CHAPTER 6 The Periodic Table and Periodic Law

BIG Idea Periodic trends in the properties of atoms allow us to predict physical and chemical properties.

6.1 Development of the Modern Periodic Table

MAIN Idea The periodic table evolved over time as scientists discovered more useful ways to compare and organize the elements.

6.2 Classification of the Elements

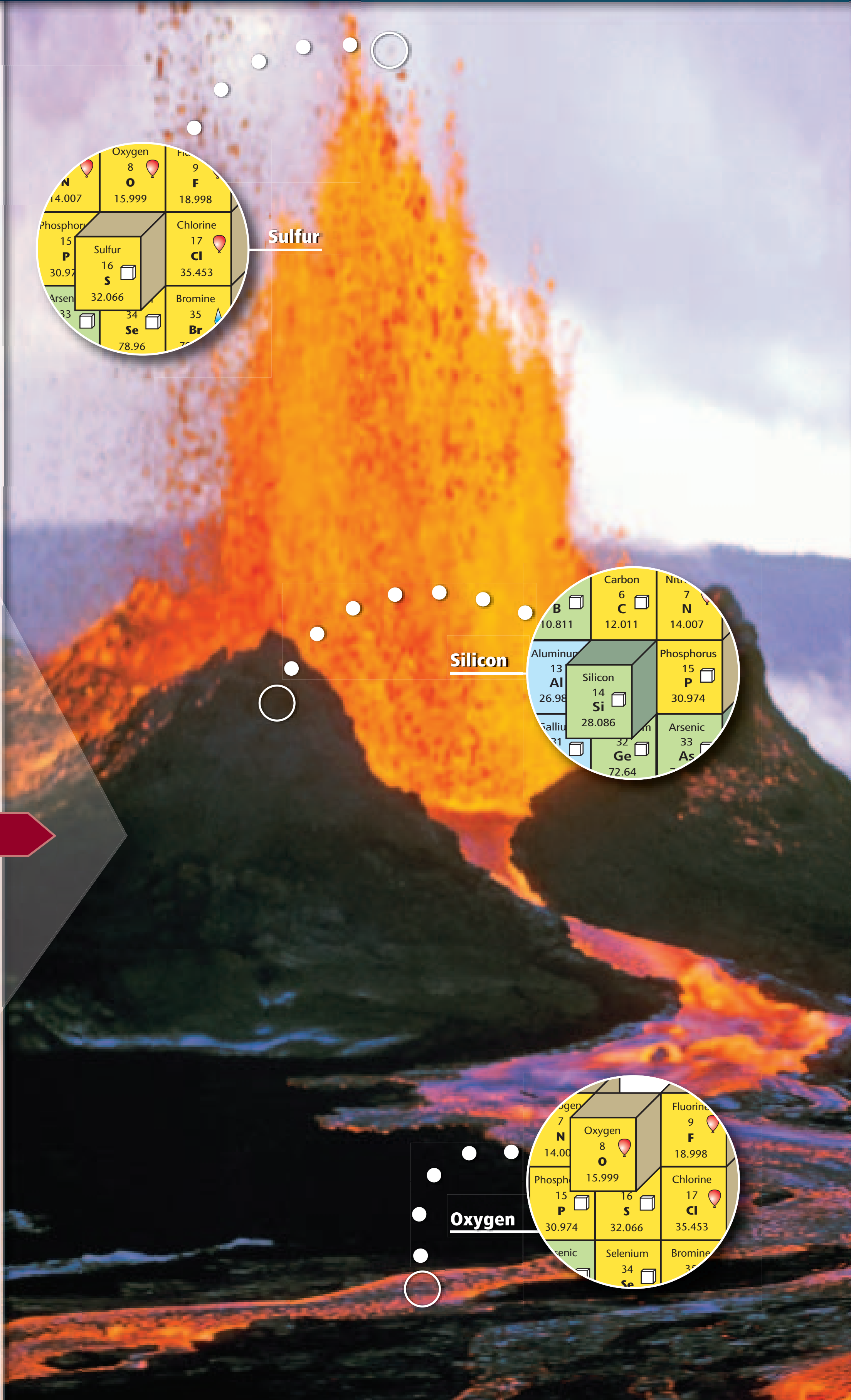
MAIN Idea Elements are organized into different blocks in the periodic table according to their electron configurations.

6.3 Periodic Trends

MAIN Idea Trends among elements in the periodic table include their size and their ability to lose or attract electrons.

ChemFacts

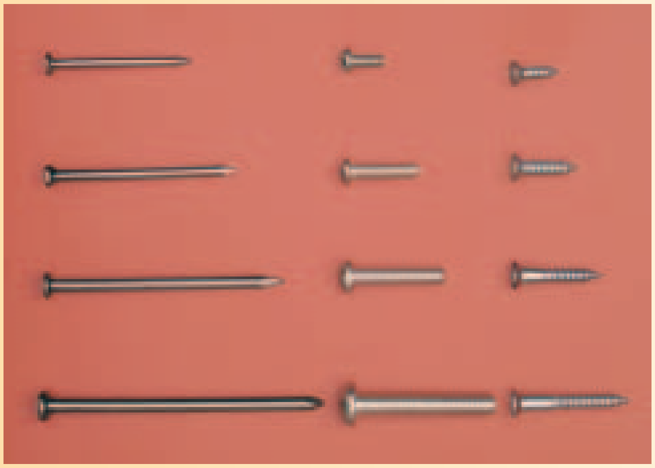
- There are 117 elements in the current periodic table. Only 90 of them occur naturally.
- Hydrogen is the most abundant element in the universe (75%) and oxygen is the most abundant element on Earth (50%).
- A 70-kg human body contains approximately 43 kg of oxygen.
- The total amount of astatine in the Earth's crust is less than 30 g, which makes it the least abundant element on Earth.



LAUNCH Lab

How can you recognize trends?

The periodic table of the elements is arranged so that the properties of the elements repeat in a regular way. Such an arrangement can also be used for common items.



Procedure

1. Read and complete the lab safety form.
2. Obtain a sample of **fasteners,** including **bolts, screws,** and **nails.**
3. Measure the length of each fastener with a **ruler.**
4. Use a **balance** to measure the mass of each fastener.
5. Place the nails in a series from smallest to largest.
6. Continue to arrange a series of screws and a series of bolts that also correspond to the series of nails created in Step 5.

Analysis

1. **Make a table** listing the length and mass of each fastener.
2. **Describe** the trend in mass as you go from left to right across each row of the table.
3. **Describe** the trend in mass as you go down each column of the table.
4. **Analyze** your organization of the fasteners, and explain any other trends that you find in the table.

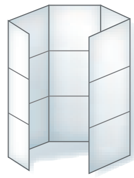
Inquiry Create a periodic table of carbonated beverages in a manner similar to this lab. What properties did you use?

FOLDABLES™ Study Organizer

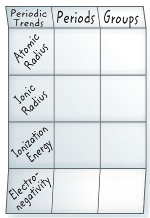
Periodic Trends Make the following Foldable to organize information about periodic trends.

STEP 1 Fold a sheet of paper into thirds lengthwise.

STEP 2 Make a 2-cm fold along one narrow edge and then fold the sheet in half below this line, and then half again.



STEP 3 Unfold the sheet and draw lines along all fold lines. Label as follows: *Periodic Trends, Periods,* and *Groups* in the first row, and *Atomic Radius, Ionic Radius, Ionization Energy,* and *Electronegativity* in the first column.



FOLDABLES **Use this Foldable with Section 6.3.** As you read the section, summarize the period and group trends of several properties of elements.

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Section 6.1

Objectives

- **Trace** the development of the periodic table.
- **Identify** key features of the periodic table.

Review Vocabulary

atomic number: the number of protons in an atom

New Vocabulary

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

Development of the Modern Periodic Table

MAIN Idea The periodic table evolved over time as scientists discovered more useful ways to compare and organize the elements.

Real-World Reading Link Imagine grocery shopping if all the apples, pears, oranges, and peaches were mixed into one bin at the grocery store. Organizing things according to their properties is often useful. Scientists organize the many different types of chemical elements in the periodic table.

Development of the Periodic Table

In the late 1700s, French scientist Antoine Lavoisier (1743–1794) compiled a list of all elements that were known at the time. The list, shown in **Table 6.1,** contained 33 elements organized in four categories. Many of these elements, such as silver, gold, carbon, and oxygen, have been known since prehistoric times. The 1800s brought a large increase in the number of known elements. The advent of electricity, which was used to break down compounds into their components, and the development of the spectrometer, which was used to identify the newly isolated elements, played major roles in the advancement of chemistry. The industrial revolution of the mid-1800s also played a major role, 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.

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, and a way to organize the elements began in earnest.

Table 6.1	Lavoisier's Table of Simple Substances (Old English Names)
Gases	light, heat, dephlogisticated air, phlogisticated gas, inflammable air
Metals	antimony, silver, arsenic, bismuth, cobalt, copper, tin, iron, manganese, mercury, molybdena, nickel, gold, platina, lead, tungsten, zinc
Nonmetals	sulphur, phosphorus, pure charcoal, radical muriatique*, radical fluorique*, radical boracique*
Earths	chalk, magnesia, barote, clay, siliceous earth

* no English name

John Newlands In 1864, English chemist John Newlands (1837–1898) proposed an organizational scheme for the elements. He noticed that when the elements were arranged by increasing atomic mass, their properties repeated every eighth element. 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,* after the musical octave in which notes repeat every eighth tone. **Figure 6.1** shows how Newlands organized 14 of the elements known in the mid-1860s. Acceptance of the law of octaves was hampered because the law did not work for all of the known elements. Also, the use of the word *octave* was harshly criticized by fellow scientists, who thought that the musical analogy was unscientific. While his law was not generally accepted, the passage of a few years would show that Newlands was basically correct; the properties of elements do repeat in a periodic way.

Meyer and Mendeleev 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 organizational scheme first. Like Newlands several years earlier, Mendeleev noticed that when the elements were ordered by increasing atomic mass, there was a 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 a periodic table. Mendeleev's table, shown in **Figure 6.2,** became widely accepted because he predicted the existence and properties of undiscovered elements that were later found. 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.

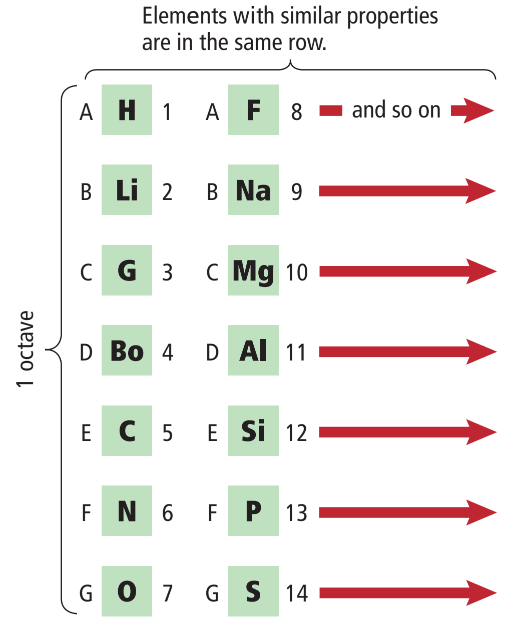
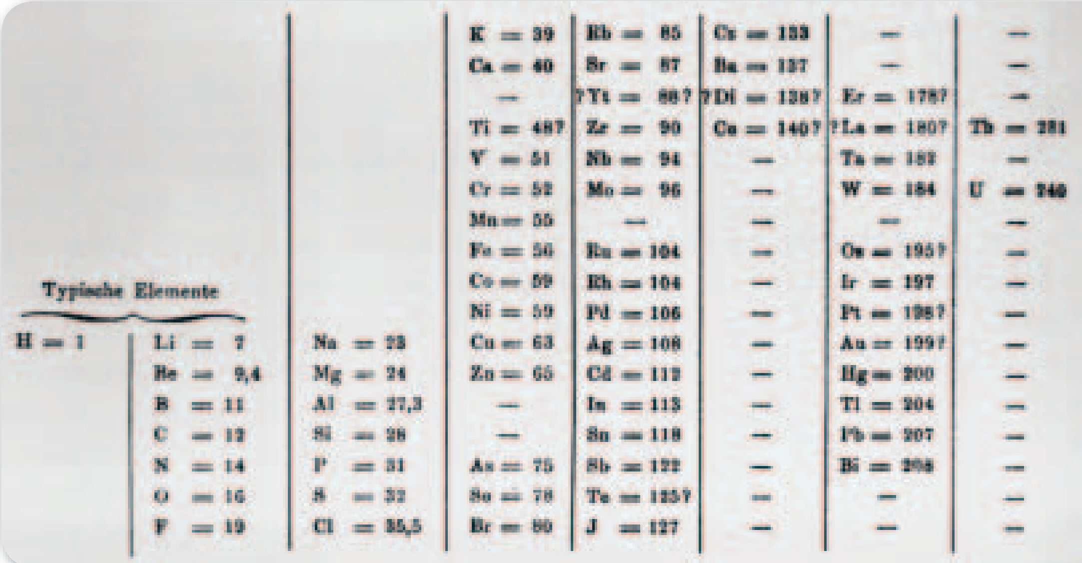


Figure 6.1 John Newlands noticed that the properties of elements repeated every eighth element, in the same way musical notes repeat every eighth note and form octaves.



Typische Elemente

			K = 39	Rb = 85	Cs = 133	—	—
			Ca = 40	Sr = 87	Ba = 137	—	—
			—	?Yt = 88?	?Di = 138?	Er = 178?	—
			Ti = 48?	Zr = 90	Ce = 140?	?La = 180?	Th = 231
			V = 51	Nb = 94	—	Ta = 182	—
			Cr = 52	Mo = 96	—	W = 184	U = 240
			Mn = 55	—	—	—	—
			Fe = 56	Ru = 104	—	Os = 195?	—
			Co = 59	Rh = 104	—	Ir = 197	—
			Ni = 59	Pd = 106	—	Pt = 198?	—
H = 1	Li = 7	Na = 23	Cu = 63	Ag = 108	—	Au = 199?	—
	Be = 9,4	Mg = 24	Zn = 65	Cd = 112	—	Hg = 200	—
	B = 11	Al = 27,3	—	In = 113	—	Tl = 204	—
	C = 12	Si = 28	—	Sn = 118	—	Pb = 207	—
	N = 14	P = 31	As = 75	Sb = 122	—	Bi = 208	—
	O = 16	S = 32	Se = 78	Te = 125?	—	—	—
	F = 19	Cl = 35,5	Br = 80	J = 127	—	—	—

Figure 6.2 In the first version of his table, published in 1869, Mendeleev arranged elements with similar chemical properties horizontally. He left empty spaces for elements that were not yet discovered.

Moseley Mendeleev's table, however, was not completely correct. After several new elements were discovered and the 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 (1887–1915). As you might 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, 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.**

VOCABULARY
WORD ORIGIN
Periodic
comes from the Greek word *periodos,* meaning *way around, circuit*

Reading Check **Compare and contrast** the ways in which Mendeleev and Moseley organized the elements.

Table 6.2 summarizes the contributions of Newlands, Meyer, Mendeleev, and Moseley to the development of the periodic table. The periodic table brought order to seemingly unrelated facts and became a significant tool for chemists. It is a useful reference for understanding and predicting the properties of elements and for organizing knowledge of atomic structure. Do the Problem-Solving Lab later in this chapter to see how the periodic law can be used to predict unknown elemental properties.

Table 6.2 Contributions to the Classification of Elements

John Newlands (1837–1898) • arranged elements by increasing atomic mass • noticed the repetition of properties every eighth element • created the law of octaves
Lothar Meyer (1830–1895) • demonstrated a connection between atomic mass and elemental properties • arranged the elements in order of increasing atomic mass
Dmitri Mendeleev (1834–1907) • demonstrated a connection between atomic mass and elemental properties • arranged the elements in order of increasing atomic mass • predicted the existence and properties of undiscovered elements
Henry Moseley (1887–1915) • discovered that atoms contain a unique number of protons called the atomic number • arranged elements in order of increasing atomic number, which resulted in a periodic pattern of properties

The Modern Periodic Table

The modern periodic 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 6.3.** The boxes are arranged in order of increasing atomic number into a series of columns, called **groups** or families, and rows, called **periods.** The table is shown in **Figure 6.5** on the next page and on the inside back cover of your textbook.

Reading Check **Define** *groups* and *periods.*

Beginning with hydrogen in period 1, there are a total of seven periods. Each group is numbered 1 through 18. For example, period 4 contains potassium and calcium. Scandium (Sc) is in the third column from the left, which is group 3. Oxygen is in group 16. The elements in groups 1, 2, and 13 to 18 possess a wide range of chemical and physical properties. For this reason, they are often referred to as the main group, or **representative elements.** The elements in groups 3 to 12 are referred to as the **transition elements.** Elements are classified as metals, nonmetals, and metalloids.

Metals Elements that are generally shiny when smooth and clean, solid at room temperature, and good conductors of heat and electricity are called **metals.** Most metals are also malleable and ductile, meaning that they can be pounded into thin sheets and drawn into wires, respectively. Most representative elements and all transition elements are metals. If you look at boron (B) in column 13, you will see a heavy stairstep line that zigzags down to astatine (At) at the bottom of group 17. This stairstep line is a visual divider between the metals and the nonmetals on the table. In **Figure 6.5,** metals are represented by the blue boxes.

Alkali metals Except for hydrogen, all of the elements on the left side of the table are metals. The group 1 elements (except for hydrogen) are known as the **alkali metals.** Because they are so reactive, alkali metals usually exist as compounds with other elements. Two familiar alkali metals are sodium (Na), one of the components of salt, and lithium (Li), often used in batteries.

Alkaline earth metals The **alkaline earth metals** are in group 2. They are also highly reactive. Calcium (Ca) and magnesium (Mg), two minerals important for your health, are examples of alkaline earth metals. Because magnesium is solid and relatively light, it is used in the fabrication of electronic devices, such as the laptop shown in **Figure 6.4.**

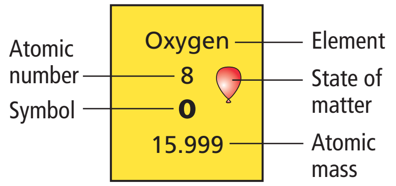
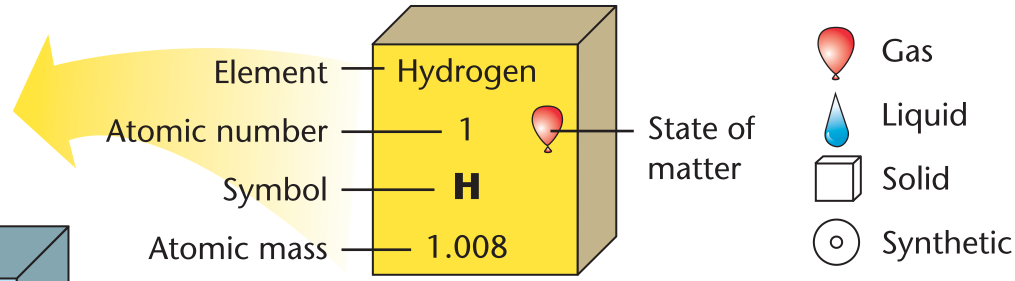


Figure 6.3 A typical box from the periodic table contains the element's name, its chemical symbol, its atomic number, its atomic mass, and its state.

Figure 6.4 Because magnesium is light and strong, it is often used in the production of electronic devices. For instance, this laptop case is made of magnesium.

Figure 6.5

PERIODIC TABLE OF THE ELEMENTS



	1	2	3	4	5	6	7	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** 55.847	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** 138.905	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** 140.115	Praseodymium 59 **Pr** 140.908	Neodymium 60 **Nd** 144.242	Promethium 61 **Pm** (145)	Samarium 62 **Sm** 150.36	Europium 63 **Eu** 151.965
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)

Concepts In Motion

Interactive Figure To see an animation of the periodic table, visit glencoe.com.

Metal

Metalloid

Nonmetal

Recently observed

	10	11	12	13	14	15	16	17	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** 32.066	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** 72.61	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** 114.82	Tin 50 **Sn** 118.710	Antimony 51 **Sb** 121.757	Tellurium 52 **Te** 127.60	Iodine 53 **I** 126.904	Xenon 54 **Xe** 131.290
	Platinum 78 **Pt** 195.08	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** 208.982	Astatine 85 **At** 209.987	Radon 86 **Rn** 222.018
	Darmstadtium 110 **Ds** (281)	Roentgenium 111 **Rg** (272)	Ununbium * 112 **Uub** (285)	Ununtrium * 113 **Uut** (284)	Ununquadium * 114 **Uuq** (289)	Ununpentium * 115 **Uup** (288)	Ununhexium * 116 **Uuh** (291)		Ununoctium * 118 **Uuo** (294)

* The names and symbols for elements 112, 113, 114, 115, 116, and 118 are temporary. Final names will be selected when the elements' discoveries are verified.

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)

PROBLEM-SOLVING LAB

Analyze Trends

Francium—solid, liquid, or gas?
Francium was discovered in 1939, but its existence was predicted by Mendeleev in the 1870s. It is the least stable of the first 101 elements: 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 Dmitri Mendeleev's prediction of the properties 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.

Think Critically

1. **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. Use the periodic law as a guide.

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	248
Francium	?	?	?

2. **Predict** whether francium is a solid, a liquid, or a gas. How can you support your prediction?
3. **Infer** which column of data presents the greatest possible error in making a prediction. Explain.
4. **Determine** why producing 1 million francium atoms per second is not enough to make measurements, such as density or melting point.

Transition and inner transition metals The transition elements are divided into **transition metals** and **inner transition metals.** The two sets of inner transition metals, known as the **lanthanide series** and **actinide series,** are located along the bottom of the periodic table. The rest of the elements in groups 3 to 12 make up the transition metals. Elements from the lanthanide series are used extensively as phosphors, substances that emit light when struck by electrons. Because it is strong and light, the transition metal titanium is used to make frames for bicycles and and eyeglasses.

Connection to Biology **Nonmetals** Nonmetals occupy the upper-right side of the periodic table. They are represented by the yellow boxes in **Figure 6.5.** **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 most abundant element in the human body is the nonmetal oxygen, which constitutes 65% of the body mass. Group 17 is comprised of highly reactive elements that are known as **halogens.** Like the group 1 and group 2 elements, the halogens are often part of compounds. Compounds made with the halogen fluorine (F) are commonly added to toothpaste and drinking water to prevent tooth decay. The extremely unreactive group 18 elements are commonly called the **noble gases** and are used in neon signs.

VOCABULARY
SCIENCE USAGE V. COMMON USAGE
Conductor
Science usage: a substance or body capable of transmitting electricity, heat, or sound
Copper is a good conductor of heat.

Common usage: a person who conducts an orchestra, chorus, or other group of musical performers
The new conductor helped the orchestra perform at its best.

■ **Figure 6.6** Scientists developing submarine technology created a robot that looks and swims like a real fish. Its body is made of a silicon resin that softens in water.

Metalloids The elements in the green boxes bordering the stairstep line in **Figure 6.5** are called metalloids, or semimetals. **Metalloids** have physical and chemical properties of both metals and nonmetals. Silicon (Si) and germanium (Ge) are two important metalloids, used extensively in computer chips and solar cells. Silicon is also used to make prosthetics or in lifelike applications, as shown in **Figure 6.6.**

This introduction to the periodic table touches only the surface of its usefulness. You can refer to the Elements Handbook at the end of your textbook to learn more about the elements in the various groups.

Section 6.1 Assessment

Section Summary

- The elements were first organized by increasing atomic mass, which led to inconsistencies. Later, they were organized by increasing atomic number.
- The periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- The periodic table organizes the elements into periods (rows) and groups (columns); elements with similar properties are in the same group.
- Elements are classified as either metals, nonmetals, or metalloids.

1. MAIN Idea **Describe** the development of the modern periodic table. Include contributions made by Lavoisier, Newlands, Mendeleev, and Moseley.
2. **Sketch** a simplified version of the periodic table, and indicate the location of 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. **Compare** For each of the given elements, list two other elements with similar chemical properties.
   a. iodine (I) **b.** barium (Ba) **c.** iron (Fe)
6. **Compare** According to the periodic table, which two elements have an atomic mass less than twice their atomic number?
7. **Interpret Data** A company plans to make an electronic device. They need to use an element that has chemical behavior similar to that of silicon (Si) and lead (Pb). The element must have an atomic mass greater than that of sulfur (S), but less than that of cadmium (Cd). Use the periodic table to determine which element the company could use.

Section 6.2

Objectives

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

Review Vocabulary

valence electron: electron in an atom's outermost orbitals; determines the chemical properties of an atom

Classification of the Elements

MAIN Idea Elements are organized into different blocks in the periodic table according to their electron configurations.

Real-World Reading Link A house number is not enough to deliver a letter to the correct address. More information, such as street name, city, and state, is necessary to deliver the letter. Similarly, chemical elements are identified according to details about the arrangement of their electrons.

Organizing the Elements by Electron Configuration

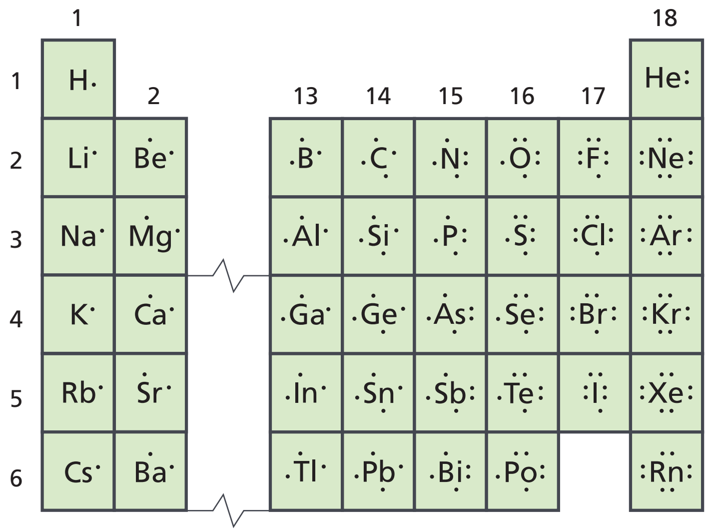
As you read in Chapter 5, electron configuration determines the chemical properties of an element. Writing out electron configurations using the aufbau diagram can be tedious. Fortunately, you can determine an atom's electron configuration and its number of valence electrons from its positon on the periodic table. The electron configurations for some of the group 1 elements are listed in **Table 6.3.** All four configurations have a single electron in their outermost orbitals.

Valence electrons Recall that electrons in the highest principal energy level of an atom are called valence electrons. Each of the group 1 elements has one electron in its highest energy level; thus, each element has one valence electron. The group 1 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 1 element has a valence electron configuration of s^1. Each group 2 element has a valence electron configuration of s^2. Each column in groups 1, 2, and 13 to 18 on the periodic table has its own 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 levels and lithium is found in period 2. Now look at gallium, with its electron configuration of $[Ar]4s^2 3d^{10} 4p^1$. Gallium's valence electrons are in the fourth energy level, and gallium is found in the fourth period.

Table 6.3 Electron Configuration for the Group 1 Elements

Period 1	hydrogen	$1s^1$	$1s^1$
Period 2	lithium	$1s^2 2s^1$	$[He]2s^1$
Period 3	sodium	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Period 4	potassium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	$[Ar]4s^1$



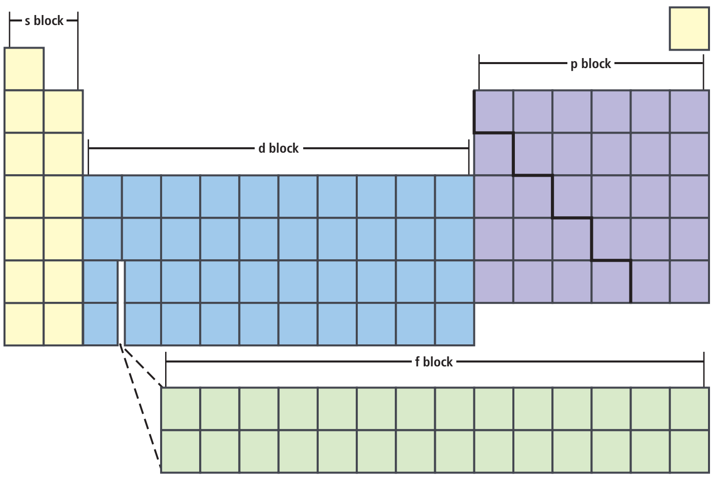
■ **Figure 6.7** The figure shows the electron-dot structure of most representative elements.

Observe ***How does the number of valence electrons vary within a group?***

Valence electrons of the representative elements Elements in group 1 have one valence electron; group 2 elements have two valence electrons. Group 13 elements have three valence electrons, group 14 elements have four, and so on. The noble gases in group 18 each have eight valence electrons, with the exception of helium, which has only two valence electrons. **Figure 6.7** shows how the electron-dot structures you learned in Chapter 5 illustrate the connection between group number and number of valence electrons. Notice that the number of valence electrons for the elements in group 13 to 18 is ten less than their group number.

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 6.8.**



■ **Figure 6.8** The periodic table is divided into four blocks—s, p, d, and f.

Analyze ***What is the relationship between the maximum number of electrons an energy sublevel can hold and the size of that block on the diagram?***

Concepts In Motion

Interactive Table Explore noble gas electron configurations at glencoe.com.

Table 6.4 Noble Gas Electron Configuration

Period	Principal Energy Level	Element	Electron Configuration
1	$n = 1$	helium	$1s^2$
2	$n = 2$	neon	$[He]2s^22p^6$
3	$n = 3$	argon	$[Ne]3s^23p^6$
4	$n = 4$	krypton	$[Ar]4s^24p^6$

s-Block elements The s-block consists of groups 1 and 2, and the element helium. Group 1 elements have partially filled s orbitals containing one valence electron and electron configurations ending in s^1. Group 2 elements have completely filled s orbitals containing two valence electrons and electron configurations ending in s^2. Because s orbitals hold two electrons at most, the s-block spans two groups.

p-Block elements After the s sublevel is filled, the valence electrons next occupy the p sublevel. The p-block, comprised of groups 13 through 18, contains elements with filled or partially filled p orbitals. There are no p-block elements in period 1 because the p sublevel does not exist for the first principal energy level ($n = 1$). The first p-block element is boron (B), in the second period. The p-block spans six groups because the three p orbitals can hold a maximum of six electrons. The group 18 elements (noble gases) are unique members of the p-block. Their atoms are so stable that they undergo virtually no chemical reactions. The electron configurations of the first four noble gas elements is shown in **Table 6.4.** 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. Together, the s- and p-blocks comprise the representative elements.

VOCABULARY

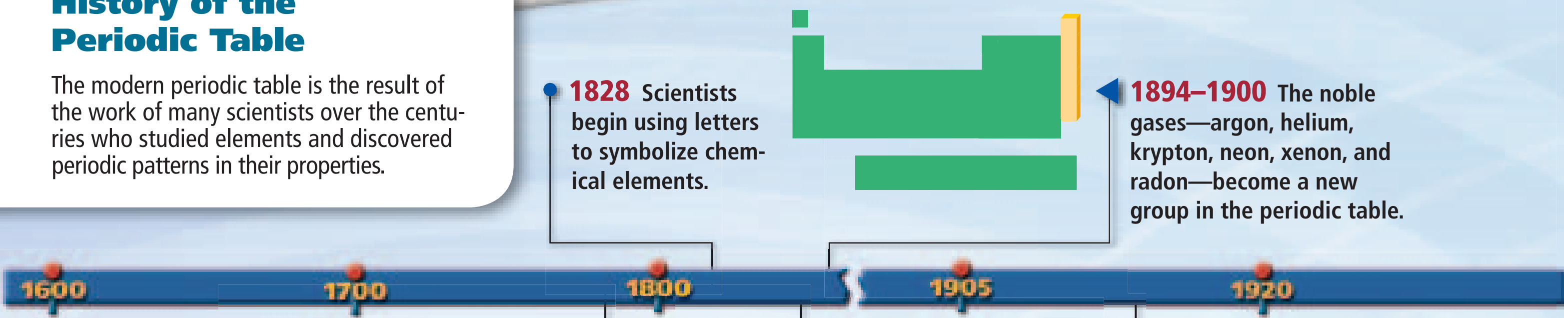
ACADEMIC VOCABULARY

Structure
something made up of more-or-less interdependent elements or parts
Many scientists were involved in the discovery of the structure of the atom.

■ **Figure 6.9**

History of the Periodic Table

The modern periodic table is the result of the work of many scientists over the centuries who studied elements and discovered periodic patterns in their properties.



1828 Scientists begin using letters to symbolize chemical elements.

1894–1900 The noble gases—argon, helium, krypton, neon, xenon, and radon—become a new group in the periodic table.

1789 Antoine Lavoisier defines the chemical element, develops a list of all known elements, and distinguishes between metals and nonmetals.

1869 Lothar Meyer and Dmitri Mendeleev independently develop tables based on element characteristics and predict the properties of unknown elements.

1913 Henry Moseley determines the atomic number of known elements and establishes that element properties vary periodically with atomic number.

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 usually 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 a 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 d orbital, which is partially filled, has an energy level of $n = 3$. As you read in Chapter 5, the aufbau Principle states that the 4s orbital has a lower energy level than the 3d orbital. Therefore, the 4s orbital is filled before the 3d orbital. The five d orbitals can hold a total of ten electrons; thus, the d-block spans ten groups on the periodic table.

f-Block elements The f-block contains the inner transition metals. Its 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.

Therefore, 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 orbitals containing electrons. Note that 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.

The development of the periodic table took many years and is still an ongoing project as new elements are synthetized. Refer to **Figure 6.9** to learn more about the history of the periodic table and the work of the many scientists who contributed to its development.

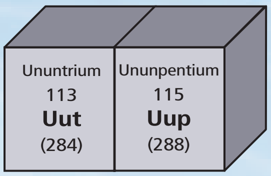
Reading Check **Summarize** how each block of the periodic table is defined.

CAREERS IN CHEMISTRY

Research Chemist Some nuclear chemists specialize in studying the newest and heaviest elements. To produce heavy elements, a nuclear chemist works with a large team, including physicists, engineers, and technicians. Heavy elements are produced by collisions in a particle accelerator. The nuclear chemist analyzes the data from these collisions to identify the elements and understand their properties. For more information on chemistry careers, visit glencoe.com.

1940 Synthesized elements with an atomic number larger than 92 become part of a new block of the periodic table called the actinides.

1985 The International Union of Pure and Applied Chemistry adopts the form of the periodic table currently used by scientists worldwide.



2004 Scientists in Russia report the discovery of elements 113 and 115.

1950 1965 1980 1995 2010

1969 Researchers at the University of Berkeley synthesize the first element heavier than the actinides. It has a half-life of 4.7 seconds and is named rutherfordium.

1999 Researchers report the discovery of element 114, ununquadium. Scientists believe this element might be the first of a series of relatively stable synthetic elements.

Concepts In Motion

Interactive Time Line To learn more about these discoveries and others, visit glencoe.com.

Chemistry Online

EXAMPLE Problem 6.1

Electron Configuration and the Periodic Table Strontium, which is used to produce red fireworks, has an electron configuration of $[Kr]5s^2$. Without using the periodic table, determine the group, period, and block of strontium.

1 Analyze the Problem

You are given the electron configuration of strontium.

Known

Electron configuration = $[Kr]5s^2$

Unknown

Group = **?**

Period = **?**

Block = **?**

2 Solve for the Unknown

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

For representative elements, the number of valence electrons can indicate the group number.

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

The number of the highest energy level indicates the period number.

3 Evaluate the Answer

The relationships among electron configuration and position on the periodic table have been correctly applied.

PRACTICE Problems

Extra Practice Page 979 and glencoe.com

8. 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.** $[Kr]5s^2 4d^{10} 5p^5$

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

10. Challenge Write the electron configuration of the following elements.

a. the group 2 element in the fourth period

b. the group 12 element in the fourth period

c. the noble gas in the fifth period

d. the group 16 element in the second period

Section 6.2 Assessment

Section Summary

- The periodic table has four blocks (s, p, d, f).
- Elements within a group have similar chemical properties.
- The group number for elements in groups 1 and 2 equals the element's number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.

11. MAIN Idea **Explain** what determines the blocks in the periodic table.

12. Determine in which block of the periodic table are the elements having the following valence electron configurations.

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

13. Infer Xenon, a nonreactive gas used in strobe lights, is a poor conductor of heat and electricity. Would you expect xenon to be a metal, a nonmetal, or a metalloid? Where would you expect it to be on the periodic table? Explain.

14. Explain why elements within a group have similar chemical properties.

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

Chemistry Online Self-Check Quiz glencoe.com

Section 6.3

Objectives

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

Review Vocabulary

principal energy level: the major energy level of an atom

New Vocabulary

ion
ionization energy
octet rule
electronegativity

Periodic Trends

MAIN Idea Trends among elements in the periodic table include their size and their ability to lose or attract electrons.

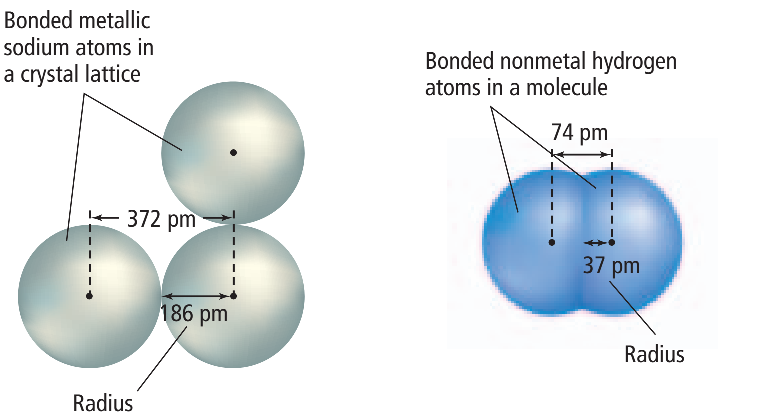
Real-World Reading Link A calendar is a useful tool for keeping track of activities. The pattern of days, from Sunday to Saturday, is repeated week after week. If you list an activity many weeks ahead, you can tell from the day of the week what else might happen on that day. In much the same way, the organization of the periodic table tells us about the behavior of many of the elements.

Atomic Radius

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. Atomic size is a periodic trend influenced by electron configuration. Recall from Chapter 5, the electron cloud surrounding a nucleus does not have a clearly defined edge. 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 as shown in **Figure 6.10.** For elements that commonly occur as molecules, such as many nonmetals, the 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 6.10.**

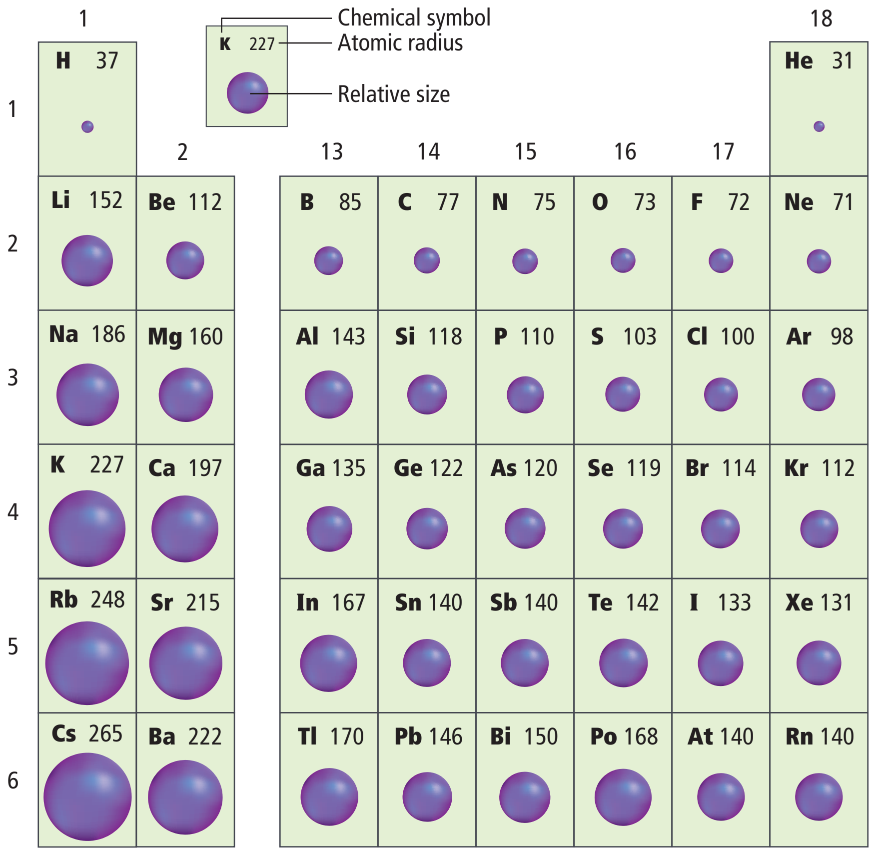
Figure 6.10 Atomic radii depend on the type of bonds that atoms form.



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

The radius of a nonmetal atom is often determined from a molecule of two identical atoms.

■ **Figure 6.11** The atomic radii of the representative elements, given in picometers (10^{-12} m), vary as you move from left to right within a period and down a group.
Infer *why the atomic radii increase as you move down a group.*



Concepts In Motion

Interactive Figure To see an animation of the trends in atomic radii, visit glencoe.com.

FOLDABLES Incorporate information from this section into your Foldable.

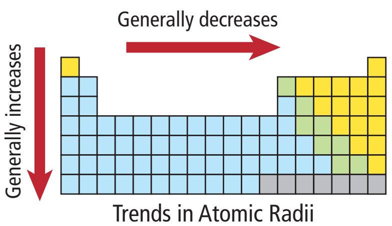
Trends within periods In general, there is a decrease in atomic radii as you move from left to right across a period. This trend, shown in **Figure 6.11,** is caused by the increasing positive charge in the nucleus and the fact that the principal energy level within a period remains the same. Each successive element has one additional proton and electron, and each additional electron is added to orbitals corresponding 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, which pulls the outermost electrons closer to the nucleus.

Reading Check **Discuss** how the fact that the principal energy level remains the same within a period explains the decrease in the atomic radii across a period.

Trends within groups Atomic radii generally increase as you move down a group. The nuclear charge increases, and electrons are added to orbitals corresponding to successively higher principal energy levels. However, the increased nuclear charge does not pull the outer electrons toward the nucleus to make the atom smaller.

Moving down a group, the outermost orbital increases in size along with the increasing principal energy level; thus, the atom becomes larger. The larger orbital means that the outer electrons are farther from the nucleus. This increased distance offsets the 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 nucleus. **Figure 6.12** summarizes the group and period trends.

■ **Figure 6.12** Atomic radii generally decrease from left to right in a period and generally increase as you move down a group.



EXAMPLE Problem 6.2

Interpret Trends in Atomic Radii Which has the largest atomic radius: carbon (C), fluorine (F), beryllium (Be), or lithium (Li)? Answer without referring to **Figure 6.10.** 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. **Determine the periods.**

Ordering the elements from left-to-right across the period yields: Li, Be, C, and F.

The first element in period 2, lithium, has the largest radius. **Apply the trend of decreasing radii across a period.**

3 Evaluate the Answer

The period trend in atomic radii has been correctly applied. Checking radii values in **Figure 6.10** verifies the answer.

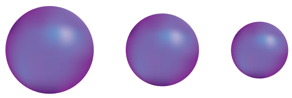
PRACTICE Problems

Extra Practice Page 979 and glencoe.com

Answer the following questions using your knowledge of group and period trends in atomic radii. Do not use the atomic radii values in Figure 6.10 to answer the questions.

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

17. The figure on the right shows helium, krypton, and radon. Which one is krypton? How can you tell?



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? Explain.

19. Challenge Determine which element in each pair has the largest atomic radius:

a. the element in period 2, group 1; or the element in period 3, group 18

b. the element in period 5, group 2; or the element in period 3, group 16

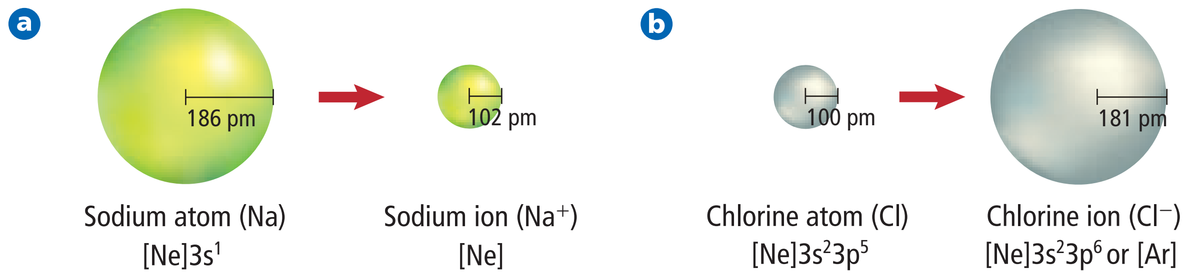
c. the element in period 3, group 14; or the element in period 6, group 15

d. the element in period 4, group 18; or the element in period 2, group 16

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 will learn about ions in Chapter 7, but for now, consider 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. The reason for the decrease in size is twofold. The electron lost from the atom will almost always be a valence electron. The loss of a valence electron can leave a completely empty outer orbital, which results in a smaller radius. Furthermore, the electrostatic repulsion between the now-fewer number of remaining electrons and the positively charged nucleus decreases, allowing the electrons to be pulled closer to the nucleus.



■ **Figure 6.13** The size of atoms varies greatly when they form ions.
a. Positive ions are smaller than the neutral atoms from which they form.
b. Negative ions are larger than the neutral atoms from which they form.

FOLDABLES Incorporate information from this section into your Foldable.

When atoms gain electrons and form negatively charged ions, they become larger. The addition of an electron to an 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.

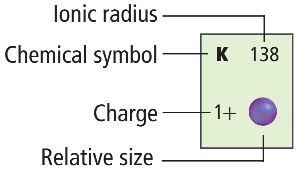
Figure 6.13a illustrates how the radius of sodium decreases when sodium atoms form positive ions, and **Figure 6.13b** shows how the radius of chlorine increases when chlorine atoms form negative ions.

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 from left to right across a period, the size of the positive ions gradually decreases. Then, beginning in group 15 or 16, the size of the much-larger negative ions also gradually decreases.

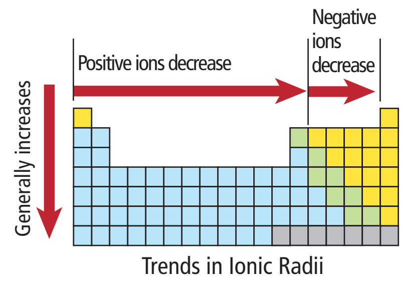
■ **Figure 6.14** The ionic radii of most of the representative elements are shown in picometers (10^{-12} m).

Explain ***why the ionic radii increase for both positive and negative ions as you move down a group.***

Period	1	2	13	14	15	16	17
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+		



Trends within groups As you move down a group, an ion's outer electrons are in orbitals corresponding to 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. The group and period trends in ionic radii are summarized in **Figure 6.15.**



■ **Figure 6.15** The diagram summarizes the general trends in ionic radii.

Ionization Energy

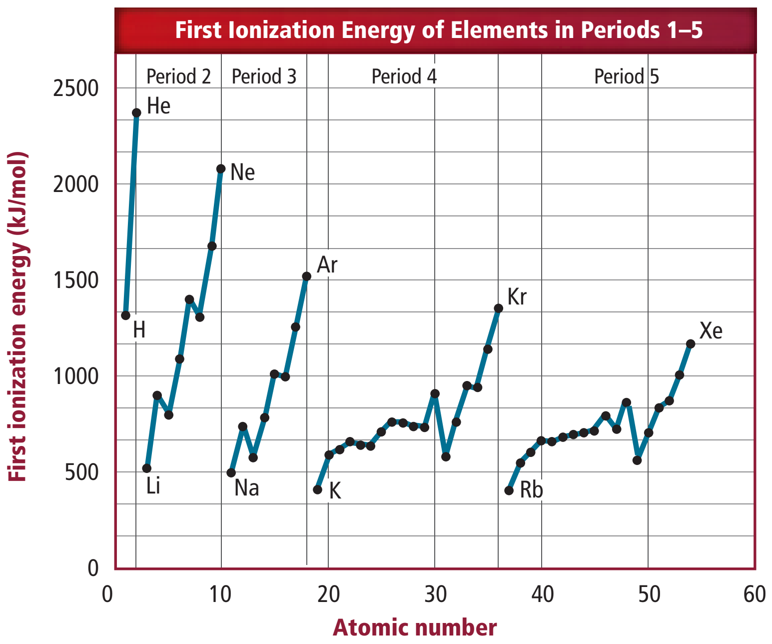
To form a positive ion, an electron must be removed from a neutral atom. This requires energy. The energy is needed to overcome the attraction between the positive charge of the nucleus and the negative charge of the electron. **Ionization energy** is defined as the energy required to remove an electron 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.**

Reading Check **Define** *ionization energy.*

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. Lithium's low ionization energy, for example, is important for its use in lithium-ion computer backup batteries where the ability to lose electrons easily makes a battery that can quickly provide a large amount of electrical power.

ChemistryOnline

Personal Tutor For an online tutorial on identifying trends, visit glencoe.com.



■ **Figure 6.16** The first ionization energies for elements in periods 1 through 5 are shown as a function of the atomic number.

Graph Check

Describe the trend in first ionization energies within a group.

Table 6.5 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.

Each set of connected points on the graph in **Figure 6.16** represents the elements in a period. The group 1 metals have low ionization energies. Thus, group 1 metals (Li, Na, K, Rb) are likely to form positive ions. The group 18 elements (He, Ne, Ar, Kr, Xe) have high ionization energies and are unlikely to form ions. The stable electron configuration of gases of group 18 greatly limits their reactivity.

Removing more than one electron 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 6.5** lists the-first-through ninth ionization energies for elements in period 2.

Reading across **Table 6.5** from left to right, you will 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 increases 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 that a lithium atom is likely to lose its first valence electron but extremely unlikely to lose its second.

Reading Check **Infer** how many electrons carbon is likely to lose.

If you examine the table, you will notice that the ionization at which the large increase in energy occurs is related to the atom's number of valence electrons. Lithium has one valence electron and the increase occurs after the first ionization energy. Lithium easily forms the common lithium 1+ ion but is unlikely to form a lithium 2+ ion. The increase in ionization energy shows that atoms hold onto their inner core electrons much more strongly than they hold onto their valence electrons.

Real-World Chemistry
Ionization Energy

Scuba diving The increased pressure that scuba divers experience far below the water's surface can cause too much oxygen to enter their blood, which would result in confusion and nausea. To avoid this, divers sometimes use a gas mixture called *heliox*—oxygen diluted with helium. Helium's high ionization energy ensures that it will not react chemically in the bloodstream.

Trends within periods As shown in **Figure 6.16** and by the values in **Table 6.5,** first ionization energies generally increase as you move from 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. Less energy is required to remove the valence electrons farther from the nucleus. **Figure 6.17** summarizes the group and period 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.

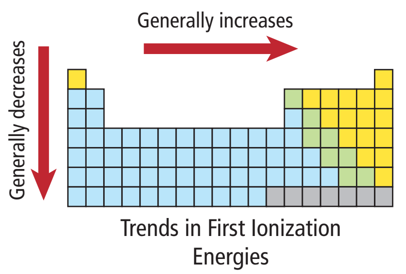


Figure 6.17 Ionization energies generally increase from left to right in a period and generally decrease as you move down a group.

FOLDABLES Incorporate information from this section into your Foldable.

MiniLab

Organize Elements

Can you find the pattern?

Procedure

1. Read and complete the lab safety form.
2. Make a set of element cards based on the information in the chart at right.
3. Organize the cards by increasing mass, and start placing them into a 4 × 3 grid.
4. Place each card based on its properties, and leave gaps when necessary.

Symbol	Mass (g)	State	Color
Ad	52.9	solid/liquid	orange
Ax	108.7	ductile solid	light blue
Bp	69.3	gas	red
Cx	112.0	brittle solid	light green
Lq	98.7	ductile solid	blue
Pd	83.4	brittle solid	green
Qa	68.2	ductile solid	dark blue
Rx	106.9	liquid	yellow
Tu	64.1	brittle solid	hunter
Xn	45.0	gas	crimson

Analysis

1. **Make a table** listing the placement of each element.
2. **Describe** the period (across) and group (down) trends for the color in your new table.
3. **Describe** the period and group trends for the mass in your new table. Explain your placement of any elements that do not fit the trends.
4. **Predict** the placement of a newly found element, Ph, that is a fuchsia gas. What would be an expected range for the mass of Ph?
5. **Predict** the properties for the element that would fill the last remaining gap in the table.

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 Ds	111 Uuu	112 Uub	113 Uut	114 Uuq	115 Uup	116 Uuh		118 Uuo

Electronegativity Values in Paulings

Concepts In Motion

Interactive Figure To see an animation of the trends in electronegativity, visit glencoe.com.

Figure 6.18 The electronegativity values for most of the elements are shown. The values are given in Paulings.

Infer *why electronegativity values are not listed for the noble gases.*

Electronegativity

The **electronegativity** of an element indicates the relative ability of its atoms to attract electrons in a chemical bond. As shown in **Figure 6.18,** electronegativity generally decreases as you move down a group, and increases as you move from left to right across a period. Electronegativity values 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). 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. Note that because the noble gases form very few compounds, they do not have an electronegativity value.

Section 6.3 Assessment

Section Summary

- Atomic and ionic radii decrease from left to right across a period, and increase as you move down a group.
- Ionization energies generally increase from 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 to acquire a full set of eight valence electrons.
- Electronegativity generally increases from left to right across a period, and decreases as you move down a group.

20. **MAIN Idea** **Explain** how the period and group trends in atomic radii are related to electron configuration.
21. **Indicate** whether fluorine or bromine has a larger value for each of the following properties.
    - **a.** electronegativity
    - **b.** ionic radius
    - **c.** atomic radius
    - **d.** ionization energy
22. **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.
23. **Calculate** Determine the differences in electronegativity, ionic radius, atomic radius, and first ionization energy for oxygen and beryllium.
24. **Make and Use Graphs** Graph the atomic radii of the representative 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.

Elements of the Body

Every time you eat a sandwich or take a breath, you are taking in elements your body needs to function normally. These elements have specific properties, depending on their location on the periodic table. **Figure 1** shows the percent by mass composition of cells in the human body.

Oxygen In an adult body, there are more than 14 billion billion billion oxygen atoms! Without a constant input of oxygen into the blood, the human body could die in just a few minutes.

Carbon Carbon can form strong bonds with itself and other elements. Carbon forms the long-chained carbon backbones that are an essential part of organic molecules such as carbohydrates, proteins, and lipids. The DNA molecule that determines your physical features relies on the versatility of carbon and its ability to bond with many different elements.

Hydrogen There are more hydrogen atoms in the body than atoms of all the other elements combined, although hydrogen represents only 10% of the composition by mass because of their significantly lower mass. The human body, requires hydrogen not in its elemental form, but in a variety of essential compounds, like water. With oxygen and carbon, hydrogen is also a crucial part of carbohydrates and other organic molecules that your body needs for energy.

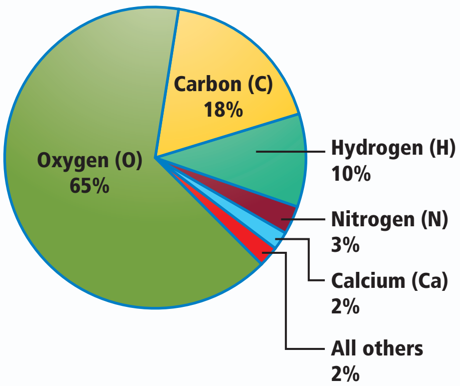


Figure 1 The human body is composed of many different elements.

Figure 2 The entire human body is covered with muscles.

Nitrogen As shown in **Figure 2,** the human body is entirely covered with muscle. Nitrogen atoms are found in compounds that make up the proteins your body needs to build muscle.

Other elements in the body Oxygen, carbon, hydrogen, and nitrogen are the most abundant elements in your body but only a few of the elements that your body needs to live and grow. Trace elements, which together make up less than 2% of the body's mass, are a critical part of your body. Your bones and teeth could not grow without the constant intake of calcium. Although sulfur comprises less than 1 percent of the human body by mass, it is an essential component and is found in the proteins in your fingernails for instance. Sodium and potassium are crucial for the transmission of electrical signals in your brain.

WRITING in Chemistry

Can you get all of the trace elements you need by eating only pre-packaged food? Why are trace elements necessary, despite the fact that they are present only in such small amounts? Discuss these issues with your classmates. For more information about elements of the body, visit glencoe.com.

CHEMLAB

INVESTIGATE DESCRIPTIVE CHEMISTRY

Background: 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.

Question: *What is the pattern of properties of the representative elements?*

Materials

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

Safety Precautions

WARNING: ***Never test chemicals by tasting. 1.0M HCl is harmful to eyes and clothing. Brittle samples might shatter into sharp pieces.***

Procedure

1. Read and complete the lab safety form.
2. Observe and record the appearance (physical state, color, luster, texture, and so on) of the element sample in each test tube without removing the stoppers.
3. Remove a small sample of each of the elements contained in a plastic dish and place it on a hard surface. Gently tap each element sample with a small hammer. If the element is malleable, it will flatten. If it is brittle, it will shatter. Record your observations.
4. Use the conductivity tester to determine which elements conduct electricity. Clean the electrodes with water, and dry them before testing each element.
5. 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.
6. Use a spatula to put a small amount of each element into the corresponding test tubes. Using a graduated cylinder, add 5 mL of 1.0*M* HCl to each test tube. Observe each tube for at least 1 minute. The formation of bubbles is evidence of a reaction between the acid and the element. Record your observations.

Observation of Elements

Classification	Properties
Metals	• malleable • good conductor of electricity • lustrous • silver or white in color • many react with acids
Nonmetals	• solids, liquids, or gases • do not conduct electricity • do not react with acids • likely brittle if solid
Metalloids	• combine properties of metals and nonmetals

7. **Cleanup and Disposal** Dispose of all materials as instructed by your teacher.

Analyze and Conclude

1. **Interpret Data** Using the table above and your observations, list the element samples that display the general characteristics of metals.
2. **Interpret Data** Using the table above and your observations, list the element samples that display the general characteristics of nonmetals.
3. **Interpret Data** Using the table above and your observations, list the element samples that display the general characteristics of metalloids.
4. **Model** Construct a periodic table, and label the representative elements by group (1 through 17). Using your results and the periodic table presented in this chapter, record the identities of elements observed during the lab in periodic table you have constructed.
5. **Infer** Describe any trends among the elements you observed in the lab.

INQUIRY EXTENSION

Investigate Were there any element samples that did not fit into one of the three categories? What additional investigations could you conduct to learn even more about these elements' characteristics?

CHAPTER 6

Study Guide

BIG Idea Periodic trends in the properties of atoms allow us to predict physical and chemical properties.

Section 6.1 Development of the Modern Periodic Table

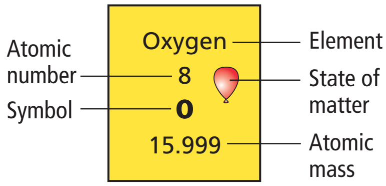
MAIN Idea The periodic table evolved over time as scientists discovered more useful ways to compare and organize the elements.

Vocabulary

actinide series (p. 180)
alkali metal (p. 177)
alkaline earth metal (p. 177)
group (p. 177)
halogen (p. 180)
inner transition metal (p. 180)
lanthanide series (p. 180)
metal (p. 177)
metalloid (p. 181)
noble gas (p. 180)
nonmetal (p. 180)
period (p. 177)
periodic law (p. 176)
representative element (p. 177)
transition element (p. 177)
transition metal (p. 180)

Key Concepts

- The elements were first organized by increasing atomic mass, which led to inconsistencies. Later, they were organized by increasing atomic number.
- The periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- The periodic table organizes the elements into periods (rows) and groups (columns); elements with similar properties are in the same group.
- Elements are classified as either metals, nonmetals, or metalloids.



Section 6.2 Classification of the Elements

MAIN Idea Elements are organized into different blocks in the periodic table according to their electron configurations.

Key Concepts

- The periodic table has four blocks (s, p, d, f).
- Elements within a group have similar chemical properties.
- The group number for elements in groups 1 and 2 equals the element's number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.

Section 6.3 Periodic Trends

MAIN Idea Trends among elements in the periodic table include their size and their ability to lose or attract electrons.

Vocabulary

electronegativity (p. 194)
ion (p. 189)
ionization energy (p. 191)
octet rule (p. 193)

Key Concepts

- Atomic and ionic radii decrease from left to right across a period, and increase as you move down a group.
- Ionization energies generally increase from 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 to acquire a full set of eight valence electrons.
- Electronegativity generally increases from left to right across a period, and decreases as you move down a group.

CHAPTER 6 Assessment

Section 6.1

Mastering Concepts

25. Explain how Mendeleev's periodic table was in error.

26. Explain the contribution of Newlands's law of octaves to the development of the modern periodic table.

27. Lothar Meyer and Dmitri Mendeleev both proposed similar periodic tables in 1869. Why is Mendeleev generally given credit for the periodic table?

28. What is the periodic law?

29. Describe the general characteristics of metals.

30. What are the general properties of a metalloid?

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

a. oxygen
b. barium
c. germanium
d. iron

32. Match each item on the left with its corresponding group on the right.

a. alkali metals — **1.** group 18
b. halogens — **2.** group 1
c. alkaline earth metals — **3.** group 2
d. noble gases — **4.** group 17

33. 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.

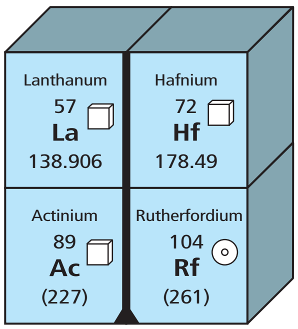


Figure 6.19

34. Explain what the dark line running down the middle of **Figure 6.19** indicates.

35. Give the chemical symbol of each of the following elements.

a. a metal used in thermometers
b. a radioactive gas used to predict earthquakes; the noble gas with the greatest atomic mass
c. a coating for food cans; it is the metal in group 14 with the lowest atomic mass
d. an inner transition metal that is used to make burglar-proof vaults; also the name of a coin

36. If a new halogen and a new noble gas were discovered, what would be their atomic numbers?

Mastering Problems

37. If the periodic table were arranged by atomic mass, which of the first 55 elements would be ordered differently than they are in the existing table?

38. **New Heavy Element** If scientists discovered an element with 117 protons, what would be the its group and period? Would it be a metal, a metalloid, or a nonmetal?

39. **Naming New Elements** Recently discovered elements that have not been fully verified are given temporary names using the prefixe words in **Table 6.6.** Based on this system, write names for elements 117 to 120.

Table 6.6 Prefixes

0	1	2	3	4
nil	un	b(i)	tr(i)	quad
5	6	7	8	9
pent	hex	sept	oct	en(n)

40. Give the chemical symbol for each element.

a. the element in period 3 that can be used in making computer chips because it is a metalloid
b. the group 13, period 5 metal used in making flat screens for televisions
c. an element used as a filament in lightbulbs; has the highest atomic mass natural elements in group 6

Section 6.2

Mastering Concepts

41. **Household Products** Why do the elements chlorine, used in laundry bleach, and iodine, a nutrient added to table salt, have similar chemical properties?

42. How is the energy level of an atom's valence electrons related to its period in the periodic table?

43. How many valence electrons does each noble gas have?

44. What are the four blocks of the periodic table?

45. What electron configuration has the greatest stability?

46. Explain how an atom's valence electron configuration determines its place in the periodic table.

47. Write the electron configuration for the element fitting each of the following descriptions.

a. the metal in group 15 that is part of compounds often found in cosmetics
b. the halogen in period 3 that is part of a bleaching compound used in paper production
c. the transition metal that is a liquid at room temperature; is sometimes used in outdoor security lights

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48. Determine the group, period, and block in which each of the following elements is located in the periodic table.
   a. $[Kr]5s^2 4d^1$
   b. $[Ar]4s^2 3d^{10} 4p^3$
   c. $[He]2s^2 2p^6$
   d. $[Ne]3s^2 3p^1$
49. 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?
50. **Table 6.7** shows the number of elements in the first five periods of the periodic table. Explain why some of the periods have different numbers of elements.

Table 6.7 Number of Elements in Periods 1–5

Period	1	2	3	4	5
Number of elements	2	8	8	18	18

51. **Coins** One of the transition groups is often called the coinage group because at one time many coins are made of these metals. Which group is this? What elements in this group is still used in many U.S. coins today?
52. Do any of the halogens have their valence electrons in orbitals of the same energy level? Explain.
53. The transition elements have their valence electrons in orbitals of more than one energy level, but the representative elements have their valence electrons in orbitals of only one energy level. Show this by using the electron configurations of a transition element and a representative element as examples.

Mastering Problems

54. **Fireworks** Barium is a metal that gives a green color to fireworks. Write the electron configuration for barium. Classify it according to group, period, and block in the periodic table.
55. **Headphones** Neodymium magnets can be used in stereo headphones because they are powerful and lightweight. Write the electron configuration for neodymium. In which block of the periodic table is it?
56. **Soda Cans** The metal used to make soda cans has the electron configuration $[Ne]3s^2 3p^1$. Identify the metal and give its group, period, and block.
57. Identify each missing part of **Table 6.8.**

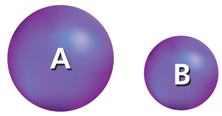
Table 6.8 Electron Configuration

Period	Group	Element	Electron Configuration
3		Mg	$[Ne]3s^2$
4	14	Ge	
	12	Cd	$[Kr]5s^2 4d^{10}$
2	1		$[He]2s^1$

Section 6.3

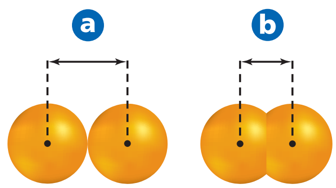
Mastering Concepts

58. What is ionization energy?
59. An element forms a negative ion when ionized. On what side of the periodic table is the element located? Explain.
60. Of the elements magnesium, calcium, and barium, which forms the ion with the largest radius? The smallest? What periodic trend explains this?
61. Explain why each successive ionization of an electron requires a greater amount of energy.
62. How does the ionic radius of a nonmetal compare with its atomic radius? Explain the change in radius.
63. Explain why atomic radii decrease as you move from left to right across a period.
64. Which element has the larger ionization energy?
   a. Li, N
   b. Kr, Ne
   c. Cs, Li
65. Explain the octet rule. Why are hydrogen and helium exceptions to the octet rule?



■ Figure 6.20

66. Use **Figure 6.20** to answer each of the following questions. Explain your reasoning for each answer.
   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 order?
   c. If A and B represent the ionic radii of two elements in the same group, what is their order?
67. How many valence electrons do elements in group 1 have? In group 18?



■ Figure 6.21

68. **Figure 6.21** shows two ways to define an atomic radius. Describe each method. When is each method used?
69. **Chlorine** The electron configuration of a chlorine atom is $[Ne]3s^2 3p^5$. When it gains an electron and becomes an ion, its electron configuration changes to $[Ne]3s^2 3p^6$, or [Ar], the electron configuration for argon. Has the chlorine atom changed to an argon atom? Explain.

Mastering Problems

70. Sport Bottles Some sports bottles are made of Lexan, a plastic containing a compound of the elements chlorine, carbon, and oxygen. Order these elements from greatest to least according to atomic radius and ionic radius.

71. Contact Lenses Soft contact lenses are made of silicon and oxygen atoms bonded together. Create a table listing the atomic and ionic electron configurations, and the atomic and ionic radii for silicon and oxygen. When silicon bonds with oxygen, which atoms become larger and which become smaller? Why?

72. Artificial Sweetener Some diet sodas contain the artificial sweetener aspartame, a compound containing carbon, nitrogen, oxygen, and other atoms. Create a table showing the atomic and ionic radii of carbon, nitrogen, and oxygen. (Assume the ionization states shown in **Figure 6.14.**) Use the table to predict whether the sizes of carbon, nitrogen, and oxygen atoms increase or decrease in size when they form bonds in aspartame.

Mixed Review

73. Define an ion.

74. Explain why the radius of an atom cannot be measured directly.

75. What is the metalloid in period 2 of the periodic table that is part of compounds used as water softeners?

76. Do you expect cesium, a group 1 element used in infrared lamps, or bromine, a halogen used in firefighting compounds to have the greatest electronegativity? Why?

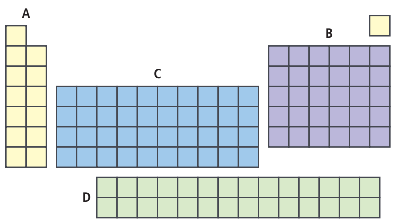


Figure 6.22

77. Figure 6.22 shows different sections of the periodic table. Give the name of each section, and explain what the elements in each section have in common.

78. Which element in each pair is more electronegative?
a. K, As **b.** N, Sb **c.** Sr, Be

79. 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.

80. Most of the atomic masses in Mendeleev's table are different from today's values. Explain why.

81. 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?

82. Milk The element with the electron configuration $[Ar]4s^2$ is an important mineral in milk. Identify this element's group, period, and block in the periodic table.

83. Why are there no p-block elements in the first period?

84. Jewelry What are the two transition metals that are used in making jewelry and are the group 11 elements with the lowest atomic masses?

85. Which has the largest ionization energy, platinum, an element sometimes used in dental crowns, or cobalt, an element that provides a bright blue color to pottery?

Think Critically

86. Apply 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?

87. Make and Use Graphs The densities of the group 15 elements are given in **Table 6.9.** Plot density versus atomic number, and state any trends you observe.

Table 6.9 Group 15 Density Data

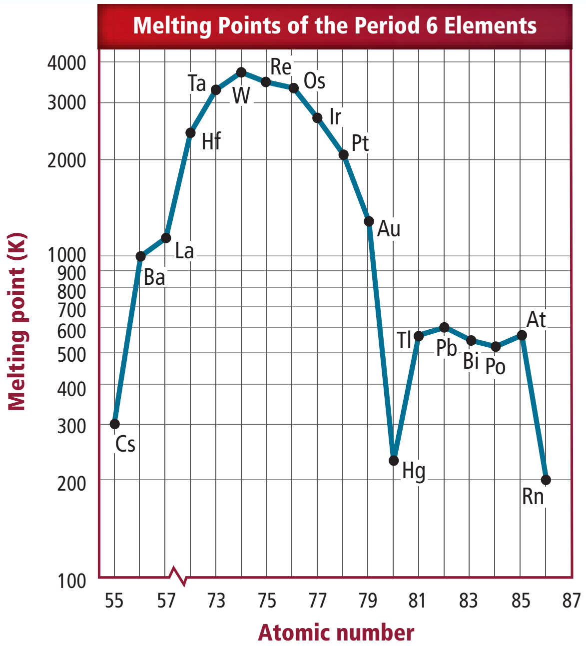
Element	Atomic Number	Density (g/cm^3)
Nitrogen	7	1.25×10^{-3}
Phosphorus	15	1.82
Arsenic	33	5.73
Antimony	51	6.70
Bismuth	83	9.78

88. Generalize The outer-electron configurations of elements in group 1 can be written as ns^1, where n refers to the element's period and its principal energy level. Develop a similar notation for all the other groups of the representative elements.

89. Identify A period 3 representative element is part of the rough material on the side of a match box used for lighting matches. **Table 6.10** shows the ionization energies for this element. Use the information in the table to infer the identity of the element. Explain.

Table 6.10 Ionization Energies in kJ/mol

Number	1st	2nd	3rd	4th	5th	6th
Ionization energy	1010	1905	2910	4957	6265	21,238



■ **Figure 6.23**

90. Interpret Data The melting points of the period 6 elements are plotted versus atomic number in **Figure 6.23.** Determine the trends in melting point and the orbital configurations of the elements. Form a hypothesis that explains the trends.

Challenge Problem

91. Ionization energies are expressed in kilojoules per mole, but the energy to remove an electron from a gaseous atom is expressed in joules. Use the values in **Table 6.6** to calculate the energy, in joules, required to remove the first electron from an atom of Li, Be, B, and C. Then, use the relationship 1 eV $= 1.60 \times 10^{-19}$ J to convert the values to electron volts.

Cumulative Review

92. 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

93. 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 μg to kilograms

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

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

Additional Assessment

WRITING in Chemistry

96. Triads 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?

97. Affinity Electron affinity is another periodic property of the elements. Write a report on what electron affinity is, and describe its group and period trends.

Document-Based Questions

Mendeleev's original periodic table is remarkable given the knowledge of elements at that time, and yet it is different from the modern version. Compare Mendeleev's table, shown in **Table 6.12,** *with the modern periodic table shown in* **Figure 6.5.**

Data obtained from: Dmitrii Mendeleev, *The Principles of Chemistry,* 1891.

Table 6.11 Groups of Elements

Series	0	I	II	III	IV	V	VI	VII	VIII
1	—	H	—	—	—		—	—	
2	He	Li	Be	B	C	N	O	F	
3	Ne	Na	Mg	Al	Si	P	S	Cl	
4	Ar	K	Ca	So	Ti	V	Cr	Mn	Fe
5		Cu	Zn	Ga	Ge	As	Se	Br	Co Ni (Cu)
6	Kr	Rb	Sr	Y	Zr	Nb	Mo	—	Ru
7		Ag	Cad	In	Sn	Sb	Te	I	Rh Pd (Ag)
8	Xe	Cs	Ba	La	—	—	—	—	—
9		—	—	—	—	—	—	—	—
10	—	—	—	Yb	—	Ta	W	—	Os
11		Au	Hg	Tl		Bi	—	—	Ir Pt (Au)
12	—	—	Rd	—	Th	—	U		

98. Mendeleev placed the noble gases on the left of his table. Why does placement on the right of the modern table make more sense?

99. Which block on Mendeleev's table was most like today's placement? Which block was least like today's placement? Why?

100. Most of the atomic masses in Mendeleev's table differ from today's values. Why do you think this is so?

Cumulative

Standardized Test Practice

Multiple Choice

1. 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.

2. Which statement is NOT true?
A. The atomic radius of Na is less than the atomic radius of Mg.
B. The electronegativity of C is greater than the electronegativity of B.
C. The ionic radius of Br^- is greater than the atomic radius of Br.
D. The first ionization energy of K is greater than the first ionization energy of Rb.

3. What is the group, period, and block of an atom with the electron configuration $[Ar]4s^23d^{10}4p^4$?
A. group 14, period 4, d-block
B. group 16, period 3, p-block
C. group 14, period 4, p-block
D. group 16, period 4, p-block

Use the table below to answer Questions 4 and 5.

Characteristics of Elements		
Element	Block	Characteristic
X	s	soft solid; reacts readily with oxygen
Y	p	gas at room temperature; forms salts
Z	—	inert gas

4. In which group does Element X most likely belong?
A. 1
B. 17
C. 18
D. 4

5. In which block is Element Z most likely found?
A. s-block
B. p-block
C. d-block
D. f-block

Use the table below to answer Questions 6 and 7.

Percent Composition By Mass of Selected Nitrogen Oxides		
Compound	Percent Nitrogen	Percent Oxygen
N_2O_4	30.4%	69.6%
N_2O_3	?	?
N_2O	63.6%	36.4%
N_2O_5	25.9%	74.1%

6. What is the percent nitrogen in the compound N_2O_3?
A. 44.75%
B. 46.7%
C. 28.1%
D. 36.8%

7. A sample of a nitrogen oxide contains 1.29 g of nitrogen and 3.71 g of oxygen. Which compound is this most likely to be?
A. N_2O_4
B. N_2O_3
C. N_2O
D. N_2O_5

8. On the modern periodic table, metalloids are found only in
A. the d-block.
B. groups 13 through 17.
C. the f-block.
D. groups 1 and 2.

9. Which group is composed entirely of nonmetals?
A. 1
B. 13
C. 15
D. 18

10. It can be predicted that element 118 would have properties similar to a(n)
A. alkali earth metal.
B. halogen.
C. metalloid.
D. noble gas.

Short Answer

11. Write the electron configuration for the element arsenic (As).

12. Write the nuclear decay equation for the beta decay of iodine-131.

13. Two students are identifying a sample of tap water. Student A says that tap water is a mixture, while Student B says that it is a compound. Which student is correct? Justify your answer.

Extended Response

Use the table below to answer Questions 14 and 15.

Successive Ionization Energies for Selected Period 2 Elements, in kJ/mol

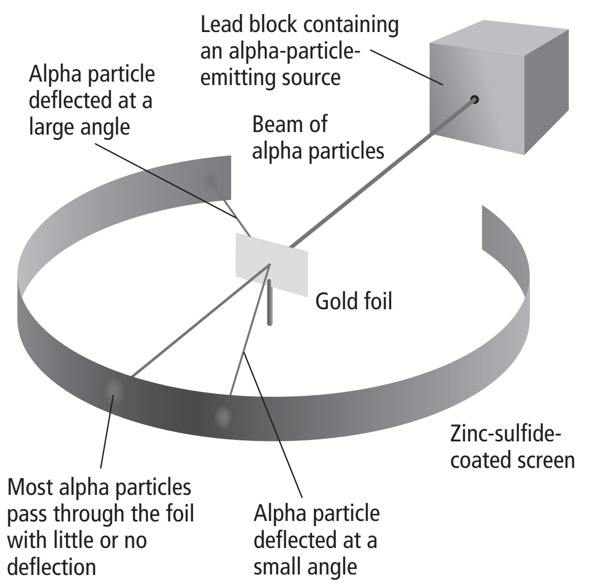
Element	Li	Be	B	C
Valence e-	1	2	3	4
First ionization energy	520	900	800	1090
Second ionization energy	7300	1760	2430	2350
Third ionization energy		14,850	3660	4620
Fourth ionization energy			25,020	6220
Fifth ionization energy				37,830

14. Correlate the biggest jump in ionization energy to the number of valence electrons in each atom.

15. Predict which ionization energy will show the largest jump for magnesium. Explain your answer.

SAT Subject Test: Chemistry

For Questions 16 to 19, answer true or false for the first statement, and true or false for the second statement. If the second statement is a correct explanation of the first statement, write CE.



Statement I		Statement II
16. Some particles bounce off the gold foil	BECAUSE	the nucleus is negatively charged.
17. Some particles bounce off the gold foil	BECAUSE	they hit protons in the nucleus.
18. Many particles pass through the gold foil	BECAUSE	atoms are made of protons, neutrons, and electrons.
19. The symbol for an alpha particle is $^{4}_{2}He$	BECAUSE	protons and neutrons have about the same mass.

NEED EXTRA HELP?																			
If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19
Review Section . . .	6.2	6.3	6.2	6.2	6.2	3.4	3.4	6.2	6.2	6.3	5.3	4.4	3.3	6.3	6.3	4.2	4.2	4.2	4.4