6 The Periodic Table

INSIDE:

- **6.1** Organizing the Elements
- **6.2** Classifying the Elements
- **6.3** Periodic Trends

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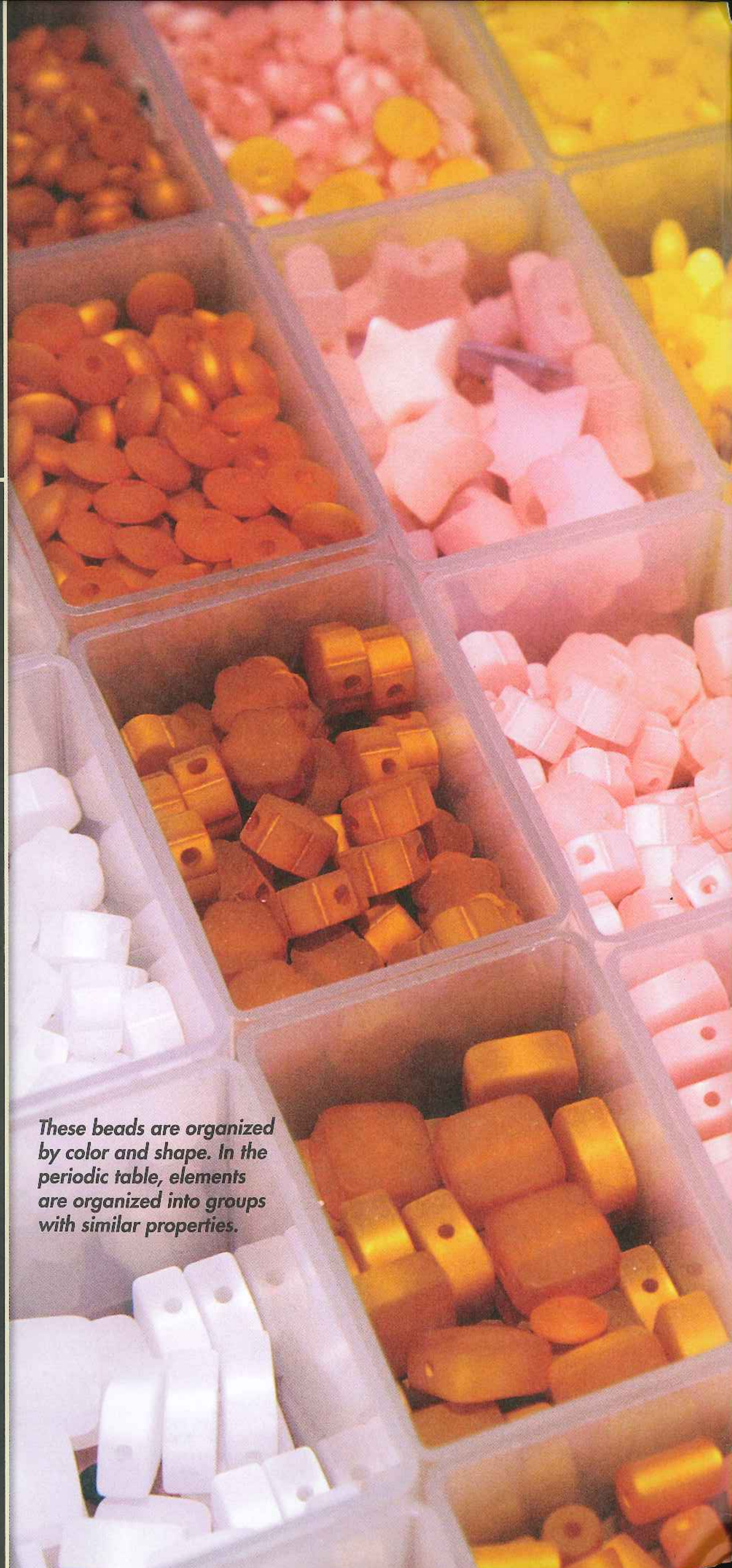
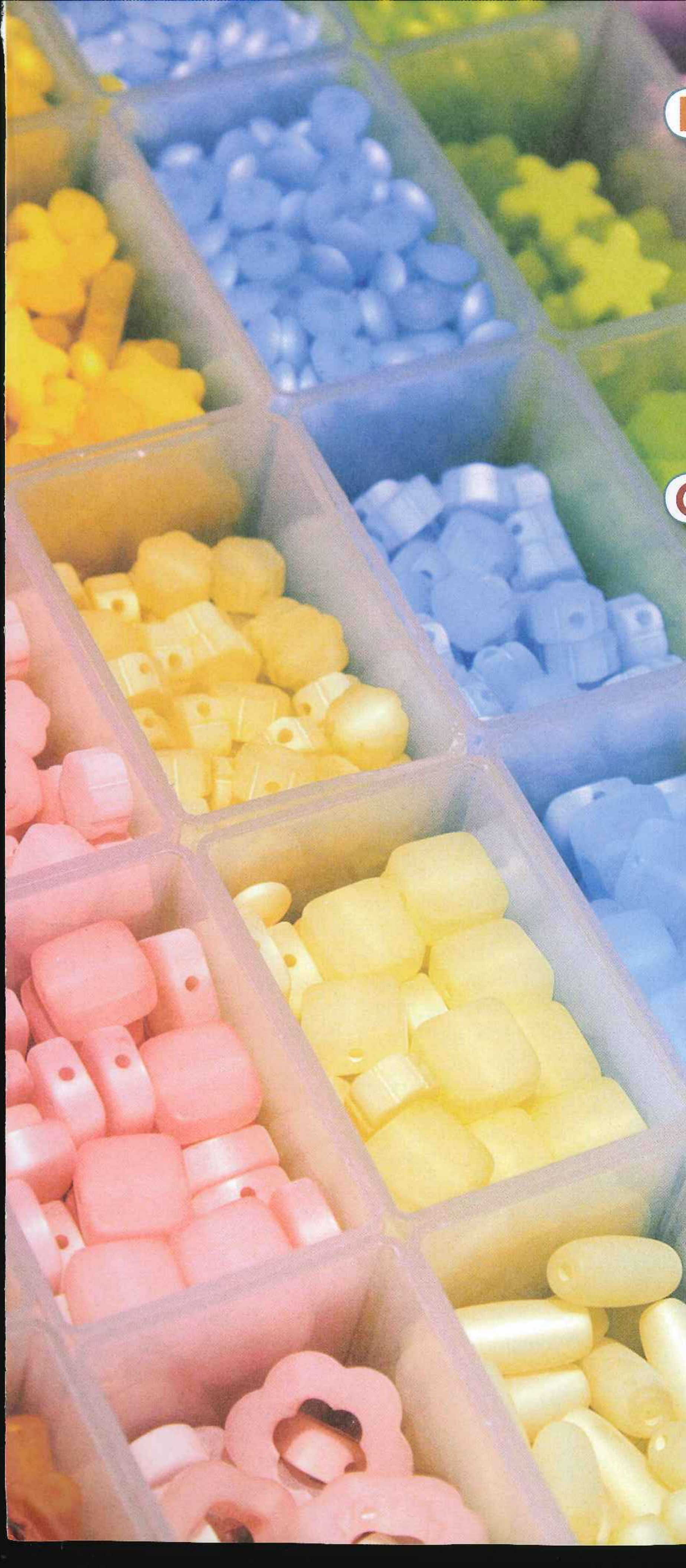


These beads are organized by color and shape. In the periodic table, elements are organized into groups with similar properties.



BIGIDEA

ELECTRONS AND THE STRUCTURE OF ATOMS

Essential Questions:

1. *What information does the periodic table provide?*
2. *How can periodic trends be explained?*

CHEMYSTERY

Made in the USA

Only 90 elements in the periodic table are known to occur naturally. You might ask then, "Why are there more than a hundred elements listed in the periodic table?"

Between the years of 1940 and 1958, nine of these "unnatural" elements were discovered by the American scientist Glenn Seaborg and his colleagues at the University of California at Berkeley. Later in 1974, Seaborg and his team discovered yet another element. Several other "unnatural" elements were discovered by Russian and German scientists.

How do these elements compare to other elements in the periodic table?

▶ Connect to the **BIG**IDEA As you read about the periodic table, keep an eye out for elements that do not occur naturally.

NATIONAL SCIENCE EDUCATION STANDARDS

A-1, A-2, B-1, B-2, E-2, G-1, G-2, G-3

6.1 Organizing the Elements

CHEMISTRY & YOU

Q: *How can you organize and classify elements?* If you have ever played a card game, then you have probably organized your cards. Maybe you classified them by color or number. Elements can also be classified. In this lesson, you will learn how elements are arranged in the periodic table and what that arrangement reveals about the elements.

Key Questions

How did chemists begin to organize the known elements?

How did Mendeleev organize his periodic table?

How is the modern periodic table organized?

What are three broad classes of elements?

Vocabulary

- periodic law
- metal
- nonmetal
- metalloid

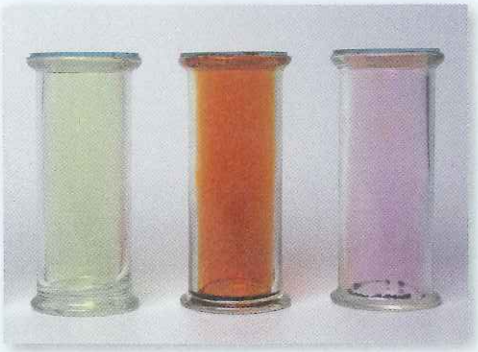


Figure 6.1 Triad in Dobereiner's System
Chlorine, bromine, and iodine formed one triad. These elements have similar chemical properties.

Searching for an Organizing Principle

How did chemists begin to organize the known elements?

A few elements, including copper, silver, and gold, have been known for thousands of years. Yet, there were only 13 elements identified by the year 1700. Chemists suspected that other elements existed. They had even assigned names to some of these elements, but they were unable to isolate the elements from their compounds. As chemists began to use scientific methods to search for elements, the rate of discovery increased. In one decade (1765–1775), chemists identified five new elements, including three colorless gases, hydrogen, nitrogen, and oxygen. Was there a limit to the number of elements? How would chemists know when they had discovered all the elements? To begin to answer these questions, chemists needed to find a logical way to organize the elements.

Early chemists used the properties of elements to sort them into groups. In 1829, a German chemist, J. W. Dobereiner (1780–1849), published a classification system. In his system, the known elements were grouped into triads. A triad is a set of three elements with similar properties. The elements in Figure 6.1 formed one triad. Chlorine, bromine, and iodine may look different, but they have very similar chemical properties. For example, they react easily with metals. Unfortunately, all the known elements could not be grouped into triads.

Dobereiner noted a pattern in his triads. One element in each triad tended to have properties with values that fell midway between those of the other two elements. For example, the average of the atomic masses of chlorine and iodine is [(35.453 1 + 126.90)/2] or 81.18 amu. This value is close to the atomic mass of bromine, which is 79.904 amu.

но въ ней, мнѣ кажется, уже ясно выражается примѣнимость выставляемаго мною начала ко всей совокупности элементовъ, пай которыхъ извѣстенъ съ достовѣрностію. На этотъ разъ я и желалъ преимущественно найдти общую систему элементовъ. Вотъ этотъ опытъ:

			Ti=50	Zr=90	?=180.
			V=51	Nb=94	Ta=182.
			Cr=52	Mo=96	W=186.
			Mn=55	Rh=104,4	Pt=197,4
			Fe=56	Ru=104,4	Ir=198.
			Ni=Co=59	Pl=106,6	Os=199.
H=1			Cu=63,4	Ag=108	Hg=200.
	Be=9,4	Mg=24	Zn=65,2	Cd=112	
	B=11	Al=27,4	?=68	Ur=116	Au=197?
	C=12	Si=28	?=70	Sn=118	
	N=14	P=31	As=75	Sb=122	Bi=210
	O=16	S=32	Se=79,4	Te=128?	
	F=19	Cl=35,5	Br=80	I=127	
Li=7	Na=23	K=39	Rb=85,4	Cs=133	Tl=204
		Ca=40	Sr=87,6	Ba=137	Pb=207.
		?=45	Ce=92		
		?Er=56	La=94		
		?Yt=60	Di=95		
		?In=75,6	Th=118?		

Figure 6.2
Mendeleev's Periodic Table
In this early version of the periodic table, Mendeleev (shown on the stamp below) arranged elements with similar properties in the same row.
Identify ***Which element is grouped with chlorine (Cl), bromine (Br), and (I) iodine?***

Mendeleev's Periodic Table

How did Mendeleev organize his periodic table?

From 1829 to 1869, different systems for organizing the elements were proposed, but none of them gained wide acceptance. In 1869, a Russian chemist and teacher, Dmitri Mendeleev, published a table of the elements. Later that year, a German chemist, Lothar Meyer, published a nearly identical table. Mendeleev was given more credit than Meyer because he published his table first and because he was better able to explain its usefulness.

Mendeleev developed his table while working on a textbook for his students. He needed a way to show the relationships among more than 60 elements. He wrote the properties of each element on a separate note card. This approach allowed him to move the cards around until he found an organization that worked. The organization he chose was a periodic table. Elements in a periodic table are arranged into groups based on a set of repeating properties. **Mendeleev arranged the elements in his periodic table in order of increasing atomic mass.**

Figure 6.2 is an early version of Mendeleev's periodic table. Look at the column that starts with Ti = 50. Notice the two question marks between the entries for zinc (Zn) and arsenic (As). Mendeleev left these spaces in his table because he knew that bromine belonged with chlorine and iodine. He predicted that elements would be discovered to fill those spaces, and he predicted what their properties would be based on their locations in the table. The elements between zinc and arsenic were gallium and germanium, which were discovered in 1875 and 1886, respectively. There was a close match between the predicted properties and the actual properties of these elements. This match helped convince scientists that Mendeleev's periodic table was a powerful tool.

See more about organizing information. CONCEPTS IN ACTION

1	1 H																															2 He
2	3 Li	4 Be																									5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg																									13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca															21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr															39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup	116 Uuh		118 Uuo

Figure 6.3 Modern Periodic Table
In the modern periodic table, the elements are arranged in order of increasing atomic number.
Interpret Diagrams ***How many elements are in the second period?***

Today's Periodic Table

How is the modern periodic table organized?

The atomic mass of iodine (I) is 126.90. The atomic mass of tellurium (Te) is 127.60. In a periodic table based on atomic mass, iodine should come before tellurium since iodine has a smaller atomic mass than tellurium does. However, based on its chemical properties, iodine belongs in a group with bromine and chlorine. Mendeleev broke his rule about placing elements in strict order of atomic mass and placed tellurium before iodine in his periodic table. He assumed that the atomic masses for iodine and tellurium were incorrect, but they were not. A similar problem occurred with other pairs of elements. The problem wasn't with the atomic masses but with using atomic mass to organize the periodic table.

Mendeleev developed his table before scientists knew about the structure of atoms. He didn't know that the atoms of each element contain a unique number of protons. Recall that the number of protons is the atomic number. In 1913, a British physicist, Henry Moseley, determined an atomic number for each known element. Tellurium's atomic number is 52 and iodine's is 53, so it makes sense for iodine to come after tellurium in the periodic table. **In the modern periodic table, elements are arranged in order of increasing atomic number.**

The elements in Figure 6.3 are arranged in order of atomic number, starting with hydrogen, which has atomic number 1. There are seven rows, or periods, in the table. Period 1 has 2 elements, Period 2 has 8 elements, Period 4 has 18 elements, and Period 6 has 32 elements. Each period corresponds to a principal energy level. There are more elements in higher numbered periods because there are more orbitals in higher energy levels. (Recall the rules you studied in Chapter 5 for how electrons fill orbitals.)

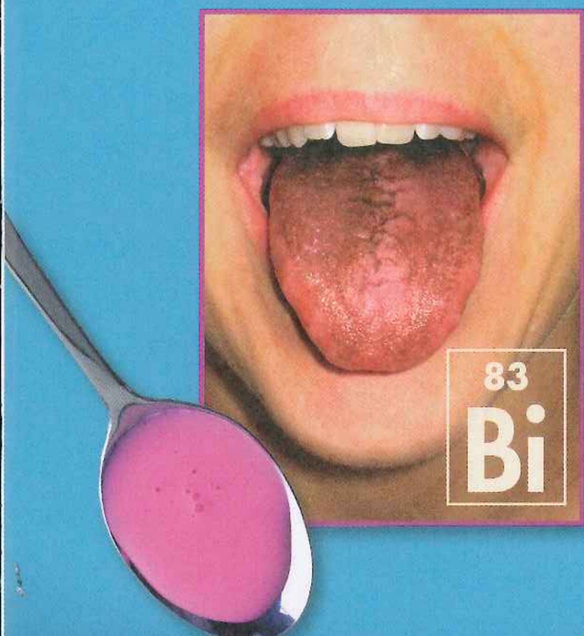
The properties of the elements within a period change as you move across a period from left to right. However, the pattern of properties within a period repeats as you move from one period to the next. This pattern gives rise to the **periodic law:** When elements are arranged in order of increasing atomic number, there is a periodic repetition of their physical and chemical properties. The arrangement of the elements into periods has an important consequence. Elements that have similar chemical and physical properties end up in the same column in the periodic table.

READING SUPPORT

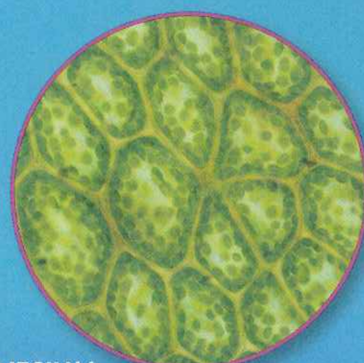
Build Vocabulary: *Word Origins*
Periodic comes from the Greek roots *peri,* meaning "around" and *hodos,* meaning "path." In a periodic table, properties repeat from left to right across each period. ***If the Greek word* metron *means "measure," what does* perimeter *mean?***

"Elemental" Trivia

Did you know that the stench of a skunk's spray is largely due to compounds that contain the element sulfur, or that rubies are red because of small amounts of chromium? Discover more fun facts about other elements as you read this page.



BISMUTH A compound containing bismuth is commonly used to treat indigestion. The bismuth can combine with sulfur in saliva and temporarily turn a person's tongue black!



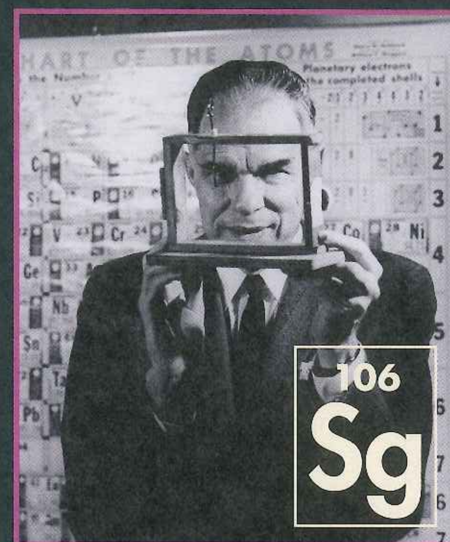
12
Mg

MAGNESIUM Magnesium is a component of chlorophyll, the green pigment in plants that allows photosynthesis to occur.

1
H

HYDROGEN Hydrogen is the most abundant element in the universe and, among other places, is found in stars and auroras.

ANTIMONY The element antimony is often used to increase the hardness and strength of pewter figurines.



SEABORGIUM In 1974, this element was created by a team of scientists that included Glenn T. Seaborg. It was the first element to be named after a living person.

Take It Further

1. Explain In elemental form, antimony (Sb) and bismuth (Bi) are both brittle, crystalline solids at room temperature. They are also poor conductors of heat and electricity. How does the periodic law support this observation? Use Figure 6.3 to explain your answer.

2. Classify Look ahead to Figure 6.4. Use the figure to classify the five elements above as metals, nonmetals, or metalloids.

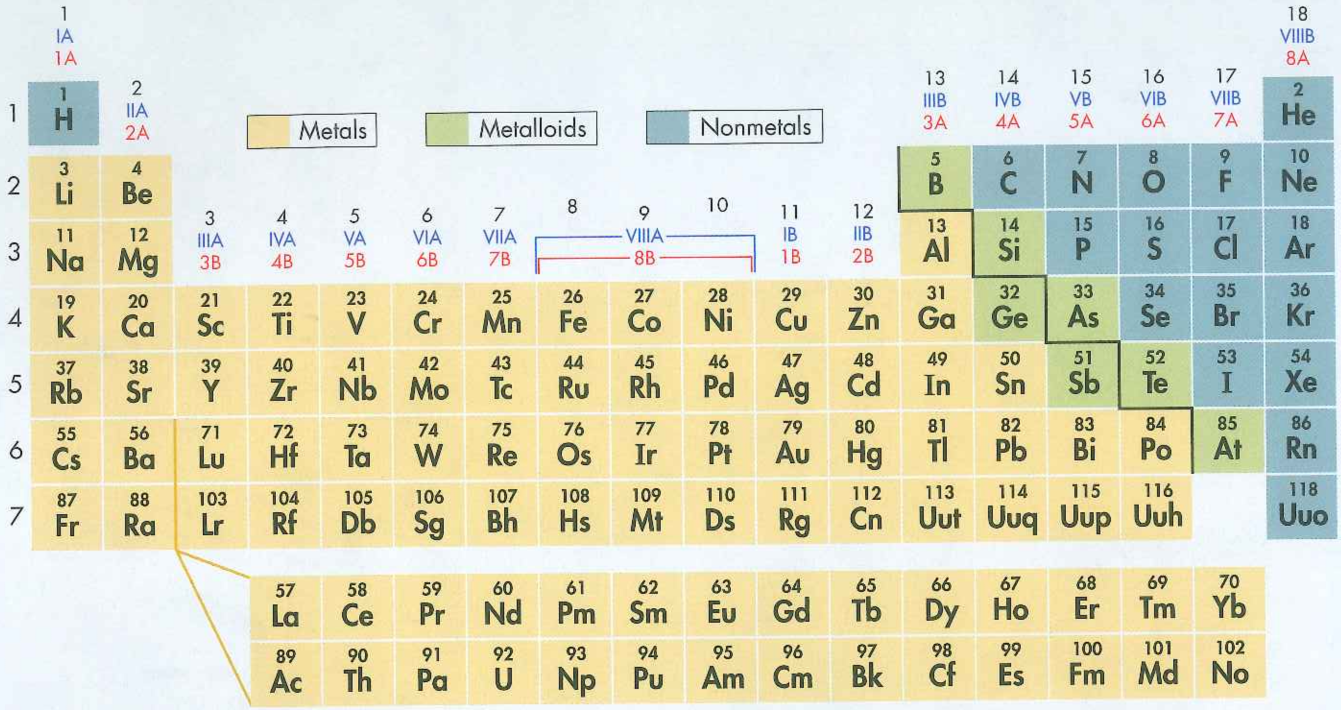


Figure 6.4 Classifying Elements
Periodic tables are sometimes color-coded to classify certain types of elements. This periodic table classifies elements as metals (yellow), nonmetals (blue), and metalloids (green).

Metals, Nonmetals, and Metalloids

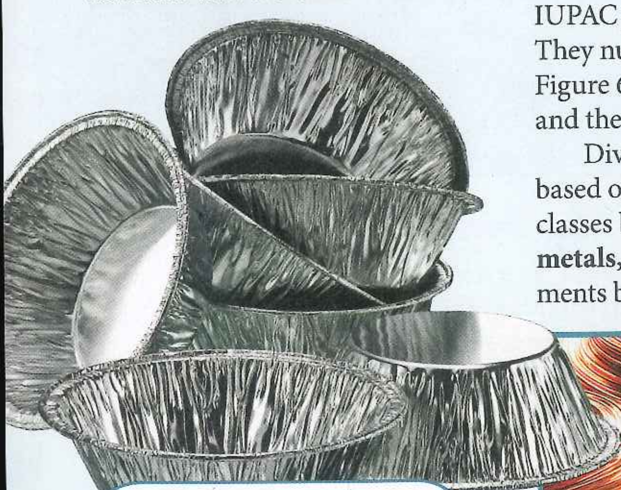
What are three broad classes of elements?

Most periodic tables are laid out like the one in Figure 6.4. Notice some elements from Periods 6 and 7 are placed beneath the table. This arrangement makes the periodic table more compact. It also reflects an underlying structure of the periodic table, which you will study in Lesson 6.2. Each column, or group, in this table has three labels. Scientists in the United States primarily use the labels shown in red. Scientists in Europe use the labels shown in blue.

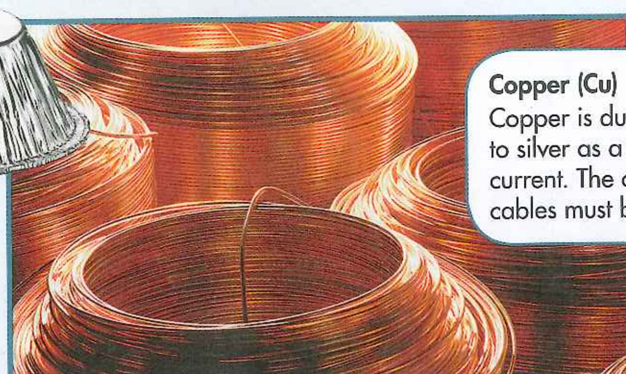
For scientists to communicate clearly, they need to agree on the standards they will use. The International Union of Pure and Applied Chemistry (IUPAC) is an organization that sets standards for chemistry. In 1985, IUPAC proposed a new system for labeling groups in the periodic table. They numbered the groups from left to right 1 through 18 (the black labels in Figure 6.4). The large periodic table in Figure 6.9 includes the IUPAC system and the system used in the United States.

Dividing the elements into groups is not the only way to classify them based on their properties. The elements can be grouped into three broad classes based on their general properties. **Three classes of elements are metals, nonmetals, and metalloids.** Across a period, the properties of elements become less metallic and more nonmetallic.

Figure 6.5 Metals
The metals aluminum, copper, and iron have many important uses. The properties of the metal determine how it is used.



Aluminum (Al)
Aluminum is one of the metals that can be shaped into a thin sheet, or foil.



Copper (Cu)
Copper is ductile and second only to silver as a conductor of electric current. The copper used in electrical cables must be 99.99 percent pure.

Metals The number of yellow squares in Figure 6.4 shows that most elements are metals—about 80 percent. **Metals** are generally good conductors of heat and electric current. A freshly cleaned or cut surface of a metal will have a high luster, or sheen. The sheen is caused by the metal's ability to reflect light. All metals are solids at room temperature, except for mercury (Hg). Many metals are ductile, meaning that they can be drawn into wires. Most metals are malleable, meaning that they can be hammered into thin sheets without breaking. Figure 6.5 shows how the properties of metals can determine how metals are used.

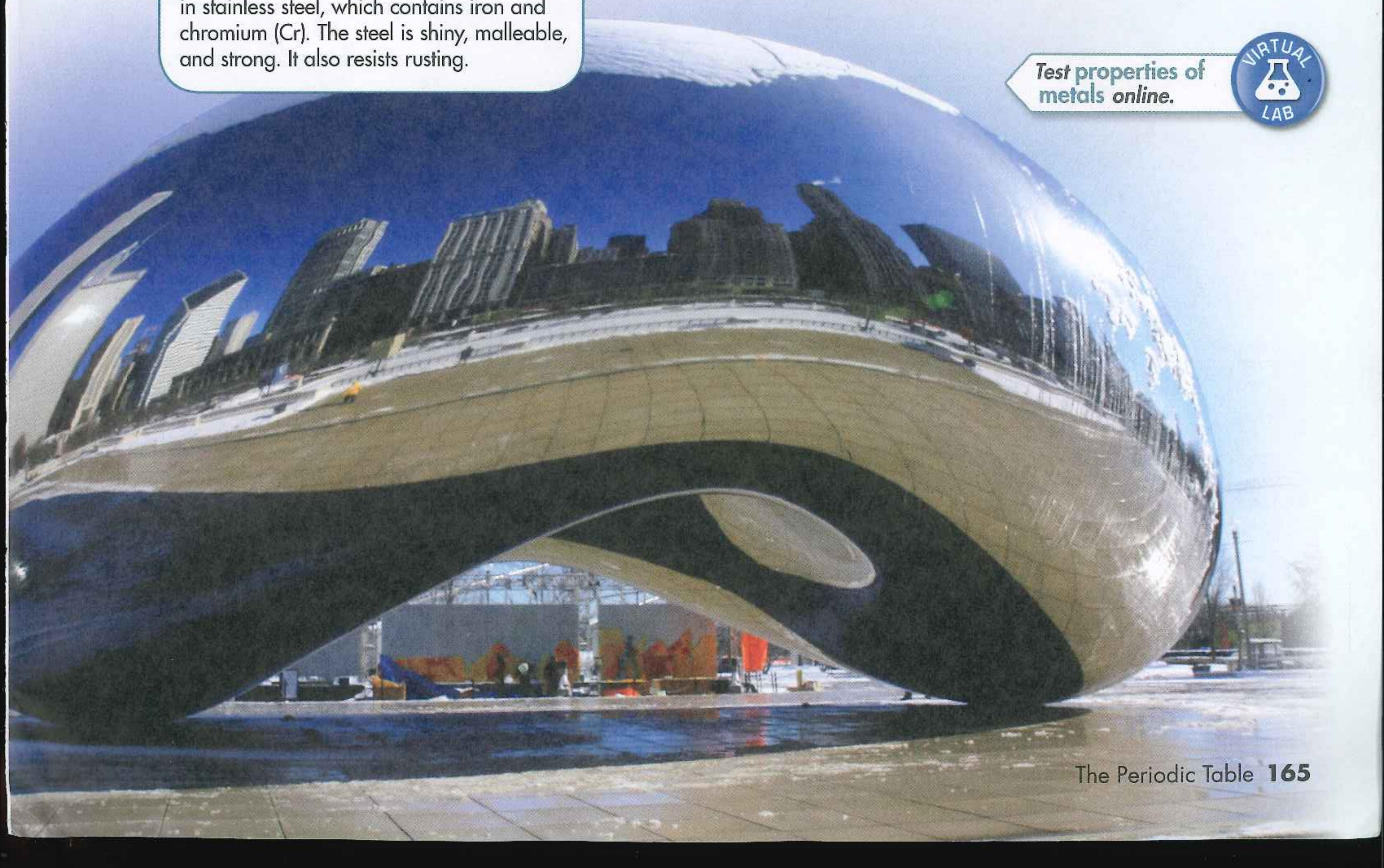
Nonmetals In Figure 6.4, blue is used to identify the nonmetals. With the exception of hydrogen, these elements are in the upper-right corner of the periodic table. There is a greater variation in physical properties among nonmetals than among metals. Most nonmetals are gases at room temperature, including the main components of air—nitrogen and oxygen. A few are solids, such as sulfur and phosphorus. One nonmetal, bromine, is a dark-red liquid. Some examples of nonmetals are shown in Figure 6.6.

The variation among nonmetals makes it difficult to describe one set of general properties that will apply to all nonmetals. However, nonmetals tend to have properties that are opposite to those of metals. In general, **nonmetals** are poor conductors of heat and electric current. Carbon, in the form of graphite, is an exception to this rule. Solid nonmetals tend to be brittle, meaning that they will shatter if hit with a hammer.

Figure 6.6 Nonmetals
The properties of nonmetals vary.

Carbon (C) and Phosphorus (P)
A diamond, which is composed of carbon, is very hard. Some match heads are coated with phosphorus, a brittle solid.

Iron (Fe)
Cloud Gate in Chicago, Illinois, is covered in stainless steel, which contains iron and chromium (Cr). The steel is shiny, malleable, and strong. It also resists rusting.



Test properties of metals *online.* VIRTUAL LAB

CHEMISTRY & YOU

Q: *All of the known elements are listed in the periodic table. What are different ways you could use the periodic table to classify elements?*

Metalloids There is a heavy stair-step line in Figure 6.4 that separates the metals from the nonmetals. Most of the elements that border this line are shaded green. These elements are metalloids. A **metalloid** generally has properties that are similar to those of metals and nonmetals. Under some conditions, a metalloid may behave like a metal. Under other conditions, it may behave like a nonmetal. The behavior often can be controlled by changing the conditions. For example, like most nonmetals, pure silicon is a poor conductor of electric current. However, if a small amount of boron is mixed with silicon, the mixture is a good conductor of electric current. Silicon can be cut into wafers and used to make computer chips. Silicon is also present as the compound silicon dioxide in glass items like the one in Figure 6.7.

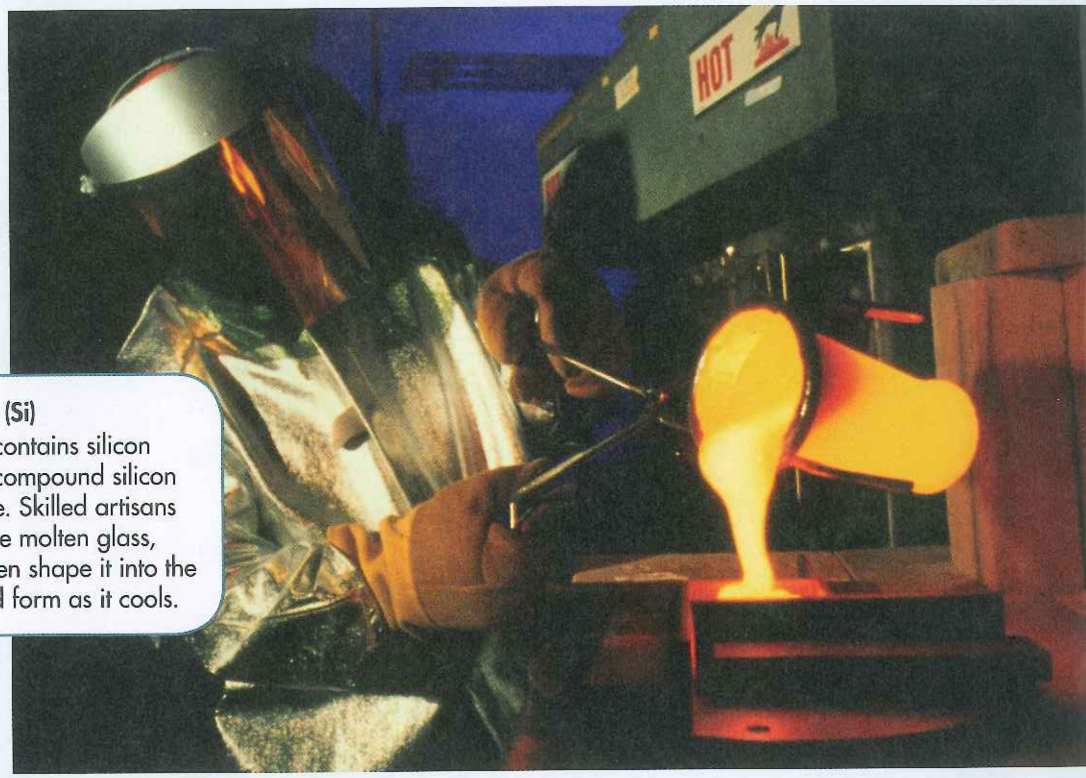


Figure 6.7 Metalloid
Computer chips and glass are two common items that contain metalloids.

Silicon (Si)
Glass contains silicon in the compound silicon dioxide. Skilled artisans prepare molten glass, and then shape it into the desired form as it cools.

6.1 LessonCheck

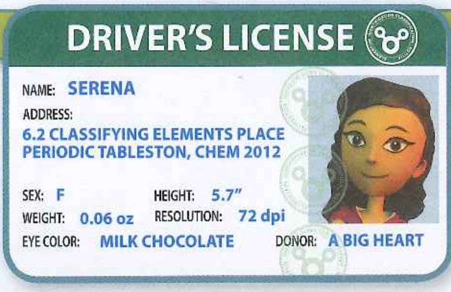
ONLINE PROBLEMS

1. **Explain** How did chemists begin the process of organizing elements?
2. **Identify** What property did Mendeleev use to organize his periodic table?
3. **Explain** How are elements arranged in the modern periodic table?
4. **List** Name the three broad classes of elements.
5. **Classify** Identify each element as a metal, a metalloid, or a nonmetal.
   a. gold **c.** sulfur
   b. silicon **d.** barium
6. **Compare** Which of the following sets of elements have similar physical and chemical properties?
   a. oxygen, nitrogen, carbon, boron
   b. strontium, magnesium, calcium, beryllium
   c. nitrogen, neon, nickel, niobium
7. **Identify** Name two elements that have properties similar to those of the element sodium.

BIG IDEA
ELECTRONS AND THE STRUCTURE OF ATOMS

8. Why is atomic number better than atomic mass for organizing the elements in the periodic table?

6.2 Classifying the Elements



CHEMISTRY & YOU

Q: *What can you learn about each element from the periodic table?* Many people carry a form of identification, such as a driver's license. An ID contains information specific to a particular person, such as the person's name, address, height, eye color, and weight. The periodic table contains a square for each element that supplies information about that element. In this lesson, you will learn the types of information that are usually listed in a periodic table.

Key Questions

What information can be displayed in a periodic table?

How can elements be classified based on electron configurations?

Vocabulary

- alkali metal
- alkaline earth metal
- halogen
- noble gas
- representative element
- transition metal
- inner transition metal

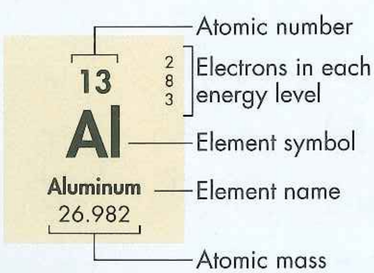


Figure 6.8 Periodic Table Square
This is the element square for aluminum from the periodic table in Figure 6.9.
Interpret Diagrams ***What does the data in the square tell you about the structure of an aluminum atom?***

Reading the Periodic Table

What information can be displayed in a periodic table?

The periodic table is a very useful tool in chemistry. **The periodic table usually displays the symbols and names of the elements, along with information about the structure of their atoms.** Figure 6.8 shows one square from the detailed periodic table of the elements in Figure 6.9 on page 168. In the center of the square is the symbol for aluminum (Al). The atomic number for aluminum (13) is above the symbol. The element name and atomic mass are below the symbol. There is also a vertical column with the numbers 2, 8, and 3, which indicate the number of electrons in each occupied energy level of an aluminum atom.

The symbol for aluminum is printed in black because aluminum is a solid at room temperature. In Figure 6.9, the symbols for gases are in red. The symbols for the two elements that are liquids at room temperature, mercury and bromine, are in blue. The symbols for some elements in Figure 6.9 are printed in grey. These elements are not found in nature. In Chapter 25, you will learn how scientists produce these elements.

The background colors in the squares are used to distinguish groups of elements in the periodic table. For example, two shades of orange are used for the metals in Groups 1A and 2A. The elements in Group 1A are called **alkali metals.** The elements in Group 2A are called **alkaline earth metals.** The name alkali comes from the Arabic *al aqali,* meaning "the ashes." Wood ashes are rich in compounds of the alkali metals sodium and potassium. Some groups of nonmetals also have special names. The nonmetals of Group 7A are called **halogens.** The name *halogen* comes from the combination of the Greek word *hals,* meaning "salt," and the Latin word *genesis,* meaning "to be born." There is a general class of compounds called salts, which include the compound called table salt. Chlorine, bromine, and iodine, the most common halogens, can be prepared from their salts.

Periodic Table of the Elements

Representative Elements

- Alkali Metals
- Alkaline Earth Metals
- Other Metals
- Metalloids
- Nonmetals
- Noble Gases

Transition Elements

- Transition metals
- Inner transition metals

C	Solid
Br	Liquid
He	Gas
Tc	Not found in nature

Period	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B
1	1 H Hydrogen 1.0079 1								
2	3 Li Lithium 6.941 2, 1	4 Be Beryllium 9.0122 2, 2							
3	11 Na Sodium 22.990 2, 8, 1	12 Mg Magnesium 24.305 2, 8, 2							
4	19 K Potassium 39.098 2, 8, 8, 1	20 Ca Calcium 40.08 2, 8, 8, 2	21 Sc Scandium 44.956 2, 8, 9, 2	22 Ti Titanium 47.90 2, 8, 10, 2	23 V Vanadium 50.941 2, 8, 11, 2	24 Cr Chromium 51.996 2, 8, 13, 1	25 Mn Manganese 54.938 2, 8, 13, 2	26 Fe Iron 55.847 2, 8, 14, 2	27 Co Cobalt 58.933 2, 8, 15, 2
5	37 Rb Rubidium 85.468 2, 8, 18, 8, 1	38 Sr Strontium 87.62 2, 8, 18, 8, 2	39 Y Yttrium 88.906 2, 8, 18, 9, 2	40 Zr Zirconium 91.22 2, 8, 18, 10, 2	41 Nb Niobium 92.906 2, 8, 18, 12, 1	42 Mo Molybdenum 95.94 2, 8, 18, 13, 1	43 Tc Technetium (98) 2, 8, 18, 14, 1	44 Ru Ruthenium 101.07 2, 8, 18, 15, 1	45 Rh Rhodium 102.91 2, 8, 18, 16, 1
6	55 Cs Cesium 132.91 2, 8, 18, 18, 8, 1	56 Ba Barium 137.33 2, 8, 18, 18, 8, 2	71 Lu Lutetium 174.97 2, 8, 18, 32, 9, 2	72 Hf Hafnium 178.49 2, 8, 18, 32, 10, 2	73 Ta Tantalum 180.95 2, 8, 18, 32, 11, 2	74 W Tungsten 183.85 2, 8, 18, 32, 12, 2	75 Re Rhenium 186.21 2, 8, 18, 32, 13, 2	76 Os Osmium 190.2 2, 8, 18, 32, 14, 2	77 Ir Iridium 192.22 2, 8, 18, 32, 15, 2
7	87 Fr Francium (223) 2, 8, 18, 32, 18, 8, 1	88 Ra Radium (226) 2, 8, 18, 32, 18, 8, 2	103 Lr Lawrencium (262) 2, 8, 18, 32, 32, 9, 2	104 Rf Rutherfordium (261) 2, 8, 18, 32, 32, 10, 2	105 Db Dubnium (262) 2, 8, 18, 32, 32, 11, 2	106 Sg Seaborgium (263) 2, 8, 18, 32, 32, 12, 2	107 Bh Bohrium (264) 2, 8, 18, 32, 32, 13, 2	108 Hs Hassium (265) 2, 8, 18, 32, 32, 14, 2	109 Mt Meitnerium (268) 2, 8, 18, 32, 32, 15, 2

Elements 104–118 are the transactinide elements.

57 La Lanthanum 138.91 2, 8, 18, 18, 9, 2	58 Ce Cerium 140.12 2, 8, 18, 20, 8, 2	59 Pr Praseodymium 140.91 2, 8, 18, 21, 8, 2	60 Nd Neodymium 144.24 2, 8, 18, 22, 8, 2	61 Pm Promethium (145) 2, 8, 18, 23, 8, 2	62 Sm Samarium 150.4 2, 8, 18, 24, 8, 2

Lanthanide Series

89 Ac Actinium (227) 2, 8, 18, 32, 18, 9, 2	90 Th Thorium 232.04 2, 8, 18, 32, 18, 10, 2	91 Pa Protactinium 231.04 2, 8, 18, 32, 20, 9, 2	92 U Uranium 238.03 2, 8, 18, 32, 21, 9, 2	93 Np Neptunium (237) 2, 8, 18, 32, 22, 9, 2	94 Pu Plutonium (244) 2, 8, 18, 32, 24, 8, 2

Actinide Series

Figure 6.9 Periodic Table
In this periodic table, the colors of the squares are used to classify the elements.

KINETIC ART
Take a tour of the periodic table online.

13 Al (2, 8, 3) Aluminum 26.982	
13	Atomic number
2, 8, 3	Electrons in each energy level
Al	Element symbol
Aluminum	Element name
26.982	Atomic mass†

†*The atomic masses in parentheses are the mass numbers of the longest-lived isotope of elements for which a standard atomic mass cannot be defined.*

10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
								2 He Helium 4.0026 (2)
			5 B Boron 10.81 (2, 3)	6 C Carbon 12.011 (2, 4)	7 N Nitrogen 14.007 (2, 5)	8 O Oxygen 15.999 (2, 6)	9 F Fluorine 18.998 (2, 7)	10 Ne Neon 20.179 (2, 8)
			13 Al Aluminum 26.982 (2, 8, 3)	14 Si Silicon 28.086 (2, 8, 4)	15 P Phosphorus 30.974 (2, 8, 5)	16 S Sulfur 32.06 (2, 8, 6)	17 Cl Chlorine 35.453 (2, 8, 7)	18 Ar Argon 39.948 (2, 8, 8)
28 Ni Nickel 58.71 (2, 8, 16, 2)	29 Cu Copper 63.546 (2, 8, 18, 1)	30 Zn Zinc 65.38 (2, 8, 18, 2)	31 Ga Gallium 69.72 (2, 8, 18, 3)	32 Ge Germanium 72.59 (2, 8, 18, 4)	33 As Arsenic 74.922 (2, 8, 18, 5)	34 Se Selenium 78.96 (2, 8, 18, 6)	35 Br Bromine 79.904 (2, 8, 18, 7)	36 Kr Krypton 83.80 (2, 8, 18, 8)
46 Pd Palladium 106.4 (2, 8, 18, 18)	47 Ag Silver 107.87 (2, 8, 18, 18, 1)	48 Cd Cadmium 112.41 (2, 8, 18, 18, 2)	49 In Indium 114.82 (2, 8, 18, 18, 3)	50 Sn Tin 118.69 (2, 8, 18, 18, 4)	51 Sb Antimony 121.75 (2, 8, 18, 18, 5)	52 Te Tellurium 127.60 (2, 8, 18, 18, 6)	53 I Iodine 126.90 (2, 8, 18, 18, 7)	54 Xe Xenon 131.30 (2, 8, 18, 18, 8)
78 Pt Platinum 195.09 (2, 8, 18, 32, 17, 1)	79 Au Gold 196.97 (2, 8, 18, 32, 18, 1)	80 Hg Mercury 200.59 (2, 8, 18, 32, 18, 2)	81 Tl Thallium 204.37 (2, 8, 18, 32, 18, 3)	82 Pb Lead 207.2 (2, 8, 18, 32, 18, 4)	83 Bi Bismuth 208.98 (2, 8, 18, 32, 18, 5)	84 Po Polonium (209) (2, 8, 18, 32, 18, 6)	85 At Astatine (210) (2, 8, 18, 32, 18, 7)	86 Rn Radon (222) (2, 8, 18, 32, 18, 8)
110 Ds Darmstadtium (269) (2, 8, 18, 32, 32, 17, 1)	111 Rg Roentgenium (272) (2, 8, 18, 32, 32, 18, 1)	112 Cn Copernicium (277) (2, 8, 18, 32, 32, 18, 2)	113 *Uut Ununtrium (284) (2, 8, 18, 32, 32, 18, 3)	114 *Uuq Ununquadium (289) (2, 8, 18, 32, 32, 18, 4)	115 *Uup Ununpentium (288) (2, 8, 18, 32, 32, 18, 5)	116 *Uuh Ununhexium (293) (2, 8, 18, 32, 32, 18, 6)	117 *Uus Ununseptium Classification pending	118 *Uuo Ununoctium (299) (2, 8, 18, 32, 32, 18, 8)

*Discovery reported but not verified

63 Eu Europium 151.96 (2, 8, 18, 25, 8, 2)	64 Gd Gadolinium 157.25 (2, 8, 18, 25, 9, 2)	65 Tb Terbium 158.93 (2, 8, 18, 27, 8, 2)	66 Dy Dysprosium 162.50 (2, 8, 18, 28, 8, 2)	67 Ho Holmium 164.93 (2, 8, 18, 29, 8, 2)	68 Er Erbium 167.26 (2, 8, 18, 30, 8, 2)	69 Tm Thulium 168.93 (2, 8, 18, 31, 8, 2)	70 Yb Ytterbium 173.04 (2, 8, 18, 32, 8, 2)
95 Am Americium (243) (2, 8, 18, 32, 25, 8, 2)	96 Cm Curium (247) (2, 8, 18, 32, 25, 9, 2)	97 Bk Berkelium (247) (2, 8, 18, 32, 27, 8, 2)	98 Cf Californium (251) (2, 8, 18, 32, 28, 8, 2)	99 Es Einsteinium (252) (2, 8, 18, 32, 29, 8, 2)	100 Fm Fermium (257) (2, 8, 18, 32, 30, 8, 2)	101 Md Mendelevium (258) (2, 8, 18, 32, 31, 8, 2)	102 No Nobelium (259) (2, 8, 18, 32, 32, 8, 2)

Q: *What can you learn about each element from the periodic table?*

Electron Configurations in Groups

How can elements be classified based on electron configurations?

Electrons play a key role in determining the properties of elements, so there should be a connection between an element's electron configuration and its location in the periodic table. **Elements can be sorted into noble gases, representative elements, transition metals, or inner transition metals based on their electron configurations.** You may want to refer to the periodic table as you read about these classes of elements.

The Noble Gases The photos in Figure 6.10 show uses of helium, neon, and argon. Helium, neon, and argon are examples of **noble gases,** the elements in Group 8A of the periodic table. These nonmetals are sometimes called the inert gases because they rarely take part in a reaction. The electron configurations for the first four noble gases in Group 8A are listed below.

Helium (He)	$1s^2$
Neon (Ne)	$1s^2 2s^2 2p^6$
Argon (Ar)	$1s^2 2s^2 2p^6 3s^2 3p^6$
Krypton (Kr)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

Look at the description of the highest occupied energy level for each element, which is highlighted in yellow. The *s* and *p* sublevels are completely filled with electrons—two electrons in the *s* sublevel and six electrons in the *p* sublevel. Chapter 7 will explain how this arrangement of electrons is related to the relative inactivity of the noble gases.

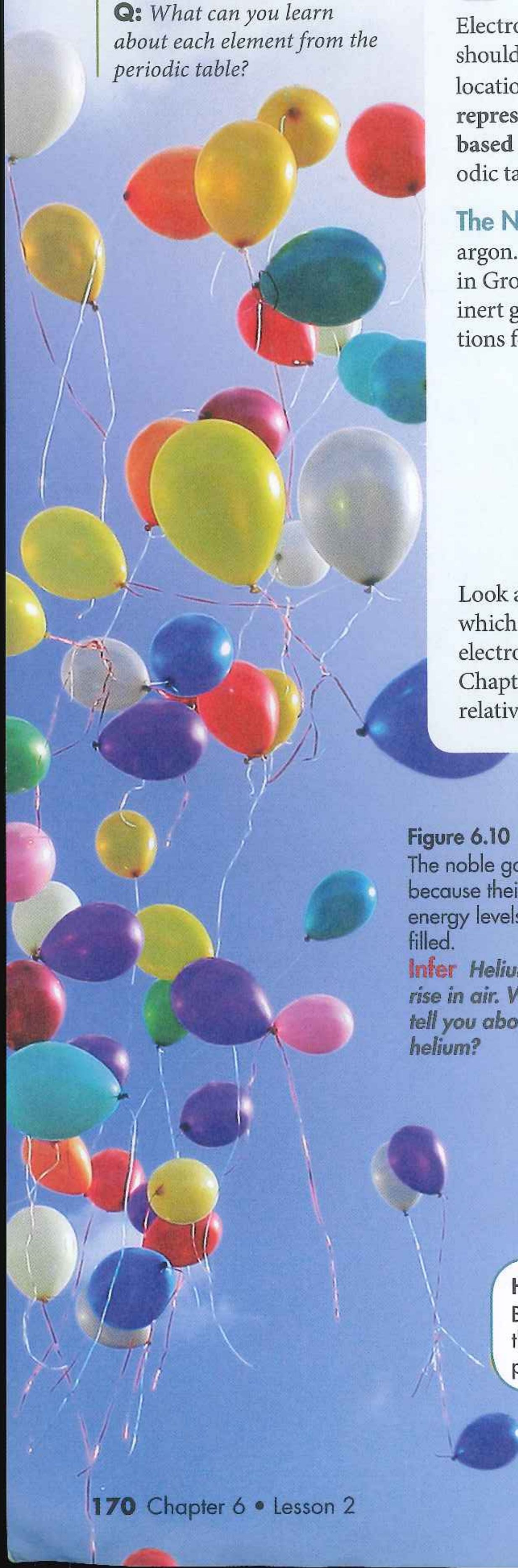
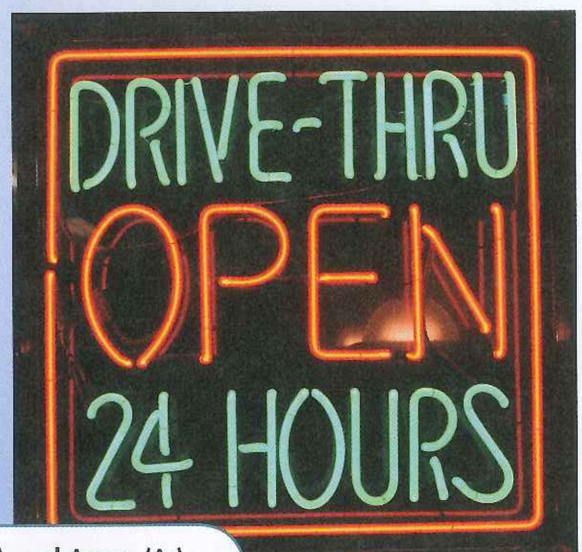


Figure 6.10 Noble Gases
The noble gases are unique because their highest occupied energy levels are completely filled.
Infer ***Helium-filled balloons rise in air. What does this tell you about the density of helium?***



Helium (He), Neon (Ne), and Argon (Ar)
Balloons are often filled with helium to give them "lift." The noble gases neon and argon produce the colors in this neon sign.

The Representative Elements Figure 6.11 shows the portion of the periodic table containing Groups 1A through 7A. Elements in Groups 1A through 7A are often referred to as **representative elements** because they display a wide range of physical and chemical properties. Some elements in these groups are metals, some are nonmetals, and some are metalloids. Most of them are solids, but a few are gases at room temperature, and one, bromine, is a liquid.

In atoms of representative elements, the *s* and *p* sublevels of the highest occupied energy level are not filled. Look at the electron configurations for lithium, sodium, and potassium below. In atoms of these Group 1A elements, there is only one electron in the highest occupied energy level. The electron is in an *s* sublevel.

Lithium (Li)	$1s^2 2s^1$
Sodium (Na)	$1s^2 2s^2 2p^6 3s^1$
Potassium (K)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

In atoms of the Group 4A elements carbon, silicon, and germanium, there are four electrons in the highest occupied energy level.

Carbon (C)	$1s^2 2s^2 2p^2$
Silicon (Si)	$1s^2 2s^2 2p^6 3s^2 3p^2$
Germanium (Ge)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$

For any representative element, its group number equals the number of electrons in the highest occupied energy level.

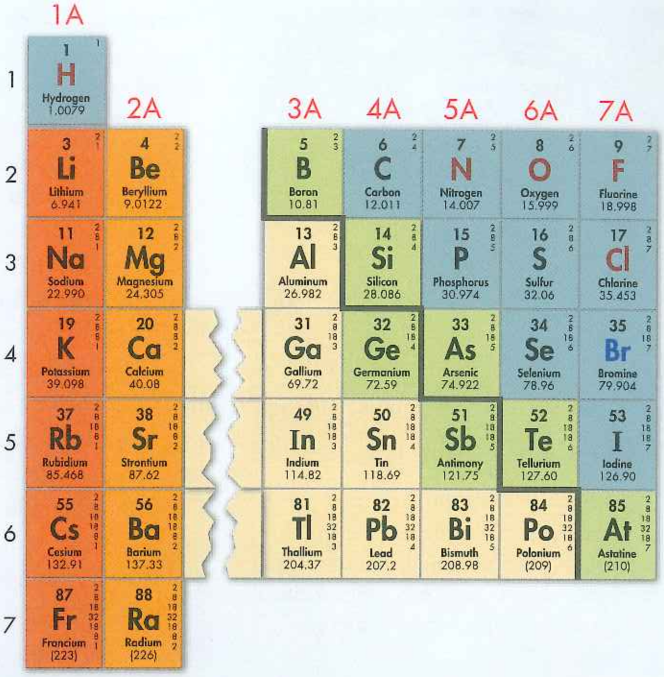
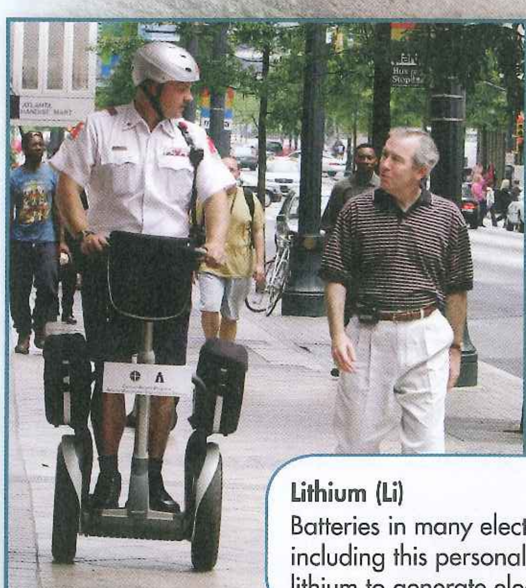
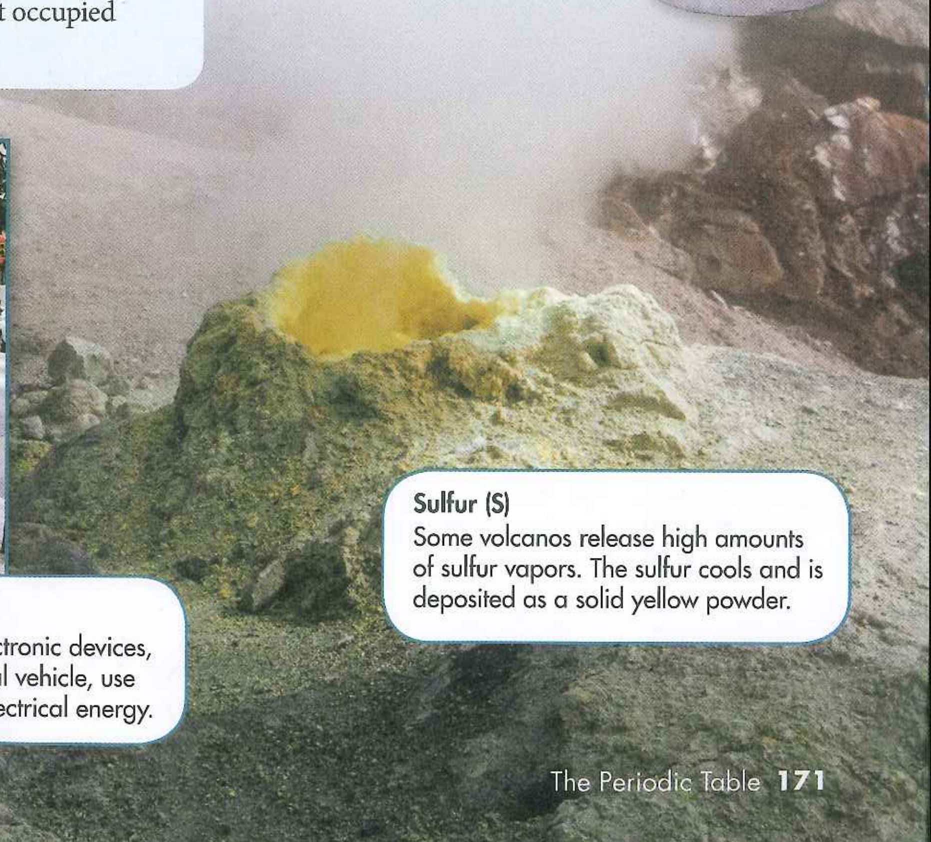


Figure 6.11 Representative Elements
Some of the representative elements exist in nature as elements. Others are found only in compounds.

Tin (Sn)
Artisans often coat objects made of other metals with tin because tin resists corrosion.





Sulfur (S)
Some volcanos release high amounts of sulfur vapors. The sulfur cools and is deposited as a solid yellow powder.

Lithium (Li)
Batteries in many electronic devices, including this personal vehicle, use lithium to generate electrical energy.

Figure 6.12 Uranium
Nuclear power plants use the inner transition metal uranium as fuel. The material shown is called yellowcake, an impure compound of uranium.

Transition Elements In the periodic table, the B group elements separate the A groups on the left side of the table from the A groups on the right side. Elements in the B groups are referred to as transition elements. There are two types of transition elements—transition metals and inner transition metals. They are classified based on their electron configurations.

The **transition metals** are the Group B elements that are usually displayed in the main body of a periodic table. Copper, silver, gold, and iron are transition metals. In atoms of a transition metal, the highest occupied *s* sublevel and a nearby *d* sublevel contain electrons. These elements are characterized by the presence of electrons in *d* orbitals.

The **inner transition metals** are the elements that appear below the main body of the periodic table. In atoms of these elements, the highest occupied *s* sublevel and a nearby *f* sublevel generally contain electrons. The inner transition metals are characterized by the presence of electrons in *f* orbitals. Uranium, an example of an inner transition metal, is shown in Figure 6.12.

Before scientists knew much about inner transition metals, people referred to them as rare-earth elements. This name is misleading because some inner transition metals are more abundant than other elements. Notice that some of the inner transition metals are not found in nature. These elements were prepared in laboratories using methods presented in Chapter 25.

Blocks of Elements If you consider both the electron configurations and the positions of the elements in the periodic table, another pattern emerges. In Figure 6.13, the periodic table is divided into sections, or blocks, that correspond to the highest occupied sublevels. The *s* block contains the elements in Groups 1A and 2A and the noble gas helium. The *p* block contains the elements in Groups 3A, 4A, 5A, 6A, 7A, and 8A, with the exception of helium. The transition metals belong to the *d* block, and the inner transition metals belong to the *f* block.

You can use Figure 6.13 to help determine electron configurations of elements. Each period on the periodic table corresponds to a principal energy level. Suppose an element is located in Period 3. You know that the *s* and *p* sublevels in energy levels 1 and 2 are filled with electrons. You then read across Period 3 from left to right to complete the configuration. For transition elements, electrons are added to a *d* sublevel with a principal energy level that is one less than the period number. For the inner transition metals, the principal energy level of the *f* sublevel is two less than the period number. This procedure gives the correct electron configurations for most atoms.

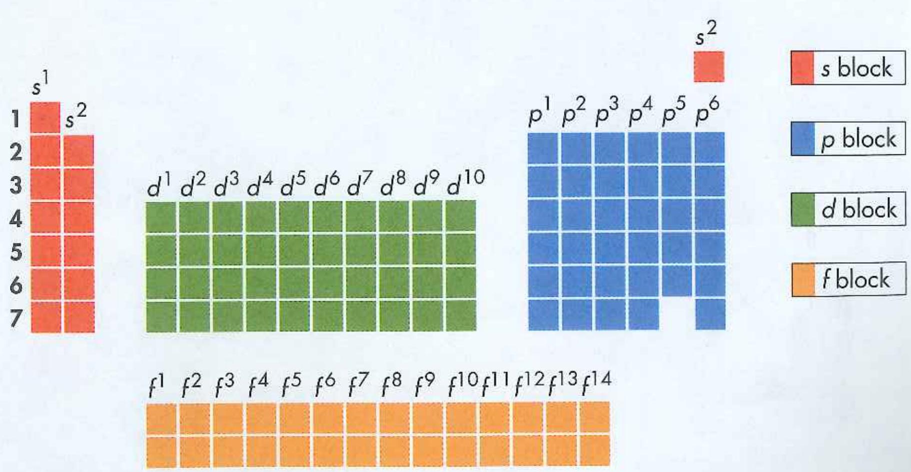


Figure 6.13 Electron Configurations
This diagram classifies elements into blocks according to sublevels that are filled or are filling with electrons.
Interpret Diagrams ***In the highest occupied energy level of a halogen atom, how many electrons are in the p sublevel?***

Sample Problem 6.1

Using Energy Sublevels to Write Electron Configurations

Use Figure 6.9 and Figure 6.13 to write the electron configuration for nickel (Ni).

1 Analyze Identify the relevant concepts. For all elements, the atomic number is equal to the total number of electrons. For a representative element, the highest occupied energy level is the same as the number of the period in which the element is located. You can tell how many electrons are in this energy level from the group in which the element is located.

2 Solve Apply the concepts to this problem.

Use Figure 6.9 to identify where the atom is in the periodic table and the number of electrons in the atom. → Nickel is located in the fourth period and has 28 electrons.

Use Figure 6.13 to determine the electron configuration. → In nickel, the *s* and *p* sublevels in the first three energy levels are full, so the configuration begins with $1s^22s^22p^63s^23p^6$. Next is $4s^2$ and $3d^8$. Put it all together: $1s^22s^22p^63s^23p^63d^84s^2$

9. Use Figure 6.9 and Figure 6.13 to write the electron configurations of the following elements:

a. carbon
b. strontium
c. vanadium

10. List the symbols for all the elements whose electron configurations end as follows. *Note*: Each *n* represents an energy level.

a. ns^2np^1
b. ns^2np^5
c. $ns^2np^6nd^2(n+1)s^2$

6.2 LessonCheck

11. **Identify** What types of information can be included in a periodic table?

12. **List** Into what four classes can elements be sorted based on their electron configurations?

13. **Explain** Why do the elements potassium and sodium have similar chemical properties?

14. **Classify** Identify each element as an alkali metal, an alkaline earth metal, or a halogen:

a. barium
b. chlorine
c. lithium
d. beryllium

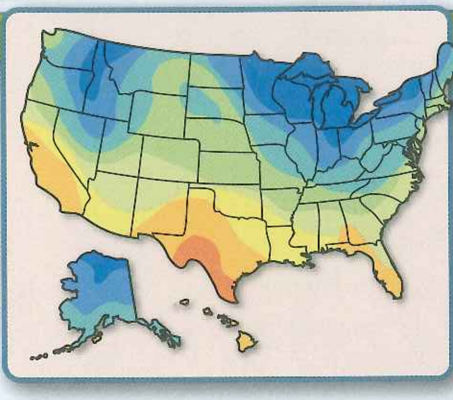
15. **Classify** Based on the following electron configurations, identify each element as a representative element, transition metal, or noble gas.

a. $1s^22s^22p^63s^23p^63d^{10}4s^24p^6$
b. $1s^22s^22p^63s^23p^63d^64s^2$
c. $1s^22s^22p^63s^23p^2$

16. **Describe** How many electrons are in the highest occupied energy level of an element in Group 5A?

17. **Identify** Which of these elements are transition metals: Cu, Sr, Cd, Au, Al, Ge, Co?

6.3 Periodic Trends



CHEMISTRY & YOU

Q: *How are trends in the weather similar to trends in the properties of elements?* Although the weather changes from day to day. The weather you experience is related to your location on the globe. For example, Florida has an average temperature that is higher than Minnesota's. Similarly, a rain forest receives more rain than a desert. These differences are attributable to trends in the weather. In this lesson, you will learn how a property such as atomic size is related to the location of an element in the periodic table.

Key Questions

What are the trends among the elements for atomic size?

How do ions form?

What are the trends among the elements for first ionization energy, ionic size, and electronegativity?

Vocabulary

- atomic radius
- ion
- cation
- anion
- ionization energy
- electronegativity

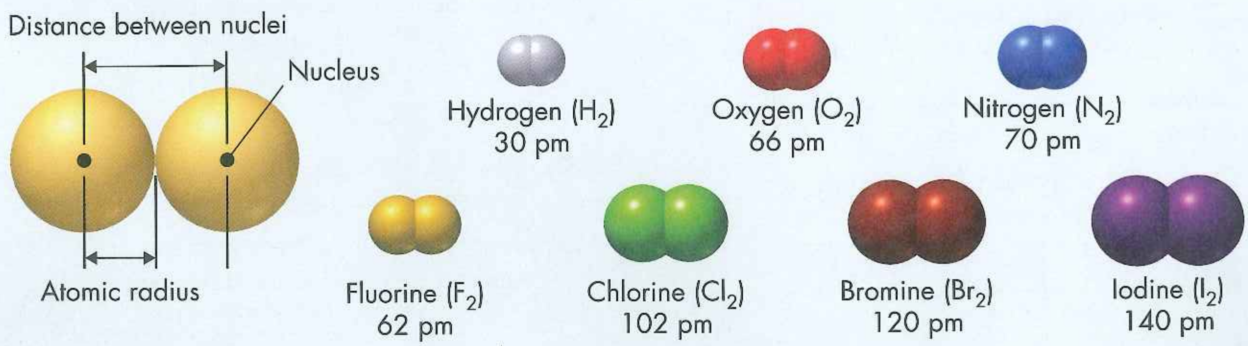
Trends in Atomic Size

What are the trends among the elements for atomic size?

One way to think about atomic size is to look at the units that form when atoms of the same element are joined to one another. These units are called molecules. Figure 6.14 shows models of molecules (molecular models) for seven nonmetals. Because the atoms in each molecule are identical, the distance between the nuclei of these atoms can be used to estimate the size of the atoms. This size is expressed as an atomic radius. The **atomic radius** is one half of the distance between the nuclei of two atoms of the same element when the atoms are joined.

The distances between atoms in a molecule are extremely small. So the atomic radius is often measured in picometers (pm). Recall that there are one trillion, or 10^{12}, picometers in a meter. The molecular model of iodine in Figure 6.14 is the largest. The distance between the nuclei in an iodine molecule is 280 pm. Because the atomic radius is one half the distance between the nuclei, a value of 140 pm (280/2) is assigned as the radius of the iodine atom. **In general, atomic size increases from top to bottom within a group and decreases from left to right across a period.**

Figure 6.14 Atomic Radii
This diagram compares the atomic radii of seven nonmetals.



Interpret Graphs

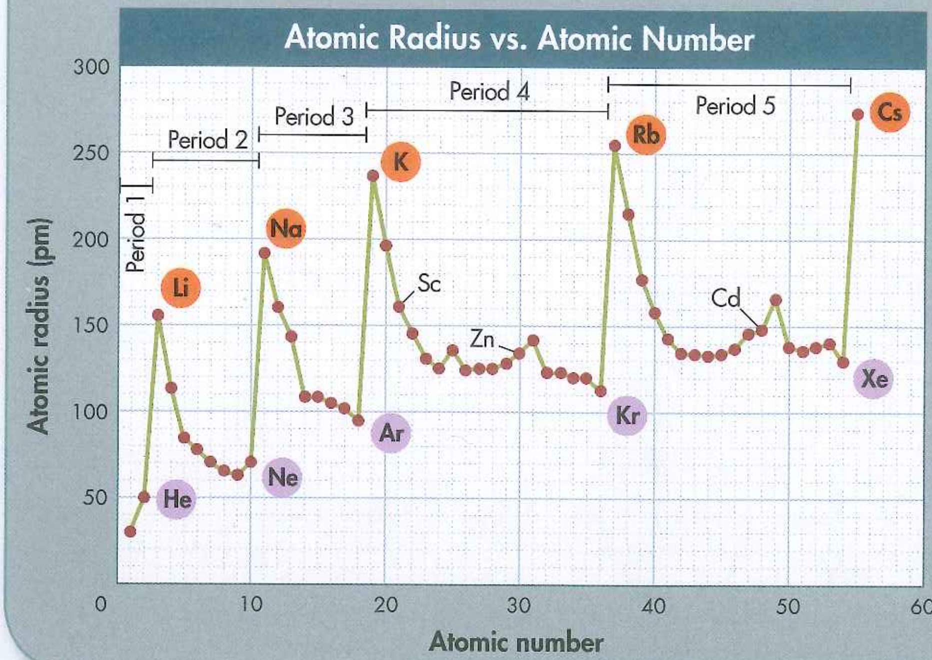


Figure 6.15 This graph plots atomic radius versus atomic number for elements with atomic numbers from 1 to 55.

a. Read Graphs Which alkali metal has an atomic radius of 238 pm?

b. Draw Conclusions Based on the data for alkali metals and noble gases, how does atomic size change within a group?

c. Predict Is an atom of barium, atomic number 56, smaller or larger than an atom of cesium (Cs)?

Group Trends in Atomic Size Look at the data for the alkali metals and noble gases in Figure 6.15. The atomic radius within these groups increases as the atomic number increases. This increase is an example of a trend.

As the atomic number increases within a group, the charge on the nucleus increases and the number of occupied energy levels increases. These variables affect atomic size in opposite ways. The increase in positive charge draws electrons closer to the nucleus. The increase in the number of occupied orbitals shields electrons in the highest occupied energy level from the attraction of protons in the nucleus. The shielding effect is greater than the effect of the increase in nuclear charge, so the atomic size increases.

Period Trends in Atomic Size Look again at Figure 6.15. With increasing atomic number, each element has one more proton and one more electron than the preceding element. Across a period, the electrons are added to the same principal energy level. The shielding effect is constant for all the elements in a period. The increasing nuclear charge pulls the electrons in the highest occupied energy level closer to the nucleus, and the atomic size decreases. Figure 6.16 summarizes the group and period trends in atomic size.

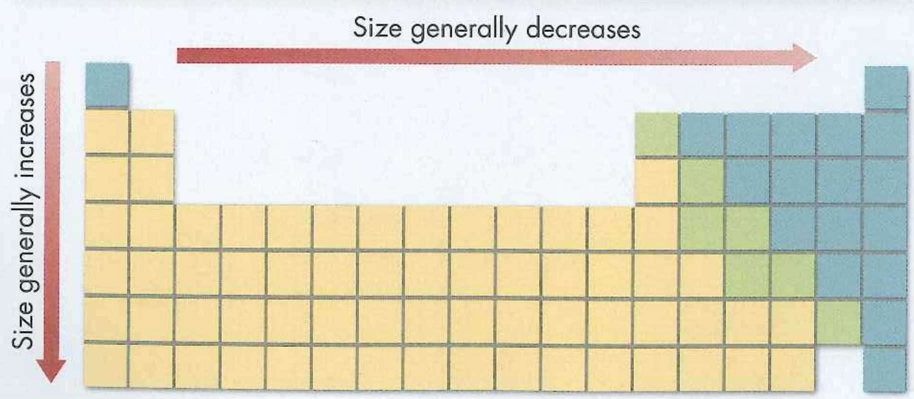
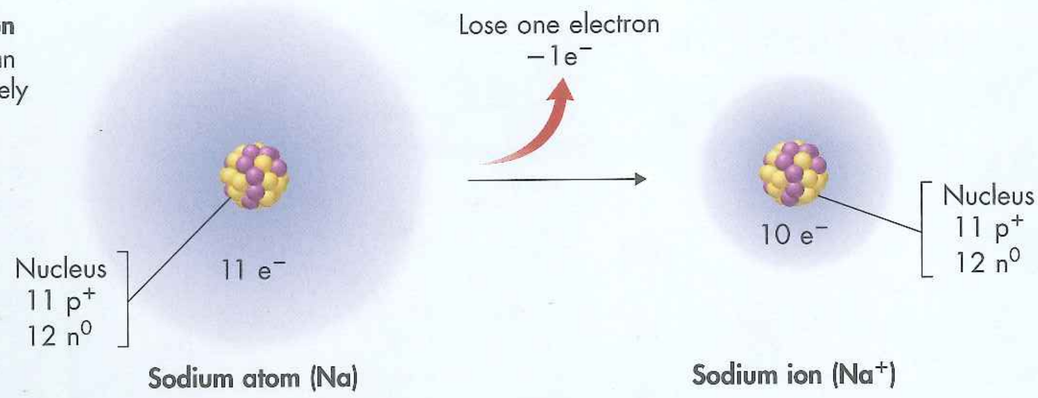


Figure 6.16 Trends in Atomic Size
The size of atoms tends to decrease from left to right across a period and increase from top to bottom within a group.
Predict ***If a halogen and an alkali metal are in the same period, which one will have the larger radius?***

See periodic trends animated online. KINETIC ART

Figure 6.17 **Cation Formation** When a sodium atom loses an electron, it becomes a positively charged ion.



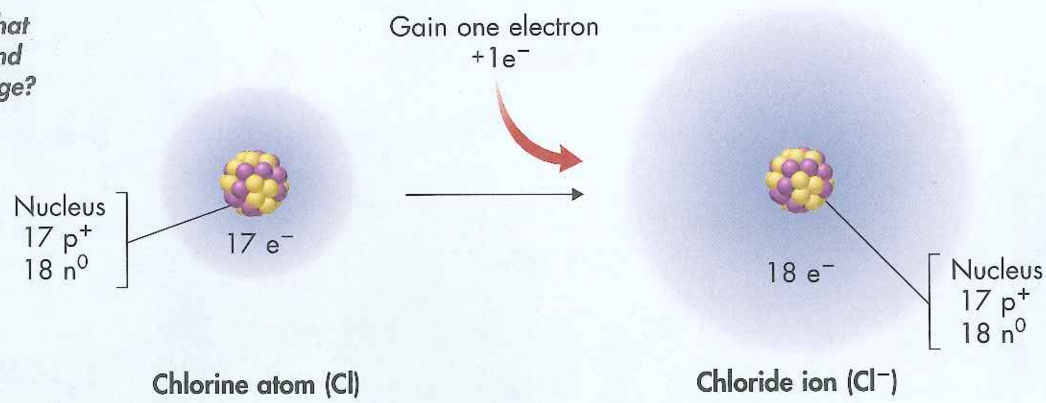
Ions

How do ions form?

Some compounds are composed of particles called ions. An **ion** is an atom or group of atoms that has a positive or negative charge. An atom is electrically neutral because it has equal numbers of protons and electrons. For example, an atom of sodium (Na) has 11 positively charged protons and 11 negatively charged electrons. The net charge on a sodium atom is zero $[(+11) + (-11) = 0]$.

Positive and negative ions form when electrons are transferred between atoms. Atoms of metals, such as sodium, tend to form ions by losing one or more electrons from their highest occupied energy levels. Figure 6.17 compares the atomic structure of a sodium atom and a sodium ion. In the sodium ion, the number of electrons (10) is not equal to the number of protons (11). Because there are more positively charged protons than negatively charged electrons, the sodium ion has a net positive charge. An ion with a positive charge is called a **cation.** The charge for a cation is written as a number followed by a plus sign. If the charge is 1, the number in 1+ is usually omitted from the symbol for the ion. For example, Na^{1+} is written as Na^+. Atoms of nonmetals, such as chlorine, tend to form ions by gaining one or more electrons. Figure 6.18 compares the atomic structure of a chlorine atom and a chloride ion. In a chloride ion, the number of electrons (18) is not equal to the number of protons (17). Because there are more negatively charged electrons than positively charged protons, the chloride ion has a net negative charge. An ion with a negative charge is called an **anion.** The charge for an anion is written as a number followed by a minus sign.

Figure 6.18 Anion Formation When a chlorine atom gains an electron, it becomes a negatively charged ion.
Interpret Diagrams *What happens to the protons and neutrons during this change?*



Trends in Ionization Energy

What are the trends among the elements for first ionization energy?

Recall that electrons can move to higher energy levels when atoms absorb energy. Sometimes the electron has enough energy to overcome the attraction of the protons in the nucleus. The energy required to remove an electron from an atom is called **ionization energy.** This energy is measured when an element is in its gaseous state. The energy required to remove the first electron from an atom is called the first ionization energy. The cation produced has a 1+ charge. **First ionization energy tends to decrease from top to bottom within a group and increase from left to right across a period.**

Ionization energies can help you predict what ions an element will form. Look at the data in Table 6.1 for lithium (Li), sodium (Na), and potassium (K). The increase in energy between the first and second ionization energies is large. It is relatively easy to remove one electron from a Group 1A metal atom, but it is difficult to remove a second electron. This difference indicates that Group 1A metals tend to form ions with a 1+ charge.

Interpret Data

Ionization Energies of First 20 Elements (kJ/mol*)

Symbol	First	Second	Third
H	1312		
He (noble gas)	2372	5247	
Li	520	7297	11,810
Be	899	1757	14,840
B	801	2430	3659
C	1086	2352	4619
N	1402	2857	4577
O	1314	3391	5301
F	1681	3375	6045
Ne (noble gas)	2080	3963	6276
Na	496	4565	6912
Mg	738	1450	7732
Al	578	1816	2744
Si	786	1577	3229
P	1012	1896	2910
S	999	2260	3380
Cl	1256	2297	3850
Ar (noble gas)	1520	2665	3947
K	419	3069	4600
Ca	590	1146	4941

*An amount of matter equal to the atomic mass in grams

Table 6.1 The table compares ionization energies for elements with atomic numbers 1 through 20.

a. Read Tables What are the values for the first, second, and third ionization energies for sodium and aluminum?

b. Compare Is it easier to remove an electron from a sodium (Na) or aluminum (Al) atom? From Na^+ or Al^+? From Na^{2+} or Al^{2+}?

c. Draw Conclusions Which ion is more common—Na^{3+} or Al^{3+}?

Note: The second ionization energy is the energy needed to remove an electron from an ion with a 1+ charge. This produces an ion with a 2+ charge. The third ionization energy is the energy needed to remove an electron from an ion with a 2+ charge. This produces an ion with a 3+ charge.

Interpret Graphs

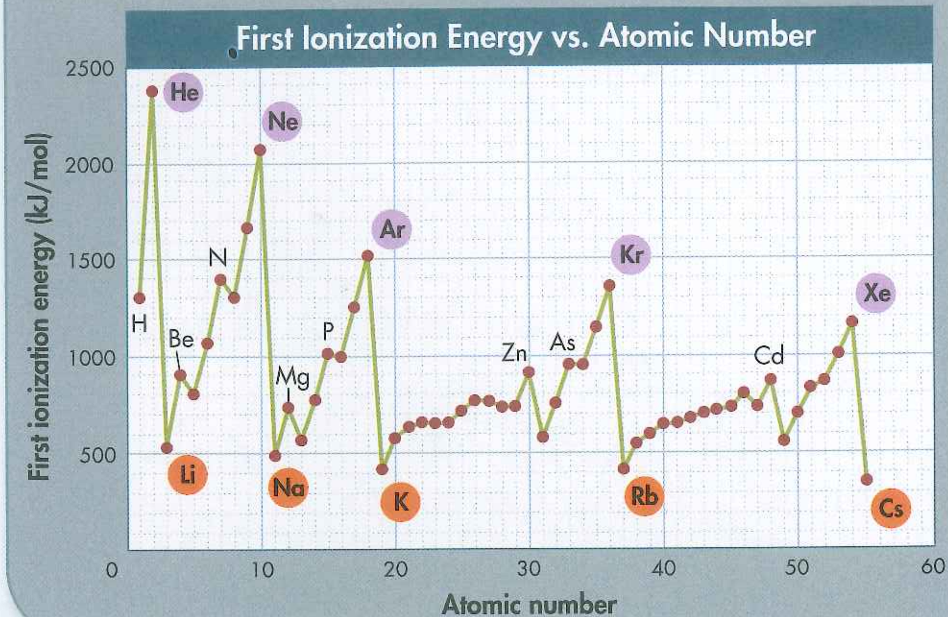
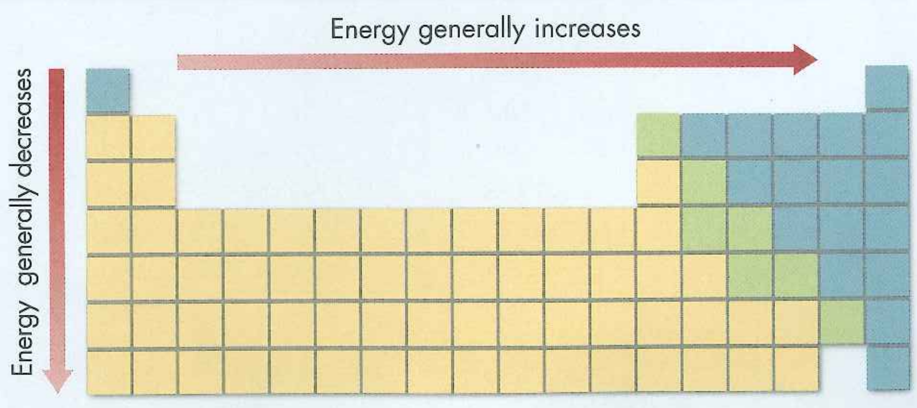


Figure 6.19 This graph reveals group and period trends for ionization energy.
a. Read Graphs Which element in Period 2 has the lowest first ionization energy? In Period 3?
b. Describe What are the group trends for first ionization energy for noble gases and alkali metals?
c. Predict If you drew a graph for second ionization energy, which element would you have to omit? Explain.

Group Trends in Ionization Energy Figure 6.19 is a graph of first ionization energy versus atomic number. Look at the data for the noble gases and the alkali metals. In general, first ionization energy decreases from top to bottom within a group. Recall that the atomic size increases as the atomic number increases within a group. As the size of the atom increases, nuclear charge has a smaller effect on the electrons in the highest occupied energy level. Less energy is required to remove an electron from this energy level, and the first ionization energy is lower.

Period Trends in Ionization Energy In general, the first ionization energy of representative elements tends to increase from left to right across a period. This trend can be explained by the nuclear charge and the shielding effect. The nuclear charge increases across the period, but the shielding effect remains constant. As a result, there is an increase in the attraction of the nucleus for an electron. Thus, it takes more energy to remove an electron from an atom. Figure 6.20 summarizes the group and period trends for first ionization energy.

Figure 6.20
Trends in First Ionization Energy
First ionization energy tends to increase from left to right across a period and decrease from top to bottom within a group.
Predict *Which element would have the larger first ionization energy—an alkali metal in Period 2 or an alkali metal in Period 4?*



Trends in Ionic Size

What are the trends among the elements for ionic size?

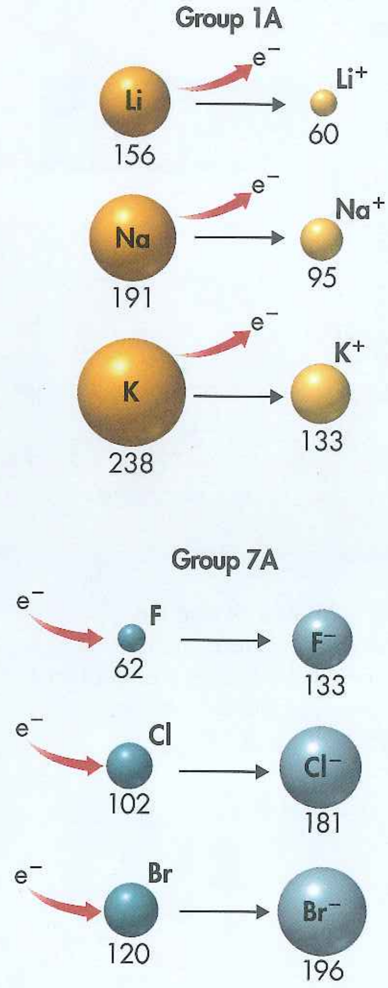
During reactions between metals and nonmetals, metal atoms tend to lose electrons and nonmetal atoms tend to gain electrons. This transfer of electron has a predictable effect on the size of the ions that form. Cations are always smaller than the atoms from which they form. Anions are always larger than the atoms from which they form. **Ionic size tends to increase from top to bottom within a group. Generally, the size of cations and anions decrease from left to right across a period.**

Group Trends in Ionic Size Figure 6.21 compares the relative sizes of the atoms and ions for three metals in Group 1A—lithium (Li), sodium (Na), and potassium (K). For each of these elements, the ion is much smaller than the atom. For example, the radius of a sodium ion (95 pm) is about half the radius of a sodium atom (191 pm). When a sodium atom loses an electron, the attraction between the remaining electrons and the nucleus is increased. As a result, the electrons are drawn closer to the nucleus. Metals that are representative elements tend to lose all their outermost electrons during ionization. Therefore, the ion has one fewer occupied energy level.

The trend is the opposite for nonmetals, like the halogens in Group 7A. Look at Figure 6.21, and compare the relative sizes of the atoms and ions for fluorine (F), chlorine (Cl), and bromine (Br). For each of these elements, the ion is much larger than the atom. For example, the radius of a fluoride ion (133 pm) is more than twice the radius of a fluorine atom (62 pm). As the number of electrons increases, the attraction of the nucleus for any one electron decreases.

Period Trends in Ionic Size Look ahead at Figure 6.23. From left to right across a period, two trends are visible—a gradual decrease in the size of the positive ions (cations), followed by a gradual decrease in the size of the negative ions (anions). Figure 6.22 summarizes the group and period trends in ionic size.

Figure 6.21
Comparing Atomic and Ionic Sizes
This diagram compares the relative sizes of atoms and ions for selected alkali metals (Group 1A) and halogens (Group 7A). The numbers are measurements of the radii given in picometers (pm).



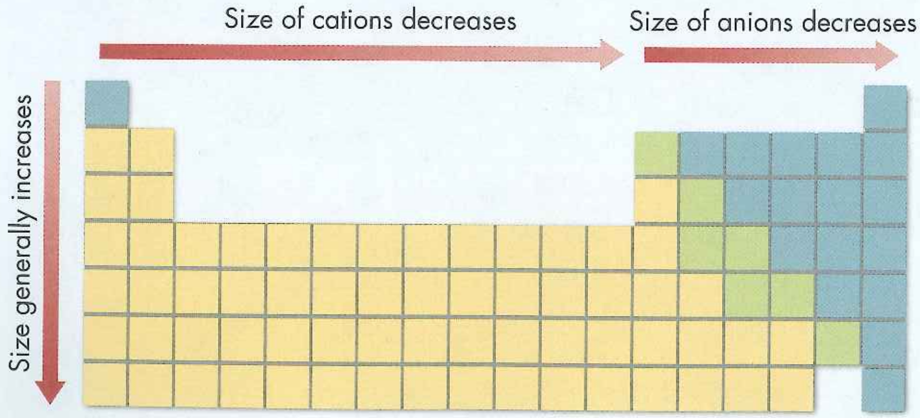


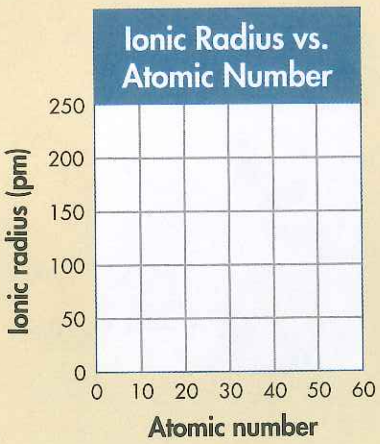
Figure 6.22 Trends in Ionic Size
The ionic radii for cations and anions decrease from left to right across periods and increase from top to bottom within groups.

Quick Lab

Purpose To use a graph to identify period and group trends

Materials

- graph paper
- pencil



Periodic Trends in Ionic Radii

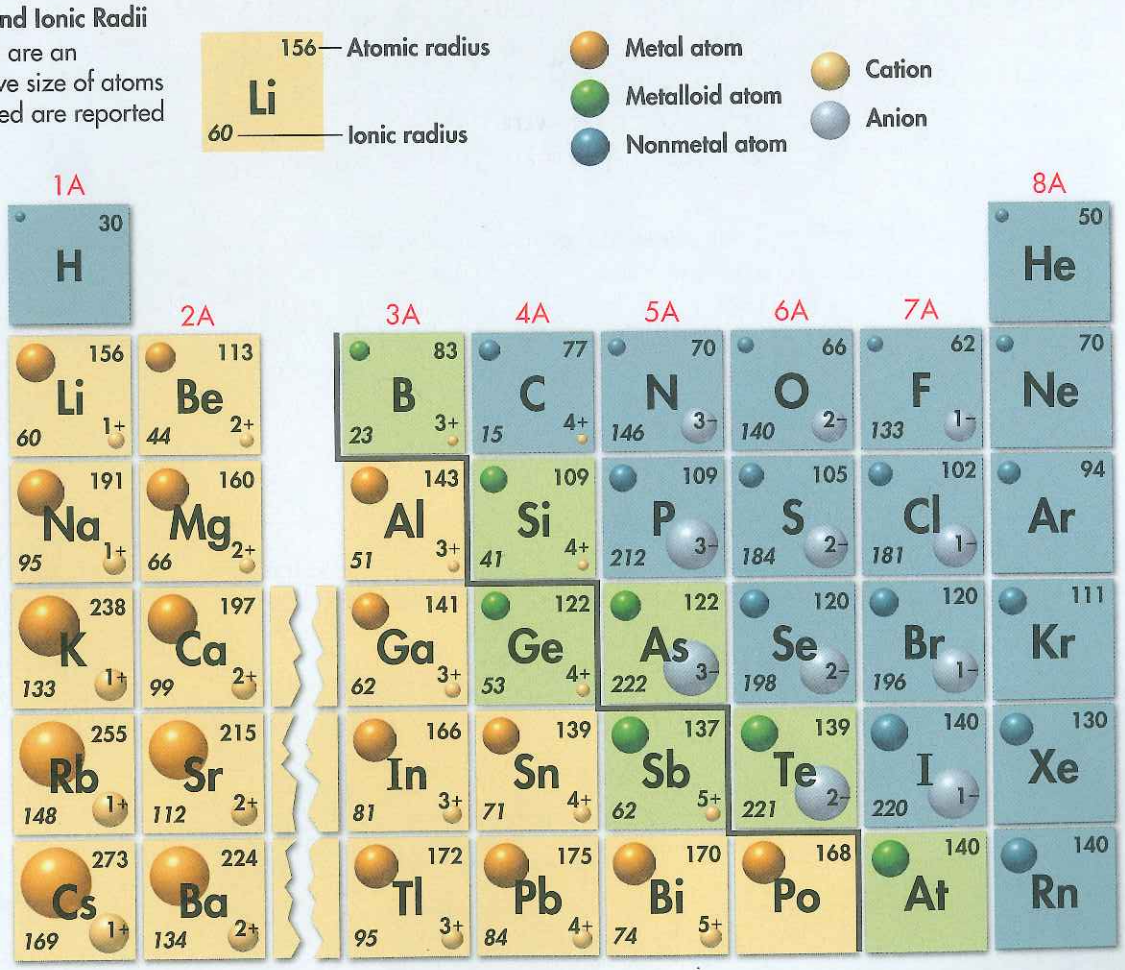
Procedure

Use the data presented in Figure 6.23 to plot ionic radius versus atomic number.

Analyze and Conclude

1. **Compare** How does the size change when an atom forms a cation and when an atom forms an anion?
2. **Describe** How do the ionic radii vary within a group of metals? How do they vary within a group of nonmetals?
3. **Describe** What is the shape of a portion of the graph that corresponds to one period?
4. **Compare and Contrast** Is the trend across a period similar or different for Periods 2, 3, 4, and 5?
5. **Explain** Propose explanations for the trends you have described for ionic radii within groups and across periods.

Figure 6.23 Atomic and Ionic Radii
Atomic and ionic radii are an indication of the relative size of atoms and ions. The data listed are reported in picometers (pm).



Trends in Electronegativity

What are the trends among the elements for electronegativity?

In Chapters 7 and 8, you will study two types of bonds that can exist in compounds. Electrons are involved in both types of bonds. There is a property that can be used to predict the type of bond that will form during a reaction. This property is called electronegativity. **Electronegativity** is the ability of an atom of an element to attract electrons when the atom is in a compound. Scientists use factors such as ionization energy to calculate values for electronegativity.

Table 6.2 lists electronegativity values for representative elements in Groups 1A through 7A. The elements are arranged in the same order as in the periodic table. The noble gases are omitted because they do not form many compounds. The data in Table 6.2 is expressed in Pauling units. Linus Pauling won a Nobel Prize in Chemistry for his work on chemical bonds. He was the first to define electronegativity.

In general, electronegativity values decrease from top to bottom within a group. For representative elements, the values tend to increase from left to right across a period. Metals at the far left of the periodic table have low values. By contrast, nonmetals at the far right (excluding noble gases) have high values. The electronegativity values among the transition metals are not as regular.

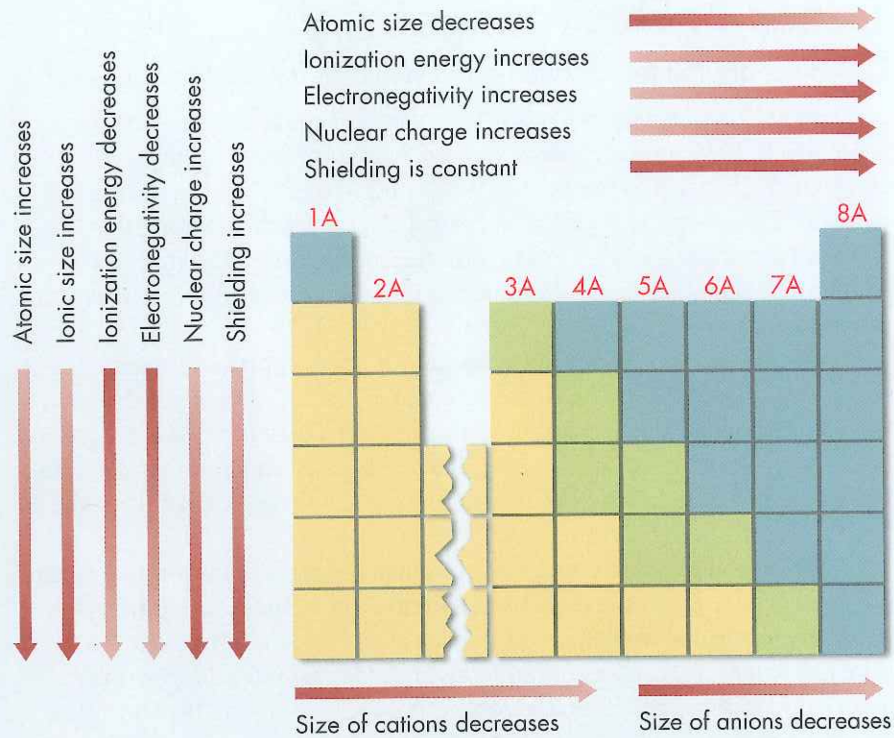
The least electronegative element in the table is cesium, with an electronegativity value of 0.7. It has the least tendency to attract electrons. When it reacts, it tends to lose electrons and form cations. The most electronegative element is fluorine, with a value of 4.0. Because fluorine has such a strong tendency to attract electrons, when it is bonded to any other element it either attracts the shared electrons or forms an anion.

Figure 6.24, on the next page, summarizes several trends that exist among the elements. Refer to this figure as you study the periodic trends presented in this chapter.

Table 6.2

Electronegativity Values for Selected Elements

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9		



CHEMISTRY & YOU

Q: *You are familiar with using a weather map to identify trends in the weather. For example, certain areas are typically warmer than other areas. What trends in the properties of elements can you identify with the help of the periodic table?*

Figure 6.24 Summary of Periodic Trends
Trends for atomic size, ionization energy, ionic size, and electronegativity vary within groups and across periods. The trends that exist among these properties can be explained by variations in atomic structure. The increase in nuclear charge within groups and across periods explains many trends. Within groups, an increase in the number of occupied energy levels and an increase in shielding both have a significant effect on each trend.
Interpret Diagrams ***Which properties tend to decrease across a period? Which properties tend to decrease down a group?***

6.3 LessonCheck

ONLINE PROBLEMS

18. **Review** How does atomic size change within groups and across periods?

19. **Explain** When do ions form?

20. **Summarize** How do first ionization energies vary within groups and across periods?

21. **Describe** Compare the size of ions to the size of the atoms from which they form.

22. **Review** How do electronegativity values vary within groups and across periods?

23. **Explain** In general, how can the periodic trends displayed by elements be explained?

24. **Sequence** Arrange these elements in order of decreasing atomic size: sulfur, chlorine, aluminum, and sodium. Does your arrangement demonstrate a periodic trend or a group trend?

25. **Identify** Which element in each pair has the larger first ionization energy?

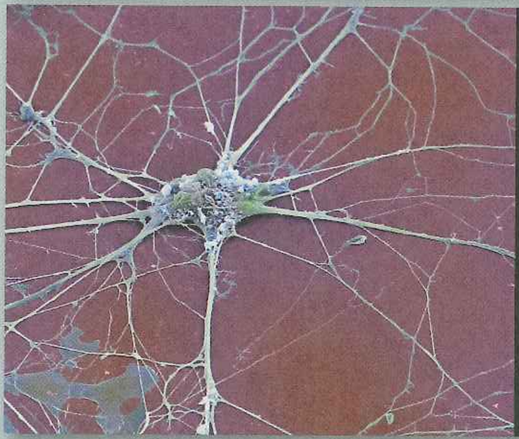
a. sodium, potassium

b. magnesium, phosphorus

Elements of Life

Like everything else in the universe, your body is made up of elements. Your body uses these elements for different functions. Roughly 97 percent of the human body consists of just four elements: oxygen, carbon, hydrogen, and nitrogen. The remaining 3 percent contains about 20 other elements that are essential to life.

CIRCULATORY SYSTEM Iron and oxygen are critical to the circulatory system—the system that carries blood throughout the body. Iron, which is contained in red blood cells, helps transport oxygen from the lungs to other cells in your body. Two other elements—copper and cobalt—are necessary for the formation of red blood cells.



NERVOUS SYSTEM Sodium and potassium are essential to the nervous system, in particular the nerve cells. These elements allow your brain to communicate with other tissues in your body. Other elements that are important for proper nervous system function include calcium, chlorine, zinc, and magnesium.

SKELETAL SYSTEM Your bones and teeth—two components of the skeletal system—are largely comprised of calcium and phosphorus, which give bones and teeth their strength. Fluorine, boron, magnesium, and silicon are also important for bone growth and for maintaining bone strength.

Take It Further

1. Describe Use the information provided on page R1 to estimate the composition of the human body in terms of metals, nonmetals, and metalloids.

2. Predict The elements sodium, magnesium, potassium, and calcium are the most abundant metals in the human body and are present as ions. What is the charge of each of these ions?

3. Sequence Use Figure 6.23 to list the ions in Question 2 from smallest to largest.

Small-Scale Lab

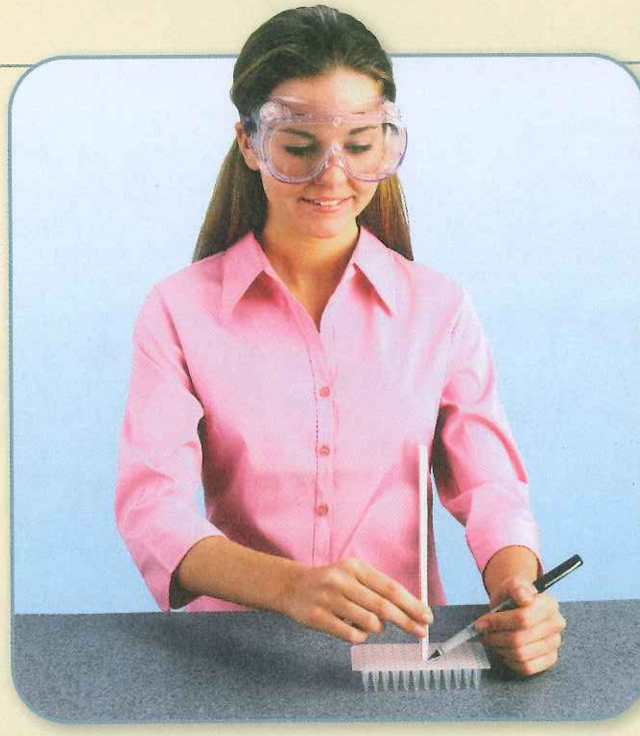
Periodicity in Three Dimensions

Purpose

To build three-dimensional models for periodic trends

Materials

- 96-well spot plate
- straws
- scissors
- metric ruler
- permanent fine-line marker



Procedure

1. Measure the depth of a well in the spot plate by inserting a straw into a well and holding the straw upright as shown in the photograph. Make a mark on the straw at the point where the straw meets the surface of the plate. Measure the distance from the end of the straw to the mark in centimeters (cm). Record this distance as well depth.

2. Cut the straw to a length that is 4.0 cm plus well depth. The straw will extend exactly 4.0 cm above the surface of the plate.

3. Fluorine has an electronegativity value of 4.0. On a scale of 1.0 cm equals 1.0 unit of electronegativity, the portion of the straw that extends above the surface of the plate represents the electronegativity value for fluorine. Using the same scale, cut straws to represent the electronegativity values for all the elements listed in Table 6.2. Remember to add the well depth to the electronegativity value before cutting a straw. As you cut the straws, mark each straw with the chemical symbol of the element that the straw represents.

4. Arrange the straws in the spot plate in rows and columns to match the locations of the elements in the periodic table.

Analyze and Conclude

1. Use Models Which element represented in your model is the most electronegative?

2. Use Models Based on your model, what is the general trend in electronegativity from left to right across a period?

3. Interpret Diagrams Relate the trend in electronegativity across a period to the location of metals and nonmetals in the periodic table.

4. Use Models Based on your model, what is the general trend in electronegativity within a group? Are there any notable exceptions?

5. Explain Why do you think that the electronegativity value for hydrogen is so high given its location in the periodic table?

You're the Chemist

1. Design an Experiment Construct a similar three-dimensional model for first ionization energies. Use the data in Table 6.1 to construct the model. Use a scale of 1.0 cm equals 300 kJ/mol.

2. Design an Experiment Design and construct a three-dimensional model that shows trends in atomic and ionic radii for the elements in Groups 1A and 7A. Devise a way to display both ionic and atomic radii in the same model.

3. Analyze Data Xenon has an electronegativity value of 2.6. Cut and place a straw in your first model to represent xenon. Does xenon support the trend for electronegativity across a period? Is xenon likely to form compounds? Explain your answers.

6 Study Guide

BIGIDEA ELECTRONS AND THE STRUCTURE OF ATOMS

Periodic tables may contain each element's name, symbol, atomic number, atomic mass, and number of electrons in each energy level. The electron configuration of an element can be determined based on the location of an element in the periodic table. Atomic size, ionization energy, ionic size, and electronegativity are trends that vary across periods and groups of the periodic table. These trends can be explained by variations in atomic structure. The increase in nuclear charge within groups and across periods explains many trends. Within groups, an increase in electron shielding has a significant effect on these trends.

6.1 Organizing the Elements

Early chemists used the properties of elements to sort them into groups.

Mendeleev arranged the elements in his periodic table in order of increasing atomic mass.

In the modern periodic table, elements are arranged in order of increasing atomic number.

Three classes of elements are metals, nonmetals, and metalloids.

- periodic law (162)
- metal (165)
- nonmetal (165)
- metalloid (166)

6.2 Classifying the Elements

The periodic table usually displays the symbols and names of elements, along with information about the structure of their atoms.

Elements can be sorted into noble gases, representative elements, transition metals, or inner transition metals based on their electron configurations.

- alkali metal (167)
- alkaline earth metal (167)
- halogen (167)
- noble gas (170)
- representative element (171)
- transition metal (172)
- inner transition metal (172)

6.3 Periodic Trends

In general, atomic size increases from top to bottom within a group and decreases from left to right across a period.

Positive and negative ions form when electrons are transferred between atoms.

First ionization energy tends to decrease from top to bottom within a group and increase from left to right across a period.

Ionic size tends to increase from top to bottom within a group. Generally, the size of cations and anions decrease from left to right across a period.

In general, electronegativity values decrease from top to bottom within a group. For representative elements, the values tend to increase from left to right across a period.

- atomic radius (174)
- ion (176)
- cation (176)
- anion (176)
- ionization energy (177)
- electronegativity (181)

6 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

6.1 Organizing the Elements

26. Why did Mendeleev leave spaces in his periodic table?

★ **27.** What effect did the discovery of gallium have on the acceptance of Mendeleev's table?

28. What pattern is revealed when the elements are arranged in a periodic table in order of increasing atomic number?

29. Based on their locations in the periodic table, would you expect carbon and silicon to have similar properties? Explain your answer.

30. Identify each property below as more characteristic of a metal or a nonmetal.

a. a gas at room temperature
b. brittle
c. malleable
d. poor conductor of electric current
e. shiny

31. In general, how are metalloids different from metals and nonmetals?

6.2 Classifying the Elements

32. Where are the alkali metals, the alkaline earth metals, the halogens, and the noble gases located in the periodic table?

33. Which of the following are symbols for representative elements: Na, Mg, Fe, Ni, Cl?

★ **34.** Which noble gas does not have eight electrons in its highest occupied energy level?

35. Which of these metals isn't a transition metal?

a. aluminum
b. silver
c. iron
d. zirconium

36. Use Figure 6.13 to write the electron configurations of these elements.

a. boron
b. arsenic
c. fluorine
d. zinc
e. aluminum

37. Write the electron configurations of these elements.

a. the noble gas in Period 3
b. the metalloid in Period 3
c. the alkali earth metal in Period 3

6.3 Periodic Trends

★ **38.** Which element in each pair has atoms with a larger atomic radius?

a. sodium, lithium
b. strontium, magnesium
c. carbon, germanium
d. selenium, oxygen

39. Explain the difference between the first and second ionization energy of an element.

40. Which element in each pair has a greater first ionization energy?

a. lithium, boron
b. magnesium, strontium
c. cesium, aluminum

41. Arrange the following groups of elements in order of increasing ionization energy:

a. Be, Mg, Sr **b.** Bi, Cs, Ba **c.** Na, Al, S

42. Why is there a large increase between the first and second ionization energies of the alkali metals?

★ **43.** How does the ionic radius of a typical metal compare with its atomic radius?

44. Which particle has the larger radius in each atom/ion pair?

a. Na, Na^+
b. S, S^{2-}
c. I, I^-
d. Al, Al^{3+}

45. Which element in each pair has a higher electronegativity value?

a. Cl, F
b. C, N
c. Mg, Ne
d. As, Ca

46. Why are noble gases not included in Table 6.2?

★ **47.** When the elements in each pair are chemically combined, which element in each pair has a greater attraction for electrons?

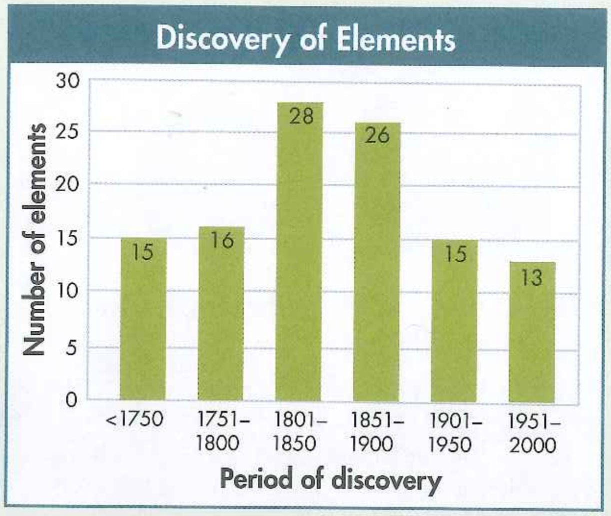
a. Ca or O
b. O or F
c. H or O
d. K or S

48. For which of these properties does lithium have a larger value than potassium?

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

Understand Concepts

49. The bar graph shows how many elements were discovered before 1750 and in each 50-year period between 1750 and 2000.



a. In which 50-year period were the most elements discovered?
b. How did Mendeleev's work contribute to the discovery of elements?
c. What percent of these elements were discovered by 1900?

50. Write the symbol of the element or elements that fit each description.
a. a nonmetal in Group 4A
b. the inner transition metal with the lowest atomic number
c. all of the nonmetals for which the atomic number is a multiple of five
d. a metal in Group 5A

***51.** In which pair of elements are the chemical properties of the elements most similar? Explain your reasoning.
a. sodium and chlorine
b. nitrogen and phosphorus
c. boron and oxygen

52. Explain why fluorine has a smaller atomic radius than both oxygen and chlorine.

53. Would you expect metals or nonmetals in the same period to have higher ionization energies? Give a reason for your answer.

54. In each pair, which ion is larger?
a. Ca^{2+}, Mg^{2+} **b.** Cl^{-}, P^{3-} **c.** Cu^{+}, Cu^{2+}

55. Use the graph in Figure 6.15 to estimate the atomic radius of the indium atom.

***56.** List the symbols for all the elements with electron configurations that end as follows. *Note:* Each *n* represents an energy level.
a. ns^1 **b.** ns^2np^4 **c.** ns^2nd^{10}

57. Explain why there should be a connection between an element's electron configuration and its location on the periodic table.

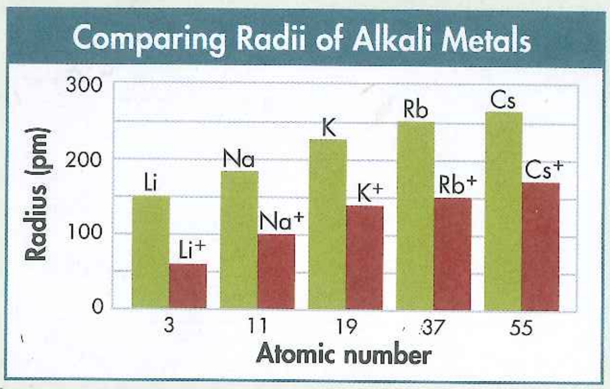
58. Which equation represents the first ionization of an alkali metal atom?
a. $Cl \longrightarrow Cl^+ + e^-$
b. $Ca \longrightarrow Ca^+ + e^-$
c. $K \longrightarrow K^+ + e^-$
d. $H \longrightarrow H^+ + e^-$

59. What trend is demonstrated by the following series of equations?

$$Li + 520\ kJ/mol \longrightarrow Li^+ + e^-$$
$$O + 1314\ kJ/mol \longrightarrow O^+ + e^-$$
$$F + 1681\ kJ/mol \longrightarrow F^+ + e^-$$
$$Ne + 2080\ kJ/mol \longrightarrow Ne^+ + e^-$$

***60.** There is a large jump between the second and third ionization energies of magnesium. There is a large jump between the third and fourth ionization energies of aluminum. Explain these observations.

61. The bar graph shows the relationship between atomic and ionic radii for Group 1A elements.



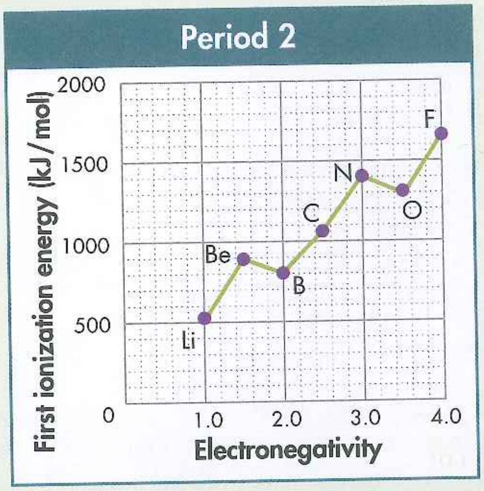
a. Describe and explain the trend in atomic radius within the group.
b. Explain the difference between the size of the atoms and the size of the ions.

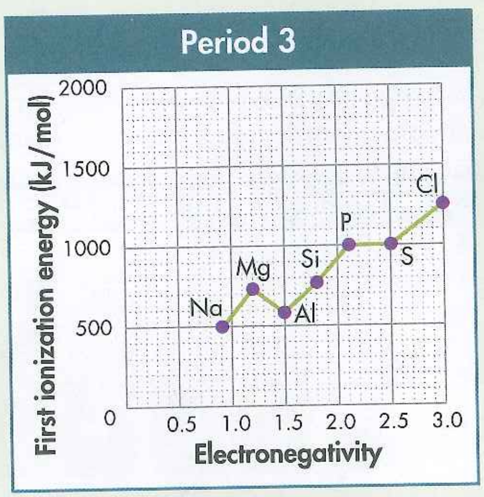
62. Locate each of the following elements in the periodic table and decide whether its atoms are likely to form anions or cations.
a. sodium
b. fluorine
c. calcium
d. potassium
e. iodine
f. beryllium
g. oxygen
h. lithium

Think Critically

63. **Predict** Do you think there are more elements left to discover? If so, what is the lowest atomic number a new element could have? Explain.

64. **Interpret Graphs** The graphs show the relationship between the electronegativities and first ionization energies for Period 2 and Period 3 elements.



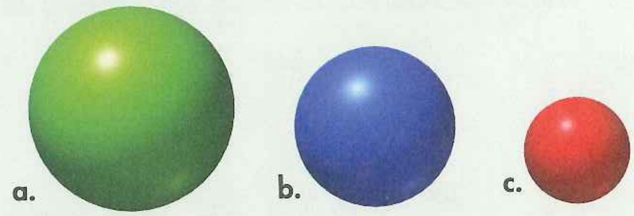


a. Based on data for these two periods, what is the general trend between these two values?
b. Use nuclear charge and shielding effect to explain this trend.

*65. **Explain** Give a reason for each of the following comparisons:
a. Calcium has a smaller second ionization energy than does potassium.
b. Lithium has a larger first ionization energy than does cesium.
c. Magnesium has a larger third ionization energy than does aluminum.

*66. **Explain** Why does it take more energy to remove a 4*s* electron from zinc than from calcium?

67. **Sequence** The following spheres represent Ca, Ca^{2+}, and Mg^{2+}. Which one is which? Explain your reasoning.

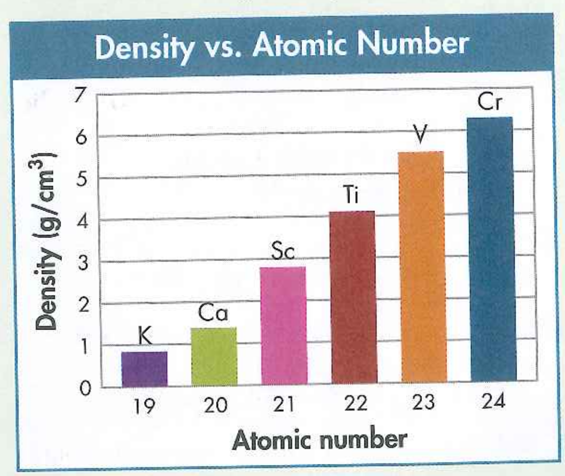


*68. **Apply Concepts** Write the electron configurations of the following ions:
a. the liquid in Group 7A with a 1− charge
b. the metalloid in Period 3 with a 4+ charge
c. the gas in Group 6A with a 2− charge
d. the alkali earth metal in Period 3 with a 2+ charge

69. **Interpret Diagrams** Use the periodic table and Figure 6.13 to identify the following elements:
a. has its outermost electron in $7s^1$
b. contains only one electron in a *d* orbital

70. **Make Generalizations** Why is the first ionization energy of a nonmetal much higher than that of an alkali metal?

71. **Infer** The bar graph shows the densities for the first six elements in Period 4. The density increases across this period from potassium to chromium. Use trends in the periodic table to explain this behavior. *Hint*: What is the equation for determining density?



72. **Explain** Why are cations smaller and anions larger than the corresponding atoms?

Enrichment

73. Analyze Data Make a graph of atomic mass versus atomic number. Choose eleven points (atomic numbers 1, 10, 20, and so forth up to atomic number 100) to make your graph. Use the graph to describe the relationship between atomic mass and atomic number. Is there a 1:1 correlation between atomic mass and atomic number? Explain.

74. Compare and Contrast The Mg^{2+} and Na^+ ions each have ten electrons. Which ion would you expect to have the smaller radius? Why?

***75. Predict** Electron affinity is a measure of an atom's ability to gain electrons. Predict the trend for electron affinity across a period. Explain your answer.

76. Explain The ions S^{2-}, Cl^-, K^+, Ca^{2+}, and Sc^{3+} have the same total number of electrons as the noble gas argon. How would you expect the radii of these ions to vary? Would you expect to see the same variation in the series O^{2-}, F^-, Na^+, Mg^{2+}, and Al^{3+}, in which each ion has the same total number of electrons as the noble gas neon? Explain your answer.

77. Graph The ionization energies for the removal of the first six electrons in carbon are, starting with the first electron, 1086 kJ/mol, 2352 kJ/mol, 4619 kJ/mol, 6220 kJ/mol, 37,820 kJ/mol, and 47,260 kJ/mol.

- **a.** Use these data to construct a graph of ionization energy versus ionization number. *Note*: The ionization number indicates which electron is lost.
- **b.** Between which two ionization numbers does the ionization energy have the largest increase? Why is this behavior predictable?

***78. Infer** Atoms and ions with the same number of electrons are described as *isoelectronic*.

- **a.** Write the symbol for a cation and an anion that are isoelectronic with krypton.
- **b.** Is it possible for a cation to be isoelectronic with an anion in the same period? Explain.

79. Predict Estimate the atomic radius of praseodymium based on the following data for atomic radii of neighboring elements: La (187.9 pm), Ce (183.2 pm), Nd (182.1 pm), and Pm (181.1 pm). Compare your prediction to the value given in a chemistry handbook.

Write About Science

80. Explain Why does the size of an atom tend to increase from top to bottom within a group? Why does the size of an atom tend to decrease from left to right across a period?

81. Connect to the BIGIDEA The ion Zn^{2+} is important in several biological processes. One process depends on Zn^{2+} temporarily binding to a molecule in red blood cells. When Zn^{2+} is absent, Cd^{2+} can bind to the molecule. However, Cd^{2+} binds more strongly and adversely affects the process. The differences in ionic size is one important cause of the difference in biological activity. How would you expect Hg^{2+} to affect this process? Why?

CHEMYSTERY

Made in the USA

Several of the "unnatural" elements were in fact "made in the USA." For example, elements with atomic numbers 94 through 102 were first artificially prepared in California. Three of these elements have names to prove it—americium, berkelium, and californium. Elements such as these are labeled on the periodic table as "Not found in nature" in this book and as "Artificially prepared" in some others.

Most of the artificially prepared elements are actinides or transactinides. Each of these elements has an unstable nucleus. As a result, these elements undergo radioactive decay, which means their nuclei spontaneously break down into smaller parts in the attempt to gain stability.

82. Infer The elements with atomic numbers 99, 101, 104, and 107 were named to honor past influential scientists. Identify the scientist that each element is meant to honor.

83. Connect to the BIGIDEA Many smoke detectors use the artificially prepared element americium. For a challenge, write the electron configuration of americium.

Cumulative Review

84. Explain why science today depends less on chance discoveries than it did in the past.

★85. Identify each process as a chemical change or a physical change.

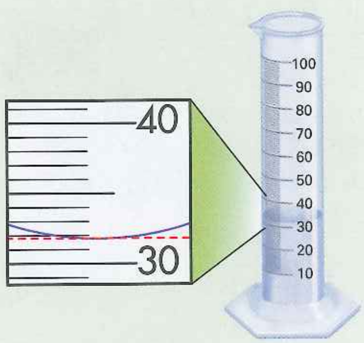
a. melting of iron
b. lighting a match
c. grinding corn
d. souring of milk

86. Describe at least two methods to separate a mixture of small copper and iron beads.

87. In the United States, a typical can of cola holds 355 mL. How many 2.00-L bottles could be filled from a 24-can case of cola?

88. The volume of the liquid in the graduated cylinder is reported as 31.8 mL.

a. How many significant figures are there in the measurement?
b. In which digit is there uncertainty?



89. A cube of plastic 1.20×10^{-5} km on a side has a mass of 1.70 g. Show by calculation whether this plastic cube will sink or float in pure water.

★90. Convert the measurements to meters. Express your answers in scientific notation.

a. 2.24 nm
b. 8.13 cm
c. 7.4 pm
d. 9.37 mm

91. An apprentice jeweler determines the density of a sample of pure gold to be 20.3 g/cm^3. The accepted value is 19.3 g/cm^3. What is the percent error of the jeweler's density measurement?

92. What is the mass of 7.7 L of gasoline at 20°C? Assume the density of gasoline to be 0.68 g/cm^3.

★93. A black olive containing its seed has a mass of 4.5 g and a volume of 4.3 cm^3. Will the olive sink or float on water?

94. The distance from the sun to Earth is 1.50×10^8 km. The speed of light is 3.00×10^8 m/s. How many round trips between Earth and the sun could a beam of light make in one day?

95. The table shows how the volume of sulfur varies with mass. How does the density of sulfur vary with mass?

Mass of Sulfur vs. Volume of Sulfur

Mass of sulfur (g)	Volume of sulfur (cm^3)
23.5	11.4
60.8	29.2
115	55.5
168	81.1

96. Calculate the volume of acetone with the same mass as 15.0 mL of mercury. The density of mercury is 13.59 g/mL. The density of acetone is 0.792 g/mL.

97. A rectangular container has inside dimensions of 15.2 cm by 22.9 cm and is about 1 meter tall. Water is poured into the container to a height of 55.0 cm. When a jagged rock with a mass of 5.21 kg is placed in the container, it sinks to the bottom. The water level rises to 58.3 cm and completely covers the rock. What is the density of the rock?

★98. How many neutrons does an atom of each isotope contain?

a. ${}^{84}_{36}Kr$ **b.** ${}^{79}_{35}Br$ **c.** ${}^{190}_{76}Os$ **d.** ${}^{185}_{75}Re$

99. Name the element and calculate the number of requested subatomic particles in each isotope.

a. neutrons in ${}^{109}_{47}Ag$
b. protons in ${}^{118}_{50}Sn$
c. electrons in ${}^{96}_{42}Mo$
d. electrons in ${}^{45}_{21}Sc$

★100. How many filled *p* orbitals do atoms of these elements contain?

a. carbon
b. phosphorus
c. oxygen
d. nitrogen

If You Have Trouble With . . .

Question	84	85	86	87	88	89	90	91	92	93	94	95	96	97	98	99	100
See Chapter	1	2	2	3	3	3	3	3	3	3	3	3	3	3	4	4	5

Standardized Test Prep

Select the choice that best answers each question or completes each statement.

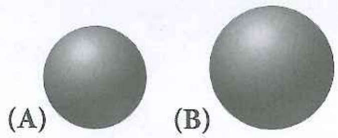
1. Which of the following properties increases as you move across a period from left to right?
   I. electronegativity
   II. ionization energy
   III. atomic radius

   (A) I and II only **(C)** II and III only
   (B) I and III only **(D)** I, II, and III

2. List the symbols for sodium, sulfur, and cesium in order of increasing atomic radii.
   (A) Na, S, Cs **(C)** S, Na, Cs
   (B) Cs, Na, S **(D)** Cs, S, Na

3. The electron configuration for an element in the halogen group should always end with
   (A) ns^2np^6. **(C)** ns^2np^4.
   (B) ns^2np^5. **(D)** ns^2np^2.

Use the spheres to answer Questions 4 and 5.



4. If the spheres represent a potassium atom and a potassium ion, which best represents the ion?

5. If the spheres represent an atom and an anion of the same element, which sphere represents the atom and which represents the anion?

Tips for Success

Interpreting Data Tables Tables can present a large amount of data in a small space. Before you try to answer questions based on a table, look at the table. Read the title, if there is one, and the column headings. Then read the questions. As you read each question, decide which data you will need to use to answer the question.

Use the data table to answer Questions 6–8.

Alkali metal	Atomic radius (pm)	First ionization energy (kJ/mol)	Electronegativity value
Li	152	520	1.0
Na	186	495.8	0.9
K	227	418.8	0.8
Rb	244	250	0.8
Cs	262	210	0.7

6. If you plot atomic radius versus first ionization energy, would the graph reveal a direct or inverse relationship?

7. If you plot atomic radius versus electronegativity, would the graph reveal a direct or inverse relationship?

8. If you plot first ionization energy versus electronegativity, would the graph reveal a direct or inverse relationship?

For each question there are two statements. Decide whether each statement is true or false. Then decide whether Statement II is a correct explanation for Statement I.

	Statement I		Statement II
9.	Electronegativity values are higher for metals than for nonmetals.	**BECAUSE**	Atoms of nonmetals are among the largest atoms.
10.	A calcium atom is larger than a calcium ion.	**BECAUSE**	Ions are always larger than the atoms from which they are formed.
11.	The element hydrogen is a metal.	**BECAUSE**	Hydrogen is on the left in the periodic table.
12.	Among all the elements in a period, the noble gas always has the smallest ionization energy.	**BECAUSE**	Within any period, atomic radii tend to decrease moving from right to left.

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10	11	12
See Lesson	6.3	6.3	6.2	6.3	6.3	6.3	6.3	6.3	6.3	6.3	6.1	6.3