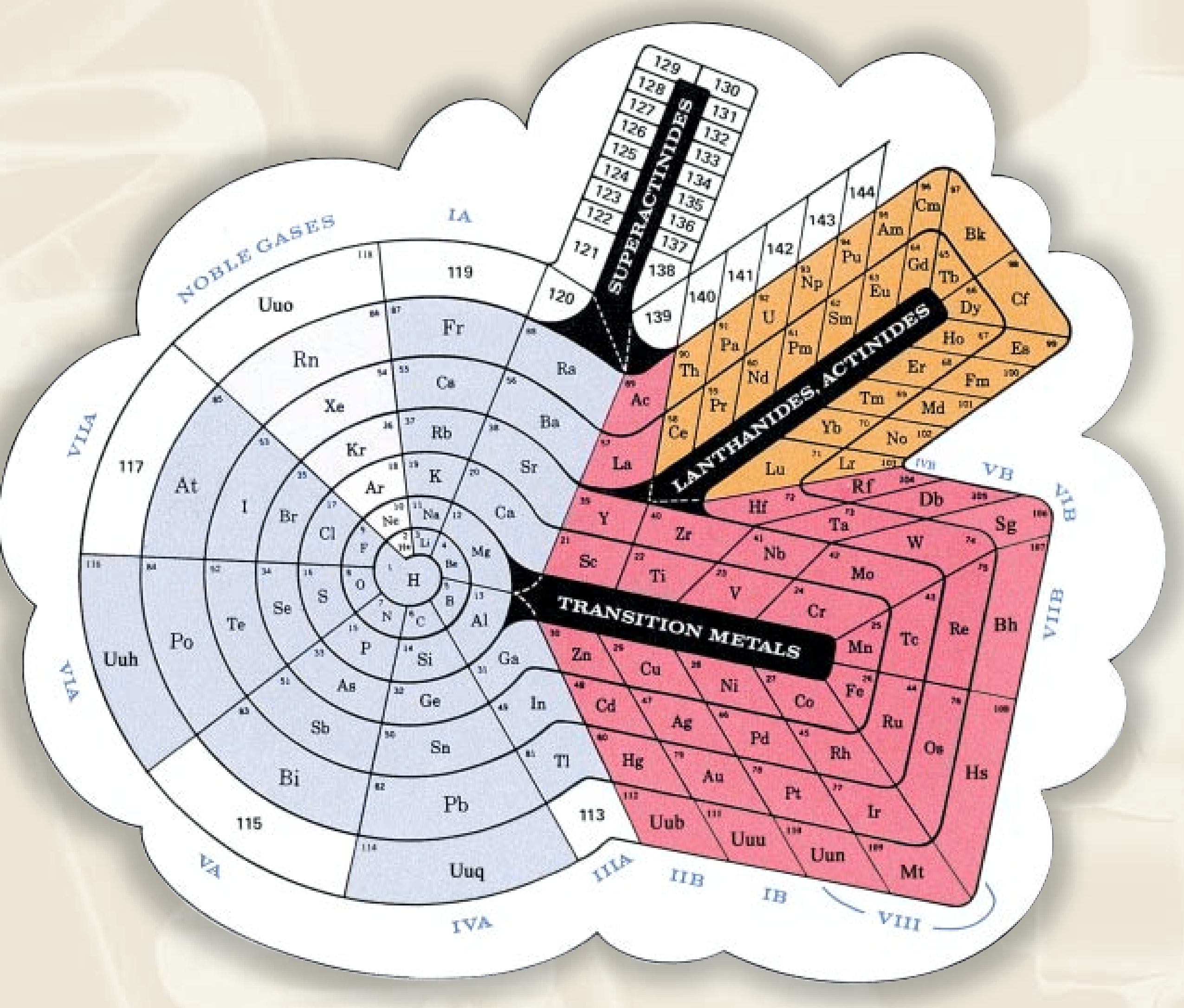
CHAPTER 5

The Periodic Law



The physical and chemical properties of the elements are periodic functions of their atomic numbers.

History of the Periodic Table

SECTION 5-1

OBJECTIVES

- Explain the roles of Mendeleev and Moseley in the development of the periodic table.
- Describe the modern periodic table.
- Explain how the periodic law can be used to predict the physical and chemical properties of elements.
- Describe how the elements belonging to a group of the periodic table are interrelated in terms of atomic number.

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.

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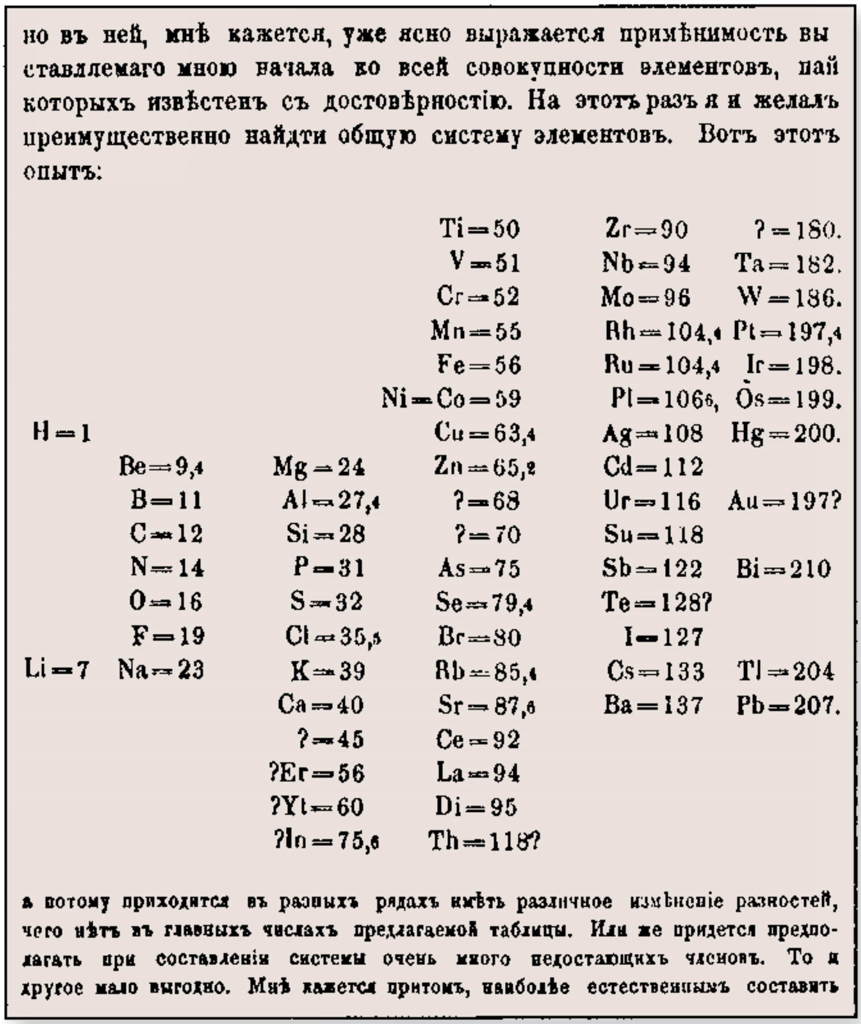
Mendeleev and Chemical Periodicity

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 that 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 are also periodic.

FIGURE 5-1 The regularly spaced water waves represent a simple periodic pattern.



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

			Ti=50	Zr=90	?=180.
			V=51	Nb=94	Ta=182.
			Cr=52	Mo=96	W=186.
			Mn=55	Rh=104,4	Pt=197,4
			Fe=56	Ru=104,4	Ir=198.
			Ni=Co=59	Pl=106,6,	Os=199.
H=1			Cu=63,4	Ag=108	Hg=200.
	Be=9,4	Mg=24	Zn=65,2	Cd=112	
	B=11	Al=27,4	?=68	Ur=116	Au=197?
	C=12	Si=28	?=70	Su=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

FIGURE 5-2 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.

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 5-2, 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 5-2). 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?

Moseley and the Periodic Law

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.

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The Modern Periodic Table

The periodic table has undergone extensive change since Mendeleev's time (see Figure 5-6 on pages 130–131). 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,

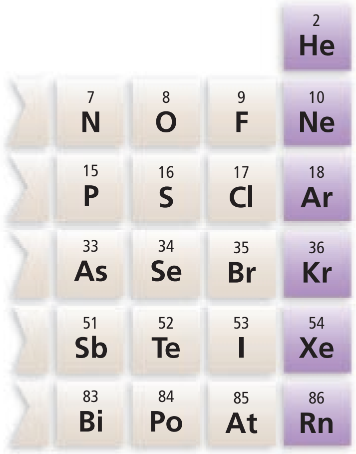


FIGURE 5-3 The noble gases, also known as the Group 18 elements, are all rather unreactive. As you will read, the reason for this low reactivity also accounts for the special place occupied by the noble gases in the periodic table.

another noble gas, helium, He, had been discovered as a component of the sun, based on the emission spectrum of sunlight. 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. 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 to place in his new group, krypton, Kr, and xenon, Xe. The final noble gas, radon, Rn, was discovered in 1900 by the German scientist Friedrich Ernst Dorn.

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, as shown in Figure 5-6 on pages 130–131.

Periodicity

Periodicity with respect to atomic number can be observed in any group of elements in the periodic table. Consider the noble gases of Group 18. The first noble gas is helium, He. It has an atomic number of 2. The elements following helium in atomic number have completely different properties until the next noble gas, neon, Ne, which has an atomic number of 10, is reached. 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). The differences in atomic number between successive noble gases are shown in Figure 5-4. Also shown in Figure 5-4 are atomic-number differences between the elements of Group 1, which 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.

Starting with the first member of Groups 13–17, a similar periodic pattern 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 5-2, 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.

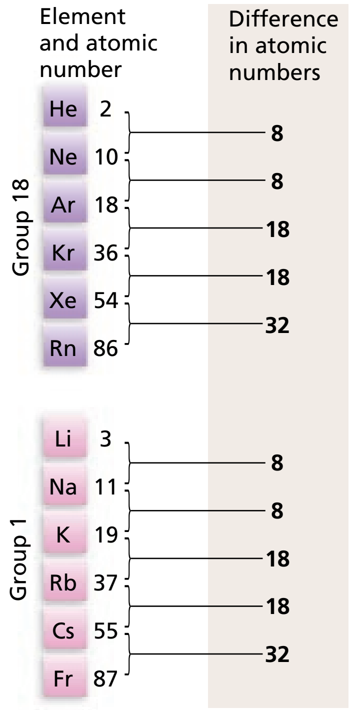


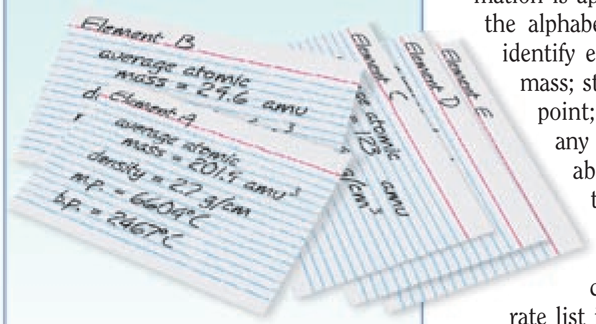
FIGURE 5-4 In each of Groups 1 and 18, the differences between the atomic numbers of successive elements are 8, 8, 18, 18, and 32, respectively. Groups 2 and 13–17 follow a similar pattern.

Quick LAB

Designing Your Own Periodic Table

Materials

- index cards



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 letters 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.

SECTION REVIEW

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 added to the periodic table after 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 5-4 as a guide.)

SECTION 5-2

OBJECTIVES

- Describe the relationship between electrons in sublevels and the length of each period of the periodic table.
- Locate and name the four blocks of the periodic table. Explain the reasons for these names.
- Discuss the relationship between group configurations and group numbers.
- Describe the locations in the periodic table and the general properties of the alkali metals, the alkaline-earth metals, the halogens, and the noble gases.

Electron Configuration and the Periodic Table

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.

Periods and Blocks of the Periodic Table

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*. (As shown in Figure 5-6, there are a total of seven periods of elements in the modern periodic table.) As can be seen in Table 5-1, the length of each period is determined by the number of electrons that can occupy the sublevels being filled in that period.

TABLE 5-1 ***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*, etc.

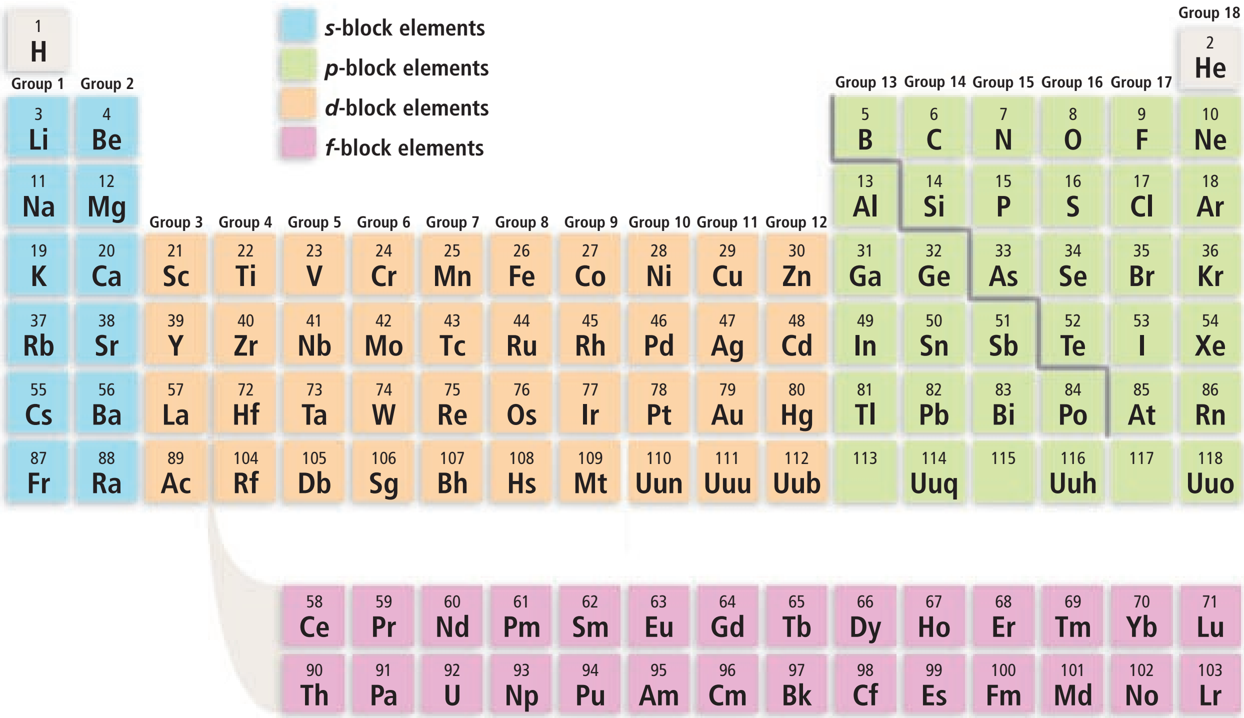
In the first period, the 1*s* sublevel is being filled. The 1*s* sublevel can hold a total of two electrons. Therefore, the first period consists of two elements—hydrogen and helium. In the second period, the 2*s* sublevel, which can hold two electrons, and the 2*p* sublevel, which can hold six electrons, 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 29 known 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]$3d^{10}4s^{2}4p^{3}$. The 4 in $4p^{3}$ indicates that arsenic's highest occupied energy level is the fourth energy level. Arsenic is therefore in the fourth period in the periodic table. The period and electron configuration for each element can be found in the periodic table on pages 130–131.

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 5-5. 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 5-5 Based on the electron configurations of the elements, the periodic table can be subdivided into four sublevel blocks.

Sublevel Blocks of the Periodic Table



🟦 *s*-block elements
🟩 *p*-block elements
🟧 *d*-block elements
🟪 *f*-block elements

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 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub	113	114 Uuq	115	116 Uuh	117	118 Uuo
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Periodic Table of the Elements

Key:

6 — Atomic number

C — Symbol

Carbon — Name

12.011 — Average atomic mass

[He]$2s^22p^2$ — Electron configuration

Period	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9
1	1 **H** Hydrogen 1.00794 $1s^1$								
2	3 **Li** Lithium 6.941 [He]$2s^1$	4 **Be** Beryllium 9.012182 [He]$2s^2$							
3	11 **Na** Sodium 22.989768 [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 44.955910 [Ar]$3d^14s^2$	22 **Ti** Titanium 47.88 [Ar]$3d^24s^2$	23 **V** Vanadium 50.9415 [Ar]$3d^34s^2$	24 **Cr** Chromium 51.9961 [Ar]$3d^54s^1$	25 **Mn** Manganese 54.93805 [Ar]$3d^54s^2$	26 **Fe** Iron 55.847 [Ar]$3d^64s^2$	27 **Co** Cobalt 58.93320 [Ar]$3d^74s^2$
5	37 **Rb** Rubidium 85.4678 [Kr]$5s^1$	38 **Sr** Strontium 87.62 [Kr]$5s^2$	39 **Y** Yttrium 88.90585 [Kr]$4d^15s^2$	40 **Zr** Zirconium 91.224 [Kr]$4d^25s^2$	41 **Nb** Niobium 92.90638 [Kr]$4d^45s^1$	42 **Mo** Molybdenum 95.94 [Kr]$4d^55s^1$	43 **Tc** Technetium (97.9072) [Kr]$4d^65s^1$	44 **Ru** Ruthenium 101.07 [Kr]$4d^75s^1$	45 **Rh** Rhodium 102.906 [Kr]$4d^85s^1$
6	55 **Cs** Cesium 132.90543 [Xe]$6s^1$	56 **Ba** Barium 137.327 [Xe]$6s^2$	57 **La** Lanthanum 138.9055 [Xe]$5d^16s^2$	72 **Hf** Hafnium 178.49 [Xe]$4f^{14}5d^26s^2$	73 **Ta** Tantalum 180.9479 [Xe]$4f^{14}5d^36s^2$	74 **W** Tungsten 183.84 [Xe]$4f^{14}5d^46s^2$	75 **Re** Rhenium 186.207 [Xe]$4f^{14}5d^56s^2$	76 **Os** Osmium 190.23 [Xe]$4f^{14}5d^66s^2$	77 **Ir** Iridium 192.22 [Xe]$4f^{14}5d^76s^2$
7	87 **Fr** Francium (223.0197) [Rn]$7s^1$	88 **Ra** Radium (226.0254) [Rn]$7s^2$	89 **Ac** Actinium (227.0278) [Rn]$6d^17s^2$	104 **Rf** Rutherfordium (261.11) [Rn]$5f^{14}6d^27s^2$	105 **Db** Dubnium (262.114) [Rn]$5f^{14}6d^37s^2$	106 **Sg** Seaborgium (263.118) [Rn]$5f^{14}6d^47s^2$	107 **Bh** Bohrium (262.12) [Rn]$5f^{14}6d^57s^2$	108 **Hs** Hassium (265)† [Rn]$5f^{14}6d^67s^2$	109 **Mt** Meitnerium (266)† [Rn]$5f^{14}6d^77s^2$

58 **Ce** Cerium 140.115 [Xe]$4f^15d^16s^2$	59 **Pr** Praseodymium 140.908 [Xe]$4f^36s^2$	60 **Nd** Neodymium 144.24 [Xe]$4f^46s^2$	61 **Pm** Promethium (144.9127) [Xe]$4f^56s^2$	62 **Sm** Samarium 150.36 [Xe]$4f^66s^2$
90 **Th** Thorium 232.0381 [Rn]$6d^27s^2$	91 **Pa** Protactinium 231.03588 [Rn]$5f^26d^17s^2$	92 **U** Uranium 238.0289 [Rn]$5f^36d^17s^2$	93 **Np** Neptunium (237.0482) [Rn]$5f^46d^17s^2$	94 **Pu** Plutonium 244.0642 [Rn]$5f^67s^2$

* Temporary name not yet approved by IUPAC.

† Estimated from currently available IUPAC data.

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FIGURE 5-6 In the common periodic table, the elements are arranged in vertical groups and in horizontal periods.

Metals

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

Metalloids

- Semiconductors

Nonmetals

- Halogens
- Other nonmetals
- Noble gases

Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18	Period
								2 **He** Helium 4.002602 $1s^2$	1
			5 **B** Boron 10.811 [He]$2s^22p^1$	6 **C** Carbon 12.011 [He]$2s^22p^2$	7 **N** Nitrogen 14.00674 [He]$2s^22p^3$	8 **O** Oxygen 15.9994 [He]$2s^22p^4$	9 **F** Fluorine 18.9984032 [He]$2s^22p^5$	10 **Ne** Neon 20.1797 [He]$2s^22p^6$	2
			13 **Al** Aluminum 26.981539 [Ne]$3s^23p^1$	14 **Si** Silicon 28.0855 [Ne]$3s^23p^2$	15 **P** Phosphorus 30.9738 [Ne]$3s^23p^3$	16 **S** Sulfur 32.066 [Ne]$3s^23p^4$	17 **Cl** Chlorine 35.4527 [Ne]$3s^23p^5$	18 **Ar** Argon 39.948 [Ne]$3s^23p^6$	3
28 **Ni** Nickel 58.6934 [Ar]$3d^84s^2$	29 **Cu** Copper 63.546 [Ar]$3d^{10}4s^1$	30 **Zn** Zinc 65.39 [Ar]$3d^{10}4s^2$	31 **Ga** Gallium 69.723 [Ar]$3d^{10}4s^24p^1$	32 **Ge** Germanium 72.61 [Ar]$3d^{10}4s^24p^2$	33 **As** Arsenic 74.92159 [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 83.80 [Ar]$3d^{10}4s^24p^6$	4
46 **Pd** Palladium 106.42 [Kr]$4d^{10}5s^0$	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 121.757 [Kr]$4d^{10}5s^25p^3$	52 **Te** Tellurium 127.60 [Kr]$4d^{10}5s^25p^4$	53 **I** Iodine 126.904 [Kr]$4d^{10}5s^25p^5$	54 **Xe** Xenon 131.29 [Kr]$4d^{10}5s^25p^6$	5
78 **Pt** Platinum 195.08 [Xe]$4f^{14}5d^96s^1$	79 **Au** Gold 196.96654 [Xe]$4f^{14}5d^{10}6s^1$	80 **Hg** Mercury 200.59 [Xe]$4f^{14}5d^{10}6s^2$	81 **Tl** Thallium 204.3833 [Xe]$4f^{14}5d^{10}6s^26p^1$	82 **Pb** Lead 207.2 [Xe]$4f^{14}5d^{10}6s^26p^2$	83 **Bi** Bismuth 208.98037 [Xe]$4f^{14}5d^{10}6s^26p^3$	84 **Po** Polonium (208.9824) [Xe]$4f^{14}5d^{10}6s^26p^4$	85 **At** Astatine (209.9871) [Xe]$4f^{14}5d^{10}6s^26p^5$	86 **Rn** Radon (222.0176) [Xe]$4f^{14}5d^{10}6s^26p^6$	6
110 **Uun*** Ununnilium (269)† [Rn]$5f^{14}6d^97s^1$	111 **Uuu*** Unununium (272)† [Rn]$5f^{14}6d^{10}7s^1$	112 **Uub*** Ununbium (277)† [Rn]$5f^{14}6d^{10}7s^2$	113	114 **Uuq*** Ununquadium (285)† [Rn]$5f^{14}6d^{10}7s^27p^2$	115	116 **Uuh*** Ununhexium (289)† [Rn]$5f^{14}6d^{10}7s^27p^4$	117	118 **Uuo*** Ununoctium (293)† [Rn]$5f^{14}6d^{10}7s^27p^6$	7

63 **Eu** Europium 151.966 [Xe]$4f^76s^2$	64 **Gd** Gadolinium 157.25 [Xe]$4f^75d^16s^2$	65 **Tb** Terbium 158.92534 [Xe]$4f^96s^2$	66 **Dy** Dysprosium 162.50 [Xe]$4f^{10}6s^2$	67 **Ho** Holmium 164.930 [Xe]$4f^{11}6s^2$	68 **Er** Erbium 167.26 [Xe]$4f^{12}6s^2$	69 **Tm** Thulium 168.93421 [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 (243.0614) [Rn]$5f^77s^2$	96 **Cm** Curium (247.0703) [Rn]$5f^76d^17s^2$	97 **Bk** Berkelium (247.0703) [Rn]$5f^97s^2$	98 **Cf** Californium (251.0796) [Rn]$5f^{10}7s^2$	99 **Es** Einsteinium (252.083) [Rn]$5f^{11}7s^2$	100 **Fm** Fermium (257.0951) [Rn]$5f^{12}7s^2$	101 **Md** Mendelevium (258.10) [Rn]$5f^{13}7s^2$	102 **No** Nobelium (259.1009) [Rn]$5f^{14}7s^2$	103 **Lr** Lawrencium (262.11) [Rn]$5f^{14}6d^17s^2$

The atomic masses listed in this table reflect the precision of current measurements. (Values listed in parentheses are those of the element's most stable or most common isotope.) In calculations throughout the text, however, atomic masses have been rounded to two places to the right of the decimal.

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 5-3, the ease with which the single electron is lost helps to 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.

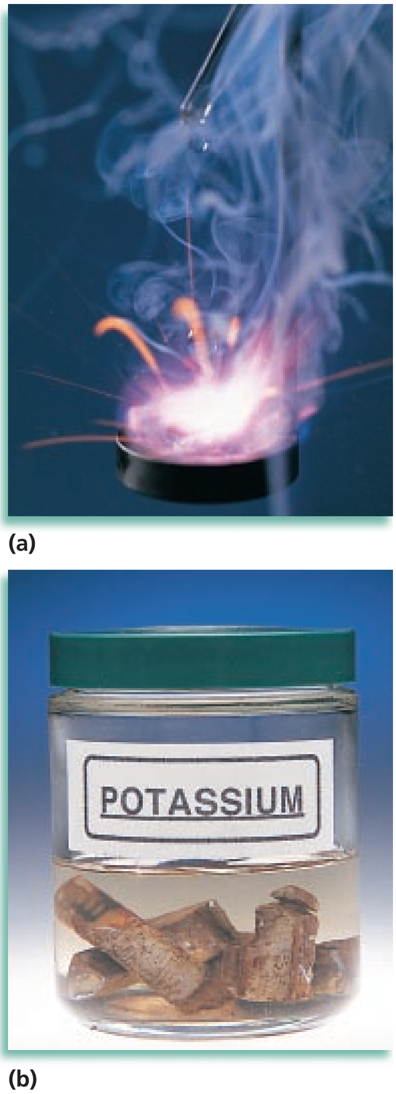

FIGURE 5-7 (a) Like other alkali metals, potassium reacts so strongly with water that (b) it must be stored in kerosene or oil to prevent it from reacting with moisture in the air.

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

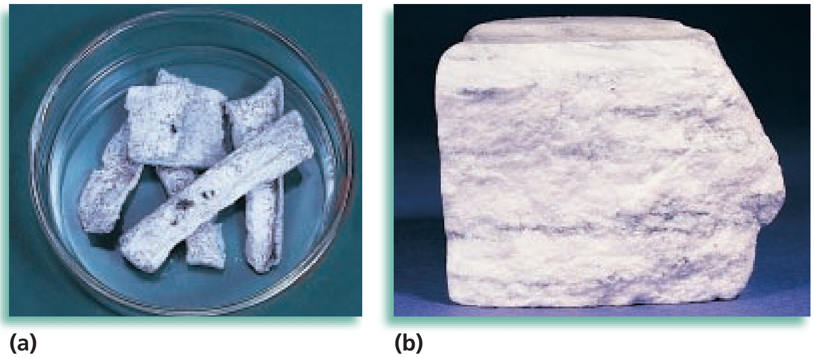

FIGURE 5-8 Calcium, an alkaline-earth metal, is too reactive to be found in nature in its pure state (a). Instead, it exists in compounds, such as in the minerals that make up marble (b).

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.

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SAMPLE PROBLEM 5-1

a. Without looking at the periodic table, give the group, period, and block in which the element with 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 active or less active than the element described in (a)?

SOLUTION

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 1*s*, 2*s*, and 2*p* sublevels are completely filled (see Table 5-1). 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.

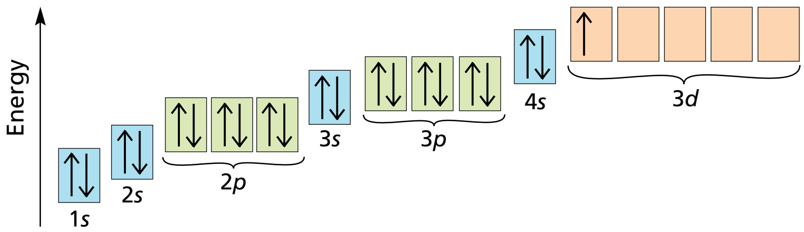
$$1s^2 2s^2 2p^6 3s^1 \quad \text{or} \quad [\text{Ne}]3s^1$$

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

	Answer
1. Without looking at the periodic table, give the group, period, and block in which the element with the electron configuration [Kr]5s^1 is located.	1. Group 1, fifth period, *s* block
2. a. Without looking at the periodic table, write the group configuration for the Group 2 elements.	2. a. ns^2
b. Without looking at the periodic table, write the complete electron configuration for the Group 2 element in the fourth period.	b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
c. Refer to Figure 5-6 to identify the element described in (b). Then write the element's noble-gas notation.	c. Ca, [Ar]4s^2
d. How does the reactivity of the element in (b) compare with the reactivity of the element in Group 1 of the same period?	d. The element is in Group 2, so it is probably less reactive than the Group 1 element of the same period.

FIGURE 5-9 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 3*d* sublevel is slightly higher in energy than the 4*s* sublevel, so these are filled in the order 4*s*3*d* (see Figure 5-9). 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 $(n - 1)d^{10}ns^2$.

As you read in Chapter 4, 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 unreactive that they do not easily form compounds, existing 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 5-10.

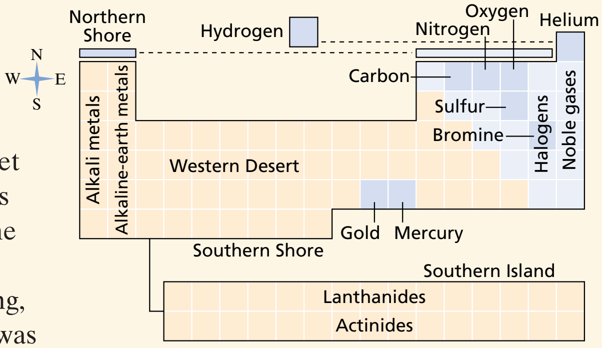
FIGURE 5-10 Mercury, tungsten, and vanadium are transition elements. Locate them in the *d* block of the periodic table on page 129.


Mercury Tungsten Vanadium

CHEMICAL COMMENTARY

The Wild Kingdom

From *The Periodic Kingdom: A Journey Into the World of the Chemical Elements* by P. W. Atkins

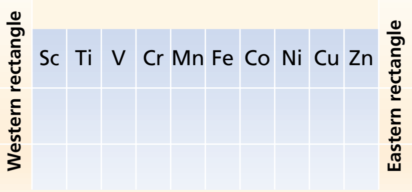


The general layout of the Periodic Kingdom.

Welcome to the Periodic Kingdom. This is a land of the imagination, but it is closer to reality than it appears to be. This is the kingdom of the chemical elements, the substances from which everything tangible is made. It is not an extensive country, for it consists of only a hundred or so regions . . . yet it accounts for everything material in our actual world. From the hundred elements that are at the center of our story, all planets, rocks, vegetation, and animals are made. These elements are the basis of the air, the oceans, and the Earth itself. We stand on the elements, we eat the elements, we *are* the elements. Because our brains are made up of elements, even our opinions are, in a sense, properties of the elements and hence inhabitants of the kingdom.

Even from . . . far above the country, we can see broad features of the landscape. There are the glittering, lustrous regions made up of metals and lying together in what we shall call the Western Desert. This desert is broadly uniform, but there is a subtlety of shades, indicating a variety of characteristics. Here and there are gentle splashes of color, such as the familiar glint of gold and the blush of copper. How remarkable it is that these desert lands make up so much of the kingdom . . . yet the kingdom supplies such luxuriance in the real world!

Generally speaking, the Western Desert was explored and exploited from east to west; that is, technology and industry made use of them in that order. Copper displaced stone to give us the Bronze Age. Then, as the explorers pressed westward, applying ever more vigorous means of discovery, they encountered iron and used it to fabricate more effective weaponry . . . The strongest states enjoyed freedom from constant aggression, and this gave them time for scholarship; thus in due course explorers were able to penetrate into more distant western regions of the desert.



The Isthmus

Deep in the Western Desert, in an isthmuslike zone running from zinc on the east to scandium on the west, they finally stumbled upon titanium, a remarkable prize indeed. Titanium has exactly the properties that a society bent on high technology needs if it is to take to the skies: this is a metal that is tough and resistant to corrosion, yet light, and it is typical of its part of the Western Desert. Titanium and its neighbors vanadium and molybdenum, in alliance with iron, form the durable steels that enable us to chop through stone and build on a massive scale. It is a remarkable feature of the Periodic Kingdom that the Isthmus . . . has provided so many of the workhorses of our society, and that its members so readily form alliances with one another.

Reading for Meaning

What does Atkins mean when he says that metals of the Isthmus form “alliances” with one another? What does this imply about these elements?

Read Further

Atkins remarks that titanium has properties that are well-suited for a high-tech society. Find out how we obtain titanium, and list five ways that it is used in our society.

SAMPLE PROBLEM 5-2

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.

SOLUTION The number of the highest occupied energy level is five, so the element is in the fifth period. There are five electrons in the *d* sublevel. This means that the *d* sublevel is incompletely filled because it 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 is the Group 6 element in the fifth period. The element is molybdenum. The others in Group 6 are chromium, tungsten, and seaborgium.

PRACTICE

	Answer
1. Without looking at the periodic table, identify the period, block, and group in which the element with the electron configuration [Ar]$3d^8 4s^2$ is located.	**1.** fourth period, *d* block, Group 10
2. a. Without looking at the periodic table, write the outer electron configuration for the Group 12 element in the fifth period.	**2. a.** $4d^{10} 5s^2$
b. Refer to the periodic table to identify the element described in (a) and to write the element's noble-gas notation.	**b.** Cd, [Kr]$4d^{10} 5s^2$

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 Table 5-2.

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,

TABLE 5-2 ***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

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 5-6 and in the periodic table printed on the inside back cover of this textbook.

The elements of Group 17 (fluorine, chlorine, bromine, iodine, and astatine) are known as the **halogens.** The halogens are the most reactive nonmetals. They react vigorously with most metals to form examples of the type of compound known as salts. As you will see later, 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 only very small quantities. Most of its properties are estimated, although it is known to be a solid.

The metalloids, or semiconducting elements, fall on both sides of a line separating 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.

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FIGURE 5-11 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 on page 129.

Fluorine

Chlorine

Bromine

Iodine

SAMPLE PROBLEM 5-3

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.

SOLUTION 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 ten, 14 – 10 = 4. With two electrons in the *s* sublevel, two electrons must also be present in the 2*p* sublevel, giving an outer electron configuration of $2s^22p^2$. The element is carbon, C, which is a nonmetal.

PRACTICE

	Answer
1. a. Without looking at the periodic table, write the outer electron configuration for the Group 17 element in the third period.	**1.** a. $3s^23p^5$
b. Name the element described in (a), and identify it as a metal, nonmetal, or metalloid.	b. chlorine, nonmetal
2. a. Without looking at the periodic table, identify the period, block, and group of an element with the electron configuration $[Ar]3d^{10}4s^24p^3$.	**2.** a. fourth period, *p* block, Group 15
b. Name the element described in (a), and identify it as a metal, nonmetal, or metalloid.	b. arsenic, 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. Their position 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.

SAMPLE PROBLEM 5-4

The electron configurations of atoms of four elements are written at the top of page 139. For each element, name the block and group in the periodic table in which it is located. Then name the element by consulting the periodic table on pages 130–131. Identify each

element as a metal, nonmetal, or metalloid. Finally, describe it as likely to be of high reactivity or of low reactivity.

a. $[Xe]4f^{14}5d^{9}6s^{1}$

b. $[Ne]3s^{2}3p^{5}$

c. $[Ne]3s^{2}3p^{6}$

d. $[Xe]4f^{6}6s^{2}$

SOLUTION

a. The $4f$ sublevel is filled with 14 electrons. The $5d$ sublevel is partially filled with nine electrons. Therefore, this is a d-block element. 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 this to be a p-block element. With a total of seven electrons in the ns and np sublevels, this is a Group 17 element, a halogen. 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, an unreactive nonmetal and a noble gas.

d. The incomplete $4f$ sublevel shows that it is an f-block element and a lanthanide. Group numbers are not assigned to the f block. The element is samarium, Sm. The lanthanides are all reactive metals.

PRACTICE

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

Answer

1\. a. p block, second period, Group 17, halogens, fluorine, nonmetal, high reactivity
b. d block, fourth period, Group 11, transition elements, copper, metal, low reactivity
c. s block, fifth period, Group 1, alkali metals, rubidium, metal, high reactivity

SECTION REVIEW

1. Into what four blocks can the periodic table be divided to illustrate the relationship between the elements' electron configurations and their placement in the periodic table?

2. What name is given to each of the following groups of elements on 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 with the electron configuration $[Ar]3d^{7}4s^{2}$.

SECTION 5-3

Electron Configuration and Periodic Properties

OBJECTIVES

- Define *atomic* and *ionic radii, ionization energy, electron affinity,* and *electronegativity.*
- Compare the periodic trends of atomic radii, ionization energy, and electronegativity, and state the reasons for these variations.
- Define *valence electrons,* and state how many are present in atoms of each main-group element.
- Compare the atomic radii, ionization energies, and electronegativities of the *d*-block elements with those of the main-group elements.

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.

Atomic Radii

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, to estimate the size of an atom, the conditions under which the atom exists must be specified. 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, then divide this distance by two. As illustrated in Figure 5-12, **atomic radius** *may be defined as one-half the distance between the nuclei of identical atoms that are bonded together.*

Period Trends

Figure 5-13 gives the atomic radii of the elements and Figure 5-14 presents this information graphically. Note that there is a gradual decrease in atomic radii across the second period from lithium, Li, to neon, Ne.

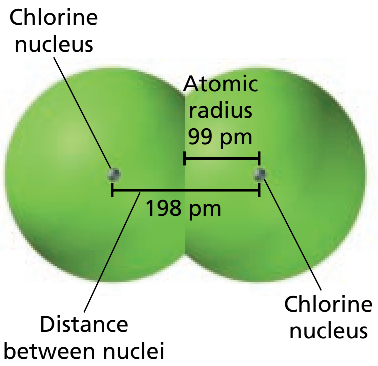
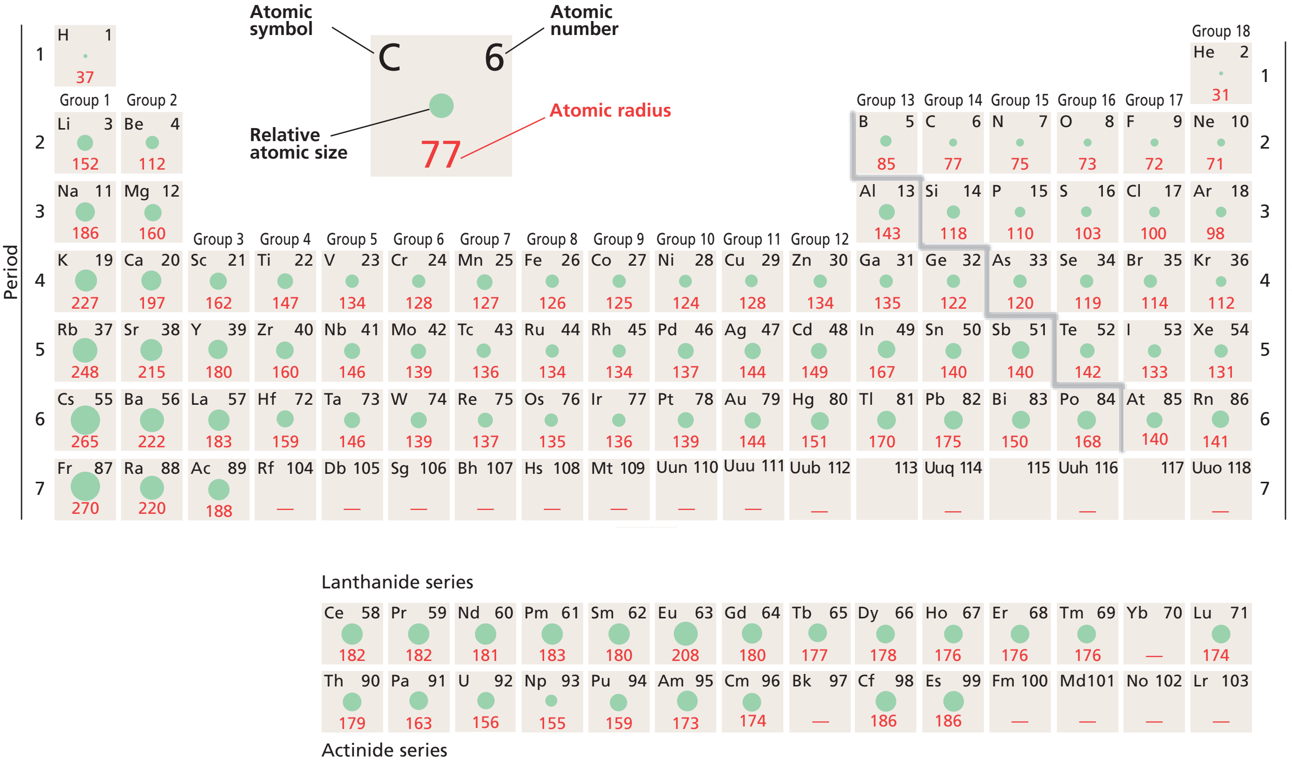


FIGURE 5-12 One method of determining atomic radius is to measure the distance between the nuclei of two identical atoms that are bonded together in an element or compound, then divide this distance by two. The atomic radius of a chlorine atom, for example, is 99 picometers (pm).

Periodic Table of Atomic Radii (pm)



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	H 1 37																	He 2 31
2	Li 3 152	Be 4 112											B 5 85	C 6 77	N 7 75	O 8 73	F 9 72	Ne 10 71
3	Na 11 186	Mg 12 160											Al 13 143	Si 14 118	P 15 110	S 16 103	Cl 17 100	Ar 18 98
4	K 19 227	Ca 20 197	Sc 21 162	Ti 22 147	V 23 134	Cr 24 128	Mn 25 127	Fe 26 126	Co 27 125	Ni 28 124	Cu 29 128	Zn 30 134	Ga 31 135	Ge 32 122	As 33 120	Se 34 119	Br 35 114	Kr 36 112
5	Rb 37 248	Sr 38 215	Y 39 180	Zr 40 160	Nb 41 146	Mo 42 139	Tc 43 136	Ru 44 134	Rh 45 134	Pd 46 137	Ag 47 144	Cd 48 149	In 49 167	Sn 50 140	Sb 51 140	Te 52 142	I 53 133	Xe 54 131
6	Cs 55 265	Ba 56 222	La 57 183	Hf 72 159	Ta 73 146	W 74 139	Re 75 137	Os 76 135	Ir 77 136	Pt 78 139	Au 79 144	Hg 80 151	Tl 81 170	Pb 82 175	Bi 83 150	Po 84 168	At 85 140	Rn 86 141
7	Fr 87 270	Ra 88 220	Ac 89 188	Rf 104 —	Db 105 —	Sg 106 —	Bh 107 —	Hs 108 —	Mt 109 —	Uun 110 —	Uuu 111 —	Uub 112 —	113	Uuq 114 —	115	Uuh 116 —	117	Uuo 118 —

Lanthanide series

Ce 58 182	Pr 59 182	Nd 60 181	Pm 61 183	Sm 62 180	Eu 63 208	Gd 64 180	Tb 65 177	Dy 66 178	Ho 67 176	Er 68 176	Tm 69 176	Yb 70 —	Lu 71 174
Th 90 179	Pa 91 163	U 92 156	Np 93 155	Pu 94 159	Am 95 173	Cm 96 174	Bk 97 —	Cf 98 186	Es 99 186	Fm 100 —	Md101 —	No 102 —	Lr 103 —

Actinide series

FIGURE 5-13 Atomic radii decrease from left to right across a period and increase down a group.

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 5-13.

Group Trends

Examine the atomic radii of the Group 1 elements in Figure 5-13. Notice that the radii of the elements increase as you read down the group. As electrons occupy sublevels in successively higher main energy levels located 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.

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Module 3: Periodic Properties

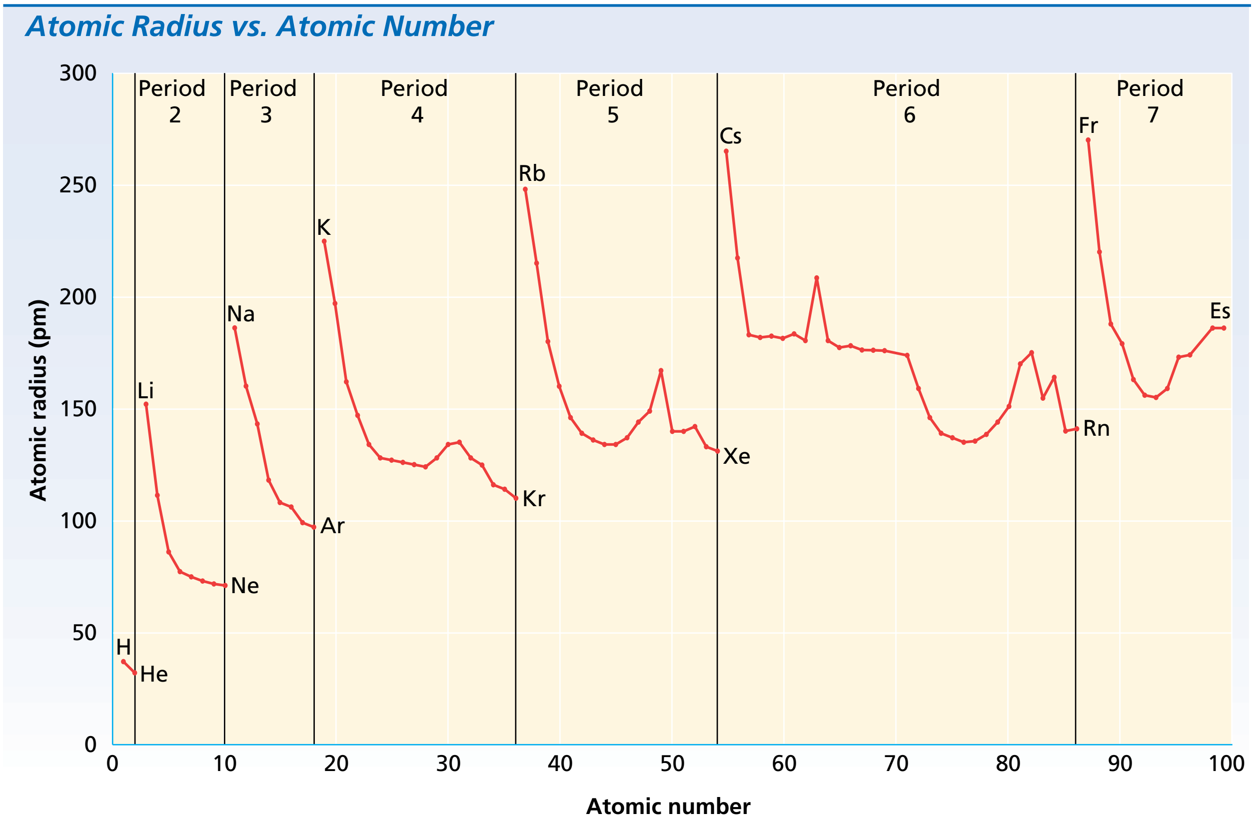


FIGURE 5-14 The plot of atomic radius versus atomic number shows period and group trends.

SAMPLE PROBLEM 5-5

a. 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.

b. 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.

SOLUTION

a. 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.

b. All of the elements are in Group 2. Of the four, barium has the highest atomic number and is farthest down the group. Therefore, barium has the largest atomic radius because atomic radii increase down a group.

PRACTICE

1. Of the elements Li, O, C, and F, identify the one with the largest atomic radius and the one with the smallest atomic radius. *Answer* Li, F
2. Of the elements Br, At, F, I, and Cl, identify the one with the smallest atomic radius and the one with the largest atomic radius. *Answer* F, At

Ionization 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.

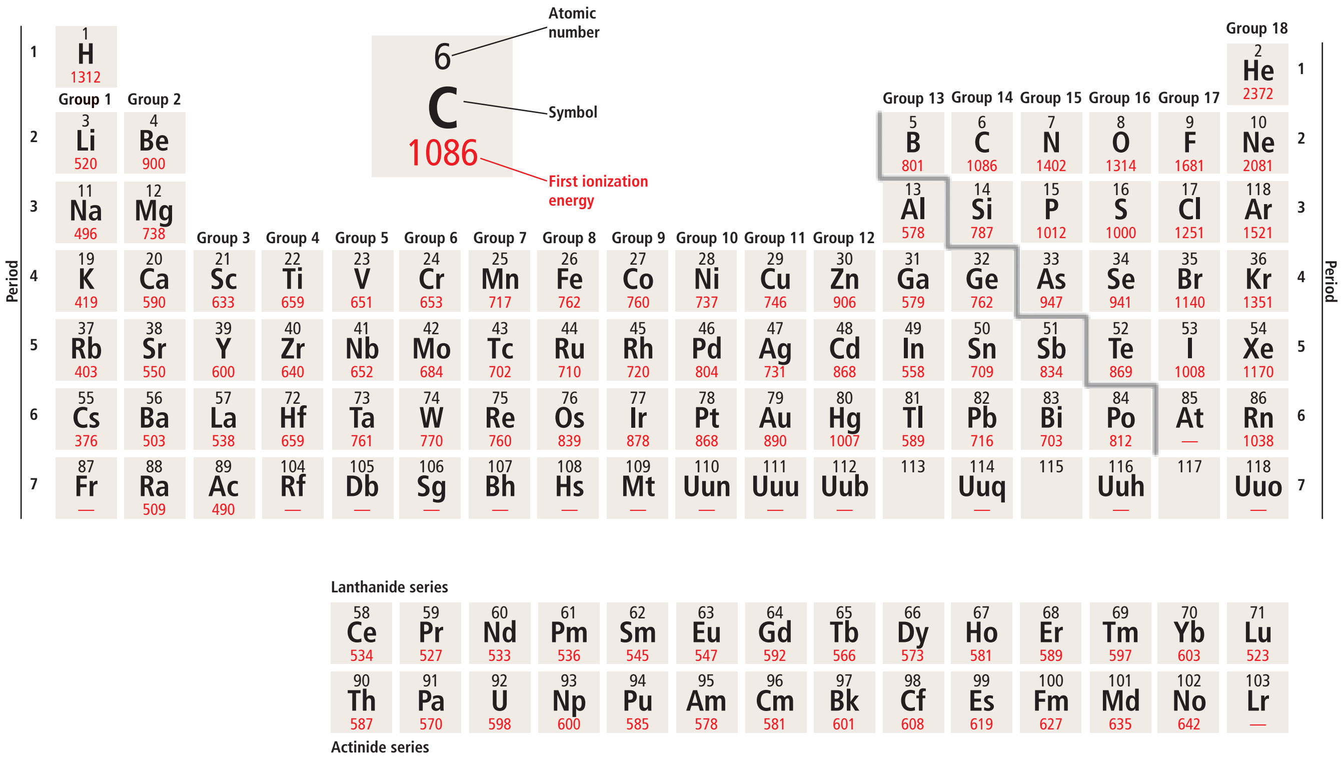
$$A + \text{energy} \rightarrow A^{+} + e^{-}$$

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 5-15 gives the first ionization energies for the elements in kilojoules per mole (kJ/mol). Figure 5-16 presents this information graphically.

FIGURE 5-15 In general, first ionization energies increase across a period and decrease down a group.

Periodic Table of Ionization Energies (kJ/mol)



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	Period
1	1 H 1312																	2 He 2372	1
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	2
3	11 Na 496	12 Mg 738											13 Al 578	14 Si 787	15 P 1012	16 S 1000	17 Cl 1251	118 Ar 1521	3
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	4
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	5
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	6
7	87 Fr —	88 Ra 509	89 Ac 490	104 Rf —	105 Db —	106 Sg —	107 Bh —	108 Hs —	109 Mt —	110 Uun —	111 Uuu —	112 Uub —	113	114 Uuq —	115	116 Uuh —	117	118 Uuo —	7

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 —

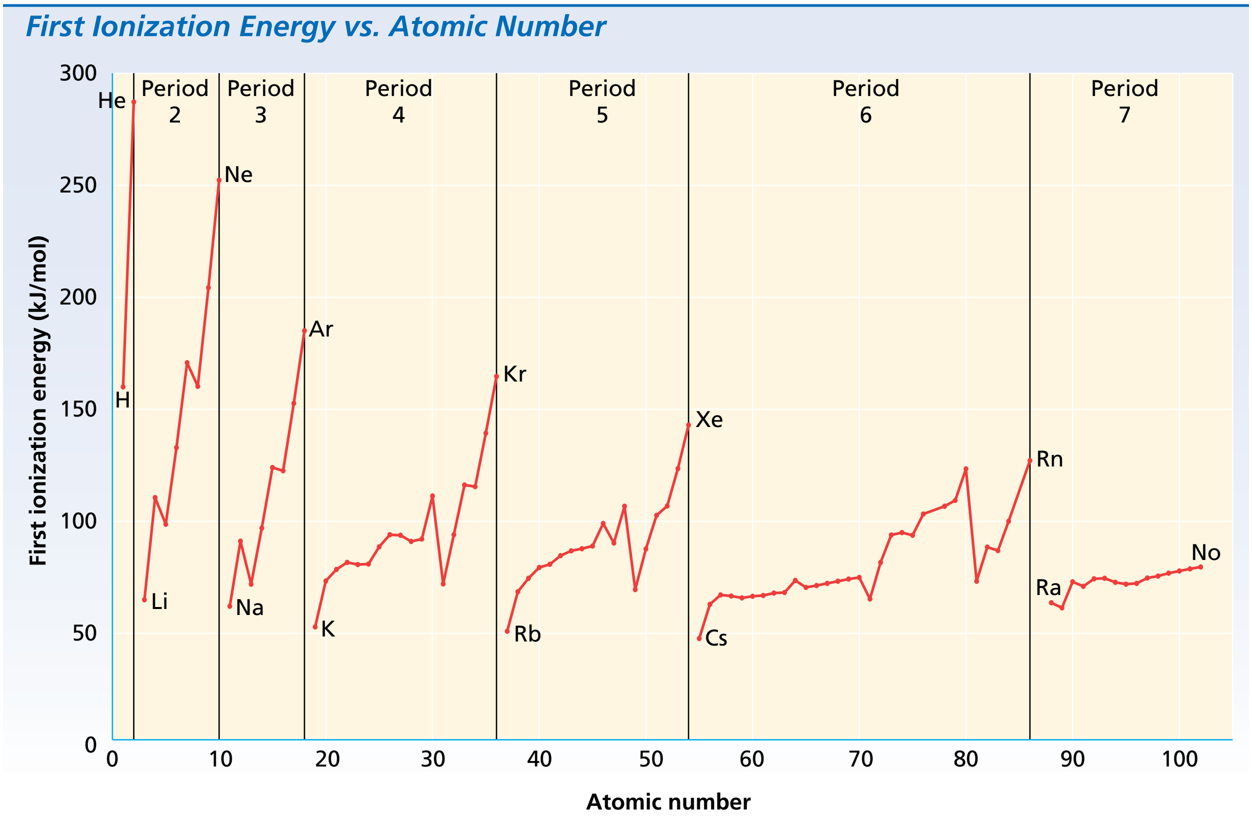


FIGURE 5-16 Plot of first ionization energy, IE_1, versus atomic number. As atomic number increases, both the period and the group trends become less pronounced.

Period Trends

In Figures 5-15 and 5-16, 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.

Table 5-3 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. 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 Table 5-3 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 Table 5-3 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

TABLE 5-3 ***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

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 ($[\mathrm{He}]2s^1$) leaves the helium noble-gas configuration. The removal of four electrons from a carbon atom ($[\mathrm{He}]2s^2 2p^2$) also leaves the helium configuration. A bigger table would show that this trend continues across the entire periodic system.

SAMPLE PROBLEM 5-6

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. For each element, decide if it is more likely to be in the *s* block or *p* block. Which element is more likely to form a positive ion?

SOLUTION

Element A has a very low ionization energy, meaning that atoms of A lose electrons easily. Therefore, element A is most likely to be a metal of the *s* block because ionization energies increase across the periods. Element B has a very high ionization energy, meaning that atoms of B lose electrons with difficulty. We would expect element B to 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 does element B.

PRACTICE

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

Q: $3s^2 3p^5$ R: $3s^1$ T: $4d^{10}5s^2 5p^5$ X: $4d^{10}5s^2 5p^1$

	Answer
a. Identify the block location of each hypothetical main-group element.	1. a. Q is in the *p* block, R is in the *s* block, T is in the *p* block, and X is in the *p* block.
b. Which of these elements are in the same period? Which are in the same group?	b. Q and R, and X and T are in the same period. Q and T 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?	c. Q would have the highest ionization energy, and R would have the lowest.
d. Which element would you expect to have the highest second ionization energy?	d. R
e. Which of the elements is most likely to form a 1+ ion?	e. R

Electron Affinity

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.

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

In this book, the quantity of energy released is represented by a negative number. 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 is represented by a positive number. An ion produced in this way will be unstable and will lose the added electron spontaneously.

Figure 5-17 shows the electron affinity in kilojoules per mole for the elements. Figure 5-18 presents these data graphically.

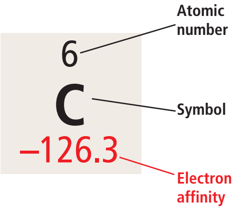
Period Trends

Among the elements of each period, the halogens (Group 17) gain electrons most readily. This is indicated in Figure 5-17 by the large negative values of halogens' electron affinities. The ease with which halogen atoms gain electrons is a major reason for the high reactivities of the 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. 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 much more easily

CHEMISTRY INTERACTIVE TUTOR
Module 3: Periodic Properties

FIGURE 5-17 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.

Periodic Table of Electron Affinities (kJ/mol)



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 Uun —	111 Uuu —	112 Uub —	113	114 Uuq —	115	116 Uuh —	117	118 Uuo —

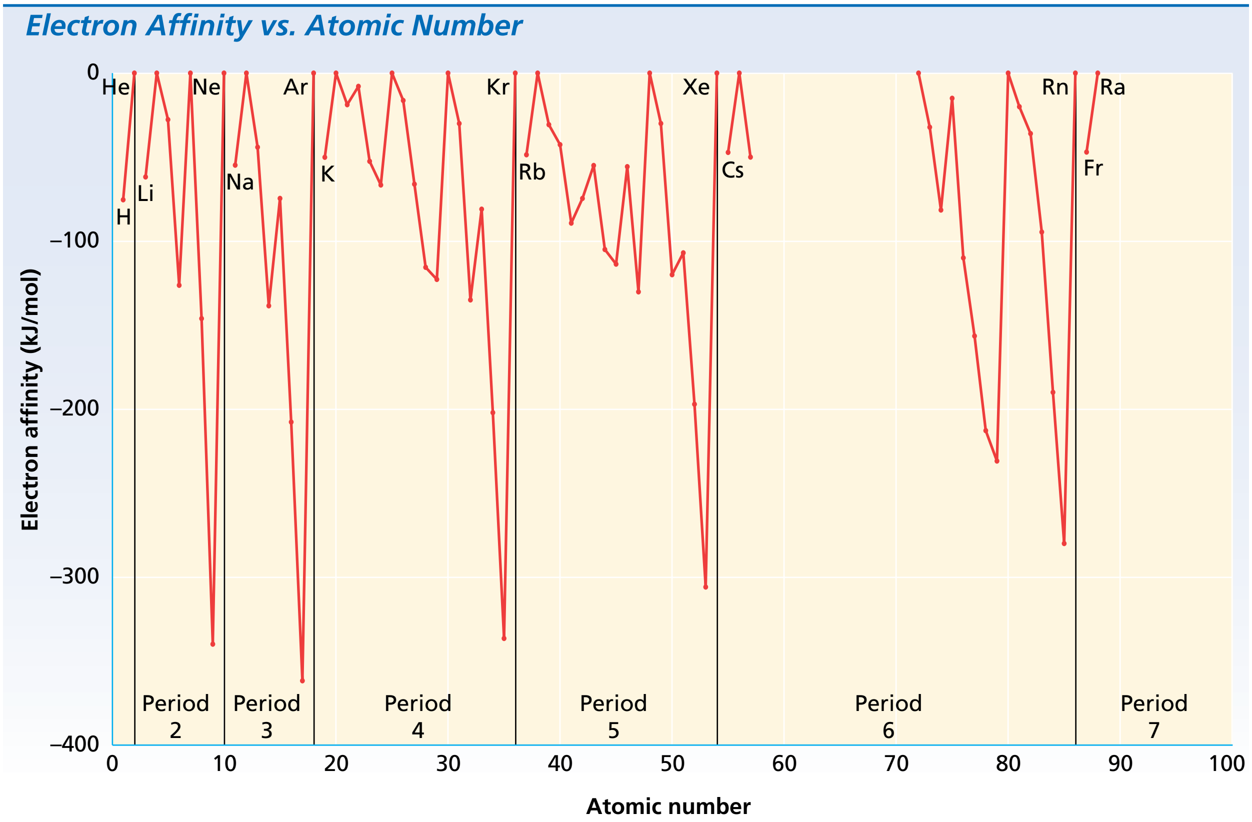


FIGURE 5-18 The plot of electron affinity versus atomic number shows that most atoms release energy when they acquire an electron, as indicated by negative values.

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 add 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 always 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]$2s^2 2p^4$) achieves the configuration of the noble gas neon by adding two electrons to form the ion O^{2-}([He]$2s^2 2p^6$). Nitrogen achieves a neon configuration by adding three electrons to form the ion N^{3-}.

Ionic Radii

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Module 3: Periodic Properties

Figure 5-19 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.

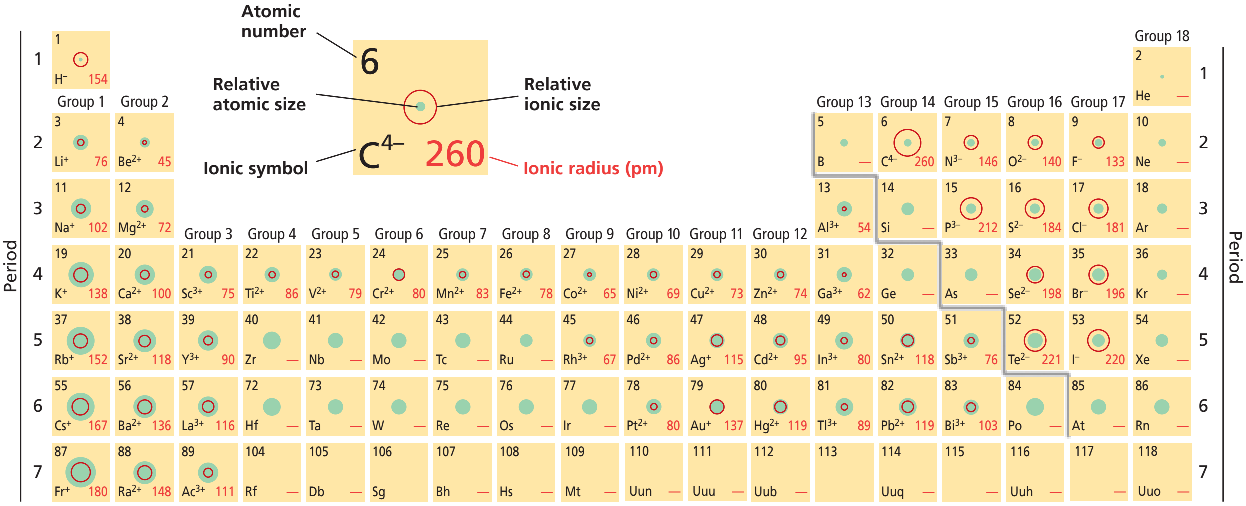
A negative ion is known as an **anion.** The formation of an anion by the addition of one or more electrons always leads to an increase in atomic radius. This is because the total positive charge of the nucleus remains unchanged when an electron is added to an atom or an ion. So the electrons are not 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

FIGURE 5-19 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.

Periodic Table of Ionic Radii (pm)



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^- 154																	2 He —
2	3 Li^+ 76	4 Be^{2+} 45											5 B —	6 C^{4-} 260	7 N^{3-} 146	8 O^{2-} 140	9 F^- 133	10 Ne —
3	11 Na^+ 102	12 Mg^{2+} 72											13 Al^{3+} 54	14 Si —	15 P^{3-} 212	16 S^{2-} 184	17 Cl^- 181	18 Ar —
4	19 K^+ 138	20 Ca^{2+} 100	21 Sc^{3+} 75	22 Ti^{2+} 86	23 V^{2+} 79	24 Cr^{2+} 80	25 Mn^{2+} 83	26 Fe^{2+} 78	27 Co^{2+} 65	28 Ni^{2+} 69	29 Cu^{2+} 73	30 Zn^{2+} 74	31 Ga^{3+} 62	32 Ge —	33 As —	34 Se^{2-} 198	35 Br^- 196	36 Kr —
5	37 Rb^+ 152	38 Sr^{2+} 118	39 Y^{3+} 90	40 Zr —	41 Nb —	42 Mo —	43 Tc —	44 Ru —	45 Rh^{3+} 67	46 Pd^{2+} 86	47 Ag^+ 115	48 Cd^{2+} 95	49 In^{3+} 80	50 Sn^{2+} 118	51 Sb^{3+} 76	52 Te^{2-} 221	53 I^- 220	54 Xe —
6	55 Cs^+ 167	56 Ba^{2+} 136	57 La^{3+} 116	72 Hf —	73 Ta —	74 W —	75 Re —	76 Os —	77 Ir —	78 Pt^{2+} 80	79 Au^+ 137	80 Hg^{2+} 119	81 Tl^{3+} 89	82 Pb^{2+} 119	83 Bi^{3+} 103	84 Po —	85 At —	86 Rn —
7	87 Fr^+ 180	88 Ra^{2+} 148	89 Ac^{3+} 111	104 Rf —	105 Db —	106 Sg	107 Bh —	108 Hs —	109 Mt —	110 Uun —	111 Uuu —	112 Uub —	113	114 Uuq —	115 —	116 Uuh —	117 —	118 Uuo —

the same main energy level. 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

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

Valence Electrons

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 Table 5-4. 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. In other cases, only the electrons from the *p* sublevel are involved.

TABLE 5-4 ***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

Electronegativity

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.

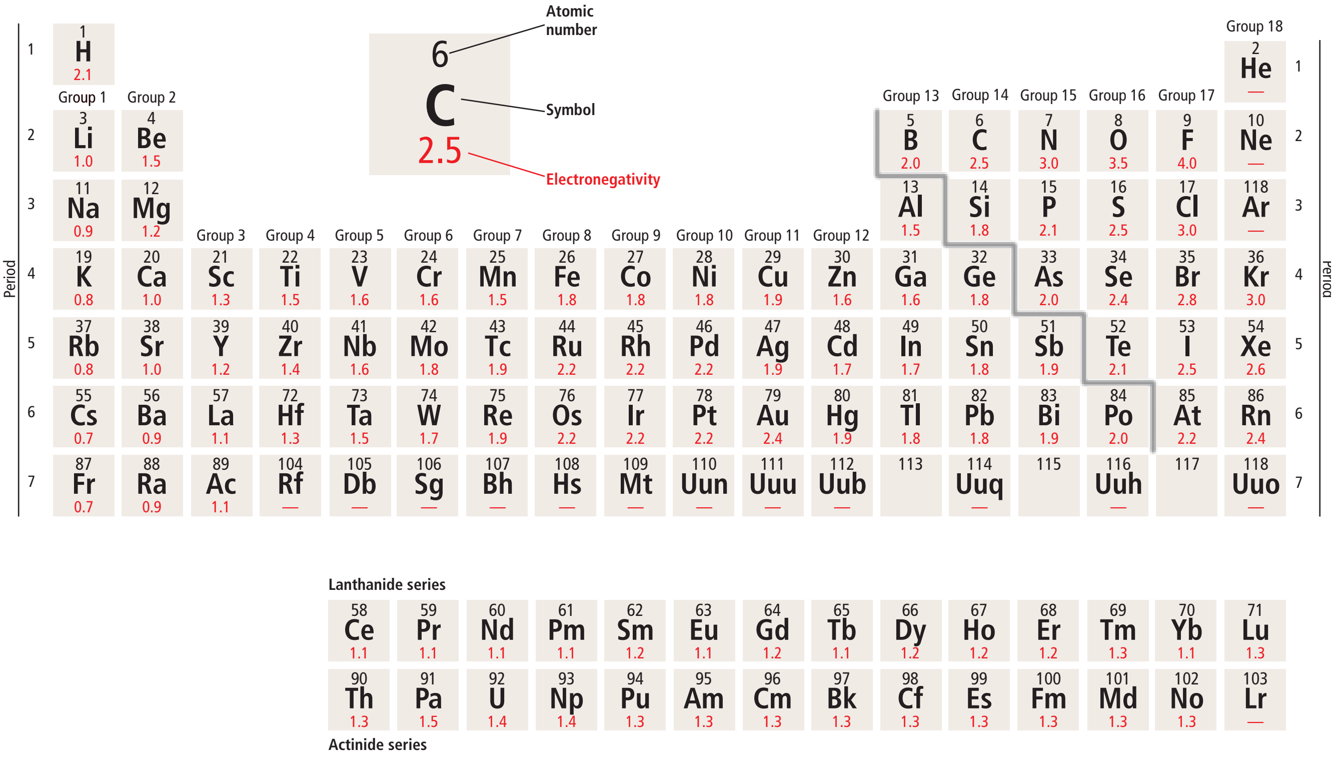
Linus Pauling, one of America's most famous chemists, devised a scale of numerical values reflecting the tendency of an atom to attract electrons. **Electronegativity** *is a measure of the ability of an atom in a chemical compound to attract electrons.* The most electronegative element, fluorine, is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

Period Trends

As shown in Figure 5–20, 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 in compounds. *Electronegativities tend to either decrease down a group or*

FIGURE 5-20 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.

Periodic Table of Electronegativities



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	118 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 Uun —	111 Uuu —	112 Uub —	113	114 Uuq —	115	116 Uuh —	117	118 Uuo —

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	70 Yb 1.1	71 Lu 1.3
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	102 No 1.3	103 Lr —

Actinide series

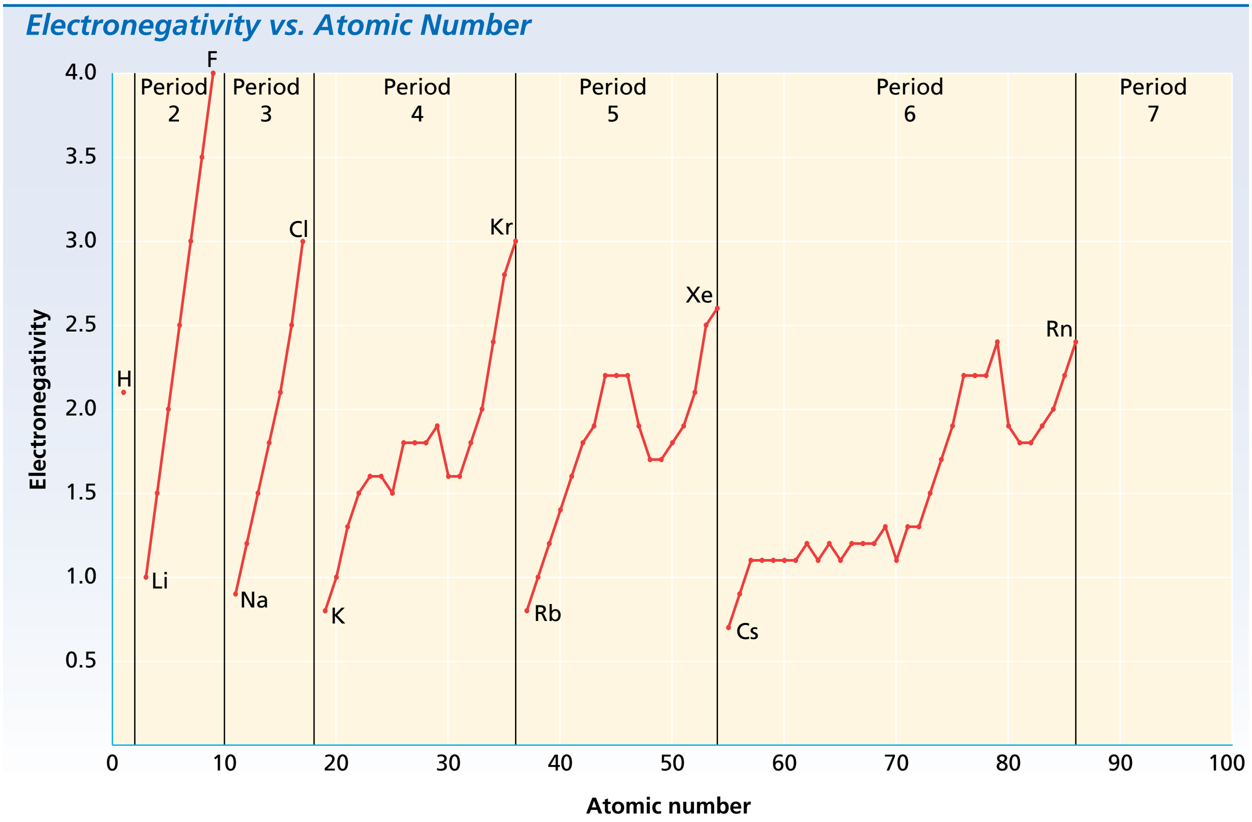


FIGURE 5-21 The plot shows electronegativity versus atomic number for Periods 1–6.

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 5-21.

SAMPLE PROBLEM 5-7

Among the elements gallium, Ga, bromine, Br, and calcium, Ca, which has the highest electronegativity? Explain why in terms of periodic trends.

SOLUTION

The elements are all in the fourth period. Bromine has the highest atomic number and is farthest to the right in the period. Therefore, it should have the highest electronegativity because electronegativity increases across the periods.

PRACTICE

1. Consider the five hypothetical main-group elements E, G, J, L, and M with the outer electron configurations shown at the top of page 153.

$E = 2s^22p^5 \quad G = 4d^{10}5s^25p^5 \quad J = 2s^22p^2$

$L = 5d^{10}6s^26p^5 \quad M = 2s^22p^4$

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 its neutral atom.

d. Which element(s) contains seven valence electrons?

Answer

1. a. All are in the *p* block. E, J, and M are in the same period, and E, G, and L are in the same group.

   b. E should have the highest electron affinity; E, G, and L are most likely to form 1– ions; E should have the highest electronegativity.

   c. The ionic radius would be larger.

   d. E, G, and L

Periodic Properties of the *d*- and *f*-Block Elements

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 5-14 and 5-16, 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. 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 5-14 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

The order in which electrons are removed from all atoms of the *d*-block and *f*-block elements is *exactly the reverse of the order given by the electron-configuration notation.* In other words, 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 $[Ar]3d^64s^2$. It loses a 4*s* electron first to form Fe^+ ($[Ar]3d^64s^1$). Fe^+ can then lose the second 4*s* electron 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 REVIEW

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. Among the main-group elements, what is the relationship between group number and the number of valence electrons among group members?

3. a. In general, how do the periodic properties of the *d*-block elements compare with those of the main-group elements?
b. Explain the comparisons made in (a).

CHAPTER 5 REVIEW

CHAPTER SUMMARY

5-1

- 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.

Vocabulary

actinide (126)
lanthanide (126)
periodic law (125)
periodic table (125)

5-2

- 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.

Vocabulary

alkali metals (132)
alkaline-earth metals (132)
halogens (137)
main-group elements (136)
transition elements (134)

5-3

- 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.

Vocabulary

anion (149)
atomic radius (140)
cation (149)
electron affinity (147)
electronegativity (151)
ion (143)
ionization (143)
ionization energy (143)
valence electrons (150)

REVIEWING CONCEPTS

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 (5-1)

2. State the periodic law. (5-1)

3. How is the periodic law demonstrated within the groups of the periodic table? (5-1)

4. a. How do the electron configurations within the same group of elements compare?
   b. Why are the noble gases relatively unreactive? (5-2)

5. What determines the length of each period in the periodic table? (5-2)

6. What is the relationship between the electron configuration of an element and the period in which that element appears in the periodic table? (5-2)

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. (5-2)

8. a. Which elements are designated as the alkali metals?
b. List four of their characteristic properties. (5-2)

9. a. Which elements are designated as the alkaline-earth metals?
b. How do their characteristic properties compare with those of the alkali metals? (5-2)

10. a. Write the usual 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? (5-2)

11. What name is sometimes used to refer to the entire set of *d*-block elements? (5-2)

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? (5-2)

13. a. Which elements are designated as the halogens?
b. List three of their characteristic properties. (5-2)

14. a. Which elements are metalloids?
b. Describe their characteristic properties. (5-2)

15. Which elements make up the *f* block in the periodic table? (5-2)

16. a. What are the main-group elements?
b. What trends can be observed across the various periods within the main-group elements? (5-2)

17. a. What is meant by atomic radius?
b. What trend is observed among the atomic radii of main-group elements across a period?
c. How can this trend be explained? (5-3)

18. a. What trend is observed among the atomic radii of main-group elements down a group?
b. How can this trend be explained? (5-3)

19. Define each of the following terms:
a. ion
b. ionization
c. first ionization energy
d. second ionization energy (5-3)

20. 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. (5-3)

21. a. What is electron affinity?
b. What signs are associated with electron affinity values, and what is the significance of each sign? (5-3)

22. 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? (5-3)

23. a. What are valence electrons?
b. Where are such electrons located? (5-3)

24. 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
b. Group 2
c. Group 13
d. Group 16
e. Group 17
f. Group 18 (5-3)

25. a. What is electronegativity?
b. Why is fluorine special in terms of electronegativity? (5-3)

26. Identify the most- and least-electronegative groups of elements in the periodic table. (5-3)

PROBLEMS

Electron Configuration and Periodic Properties

27. 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
b. O
c. Cu
d. Br
e. Sn

28. 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 5-1.)
a. $[Ne]3s^2 3p^4$
b. $[Kr]4d^{10}5s^2 5p^2$
c. $[Xe]4f^{14}5d^{10}6s^2 6p^5$

29. Based on the information given below, give the group, period, block, and identity of each element described. (Hint: See Sample Problem 5-2.)
a. $[He]2s^2$
b. $[Ne]3s^1$
c. $[Kr]5s^2$
d. $[Ar]4s^2$
e. $[Ar]3d^54s^1$

30. Without looking at the periodic table, write the expected outer electron configuration for each of the following elements. (Hint: See Sample Problem 5-3.)
a. Group 7, fourth period
b. Group 3, fifth period
c. Group 12, sixth period

31. 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 5-4.)
a. $[Ne]3s^23p^1$
b. $[Ar]3d^{10}4s^24p^6$
c. $[Kr]4d^{10}5s^1$
d. $[Xe]4f^15d^16s^2$

Atomic Radius, Ionization, Electron Affinity, and Electronegativity

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 5-5.)

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 5-7.)

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

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. $[Ar]3d^24s^2$
c. $[Kr]4d^{10}5s^1$
d. $[Xe]4f^{14}5d^96s^1$

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

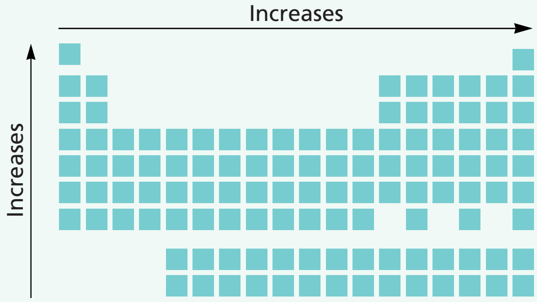
41. Which of the following does not have a noble-gas configuration: Na^+, Rb^+, O^{2-}, Br^-, Ca^+, Al^{3+}, S^{2-}?

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?

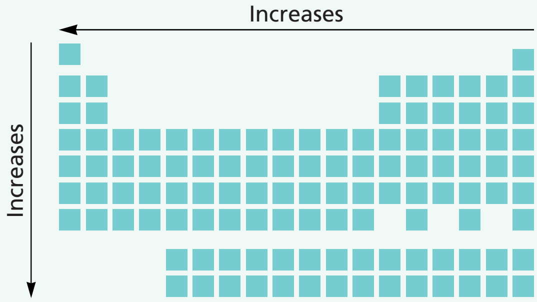
43. Write the noble-gas notation for the electron configuration of each of the following elements, and indicate the period and group in which each belongs:
a. Mg
b. P
c. Sc
d. Y

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

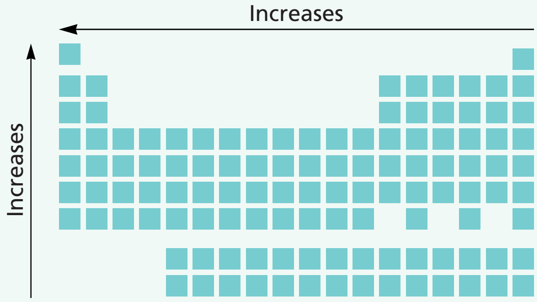
45. Identify which trends in the diagrams below describe atomic radius, ionization energy, electron affinity, and electronegativity.



a.



b.



c.

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

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

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

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

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 Using the data on the five new elements, create a periodic table based on their properties.

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 periodic trends, and describe the trends that you identify.

TECHNOLOGY & LEARNING

54. **Graphing Calculator** Graphing Atomic Radius vs. Atomic Number

The graphing calculator can run a program that graphs data such as atomic radius versus atomic number. Graphing the data within the different periods will allow you to discover trends. Begin by creating a table of data. Then use the program to plot the data.

Go to Appendix C. If you are using a TI 83 Plus, you can download the program and data sets and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. Remember that after creating the lists, you will need to name the program and check the display, as explained in Appendix C. You will then be ready to run the program. After you have graphed the data, answer these questions.

a. Would you expect any atomic number to have an atomic radius of 20? Why or why not?

b. A relationship is considered a function if it can pass a vertical line test. That is, if a vertical line can be drawn anywhere on the graph and only pass through one point, the relationship is a function. Do either of these sets of data represent a function? If so, which one(s)?

c. How would you describe the graphical relationship between the atomic numbers and atomic radii?

HANDBOOK SEARCH

55. Review the boiling point and melting point data in the tables of the *Elements Handbook* (beginning on page 730). Make a list of the elements that exist as liquids or gases at the boiling point of water, 100°C.

56. 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* on transition metals (pages 740–749) 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 & WRITING

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

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

ALTERNATIVE ASSESSMENT

59. Your teacher will give you an index card that identifies the electronegativity, ionization energy, and electron affinity of an element. Identify the element by analyzing these properties in relation to periodic trends.

60. 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.