CHAPTER 6

The Periodic Table and Periodic Law

What You'll Learn

- You will explain why elements in a group have similar properties.
- You will relate the group and period trends seen in the periodic table to the electron configuration of atoms.
- You will identify the s-, p-, d-, and f-blocks of the periodic table.

Why It's Important

The periodic table is the single most powerful chemistry reference tool available to you. Understanding its organization and interpreting its data will greatly aid you in your study of chemistry.

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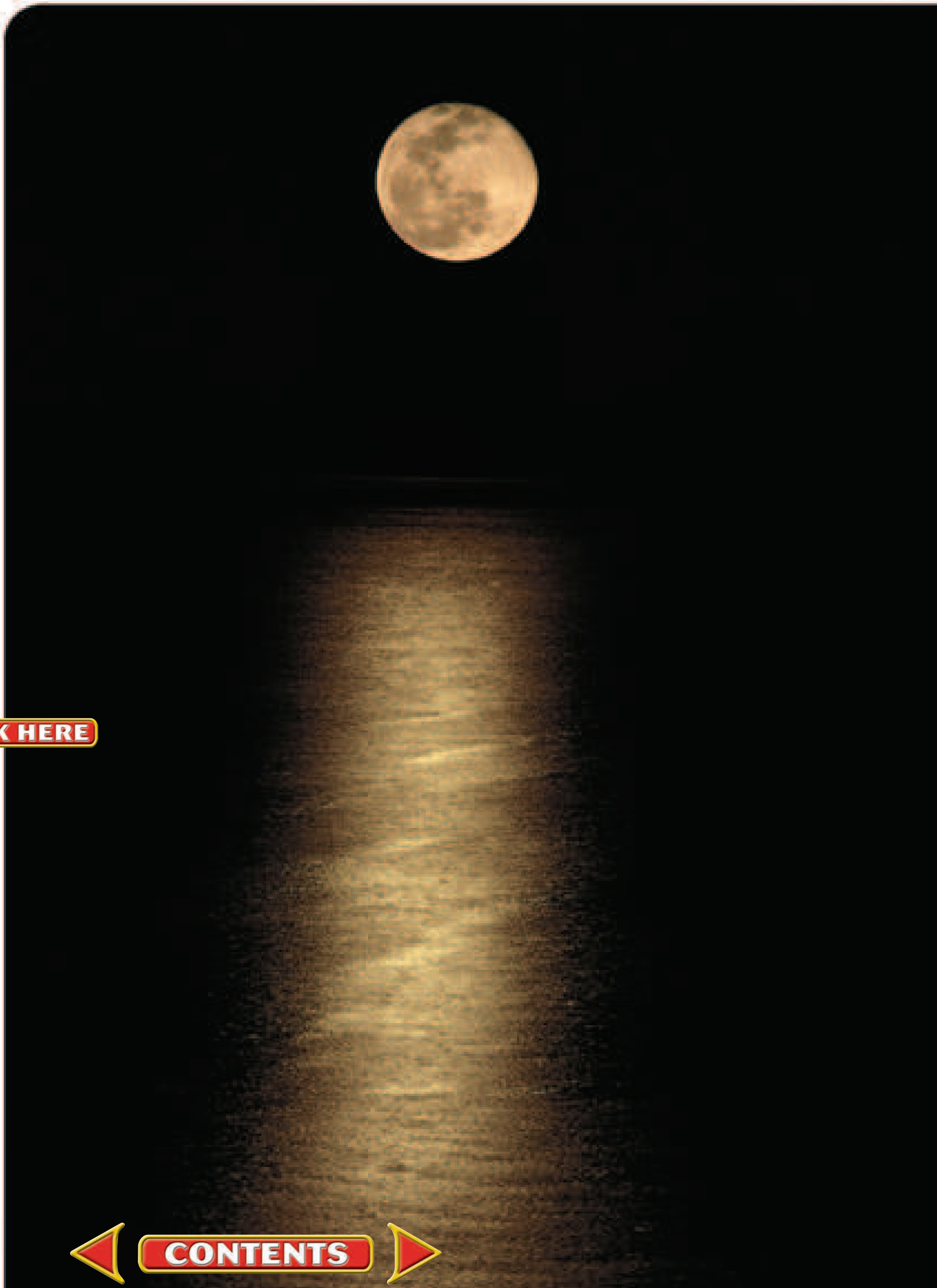
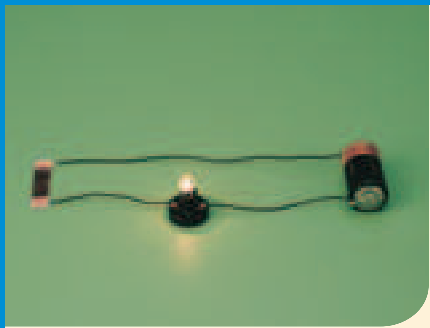
CHEMISTRY Online

Visit the Chemistry Web site at **science.glencoe.com** to find links about the periodic table.



The phases of the moon and the cycle of ocean tides are both periodic events, that is, they repeat in a regular manner.

DISCOVERY LAB



Materials

tape
samples of copper
light socket with bulb, wires, and battery

Versatile Metals

A variety of processes can be used to shape metals into different forms. Because of their physical properties, metals are used in a wide range of applications.

Safety Precautions

Be careful when bending the copper samples, as they may have sharp edges.

Procedure

1. Observe the different types of copper metal that your teacher gives you. Write down as many observations as you can about each of the copper samples.
2. Try gently bending each copper sample (do not break the samples). Record your observations.
3. Connect each copper sample to the circuit as shown in the photo. Record your observations.

Analysis

What properties of copper are similar in all of the samples? How do the samples of copper differ? List several common applications of copper. What properties make metals such as copper so versatile?

Section 6.1

Development of the Modern Periodic Table

Objectives

- **Trace** the development and **identify** key features of the periodic table.

Vocabulary

periodic law
group
period
representative element
transition element
metal
alkali metal
alkaline earth metal
transition metal
inner transition metal
nonmetal
halogen
noble gas
metalloid

You have already learned much in your study of chemistry. Wouldn't it be nice if you could easily organize the chemistry knowledge you are acquiring? You can, with the help of the *periodic* table. It is called a periodic table because, much like the phases of the moon, one of which is shown in the chapter opening photo, the properties of the elements in the table repeat in a periodic way. The periodic table will be an invaluable tool as you continue this course in chemistry. However, before you learn about the modern periodic table, a recounting of the history behind the table's development will help you understand its significance.

History of the Periodic Table's Development

In the late 1790s, French scientist Antoine Lavoisier compiled a list of elements known at the time. The list contained 23 elements. Many of these elements, such as silver, gold, carbon, and oxygen, were known since prehistoric times. The 1800s brought many changes to the world, including an explosion in the number of known elements. The advent of electricity, which was used to break compounds down into their component elements, and the development of the spectrometer, which was used to identify the newly isolated elements, played major roles in the advancement of chemistry. So did the industrial revolution

Figure 6-1

A resident of London, England invented the word smog to describe the city's filthy air, a combination of smoke and natural fog. The quality of London's air became so poor that in 1952 about 4000 Londoners died during a four-day period. This incident led to the passage of England's Clean Air Act in 1956.

CHEMISTRY CD-ROM

Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Astronomy CONNECTION

The element technetium does not occur naturally on Earth. It has been found in stars. Astronomers analyze the chemical composition of stellar matter by using an instrument called a spectroscope, which separates the light from a star into individual colors, much as a prism does. Although each star has a unique composition of elements, all stars are composed mainly of the gases hydrogen and helium. The Sun, for example, is estimated to be about 70 percent hydrogen and 28 percent helium. A tiny fraction of a star's mass may come from heavier elements such as oxygen, carbon, nitrogen, calcium, or sodium. Two percent of our Sun's mass comes from these heavier elements.

of the mid-1800s, which led to the development of many new chemistry-based industries, such as the manufacture of petrochemicals, soaps, dyes, and fertilizers. By 1870, there were approximately 70 known elements—almost triple the number known in Lavoisier's time. As you can see in **Figure 6-1,** the industrial revolution also created problems, such as increased chemical pollution.

Along with the discovery of new elements came volumes of new scientific data related to the elements and their compounds. Chemists of the time were overwhelmed with learning the properties of so many new elements and compounds. What chemists needed was a tool for organizing the many facts associated with the elements. A significant step toward this goal came in 1860, when chemists agreed upon a method for accurately determining the atomic masses of the elements. Until this time, different chemists used different mass values in their work, making the results of one chemist's work hard to reproduce by another. With newly agreed upon atomic masses for the elements, the search for relationships between atomic mass and elemental properties began in earnest.

John Newlands In 1864, English chemist John Newlands (1837–1898), who is shown in **Figure 6-2,** proposed an organization scheme for the elements. Newlands noticed that when the elements were arranged by increasing atomic mass, their properties repeated every eighth element. In other words, the first and eighth elements had similar properties, the second and ninth elements had similar properties, and so on. A pattern such as this is called periodic because it repeats in a specific manner. Newlands named the periodic relationship that he observed in chemical properties the law of octaves, because an octave is a group of musical notes that repeats every eighth tone. **Figure 6-2** also shows how Newlands organized the first 14 "known" elements (as of the mid-1860s). If you compare Newlands's arrangement of the elements with the modern periodic table on the inside back cover of your textbook, you'll see that some of his rows correspond to columns on the modern periodic table. Acceptance of the law of octaves was hampered because the law did not work for all of the known elements. Also, unfortunately for Newlands, the use of the word octave was harshly criticized by fellow scientists who thought that the musical analogy was unscientific. While Newlands's law was not generally accepted, the passage of a few years would show that he was basically correct; the properties of elements do repeat in a periodic way.

Meyer, Mendeleev, and Moseley In 1869, German chemist Lothar Meyer (1830–1895) and Russian chemist Dmitri Mendeleev (1834–1907) each demonstrated a connection between atomic mass and elemental properties. Mendeleev, however, is generally given more credit than Meyer because he published his organization scheme first and went on to better demonstrate its usefulness. Like Newlands several years earlier, Mendeleev noticed that when the elements were ordered by increasing atomic mass, there was a repetition, or periodic pattern, in their properties. By arranging the elements in order of increasing atomic mass into columns with similar properties, Mendeleev organized the elements into the first periodic table. Mendeleev and part of his periodic table are shown in **Figure 6-3.** Part of the reason Mendeleev's table was widely accepted was that he predicted the existence and properties of undiscovered elements. Mendeleev left blank spaces in the table where he thought the undiscovered elements should go. By noting trends in the properties of known elements, he was able to predict the properties of the yet-to-be discovered elements scandium, gallium, and germanium.

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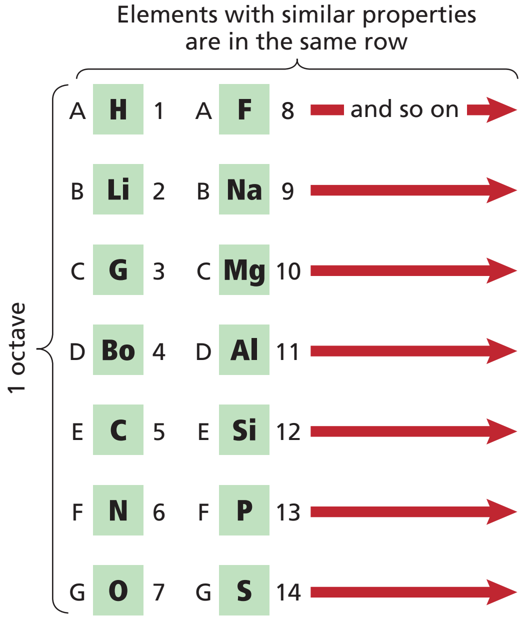

Figure 6-2

John Newlands noticed that the properties of elements repeated in a manner similar to an octave on a musical scale (A, B, C, D, E, F, G, A, and so on). While there are some similarities between the law of octaves and the modern periodic table, there also are significant differences. You'll notice that some of the chemical symbols do not match. For example, beryllium (Be) was also known as glucinum (G). What similarities and differences can you identify?

Mendeleev's table, however, was not completely correct. After several new elements were discovered and atomic masses of the known elements were more accurately determined, it became apparent that several elements in his table were not in the correct order. Arranging the elements by mass resulted in several elements being placed in groups of elements with differing properties. The reason for this problem was determined in 1913 by English chemist Henry Moseley. As you may recall from Chapter 4, Moseley discovered that atoms of each element contain a unique number of protons in their nuclei—the number of protons being equal to the atom's atomic number. By arranging the elements in order of increasing atomic number instead of increasing atomic mass, as Mendeleev had done, the problems with the order of the elements in the periodic table were solved. Moseley's arrangement of elements by atomic number resulted in a clear periodic pattern of properties. The statement that there is a periodic repetition of chemical and physical properties of the elements when they are arranged by increasing atomic number is called the **periodic law.**

Figure 6-3

Dmitri Mendeleev produced the first useful and widely accepted periodic table. The monument shown on the right is located in St. Petersburg, Russia, and shows an early version of Mendeleev's periodic table. The blank areas on the table show the positions of elements that had not yet been discovered.

Figure 6-4

The modern periodic table arranges the elements by increasing atomic number. The columns are known as groups or families, and the rows are known as periods.

The periodic table became a significant tool for chemists working in the new industries created during the industrial revolution. The table brought order to seemingly unrelated facts. You, too, will find the periodic table a valuable tool. Among other things, it is a useful reference for understanding and predicting the properties of elements and for organizing your knowledge of atomic structure. Do the **problem-solving LAB** on the next page to see how the periodic law can be used to predict unknown elemental properties.

The Modern Periodic Table

The modern periodic table is shown in **Figure 6-4** and on the inside back cover of your textbook. A larger, two-page version of the table appears in **Figure 6-7** on pages 156-157. The table consists of boxes, each containing an element name, symbol, atomic number, and atomic mass. A typical box from the table is shown in **Figure 6-5.** The boxes are arranged in order of increasing atomic number into a series of columns, called **groups** or families, and rows, called **periods.** Beginning with hydrogen in period 1, there are a total of seven periods. Each group is numbered 1 through 8, followed by the letter A or B. For example, scandium (Sc) is in the third column from the left, group 3B. What group is oxygen in? What period contains potassium and calcium? The groups designated with an A (1A through 8A) are often referred to as the main group, or **representative elements** because they possess a wide range of chemical and physical properties. The groups designated with a B (1B through 8B) are referred to as the **transition elements.** A more recent numbering system, which uses the numbers 1 through 18, also appears above each group. The number-and-letter system is used throughout this textbook.

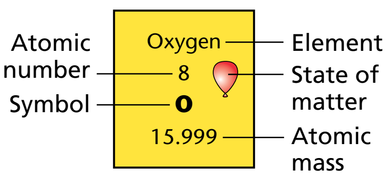

Figure 6-5

A typical box from the periodic table contains important information about an element.

Classifying the elements There are three main classifications for the elements—metals, nonmetals, and metalloids. **Metals** are elements that are generally shiny when smooth and clean, solid at room temperature, and good conductors of heat and electricity. Most metals also are ductile and malleable, meaning that they can be pounded into thin sheets and drawn into wires, respectively. **Figure 6-6** shows several applications that make use of the physical properties of metals.

Most group A elements and all group B elements are metals. If you look at boron (B) in column 3A, you see a heavy stair-step line that zigzags down to astatine (At) at the bottom of group 7A. This stair-step line serves as a visual divider between the metals and the nonmetals on the table. Metals are represented by the light blue boxes in **Figure 6-7.** Except for hydrogen, all of the elements on the left side of the table are metals. The group 1A elements (except for hydrogen) are known as the **alkali metals;** the group 2A elements are known as the **alkaline earth metals.** Both the alkali metals and the alkaline earth metals are chemically reactive, with the alkali metals being the more reactive of the two groups.

Figure 6-6

Metals are used in a wide variety of applications. The excellent electrical conductivity of metals such as copper, makes them a good choice for transmitting electrical power. Ductility and malleability allow metals to be formed into coins, tools, fasteners, and wires.

problem-solving LAB

Francium—solid, liquid or gas?

Predicting Of the first 101 elements, francium is the least stable. Its most stable isotope has a half-life of just 22 minutes! Use your knowledge about the properties of other alkali metals to predict some of francium's properties.

Analysis

In the spirit of Dimitri Mendeleev's prediction of the properties of several, as of then, undiscovered elements, use the given information about the known properties of the alkali metals to devise a method for determining the corresponding property of francium.

Alkali Metals Data

Element	Melting point (°C)	Boiling point (°C)	Radius (pm)
lithium	180.5	1347	152
sodium	97.8	897	186
potassium	63.3	766	227
rubidium	39.31	688	248
cesium	28.4	674.8	265
francium	?	?	?

Thinking Critically

1. Using the periodic law as a guide, devise an approach that clearly displays the trends for each of the properties given in the table and allows you to extrapolate a value for francium.
2. Predict whether francium is a solid, liquid, or gas. How can you support your prediction?
3. Which of the given columns of data presents the greatest possible error in making a prediction? Explain.
4. Currently, scientists can produce about one million francium atoms per second. Explain why this is still not enough to make basic measurements such as density or melting point.

Figure 6-7

PERIODIC TABLE OF THE ELEMENTS

Element — Hydrogen
Atomic number — 1
Symbol — H
Atomic mass — 1.008
State of matter

Gas
Liquid
Solid
Synthetic

Period	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	9
1	Hydrogen 1 H 1.008								
2	Lithium 3 Li 6.941	Beryllium 4 Be 9.012							
3	Sodium 11 Na 22.990	Magnesium 12 Mg 24.305							
4	Potassium 19 K 39.098	Calcium 20 Ca 40.078	Scandium 21 Sc 44.956	Titanium 22 Ti 47.867	Vanadium 23 V 50.942	Chromium 24 Cr 51.996	Manganese 25 Mn 54.938	Iron 26 Fe 55.845	Cobalt 27 Co 58.933
5	Rubidium 37 Rb 85.468	Strontium 38 Sr 87.62	Yttrium 39 Y 88.906	Zirconium 40 Zr 91.224	Niobium 41 Nb 92.906	Molybdenum 42 Mo 95.94	Technetium 43 Tc (98)	Ruthenium 44 Ru 101.07	Rhodium 45 Rh 102.906
6	Cesium 55 Cs 132.905	Barium 56 Ba 137.327	Lanthanum 57 La 138.906	Hafnium 72 Hf 178.49	Tantalum 73 Ta 180.948	Tungsten 74 W 183.84	Rhenium 75 Re 186.207	Osmium 76 Os 190.23	Iridium 77 Ir 192.217
7	Francium 87 Fr (223)	Radium 88 Ra (226)	Actinium 89 Ac (227)	Rutherfordium 104 Rf (261)	Dubnium 105 Db (262)	Seaborgium 106 Sg (266)	Bohrium 107 Bh (264)	Hassium 108 Hs (277)	Meitnerium 109 Mt (268)

The number in parentheses is the mass number of the longest lived isotope for that element.

Lanthanide series	Cerium 58 Ce 140.116	Praseodymium 59 Pr 140.908	Neodymium 60 Nd 144.24	Promethium 61 Pm (145)	Samarium 62 Sm 150.36	Europium 63 Eu 151.964
Actinide series	Thorium 90 Th 232.038	Protactinium 91 Pa 231.036	Uranium 92 U 238.029	Neptunium 93 Np (237)	Plutonium 94 Pu (244)	Americium 95 Am (243)

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CHEMISTRY Online

Visit the Chemistry Web site at **science.glencoe.com** to find updates on the periodic table.

- Metal
- Metalloid
- Nonmetal
- Recently discovered

10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
								Helium 2 **He** 4.003
			Boron 5 **B** 10.811	Carbon 6 **C** 12.011	Nitrogen 7 **N** 14.007	Oxygen 8 **O** 15.999	Fluorine 9 **F** 18.998	Neon 10 **Ne** 20.180
			Aluminum 13 **Al** 26.982	Silicon 14 **Si** 28.086	Phosphorus 15 **P** 30.974	Sulfur 16 **S** 32.065	Chlorine 17 **Cl** 35.453	Argon 18 **Ar** 39.948
Nickel 28 **Ni** 58.693	Copper 29 **Cu** 63.546	Zinc 30 **Zn** 65.39	Gallium 31 **Ga** 69.723	Germanium 32 **Ge** 72.64	Arsenic 33 **As** 74.922	Selenium 34 **Se** 78.96	Bromine 35 **Br** 79.904	Krypton 36 **Kr** 83.80
Palladium 46 **Pd** 106.42	Silver 47 **Ag** 107.868	Cadmium 48 **Cd** 112.411	Indium 49 **In** 114.818	Tin 50 **Sn** 118.710	Antimony 51 **Sb** 121.760	Tellurium 52 **Te** 127.60	Iodine 53 **I** 126.904	Xenon 54 **Xe** 131.293
Platinum 78 **Pt** 195.078	Gold 79 **Au** 196.967	Mercury 80 **Hg** 200.59	Thallium 81 **Tl** 204.383	Lead 82 **Pb** 207.2	Bismuth 83 **Bi** 208.980	Polonium 84 **Po** (209)	Astatine 85 **At** (210)	Radon 86 **Rn** (222)
Ununnilium * 110 **Uun** (281)	Unununium * 111 **Uuu** (272)	Ununbium * 112 **Uub** (285)		Ununquadium * 114 **Uuq** (289)		Ununhexium * 116 **Uuh** (289)		Ununoctium * 118 **Uuo** (293)

* Names not officially assigned. Discovery of elements 114, 116, and 118 recently reported. Further information not yet available.

Gadolinium 64 **Gd** 157.25	Terbium 65 **Tb** 158.925	Dysprosium 66 **Dy** 162.50	Holmium 67 **Ho** 164.930	Erbium 68 **Er** 167.259	Thulium 69 **Tm** 168.934	Ytterbium 70 **Yb** 173.04	Lutetium 71 **Lu** 174.967
Curium 96 **Cm** (247)	Berkelium 97 **Bk** (247)	Californium 98 **Cf** (251)	Einsteinium 99 **Es** (252)	Fermium 100 **Fm** (257)	Mendelevium 101 **Md** (258)	Nobelium 102 **No** (259)	Lawrencium 103 **Lr** (262)

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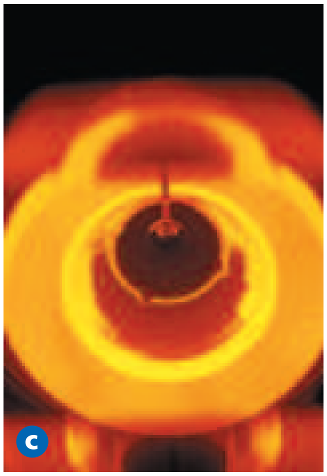


Figure 6-8

a A mountain climber breathes from a container of compressed oxygen gas, a nonmetal. b This Persian brass bowl contains inlays of the transition metals silver and gold. c Silicon crystals, a metalloid, are grown in an inert atmosphere of argon, a nonmetal. The crystals are used in the manufacture of computer chips.

The group B elements, or transition elements, are divided into **transition metals** and **inner transition metals.** The two sets of inner transition metals, known as the lanthanide and actinide series, are located along the bottom of the periodic table. The rest of the group B elements make up the transition metals. Elements from the lanthanide series are used extensively as phosphors, substances that emit light when struck by electrons. The **How It Works** at the end of the chapter explains more about phosphors and how images are formed on a television screen.

Nonmetals occupy the upper right side of the periodic table. They are represented by the yellow boxes in **Figure 6-7. Nonmetals** are elements that are generally gases or brittle, dull-looking solids. They are poor conductors of heat and electricity. The only nonmetal that is a liquid at room temperature is bromine (Br). The highly reactive group 7A elements are known as **halogens,** and the extremely unreactive group 8A elements are commonly called the **noble gases.**

Examine the elements in green boxes bordering the stair-step line in **Figure 6-7.** These elements are called metalloids, or semimetals. **Metalloids** are elements with physical and chemical properties of both metals and nonmetals. Silicon and germanium are two of the most important metalloids, as they are used extensively in computer chips and solar cells. Applications that make use of the properties of nonmetals, transition metals, and metalloids are shown in **Figure 6-8.** Do the **CHEMLAB** at the end of this chapter to observe trends among various elements.

This introduction to the periodic table only touches the surface of its usefulness. In the next section, you will discover how an element's electron configuration, which you learned about in Chapter 5, is related to its position on the periodic table.

Section 6.1 Assessment

1. Describe the development of the modern periodic table. Include contributions made by Newlands, Mendeleev, and Moseley.
2. Sketch a simplified version of the periodic table and indicate the location of groups, periods, metals, nonmetals, and metalloids.
3. Describe the general characteristics of metals, nonmetals, and metalloids.
4. Identify each of the following as a representative element or a transition element.
   - **a.** lithium (Li)
   - **b.** platinum (Pt)
   - **c.** promethium (Pm)
   - **d.** carbon (C)
5. **Thinking Critically** For each of the given elements, list two other elements with similar chemical properties.
   - **a.** iodine (I)
   - **b.** barium (Ba)
   - **c.** iron (Fe)
6. **Interpreting Data** An unknown element has chemical behavior similar to that of silicon (Si) and lead (Pb). The unknown element has a mass greater than that of sulfur (S), but less than that of cadmium (Cd). Use the periodic table to determine the identity of the unknown element.

Classification of the Elements

Objectives

- **Explain** why elements in the same group have similar properties.
- **Identify** the four blocks of the periodic table based on electron configuration.

In Chapter 5, you learned how to write the electron configuration for an atom. This is an important skill because the electron configuration determines the chemical properties of the element. However, the process of writing out electron configurations using the aufbau diagram can be tedious. Fortunately, by noting an atom's position on the periodic table, you can determine its electron configuration and its number of valence electrons.

Organizing the Elements by Electron Configuration

Take a look at the electron configurations for the group 1A elements listed below. These elements comprise the first four periods of group 1A.

Period 1	hydrogen	$1s^1$	$1s^1$
Period 2	lithium	$1s^22s^1$	$[He]2s^1$
Period 3	sodium	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Period 4	potassium	$1s^22s^22p^63s^23p^64s^1$	$[Ar]4s^1$

What do the four configurations have in common? The answer is that they all have a single electron in their outermost energy level.

Valence electrons Recall from Chapter 5 that electrons in the highest principal energy level of an atom are called valence electrons. Each of the group 1A elements has one electron in its highest energy level; thus, each element has one valence electron. This is no coincidence. The group 1A elements have similar chemical properties because they all have the same number of valence electrons. This is one of the most important relationships in chemistry; *atoms in the same group have similar chemical properties because they have the same number of valence electrons.* Each group 1A element has a valence electron configuration of s^1. Likewise, each group 2A element has a valence electron configuration of s^2. Each column of group A elements on the periodic table has its own unique valence electron configuration.

Valence electrons and period The energy level of an element's valence electrons indicates the period on the periodic table in which it is found. For example, lithium's valence electron is in the second energy level and lithium is found in period 2. Now look at gallium, with its electron configuration of $[Ar]4s^23d^{10}4p^1$. Gallium's valence electrons are in the fourth energy level, and gallium is found in the fourth period. What is the electron configuration for the group 1A element in the sixth period?

Valence electrons and group number A representative element's group number and the number of valence electrons it contains also are related. Group 1A elements have one valence electron, group 2A elements have two valence electrons, and so on. There are several exceptions to this rule, however. The noble gases in group 8A each have eight valence electrons, with the exception of helium, which has only two valence electrons. Also, the group number rule applies only to the representative elements (the group A elements). See **Figure 6-9** on the next page. The electron-dot structures you learned in Chapter 5 illustrate the connection between group number and number of valence electrons.

The s-, p-, d-, and f-Block Elements

The periodic table has columns and rows of varying sizes. The reason behind the table's odd shape becomes clear if it is divided into sections, or blocks, representing the atom's energy sublevel being filled with valence electrons. Because there are four different energy sublevels (s, p, d, and f), the periodic table is divided into four distinct blocks, as shown in **Figure 6-10.**

s-block elements The s-block consists of groups 1A and 2A, and the elements hydrogen and helium. In this block, the valence electrons, represented in **Figure 6-9,** occupy only s orbitals. Group 1A elements have partially filled s orbitals containing one valence electron and electron configurations ending in s^1. Group 2A elements have completely filled s orbitals containing two valence electrons and electron configurations ending in s^2. Because s orbitals hold a maximum of two electrons, the s-block portion of the periodic table spans two groups.

p-block elements After the s sublevel is filled, the valence electrons, represented in **Figure 6-9,** next occupy the p sublevel and its three p orbitals. The p-block of the periodic table, comprised of groups 3A through 8A, contains elements with filled or partially filled p orbitals. Why are there no p-block elements in period 1? The answer is that the p sublevel does not exist for the first principal energy level ($n = 1$). Thus, the first p-block element is boron (B), in the second period. The p-block spans six groups on the periodic table because the three p orbitals can hold a maximum of six electrons. Together, the s- and p-blocks comprise the representative, or group A, elements.

The group 8A, or noble gas, elements are unique members of the p-block because of their incredible stability. Noble gas atoms are so stable that they undergo virtually no chemical reactions. The reason for their stability lies in their electron configurations. Look at the electron configurations of the first four noble gas elements shown in **Table 6-1.** Notice that *both* the s and p orbitals corresponding to the period's principal energy level are *completely filled*. This arrangement of electrons results in an unusually stable atomic structure. You soon will learn that this stable configuration plays an important role in the formation of ions and chemical bonds.

d-block elements The d-block contains the transition metals and is the largest of the blocks. Although there are a number of exceptions, d-block elements are characterized by a filled outermost s orbital of energy level n, and filled or partially filled d orbitals of energy level $n - 1$. As you move across the period, electrons fill the d orbitals. For example, scandium (Sc), the first d-block element, has an electron configuration of $[Ar]4s^2 3d^1$. Titanium, the next element on the table, has an electron configuration of $[Ar]4s^2 3d^2$. Note that titanium's filled outermost s orbital has an energy level of $n = 4$, while the partially filled d orbital has an energy level of $n - 1$, or 3. The five d orbitals can hold a total of ten electrons; thus, the d-block spans ten groups on the periodic table.

Careers Using Chemistry

Medical Lab Technician

Would you like to analyze blood and tissue samples? How about determining the chemical content of body fluids? If so, you might enjoy being a medical lab technician.

Medical or clinical lab technicians work in large hospitals or independent labs. Under the direction of a technologist, they prepare specimens, conduct tests, and operate computerized analyzers. Technicians need to pay close attention to detail, have good judgement, and be skilled in using computers.

Figure 6-9

The electron-dot structures of most of the representative elements are shown here. The number of valence electrons is the same for all members of a group. For the group A elements, an atom's number of valence electrons is equal to its group number (in the 1A, 2A, . . . numbering system).

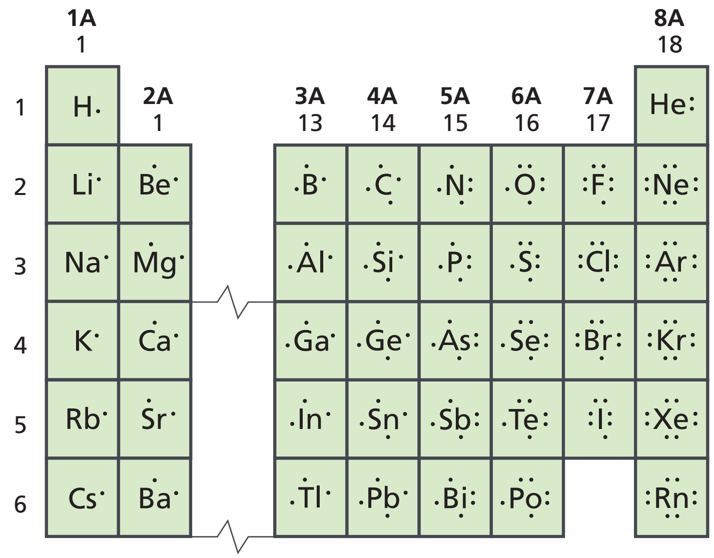


Table 6-1

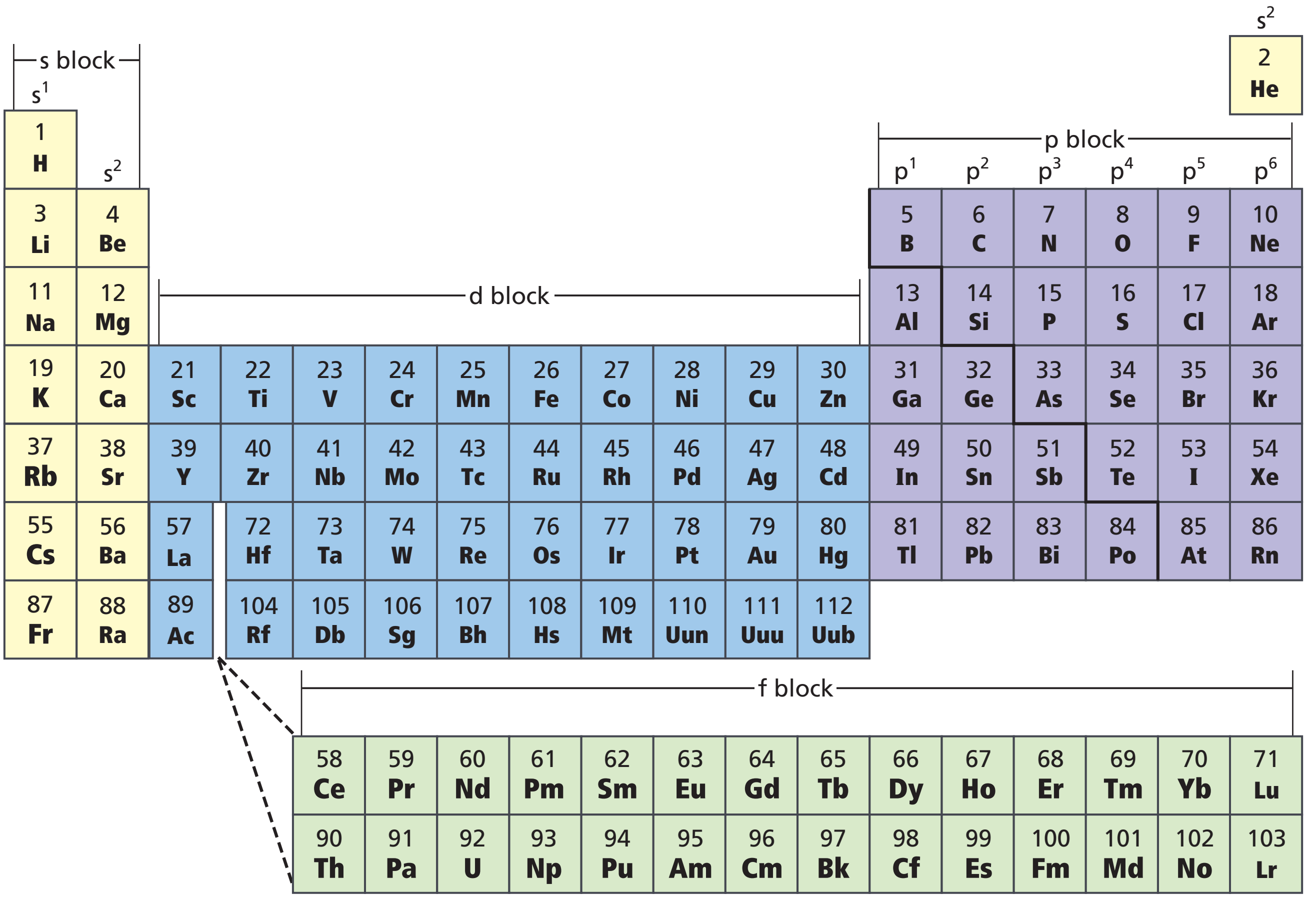
Electron Configurations of Helium, Neon, Argon, and Krypton				
Period	**Principal energy level**	**Element**	**Electron configuration**	**Electron dot structure**
1	$n = 1$	helium	$1s^2$	He:
2	$n = 2$	neon	$[He]2s^22p^6$	:Ne:
3	$n = 3$	argon	$[Ne]3s^23p^6$	:Ar:
4	$n = 4$	krypton	$[Ar]4s^23d^{10}4p^6$	:Kr:

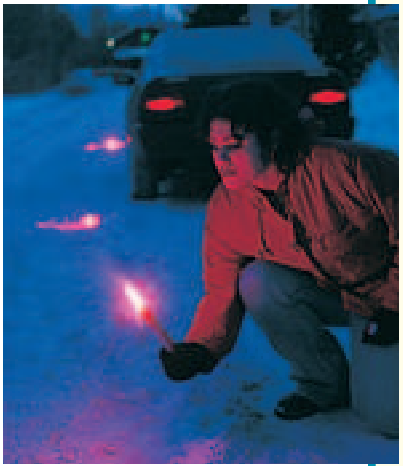
f-block elements The f-block contains the inner transition metals. The f-block elements are characterized by a filled, or partially filled outermost s orbital, and filled or partially filled 4f and 5f orbitals. The electrons of the f sublevel do not fill their orbitals in a predictable manner. Because there are seven f orbitals holding up to a maximum of 14 electrons, the f-block spans 14 columns of the periodic table.

Thus, the s-, p-, d-, and f-blocks determine the shape of the periodic table. As you proceed down through the periods, the principal energy level increases, as does the number of energy sublevels containing electrons. Period 1 contains only s-block elements, periods 2 and 3 contain both s- and p-block elements, periods 4 and 5 contain s-, p-, and d-block elements, and periods 6 and 7 contain s-, p-, d-, and f-block elements.

Figure 6-10

Although electrons fill the orbitals of s- and p-block elements in a predictable manner, there are a number of exceptions in the d- and f-block elements. What is the relationship between the maximum number of electrons an energy sublevel can hold and the size of that block on the diagram?





Strontium-containing compounds are used to produce the bright red seen in these road flares.

EXAMPLE PROBLEM 6-1

Electron Configuration and the Periodic Table

Strontium has an electron configuration of [Kr]$5s^2$. Without using the periodic table, determine the group, period, and block in which strontium is located on the periodic table.

1. Analyze the Problem

You are given the electron configuration of strontium. The energy level of the valence electrons can be used to determine the period in which strontium is located. The electron configuration of the valence electrons can be used to determine the group and the block in which strontium is located.

2. Solve for the Unknown

Group The valence electron configuration of s^2 indicates that strontium is in group 2A. All group 2A elements have the s^2 configuration.

Period The 5 in $5s^2$ indicates that strontium is in period 5.

Block The s^2 indicates that strontium's valence electrons fill the s sublevel. Thus, strontium is in the s-block.

3. Evaluate the Answer

The relationships among electron configuration and position on the periodic table have been correctly applied. The given information identifies a unique position on the table, as it must.

Practice!

For more practice with electron configuration problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

7. Without using the periodic table, determine the group, period, and block of an atom with the following electron configurations.
   a. [Ne]$3s^2$ **b.** [He]$2s^2$ **c.** [Kr]$5s^24d^{10}5p^5$
8. Write the electron configuration of the element fitting each of the following descriptions.
   a. The group 2A element in the fourth period
   b. The noble gas in the fifth period
   c. The group 2B element in the fourth period
   d. The group 6A element in the second period
9. What are the symbols for the elements with the following valence electron configurations?
   a. s^2d^1 **b.** s^2p^3 **c.** s^2p^6

Section 6.2 Assessment

10. Explain why elements in the same group on the periodic table have similar chemical properties.

11. Given each of the following valence electron configurations, determine which block of the periodic table the element is in.
a. s^2p^4 **b.** s^1 **c.** s^2d^1 **d.** s^2p^1

12. Describe how each of the following are related.
a. Group number and number of valence electrons for representative elements
b. Principal energy level of valence electrons and period number

13. Without using the periodic table, determine the group, period, and block of an atom with an electron configuration of [Ne]$3s^23p^4$.

14. Thinking Critically A gaseous element is a poor conductor of heat and electricity, and is extremely nonreactive. Is the element likely to be a metal, nonmetal, or metalloid? Where would the element be located on the periodic table? Explain.

15. Formulating Models Make a simplified sketch of the periodic table and label the s-, p-, d-, and f-blocks.

Section 6.3 Periodic Trends

Many properties of the elements tend to change in a predictable way, known as a trend, as you move across a period or down a group. You will explore several periodic trends in this section. Do the **miniLAB** on the next page to explore several properties that behave periodically.

Objectives

- **Compare** period and group trends of several properties.
- **Relate** period and group trends in atomic radii to electron configuration.

Vocabulary

ion
ionization energy
octet rule
electronegativity

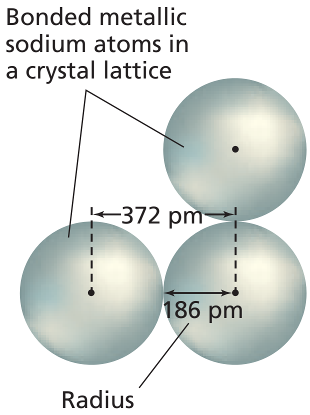
Atomic Radius

The electron cloud surrounding a nucleus is based on probability and does not have a clearly defined edge. It is true that the outer limit of an electron cloud is defined as the spherical surface within which there is a 90% probability of finding an electron. However, this surface does not exist in a physical way, as the outer surface of a golf ball does. Atomic size is defined by how closely an atom lies to a neighboring atom. Because the nature of the neighboring atom can vary from one substance to another, the size of the atom itself also tends to vary somewhat from substance to substance.

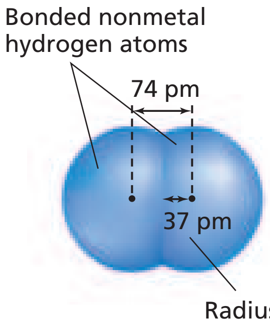
For metals such as sodium, the atomic radius is defined as half the distance between adjacent nuclei in a crystal of the element. See **Figure 6-11a.** For elements that commonly occur as molecules, such as many nonmetals, the

Figure 6-11

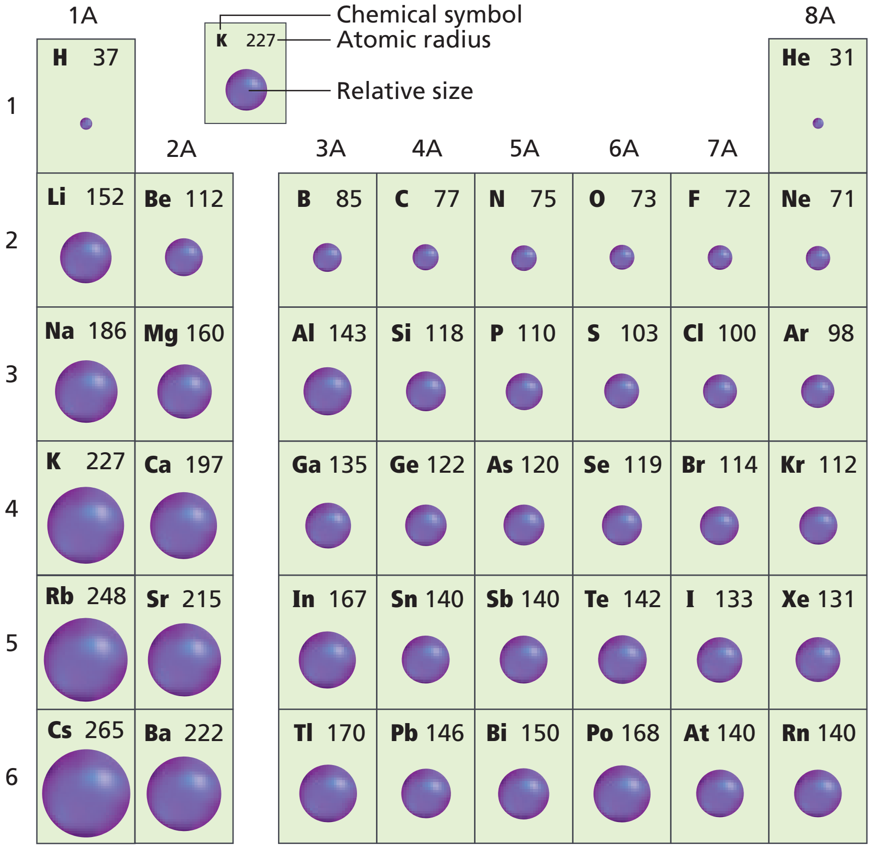
The table gives atomic radii of the representative elements.



a The radius of a metal atom in a metallic crystal is one-half the distance between two adjacent atoms in the crystal.



b The radius of a nonmetal atom is often determined from a diatomic molecule of an element.



Period	1A	2A	3A	4A	5A	6A	7A	8A
1	H 37							He 31
2	Li 152	Be 112	B 85	C 77	N 75	O 73	F 72	Ne 71
3	Na 186	Mg 160	Al 143	Si 118	P 110	S 103	Cl 100	Ar 98
4	K 227	Ca 197	Ga 135	Ge 122	As 120	Se 119	Br 114	Kr 112
5	Rb 248	Sr 215	In 167	Sn 140	Sb 140	Te 142	I 133	Xe 131
6	Cs 265	Ba 222	Tl 170	Pb 146	Bi 150	Po 168	At 140	Rn 140

c The atomic radii of the representative elements are given in picometers (1×10^{-12} meters) and their relative sizes are shown. The radii for the transition metals have been omitted because they exhibit many exceptions to the general trends shown here. What causes the increase in radii as you move down a group?

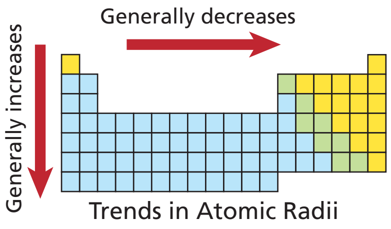


Figure 6-12

This small table provides a summary of the general trends in atomic radii.

atomic radius is defined as half the distance between nuclei of identical atoms that are chemically bonded together. The atomic radius of a nonmetal diatomic hydrogen molecule (H_2) is shown in **Figure 6-11b.**

Trends within periods A pattern in atomic size emerges as you look across a period in **Figure 6-11c.** In general, there is a decrease in atomic radii as you move left-to-right across a period. This trend is caused by the increasing positive charge in the nucleus and the fact that the principal energy level within a period remains the same. Each successive element has one additional proton and electron, and each additional electron is added to the same principal energy level. Moving across a period, no additional electrons come between the valence electrons and the nucleus. Thus, the valence electrons are not shielded from the increased nuclear charge. The result is that the increased nuclear charge pulls the outermost electrons closer to the nucleus.

Trends within groups Atomic radii generally increase as you move down a group. The nuclear charge increases and electrons are added to successively higher principal energy levels. Although you might think the increased nuclear charge would pull the outer electrons toward the nucleus and make the atom smaller, this effect is overpowered by several other factors. Moving down a group, the outermost orbital increases in size along with the increasing principal energy level; thus, making the atom larger. The larger orbital means that the outer electrons are farther from the nucleus. This increased distance offsets the greater pull of the increased nuclear charge. Also, as additional orbitals between the nucleus and the outer electrons are occupied, these electrons shield the outer electrons from the pull of the nucleus. **Figure 6-12** summarizes the group and period trends in atomic radii.

miniLAB

Periodicity of Molar Heats of Fusion and Vaporization

Making and Using Graphs The heats required to melt or to vaporize a mole (a specific amount of matter) of matter are known as the molar heat of fusion (H_f) and the molar heat of vaporization (H_v), respectively. These heats are unique properties of each element. You will investigate if the molar heats of fusion and vaporization for the period 2 and 3 elements behave in a periodic fashion.

Materials either a graphing calculator, a computer graphing program, or graph paper; Appendix **Table C-6** or access to comparable element data references

Procedure

Use **Table C-6** in Appendix C to look up and record the molar heat of fusion and the molar heat of vaporization for the period 3 elements listed in the table. Then, record the same data for the period 2 elements.

Molar Heat Data

Element	Atomic number	H_f (kJ/mol)	H_v (kJ/mol)
Na	11		
Mg	12		
Al	13		
Si	14		
P	15		
S	16		
Cl	17		
Ar	18		

Analysis

1. Graph molar heats of fusion versus atomic number. Connect the points with straight lines and label the curve. Do the same for molar heats of vaporization.
2. Do the graphs repeat in a periodic fashion? Describe the graphs to support your answer.

EXAMPLE PROBLEM 6-2

Interpreting Trends in Atomic Radii

Which has the largest atomic radius: carbon (C), fluorine (F), beryllium (Be), or lithium (Li)? Do not use **Figure 6-11** to answer the question. Explain your answer in terms of trends in atomic radii.

1. Analyze the Problem

You are given four elements. First, determine the groups and periods the elements occupy. Then apply the general trends in atomic radii to determine which has the largest atomic radius.

2. Solve for the Unknown

From the periodic table, all the elements are found to be in period 2. Ordering the elements from left-to-right across the period yields:
Li, Be, C, F

Applying the trend of decreasing radii across a period means that lithium, the first element in period 2, has the largest radius.

3. Evaluating the Answer

The group trend in atomic radii has been correctly applied. Checking radii values from **Figure 6-11** verifies the answer.

PRACTICE PROBLEMS

Answer the following questions using your knowledge of group and period trends in atomic radii. Do not use the atomic radii values in **Figure 6-11** to answer the questions.

16. Which has the largest radius: magnesium (Mg), silicon (Si), sulfur (S), or sodium (Na)? The smallest?

17. Which has the largest radius: helium (He), xenon (Xe), or argon (Ar)? The smallest?

18. Can you determine which of two unknown elements has the larger radius if the only known information is that the atomic number of one of the elements is 20 greater than the other?

Practice!

For more practice with periodic trend problems, go to **Supplemental Practice Problems** in Appendix A.

Ionic Radius

Atoms can gain or lose one or more electrons to form ions. Because electrons are negatively charged, atoms that gain or lose electrons acquire a net charge. Thus, an **ion** is an atom or a bonded group of atoms that has a positive or negative charge. You'll learn about ions in detail in Chapter 8, but for now, let's look at how the formation of an ion affects the size of an atom.

When atoms lose electrons and form positively charged ions, they always become smaller. For example, as shown in **Figure 6-13a** on the next page a sodium atom with a radius of 186 pm shrinks to a radius of 95 pm when it forms a positive sodium ion. The reason for the decrease in size is twofold. The electron lost from the atom will always be a valence electron. The loss of a valence electron may leave a completely empty outer orbital, which results in a smaller radius. Furthermore, the electrostatic repulsion between the now fewer number of remaining electrons decreases, allowing them to be pulled closer to the nucleus.

When atoms gain electrons and form negatively charged ions, they always become larger, as shown in **Figure 6-13b.** The addition of an electron to an

CONTENTS

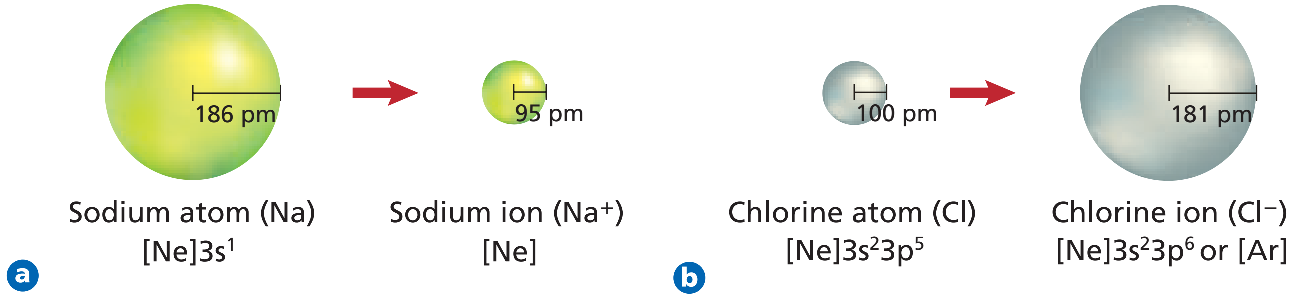


Figure 6-13

Atoms undergo significant changes in size when forming ions. **a** The sodium atom loses an electron and becomes smaller. **b** The chlorine ion gains an electron and becomes larger. How is each ion's electron configuration related to those of the noble gas elements?

atom increases the electrostatic repulsion between the atom's outer electrons, forcing them to move farther apart. The increased distance between the outer electrons results in a larger radius.

Trends within periods The ionic radii of most of the representative elements are shown in **Figure 6-14.** Note that elements on the left side of the table form smaller positive ions, and elements on the right side of the table form larger negative ions. In general, as you move left-to-right across a period, the size of the positive ions gradually decreases. Then, beginning in group 5A or 6A, the size of the much larger negative ions also gradually decreases.

Trends within groups As you move down a group, an ion's outer electrons are in higher principal energy levels, resulting in a gradual increase in ionic size. Thus, the ionic radii of both positive and negative ions increase as you move down a group. **Figure 6-15** on the next page summarizes the group and period trends in ionic radii.

Chemical symbol — **K** 138 — Ionic radius
Charge — 1+ — Relative size

Period	1A	2A	3A	4A	5A	6A	7A
2	**Li** 76, 1+	**Be** 31, 2+	**B** 20, 3+	**C** 15, 4+	**N** 146, 3–	**O** 140, 2–	**F** 133, 1–
3	**Na** 102, 1+	**Mg** 72, 2+	**Al** 54, 3+	**Si** 41, 4+	**P** 212, 3–	**S** 184, 2–	**Cl** 181, 1–
4	**K** 138, 1+	**Ca** 100, 2+	**Ga** 62, 3+	**Ge** 53, 4+	**As** 222, 3–	**Se** 198, 2–	**Br** 195, 1–
5	**Rb** 152, 1+	**Sr** 118, 2+	**In** 81, 3+	**Sn** 71, 4+	**Sb** 62, 5+	**Te** 221, 2–	**I** 220, 1–
6	**Cs** 167, 1+	**Ba** 135, 2+	**Tl** 95, 3+	**Pb** 84, 4+	**Bi** 74, 5+		

Figure 6-14

The table shows the ionic radii of most of the representative elements. The ion sizes are shown relative to one another, while the actual radii are given in picometers (1×10^{-12} meters). Note that the elements on the left side of the table form positive ions, and those on the right form negative ions.

Ionization Energy

To form a positive ion, an electron must be removed from a neutral atom. This requires energy. The energy is needed to overcome the attraction between the positive charge in the nucleus and the negative charge of the electron. **Ionization energy** is defined as the energy required to remove an electron from a gaseous atom. For example, 8.64×10^{-19} J is required to remove an electron from a gaseous lithium atom. The energy required to remove the first electron from an atom is called the first ionization energy. Therefore, the first ionization energy of lithium equals 8.64×10^{-19} J. The loss of the electron results in the formation of a Li^+ ion. The first ionization energies of the elements in periods 1 through 5 are plotted on the graph in **Figure 6-16.**

Think of ionization energy as an indication of how strongly an atom's nucleus holds onto its valence electrons. A high ionization energy value indicates the atom has a strong hold on its electrons. Atoms with large ionization energy values are less likely to form positive ions. Likewise, a low ionization energy value indicates an atom loses its outer electron easily. Such atoms are likely to form positive ions.

Take a close look at the graph in **Figure 6-16.** Each set of connected points represents the elements in a period. From the graph, it is clear that the group 1A metals have low ionization energies. Thus, group 1A metals (Li, Na K, Rb) are likely to form positive ions. It also is clear that the group 8A elements (He, Ne, Ar, Kr, Xe) have high ionization energies and are unlikely to form ions. Gases of group 8A are extremely unreactive—their stable electron configuration greatly limits their reactivity.

After removing the first electron from an atom, it is possible to remove additional electrons. The amount of energy required to remove a second electron from a 1+ ion is called the second ionization energy, the amount of energy required to remove a third electron from a 2+ ion is called the third ionization energy, and so on. **Table 6-2** on the next page lists the first through ninth ionization energies for elements in period 2.

Reading across **Table 6-2** from left-to-right, you see that the energy required for each successive ionization always increases. However, the increase in energy does not occur smoothly. Note that for each element there is an ionization for which the required energy jumps dramatically. For example, the second ionization energy of lithium (7300 kJ/mol) is much greater than its first ionization energy (520 kJ/mol). This means a lithium atom is relatively likely to lose its first valence electron, but extremely unlikely to lose its second.

If you examine the table, you'll see that the ionization at which the large jump in energy occurs is related to the atom's number of valence electrons. Lithium has one valence electron and the jump occurs after the first ionization energy. Lithium easily forms the common lithium 1+ ion, but is unlikely to form a lithium 2+ ion. The jump in ionization energy shows that atoms hold

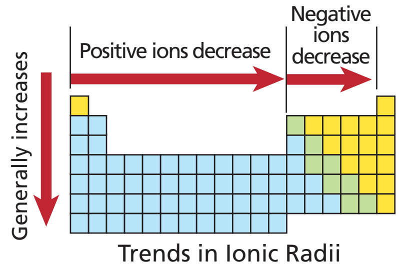


Figure 6-15

This small table provides a summary of the general trends in ionic radii.

Figure 6-16

The graph shows the first ionization energies for elements in periods 1 through 5. Note the high energies required to remove an electron from a noble gas element. What trend in first ionization energies do you observe as you move down a group?

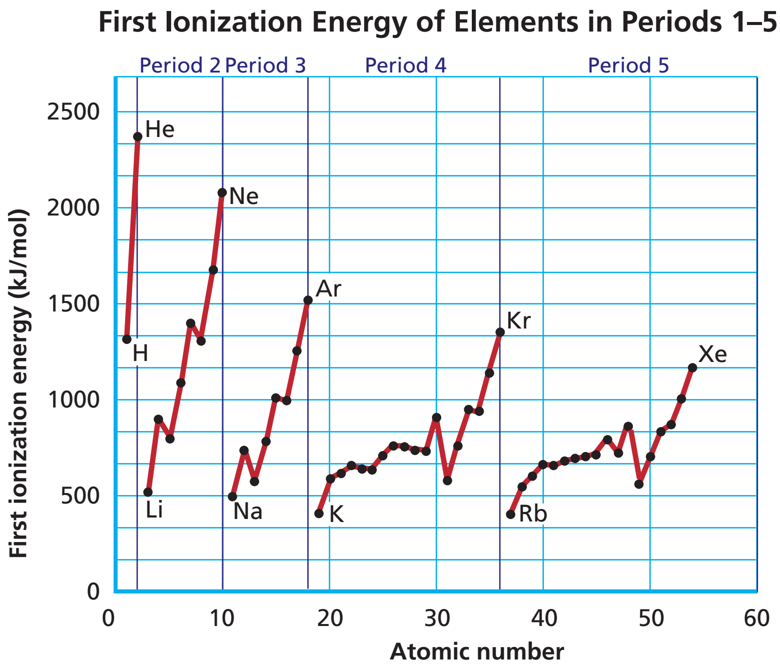


Table 6-2

Successive Ionization Energies for the Period 2 Elements										
Element	Valence electrons	Ionization energy (kJ/mol)* 1st	2nd	3rd	4th	5th	6th	7th	8th	9th
Li	1	520	7300							
Be	2	900	1760	14 850						
B	3	800	2430	3660	25 020					
C	4	1090	2350	4620	6220	37 830				
N	5	1400	2860	4580	7480	9440	53 270			
O	6	1310	3390	5300	7470	10 980	13 330	71 330		
F	7	1680	3370	6050	8410	11 020	15 160	17 870	92 040	
Ne	8	2080	3950	6120	9370	12 180	15 240	20 000	23 070	115 380

* mol is an abbreviation for mole, a quantity of matter.

onto their inner core electrons much more strongly than they hold onto their valence electrons. Where does the jump in ionization energy occur for oxygen, an atom with six valence electrons?

Trends within periods As shown in **Figure 6-16** and by the values in **Table 6-2,** first ionization energies generally increase as you move left-to-right across a period. The increased nuclear charge of each successive element produces an increased hold on the valence electrons.

Trends within groups First ionization energies generally decrease as you move down a group. This decrease in energy occurs because atomic size increases as you move down the group. With the valence electrons farther from the nucleus, less energy is required to remove them. **Figure 6-17** summarizes the group and period trends in first ionization energies.

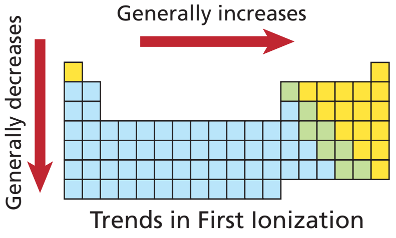


Figure 6-17

This small table provides a summary of the general trends in first ionization energies.

Octet rule When a sodium atom loses its single valence electron to form a 1+ sodium ion, its electron configuration changes as shown below.

Sodium atom $1s^22s^22p^63s^1$ Sodium ion $1s^22s^22p^6$

Note that the sodium ion has the same electron configuration as neon ($1s^22s^22p^6$), a noble gas. This observation leads to one of the most important principles in chemistry, the octet rule. The **octet rule** states that atoms tend to gain, lose, or share electrons in order to acquire a full set of eight valence electrons. This reinforces what you learned earlier that the electron configuration of filled s and p orbitals of the same energy level (consisting of eight valence electrons) is unusually stable. Note that the first period elements are an exception to the rule, as they are complete with only two valence electrons.

The octet rule is useful for determining the type of ions likely to form. Elements on the right side of the periodic table tend to gain electrons in order to acquire the noble gas configuration; therefore, these elements tend to form negative ions. In a similar manner, elements on the left side of the table tend to lose electrons and form positive ions.

Electronegativity

The **electronegativity** of an element indicates the relative ability of its atoms to attract electrons in a chemical bond. **Figure 6-18** lists the electronegativity values for most of the elements. These values are calculated based upon a number of factors, and are expressed in terms of a numerical value of 3.98 or less. The units of electronegativity are arbitrary units called Paulings, named after American scientist Linus Pauling (1901–1994).

Note that because the noble gases form very few compounds, they have been left out of **Figure 6-18.** Fluorine is the most electronegative element, with a value of 3.98, and cesium and francium are the least electronegative elements, with values of 0.79 and 0.7, respectively. In a chemical bond, the atom with the greater electronegativity more strongly attracts the bond's electrons. You will use electronegativity values in upcoming chapters to help determine the types of bonds that exist between elements in a compound.

Trends within periods and groups Electronegativity generally decreases as you move down a group, and increases as you move left-to-right across a period; therefore, the lowest electronegativities are found at the lower left side of the periodic table, while the highest electronegativities are found at the upper right.

Figure 6-18

The table shows the electronegativity values for most of the elements. In which areas of the periodic table do the highest electronegativities tend to occur? The lowest?

Increasing electronegativity

Decreasing electronegativity

- electronegativity < 1.0
- 1.0 ≥ electronegativity < 2.0
- 2.0 ≥ electronegativity < 3.0
- 3.0 ≥ electronegativity < 4.0

1 **H** 2.20																	2 **He**
3 **Li** 0.98	4 **Be** 1.57											5 **B** 2.04	6 **C** 2.55	7 **N** 3.04	8 **O** 3.44	9 **F** 3.98	10 **Ne**
11 **Na** 0.93	12 **Mg** 1.31											13 **Al** 1.61	14 **Si** 1.90	15 **P** 2.19	16 **S** 2.58	17 **Cl** 3.16	18 **Ar**
19 **K** 0.82	20 **Ca** 1.00	21 **Sc** 1.36	22 **Ti** 1.54	23 **V** 1.63	24 **Cr** 1.66	25 **Mn** 1.55	26 **Fe** 1.83	27 **Co** 1.88	28 **Ni** 1.91	29 **Cu** 1.90	30 **Zn** 1.65	31 **Ga** 1.81	32 **Ge** 2.01	33 **As** 2.18	34 **Se** 2.55	35 **Br** 2.96	36 **Kr**
37 **Rb** 0.82	38 **Sr** 0.95	39 **Y** 1.22	40 **Zr** 1.33	41 **Nb** 1.6	42 **Mo** 2.16	43 **Tc** 2.10	44 **Ru** 2.2	45 **Rh** 2.28	46 **Pd** 2.20	47 **Ag** 1.93	48 **Cd** 1.69	49 **In** 1.78	50 **Sn** 1.96	51 **Sb** 2.05	52 **Te** 2.1	53 **I** 2.66	54 **Xe**
55 **Cs** 0.79	56 **Ba** 0.89	57 **La** 1.1	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 1.7	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.2	78 **Pt** 2.2	79 **Au** 2.4	80 **Hg** 1.9	81 **Tl** 1.8	82 **Pb** 1.8	83 **Bi** 1.9	84 **Po** 2.0	85 **At** 2.2	86 **Rn**
87 **Fr** 0.70	88 **Ra** 0.90	89 **Ac** 1.1	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Uun**	111 **Uuu**	112 **Uub**						

Electronegativity Values in Paulings

Section 6.3 Assessment

19. Sketch a simplified periodic table and use arrows and labels to compare period and group trends in atomic and ionic radii, ionization energies, and electronegativities.

20. Explain how the period and group trends in atomic radii are related to electron configuration.

21. Which has the largest atomic radius: nitrogen (N), antimony (Sb), or arsenic (As)? The smallest?

22. For each of the following properties, indicate whether fluorine or bromine has a larger value.
a. electronegativity
b. ionic radius
c. atomic radius
d. ionization energy

23. Thinking Critically Explain why it takes more energy to remove the second electron from a lithium atom than it does to remove the fourth electron from a carbon atom.

24. Making and Using Graphs Graph the atomic radii of the group A elements in periods 2, 3, and 4 versus their atomic numbers. Connect the points of elements in each period, so that there are three separate curves on the graph. Summarize the trends in atomic radii shown on your graph. Explain.

CHEMLAB 6

Descriptive Chemistry of the Elements

What do elements look like? How do they behave? Can periodic trends in the properties of elements be observed? You cannot examine all of the elements on the periodic table because of limited availability, cost, and safety concerns. However, you can observe several of the representative elements, classify them, and compare their properties. The observation of the properties of elements is called descriptive chemistry.

Problem

What is the pattern of properties of the representative elements?

Objectives

- **Observe** properties of various elements.
- **Classify** elements as metals, nonmetals, and metalloids.
- **Examine** general trends within the periodic table.

Materials

stoppered test tubes containing small samples of elements
plastic dishes containing samples of elements
conductivity apparatus
1.0*M* HCl
test tubes (6)
test tube rack
10-mL graduated cylinder
spatula
small hammer
glass marking pencil

Safety Precautions

- **Wear safety goggles and a lab apron at all times.**
- **Do not handle elements with bare hands.**
- **1.0*M* HCl is harmful to eyes and clothing.**
- **Never test chemicals by tasting.**
- **Follow any additional safety precautions provided by your teacher.**

Pre-Lab

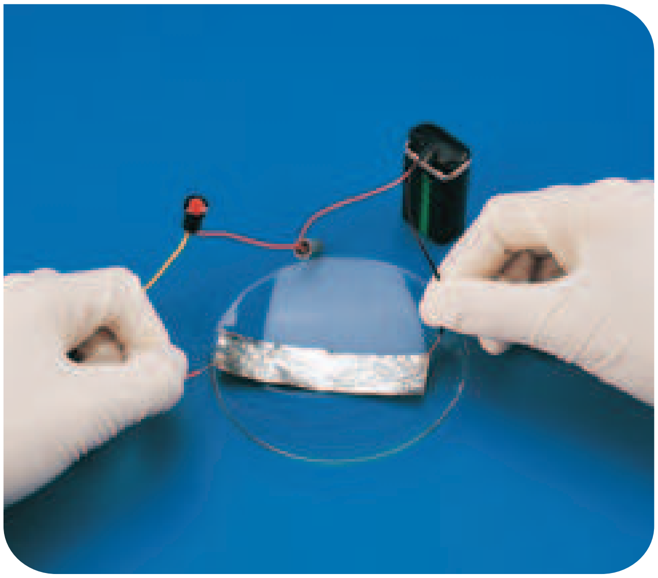
1. Read the entire **CHEMLAB.**
2. Prepare a data table similar to the one below to record the observations you make during the lab.
3. Examine the periodic table. What is the physical state of most metals? Nonmetals? Metalloids?
4. Look up the definitions of the terms luster, malleability, and electrical conductivity. To what elements do they apply?

Observation of Elements

Element	Appearance and physical state	Malleable or brittle?	Reactivity with HCl	Electrical conductivity	Classification

Procedure

1. Observe and record the appearance of the element sample in each test tube. Observations should include physical state, color, and other characteristics such as luster and texture. **CAUTION:** *Do not remove the stoppers from the test tubes.*
2. Remove a small sample of each of the elements contained in a dish and place it on a hard surface designated by your teacher. Gently tap each element sample with a small hammer. **CAUTION:** *Safety goggles must be worn.* If the element is malleable it will flatten. If it is brittle, it will shatter. Record your observations.
3. Use the conductivity tester to determine which elements conduct electricity. An illuminated light bulb is evidence of electrical conductivity. Record your results in your data table. Clean the electrodes with water and make sure they are dry before testing each element.



4. Label each test tube with the symbol for one of the elements in the plastic dishes. Using a graduated cylinder, add 5 mL of water to each test tube.
5. Use a spatula to put a small amount of each of the six elements (approximately 0.2 g or a 1-cm long ribbon) into the test tube labeled with its chemical symbol. Using a graduated cylinder, add 5 mL of 1.0*M* HCl to each test tube. Observe each test tube for at least one minute. The formation of bubbles is evidence of a reaction between the acid and the element. Record your observations.

Cleanup and Disposal

Dispose of all materials as instructed by your teacher.

Analyze and Conclude

1. **Interpreting Data** Metals are usually malleable and good conductors of electricity. They are generally lustrous and silver or white in color. Many react with acids. Write the word "metal" beneath the Classification heading in the data table for those element samples that display the general characteristics of metals.
2. **Interpreting Data** Nonmetals can be solids, liquids, or gases. They do not conduct electricity and do not react with acids. If a nonmetal is a solid, it is likely to be brittle and have color (other than white or silver). Write the word "nonmetal" beneath the Classification heading in the data table for those element samples that display the general characteristics of nonmetals.
3. **Interpreting Data** Metalloids combine some of the properties of both metals and nonmetals. Write the word "metalloid" beneath the Classification heading in the data table for those element samples that display the general characteristics of metalloids.
4. **Making a Model** Construct a periodic table and label the representative elements by group (1A through 7A). Using the information in your data table and the periodic table, record the identities of elements observed during the lab in your periodic table.
5. **Interpreting** Describe any trends among the elements you observed in the lab.

Real-World Chemistry

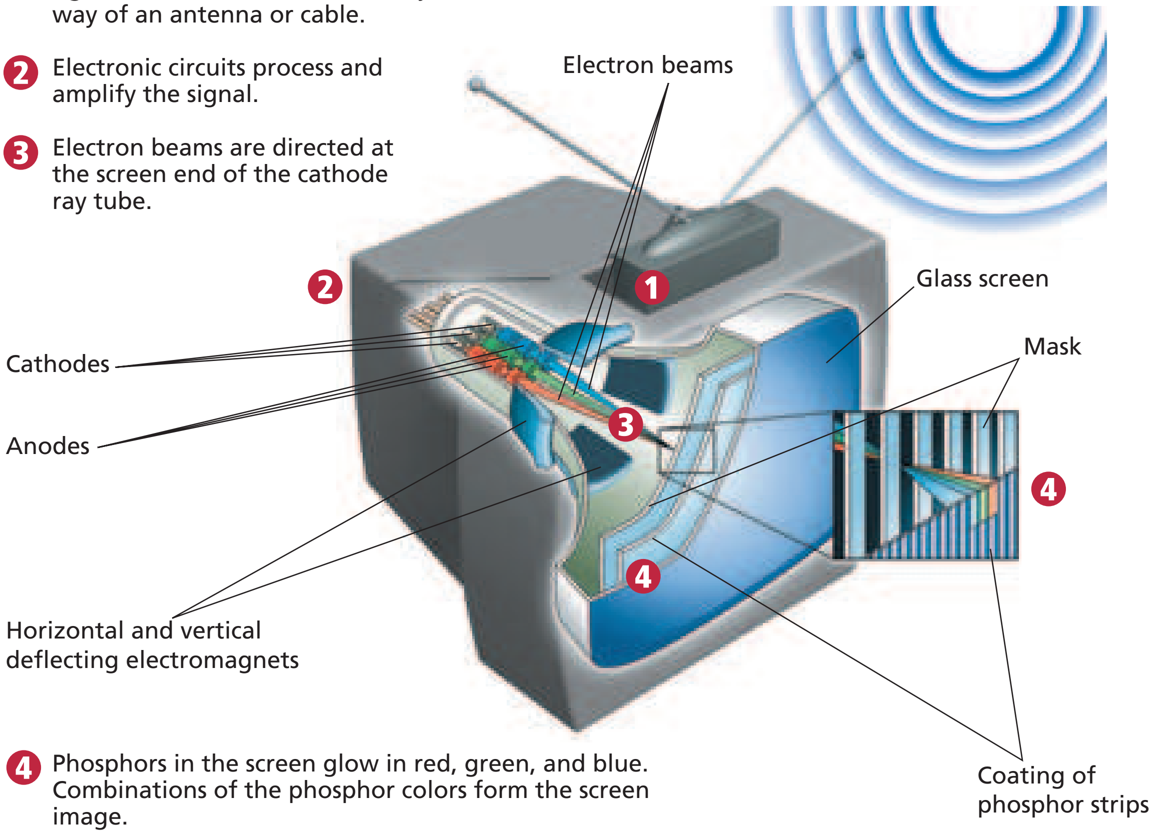
1. Why did it take so long to discover the first noble gas element?
2. Research one of the most recently discovered elements. New elements are created in particle accelerators and tend to be very unstable. Because of this, many of the properties of a new element can not be determined. Using periodic group trends in melting and boiling point, predict whether the new element you selected is likely to be a solid, liquid, or gas.

How It Works

Television Screen

Most television screens are part of a cathode ray tube. As you know, a cathode ray tube is an evacuated chamber which produces a beam of electrons, known as a cathode ray. Electronic circuitry inside the television processes an electronic signal received from the television station. The processed signal is used to vary the strength of several electron beams, while magnetic fields are used to direct the beams to different parts of the screen.

1. The television receives an electronic signal from a television station by way of an antenna or cable.
2. Electronic circuits process and amplify the signal.
3. Electron beams are directed at the screen end of the cathode ray tube.



4. Phosphors in the screen glow in red, green, and blue. Combinations of the phosphor colors form the screen image.

Thinking Critically

1. **Relating Cause and Effect** Why don't the phosphors in a television screen glow when the television is turned off?
2. **Inferring** Why is the length of time over which a phosphor emits light an important factor to consider when designing a television screen?

CHAPTER 6 STUDY GUIDE

Summary

6.1 Development of the Modern Periodic Table

- Periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- Newlands's law of octaves, which was never accepted by fellow scientists, organized the elements by increasing atomic mass. Mendeleev's periodic table, which also organized elements by increasing atomic mass, became the first widely accepted organization scheme for the elements. Moseley fixed the errors inherent in Mendeleev's table by organizing the elements by increasing atomic number.
- The periodic table organizes the elements into periods (rows) and groups (columns) by increasing atomic number. Elements with similar properties are in the same group.
- Elements are classified as either metals, nonmetals, or metalloids. The stair-step line on the table separates metals from nonmetals. Metalloids border the stair-step line.

6.2 Classification of the Elements

- Elements in the same group on the periodic table have similar chemical properties because they have the same valence electron configuration.
- The four blocks of the periodic table can be characterized as follows:

  s-block: filled or partially filled s orbitals.
  p-block: filled or partially filled p orbitals.
  d-block: filled outermost s orbital of energy level *n*, and filled or partially filled d orbitals of energy level $n - 1$.
  f-block: filled outermost s orbital, and filled or partially filled 4f and 5f orbitals.
- For the group A elements, an atom's group number equals its number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.
- The s^2p^6 electron configuration of the group 8A elements (noble gases) is exceptionally stable.

6.3 Periodic Trends

- Atomic radii generally decrease as you move left-to-right across a period, and increase as you move down a group.
- Positive ions are smaller than the neutral atoms from which they form. Negative ions are larger than the neutral atoms from which they form.
- Ionic radii of both positive and negative ions decrease as you move left-to-right across a period. Ionic radii of both positive and negative ions increase as you move down a group.
- Ionization energy indicates how strongly an atom holds onto its electrons. After the valence electrons have been removed from an atom, there is a tremendous jump in the ionization energy required to remove the next electron.
- Ionization energies generally increase as you move left-to-right across a period, and decrease as you move down a group.
- The octet rule states that atoms gain, lose, or share electrons in order to acquire the stable electron configuration of a noble gas.
- Electronegativity, which indicates the ability of atoms of an element to attract electrons in a chemical bond, plays a role in determining the type of bond formed between elements in a compound.
- Electronegativity values range from 0.7 to 3.96, and generally increase as you move left-to-right across a period, and decrease as you move down a group.

Vocabulary

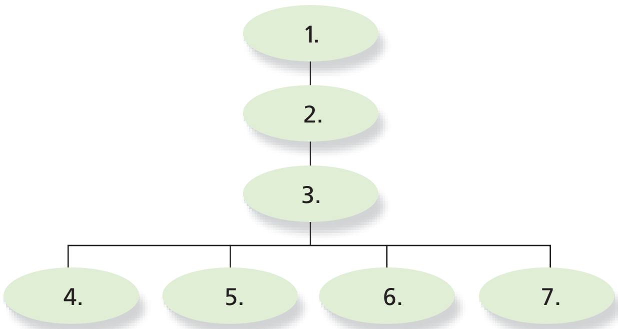
- alkali metal (p. 155)
- alkaline earth metal (p. 155)
- electronegativity (p. 168)
- group (p. 154)
- halogen (p. 158)
- inner transition metal (p. 158)
- ion (p. 165)
- ionization energy (p. 167)
- metal (p. 155)
- metalloid (p. 158)
- noble gas (p. 158)
- nonmetal (p. 158)
- octet rule (p. 168)
- period (p. 154)
- periodic law (p. 153)
- representative element (p. 154)
- transition element (p. 154)
- transition metal (p. 158)

CLICK HERE

Go to the Chemistry Web site at **science.glencoe.com** *or use the* **Chemistry CD-ROM** *for additional Chapter 6 Assessment.*

Concept Mapping

25. Complete the concept map using the following terms: electronegativity, electron configuration, periodic trends, ionic radius, atomic radius, ionization energy, and periodic table.



Mastering Concepts

26. Explain how Mendeleev's periodic table was in error. How was this error fixed? (6.1)

27. Explain the contribution of Newlands's law of octaves to the development of the modern periodic table. (6.1)

28. German chemist Lothar Meyer and Russian chemist Dmitri Mendeleev both proposed similar periodic tables in 1869. Why is Mendeleev generally given credit for the periodic table? (6.1)

29. How was Mendeleev's periodic table organized? (6.1)

30. What is the periodic law? (6.1)

31. Identify each of the following as a metal, nonmetal, or metalloid. (6.1)

a. oxygen
b. barium
c. germanium
d. iron
e. neon
f. praseodymium

32. Describe the general characteristics of metals. (6.1)

33. Match each numbered item on the right with the lettered item that it is related to on the left. (6.1)

a. alkali metals	**1.** group 8A
b. halogens	**2.** group 1A
c. alkaline earth metals	**3.** group 2A
d. noble gases	**4.** group 7A

34. Identify each of the elements in problem 31 as a representative element or a transition element. (6.1)

35. Sketch a simplified periodic table and use labels to identify the alkali metals, alkaline earth metals, transition metals, inner transition metals, noble gases, and halogens. (6.1)

36. A shiny solid element also is ductile. What side of the periodic table is it likely to be found? (6.1)

37. What are the general properties of a metalloid? List three metalloid elements. (6.1)

38. What is the purpose of the heavy stair-step line on the periodic table? (6.1)

39. Describe the two types of numbering used to identify groups on the periodic table. (6.1)

40. Give the chemical symbol of each of the following elements. (6.1)

a. the two elements that are liquids at room temperature
b. the noble gas with the greatest atomic mass
c. any metal from group 4A
d. any inner transition metal

41. Why do the elements chlorine and iodine have similar chemical properties? (6.2)

42. How are the numbers of valence electrons of the group A elements related to the group number? (6.2)

43. How is the energy level of an atom's valence electrons related to the period it is in on the periodic table? (6.2)

44. How many valence electrons do each of the noble gases have? (6.2)

45. What are the four blocks of the periodic table? (6.2)

46. In general, what electron configuration has the greatest stability? (6.2)

47. Determine the group, period, and block in which each of the following elements is located on the periodic table. (6.2)

a. $[Kr]5s^2 4d^1$
b. $[Ar]4s^2 3d^{10} 4p^3$
c. $[He]2s^2 2p^6$
d. $[Ne]3s^2 3p^1$

48. Categorize each of the elements in problem 47 as a representative element or a transition metal. (6.2)

49. Explain how an atom's valence electron configuration determines its place on the periodic table. (6.2)

50. Write the electron configuration for the element fitting each of the following descriptions. (6.2)

a. the metal in group 5A
b. the halogen in period 3
c. the alkali metal in period 2
d. the transition metal that is a liquid at room temperature

CONTENTS

51. Explain why the radius of an atom cannot be measured directly. (6.3)

52. Given any two elements within a group, is the element with the larger atomic number likely to have a larger or smaller atomic radius than the other element? (6.2)

53. Which elements are characterized as having their d orbitals fill with electrons as you move left-to-right across a period? (6.2)

54. Explain why is it harder to remove an inner shell electron than a valence electron from an atom. (6.3)

55. An element forms a negative ion when ionized. On what side of the periodic table is the element located? Explain. (6.3)

56. Of the elements magnesium, calcium, and barium, which forms the ion with the largest radius? The smallest? What periodic trend explains this? (6.3)

57. What is ionization energy? (6.3)

58. Explain why each successive ionization of an electron requires a greater amount of energy. (6.3)

59. Which group has the highest ionization energies? Explain why. (6.3)

60. Define an ion. (6.3)

61. How does the ionic radius of a nonmetal compare with its atomic radius? Explain why the change in radius occurs. (6.3)

62. Explain why atomic radii decrease as you move left-to-right across a period. (6.3)

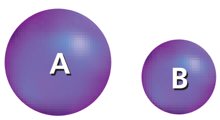
63. Which element in each pair has the larger ionization energy? (6.3)

a. Li, N
b. Kr, Ne
c. Cs, Li

64. Explain the octet rule. (6.3)

65. Use the illustration of spheres A and B to answer each of the following questions. Explain your reasoning for each answer. (6.3)

a. If A is an ion and B is an atom of the same element, is the ion a positive or negative ion?



b. If A and B represent the atomic radii of two elements in the same period, what is their correct order (left-to-right)?

c. If A and B represent the ionic radii of two elements in the same group, what is their correct order (top-to-bottom)?

66. How many valence electrons do elements in each of the following groups have? (6.3)

a. group 8A
b. group 3A
c. group 1A

67. Na^+ and Mg^{2+} ions each have ten electrons surrounding their nuclei. Which ion would you expect to have the larger radius? Why? (6.3)

Mixed Review

Sharpen your problem-solving skills by answering the following.

68. Match each numbered item on the right with the lettered item that it is related to on the left.

a. group A elements	**1.** periods
b. columns	**2.** representative elements
c. group B elements	**3.** groups
d. rows	**4.** transition elements

69. Which element in each pair is more electronegative?

a. K, As
b. N, Sb
c. Sr, Be

70. Explain why the s-block of the periodic table is two groups wide, the p-block is six groups wide, and the d-block is ten groups wide.

71. Arrange the elements oxygen, sulfur, tellurium, and selenium in order of increasing atomic radii. Is your order an example of a group trend or a period trend?

72. Identify the elements with the following valence electron configurations.

a. $5s^1$	**c.** $3s^2$
b. $4s^2 3d^2$	**d.** $4s^2 4p^3$

73. Which of the following is not a reason why atomic radii increase as you move down a group?
a. shielding of inner electrons
b. valence electrons in larger orbitals
c. increased charge in the nucleus

74. Explain why there are no p-block elements in the first period of the periodic table.

75. Identify each of the following as an alkali metal, alkaline earth metal, transition metal, or inner transition metal.

a. cesium	**d.** ytterbium
b. zirconium	**e.** uranium
c. gold	**f.** francium

76. An element is a brittle solid that does not conduct electricity well. Is the element a metal, nonmetal, or metalloid?

Thinking Critically

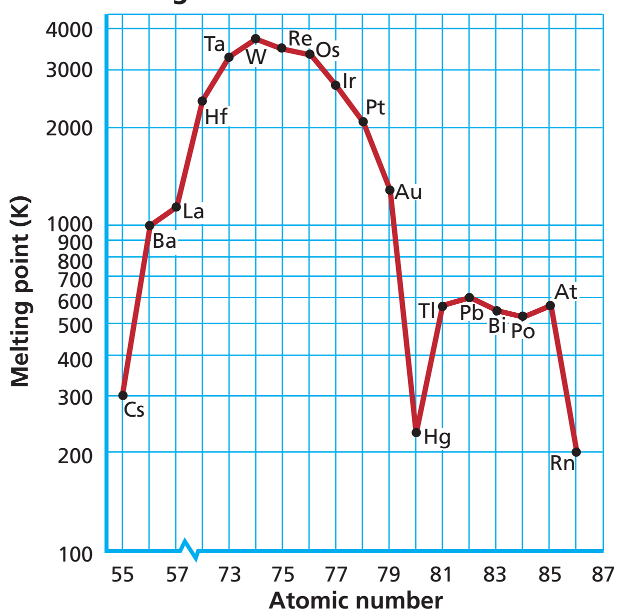
77. Interpreting Data Given the following data about an atom's ionization energies, predict its valence electron configuration. Explain your reasoning.

Ionization Data	
Ionization	**Ionization Energy (kJ/mol)**
First	734
Second	1850
Third	16 432

78. Applying Concepts Sodium forms a 1+ ion, while fluorine forms a 1− ion. Write the electron configuration for each ion. Why don't these two elements form 2+ and 2− ions, respectively?

79. Interpreting Data The melting points of the period 6 elements are plotted versus atomic number in the graph shown below. Determine the trends in melting point by analyzing the graph and the orbital configurations of the elements. Form a hypothesis that explains the trends. (Hint: In Chapter 5, you learned that half-filled sets of orbitals are more stable than other configurations of partially filled orbitals.)

Melting Points of the Period 6 Elements



Group 5A Density Data		
Element	**Atomic Number**	**Density (g/cm^3)**
nitrogen	7	1.25×10^{-3}
phosphorus	15	1.82
arsenic	33	5.73
antimony	51	6.70
bismuth	83	9.78

80. Making and Using Graphs The densities of the group 5A elements are given in the table above. Plot density versus atomic number and state any trends you observe.

Writing in Chemistry

81. In the early 1800s, German chemist J. W. Dobereiner proposed that some elements could be classified into sets of three, called triads. Research and write a report on Dobereiner's triads. What elements comprised the triads? How were the properties of elements within a triad similar?

82. Electron affinity is another periodic property of the elements. Research and write a report on what electron affinity is and describe its group and period trends.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

83. Define matter. Identify whether or not each of the following is a form of matter. (Chapter 1)

a. microwaves
b. helium inside a balloon
c. heat from the Sun
d. velocity
e. a speck of dust
f. the color blue

84. Convert the following mass measurements as indicated. (Chapter 2)

a. 1.1 cm to meters
b. 76.2 pm to millimeters
c. 11 Mg to kilograms
d. 7.23 micrograms to kilograms

85. How is the energy of a quantum of emitted radiation related to the frequency of the radiation? (Chapter 5)

86. What element has the ground-state electron configuration of $[Ar]4s^23d^6$? (Chapter 5).

Use these questions and the test-taking tip to prepare for your standardized test.

1. Periodic law states that elements show a _____.

a. repetition of their physical properties when arranged by increasing atomic radius
b. repetition of their chemical properties when arranged by increasing atomic mass
c. periodic repetition of their properties when arranged by increasing atomic number
d. periodic repetition of their properties when arranged by increasing atomic mass

2. Elements in the same group of the periodic table have the same _____.

a. number of valence electrons
b. physical properties
c. number of electrons
d. electron configuration

3. All of the following are true EXCEPT _____.

a. atomic radius of Na < atomic radius of Mg
b. electronegativity of C > electronegativity of B
c. ionic radius of Br^- > atomic radius of Br
d. first ionization energy of K > first ionization energy of Rb

4. Which of the following is NOT true of an atom obeying the octet rule?

a. obtains a full set of eight valence electrons
b. acquires the valence configuration of a noble gas
c. possesses eight electrons in total
d. has a s^2p^6 valence configuration

5. What is the group, period, and block of an atom with the electron configuration $[Ar]4s^23d^{10}4p^4$?

a. group 4A, period 4, d-block
b. group 6A, period 3, p-block
c. group 4A, period 4, p-block
d. group 6A, period 4, p-block

6. Moving down a group on the periodic table, which two atomic properties follow the same trend?

a. atomic radius and ionization energy
b. ionic radius and atomic radius
c. ionization energy and ionic radius
d. ionic radius and electronegativity

Interpreting Tables Use the periodic table and the table at the bottom of the page to answer questions 7 and 8.

7. It can be predicted that silicon will experience a large jump in ionization energy after its _____.

a. second ionization
b. third ionization
c. fourth ionization
d. fifth ionization

8. Which of the following requires the most energy?

a. second ionization of Li
b. fourth ionization of N
c. first ionization of Ne
d. third ionization of Be

9. Niobium (Nb) is a(n) _____.

a. nonmetal **c.** alkali metal
b. transition metal **d.** halogen

10. It can be predicted that element 118 would have properties similar to a(n) _____.

a. alkali earth metal **c.** metalloid
b. halogen **d.** noble gas

THE PRINCETON REVIEW **TEST-TAKING TIP**

Practice, Practice, Practice Practice to improve your performance on standardized tests. Don't compare yourself to anyone else.

Successive Ionization Energies for the Period 2 Elements

Element	Valence electrons	Ionization energy (kJ/mol)* 1st	2nd	3rd	4th	5th	6th	7th	8th	9th
Li	1	520	7300							
Be	2	900	1760	14 850						
B	3	800	2430	3660	25 020					
C	4	1090	2350	4620	6220	37 830				
N	5	1400	2860	4580	7480	9440	53 270			
O	6	1310	3390	5300	7470	10 980	13 330	71 330		
F	7	1680	3370	6050	8410	11 020	15 160	17 870	92 040	
Ne	8	2080	3950	6120	9370	12 180	15 240	20 000	23 070	115 380

* mol is an abbreviation for mole, a quantity of matter.