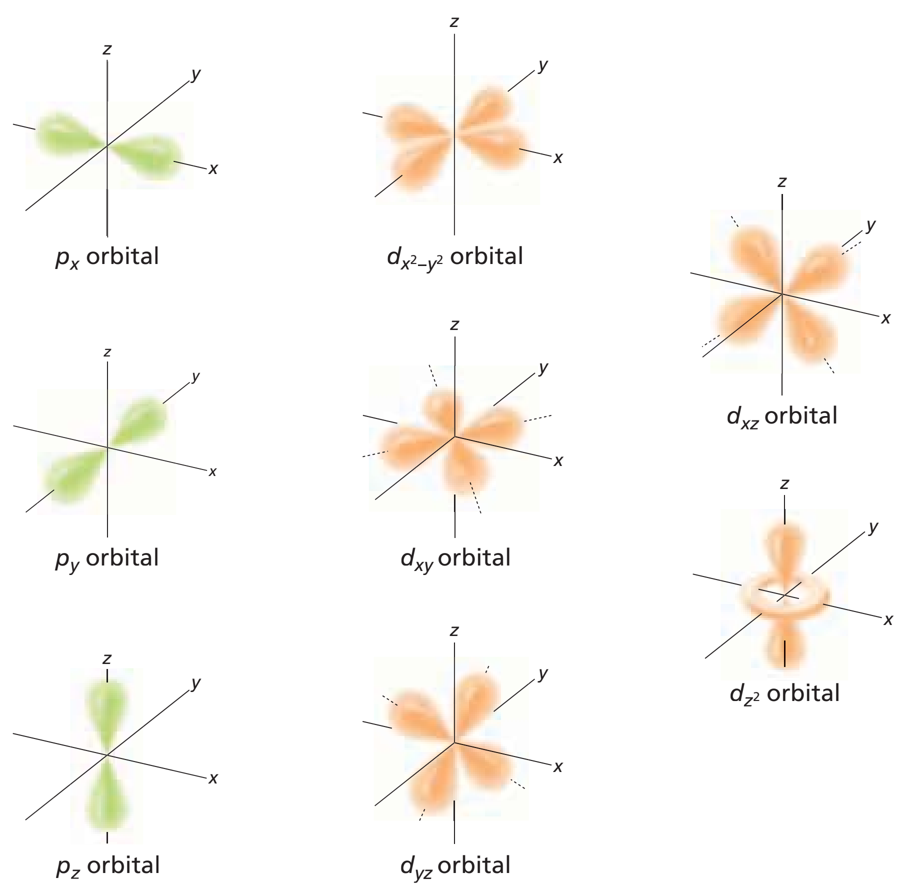


Figure 21
a The *s* orbital is spherically shaped. There is one *s* orbital for each value n = 1, 2, 3…of the principal number.

b For each of the values n = 2, 3, 4…, there are three *p* orbitals. All are dumbbell shaped, but they differ in orientation.

c For each of the values n = 3, 4, 5…, there are five *d* orbitals. Four of the five have similar shapes, but differ in orientation.

An Electron Occupies the Lowest Energy Level Available

The Pauli exclusion principle is one rule to help you write an electron configuration for an atom. Another rule is the aufbau principle. *Aufbau* is the German word for "building up." The **aufbau principle** states that electrons fill orbitals that have the lowest energy first.

aufbau principle

the principle that states that the structure of each successive element is obtained by adding one proton to the nucleus of the atom and one electron to the lowest-energy orbital that is available

Recall that the smaller the principal quantum number, the lower the energy. But within an energy level, the smaller the l quantum number, the lower the energy. Recall that chemists use letters to represent the l quantum number. So, the order in which the orbitals are filled matches the order of energies, which starts out as follows:

$$1s < 2s < 2p < 3s < 3p$$

After this point, the order is less obvious. **Figure 22** shows that the energy of the 3*d* orbitals is slightly higher than the energy of the 4*s* orbitals. As a result, the order in which the orbitals are filled is as follows:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d$$

Additional irregularities occur at higher energy levels.

Can you determine which orbitals electrons of a carbon atom occupy? Two electrons occupy the 1*s* orbital, two electrons occupy the 2*s* orbital, and two electrons occupy the 2*p* orbitals. Now try the same exercise for titanium. Two electrons occupy the 1*s* orbital, two electrons occupy the 2*s* orbital, six electrons occupy the 2*p* orbitals, two electrons occupy the 3*s* orbital, six electrons occupy the 3*p* orbitals, two electrons occupy the 3*d* orbitals, and two electrons occupy the 4*s* orbital.

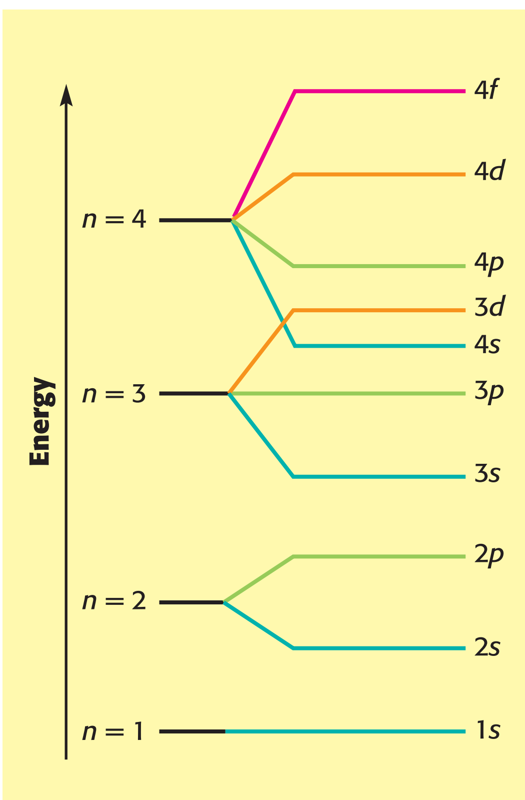


Figure 22
This diagram illustrates how the energy of orbitals can overlap such that 4*s* fills before 3*d*.

An Electron Configuration Is a Shorthand Notation

Based on the quantum model of the atom, the arrangement of the electrons around the nucleus can be shown by the nucleus's electron configuration. For example, sulfur has sixteen electrons. Its electron configuration is written as $1s^22s^22p^63s^23p^4$. This line of symbols tells us about these sixteen electrons. Two electrons are in the 1*s* orbital, two electrons are in the 2*s* orbital, six electrons are in the 2*p* orbitals, two electrons are in the 3*s* orbital, and four electrons are in the 3*p* orbitals.

Each element's configuration builds on the previous elements' configurations. To save space, one can write this configuration by using a configuration of a noble gas. The noble gas electron configurations that are often used are the configurations of neon, argon, krypton, and xenon. The neon atom's configuration is $1s^22s^22p^6$, so the electron configuration of sulfur is written as shown below.

$$[\mathrm{Ne}]3s^23p^4$$

Hund's rule

the rule that states that for an atom in the ground state, the number of unpaired electrons is the maximum possible and these unpaired electrons have the same spin

Does an electron enter the first 3*p* orbital to pair with a single electron that is already there? Or does the electron fill another 3*p* orbital? According to Hund's rule, the second answer is correct. **Hund's rule** states that orbitals of the same *n* and *l* quantum numbers are each occupied by one electron before any pairing occurs. For example, sulfur's configuration is shown by the *orbital diagram* below. Electrons are represented by arrows. Note that an electron fills each 3*p* orbital before an electron enters an orbital that is already occupied.

1*s*	2*s*	2*p*	3*s*	3*p*
↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑ ↑

SAMPLE PROBLEM C

Writing Electron Configurations

PRACTICE HINT

Remember that an *s* orbital holds 2 electrons, three *p* orbitals hold 6 electrons, and five *d* orbitals hold 10 electrons.

Write the electron configuration for an atom whose atomic number is 20.

1 Gather information.

- The atomic number of the element is 20.

2 Plan your work.

- The atomic number represents the number of protons in an atom.
- The number of protons must equal the number of electrons in an atom.
- Write the electron configuration for that number of electrons by following the Pauli exclusion principle and the aufbau principle.
- A noble gas configuration can be used to write this configuration.

3 Calculate.

- *atomic number = number of protons = number of electrons* = 20
- According to the aufbau principle, the order of orbital filling is 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, and so on.
- The electron configuration for an atom of this element is written as follows:

$$1s^22s^22p^63s^23p^64s^2$$

- This electron configuration can be abbreviated as follows:

$$[Ar]4s^2$$

4 Verify your results.

- The sum of the superscripts is (2 + 2 + 6 + 2 + 6 + 2) = 20. Therefore, all 20 electrons are included in the electron configuration.

PRACTICE

PROBLEM SOLVING SKILL

1. Write the electron configuration for an atom of an element whose atomic number is 8.
2. Write the electron configuration for an atom that has 17 electrons.

3 Section Review

UNDERSTANDING KEY IDEAS

1. How does Bohr's model of the atom differ from Rutherford's?

2. What happens when an electron returns to its ground state from its excited state?

3. What does *n* represent in the quantum model of electrons in atoms?

PRACTICE PROBLEMS

4. What is the atomic number of an element whose atom has the following electron configuration: $1s^22s^22p^63s^23p^63d^24s^2$?

5. Write the electron configuration for an atom that has 13 electrons.

6. Write the electron configuration for an atom that has 33 electrons.

7. How many orbitals are completely filled in an atom whose electron configuration is $1s^22s^22p^63s^1$?

CRITICAL THINKING

8. Use the Pauli exclusion principle or the aufbau principle to explain why the following electron configurations are incorrect:

a. $1s^22s^32p^63s^1$

b. $1s^22s^22p^53s^1$

9. Why is a shorter wavelength of light emitted when an electron "falls" from $n = 4$ to $n = 1$ than when an electron "falls" from $n = 2$ to $n = 1$?

10. Calculate the maximum number of electrons that can occupy the third principal energy level.

11. Why do electrons fill the 4*s* orbital before they start to occupy the 3*d* orbital?

SECTION

4 Counting Atoms

KEY TERMS

- **atomic mass**
- **mole**
- **molar mass**
- **Avogadro's number**

OBJECTIVES

1. **Compare** the quantities and units for atomic mass with those for molar mass.
2. **Define** *mole,* and explain why this unit is used to count atoms.
3. **Calculate** either mass with molar mass or number with Avogadro's number given an amount in moles.

Atomic Mass

You would not expect something as small as an atom to have much mass. For example, copper atoms have an average mass of only 1.0552×10^{-25} kg.

Each penny in **Figure 23** has an average mass of 3.13×10^{-3} kg and contains copper. How many copper atoms are there in one penny? Assuming that a penny is pure copper, you can find the number of copper atoms by dividing the mass of the penny by the average mass of a single copper atom or by using the following conversion factor:

$$1 \text{ atom Cu}/1.0552 \times 10^{-25} \text{ kg}$$

$$3.13 \times 10^{-3} \cancel{\text{kg}} \times \frac{1 \text{ atom Cu}}{1.0552 \times 10^{-25} \cancel{\text{kg}}} = 2.97 \times 10^{22} \text{ Cu atoms}$$

Masses of Atoms Are Expressed in Atomic Mass Units

atomic mass
the mass of an atom expressed in atomic mass units

Obviously, atoms are so small that the gram is not a very convenient unit for expressing their masses. Even the picogram (10^{-12} g) is not very useful. A special mass unit is used to express **atomic mass.** This unit has two names—the atomic mass unit (amu) and the Dalton (Da). In this book, *atomic mass unit* will be used.

But how can you tell what the atomic mass of a specific atom is? When the atomic mass unit was first set up, an atom's mass number was supposed to be the same as the atom's mass in atomic mass units. Mass number and atomic mass units would be the same because a proton and a neutron each have a mass of about 1.0 amu.

For example, a copper-63 atom has an atomic mass of 62.940. A copper-65 atom has an atomic mass of 64.928. (The slight differences from exact values will be discussed in later chapters.)

Another way to determine atomic mass is to check a periodic table, such as the one on the inside cover of this book. The mass shown is an average of the atomic masses of the naturally occurring isotopes. For this reason, copper is listed as 63.546 instead of 62.940 or 64.928.

Figure 23
These pennies are made mostly of copper atoms. Each copper atom has an average mass of 1.0552×10^{-25} kg.

Introduction to the Mole

Most samples of elements have great numbers of atoms. To make working with these numbers easier, chemists created a new unit called the *mole* (mol). A **mole** is defined as the number of atoms in exactly 12 grams of carbon-12. The mole is the SI unit for the amount of a substance.

mole
the SI base unit used to measure the amount of a substance whose number of particles is the same as the number of atoms in 12 g of carbon-12

Chemists use the mole as a counting unit, just as you use the dozen as a counting unit. Instead of asking for 12 eggs, you ask for 1 dozen eggs. Similarly, chemists refer to 1 mol of carbon or 2 mol of iron.

To convert between moles and grams, chemists use the molar mass of a substance. The **molar mass** of an element is the mass in grams of one mole of the element. Molar mass has the unit grams per mol (g/mol). The mass in grams of 1 mol of an element is numerically equal to the element's atomic mass from the periodic table in atomic mass units. For example, the atomic mass of copper to two decimal places is 63.55 amu. Therefore, the molar mass of copper is 63.55 g/mol. **Skills Toolkit 1** shows how to convert between moles and mass in grams using molar mass.

molar mass
the mass in grams of 1 mol of a substance

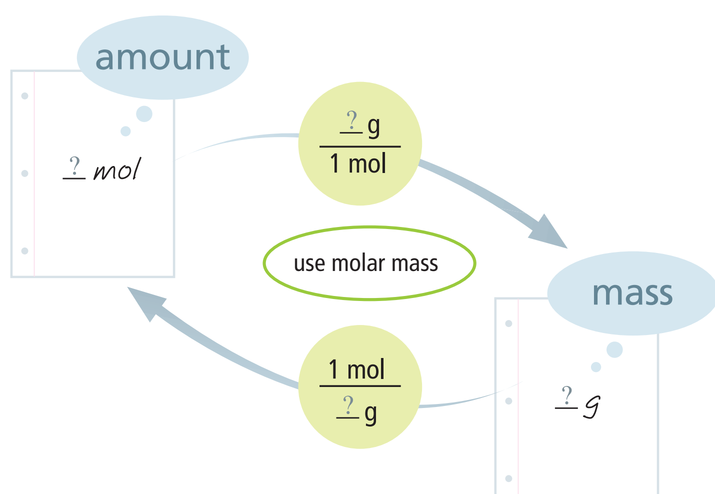
Scientists have also determined the number of particles present in 1 mol of a substance, called **Avogadro's number.** One mole of pure substance contains 6.022 1367 $\times$ 10^{23} particles. To get some idea of how large Avogadro's number is, imagine that every living person on Earth (about 6 billion people) started counting the number of atoms of 1 mol C. If each person counted nonstop at a rate of one atom per second, it would take over 3 million years to count every atom.

Avogadro's number
6.022 $\times$ 10^{23}, the number of atoms or molecules in 1 mol

Avogadro's number may be used to count any kind of particle, including atoms and molecules. For calculations in this book, Avogadro's number will be rounded to 6.022 $\times$ 10^{23} particles per mole. **Skills Toolkit 2** shows how to use Avogadro's number to convert between amount in moles and the number of particles.

SKILLS Toolkit 1

Determining the Mass from the Amount in Moles



SAMPLE PROBLEM D

Converting from Amount in Moles to Mass

Determine the mass in grams of 3.50 mol of copper.

1 Gather information.

- amount of Cu = 3.50 mol
- mass of Cu = ? g Cu
- molar mass of Cu = 63.55 g

2 Plan your work.

- First, make a set-up that shows what is given and what is desired.

$$3.50 \text{ mol Cu} \times ? = ? \text{ g Cu}$$

- Use a conversion factor that has g Cu in the numerator and mol Cu in the denominator.

$$3.50 \cancel{\text{mol Cu}} \times \frac{? \text{ g Cu}}{1 \cancel{\text{mol Cu}}} = ? \text{ g Cu}$$

PRACTICE HINT

For elements and compounds, the mass will always be a number that is greater than the number of moles.

3 Calculate.

- The correct conversion factor is the molar mass of Cu, 63.55 g/mol. Place the molar mass in the equation, and calculate the answer. Use the periodic table in this book to find mass numbers of elements.

$$3.50 \cancel{\text{mol Cu}} \times \frac{63.55 \text{ g Cu}}{1 \cancel{\text{mol Cu}}} = 222 \text{ g Cu}$$

4 Verify your results.

- To verify that the answer of 222 g is correct, find the number of moles of 222 g of copper.

$$222 \text{ g of Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} = 3.49 \text{ mol Cu}$$

The amount of 3.49 mol is close to the 3.50 mol, so the answer of 222 g is reasonable.

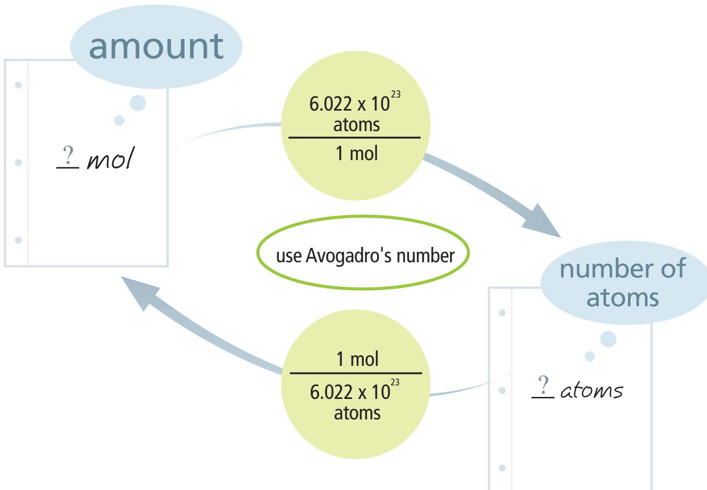
PRACTICE

PROBLEM SOLVING SKILL

1. What is the mass in grams of 1.00 mol of uranium?
2. What is the mass in grams of 0.0050 mol of uranium?
3. Calculate the number of moles of 0.850 g of hydrogen atoms. What is the mass in grams of 0.850 mol of hydrogen atoms?
4. Calculate the mass in grams of 2.3456 mol of lead. Calculate the number of moles of 2.3456 g of lead.

SKILLS Toolkit 2

Determining the Number of Atoms from the Amount in Moles



SAMPLE PROBLEM E

Converting from Amount in Moles to Number of Atoms

Determine the number of atoms in 0.30 mol of fluorine atoms.

1 Gather information.

- amount of F = 0.30 mol
- number of atoms of F = ?

2 Plan your work.

- To determine the number of atoms, select the conversion factor that will take you from the amount in moles to the number of atoms.

$$\text{amount (mol)} \times 6.022 \times 10^{23} \text{ atoms/mol} = \text{number of atoms}$$

3 Calculate.

$$0.30 \cancel{\text{mol F}} \times \frac{6.022 \times 10^{23} \text{ F atoms}}{1 \cancel{\text{mol F}}} = 1.8 \times 10^{23} \text{ F atoms}$$

4 Verify your results.

- The answer has units that are requested in the problem. The answer is also less than 6.022×10^{23} atoms, which makes sense because you started with less than 1 mol.

PRACTICE HINT

Make sure to select the correct conversion factor so that units cancel to give the unit required in the answer.

PRACTICE

1. How many atoms are in 0.70 mol of iron?
2. How many moles of silver are represented by 2.888×10^{23} atoms?
3. How many moles of osmium are represented by 3.5×10^{23} atoms?

PROBLEM SOLVING SKILL

Figure 24
Carbon, which composes diamond, is the basis for the atomic mass scale that is used today.

Chemists and Physicists Agree on a Standard

The atomic mass unit has been defined in a number of different ways over the years. Originally, atomic masses expressed the ratio of the mass of an atom to the mass of a hydrogen atom. Using hydrogen as the standard turned out to be inconvenient because hydrogen does not react with many elements. Early chemists determined atomic masses by comparing how much of one element reacted with another element.

Because oxygen combines with almost all other elements, oxygen became the standard of comparison. The atomic mass of oxygen was defined as exactly 16, and the atomic masses of the other elements were based on this standard. But this choice also led to difficulties. Oxygen exists as three isotopes. Physicists based their atomic masses on the assignment of 16.0000 as the mass of the most common oxygen isotope. Chemists, on the other hand, decided that 16.0000 should be the average mass of all oxygen isotopes, weighted according to the abundance of each isotope. So, to a physicist, the atomic mass of fluorine was 19.0044, but to a chemist, it was 18.9991.

Finally, in 1960, a conference of chemists and physicists agreed on a scale based on an isotope of carbon. Carbon is shown in **Figure 24.** Used by all scientists today, this scale defines the atomic mass unit as exactly one-twelfth of the mass of one carbon-12 atom. As a result, one atomic mass unit is equal to $1.600\,5402 \times 10^{-27}$ kg. The mass of an atom is indeed quite small.

4 Section Review

UNDERSTANDING KEY IDEAS

1. What is atomic mass?
2. What is the SI unit for the amount of a substance that contains as many particles as there are atoms in exactly 12 grams of carbon-12?
3. Which atom is used today as the standard for the atomic mass scale?
4. What unit is used for molar mass?
5. How many particles are present in 1 mol of a pure substance?

PRACTICE PROBLEMS

6. Convert 3.01×10^{23} atoms of silicon to moles of silicon.
7. How many atoms are present in 4.0 mol of sodium?
8. How many moles are represented by 118 g of cobalt? Cobalt has an atomic mass of 58.93 amu.
9. How many moles are represented by 250 g of platinum?
10. Convert 0.20 mol of boron into grams of boron. How many atoms are present?

CRITICAL THINKING

11. What is the molar mass of an element?
12. How is the mass in grams of an element converted to amount in moles?
13. How is the mass in grams of an element converted to number of atoms?

4
Be
Beryllium
9.012 182
$[He]2s^2$

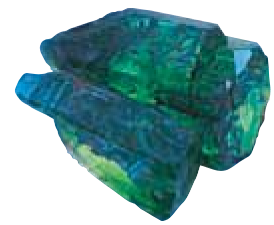
Where Is Be?

Earth's crust:
0.005% by mass

Element Spotlight

Beryllium: An Uncommon Element

Although it is an uncommon element, beryllium has a number of properties that make it very useful. Beryllium has a relatively high melting point (1278°C) and is an excellent conductor of energy as heat and electrical energy. Beryllium transmits X rays extremely well and is therefore used to make "windows" for X-ray devices. All compounds of beryllium are toxic to humans. People who experience prolonged exposure to beryllium dust may contract berylliosis, a disease that can lead to severe lung damage and even death.



Crystals of pure beryllium look very different from the combined form of beryllium in an emerald.

Industrial Uses

- The addition of 2% beryllium to copper forms an alloy that is six times stronger than copper is. This alloy is used for nonsparking tools, critical moving parts in jet engines, and components in precision equipment.
- Beryllium is used in nuclear reactors as a neutron reflector and as an alloy with the fuel elements.

Real-World Connection Emerald and aquamarine are precious forms of the mineral beryl, $Be_3Al_2(SiO_3)_6$.

A Brief History

1800 — 1900

1798: R. J. Haüy, a French mineralogist, observes that emeralds and beryl have the same optical properties and therefore the same chemical composition.

1828: F. Wöhler of Germany gives beryllium its name after he and W. Bussy of France simultaneously isolate the pure metal.

1898: P. Lebeau discovers a method of extracting high-purity beryllium by using an electrolytic process.

1926: M. G. Corson of the United States discovers that beryllium can be used to age-harden copper-nickel alloys.

1942: A Ra-Be source provides the neutrons for Fermi's studies. These studies lead to the construction of a nuclear reactor.

Questions

1. Research how the beryllium and copper alloy is made and what types of equipment are made of this alloy.
2. Research how beryllium is used in nuclear reactors.
3. Research berylliosis and use the information to make a medical information brochure. Be sure to include symptoms, causes, and risk factors in your report.

internet connect
www.scilinks.org
Topic : Beryllium
SciLinks code: HW4021
SCILINKS Maintained by the National Science Teachers Association

3 CHAPTER HIGHLIGHTS

KEY IDEAS

SECTION ONE Substances Are Made of Atoms

- Three laws support the existence of atoms: the law of definite proportions, the law of conservation of mass, and the law of multiple proportions.
- Dalton's atomic theory contains five basic principles, some of which have been modified.

SECTION TWO Structure of Atoms

- Protons, particles that have a positive charge, and neutrons, particles that have a neutral charge, make up the nuclei of most atoms.
- Electrons, particles that have a negative charge and very little mass, occupy the region around the nucleus.
- The atomic number of an atom is the number of protons the atom has. The mass number of an atom is the number of protons plus the number of neutrons.
- Isotopes are atoms that have the same number of protons but different numbers of neutrons.

SECTION THREE Electron Configuration

- The quantum model describes the probability of locating an electron at any place.
- Each electron is assigned four quantum numbers that describe it. No two electrons of an atom can have the same four quantum numbers.
- The electron configuration of an atom reveals the number of electrons an atom has.

SECTION FOUR Counting Atoms

- The masses of atoms are expressed in atomic mass units (amu). The mass of an atom of the carbon-12 isotope is defined as exactly 12 atomic mass units.
- The mole is the SI unit for the amount of a substance that contains as many particles as there are atoms in exactly 12 grams of carbon-12.
- Avogadro's number, 6.022×10^{23} particles per mole, is the number of particles in a mole.

KEY TERMS

law of definite proportions
law of conservation of mass
law of multiple proportions

electron
nucleus
proton
neutron
atomic number
mass number
isotope

orbital
electromagnetic spectrum
ground state
excited state
quantum number
Pauli exclusion principle
electron configuration
aufbau principle
Hund's rule

atomic mass
mole
molar mass
Avogadro's number

KEY SKILLS

Determining the Number of Particles in an Atom
Sample Problem A p. 86

Determining the Number of Particles in Isotopes
Sample Problem B p. 89

Writing Electron Configurations
Sample Problem C p. 98

Converting Amount in Moles to Mass
Skills Toolkit 1 p. 101
Sample Problem D p. 102

Converting Amount in Moles to Number of Atoms
Skills Toolkit 2 p. 103
Sample Problem E p. 103

Physical Setting/Chemistry
REGENTS EXAM PRACTICE

3

PART A

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

1. Any sample of sodium chloride contains *exactly* 60.7% chlorine and 39.3% sodium. This is best described by the
(1) law of conservation of mass.
(2) law of multiple proportions.
(3) law of definite proportions.
(4) law of conservation of energy.

2. Which statement about the mass and charge of an electron is correct?
(1) An electron has a mass greater than a proton and has a positive charge.
(2) An electron has a mass greater than a proton and has a negative charge.
(3) An electron has a mass smaller than a proton and has a positive charge.
(4) An electron has a mass smaller than a proton and has a negative charge.

3. What is the nuclear charge of an atom of strontium?
(1) +38 **(3)** 0
(2) +88 **(4)** +50

4. In Rutherford's gold foil experiment most of the alpha particles passed straight through the gold foil. As a result of this experiment, he concluded that atoms consist mainly of
(1) a large nucleus. **(3)** electrons.
(2) empty space. **(4)** protons.

5. The nucleus of an atom consists of
(1) protons and neutrons.
(2) protons, neutrons, and electrons.
(3) protons only.
(4) empty space.

6. What is the total number of protons found in the nucleus of an atom of sodium-23?
(1) 23 **(3)** 16
(2) 11 **(4)** 12

7. An atom of hydrogen-2 and an atom of hydrogen-3 differ in their
(1) number of protons.
(2) atomic number.
(3) number of electrons.
(4) number of neutrons.

8. Which subatomic particle is electrically neutral?
(1) proton **(3)** neutron
(2) electron **(4)** positron

9. What is the mass number of an atom that contains 17 protons, 17 electrons, and 18 neutrons?
(1) 35 **(3)** 17
(2) 52 **(4)** 34

10. Which of the following is an isotope of carbon-12?
(1) $^{13}_{6}X$ **(3)** $^{13}_{7}X$
(2) $^{12}_{7}X$ **(4)** $^{12}_{6}X$

11. Which symbol represents an atom with 19 protons and 21 neutrons?
(1) ^{40}Ca **(3)** ^{21}Sc
(2) ^{40}K **(4)** ^{39}K

Regents Test-Taking Tip

Choose the best possible answer for each question, even if you think there is another possible answer that is not given.

12. All isotopes of an atom have
(1) the same number of protons and the same number of neutrons.
(2) the same number of protons and a different number of neutrons.
(3) a different number of protons and the same number of neutrons.
(4) a different number of protons and a different number of neutrons.

13. Which element has an electron configuration of $1s^2 2s^2 2p^6 3s^2$ in the ground state?
(1) Ca (3) Cl
(2) F (4) Mg

14. Which electron configuration represents an atom in the ground state that has five electrons in its outermost energy level?
(1) $1s^2 2s^5$
(2) $1s^2 2s^2 2p^5$
(3) $1s^2 2s^2 2p^3$
(4) $1s^2 2s^2 2p^3 3s^2$

15. As electrons fall from the excited state back to the ground state, energy is emitted in the form of
(1) light.
(2) gamma radiation.
(3) heat.
(4) alpha particles.

16. Which electron transition is accompanied by the greatest emission of energy (n = the principal energy level)?
(1) $n = 3$ to $n = 4$
(2) $n = 5$ to $n = 2$
(3) $n = 2$ to $n = 5$
(4) $n = 4$ to $n = 3$

17. Electrons are most likely found in the region of the atom known as a(n)
(1) quantum. (3) nucleus.
(2) photon. (4) energy level.

18. One atomic mass unit (amu) is equal to
(1) $\frac{1}{16}$ the mass of a ^{16}O atom.
(2) $\frac{1}{8}$ the mass of an ^{8}O atom.
(3) $\frac{1}{12}$ the mass of a ^{12}C atom.
(4) $\frac{1}{6}$ the mass of a ^{12}C atom.

Note that questions 19 through 22 have only three choices.

19. As an electron moves from the ground state to the excited state, its energy
(1) increases.
(2) decreases.
(3) remains the same.

20. As the mass number of the isotopes of carbon increases, the number of protons
(1) increases.
(2) decreases.
(3) remains the same.

21. Whenever a sodium atom loses an electron, the number of protons in its nucleus
(1) increases.
(2) decreases.
(3) remains the same.

22. What is needed to convert between mass and the amount in moles?
(1) molar mass
(2) Avogadro's number
(3) four quantum numbers

PART B–1

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

23. In which outermost energy level are electrons of the elements in Period 3 found?
(1) 1 (3) 3
(2) 2 (4) 4

24. Which pair contains the same number of neutrons?
(1) $^{13}_{6}$C and $^{13}_{7}$N
(2) $^{13}_{7}$N and $^{14}_{7}$N
(3) $^{13}_{6}$C and $^{14}_{7}$N
(4) $^{13}_{6}$C and $^{12}_{6}$C

25. All isotopes of neutral atoms of boron have
(1) 5 protons and 6 neutrons.
(2) 5 protons and 5 electrons.
(3) 35 protons and 45 neutrons.
(4) 35 protons and 35 electrons.

26. The total number of electrons in a neutral atom of any element is equal to the element's
(1) number of protons.
(2) number of neutrons.
(3) number of protons plus neutrons.
(4) mass number.

27. Which pair has the same electron configuration?
(1) Na and Ar **(3)** Na^+ and Ne
(2) Na and Ne **(4)** Na^+ and Ar

28. There are 3 isotopes of hydrogen, H-1, H-2, and H-3. All of three isotopes have
(1) a mass of 1.00794 amu.
(2) an atomic number of 2.
(3) the same number of neutrons.
(4) the same number of protons.

29. Carbon exists as three isotopes: C-12, C-13, and C-14. Carbon has an average atomic mass of 12.0107 amu. This indicates that
(1) any sample of carbon consists of an equal amount of each type of isotopes.
(2) any sample of carbon consists of mostly the C-13 isotope.
(3) any sample of carbon consists of mostly the C-12 isotope.
(4) all carbon atoms have a mass of 12.0107 amu.

30. What is the mass of 0.25 mol of calcium?
(1) 5.0 grams **(3)** 25 grams
(2) 40 grams **(4)** 10 grams

31. What is the mass of 4 moles of neon?
(1) 40 grams **(3)** 20 grams
(2) 10 grams **(4)** 80 grams

32. What is the average atomic mass of the element whose atoms have 7 protons?
(1) 7.00 amu **(3)** 21.0 amu
(2) 14.0 amu **(4)** 8.0 amu

33. What is the atomic mass of an atom with 8 protons, 9 neutrons, and 8 electrons?
(1) 8 amu **(3)** 17 amu
(2) 16 amu **(4)** 25 amu

34. How many moles is 788 grams of Au?
(1) 4 mol
(2) 0.25 mol
(3) 1.6×10^5 mol
(4) 6.02×10^{23} mol

35. A hypothetical element consists of two isotopes that have masses of 35.0 amu and 37.0 amu. The isotope that has a mass of 35.0 amu represents 75% of the element. What is the average atomic mass of this hypothetical element?
(1) 35.0 amu **(3)** 36 amu
(2) 35.5 amu **(4)** 37 amu

36. Which could be the atomic number of an atom with four electrons in its outer energy level?
(1) 22 **(3)** 14
(2) 4 **(4)** 10

37. Which of the following is the electron configuration for an atom of iodine in the ground state?
(1) $[Ar]4s^24p^1$ **(3)** $[Kr]5s^25p^5$
(2) $[Kr]4s^24p^5$ **(4)** $[Xe]5s^25p^5$

38. Which of the following elements has seven electrons in its outermost energy level?
(1) oxygen **(3)** neon
(2) iodine **(4)** nitrogen

39. The neutral atom with an electron configuration of $1s^22s^22p^63s^23p^6$ is
(1) fluorine.
(2) neon.
(3) argon.
(4) potassium.

40. How many protons does a neutral atom with an electron configuration of $1s^22s^22p^63s^2$ contain in its nucleus?
(1) 12 **(3)** 36
(2) 24 **(4)** 48

41. Which is the electron configuration of an atom that has electrons in an excited state?
(1) $1s^2 2s^2 2p^1$
(2) $1s^2 2s^2 2p^5$
(3) $1s^2 2s^2 2p^3$
(4) $1s^2 2s^2 3s^2$

42. As an electron moves from the ground state to the excited state, the energy of the atom
(1) increases.
(2) decreases.
(3) remains the same.
(4) goes to zero.

43. Spectral lines are produced by electrons when
(1) they move from higher energy levels to lower energy levels.
(2) they move from lower energy levels to higher energy levels.
(3) they are removed from an atom.
(4) they move into the nucleus.

44. Which electron transition is accompanied by the release of energy?
(1) $3p$ to $2s$
(2) $1s$ to $3p$
(3) $2s$ to $2p$
(4) $3s$ to $4s$

45. Which principal energy level has electrons with the lowest energy?
(1) $n = 1$
(2) $n = 2$
(3) $n = 3$
(4) $n = 4$

PART B–2

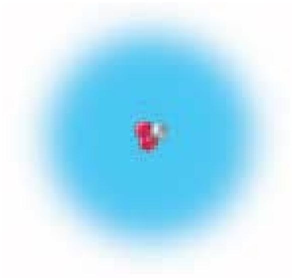
Answer the following items.

46. An atom contains 11 protons, 11 electrons, and 12 neutrons. Answer the following questions based on this information.
a. Write the symbol for this atom.
b. What is the atomic number of this atom?
c. What is the mass number of this atom?
d. Write the electron configuration for this atom.
e. Draw the symbol for an isotope of this atom.

47. When a student performed a flame test in the lab with a sample of sodium chloride, a bright yellow flame was observed.
a. Write the electron configuration of sodium in the ground state.
b. Write a possible electron configuration of sodium in the excited state.
c. What is the difference in energy of the atom in the ground state versus the excited state?
d. Explain why a sample of potassium chloride produces a different color (violet) when subjected to the same conditions as the flame test with a sample of sodium chloride.

48. Calculate the number of atoms present in each of the following:
a. 2 mol Fe
b. 40.1 g Ca, which has an atomic mass of 40.08 amu
c. 4.5 mol of boron-11
d. 0.03 g of radon

49. The diagram below represents the model of an atom.

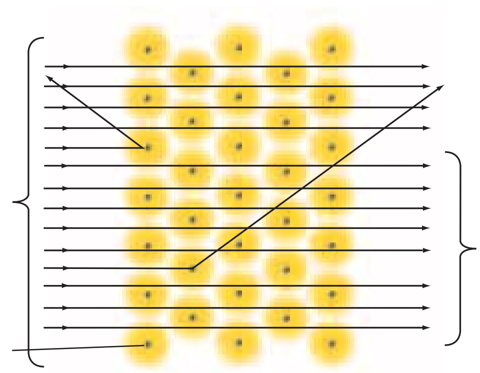


a. What is the atomic number of the atom depicted in the figure above?
b. What is the mass number of the atom depicted in the figure above?
c. What element is represented by the figure above?
d. What is the nuclear charge of the atom depicted in the figure above?
e. Draw a model of an isotope of this element.

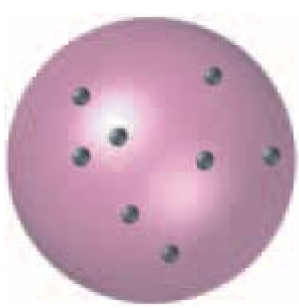
PART C

Answer the following items.

50. In 1909, observations by Ernest Rutherford were not consistent with the plum-pudding model of the atom. The diagram to the right represents Ernest Rutherford's gold foil experiment.



a. Most of the positively charged particles passed undeflected through the gold foil. What did this illustrate about the structure of the atom?
b. Some particles were greatly or slightly deflected. What did this illustrate about the atom?
c. Below is an illustration of J.J. Thomson's plum-pudding model of an atom which proposed that the electrons of an atom were embedded in a positively-charged ball of matter. Discuss two ways in which this model differs from the current theory of atomic structure.



51. Complete the following table based on isotopes of silicon.

Isotope	Number of protons	Number of electrons	Number of neutrons
Si-28			
Si-29			
Si-30			

52. Define and state the importance of each of the following terms which deal with electron configuration of an atom.
a. Pauli exclusion principle
b. Aufbau principle
c. Hund's rule

53. Write the symbol and electron configuration for each of the following atoms.
a. an atom with 20 protons, 20 neutrons, and 20 electrons
b. an atom with 17 protons, 18 neutrons, and 17 electrons
c. an atom with 18 protons, 21 neutrons, and 18 electrons

54. Answer the following questions based on a 48 gram sample of magnesium.
a. How many moles of magnesium does the sample have?
b. How many atoms of magnesium does the sample have?
c. How many protons compose the nucleus of each magnesium atom?

55. Answer the following questions based on a 2.5 mole sample of solid strontium.
a. How many grams of strontium does the sample have?
b. How many atoms of strontium does the sample have?
c. How many protons, neutrons, and electrons compose a neutral atom of strontium-88?
d. Write the electron configuration for strontium in the ground state.
e. How many electrons are present in the outermost energy level?
f. How many occupied principal energy levels does strontium have?
g. What is the net charge of the strontium nucleus?

56. Calculate the number of atoms present in the following:
a. 2 moles of copper
b. 10 grams of neon

57. Write the ground state electron configuration for a neutral atom that has 17 protons in its nucleus.

FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."

58. What represents the ground state in this diagram?

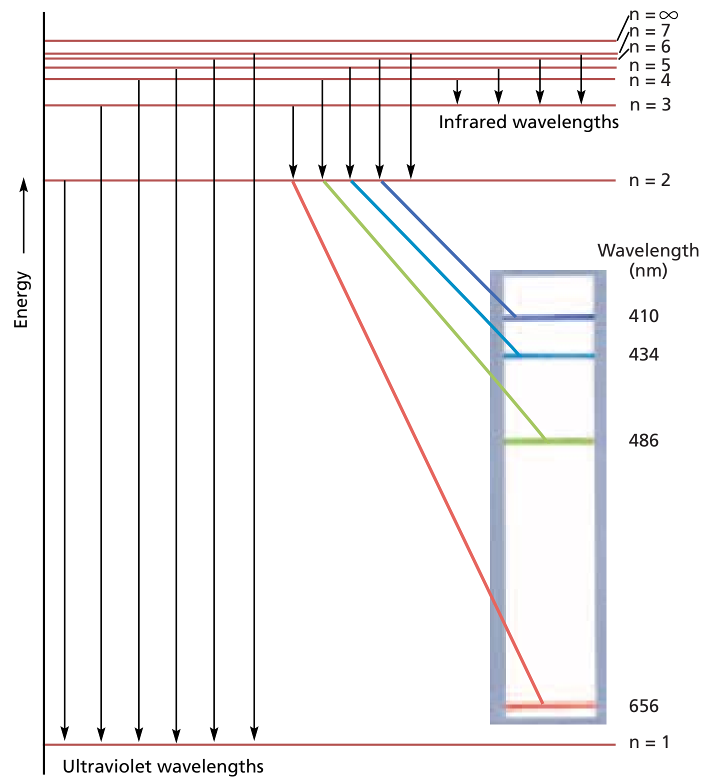
59. Which energy-level changes can be detected by the unaided eye?

60. Does infrared light have more energy than ultraviolet light? Why or why not?

61. Which energy levels represent a hydrogen electron in an excited state?

62. What does the energy level labeled "$n = \infty$" represent?

63. If an electron is beyond the $n = \infty$ level, is the electron a part of the hydrogen atom?



TECHNOLOGY AND LEARNING

64. Graphing Calculator

Calculate Numbers of Protons, Electrons, and Neutrons.

A graphing calculator can run a program that calculates the numbers of protons, electrons, and neutrons given the atomic mass and numbers for an atom. For example, given a calcium-40 atom, you will calculate the numbers of protons, electrons, and neutrons in the atom.

Go to Appendix C. If you are using a TI-83 Plus, you can download the program NUMBER and data and can run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. After you have run the program, answer the questions below.

a. Which element has the most protons?

b. How many neutrons does mercury-201 have?

c. Carbon-12 and carbon-14 have the same atomic number. Do they have the same number of neutrons? Why or why not?