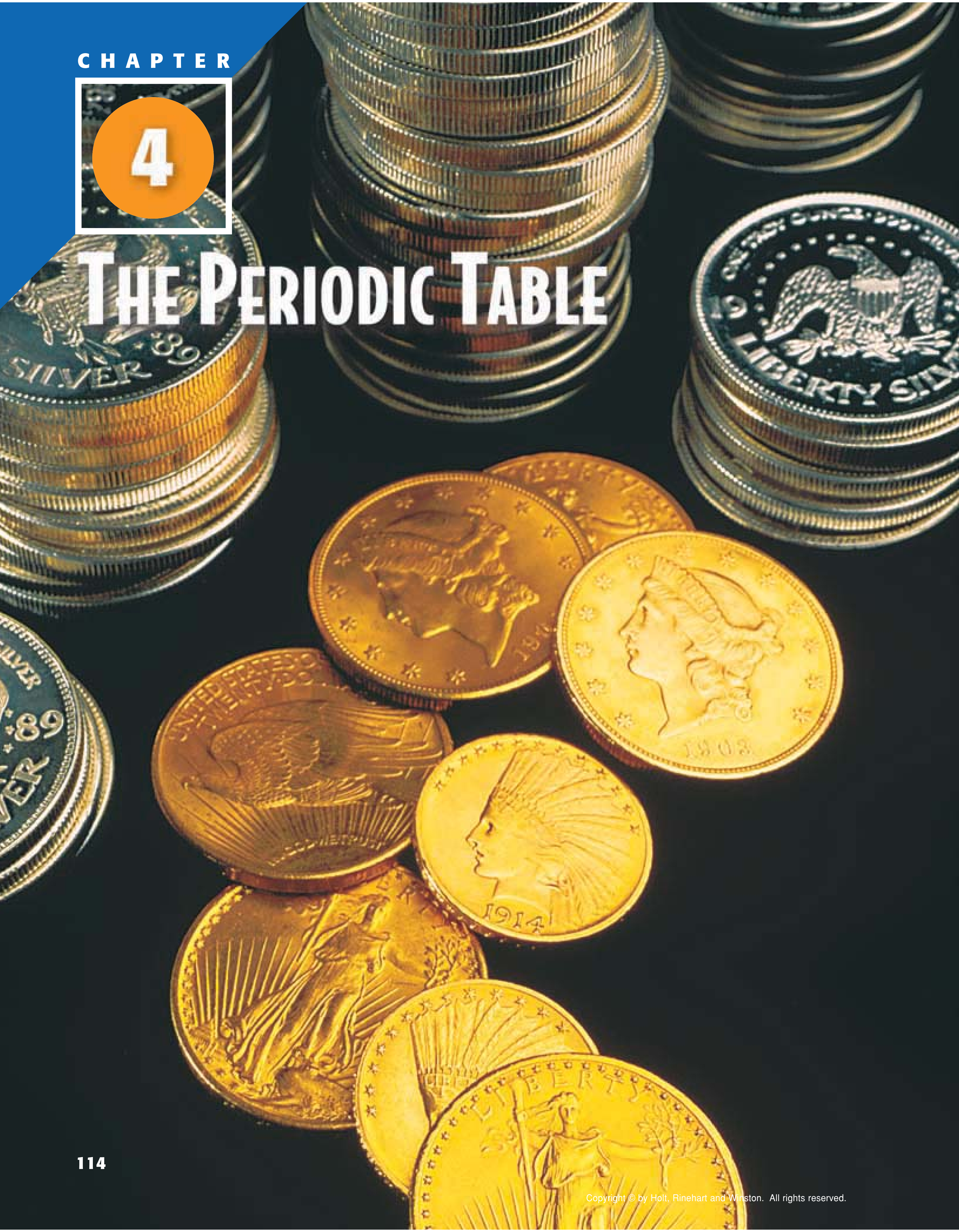
CHAPTER

4

THE PERIODIC TABLE



Copyright © by Holt, Rinehart and Winston. All rights reserved.

The United States established its first mint to make silver and gold coins in Philadelphia in 1792. Some of these old gold and silver coins have become quite valuable as collector's items. An 1804 silver dollar recently sold for more than $4 million. A silver dollar is actually 90% silver and 10% copper. Because the pure elements gold and silver are too soft to be used alone in coins, other metals are mixed with them to add strength and durability. These metals include platinum, copper, zinc, and nickel. Metals make up the majority of the elements in the periodic table.

START-UP ACTIVITY

SAFETY PRECAUTIONS

What Is a Periodic Table?

PROCEDURE

1. Sit in your assigned desk according to the seating chart your teacher provides.
2. On the blank chart your teacher gives you, jot down information about yourself—such as name, date of birth, hair color, and height—in the space that represents where you are seated.
3. Find out the same information from as many people sitting around you as possible, and write that information in the corresponding spaces on the seating chart.

ANALYSIS

1. Looking at the information you gathered, try to identify patterns that could explain the order of people in the seating chart. If you cannot yet identify a pattern, collect more information and look again for a pattern.
2. Test your pattern by gathering information from a person you did not talk to before.
3. If the new information does not fit in with your pattern, reevaluate your data to come up with a new hypothesis that explains the patterns in the seating chart.

Pre-Reading Questions

1. Define *element*.
2. What is the relationship between the number of protons and the number of electrons in a neutral atom?
3. As electrons fill orbitals, what patterns do you notice?

CONTENTS 4

internet connect
www.scilinks.org
Topic: The Periodic Table
SciLinks code: HW4094
SciLinks. Maintained by the National Science Teachers Association

Copyright © by Holt, Rinehart and Winston. All rights reserved.

SECTION

1 How Are Elements Organized?

KEY TERMS

- **periodic law**
- **valence electron**
- **group**
- **period**

OBJECTIVES

1. **Describe** the historical development of the periodic table.
2. **Describe** the organization of the modern periodic table according to the periodic law.

Patterns in Element Properties

Pure elements at room temperature and atmospheric pressure can be solids, liquids, or gases. Some elements are colorless. Others, like the ones shown in **Figure 1,** are colored. Despite the differences between elements, groups of elements share certain properties. For example, the elements lithium, sodium, potassium, rubidium, and cesium can combine with chlorine in a 1:1 ratio to form LiCl, NaCl, KCl, RbCl, and CsCl. All of these compounds are white solids that dissolve in water to form solutions that conduct electricity.

Similarly, the elements fluorine, chlorine, bromine, and iodine can combine with sodium in a 1:1 ratio to form NaF, NaCl, NaBr, and NaI. These compounds are also white solids that can dissolve in water to form solutions that conduct electricity. These examples show that even though each element is different, groups of them have much in common.

Topic Link

Refer to the chapter "The Science of Chemistry" for a definition and discussion of elements.

Figure 1
The elements chlorine, bromine, and iodine, pictured from left to right, look very different from each other. But each forms a similar-looking white solid when it reacts with sodium.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

John Newlands Noticed a Periodic Pattern

Elements vary widely in their properties, but in an orderly way. In 1865, the English chemist John Newlands arranged the known elements according to their properties and in order of increasing atomic mass. He placed the elements in a table.

As he studied his arrangement, Newlands noticed that all of the elements in a given row had similar chemical and physical properties. Because these properties seemed to repeat every eight elements, Newlands called this pattern the *law of octaves*.

This proposed law met with some skepticism when it was first presented, partly because chemists at the time did not know enough about atoms to be able to suggest a physical basis for any such law.

Dmitri Mendeleev Invented the First Periodic Table

In 1869, the Russian chemist Dmitri Mendeleev used Newlands's observation and other information to produce the first orderly arrangement, or periodic table, of all 63 elements known at the time. Mendeleev wrote the symbol for each element, along with the physical and chemical properties and the relative atomic mass of the element, on a card. Like Newlands, Mendeleev arranged the elements in order of increasing atomic mass. Mendeleev started a new row each time he noticed that the chemical properties of the elements repeated. He placed elements in the new row directly below elements of similar chemical properties in the preceding row. He arrived at the pattern shown in **Figure 2.**

Two interesting observations can be made about Mendeleev's table. First, Mendeleev's table contains gaps that elements with particular properties should fill. He predicted the properties of the missing elements.

ПЕРИОДИЧЕСКАЯ СИСТЕМА ЭЛЕМЕНТОВ

ГРУППЫ ЭЛЕМЕНТОВ

Ряды	I	II	III	IV	V	VI	VII	VIII		
I	H 1									
II	Li 7	Be 9,4	B 11	C 12	N 14	O 16	F 19			
III	Na 23	Mg 24	Al 27,4	Si 28	P 31	S 32	Cl 35,5			
IV	K 39	Ca 40	? 45	Ti 50	V 51	Cr 52	Mn 55	Fe 56	Co 59	Ni 59
V	Cu 63,4	Zn 65,2	? 68	? 70	As 75	Se 79,4	Br 80			
VI	Rb 85,4	Sr 87,6	Yt ? 88	Zr 90	Nb 94	Mo 96	? 100	Ru 104,4	Rh 104,4	Pd 106,6
VII	Ag 108	Cd 112	In 113	Sn 118	Sb 122	Te 128?	J 127			
VIII	Cs 133	Ba 137	Di ? 138	Ce ? 140						
IX										
X			Er ? 178	La ? 180	Ta 182	W 186		Pt 197,4	Ir 198	Os 199
XI	Au 197?	Hg 200	Tl 204	Pb 207	Bi 210					
XII				Th 231		U 240				

Figure 2
Mendeleev's table grouped elements with similar properties into vertical columns. For example, he placed the elements highlighted in red in the table—fluorine, chlorine, bromine, and iodine—into the column that he labeled "VII."

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Table 1 Predicted Versus Actual Properties for Three Elements

Properties	Ekaaluminum (gallium, discovered 1875)		Ekaboron (scandium, discovered 1877)		Ekasilicon (germanium, discovered 1886)	
	Predicted	**Observed**	**Predicted**	**Observed**	**Predicted**	**Observed**
Density	6.0 g/cm^3	5.96 g/cm^3	3.5 g/cm^3	3.5 g/cm^3	5.5 g/cm^3	5.47 g/cm^3
Melting point	low	30ºC	*	*	high	900ºC
Formula of oxide	Ea_2O_3	Ga_2O_3	Eb_2O_3	Sc_2O_3	EsO_2	GeO_2
Solubility of oxide	*	*	dissolves in acid	dissolves in acid	*	*
Density of oxide	*	*	*	*	4.7 g/cm^3	4.70 g/cm^3
Formula of chloride	*	*	*	*	$EsCl_4$	$GeCl_4$
Color of metal	*	*	*	*	dark gray	grayish white

He also gave these elements provisional names, such as "Ekaaluminum" (the prefix *eka-* means "one beyond") for the element that would come below aluminum. These elements were eventually discovered. As **Table 1** illustrates, their properties were close to Mendeleev's predictions. Although other chemists, such as Newlands, had created tables of the elements, Mendeleev was the first to use the table to predict the existence of undiscovered elements. Because Mendeleev's predictions proved true, most chemists accepted his periodic table of the elements.

Second, the elements do not always fit neatly in order of atomic mass. For example, Mendeleev had to switch the order of tellurium, Te, and iodine, I, to keep similar elements in the same column. At first, he thought that their atomic masses were wrong. However, careful research by others showed that they were correct. Mendeleev could not explain why his order was not always the same.

The Physical Basis of the Periodic Table

About 40 years after Mendeleev published his periodic table, an English chemist named Henry Moseley found a different physical basis for the arrangement of elements. When Moseley studied the lines in the X-ray spectra of 38 different elements, he found that the wavelengths of the lines in the spectra decreased in a regular manner as atomic mass increased. With further work, Moseley realized that the spectral lines correlated to atomic number, not to atomic mass.

When the elements were arranged by increasing atomic number, the discrepancies in Mendeleev's table disappeared. Moseley's work led to both the modern definition of atomic number, and showed that atomic number, not atomic mass, is the basis for the organization of the periodic table.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

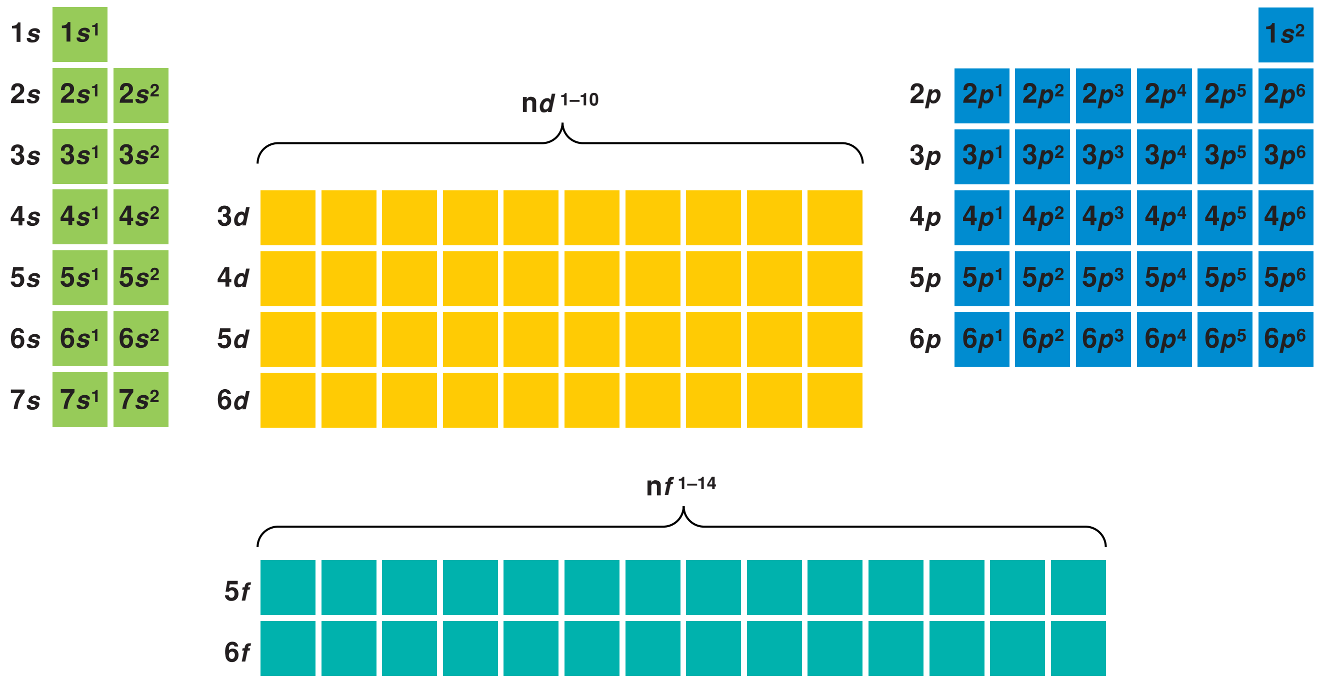


Figure 3
The shape of the periodic table is determined by how electrons fill orbitals. Only the *s* and *p* electrons are shown individually because unlike the *d* and *f* electrons, they fill orbitals sequentially.

The Periodic Law

According to Moseley, tellurium, whose atomic number is 52, belongs before iodine, whose atomic number is 53. Mendeleev had placed these elements in the same order based on their properties. Today, Mendeleev's principle of chemical periodicity is known as the **periodic law,** which states that when the elements are arranged according to their atomic numbers, elements with similar properties appear at regular intervals.

periodic law
the law that states that the repeating physical and chemical properties of elements change periodically with their atomic number

Organization of the Periodic Table

To understand why elements with similar properties appear at regular intervals in the periodic table, you need to examine the electron configurations of the elements. As shown in **Figure 3,** elements in each column of the table have the same number of electrons in their outer energy level. These electrons are called **valence electrons.** It is the valence electrons of an atom that participate in chemical reactions with other atoms, so elements with the same number of valence electrons tend to react in similar ways. Because *s* and *p* electrons fill sequentially, the number of valence electrons in *s*- and *p*-block elements are predictable. For example, atoms of elements in the column on the far left have one valence electron. Atoms of elements in the column on the far right have eight valence electrons (except for helium, which has two). A vertical column on the periodic table is called a **group.** A complete version of the modern periodic table is shown in **Figure 4** on the next two pages.

valence electron
an electron that is found in the outermost shell of an atom and that determines the atom's chemical properties

group
a vertical column of elements in the periodic table; elements in a group share chemical properties

Topic Link
Refer to the chapter "Atoms and Moles" for a discussion of electron configuration.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Periodic Table of the Elements

Key:

6	Atomic number
C	Symbol
Carbon	Name
12.0107	Average atomic mass
$[He]2s^2 2p^2$	Electron configuration

<table>
<tr><th>Period</th><th>Group 1</th><th>Group 2</th><th>Group 3</th><th>Group 4</th><th>Group 5</th><th>Group 6</th><th>Group 7</th><th>Group 8</th><th>Group 9</th></tr>
<tr><td>1</td><td>1
H
Hydrogen
1.007 94
$1s^1$</td><td></td><td></td><td></td><td></td><td></td><td></td><td></td><td></td></tr>
<tr><td>2</td><td>3
Li
Lithium
6.941
$[He]2s^1$</td><td>4
Be
Beryllium
9.012 182
$[He]2s^2$</td><td></td><td></td><td></td><td></td><td></td><td></td><td></td></tr>
<tr><td>3</td><td>11
Na
Sodium
22.989 770
$[Ne]3s^1$</td><td>12
Mg
Magnesium
24.3050
$[Ne]3s^2$</td><td></td><td></td><td></td><td></td><td></td><td></td><td></td></tr>
<tr><td>4</td><td>19
K
Potassium
39.0983
$[Ar]4s^1$</td><td>20
Ca
Calcium
40.078
$[Ar]4s^2$</td><td>21
Sc
Scandium
44.955 910
$[Ar]3d^1 4s^2$</td><td>22
Ti
Titanium
47.867
$[Ar]3d^2 4s^2$</td><td>23
V
Vanadium
50.9415
$[Ar]3d^3 4s^2$</td><td>24
Cr
Chromium
51.9961
$[Ar]3d^5 4s^1$</td><td>25
Mn
Manganese
54.938 049
$[Ar]3d^5 4s^2$</td><td>26
Fe
Iron
55.845
$[Ar]3d^6 4s^2$</td><td>27
Co
Cobalt
58.933 200
$[Ar]3d^7 4s^2$</td></tr>
<tr><td>5</td><td>37
Rb
Rubidium
85.4678
$[Kr]5s^1$</td><td>38
Sr
Strontium
87.62
$[Kr]5s^2$</td><td>39
Y
Yttrium
88.905 85
$[Kr]4d^1 5s^2$</td><td>40
Zr
Zirconium
91.224
$[Kr]4d^2 5s^2$</td><td>41
Nb
Niobium
92.906 38
$[Kr]4d^4 5s^1$</td><td>42
Mo
Molybdenum
95.94
$[Kr]4d^5 5s^1$</td><td>43
Tc
Technetium
(98)
$[Kr]4d^6 5s^1$</td><td>44
Ru
Ruthenium
101.07
$[Kr]4d^7 5s^1$</td><td>45
Rh
Rhodium
102.905 50
$[Kr]4d^8 5s^1$</td></tr>
<tr><td>6</td><td>55
Cs
Cesium
132.905 45
$[Xe]6s^1$</td><td>56
Ba
Barium
137.327
$[Xe]6s^2$</td><td>57
La
Lanthanum
138.9055
$[Xe]5d^1 6s^2$</td><td>72
Hf
Hafnium
178.49
$[Xe]4f^{14} 5d^2 6s^2$</td><td>73
Ta
Tantalum
180.9479
$[Xe]4f^{14} 5d^3 6s^2$</td><td>74
W
Tungsten
183.84
$[Xe]4f^{14} 5d^4 6s^2$</td><td>75
Re
Rhenium
186.207
$[Xe]4f^{14} 5d^5 6s^2$</td><td>76
Os
Osmium
190.23
$[Xe]4f^{14} 5d^6 6s^2$</td><td>77
Ir
Iridium
192.217
$[Xe]4f^{14} 5d^7 6s^2$</td></tr>
<tr><td>7</td><td>87
Fr
Francium
(223)
$[Rn]7s^1$</td><td>88
Ra
Radium
(226)
$[Rn]7s^2$</td><td>89
Ac
Actinium
(227)
$[Rn]6d^1 7s^2$</td><td>104
Rf
Rutherfordium
(261)
$[Rn]5f^{14} 6d^2 7s^2$</td><td>105
Db
Dubnium
(262)
$[Rn]5f^{14} 6d^3 7s^2$</td><td>106
Sg
Seaborgium
(263)
$[Rn]5f^{14} 6d^4 7s^2$</td><td>107
Bh
Bohrium
(264)
$[Rn]5f^{14} 6d^5 7s^2$</td><td>108
Hs
Hassium
(265)
$[Rn]5f^{14} 6d^6 7s^2$</td><td>109
Mt
Meitnerium
(268)†
$[Rn]5f^{14} 6d^7 7s^2$</td></tr>
</table>

<table>
<tr><td>58
Ce
Cerium
140.116
$[Xe]4f^1 5d^1 6s^2$</td><td>59
Pr
Praseodymium
140.907 65
$[Xe]4f^3 6s^2$</td><td>60
Nd
Neodymium
144.24
$[Xe]4f^4 6s^2$</td><td>61
Pm
Promethium
(145)
$[Xe]4f^5 6s^2$</td><td>62
Sm
Samarium
150.36
$[Xe]4f^6 6s^2$</td></tr>
<tr><td>90
Th
Thorium
232.0381
$[Rn]6d^2 7s^2$</td><td>91
Pa
Protactinium
231.035 88
$[Rn]5f^2 6d^1 7s^2$</td><td>92
U
Uranium
238.0289
$[Rn]5f^3 6d^1 7s^2$</td><td>93
Np
Neptunium
(237)
$[Rn]5f^4 6d^1 7s^2$</td><td>94
Pu
Plutonium
(244)
$[Rn]5f^6 7s^2$</td></tr>
</table>

† Estimated from currently available IUPAC data.

* The systematic names and symbols for elements greater than 109 will be used until the approval of trivial names by IUPAC.

internet connect

go.hrw.com

Topic: Periodic Table
Go To: go.hrw.com
Keyword: HN0 PERIODIC

Visit the HRW Web site for updates on the periodic table.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Figure 4

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

Nonmetals
- Hydrogen
- Semiconductors (also known as *metalloids*)
- Other nonmetals
- Halogens
- Noble gases

Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
								2 **He** Helium 4.002 602 $1s^2$
			5 **B** Boron 10.811 [He]$2s^22p^1$	6 **C** Carbon 12.0107 [He]$2s^22p^2$	7 **N** Nitrogen 14.006 74 [He]$2s^22p^3$	8 **O** Oxygen 15.9994 [He]$2s^22p^4$	9 **F** Fluorine 18.998 4032 [He]$2s^22p^5$	10 **Ne** Neon 20.1797 [He]$2s^22p^6$
			13 **Al** Aluminum 26.981 538 [Ne]$3s^23p^1$	14 **Si** Silicon 28.0855 [Ne]$3s^23p^2$	15 **P** Phosphorus 30.973 761 [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]$3s^23p^6$
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 74.921 60 [Ar]$3d^{10}4s^24p^3$	34 **Se** Selenium 78.96 [Ar]$3d^{10}4s^24p^4$	35 **Br** Bromine 79.904 [Ar]$3d^{10}4s^24p^5$	36 **Kr** Krypton 83.80 [Ar]$3d^{10}4s^24p^6$
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 121.760 [Kr]$4d^{10}5s^25p^3$	52 **Te** Tellurium 127.60 [Kr]$4d^{10}5s^25p^4$	53 **I** Iodine 126.904 47 [Kr]$4d^{10}5s^25p^5$	54 **Xe** Xenon 131.29 [Kr]$4d^{10}5s^25p^6$
78 **Pt** Platinum 195.078 [Xe]$4f^{14}5d^96s^1$	79 **Au** Gold 196.966 55 [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 208.980 38 [Xe]$4f^{14}5d^{10}6s^26p^3$	84 **Po** Polonium (209) [Xe]$4f^{14}5d^{10}6s^26p^4$	85 **At** Astatine (210) [Xe]$4f^{14}5d^{10}6s^26p^5$	86 **Rn** Radon (222) [Xe]$4f^{14}5d^{10}6s^26p^6$
110 **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]$5f^{14}6d^{10}7s^2$		114 **Uuq*** Ununquadium (285)† [Rn]$5f^{14}6d^{10}7s^27p^2$				

A team at Lawrence Berkeley National Laboratories reported the discovery of elements 116 and 118 in June 1999. The same team retracted the discovery in July 2001. The discovery of element 114 has been reported but not confirmed.

63 **Eu** Europium 151.964 [Xe]$4f^76s^2$	64 **Gd** Gadolinium 157.25 [Xe]$4f^75d^16s^2$	65 **Tb** Terbium 158.925 34 [Xe]$4f^96s^2$	66 **Dy** Dysprosium 162.50 [Xe]$4f^{10}6s^2$	67 **Ho** Holmium 164.930 32 [Xe]$4f^{11}6s^2$	68 **Er** Erbium 167.26 [Xe]$4f^{12}6s^2$	69 **Tm** Thulium 168.934 21 [Xe]$4f^{13}6s^2$	70 **Yb** Ytterbium 173.04 [Xe]$4f^{14}6s^2$	71 **Lu** Lutetium 174.967 [Xe]$4f^{14}5d^16s^2$
95 **Am** Americium (243) [Rn]$5f^77s^2$	96 **Cm** Curium (247) [Rn]$5f^76d^17s^2$	97 **Bk** Berkelium (247) [Rn]$5f^97s^2$	98 **Cf** Californium (251) [Rn]$5f^{10}7s^2$	99 **Es** Einsteinium (252) [Rn]$5f^{11}7s^2$	100 **Fm** Fermium (257) [Rn]$5f^{12}7s^2$	101 **Md** Mendelevium (258) [Rn]$5f^{13}7s^2$	102 **No** Nobelium (259) [Rn]$5f^{14}7s^2$	103 **Lr** Lawrencium (262) [Rn]$5f^{14}6d^17s^2$

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.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

period
a horizontal row of elements in the periodic table

A horizontal row on the periodic table is called a **period.** Elements in the same period have the same number of occupied energy levels. For example, all elements in Period 2 have atoms whose electrons occupy two principal energy levels, including the 2*s* and 2*p* orbitals. Elements in Period 5 have outer electrons that fill the 5*s*, 5*d*, and 5*p* orbitals.

This correlation between period number and the number of occupied energy levels holds for all seven periods. So a periodic table is not needed to tell to which period an element belongs. All you need to know is the element's electron configuration. For example, germanium has the electron configuration $[Ar]3d^{10}4s^{2}4p^{2}$. The largest principal quantum number it has is 4, which means germanium has four occupied energy levels. This places it in Period 4.

The periodic table provides information about each element, as shown in the key for **Figure 4.** This periodic table lists the atomic number, symbol, name, average atomic mass, and electron configuration in shorthand form for each element.

In addition, some of the categories of elements are designated through a color code. You may notice that many of the color-coded categories shown in **Figure 4** are associated with a certain group or groups. This shows how categories of elements are grouped by common properties which result from their common number of valence electrons. The next section discusses the different kinds of elements on the periodic table and explains how their electron configurations give them their characteristic properties.

1 Section Review

UNDERSTANDING KEY IDEAS

1. How can one show that elements that have different appearances have similar chemical properties?
2. Why was the pattern that Newlands developed called the *law of octaves*?
3. What led Mendeleev to predict that some elements had not yet been discovered?
4. What contribution did Moseley make to the development of the modern periodic table?
5. State the periodic law.
6. What do elements in the same period have in common?
7. What do elements in the same group have in common?

CRITICAL THINKING

8. Why can Period 1 contain a maximum of two elements?
9. In which period and group is the element whose electron configuration is $[Kr]5s^{1}$?
10. Write the outer electron configuration for the Group 2 element in Period 6.
11. What determines the number of elements found in each period in the periodic table?
12. Are elements with similar chemical properties more likely to be found in the same period or in the same group? Explain your answer.
13. How many valence electrons does phosphorus have?
14. What would you expect the electron configuration of element 113 to be?

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Consumer Focus

Essential Elements

Four elements—hydrogen, oxygen, carbon, and nitrogen—account for more than 99% of all atoms in the human body.

Good Health Is Elementary

Hydrogen, oxygen, carbon, and nitrogen are the major components of the many different molecules that our bodies need. Likewise, these elements are the major elements in the molecules of the food that we eat.

Another seven elements, listed in **Table 2,** are used by our bodies in substantial quantities, more than 0.1 g per day. These elements are known as *macronutrients* or, more commonly, as *minerals*.

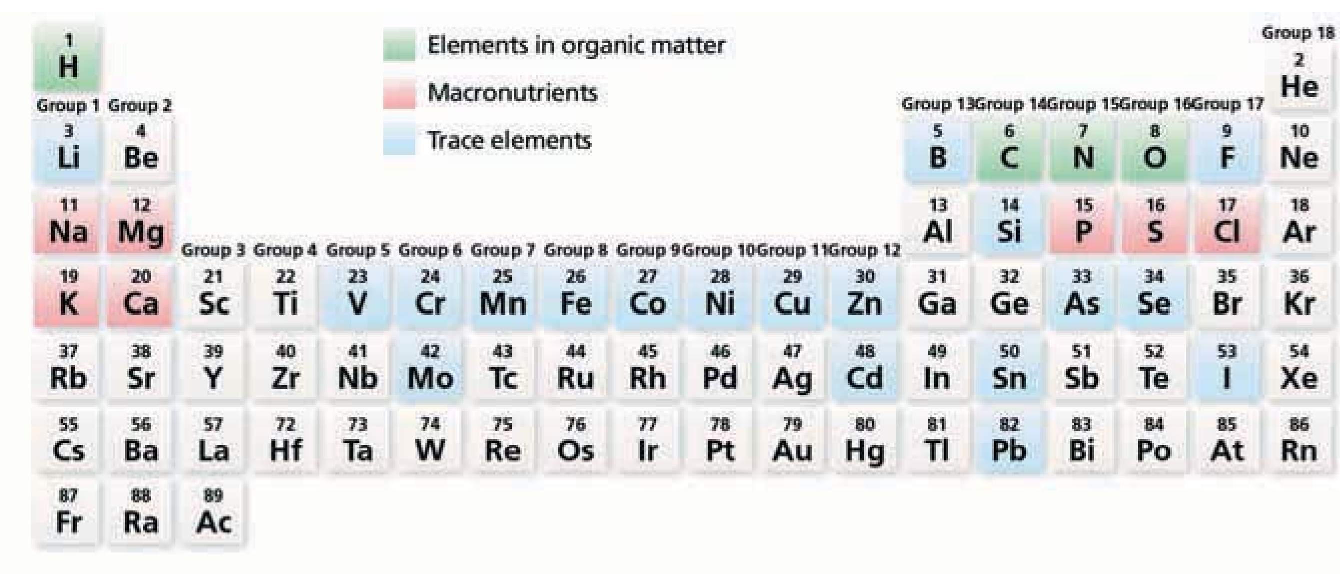
Some elements, known as *trace elements* or *micronutrients,* are necessary for healthy human life, but only in very small amounts. In many cases, humans need less than 15 nanograms, or 15×10^{-9} g, of a particular trace element per day to maintain good health. This means that you need less than 0.0004 g of such trace elements during your entire lifetime!

Table 2 Macronutrients

Element	Symbol	Role in human body chemistry
Calcium	Ca	bones, teeth; essential for blood clotting and muscle contraction
Phosphorus	P	bones, teeth; component of nucleic acids, including DNA
Potassium	K	present as K^+ in all body fluids; essential for nerve action
Sulfur	S	component of many proteins; essential for blood clotting
Chlorine	Cl	present as Cl^- in all body fluids; important to maintaining salt balance
Sodium	Na	present as Na^+ in all body fluids; essential for nerve and muscle action
Magnesium	Mg	in bones and teeth; essential for muscle action

Questions

1. What do the two macronutrients involved in nerve action have in common?
2. You may recognize elements such as arsenic and lead as toxic. Explain how these elements can be nutrients even though they are toxic.



Copyright © by Holt, Rinehart and Winston. All rights reserved.

SECTION

2 Tour of the Periodic Table

KEY TERMS

- main-group element
- alkali metal
- alkaline-earth metal
- halogen
- noble gas
- transition metal
- lanthanide
- actinide
- alloy

OBJECTIVES

1. **Locate** the different families of main-group elements on the periodic table, describe their characteristic properties, and relate their properties to their electron configurations.
2. **Locate** metals on the periodic table, describe their characteristic properties, and relate their properties to their electron configurations.

The Main-Group Elements

Elements in groups 1, 2, and 13–18 are known as the **main-group elements.** As shown in **Figure 5,** main-group elements are in the *s*- and *p*-blocks of the periodic table. The electron configurations of the elements in each main group are regular and consistent: the elements in each group have the same number of valence electrons. For example, Group 2 elements have two valence electrons. The configuration of their valence electrons can be written as ns^2, where n is the period number. Group 16 elements have a total of six valence electrons in their outermost *s* and *p* orbitals. Their valence electron configuration can be written as ns^2np^4.

main-group elements
an element in the *s*-block or *p*-block of the periodic table

The main-group elements are sometimes called the *representative elements* because they have a wide range of properties. At room temperature and atmospheric pressure, many are solids, while others are liquids or gases. About half of the main-group elements are metals. Many are extremely reactive, while several are nonreactive. The main-group elements silicon and oxygen account for four of every five atoms found on or near Earth's surface.

Four groups within the main-group elements have special names. These groups are the *alkali metals* (Group 1), the *alkaline-earth metals* (Group 2), the *halogens* (Group 17), and the *noble gases* (Group 18).

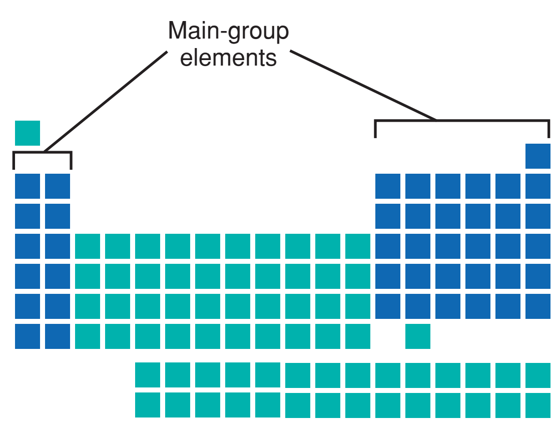


Figure 5
Main-group elements have diverse properties and uses. They are highlighted in the groups on the left and right sides of the periodic table.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

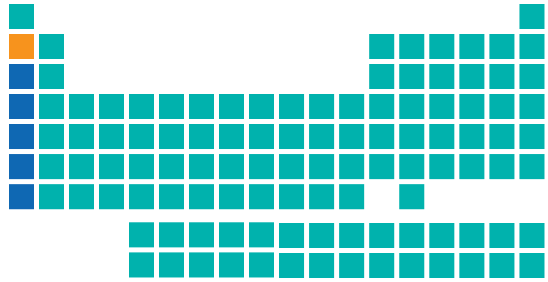


Figure 6
The alkali metals make up the first group of the periodic table. Lithium, pictured here, is an example of an alkali metal.

The Alkali Metals Make Up Group 1

Elements in Group 1, which is highlighted in **Figure 6,** are called **alkali metals.** Alkali metals are so named because they are metals that react with water to make alkaline solutions. For example, potassium reacts vigorously with cold water to form hydrogen gas and the compound potassium hydroxide, KOH. Because the alkali metals have a single valence electron, they are very reactive. In losing its one valence electron, potassium achieves a stable electron configuration.

Alkali metals are usually stored in oil to keep them from reacting with the oxygen and water in the air. Because of their high reactivity, alkali metals are never found in nature as pure elements but are found combined with other elements as compounds. For instance, the salt sodium chloride, NaCl, is abundant in sea water.

Some of the physical properties of the alkali metals are listed in **Table 3.** All these elements are so soft that they can be easily cut with a knife. The freshly cut surface of an alkali metal is shiny, but it dulls quickly as the metal reacts with oxygen and water in the air. Like other metals, the alkali metals are good conductors of electricity.

alkali metal

one of the elements of Group 1 of the periodic table (lithium, sodium, potassium, rubidium, cesium, and francium)

internet connect
www.scilinks.org
Topic: Alkali Metals
SciLinks code: HW4007
SCILINKS Maintained by the National Science Teachers Association

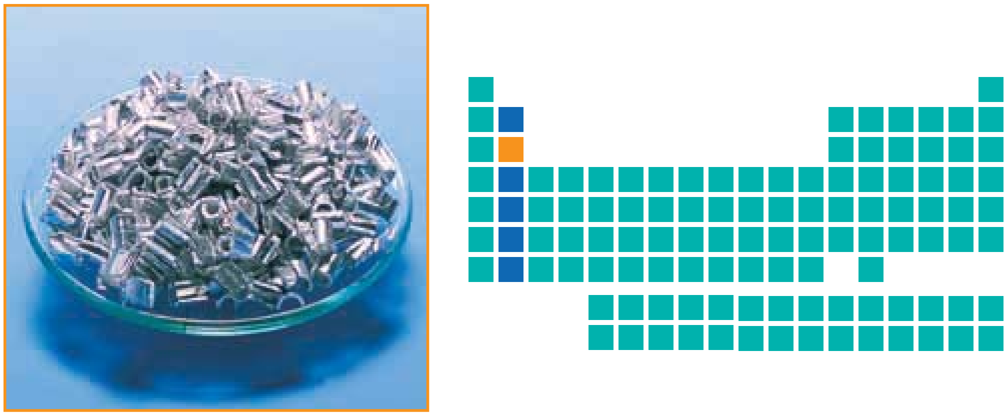
Table 3 Physical Properties of Alkali Metals

Element	Flame test	Hardness (Mohs' scale)	Melting Point (°C)	Boiling Point (°C)	Density (g/cm^3)	Atomic radius (pm)
Lithium	red	0.6	180.5	1342	0.53	134
Sodium	yellow	0.4	97.7	883	0.97	154
Potassium	violet	0.5	63.3	759	0.86	196
Rubidium	yellowish violet	0.3	39.3	688	1.53	(216)
Cesium	reddish violet	0.2	28.4	671	1.87	(233)

Refer to Appendix A for more information about the properties of elements, including alkali metals.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Figure 7
The alkaline-earth metals make up the second group of the periodic table. Magnesium, pictured here, is an example of an alkaline-earth metal.



The Alkaline-Earth Metals Make Up Group 2

alkaline-earth metal

one of the elements of Group 2 of the periodic table (beryllium, magnesium, calcium, strontium, barium, and radium)

internet connect
www.scilinks.org
Topic: Alkaline-Earth Metals
SciLinks code: HW4008
SciLINKS Maintained by the National Science Teachers Association

Group 2 elements, which are highlighted in **Figure 7,** are called **alkaline-earth metals.** Like the alkali metals, the alkaline-earth metals are highly reactive, so they are usually found as compounds rather than as pure elements. For example, if the surface of an object made from magnesium is exposed to the air, the magnesium will react with the oxygen in the air to form the compound magnesium oxide, MgO, which eventually coats the surface of the magnesium metal.

The alkaline-earth metals are slightly less reactive than the alkali metals. The alkaline-earth metals have two valence electrons and must lose both their valence electrons to get to a stable electron configuration. It takes more energy to lose two electrons than it takes to lose just the one electron that the alkali metals must give up to become stable. Although the alkaline-earth metals are not as reactive, they are harder and have higher melting points than the alkali metals.

Beryllium is found in emeralds, which are a variety of the mineral beryl. Perhaps the best-known alkaline-earth metal is calcium, an important mineral nutrient found in the human body. Calcium is essential for muscle contraction. Bones are made up of calcium phosphate. Calcium compounds, such as limestone and marble, are common in the Earth's crust. Marble is made almost entirely of pure calcium carbonate. Because marble is hard and durable, it is used in sculptures.

The Halogens, Group 17, Are Highly Reactive

halogen

one of the elements of Group 17 of the periodic table (fluorine, chlorine, bromine, iodine, and astatine); halogens combine with most metals to form salts

internet connect
www.scilinks.org
Topic: Halogens
SciLinks code: HW4065
SciLINKS Maintained by the National Science Teachers Association

Elements in Group 17 of the periodic table, which are highlighted in **Figure 8** on the next page, are called the **halogens.** The halogens are the most reactive group of nonmetal elements because of their electron configuration. Halogens have seven valence electrons—just one short of a stable configuration. When halogens react, they often gain the one electron needed to have eight valence electrons, a filled outer energy level. Because the alkali metals have one valence electron, they are ideally suited to react with the halogens. For example, the alkali metal sodium easily loses its one valence electron to the halogen chlorine to form the compound sodium chloride, NaCl, which is table salt. The halogens react with most metals to produce salts. In fact, the word *halogen* comes from Greek and means "salt maker."

Copyright © by Holt, Rinehart and Winston. All rights reserved.

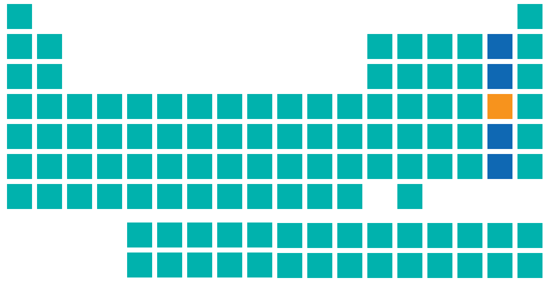


Figure 8
The halogens make up Group 17 of the periodic table. Bromine, one of only two elements that are liquids at room temperature, is an example of a halogen.

The halogens have a wide range of physical properties. Fluorine and chlorine are gases at room temperature, but bromine, depicted in **Figure 8,** is a liquid, and iodine and astatine are solids. The halogens are found in sea water and in compounds found in the rocks of Earth's crust. Astatine is one of the rarest of the naturally occurring elements.

internet connect
www.scilinks.org
Topic: Noble Gases
SciLinks code: HW4083
SCILINKS Maintained by the National Science Teachers Association

The Noble Gases, Group 18, Are Unreactive

Group 18 elements, which are highlighted in **Figure 9,** are called the **noble gases.** The noble gas atoms have a full set of electrons in their outermost energy level. Except for helium ($1s^2$), noble gases have an outer-shell configuration of ns^2np^6. From the low chemical reactivity of these elements, chemists infer that this full shell of electrons makes these elements very stable. The low reactivity of noble gases leads to some special uses. Helium, a noble gas, is used to fill blimps because it has a low density and is not flammable.

The noble gases were once called *inert gases* because they were thought to be completely unreactive. But in 1962, chemists were able to get xenon to react, making the compound $XePtF_6$. In 1979, chemists were able to form the first xenon-carbon bonds.

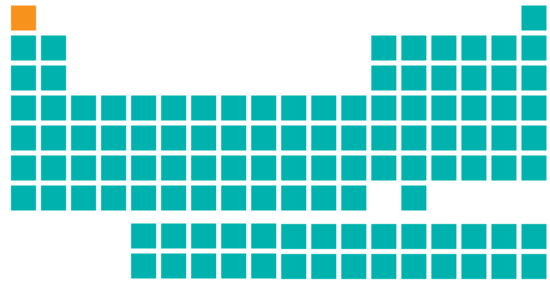
noble gas

an unreactive element of Group 18 of the periodic table (helium, neon, argon, krypton, xenon, or radon) that has eight electrons in its outer level (except for helium, which has two electrons)

Figure 9
The noble gases make up Group 18 of the periodic table. Helium, whose low density makes it ideal for use in blimps, is an example of a noble gas.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Figure 10
Hydrogen sits apart from all other elements in the periodic table. Hydrogen is extremely flammable and is used as fuel for space shuttle launches.



Hydrogen Is in a Class by Itself

Hydrogen is the most common element in the universe. It is estimated that about three out of every four atoms in the universe are hydrogen. Because it consists of just one proton and one electron, hydrogen behaves unlike any other element. As shown in **Figure 10,** hydrogen is in a class by itself in the periodic table.

With its one electron, hydrogen can react with many other elements, including oxygen. Hydrogen gas and oxygen gas react explosively to form water. Hydrogen is a component of the organic molecules found in all living things. The main industrial use of hydrogen is in the production of ammonia, NH_3. Large quantities of ammonia are used to make fertilizers.

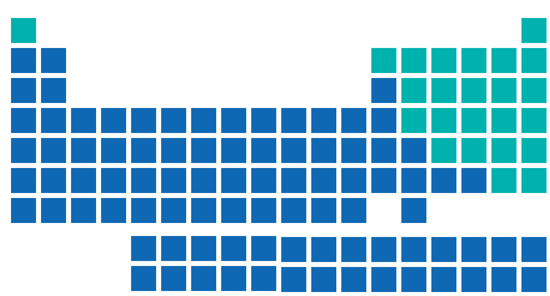
Most Elements Are Metals

internet connect
www.scilinks.org
Topic: Metals
SciLinks code: HW4079
SCILINKS. Maintained by the National Science Teachers Association

Figure 11 shows that the majority of elements, including many main-group ones, are metals. But what exactly is a metal? You can often recognize a metal by its shiny appearance, but some nonmetal elements, plastics, and minerals are also shiny. For example, a diamond usually has a brilliant luster. However, diamond is a mineral made entirely of the nonmetal element carbon.

Conversely, some metals appear black and dull. An example is iron, which is a very strong and durable metal. Iron is a member of Group 8 and is therefore not a main-group element. Iron belongs to a class of elements called *transition metals*. However, wherever metals are found on the periodic table, they tend to share certain properties.

Figure 11
The regions highlighted in blue indicate the elements that are metals.



Copyright © by Holt, Rinehart and Winston. All rights reserved.

Metals Share Many Properties

All metals are excellent conductors of electricity. Electrical conductivity is the one property that distinguishes metals from the nonmetal elements. Even the least conductive metal conducts electricity 100 000 times better than the best nonmetallic conductor does.

Metals also exhibit other properties, some of which can also be found in certain nonmetal elements. For example, metals are excellent conductors of heat. Some metals, such as manganese and bismuth, are very brittle. Other metals, such as gold and copper, are ductile and malleable. *Ductile* means that the metal can be squeezed out into a wire. *Malleable* means that the metal can be hammered or rolled into sheets. Gold, for example, can be hammered into very thin sheets, called "gold leaf," and applied to objects for decoration.

Transition Metals Occupy the Center of the Periodic Table

The **transition metals** constitute Groups 3 through 12 and are sometimes called the *d*-block elements because of their position in the periodic table, shown in **Figure 12.** Unlike the main-group elements, the transition metals in each group do not have identical outer electron configurations. For example, nickel, Ni, palladium, Pd, and platinum, Pt, are Group 10 metals. However, Ni has the electron configuration $[Ar]3d^84s^2$, Pd has the configuration $[Kr]4d^{10}$, and Pt has the configuration $[Xe]4f^{14}5d^96s^1$. Notice, however, that in each case the sum of the outer *d* and *s* electrons is equal to the group number, 10.

A transition metal may lose different numbers of valence electrons depending on the element with which it reacts. Generally, the transition metals are less reactive than the alkali metals and the alkaline-earth metals are. In fact, some transition metals are so unreactive that they seldom form compounds with other elements. Palladium, platinum, and gold are among the least reactive of all the elements other than the noble gases. These three transition metals can be found in nature as pure elements.

Transition metals, like other metals, are good conductors of heat and electricity. They are also ductile and malleable, as shown in **Figure 12.**

transition metal

one of the metals that can use the inner shell before using the outer shell to bond

internet connect

www.scilinks.org
Topic: Transition Metals
SciLinks code: HW4168
SCILINKS Maintained by the National Science Teachers Association

Figure 12
Copper, a transition metal, is used in wiring because it conducts electricity well. Because of its ductility and malleability, it can be formed into wires that bend easily.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

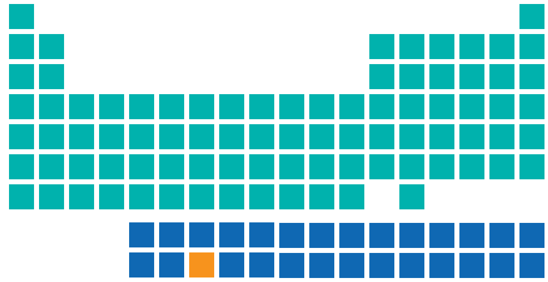


Figure 13
The lanthanides and actinides are placed at the bottom of the periodic table. Uranium, an actinide, is used in nuclear reactors. The collection of uranium-238 kernels is shown here.

Lanthanides and Actinides Fill *f*-orbitals

Part of the last two periods of transition metals are placed toward the bottom of the periodic table to keep the table conveniently narrow, as shown in **Figure 13.** The elements in the first of these rows are called the **lanthanides** because their atomic numbers follow the element lanthanum. Likewise, elements in the row below the lanthanides are called **actinides** because they follow actinium. As one moves left to right along these rows, electrons are added to the 4*f* orbitals in the lanthanides and to the 5*f* orbitals in the actinides. For this reason, the lanthanides and actinides are sometimes called the *f*-block of the periodic table.

lanthanide
a member of the rare-earth series of elements, whose atomic numbers range from 58 (cerium) to 71 (lutetium)

actinide
any of the elements of the actinide series, which have atomic numbers from 89 (actinium, Ac) through 103 (lawrencium, Lr)

The lanthanides are shiny metals similar in reactivity to the alkaline-earth metals. Some lanthanides have practical uses. Compounds of some lanthanide metals are used to produce color television screens.

The actinides are unique in that their nuclear structures are more important than their electron configurations. Because the nuclei of actinides are unstable and spontaneously break apart, all actinides are radioactive. The best-known actinide is uranium.

Other Properties of Metals

The melting points of metals vary widely. Tungsten has the highest melting point, 4322°C, of any element. In contrast, mercury melts at –39°C, so it is a liquid at room temperature. This low melting point, along with its high density, makes mercury useful for barometers.

alloy
a solid or liquid mixture of two or more metals

Metals can be mixed with one or more other elements, usually other metals, to make an **alloy.** The mixture of elements in an alloy gives the alloy properties that are different from the properties of the individual elements. Often these properties eliminate some disadvantages of the pure metal. A common alloy is brass, a mixture of copper and zinc, which is harder than copper and more resistant to corrosion. Brass has a wide range of uses, from inexpensive jewelry to plumbing hardware. Another alloy made from copper is sterling silver. A small amount of copper is mixed with silver to produce sterling silver, which is used for both jewelry and flatware.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Figure 14
Steel is an alloy made of iron and carbon. When heated, steel can be worked into many useful shapes.

Many iron alloys, such as the steel shown in **Figure 14,** are harder, stronger, and more resistant to corrosion than pure iron. Steel contains between 0.2% and 1.5% carbon atoms and often has tiny amounts of other elements such as manganese and nickel. Stainless steel also incorporates chromium. Because of its hardness and resistance to corrosion, stainless steel is an ideal alloy for making knives and other tools.

2 Section Review

UNDERSTANDING KEY IDEAS

1. Which group of elements is the most unreactive? Why?
2. Why do groups among the main-group elements display similar chemical behavior?
3. What properties do the halogens have in common?
4. Why is hydrogen set apart by itself?
5. How do the valence electron configurations of the alkali metals compare with each other?
6. Why are the alkaline-earth metals less reactive than the alkali metals?
7. In which groups of the periodic table do the transition metals belong?
8. Why are the nuclear structures of the actinides more important than the electron configurations of the actinides?
9. What is an alloy?

CRITICAL THINKING

10. Noble gases used to be called *inert gases.* What discovery changed that term, and why?
11. If you find an element in nature in its pure elemental state, what can you infer about the element's chemical reactivity?
12. Explain why the transition metals are sometimes referred to as the *d*-block elements.
13. Can an element that conducts heat, is malleable, and has a high melting point be classified as a metal? Explain your reasoning.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

SECTION 3

Trends in the Periodic Table

KEY TERMS

- **ionization energy**
- **electron shielding**
- **bond radius**
- **electronegativity**

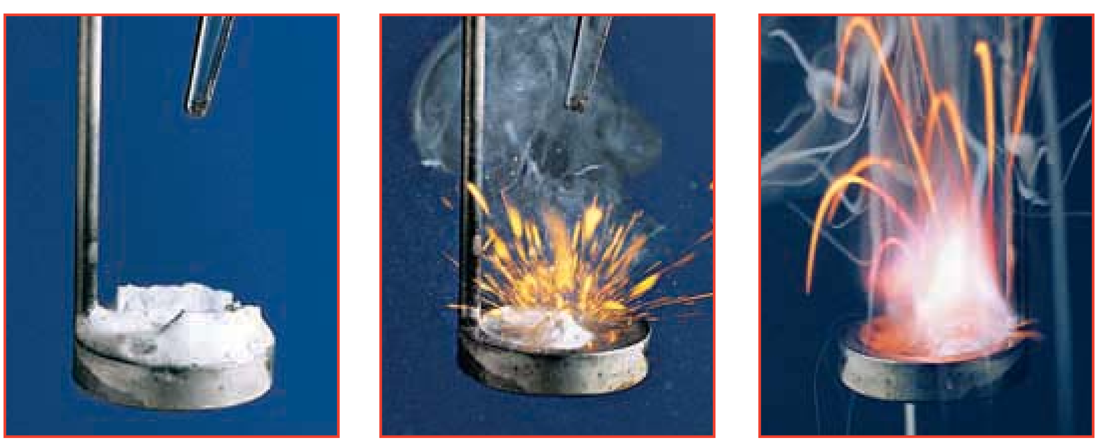
OBJECTIVES

1. **Describe** periodic trends in ionization energy, and relate them to the atomic structures of the elements.
2. **Describe** periodic trends in atomic radius, and relate them to the atomic structures of the elements.
3. **Describe** periodic trends in electronegativity, and relate them to the atomic structures of the elements.
4. **Describe** periodic trends in ionic size, electron affinity, and melting and boiling points, and relate them to the atomic structures of the elements.

Periodic Trends

The arrangement of the periodic table reveals trends in the properties of the elements. A *trend* is a predictable change in a particular direction. For example, there is a trend in the reactivity of the alkali metals as you move down Group 1. As **Figure 15** illustrates, each of the alkali metals reacts with water. However, the reactivity of the alkali metals varies. At the top of Group 1, lithium is the least reactive, sodium is more reactive, and potassium is still more reactive. In other words, there is a trend toward greater reactivity as you move down the alkali metals in Group 1.

Understanding a trend among the elements enables you to make predictions about the chemical behavior of the elements. These trends in properties of the elements in a group or period can be explained in terms of electron configurations.



Lithium Sodium Potassium

Figure 15
Chemical reactivity with water increases from top to bottom for Group 1 elements. Reactions of lithium, sodium, and potassium with water are shown.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

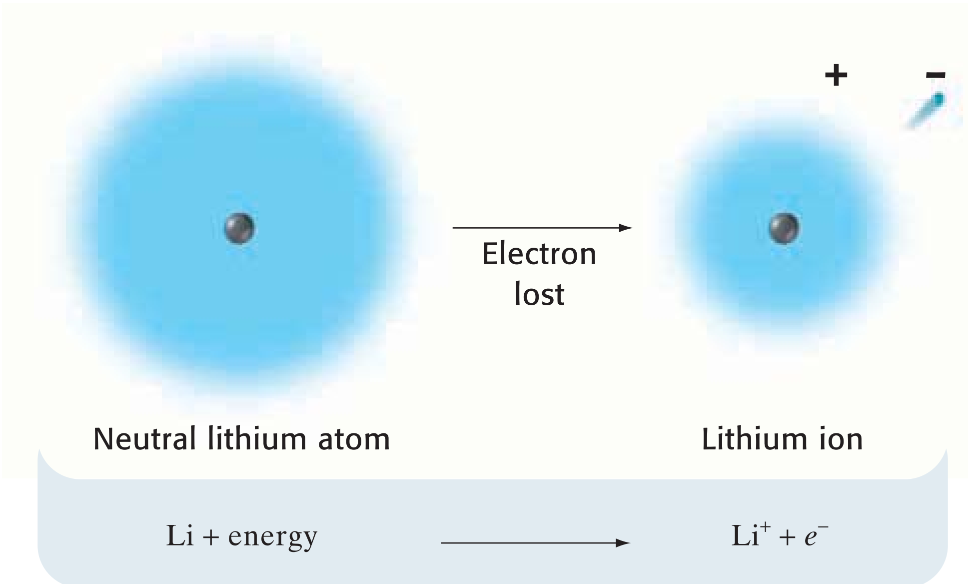


Figure 16
When enough energy is supplied, a lithium atom loses an electron and becomes a positive ion. The ion is positive because its number of protons now exceeds its number of electrons by one.

Ionization Energy

When atoms have equal numbers of protons and electrons, they are electrically neutral. But when enough energy is added, the attractive force between the protons and electrons can be overcome. When this happens, an electron is removed from an atom. The neutral atom then becomes a positively charged ion.

Figure 16 illustrates the removal of an electron from an atom. The energy that is supplied to remove an electron is the **ionization energy** of the atom. This process can be described as shown below.

$$\underset{\text{neutral atom}}{A} + \text{ionization energy} \rightarrow \underset{\text{ion}}{A^+} + \underset{\text{electron}}{e^-}$$

ionization energy
the energy required to remove an electron from an atom or ion

Ionization Energy Decreases as You Move Down a Group

Ionization energy tends to decrease down a group, as **Figure 17** on the next page shows. Each element has more occupied energy levels than the one above it has. Therefore, the outermost electrons are farthest from the nucleus in elements near the bottom of a group.

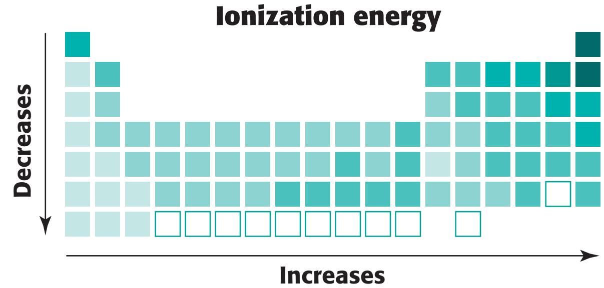
Similarly, as you move down a group, each successive element contains more electrons in the energy levels between the nucleus and the outermost electrons. These inner electrons shield the outermost electrons from the full attractive force of the nucleus. This **electron shielding** causes the outermost electrons to be held less tightly to the nucleus.

electron shielding
the reduction of the attractive force between a positively charged nucleus and its outermost electrons due to the cancellation of some of the positive charge by the negative charges of the inner electrons

Notice in **Figure 18** on the next page that the ionization energy of potassium is less than that of lithium. The outermost electrons of a potassium atom are farther from its nucleus than the outermost electrons of a lithium atom are from their nucleus. So, the outermost electrons of a lithium atom are held more tightly to its nucleus. As a result, removing an electron from a potassium atom takes less energy than removing one from a lithium atom.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Figure 17
Ionization energy generally decreases down a group and increases across a period, as shown in this diagram. Darker shading indicates higher ionization energy.

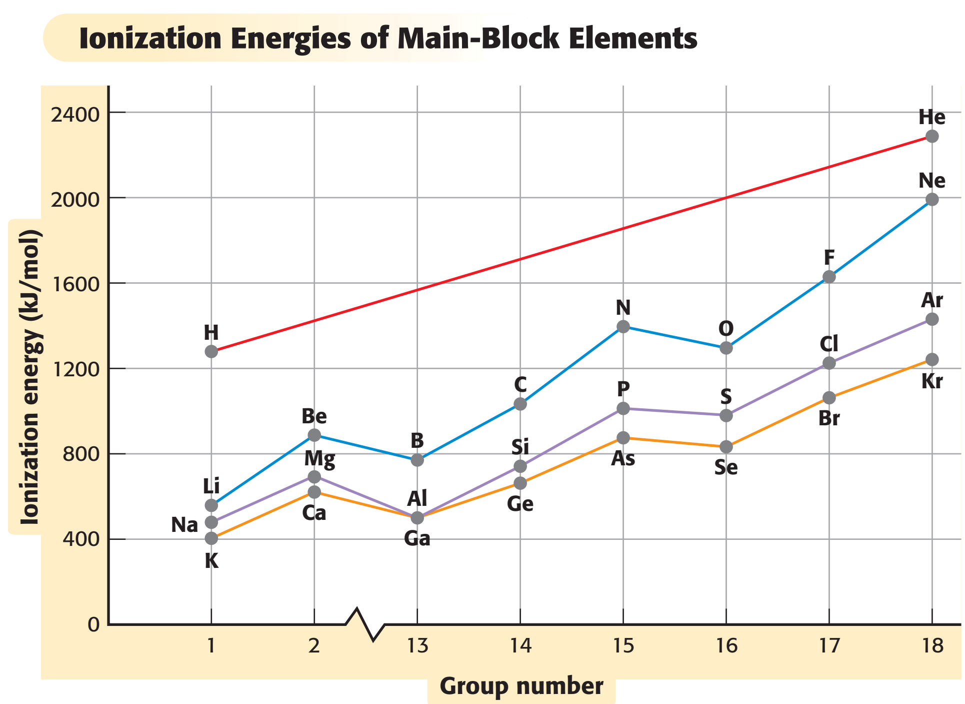


Ionization Energy Increases as You Move Across a Period

Ionization energy tends to increase as you move from left to right across a period, as **Figure 17** shows. From one element to the next in a period, the number of protons and the number of electrons increase by one each. The additional proton increases the nuclear charge. The additional electron is added to the same outer energy level in each of the elements in the period. A higher nuclear charge more strongly attracts the outer electrons in the same energy level, but the electron-shielding effect from inner-level electrons remains the same. Thus, more energy is required to remove an electron because the attractive force on them is higher.

Figure 18 shows that the ionization energy of neon is almost four times greater than that of lithium. A neon atom has 10 protons in its nucleus and 10 electrons filling two energy levels. In contrast, a lithium atom has 3 protons in its nucleus and 3 electrons distributed in the same two energy levels as those of neon. The attractive force between neon's 10 protons and 10 electrons is much greater than that between lithium's 3 protons and 3 electrons. As a result, the ionization energy of neon is much higher than that of lithium.

Figure 18
Ionization energies for hydrogen and for the main-group elements of the first four periods are plotted on this graph.



Copyright © by Holt, Rinehart and Winston. All rights reserved.

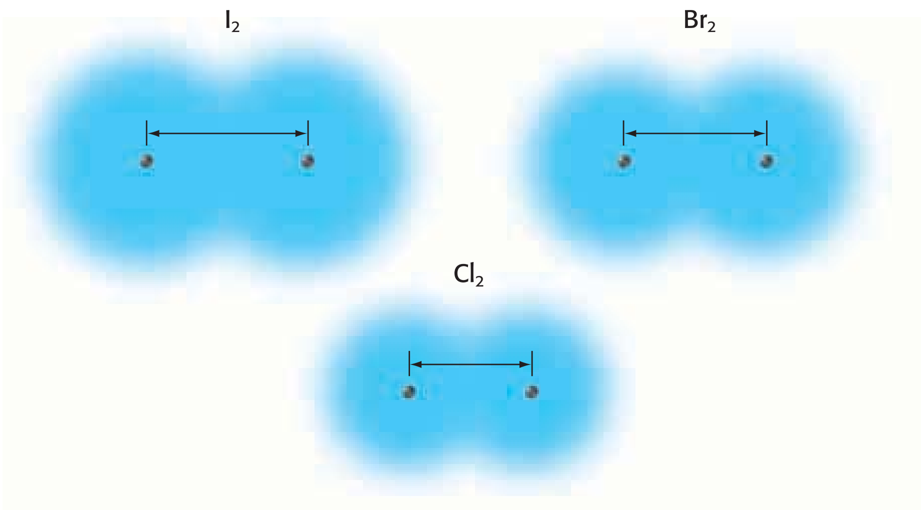


Figure 19
In each molecule, half the distance of the line represents the bond radius of the atom.

Atomic Radius

The exact size of an atom is hard to determine. An atom's size depends on the volume occupied by the electrons around the nucleus, and the electrons do not move in well-defined paths. Rather, the volume the electrons occupy is thought of as an electron cloud, with no clear-cut edge. In addition, the physical and chemical state of an atom can change the size of an electron cloud.

Figure 19 shows one way to measure the size of an atom. This method involves calculating the **bond radius,** the length that is half the distance between the nuclei of two bonded atoms. The bond radius can change slightly depending on what atoms are involved.

bond radius
half the distance from center to center of two like atoms that are bonded together

Atomic Radius Increases as You Move Down a Group

Atomic radius increases as you move down a group, as **Figure 20** shows. As you proceed from one element down to the next in a group, another principal energy level is filled. The addition of another level of electrons increases the size, or atomic radius, of an atom.

Electron shielding also plays a role in determining atomic radius. Because of electron shielding, the effective nuclear charge acting on the outer electrons is almost constant as you move down a group, regardless of the energy level in which the outer electrons are located. As a result, the outermost electrons are not pulled closer to the nucleus. For example, the effective nuclear charge acting on the outermost electron in a cesium atom is about the same as it is in a sodium atom.

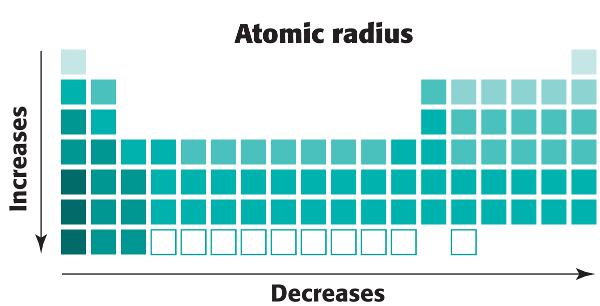


Figure 20
Atomic radius generally increases down a group and decreases across a period, as shown in this diagram. Darker shading indicates higher atomic radius.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

As a member of Period 6, cesium has six occupied energy levels. As a member of Period 3, sodium has only three occupied energy levels. Although cesium has more protons and electrons, the effective nuclear charge acting on the outermost electrons is about the same as it is in sodium because of electron shielding. Because cesium has more occupied energy levels than sodium does, cesium has a larger atomic radius than sodium has. **Figure 21** shows that the atomic radius of cesium is about 230 pm, while the atomic radius of sodium is about 150 pm.

Atomic Radius Decreases as You Move Across a Period

Topic Link

Refer to Appendix A for a chart of relative atomic radii of the elements.

As you move from left to right across a period, each atom has one more proton and one more electron than the atom before it has. All additional electrons go into the same principal energy level—no electrons are being added to the inner levels. As a result, electron shielding does not play a role as you move across a period. Therefore, as the nuclear charge increases across a period, the effective nuclear charge acting on the outer electrons also increases. This increasing nuclear charge pulls the outermost electrons closer and closer to the nucleus and thus reduces the size of the atom.

Figure 21 shows how atomic radii decrease as you move across a period. Notice that the decrease in size is significant as you proceed across groups going from Group 1 to Group 14. The decrease in size then tends to level off from Group 14 to Group 18. As the outermost electrons are pulled closer to the nucleus, they also get closer to one another.

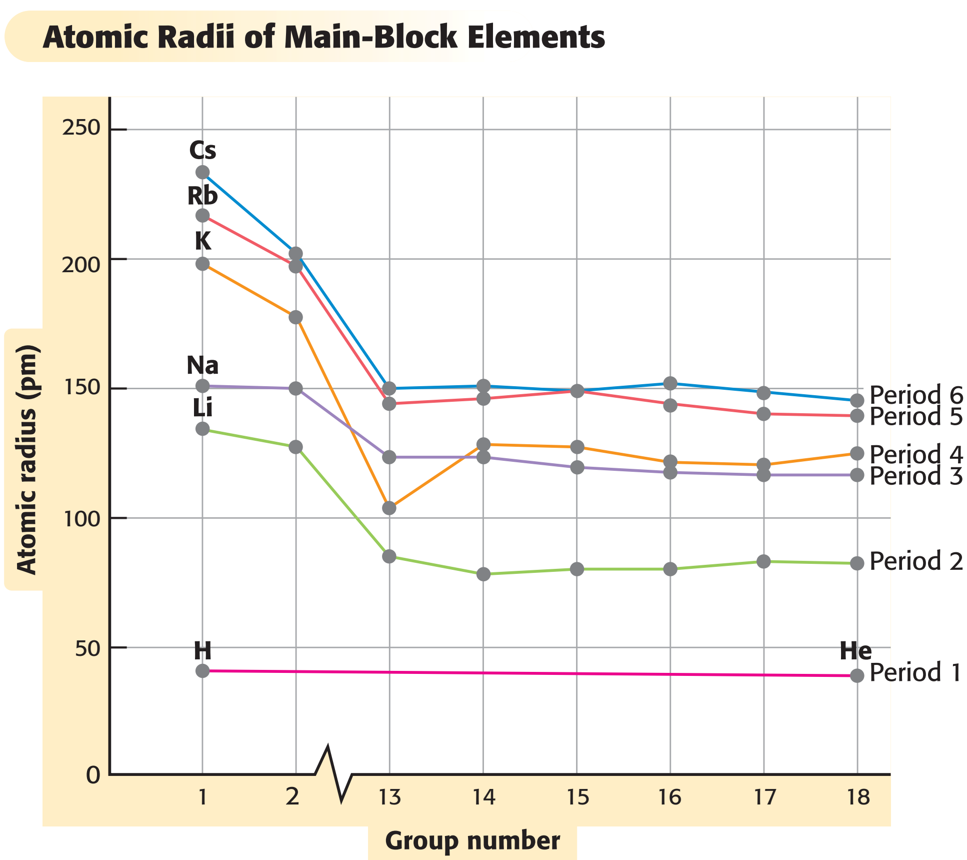


Figure 21
Atomic radii for hydrogen and the main-group elements in Periods 1 through 6 are plotted on this graph.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

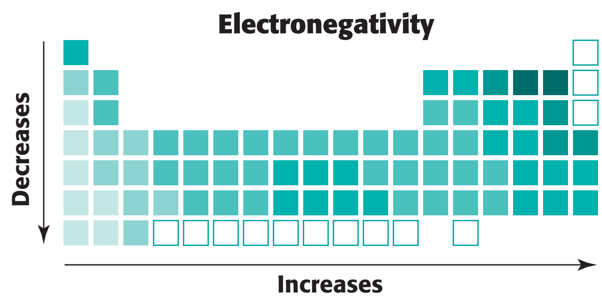


Figure 22
Electronegativity tends to decrease down a group and increase across a period, as shown in this diagram. Darker shading indicates higher electronegativity.

Repulsions between these electrons get stronger. Finally, a point is reached where the electrons will not come closer to the nucleus because the electrons would have to be too close to each other. Therefore, the sizes of the atomic radii level off as you approach the end of each period.

Electronegativity

Atoms often bond to one another to form a compound. These bonds can involve the sharing of valence electrons. Not all atoms in a compound share electrons equally. Knowing how strongly each atom attracts bonding electrons can help explain the physical and chemical properties of the compound.

Linus Pauling, one of America's most famous chemists, made a scale of numerical values that reflect how much an atom in a molecule attracts electrons, called **electronegativity** values. Chemical bonding that comes from a sharing of electrons can be thought of as a tug of war. The atom with the higher electronegativity will pull on the electrons more strongly than the other atom will.

electronegativity
a measure of the ability of an atom in a chemical compound to attract electrons

Fluorine is the element whose atoms most strongly attract shared electrons in a compound. Pauling arbitrarily gave fluorine an electronegativity value of 4.0. Values for the other elements were calculated in relation to this value.

Electronegativity Decreases as You Move Down a Group

Electronegativity values generally decrease as you move down a group, as **Figure 22** shows. Recall that from one element to the next one in a group, the principal quantum number increases by one, so another principal energy level is occupied. The more protons an atom has, the more strongly it should attract an electron. Therefore, you might expect that electronegativity increases as you move down a group.

However, electron shielding plays a role again. Even though cesium has many more protons than lithium does, the effective nuclear charge acting on the outermost electron is almost the same in both atoms. But the distance between cesium's sixth principal energy level and its nucleus is greater than the distance between lithium's third principal energy level and its nucleus. This greater distance means that the nucleus of a cesium atom cannot attract a valence electron as easily as a lithium nucleus can. Because cesium does not attract an outer electron as strongly as lithium, it has a smaller electronegativity value.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

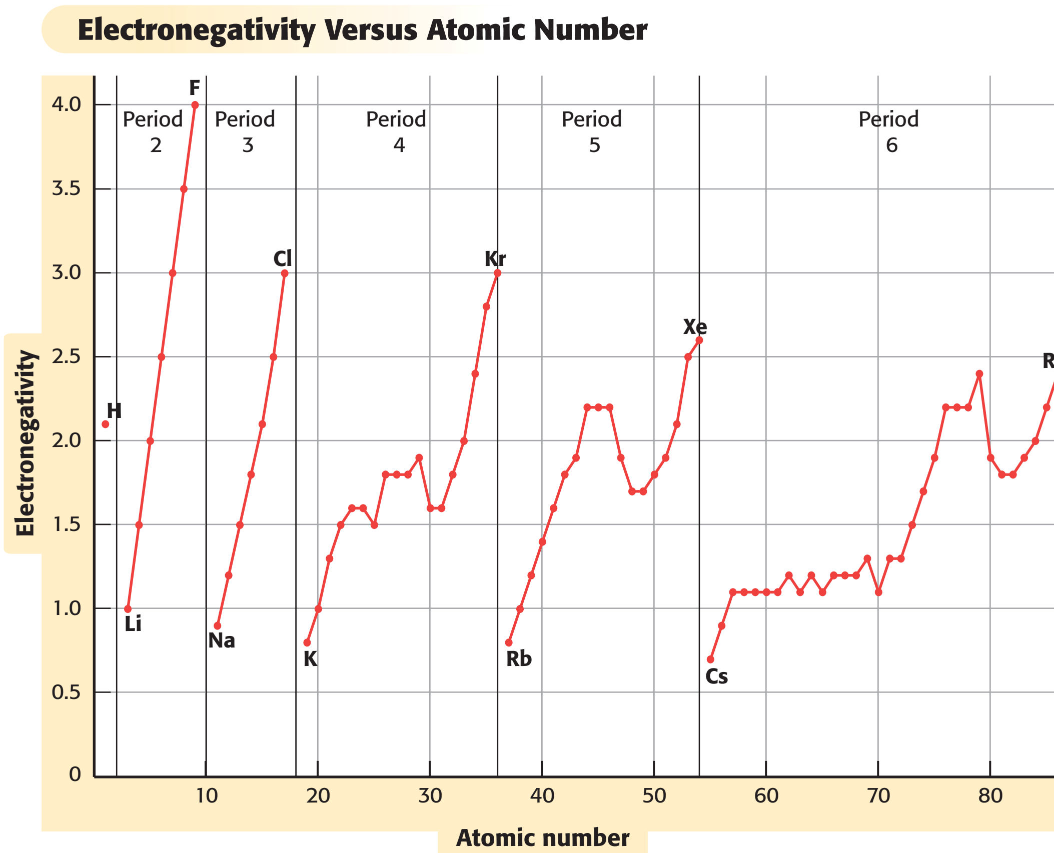


Figure 23
This graph shows electronegativity compared to atomic number for Periods 1 through 6. Electronegativity tends to increase across a period because the effective nuclear charge becomes greater as protons are added.

Electronegativity Increases as You Move Across a Period

As **Figure 23** shows, electronegativity usually increases as you move left to right across a period. As you proceed across a period, each atom has one more proton and one more electron—in the same principal energy level—than the atom before it has. Recall that electron shielding does not change as you move across a period because no electrons are being added to the inner levels. Therefore, the effective nuclear charge increases across a period. As this increases, electrons are attracted much more strongly, resulting in an increase in electronegativity.

Notice in **Figure 23** that the increase in electronegativity across a period is much more dramatic than the decrease in electronegativity down a group. For example, if you go across Period 3, the electronegativity more than triples, increasing from 0.9 for sodium, Na, to 3.2 for chlorine, Cl. In contrast, if you go down Group 1 the electronegativity decreases only slightly, dropping from 0.9 for sodium to 0.8 for cesium, Cs.

This difference can be explained if you look at the changes in atomic structure as you move across a period and down a group. Without the addition of any electrons to inner energy levels, elements from left to right in a period experience a significant increase in effective nuclear charge. As you move down a group, the addition of electrons to inner energy levels causes the effective nuclear charge to remain about the same. The electronegativity drops slightly because of the increasing distance between the nucleus and the outermost energy level.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Other Periodic Trends

You may have noticed that effective nuclear charge and electron shielding are often used in explaining the reasons for periodic trends. Effective nuclear charge and electron shielding also account for two other periodic trends that are related to the ones already discussed: ionic size and electron affinity. Still other trends are seen by examining how melting point and boiling point change as you move across a period or down a group. The trends in melting and boiling points are determined by how electrons form pairs as *d* orbitals fill.

Periodic Trends in Ionic Size and Electron Affinity

Recall that atoms form ions by either losing or gaining electrons. Like atomic size, ionic size has periodic trends. As you proceed down a group, the outermost electrons in ions are in higher energy levels. Therefore, just as atomic radius increases as you move down a group, usually the ionic radius increases as well, as shown in **Figure 24a.** These trends hold for both positive and negative ions.

Metals tend to lose one or more electrons and form a positive ion. As you move across a period, the ionic radii of metal cations tend to decrease because of the increasing nuclear charge. As you come to the nonmetal elements in a period, their atoms tend to gain electrons and form negative ions. **Figure 24a** shows that as you proceed through the anions on the right of a period, ionic radii still tend to decrease because of the anions' increasing nuclear charge.

Neutral atoms can also gain electrons. The energy change that occurs when a neutral atom gains an electron is called the atom's *electron affinity*. This property of an atom is different from electronegativity, which is a measure of an atom's attraction for an electron when the atom is bonded to another atom. **Figure 24b** shows that electron affinity tends to decrease as you move down a group. This trend is due to the increasing effect of electron shielding. In contrast, electron affinity tends to increase as you move across a period because of the increasing nuclear charge.

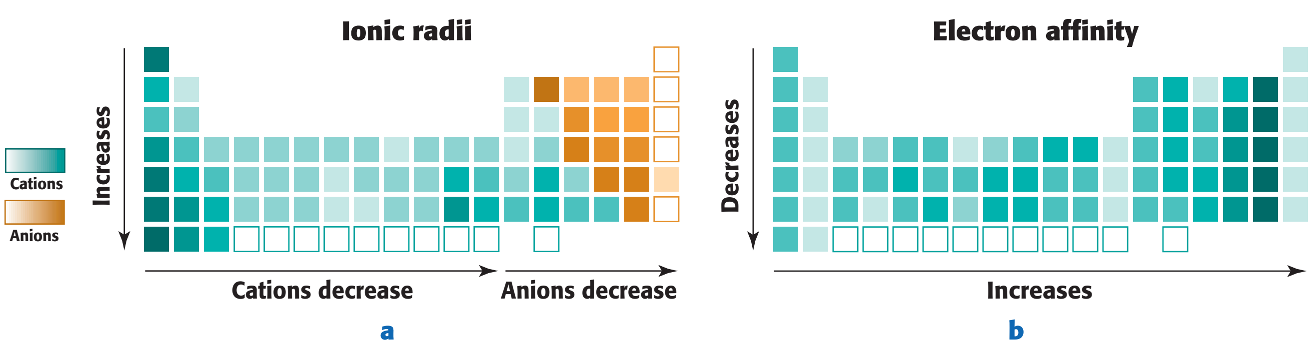


Figure 24
Ionic size tends to increase down groups and decrease across periods. Electron affinity generally decreases down groups and increases across periods.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

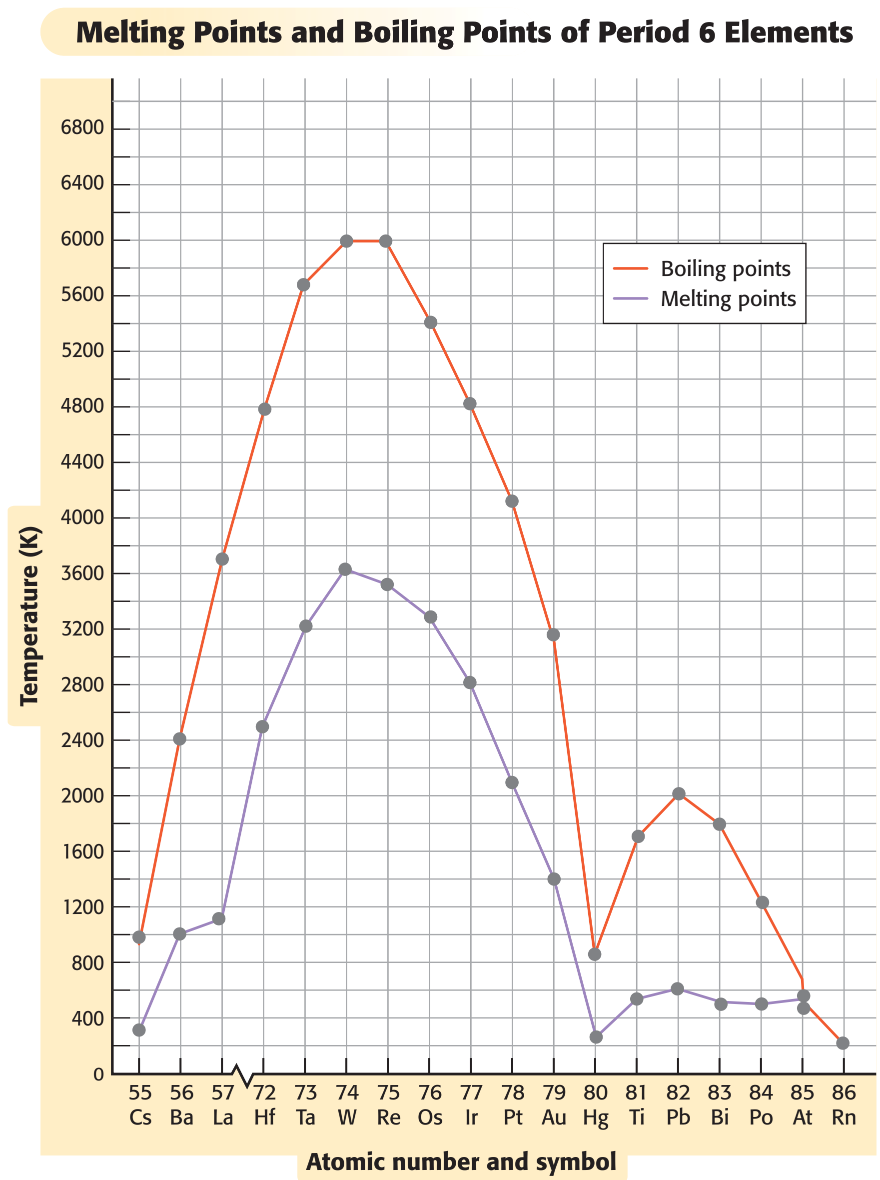
Periodic Trends in Melting and Boiling Points

The melting and boiling points for the elements in Period 6 are shown in **Figure 25.** Notice that instead of a generally increasing or decreasing trend, melting and boiling points reach two different peaks as *d* and *p* orbitals fill.

Cesium, Cs, has low melting and boiling points because it has only one valence electron to use for bonding. From left to right across the period, the melting and boiling points at first increase. As the number of electrons in each element increases, stronger bonds between atoms can form. As a result, more energy is needed for melting and boiling to occur.

Near the middle of the *d*-block, the melting and boiling points reach a peak. This first peak corresponds to the elements whose *d* orbitals are almost half filled. The atoms of these elements can form the strongest bonds, so these elements have the highest melting and boiling points in this period. For Period 6, the elements with the highest melting and boiling points are tungsten, W, and rhenium, Re.

Figure 25
As you move across Period 6, the periodic trend for melting and boiling points goes through two cycles of first increasing, reaching a peak, and then decreasing.



Copyright © by Holt, Rinehart and Winston. All rights reserved.

As more electrons are added, they begin to form pairs within the *d* orbitals. Because of the decrease in unpaired electrons, the bonds that the atoms can form with each other become weaker. As a result, these elements have lower melting and boiling points. The lowest melting and boiling points are reached at mercury, whose *d* orbitals are completely filled. Mercury, Hg, has the second-lowest melting and boiling points in this period. The noble gas radon, Rn, is the only element in Period 6 with a lower boiling point than that of mercury.

As you proceed past mercury, the melting and boiling points again begin to rise as electrons are now added to the *p* orbital. The melting and boiling points continue to rise until they peak at the elements whose *p* orbitals are almost half filled. Another decrease is seen as electrons pair up to fill *p* orbitals. When the noble gas radon, Rn, is reached, the *p* orbitals are completely filled. The noble gases are monatomic and have no bonding forces between atoms. Therefore, their melting and boiling points are unusually low.

3 Section Review

UNDERSTANDING KEY IDEAS

1. What is ionization energy?

2. Why is measuring the size of an atom difficult?

3. What can you tell about an atom that has high electronegativity?

4. How does electron shielding affect atomic size as you move down a group?

5. What periodic trends exist for ionization energy?

6. Describe one way in which *atomic radius* is defined.

7. Explain how the trends in melting and boiling points differ from the other periodic trends.

8. Why do both atomic size and ionic size increase as you move down a group?

9. How is electron affinity different from electronegativity?

10. What periodic trends exist for electronegativity?

11. Why is electron shielding not a factor when you examine a trend across a period?

CRITICAL THINKING

12. Explain why the noble gases have high ionization energies.

13. What do you think happens to the size of an atom when the atom loses an electron? Explain.

14. With the exception of the noble gases, why is an element with a high ionization energy likely to have high electron affinity?

15. Explain why atomic radius remains almost unchanged as you move through Period 2 from Group 14 to Group 18.

16. Helium and hydrogen have almost the same atomic size, yet the ionization energy of helium is almost twice that of hydrogen. Explain why hydrogen has a much higher ionization energy than any element in Group 1 does.

17. Why does mercury, Hg, have such a low melting point? How would you expect mercury's melting point to be different if the *d*-block contained more groups than it does?

18. What exceptions are there in the increase of ionization energies across a period?

Copyright © by Holt, Rinehart and Winston. All rights reserved.

SECTION

4 Where Did the Elements Come From?

KEY TERMS

- **nuclear reaction**
- **superheavy element**

OBJECTIVES

1. **Describe** how the naturally occurring elements form.
2. **Explain** how a transmutation changes one element into another.
3. **Describe** how particle accelerators are used to create synthetic elements.

Natural Elements

Of all the elements listed in the periodic table, 93 are found in nature. Three of these elements, technetium, Tc, promethium, Pm, and neptunium, Np, are not found on Earth but have been detected in the spectra of stars. The nebula shown in **Figure 26** is one of the regions in the galaxy where new stars are formed and where elements are made.

Most of the atoms in living things come from just six elements. These elements are carbon, hydrogen, oxygen, nitrogen, phosphorus, and sulfur. Scientists theorize that these elements, along with all 93 natural elements, were created in the centers of stars billions of years ago, shortly after the universe formed in a violent explosion.

Figure 26
Three natural elements—technetium, promethium, and neptunium—have been detected only in the spectra of stars.

internet connect
www.scilinks.org
Topic: Origin of Elements
SciLinks code: HW4093
SCILINKS. Maintained by the National Science Teachers Association

Copyright © by Holt, Rinehart and Winston. All rights reserved.

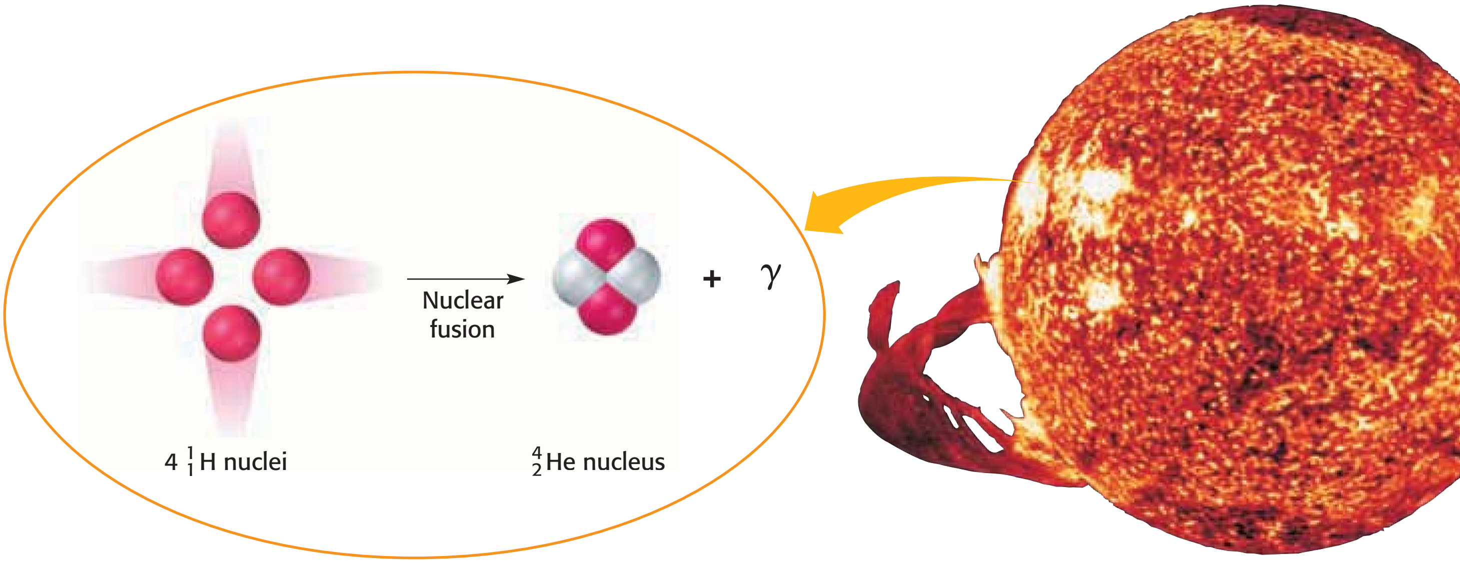


Figure 27
Nuclear reactions like those in the sun can fuse four hydrogen nuclei into one helium nucleus, releasing gamma radiation, γ.

Hydrogen and Helium Formed After the Big Bang

Much of the evidence about the universe's origin points toward a single event: an explosion of unbelievable violence, before which all matter in the universe could fit on a pinhead. This event is known as the *big bang*. Most scientists currently accept this model about the universe's beginnings. Right after the big bang, temperatures were so high that matter could not exist; only energy could. As the universe expanded, it cooled and some of the energy was converted into matter in the form of electrons, protons, and neutrons. As the universe continued to cool, these particles started to join and formed hydrogen and helium atoms.

Over time, huge clouds of hydrogen accumulated. Gravity pulled these clouds of hydrogen closer and closer. As the clouds grew more dense, pressures and temperatures at the centers of the hydrogen clouds increased, and stars were born. In the centers of stars, **nuclear reactions** took place. The simplest nuclear reaction, as shown in **Figure 27,** involves fusing hydrogen nuclei to form helium. Even now, these same nuclear reactions are the source of the energy that we see as the stars' light and feel as the sun's warmth.

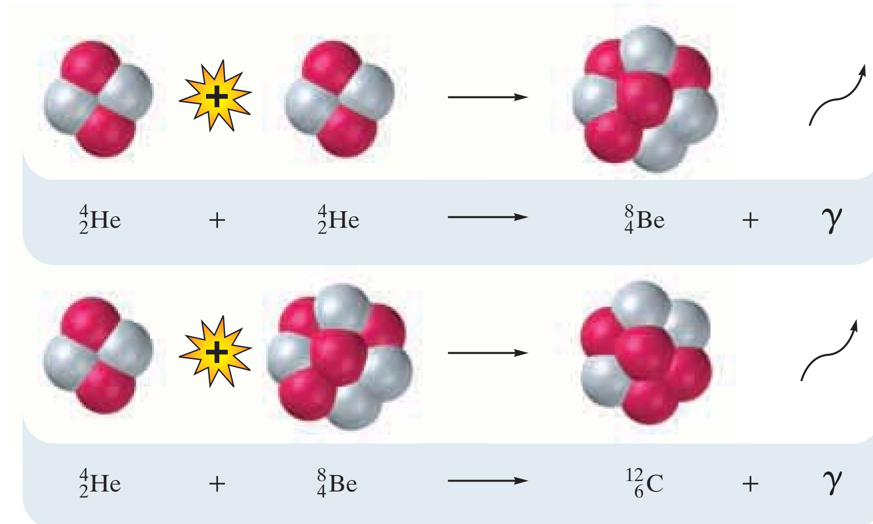
nuclear reaction
a reaction that affects the nucleus of an atom

Other Elements Form by Nuclear Reactions in Stars

The mass of a helium nucleus is less than the total mass of the four hydrogen nuclei that fuse to form it. The mass is not really "lost" in this nuclear reaction. Rather, the missing mass is converted into energy. Einstein's equation $E = mc^2$ describes this mass-energy relationship quantitatively. The mass that is converted to energy is represented by m in this equation. The constant c is the speed of light. Einstein's equation shows that fusion reactions release very large amounts of energy. The energy released by a fusion reaction is so great it keeps the centers of the stars at very high temperatures.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Figure 28
Nuclear reactions can form a beryllium nucleus by fusing helium nuclei. The beryllium nucleus can then fuse with another helium nucleus to form a carbon nucleus.



The temperatures in stars get high enough to fuse helium nuclei with one another. As helium nuclei fuse, elements of still higher atomic numbers form. **Figure 28** illustrates such a process: two helium nuclei fuse to form a beryllium nucleus, and gamma radiation is released. The beryllium nucleus can then fuse with another helium nucleus to form a carbon nucleus. Such repeated fusion reactions can form atoms as massive as iron.

Very massive stars (stars whose masses are more than 100 times the mass of our sun) are the source of heavier elements. When such a star has converted almost all of its core hydrogen and helium into the heavier elements up to iron, the star collapses and then blows apart in an explosion called a *supernova*. All of the elements heavier than iron on the periodic table are formed in this explosion. The star's contents shoot out into space, where they can become part of newly forming star systems.

Transmutations

internet connect
www.scilinks.org
Topic: Alchemy
SciLinks code: HW4006
SCILINKS. Maintained by the National Science Teachers Association

In the Middle Ages, many early chemists tried to change, or transmute, ordinary metals into gold. Although they made many discoveries that contributed to the development of modern chemistry, their attempts to transmute metals were doomed from the start. These early chemists did not realize that a *transmutation*, whereby one element changes into another, is a nuclear reaction. It changes the nucleus of an atom and therefore cannot be achieved by ordinary chemical means.

Transmutations Are a Type of Nuclear Reaction

Although nuclei do not change into different elements in ordinary chemical reactions, transmutations can happen. Early chemists such as John Dalton had insisted that atoms never change into other elements, so when scientists first encountered transmutations in the 1910s, their results were not always believed.

While studying the passage of high-speed alpha particles (helium nuclei) through water vapor in a cloud chamber, Ernest Rutherford observed some long, thin particle tracks. These tracks matched the ones caused by protons in experiments performed earlier by other scientists.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

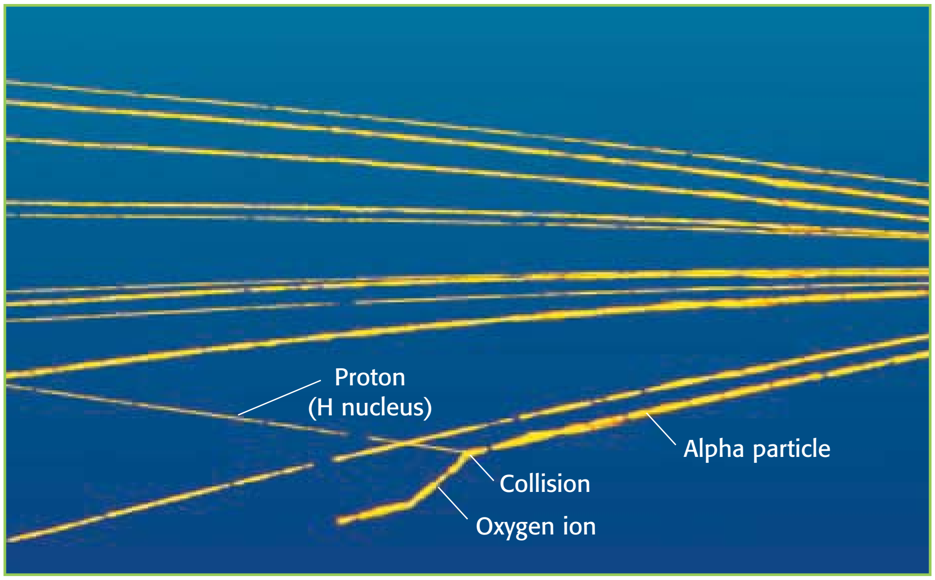


Figure 29
Observe the spot in this cloud-chamber photo where an alpha particle collided with the nucleus of a nitrogen atom. The left track was made by an oxygen atom; the right track, by a proton.

Rutherford reasoned correctly that the atomic nuclei in air were disintegrating upon being struck by alpha particles. He believed that the nuclei in air had disintegrated into the nuclei of hydrogen (protons) plus the nuclei of some other atom.

Two chemists, an American named W. D. Harkins and an Englishman named P.M.S. Blackett, studied this strange phenomenon further. Blackett took photos of 400 000 alpha particle tracks that formed in cloud chambers. He found that 8 of these tracks forked to form a **Y**, as shown in **Figure 29.** Harkins and Blackett concluded that the **Y** formed when an alpha particle collided with a nitrogen atom in air to produce an oxygen atom and a proton, and that a transmutation had thereby occurred.

Synthetic Elements

The discovery that a transmutation had happened started a flood of research. Soon after Harkins and Blackett had observed a nitrogen atom forming oxygen, other transmutation reactions were discovered by bombarding various elements with alpha particles. As a result, chemists have synthesized, or created, more elements than the 93 that occur naturally. These are *synthetic elements*. All of the transuranium elements, or those with more than 92 protons in their nuclei, are synthetic elements. To make them, one must use special equipment, called *particle accelerators,* described below.

The Cyclotron Accelerates Charged Particles

Many of the first synthetic elements were made with the help of a cyclotron, a particle accelerator invented in 1930 by the American scientist E.O. Lawrence. In a cyclotron, charged particles are given one pulse of energy after another, speeding them to very high energies. The particles then collide and fuse with atomic nuclei to produce synthetic elements that have much higher atomic numbers than naturally occurring elements do. However, there is a limit to the energies that can be reached with a cyclotron and therefore a limit to the synthetic elements that it can make.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

The Synchrotron Is Used to Create Superheavy Elements

As a particle reaches a speed of about one-tenth the speed of light, it gains enough energy such that the relation between energy and mass becomes an obstacle to any further acceleration. According to the equation $E = mc^2$, the increase in the particle's energy also means an increase in its mass. This makes the particle accelerate more slowly so that it arrives too late for the next pulse of energy from the cyclotron, which is needed to make the particle go faster.

The solution was found with the synchrotron, a particle accelerator that times the pulses to match the acceleration of the particles. A synchrotron can accelerate only a few types of particles, but those particles it can accelerate reach enormous energies. Synchrotrons are now used in many areas of basic research, including explorations into the foundations of matter itself. The Fermi National Accelerator Laboratory in Batavia, IL has a circular accelerator which has a circumference of 4 mi! Subatomic particles are accelerated through this ring to 99.9999% of the speed of light.

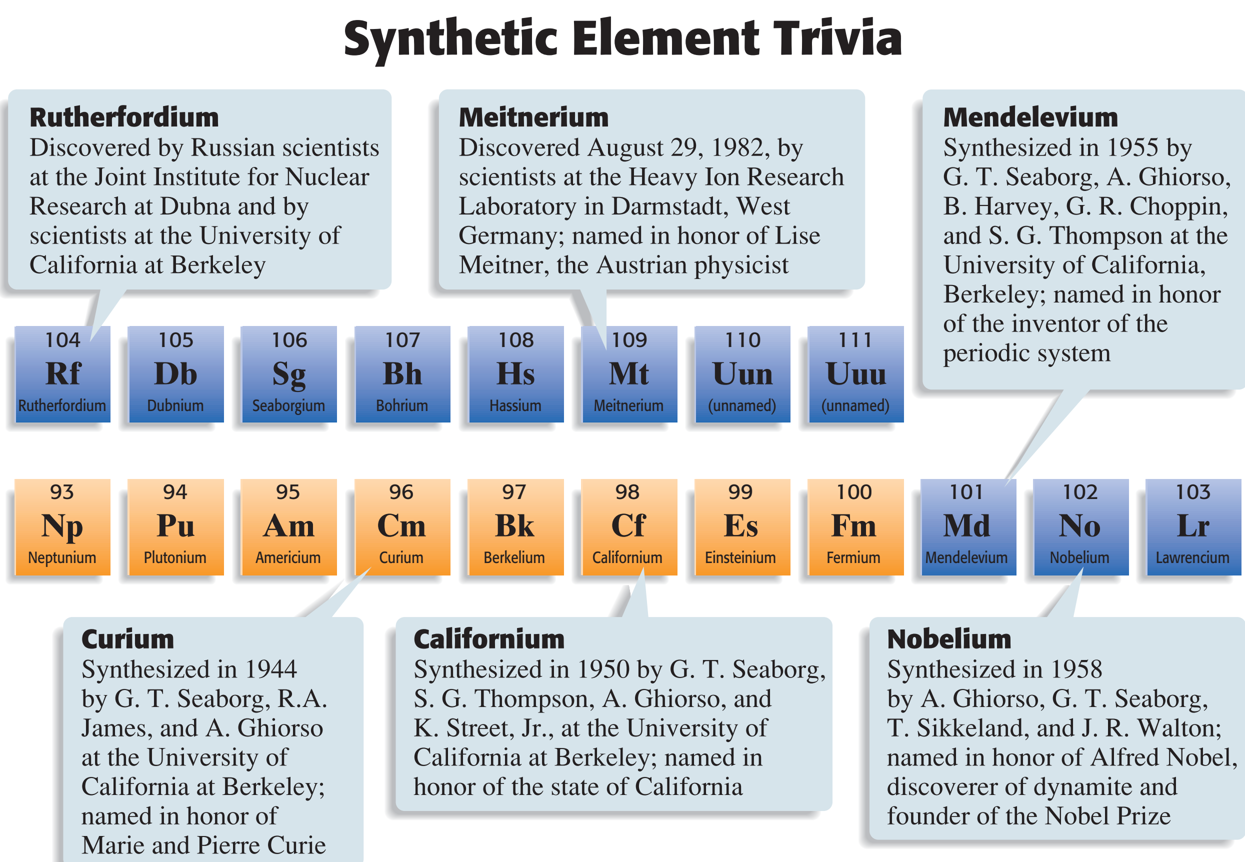


Figure 30
All of the highlighted elements are synthetic. Those shown in orange were created by making moving particles collide with stationary targets. The elements shown in blue were created by making nuclei collide.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Once the particles have been accelerated, they are made to collide with one another. **Figure 30** shows some of the **superheavy elements** created with such collisions. When a synchrotron is used to create an element, only a very small number of nuclei actually collide. As a result, only a few nuclei may be created in these collisions. For example, only three atoms of meitnerium were detected in the first attempt, and these atoms lasted for only 0.0034 s. Obviously, identifying elements that last for such a short time is a difficult task. Scientists in only a few nations have the resources to carry out such experiments. The United States, Germany, Russia, and Sweden are the locations of the largest such research teams.

superheavy element

an element whose atomic number is greater than 106

One of the recent superheavy elements that scientists report is element 114. To create element 114, Russian scientists took plutonium-244, supplied by American scientists, and bombarded it with accelerated calcium-40 atoms for 40 days. In the end, only a single nucleus was detected. It lasted for 30 seconds before decaying into element 112.

Most superheavy elements exist for only a tiny fraction of a second. Thirty seconds is a very long life span for a superheavy element. This long life span of element 114 points to what scientists have long suspected: that an "island of stability" would be found beginning with element 114. Based on how long element 114 lasted, their predictions may have been correct. However, scientists still must try to confirm that element 114 was in fact created. The results of a single experiment are never considered valid unless the experiments are repeated and produce the same results.

4 Section Review

UNDERSTANDING KEY IDEAS

1. How and where did the natural elements form?
2. What element is the building block for all other natural elements?
3. What is a synthetic element?
4. What is a transmutation?
5. Why is transmutation classified as a nuclear reaction?
6. How did Ernest Rutherford deduce that he had observed a transmutation in his cloud chamber?
7. How are cyclotrons used to create synthetic elements?
8. How are superheavy elements created?

CRITICAL THINKING

9. Why is the following statement *not* an example of a transmutation? Zinc reacts with copper sulfate to produce copper and zinc sulfate.
10. Elements whose atomic numbers are greater than 92 are sometimes referred to as the transuranium elements. Why?
11. Why must an extremely high energy level be reached before a fusion reaction can take place?
12. If the synchrotron had not been developed, how would the periodic table look?
13. What happens to the mass of a particle as the particle approaches the speed of light?
14. How many different kinds of nuclear reactions must protons go through to produce a carbon atom?

Copyright © by Holt, Rinehart and Winston. All rights reserved.

SCIENCE AND TECHNOLOGY

CAREER APPLICATION

Materials Scientist

A materials scientist is interested in discovering materials that can last through harsh conditions, have unusual properties, or perform unique functions. These materials might include the following: a lightweight plastic that conducts electricity; extremely light but strong materials to construct a space platform; a plastic that can replace iron and aluminum in building automobile engines; a new building material that expands and contracts very little, even in extreme temperatures; or a strong, flexible, but extremely tough material that can replace bone or connective tissue in surgery. Materials engineers develop such materials and discover ways to mold or shape these materials into usable forms. Many materials scientists work in the aerospace industry and develop new materials that can lower the mass of aircraft, rockets, and space vehicles.

internet connect

www.scilinks.org
Topic: Superconductors
SciLinks code: HW4170

SCILINKS Maintained by the National Science Teachers Association

Superconductors

The strong magnetic field produced by these superconducting electromagnets can suspend this 8 cm disk.

Superconductivity Discovered

It has long been known that a metal becomes a better conductor as its temperature is lowered. In 1911, Heike Kamerlingh Onnes, a Dutch physicist, was studying this effect on mercury. When he used liquid helium to cool the metal to about −269°C, an unexpected thing happened—the mercury lost all resistance and became a superconductor. Scientists were excited about this new discovery, but the use of superconductors was severely limited by the huge expense of cooling them to near absolute zero. Scientists began research to find a material that would superconduct at temperatures above −196°C, the boiling point of cheap-to-produce liquid nitrogen.

"High-Temperature" Superconductors

Finally, in 1987 scientists discovered materials that became superconductors when cooled to only −183°C. These "high-temperature" superconductors were not metals but ceramics; usually copper oxides combined with elements such as yttrium or barium.

High-temperature superconductors are used in building very powerful electromagnets that are not limited by resistance or heat. These magnets can be used to build powerful particle accelerators and high-efficiency electric motors and generators. Engineers are working to build a system that will use superconducting electromagnets to levitate a passenger train above its guide rail so that the train can move with little friction and thus save fuel.

Questions

1. How does temperature normally affect electrical conductivity in metals?
2. What happened unexpectedly when mercury was cooled to near absolute zero?
3. How might consumers benefit from the use of superconducting materials?

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Chapter Highlights 4

KEY TERMS / KEY IDEAS

SECTION ONE How Are Elements Organized?

Key terms: **periodic law**, **valence electron**, **group**, **period**

- John Newlands, Dmitri Mendeleev, and Henry Moseley contributed to the development of the periodic table.
- The periodic law states that the properties of elements are periodic functions of the elements' atomic numbers.
- In the periodic table, elements are ordered by increasing atomic number. Rows are called *periods*. Columns are called *groups*.
- Elements in the same period have the same number of occupied energy levels. Elements in the same group have the same number of valence electrons.

SECTION TWO Tour of the Periodic Table

Key terms: **main-group element**, **alkali metal**, **alkaline-earth metal**, **halogen**, **noble gas**, **transition metal**, **lanthanide**, **actinide**, **alloy**

- The main-group elements are Group 1 (alkali metals), Group 2 (alkaline-earth metals), Groups 13–16, Group 17 (halogens), and Group 18 (noble gases).
- Hydrogen is in a class by itself.
- Most elements are metals, which conduct electricity. Metals are also ductile and malleable.
- Transition metals, including the lanthanides and actinides, occupy the center of the periodic table.

SECTION THREE Trends in the Periodic Table

Key terms: **ionization energy**, **electron shielding**, **bond radius**, **electronegativity**

- Periodic trends are related to the atomic structure of the elements.
- Ionization energy, electronegativity, and electron affinity generally increase as you move across a period and decrease as you move down a group.
- Atomic radius and ionic size generally decrease as you move across a period and increase as you move down a group.
- Melting points and boiling points pass through two cycles of increasing, peaking, and then decreasing as you move across a period.

SECTION FOUR Where Did the Elements Come From?

Key terms: **nuclear reaction**, **superheavy element**

- The 93 natural elements were formed in the interiors of stars. Synthetic elements (elements whose atomic numbers are greater than 93) are made using particle accelerators.
- A transmutation is a nuclear reaction in which one nucleus is changed into another nucleus.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

4

Physical Setting/Chemistry
REGENTS EXAM PRACTICE

PART A

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

1. In the modern periodic table, elements are arranged in order of
(1) decreasing atomic mass.
(2) Mendeleev's original model.
(3) increasing atomic number.
(4) year of discovery.

2. The outer energy level of an atom contains particles known as
(1) kernel electrons.
(2) valence protons.
(3) kernel protons.
(4) valence electrons.

3. The periodic law states that
(1) repeating properties of elements change periodically with their atomic number.
(2) periodic properties of the elements are functions of their atomic masses.
(3) elements whose atoms have the same number of electrons in their outermost energy level must be in the same period.
(4) elements whose atoms have the same number of occupied energy levels must be in the same group.

4. A horizontal row on a periodic table is called a
(1) family.
(2) period.
(3) group.
(4) quantum.

5. Most of the elements on a periodic table are
(1) noble gases.
(2) nonmetals.
(3) metals.
(4) halogens.

6. Which applies to all the elements in Group 18?
(1) They are known as alkali metals.
(2) Their atoms have a full set of electrons in their outermost energy level.
(3) Their atoms have a total of 18 electrons.
(4) They are highly reactive and combine with other elements very easily.

7. Which is a physical property of all metals?
(1) strong resistance to abrasion
(2) low melting point
(3) good conductor of electricity
(4) high solubility in water

8. All elements in Group 2
(1) are also members of Period 2 in the periodic table.
(2) have the same atomic mass.
(3) are known as the alkaline-earth metals.
(4) have the same chemical stability as the elements in Group 18.

9. Group 17 elements, the halogens, are the most reactive of the nonmetal elements because
(1) their atoms require only one electron to fill their outermost energy level.
(2) they are the nonmetals farthest to the right in a periodic table.
(3) their atoms have the same electron configuration as the alkali metals.
(4) their atoms have the same electron configuration as the metals in Group 2.

Regents Test-Taking Tip

If time permits, take short mental breaks during the test to improve your concentration.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

10. Which is a transition element?
(1) Fe
(2) Ba
(3) C
(4) Xe

11. Compared to the alkali metals, the alkaline-earth metals
(1) are more reactive.
(2) are less reactive.
(3) have fewer valence electrons.
(4) cannot react with nonmetals such as the halogens.

12. Which group on the periodic table includes elements that represent all three states of matter at room temperature?
(1) Group 1
(2) Group 16
(3) Group 17
(4) Group 18

13. Which element has the highest electronegativity value?
(1) carbon
(2) oxygen
(3) hydrogen
(4) fluorine

14. What are two properties of most nonmetals?
(1) high ionization energy and poor electrical conductivity
(2) high ionization energy and good electrical conductivity
(3) low ionization energy and poor electrical conductivity
(4) low ionization energy and good electrical conductivity

15. Which of the following atoms requires the least energy to remove its most loosely bound electron?
(1) K
(2) Cu
(3) Ca
(4) C

16. According to the succession of elements within a period, their atoms generally have
(1) increasing radii and increasing ionization energies.
(2) increasing radii and decreasing ionization energies.
(3) decreasing radii and decreasing ionization energies.
(4) decreasing radii and increasing ionization energies.

17. According to the succession of elements within a group, atomic radius
(1) generally increases.
(2) generally decreases.
(3) does not change.
(4) varies unpredictably.

18. Elements from which area of the periodic table generally have the lowest electronegativity values?
(1) upper right
(2) upper left
(3) lower left
(4) lower right

19. The energy change that occurs when a neutral atom gains an electron is called the atom's
(1) ionization energy.
(2) electronegativity.
(3) electron affinity.
(4) potential energy.

20. What is the name of the process by which one element is changed into another element during a nuclear reaction?
(1) electron shielding
(2) transmutation
(3) particle acceleration
(4) orbital hybridization

Copyright © by Holt, Rinehart and Winston. All rights reserved.

PART B–1

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

21. Identify the period containing the element whose atoms have the following electron configuration.

$$[Kr]4d^{10}5s^2 5p^1$$

(1) Period 2
(2) Period 3
(3) Period 4
(4) Period 5

22. Which statement applies to an element whose atoms have the following electron configuration?

$$[Xe]6s^2$$

(1) This element is found in Group 6 of a periodic table.
(2) This element is found in Group 2 of a periodic table.
(3) This element has the same chemical properties as the element whose atom has the electron configuration $[Xe]6s^1$.
(4) An atom of this element has six electrons in its outermost energy level.

23. Which is the outermost, occupied energy level in atoms of the elements in Period 4?
(1) 1
(2) 2
(3) 3
(4) 4

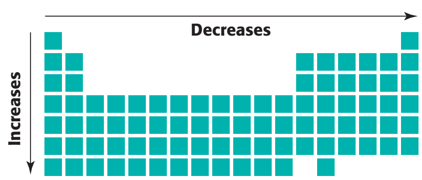
24. Which of the following series of elements contains two transition metals?
(1) N, O, F, and N
(2) Ar, Kr, Cr, and Sr
(3) Co, P, Ne, and Cu
(4) Be, Mg, Ca, and Sr

25. An element that is an excellent conductor of electricity
(1) can be classified as a halogen.
(2) can be a member of Period 7.
(3) has either six or seven valence electrons.
(4) can be found in Group 18.

26. Which of the following physical properties could be the basis for separating the metals that are mixed to make an alloy?
(1) color
(2) melting point
(3) atomic mass
(4) atomic number

27. Although it is in a class by itself, the element hydrogen is placed nearest to Group 1 because its atom has
(1) one electron in its outermost energy level like those of all the alkali metals.
(2) the same atomic mass as the alkali metals.
(3) the same electron configuration as the atoms of all the alkali metals.
(4) a low ionization energy like all the elements in Group 1.

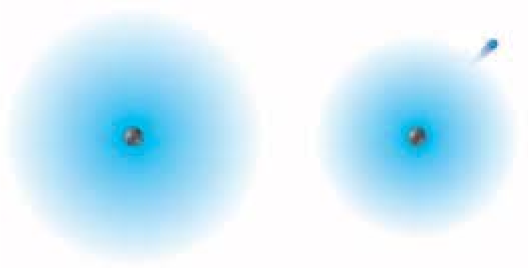
28. Which periodic trend is illustrated by the diagram below?



(1) ionization energy
(2) atomic radius
(3) electronegativity
(4) electron affinity

Copyright © by Holt, Rinehart and Winston. All rights reserved.

Question 29 is based on the diagram below which illustrates an atom on the left.



29. Compared to the atom, the structure shown on the right has a smaller radius because it has
(1) a stable electron configuration like that of a noble gas.
(2) the same number of electrons and protons and is therefore neutral.
(3) one less electron than its parent atom.
(4) one more electron than its parent atom.

30. The ionization energy for iodine (I) is about twice that for rubidium (Rb) even though they are both in the same period. The reason for this difference in their ionization energies is that an atom of iodine
(1) has a radius that is about twice that of rubidium.
(2) has seven electrons in its outermost energy level whereas rubidium has only one.
(3) has twice as many electrons in its outermost energy level as compared to rubidium.
(4) cannot achieve a stable electron configuration.

31. In contrast to natural elements, synthetic elements
(1) include all the elements in period 6.
(2) are formed by nuclear reactions that occur in the stars.
(3) include those elements that have atomic numbers greater than 93.
(4) are created with the help of high-temperature superconductors.

PART B–2

Answer the following items.

32. Are elements with similar chemical properties more likely to be found in the same period or in the same group? Explain our answer.

33. Why is barium (Ba) placed in Group 2 and in Period 6?

34. What determines the horizontal arrangement of the periodic table?

35. How do tellurium (Te) and iodine (I) show that the periodic table is not arranged on the basis of increasing atomic mass?

36. Explain how the electron configurations of the atoms of the transition metals differ from the atoms of metals in Groups 1 and 2.

37. Why are the lanthanides and the actinides placed at the bottom of a periodic table?

38. The atomic number of yttrium, which follows strontium in a periodic table, exceeds the atomic number of strontium by one. In contrast, barium is 18 atomic numbers after strontium, but it falls directly beneath strontium on a periodic table. Does strontium share more properties with yttrium or barium? Explain your answer.

Copyright © by Holt, Rinehart and Winston. All rights reserved.

39. In the middle ages, many alchemists tried to transform lead into gold. The current periodic table places lead and gold in the same period but in different groups. Explain why lead and gold are in the same period but in different groups.

40. If an element breaks when it is struck with a hammer, must it be a metal? Explain your answer.

41. Consider two main-group elements, A and B. Element A has an ionization energy of 419 kJ/mol. Element B has an ionization energy of 1000 kJ/mol. Which element is more likely to form a cation?

42. Argon differs from both chlorine and potassium by one proton each. Compare the electron configurations of these three elements to explain the reactivity of these elements.

43. Why is it highly unlikely for calcium to form a Ca^{+} cation and for sodium to form a Na^{2+} cation?

44. How would you prove that an alloy is a mixture and not a compound?

45. Use the periodic table to describe the chemical properties of the following elements:

a. iodine, I
b. krypton, Kr
c. rubidium, Rb

PART C

Answer the following items.

46. In his periodic table, Mendeleev placed fluorine, chlorine, bromine, and iodine in the same group that he labeled “VII.” Examine the electron configurations of these elements, and explain why Mendeleev grouped the elements this way.

47. While at an amusement park, you inhale helium from a balloon to make your voice higher pitched. A friend says that the helium reacts with and tightens your vocal cords. As a result, your vocal cords might get damaged. Explain why your friend cannot be correct?

48. In trying to determine the atomic radii of various atoms, a scientist measures the radius of a bonded atom five different times and gets five different results. The method used is correct, and the instrumentation is working correctly. How are different results possible?

49. In 1934, Irène Joliot-Curie created the first artificial radioactive isotope-phosphorus-30. She created this artificial isotope by bombarding aluminum-27, a shiny metal with conductive properties, with helium nuclei. The resulting product was a nonmetal with completely different properties from aluminum. What caused the change in properties?

Copyright © by Holt, Rinehart and Winston. All rights reserved.

FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."

50. What relationship is represented in the graph shown?

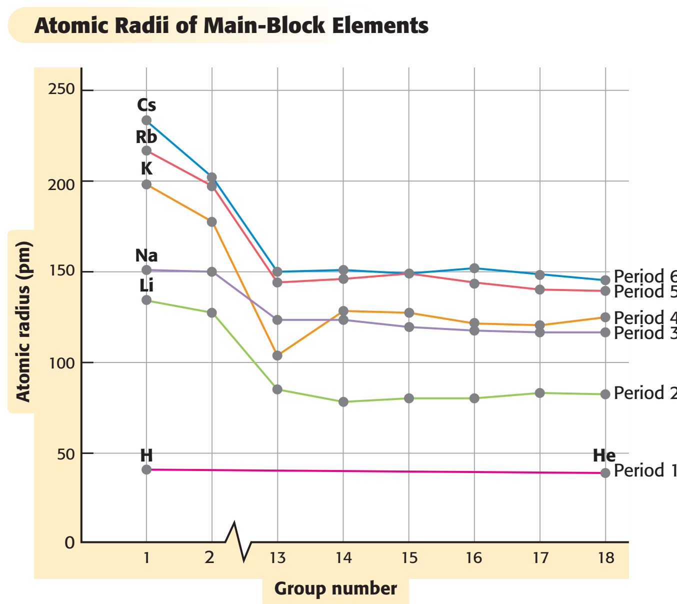
51. What do the numbers on the *y*-axis represent?

52. In every Period, which Group contains the element with the greatest atomic radius?

53. Why is the axis representing group number drawn the way it is in going from Group 2 to Group 13?

54. Which period shows the greatest change in atomic radius?

55. Notice that the points plotted for the elements in Periods 5 and 6 of Group 2 overlap. What does this overlap indicate?



TECHNOLOGY AND LEARNING

56. 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.

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. Press the APPS key on your calculator, then choose the application CHEMAPPS. Press 8, then highlight ALL on the screen, press 1, then highlight LOAD and press 2 to load the data into your calculator. Quit the application, and then run the program RADIUS. For L_1, press 2nd and LIST, and choose ATNUM. For L_2, press 2nd and LIST and choose ATRAD.

If you are using another calculator, your teacher will provide you with keystrokes and data sets to use.

a. Would you expect any atomic number to have an atomic radius of 20 pm? Explain.

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. Does this set of data represent a function? Explain.

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

Copyright © by Holt, Rinehart and Winston. All rights reserved.