CHAPTER OUTLINE

CHAPTER THREE

Atomic Structure, the Periodic Table, and Radioactivity

Music consists of a series of tones that build octave after octave. Similarly, elements have properties that recur period after period.

In Chapter 1 we learned that all matter is made up of small particles called atoms and that 113 different types of atoms are known, each type of atom corresponding to a different element. Furthermore, we found that compounds result from the chemical combination of different types of atoms in various ratios and arrangements.

Until the last two decades of the nineteenth century, scientists believed that atoms were solid, indivisible spheres without an internal structure. Today, this model of the atom is known to be incorrect. Evidence from a variety of sources indicates that atoms are made up of even smaller particles called *subatomic particles*. In this chapter we consider the fundamental types of subatomic particles, how they arrange themselves within an atom, and the relationship between an atom's subatomic makeup and its chemical identity.

Learning Focus

Describe the internal structure of an atom in terms of subatomic particles present and give the fundamental properties for each of the types of subatomic particles.

3.1 Internal Structure of an Atom

Atoms possess internal structure; that is, they are made up of even smaller particles, which are called subatomic particles. **Subatomic particles** *are very small particles that are the building blocks from which atoms are made*. Three types of subatomic particles are found within atoms: electrons, protons, and neutrons. Key properties of these three types of particles are summarized in Table 3.1. An **electron** *is a subatomic particle that possesses a negative (−) electrical charge*. It is the smallest, in terms of mass, of the three types of subatomic particles. A **proton** *is a subatomic particle that possesses a positive (+) electrical charge*. Protons and electrons carry the *same amount*

Table 3.1
Charge and Mass Characteristics of Electrons, Protons, and Neutrons

	Electron	Proton	Neutron
Charge	-1	$+1$	0
Actual mass (g)	9.109×10^{-28}	1.673×10^{-24}	1.675×10^{-24}
Relative mass (based on the electron being 1 unit)	1	1837	1839

▶ Atoms of all 113 elements contain the same three types of subatomic particles. Different elements differ only in the numbers of the various subatomic particles they contain.

of charge; the charges, however, are opposite (positive versus negative). A **neutron** *is a subatomic particle that has no charge associated with it; that is, it is neutral.* Both protons and neutrons are massive particles compared to electrons; they are almost 2000 times heavier.

Arrangement of Subatomic Particles Within an Atom

The arrangement of subatomic particles within an atom is not haphazard. *All* protons and *all* neutrons present are found at the center of an atom in a very tiny volume called the *nucleus* (Figure 3.1). The **nucleus** *is the very small, dense, positively charged center of an atom.* A nucleus is always positively charged because it contains positively charged protons. Because the nucleus houses the heavy subatomic particles (protons and neutrons), almost all (over 99.9%) of the mass of an atom is concentrated in its nucleus. The small size of the nucleus, coupled with its large amount of mass, causes nuclear material to be extremely dense.

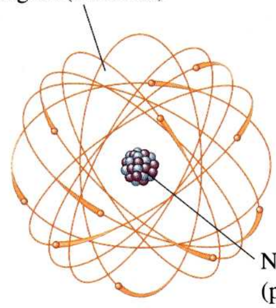


Figure 3.1 The protons and neutrons of an atom are found in the central nuclear region, or nucleus, and the electrons are found in an electron cloud outside the nucleus. Note that this figure is not drawn to scale; the correct scale would be comparable to a penny (the nucleus) in the center of a baseball field (the atom).

Closely resembling the term *nucleus* is the term *nucleon.* A **nucleon** *is any subatomic particle found in the nucleus of an atom.* Thus both protons and neutrons are nucleons, and the nucleus can be regarded as containing a collection of nucleons (protons and neutrons).

The outer (extranuclear) region of an atom contains all of the electrons. In this region, which accounts for most of the volume of an atom, the electrons move rapidly about the nucleus. The electrons are attracted to the positively charged protons of the nucleus by the forces that exist between particles of opposite charge. The motion of the electrons in the extranuclear region determines the volume (size) of the atom in the same way that the blade of a fan determines a volume by its circular motion. The volume occupied by the electrons is sometimes referred to as the *electron cloud.* Because electrons are negatively charged, the electron cloud is also negatively charged. Figure 3.1 illustrates the nuclear and extranuclear regions of an atom.

Charge Neutrality of an Atom

An atom as a whole is electrically neutral; that is, it has no *net* electrical charge. For this to be the case, the same amount of positive and negative charge must be present in the atom. Equal amounts of positive and negative charges cancel one another. Thus equal numbers of protons and electrons are present in an atom.

$$\text{Number of protons} = \text{number of electrons}$$

Size Relationships Within an Atom

To help you visualize the size relationships among the parts of an atom, imagine enlarging (magnifying) the nucleus until it is the size of a baseball (about 2.9 inches in diameter). If the nucleus were this large, the whole atom would have a diameter of approximately 2.5 miles. The electrons would still be smaller than the periods used to end sentences in this text, and they would move about at random within that 2.5-mile region.

The concentration of nearly all of the mass of an atom in the nucleus can also be illustrated by using our imagination. If a coin the same size as a copper penny contained copper nuclei (copper atoms stripped of their electrons) rather than copper atoms (which are mostly empty space), the coin would weigh 190,000,000 tons! Nuclei are indeed very dense matter.

Despite the existence of subatomic particles, we will continue to use the concept of atoms as the fundamental building blocks for all types of matter. Subatomic particles do not lead an independent existence for any appreciable length of time; they gain stability by joining together to form atoms.

Practice Questions and Problems

3.1 Indicate whether each of the following statements describes a *proton,* an *electron,* or a *neutron.*

a. Possesses a negative charge
b. Has no charge
c. Has a charge equal to, but opposite in sign from, that of an electron
d. Has a mass slightly less than that of a neutron

3.2 Describe the composition of the nucleus of an atom and its location relative to the atom as a whole.

3.3 Why does the nucleus of every atom have a positive charge?

3.4 What is the relationship between the number of protons and the number of electrons in an atom?

3.5 Indicate whether each of the following statements about the nucleus of an atom is *true* or *false.*

a. The nucleus of an atom contains all of the "heavy" subatomic particles.
b. The nucleus of an atom contains only neutrons.
c. The nucleus of an atom accounts for almost all of the volume of the atom.
d. The nucleus of an atom contains all of the subatomic particles present in an atom.

3.2 Atomic Number and Mass Number

Learning Focus

Define the terms *atomic number* and *mass number* and, given these two numbers, know how to determine the number of protons, neutrons, and electrons present in an atom.

The **atomic number** *of an atom is the number of protons in its nucleus.* Because an atom has the same number of electrons as protons (Section 3.1), the atomic number also specifies the number of electrons present. The symbol Z is used as a general designation for the atomic number.

$$\text{Atomic number} = \text{number of protons} = \text{number of electrons} = Z$$

The **mass number** *of an atom is the sum of the numbers of protons and neutrons in its nucleus.* Thus the mass number gives the number of subatomic particles present in the nucleus. The mass of an atom is almost totally accounted for by the protons and neutrons present—hence the term *mass number.* The symbol A is used as a general designation for the mass number.

► Atomic number and mass number are always *whole* numbers because they are obtained by counting whole objects (protons, neutrons, and electrons).

$$\text{Mass number} = \text{number of protons} + \text{number of neutrons} = A$$

The number and identity of subatomic particles present in an atom can be calculated from its atomic and mass numbers in the following manner.

$$\text{Number of protons} = \text{atomic number} = Z$$
$$\text{Number of electrons} = \text{atomic number} = Z$$
$$\text{Number of neutrons} = \text{mass number} - \text{atomic number} = A - Z$$

Note that neutron count is obtained by subtracting atomic number from mass number.

► The *sum* of the mass number and the atomic number for an atom corresponds to the total number of subatomic particles present in the atom (protons, neutrons, and electrons).

Example 3.1 Determining the Subatomic Particle Makeup of an Atom Given Its Atomic Number and Mass Number

An atom has an atomic number of 9 and a mass number of 19.

a. Determine the number of protons present.
b. Determine the number of neutrons present.
c. Determine the number of electrons present.

Solution

a. There are 9 protons because the atomic number is always equal to the number of protons present.
b. There are 10 neutrons because the number of neutrons is always obtained by subtracting the atomic number from the mass number.

$$\underbrace{\text{(Protons + neutrons)}}_{\text{Mass number}} - \underbrace{\text{protons}}_{\text{Atomic number}} = \text{neutrons}$$

c. There are 9 electrons because the number of protons and the number of electrons are always the same in an atom.

Electrons and Chemical Properties

The chemical properties of an atom, which are the basis for its identification, are determined by the number and arrangement of the electrons about the nucleus. When two atoms interact, the outer part (electrons) of one interacts with the outer part (electrons) of the other. The small nuclear centers never come in contact with each other in a chemical reaction. The number of electrons about a nucleus may be considered to be determined by the number of protons in the nucleus; charge balance requires an equal number of the two (Section 3.1). Hence the number of protons (which is the atomic number) characterizes an atom. All atoms with the same atomic number have the same chemical properties and are atoms of the same element.

In Section 1.6, an element was defined as a pure substance that cannot be broken down into simpler substances by ordinary chemical means. Although this is a good historical definition for an element, we can now give a more rigorous definition by using the concept of atomic number. An **element** *is a pure substance in which all atoms present have the same atomic number.*

An alphabetical listing of the 113 known elements, with their atomic numbers as well as other information, is found on the inside front cover of this text. If you check the atomic number column in this tabulation, you will find an entry for each of the numbers in the sequence 1 to 112 plus 114. The highest-atomic-numbered element that occurs naturally is uranium (element 92); elements 93 to 112 and 114 have been made in the laboratory but are not found in nature (Section 1.7). The fact that there are no gaps in the numerical sequence 1 to 92 is interpreted by scientists to mean that there are no "missing elements" yet to be discovered in nature.

Practice Questions and Problems

3.6 Determine the atomic number and mass number for atoms with the following subatomic makeups.
a. 2 protons, 2 neutrons, and 2 electrons
b. 4 protons, 5 neutrons, and 4 electrons
c. 5 protons, 4 neutrons, and 5 electrons
d. 28 protons, 30 neutrons, and 28 electrons

3.7 Determine the number of protons, neutrons, and electrons present in atoms with the following characteristics.
a. Atomic number = 8 and mass number = 16
b. Mass number = 18 and $Z = 8$
c. Atomic number = 20 and $A = 44$
d. $A = 257$ and $Z = 100$

3.8 Indicate whether the *atomic number,* the *mass number,* or *both the atomic number and mass number* are needed to determine each of the following.
a. Number of protons in an atom
b. Number of neutrons in an atom

c. Number of nucleons in an atom
d. Total number of subatomic particles in an atom

3.9 What information about the subatomic particles present in an atom is obtained from each of the following?
a. Atomic number
b. Mass number
c. Mass number − atomic number
d. Mass number + atomic number

3.10 What is the definition for an element in terms of atomic number?

Learning Focus

Be able to define the term *isotope,* write the symbol for an isotope, and calculate the atomic mass for an element from isotopic masses and percent abundances.

3.3 Isotopes and Atomic Masses

Charge neutrality (Section 3.1) requires the presence in an atom of an equal number of protons and electrons. However, because neutrons have no electrical charge, their numbers in atoms do not have to be the same as the number of protons or electrons. Most atoms contain more neutrons than either protons or electrons.

Studies of atoms of various elements also show that the number of neutrons present in atoms of an element is not constant; it varies over a small range. This means that not all atoms of an element have to be identical. They must have the same number of protons and electrons, but they can differ in the number of neutrons.

Atoms of an element that differ in neutron count are called isotopes. **Isotopes** *are atoms of an element that have the same number of protons and electrons but different numbers of neutrons.* Different isotopes always have the same atomic number and different mass numbers.

▶ The word *isotope* comes from the Greek *iso,* meaning "equal," and *topos,* meaning "place." Isotopes occupy an equal place (location) in listings of elements because all isotopes of an element have the same atomic number.

Most elements found in nature exist in isotopic forms, with the number of naturally occurring isotopes ranging from two to ten. For example, all silicon atoms have 14 protons and 14 electrons. Most silicon atoms also contain 14 neutrons. However, some silicon atoms contain 15 neutrons and others contain 16 neutrons. Thus three different kinds of silicon atoms exist, that is, three silicon isotopes exist.

Isotopes of an element have the same chemical properties, but their physical properties are often slightly different. Isotopes of an element have the same chemical properties because they have the same number of electrons. They have slightly different physical properties because they have different numbers of neutrons and therefore different masses.

When it is necessary to distinguish between isotopes of an element, the following notation is used:

Mass number A
Atomic number Z Symbol Chemical symbol

The atomic number is written as a *subscript* to the left of the elemental symbol for the atom. The mass number is written as a *superscript* to the left of the elemental symbol. Thus the three silicon isotopes are designated, respectively, as

$$^{28}_{14}\text{Si}, \quad ^{29}_{14}\text{Si}, \quad \text{and} \quad ^{30}_{14}\text{Si}$$

▶ There are a few elements for which all naturally occurring atoms have the same number of neutrons—that is, for which all atoms are identical. They include the elements Be, F, Na, Al, P, and Au.

▶ A mass number, in contrast to an atomic number, lacks uniqueness. Atoms of different elements can have the same mass number. For example, carbon-14 and nitrogen-14 have the same mass numbers. Atoms of different elements, however, cannot have the same atomic number.

Names for isotopes include the mass number. $^{28}_{14}\text{Si}$ is called silicon-28, and $^{29}_{14}\text{Si}$ is called silicon-29. The atomic number is not included in the name because it is the same for all isotopes of an element.

The various isotopes of a given element are of varying abundance; usually one isotope is predominant. Silicon is typical of this situation. The percentage abundances for its three isotopes are 92.21% ($^{28}_{14}\text{Si}$), 4.70% ($^{29}_{14}\text{Si}$), and 3.09% ($^{30}_{14}\text{Si}$). Percentage abundances are number percentages (numbers of atoms) rather than mass percentages. A sample of 10,000 silicon atoms contains 9221 $^{28}_{14}\text{Si}$ atoms, 470 $^{29}_{14}\text{Si}$ atoms, and 309 $^{30}_{14}\text{Si}$ atoms.

There are 286 isotopes that occur naturally. In addition, over 2000 more have been synthesized in the laboratory via nuclear rather than chemical reactions. All these synthetic isotopes are unstable (radioactive). Despite their instability, many are used in chemical and biological research, as well as in medicine. The topic of radioactivity is considered in the last few sections of this chapter.

▶ An analogy involving isotopes and identical twins may be helpful: Identical twins need not weigh the same, even though they have identical "gene packages." Likewise, isotopes, even though they have different masses, have the same number of protons.

Atomic Masses

The existence of isotopes means that atoms of an element can have several different masses. For example, silicon atoms can have any one of three masses because there are three silicon isotopes. Which of these three silicon isotopic masses is used in situations in which the mass of the element silicon needs to be specified? The answer is none of them. Instead we use a *weighted-average mass* that takes into account the existence of isotopes and their relative abundances.

The *weighted-average mass* of the isotopes of an element is known as the element's atomic mass. An **atomic mass** *is the calculated average mass for the isotopes of an element, expressed on a scale using atoms of* $^{12}_{6}C$ *as the reference.* What we need to calculate an atomic mass are the masses of the various isotopes on the $^{12}_{6}C$ reference scale and the percentage abundance of each isotope.

▶ The terms *atomic mass* and *atomic weight* are often used interchangeably. Atomic mass, however, is the correct term.

The $^{12}_{6}C$ reference scale mentioned in the definition of *atomic mass* is a scale scientists have set up for comparing the masses of atoms. On this scale, the mass of a $^{12}_{6}C$ atom is defined to be exactly 12 atomic mass units (amu). The masses of all other atoms are then determined relative to that of $^{12}_{6}C$. For example, if an atom is twice as heavy as $^{12}_{6}C$, its mass is 24 amu, and if an atom weighs half as much as an atom of $^{12}_{6}C$, its mass is 6 amu.

Example 3.2 shows how an atomic mass is calculated by using the amu ($^{12}_{6}C$) scale, the percentage abundances of isotopes, and the number of isotopes of an element.

Example 3.2 Calculation of an Element's Atomic Mass

Naturally occurring chlorine exists in two isotopic forms, $^{35}_{17}Cl$ and $^{37}_{17}Cl$. The relative mass of $^{35}_{17}Cl$ is 34.97 amu, and its abundance is 75.53%; the relative mass of $^{37}_{17}Cl$ is 36.97 amu, and its abundance is 24.47%. What is the atomic mass of chlorine?

Solution

An element's atomic mass is calculated by multiplying the relative mass of each isotope by its fractional abundance and then totaling the products. The fractional abundance for an isotope is its percentage abundance converted to decimal form (divided by 100).

$$^{35}_{17}Cl: \quad \left(\frac{75.53}{100}\right) \times 34.97 \text{ amu} = (0.7553) \times 34.97 \text{ amu} = 26.41 \text{ amu}$$

$$^{37}_{17}Cl: \quad \left(\frac{24.47}{100}\right) \times 36.97 \text{ amu} = (0.2447) \times 36.97 \text{ amu} = 9.047 \text{ amu}$$

$$\begin{aligned} \text{Atomic mass of Cl} &= (26.41 + 9.047) \text{ amu} \\ &= 35.46 \text{ amu} \end{aligned}$$

This calculation involved an element containing just two isotopes. A similar calculation for an element having three isotopes would be carried out the same way, but it would have three terms in the final sum; an element possessing four isotopes would have four terms in the final sum.

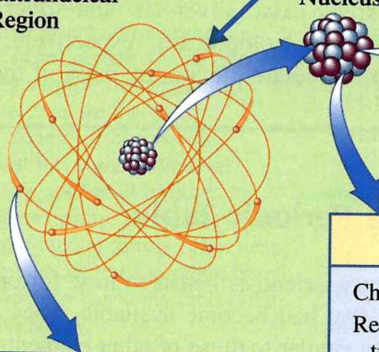
The alphabetical listing of the known elements printed inside the front cover of this text gives the calculated atomic mass for each of the elements; it is the last column of numbers.

The accompanying Chemistry at a Glance summarizes all that we have said about atoms thus far.

Atomic Structure

ATOMIC STRUCTURE

There are 113 known kinds of atoms, one kind for each element. All atoms have a nucleus (a small dense center containing all protons and neutrons) and an extranuclear region containing all electrons.



Protons

Charge = +1
Relative mass = 1837 times that of an electron

Atomic Number

The number of protons in the nucleus identifies the atom and equals the atomic number.

Neutrons

Charge = 0
Relative mass = 1839 times that of an electron

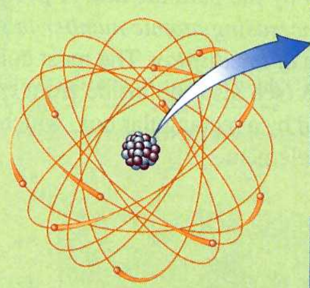
Mass Number

The number of protons plus the number of neutrons in an atom equals the mass number.

Electrons

Charge = –1
Relative mass = 1
Electrons determine the chemical properties of an atom.

VARIATIONS IN ATOMIC STRUCTURE



Isotopes

The atoms of various isotopes of an element contain the same number of protons but differ in the number of neutrons in the nucleus. Isotopes have the same atomic number and different mass numbers.

Atomic Mass

The average mass of all isotopes of an element weighted according to natural abundance is the atomic mass.

Practice Questions and Problems

3.11 What are the atomic number and the mass number for an atom denoted as $^{24}_{12}X$?

3.12 Determine the number of protons, electrons, and neutrons for the following atoms.
a. $^{12}_{6}C$ b. $^{16}_{8}O$ c. $^{26}_{12}Mg$ d. $^{197}_{79}Au$

3.13 How many protons, electrons, and neutrons are present in the following isotopes of zirconium?
a. $^{90}_{40}Zr$ b. $^{92}_{40}Zr$ c. $^{95}_{40}Zr$ d. $^{96}_{40}Zr$

3.14 Indicate whether the isotopes of an element differ or are the same in each of the following properties.
a. Atomic number b. Mass number
c. Number of neutrons d. Number of electrons

3.15 The atomic number of the element carbon (C) is 6. Write the complete symbol for each of the following carbon isotopes: carbon-12, carbon-13, and carbon-14.

3.16 What information about the isotopes of an element is needed before the atomic mass of the element can be calculated?

3.17 Most (92.58%) naturally occurring atoms of the element lithium have a mass of 7.02 amu. The remainder (7.42%) have a mass of 6.01 amu. Calculate the atomic mass of the element lithium.

3.18 Two naturally occurring isotopes of silver exist: silver-107 and silver-109. The percentage abundances and masses of these isotopes are, respectively, (51.82%, 106.9 amu) and (48.18%, 108.9 amu). Calculate the atomic mass of the element silver.

3.4 The Periodic Law and the Periodic Table

Learning Focus

Understand and state the periodic law. Understand the rationale behind the organization of the periodic table, relate the terms *group* and *period* to the periodic law, and indicate what general information about each element is given in the periodic table.

During the early part of the nineteenth century, scientists began to look for order in the increasing amount of chemical information that had become available. They knew that certain elements had properties that were very similar to those of other elements, and they sought reasons for these similarities in the hope that these similarities would suggest a method for arranging or classifying the elements.

In 1869, these efforts culminated in the discovery of what is now called the *periodic law,* proposed independently by the Russian chemist Dmitri Mendeleev (Figure 3.2) and the German chemist Julius Lothar Meyer. Given in its modern form, the **periodic law** *states that when elements are arranged in order of increasing atomic number, elements with similar properties occur at periodic (regularly recurring) intervals.*

A periodic table represents a compact graphical method for representing the behavior described by the periodic law. A **periodic table** *is a graphical display of the elements in order of increasing atomic number in which elements with similar properties fall in the same column of the display.* The most commonly used form of the periodic table is shown in Figure 3.3 (see also the inside front cover of the text). Within the table, each element is represented by a rectangular box, which contains the symbol, atomic number, and atomic mass of the element.

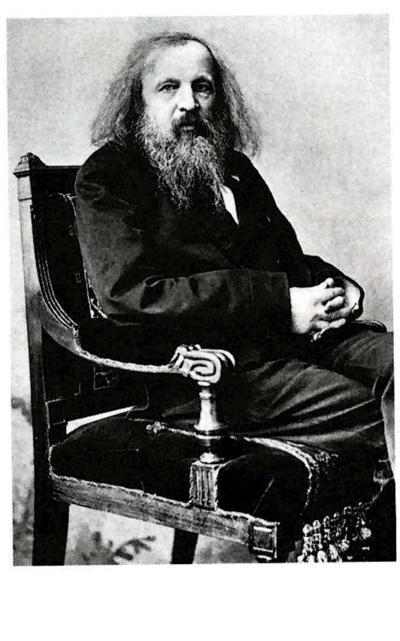


Figure 3.2 Dmitri Ivanovich Mendeleev (1834–1907). Mendeleev constructed a periodic table as part of his effort to systematize chemistry. He received many international honors for his work, but his reception at home in czarist Russia was mixed. Element 101 carries his name.

Groups and Periods of Elements

▶ Using the information on a periodic table, you can quickly determine the number of protons and electrons for atoms of an element. However, no information concerning neutrons is available from a periodic table; mass numbers are not part of the information given, because they are not unique to an element.

▶ The elements within a given periodic-table group show numerous similarities in properties, the degree of similarity varying from group to group. In no case are the group members "clones" of one another. Each element has some individual characteristics not found in other elements of the group. By analogy, the members of a human family often bear many resemblances to each other, but each member also has some (and often much) individuality.

The location of an element within the periodic table is specified by giving its group number and period number.

A **group** *in the periodic table is a vertical column of elements.* There are two notations in use for designating individual periodic-table groups. In the first notation, which has been in use for many years, groups are designated by using Roman numerals and the letters A and B. In the second notation, which has recently been recommended for use by an international scientific commission, the Arabic numbers 1 through 18 are used. Note that in Figure 3.3 both group notations are given at the top of each group. The elements with atomic numbers 8, 16, 34, 52, and 84 (O, S, Se, Te, and Po) constitute Group VIA (old notation) or Group 16 (new notation).

Several groups of elements have common (non-numerical) names that are used so frequently that they should be learned. The Group IA elements, except for hydrogen, are called the *alkali metals,* and the Group IIA elements the *alkaline earth metals.* On the opposite side of the periodic table from the IA and IIA elements are found the *halogens* (the Group VIIA elements) and the *noble gases* (Group VIIIA).

A **period** *in the periodic table is a horizontal row of elements.* For identification purposes, the periods are numbered sequentially with Arabic numbers, starting at the top of the periodic table. In Figure 3.3 period numbers are found on the left side of the table. The elements Na, Mg, Al, Si, P, S, Cl, and Ar are all members of Period 3, the third row of elements. Period 4 is the fourth row of elements, and so on. There are only two elements in Period 1, H and He.

24 Cr 52.00 — Atomic number, Symbol, Atomic mass

Period	1 Group IA	2 Group IIA	3 Group IIIB	4 Group IVB	5 Group VB	6 Group VIB	7 Group VIIB	8 Group VIIIB	9 Group VIIIB	10 Group VIIIB	11 Group IB	12 Group IIB	13 Group IIIA	14 Group IVA	15 Group VA	16 Group VIA	17 Group VIIA	18 Group VIIIA
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (269)	109 Mt (268)	110 (271)	111 (272)	112 (277)		114 (289)				

Metals — Nonmetals

58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
90 Th (232)	91 Pa (231)	92 U (238)	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (248)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (260)	102 No (259)	103 Lr (262)

Figure 3.3 The periodic table of the elements is a graphical way to show relationships among the elements. Elements with similar chemical properties fall in the same vertical column.

1																															2
3	4																									5	6	7	8	9	10
11	12																									13	14	15	16	17	18
19	20	21															22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
37	38	39															40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105	106	107	108	109	110	111	112		114				

Figure 3.4 In this periodic table, elements 58 through 71 and 90 through 103 (in color) are shown in their proper positions.

The location of any element in the periodic table is specified by giving its group number and its period number. The element gold, with an atomic number of 79, belongs to Group IB (or 11) and is in Period 6. The element nitrogen, with an atomic number of 7, belongs to Group VA (or 15) and is in Period 2.

▶ When the statement "the first ten elements" is used, it means the first ten elements in the periodic table, the elements with atomic numbers 1 through 10.

The Shape of the Periodic Table

Within the periodic table of Figure 3.3 the practice of arranging the elements according to increasing atomic number is violated in Groups IIIB and IVB. Element 72 follows element 57, and element 104 follows element 89. The missing elements, elements 58 through 71 and 90 through 103 are located in two rows at the bottom of the periodic table. Technically, the elements at the bottom of the table should be included in the body of the table, as shown in Figure 3.4. However, in order to have a more compact table, we place them at the bottom of the table as shown in Figure 3.3.

Practice Questions and Problems

3.19 Give the symbol of the element that occupies each of the following positions in the periodic table.
a. Period 4, Group IIA
b. Period 5, Group VIB
c. Group IA, Period 2
d. Group IVA, Period 5

3.20 For each of the following sets of elements, choose the two that would be expected to have similar chemical properties.
a. ${}_{19}K$, ${}_{29}Cu$, ${}_{37}Rb$, ${}_{41}Nb$
b. ${}_{13}Al$, ${}_{14}Si$, ${}_{15}P$, ${}_{33}As$
c. ${}_{9}F$, ${}_{40}Zr$, ${}_{50}Sn$, ${}_{53}I$
d. ${}_{11}Na$, ${}_{12}Mg$, ${}_{54}Xe$, ${}_{55}Cs$

3.21 Using the periodic table, determine
a. the atomic number of the element carbon
b. the atomic mass of the element silicon
c. the atomic number of the element with an atomic mass of 88.91 amu
d. the atomic mass of the element located in Period 2 and Group IIA

3.22 The following statements either define or are closely related to the terms *periodic law, period,* or *group*. Match each statement with the appropriate term.
a. This is a vertical arrangement of elements in the periodic table.
b. The properties of the elements repeat in a regular way as atomic numbers increase.
c. Elements 10, 18, and 36 belong to this type of arrangement.
d. Elements 24 and 33 belong to this type of arrangement.

Learning Focus

Be able to classify an element as a metal or a nonmetal, and know the general physical properties for metallic and nonmetallic elements.

3.5 Metals and Nonmetals

In the previous section, we noted that the Group IA and IIA elements are known, respectively, as the alkali metals and the alkaline earth metals. Both of these designations contain the word *metal.* But what is a metal?

On the basis of selected physical properties, elements are classified into the categories metal and nonmetal. A **metal** *is an element that has the characteristic properties of luster,*

(a) (b)

Figure 3.5 (a) Some familiar metals are aluminum, lead, tin, and zinc. (b) Some familiar nonmetals are sulfur (yellow), phosphorus (dark red), and bromine (reddish-brown).

thermal conductivity, electrical conductivity, and malleability. With the exception of mercury, all metals are solids at room temperature (25°C). Among the more familiar metals are the elements iron, aluminum, copper, silver, gold, lead, tin, and zinc (see Figure 3.5a). A **nonmetal** *is an element characterized by the absence of the properties of luster, thermal conductivity, electrical conductivity, and malleability*. Many of the nonmetals, such as hydrogen, oxygen, nitrogen, and the noble gases, are gases. The only nonmetal found as a liquid at room temperature is bromine. Solid nonmetals include carbon, iodine, sulfur, and phosphorus (Figure 3.5b).

Table 3.2 contrasts selected physical properties of metals and nonmetals. In many ways, the general properties of metals and nonmetals are opposites. Metals generally are lustrous (shine, reflect light), malleable (can be drawn into wires), ductile (can be rolled into sheets), and good thermal and electrical conductors. Nonmetals tend to lack these properties. Generally, the nonmetals have lower densities and lower melting points than metals.

Periodic Table Locations for Metals and Nonmetals

The majority of the elements are metals. Only 22 elements are nonmetals. It is not necessary to memorize which elements are nonmetals and which are metals; this information is obtainable from a periodic table (Figure 3.6). The steplike heavy line that runs through the right third of the periodic table separates the metals on the left from the nonmetals on the right. Note also that the element hydrogen is a nonmetal.

The fact that the vast majority of elements are metals in no way indicates that metals are more important than nonmetals. Most nonmetals are relatively common and are found in many important compounds. For example, water (H_2O) is a compound involving two nonmetals.

Table 3.2
Selected Physical Properties of Metals and Nonmetals

Metals	Nonmetals
1. High electrical conductivity that decreases with increasing temperature	1. Poor electrical conductivity (except carbon in the form of graphite)
2. High thermal conductivity	2. Good heat insulators (except carbon in the form of diamond)
3. Metallic gray or silver luster[a]	3. No metallic luster
4. Almost all are solids[b]	4. Solids, liquids, or gases
5. Malleable (can be hammered into sheets)	5. Brittle in solid state
6. Ductile (can be drawn into wires)	6. Nonductile

[a]Except copper and gold.
[b]Except mercury; cesium and gallium melt on a hot summer day (85°F) or when held in a person's hand.

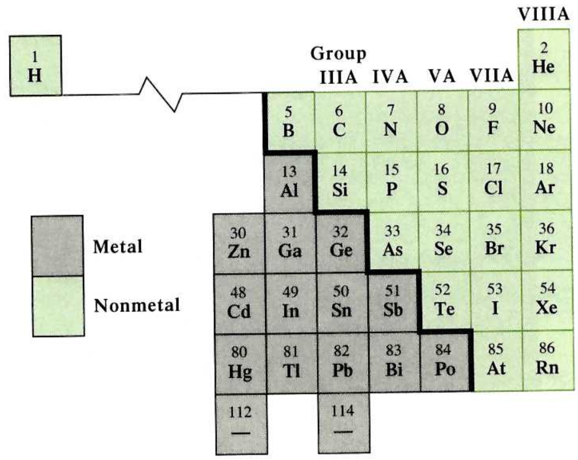


Figure 3.6 This portion of the periodic table shows the dividing line between metals and nonmetals. All elements that are not shown are metals.

An analysis of the abundance of the elements in Earth's crust (Figure 1.11) in terms of metals and nonmetals shows that the two most abundant elements, which account for 80.2% of all atoms, are nonmetals—oxygen and silicon. The four most abundant elements in the human body (Table 1.1), which comprise over 99% of all atoms in the body, are nonmetals—hydrogen, oxygen, carbon, and nitrogen.

The ranking of metals according to abundance does not reflect their importance in today's world. Abundance and importance (use) are two entirely different things. Some metals with low abundances in Earth's crust are more widely used than some of the more abundant metals. Why is this so? Obviously, the chemical properties of the various metals help determine their use. Another major factor is the availability of a metal's deposits

Chemical Portraits 4 Metals of Importance as Determined by Extent of Use

Iron (Fe)

Profile: Iron is the most used of all metals. In fact, more Fe is consumed, both worldwide and in the United States, than all other metals combined. When pure, Fe is lustrous, silvery, and soft. However, pure Fe is seldom encountered; rather the end form for nearly all consumption of Fe is steel, an Fe alloy. When exposed to air and moisture, iron *rusts*; the reddish-brown compound formed is a hydrated form of Fe_2O_3.

Uses: Two types of steel exist: carbon steel and alloy steel. Both are metal alloys that contain iron and carbon. Alloy steels also contain small amounts of other metals added to improve specific characteristics of the steel.

What is the major function for iron present in the human body?

Aluminum (Al)

Profile: Aluminum, the most abundant metal on Earth, is the second-most-used metal (although its use is just 10% that of iron). Freshly cut aluminum appears silvery. However, its surface quickly changes to a dull gray as a thin film of aluminum oxide (Al_2O_3) forms, preventing further corrosion.

Uses: Since 1994, the transportation sector has been the largest U.S. market for aluminum. The second-largest market is containers and packaging—beverage cans, food containers, and household and institutional foil. In addition, 90% of all overhead electrical transmission lines in the United States are made of aluminum alloys.

What is the "driving force" for the recycling of aluminum-containing products?

Copper (Cu)

Profile: With a reddish-yellow color, copper is the only metal besides gold to have a natural color other than gray-silver. Usage for Cu, the third-most-used metal, is just 5% that of iron. Weathering coats Cu with a green film that protects it from further corrosion. This green film is a compound formed from the reaction of Cu with three components of the air: oxygen, carbon dioxide, and water.

Uses: Copper is one of the few elements that is mainly used in pure form rather than as an alloy or compound. Building wire (16%) and plumbing (14%) are Cu's top two markets. Important copper alloys include *bronze* (up to 33% zinc present) and *brass* (up to 18% tin present).

How much copper is present in the various types of United States coinage?

See the text web site at **http://chemistry.college.hmco.com/students** for answers to the above questions and for further information.

in Earth's crust. Some metals have been concentrated by natural processes into localized areas. Less abundant metals that are found in "concentrated form" are generally more used than more abundant metals that are not "concentrated." The difficulty and cost of isolating a metal from its natural sources also affect the extent of its use. Chemical Portraits 4 profiles the three "most used" metals: iron, aluminum, and copper.

Practice Questions and Problems

3.23 In which of the following pairs of elements are both members of the pair metals?
a. $_{17}Cl$ and $_{35}Br$ b. $_{13}Al$ and $_{14}Si$ c. $_{29}Cu$ and $_{42}Mo$ d. $_{30}Zn$ and $_{83}Bi$

3.24 Identify the nonmetal in each of the following sets of elements.
a. S, Na, K b. Cu, Li, P c. Be, I, Ca d. Fe, Cl, Ga

3.25 Classify each of the following general physical properties as a property of metallic elements or of nonmetallic elements.
a. Malleable and ductile
b. Low electrical conductivity
c. High thermal conductivity
d. Good heat insulator

3.6 Electron Arrangements Within Atoms

Learning Focus

Understand how the terms *electron shell, electron subshell,* and *electron orbital* are related and how they are used in describing electron arrangements within an atom.

Current chemical theory indicates that as electrons move about an atom's nucleus, they are restricted to specific regions within the extranuclear portion of the atom. Such restrictions are determined by the amount of energy the electrons possess. Furthermore, chemical theory indicates that electron energies are limited to certain values and that a specific "behavior" is associated with each allowed energy value.

The space in which electrons move rapidly about a nucleus is divided into subspaces called *shells, subshells,* and *orbitals.*

Electron Shells

Electrons within an atom are grouped into main energy levels called electron shells. An **electron shell** *is a region of space about a nucleus that contains electrons that have approximately the same energy and that spend most of their time approximately the same distance from the nucleus.*

Electron shells are numbered 1, 2, 3, and so on, outward from the nucleus. Electron energy increases as the distance of the electron shell from the nucleus increases. An electron in shell 1 has the minimum amount of energy that an electron can have.

▶ Electrons that occupy the first electron shell are closer to the nucleus and have a lower energy than electrons in the second electron shell.

The maximum number of electrons that an electron shell can accommodate varies; the higher the shell number (n), the more electrons that can be present. In higher-energy shells the electrons are farther from the nucleus, and a greater volume of space is available for them; hence more electrons can be accommodated. (Conceptually, electron shells may be considered to be nested one inside another, somewhat like the layers of flavors inside a jawbreaker or similar type of candy.)

The lowest-energy shell ($n = 1$) accommodates a maximum of 2 electrons. In the second, third, and fourth shells, 8, 18, and 32 electrons, respectively, are allowed. The relationship among these numbers is given by the formula $2n^2$, where n is the shell number. For example, when $n = 4$, the quantity $2n^2 = 2(4)^2 = 32$.

Electron Subshells

Within each electron shell, electrons are further grouped into energy sublevels called electron subshells. An **electron subshell** *is a region of space within an electron shell that contains electrons that have the same energy.* We can draw an analogy between the relationship of shells and subshells and the physical layout of a high-rise apartment complex.

The shells are analogous to the floors of the apartment complex, and the subshells are the counterparts of the various apartments on each floor.

The number of subshells within a shell is the same as the shell number. Shell 1 contains one subshell, shell 2 contains two subshells, shell 3 contains three subshells, and so on.

▶ The letters used to label the different types of subshells come from old spectroscopic terminology associated with the lines in the spectrum of the element hydrogen. These lines were denoted as *s*harp, *p*rincipal, *d*iffuse, and *f*undamental. Relationships exist between such lines and the arrangement of electrons in an atom.

Subshells within a shell differ in size (that is, the maximum number of electrons they can accommodate) and energy. The higher the energy of the contained electrons, the larger the subshell.

Subshell size (type) is designated using the letters *s*, *p*, *d*, and *f*. Listed in this order, these letters denote subshells of increasing energy and size. The lowest-energy subshell within a shell is always the *s* subshell, the next highest is the *p* subshell, then the *d* subshell, and finally the *f* subshell. An *s* subshell can accommodate 2 electrons, a *p* subshell 6 electrons, a *d* subshell 10 electrons, and an *f* subshell 14 electrons.

Both a number and a letter are used in identifying subshells. The number gives the shell within which the subshell is located, and the letter gives the type of subshell. Shell 1 has only one subshell, the 1*s*. Shell 2 has two subshells, the 2*s* and 2*p*. Shell 3 has three subshells, the 3*s*, 3*p*, and 3*d*; and so on. Figure 3.7 summarizes the relationships between electron shells and electron subshells for the first four shells.

Electron Orbitals

Electron subshells have within them a certain, definite number of locations (regions of space), called electron orbitals, where electrons may be found. In our apartment complex analogy, if shells are the counterparts of floor levels and subshells are the apartments, then electron orbitals are the rooms of the apartments. An **electron orbital** *is a region of space within an electron subshell where an electron with a specific energy is most likely to be found.*

▶ An electron orbital is also often called an atomic orbital.

An electron orbital, independent of all other considerations, can accommodate a maximum of 2 electrons. Thus an *s* subshell (2 electrons) contains one orbital, a *p* subshell (6 electrons) contains three orbitals, a *d* subshell (10 electrons) contains five orbitals, and an *f* subshell (14 electrons) contains seven orbitals.

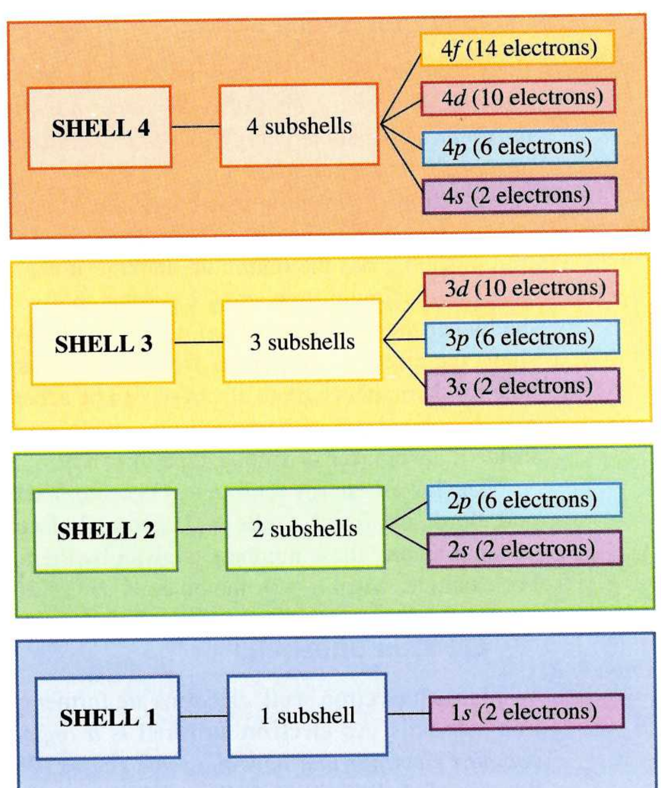
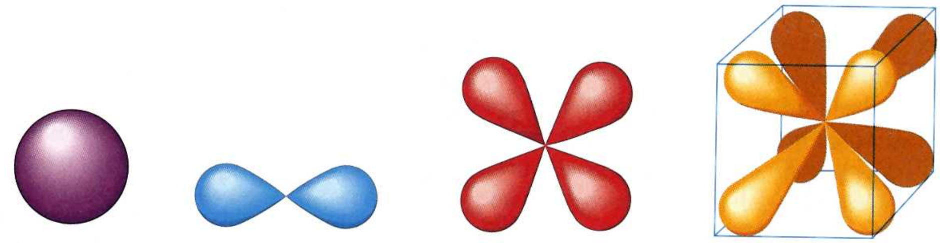


Figure 3.7 The number of subshells within a shell is equal to the shell number, as shown here for the first four shells. Each individual subshell is denoted with both a number (its shell) and a letter (the type of subshell it is in).

Figure 3.8 An *s* orbital has a spherical shape, a *p* orbital has two lobes, a *d* orbital has four lobes, and an *f* orbital has eight lobes. The *f* orbital is shown within a cube to illustrate that its lobes are directed toward the corners of a cube. Some *d* and *f* orbitals have shapes related to, but not identical to, those shown.



(a) *s* orbital **(b)** *p* orbital **(c)** *d* orbital **(d)** *f* orbital

Orbitals have distinct shapes that are related to the type of subshell in which they are found. Note that we are talking not about the shape of an electron but, rather, about the shape of the region in which the electron is found. An orbital in an *s* subshell, which is called an *s* orbital, has a spherical shape (Figure 3.8a). Orbitals found in *p* subshells—*p* orbitals—have shapes similar to the "figure 8" of an ice skater (Figure 3.8b). More complex shapes involving four and eight lobes, respectively, are associated with *d* and *f* orbitals (Figures 3.8c and 3.8d). Some *d* and *f* orbitals have shapes related to, but not identical to, those shown in Figure 3.8.

Figure 3.9 which is an extension of Figure 3.7, summarizes the important relationships among electron shells, electron subshells, and electron orbitals.

Practice Questions and Problems

3.26 What is the maximum number of electrons that can occupy each of the following *electron shells*?
a. First shell b. Second shell c. Third shell d. Fifth shell

3.27 How many electron subshells are present in each of the following electron shells?
a. First shell b. Second shell c. Third shell d. Fourth shell

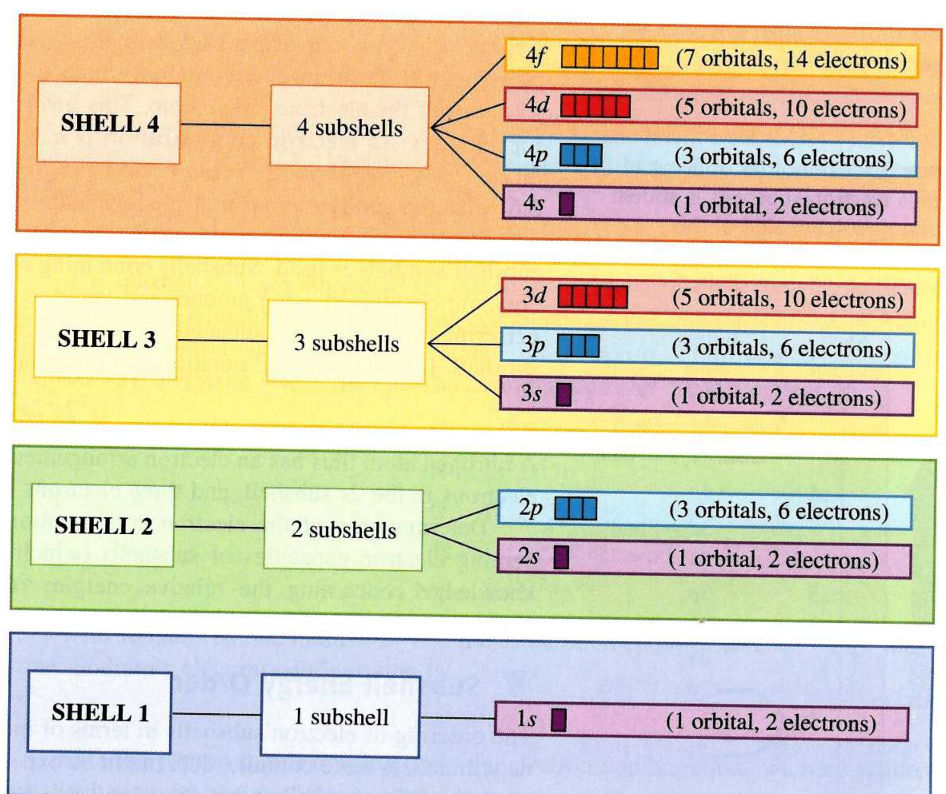


Figure 3.9 A summary of the interrelationships among electron shells, electron subshells, and electron orbitals for the first four shells. Similar relationship patterns exist for higher-numbered shells.

3.28 What is the maximum number of electrons that can occupy each of the following types of *electron subshells*?
a. *s* subshell b. *p* subshell c. *d* subshell d. *f* subshell

3.29 What is the maximum number of electrons that can occupy each of the following *electron orbitals*?
a. 2*s* b. 3*p* c. 3*d* d. 4*s*

3.30 The following statements define or are closely related to the terms *electron shell, electron subshell,* and *electron orbital.* Match each statement with the appropriate term.
a. In terms of electron capacity, this unit is the smallest of the three.
b. This unit can contain a maximum of 2 electrons.
c. This unit is designated using just a number.
d. The term *energy sublevel* is closely associated with this unit.

3.31 Indicate whether each of the following statements is *true* or *false*.
a. An orbital has a definite size and shape, which are related to the energy of the electrons it could contain.
b. All the orbitals in a subshell have the same energy.
c. All subshells accommodate the same number of electrons.
d. A 2*p* subshell and a 3*p* subshell can accommodate the same number of electrons.

3.32 Give the maximum number of electrons that can occupy each of the following electron-accommodating units.
a. One of the orbitals in the 2*p* subshell
b. One of the orbitals in the 3*d* subshell
c. The 4*p* subshell
d. The third shell

3.7 Electron Configurations

Learning Focus

Write the electron configuration for atoms of any element on the basis of the relative energies of electron subshells.

Electron shells, electron subshells, and electron orbitals describe "permissible" locations for electrons—that is, where electrons *can* be found. Electrons do not occupy these "locations" in a random, haphazard fashion; a very predictable pattern exists for the arrangement of electrons about an atom's nucleus. We are now ready to discuss the *actual* locations of the electrons in an atom. This involves specifying the *electron configuration* for an atom. An **electron configuration** *is a statement of how many electrons an atom has in each of its subshells.* Because subshells group electrons according to energy (Section 3.6), electron configurations indicate how many electrons an atom has of various energies.

Electron configurations are not written out in words; a shorthand system based on subshell symbols is used. Subshells containing electrons, listed in order of increasing energy, are designated using number-and-letter combinations (1*s*, 2*s*, 2*p*, and so on). A superscript following each subshell designation indicates the number of electrons in that subshell. In this shorthand notation, the electron configuration for nitrogen is

$$1s^2 2s^2 2p^3$$

A nitrogen atom thus has an electron arrangement of two electrons in the 1*s* subshell, two electrons in the 2*s* subshell, and three electrons in the 2*p* subshell.

Determination of the electron configuration for an atom requires knowledge concerning electron capacities of subshells (which we already have; see Section 3.6) and knowledge concerning the relative energies of subshells (which we now consider). *Electrons occupy electron subshells in an atom in order of increasing subshell energy.*

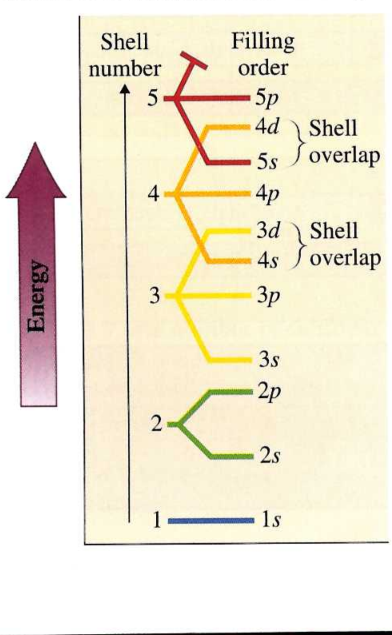


Figure 3.10 The order of filling of various electron subshells is shown on the right-hand side of this diagram. Above the 3*p* subshell, subshells of different shells "overlap."

Subshell Energy Order

The ordering of electron subshells in terms of increasing energy, which is experimentally determined, is more complex than might be expected. This is because the energies of subshells in different shells often "overlap," as shown in Figure 3.10. This diagram shows, for example, that the 4*s* subshell has lower energy than the 3*d* subshell.

A useful mnemonic (memory) device for remembering subshell filling order, which incorporates "overlap" situations such as those in Figure 3.10, is given in Figure 3.11. This diagram, which lists all subshells needed to specify the electron arrangements for all 113 elements, is constructed by locating all *s* subshells in column 1, all *p* subshells in column 2, and so on. Subshells that belong to the same shell are found in the same row. The order of subshell filling is given by following the diagonal arrows, starting at the top. The 1*s* subshell fills first. The second arrow points to (goes through) the 2*s* subshell, which fills next. The third arrow points to both the 2*p* and the 3*s* subshells. The 2*p* fills first, followed by the 3*s*. Any time a single arrow points to more than one subshell, we start at the tail of the arrow and work to its tip to determine the proper filling sequence.

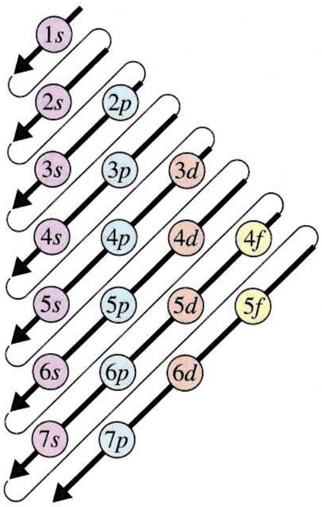


Figure 3.11 The order for filling electron subshells with electrons follows the order given by the arrows in this diagram. Start with the arrow at the top of the diagram and work toward the bottom of the diagram, moving from the bottom of one arrow to the top of the next-lower arrow.

Writing Electron Configurations

We are now ready to write electron configurations. Let us systematically consider electron configurations for the first few elements in the periodic table.

Hydrogen (atomic number = 1) has only one electron, which goes into the 1*s* subshell, which has the lowest energy of all subshells. Hydrogen's electron configuration is written as

$$1s^1$$

Helium (atomic number = 2) has two electrons, both of which occupy the 1*s* subshell. (Remember, an *s* subshell contains one orbital, and an orbital can accommodate two electrons.) Helium's electron configuration is

$$1s^2$$

Lithium (atomic number = 3) has three electrons, and the third electron cannot enter the 1*s* subshell because its maximum capacity is two electrons. (All *s* subshells are completely filled with two electrons.) The third electron is placed in the next-higher-energy subshell, the 2*s*. The electron configuration for lithium is

$$1s^22s^1$$

For *beryllium* (atomic number = 4), the additional electron is placed in the 2*s* subshell, which is now completely filled, giving beryllium the electron configuration

$$1s^22s^2$$

For *boron* (atomic number = 5), the 2*p* subshell, which is the subshell of next-higher energy (Figures 3.10 and 3.11) becomes occupied for the first time. Boron's electron configuration is

$$1s^22s^22p^1$$

▶ The symbols $1s^2$, $2s^2$, and $2p^3$ are read as "one s two," "two s two," and "two p three," not as "one s squared," "two s squared," and "two p cubed."

A *p* subshell can accommodate six electrons because there are three orbitals within it (Section 3.6). The 2*p* subshell can thus accommodate the additional electrons found in the elements with atomic numbers 6 through 10: *carbon* (C), *nitrogen* (N), *oxygen* (O), *fluorine* (F), and *neon* (Ne). The electron configurations for these elements are

$$\begin{aligned} \text{C:}\ & 1s^22s^22p^2 \\ \text{N:}\ & 1s^22s^22p^3 \\ \text{O:}\ & 1s^22s^22p^4 \\ \text{F:}\ & 1s^22s^22p^5 \\ \text{Ne:}\ & 1s^22s^22p^6 \end{aligned}$$

With *sodium* (atomic number = 11), the 3*s* subshell acquires an electron for the first time. Sodium's electron configuration is

$$1s^22s^22p^63s^1$$

Note the pattern that is developing in the electron configurations we have written so far. Each element has an electron configuration that is the same as the one just before it except for the addition of one electron.

▶ An *electron configuration* is a shorthand notation designating the subshells in an atom that are occupied by electrons. The sum of the superscripts in an electron configuration equals the total number of electrons present and hence must equal the atomic number of the element.

Electron configurations for other elements are obtained by simply extending the principles we have just illustrated. A subshell of lower energy is always filled before electrons are added to the next highest subshell; this continues until the correct number of electrons have been accommodated.

For a few elements in the middle of the periodic table, the actual distribution of electrons within subshells differs slightly from that obtained by using the procedures outlined in this section. These exceptions are caused by very small energy differences between some subshells and are not important in the uses we shall make of electron configurations.

Example 3.3 Writing an Electron Configuration

Write the electron configurations for the following elements.

a. Strontium (atomic number = 38) **b.** Lead (atomic number = 82)

Solution

a. The number of electrons in a strontium atom is 38. Remember that the atomic number gives the number of electrons (Section 3.2). We will need to fill subshells, in order of increasing energy, until 38 electrons have been accommodated.

The 1*s*, 2*s*, and 2*p* subshells fill first, accommodating a total of 10 electrons among them.

$$1s^2 2s^2 2p^6 \ldots$$

Next, according to Figures 3.10 and 3.11 the 3*s* subshell fills and then the 3*p* subshell.

$$1s^2 2s^2 2p^6 3s^2 3p^6 \ldots$$

We have accommodated 18 electrons at this point. We still need to add 20 more electrons to get our desired number of 38.

The 4*s* subshell fills next, followed by the 3*d* subshell, giving us 30 electrons at this point.

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} \ldots$$

Note that the maximum electron population for *d* subshells is 10 electrons.

Eight more electrons are needed, which are added to the next two higher subshells, the 4*p* and the 5*s*. The 4*p* subshell can accommodate 6 electrons, and the 5*s* can accommodate 2 electrons.

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$$

To double-check that we have the correct number of electrons, 38, we add the superscripts in our final electron configuration.

$$2 + 2 + 6 + 2 + 6 + 2 + 10 + 6 + 2 = 38$$

The sum of the superscripts in any electron configuration should add up to the atomic number if the configuration is for a neutral atom.

b. To write this configuration, we continue along the same lines as in part **a**, remembering that the maximum electron subshell populations are $s = 2$, $p = 6$, $d = 10$, and $f = 14$.

Lead, with an atomic number of 82, contains 82 electrons, which are added to subshells in the following order. (The line of numbers beneath the electron configuration is a running total of added electrons and is obtained by adding the superscripts up to that point. We stop when we have 82 electrons.)

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$$

2 4 10 12 18 20 30 36 38 48 54 56 70 80 82

Running total of electrons added

Note in this electron configuration that the 6*p* subshell contains only 2 electrons, even though it can hold a maximum of 6. We put only 2 electrons in this subshell because that is sufficient to give 82 total electrons. If we had completely filled this subshell, we would have had 86 total electrons, which is too many.

Practice Questions and Problems

3.33 On the basis of the total number of electrons present, identify the elements whose electron configurations are
a. $1s^22s^22p^4$ b. $1s^22s^22p^6$ c. $1s^22s^22p^63s^23p^1$ d. $1s^22s^22p^63s^23p^64s^2$

3.34 Write the electron configurations for atoms of the following elements.
a. Carbon b. Scandium c. Arsenic d. Element with atomic number of 17

3.35 Indicate whether each of the following statements concerning the element whose electron configuration is $1s^22s^22p^63s^23p^5$ is *true* or *false*.
a. Atoms of the element contain 17 electrons.
b. Atoms of the element have electrons in five different subshells.
c. Atoms of the element have electrons in subshells in three different shells.
d. The identity of the element is sulfur.

Learning Focus

Be able to relate the periodic law and the shape of the periodic table to electron configurations for atoms of the elements.

3.8 The Electronic Basis for the Periodic Law and the Periodic Table

For many years, there was no explanation available for either the periodic law or why the periodic table has the shape that it has. We now know that the theoretical basis for both the periodic law and the periodic table is found in electronic theory. As we saw earlier in the chapter (Section 3.2), when two atoms interact, it is their electrons that interact. Thus the number and arrangement of electrons determine how an atom reacts with other atoms—that is, what its chemical properties are. The properties of the elements repeat themselves in a periodic manner because the arrangement of electrons about the nucleus of an atom follows a periodic pattern, as we saw in Section 3.7.

Electron Configurations and the Periodic Law

The periodic law (Section 3.4) points out that the properties of the elements repeat themselves in a regular manner when the elements are arranged in order of increasing atomic number. The elements that have similar chemical properties are placed under one another in vertical columns (groups) in the periodic table.

Groups of elements have similar chemical properties because of similarities in their electron configuration. *Chemical properties repeat themselves in a regular manner among the elements because electron configurations repeat themselves in a regular manner among the elements.*

To illustrate this correlation between similar chemical properties and similar electron configurations, let us look at the electron configurations of two groups of elements known to have similar chemical properties.

We begin with the elements lithium, sodium, potassium, and rubidium, all members of Group IA of the periodic table. The electron configurations for these elements are

$$
\begin{aligned}
{}_{3}\text{Li}&: \ 1s^2 2s^1 \\
{}_{11}\text{Na}&: \ 1s^22s^22p^6 3s^1 \\
{}_{19}\text{K}&: \ 1s^22s^22p^63s^23p^6 4s^1 \\
{}_{37}\text{Rb}&: \ 1s^22s^22p^63s^23p^64s^23d^{10}4p^6 5s^1
\end{aligned}
$$

▶ The electron arrangement in the outermost shell is the same for elements in the same group. This is why elements in the same group have similar chemical properties.

Note that each of these elements has one electron in its outermost shell. (The outermost shell is the shell with the highest number.) This similarity in outer-shell electron arrangements causes these elements to have similar chemical properties. In general, elements with similar outer-shell electron configurations have similar chemical properties.

Let us consider another group of elements known to have similar chemical properties: fluorine, chlorine, bromine, and iodine of Group VIIA of the periodic table. The electron configurations for these four elements are

$$_{9}\text{F}: \quad 1s^2 2s^2 2p^5$$
$$_{17}\text{Cl}: \quad 1s^2 2s^2 2p^6 3s^2 3p^5$$
$$_{35}\text{Br}: \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$$
$$_{53}\text{I}: \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$$

Once again, similarities in electron configuration are readily apparent. This time, the repeating pattern involves an outermost *s* and *p* subshell containing seven electrons (shown in color). Remember that for Br and I, shell numbers 4 and 5 designate, respectively, electrons in the outermost shells.

Electron Configurations and the Periodic Table

One of the strongest pieces of supporting evidence for the assignment of electrons to shells, subshells, and orbitals is the periodic table itself. The basic shape and structure of this table, which was determined many years before electrons were even discovered, is consistent with and can be explained by electron configurations. Indeed, the specific location of an element in the periodic table can be used to obtain information about its electron configuration.

As the first step in linking electron configurations to the periodic table, let us analyze the general shape of the periodic table in terms of columns of elements. As shown in Figure 3.12, on the extreme left of the table, there are 2 columns of elements; in the center, there is a region containing 10 columns of elements; to the right there is a block of 6 columns of elements; and in the two rows at the bottom of the table, there are 14 columns of elements.

The number of columns of elements in the various regions of the periodic table—2, 6, 10, and 14—is the same as the maximum number of electrons that the various types of subshells can accommodate. We will see shortly that this is a very significant observation; the number matchup is no coincidence. The various columnar regions of the periodic table are called the *s* area (2 columns), the *p* area (6 columns), the *d* area (10 columns), and the *f* area (14 columns), as shown in Figure 3.12.

The concept of *distinguishing electrons* is the key to obtaining electron configuration information from the periodic table. The **distinguishing electron** *for an element is*

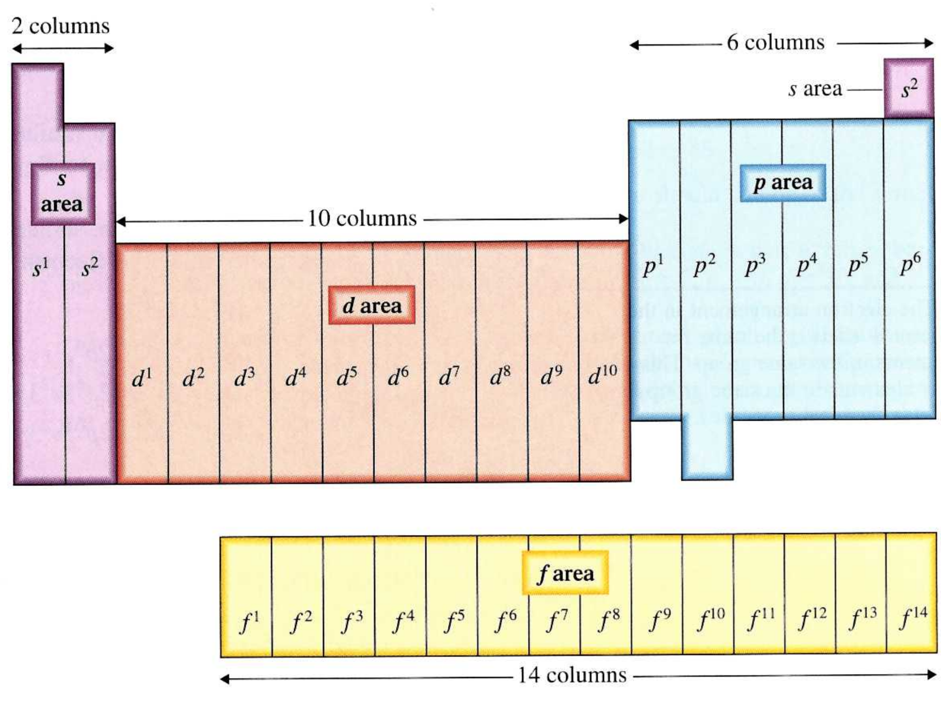


Figure 3.12 Electron configurations and the positions of elements in the periodic table. The periodic table can be divided into four areas that are 2, 6, 10, and 14 columns wide. The four areas contain elements whose distinguishing electron is located, respectively, in *s*, *p*, *d*, and *f* subshells. The extent of filling of the subshell that contains an element's distinguishing electron can be determined from the element's position in the periodic table.

the last electron that is added to its electron configuration when subshells are filled in order of increasing energy. This last electron is the one that causes an element's electron configuration to differ from that of the element immediately preceding it in the periodic table.

For all elements that are located in the *s* area of the periodic table, the distinguishing electron is always found in an *s* subshell. All *p* area elements have distinguishing electrons in *p* subshells. Similarly, elements in the *d* and *f* areas of the periodic table have distinguishing electrons located in *d* and *f* subshells, respectively. Thus the area location of an element in the periodic table can be used to determine the type of subshell that contains the distinguishing electron. Note that the element helium belongs to the *s* rather than the *p* area of the periodic table, even though its table position is on the right-hand side. (The reason for this placement of helium will be explained in Section 4.3).

The extent to which the subshell containing an element's distinguishing electron is filled can also be determined from the element's position in the periodic table. All elements in the first column of a specific area contain only one electron in the subshell; all elements in the second column contain two electrons in the subshell; and so on. Thus all elements in the first column of the *p* area (Group IIIA) have an electron configuration ending in p^1. Elements in the second column of the *p* area (Group IVA) have electron configurations ending in p^2; and so on. Similar relationships hold in other areas of the table, as shown in Figure 3.12.

Practice Questions and Problems

3.36 Indicate whether the elements represented by the given pairs of electron configurations have similar chemical properties.
a. $1s^22s^1$ and $1s^22s^2$
b. $1s^22s^22p^6$ and $1s^22s^22p^63s^23p^6$
c. $1s^22s^22p^3$ and $1s^22s^22p^63s^23p^64s^23d^3$
d. $1s^22s^22p^63s^23p^4$ and $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$

3.37 Specify the location of each of the following elements in the periodic table in terms of *s* area, *p* area, *d* area, or *f* area.
a. Magnesium b. Copper c. Uranium d. Bromine

3.38 With the help of the periodic table, for each of the following elements, specify the extent to which the subshell containing the distinguishing electron is filled (s^2, p^3, d^5, etc.).
a. $_{13}Al$ b. $_{23}V$ c. $_{20}Ca$ d. $_{36}Kr$

3.9 Classification of the Elements

Learning Focus

Understand the classification system for elements that is based on electron configuration.

The elements can be classified in several ways. The two most common classification systems are

1. A system based on selected physical properties of the elements, in which they are described as metals or nonmetals. This classification scheme was discussed in Section 3.5.
2. A system based on the electron configurations of the elements, in which elements are described as *noble gas, representative, transition,* or *inner transition elements.*

The classification scheme based on electron configurations of the elements is depicted in Figure 3.13.

▶ The electron configurations of the noble gases will be an important focal point when we consider chemical bonding theory in Chapter 4.

The **noble-gas elements** *are found in the far right column of the periodic table*. They are all gases at room temperature, and they have little tendency to form chemical compounds. With one exception, the distinguishing electron for a noble gas completes the *p* subshell; therefore, noble gases have electron configurations ending in p^6. The exception is helium, in which the distinguishing electron completes the first shell—a shell that has only two electrons. Helium's electron configuration is $1s^2$.

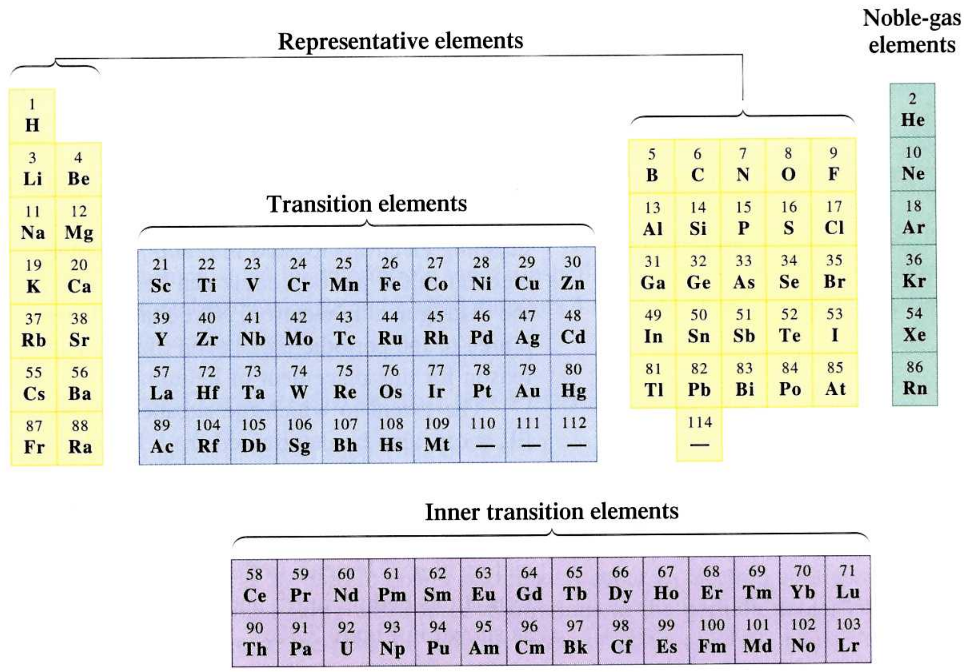


Figure 3.13 A classification scheme for the elements based on their electron configurations. Representative elements occupy the *s* area and most of the *p* area shown in Figure 3.12. The noble-gas elements occupy the last column of the *p* area. The transition elements are found in the *d* area, and the inner transition elements are found in the *f* area.

The unreactive nature of the noble gases is reflected in their occurrence in nature in the form of individual uncombined atoms. There are no known compounds of the lighter noble gases—helium, neon, and argon—and only a very few compounds of the heavier noble gases—krypton, xenon, and radon. Chemical Portraits 5 profiles helium, argon, and krypton, the three noble gases that are totally unreactive (no compounds are known).

Chemical Portraits 5 — Noble Gases: The Most Unreactive of All Elements

Helium (He)

Profile: Helium is the only element that was discovered extraterrestrially before being found on Earth. It was first detected in solar flares associated with the sun's outer surface. On Earth, almost all He is obtained from natural gas, which contains up to 0.3% He. Because of dwindling natural gas supplies, the U.S. government now stockpiles He to ensure its availability in the future. The amount of He present in the Earth's atmosphere is very low.

Uses: Although hydrogen is less dense than helium, He is preferred for use in balloons and lighter-than-air craft such as blimps because it is nonflammable.

Why is He substituted for N_2 in the breathing mixtures used by "deep-sea" divers?

Neon (Ne)

Profile: Neon, like all noble gases, is a colorless, odorless, nonflammable, nontoxic gas. It is three times more abundant than helium and is extracted from liquid air for commercial purposes. A helium-neon laser, emitting red light, was the first continuously operating gas laser.

Uses: Ne gas emits a bright red glow when an electric current is passed through it. This is the basis for its use in luminescent advertising signs and the designation of such signs as "neon signs." (The term *neon sign* has become a generic term that denotes all such advertising signs, even though many of them contain other noble gases than neon.).

What is the relationship between the element neon and bar-code scanners used in grocery stores?

Argon (Ar)

Profile: Argon, with a 1% by volume concentration, is the third-most-abundant atmospheric gas. This concentration is 400 times greater than that of all the other noble gases combined.

Uses: A 93% Ar-7% N_2 mixture is used to fill metal filament light bulbs. This mixture extends the life of the metal filament by slowing its vaporization, a process that causes the light bulb to "darken." Argon also finds use as a "blanketing agent" in high temperature metallurgical welding processes; the blanketing agent prevents the hot metal from reacting with oxygen. Ar is preferred over He for this purpose because its greater density provides better protection.

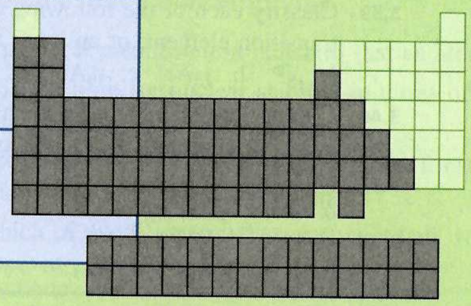
Why is 93% argon rather than 100% argon gas used in metal filament light bulbs?

See the text web site at **http://chemistry.college.hmco.com/students** for answers to the above questions and for further information.

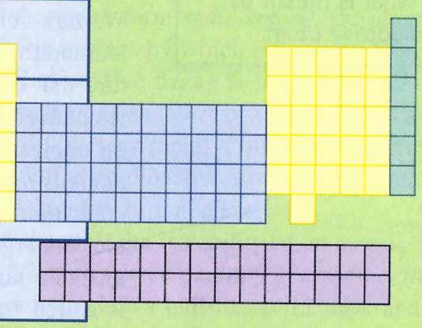
Chemistry *at a Glance*

Element Classification Schemes and the Periodic Table

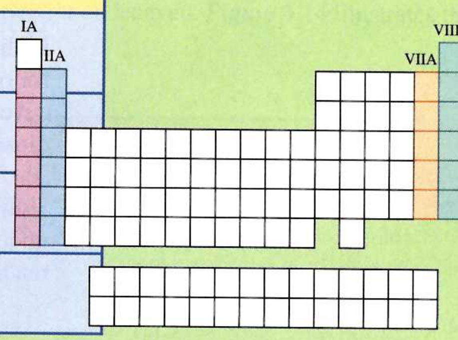
CLASSIFICATION BY PHYSICAL PROPERTIES	
Nonmetals	■ No metallic luster ■ Poor electrical conductivity ■ Good heat insulators ■ Brittle and nonmalleable
Metals	■ Metallic gray or silver luster ■ High electrical and thermal conductivity ■ Malleable and ductile



CLASSIFICATION BY ELECTRONIC PROPERTIES	
Representative elements	■ Found in *s* area and first five columns of the *p* area ■ Some are metals, some nonmetals
Noble-gas elements	■ Found in last column of *p* area plus He (*s* area) ■ All are nonmetals
Transition elements	■ Found in *d* area ■ All are metals
Inner transition elements	■ Found in *f* area ■ All are metals



PERIODIC TABLE GROUPS WITH SPECIAL NAMES	
Alkali metals	■ Group IA elements (except for H, a nonmetal) ■ Electron configurations end in s^1
Alkaline earth metals	■ Group IIA elements ■ Electron configurations end in s^2
Halogens	■ Group VIIA ■ Electron configurations end in p^5
Noble gases	■ Group VIIIA elements ■ Electron configurations end in p^6, except for He, which ends in s^2



The **representative elements** *are all the elements of the* s *and* p *areas of the periodic table, with the exception of the noble gases*. The distinguishing electron in these elements partially or completely fills an *s* subshell or partially fills a *p* subshell. The representative elements include most of the more common elements.

The **transition elements** *are all the elements of the* d *area of the periodic table*. Each has its distinguishing electron in a *d* subshell.

The **inner transition elements** *are all the elements of the* f *area of the periodic table*. Each has its distinguishing electron in an *f* subshell. There is very little variance in the properties of either the 4*f* or the 5*f* series of inner transition elements.

The Chemistry at a Glance above contrasts the three element classification schemes that have been considered so far in this chapter: by physical properties (Section 3.5), by electron configuration (Section 3.9), and by non-numerical periodic table group names (Section 3.4).

Practice Questions and Problems

3.39 Classify each of the following elements as a noble gas, a representative element, a transition element, or an inner transition element.
a. $_{15}P$ b. $_{18}Ar$ c. $_{79}Au$ d. $_{92}U$

3.40 Classify the element with each of the following electron configurations as a noble gas, a representative element, transition element, or inner transition element.
a. $1s^22s^22p^6$
b. $1s^22s^22p^63s^23p^4$
c. $1s^22s^22p^63s^23p^64s^23d^1$
d. $1s^22s^22p^63s^23p^64s^2$

Learning Focus

Understand what is meant by the term *radioactive atom.*

3.10 Nuclear Stability and Radioactivity

We now return to further discussion about the nucleus of an atom. In Section 3.1 we learned that an atomic nucleus, located at the center of an atom, is a small, dense structure that contains all of the protons and neutrons present in an atom. Numerous studies concerning atomic nuclei show that they may be divided into two categories based on nuclear stability. Some nuclei are stable and others are not. A **stable nucleus** *is a nucleus that does not easily undergo change.* Conversely, an **unstable nucleus** *is a nucleus that spontaneously undergoes change.* The spontaneous change that unstable nuclei undergo involves emission of radiation from the nucleus, a process by which an unstable nucleus can become more stable. The radiation emitted from unstable nuclei is called *radioactivity.* **Radioactivity** *is the radiation spontaneously emitted from an unstable nucleus.* Atoms that possess unstable nuclei are said to be *radioactive.* A **radioactive atom** *is an atom with an unstable nucleus from which radiation is spontaneously emitted.*

Of the 88 elements found in nature (Section 1.7), 29 have at least one naturally occurring isotope that has an unstable nucleus and is, therefore, a *radioactive* isotope. Radioactive isotopes are known for *all* 113 elements, even though they occur naturally for only the aforementioned 29 elements. This is because laboratory procedures have been developed by which scientists convert nonradioactive isotopes (stable nucleus) into radioactive isotopes (unstable nucleus).

No simple rule exists for predicting whether a particular nucleus is radioactive. However, considering some observations about those nuclei that are stable is helpful in understanding why some nuclei are stable and others are not. Two generalizations are readily apparent from a study of the properties of stable nuclei found in nature.

1. *There is a correlation between nuclear stability and the total number of nucleons found in a nucleus.* All nuclei with 84 or more protons are unstable. The largest stable nucleus known is that of $^{209}_{83}Bi$, a nucleus that contains 209 nucleons. It thus appears that there is a limit to the number of nucleons that can be packed into a stable nucleus.
2. *There is a correlation between nuclear stability and neutron-to-proton ratio in a nucleus.* The number of neutrons necessary to create a stable nucleus increases as the number of protons increases. For elements of low atomic number, neutron-to-proton ratios for stable nuclei are very close to 1. For heavier elements, stable nuclei have higher neutron-to-proton ratios, and the ratio reaches approximately 1.5 for the heaviest stable elements. These observations suggest that neutrons are at least partially responsible for the stability of a nucleus. It should be remembered that like charges repel each other and that most nuclei contain many protons (with identical positive charges) squeezed together into a very small volume. As the number of protons increases, the forces of repulsion between protons sharply increase.

Therefore, a greater number of neutrons is necessary to counteract the increased repulsions. Finally, at element 84, the repulsive forces become so great that nuclei are unstable regardless of the number of neutrons present.

Practice Questions and Problems

3.41 What physical manifestation indicates that an atom possesses an unstable nucleus?

3.42 What is the limit for nuclear stability in terms of number of nucleons present in a nucleus?

3.43 How do the neutron-to-proton ratios compare for stable nuclei of low atomic number and stable nuclei of high atomic number?

3.44 For which of the following elements would all isotopes be radioactive?
a. $_{76}Os$ b. $_{86}Rn$ c. $_{96}Cm$ d. $_{106}Sg$

Learning Focus

Understand the concept of half-life, and determine the amount of a radionuclide left after a given number of half-lives, or vice versa.

3.11 Half-Life

The term *radioactive isotope,* and its shortened form *radioisotope,* are alternative designations for the term *radioactive atom.* The process by which radioactive isotopes produce radiation is called *radioactive decay.* **Radioactive decay** *is the process whereby an unstable nucleus undergoes change as a result of the emission of radiation.* Not all radioactive isotopes decay at the same rate. Some decay very rapidly; others undergo change at extremely slow rates. This indicates that not all radioactive isotopes are equally unstable. The faster the decay rate, the lower the stability of a nucleus.

The concept of half-life is used to quantify the instability of an atomic nucleus. The **half-life** ($t_{1/2}$) *is the time required for* 1/2 *of any given quantity of a radioactive substance to undergo decay.* For example, if a radionuclide's half-life is 12 days and you have a 4.00-g sample of it, then after 1 half-life (12 days), only 2.00 g of the sample (1/2 of the original amount) will remain undecayed; the other half will have decayed into some other substance (Section 3.12). Similarly, during the next half-life, 1/2 of the 2.00 g remaining will decay, leaving 1/4 of the original atoms (1.00 g) unchanged. After 3 half-lives, 1/8 (1/2 × 1/2 × 1/2) of the original sample will remain undecayed. Figure 3.14 illustrates the radioactive decay curve for a radioisotope.

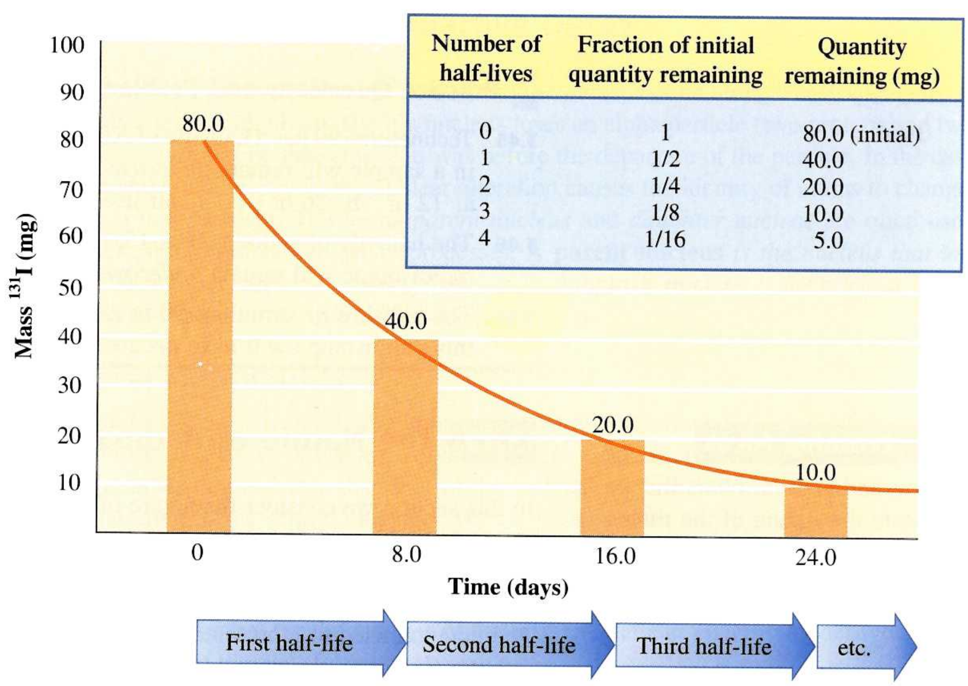


Figure 3.14 Decay of 80.0 mg of ^{131}I, which has a half-life of 8.0 days. After each half-life period, the quantity of material present at the beginning of the period is reduced by half.

▶ Most radionuclides used in diagnostic medicine have short half-lives. This limits to a short time interval the exposure of the human body to radiation.

▶ The half-life for a radionuclide is independent of external conditions such as temperature, pressure, and state of chemical combination.

There is a wide range of half-lives for radionuclides. Half-lives as long as billions of years and as short as a fraction of a second have been determined (Table 3.3). Most naturally occurring radionuclides have long half-lives. However, some radionuclides with *short* half-lives are also found in nature. Naturally occurring mechanisms exist for the continual production of the short-lived species.

The decay rate (half-life) of a radionuclide is constant. It is independent of physical conditions such as temperature, pressure, and state of chemical combination. It depends only on the identity of the radionuclide. For example, radioactive sodium-24 decays at the same rate whether it is incorporated into NaCl, NaBr, Na_2SO_4, or $NaC_2H_3O_2$. If a nuclide is radioactive, nothing will stop it from decaying and nothing will increase or decrease its decay rate.

Table 3.3
Range of Half-lives Found for Naturally Occurring Radionuclides

Element	Half-life ($t_{1/2}$)
vanadium-50	6×10^{15} yr
platinum-190	6.9×10^{11} yr
uranium-238	4.5×10^{9} yr
uranium-235	7.1×10^{8} yr
thorium-230	7.5×10^{4} yr
lead-210	22 yr
bismuth-214	19.7 min
polonium-212	3.0×10^{-7} sec

Example 3.4 Using Half-life to Calculate the Amount of Radioisotope That Remains Undecayed After a Certain Time

Iodine-131 is a radionuclide that is frequently used in nuclear medicine. Among other things, it is used to detect fluid buildup in the brain. The half-life of iodine-131 is 8.0 days. How much of a 16.0-g sample of iodine-131 will remain undecayed after a period of 32 days?

Solution

First, we must determine the number of half-lives that have elapsed.

$$32 \cancel{\text{days}} \times \left(\frac{1 \text{ half-life}}{8.0 \cancel{\text{days}}}\right) = 4 \text{ half-lives}$$

Constructing a tabular summary of the amount of sample remaining after each of the elapsed half-lives yields the following data.

Number of half-lives	Fraction decayed	Amount undecayed
0	none	original amount (16.0 g)
1	1/2	$1/2 \times (16.0\text{ g}) = 8.00$ g
2	$1/2 \times 1/2 = 1/4$	$1/4 \times (16.0\text{ g}) = 4.00$ g
3	$1/2 \times 1/2 \times 1/2 = 1/8$	$1/8 \times (16.0\text{ g}) = 2.00$ g
4	$1/2 \times 1/2 \times 1/2 \times 1/2 = 1/16$	$1/16 \times (16.0\text{ g}) = 1.00$ g

Thus 1.00 g of undecayed ^{131}I remains after 4 half-lives (32 days) have elapsed.

Practice Questions and Problems

3.45 Technetium-99 has a half-life of 6.0 hours. What fraction of the technetium-99 atoms in a sample will remain *undecayed* after the following times?
a. 12 hr b. 36 hr c. 3 half-lives d. 6 half-lives

3.46 The half-life of sodium-24 is 15.0 hr. How many grams of a 4.00-g sample of this radioisotope will remain *undecayed* after 60.0 hr?

3.47 The half-life of strontium-90 is 28 years. How many grams of a 4.00-g sample of this radioisotope will have *decayed* after 112 years?

3.12 The Nature of Radioactive Emissions

Learning Focus

Name and write symbols that indicate the nature of the three most common types of radiation given off by unstable nuclei.

In this section we consider the nature of the three most common types of radiation emanating from the nuclei of radioactive atoms. These radiation types are alpha particles, beta particles, and gamma rays.

An **alpha particle** *is a particle in which two protons and two neutrons are present.* The notation used to represent an alpha particle is $^4_2\alpha$. The numerical subscript indicates that the charge on the particle is +2 (from the two protons). The numerical superscript

Table 3.4
Characteristics of the Three Most Common Types of Radiation Given Off by Radioactive Atoms

Type of radiation	Symbol for radiation	Charge	Relative mass
Alpha	${}^{4}_{2}\alpha$	+2	heavy
Beta	${}^{0}_{-1}\beta$	−1	light
Gamma	${}^{0}_{0}\gamma$	none	none

indicates a mass of 4 amu. Alpha particles are identical to the nuclei of helium-4 (${}^{4}_{2}He$) atoms; because of this, an alternative designation for an alpha particle is ${}^{4}_{2}He$.

A **beta particle** *is a particle whose charge and mass are identical to those of an electron.* However, beta particles are not extranuclear electrons; they are particles that have been produced inside the nucleus and then ejected. We will discuss this process in Section 3.13. The symbol used to represent a beta particle is ${}^{0}_{-1}\beta$. The numerical subscript indicates that the charge on the beta particle is −1; it is the same as that of an electron. The use of the superscript zero for the mass of a beta particle should be interpreted as meaning not that a beta particle has no mass but, rather, that the mass is very close to zero amu. The actual mass of a beta particle is 0.00055 amu.

A **gamma ray** *is a form of high energy radiation without mass or charge.* Gamma rays are very high-energy radiation, somewhat like X rays. The symbol for a gamma ray is ${}^{0}_{0}\gamma$.

Table 3.4 contrasts the physical characteristics of alpha, beta, and gamma types of radiation.

Practice Questions and Problems

3.48 Supply a complete symbol, with superscript and subscript, for each of the following types of radiation.
a. Alpha particle b. Beta particle c. Gamma ray

3.49 Give the charge and the mass (in amu) of each of the following types of radiation.
a. Alpha particle b. Beta particle c. Gamma ray

3.50 State the composition of an alpha particle in terms of protons and neutrons.

3.51 What is the relationship between the characteristics of a beta particle and those of an electron?

3.13 Equations for Radioactive Decay

Learning Focus

Write equations, balanced for mass number and atomic number, to represent various alpha, beta, and gamma decay processes. Understand what is meant by the terms *parent nucleus* and *daughter nucleus*.

Alpha, beta, and gamma emissions come from the nucleus of an atom. These spontaneous emissions alter nuclei; obviously, if a nucleus loses an alpha particle (two protons and two neutrons), it will not be the same as it was before the departure of the particle. In the case of alpha and beta emissions, the nuclear alteration causes the identity of atoms to change, forming a new element. The terms *parent nucleus* and *daughter nucleus* are often used when describing radioactive decay processes. A **parent nucleus** *is the nucleus that undergoes decay in a radioactive decay process.* A **daughter nucleus** *is the nucleus that is produced as a result of a radioactive decay process.*

Alpha Particle Decay

Alpha particle decay, which is the emission of an alpha particle from a nucleus, always results in the formation of a nucleus of a different element. The product nucleus has an atomic number that is 2 less than that of the original nucleus and a mass number that is 4 less than that of the original nucleus. We can represent alpha particle decay in general terms by the equation

$$ {}^{A}_{Z}X \longrightarrow {}^{4}_{2}\alpha + {}^{A-4}_{Z-2}Y $$

▸ Loss of an alpha particle from an unstable nucleus results in (1) a decrease of 4 units in the mass number (A) and (2) a decrease of 2 units in the atomic number (Z).

where X is the symbol for the nucleus of the original element undergoing decay and Y is the symbol of the element formed as a result of the decay.

▶ Note that the symbols in nuclear equations stand for *nuclei* rather than atoms. We do not worry about electrons when writing nuclear equations.

Let us write equations for two alpha particle decay processes. Both $^{211}_{83}Bi$ and $^{238}_{92}U$ are radioisotopes that undergo alpha particle decay. The nuclear equations for these two decay processes are

$$^{211}_{83}\text{Bi} \longrightarrow {}^{4}_{2}\alpha + {}^{207}_{81}\text{Tl}$$
$$^{238}_{92}\text{U} \longrightarrow {}^{4}_{2}\alpha + {}^{234}_{90}\text{Th}$$

where thallium-207 and thorium-234 are the daughter nuclei.

In a **balanced nuclear equation,** *the sums of the subscripts (atomic numbers or particle charges) on both sides of the equation are equal, and the sums of the superscripts (mass numbers) on both sides of the equation are equal.*

Both of our example equations are balanced. In the alpha decay of $^{211}_{83}Bi$, the subscripts on both sides total 83, and the superscripts total 211. For the alpha decay of $^{238}_{92}U$, the subscripts total 92 on both sides, and the superscripts total 238.

▶ The rules for balancing nuclear equations are
1. The sum of the subscripts must be the same on both sides of the equation.
2. The sum of the superscripts must be the same on both sides of the equation.

Beta Particle Decay

Beta particle decay always results in the formation of a nucleus of a different element. The mass number of the new nucleus is the same as that of the original atom. However, the atomic number has increased by 1 unit. The general equation for beta decay is

$$^{A}_{Z}\text{X} \longrightarrow {}^{\ 0}_{-1}\beta + {}_{Z+1}^{\quad A}\text{Y}$$

▶ Loss of a beta particle from an unstable nucleus results in (1) no change in the mass number (A) and (2) an increase of 1 unit in the atomic number (Z).

Specific examples of beta particle decay are

$$^{10}_{4}\text{Be} \longrightarrow {}^{\ 0}_{-1}\beta + {}^{10}_{5}\text{B}$$
$$^{234}_{90}\text{Th} \longrightarrow {}^{\ 0}_{-1}\beta + {}^{234}_{91}\text{Pa}$$

Both of these nuclear equations are balanced; superscripts and subscripts add to the same sums on both sides of the equation.

At this point in the discussion, you may be wondering how a nucleus, which is composed only of neutrons and protons, ejects a negative particle (beta particle) when no such particle is present in the nucleus. Explained simply, a neutron in the nucleus is transformed into a proton and a beta particle through a complex series of steps; that is,

$$\text{Neutron} \longrightarrow \text{proton} + \text{beta particle}$$
$$^{1}_{0}\text{n} \longrightarrow {}^{1}_{1}\text{p} + {}^{\ 0}_{-1}\beta$$

Once it is formed within the nucleus, the beta particle is ejected with a high velocity. Note the symbols used to denote a neutron ($^{1}_{0}n$; no charge and a mass of 1 amu) and a proton ($^{1}_{1}p$; a +1 charge and a mass of 1 amu).

Gamma Ray Emission

For naturally occurring radioisotopes, gamma ray emission almost always takes place in conjunction with an alpha or a beta decay process; it never occurs independently. These gamma rays are often not included in the nuclear equation because they do not affect the balancing of the equation or the identity of the daughter nucleus. This can be seen from the following two nuclear equations.

▶ Gamma rays are to nuclear reactions what heat is to ordinary chemical reactions.

$$^{226}_{88}\text{Ra} \longrightarrow {}^{222}_{86}\text{Rn} + {}^{4}_{2}\alpha + {}^{0}_{0}\gamma$$

Balanced nuclear equation with gamma radiation included

$$^{226}_{88}\text{Ra} \longrightarrow {}^{222}_{86}\text{Rn} + {}^{4}_{2}\alpha$$

Balanced nuclear equation with gamma radiation omitted

▶ Among *synthetically* produced radioisotopes, pure "gamma emitters," radioisotopes that give off gamma rays but no alpha or beta particles, occur. These radioisotopes are important in diagnostic nuclear medicine. Pure "gamma emitters" are not found among naturally occurring radioisotopes.

The fact that gamma rays are often left out of balanced nuclear equations should not be interpreted to mean that such rays are not important in nuclear chemistry. On the contrary, gamma rays are more important than alpha and beta particles when the effects of external radiation exposure on living organisms are considered.

Example 3.5 Writing Balanced Nuclear Equations, Given the Parent Nucleus and Its Mode of Decay

Write a balanced nuclear equation for the decay of each of the following radioactive nuclei. The mode of decay is indicated in parentheses.

a. ${}^{70}_{31}\text{Ga}$ (beta emission) **b.** ${}^{144}_{60}\text{Nd}$ (alpha emission)
c. ${}^{248}_{100}\text{Fm}$ (alpha emission) **d.** ${}^{113}_{47}\text{Ag}$ (beta emission)

Solution

In each case, the atomic and mass numbers of the daughter nucleus are obtained by writing the symbols of the parent nucleus and the particle emitted by the nucleus (alpha or beta). Then the equation is balanced.

a. Let X represent the product of the radioactive decay, the daughter nucleus. Then

$$ {}^{70}_{31}\text{Ga} \longrightarrow {}^{0}_{-1}\beta + \text{X} $$

The sums of the superscripts on both sides of the equation must be equal, so the superscript for X must be 70. In order for the sums of the subscripts on both sides of the equation to be equal, the subscript for X must be 32. Then $31 = (-1) + (32)$. As soon as we determine the subscript of X, we can obtain the identity of X by looking at a periodic table. The element with an atomic number of 32 is Ge (germanium). Therefore,

$$ {}^{70}_{31}\text{Ga} \longrightarrow {}^{0}_{-1}\beta + {}^{70}_{32}\text{Ge} $$

b. Letting X represent the product of the radioactive decay, we have, for the alpha decay of ${}^{144}_{60}\text{Nd}$,

$$ {}^{144}_{60}\text{Nd} \longrightarrow {}^{4}_{2}\alpha + \text{X} $$

We balance the equation by making the superscripts on each side of the equation total 144 and the subscripts total 60. We get

$$ {}^{144}_{60}\text{Nd} \longrightarrow {}^{4}_{2}\alpha + {}^{140}_{58}\text{Ce} $$

c. Similarly, we write

$$ {}^{248}_{100}\text{Fm} \longrightarrow {}^{4}_{2}\alpha + \text{X} $$

Balancing superscripts and subscripts, we get

$$ {}^{248}_{100}\text{Fm} \longrightarrow {}^{4}_{2}\alpha + {}^{244}_{98}\text{Cf} $$

In alpha emission, the atomic number of the daughter nucleus always decreases by 2, and the mass number of the daughter nucleus always decreases by 4.

d. Finally, we write

$$ {}^{113}_{47}\text{Ag} \longrightarrow {}^{0}_{-1}\beta + \text{X} $$

In beta emission, the atomic number of the daughter nucleus always increases by 1, and the mass number does not change from that of the parent. The balancing procedure gives us the result

$$ {}^{113}_{47}\text{Ag} \longrightarrow {}^{0}_{-1}\beta + {}^{113}_{48}\text{Cd} $$

Practice Exercises and Questions

3.52 Write balanced nuclear equations for the alpha particle decay of the following radioisotopes.
a. ${}^{200}_{84}\text{Po}$ b. ${}^{244}_{96}\text{Cm}$ c. Curium-240 d. Uranium-238

3.53 Write balanced nuclear equations for the beta particle decay of the following radioisotopes.
a. ${}^{10}_{4}\text{Be}$ b. ${}^{77}_{32}\text{Ge}$ c. Iron-60 d. Sodium-25

3.54 What is the effect on the mass number and atomic number of the parent nucleus when alpha particle decay occurs?

3.55 What is the effect on the mass number and atomic number of the parent nucleus when beta particle decay occurs?

3.56 Supply the missing symbol in each of the following radioactive decay equations.
a. ${}^{34}_{14}Si \rightarrow {}^{34}_{15}P + ?$ b. $? \rightarrow {}^{28}_{13}Al + {}^{0}_{-1}\beta$
c. ${}^{252}_{99}Es \rightarrow {}^{248}_{97}Bk + ?$ d. ${}^{204}_{82}Pb \rightarrow ? + {}^{4}_{2}\alpha$

3.57 Identify the mode of decay for each of the following radioactive decays, where the parent and daughter nuclei are given.
a. Parent = platinum-190; daughter = osmium-186
b. Parent = oxygen-19; daughter = fluorine-19

Learning Focus

Contrast the biological effects of alpha, beta, and gamma radiation.

3.14 Biological Effects of Radiation

The alpha, beta, and gamma radiations produced from radioactive decay possess extremely high amounts of energy. It is this high-energy content that makes radiation dangerous to living organisms. As the radiations travel outward from their nuclear sources into the material surrounding the radioactive substance, this energy is dissipated through collisions with the atoms and molecules of the surrounding materials. In the great majority of radiation–atom and radiation–molecule interactions, electrons are knocked away from the atoms and molecules involved, producing very reactive species with chemistries much different from the species from which they were produced. It is the presence of these radiation-produced, very reactive species that leads to "radiation damage" in living cells.

The three types of naturally occurring radioactive emissions—alpha particles, beta particles, and gamma rays—differ in their ability to penetrate matter and cause cellular damage. Consequently, the extent of the biological effects of radiation depends on the type of radiation involved.

Alpha Particle Effects

Alpha particles are the most massive and also the slowest particles involved in natural radioactive decay processes. Maximum alpha particle velocities are on the order of one-tenth of the speed of light. For a given alpha-emitting radioisotope all alpha particles have the same energy; different alpha-emitting radioisotopes, however, produce alpha particles of differing energies.

▶ The speed of light, 3.0×10^8 m/sec (186,000 miles/sec), is the maximum limit of velocity. Objects cannot travel faster than the speed of light.

Because of their "slowness," alpha particles have low penetrating power and cannot penetrate the body's outer layers of skin. The major damage from alpha radiation occurs when alpha-emitting radioisotopes are ingested—for example, in contaminated food. There are no protective layers of skin within the body.

Beta Particle Effects

Unlike alpha particles, which are all emitted with the same discrete energy from a given radioisotope, beta particles emerge from a beta-emitting substance with a continuous range of energies up to a specific limit that is characteristic of the particular radioisotope. Maximum beta particle velocities are on the order of nine-tenths of the speed of light.

With their greater velocity, beta particles can penetrate much deeper than alpha particles and can cause severe skin burns if their source remains in contact with the skin for an appreciable time. Because of their much smaller size, they do not collide with as many molecules as do alpha particles. An alpha particle is approximately 8000 times heavier than a beta particle. A typical alpha particle travels about 6 cm in air and collides with 40,000 molecules and/or atoms and a typical beta particle travels 1000 cm in air and collides with about 2000 molecules and/or atoms. Internal exposure to beta radiation is as serious as internal alpha exposure.

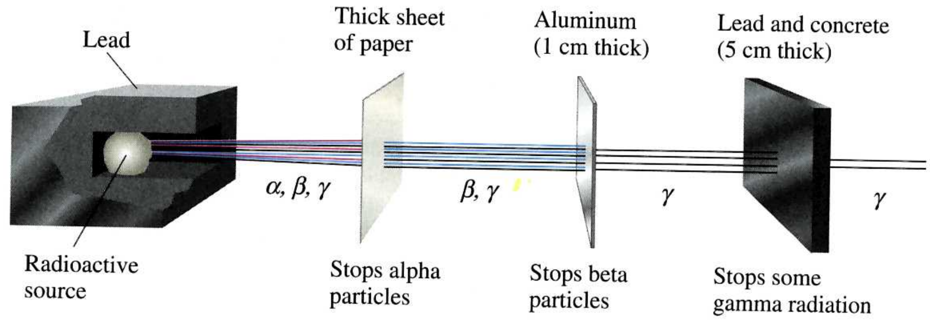


Figure 3.15 Alpha, beta, and gamma radiations differ in penetrating ability.

Gamma Radiation Effects

Gamma radiation is released at a velocity equal to that of the speed of light. Gamma rays readily penetrate deeply into organs, bone, and tissue.

Figure 3.15 contrasts the abilities of alpha, beta, and gamma radiations to penetrate paper, aluminum foil, and a thin layer of a lead–concrete mixture.

▶ X rays (which come from non-nuclear sources) and gamma rays are similar except that X rays are of lower energy. X rays used in diagnostic medicine have energies approximately 10% of that of gamma rays.

Practice Exercises and Questions

3.58 Contrast the abilities of alpha, beta, and gamma radiations to penetrate a thick sheet of paper.

3.59 Contrast the abilities of alpha, beta, and gamma radiations to penetrate human skin.

3.60 Contrast the velocities with which alpha, beta, and gamma radiations are emitted by unstable nuclei.

3.61 Contrast the distance that alpha and beta particles can travel in air, and the number of collisions they undergo in traveling that distance, before their excess energy is dissipated.

The next Chemistry at a Glance summarizes the terminology and concepts we have considered so far that are related to unstable nuclei and the radiations that they give off.

3.15 Nuclear Medicine

Learning Focus

Know the basic principles governing the use of radioisotopes in diagnostic and therapeutic nuclear medicine.

In medicine, radioisotopes are used both diagnostically and therapeutically. In diagnostic applications, technicians use small amounts of radioisotopes whose progress through the body or localization in specific organs can be followed. Larger quantities of radioisotopes are used in therapeutic applications.

▶ An additional use for radioisotopes in medicine, besides diagnostic and therapeutic uses, is as a source of power (Section 11.12). Cardiac pacemakers powered by plutonium-238 can remain in a patient for longer periods than those powered by chemical batteries without the additional surgery required to replace batteries.

Diagnostic Uses for Radioisotopes

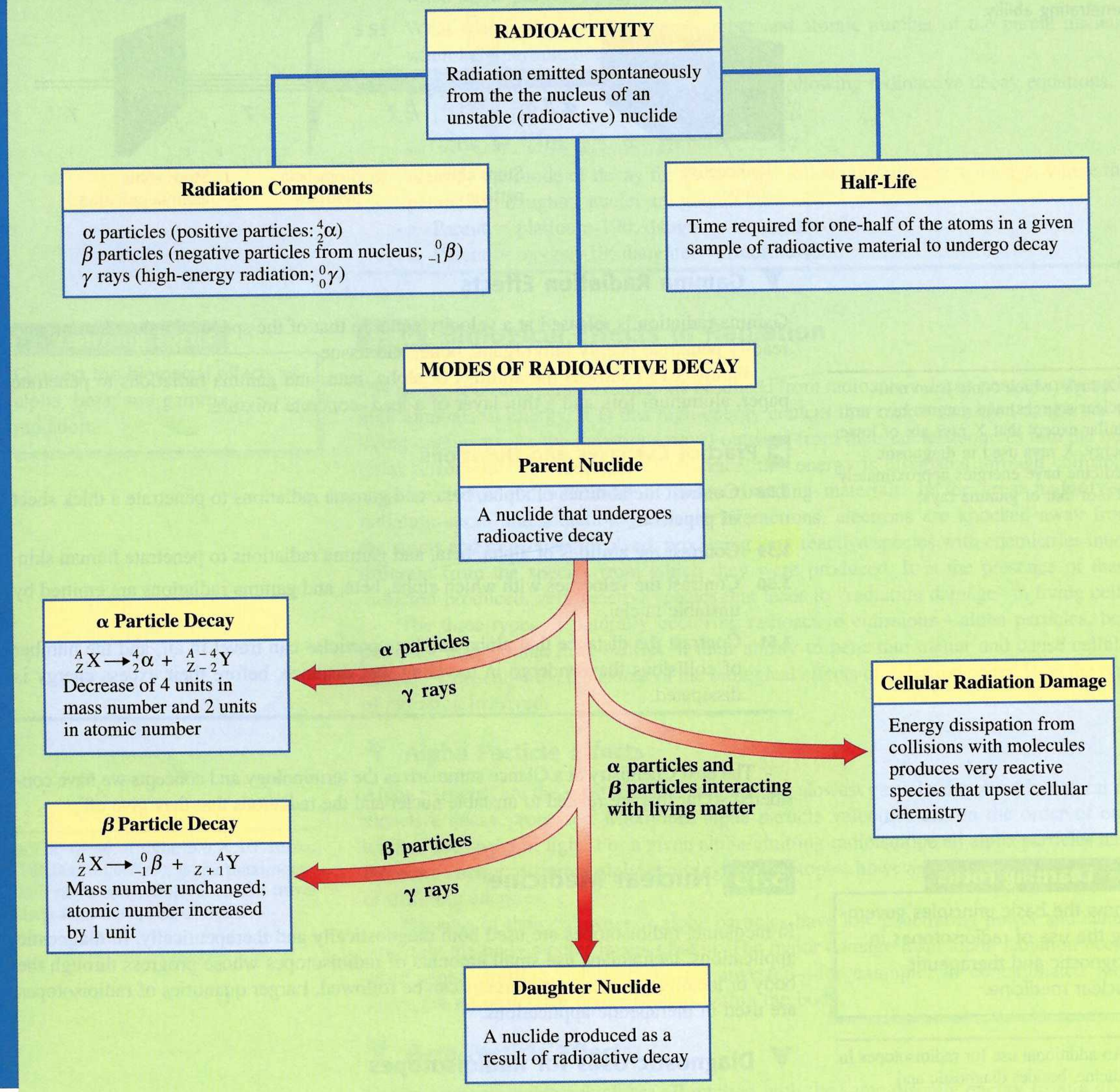
The fundamental chemical principle behind the use of radioisotopes in diagnostic medical work is the fact that a radioactive nucleus of an element has the same chemical properties as a nonradioactive nucleus of the element. Thus body chemistry is not upset by the presence of a small amount of a radioactive substance whose nonradioactive form is already present in the body.

The criteria used in selecting radioisotopes for diagnostic procedures include the following:

1. At low concentrations (to minimize radiation damage), the radioisotope must be detectable by instrumentation placed outside the body. Almost all diagnostic radioisotopes are gamma emitters, because the penetrating power of alpha and beta particles is too low.

Chemistry *at a Glance*

Terminology Associated with Nuclear Reactions



2. The radioisotope must have a short half-life so that the intensity of the radiation is sufficiently great to be detected. A short half-life also limits the time period of radiation exposure.
3. The radioisotope must have a known mechanism for elimination from the body so that the material does not remain in the body indefinitely.
4. The chemical properties of the radioisotope must be such that it is compatible with normal body chemistry. It must be able to be selectively transmitted to the part or system of the body that is under study.

The circulation of blood in the body can be followed by using radioactive sodium-24. A small amount of this isotope is injected into the bloodstream in the form of a sodium chloride solution. The movement of this radioisotope through the circulatory system can be followed easily with radiation detection equipment. If it takes longer than normal for the isotope to show up at a particular part of the body, this is an indication that the circulation is impaired at that spot.

Radiologists evaluate the functioning of the thyroid gland by administering iodine-131, usually in the form of a sodium iodide (NaI) solution. The radioactive iodine behaves in the same manner as ordinary iodine and is absorbed by the thyroid at a rate related to the activity of the gland. If a hypothyroid condition exists, then the amount accumulated is less than normal; and if a hyperthyroid condition exists, then a greater-than-average amount accumulates.

The size and shape of organs, as well as the presence of tumors, can be determined in some situations by scanning the organ in which a radioisotope tends to concentrate. Iodine-131 and technetium-99 are used to generate thyroid and brain scans, respectively. In the brain, technetium-99 concentrates in brain tumors more than in normal brain tissue; this helps radiologists determine the presence, size, and location of brain tumors. Figure 3.16 shows a brain scan obtained by using the radioisotope technetium-99. In this figure, the bright spot at the upper right indicates a tumor that has absorbed a greater amount of radioactive material than the normal brain tissue.

▼ Therapeutic Uses for Radioisotopes

The objectives in therapeutic radioisotope use are entirely different from those for diagnostic procedures. The main objective in the therapeutic use of radioisotopes is to *selectively destroy* abnormal (usually cancerous) cells. The radioisotope is often, but not always, placed within the body. Therapeutic radioisotopes implanted in the body are usually alpha or beta emitters, because an intense dose of radiation in a small localized area is needed.

A commonly used implantation radioisotope that is effective in the localized treatment of tumors is yttrium-90, a beta emitter with a half-life of 64 hr. Yttrium-90 salts are implanted by inserting small, hollow needles into the tumor.

External, high-energy beams of gamma radiation are also extensively used in the treatment of certain cancers. Cobalt-60 is frequently used for this purpose; a beam of radiation is focused on the small area of the body where the tumor is located (see Figure 3.17).

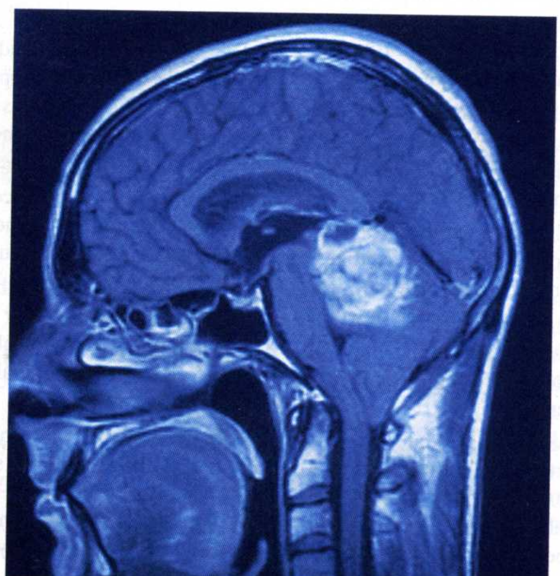


Figure 3.16 Brain scans, such as this one, are obtained using radioactive technetium-99, a laboratory-produced radioisotope.

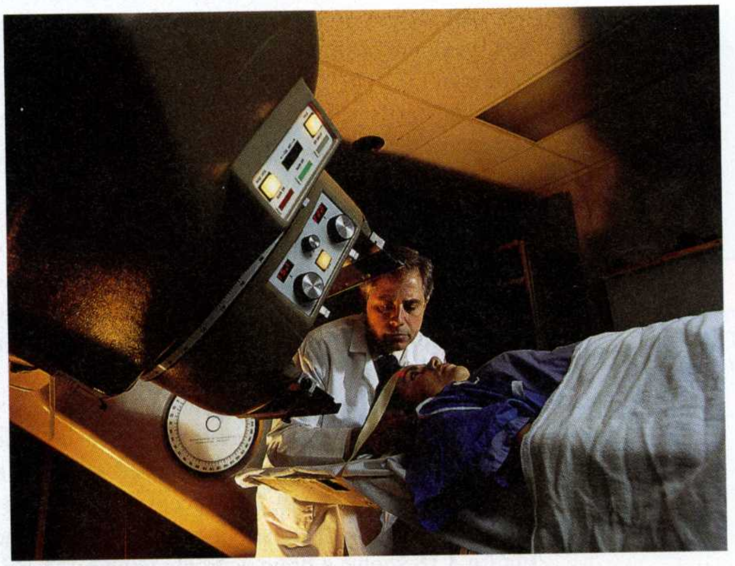


Figure 3.17 Cobalt-60 is used as a source of gamma radiation in radiation therapy.

Practice Exercises and Questions

3.62 Why are the radioisotopes used for diagnostic procedures usually gamma emitters?

3.63 Why do the radioisotopes used in diagnostic procedures nearly always have short half-lives?

3.64 How do the radioisotopes used for therapeutic purposes differ from the radioisotopes used for diagnostic purposes?

3.65 Contrast the different ways in which cobalt-60 and yttrium-90 are used in radiation therapy.

CONCEPTS TO REMEMBER

Subatomic particles. Subatomic particles, the very small building blocks from which atoms are made, are of three major types: electrons, protons, and neutrons. Electrons are negatively charged, protons are positively charged, and neutrons have no charge. All neutrons and protons are found at the center of the atom in the nucleus. The electrons occupy the region about the nucleus. Protons and neutrons have much larger masses than the electron.

Atomic number and mass number. Each atom has a characteristic atomic number and mass number. The atomic number is equal to the number of protons in the nucleus of the atom. The mass number is equal to the total number of protons and neutrons in the nucleus.

Isotopes. Isotopes are atoms that have the same number of protons and electrons but have different numbers of neutrons. The isotopes of an element always have the same atomic number and different mass numbers. Isotopes of an element have the same chemical properties.

Atomic mass. The atomic mass of an element is a calculated average mass. It depends on the percentage abundances and masses of the naturally occurring isotopes of the element.

Periodic law and periodic table. The periodic law states that when elements are arranged in order of increasing atomic number, elements with similar chemical properties occur at periodic (regularly recurring) intervals. The periodic table is a graphical representation of the behavior described by the periodic law. In a modern periodic table, vertical columns contain elements with similar chemical properties. A group in the periodic table is a vertical column of elements. A period in the periodic table is a horizontal row of elements.

Metals and nonmetals. Metals exhibit luster, thermal conductivity, electrical conductivity, and malleability. Nonmetals are characterized by the absence of the properties associated with metals. The majority of the elements are metals. The steplike heavy line that runs through the right third of the periodic table separates the metals on the left from the nonmetals on the right.

Electron shell. A shell contains electrons that have approximately the same energy and spend most of their time approximately the same distance from the nucleus.

Electron subshell. A subshell contains electrons that all have the same energy. The number of subshells in a particular shell is equal to the shell number. Each subshell can hold a specific maximum number of electrons. These values are 2, 6, 10, and 14 for *s*, *p*, *d*, and *f* subshells, respectively.

Electron orbital. An orbital is a region of space about a nucleus where an electron with a specific energy is most likely to be found. Each subshell consists of one or more orbitals. For *s*, *p*, *d*, and *f* subshells there are 1, 3, 5, and 7 orbitals, respectively. No more than two electrons may occupy any orbital.

Electron configuration. An electron configuration is a statement of how many electrons an atom has in each of its subshells. The principle that electrons normally occupy the lowest-energy subshell available is used to write electron configurations.

Electron configurations and the periodic law. Chemical properties repeat themselves in a regular manner among the elements because electron configurations repeat themselves in a regular manner among the elements.

Electron configurations and the periodic table. The groups of the periodic table consist of elements with similar electron configurations. Thus the location of an element in the periodic table can be used to obtain information about its electron configuration.

Classification system for the elements. On the basis of electron configuration, elements can be classified into four categories: noble gases (far right column of the periodic table); representative elements (*s* and *p* areas of the periodic table, with the exception of the noble gases); transition elements (*d* area of the periodic table); and inner transition elements (*f* area of the periodic table).

Nuclear stability and radioactivity. Some atoms possess nuclei that are unstable. To achieve stability, these unstable nuclei spontaneously emit energy (radiation). Such atoms are said to be radioactive.

Emissions from radioactive nuclei. The types of radiation emitted by naturally occurring radioactive nuclei are alpha, beta, and gamma. These radiations can be characterized by mass and charge values. Alpha particles carry a positive charge, beta particles a negative charge, and gamma radiation no charge.

Half-life. Every radioisotope decays at a characteristic rate given by its half-life. One half-life is the time required for half of any given quantity of a radioactive substance to undergo decay.

Balancing nuclear equations. Atomic numbers (subscripts) and mass numbers (superscripts) are always shown when nuclei are represented in a nuclear equation. The balancing of nuclear equations is based on both *charge* and *nucleon* conservation. Charge conversion means that the sum of the subscripts for the products must equal the sum of the subscripts for the reactants. Nucelon conservation means that the sum of the superscripts for the products equals the sum of the superscripts for the reactants.

Biological effects of radiation. The biological effects of radiation depend on the energy, size, and penetrating ability of the radiation. Alpha particles have the greatest size, and gamma rays have the greatest penetrating ability.

Nuclear medicine. Radioisotopes are used in medicine for both diagnosis and therapy. Diagnostic radioisotopes are generally gamma emitters, whereas therapeutic radioisotopes are often alpha and beta emitters. The choice of radioisotope is dictated by the purpose of its use as well as by the target organ.

KEY REACTIONS AND EQUATIONS

1. Relationships involving atomic number and mass number for a neutral atom (Section 3.2)
   Atomic number = number of protons = number of electrons
   Mass number = number of protons + number of neutrons
   Mass number = total number of subatomic particles in the nucleus
   Mass number − atomic number = number of neutrons
   Mass number + atomic number = total number of subatomic particles
2. Relationships involving electron shells, electron subshells, and electron orbitals (Section 3.6)
   Number of subshells in a shell = shell number
   Maximum number of electrons in an *s* subshell = 2
   Maximum number of electrons in a *p* subshell = 6
   Maximum number of electrons in a *d* subshell = 10
   Maximum number of electrons in an *f* subshell = 14
   Maximum number of electrons in an orbital = 2
3. Order of filling of subshells in terms of increasing energy (Section 3.7)
   1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*, 5*f*, 6*d*, 7*p*
4. General equation for alpha decay (Section 3.13)
   $$^{A}_{Z}\text{X} \longrightarrow {}^{A-4}_{Z-2}\text{Y} + {}^{4}_{2}\alpha$$
5. General equation for beta decay (Section 3.13)
   $$^{A}_{Z}\text{X} \longrightarrow {}^{A}_{Z+1}\text{Y} + {}^{0}_{-1}\beta$$

KEY TERMS

Alpha particle (3.12)
Atomic mass (3.3)
Atomic number (3.2)
Balanced nuclear equation (3.13)
Beta particle (3.12)
Daughter nucleus (3.13)
Distinguishing electron (3.8)
Electron (3.1)
Electron configuration (3.7)
Electron orbital (3.6)
Electron shell (3.6)
Electron subshell (3.6)
Element (3.2)
Gamma ray (3.12)
Group (3.4)
Half-life (3.11)
Inner transition elements (3.9)
Isotopes (3.3)
Mass number (3.2)
Metal (3.5)
Neutron (3.1)
Noble-gas elements (3.9)
Nonmetal (3.5)
Nucleon (3.1)
Nucleus (3.1)
Parent nucleus (3.13)
Period (3.4)
Periodic law (3.4)
Periodic table (3.4)
Proton (3.1)
Radioactive atom (3.10)
Radioactive decay (3.11)
Radioactivity (3.10)
Representative elements (3.9)
Stable nucleus (3.10)
Subatomic particles (3.1)
Transition elements (3.9)
Unstable nucleus (3.10)

ADDITIONAL PROBLEMS

3.66 With the help of the periodic table, complete the following table.

Element	Symbol	Atomic number	Mass number	Number of protons	Number of neutrons
(a) ____	$^{3}_{2}He$	____	____	____	____
(b) nickel	____	____	60	____	____
(c) ____	____	18	37	____	____
(d) ____	____	____	90	____	52
(e) ____	$^{235}_{92}U$	____	____	____	____
(f) ____	____	____	____	17	20
(g) ____	____	____	232	94	____
(h) ____	$^{32}_{16}S$	____	____	____	____
(i) iron	____	____	56	____	____
(j) calcium	____	____	____	____	20

3.67 Write the complete symbol ($^{A}_{Z}E$), with the help of the periodic table, for atoms with the following characteristics.
a. Contains 20 electrons and 24 neutrons
b. Beryllium atom with one more neutron than proton
c. Silver atom that contains 157 subatomic particles
d. Beryllium atom that contains 9 nucleons

3.68 Characterize each of the following pairs of atoms as containing (1) the same number of neutrons, (2) the same number of electrons, or (3) the same total number of subatomic particles.
a. $^{13}_{6}C$ and $^{14}_{7}N$ b. $^{18}_{8}O$ and $^{19}_{9}F$
c. $^{37}_{17}Cl$ and $^{36}_{18}Ar$ d. $^{35}_{17}Cl$ and $^{37}_{17}Cl$

3.69 Using the information given in the following table, indicate whether each of the following pairs of atoms are isotopes.

	Atom A	Atom B	Atom C	Atom D
Number of protons	9	10	10	9
Number of neutrons	10	9	10	9
Number of electrons	9	10	10	9

a. A and B b. A and C c. A and D d. B and C

3.70 Three naturally occurring isotopes of magnesium exist: magnesium-24, magnesium-25, and magnesium-26. The percentage abundances and masses of these three isotopes are, respectively, (78.99%, 23.99 amu), (10.00%, 24.99 amu), and (11.01%, 25.98 amu). Calculate the atomic mass of magnesium.

3.71 The atomic mass of fluorine is 18.998 amu, and all fluorine atoms have this mass. The atomic mass of iron is 55.847 amu, and not a single iron atom has this mass. Explain.

3.72 Which of the six elements nitrogen, beryllium, argon, aluminum, silver, and gold belong(s) in each of the following classifications?
a. Period number and Roman-numeral group number are numerically equal
b. Readily conducts electricity and heat
c. Has an atomic mass greater than its atomic number
d. All atoms have a nuclear charge greater than +20

3.73 Write electron configurations for the following elements.
a. The Group IIIA element in the same period as $_{4}Be$
b. The Period 3 element in the same group as $_{5}B$
c. The lowest-atomic-numbered metal in Group IA
d. The highest-atomic-numbered metal that contains only *s* electrons

3.74 Write electron configurations for the following atoms.
a. An atom with 13 electrons
b. An atom with 13 protons
c. An atom with a +13 charge on its nucleus
d. An atom with an atomic number of 13

3.75 With the help of the periodic table, identify the element of lowest atomic number whose electron configuration
a. contains one or more *p* electrons
b. contains one or more *d* electrons
c. contains three or more *s* electrons
d. contains nine or more *p* electrons

3.76 Write nuclear equations for each of the following radioactive decay processes.
a. Thallium-206 is formed by beta emission.
b. Palladium-109 undergoes beta emission.
c. Plutonium-241 is formed by alpha emission.
d. Fermium-249 undergoes alpha emission.

3.77 Cobalt-55 has a half-life of 18 hours. How long will it take, in hours, for the following fractions of atoms in a cobalt-55 sample to undergo decay?
a. 7/8 b. 31/32 c. 63/64 d. 127/128

PRACTICE TEST ▶ True/False

3.78 All protons and neutrons present in an atom are found in its nucleus.

3.79 A nucleon is any subatomic particle that is uncharged (neutral).

3.80 The difference between the mass number and atomic number for an atom gives to the number of electrons present in the atom.

3.81 Isotopes of an element have the same chemical properties because their electron configurations are identical.

3.82 In the standard periodic table, elements with similar chemical properties are always found in the same period.

3.83 The majority of the elements are metals rather than nonmetals.

3.84 The maximum number of electrons that any electron subshell can accommodate is six.

3.85 Electron configurations give the number of electrons that are located in various electron orbitals.

3.86 The number of columns of elements in the *s* area of the periodic table is two.

3.87 The noble-gas elements are found in the far-right column of the periodic table.

3.88 Atoms that possess unstable nuclei are said to be radioactive.

3.89 In a radioactive sample, half of the radioactive atoms undergo decay during the first half-life, and the other half undergo decay during the second half-life.

3.90 Alpha particles are much heavier than beta particles and travel much faster.

3.91 Beta particle decay always results in the formation of an atom of a different element.

3.92 The composition of an alpha particle is two protons and two electrons.

PRACTICE TEST ▶ Multiple Choice

3.93 Which of the following collections of subatomic particles would have the greatest mass?
a. 4 electrons and 1 proton
b. 2 neutrons and 1 electron
c. 1 proton and 2 neutrons
d. 1 electron, 1 proton, and 1 neutron

3.94 The nucleus of an atom
a. contains all subatomic particles present in the atom
b. is negatively charged because of the presence of electrons
c. is neutral because it contains only neutrons
d. accounts for only a small amount of the total volume of an atom

3.95 An atom contains 26 protons, 30 neutrons, and 26 electrons. The atomic number and mass number for this atom are, respectively,
a. 30 and 26
b. 26 and 30
c. 26 and 56
d. 30 and 56

3.96 An atom of $^{27}_{14}Si$ contains
a. more protons than neutrons
b. more electrons than protons
c. the same equal number of protons and neutrons
d. more neutrons than electrons

3.97 Isotopes of a given element have
a. the same mass number but different numbers of protons
b. different mass numbers but the same number of protons
c. the same atomic number but different chemical properties
d. the same mass number but different chemical properties

3.98 Chlorine, which exists in nature in two isotopic forms, has an atomic mass of 35.5 amu. This means that
a. all chlorine atoms have masses of 35.5 amu
b. some, but not all, chlorine atoms have masses of 35.5 amu
c. 35.5 amu is the upper limit for the mass of a chlorine atom
d. no chlorine atoms have masses of 35.5 amu

3.99 Which of the following statements is consistent with the electron configuration $1s^22s^22p^63s^23p^6$?
a. There are 6 electrons present in a 3*p* orbital.
b. There are 6 electrons present in a 3*p* subshell.
c. There are 6 electrons present in a 3*p* shell.
d. There are 6 electrons present in the third shell.

3.100 The elements in group IVA of the periodic table all have electron configurations ending in
a. p^2
b. p^4
c. d^4
d. s^2

3.101 The daughter nucleus produced by the beta decay of $^{234}_{90}Th$ is
a. $^{230}_{89}Ac$ b. $^{234}_{91}Pa$ c. $^{234}_{92}U$ d. $^{230}_{88}Ra$

3.102 The loss of an alpha particle by a radioactive atom causes
a. no change in mass number
b. the atomic number to decrease by 4
c. the mass number to decrease by 2
d. the atomic number to decrease by 2