The Periodic Table

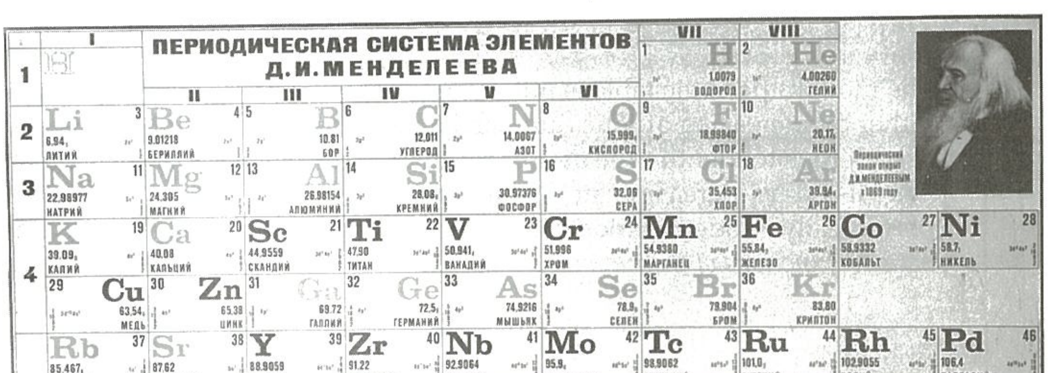
TOPIC 5

How Scientists Use the Periodic Table

Is a Russian periodic table the same as an American periodic table?

The first periodic table was developed in Russia almost 150 years ago. Look at part of the original periodic table in the picture below. Even though you might not be able to read Russian, think about what you know about the periodic table. See if you can answer these questions:

- **What portion of the table is shown? How do you know?**
- **What information is given about each element?**
- **What seems to be different from a modern periodic table?**



TOPIC

5 The Periodic Table

Vocabulary

atomic radius	ionic radius	noble gas
electronegativity	ionization energy	nonmetal
family	metal	periodic law
group	metalloid	period

Topic Overview

Look in almost any chemistry classroom or lab here in the United States, or for that matter, almost anywhere in the world, and you will find a periodic table. The table is a brilliant arrangement of the elements used by chemists everywhere. Just finding a given element in the periodic table tells you much about an element.

Currently, more than 100 elements are known. Most of them occur naturally, while others are made artificially. These elements vary greatly in their physical and chemical properties as well as in the characteristics of their compounds. It has long been recognized that if the elements could be classified, it would simplify their study. In this topic we will present the arrangement of the periodic table. Next, we will examine different types of elements. Finally, trends of important properties will be considered.

Classifying Elements

While there were early attempts to classify and arrange the elements in some orderly fashion, it was Dmitri Mendeleev, a Russian chemist, who is given credit for first arranging elements in a usable manner. Mendeleev observed that when the elements were arranged in order of increasing atomic mass, similar chemical and physical properties appeared at regular, or periodic, intervals. Mendeleev's work in the middle of the nineteenth century laid the basis for the periodic table as we know it today.

However, in Mendeleev's periodic table, the properties of several pairs of elements, such as iodine and tellurium, seemed out of order. If they were switched on the table, the properties would match better, but they would not be in order of increasing atomic mass.

The modern periodic table is not arranged by increasing atomic mass, but rather by increasing atomic number. Henry Moseley, an English scientist, used X rays to identify the atomic number of the elements. If the elements were listed by increasing atomic number, the properties repeated periodically. Modern **periodic law** states: *The properties of the elements are periodic functions of their atomic numbers.*

Table Information about the Elements

The periodic table is an arrangement of the elements, from left to right across each descending row, in order of increasing atomic number. The periodic table used in the *Reference Tables for Physical Setting/Chemistry* displays some properties of each of the elements. To fully understand some of the information the periodic table provides, refer to Figure 5-1 as the boxes on the periodic table are discussed.

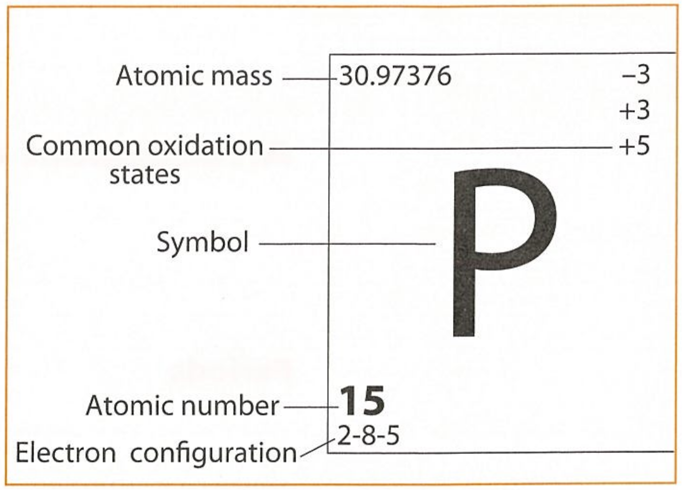


Figure 5-1. Sample information from the periodic table: Although the information about the elements can differ from one periodic table to another, some information, such as atomic mass, atomic number, and symbol, are common to almost all tables.

Chemical Symbols

The symbol of each element is usually found in the center of the box. With over 100 elements to refer to, a symbol is a short and easy way to indicate what element you're talking about, and scientists all over the world can identify an element by its symbol.

Each symbol has one, two, or three letters. The first letter is capitalized, and any other letters present are lower case.

The symbol is related to the name of the element, although sometimes the relationship is not obvious. For example, carbon has an obvious symbol of C. Sodium, however, has the symbol Na, which comes from its Latin name of *natrium*.

Symbols are assigned by an organization known as IUPAC after agreement on the name of a newly discovered element. Agreement has now been reached for a total of 114 elements including numbers 1-112, ending with the element copernicium, Cn. Additionally elements Fl (114) and Lv (116) have been added. Systematic names are assigned to identify these elements as shown in Fig. 5-2. For example # 118 has the systematic symbol of Ununoctium.

Roots Used for Naming Elements

Digit	Root	Digit	Root
0	nil	5	pent
1	un	6	hex
2	bi	7	sept
3	tri	8	oct
4	quad	9	enn

Un un oct ium

118

Figure 5-2. Prefixes used in systematic names: To name elements with atomic numbers greater than 108, use the prefix for each digit, then add *-ium* to the end of the third prefix. For example, Element 118 would have the systematic symbol of Ununoctium.

Other Information about the Elements

In the boxes on the periodic table in this book, the atomic number is located below and to the left of the symbol. Below the atomic number is the electron configuration showing how the electrons are arranged according to their energy levels. The atomic mass is above the symbol and to the left. Selected oxidation states are in the upper right-hand corner of the box.

Often, a periodic table will include the name of the element and its state at room temperature. Not all tables include electron configuration and common oxidation states. The information provided by a periodic table will often depend on what information is needed. A periodic table used to

MemoryJogger

The decimal atomic masses given for each element on the periodic table are the weighted average of the masses of the isotopes of that element.

determine bonding types, for example, might just include the symbols and electronegativities of the elements.

Arrangement of the Periodic Table

You now know that the elements on the periodic table are listed according to increasing atomic number. But how does this arrangement allow periodic properties to be seen? As you will learn, the columns and rows of the table have special significance.

Periods

The horizontal rows of the table are called **periods.** The number at the beginning of the period indicates the principal energy level in which the valence electrons are located for the atoms of that period. Potassium (K) and bromine (Br) are members of Period 4, so they have valence electrons in the fourth principal energy level.

In each period, the number of valence electrons increases from left to right, and the properties of the elements change systematically across a period. For example, the elements on the left side of the table have common properties that are described in the next section; these elements are called **metals.** Metals comprise about 75% of all the elements. To the right of the middle of the table are elements called **metalloids,** which have some properties of both metals and nonmetals. On most periodic tables, the metalloids are located adjacent to a diagonal, stair-step line. To the right of

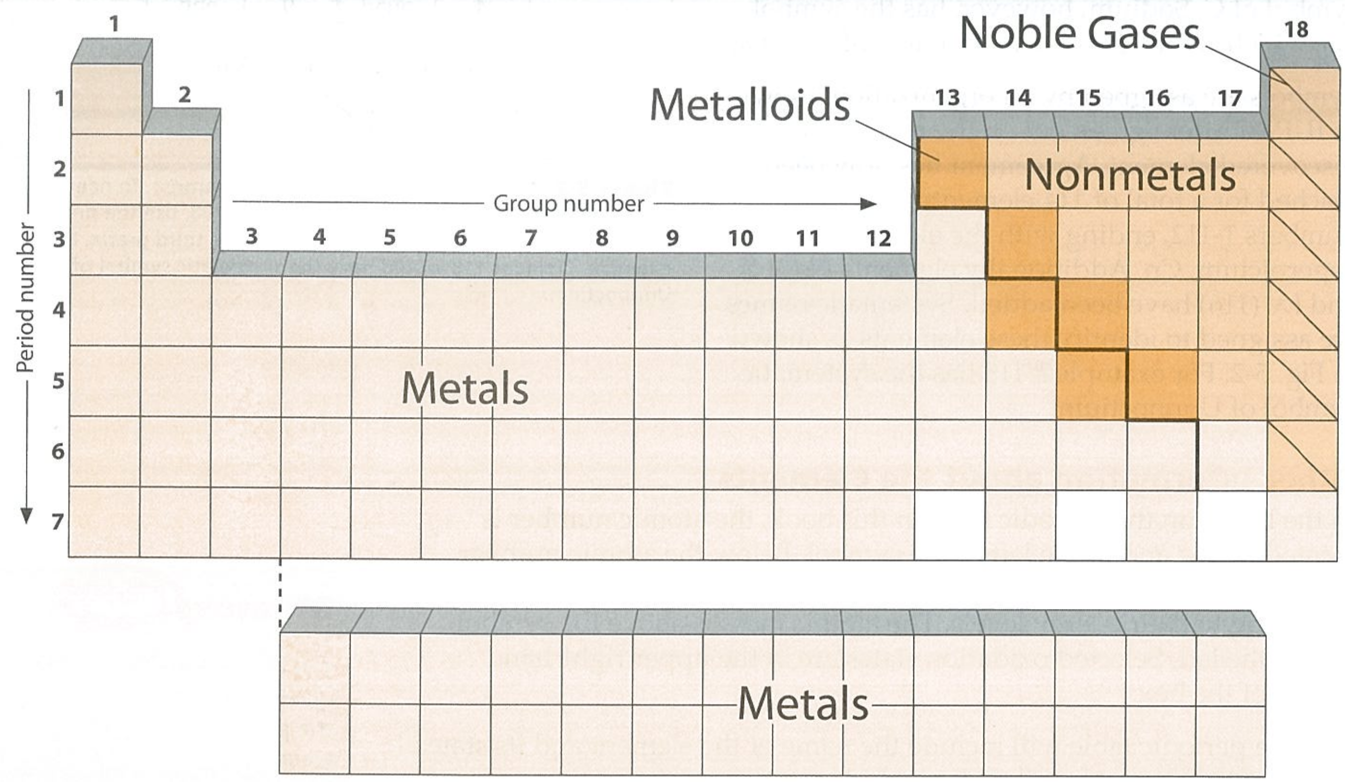


Figure 5-3. A trend from left to right across a period: Metals are found on the left of the periodic table; metalloids, on the staircase; and nonmetals and noble gases on the right side.

the metalloids, each period contains one or more elements with properties also described in the next section; these elements are known as **nonmetals**. Each period of the table ends with a noble gas. The metals and metalloids are all solid with the exception of mercury, a liquid. Nonmetals include solids, a liquid (bromine), and gases. Figure 5-3 summarizes this trend on the periodic table.

Groups, or Families

The vertical columns of the periodic table are called **groups,** or **families.** With a few exceptions, each member of a given group contains the same number of valence electrons. The number of valence electrons for each element is shown as the last number in the electron configuration. Phosphorus, in Group 15, has an electron configuration of 2-8-5. Thus, phosphorus has five valence electrons. All the members of Group 15 have five valence electrons. The elements in Group 18 have eight valence electrons. Helium, in Group 18, is an exception, having only two valence electrons. Because it is the number of valence electrons that determines much of the chemical reactivity of the element, the members of a given group have similar chemical properties.

Types of Elements

You've just seen that most elements are metals, and some are metalloids or nonmetals. An element can be classified as one of these types according to where it is located on the periodic table. The properties of each type of element are quite important in determining how it can be used.

Metals

Remember that most known elements are metals. The most active metals are located in Groups 1 and 2. In any group of the periodic table, the metallic properties of the elements increase from the top to the bottom of the group.

General Properties of Metals

- Metals are solids at room temperature, with the exception of mercury, which is a liquid.
- Most of the metals have densities greater than water, but the alkali metals (Group 1) will float.
- Metals are malleable, which means that they can be hammered into a shape.
- Metals are ductile, which means that they can be drawn or pulled into a wire.
- Metals have luster, which means that they are shiny.
- Metals are good conductors of heat and electricity. This property stems from the mobility of their valence electrons.
- Metals have relatively low ionization energy and electronegativity values.
- Metals tend to lose electrons to form positive ions with smaller radii. (See Figure 5-4A.)

Transition Elements The elements of Groups 3 through 12 are called the transition elements, or sometimes the transition metals. The transition elements in each period represent a series of elements in which the outermost *d* orbitals are being filled. For example, in Period 4, the transition elements proceed from scandium to zinc. Going from left to right across these elements, the 3*d* orbitals are being filled.

Transition elements are typically hard solids with high melting points, with the notable exception of mercury.

Transition elements are characterized by multiple oxidation states. When the transition elements react, they may use electrons from both *s* and *d* sublevels. The ionization energies of the *d*-orbital electrons have values close to those of their *s* electrons, and different numbers of electrons can be removed, resulting in different oxidation states. In general, the transition elements are far less reactive than the metals of Groups 1 and 2.

Another rather unique property of transition elements is that they often form ions that have color. The transition elements have several empty or half-filled *d* orbitals of nearly equal energy content. White light shining on these elements can excite electrons to slightly higher orbitals by the absorption of energy. When the electrons return to their ground state, they emit energy with the frequencies of visible colors.

Metalloids

The metalloids (B, Si, Ge, As, Sb, and Te), sometimes called semimetals, are sandwiched between the metals and the nonmetals. They can be found adjacent to the diagonal, stair-step line on the periodic table used in this book. Metalloids represent an intermediate type of element, displaying both metallic and nonmetallic properties.

Nonmetals

Although the properties of nonmetals vary more than those of metals, some standard properties can be observed.

General Properties of Nonmetals

- Many nonmetals are gases or molecular or network solids at room temperature. Bromine is an exception, being a liquid at room temperature.
- Nonmetals are not malleable or ductile; they tend to be brittle in the solid phase.
- Solid nonmetals lack luster, and their surface appears dull.
- Nonmetals have high ionization energy and high electronegativity values.
- Nonmetals are poor conductors of heat and electricity.
- Nonmetals tend to gain electrons to become negative ions with radii larger than their atoms. (See Figure 5-4B.)

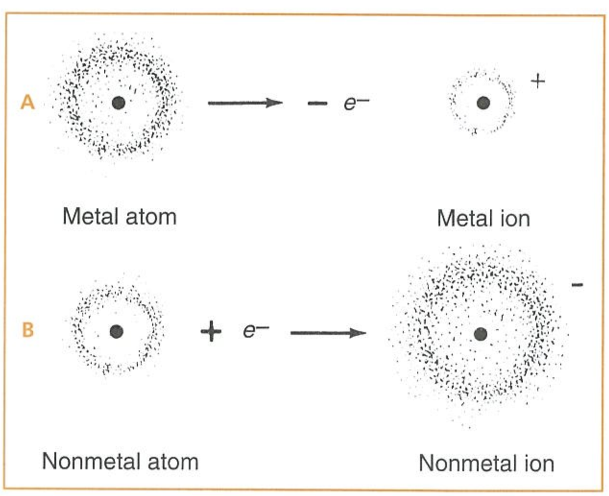


Figure 5-4. Ion formation: (A) Metals have three or fewer valence electrons. They tend to lose these electrons during chemical reactions, forming positive ions. (B) Nonmetals have four or more valence electrons. The more valence electrons a nonmetal has, up to seven, the more likely it is to gain electrons, forming a negative ion.

Noble Gases

The elements in Group 18 are called the **noble gases.** Noble gases don't have all the properties of nonmetals because they are generally unreactive. Only a few stable compounds containing noble gases have been formed.

Each of these elements has a completely filled outer energy level (valence level) of electrons. The valence level of helium is filled with only two electrons, whereas the rest of the noble gases have a complete octet (eight electrons), which is an extremely stable electron configuration.

Allotropes Some nonmetals can exist in two or more forms in the same phase. These forms are called allotropes. Oxygen ($O_2(g)$) and ozone ($O_3(g)$) are examples of allotropes. Table 5-1 compares some physical properties of oxygen and ozone.

Table 5-1. Properties of Oxygen and Ozone

Property	Oxygen (O_2)	Ozone (O_3)
Molecular mass (amu)	32	48
Melting point (K)	54.3	80.5
Boiling point (K)	90.2	161.7
Density (g/cm^3)	1.14	1.61

The table shows that allotropes have different physical properties. They also differ chemically. Ozone is a much stronger oxidizing agent than molecular oxygen and can cause serious damage to organic molecules. In the upper layers of the atmosphere, ozone absorbs harmful ultraviolet rays preventing them from reaching ground level.

Other nonmetals also show allotropy. Carbon is found as graphite, diamond, and buckminsterfullerene, which has a formula of C_{60}. Phosphorus can be found as yellow (white), red, and black allotropes.

Review Questions — Set 5.1

1. The observed regularities in the properties of the elements are periodic functions of their
(1) atomic numbers
(2) mass numbers
(3) oxidation states
(4) nonvalence electrons

2. Elements in Mendeleev's periodic table were arranged according to their
(1) atomic number
(2) atomic mass
(3) relative activity
(4) relative size

3. Most of the groups in the periodic table of the elements contain
(1) nonmetals only
(2) metals only
(3) nonmetals and metals
(4) metals and metalloids

4. An element is a gas at room temperature. It could be
(1) a metal or a metalloid
(2) a metal or a nonmetal
(3) a metalloid or a nonmetal
(4) a nonmetal only

5. Atoms of metals tend to
   (1) lose electrons and form negative ions
   (2) lose electrons and form positive ions
   (3) gain electrons and form negative ions
   (4) gain electrons and form positive ions

6. A solid element that is malleable, a good conductor of electricity, and reacts with oxygen is classified as a
   (1) metalloid (3) noble gas
   (2) metal (4) nonmetal

7. When a metal atom combines with a nonmetal atom, the nonmetal atom will
   (1) lose electrons and decrease in size
   (2) lose electrons and increase in size
   (3) gain electrons and decrease in size
   (4) gain electrons and increase in size

8. A Mg atom differs from a Mg^{2+} ion in that the atom has a
   (1) smaller radius (3) smaller nucleus
   (2) larger radius (4) larger nucleus

9. Which of the following elements has an ionic radius smaller than its atomic radius?
   (1) neon (3) sodium
   (2) nitrogen (4) sulfur

10. When a potassium atom reacts with a bromine atom, the potassium atom will
    (1) lose only 1 electron (3) gain only 1 electron
    (2) lose 2 electrons (4) gain 2 electrons

11. At room temperature, potassium is classified as
    (1) a metallic solid (3) a network solid
    (2) a molecular solid (4) an ionic solid

12. At room temperature, which substance is the best conductor of electricity?
    (1) nitrogen (3) sulfur
    (2) neon (4) silver

13. The element arsenic has the properties of
    (1) metals only
    (2) nonmetals only
    (3) both metals and nonmetals
    (4) neither metals nor nonmetals

14. Which list of elements contains two metalloids?
    (1) Ga, Ge, Sn (3) C, Si, Ge
    (2) Si, P, S (4) B, C, N

15. Which set of elements contains a metalloid?
    (1) K, Mn, As, Ar (3) Ba, Ag, Sn, Xe
    (2) Li, Mg, Ca, Kr (4) Fr, F, O, Rn

16. On the periodic table, an element classified as a metalloid can be found in
    (1) Period 6, Group 5 (3) Period 3, Group 16
    (2) Period 2, Group 14 (4) Period 4, Group 15

17. Which element in Period 4 is classified as an active nonmetal?
    (1) Ga (3) Br
    (2) Ge (4) Kr

18. A property of most nonmetals in the solid state is that they are
    (1) good conductors of heat
    (2) good conductors of electricity
    (3) brittle
    (4) malleable

19. Which properties are characteristic of nonmetals?
    (1) low thermal conductivity and low electrical conductivity
    (2) low thermal conductivity and high electrical conductivity
    (3) high thermal conductivity and low electrical conductivity
    (4) high thermal conductivity and high electrical conductivity

20. Which element in Period 2 of the periodic table is the most reactive nonmetal?
    (1) carbon (3) oxygen
    (2) nitrogen (4) fluorine

21. Which element is brittle in the solid phase and is a poor conductor of heat and electricity?
    (1) calcium (3) sulfur
    (2) strontium (4) copper

22. Which element in Period 4 is classified as an active metal?
    (1) Ca (3) Br
    (2) V (4) Ge

23. The presence of which ion usually produces a colored solution?
    (1) K^+ (3) Fe^{2+}
    (2) F^- (4) S^{2-}

24. Which set of properties is most characteristic of transition elements?
    (1) colorless ions in solution, multiple positive oxidation states
    (2) colorless ions in solution, multiple negative oxidation states
    (3) colored ions in solution, multiple positive oxidation states
    (4) colored ions in solution, multiple negative oxidation states

25. Which salt contains an ion that forms a colored solution?

(1) $Mg(NO_3)_2$ (3) $Ni(NO_3)_3$
(2) $Ca(NO_3)_2$ (4) $Al(NO_3)_3$

26. Which group in the periodic table contains an element that can form a blue sulfate compound?

(1) 1 (3) 11
(2) 2 (4) 17

27. Aqueous solutions of compounds containing element X are blue. Element X could be

(1) carbon (3) sodium
(2) copper (4) potassium

28. Pure silicon is chemically classified as a metalloid because silicon

(1) is malleable and ductile
(2) is an excellent conductor of heat and electricity
(3) exhibits hydrogen bonding
(4) exhibits metallic and nonmetallic properties

29. Which compound forms a colored aqueous solution?

(1) $CaCl_2$
(2) $CrCl_3$
(3) $NaOH$
(4) KBr

Properties of Elements

Some periodic properties of elements, such as metallic character, already have been discussed. However, many other properties of elements can be predicted based on the period or group the element belongs to.

Ionization Energy

The amount of energy needed to remove the most loosely bound electron from a neutral gaseous atom is called the **ionization energy** of the element.

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

Atoms with more than one electron have more than one ionization energy, but this first ionization energy is most significant.

Trends in a Period Figure 5-5 illustrates the periodic function of ionization energy. Values from left to right across a period generally increase. Notice that the same pattern is seen in the two periods shown. As the atoms are considered from left to right of the table, an increase in the number of protons is revealed. As the nuclear charge increases, the electrons are more strongly attracted, and hence more energy is needed to remove them from the atom.

Trends in a Group Figure 5-6 shows the ionization energy from the top to the bottom of Group 2. The ionization energy decreases because valence electrons in each successive element are at a higher energy level and, thus, farther from the nucleus. It is easier to remove them when they are farther from the positive charge of the nucleus and shielded from it by other levels of electrons.

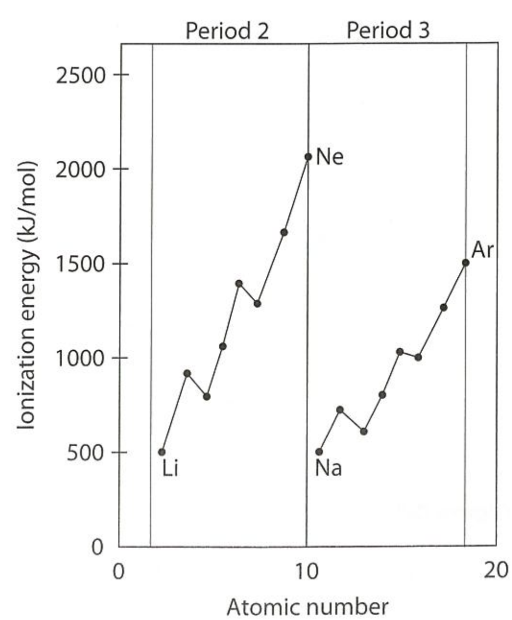


Figure 5-5. Ionization energy of Periods 2 and 3: In general, ionization energy increases from left to right across a period. The peaks in the lines result from the increased stability of filled and half-filled sublevels of electrons.

Atomic Radii

The radius of an atom is a good measure of the size of the atom. The **atomic radius** is defined as half the distance between two adjacent atoms in a crystal or half the distance between nuclei of identical atoms that are bonded together. Several methods can be used to determine the atomic radius of an element. No matter how the radius is measured, certain trends remain constant.

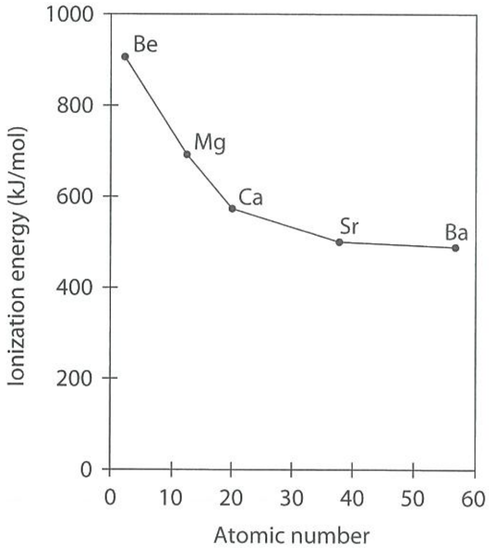


Figure 5-6. Ionization energy of Group 2: Ionization energy is less in larger atoms that have the same valence electron configuration as smaller atoms.

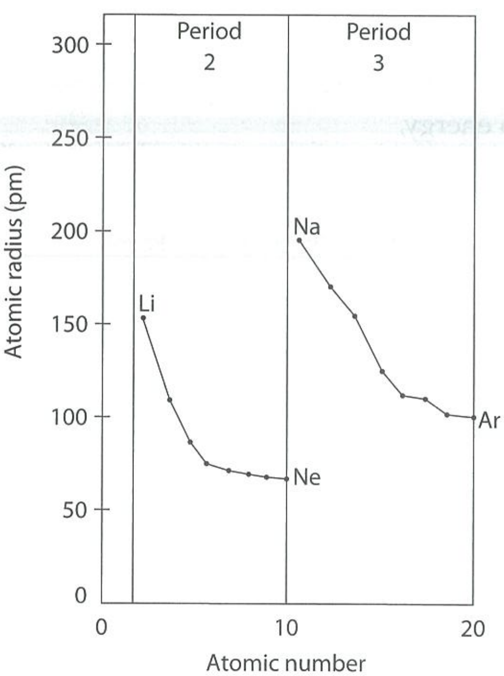


Figure 5-7. Atomic radii for Periods 2 and 3: From left to right, atomic radius decreases across a period.

Trends in a Period From left to right in a period, there is a repeating pattern of decreasing atomic radii. See Figure 5.7. In each period, metals have larger radii than nonmetals. The valence electrons of all the members of a period are at the same general energy level, but the number of protons in the nucleus attracting them increases, causing the radii to decrease.

Trends in a Group In any given group, each successive member has more inner-level electrons than the preceding member. These electrons shield the valence electrons from the nucleus, reducing the attractive force of the nucleus. Additionally, the valence electrons are more energetic than those of elements higher in the group. Thus, the atomic radius increases from top to bottom of a group. The atomic radii are shown in Table S of Appendix 1, *Reference Tables for Physical Setting/ Chemistry*. Atomic radii are often measured in picometers (10^{-12} m). R

Other Properties and Trends

You can easily see from the properties discussed so far that properties of elements follow predictable trends. Other properties of the elements also follow patterns.

Ionic Radii Atoms gain or lose electrons to become charged particles called ions. Metals tend to lose their valence electrons in chemical reactions to become positive ions. Nonmetals tend to gain electrons in chemical reactions and become negative ions. As these atoms gain or lose electrons, they complete an octet of valence electrons.

One way to compare an ion to the original atom is to compare the atomic radius to the ionic radius. **Ionic radius** is the distance from the nucleus to the outer energy level of the ion. As you might expect, when an atom loses its valence electrons and becomes positively charged it loses an energy level and its radius decreases. Therefore, the radius of a metallic ion is smaller than the radius of the atom. When a nonmetal gains electrons, the stable arrangement results in an increase in the radius. Thus, the radii of nonmetallic ions are greater than those of their atoms, as shown in Figure 5-4.

Electronegativity The **electronegativity** value of an atom is a measure of its attraction for electrons when bonded to another atom. Cesium, with an electronegativity value of 0.7, has the lowest attraction for bonded electrons. Fluorine, with a value of 4.0, has the highest. Electronegativity values are listed in Table S of Appendix 1, *Reference Tables for Physical Setting/ Chemisty*. R

Electronegativity values can be used to predict the type of bond that will be formed between two atoms. The noble gases are not generally assigned electronegativity values because they form very few stable chemical compounds.

There is a regular increase in the electronegativity value of each element when a period is considered from left to right. Metals tend to have low values, while the nonmetals of each period are higher.

In each group, the highest electronegativity value is found at the top. Attraction for bonded electrons is less toward the bottom of the group.

Reactivity of Elements Some elements can be found uncombined in nature, that is, in the atomic state. Oxygen may be simply O_2. The noble gases are always found as free elements, as they rarely react. Other elements are so reactive that they are never found in the uncombined, or free, state. Such is the case for the elements of Groups 1, 2, and 17. Some of these elements can only be obtained by the electrolysis of their fused salts.

Properties of Groups

The elements in any group of the periodic table have related chemical properties. Inspection of a group shows that all the members have the same number of valence electrons, and it is this similarity that accounts for the similarity in chemical properties. There is a regularity that can be seen quite clearly in the manner in which the members of a group react with elements of a different group.

Although each member of a group has the same number of valence electrons, there can be a change in the type of element from top to bottom. In Group 14, the top member, carbon, is clearly a nonmetal. Silicon and germanium are metalloids, while tin and lead are metals. As each group is considered from top to bottom, the metallic characteristics increase.

Examine the chemical properties of each group of the representative elements. Notice how these properties are based on the arrangements of the electrons in each group.

Hydrogen

The vertical columns of the periodic table have been previously identified as "groups." Hydrogen is unique among all of the elements as the only element not belonging to a group. Although it seems to be placed in Group 1, notice that it is separated from the rest of the group. Hydrogen does not have physical or chemical properties similar to the Group 1 metals.

In a chemical reaction, hydrogen often loses its single valence electron to become H^+. Hydrogen also can combine with metals to form metallic hydrides, such as NaH. In these compounds the metals have a positive oxidation state, while the hydrogen is assigned 1− H^-. Hydrogen will also share electrons with another atom to gain stability. An example of hydrogen and oxygen sharing electrons can be seen in water (H_2O).

Groups 1 and 2

The elements of Group 1 (Figure 5-8), the alkali metals, and of Group 2 (Figure 5-9), the alkaline earth metals, show typical metallic characteristics. The members of both these groups easily lose their electrons and are never found in nature in their atomic state. That is, they are always found in

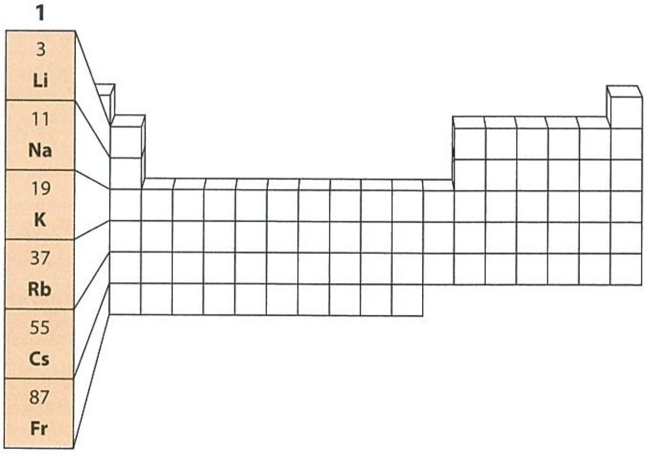


Figure 5-8. **Group 1, the alkali metals.**

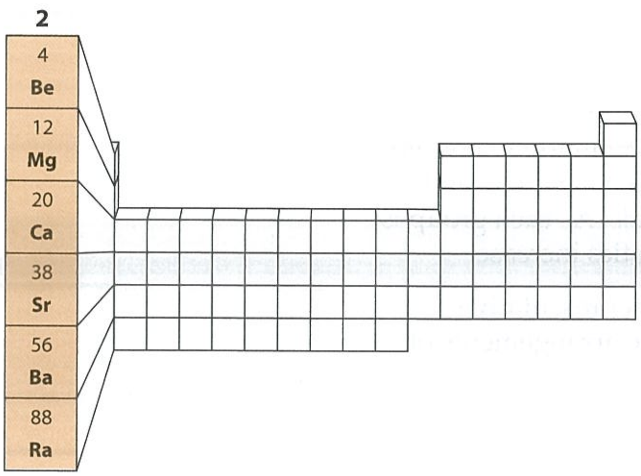


Figure 5-9. **Group 2, the alkaline earth metals.**

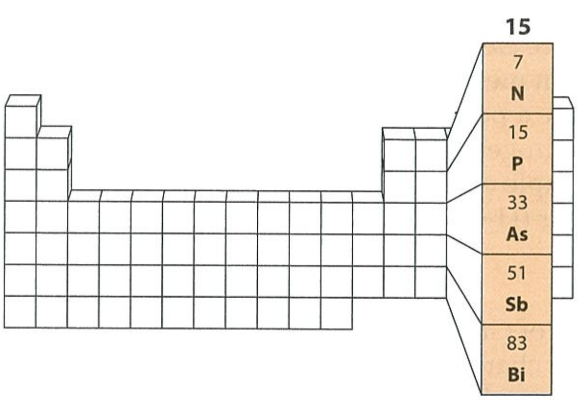


Figure 5-10. **Elements in Group 15.**

compounds. They can be reduced to their free states by electrolysis of their compounds. All of these elements have low ionization energies and electronegativity values. They typically achieve stable octets by losing electrons to form ionic bonds.

In general, from top to bottom in both groups, reactivity increases. It should be noted that each Group 1 element is more reactive than the Group 2 element of the same period. Francium is the most reactive metal.

Group 15

The members of Group 15 (Figure 5-10) show the change from nonmetallic to metallic properties from top to bottom of the group. Nitrogen and phosphorus are typical nonmetals and can accept three electrons to form 3− anions. Bismuth loses electrons to become a positive cation, which is typical of metals. Arsenic and antimony are generally classified as metalloids.

Nitrogen is a stable gas at room temperature, largely because it contains a triple bond between the two nitrogen atoms in N_2. Phosphorus forms P_4, which is far more reactive than N_2. Nitrogen is an essential component of protein synthesis and plant processes. Because atmospheric N_2 is so unreactive, it must be changed to a usable form. One way this is done is by nitrogen-fixing bacteria that release usable nitrogen compounds into the soil.

Group 16

The elements in Group 16 (Figure 5-11) also show the expected progression from nonmetal to metal with increasing atomic number. Oxygen and sulfur are typical nonmetals, while the last member, polonium, is a metal. Selenium and tellurium are metalloids.

Oxygen may well be the most important element in the group. It is a diatomic molecule at room temperature, sharing two pairs of electrons between the two atoms. Oxygen is a reactive element, easily forming compounds with other elements. Were it not for the process of photosynthesis, all of the oxygen on Earth would be in the combined state.

In a compound, oxygen normally has an oxidation value of 2−. Its negative oxidation value results from its high electronegativity as it attracts shared electrons in a bond. Only when oxygen bonds to the only element more electronegative than itself (fluorine) does oxygen have a positive oxidation value.

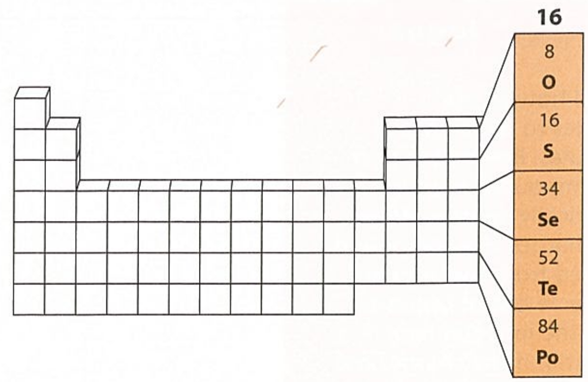


Figure 5-11. **Elements in Group 16.**

Group 17

The members of Group 17 (Figure 5-12) are also known as the halogens when they are in the free state. When atoms of elements in this group gain an electron, they become an ion with a 1– charge, and the salts formed are called halides. All the members of the group are nonmetals, but the normal trend of nonmetallic to metallic characteristics is seen even in this group. Iodine, the heaviest nonradioactive member, is a solid with some luster characteristic of metals.

The halogens are the only group of the table containing all three states of matter at room conditions. Fluorine and chlorine are gases; bromine is a liquid; and iodine, a solid. Astatine is a radioactive element with no practical uses. The half-life of At-210 is only 8.3 hours. Therefore, there are no sizeable amounts of astatine in nature.

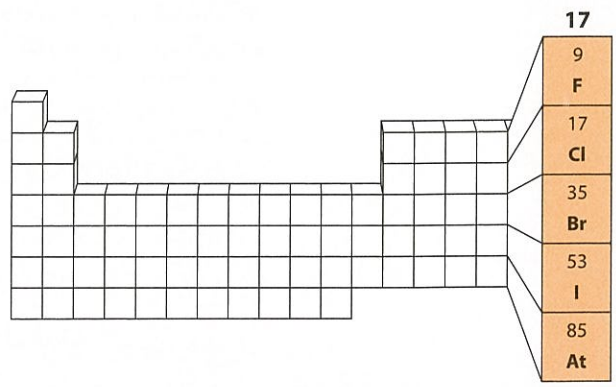


Figure 5-12. **The halogens (Group 17).**

As is the case with other nonpolar molecules, the halogens are held in the solid and liquid phases by weak van der Waals forces. Because these forces are the result of temporary dipoles caused by the movement of electrons, it follows that the more electrons a molecule possesses, the higher its melting and boiling points.

Because of their high reactivity, the halogens occur in nature only in their combined form. Fluorine, the most reactive halogen, can only be prepared from its fluoride compounds by electrolysis. The other halogens can be prepared by chemical means.

Group 18

The elements of Group 18 (Figure 5-13) are called the noble gases. They do not combine to form diatomic molecules, but rather exist as monatomic molecules. Each of these atoms has a complete outer energy level of electrons and therefore is chemically unreactive. The outer energy level of helium has two electrons, while each of the remaining noble gases has eight.

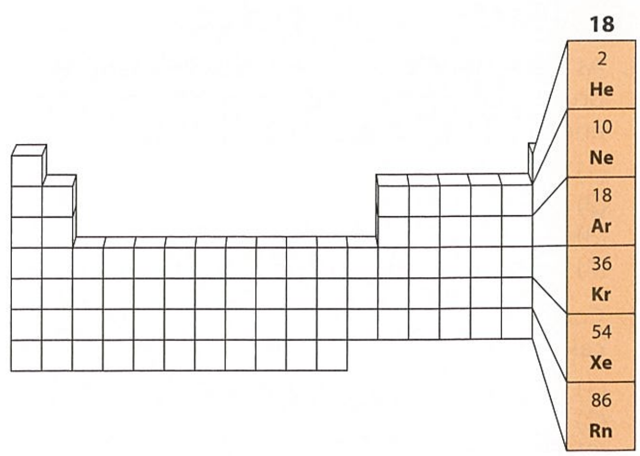


Figure 5-13. **The noble gases (Group 18).**

Memory Jogger

Recall how to use charges to write correct formulas. Remember that all compounds must be electrically neutral. When the charges in a compound formed from two ions are the same but of opposite sign, the ions combine in a 1:1 ratio. If the charges are different, write the charge of one as the subscript of the other to obtain the correct formula.

This group was once called the inert gas group because it was thought that the elements could not react to produce stable compounds. Because their valence levels are full, the noble gases normally would not add additional electrons. However, fluorine has such a high attraction for electrons that it can attract electrons from some of the noble gases, causing them to have a positive oxidation state. Thus, some reactivity is observed. Only argon and xenon have produced stable compounds, although even neon has been coaxed into reactions under severe conditions of temperature and pressure. Helium alone has not shown any reactivity at all.

Like the halogen group, the noble gases are nonpolar and are therefore held in the liquid and solid phases by van der Waals forces. As expected, helium, with the fewest electrons, has the lowest boiling point, and the boiling point increases from top to bottom in the group.

Reactions Between Groups

Consider the reactions of Group 1 (the alkali metals) with the members of Group 17 (the halogens). The members of Group 1 each have one valence electron to lose and form ions with a 1+ charge. Group 17 elements each have seven valence electrons and can accept one electron from the alkali metal, forming an ion with a 1− charge. The general formula for the combining of a Group 1 element with a Group 17 element is MX, where M represents the Group 1 element, and X represents the Group 17 element. An example is KI.

Group 2 elements have two valence electrons and form compounds with Group 17 elements having the general formula MX_2. Examples include $BeCl_2$, $MgBr_2$, and CaF_2.

Review Questions — Set 5.2

30. Which value represents the first ionization energy of a nonmetal?

(1) 497.9 kJ
(2) 577.4 kJ
(3) 811.7 kJ
(4) 1000. kJ

31. As the elements of Group 1 are considered in order from top to bottom, the first ionization energy of each successive element will

(1) decrease
(2) increase
(3) remain the same
(4) follow an unpredictable pattern

32. Compared to an atom of potassium, an atom of calcium has a

(1) larger radius and lower reactivity
(2) larger radius and higher reactivity
(3) smaller radius and lower reactivity

33. Which statement describes the elements in Period 3?

(1) Each successive element has a greater atomic radius.
(2) Each successive element has a lower electronegativity.
(3) All elements have similar chemical properties.
(4) All elements have valence electrons in the same principal energy level.

34. As atoms of elements in Group 16 are considered in order from top to bottom, the electronegativity of each successive element

(1) decreases
(2) increases
(3) remains the same

35. A nonmetal could have an electronegativity of

(1) 1.0
(2) 2.0
(3) 1.6
(4) 2.6

36. Which properties are most common in nonmetals?

(1) low ionization energy and low electronegativity
(2) low ionization energy and high electronegativity
(3) high ionization energy and low electronegativity
(4) high ionization energy and high electronegativity

37. A diatomic element with a high first ionization energy would most likely be a

(1) nonmetal with a high electronegativity
(2) nonmetal with a low electronegativity
(3) metal with a high electronegativity
(4) metal with a low electronegativity

38. Which element at room temperature is a poor conductor of electricity and has a relatively high electronegativity?

(1) Cu (3) Mg
(2) S (4) Fe

39. Within Period 2 of the periodic table, as the atomic number increases, the atomic radius generally

(1) decreases (3) remains the same
(2) increases (4) follows no pattern

40. Atoms of which element have the smallest radius?

(1) Si (3) S
(2) P (4) Cl

41. Which statement best compares the atomic radius of a potassium atom and the atomic radius of a calcium atom?

(1) The radius of the potassium atom is smaller because of its smaller nuclear charge.
(2) The radius of the potassium atom is smaller because of its larger nuclear charge.
(3) The radius of the potassium atom is larger because of its smaller nuclear charge.
(4) The radius of the potassium atom is larger because of its larger nuclear charge.

42. According to the reference table, which of the following elements has the smallest radius?

(1) nickel (3) calcium
(2) cobalt (4) potassium

43. In which area of the periodic table are the elements with the strongest nonmetallic properties located?

(1) lower left (3) lower right
(2) upper left (4) upper right

44. At which location in the periodic table would the most active metallic element be found?

(1) in Group 1 at the top
(2) in Group 1 at the bottom
(3) in Group 17 at the top
(4) in Group 17 at the bottom

45. In which section of the periodic table are the most active nonmetals located?

(1) upper right corner (3) upper left corner
(2) lower right corner (4) lower left corner

46. What is the total number of elements in Group 17 that are gases at room temperature and standard pressure?

(1) 1 (3) 3
(2) 2 (4) 4

47. Which of the following groups in the periodic table contain elements so reactive that they are never found in the free state?

(1) 1 and 2 (3) 2 and 15
(2) 1 and 11 (4) 11 and 15

48. Which halogen can only be prepared from its fused compounds?

(1) I_2 (3) Br_2
(2) Cl_2 (4) F_2

49. As the elements in Group 15 are considered in order of increasing atomic number, which sequence in properties occurs?

(1) nonmetal→metalloid→metal
(2) metalloid→metal→nonmetal
(3) metal→metalloid→nonmetal
(4) metal→nonmetal→metalloid

50. Which elements have the most similar chemical properties?

(1) K and Na (3) K and Ca
(2) K and Cl (4) K and S

51. Because of its high reactivity, which element is normally obtained by the electrolysis of its fused salts?

(1) sulfur (3) argon
(2) lithium (4) gold

52. Which element in Group 17 is the most active nonmetal?

(1) Br (3) Cl
(2) I (4) F

53. Which of the following Group 15 elements has the most nonmetallic properties?

(1) Bi (3) Sb
(2) P (4) N

54. The elements calcium and strontium have similar chemical properties because they both have the same

(1) atomic number
(2) mass number
(3) number of valence electrons
(4) number of completely filled subshells

55. Which Group 15 element exists as a diatomic molecule at room temperature and pressure?

(1) phosphorus (3) bismuth
(2) nitrogen (4) arsenic

56. Which two elements have chemical properties that are most similar?

(1) Cl and Ar (3) K and Ca
(2) Li and Na (4) C and N

57. In which set do the elements exhibit the most similar chemical properties?

(1) N, O, and F (3) Li, Na, and K
(2) Hg, Br, and Rn (4) Al, Si, and P

58. Which of these metals loses electrons most readily?

(1) calcium (3) potassium
(2) magnesium (4) sodium

59. If M represents an element in Group 2, the formula of its chloride would be

(1) MCl (3) M_2Cl
(2) MCl_2 (4) M_2Cl_2

60. Which group below contains elements with the greatest variation in chemical properties?

(1) Li, Be, B (3) B, Al, Ga
(2) Li, Na, K (4) Be, Mg, Ca

61. Which element in Group 15 would most likely have luster and good electrical conductivity?

(1) N (3) Bi
(2) P (4) As

62. Which statement best describes the Group 2 metals?

(1) They have one valence electron, and they form ions with a 1+ charge.
(2) They have one valence electron, and they form ions with a 1− charge.
(3) They have two valence electrons, and they form ions with a 2+ charge.
(4) They have two valence electrons, and they form ions with a 2− charge.

63. Which element in Group 15 has the strongest metallic character?

(1) Bi (3) P
(2) As (4) N

64. Which halogens are gases at room temperature and pressure?

(1) chlorine and fluorine
(2) chlorine and bromine
(3) iodine and fluorine
(4) iodine and bromine

65. In which group of elements do the atoms gain electrons most readily?

(1) 1 (3) 16
(2) 2 (4) 18

66. Which element is more reactive than strontium?

(1) potassium (3) iron
(2) calcium (4) copper

67. The oxide of metal X has the formula XO. Which group in the periodic table contains metal X?

(1) Group 1
(2) Group 2
(3) Group 13
(4) Group 17

68. As elements in a group of the periodic table are considered in order from top to bottom, the metallic character of each successive element generally

(1) decreases
(2) increases
(3) remains the same

69. As the elements in Period 3 are considered from left to right, they tend to

(1) lose electrons more readily and increase in metallic character
(2) lose electrons more readily and increase in nonmetallic character
(3) gain electrons more readily and increase in metallic character
(4) gain electrons more readily and increase in nonmetallic character

70. An atom of which element in the ground state has a complete outermost energy level?

(1) He (3) Hg
(2) Be (4) H

71. Which of the following Group 15 elements has the most metallic properties?

(1) Bi (3) Sb
(2) P (4) N

72. As the atoms of the elements in Group 1 of the periodic table are considered from top to bottom, the number of valence electrons in the atoms of each successive element

(1) decreases
(2) increases
(3) remains the same

73. Which element attains the structure of a noble gas when it becomes a 1^+ ion?

(1) K
(2) Ca
(3) F
(4) Ne

74. According to the reference table, which sequence correctly places the elements in order of increasing ionization energy?

(1) $H \rightarrow Li \rightarrow Na \rightarrow K$
(2) $I \rightarrow Br \rightarrow Cl \rightarrow F$
(3) $O \rightarrow S \rightarrow Se \rightarrow Te$
(4) $H \rightarrow Be \rightarrow Al \rightarrow Ga$

Answer the following questions, using complete sentences when appropriate.

75. The ionization energy of Na is 118 kcal, while the ionization energy of Mg to form Mg^+ is 175. While this is the expected result for an adjacent element, the ionization of Na^+ to Na^{2+} is 1091 kcal, while only 345 kcal is needed to ionize Mg^+ to Mg^{2+}. Explain.

76. The Na^+ ion has a smaller radius than the Ne atom. Explain why this is so.

77. Chlorine reacts with water according to the equation:

$$Cl_2 + H_2O \rightarrow HClO + H^+ + Cl^-$$

Write an equation to show how iodine reacts with water.

78. Why is the periodic table organized by increasing atomic number rather than by increasing atomic mass, as had been suggested by Mendeleev?

79. Explain why the members of Group 1 react by losing an electron, but the members of Group 17 react by gaining an electron.

80. Let the letter M stand for a member of Group 13. What is the formula of the combination of this element with bromine? With oxygen? Explain.

81. Why is hydrogen not considered to be a member of Group 1 (the alkali metals)?

82. Consider atoms of the elements boron, carbon and aluminum. Which is the largest? The smallest? Which has the highest ionization energy? The lowest?

83. Why do elements in a given group of the periodic table show similar chemical properties?

84. Why do the elements of a given period always follow the progression of metal to metalloid to nonmetal to noble gas?

85. What change in charge takes place as a metal loses one or more electrons?

86. Using the law of octets, explain why it is unlikely for sodium to form a Na^{2+} ion.

87. What would be the general formula of a Group 2 element (represented by M) combined with chlorine of Group 17?

88. What would be the general formula of a Group 17 element (represented by X) combined with magnesium of Group 2?

89. What would be the general formula of a Group 1 element combined with an oxygen of Group 16?

90. Mendeleev arranged the elements on the table in order of increasing atomic mass, but it was later learned that the correct order is by increasing atomic number. Study the periodic table and locate examples where the atomic number of two successive elements increases, but the atomic mass decreases.

Practice Questions
for the New York Regents Exam

Directions
Review the Test-Taking Strategies section of this book. Then answer the following questions. Read each question carefully and answer with a correct choice or response.

Part A

1 The elements in the modern periodic table are arranged according to their
(1) atomic number
(2) oxidation number
(3) atomic mass
(4) nuclear mass

2 Which characteristic describes most nonmetals in the solid phase?
(1) good conductors of electricity
(2) good conductors of heat
(3) malleable
(4) brittle

3 When combining with nonmetallic atoms, metallic atoms generally will
(1) lose electrons and form negative ions
(2) lose electrons and form positive ions
(3) gain electrons and form negative ions
(4) gain electrons and form positive ions

4 At STP, both diamond and graphite are solids composed of carbon atoms. These solids have
(1) the same structure and the same properties.
(2) the same structure and different properties.
(3) different structures and the same properties.
(4) different structures and different properties.

5 Which characteristic describes most metals in the solid phase?
(1) good conductors of electricity
(2) poor conductors of heat
(3) dull
(4) brittle

6 Which element is a nonmetallic liquid at room temperature?
(1) hydrogen (3) mercury
(2) oxygen (4) bromine

7 Which physical characteristic of a solution may indicate the presence of a transition element?
(1) its density
(2) its color
(3) its effect on litmus
(4) its effect on phenolphthalein

8 Compared to atoms of metals, atoms of nonmetals generally
(1) have higher electronegativity values
(2) have lower first ionization energies
(3) conduct electricity more readily
(4) lose electrons more readily

9 Compared to the atoms of nonmetals in Period 3, the atoms of metals in Period 3 have
(1) fewer electron shells.
(2) more electron shells.
(3) fewer valence electrons.
(4) more valence electrons.

10 Which part of the periodic table contains elements with the strongest metallic properties?
(1) upper right
(2) upper left
(3) lower right
(4) lower left

11 Which Group 17 element is a solid at room temperature and pressure?
(1) Br_2 (3) Cl_2
(2) F_2 (4) I_2

12 Which gas is monatomic at room temperature and pressure?
(1) nitrogen
(2) neon
(3) fluorine
(4) chlorine

13 Atoms of elements in a group of the periodic table have similar chemical properties. This similarity is most closely related to the atoms'
(1) number of principal energy levels
(2) number of valence electrons
(3) atomic numbers
(4) atomic masses

14 In which group are all the elements found naturally only in compounds?
(1) 18 (3) 11
(2) 2 (4) 14

Part B–1

15 An element is a solid at room temperature. It can be a
(1) metal only
(2) metalloid only
(3) metal or a nonmetal only
(4) metal, a metalloid, or a nonmetal

16 How does the size of a barium ion compare to the size of a barium atom?
(1) The ion is smaller because it has fewer electrons.
(2) The ion is smaller because it has more electrons.
(3) The ion is larger because it has fewer electrons.
(4) The ion is larger because it has more electrons.

17 Which element is malleable and ductile?
(1) S (3) Ge
(2) Si (4) Au

18 Which of the following Period 4 elements has the most metallic characteristics?
(1) Ca (3) As
(2) Ge (4) Br

19 Which three groups of the periodic table contain the most elements classified as metalloids?
(1) 1, 2, and 13
(2) 2, 13, and 14
(3) 14, 15, and 16
(4) 16, 17, and 18

20 An aqueous solution of XCl_2 contains colored ions. Element X could be
(1) Ba (3) Ni
(2) Ca (4) Bi

21 Which element has the highest first ionization energy?
(1) sodium
(2) aluminum
(3) calcium
(4) phosphorus

22 An element has a first ionization energy of 1314 kJ/mol and an electronegativity of 3.5. It is classified as a
(1) metal (3) metalloid
(2) nonmetal (4) halogen

23 Which sequence of elements is arranged in order of decreasing atomic radii?
(1) Al, Si, P
(2) Li, Na, K
(3) Cl, Br, I
(4) N, C, B

24 Which ion has the largest radius?
(1) Na^+ (3) K^+
(2) Mg^{2+} (4) Ca^{2+}

25 In the ground state, atoms of the elements in Group 15 of the periodic table all have the same number of
(1) filled principal energy levels
(2) occupied principal energy levels
(3) neutrons in the nucleus
(4) electrons in the valence energy level

26 Which of the following Group 18 elements would be most likely to form a compound with fluorine?
(1) He (3) Ar
(2) Ne (4) Kr

27 The properties of carbon are expected to be most similar to those of
(1) boron
(2) aluminum
(3) silicon
(4) phosphorus

28 If M represents an alkali metal of Group 1, what is the formula for the compound formed by M and oxygen?
(1) MO_2 (3) M_2O_3
(2) M_2O (4) M_3O_2

29 An atom of an element has 28 innermost electrons and 7 valence electrons. In which period of the periodic table is this element located?
(1) 5 (3) 3
(2) 2 (4) 4

Parts B–2 and C

30 Mendeleev arranged the periodic table in order of increasing atomic masses. Locate iodine and tellurium on the table and note that they are not arranged by increasing mass, and yet Mendeleev placed iodine in Group 17 and tellurium in Group 16. What is the likely reason that he did not arrange them by increasing mass?

Base your answers to questions 31 through 33 on the following information.

In the 19th century, Dmitri Mendeleev predicted the existence of a then unknown element *X* with a mass of 68. He also predicted that an oxide of *X* would have the formula X_2O_3.

31 Determine the number of valence electrons in a neutral atom of element *X*.

32 On the modern periodic table, what are the group number and period number of the element *X*?

33 Name an element in the same group as element *X* that would have an atomic radius smaller than that of element *X*.

Base your answers to questions 34 through 37 on the information below.

A metal, M, was obtained from a compound in a rock sample. Experiments have determined that the element is a member of Group 2 of the periodic table.

34 What is the phase of element *M* at STP?

35 Explain, in terms of electrons, why element *M* is a good conductor of electricity.

36 Explain why the radius of a positive ion of the element *M* is smaller than an atom of element *M*.

37 Using the symbol *M* for the element, write the chemical formula for the compound that forms when element *M* reacts with iodine.

Base your answers to questions 38 through 41 on the following information.

The atomic radius and the ionic radius for some Group 1 and some Group 17 elements are given in the tables below.

Atomic and Ionic Radii of Some Elements

Group 1

Particle	Radius (pm)
Li atom Li^+ ion	130. 78
Na atom Na^+ ion	160. 98
K atom K^+ ion	200. 133
Rb atom Rb^+ ion	215 148

Group 17

Particle	Radius (pm)
F atom F^- ion	60. 133
Cl atom Cl^- ion	100. 181
Br atom Br^- ion	117 ?
I atom I^- ion	136 220.

38 Estimate the radius of a Br^- ion.

39 Explain, in terms of electron shells, why the radius of a K^+ ion is greater than the radius of an Na^+ ion.

40 Write both the name and charge of the particle that is gained by an F atom when the atom becomes an F^- ion.

41 State the relationship between atomic number and first ionization energy as the elements in Group 1 are considered in order of increasing atomic number.

Base your answers to questions 42 through 44 on the following information.

The ionic radii of some Group 2 elements are given in the table below.

Ionic Radii of Some Group 2 Elements

Symbol	Atomic Number	Ionic Radius (pm)
Be	4	44
Mg	12	66
Ca	20	99
Ba	56	134

42 Estimate the ionic radius of the element strontium.

43 State the trend in ionic radius as the elements of Group 2 are considered in order of increasing atomic number.

44 Explain, in terms of electrons, why the ionic radius of a Group 2 element is smaller than its atomic radius.