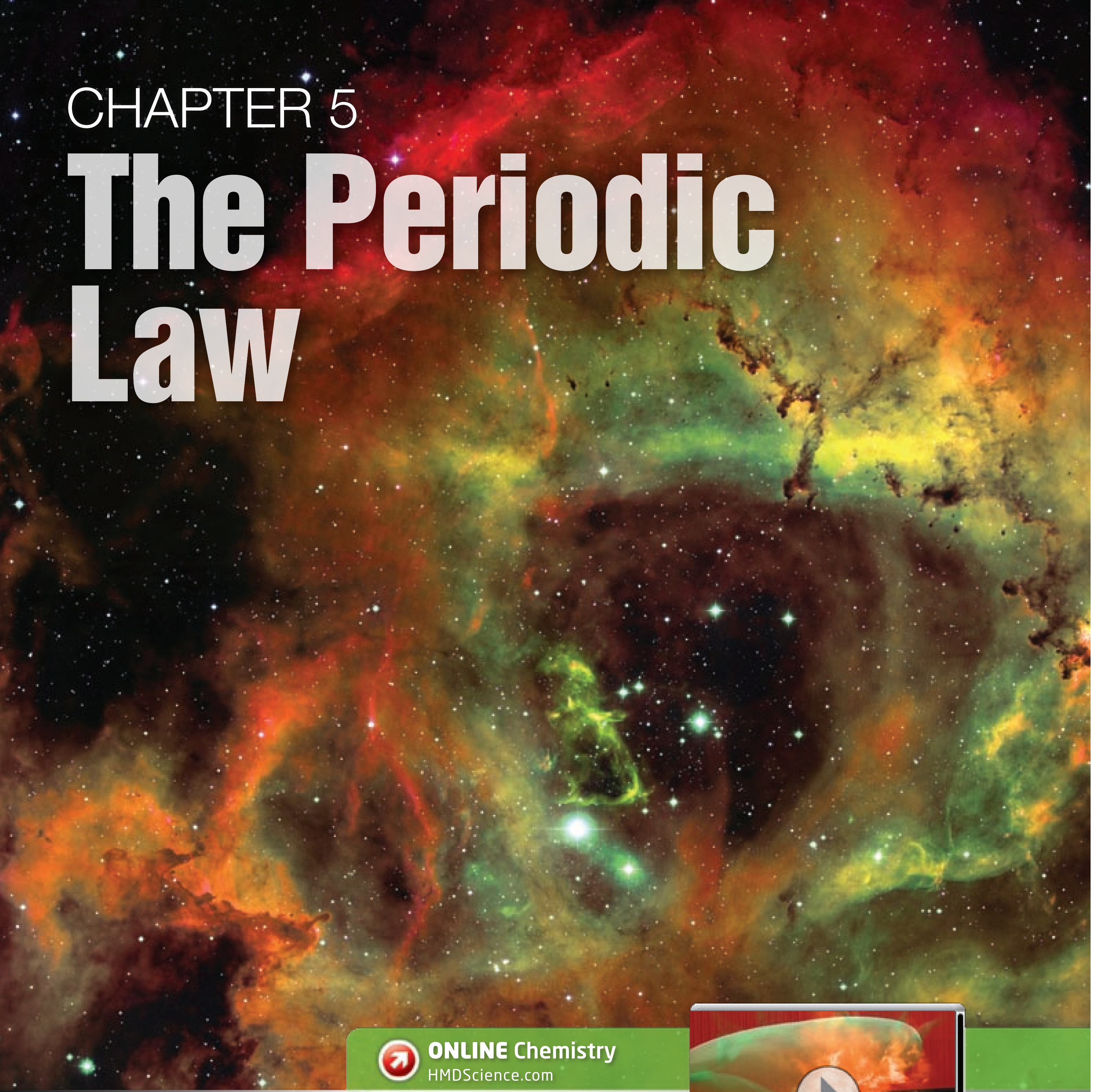


CHAPTER 5
The Periodic Law

ONLINE Chemistry
HMDScience.com

SECTION 1
History of the Periodic Table

SECTION 2
Electron Configuration and the Periodic Table

SECTION 3
Electron Configuration and Periodic Properties

ONLINE LABS include:
The Mendeleev Lab of 1869
Reactivity of Halide Ions
Periodicity of Properties of Oxides
Exploring the Periodic Table

PREMIUM CONTENT
Why It Matters Video
HMDScience.com
Periodic Law

SECTION 1

Main Ideas

- Mendeleev's periodic table grouped elements by their properties.
- Moseley arranged elements by their atomic numbers.
- Modern periodic tables arrange the elements by both atomic number and properties.

History of the Periodic Table

Key Terms

periodic law
periodic table
lanthanide
actinide

Imagine the confusion among chemists during the middle of the nineteenth century. By 1860, more than 60 elements had been discovered. Chemists had to learn the properties of these elements as well as those of the many compounds that they formed—a difficult task. And to make matters worse, there was no method for accurately determining an element's atomic mass or the number of atoms of an element in a particular chemical compound. Different chemists used different atomic masses for the same elements, resulting in different compositions being proposed for the same compounds. This made it nearly impossible for one chemist to understand the results of another.

In September 1860, a group of chemists assembled at the First International Congress of Chemists in Karlsruhe, Germany, to settle the issue of atomic mass as well as some other matters that were making communication difficult. At the Congress, Italian chemist Stanislao Cannizzaro presented a convincing method for accurately measuring the relative masses of atoms. Cannizzaro's method enabled chemists to agree on standard values for atomic mass and initiated a search for relationships between atomic mass and other properties of the elements.

VIRGINIA STANDARDS

CH.1.i The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include: construction and defense of a scientific viewpoint.
CH.1.EKS-26, CH.2.EKS-16

MAIN IDEA

Mendeleev's periodic table grouped elements by their properties.

When the Russian chemist Dmitri Mendeleev heard about the new atomic masses discussed at Karlsruhe, he decided to include the new values in a chemistry textbook he was writing. In the book, Mendeleev hoped to organize the elements according to their properties. He went about this much as you might organize information for a research paper. He placed the name of each known element on a card, together with the atomic mass of the element and a list of its observed physical and chemical properties. He then arranged the cards according to various properties and looked for trends or patterns.

Mendeleev noticed that when the elements were arranged in order of increasing atomic mass, certain similarities in their chemical properties appeared at regular intervals. Such a repeating pattern is referred to as *periodic*. The second hand of a watch, for example, passes over any given mark at periodic, 60-second intervals. The circular waves created by a drop of water hitting a water surface, as shown in **Figure 1.1**, are also periodic.

FIGURE 1.1

Periodic Patterns The regularly spaced water waves represent a simple periodic pattern.

CRITICAL THINKING

Relate Why are regularly-spaced water waves referred to as a periodic pattern?

FIGURE 1.2

Mendeleev's First Periodic Table In his first published periodic table, Mendeleev arranged the elements in vertical periods according to relative atomic mass. The atomic mass for each element is indicated by the number following the element's symbol. The unknown elements indicated by question marks at estimated atomic masses 45, 68, and 70 were later identified as scandium, Sc;gallium, Ga, and germanium, Ge.

но въ ней, мнѣ кажется, уже ясно выражается примѣнимость вы ставляемаго 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?		

а потому приходится въ разныхъ рядахъ имѣть различное измѣненіе разностей, чего нѣтъ въ главныхъ числахъ предлагаемой таблицы. Или же придется предполагать при составленіи системы очень много недостающихъ членовъ. То и другое мало выгодно. Мнѣ кажется притомъ, наиболѣе естественнымъ составить

Mendeleev created a table in which elements with similar properties were grouped together—a periodic table of the elements. His first periodic table, shown in **Figure 1.2**, was published in 1869. Note that Mendeleev placed iodine, I (atomic mass 127), after tellurium, Te (atomic mass 128). Although this contradicted the pattern of listing the elements in order of increasing atomic mass, it allowed Mendeleev to place tellurium in a group of elements with which it shares similar properties. Reading horizontally across Mendeleev's table, this group includes oxygen, O; sulfur, S; and selenium, Se. Iodine could also then be placed in the group it resembles chemically, which includes fluorine, F, chlorine, Cl, and bromine, Br.

Mendeleev's procedure left several empty spaces in his periodic table (see **Figure 1.2**). In 1871, the Russian chemist boldly predicted the existence and properties of the elements that would fill three of the spaces. By 1886, all three elements had been discovered. Today these elements are known as scandium, Sc, gallium, Ga; and germanium, Ge. Their properties are strikingly similar to those predicted by Mendeleev.

The success of Mendeleev's predictions persuaded most chemists to accept his periodic table and earned him credit as the discoverer of the periodic law. Two questions remained, however. (1) Why could most of the elements be arranged in the order of increasing atomic mass, but a few could not? (2) What was the reason for chemical periodicity?

MAIN IDEA

Moseley arranged elements by their atomic numbers.

The first question was not answered until more than 40 years after Mendeleev's first periodic table was published. In 1911, the English scientist Henry Moseley, who was working with Ernest Rutherford, examined the spectra of 38 different metals. When analyzing his data, Moseley discovered a previously unrecognized pattern. The elements in the periodic table fit into patterns better when they were arranged in increasing order according to nuclear charge, or the number of protons in the nucleus. Moseley's work led to both the modern definition of atomic number and the recognition that atomic number, not atomic mass, is the basis for the organization of the periodic table.

Moseley's discovery was consistent with Mendeleev's ordering of the periodic table by properties rather than strictly by atomic mass. For example, according to Moseley, tellurium, with an atomic number of 52, belongs before iodine, which has an atomic number of 53. Today, **Mendeleev's principle of chemical periodicity is correctly stated in what is known as the periodic law: The physical and chemical properties of the elements are periodic functions of their atomic numbers.** In other words, when the elements are arranged in order of increasing atomic number, elements with similar properties appear at regular intervals.

MAIN IDEA

Modern periodic tables arrange the elements by both atomic number and properties.

The periodic table has undergone extensive change since Mendeleev's time. Chemists have discovered new elements and, in more recent years, synthesized new ones in the laboratory. Each of the more than 40 new elements, however, can be placed in a group of other elements with similar properties. **The periodic table is an arrangement of the elements in order of their atomic numbers so that elements with similar properties fall in the same column, or group.**

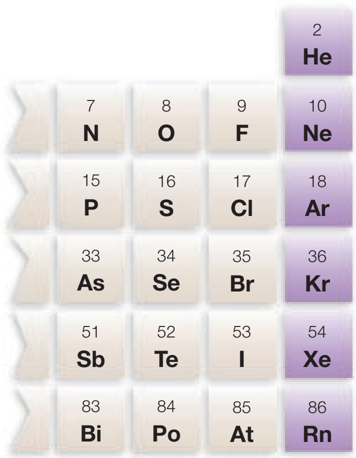
The Noble Gases

Perhaps the most significant addition to the periodic table came with the discovery of the noble gases. In 1894, English physicist John William Strutt (Lord Rayleigh) and Scottish chemist Sir William Ramsay discovered argon, Ar, a gas in the atmosphere that had previously escaped notice because of its total lack of chemical reactivity. Back in 1868 another noble gas, helium, He, had been discovered as a component of the sun. In 1895, Ramsay showed that helium also exists on Earth.

In order to fit argon and helium into the periodic table, Ramsay proposed a new group (see **Figure 1.3**). He placed this group between the groups now known as Group 17 (the fluorine family) and Group 1 (the lithium family). In 1898, Ramsay discovered two more noble gases, krypton, Kr, and xenon, Xe. The final noble gas, radon, Rn, was discovered in 1900 by the German scientist Friedrich Ernst Dorn.

FIGURE 1.3

Noble Gases The noble gases, also known as the Group 18 elements, are all rather unreactive. As you will read, the reason for this low reactivity also accounts for the special place occupied by the noble gases in the periodic table.



The Lanthanides

The next step in the development of the periodic table was completed in the early 1900s. It was then that the puzzling chemistry of the lanthanides was finally understood. **The lanthanides are the 14 elements with atomic numbers from 58 (cerium, Ce) to 71 (lutetium, Lu).** Because these elements are so similar in chemical and physical properties, the process of separating and identifying them was a tedious task that required the effort of many chemists.

The Actinides

Another major step in the development of the periodic table was the discovery of the actinides. **The actinides are the 14 elements with atomic numbers from 90 (thorium, Th) to 103 (lawrencium, Lr).** The lanthanides and actinides belong in Periods 6 and 7, respectively, of the periodic table, between the elements of Groups 3 and 4. To save space, the lanthanides and actinides are usually set off below the main portion of the periodic table.

QuickLAB — DESIGNING YOUR OWN PERIODIC TABLE

QUESTION

Can you design your own periodic table using information similar to that available to Mendeleev?

PROCEDURE

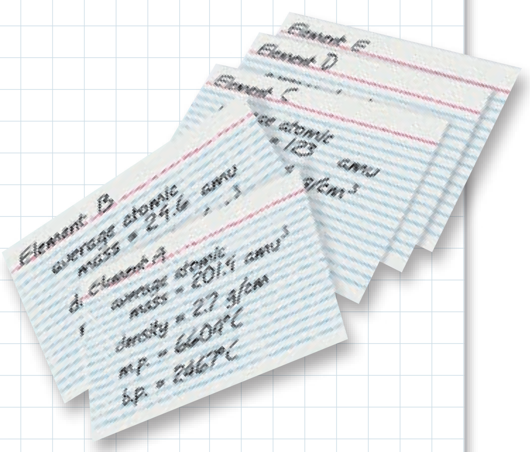
1. Write down the information available for each element on separate index cards. The following information is appropriate: a letter of the alphabet (A, B, C, etc.) to identify each element; atomic mass; state; density; melting point; boiling point; and any other readily observable physical properties. Do not write the name of the element on the index card, but keep a separate list indicating the letter you have assigned to each element.
2. Organize the cards for the elements in a logical pattern as you think Mendeleev might have done.

DISCUSSION

1. Keeping in mind that the information you have is similar to that available to Mendeleev in 1869, answer the following questions.
   a. Why are atomic masses given instead of atomic numbers?
   b. Can you identify each element by name?
2. How many groups of elements, or families, are in your periodic table? How many periods, or series, are in the table?
3. Predict the characteristics of any missing elements. When you have finished, check your work using your separate list of elements and a periodic table.

MATERIALS

- index cards



Periodicity

Periodicity with respect to atomic number can be observed in any group of elements in the periodic table. For example, consider the noble gases of Group 18. The first noble gas is helium, He. Helium has an atomic number of 2. The elements following helium in atomic number have completely different properties until the next noble gas, neon, Ne, is reached. Neon has an atomic number of 10. The remaining noble gases in order of increasing atomic number are argon (Ar, atomic number 18), krypton (Kr, atomic number 36), xenon (Xe, atomic number 54), and radon (Rn, atomic number 86). Thus, the differences in atomic number between successive noble gases are 8, 8, 18, 18, and 32, as shown in **Figure 1.4**.

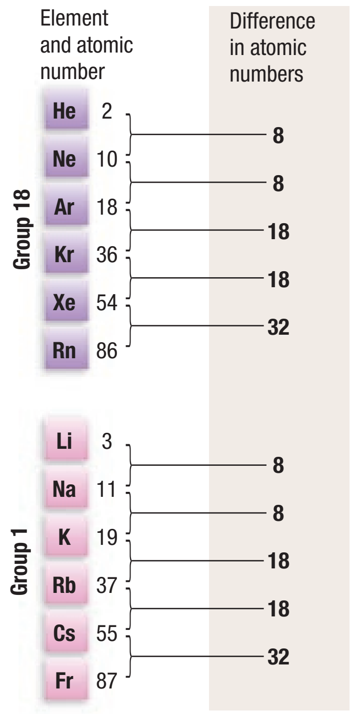
Also shown in **Figure 1.4** are atomic-number differences between the elements of Group 1. These elements are all solid, silvery metals. As you can see, the differences in atomic number between the Group 1 metals follow the same pattern as the differences in atomic number between the noble gases: 8, 8, 18, 18, and 32.

Starting with the first member of Groups 13–17, a similar periodic tern is repeated. The atomic number of each successive element is 8, 18, 18, and 32 higher than the atomic number of the element above it. In Section 2, you will see that the second mystery presented by Mendeleev's periodic table—the reason for periodicity—is explained by the arrangement of the electrons around the nucleus.

FIGURE 1.4

Patterns in the Periodic Table

In each of Groups 1 and 18, the differences between the atomic numbers of successive elements are 8, 8, 18, 18, and 32, respectively. Groups 2 and 13–17 follow a similar pattern.



SECTION 1 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. a. Who is credited with developing a method that led to the determination of standard relative atomic masses?

b. Who discovered the periodic law?

c. Who established atomic numbers as the basis for organizing the periodic table?

2. State the periodic law.

3. Name three sets of elements that have been added to the periodic table since Mendeleev's time.

4. How do the atomic numbers of the elements within each of Groups 1, 2, and 13–18 of the periodic table vary? (Refer to **Figure 1.4** as a guide.)

Critical Thinking

5. RELATING IDEAS Why are elements' atomic masses not in strict increasing order in the periodic table, even though the properties of the elements are similar? For example, by atomic mass, tellurium, Te, should be in group 17, and iodine, I, should be in Group 16, but grouping by properties has Te in Group 16 and I in Group 17.

SECTION 2

Main Ideas

- The period of an element is determined by its electron configuration.

VIRGINIA STANDARDS

CH.2 The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of:
CH.2.d families or groups.
CH.2.e periods.
CH.2.g electron configurations, valence electrons, and oxidation numbers.
CH.2.EKS-1; CH.2.EKS-8; CH.2.EKS-9; CH.2.EKS-11

Electron Configuration and the Periodic Table

Key Terms

alkali metals
alkaline-earth metals
transition elements
main-group elements
halogens

The Group 18 elements of the periodic table (the noble gases) undergo few chemical reactions. This stability results from the gases' special electron configurations. Helium's highest occupied level, the 1*s* orbital, is completely filled with electrons. And the highest occupied levels of the other noble gases contain stable octets. Generally, the electron configuration of an atom's highest occupied energy level governs the atom's chemical properties.

MAIN IDEA

The period of an element is determined by its electron configuration.

While the elements are arranged vertically in the periodic table in groups that share similar chemical properties, they are also organized horizontally in rows, or *periods*. (There are a total of seven periods of elements in the modern periodic table.) As can be seen in **Figure 2.1**, the length of each period is determined by the number of electrons that can occupy the sublevels being filled in that period.

FIGURE 2.1

RELATIONSHIP BETWEEN PERIOD LENGTH AND SUBLEVELS BEING FILLED IN THE PERIODIC TABLE

Period number	Number of elements in period	Sublevels in order of filling
1	2	1*s*
2	8	2*s* 2*p*
3	8	3*s* 3*p*
4	18	4*s* 3*d* 4*p*
5	18	5*s* 4*d* 5*p*
6	32	6*s* 4*f* 5*d* 6*p*
7	32	7*s* 5*f* 6*d* 7*p*

In the first period, the 1*s* sublevel is being filled. The 1*s* sublevel can hold a total of two electrons. Therefore, the first period consists of two elements—hydrogen and helium. In the second period, the 2*s* sublevel, which can hold two electrons, and the 2*p* sublevel, which can hold six electrons, are being filled. Consequently, the second period totals eight elements. Similarly, filling of the 3*s* and 3*p* sublevels accounts for the eight elements of the third period. Filling 3*d* and 4*d* sublevels in addition to the *s* and *p* sublevels adds 10 elements to both the fourth and fifth periods. Therefore, each of these periods totals 18 elements. Filling 4*f* sublevels in addition to *s, p,* and *d* sublevels adds 14 elements to the sixth period, which totals 32 elements. And as new elements are created, the 25 named elements in Period 7 could, in theory, be extended to 32.

The period of an element can be determined from the element's electron configuration. For example, arsenic, As, has the electron configuration $[Ar]3d^{10}4s^24p^3$. The 4 in $4p^3$ indicates that arsenic's highest occupied energy level is the fourth energy level. Arsenic is thus in the fourth period in the periodic table. The period and electron configuration for each element can be found in the periodic table (on the next two pages). Based on the electron configurations of the elements, the periodic table can be divided into four blocks, the *s, p, d,* and *f* blocks. This division is illustrated in **Figure 2.2.** The name of each block is determined by whether an *s, p, d,* or *f* sublevel is being filled in successive elements of that block.

FIGURE 2.2

Sublevel Blocks Based on the electron configurations of the elements, the periodic table can be subdivided into four sublevel blocks.

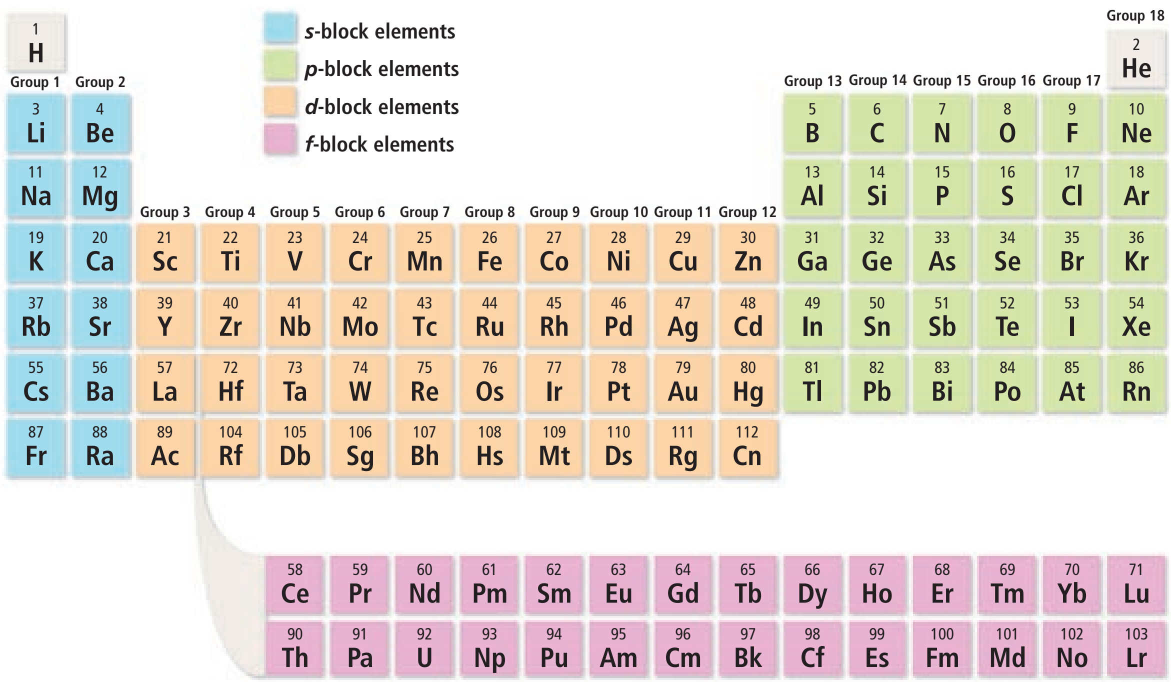


FIGURE 2.3

The Periodic Table of the Elements In the modern periodic table, the elements are arranged by atomic number and form vertical groups and horizontal periods.

Key:

- Atomic number: 13
- Symbol: Al
- Name: Aluminum
- Average atomic mass: 26.981 5386
- Electron configuration: [Ne]$3s^2 3p^1$

Period	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9
1	1 **H** Hydrogen ***1.008*** $1s^1$								
2	3 **Li** Lithium ***6.94*** [He]$2s^1$	4 **Be** Beryllium 9.012 182 [He]$2s^2$							
3	11 **Na** Sodium 22.989 769 28 [Ne]$3s^1$	12 **Mg** Magnesium 24.3050 [Ne]$3s^2$							
4	19 **K** Potassium 39.0983 [Ar]$4s^1$	20 **Ca** Calcium 40.078 [Ar]$4s^2$	21 **Sc** Scandium 44.955 912 [Ar]$3d^1 4s^2$	22 **Ti** Titanium 47.867 [Ar]$3d^2 4s^2$	23 **V** Vanadium 50.9415 [Ar]$3d^3 4s^2$	24 **Cr** Chromium 51.9961 [Ar]$3d^5 4s^1$	25 **Mn** Manganese 54.938 045 [Ar]$3d^5 4s^2$	26 **Fe** Iron 55.845 [Ar]$3d^6 4s^2$	27 **Co** Cobalt 58.933 195 [Ar]$3d^7 4s^2$
5	37 **Rb** Rubidium 85.4678 [Kr]$5s^1$	38 **Sr** Strontium 87.62 [Kr]$5s^2$	39 **Y** Yttrium 88.905 85 [Kr]$4d^1 5s^2$	40 **Zr** Zirconium 91.224 [Kr]$4d^2 5s^2$	41 **Nb** Niobium 92.906 38 [Kr]$4d^4 5s^1$	42 **Mo** Molybdenum 95.94 [Kr]$4d^5 5s^1$	43 **Tc** Technetium (98) [Kr]$4d^6 5s^1$	44 **Ru** Ruthenium 101.07 [Kr]$4d^7 5s^1$	45 **Rh** Rhodium 102.905 50 [Kr]$4d^8 5s^1$
6	55 **Cs** Cesium 132.905 4519 [Xe]$6s^1$	56 **Ba** Barium 137.327 [Xe]$6s^2$	57 **La** Lanthanum 138.905 47 [Xe]$5d^1 6s^2$	72 **Hf** Hafnium 178.49 [Xe]$4f^{14} 5d^2 6s^2$	73 **Ta** Tantalum 180.947 88 [Xe]$4f^{14} 5d^3 6s^2$	74 **W** Tungsten 183.84 [Xe]$4f^{14} 5d^4 6s^2$	75 **Re** Rhenium 186.207 [Xe]$4f^{14} 5d^5 6s^2$	76 **Os** Osmium 190.23 [Xe]$4f^{14} 5d^6 6s^2$	77 **Ir** Iridium 192.217 [Xe]$4f^{14} 5d^7 6s^2$
7	87 **Fr** Francium (223) [Rn]$7s^1$	88 **Ra** Radium (226) [Rn]$7s^2$	89 **Ac** Actinium (227) [Rn]$6d^1 7s^2$	104 **Rf** Rutherfordium (261) [Rn]$5f^{14} 6d^2 7s^2$	105 **Db** Dubnium (262) [Rn]$5f^{14} 6d^3 7s^2$	106 **Sg** Seaborgium (266) [Rn]$5f^{14} 6d^4 7s^2$	107 **Bh** Bohrium (264) [Rn]$5f^{14} 6d^5 7s^2$	108 **Hs** Hassium (277) [Rn]$5f^{14} 6d^6 7s^2$	109 **Mt** Meitnerium (268) [Rn]$5f^{14} 6d^7 7s^2$

58 **Ce** Cerium 140.116 [Xe]$4f^1 5d^1 6s^2$	59 **Pr** Praseodymium 140.907 65 [Xe]$4f^3 6s^2$	60 **Nd** Neodymium 144.242 [Xe]$4f^4 6s^2$	61 **Pm** Promethium (145) [Xe]$4f^5 6s^2$	62 **Sm** Samarium 150.36 [Xe]$4f^6 6s^2$
90 **Th** Thorium 232.038 06 [Rn]$6d^2 7s^2$	91 **Pa** Protactinium 231.035 88 [Rn]$5f^2 6d^1 7s^2$	92 **U** Uranium 238.028 91 [Rn]$5f^3 6d^1 7s^2$	93 **Np** Neptunium (237) [Rn]$5f^4 6d^1 7s^2$	94 **Pu** Plutonium (244) [Rn]$5f^6 7s^2$

* The systematic names and symbols for elements greater than 112 will be used until the approval of trivial names by IUPAC.

Elements whose average atomic masses appear bolded and italicized are recognized by the International Union of Pure and Applied Chemistry (IUPAC) to have several stable isotopes. Thus, the average atomic mass for each of these elements is officially expressed as a range of values. A range of values expresses that the average atomic mass of a sample of one of these elements is not a constant in nature but varies depending on the physical, chemical, and nuclear history of the material in which the sample is found. However, the values in this table are appropriate for everyday calculations. A value given in parentheses is not an average atomic mass but is the mass number of that element's most stable or most common isotope.

- Hydrogen
- Semiconductors (also known as metalloids)

Metals
- Alkali metals
- Alkaline-earth metals
- Transition metals
- Other metals

Nonmetals
- Halogens
- Noble gases
- Other nonmetals

Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
								2 **He** Helium 4.002 602 $1s^2$
			5 **B** Boron ***10.81*** [He]$2s^22p^1$	6 **C** Carbon ***12.011*** [He]$2s^22p^2$	7 **N** Nitrogen ***14.007*** [He]$2s^22p^3$	8 **O** Oxygen ***15.999*** [He]$2s^22p^4$	9 **F** Fluorine 18.998 4032 [He]$2s^22p^5$	10 **Ne** Neon 20.1797 [He]$2s^22p^6$
			13 **Al** Aluminum 26.981 5386 [Ne]$3s^23p^1$	14 **Si** Silicon ***28.085*** [Ne]$3s^23p^2$	15 **P** Phosphorus 30.973 762 [Ne]$3s^23p^3$	16 **S** Sulfur ***32.06*** [Ne]$3s^23p^4$	17 **Cl** Chlorine ***35.45*** [Ne]$3s^23p^5$	18 **Ar** Argon 39.948 [Ne]$3s^23p^6$
28 **Ni** Nickel 58.6934 [Ar]$3d^84s^2$	29 **Cu** Copper 63.546 [Ar]$3d^{10}4s^1$	30 **Zn** Zinc 65.409 [Ar]$3d^{10}4s^2$	31 **Ga** Gallium 69.723 [Ar]$3d^{10}4s^24p^1$	32 **Ge** Germanium 72.63 [Ar]$3d^{10}4s^24p^2$	33 **As** Arsenic 74.921 60 [Ar]$3d^{10}4s^24p^3$	34 **Se** Selenium 78.96 [Ar]$3d^{10}4s^24p^4$	35 **Br** Bromine 79.904 [Ar]$3d^{10}4s^24p^5$	36 **Kr** Krypton 83.798 [Ar]$3d^{10}4s^24p^6$
46 **Pd** Palladium 106.42 [Kr]$4d^{10}$	47 **Ag** Silver 107.8682 [Kr]$4d^{10}5s^1$	48 **Cd** Cadmium 112.411 [Kr]$4d^{10}5s^2$	49 **In** Indium 114.818 [Kr]$4d^{10}5s^25p^1$	50 **Sn** Tin 118.710 [Kr]$4d^{10}5s^25p^2$	51 **Sb** Antimony 121.760 [Kr]$4d^{10}5s^25p^3$	52 **Te** Tellurium 127.60 [Kr]$4d^{10}5s^25p^4$	53 **I** Iodine 126.904 47 [Kr]$4d^{10}5s^25p^5$	54 **Xe** Xenon 131.293 [Kr]$4d^{10}5s^25p^6$
78 **Pt** Platinum 195.084 [Xe]$4f^{14}5d^96s^1$	79 **Au** Gold 196.966 569 [Xe]$4f^{14}5d^{10}6s^1$	80 **Hg** Mercury 200.59 [Xe]$4f^{14}5d^{10}6s^2$	81 **Tl** Thallium ***204.38*** [Xe]$4f^{14}5d^{10}6s^26p^1$	82 **Pb** Lead 207.2 [Xe]$4f^{14}5d^{10}6s^26p^2$	83 **Bi** Bismuth 208.980 40 [Xe]$4f^{14}5d^{10}6s^26p^3$	84 **Po** Polonium (209) [Xe]$4f^{14}5d^{10}6s^26p^4$	85 **At** Astatine (210) [Xe]$4f^{14}5d^{10}6s^26p^5$	86 **Rn** Radon (222) [Xe]$4f^{14}5d^{10}6s^26p^6$
110 **Ds** Darmstadtium (271) [Rn]$5f^{14}6d^97s^1$	111 **Rg** Roentgenium (272) [Rn]$5f^{14}6d^{10}7s^1$	112 **Cn** Copernicium (285) [Rn]$5f^{14}6d^{10}7s^2$	113 **Uut*** Ununtrium (284) [Rn]$5f^{14}6d^{10}7s^27p^1$	114 **Uuq*** Ununquadium (289) [Rn]$5f^{14}6d^{10}7s^27p^2$	115 **Uup*** Ununpentium (288) [Rn]$5f^{14}6d^{10}7s^27p^3$	116 **Uuh*** Ununhexium (292) [Rn]$5f^{14}6d^{10}7s^27p^4$	117 **Uus*** Ununseptium (294) [Rn]$5f^{14}6d^{10}7s^27p^5$	118 **Uuo*** Ununoctium (294) [Rn]$5f^{14}6d^{10}7s^27p^6$

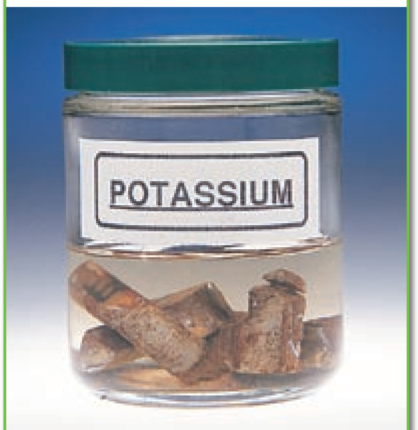
The discoveries of elements with atomic numbers 113–118 have been reported but not fully confirmed.

63 **Eu** Europium 151.964 [Xe]$4f^76s^2$	64 **Gd** Gadolinium 157.25 [Xe]$4f^75d^16s^2$	65 **Tb** Terbium 158.925 35 [Xe]$4f^96s^2$	66 **Dy** Dysprosium 162.500 [Xe]$4f^{10}6s^2$	67 **Ho** Holmium 164.930 32 [Xe]$4f^{11}6s^2$	68 **Er** Erbium 167.259 [Xe]$4f^{12}6s^2$	69 **Tm** Thulium 168.934 21 [Xe]$4f^{13}6s^2$	70 **Yb** Ytterbium 173.04 [Xe]$4f^{14}6s^2$	71 **Lu** Lutetium 174.967 [Xe]$4f^{14}5d^16s^2$
95 **Am** Americium (243) [Rn]$5f^77s^2$	96 **Cm** Curium (247) [Rn]$5f^76d^17s^2$	97 **Bk** Berkelium (247) [Rn]$5f^97s^2$	98 **Cf** Californium (251) [Rn]$5f^{10}7s^2$	99 **Es** Einsteinium (252) [Rn]$5f^{11}7s^2$	100 **Fm** Fermium (257) [Rn]$5f^{12}7s^2$	101 **Md** Mendelevium (258) [Rn]$5f^{13}7s^2$	102 **No** Nobelium (259) [Rn]$5f^{14}7s^2$	103 **Lr** Lawrencium (262) [Rn]$5f^{14}6d^17s^2$

FIGURE 2.4

Group 1: Alkali Metals

(a) Like other alkali metals, potassium reacts strongly with water.



(b) Potassium must be stored in kerosene or oil to prevent it from reacting with moisture in the air.

The *S*-Block Elements: Groups 1 and 2

The elements of the *s*-block are chemically reactive metals. The Group 1 metals are more reactive than those of Group 2. The outermost energy level in an atom of each Group 1 element contains a single s electron. For example, the configurations of lithium and sodium are $[He]2s^1$ and $[Ne]3s^1$, respectively. As you will learn in Section 3, the ease with which the single electron is lost helps make the Group 1 metals extremely reactive. Using *n* for the number of the highest occupied energy level, the outer, or group, configurations of the Group 1 and 2 elements are written ns^1 and ns^2, respectively. For example, the configuration of Na is $[Ne]3s^1$, so the group configuration is written ns^1, where $n = 3$.

The elements of Group 1 of the periodic table (lithium, sodium, potassium, rubidium, cesium, and francium) are known as the alkali metals. In their pure state, all of the alkali metals have a silvery appearance and are soft enough to cut with a knife. However, because they are so reactive, alkali metals are not found in nature as free elements. They combine vigorously with most nonmetals. And they react strongly with water to produce hydrogen gas and aqueous solutions of substances known as alkalis. Because of their extreme reactivity with air or moisture, alkali metals are usually stored in kerosene. Proceeding down the column, the elements of Group 1 melt at successively lower temperatures.

The elements of Group 2 of the periodic table (beryllium, magnesium, calcium, strontium, barium, and radium) are called the alkaline-earth metals. Atoms of alkaline-earth metals contain a pair of electrons in their outermost *s* sublevel. Consequently, the group configuration for Group 2 is ns^2. The Group 2 metals are harder, denser, and stronger than the alkali metals. They also have higher melting points. Although they are less reactive than the alkali metals, the alkaline-earth metals are also too reactive to be found in nature as free elements.

FIGURE 2.5

Group 2: Alkaline-Earth Metals

(a) Calcium, an alkaline-earth metal, is too reactive to be found in nature in its pure state.

(b) Instead, it exists in compounds, such as in the minerals that make up marble.

Hydrogen and Helium

Before discussing the other blocks of the periodic table, let's consider two special cases in the classification of the elements—hydrogen and helium. Hydrogen has an electron configuration of $1s^1$, but despite the ns^1 configuration, it does not share the same properties as the elements of Group 1. Although it is located above the Group 1 elements in many periodic tables, hydrogen is a unique element, with properties that do not closely resemble those of any group.

Like the Group 2 elements, helium has an ns^2 group configuration. Yet it is part of Group 18. Because its highest occupied energy level is filled by two electrons, helium possesses special chemical stability, exhibiting the unreactive nature of a Group 18 element. By contrast, the Group 2 metals have no special stability; their highest occupied energy levels are not filled because each metal has an empty available *p* sublevel.

CHECK FOR UNDERSTANDING

Apply Which is more important in determining an element's group: the electron configuration or the element's properties? Explain.

The Periodic Table and Electron Configurations

PREMIUM CONTENT

Learn It! Video
HMDScience.com

Solve It! Cards
HMDScience.com

Sample Problem A **(a) Without looking at the periodic table, identify the group, period, and block in which the element that has the electron configuration [Xe]$6s^2$ is located. (b) Without looking at the periodic table, write the electron configuration for the Group 1 element in the third period. Is this element likely to be more reactive or less reactive than the element described in (a)?**

SOLVE

a. The element is in Group 2, as indicated by the group configuration of ns^2. It is in the sixth period, as indicated by the highest principal quantum number in its configuration, 6. The element is in the s-block.

b. In a third-period element, the highest occupied energy level is the third main energy level, $n = 3$. The $1s$, $2s$, and $2p$ sublevels are completely filled (see **Figure 2.1**). A Group 1 element has a group configuration of ns^1, which indicates a single electron in its highest *s* sublevel. Therefore, this element has the following configuration:

$$1s^22s^22p^63s^1 \quad \text{or} \quad [\text{Ne}]3s^1$$

Because it is in Group 1 (the alkali metals), this element is likely to be more reactive than the element described in (a), which is in Group 2 (the alkaline-earth metals).

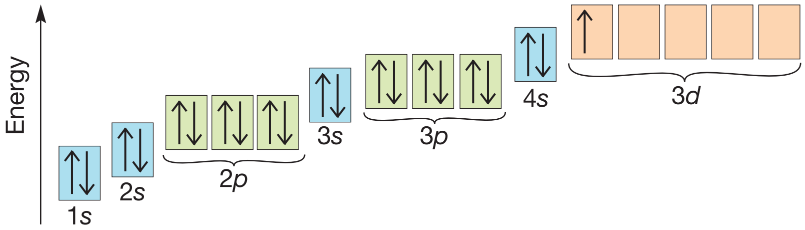
Practice Answers in Appendix E

1. Without looking at the periodic table, identify the group, period, and block in which the element that has the electron configuration [Kr]$5s^1$ is located.
2. **a.** Without looking at the periodic table, write the group configuration for the Group 2 elements.
   b. Without looking at the periodic table, write the complete electron configuration for the Group 2 element in the fourth period.
   c. Refer to **Figure 2.3** to identify the element described in (b). Then, write the element's noble-gas notation.

FIGURE 2.6

Group 3: Electron Configuration

The diagram shows the electron configuration of scandium, Sc, the Group 3 element of the fourth period. In general, the (n–1)d sublevel in Groups 3–12 is occupied by electrons after the ns sublevel is filled.



The *d*-Block Elements: Groups 3–12

For energy level n, there are n possible sublevels, so the d sublevel first appears when $n = 3$. This $3d$ sublevel is slightly higher in energy than the $4s$ sublevel, so these are filled in the order $4s3d$ (see **Figure 2.6**). This order of filling is also seen for higher values of n. Each d sublevel consists of five orbitals with a maximum of two electrons each, or up to 10 electrons possible in each d sublevel. In addition to the two ns electrons of Group 2, atoms of the Group 3 elements each have one electron in the d sublevel of the (n–1) energy level. The group configuration for Group 3 is therefore $(n–1)d^1ns^2$. Atoms of the Group 12 elements have 10 electrons in the d sublevel plus two electrons in the ns sublevel. The group configuration for Group 12 is $(n–1)d^{10}ns^2$.

Some deviations from orderly d sublevel filling occur in Groups 4–11. As a result, elements in these d-block groups, unlike those in s-block and p-block groups, do not necessarily have identical outer electron configurations. For example, in Group 10, nickel, Ni, has the electron configuration $[Ar]3d^84s^2$. Palladium, Pd, has the configuration $[Kr]4d^{10}5s^0$. And platinum, Pt, has the configuration $[Xe]4f^{14}5d^96s^1$. Notice, however, that in each case the sum of the outer s and d electrons is equal to the group number.

The *d*-block elements are metals with typical metallic properties and are often referred to as transition elements. They are good conductors of electricity and have a high luster. They are typically less reactive than the alkali metals and the alkaline-earth metals. Some are so nonreactive that they do not easily form compounds and exist in nature as free elements. Palladium, platinum, and gold are among the least reactive of all the elements. Some d-block elements are shown in **Figure 2.7**.

CHECK FOR UNDERSTANDING

Explain The word *transition* refers to change. What would be a reasonable explanation for referring to the *d*-block elements as transition elements?

FIGURE 2.7

Transition Elements

Mercury

Tungsten

Vanadium

The Periodic Table and Electron Configurations

Sample Problem B **An element has the electron configuration [Kr]$4d^5 5s^1$. Without looking at the periodic table, identify the period, block, and group in which this element is located. Then, consult the periodic table to identify this element and the others in its group.**

SOLVE

The number of the highest occupied energy level is 5, so the element is in the fifth period. There are five electrons in the *d* sublevel, which means that it is incompletely filled. The *d* sublevel can hold 10 electrons. Therefore, the element is in the d-block. For *d*-block elements, the number of electrons in the *ns* sublevel (1) plus the number of electrons in the $(n–1)d$ sublevel (5) equals the group number, 6. This Group 6 element is molybdenum. The others in Group 6 are chromium, tungsten, and seaborgium.

Practice Answers in Appendix E

1. Without looking at the periodic table, identify the period, block, and group in which the element that has the electron configuration [Ar]$3d^8 4s^2$ is located.
2. Without looking at the periodic table, write the outer electron configuration for the Group 12 element in the fifth period.

The *p*-Block Elements: Groups 13–18

The *p*-block elements consist of all the elements of Groups 13–18 except helium. Electrons add to a *p* sublevel only after the *s* sublevel in the same energy level is filled. Therefore, atoms of all *p*-block elements contain two electrons in the *ns* sublevel. **The *p*-block elements together with the *s*-block elements are called the main-group elements.** For Group 13 elements, the added electron enters the *np* sublevel, giving a group configuration of ns^2np^1. Atoms of Group 14 elements contain two electrons in the *p* sublevel, giving ns^2np^2 for the group configuration. This pattern continues in Groups 15–18. In Group 18, the stable noble-gas configuration of ns^2np^6 is reached. The relationships among group numbers and electron configurations for all the groups are summarized in **Figure 2.8,** on the next page.

For atoms of *p*-block elements, the total number of electrons in the highest occupied level is equal to the group number minus 10. For example, bromine is in Group 17. It has $17 - 10 = 7$ electrons in its highest energy level. Because atoms of *p*-block elements contain two electrons in the *ns* sublevel, we know that bromine has five electrons in its outer *p* sublevel. The electron configuration of bromine is [Ar]$3d^{10}4s^24p^5$.

The properties of elements of the *p*-block vary greatly. At its right-hand end, the *p*-block includes all of the *nonmetals* except hydrogen and helium. All six of the *metalloids* (boron, silicon, germanium, arsenic, antimony, and tellurium) are also in the *p*-block. At the left-hand side and bottom of the block, there are eight *p*-block metals. The locations of the nonmetals, metalloids, and metals in the *p*-block are shown with distinctive colors in **Figure 2.3.**

FIGURE 2.8

RELATIONSHIPS AMONG GROUP NUMBERS, BLOCKS, AND ELECTRON CONFIGURATIONS

Group number	Group configuration	Block	Comments
1, 2	$ns^{1, 2}$	*s*	One or two electrons in *ns* sublevel
3–12	$(n{-}1)d^{1-10}ns^{0-2}$	*d*	Sum of electrons in *ns* and $(n-1)d$ levels equals group number
13–18	ns^2np^{1-6}	*p*	Number of electrons in *np* sublevel equals group number minus 12

The elements of Group 17 (fluorine, chlorine, bromine, iodine, and astatine) are known as the halogens (See **Figure 2.9).** The halogens are the most reactive nonmetals. They react vigorously with most metals to form salts. As you will see, the reactivity of the halogens is based on the presence of seven electrons in their outer energy levels—one electron short of the stable noble-gas configuration. Fluorine and chlorine are gases at room temperature, bromine is a reddish liquid, and iodine is a dark purple solid. Astatine is a synthetic element prepared in very small quantities. Most of its properties are estimated, although it is known to be a solid.

The metalloids, or semiconducting elements, are located between nonmetals and metals in the *p*-block. They are mostly brittle solids with some properties of metals and some of nonmetals. The metalloid elements have electrical conductivity intermediate between that of metals, which are good conductors, and nonmetals, which are nonconductors.

The metals of the *p*-block are generally harder and denser than the *s*-block alkaline-earth metals, but softer and less dense than the *d*-block metals. With the exception of bismuth, these metals are sufficiently reactive to be found in nature only in the form of compounds. Once obtained as free metals, however, they are stable in the presence of air.

FIGURE 2.9

The Halogens Fluorine, chlorine, bromine, and iodine are members of Group 17 of the periodic table, also known as the halogens. Locate the halogens in the *p*-block of the periodic table.

Fluorine **Chlorine** **Bromine** **Iodine**

The Periodic Table and Electron Configurations

Sample Problem C **Without looking at the periodic table, write the outer electron configuration for the Group 14 element in the second period. Then, name the element and identify it as a metal, nonmetal, or metalloid.**

SOLVE

The group number is higher than 12, so the element is in the *p*-block. The total number of electrons in the highest occupied *s* and *p* sublevels is therefore equal to the group number minus 10 ($14 - 10 = 4$). Two electrons are in the *s* sublevel, so two electrons must also be present in the 2*p* sublevel, which means that the outer electron configuration is $2s^22p^2$. The element is carbon, C, which is a nonmetal.

Practice Answers in Appendix E

1. **a.** Without looking at the periodic table, write the outer electron configuration for the Group 17 element in the third period.
   b. Name the element described in (a), and identify it as a metal, nonmetal, or metalloid.
2. **a.** Without looking at the periodic table, identify the period, block, and group of an element that has the electron configuration $[Ar]3d^{10}4s^24p^3$.
   b. Name the element described in (a), and identify it as a metal, nonmetal, or metalloid.

The *f*-Block Elements: Lanthanides and Actinides

In the periodic table, the *f*-block elements are wedged between Groups 3 and 4 in the sixth and seventh periods. The position of these inner transition elements reflects the fact that they involve the filling of the 4*f* sublevel. With seven 4*f* orbitals to be filled with two electrons each, there are a total of 14 *f*-block elements between lanthanum, La, and hafnium, Hf, in the sixth period. The lanthanides are shiny metals similar in reactivity to the Group 2 alkaline-earth metals.

There are also 14 *f*-block elements, the actinides, between actinium, Ac, and element 104, Rf, in the seventh period. In these elements the 5*f* sublevel is being filled with 14 electrons. The actinides are all radioactive. The first four actinides (thorium, Th, through neptunium, Np) have been found naturally on Earth. The remaining actinides are known only as laboratory-made elements.

The Periodic Table and Electron Configurations

Sample Problem D **The electron configurations of atoms of four elements are written below. Name the block and group in which each of these elements is located in the periodic table. Then, use the periodic table to name each element. Identify each element as a metal, nonmetal, or metalloid.**

Finally, describe whether each element has high reactivity or low reactivity.

a. $[Xe]4f^{14}5d^96s^1$ **c.** $[Ne]3s^23p^6$
b. $[Ne]3s^23p^5$ **d.** $[Xe]4f^66s^2$

Continued

SOLVE

a. The 4*f* sublevel is filled with 14 electrons. The 5*d* sublevel is partially filled with nine electrons. Therefore, this element is in the *d*-block. The element is the transition metal platinum, Pt, which is in Group 10 and has a low reactivity.

b. The incompletely filled *p* sublevel shows that this element is in the *p*-block. A total of seven electrons are in the *ns* and *np* sublevels, so this element is in Group 17, the halogens. The element is chlorine, Cl, and is highly reactive.

c. This element has a noble-gas configuration and thus is in Group 18 in the *p*-block. The element is argon, Ar, which is a nonreactive nonmetal and a noble gas.

d. The incomplete 4*f* sublevel shows that the element is in the *f*-block and is a lanthanide. Group numbers are not assigned to the *f*-block. The element is samarium, Sm. All of the lanthanides are reactive metals.

Practice Answers in Appendix E

1. For each of the following, identify the block, period, group, group name (where appropriate), element name, element type (metal, nonmetal, or metalloid), and relative reactivity (high or low):
   a. $[He]2s^22p^5$ **b.** $[Ar]3d^{10}4s^1$

SECTION 2 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. To illustrate the relationship between the elements' electron configurations and their placement in the periodic table, into what four blocks can the periodic table be divided?
2. What name is given to each of the following groups of elements in the periodic table?
   a. Group 1
   b. Group 2
   c. Groups 3–12
   d. Group 17
   e. Group 18
3. What are the relationships between group configuration and group number for elements in the *s, p,* and *d*-blocks?
4. Without looking at the periodic table, write the outer electron configuration for the Group 15 element in the fourth period.
5. Without looking at the periodic table, identify the period, block, and group of the element that has the electron configuration $[Ar]3d^74s^2$.

Critical Thinking

6. **APPLYING MODELS** Period 7 contains elements in the *s, p, d,* and *f*-blocks. Suppose that there were a Period 8 and it contained elements in the "g" block, where "g" had the angular momentum quantum number $\ell = 4$. If a hypothetical element in Period 8 had an atomic number of 120, into what group in the periodic table would the element fit, and what properties might it have (assuming it does not radioactively decay)?

SECTION 3

Main Ideas

- Atomic radii are related to electron configuration.
- Removing electrons from atoms to form ions requires energy.
- Adding electrons to atoms to form ions also requires energy.
- When atoms become ions, their radii change.
- Only the outer electrons are involved in forming compounds.
- Atoms have different abilities to capture electrons.
- The properties of d-block metals do not vary much.

VIRGINIA STANDARDS

CH.2 The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of:
CH.2.f trends including atomic radii, electronegativity, shielding effect, and ionization energy.
CH.2.g electron configurations, valence electrons, and oxidation numbers.
CH.2.EKS-9; CH.2.EKS-10; CH.2.EKS-12

PREMIUM CONTENT
Animated Chemistry
HMDScience.com
Period Trends (Interaction)

Electron Configuration and Periodic Properties

Key Terms

atomic radius	ionization energy	anion
ion	electron affinity	valence electron
ionization	cation	electronegativity

So far, you have learned that the elements are arranged in the periodic table according to their atomic number and that there is a rough correlation between the arrangement of the elements and their electron configurations. In this section, the relationship between the periodic law and electron configurations will be further explored.

MAIN IDEA

Atomic radii are related to electron configuration.

Ideally, the size of an atom is defined by the edge of its orbital. However, this boundary is fuzzy and varies under different conditions. Therefore, the conditions under which the atom exists must be specified to estimate its size. One way to express an atom's radius is to measure the distance between the nuclei of two identical atoms that are chemically bonded together and then divide this distance by two. **Atomic radius may be defined as one-half the distance between the nuclei of identical atoms that are bonded together.** This can be seen in **Figure 3.1** on the next page.

Period Trends

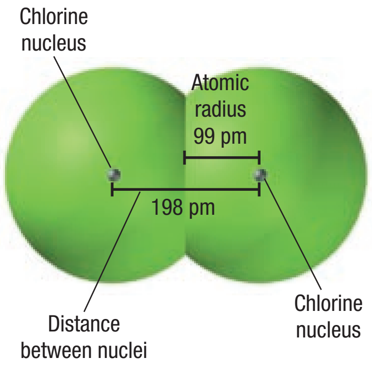
Figure 3.2 gives the atomic radii of the elements, and **Figure 3.3** (on the next spread) presents this information graphically. Note that there is a gradual decrease in atomic radii across the second period from lithium, Li, to neon, Ne. *The trend to smaller atoms across a period is caused by the increasing positive charge of the nucleus.* As electrons add to *s* and *p* sublevels in the same main energy level, they are gradually pulled closer to the more highly charged nucleus. This increased pull results in a decrease in atomic radii. The attraction of the nucleus is somewhat offset by repulsion among the increased number of electrons in the same outer energy level. As a result, the difference in radii between neighboring atoms in each period grows smaller, as shown in **Figure 3.2.**

Group Trends

Examine the atomic radii of the Group 1 elements in **Figure 3.2.** Notice that the radii of the elements increase as you read down the group.

FIGURE 3.1

Atomic Radii One method of determining atomic radius is to measure the distance between the nuclei of two identical atoms that are bonded together in an element or compound and then divide this distance by two. The atomic radius of a chlorine atom, for example, is 100 picometers (pm).



As electrons occupy sublevels in successively higher main energy levels farther from the nucleus, the sizes of the atoms increase. *In general, the atomic radii of the main-group elements increase down a group.*

Now examine the radii of the Group 13 elements. Although gallium, Ga, follows aluminum, Al, it has a slightly smaller atomic radius than does aluminum. This is because gallium, unlike aluminum, is preceded in its period by the 10 *d*-block elements. The expected increase in gallium's radius caused by the filling of the fourth main-energy level is outweighed by a shrinking of the electron cloud caused by a nuclear charge that is considerably higher than that of aluminum.

CHECK FOR UNDERSTANDING

Explain Why is it necessary for the radii of atoms to be expressed as the distance between the nuclei of two identical atoms bonded together? (Refer to **Figure 3.1.**)

FIGURE 3.2

Decreasing Radii Atomic radii decrease from left to right across a period and increase down a group. Values are given in picometers (pm).

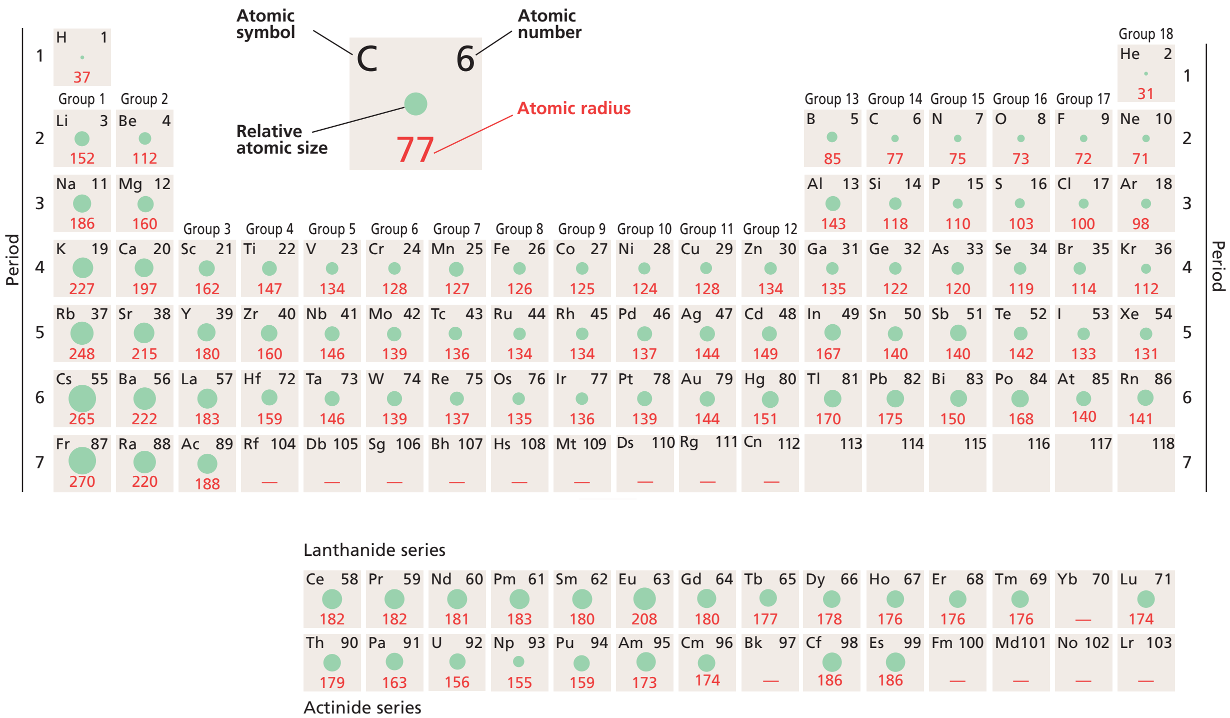
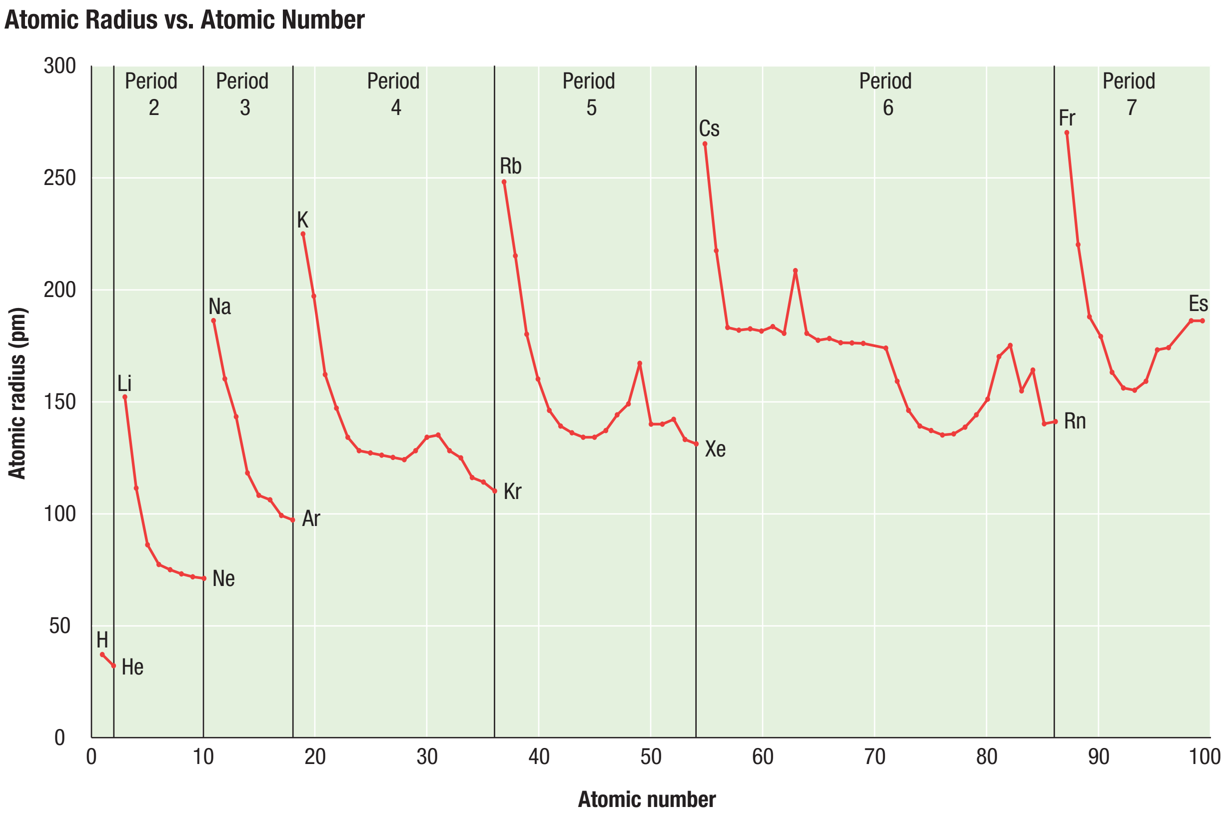


FIGURE 3.3

Periodic Trends The plot of atomic radius versus atomic number shows period and group trends.



Atomic Radius

Sample Problem E **Of the elements magnesium, Mg, chlorine, Cl, sodium, Na, and phosphorus, P, which has the largest atomic radius? Explain your answer in terms of trends in the periodic table.**

SOLVE All of the elements are in the third period. Of the four, sodium has the lowest atomic number and is the first element in the period. Therefore, sodium has the largest atomic radius, because atomic radii decrease across a period.

Practice Answers in Appendix E

1. Which of the following elements has the largest atomic radius: Li, O, C, or F? Which has the smallest atomic radius?
2. Of the elements calcium, Ca, beryllium, Be, barium, Ba, and strontium, Sr, which has the largest atomic radius? Explain your answer in terms of trends in the periodic table.
3. Of the elements aluminum, Al, magnesium, Mg, silicon, Si, and sodium, Na, which has the smallest atomic radius? Explain your answer in terms of trends in the periodic table.

MAIN IDEA

Removing electrons from atoms to form ions requires energy.

An electron can be removed from an atom if enough energy is supplied. Using A as a symbol for an atom of any element, the process can be expressed as follows.

$$A + \text{energy} \rightarrow A^+ + e^-$$

The A^+ represents an ion of element A with a single positive charge, referred to as a 1+ ion. **An ion is an atom or group of bonded atoms that has a positive or negative charge.** Sodium, for example, forms an Na^+ ion. **Any process that results in the formation of an ion is referred to as ionization.**

To compare the ease with which atoms of different elements give up electrons, chemists compare ionization energies. **The energy required to remove one electron from a neutral atom of an element is the ionization energy, IE (or first ionization energy, IE_1).** To avoid the influence of nearby atoms, measurements of ionization energies are made on isolated atoms in the gas phase. **Figure 3.4** gives the first ionization energies for the elements in kilojoules per mole (kJ/mol). **Figure 3.5**, on the next page, presents this information graphically.

FIGURE 3.4

Ionization Energy In general, first ionization energies increase across a period and decrease down a group. Values are given in KJ/mol.

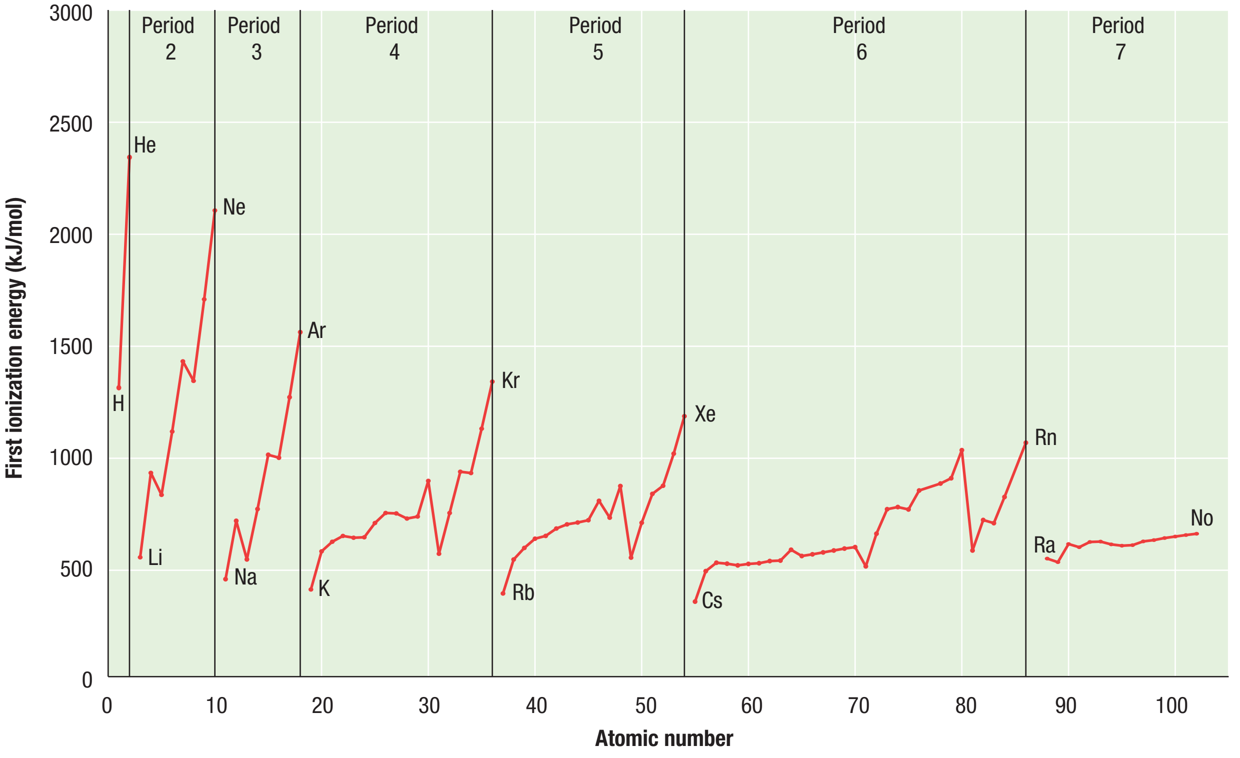
Key: 6 (Atomic number) — C (Symbol) — 1086 (First ionization energy)

Period	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1	1 H 1312																	2 He 2372
2	3 Li 520	4 Be 900											5 B 801	6 C 1086	7 N 1402	8 O 1314	9 F 1681	10 Ne 2081
3	11 Na 496	12 Mg 738											13 Al 578	14 Si 787	15 P 1012	16 S 1000	17 Cl 1251	18 Ar 1521
4	19 K 419	20 Ca 590	21 Sc 633	22 Ti 659	23 V 651	24 Cr 653	25 Mn 717	26 Fe 762	27 Co 760	28 Ni 737	29 Cu 746	30 Zn 906	31 Ga 579	32 Ge 762	33 As 947	34 Se 941	35 Br 1140	36 Kr 1351
5	37 Rb 403	38 Sr 550	39 Y 600	40 Zr 640	41 Nb 652	42 Mo 684	43 Tc 702	44 Ru 710	45 Rh 720	46 Pd 804	47 Ag 731	48 Cd 868	49 In 558	50 Sn 709	51 Sb 834	52 Te 869	53 I 1008	54 Xe 1170
6	55 Cs 376	56 Ba 503	57 La 538	72 Hf 659	73 Ta 761	74 W 770	75 Re 760	76 Os 839	77 Ir 878	78 Pt 868	79 Au 890	80 Hg 1007	81 Tl 589	82 Pb 716	83 Bi 703	84 Po 812	85 At —	86 Rn 1038
7	87 Fr —	88 Ra 509	89 Ac 490	104 Rf —	105 Db —	106 Sg —	107 Bh —	108 Hs —	109 Mt —	110 Ds —	111 Rg —	112 Cn —	113	114	115	116	117	118

Series														
Lanthanide series	58 Ce 534	59 Pr 527	60 Nd 533	61 Pm 536	62 Sm 545	63 Eu 547	64 Gd 592	65 Tb 566	66 Dy 573	67 Ho 581	68 Er 589	69 Tm 597	70 Yb 603	71 Lu 523
Actinide series	90 Th 587	91 Pa 570	92 U 598	93 Np 600	94 Pu 585	95 Am 578	96 Cm 581	97 Bk 601	98 Cf 608	99 Es 619	100 Fm 627	101 Md 635	102 No 642	103 Lr —

FIGURE 3.5

Ionization Energy Plot of first ionization energy, IE_1, versus atomic number. As atomic number increases, both the period and the group trends become less pronounced.



PREMIUM CONTENT
Animated Chemistry
HMDScience.com
Period Trends

Period Trends

In **Figures 3.4** and **3.5,** examine the ionization energies for the first and last elements in each period. You can see that the Group 1 metals have the lowest first ionization energies in their respective periods. Therefore, they lose electrons most easily. This ease of electron loss is a major reason for the high reactivity of the Group 1 (alkali) metals. The Group 18 elements, the noble gases, have the highest ionization energies. They do not lose electrons easily. The low reactivity of the noble gases is partly based on this difficulty of electron removal.

In general, ionization energies of the main-group elements increase across each period. This increase is caused by increasing nuclear charge. A higher charge more strongly attracts electrons in the same energy level. Increasing nuclear charge is responsible for both increasing ionization energy and decreasing radii across the periods. Note that, in general, nonmetals have higher ionization energies than metals do. In each period, the element of Group 1 has the lowest ionization energy and the element of Group 18 has the highest ionization energy.

Group Trends

Among the main-group elements, ionization energies generally decrease down the groups. Electrons removed from atoms of each succeeding element in a group are in higher energy levels, farther from the nucleus. Therefore, they are removed more easily. Also, as atomic number increases going down a group, more electrons lie between the nucleus and the electrons in the highest occupied energy levels. This partially shields the outer electrons from the effect of the nuclear charge. Together, these influences overcome the attraction of the electrons to the increasing nuclear charge.

Removing Electrons from Positive Ions

With sufficient energy, electrons can be removed from positive ions as well as from neutral atoms. The energies for removal of additional electrons from an atom are referred to as *the second ionization energy* (IE_2), *third ionization energy* (IE_3), and so on. **Figure 3.6** shows the first five ionization energies for the elements of the first, second, and third periods. You can see that the second ionization energy is always higher than the first, the third is always higher than the second, and so on.

FIGURE 3.6

IONIZATION ENERGIES (IN KJ/MOL) FOR ELEMENTS OF PERIODS 1–3

	Period 1		Period 2							
	H	He	Li	Be	B	C	N	O	F	Ne
IE_1	1312	**2372**	520	900	801	1086	1402	1314	1681	2081
IE_2		5250	**7298**	1757	2427	2353	2856	3388	3374	3952
IE_3			11 815	**14 849**	3660	4621	4578	5300	6050	6122
IE_4				21 007	**25 026**	6223	7475	7469	8408	9370
IE_5					32 827	**37 830**	9445	10 990	11 023	12 178

	Period 3							
	Na	Mg	Al	Si	P	S	Cl	Ar
IE_1	496	738	578	787	1012	1000	1251	1521
IE_2	**4562**	1451	1817	1577	1903	2251	2297	2666
IE_3	6912	**7733**	2745	3232	2912	3361	3822	3931
IE_4	9544	10 540	**11 578**	4356	4957	4564	5158	5771
IE_5	13 353	13 628	14 831	**16 091**	6274	7013	6540	7238

CHECK FOR UNDERSTANDING
Explain Explain in your own words what is meant by the italicized words in the text that state, "*each successive electron removed from an ion feels an increasingly stronger effective nuclear charge.*"

This is because as electrons are removed in successive ionizations, fewer electrons remain within the atom to shield the attractive force of the nucleus. *Thus, each successive electron removed from an ion feels an increasingly stronger effective nuclear charge (the nuclear charge minus the electron shielding).*

The first ionization energies in **Figure 3.6** show that removing a single electron from an atom of a Group 18 element is more difficult than removing an electron from atoms of other elements in the same period. This special stability of the noble-gas configuration also applies to ions that have noble-gas configurations. Notice in **Figure 3.6** the large increases between the first and second ionization energies of lithium, Li, and between the second and third ionization energies of beryllium, Be. Even larger increases in ionization energy exist between the third and fourth ionization energies of boron, B, and between the fourth and fifth ionization energies of carbon, C. In each case, the jump in ionization energy occurs when an ion assumes a noble-gas configuration. For example, the removal of one electron from a lithium atom ($[He]2s^1$) leaves the helium noble-gas configuration. The removal of four electrons from a carbon atom ($[He]2s^22p^2$) also leaves the helium configuration. This trend continues across the entire periodic system.

Periodic Trends in Ionization Energy

Sample Problem F **Consider two main-group elements, A and B. Element A has a first ionization energy of 419 kJ/mol. Element B has a first ionization energy of 1000 kJ/mol. Decide if each element is more likely to be in the *s*-block or *p*-block. Which element is more likely to form a positive ion?**

SOLVE

Element A has a very low ionization energy, which means that atoms of A lose electrons easily. Therefore, element A is most likely to be an s-block metal, because ionization energies increase across the periods.

Element B has a very high ionization energy, which means that atoms of B have difficulty losing electrons. Element B would most likely lie at the end of a period in the p-block.

Element A is more likely to form a positive ion, because it has a much lower ionization energy than element B does.

Practice Answers in Appendix E

1. Consider four hypothetical main-group elements, Q, R, T, and X, that have the outer electron configurations indicated below. Then, answer the questions that follow.

   Q: $3s^23p^5$ R: $3s^1$ T: $4d^{10}5s^25p^5$ X: $4d^{10}5s^25p^1$

   a. Identify the block location of each hypothetical main-group element.
   b. Which of these elements are in the same period? Which are in the same group?
   c. Which element would you expect to have the highest first ionization energy? Which would have the lowest first ionization energy?
   d. Which element would you expect to have the highest second ionization energy?
   e. Which of the elements is most likely to form a 1+ ion?

MAIN IDEA

Adding electrons to atoms to form ions also requires energy.

Neutral atoms can also acquire electrons. **The energy change that occurs when an electron is acquired by a neutral atom is called the atom's electron affinity.** Most atoms release energy when they acquire an electron.

$$A + e^- \longrightarrow A^- + \text{energy}$$

On the other hand, some atoms must be "forced" to gain an electron by the addition of energy.

$$A + e^- + \text{energy} \longrightarrow A^-$$

The quantity of energy absorbed would be represented by a positive number, but ions produced in this way are very unstable and hence the electron affinity for them is very difficult to determine. An ion produced in this way will be unstable and will lose the added electron spontaneously.

Figure 3.7 shows the electron affinity in kilojoules per mole for the elements. Positive electron affinities, because they are so difficult to determine with any accuracy, are denoted in **Figure 3.7** by "(0)." **Figure 3.8,** on the next page, presents these data graphically.

Period Trends

Among the elements of each period, the halogens (Group 17) gain electrons most readily. This is shown in **Figure 3.7** by the large negative values of halogens' electron affinities and is a major reason for the high reactivity levels of Group 17 elements. In general, as electrons add to the same *p* sublevel of atoms with increasing nuclear charge, electron affinities become more negative across each period within the *p*-block.

FIGURE 3.7

Periodic Table of Electron Affinities (kJ/mol) The values listed in parentheses in this periodic table of electron affinities are approximate. Electron affinity is estimated to be −50 kJ/mol for each of the lanthanides and 0 kJ/mol for each of the actinides.

6 — Atomic number
C — Symbol
−126.3 — Electron affinity

Period	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1	1 H −75.4																	2 He (0)
2	3 Li −61.8	4 Be (0)											5 B −27.7	6 C −126.3	7 N (0)	8 O −146.1	9 F −339.9	10 Ne (0)
3	11 Na −54.8	12 Mg (0)											13 Al −44.1	14 Si −138.5	15 P −74.6	16 S −207.7	17 Cl −361.7	18 Ar (0)
4	19 K −50.1	20 Ca (0)	21 Sc −18.8	22 Ti −7.9	23 V −52.5	24 Cr −66.6	25 Mn (0)	26 Fe −16.3	27 Co −66.1	28 Ni −115.6	29 Cu −122.8	30 Zn (0)	31 Ga −30	32 Ge −135	33 As −81	34 Se −202.1	35 Br −336.5	36 Kr (0)
5	37 Rb −48.6	38 Sr (0)	39 Y −30.7	40 Zr −42.6	41 Nb −89.3	42 Mo −74.6	43 Tc −55	44 Ru −105	45 Rh −113.7	46 Pd −55.7	47 Ag −130.2	48 Cd (0)	49 In −30	50 Sn −120	51 Sb −107	52 Te −197.1	53 I −305.9	54 Xe (0)
6	55 Cs −47.2	56 Ba (0)	57 La −50	72 Hf (0)	73 Ta −32.2	74 W −81.5	75 Re −15	76 Os −110	77 Ir −156.5	78 Pt −212.8	79 Au −230.9	80 Hg (0)	81 Tl −20	82 Pb −36	83 Bi −94.6	84 Po −190	85 At −280	86 Rn (0)
7	87 Fr −47.0	88 Ra (0)	89 Ac —	104 Rf —	105 Db —	106 Sg —	107 Bh —	108 Hs —	109 Mt —	110 Ds —	111 Rg —	112 Cn —	113	114	115	116	117	118

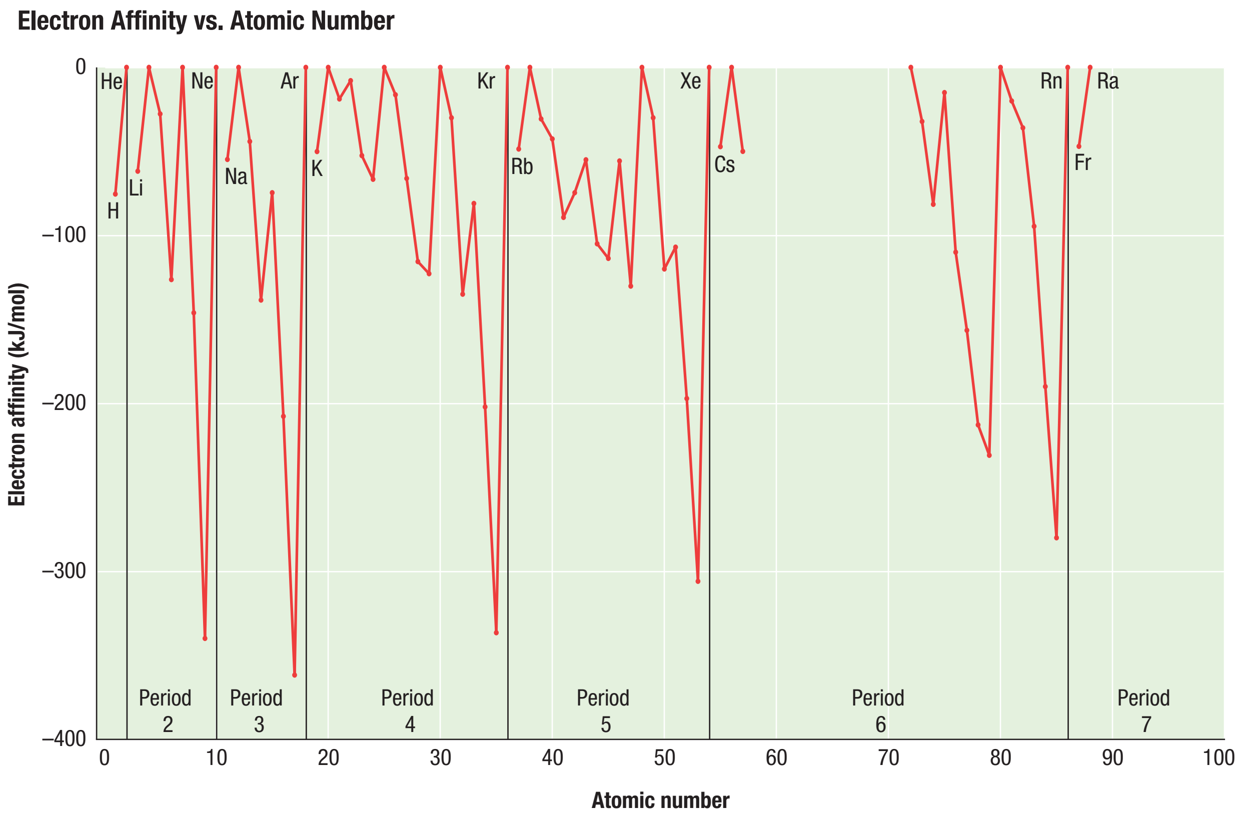


FIGURE 3.8

Electron Affinity and Atomic Numbers The plot of electron affinity versus atomic number shows that most atoms release energy when they acquire an electron, as indicated by negative values.

An exception to this trend occurs between Groups 14 and 15. Compare the electron affinities of carbon ([He]$2s^22p^2$) and nitrogen ([He]$2s^22p^3$). Adding an electron to a carbon atom gives a half-filled *p* sublevel. This occurs more easily than forcing an electron to pair with another electron in an orbital of the already half-filled *p* sublevel of a nitrogen atom.

Group Trends

Trends for electron affinities within groups are not as regular as trends for ionization energies. As a general rule, electrons are added with greater difficulty down a group. This pattern is a result of two competing factors. The first is a slight increase in effective nuclear charge down a group, which increases electron affinities. The second is an increase in atomic radius down a group, which decreases electron affinities. In general, the size effect predominates. But there are exceptions, especially among the heavy transition metals, which tend to be the same size or even decrease in radius down a group.

Adding Electrons to Negative Ions

For an isolated ion in the gas phase, it is more difficult to add a second electron to an already negatively charged ion. Therefore, second electron affinities are all positive. Certain *p*-block nonmetals tend to form negative ions that have noble gas configurations. The halogens do so by adding one electron. For example, chlorine has the configuration [Ne]$3s^23p^5$.

An atom of chlorine achieves the configuration of the noble gas argon by adding an electron to form the ion Cl^- ([Ne]$3s^23p^6$). Adding another electron is so difficult that Cl^{2-} never occurs. Atoms of Group 16 elements are present in many compounds as 2− ions. For example, oxygen ([He]$2s^22p^4$) achieves the configuration of the noble gas neon by adding two electrons to form the ion O^{2-}([He]$2s^22p^6$).

MAIN IDEA

When atoms become ions, their radii change.

Figure 3.9 shows the radii of some of the most common ions of the elements. Positive and negative ions have specific names.

A positive ion is known as a cation. The formation of a cation by the loss of one or more electrons always leads to a decrease in atomic radius, because the removal of the highest-energy-level electrons results in a smaller electron cloud. Also, the remaining electrons are drawn closer to the nucleus by its unbalanced positive charge.

A negative ion is known as an anion. The formation of an anion by the addition of one or more electrons always leads to an increase in atomic radius. This is because the total positive charge of the nucleus remains unchanged when an electron is added to an atom or an ion. So the electrons are not drawn to the nucleus as strongly as they were before the addition of the extra electron. The electron cloud also spreads out because of greater repulsion between the increased number of electrons.

Period Trends

Within each period of the periodic table, the metals at the left tend to form cations, and the nonmetals at the upper right tend to form anions. Cationic radii decrease across a period, because the electron cloud shrinks due to the increasing nuclear charge acting on the electrons in the same main energy level.

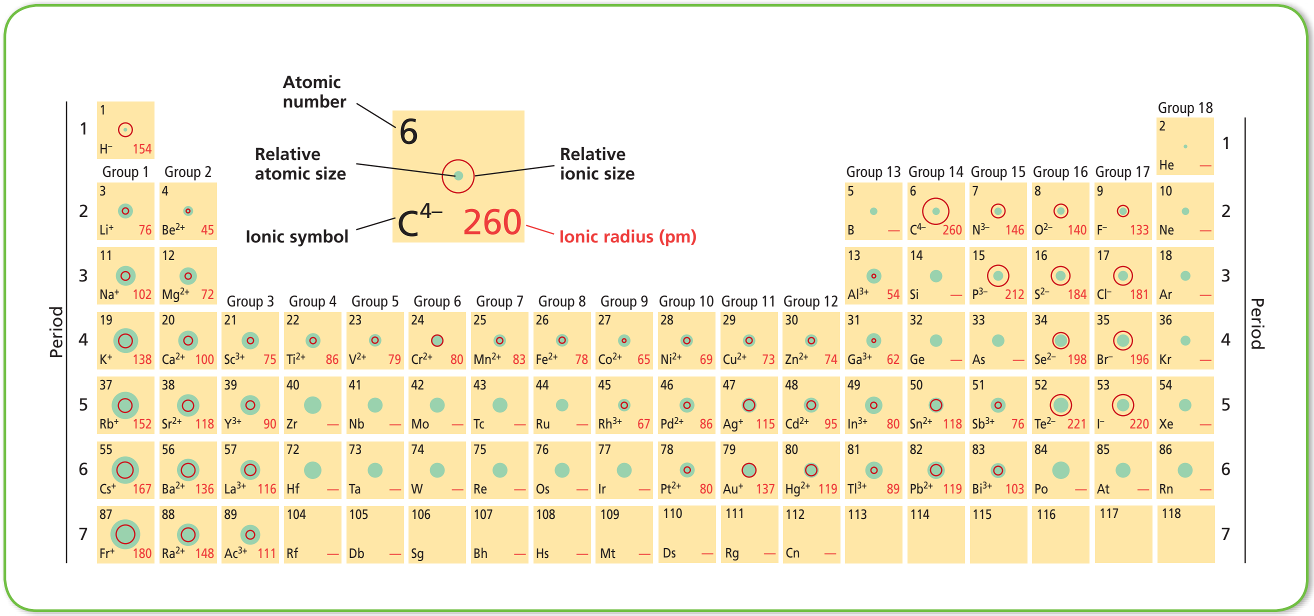


FIGURE 3.9

Periodic Table of Ionic Radii (pm) The ionic radii of the ions most common in chemical compounds are shown. Cations are smaller, and anions are larger than the atoms from which they are formed.

Starting with Group 15, in which atoms assume stable noble-gas configurations by gaining three electrons, anions are more common than cations. Anionic radii decrease across each period for the elements in Groups 15–18. The reasons for this trend are the same as the reasons that cationic radii decrease from left to right across a period.

Group Trends

The outer electrons in both cations and anions are in higher energy levels as one reads down a group. Thus, just as there is a gradual increase of atomic radii down a group, there is also a gradual increase of ionic radii.

MAIN IDEA

Only the outer electrons are involved in forming compounds.

Chemical compounds form because electrons are lost, gained, or shared between atoms. The electrons that interact in this manner are those in the highest energy levels. These are the electrons most subject to the influence of nearby atoms or ions. **The electrons available to be lost, gained, or shared in the formation of chemical compounds are referred to as valence electrons.** Valence electrons are often located in incompletely filled main-energy levels. For example, the electron lost from the 3*s* sublevel of Na to form Na^+ is a valence electron.

For main-group elements, the valence electrons are the electrons in the outermost *s* and *p* sublevels. The inner electrons are in filled energy levels and are held too tightly by the nucleus to be involved in compound formation. The Group 1 and Group 2 elements have one and two valence electrons, respectively, as shown in **Figure 3.10.** The elements of Groups 13–18 have a number of valence electrons equal to the group number minus 10. In some cases, both the *s* and *p* sublevel valence electrons of the *p*-block elements are involved in compound formation (**Figure 3.10**). In other cases, only the electrons from the *p* sublevel are involved.

FIGURE 3.10

VALENCE ELECTRONS IN MAIN-GROUP ELEMENTS

Group number	Group configuration	Number of valence electrons
1	ns^1	1
2	ns^2	2
13	ns^2p^1	3
14	ns^2p^2	4
15	ns^2p^3	5
16	ns^2p^4	6
17	ns^2p^5	7
18	ns^2p^6	8

MAIN IDEA

Atoms have different abilities to capture electrons.

Valence electrons hold atoms together in chemical compounds. In many compounds, the negative charge of the valence electrons is concentrated closer to one atom than to another. This uneven concentration of charge has a significant effect on the chemical properties of a compound. It is therefore useful to have a measure of how strongly one atom attracts the electrons of another atom within a compound.

Linus Pauling, one of America's most famous chemists, devised a scale of numerical values reflecting the tendency of an atom to attract electrons. **Electronegativity is a measure of the ability of an atom in a chemical compound to attract electrons from another atom in the compound.** The most electronegative element, fluorine, is arbitrarily assigned an electronegativity of four. Other values are then calculated in relation to this value.

Period Trends

As shown in **Figure 3.11**, *electronegativities tend to increase across each period, although there are exceptions.* The alkali and alkaline-earth metals are the least electronegative elements. In compounds, their atoms have a low attraction for electrons. Nitrogen, oxygen, and the halogens are the most electronegative elements. Their atoms attract electrons strongly.

FIGURE 3.11

Periodic Table of Electronegativities Shown are the electronegativities of the elements according to the Pauling scale. The most-electronegative elements are located in the upper right of the *p*-block. The least-electronegative elements are located in the lower left of the *s*-block.

- 6 — Atomic number
- C — Symbol
- 2.5 — Electronegativity

Period	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1	1 H 2.1																	2 He —
2	3 Li 1.0	4 Be 1.5											5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne —
3	11 Na 0.9	12 Mg 1.2											13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar —
4	19 K 0.8	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.8	28 Ni 1.8	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8	36 Kr 3.0
5	37 Rb 0.8	38 Sr 1.0	39 Y 1.2	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.7	49 In 1.7	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5	54 Xe 2.6
6	55 Cs 0.7	56 Ba 0.9	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 2.4
7	87 Fr 0.7	88 Ra 0.9	89 Ac 1.1	104 Rf —	105 Db —	106 Sg —	107 Bh —	108 Hs —	109 Mt —	110 Ds —	111 Rg —	112 Cn —	113	114	115	116	117	118

Series														
Lanthanide series	58 Ce 1.1	59 Pr 1.1	60 Nd 1.1	61 Pm 1.1	62 Sm 1.2	63 Eu 1.1	64 Gd 1.2	65 Tb 1.1	66 Dy 1.2	67 Ho 1.2	68 Er 1.2	69 Tm 1.3		71 Lu 1.3
Actinide series	90 Th 1.3	91 Pa 1.5	92 U 1.4	93 Np 1.4	94 Pu 1.3	95 Am 1.3	96 Cm 1.3	97 Bk 1.3	98 Cf 1.3	99 Es 1.3	100 Fm 1.3	101 Md 1.3		103 Lr —

CHECK FOR UNDERSTANDING
Compare The text states that some noble gases "do not form compounds and therefore cannot be assigned electronegativities." How can some of them form compounds, while others do not?

Electronegativities tend to either decrease down a group or remain about the same. The noble gases are unusual in that some of them do not form compounds and therefore cannot be assigned electronegativities. When a noble gas does form a compound, its electronegativity is rather high, similar to the values for the halogens. The combination of the period and group trends in electronegativity results in the highest values belonging to the elements in the upper right of the periodic table. The lowest values belong to the elements in the lower left of the table. These trends are shown graphically in **Figure 3.12.**

MAIN IDEA

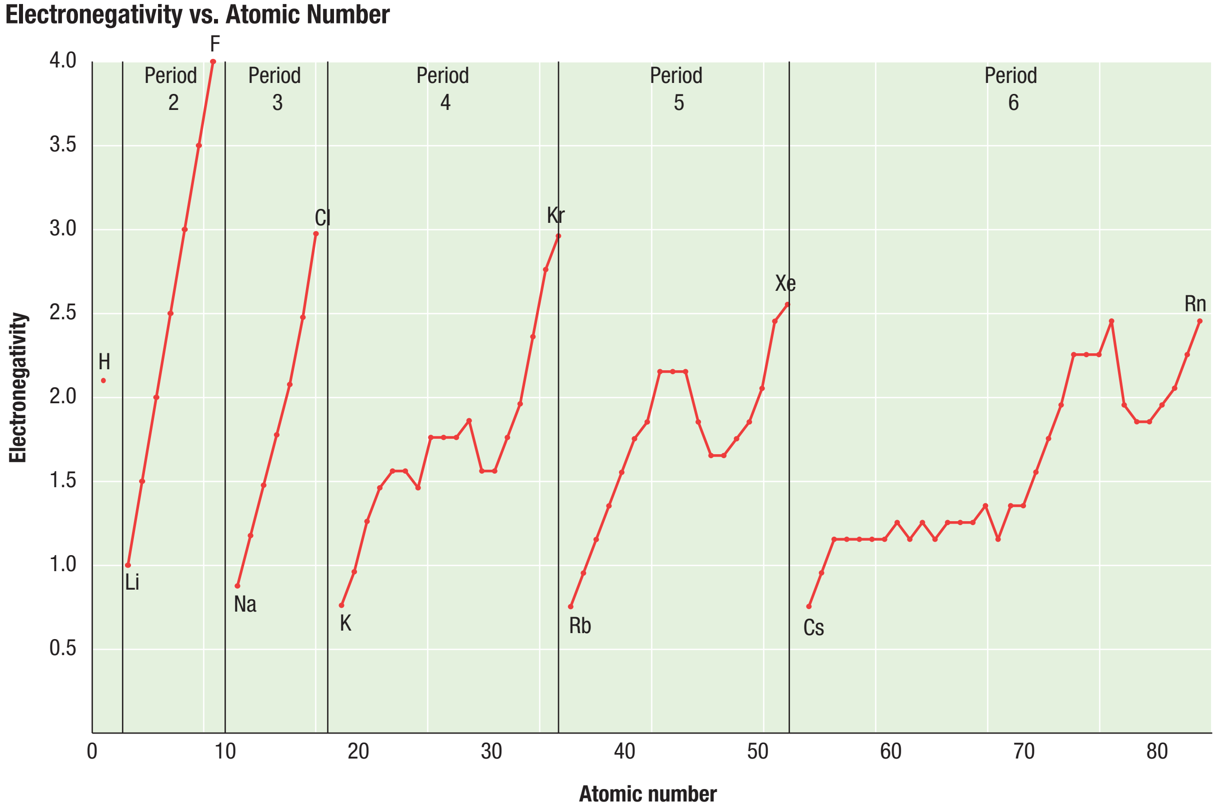
The properties of *d*-block metals do not vary much.

The properties of the *d*-block elements (which are all metals) vary less and with less regularity than those of the main-group elements. This trend is indicated by the curves in **Figures 3.3** and **3.5,** which flatten where the *d*-block elements fall in the middle of Periods 4–6.

Recall that atoms of the *d*-block elements contain from zero to two electrons in the *s* orbital of their highest occupied energy level and one to ten electrons in the *d* sublevel of the next-lower energy level.

FIGURE 3.12

Electronegativity and Atomic Number The plot shows electronegativity versus atomic number for Periods 1–6.



PREMIUM CONTENT

Solve It! Cards
HMDScience.com

Periodic Trends in Electronegativity

Sample Problem G **Of the elements gallium, Ga, bromine, Br, and calcium, Ca, which has the highest electronegativity? Explain your answer in terms of periodic trends.**

SOLVE All of these elements are in the fourth period. Bromine has the highest atomic number and is farthest to the right in the period. Therefore, bromine should have the highest electronegativity because electronegativity increases across the periods.

Practice

Answers in Appendix E

1. Consider five hypothetical main-group elements, E, G, J, L, and M, that have the outer electron configurations shown at the top of the next page.

$$E = 2s^2 2p^5 \qquad G = 4d^{10}5s^2 5p^5 \qquad J = 2s^2 2p^2$$

$$L = 5d^{10}6s^2 6p^5 \qquad M = 2s^2 2p^4$$

   a. Identify the block location for each element. Then, determine which elements are in the same period and which are in the same group.
   b. Which element would you expect to have the highest electron affinity? Which would you expect to form a 1− ion? Which should have the highest electronegativity?
   c. Compare the ionic radius of the typical ion formed by the element G with the radius of the atom from which the ion was formed.
   d. Which element(s) contain seven valence electrons?

Therefore, electrons in both the *ns* sublevel and the $(n–1)d$ sublevel are available to interact with their surroundings. As a result, electrons in the incompletely filled *d* sublevels are responsible for many characteristic properties of the *d*-block elements.

Atomic Radii

The atomic radii of the *d*-block elements generally decrease across the periods. However, this decrease is less than that for the main-group elements, because the electrons added to the $(n–1)d$ sublevel shield the outer electrons from the nucleus.

Also, note in **Figure 3.3** that the radii dip to a low and then increase slightly across each of the four periods that contain *d*-block elements. As the number of electrons in the *d* sublevel increases, the radii increase because of repulsion among the electrons.

In the sixth period, the *f*-block elements fall between lanthanum (Group 3) and hafnium (Group 4). Because of the increase in atomic number that occurs from lanthanum to hafnium, the atomic radius of hafnium is actually slightly less than that of zirconium, Zr, the element immediately above it. The radii of elements following hafnium in the sixth period vary with increasing atomic number in the usual manner.

Ionization Energy

As they do for the main-group elements, ionization energies of the *d*-block and *f*-block elements generally increase across the periods. In contrast to the decrease down the main groups, however, the first ionization energies of the *d*-block elements generally increase down each group. This is because the electrons available for ionization in the outer *s* sublevels are less shielded from the increasing nuclear charge by electrons in the incomplete $(n–1)d$ sublevels.

Ion Formation and Ionic Radii

Among all atoms of the *d*-block and *f*-block elements, electrons in the highest occupied sublevel are always removed first. For the *d*-block elements, this means that although newly added electrons occupy the *d* sublevels, the first electrons to be removed are those in the outermost *s* sublevels. For example, iron, Fe, has the electron configuration $[Ar]3d^64s^2$. First, it loses two 4*s* electrons to form Fe^{2+} ($[Ar]3d^6$). Fe^{2+} can then lose a 3*d* electron to form Fe^{3+} ($[Ar]3d^5$).

Most *d*-block elements commonly form 2+ ions in compounds. Some, such as iron and chromium, also commonly form 3+ ions. The Group 3 elements form only ions with a 3+ charge. Copper forms 1+ and 2+ ions, and silver usually forms only 1+ ions. As expected, the cations have smaller radii than the atoms do. Comparing 2+ ions across the periods shows a decrease in size that parallels the decrease in atomic radii.

Electronegativity

The *d*-block elements all have electronegativities between 1.1 and 2.54. Only the active metals of Groups 1 and 2 have lower electronegativities. The *d*-block elements also follow the general trend for electronegativity values to increase as radii decrease, and vice versa. The *f*-block elements all have similar electronegativities, which range from 1.1 to 1.5.

SECTION 3 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. State the general period and group trends among main-group elements with respect to each of the following properties:

a. atomic radii

b. first ionization energy

c. electron affinity

d. ionic radii

e. electronegativity

2. a. In general, how do the periodic properties of the *d*-block elements compare with those of the main-group elements?

b. Explain the comparison made in (a).

3. For each main-group element, what is the relationship between its group number and the number of valence electrons that the group members have?

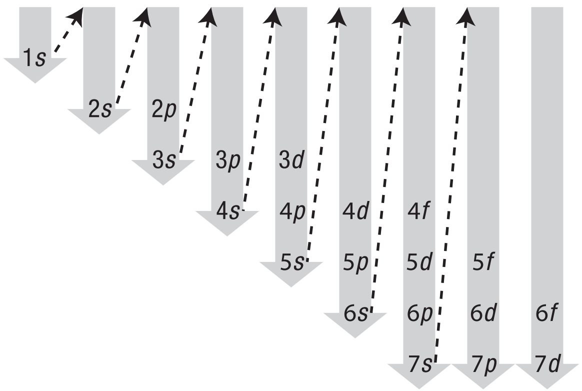
Critical Thinking

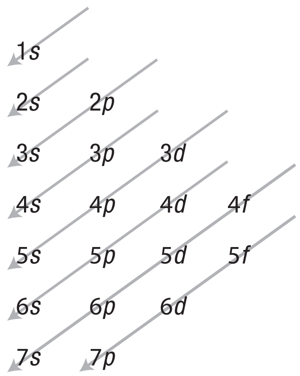
4. RELATING IDEAS Graph the general trends (left to right and top to bottom) in the second ionization energy (IE_2) of an element as a function of its atomic number, over the range Z = 1–20. Label the minima and maxima on the graph with the appropriate element symbol.

Math Tutor Writing Electron Configurations

The arrangement of elements in the periodic table reflects the arrangement of electrons in an atom. Each period begins with an atom that has an electron in a new energy level and with the exception of the first period, each period ends with an atom that has a filled set of *p* orbitals.

To write the electron configuration of an element, you must fill the sublevels in order of increasing energy. If you follow the arrows in either of the two types of mnemonics shown below, you will get correct configurations for most elements.





You also need to know how many orbitals are in each sublevel and that each orbital can contain two electrons of opposite spins. As shown in the following table, the sublevels *s, p, d,* and *f* have 1, 3, 5, and 7 available orbitals, respectively.

SUBLEVEL	*s*	*p*	*d*	*f*
No. of orbitals	1	3	5	7
No. of electrons	2	6	10	14

Sample Problem

Write the full electron configuration for phosphorus.

The atomic number of phosphorus is 15, so a phosphorus atom has 15 protons and electrons. Assign each of the 15 electrons to the appropriate sublevels. The final sublevel can be unfilled and will contain the number of valence electrons.

$$\underbrace{1s^2}_{2e^-} + \underbrace{2s^2}_{2e^-} + \underbrace{2p^6}_{6e^-} + \underbrace{3s^2}_{2e^-} + \underbrace{3p^3}_{3e^-} = 15e^-$$

So, the full electron configuration of phosphorus is $1s^22s^22p^63s^23p^3$.

Practice

Answers in Appendix E

1. Write full electron configurations for the following elements.
a. aluminum
b. neon
c. tin
d. potassium

2. Use noble gas symbols to write shorthand electron configurations for the following elements.
a. silicon
b. rubidium
c. antimony
d. arsenic

CHAPTER 5 **Summary**

PREMIUM CONTENT

Interactive Review
HMDScience.com

Review Games
Concept Maps

SECTION 1 History of the Periodic Table

- The periodic law states that the physical and chemical properties of the elements are periodic functions of their atomic numbers.
- The periodic table is an arrangement of the elements in order of their atomic numbers so that elements with similar properties fall in the same column.
- The columns in the periodic table are referred to as groups.

KEY TERMS

periodic law
periodic table
lanthanide
actinide

SECTION 2 Electron Configuration and the Periodic Table

- The rows in the periodic table are called periods.
- Many chemical properties of the elements can be explained by the configurations of the elements' outermost electrons.
- The noble gases exhibit unique chemical stability because their highest occupied levels have an octet of electrons, ns^2np^6 (with the exception of helium, whose stability arises from its highest occupied level being completely filled with two electrons, $1s^2$).
- Based on the electron configurations of the elements, the periodic table can be divided into four blocks: the *s*-block, the *p*-block, the *d*-block, and the *f*-block.

KEY TERMS

alkali metals
alkaline-earth metals
transition elements
main-group elements
halogens

SECTION 3 Electron Configuration and Periodic Properties

- The groups and periods of the periodic table display general trends in the following properties of the elements: electron affinity, electronegativity, ionization energy, atomic radius, and ionic radius.
- The electrons in an atom that are available to be lost, gained, or shared in the formation of chemical compounds are referred to as valence electrons.
- In determining the electron configuration of an ion, the order in which electrons are removed from the atom is the reverse of the order given by the atom's electron-configuration notation.

KEY TERMS

atomic radius
ion
ionization
ionization energy
electron affinity
cation
anion
valence electron
electronegativity

CHAPTER 5 Review

SECTION 1

History of the Periodic Table

REVIEWING MAIN IDEAS

1. Describe the contributions made by the following scientists to the development of the periodic table:
   a. Stanislao Cannizzaro
   b. Dmitri Mendeleev
   c. Henry Moseley

2. State the periodic law.

3. How is the periodic law demonstrated within the groups of the periodic table?

SECTION 2

Electron Configuration and the Periodic Table

REVIEWING MAIN IDEAS

4. a. How do the electron configurations within the same group of elements compare?
   b. Why are the noble gases relatively unreactive?

5. What determines the length of each period in the periodic table?

6. What is the relationship between the electron configuration of an element and the period in which that element appears in the periodic table?

7. a. What information is provided by the specific block location of an element?
   b. Identify, by number, the groups located within each of the four block areas.

8. a. Which elements are designated as the alkali metals?
   b. List four of their characteristic properties.

9. a. Which elements are designated as the alkaline-earth metals?
   b. How do their characteristic properties compare with those of the alkali metals?

10. a. Write the group configuration notation for each *d*-block group.
    b. How do the group numbers of those groups relate to the number of outer *s* and *d* electrons?

11. What name is sometimes used to refer to the entire set of *d*-block elements?

12. a. What types of elements make up the *p*-block?
    b. How do the properties of the *p*-block metals compare with those of the metals in the *s*- and *d*-blocks?

13. a. Which elements are designated as the halogens?
    b. List three of their characteristic properties.

14. a. Which elements are metalloids?
    b. Describe their characteristic properties.

15. Which elements make up the *f*-block in the periodic table?

16. a. What are the main-group elements?
    b. What trends can be observed across the various periods within the main-group elements?

PRACTICE PROBLEMS

17. Write the noble-gas notation for the electron configuration of each of the following elements, and indicate the period in which each belongs.
    a. Li c. Cu e. Sn
    b. O d. Br

18. Without looking at the periodic table, identify the period, block, and group in which the elements with the following electron configurations are located. (Hint: See Sample Problem A.)
    a. $[Ne]3s^2 3p^4$
    b. $[Kr]4d^{10} 5s^2 5p^2$
    c. $[Xe]4f^{14} 5d^{10} 6s^2 6p^5$

19. Based on the information given below, give the group, period, block, and identity of each element described. (Hint: See Sample Problem B.)
    a. $[He]2s^2$
    b. $[Ne]3s^1$
    c. $[Kr]5s^2$
    d. $[Ar]4s^2$
    e. $[Ar]3d^5 4s^1$

20. Without looking at the periodic table, write the expected outer electron configuration for each of the following elements. (Hint: See Sample Problem C.)
    a. Group 7, fourth period
    b. Group 3, fifth period
    c. Group 12, sixth period

21. Identify the block, period, group, group name (where appropriate), element name, element type, and relative reactivity for the elements with the following electron configurations. (Hint: See Sample Problem D.)
- **a.** $[Ne]3s^2 3p^1$
- **b.** $[Ar]3d^{10} 4s^2 4p^6$
- **c.** $[Kr]4d^{10} 5s^1$
- **d.** $[Xe]4f^1 5d^1 6s^2$

SECTION 3

Electron Configuration and Periodic Properties

REVIEWING MAIN IDEAS

22. **a.** What is meant by *atomic radius*?
b. What trend is observed among the atomic radii of main-group elements across a period?
c. Explain this trend.

23. **a.** What trend is observed among the atomic radii of main-group elements down a group?
b. Explain this trend.

24. Define each of the following terms:
- **a.** ion
- **b.** ionization
- **c.** first ionization energy
- **d.** second ionization energy

25. **a.** How do the first ionization energies of main-group elements vary across a period and down a group?
b. Explain the basis for each trend.

26. **a.** What is electron affinity?
b. What signs are associated with electron affinity values, and what is the significance of each sign?

27. **a.** Distinguish between a cation and an anion.
b. How does the size of each compare with the size of the neutral atom from which it is formed?

28. **a.** What are valence electrons?
b. Where are such electrons located?

29. For each of the following groups, indicate whether electrons are more likely to be lost or gained in compound formation, and give the number of such electrons typically involved.

a. Group 1	**d.** Group 16
b. Group 2	**e.** Group 17
c. Group 13	**f.** Group 18

30. **a.** What is electronegativity?
b. Why is fluorine special in terms of electronegativity?

31. Identify the most- and least-electronegative groups of elements in the periodic table.

PRACTICE PROBLEMS

32. Of cesium, Cs, hafnium, Hf, and gold, Au, which element has the smallest atomic radius? Explain your answer in terms of trends in the periodic table. (Hint: See Sample Problem E.)

33. **a.** Distinguish between the first, second, and third ionization energies of an atom.
b. How do the values of successive ionization energies compare?
c. Why does this occur?

34. Without looking at the electron affinity table, arrange the following elements in order of *decreasing* electron affinities: C, O, Li, Na, Rb, and F.

35. **a.** Without looking at the ionization energy table, arrange the following elements in order of decreasing first ionization energies: Li, O, C, K, Ne, and F.
b. Which of the elements listed in (a) would you expect to have the highest second ionization energy? Why?

36. **a.** Which of the following cations is least likely to form: Sr^{2+}, Al^{3+}, K^{2+}?
b. Which of the following anions is least likely to form: I^-, Cl^-, O^{2-}?

37. Which element is the most electronegative among C, N, O, Br, and S? Which group does it belong to? (Hint: See Sample Problem G.)

38. The two ions K^+ and Ca^{2+} each have 18 electrons surrounding the nucleus. Which would you expect to have the smaller radius? Why?

Mixed Review

REVIEWING MAIN IDEAS

39. Without looking at the periodic table, identify the period, block, and group in which each of the following elements is located.
- **a.** [Rn]$7s^1$
- **b.** [Ar]$3d^2 4s^2$
- **c.** [Kr]$4d^{10} 5s^1$
- **d.** [Xe]$4f^{14} 5d^9 6s^1$

40. **a.** Which elements are designated as the noble gases?
b. What is the most significant property of these elements?

41. Which of the following does not have a noble-gas configuration: Na^+, Rb^+, O^{2-}, Br^- Ca^+, Al^{3+}, S^{2-}?

42. **a.** How many groups are in the periodic table?
b. How many periods are in the periodic table?
c. Which two blocks of the periodic table make up the main-group elements?

43. Write the noble-gas notation for the electron configuration of each of the following elements, and indicate the period and group in which each belongs.
- **a.** Mg
- **b.** P
- **c.** Sc
- **d.** Y

44. Use the periodic table to describe the chemical properties of the following elements:
- **a.** fluorine, F
- **b.** xenon, Xe
- **c.** sodium, Na
- **d.** gold, Au

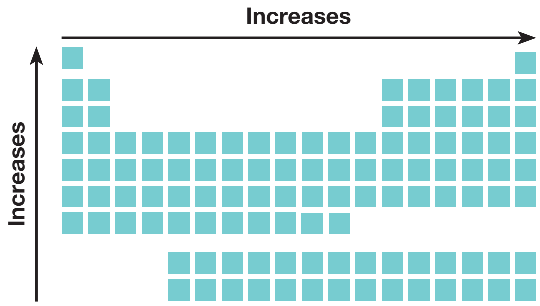
45. For each element listed below, determine the charge of the ion that is most likely to be formed and the identity of the noble gas whose electron configuration is thus achieved.

a. Li	**e.** Mg	**i.** Br
b. Rb	**f.** Al	**j.** Ba
c. O	**g.** P	
d. F	**h.** S	

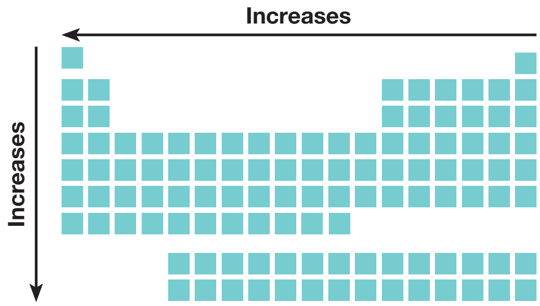
46. Describe some differences between the *s*-block metals and the *d*-block metals.

47. Why do the halogens readily form 1–ions?

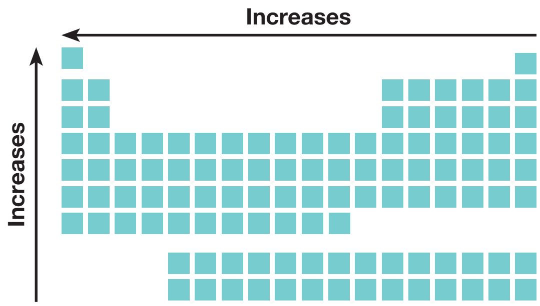
48. Identify which trends in the diagrams below describe atomic radius, ionization energy, electron affinity, and electronegativity.



a.



b.



c.

49. The electron configuration of argon differs from those of chlorine and potassium by one electron each. Compare the reactivity of these three elements.

CRITICAL THINKING

As a member on the newly-inhabited space station Alpha, you are given the task of organizing information on newly discovered elements as it comes in from the laboratory. To date, five elements have been discovered and have been assigned names and symbols from the Greek alphabet. An analysis of the new elements has yielded the following data:

Element name	Atomic no.	Atomic mass	Properties
Epsilon ε	23	47.33	nonmetal, very reactive, produces a salt when combined with a metal, gaseous state
Beta β	13	27.01	metal, very reactive, soft solid, low melting point
Gamma γ	12	25.35	nonmetal, gaseous element, extremely unreactive
Delta Δ	4	7.98	nonmetal, very abundant, forms compounds with most other elements
Lambda Λ	9	16.17	metal, solid state, good conductor, high luster, hard and dense

50. Applying Models Create a periodic table based on the properties of the five new elements.

51. Predicting Outcomes Using your newly created periodic table, predict the atomic number of an element with an atomic mass of 11.29 that has nonmetallic properties and is very reactive.

52. Predicting Outcomes Predict the atomic number of an element having an atomic mass of 15.02 that exhibits metallic properties but is softer than lambda and harder than beta.

53. Analyzing Information Analyze your periodic table for trends, and describe those trends.

USING THE HANDBOOK

54. Review the boiling point and melting point data in the tables of the *Elements Handbook* (Appendix A). Make a list of the elements that exist as liquids or gases at the boiling point of water, 100°C.

55. Because transition metals have vacant *d* orbitals, they form a greater variety of colored compounds than do the metals of Groups 1 and 2. Review the section of the *Elements Handbook* (Appendix A) on transition metals, and answer the following:

a. What colors are exhibited by chromium in its common oxidation states?

b. What gems contain chromium impurities?

c. What colors are often associated with the following metal ions: copper, cadmium, cobalt, zinc, and nickel?

d. What transition elements are considered noble metals? What are the characteristics of a noble metal?

RESEARCH AND WRITING

56. Prepare a report tracing the evolution of the current periodic table since 1900. Cite the chemists involved and their major contributions.

57. Write a report describing the contributions of Glenn Seaborg toward the discovery of many of the actinide elements.

ALTERNATIVE ASSESSMENT

58. Construct your own periodic table or obtain a poster that shows related objects, such as fruits or vegetables, in periodic arrangement. Describe the organization of the table and the trends it illustrates. Use this table to make predictions about your subject matter.

Standards-Based Assessment

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. In the modern periodic table, elements are arranged according to
A. decreasing atomic mass.
B. Mendeleev's original model.
C. increasing atomic number.
D. when they were discovered.

2. Group 17 elements, the halogens, are the most reactive of the nonmetal elements because they
A. require only one electron to fill their outer energy level.
B. have the highest ionization energies.
C. have the largest atomic radii.
D. are the farthest to the right in the periodic table.

3. The periodic law states that
A. the chemical properties of elements can be grouped according to periodicity.
B. the properties of the elements are functions of atomic mass.
C. all elements in the same group have the same number of valence electrons.
D. all elements with the same number of occupied energy levels must be in the same group.

4. As you move left to right across Period 3 from Mg to Cl, the energy needed to remove an electron from an atom
A. generally increases.
B. generally decreases.
C. does not change.
D. varies unpredictably.

5. Which of the following elements has the highest electronegativity?
A. oxygen
B. hydrogen
C. fluorine
D. carbon

6. The noble gases have
A. high ionization energies.
B. high electron affinities.
C. large atomic radii.
D. a tendency to form both cations and anions.

7. Which electron configuration is *not* correct?
A. O^{2-} $[He]2s^22p^6$
B. Mg^{2+} $[He]2s^22p^6$
C. V^{3+} $[Ar]3d^2$
D. Al^{3+} $[Ar]2s^22p^6$

8. Which two elements are more likely to have the same charge on their ions?
A. Se and As
B. Sn and Si
C. Ca and Rb
D. I and Xe

9. Using only the periodic table, choose the list that ranks the elements Sr, Te, Kr, Ru, and Cs in order of increasing ionization energy.
A. Sr < Te < Ru < Cs < Kr
B. Te < Ru < Sr < Cs < Kr
C. Cs < Sr < Ru < Te < Kr
D. Kr < Cs < Sr < Ru < Te

SHORT ANSWER

10. The second ionization energies for the elements S–Ti are listed in a scrambled order below. Assign the correct IE_2 value to each element. (Hint: S has $IE_2 = 2251$ kJ/mol, and Ti has $IE_2 = 1310$ kJ/mol.) Explain your reasoning.

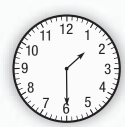
IE_2 values (kJ/mol): 2666, 2297, 3051, 1235, 2251, 1310, and 1145

11. What group most commonly forms 2– ions? Explain your reasoning.

EXTENDED RESPONSE

12. An ordered list of atomic radii for 14 consecutive elements is shown below. Without using **Figure 3.2,** make a graph of these atomic radii versus the element's atomic number. Explain your reasoning.

Atomic radii (pm): 75, 73, 72, 71, 186, 160, 143, 118, 110, 103, 100, 98, 227, and 197



Test Tip
If you are short on time, quickly scan the unanswered questions to see which might be easiest to answer.