86. A dry cleaning solvent contains chlorofluorocarbons (CFCs). Is this dry cleaning solvent an example of a homogeneous mixture or a pure substance?

87. The electronics industry manufactures semiconductor chips from silicon. Refer to the periodic table and predict an element that can be substituted for silicon.

88. The electronics industry manufactures transistors using arsenic diffusion. Refer to the periodic table and predict an element that can be substituted for arsenic.

89. Write the chemical symbol for each of the following elements.

(a) ferrum (b) plumbum
(c) stannum (d) aurum

90. State the atomic number for each of the following elements.

(a) rutherfordium (b) dubnium
(c) seaborgium (d) bohrium

91. State whether the following describes a physical or a chemical change: changing physical state but not chemical formula.

92. State whether the following describes a physical or a chemical change: changing chemical formula but not physical state.

93. The reaction of hydrogen and nitrogen gases produces 5.00 g of ammonia and releases 3250 cal of heat. How much energy is required to decompose 5.00 g of ammonia into hydrogen and nitrogen gases?

94. The decomposition of 10.0 g of ammonia requires 27,200 J of energy to give hydrogen and nitrogen gases. How much energy is released when hydrogen and nitrogen gases react to produce 10.0 g of ammonia?

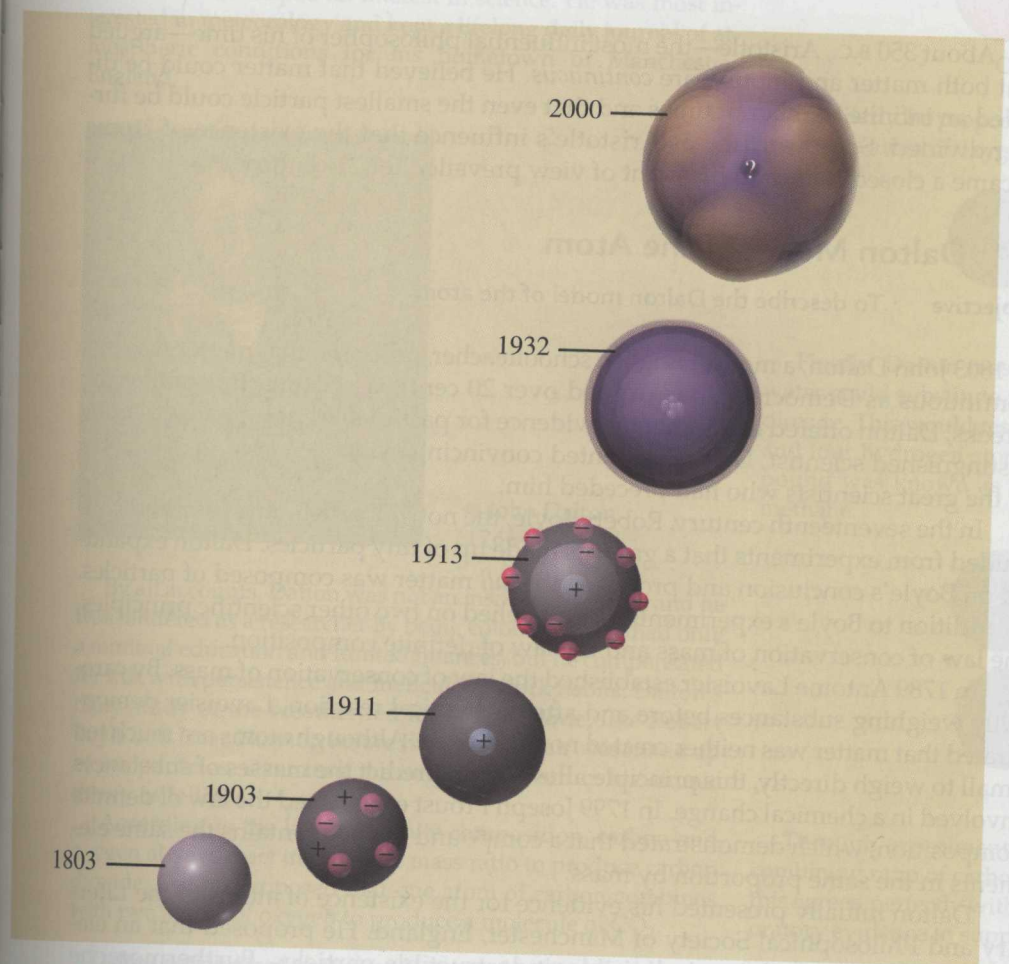
95. Iron and sulfur react to produce iron sulfide and heat energy. An experiment shows that the mass of iron and sulfur is equal to the mass of the iron sulfide. In *theory*, should the products weigh slightly more or slightly less than the reactants?

96. Hydrogen and iodine react to give hydrogen iodide while absorbing heat energy. An experiment shows that the total mass of the reactants is equal to the product. In *theory*, should the product weigh slightly more or slightly less than the reactants?

Explorer Quiz 1
Explorer Quiz 2
Explorer Quiz 3
Master Quiz

CHAPTER 5

Models of the Atom

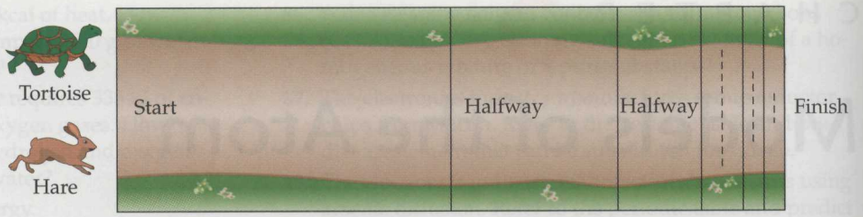


▲ In the evolving model of the atom, what does the question mark (?) represent in the 2000 model?

This chapter is carefully designed to gradually increase in sophistication of treatment for atomic theory. That is, starting with the simplistic Dalton model, the explanations evolve through the Thomson model, Rutherford model, and Bohr model. The treatment of the quantum mechanical atom, which can be considered optional, is reserved for the end of the chapter. I often choose not to cover the concepts of electron waves and orbitals in introductory chemistry as the theory may confuse many students who are trying to grasp the concepts of energy levels and sublevels. I perform demonstrations with gas discharge tubes and atomic models to help students visualize the concepts (refer to the *Instructor's Resource Manual*).

The concept of an atom was born in Greece about 450 B.C. One of the great issues of the day was whether matter and motion were continuous or discontinuous. Zeno, a Greek philosopher, pointed out that to travel any distance, you first must cover half the distance, then half the remaining distance, and so on. This paradox suggests that you will never arrive at your destination if motion is continuous (Figure 5.1). Zeno therefore reasoned that motion is *discontinuous* and occurs by a series of tiny leaps.

Democritus, another Greek philosopher, argued that matter is *discontinuous* and could not be infinitely divided. He believed that at some point a fundamental indivisible particle would emerge. He called this ultimate particle an atom, from the Greek word *atomos*, which means "indivisible."



▶ **Figure 5.1 Zeno's Paradox** Motion must occur by discontinuous jumps; otherwise, the race can never be completed. Neither the tortoise nor the hare will complete the race if they continue to move half the distance to the finish.

About 350 B.C., Aristotle—the most influential philosopher of his time—argued that both matter and motion are *continuous*. He believed that matter could be divided an infinite number of times and that even the smallest particle could be further divided. So powerful was Aristotle's influence that the existence of atoms became a closed issue and his point of view prevailed for 21 centuries.

5.1 Dalton Model of the Atom

Objective · To describe the Dalton model of the atom.

In 1803 John Dalton, a modest English schoolteacher, proposed that matter was discontinuous as Democritus had argued over 20 centuries before. But unlike the Greeks, Dalton offered experimental evidence for particles. Although he was not a distinguished scientist, Dalton presented convincing evidence based on the work of the great scientists who had preceded him.

In the seventeenth century, Robert Boyle, the noted English physicist, had concluded from experiments that a gas was made up of tiny particles. Dalton expanded on Boyle's conclusion and proposed that *all* matter was composed of particles. In addition to Boyle's experiments, Dalton relied on two other scientific principles, the law of conservation of mass and the law of definite composition.

In 1789 Antoine Lavoisier established the law of conservation of mass. By carefully weighing substances before and after a chemical reaction, Lavoisier demonstrated that matter was neither created nor destroyed. Although atoms are much too small to weigh directly, this principle allows us to predict the masses of substances involved in a chemical change. In 1799 Joseph Proust established the law of definite composition, which demonstrated that a compound always contains the same elements in the same proportion by mass.

Dalton initially presented his evidence for the existence of atoms to the Literary and Philosophical Society of Manchester, England. He proposed that an element is composed of tiny, indivisible, indestructible particles. Furthermore, he argued that compounds are simply combinations of two or more atoms of different elements. In 1808 Dalton published the atomic theory in his classic textbook *A New System of Chemical Philosophy*.

In a surprisingly short period of time, the atomic theory was generally accepted by the scientific community. The theory can be summarized as follows.

1. An element is composed of tiny, indivisible, indestructible particles called atoms.
2. All atoms of an element are identical and have the same properties.
3. Atoms of different elements combine to form compounds.
4. Compounds contain atoms in small whole-number ratios.
5. Atoms can combine in more than one ratio to form different compounds.

As we will learn, Dalton's first two proposals were incorrect, but the atomic theory was nonetheless, an important step toward understanding the nature of matter.

Chemistry Connection · John D

How was Dalton able to accept an honorary de when wearing scarlet was forbidden by his Qua

John Dalton was born the son of a weaver into a devoutly religious family. At the age of 12, he began teaching at a Quaker school and developed an interest in science. He was most interested in meteorology and kept a lifelong daily journal of atmospheric conditions for his hometown of Manchester, England.

◀ John Dalton (1766–1844)

By all accounts, Dalton was not an inspiring lecturer, and he was hindered as a researcher by being color-blind. He had only a minimal education and limited finances, but he compensated for this with persistence and meticulous work habits. Dalton's daily study of the weather led him to conclude, like Robert Boyle and Isaac Newton before him, that the air was made up of gas particles. Over time, he began to construct his atomic theory as follows.

According to the law of definite composition, carbon and oxygen always react in the same mass ratio to produce carbon dioxide. Dalton proposed that one atom of carbon combines with two atoms of oxygen to produce a molecule of CO_2.

Dalton rationalized that since he was color-blind and saw the cere

5.2 Thomson Model of the Atom

Objectives · To describe the Thomson plum-pudding model of
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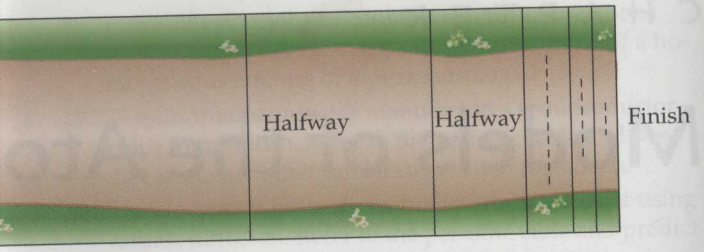

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John Dalton was born the son of a weaver into a devoutly religious family. At the age of 12, he began teaching at a Quaker school and developed an interest in science. He was most interested in meteorology and kept a lifelong daily journal of atmospheric conditions for his hometown of Manchester, England.

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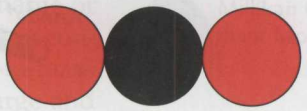
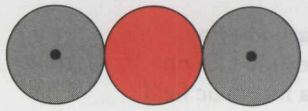


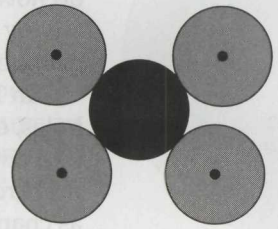
Carbon dioxide, CO_2

Similarly, he proposed that two atoms of hydrogen combine with one atom of oxygen to give a molecule of H_2O.



Water, H_2O

Finally, Dalton reasoned that two atoms of hydrogen in water could substitute for each of the oxygen atoms in carbon dioxide. This would result in a molecule with one carbon atom and four hydrogen atoms, that is, CH_4. At the time, this compound was known as marsh gas, but we now refer to it as methane.



Methane, CH_4

Through experimentation, Dalton indeed found that the combining ratio of carbon to hydrogen in methane was 1 to 4; this agreed perfectly with his prediction. Thus, Dalton had laboratory evidence to support the atomic theory.

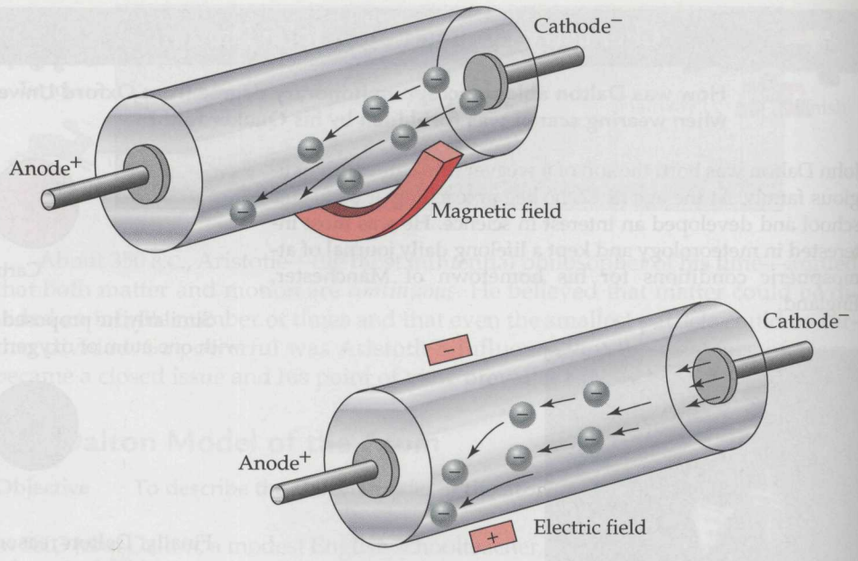
Dalton rationalized that since he was color-blind and saw the ceremonial robe as gray, he was not violating Quaker doctrine.

5.2 Thomson Model of the Atom

Multiple Proportions Movie

Objectives
- To describe the Thomson plum-pudding model of the atom.
- To state the relative charge and mass of the electron and proton.

About 50 years after Dalton's proposal, there was disturbing evidence that the atom was divisible after all. This evidence came from cathode-ray tubes, which were sealed glass tubes containing a gas at low pressure. When electricity is applied to one end, a cathode-ray tube appears to glow. This phenomenon is referred to as fluorescence, and the glowing ray is a type of light energy. Since the ray emanates from the negative cathode in the tube, the radiation is referred to as a **cathode ray**.

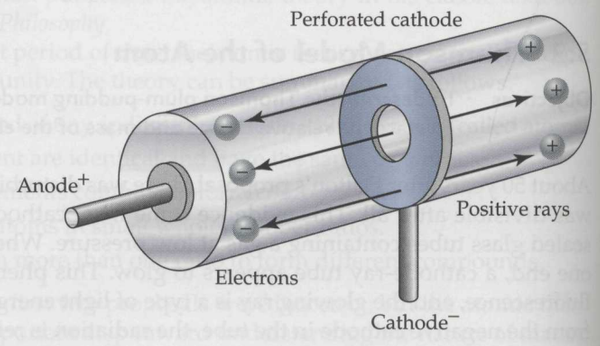


▶ **Figure 5.2 Cathode-Ray Tubes** Note the influence on cathode rays by a magnetic field and an electric field. Several English and German physicists share the credit for these experiments.

In the late 1870s, physicists observed that cathode rays were attracted by a magnetic field (Figure 5.2). This observation suggested that cathode rays are actually particles, not radiation. When different gases are put in the tubes, the results do not change. This led to the theory that cathode rays are composed of tiny, negatively charged particles. One of these subatomic particles was called an **electron** (symbol e^-).

In 1886 a German physicist experimented with cathode-ray tubes that had small holes, or channels, in the cathode. The results showed that positive rays, as well as negative rays, were produced at the cathode. The positive rays, however, moved in the direction opposite the negative cathode rays. The positive rays were referred to as channel rays, which through translation, became known as canal rays (from the German *kanal*, meaning "channel").

Further experiments revealed that canal rays were composed of small positively charged particles (Figure 5.3). The smallest particle was discovered with hydrogen gas in the tube. This particle had a charge equal, but opposite in sign, to that of an electron. One of these subatomic particles in hydrogen gas was called a **proton** (symbol p^+).



▶ **Figure 5.3 Canal Rays** This cathode-ray tube shows positive canal rays move in the direction opposite negative cathode rays.

In 1897 the English physicist J. J. Thomson demonstrated
deflected by an electric field as well as by a magnetic field (Fig
final piece of evidence confirming the notion that electrons ar
evidence had accumulated for 20 years, Thomson is usually
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Since the electron is so tiny, Thomson could not measure
mass. He was able, however, to determine its charge-to-mass
tinued his experiments and in a short time obtained a value fo
ratio for the proton as well as the electron. Table 5.1 lists the
mass of the electron and the proton.

Table 5.1 Relative Charge and Mass of the Electron and

Subatomic Particle	Symbol	Relative Charge
electron	e^-	1−
proton	p^+	1+

In 1903 J. J. Thomson proposed a subatomic model of the at
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of the atom.

Originally, Thomson was only able to determine the relat
ratio of the electron and the proton. However in 1911, after 5 y
servations, the American physicist Robert Millikan determined t
the electron. This allowed Thomson to calculate the actual mass
the proton. Thomson calculated that the mass of the electron is
that the mass of the proton is 1.67×10^{-24} g.

Note J. J. Thomson spent his academic career at Cambridge Universi
27, became director of the prestigious Cavendish Laboratory. In 1906 he
in physics for his work on the electron, and 2 years later he was knigh
tribution was far-reaching—seven of his former students and assistant
Nobel prizes.

5.3 Rutherford Model of the Atom

Objectives
- To describe the Rutherford nuclear model of the atom
- To state the relative charge and approximate mass of
ton, and neutron.

Ernest Rutherford (1871–1937) was digging potatoes on his fath
Zealand when he received the news that he had won a scholarsl
University. He had actually come in second, but the winner dec
ship to get married. The 24-year-old Rutherford postponed his ow
and immediately set out for England.

In 1897 the English physicist J. J. Thomson demonstrated that cathode rays are deflected by an electric field as well as by a magnetic field (Figure 5.2). This was the final piece of evidence confirming the notion that electrons are particles. Although evidence had accumulated for 20 years, Thomson is usually given credit for the discovery of the electron.

Since the electron is so tiny, Thomson could not measure its actual charge or mass. He was able, however, to determine its charge-to-mass ratio. Thomson continued his experiments and in a short time obtained a value for the charge-to-mass ratio for the proton as well as the electron. Table 5.1 lists the relative charge and mass of the electron and the proton.

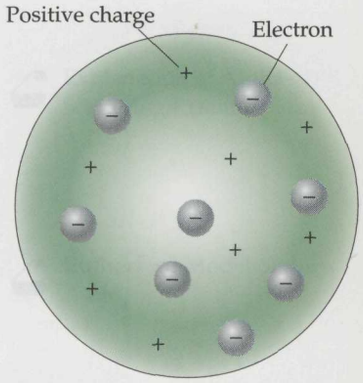
Millikan Oil Drop Experiment Movie

Table 5.1 Relative Charge and Mass of the Electron and Proton

Subatomic Particle	Symbol	Relative Charge	Relative Mass
electron	e^-	1−	1/1836
proton	p^+	1+	1

In 1903 J. J. Thomson proposed a subatomic model of the atom. The model pictured a positively charged atom containing negatively charged electrons. Thomson visualized electrons in homogeneous spheres of positive charge in a way that was analogous to raisins in English plum pudding (Figure 5.4). Thus, the Thomson proposal became popularly known as the *plum-pudding model* or *raisin-pudding model* of the atom.

Originally, Thomson was only able to determine the relative charge-to-mass ratio of the electron and the proton. However in 1911, after 5 years of tedious observations, the American physicist Robert Millikan determined the actual charge on the electron. This allowed Thomson to calculate the actual mass of the electron and the proton. Thomson calculated that the mass of the electron is 9.11×10^{-28} g and that the mass of the proton is 1.67×10^{-24} g.



▲ **Figure 5.4 Thomson Model of the Atom** Atoms are pictured as spheres of positive charge. The small negative particles in the sphere represent electrons.

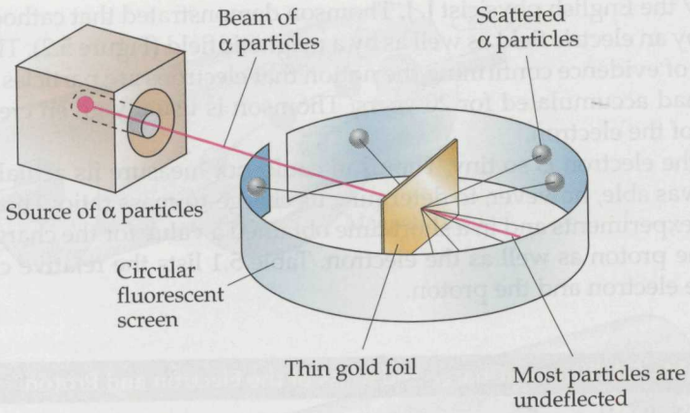
Note J. J. Thomson spent his academic career at Cambridge University and, at the age of 27, became director of the prestigious Cavendish Laboratory. In 1906 he won the Nobel prize in physics for his work on the electron, and 2 years later he was knighted. Thomson's contribution was far-reaching—seven of his former students and assistants went on to receive Nobel prizes.

5.3 Rutherford Model of the Atom

Objectives
- To describe the Rutherford nuclear model of the atom.
- To state the relative charge and approximate mass of the electron, proton, and neutron.

Ernest Rutherford (1871–1937) was digging potatoes on his father's farm in New Zealand when he received the news that he had won a scholarship to Cambridge University. He had actually come in second, but the winner declined the scholarship to get married. The 24-year-old Rutherford postponed his own wedding plans and immediately set out for England.

Alpha-Scattering Experiment



▶ **Figure 5.5 Alpha-Scattering Experiment** The diagram shows the deflection of alpha particles by a thin gold foil. Although the foil was only 0.5 μm thick, a few alpha particles actually rebounded backward.

Rutherford Experiment Movie

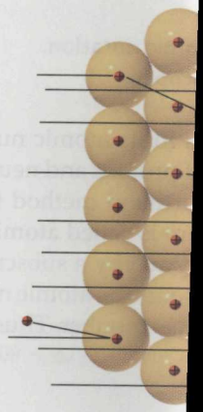
At Cambridge, Ernest Rutherford studied subatomic particles and earned his doctorate working for J. J. Thomson. Upon graduation, Rutherford went to McGill University in Canada and began to work in the field of radioactivity. After a year he went back to New Zealand, married, and returned to Manchester University in England. There he continued to study radioactivity and coined the terms "alpha ray" and "beta ray" for two types of radiation. He discovered a third type of radiation that was not affected by a magnetic field and gave it the name "gamma ray." In 1908 Rutherford received the Nobel prize in chemistry for his work on radioactivity.

In 1906 Rutherford found that alpha rays contained particles identical to those of helium atoms stripped of electrons. He experimented with alpha rays by firing them at thin gold foils. As expected, the particles passed straight through the foil or, on occasion, deflected slightly (Figure 5.5). This observation seemed reasonable since the plum-pudding model of the atom pictured homogeneous spheres.

A few years later the true picture of the atom was unveiled. The person who did the experiment was Hans Geiger, Rutherford's assistant and the inventor of the Geiger counter. Here is a description of the experiment in Rutherford's own words:

> One day Geiger came to me and said, "Don't you think that young Marsden, whom I am training in radioactive methods, ought to begin a small research?" Now I had thought that too, so I said, "Why not let him see if any alpha particles can be scattered through a large angle?" I may tell you in confidence that I did not believe that they would be, since we knew that the alpha particle was a very fast massive particle, with a great deal of energy. Then I remember two or three days later Geiger coming to me in great excitement and saying, "We have been able to get some of the alpha particles coming backwards." It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.

Rutherford interpreted the alpha-scattering results as follows. He believed that most of the alpha particles passed directly through the foil because an atom is largely empty space with electrons moving about. But in the center of the atom is the **atomic nucleus** containing protons. Rutherford reasoned that, compared to the atom, the nucleus is tiny and has a very high density. The alpha particles that



▲ **Figure 5.6 Explanation of t**
Atoms are represented by circle
An alpha particle is positively ch
heavy, positive, gold nucleus.

bounced backward recoiled
particles by the atomic nucl

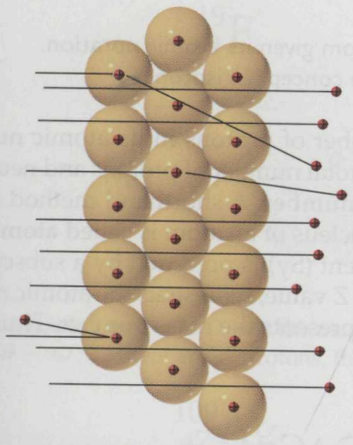
In 1911 Rutherford prop
atively charged electrons
Rutherford was able to esti
ed that an atom has a diame
ameter of about 1×10^{-13}
Figure 5.7.

Since the diameters of
comprehend, consider the
atom compared to the size
centimeter. Moreover, if the
be the size of a small marble
scattered by gold nuclei in

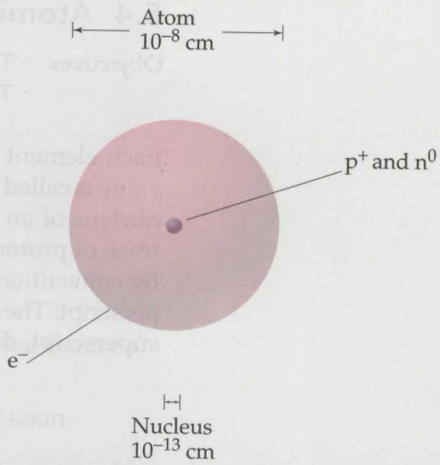
Owing to the heaviness
neutral particles in addition
mer students found the elus
the **neutron** (symbol n^0) an
We can summarize the da
Table 5.2.

Table 5.2 Subatomic Part

Subatomic Particle	Symbol
electron	e^-
proton	p^+
neutron	n^0



▲ **Figure 5.6 Explanation of the Alpha-Scattering Results** Atoms are represented by circles and nuclei by small dots. An alpha particle is positively charged and is deflected by a heavy, positive, gold nucleus.



▲ **Figure 5.7 Rutherford Model of the Atom** Most of the mass of the atom is found in the nucleus. Note that the atoms and nuclei are not to scale, as the size of an atom is about 100,000 times larger than the size of the nucleus.

bounced backward recoiled after striking the dense nucleus. The scattering of alpha particles by the atomic nuclei in the gold foil is illustrated in Figure 5.6.

Explanation of Alpha-Scattering Results

In 1911 Rutherford proposed a new model of the atom. He suggested that negatively charged electrons were distributed about a positively charged nucleus. Rutherford was able to estimate the size of the atom and its nucleus. He calculated that an atom has a diameter of about 1×10^{-8} cm and that the nucleus has a diameter of about 1×10^{-13} cm. The Rutherford model of the atom is illustrated in Figure 5.7.

Rutherford Model of the Atom

Since the diameters of atoms and nuclei are extremely small and difficult to comprehend, consider the following analogy to visualize an atom. The size of an atom compared to the size of its nucleus is similar to a kilometer compared to a centimeter. Moreover, if the atom were as large as the Astrodome, the nucleus would be the size of a small marble. This analogy clarifies why so few alpha particles were scattered by gold nuclei in the thin foil.

Owing to the heaviness of the nucleus, Rutherford predicted that it contained neutral particles in addition to protons. Twenty years later, one of Rutherford's former students found the elusive neutral particle. In 1932 James Chadwick discovered the **neutron** (symbol $\mathbf{n^0}$) and 3 years later was awarded the Nobel prize in physics. We can summarize the data for the electron, proton, and neutron as shown in Table 5.2.

Table 5.2 Subatomic Particles

Subatomic Particle	Symbol	Location	Relative Charge	Relative Mass
electron	e^-	outside nucleus	1−	1/1836
proton	p^+	inside nucleus	1+	1
neutron	n^0	inside nucleus	0	1

5.4 Atomic Notation

Objectives
- To draw a diagram of an atom given its atomic notation.
- To explain and illustrate the concept of isotopes.

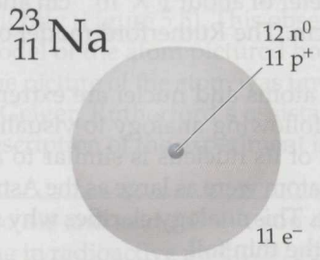
Each element has a characteristic number of protons in its atomic nucleus. This value is called the **atomic number**. The total number of protons and neutrons in the nucleus of an atom is called the **mass number**. A shorthand method for keeping track of protons and neutrons in the nucleus of an atom is called **atomic notation**. By convention, the symbol of the element (Sy) is preceded by a subscript and superscript. The subscript, designated the Z value, represents the atomic number. The superscript, designated the A value, represents the mass number. Thus,

mass number (p^+ and n^0)

$^{A}_{Z}\mathrm{Sy}$ — symbol of the element

atomic number (p^+)

As an example, an atom of sodium can be written: $^{23}_{11}\mathrm{Na}$ Here the atomic number (protons) is 11 and the mass number (protons + neutrons) is 23. From this information, we can determine the number of neutrons. The nucleus of this sodium atom contains 12 neutrons ($A - Z = 23 - 11 = 12$). Since atoms are neutral, the number of negative electrons must equal the number of positive protons. Thus, there are 11 electrons surrounding the nucleus. A simple diagram of the sodium atom is as follows.



Example Exercise 5.1 further illustrates the interpretation of atomic notation to state the composition of an atom.

Example Exercise 5.1 • Atomic Notation

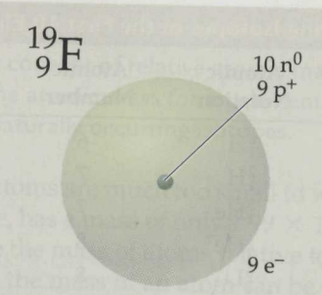
Given the atomic notation for the following atoms, draw a diagram showing the arrangement of protons, neutrons, and electrons.

(a) $^{19}_{9}\mathrm{F}$ (b) $^{109}_{47}\mathrm{Ag}$

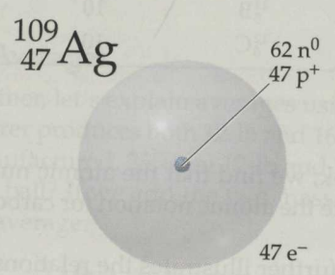
Solution

We can draw a diagram of an atom by showing protons and neutrons in the nucleus surrounded by electrons.

(a) Since the atomic number is 9 and the mass number is 19, the number of neutrons is 10 (19 − 9). If there are 9 protons, there must be 9 electrons.

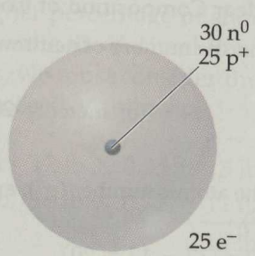


(b) Since the atomic number is 47 and the mass number is 109, the number of neutrons is 62 (109 − 47). If there are 47 protons, there must be 47 electrons.



Self-Test Exercise

Given the following diagram, indicate the nucleus using atomic notation.



Answer: $^{55}_{25}Mn$

Isotopes

Approximately 20 elements occur naturally with a fixed number of neutrons. For most elements, the number of neutrons in the nucleus varies. Atoms of the same element that have a different number of neutrons in the nucleus are called **isotopes**.

In isotopes, as the number of neutrons varies, so does the mass number. For example, hydrogen occurs naturally as two stable isotopes, protium (^1_1H) and deuterium (^2_1H). Protium has only one proton in its nucleus, whereas deuterium has a proton and a neutron. A third isotope of hydrogen, tritium (^3_1H), is unstable and radioactive. Tritium has one proton and two neutrons in its nucleus. Table 5.3 lists the naturally occurring stable isotopes of the first ten elements.

We often refer to an isotope by stating the name of the element followed by its mass number, for example, carbon-14 and cobalt-60. To write the atomic notation, we must also know the atomic number. In this textbook, the atomic number is found above the symbol of the element in the periodic table. If we refer to the inside

Table 5.3 Naturally Occurring Isotopes of the First 10 Elements

Atomic Number	Mass Number	Atomic Notation	Atomic Number	Mass Number	Atomic Notation
1	1	$^{1}_{1}H$	6	13	$^{13}_{6}C$
1	2	$^{2}_{1}H$	7	14	$^{14}_{7}N$
2	3	$^{3}_{2}He$	7	15	$^{15}_{7}N$
2	4	$^{4}_{2}He$	8	16	$^{16}_{8}O$
3	6	$^{6}_{3}Li$	8	17	$^{17}_{8}O$
3	7	$^{7}_{3}Li$	8	18	$^{18}_{8}O$
4	9	$^{9}_{4}Be$	9	19	$^{19}_{9}F$
5	10	$^{10}_{5}B$	10	20	$^{20}_{10}Ne$
5	11	$^{11}_{5}B$	10	21	$^{21}_{10}Ne$
6	12	$^{12}_{6}C$	10	22	$^{22}_{10}Ne$

front cover of this textbook, we find that the atomic number of C is 6 and that of Co is 27. Thus, we can write the atomic notation for carbon-14 as $^{14}_{6}C$ and for cobalt-60 as $^{60}_{27}Co$.

Example Exercise 5.2 further illustrates the relationship between atomic notation and the composition of an atomic nucleus.

Example Exercise 5.2 • Nuclear Composition of Isotopes

State the number of protons and the number of neutrons in an atom of each of the following isotopes.

(a) $^{37}_{17}Cl$ (b) mercury-202

Solution

The subscript value refers to the atomic number (p^+), and the superscript value refers to the mass number (p^+ and n^0).

(a) Thus, $^{37}_{17}Cl$ has 17 p^+ and 20 n^0 (37 − 17 = 20).

(b) In the periodic table, we find that the atomic number of mercury is 80. Thus, the atomic notation, $^{202}_{80}Hg$, indicates 80 p^+ and 122 n^0 (202 − 80 = 122).

Self-Test Exercise

State the number of protons and the number of neutrons in an atom of each of the following isotopes.

(a) $^{120}_{50}Sn$ (b) uranium-238

Answers: (a) 50 p^+ and 70 n^0; (b) 92 p^+ and 146 n^0

Note In general, the properties of isotopes of an element are similar. Consider a carbon-12 atom, which has six protons and six neutrons in the nucleus. If the mass number increases by one neutron, we have a carbon-13 atom. Other than a slight change in mass, the properties of carbon-12 and carbon-13 are nearly identical.

On the other hand, when the atomic number of carbon-12 increases by one proton, we have a nitrogen-13 atom. The additional proton changes the properties radically. Carbon occurs naturally as coal, diamond, and graphite, whereas nitrogen is a colorless, odorless gas in the atmosphere.

5.5 Atomic Mass

Objectives · To explain the concept c
· To calculate the atomic
dance of the naturally o

We realize, of course, that atoms are
A carbon atom, for instance, has a m
atoms, scientists determine the mass
magnetic-field instrument, the mass
carbon-12 atom. The carbon-12 isotop
is assigned a value of 12 atomic mass
(symbol **amu**) is equivalent to 1/12 t

Simple and Weighted Averages

Before proceeding any further, let's e
Let's suppose a manufacturer produc
the total shotput balls manufactured
average mass of a shotput ball? If w
2, we will find the simple average.

12 lb +

The simple average mass of a shotpu
be weighted in favor of the higher p
ages are 25% and 75%, we can write
late the weighted average mass, we n
12-lb and 16-lb balls. We can proceed

12 lb: 12
16 lb: 16

Although the weighted average r
ball weighs 15 lb. An actual shotput ba
represents the theoretical mass of an a

Atomic Mass of an Element

The **atomic mass** of an element is the w
ring isotopes. To calculate the atomic r
used in the shotput example. That is, t
atom, we must consider the mass as w

Carbon has two naturally occurri
We can calculate the atomic mass of ca
of each isotope.

Isotope	M
^{12}C	12.0
^{13}C	13.0

5.5 Atomic Mass

Objectives
- To explain the concept of relative atomic mass.
- To calculate the atomic mass for an element given the mass and abundance of the naturally occurring isotopes.

We realize, of course, that atoms are much too small to weigh directly on a balance. A carbon atom, for instance, has a mass of only 1.99×10^{-23} g. Instead of weighing atoms, scientists determine the mass of atoms relative to each other. With a special magnetic-field instrument, the mass of an atom can be compared to the mass of a carbon-12 atom. The carbon-12 isotope has been chosen as a reference standard and is assigned a value of 12 atomic mass units. Stated differently, an **atomic mass unit** (symbol **amu**) is equivalent to 1/12 the mass of a carbon-12 atom.

Simple and Weighted Averages

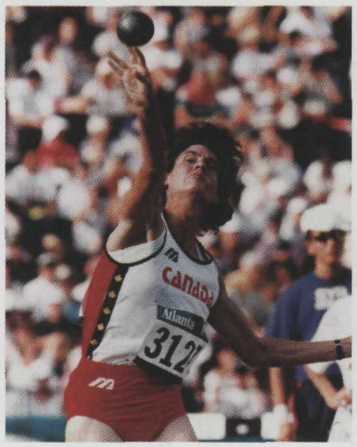
Before proceeding any further, let's explain averages using an interesting analogy. Let's suppose a manufacturer produces both 12-lb and 16-lb metal shotput balls. Of the total shotput balls manufactured, 25% are 12-lb and 75% are 16-lb. What is the average mass of a shotput ball? If we add the two masses together and divide by 2, we will find the simple average.

$$\frac{12\text{ lb} + 16\text{ lb}}{2} = 14\text{ lb}$$

The simple average mass of a shotput ball is 14 lb, but the true average mass must be weighted in favor of the higher percentage of shotput balls. Since the percentages are 25% and 75%, we can write the decimal fractions 0.25 and 0.75. To calculate the weighted average mass, we must consider the mass and percentage of the 12-lb and 16-lb balls. We can proceed as follows.

$$\begin{aligned} 12\text{ lb: } 12\text{ lb} \times 0.25 &= 3\text{ lb} \\ 16\text{ lb: } 16\text{ lb} \times 0.75 &= \underline{12\text{ lb}} \\ & \quad 15\text{ lb} \end{aligned}$$

Although the weighted average mass is 15 lb, we should note that no shotput ball weighs 15 lb. An actual shotput ball weighs either 12 lb or 16 lb. A mass of 15 lb represents the theoretical mass of an average shotput ball.



▲ **Olympic Shotput** In the Olympic shotput event, athletes toss a steel ball.

Atomic Mass of an Element

The **atomic mass** of an element is the weighted average mass of all naturally occurring isotopes. To calculate the atomic mass, we will use a method similar to the one used in the shotput example. That is, to calculate the weighted average mass of an atom, we must consider the mass as well as the percent abundance of each isotope.

Carbon has two naturally occurring stable isotopes: carbon-12 and carbon-13. We can calculate the atomic mass of carbon given the mass and natural abundance of each isotope.

Isotope	Mass	Abundance
^{12}C	12.000 amu	98.89%
^{13}C	13.003 amu	1.11%

The natural abundance of each isotope, expressed as a decimal, is 0.9889 and 0.0111. To calculate the true weighted average mass of a carbon atom, we must consider the mass and percentage of each isotope. We proceed as follows.

$$^{12}C:\ 12.000\ \text{amu} \times 0.9889 = 11.87\ \text{amu}$$
$$^{13}C:\ 13.003\ \text{amu} \times 0.0111 = 0.144\ \text{amu}$$
$$12.01\ \text{amu}$$

Even though the atomic mass of carbon is 12.01 amu, we should note that no carbon atom weighs 12.01 amu. An actual carbon atom weighs either 12.000 amu or 13.003 amu. A mass of 12.01 amu represents the mass of a hypothetical average carbon atom. Example Exercise 5.3 further illustrates the calculation of atomic mass from isotopic mass and abundance data.

Example Exercise 5.3 • Calculation of Atomic Mass

Silicon is the second most abundant element in the Earth's crust. Calculate the atomic mass of silicon given the following data for its three natural isotopes.

Isotope	Mass	Abundance
^{28}Si	27.977 amu	92.21%
^{29}Si	28.976 amu	4.70%
^{30}Si	29.974 amu	3.09%

Solution

We can find the atomic mass of silicon as follows.

$$^{28}Si:\ 27.977\ \text{amu} \times 0.9221 = 25.80\ \text{amu}$$
$$^{29}Si:\ 28.976\ \text{amu} \times 0.0470 = 1.36\ \text{amu}$$
$$^{30}Si:\ 29.974\ \text{amu} \times 0.0309 = 0.926\ \text{amu}$$
$$28.09\ \text{amu}$$

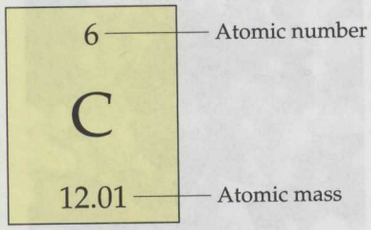
The average mass of a silicon atom is 28.09 amu, although we should note that there are no silicon atoms with a mass of 28.09 amu.

Self-Test Exercise

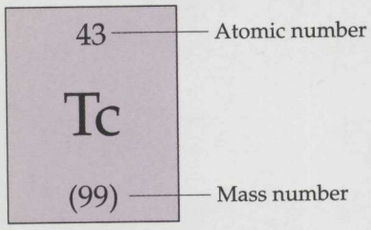
Calculate the atomic mass of copper given the following data.

Isotope	Mass	Abundance
^{63}Cu	62.930 amu	69.09%
^{65}Cu	64.928 amu	30.91%

Answer: 63.55 amu



▲ **Figure 5.8 Carbon from the Periodic Table of Elements** The atomic number designates the number of protons (6). The atomic mass designates the weighted average mass of the naturally occurring stable isotopes of the element.



▲ **Figure 5.9 Technetium from the Periodic Table of Elements** The atomic number designates the number of protons (43). The mass number for this radioactive isotope is 99. Technetium-99 has a total of 99 protons and neutrons in its nucleus.

The Periodic Table

We often need to refer to the atomic number and atomic mass of an element. For convenience, this information is listed in the periodic table. In this textbook the atomic number is indicated above the symbol, and the atomic mass is given below the symbol. Figure 5.8 illustrates carbon as it appears in the periodic table.

When we refer to the periodic table, it is not possible to determine the number of naturally occurring isotopes for a given element. Of the first 83 elements, however, 81 elements have one or more stable isotopes. Only technetium (element 43) and promethium (element 61) are radioactive and unstable. In the periodic table on the inside front cover of this textbook, the mass number of unstable elements is given in parentheses as shown in Figure 5.9.

The value in parentheses be
number of the most stable or b
refers to the total number of prot
periodic table you can easily di
that are unstable and radioactiv
in parentheses below the symbo
lustrates how to use the periodi

Example Exercise 5.4 • Nuclear

Refer to the periodic table on the i
ic number and atomic mass for iro

Solution

In the periodic table we observe

The atomic number of iron is 26, ar
table information, we should note t
topes for iron or their mass numbe

Self-Test Exercise

Refer to the periodic table on the in
number and mass number for the g

Answer: 86 and (222)

5.6 The Wave Nature of

Objectives · To explain the wave n
· To state the relations

We can use our imagination to vis
similar to an ocean wave. **Wavele**
to complete one cycle. **Frequency**
in 1 s. The velocity of light is const
3.00×10^8 m/s. Figure 5.10 illustr

As the wavelength of light de
Conversely, as the wavelength inc
the wavelength decreases from re
sualize a high-frequency light wa
ing many cycles per second. Cons
than low-frequency light, and sho
wavelength light.

Light—A Continuous Spectru

A rainbow is created when sunligl
as a miniature prism and separate

The value in parentheses below the symbol of the element indicates the mass number of the most stable or best known radioactive isotope. The mass number refers to the total number of protons and neutrons in the unstable nucleus. From the periodic table you can easily distinguish the elements that are stable from those that are unstable and radioactive. For unstable elements, there is a whole number in parentheses below the symbol of the element. Example Exercise 5.4 further illustrates how to use the periodic table as a reference.

Example Exercise 5.4 • Nuclear Composition from the Periodic Table

Refer to the periodic table on the inside cover of this textbook and determine the atomic number and atomic mass for iron.

Solution

In the periodic table we observe

26
Fe
55.85

The atomic number of iron is 26, and the atomic mass is 55.85 amu. From the periodic table information, we should note that it is not possible to determine the number of isotopes for iron or their mass numbers.

Self-Test Exercise

Refer to the periodic table on the inside cover of this text and determine the atomic number and mass number for the given radioactive isotope of radon gas.

Answer: 86 and (222)

5.6 The Wave Nature of Light

Objectives · To explain the wave nature of light.
· To state the relationship of wavelength, frequency, and energy of light.

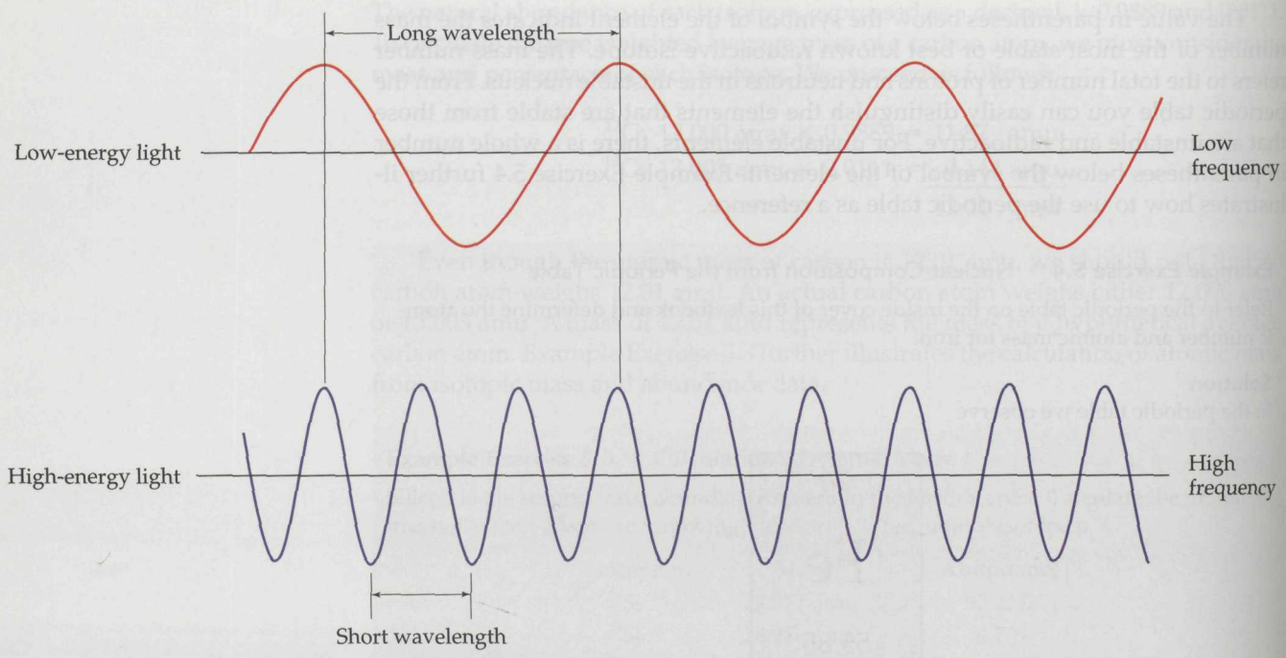
We can use our imagination to visualize light traveling through space in a fashion similar to an ocean wave. **Wavelength** refers to the distance the light wave travels to complete one cycle. **Frequency** refers to the number of wave cycles completed in 1 s. The velocity of light is constant, as all wavelengths and frequencies travel at 3.00×10^8 m/s. Figure 5.10 illustrates the wave nature of light.

As the wavelength of light decreases, the frequency of a light wave increases. Conversely, as the wavelength increases, the frequency decreases. As an example, the wavelength decreases from red to violet as the frequency increases. We can visualize a high-frequency light wave moving rapidly up and down while completing many cycles per second. Consequently, high-frequency light is more energetic than low-frequency light, and short-wavelength light is more energetic than long-wavelength light.

Light—A Continuous Spectrum

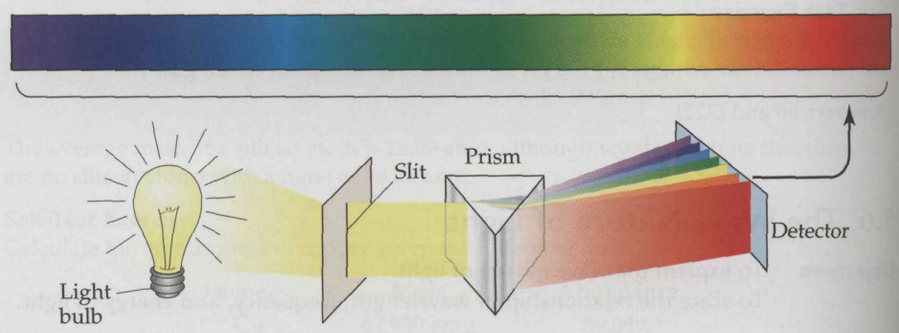
A rainbow is created when sunlight passes through raindrops. Each raindrop acts as a miniature prism and separates sunlight into various bands of color. What we

Wavelength and Frequency of Light Waves



▲ **Figure 5.10 Wavelength and Frequency of Light Waves** Notice that the wavelength is longer for low-energy light than for high-energy light. Also notice that the frequency is greater for high-energy light than for low-energy light.

White Light Passing Through a Prism



▶ **Figure 5.11 White Light Passing Through a Prism** Notice that white light separates into a rainbow of colors when it passes through a glass prism. Similarly, sunlight produces a rainbow when it passes through raindrops.

Electromagnetic Spectrum Activity

ordinarily observe as white light is actually several colors of light mixed together. When white light passes through a glass prism, it separates into all the colors of the rainbow, that is, red, orange, yellow, green, blue, and violet (Figure 5.11).

The term **light** usually refers to radiant energy that is visible. Our eyes can see light in the **visible spectrum** (400–700 nm) but not in the ultraviolet and infrared regions. The wavelength of ultraviolet radiation is too short to be visible (below 400 nm), and infrared radiation is too long (above 700 nm).

We also, on occasion, use the term "light" when referring to radiant energy that is not visible. The complete **radiant energy spectrum** is an uninterrupted band, or **continuous spectrum**, of visible and invisible light that ranges from short through long wavelengths. The radiant energy spectrum includes invisible gamma rays, X-rays, and microwaves, as well as visible light (Figure 5.12).

Example Exercise 5.5 illustrates how the wavelength, frequency, and energy of light are related.

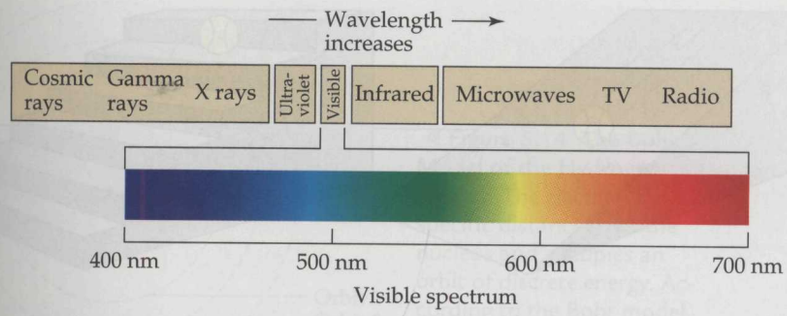


Figure 5.12 The Radiant Energy Spectrum The complete radiant energy spectrum includes short-wavelength gamma rays through long-wavelength microwaves. Notice that the visible spectrum is only a narrow window in a broad band of radiant energy.

The Radiant Energy Spectrum

Example Exercise 5.5 • Properties of Light

Considering blue light and yellow light, which has the

(a) longer wavelength? (b) higher frequency? (c) higher energy?

Solution

Referring to Figure 5.12, we notice that the wavelength of yellow light is about 600 nm and that of blue light, about 500 nm. Thus,

(a) *yellow light* has a longer wavelength than blue light.
(b) *blue light* has a higher frequency because it has a shorter wavelength.
(c) *blue light* has a higher energy because it has a higher frequency.

Self-Test Exercise

Considering infrared light and ultraviolet light, which has the

(a) longer wavelength? (b) higher frequency? (c) higher energy?

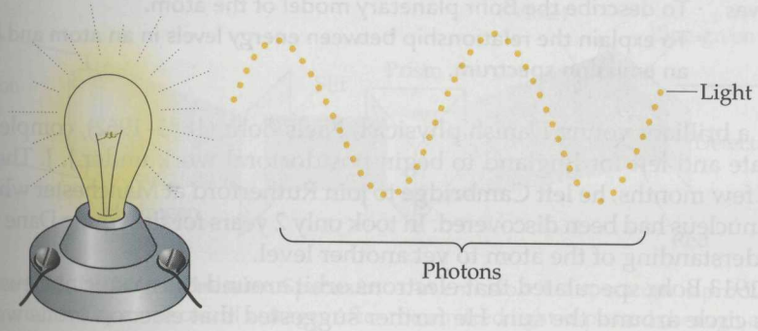
Answers: (a) infrared; (b) ultraviolet; (c) ultraviolet

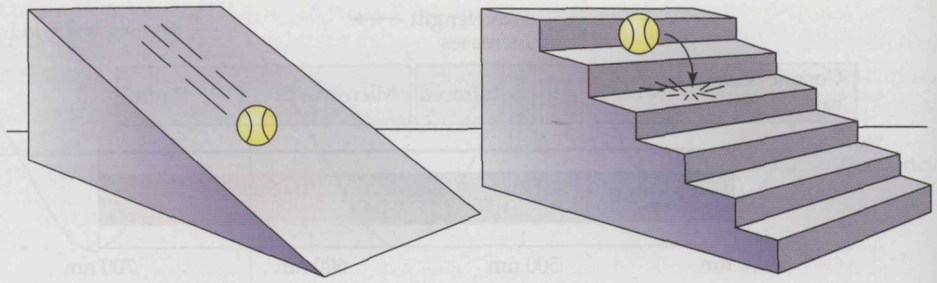
5.7 The Quantum Concept

Objective · To explain the quantum concept.

In 1900 Max Planck (1858–1947), a German physicist, introduced a revolutionary idea. Planck proposed that the energy radiated by a heated object is not continuous, but rather that the radiation is emitted in small bundles. The idea that energy is released in discrete units is referred to as the *quantum concept*.

When an object radiates light, it releases a unit of radiant energy called a **photon**. According to the quantum theory, a beam of light actually consists of a stream of individual particles. By way of example, an ordinary light bulb radiates energy in the form of tiny individual photons.





▶ **Figure 5.13 Stair Analogy for the Quantum Principle** A ball rolling down a ramp loses potential energy continuously. Conversely, a ball rolling down a flight of stairs loses potential energy in quantized amounts each time it drops from one step to another.

Stair Analogy for the Quantum Principle

We can illustrate the quantum concept with the following analogy. A ball rolling down a ramp loses potential energy continuously. Conversely, a ball rolling down a flight of stairs loses potential energy in discrete units each time it drops from one step to another. In this example, the loss of potential energy is continuous as the ball rolls down the ramp, and the loss is quantized as the ball rolls down the stairs. Figure 5.13 contrasts a continuous change in energy versus a quantized energy change.

The following Example Exercise 5.6 further illustrates practical applications of the quantum theory.

Example Exercise 5.6 • Quantum Concept

State whether each of the following scientific instruments gives a continuous or a quantized measurement of mass.

(a) triple-beam balance (b) digital electronic balance

Solution

Refer to Figure 2.3 if you have not used these balances in the laboratory.

(a) On a triple-beam balance a small metal rider is moved along a beam. Since the metal rider can be moved to any position on the beam, a triple-beam balance gives a *continuous* mass measurement.

(b) On a digital electronic balance the display indicates the mass of an object to a particular decimal place, for example, 5.015 g. Since the last digit in the display must be a whole number, a digital balance gives a *quantized* mass measurement.

Self-Test Exercise

State whether each of the following musical instruments produces continuous or quantized musical notes.

(a) acoustic guitar (b) electronic keyboard

Answers: (a) continuous; (b) quantized

5.8 Bohr Model of the Atom

Objectives
- To describe the Bohr planetary model of the atom.
- To explain the relationship between energy levels in an atom and lines in an emission spectrum.

▲ **Niels Bohr Stamp** The planetary model of the atom is illustrated for the model proposed by Niels Bohr.

In 1911 a brilliant young Danish physicist, Niels Bohr (1885–1962), completed his doctorate and left for England to begin postdoctoral work under J. J. Thomson. After a few months, he left Cambridge to join Rutherford at Manchester where the atomic nucleus had been discovered. In took only 2 years for the young Dane to raise our understanding of the atom to yet another level.

In 1913 Bohr speculated that electrons orbit around the atomic nucleus just as planets circle around the sun. He further suggested that electron orbits were at a fixed distance from the nucleus and had a definite energy. The electron was said to

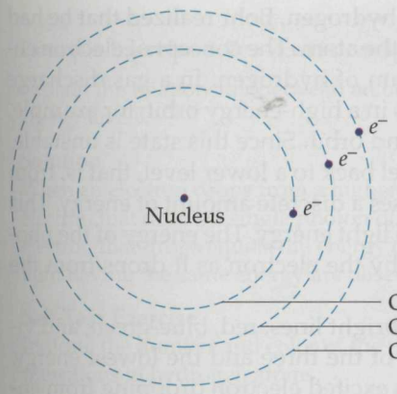


Figure 5.14 The Bohr Model of the Hydrogen Atom The electron is a specific distance from the nucleus and occupies an orbit of discrete energy. According to the Bohr model, the electron is found only in a given energy level.

The Bohr Model of the Atom

travel in a fixed-energy orbit referred to as an **energy level**. Moreover, electrons could be found only in specific energy levels and nowhere else. Figure 5.14 illustrates the model of the hydrogen atom proposed by Niels Bohr. This model is referred to as the Bohr planetary model of the atom, or simply, the **Bohr atom**.

Evidence for Energy Levels

The Bohr model was a beautiful mental picture of electrons in atoms. However, no one knew whether the model was right or wrong because there was no experimental evidence to support the theory. Coincidentally, Bohr received a paper on the emission of light from hydrogen gas. The paper showed that excited hydrogen gas emits separate emission lines of light rather than a continuous band of color.

An emission spectrum is produced when hydrogen gas is excited by an electrical voltage. To do so, hydrogen gas is sealed in a gas discharge tube and energized by electricity. The discharge tube then emits light, which separates into a series of narrow lines when passed through a prism. This collection of narrow bands of light energy is referred to as an **emission line spectrum**, and the individual bands of light are called spectral lines. The emission spectrum of hydrogen gas is shown in Figure 5.15.

Hydrogen Emission Spectrum

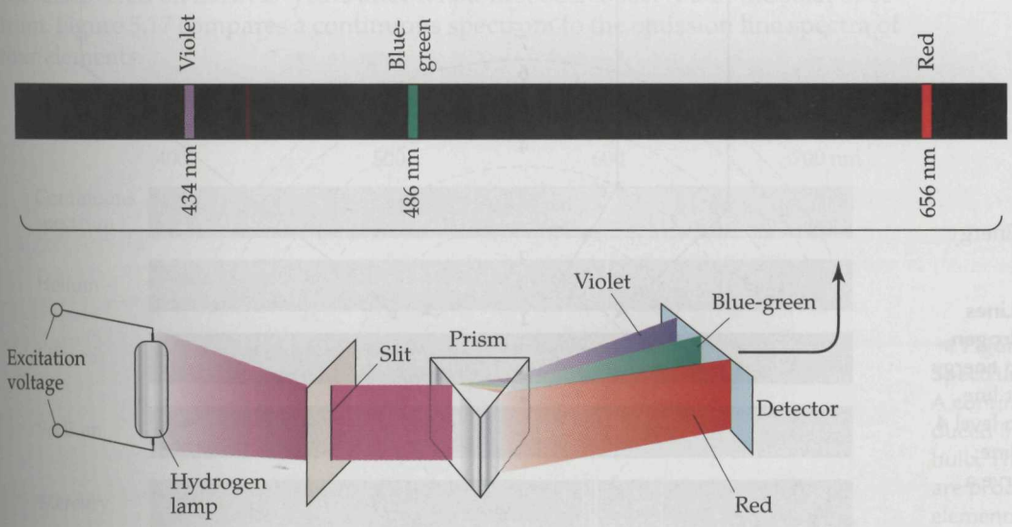
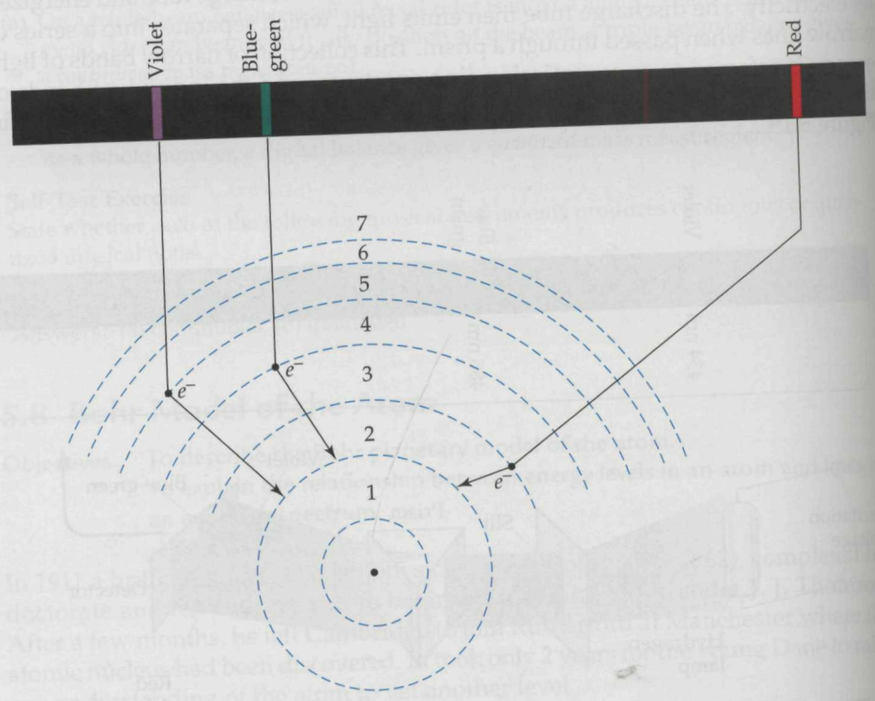


Figure 5.15 Hydrogen Emission Spectrum An emission line spectrum is produced when hydrogen gas is excited by an electrical voltage. After the emitted light is passed through a prism, three discrete vivid lines are observed.

After examining the emission spectrum of hydrogen, Bohr realized that he had experimental evidence to support his model of the atom. The concept of electron energy levels was supported by the line spectrum of hydrogen. In a gas discharge tube, excited atoms of hydrogen have electrons in a high-energy orbit; for example, the electron may temporarily occupy the second orbit. Since this state is unstable, the electron quickly drops from the higher level back to a lower level, that is, from level 2 to level 1. In the process, the electron loses a discrete amount of energy. This discrete energy loss corresponds to a photon of light energy. The energy of the photon of light equals the amount of energy lost by the electron as it drops from the higher to the lower energy level.

In the hydrogen spectrum, there are three bright lines: red, blue-green, and violet. The red line has the longest wavelength of the three and the lowest energy. Bohr found that the red line corresponds to an excited electron dropping from energy level 3 to 2. The blue-green line is more energetic than the red line and corresponds to an excited electron dropping from energy level 4 to 2. The violet line is the most energetic of the three and is produced when an electron drops from energy level 5 to 2. Figure 5.16 illustrates the correlation between Bohr's electron energy levels and the observed lines in the hydrogen spectrum.

Since a single photon is emitted each time an electron drops to a lower level, it follows that several photons are emitted when several electrons change levels. For example, if the electron drops from energy level 5 to 2 in ten hydrogen atoms, ten photons would be emitted. Each of the photons would have the same energy, and collectively they would be observed as a violet line in the emission spectrum. Example Exercise 5.7 further illustrates the relationship between energy levels and emission lines.



Spectral Lines and Energy Levels in Hydrogen

▲ **Figure 5.16 Spectral Lines and Energy Levels in Hydrogen** When electrons drop from energy level 5 to 2, we see a violet line. When electrons drop from level 4 to 2, we see a blue-green line; from level 3 to 2, we observe a red line.

Example Exercise 5.7 • E

Explain the relationship bet
energy levels.

Solution

When an electron drops fro
electron that drops, a single
electron that drops equals t
light having the same energ

Self-Test Exercise

Indicate the number and co
transitions in hydrogen ato

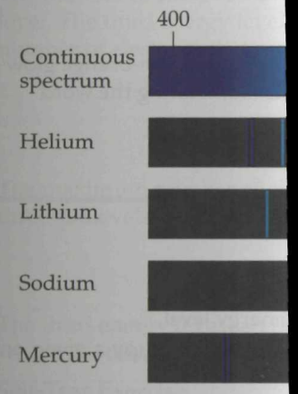
(a) 1 electron dropping fro
(b) 10 electrons dropping fr
(c) 100 electrons dropping
(d) 500 electrons dropping

Answers: (a) 1 red photon;
let photons

"Atomic Fingerprints"

Further study of emission s
of spectral lines. This obse
for atoms of each element.
"atomic fingerprint."

Atomic fingerprints a
in 1868 the atomic fingerp
from the Sun. This elemer
"sun." In 1895 an element
print identical to that obse
was discovered on Earth 27
trum. Figure 5.17 compare
four elements.



Example Exercise 5.7 • Emission Spectrum and Energy Levels

Explain the relationship between an observed emission line in a spectrum and electron energy levels.

Solution

When an electron drops from a higher to a lower energy level, light is emitted. For each electron that drops, a single photon of light energy is emitted. The energy lost by the electron that drops equals the energy of the photon that is emitted. Several photons of light having the same energy are observed as a spectral line.

Self-Test Exercise

Indicate the number and color of the photons emitted for each of the following electron transitions in hydrogen atoms.

(a) 1 electron dropping from energy level 3 to 2
(b) 10 electrons dropping from energy level 3 to 2
(c) 100 electrons dropping from energy level 4 to 2
(d) 500 electrons dropping from energy level 5 to 2

Answers: (a) 1 red photon; (b) 10 red photons; (c) 100 blue-green photons; (d) 500 violet photons

"Atomic Fingerprints"

Further study of emission spectra revealed that each element produced a unique set of spectral lines. This observation indicated that the energy levels must be unique for atoms of each element. Therefore, a line spectrum is sometimes referred to as an "atomic fingerprint."

Atomic fingerprints are useful in the identification of elements. For instance, in 1868 the atomic fingerprint of a new element was observed in the spectrum from the Sun. This element was named helium, after *helios*, the Greek word for "sun." In 1895 an element was discovered in uranium ore with an atomic fingerprint identical to that observed for helium in the Sun's spectrum. Thus, helium was discovered on Earth 27 years after it had first been observed in the solar spectrum. Figure 5.17 compares a continuous spectrum to the emission line spectra of four elements.

Experiment #6, Prentice Hall Laboratory Manual

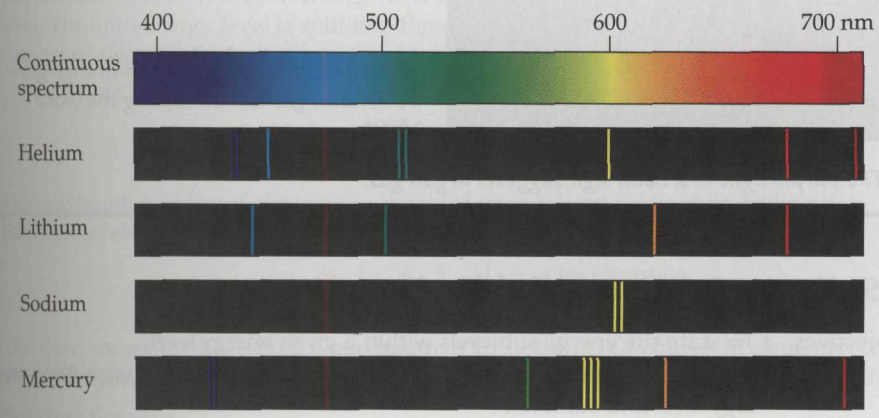


Figure 5.17 A Continuous Spectrum versus Line Spectra A continuous spectrum is produced from an ordinary light bulb. The emission line spectra are produced by excited atoms of elements In the gaseous state.

Chemistry Connection • Neon Lights

What element in the "neon" sign is emitting the purple light?

At the beginning of the twentieth century, J. J. Thomson discovered the electron using a cathode-ray tube. Thomson constructed his cathode-ray tube out of thin glass and placed a metal electrode in each end. After evacuating air from the glass tube, he introduced a small amount of gas. When the metal cathode and anode electrodes were electrically excited, he noticed that the tube glowed. He identified the glowing rays from the cathode as a stream of small negative particles. These particles were named electrons, and Thomson is given credit for their discovery. The glass tubes used by Thomson were forerunners of cathode-ray tubes, which are used today in television sets and computer monitor screens.

In 1913 Niels Bohr "explained" that exciting gases with electricity caused electrons to be temporarily promoted to higher energy states within the atom. The excited electrons, however, quickly lose energy by emitting light and returning to their original state. The light emitted by different gases varies because the energy levels within atoms vary for each element. For example, a gas discharge tube containing mercury vapor gives off a blue glow, whereas nitrogen gas gives off a yellowish-orange glow.

In 1898 the Scottish chemist William Ramsay discovered the noble gas neon. Unlike argon, which comprises about 1% of air, neon is much more rare. It is about a thousand times less concentrated in the atmosphere. When neon gas is placed in a narrow glass tube and electrically excited, it produces a reddish-orange light that is very arresting to the eye. The fact that gas discharge tubes produce an attractive array of colors led naturally to their use as advertising lights. Light from excited neon gas is very intense, and the term "neon light" has become a generic term for all lighted advertising displays.

Obviously, not all lights used in advertising are the same color. That is, "neon lights" can be red, green, blue, and so on. To produce a given color, a gas discharge tube must be filled with a specific gas. For a purple light, argon gas can be used, and for a pink light, helium gas is used. Only when we wish to produce a reddish-orange light is an advertising sign actually filled with neon gas.

◀ **Neon Light** The reddish-orange glow from neon gas is illuminating the word NEON.

The purple light in a neon sign suggests argon gas.

5.9 Energy Levels and Sublevels

Objectives
- To state the energy sublevels within a given energy level.
- To state the maximum number of electrons that can occupy a given energy level and sublevel.

As we learned in the previous section, in 1913 Niels Bohr proposed a model for the atom that pictured electrons circling around the nucleus in fixed-energy levels. His proposal was supported experimentally by the lines in the emission spectrum of

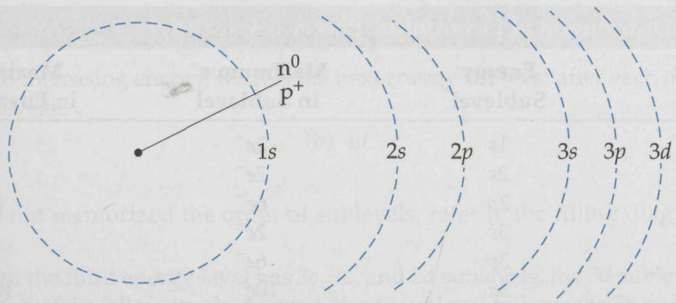


Figure 5.18 A Cross Section of an Atom The first energy level has only one sublevel (1*s*). The second energy level has two sublevels (2*s* and 2*p*). The third energy level has three sublevels (3*s*, 3*p*, and 3*d*). Although the diagram suggests that electrons travel in circular orbits, this is a simplification and is not actually the case.

Cross-Section of an Atom

hydrogen. The emission spectra of other elements, however, had far too many lines to interpret. Although Bohr could not explain the spectra of other elements, he did suggest the idea of sublevels within a main energy level. The model that eventually emerged had electrons occupying an **energy sublevel** within a main energy level. These energy sublevels were designated *s*, *p*, *d*, and *f*—in reference to the *sharp*, *principal*, *diffuse*, and *fine* lines, respectively, in the emission spectra of the elements.

The number of sublevels in each level corresponds to the number of the main energy level. That is, the first energy level (1) has one sublevel and is designated 1*s*. The second energy level (2) has two sublevels designated 2*s* and 2*p*. The third energy level (3) has three sublevels designated 3*s*, 3*p*, and 3*d*. The fourth energy level (4) is composed of 4*s*, 4*p*, 4*d*, and 4*f* sublevels (Figure 5.18).

The maximum number of electrons in each of the energy sublevels depends on the type of sublevel. That is, an *s* sublevel can hold a maximum of 2 electrons. A *p* sublevel can have a maximum of 6 electrons. A *d* sublevel can have 10 electrons, and an *f* sublevel can hold a maximum of 14 electrons.

To find the maximum number of electrons in a main energy level, we add up the electrons in each sublevel. The first energy level has one *s* sublevel; it can contain only $2e^-$. The second major energy level has two sublevels, 2*s* and 2*p*. The 2*s* can hold $2e^-$, and the 2*p* can hold $6e^-$. Thus, the second energy level can hold a maximum of $8e^-$. Example Exercise 5.8 further illustrates how energy levels, sublevels, and number of electrons are related.

Example Exercise 5.8 • Energy Levels, Sublevels, and Electrons

How many sublevels are in the third energy level. What is the maximum number of electrons that can occupy the third energy level?

Solution

The number of sublevels in an energy level corresponds to the number of the energy level. The third energy level is split into three sublevels: 3*s*, 3*p*, and 3*d*. The maximum number of electrons that can occupy a given sublevel is as follows.

$$s \text{ sublevel} = 2e^-$$
$$p \text{ sublevel} = 6e^-$$
$$d \text{ sublevel} = 10e^-$$

The maximum number of electrons in the third energy level is found by adding the three sublevels.

$$3s + 3p + 3d = \text{total electrons}$$
$$2e^- + 6e^- + 10e^- = 18e^-$$

The third energy level can hold a maximum of 18 electrons. Of course, in elements where the third energy level of an atom is not filled, there are fewer than 18 electrons.

Self-Test Exercise

How many sublevels are in the fourth energy level? What is the maximum number of electrons that can occupy the fourth energy level?

Answers: 4*s*, 4*p*, 4*d*, 4*f*; $32e^-$ ($2e^- + 6e^- + 10e^- + 14e^-$)

Table 5.4 Distribution of Electrons by Energy Level

Energy Level	Energy Sublevel	Maximum e^- in Sublevel	Maximum e^- in Energy Level
1	1*s*	$2e^-$	$2e^-$
2	2*s*	$2e^-$	
	2*p*	$6e^-$	$8e^-$
3	3*s*	$2e^-$	
	3*p*	$6e^-$	
	3*d*	$10e^-$	$18e^-$
4	4*s*	$2e^-$	
	4*p*	$6e^-$	
	4*d*	$10e^-$	
	4*f*	$14e^-$	$32e^-$

In summary, electrons are arranged around the nucleus in sublevels that have a specific fixed energy. Electrons farther from the nucleus occupy higher energy levels than those closer to the nucleus. Table 5.4 summarizes how main energy levels, sublevels, and number of electrons are related.

5.10 Electron Configuration

Objectives · To list the order of sublevels according to increasing energy.
· To write the predicted electron configurations for selected elements.

Electron Configurations Activity; Electron Configurations Movie

Electrons are arranged about the nucleus in a regular manner. The first electrons fill the energy sublevel closest to the nucleus. Additional electrons fill energy sublevels further and further from the nucleus. In other words, each energy level is filled sublevel by sublevel. The *s* sublevel is filled before a *p* sublevel, a *p* sublevel is filled before a *d* sublevel, and a *d* sublevel is filled before an *f* sublevel.

In general, sublevels are higher in energy as the energy level increases. Therefore, we would expect the order of sublevel filling to be 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 3*d*, 4*s*, and so on. This is not quite accurate because there are exceptions. For instance, the 4*s* sublevel is lower in energy than the 3*d* sublevel, and the 5*s* is lower than the 4*d* sublevel. A partial list of sublevels in order of increasing energy is $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s$.

In Chapter 6 we will learn how to predict the order in which sublevels are filled from the position of the element in the periodic table. In fact, the unusual shape of the periodic table reflects the order of sublevels according to increasing energy. For now, you should memorize the order of sublevel filling or refer to Figure 5.19.

Filling Diagram for Energy Sublevels

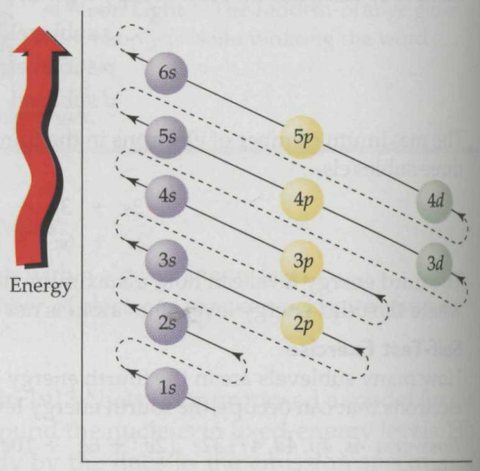


Figure 5.19 Filling Diagram for Energy Sublevels The order of sublevel filling is arranged according to increasing energy. Electrons first fill the 1*s* sublevel followed by the 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, and 6*s* sublevels.

Example Exercise 5.9 • Order of Sublevels

According to increasing energy, what is the next energy sublevel after each of the following sublevels?

(a) $3p$ (b) $4d$

Solution

If you have not memorized the order of sublevels, refer to the filling diagram in Figure 5.19.

(a) Although the third energy level has $3s$, $3p$, and $3d$ sublevels, the $3d$ sublevel does not immediately follow the $3p$. Instead, the $4s$ sublevel follows the $3p$ and precedes the $3d$. Thus,

$$3s, 3p, 4s$$

(b) Although the fourth energy level has $4s$, $4p$, $4d$, and $4f$ sublevels, the $4f$ sublevel does not immediately follow the $4d$. Instead, the $5p$ sublevel begins accepting electrons after the $4d$ is filled. Thus,

$$4p, 5s, 4d, 5p$$

Self-Test Exercise

Which sublevel gains electrons after each of the following sublevels is filled.

(a) $2s$ (b) $5p$

Answers: (a) $2p$; (b) $6s$

Electron Configuration

The **electron configuration** of an atom is a shorthand statement describing the location of electrons by sublevel. First, the sublevel is written, followed by a superscript indicating the number of electrons. For example, if the $2p$ sublevel contains two electrons, the standard notation is $2p^2$. Thus,

number of electrons

$$2p^2$$

energy level energy sublevel

Writing the electron configuration for an atom is a straightforward procedure. First, find the atomic number of the element in the periodic table; this corresponds to the number of electrons in a neutral atom. Then write the sublevels according to increasing energy. Each sublevel is filled with electrons in sequence until the total of all the superscripts equals the atomic number of the element.

As an example, let's write the electron configuration for iron. If we refer to the periodic table, we find that the atomic number of iron is 26. Given the atomic number, we know that an iron nucleus must have 26 protons and is surrounded by 26 electrons. The order in which the electron sublevels are arranged according to increasing energy is as follows.

$$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d \ldots$$

The energy sublevels in iron are filled beginning with the $1s$ sublevel and ending when there is a total of 26 electrons. The electron configuration for iron is as follows.

$$\text{Fe:}\quad 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$$

Notice that the sum of the superscripts equals the atomic number of iron (26). Example Exercise 5.10 illustrates the electron configuration for other elements.

Example Exercise 5.10 • Electron Configuration

Write the electron configuration for each of the following elements given the atomic number.

(a) Ne (b) Sr

Solution

We refer to the periodic table to find the atomic number of an element.

(a) The atomic number of neon is 10; therefore, the number of electrons is 10. We can fill sublevels until 10 electrons are present as follows.

$$\text{Ne:}\quad 1s^2\,2s^2\,2p^6$$

(b) From the periodic table, we find that the atomic number for strontium is 38. The number of electrons in a neutral atom of strontium is 38. Thus,

$$\text{Sr:}\quad 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2$$

To check your answer, find the total number of electrons by adding up the superscripts. The total is $38e^-$; this agrees with the atomic number for Sr.

Self-Test Exercise

Write the electron configuration for each of the following elements. Use standard notation, grouping electrons together according to sublevel.

(a) argon (b) cadmium

Answers: (a) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$; (b) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}$

▲ **Albert Einstein and Niels Bohr** The two famous scientists provided important insights to further our understanding of the atom.

5.11 Quantum Mechanical Model of the Atom

Objectives
- To describe the quantum mechanical model of the atom.
- To describe the relative sizes and shapes of *s* and *p* orbitals.

In the mid-1920s, a new model of the atom began to emerge. A more powerful theory evolved because the behavior of electrons could not be fully explained by the Bohr model of the atom. The German physicist Werner Heisenberg concluded that it was not possible to accurately determine both the position and energy of an electron. In his **uncertainty principle**, Heisenberg stated that it is impossible to precisely measure both the location and energy of a small particle simultaneously. In fact, the more accurately the position of an electron in an atom is known, the less precisely its energy can be determined.

In 1932 Heisenberg won the Nobel prize in physics for his uncertainty principle. Not everyone, however, subscribed to the principle of uncertainty. Some physicists found it unsettling to consider that they might live in a universe ruled by chance. Albert Einstein was sufficiently troubled by the uncertainty principle that he offered the famous quote: "It seems hard to look into God's cards but I cannot for a moment believe He plays dice as the current quantum theory alleges He does." Although the uncertainty principle was initially controversial, it was an essential contribution to the new view of the atom.

Gradually, the deeper nature of the atom came into focus. The new model retained the idea of quantized energy levels but incorporated the concept of uncertainty. The new model that emerged became known as the **quantum mechanical atom**. Recall that in the Bohr model the energy of an electron is defined in terms of a fixed-energy orbit about the nucleus. In the quantum mechanical model the en-

ergy of an electron can be describe
spatial volume surrounding the
for finding an electron of given en

Sizes and Shapes of *s* and *p* O

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We can describe the shapes of
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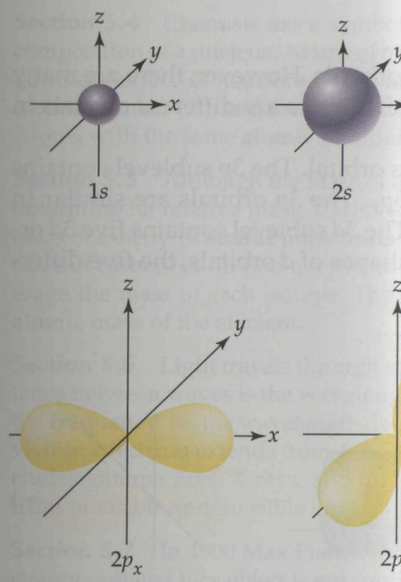
designate the size and shape of ar
its energy, and the letter that indic
2*p*, and 3*d* indicate three orbitals t
are spherical, but they are not all
a 2*s*, and a 2*s* orbital is larger than
crease as the energy level increase
s orbitals about the nucleus.

All *p* orbitals have the shape o
energy. A 3*p* orbital is larger than a
bital is said to resemble a dumbbel
can occupy either of the lobes.

There are three different 2*p* orb
in size and shape, they differ in th
orbitals intersect at the nucleus, bu
Figure 5.21 illustrates the relation



▲ **Figure 5.21 Orientation of 2*p* Orb**
The 2*p* orbitals do not have a fixed ori
tron in a $2p_x$ orbital has the same prob

ergy of an electron can be described in terms of the probability of it being within a spatial volume surrounding the nucleus. This region of high probability (~95%) for finding an electron of given energy is called an **orbital**.

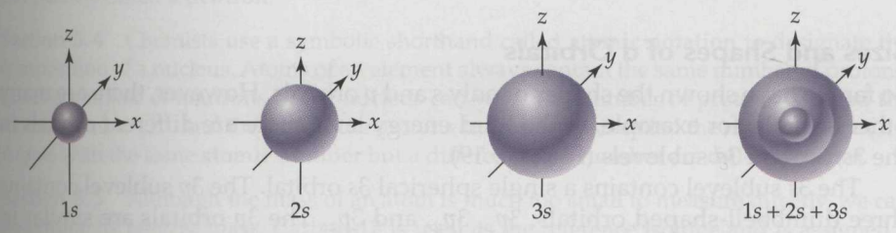
Sizes and Shapes of *s* and *p* Orbitals

In the quantum mechanical atom, orbitals are arranged about the nucleus according to their size and shape. In general, electrons having higher energy are found in larger orbitals. Similar to the energy levels in the Bohr atom, the energy of orbitals is quantized and assigned a whole-number value such as 1, 2, 3, 4, ... As the number increases, the energy and size of an orbital also increases.

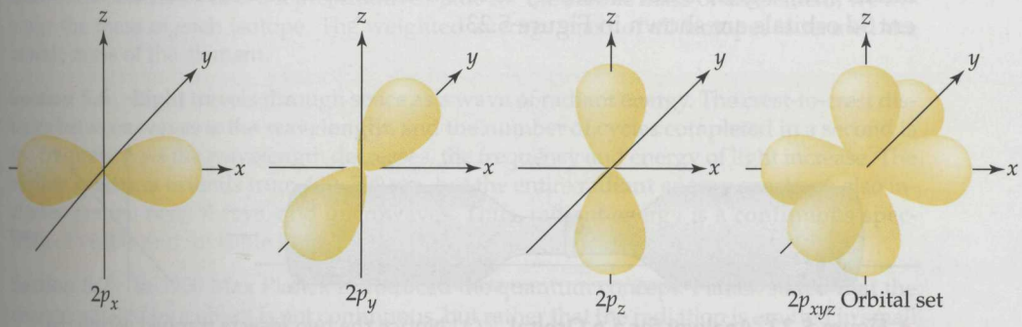
We can describe the shapes of orbitals by the letters *s*, *p*, *d*, and *f*. For example, the shape of an *s* orbital is that of a sphere, and the shape of a *p* orbital is that of a dumbbell. The shapes of *d* and *f* orbitals are too complex for our discussion. We can designate the size and shape of an orbital by combining the number that indicates its energy, and the letter that indicates its shape. For example, the designations 1*s*, 2*p*, and 3*d* indicate three orbitals that differ in size, energy, and shape. All *s* orbitals are spherical, but they are not all the same size. A 3*s* orbital is a larger sphere than a 2*s*, and a 2*s* orbital is larger than a 1*s*. That is, the size and energy of the orbital increase as the energy level increases. Figure 5.20 illustrates the relationship between *s* orbitals about the nucleus.

All *p* orbitals have the shape of a dumbbell, but they are not all equal in size or energy. A 3*p* orbital is larger than a 2*p* orbital and is at a higher energy level. A *p* orbital is said to resemble a dumbbell because it has two lobes. Electrons in a *p* orbital can occupy either of the lobes.

There are three different 2*p* orbitals. Although these three orbitals are identical in size and shape, they differ in their orientation to each other. That is, the three 2*p* orbitals intersect at the nucleus, but they are oriented at right angles to each other. Figure 5.21 illustrates the relationships between the $2p_x$, $2p_y$, and $2p_z$ orbitals. The



◀ **Figure 5.20 Relative Sizes of *s* Orbitals** The relative sizes of 1*s*, 2*s*, and 3*s* orbitals are shown. As the main energy level increases, the size and energy of the orbital increases. The nucleus of the atom is located in the center where the three axes intersect. The sketch on the far right illustrates the relationship of the 1*s*, 2*s*, and 3*s* orbitals.



▲ **Figure 5.21 Orientation of 2*p* Orbitals** The size and shape of the three 2*p* orbitals are identical. The 2*p* orbitals do not have a fixed orientation, but rather are perpendicular to each other. An electron in a $2p_x$ orbital has the same probability of occupying the $2p_y$ or $2p_z$ orbital.

p_x orbital is oriented along the x-axis of a three-dimensional axes system, and the p_y and p_z orbitals are oriented along the y-axis and z-axis, respectively.

Example Exercise 5.11 provides further practice in describing the relative sizes and shapes of orbitals.

Example Exercise 5.11 • Atomic Orbitals

Describe the relative size, energy, and shape for each of the following orbitals.

(a) 4*s* versus 3*s* and 5*s* (b) 4*p* versus 3*p* and 5*p*

Solution

The size and energy of an orbital is indicated by the number; the shape of the orbital is designated by the letter.

(a) Size and energy are greater for a 4*s* orbital than for a 3*s* orbital, but less than for a 5*s* orbital. The shape of a 4*s* orbital—and all *s* orbitals—is similar to the shape of a sphere.

(b) Size and energy are greater for a 4*p* orbital than for a 3*p* orbital but less than for a 5*p* orbital. The shape of a 4*p* orbital—and all *p* orbitals—is similar to the shape of a dumbbell.

Self-Test Exercise

Select the orbital in each of the following pairs that fits the description.

(a) the higher energy orbital: 3*p* or 4*p*

(b) the larger-size orbital: 4*d* or 5*d*

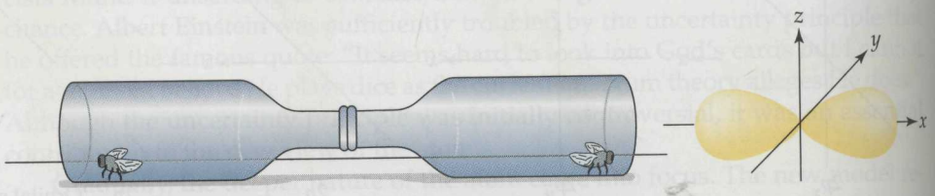
Answers: (a) 4*p*; (b) 5*d*

As an analogy, try to visualize a flying insect trapped inside two bottles with the open ends held together. The insect is free to fly about the entire inner volume of the two bottles. In this analogy, the insect represents an electron, and the two bottles represent the two lobes of a *p* orbital. Thus, there is a high probability of finding the electron anywhere within the volume of the entire *p* orbital. Figure 5.22 illustrates this analogy.

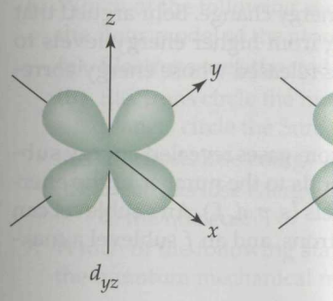
Sizes and Shapes of d Orbitals

So far we have shown the shapes of only *s* and *p* orbitals. However, there are many other orbitals; for example, in the third energy level there are different orbitals in the 3*s*, 3*p*, and 3*d* sublevels (Figure 5.19).

The 3*s* sublevel contains a single spherical 3*s* orbital. The 3*p* sublevel contains three dumbbell-shaped orbitals: $3p_x$, $3p_y$, and $3p_z$. The 3*p* orbitals are similar in shape, but 3*p* orbitals are larger than 2*p* orbitals. The 3*d* sublevel contains five 3*d* orbitals. Although it is not necessary to know the shapes of *d* orbitals, the five different 3*d* orbitals are shown in Figure 5.23.



▲ **Figure 5.22 Analogy for a *p* Orbital** (a) Notice the two insects trapped within the bottles held end to end. The two insects can both be in the left bottle, the right bottle, or one insect can be in each bottle. (b) Similarly, two electrons have a probability of being found anywhere within the two lobes of the *p* orbital.



▲ **Figure 5.23 Shapes of 3*d* C**
are similar in shape, they have a

Summary

Section 5.1 In 1803 John Dalt
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Section 5.2 Toward the end c
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Section 5.3 In 1911 Rutherfor
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Section 5.4 Chemists use a sy
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Section 5.5 Although the mas
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Section 5.6 **Light** travels throu
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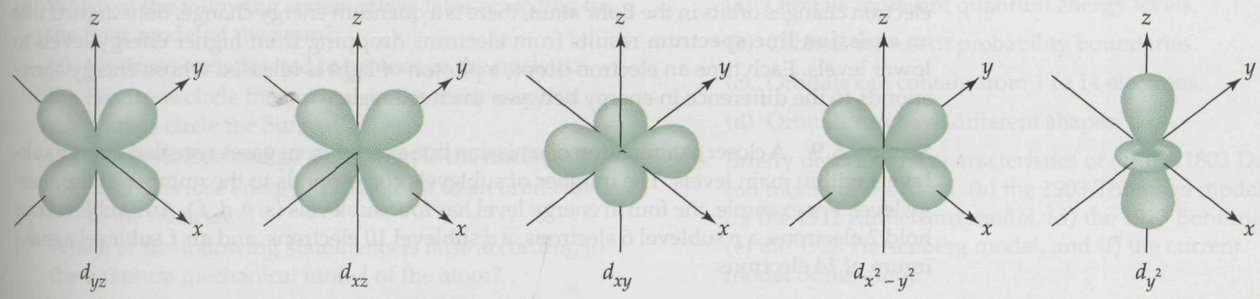


▲ **Figure 5.23 Shapes of 3*d* Orbitals** The shapes of the five 3*d* orbitals are not the same. Although three of the 3*d* orbitals are similar in shape, they have a different orientation to each other.

Summary

Section 5.1 In 1803 John Dalton proposed that matter consisted of atoms and supported the atomic theory with experiments on the behavior of gases and the laws of definite composition and conservation of mass.

Section 5.2 Toward the end of the 1800s, there was evidence that the atom was divisible. When electricity was applied to a sealed glass tube containing a gas at low pressure, negative and positive rays were observed. Scientists discovered that a **cathode ray** was composed of tiny negatively charged particles they called **electrons**. When the tube was filled with hydrogen gas, they found that a positive canal ray was composed of the smallest positively charged particles. These particles were named **protons**.

Section 5.3 In 1911 Rutherford performed a classic experiment in which alpha particles were fired at a thin sheet of gold foil. Much to his astonishment, some of the alpha particles bounced backward. He interpreted the results as evidence for a tiny, dense **atomic nucleus** at the center of the atom. The nucleus contains positively charged protons surrounded by negatively charged electrons. Twenty years later, the nucleus was also found to contain a neutral particle called a **neutron**.

Section 5.4 Chemists use a symbolic shorthand called **atomic notation** to designate the composition of a nucleus. Atoms of an element always contain the same number of protons, but the number of neutrons in the nucleus can vary. The number of protons is called the atomic number (Z), and the sum of the protons and neutrons is called the mass number (A). Atoms with the same **atomic number** but a different **mass number** are called **isotopes**.

Section 5.5 Although the mass of an atom is much too small to measure directly, we can determine its relative mass. Carbon-12 is used as the reference isotope and is assigned a mass of exactly 12 **atomic mass units (amu)**. The masses of all other atoms are related to the mass of carbon-12. To find a representative value for the atomic mass of an element, we average the mass of each isotope. The weighted average mass of all isotopes is termed the **atomic mass** of the element.

Section 5.6 **Light** travels through space as a wave of radiant energy. The crest-to-crest distance between waves is the **wavelength**, and the number of cycles completed in a second is the **frequency**. As the wavelength decreases, the frequency and energy of light increase. The **visible spectrum** extends from 400–700 nm, but the entire **radiant energy spectrum** also includes gamma rays, X rays, and microwaves. Thus, radiant energy is a **continuous spectrum** of visible and invisible light.

Section 5.7 In 1900 Max Planck introduced the quantum concept. Planck stated that the energy radiated by a object is not continuous, but rather that the radiation is emitted in small bundles. When an object radiates light, it releases a unit of radiant energy called a **photon**.

Section 5.8 In 1913 Niels Bohr suggested that electrons travel in circular orbits about the nucleus. The electron possesses a specific energy and is said to occupy an **energy level**. If an

electron changes orbits in the **Bohr atom**, there is a quantum energy change. Bohr argued that an **emission line spectrum** results from electrons dropping from higher energy levels to lower levels. Each time an electron drops, a photon of light is released whose energy corresponds to the difference in energy between the two levels.

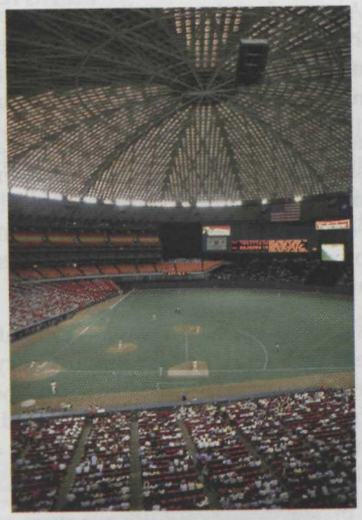
Section 5.9 A closer examination of emission line spectra from gases revealed **energy sublevels** within main levels. The number of sublevels corresponds to the number of the energy level. For example, the fourth energy level has four sublevels (s, p, d, f). An s sublevel can hold 2 electrons, a p sublevel 6 electrons, a d sublevel 10 electrons, and an f sublevel a maximum of 14 electrons.

Section 5.10 Electrons fill sublevels in order of increasing energy as follows: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s$. Notice that the $4s$ sublevel fills before the $3d$ and the $5s$ sublevel before the $4d$. A description of sublevel filling for an element is given by the **electron configuration**. A superscript following each sublevel indicates the number of electrons in a sublevel; for instance, $1s^2\,2s^2\,2p^6\,3s^1$ is the electron configuration for sodium (atomic number 11).

Section 5.11 In the 1920s our understanding of electrons in atoms became very sophisticated. In 1925 Werner Heisenberg suggested the **uncertainty principle**; that is, it is impossible to simultaneously know both the precise location and the energy of an electron. Instead, the energy of an electron can be known only in terms of its probability of being located somewhere within the atom. This description gave rise to the **quantum mechanical atom**. A location within the atom where there is a high probability of finding an electron having a certain energy is called an **orbital**. An orbital is a region about the nucleus having a given energy, size, and shape. The shape of an s orbital is spherical, and a p orbital resembles the shape of a dumbbell.

Key Concepts *

1. If an atom is magnified to the size of a golf ball, and a golf ball is equally magnified, what is the approximate size of the enlarged golf ball? (tennis ball, basketball, the Earth, the universe)
2. An atomic nucleus has been described by the analogy: "like a marble in the Astrodome." If a marble represents the atomic nucleus, what does the Astrodome represent?
3. The scattering of alpha particles by a thin gold foil has been described by the analogy: "like missiles shot through the solar system." If a missile represents an alpha particle, what do the planets represent?
4. Can atoms of different elements have the same atomic number? Can atoms of different elements have the same mass number?
5. Complete the following analogy. An ocean wave is to a drop of water as a light wave is to a _____.



◀ **Astrodome** Imagine how small a marble (nucleus) is compared to the Astrodome (atom).

* Answers to Key Concepts are in Appendix H.

6. Which of the following s
the Bohr model of the at
(a) Electrons are attracte
(b) Electrons circle the n
planets circle the Su
(c) Electrons lose energy
(d) Electrons lose energy
to the nucleus.
7. Which of the following s
the quantum mechanical

Key Terms †

Select the key term below th

_____ 1. a stream of negati
_____ 2. a negatively charg
_____ 3. a positively charg
_____ 4. a neutral subatom
_____ 5. a region in the cen
_____ 6. a value indicating
_____ 7. a value indicating
_____ 8. a symbolic metho
_____ 9. atoms of the same
_____ 10. a unit of mass exa
_____ 11. the average mass
_____ 12. the distance a ligh
_____ 13. the number of tim
_____ 14. a general term tha
_____ 15. a range of light en
_____ 16. a range of light en
_____ 17. a band of light en
_____ 18. a particle of radia
_____ 19. a model of the ato
_____ 20. a fixed-energy orb
_____ 21. a collection of nar
leasing energy
_____ 22. an electron energy
_____ 23. a shorthand descr
to increasing ener
_____ 24. the statement tha
of a particle at the
_____ 25. a sophisticated m
terms of its proba
_____ 26. a region about the
electron with a gi

† Answers to Key Terms are in A

6. Which of the following statements is false according to the Bohr model of the atom?
   (a) Electrons are attracted to protons in the nucleus.
   (b) Electrons circle the nucleus in the same way that planets circle the Sun.
   (c) Electrons lose energy as they circle the nucleus.
   (d) Electrons lose energy as they drop to an orbit closer to the nucleus.
7. Which of the following statements is false according to the quantum mechanical model of the atom?
   (a) Orbitals represent quantum energy levels.
   (b) Orbitals represent probability boundaries.
   (c) Orbitals can contain from 1 to 14 electrons.
   (d) Orbitals can have different shapes.
8. Briefly describe the characteristics of (a) the 1803 Dalton model of the atom, (b) the 1903 Thomson model, (c) the 1911 Rutherford model, (d) the 1913 Bohr model, (e) the 1926 Heisenberg model, and (f) the current model of the atom.

Key Terms†

Select the key term below that corresponds to each of the following definitions.

____ 1. a stream of negative particles produced in a cathode-ray tube
____ 2. a negatively charged subatomic particle having a negligible mass
____ 3. a positively charged subatomic particle having an approximate mass of 1 amu
____ 4. a neutral subatomic particle having an approximate mass of 1 amu
____ 5. a region in the center of an atom containing protons and neutrons
____ 6. a value indicating the number of protons in the nucleus of an atom
____ 7. a value indicating the number of protons and neutrons in the nucleus of an atom
____ 8. a symbolic method for expressing the composition of an atomic nucleus
____ 9. atoms of the same element that have a different number of neutrons
____ 10. a unit of mass exactly equal to 1/12 the mass of a carbon-12 atom
____ 11. the average mass of all the naturally occurring isotopes of an element
____ 12. the distance a light wave travels to complete one cycle
____ 13. the number of times a light wave completes a cycle in 1 second
____ 14. a general term that can refer to either visible or invisible radiant energy
____ 15. a range of light energy from violet through red, that is, 400–700 nm
____ 16. a range of light energy extending from gamma rays through microwaves
____ 17. a band of light energy that is uninterrupted
____ 18. a particle of radiant energy
____ 19. a model of the atom that describes electrons circling the nucleus in orbits
____ 20. a fixed-energy orbit that electrons occupy as they circle the nucleus
____ 21. a collection of narrow bands of light produced by atoms of a given element releasing energy
____ 22. an electron energy level resulting from splitting a main energy level
____ 23. a shorthand description of the arrangement of electrons by sublevels according to increasing energy
____ 24. the statement that it is impossible to precisely measure the location and energy of a particle at the same time
____ 25. a sophisticated model of the atom that describes the energy of an electron in terms of its probability of being found in a particular location about the nucleus
____ 26. a region about the nucleus in which there is a high probability of finding an electron with a given energy

(a) atomic mass (*Sec. 5.5*)
(b) atomic mass unit (amu) (*Sec. 5.5*)
(c) atomic notation (*Sec. 5.4*)
(d) atomic nucleus (*Sec. 5.3*)
(e) atomic number (Z) (*Sec. 5.4*)
(f) Bohr atom (*Sec. 5.8*)
(g) cathode ray (*Sec. 5.2*)
(h) continuous spectrum (*Sec. 5.6*)
(i) electron (e^-) (*Sec. 5.2*)
(j) electron configuration (*Sec. 5.10*)
(k) emission line spectrum (*Sec. 5.8*)
(l) energy level (*Sec. 5.8*)
(m) energy sublevel (*Sec. 5.9*)
(n) frequency (*Sec. 5.6*)
(o) isotopes (*Sec. 5.4*)
(p) light (*Sec. 5.6*)
(q) mass number (*A*) (*Sec. 5.4*)
(r) neutron (n^0) (*Sec. 5.3*)
(s) orbital (*Sec. 5.11*)
(t) photon (*Sec. 5.7*)
(u) proton (p^+) (*Sec. 5.2*)
(v) quantum mechanical atom (*Sec. 5.11*)
(w) radiant energy spectrum (*Sec. 5.6*)
(x) uncertainty principle (*Sec. 5.11*)
(y) visible spectrum (*Sec. 5.6*)
(z) wavelength (*Sec. 5.6*)

† Answers to Key Terms are in Appendix I.

Exercises‡

Dalton Model of the Atom (Sec. 5.1)

1. State Dalton's five proposals regarding the atomic theory.
2. State the two experimental laws Dalton used to support the atomic theory.
3. Which two of Dalton's proposals were later shown to be invalid?
4. Are atoms indestructible? Explain.

Thomson Model of the Atom (Sec. 5.2)

5. What was the simplest particle observed in cathode rays?
6. What was the simplest particle observed in canal rays?
7. What are the relative charges of the electron and the proton?
8. What are the relative masses of the electron and the proton?
9. What do the raisins represent in the plum-pudding analogy of the atom?
10. Where is the mass of an atom found according to the plum-pudding model?

Rutherford Model of the Atom (Sec. 5.3)

11. What did Rutherford conclude about the atom when alpha particles recoiled backward after striking a thin gold foil?
12. Describe an atom according to the Rutherford model.
13. State the location of electrons, protons, and neutrons in the Rutherford model of the atom.
14. State the approximate size of an atom and its nucleus in centimeters.
15. State the relative charges of the electron, proton, and neutron.
16. State the relative masses of the electron, proton, and neutron.

Atomic Notation (Sec. 5.4)

17. State the number of neutrons in an atom of each of the following isotopes.
    (a) $^{4}_{2}He$ (b) $^{32}_{16}S$
    (c) $^{10}_{5}B$ (d) $^{44}_{20}Ca$
18. State the number of neutrons in an atom of each of the following isotopes.
    (a) ^{15}N (b) ^{52}Cr
    (c) ^{26}Mg (d) ^{58}Ni
19. State the number of neutrons in an atom of each of the following isotopes.
    (a) lithium-7 (b) potassium-40
    (c) strontium-88 (d) platinum-195
20. State the number of neutrons in an atom of each of the following isotopes.
    (a) hydrogen-3 (b) cobalt-60
    (c) silicon-28 (d) iodine-131
21. Complete the following table and provide the missing information.

Atomic Notation	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
$^{4}_{2}He$					
$^{21}_{10}Ne$					
$^{50}_{22}Ti$					
$^{197}_{79}Au$					

22. Complete the following table and provide the missing information.

Atomic Notation	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
$^{A}_{Z}Se$		78			
$^{A}_{Z}X$	38			50	
$^{A}_{Z}Sn$		120			
$^{A}_{Z}X$	54			77	

23. Draw a diagram of the arrangement of protons, neutrons, and electrons in an atom of each of the following isotopes.
    (a) $^{7}_{3}Li$ (b) $^{13}_{6}C$
    (c) $^{16}_{8}O$ (d) $^{20}_{10}Ne$
24. Draw a diagram of the arrangement of protons, neutrons, and electrons in an atom of each of the following isotopes.
    (a) ^{31}P (b) ^{35}Cl
    (c) ^{40}Ar (f) ^{131}I

Atomic Mass (Sec. 5.5)

25. What is the reference isotope for the atomic mass scale?
26. What is the assigned mass for the reference isotope?
27. Why are atomic masses expressed on a *relative* atomic mass scale?
28. Distinguish between isotopic mass and atomic mass.
29. Given that the only naturally occurring isotope of aluminum is ^{27}Al, determine its mass from the periodic table.
30. Given that the only naturally occurring isotope of phosphorus is ^{31}P, determine its mass from the periodic table.
31. A marble collection has 100 large marbles with a mass of 5.0 g each and 200 small marbles with a mass of 2.0 g each. Calculate (a) the simple average mass, and (b) the weighted average mass of the marble collection.
32. The final grade in a chemistry class is based on homework (10%), quizzes (10%), tests (40%), experiments (20%), and a final exam (20%). What is the weighted av-

‡Answers to odd-numbered Exercises are in Appendix J.

erage score of a stude
homework, 95; quizze
and final exam, 68?

33. Calculate the atomic r
ing data for its natura

^{6}Li:	6.(
^{7}Li:	7.(

34. Calculate the atomic r
lowing data for its nat

^{24}Mg:	23
^{25}Mg:	24
^{26}Mg:	25

35. Calculate the atomic n
data for its natural iso

^{54}Fe:	53.
^{56}Fe:	55.
^{57}Fe:	56.
^{58}Fe:	57.

36. Calculate the atomic n
data for its natural isot

^{64}Zn:	63.
^{66}Zn:	65.
^{67}Zn:	66.
^{68}Zn:	67.
^{70}Zn:	69.

37. Chlorine has two natu
^{37}Cl. Which isotope is r
of chlorine is 35.45 amu
38. Bromine has two natur
mately equal abundan
what is the other isotop
is 79.90 amu?

The Wave Nature of Lig

39. Which of the following
or orange light?
40. Which of the following
or red light?
41. Which of the following
let, green, or orange lig

erage score of a student with the following grades: homework, 95; quizzes, 83; tests, 75; experiments, 92; and final exam, 68?

33. Calculate the atomic mass for lithium given the following data for its natural isotopes.

^{6}Li:	6.015 amu	7.42%
^{7}Li:	7.016 amu	92.58%

34. Calculate the atomic mass for magnesium given the following data for its natural isotopes.

^{24}Mg:	23.985 amu	78.70%
^{25}Mg:	24.986 amu	10.13%
^{26}Mg:	25.983 amu	11.17%

35. Calculate the atomic mass for iron given the following data for its natural isotopes.

^{54}Fe:	53.940 amu	5.82%
^{56}Fe:	55.935 amu	91.66%
^{57}Fe:	56.935 amu	2.19%
^{58}Fe:	57.933 amu	0.33%

36. Calculate the atomic mass for zinc given the following data for its natural isotopes.

^{64}Zn:	63.929 amu	48.89%
^{66}Zn:	65.926 amu	27.81%
^{67}Zn:	66.927 amu	4.11%
^{68}Zn:	67.925 amu	18.57%
^{70}Zn:	69.925 amu	0.62%

37. Chlorine has two naturally occurring isotopes: ^{35}Cl and ^{37}Cl. Which isotope is more abundant if the atomic mass of chlorine is 35.45 amu?

38. Bromine has two natural isotopes that occur in approximately equal abundance. If ^{79}Br is one of the isotopes, what is the other isotope if the atomic mass of bromine is 79.90 amu?

The Wave Nature of Light (Sec. 5.6)

39. Which of the following is most energetic: violet, green, or orange light?

40. Which of the following is least energetic: blue, yellow, or red light?

41. Which of the following has the longest wavelength: violet, green, or orange light?

42. Which of the following has the shortest wavelength: blue, yellow, or red light?

43. Which of the following wavelengths of light is most energetic: 650 nm, 550 nm, or 450 nm?

44. Which of the following wavelengths of light is least energetic: 425 nm, 525 nm, or 625 nm?

45. Which of the following wavelengths of light has the highest frequency: 650 nm, 550 nm, or 450 nm?

46. Which of the following wavelengths of light has the lowest frequency: 425 nm, 525 nm, or 625 nm?

The Quantum Concept (Sec. 5.7)

47. What is the quantum particle in light energy?

48. What is the quantum particle in electrical energy?

49. State whether each of the following is continuous or quantized.
    (a) a rainbow (b) a line spectrum

50. State whether each of the following is continuous or quantized.
    (a) a spiral staircase (b) an elevated ramp

51. State whether each of the following instruments gives a continuous or a quantized measurement of length.
    (a) a metric ruler (b) a digital laser

52. State whether each of the following instruments gives a continuous or a quantized measurement of volume.
    (a) 10-mL volumetric pipet
    (b) 10-mL graduated cylinder

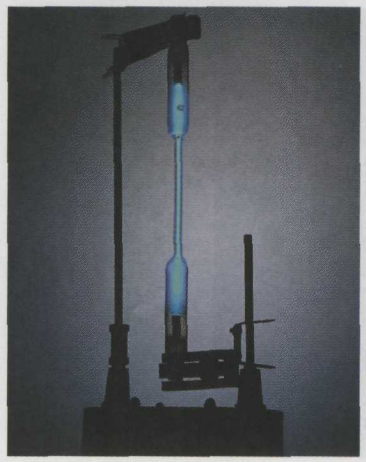
Bohr Model of the Atom (Sec. 5.8)

53. Draw the Bohr model of the atom.

54. What is the experimental evidence for electron energy levels in an atom?

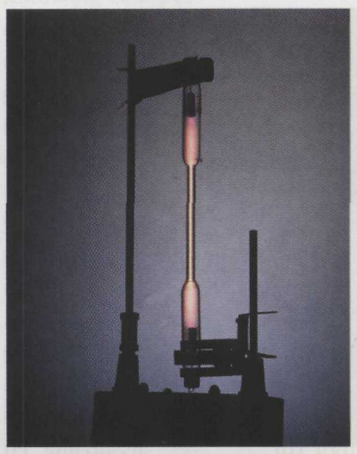
55. Which of the following energy level changes for an electron is most energetic: $5 \rightarrow 2$, $4 \rightarrow 2$, or $3 \rightarrow 2$?

56. Which of the following energy level changes for an electron is least energetic: $4 \rightarrow 1$, $3 \rightarrow 1$, or $\rightarrow 1$?



▲ **Mercury Light** The blue glow from mercury vapor illuminates the gas discharge tube.

57. In a hydrogen atom, what color is the emission line observed when the electron drops from the fourth to the second level?
58. In a hydrogen atom, what color is the emission line observed when the electron drops from the fifth to the second level?
59. An electron in a hydrogen atom drops from the fifth energy level to which lower level to emit an ultraviolet photon?
60. An electron in a hydrogen atom drops from the fifth energy level to which lower level to emit an infrared photon?
61. Which of the following lines in the emission spectrum of hydrogen is most energetic: red, blue-green, or violet?
62. Which of the following lines in the emission spectrum of hydrogen has the longest wavelength: red, blue-green, or violet?
63. How many photons of light are emitted for each of the following?
    (a) $1e^-$ drops from energy level 3 to 1
    (b) $1e^-$ drops from energy level 3 to 2
64. How many photons of light are emitted for each of the following?
    (a) $100e^-$ drop from energy level 3 to 2
    (b) $100e^-$ drop from energy level 4 to 2
65. What is the color of the spectral line emitted for each of the following electron energy changes in excited hydrogen gas?
    (a) Electrons drop from energy level 2 to 1.
    (b) Electrons drop from energy level 3 to 2.
    (c) Electrons drop from energy level 4 to 3.
66. What is the color of the spectral line emitted for each of the following electron energy changes in excited hydrogen gas?
    (a) Electrons drop from energy level 5 to 1.
    (b) Electrons drop from energy level 5 to 2.
    (c) Electrons drop from energy level 5 to 3.



▲ **Nitrogen Light** The yellowish-orange glow from nitrogen gas illuminates the gas discharge tube.

Energy Levels and Sublevels (Sec. 5.9)

67. What experimental evidence suggests the concept of electrons in energy levels?
68. What experimental evidence suggests main energy levels split into sublevels?
69. Designate all the sublevels within each of the following energy levels.
    (a) 1st (b) 2nd
    (c) 3rd (d) 4th
70. State the number of sublevels in each of the following energy levels.
    (a) 1st (b) 3rd
    (c) 5th (d) 6th
71. What is the maximum number of electrons in each of the following sublevels?
    (a) $2s$ (b) $4p$
    (c) $3d$ (d) $5f$
72. What is the maximum number of electrons in each of the following?
    (a) an s sublevel (b) a p sublevel
    (c) a d sublevel (d) an f sublevel
73. What is the maximum number of electrons in the second energy level?
74. What is the maximum number of electrons in the fourth energy level?

Electron Configuration (Sec. 5.10)

75. List the order of sublevels from $1s$ through $5p$ according to increasing energy. (*Hint:* Draw a filling diagram.)
76. Draw a filling diagram and predict the sublevel that follows $5p$.
77. Write the predicted electron configuration for each of the following elements.
    (a) He (b) Be
    (c) Co (d) Cd
78. Write the predicted electron configuration for the following elements.
    (a) boron (b) argon
    (c) manganese (d) nickel
79. Which element corresponds to each of the following electron configurations?
    (a) $1s^2\,2s^1$
    (b) $1s^2\,2s^2\,2p^6\,3s^2\,3p^2$
    (c) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^2$
    (d) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2$
80. Which element corresponds to each of the following electron configurations?
    (a) $1s^2\,2s^2\,2p^5$
    (b) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$
    (c) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^5$
    (d) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^5$

Quantum Mechanical Mod

81. What is the distinction be
82. What are two significant c model of the atom and the
83. Sketch a three-dimensiona the following orbitals. Lab
    (a) $1s$
    (c) $3p_y$
84. Sketch a three-dimensiona the following orbital sets.
    (a) $1s, 2s, 2p_x$
85. Which orbital in each of th higher energy?
    (a) $2s$ or $3s$
    (c) $2p_x$ or $2p_y$
86. Which orbital in each of th larger size?
    (a) $2s$ or $3s$
    (c) $2p_x$ or $2p_y$
87. Designate the orbital that descriptions.
    (a) a spherical orbital in th
    (b) a dumbbell-shaped orb
88. Designate the orbital that descriptions.
    (a) a spherical orbital in th
    (b) a dumbbell-shaped orb
89. State the maximum number each of the following orbita
    (a) $1s$
    (c) $3d$
90. State the maximum numbe py each of the following su
    (a) $1s$
    (c) $3d$

(Sec. 5.9)
e suggests the concept of
e suggests main energy lev-
within each of the following
) 2nd
) 4th
ls in each of the following
) 3rd
) 6th
ber of electrons in each of
) $4p$
) $5f$
ber of electrons in each of
) a p sublevel
) an f sublevel
ber of electrons in the sec-
ber of electrons in the fourth
5.10)
om 1s through 5p according
Draw a filling diagram.)
predict the sublevel that
configuration for each of
) Be
) Cd
configuration for the fol-
) argon
) nickel
to each of the following
$4p^6\ 5s^2$
to each of the following
$4p^6\ 5s^2\ 4d^5$
$4p^6\ 5s^2\ 4d^{10}\ 5p^5$

Quantum Mechanical Model of the Atom (Sec. 5.11)

81. What is the distinction between an orbit and an orbital?
82. What are two significant differences between the Bohr model of the atom and the quantum mechanical model?
83. Sketch a three-dimensional representation for each of the following orbitals. Label the x-, y-, and z-axes.
    (a) $1s$ (b) $2p_x$
    (c) $3p_y$ (d) $4p_z$
84. Sketch a three-dimensional representation for each of the following orbital sets. Label the x-, y-, and z-axes.
    (a) $1s, 2s, 2p_x$ (b) $3p_x, 3p_y, 3p_z$
85. Which orbital in each of the following pairs has the higher energy?
    (a) $2s$ or $3s$ (b) $2p_x$ or $3p_x$
    (c) $2p_x$ or $2p_y$ (d) $4p_y$ or $4p_z$
86. Which orbital in each of the following pairs has the larger size?
    (a) $2s$ or $3s$ (b) $2p_x$ or $3p_x$
    (c) $2p_x$ or $2p_y$ (d) $4p_y$ or $4p_z$
87. Designate the orbital that fits each of the following descriptions.
    (a) a spherical orbital in the fifth energy level
    (b) a dumbbell-shaped orbital in the fourth energy level
88. Designate the orbital that fits each of the following descriptions.
    (a) a spherical orbital in the sixth energy level
    (b) a dumbbell-shaped orbital in the third energy level
89. State the maximum number of electrons that can occupy each of the following orbitals.
    (a) $1s$ (b) $2p$
    (c) $3d$ (d) $4f$
90. State the maximum number of electrons that can occupy each of the following sublevels.
    (a) $1s$ (b) $2p$
    (c) $3d$ (d) $4f$

General Exercises

91. If the electron charge-to-mass ratio is 1.76×10^8 coulomb/g and the absolute charge is 1.60×10^{-19} coulomb, what is the mass of an electron in grams?
92. If the electron charge-to-mass ratio is 9.57×10^4 coulomb/g and the absolute charge is 1.60×10^{-19} coulomb, what is the mass of a proton in grams?
93. Gallium occurs naturally as ^{69}Ga and ^{71}Ga. Given the mass and abundance of ^{69}Ga (68.926 amu and 60.10%), what is the mass of ^{71}Ga?
94. Boron (atomic mass 10.811 amu) occurs naturally as ^{10}B and ^{11}B. Given the mass of the two isotopes (10.013 amu and 11.009 amu), what is the percentage abundance of each isotope?
95. Element 61 was named for the mythological Prometheus who stole fire from heaven. Refer to the periodic table and state whether Pm has any stable isotopes.
96. Element 84 was discovered by Marie Curie and named polonium for her native Poland. Refer to the periodic table and state whether Po is stable or radioactive.
97. Indicate the region of the spectrum (infrared, visible, or ultraviolet) for each of the following wavelengths of light.
    (a) 200 nm (b) 500 nm
    (c) 1200 nm
98. Which of the following light frequencies has the higher energy, 5×10^{10} cycles/s or 5×10^{11} cycles/s?
99. Explain why the electron configuration for copper is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$ rather than the predicted $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$.
100. Explain why the electron configuration for silver is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1\ 4d^{10}$ rather than the predicted $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^9$.

Explorer Quiz 1
Explorer Quiz 2
Explorer Quiz 3
Master Quiz