Example 3.7

Write abbreviated electronic configurations for the following:

a. An atom that contains 7 electrons
b. An atom that contains 17 electrons
c. An atom of element number 22
d. An atom of arsenic (As)

Solution

All these configurations were written in conventional form in Example 3.6.

a. From Example 3.6, the configuration is $1s^22s^22p^3$. We see that the two electrons in the $1s$ subshell represent the noble gas configuration of helium (He), so we can write the configuration as $[He]2s^22p^3$.
b. From Example 3.6, the configuration is $1s^22s^22p^63s^23p^5$. We see that the $1s^22s^22p^6$ portion is the configuration of neon (Ne), so we can write the configuration as $[Ne]3s^23p^5$.
c. From Example 3.6, the configuration is $1s^22s^22p^63s^23p^64s^23d^2$. We see that the first 18 electrons represented by $1s^22s^22p^63s^23p^6$ correspond to the electrons of argon (Ar). Thus, we can write the configuration as $[Ar]4s^23d^2$.
d. From Example 3.6, the configuration of arsenic is $1s^22s^22p^63s^23p^64s^23d^{10}4p^3$. Once again, the first 18 electrons can be represented by the symbol for argon. The configuration is $[Ar]4s^23d^{10}4p^3$.

Learning Check 3.7 Write abbreviated electronic configurations for the following. These are the same elements used in Learning Check 3.6.

a. An atom of element number 9
b. An atom of magnesium (Mg)
c. An atom of the element found in group VIA(16) and period 3
d. An atom that contains 23 protons

3.5 Another Look at the Periodic Table

Learning Objective

5. Determine the shell and subshell locations of the distinguishing electrons in elements, and based on their location in the periodic table, classify elements into the categories given in Figures 3.10 and 3.12.

Now that you know more about electronic configurations, you can better understand the periodic law and table. First, consider the electronic configurations of the elements belonging to the same group of the periodic table. Group IA(1), for example, contains Li, Na, K, Rb, Cs, and Fr, with electronic configurations for the first four elements as shown below:

Element Symbol	Conventional Form	Abbreviated Form
Li	$1s^22s^1$	$[He]2s^1$
Na	$1s^22s^22p^63s^1$	$[Ne]3s^1$
K	$1s^22s^22p^63s^23p^64s^1$	$[Ar]4s^1$
Rb	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$	$[Kr]5s^1$

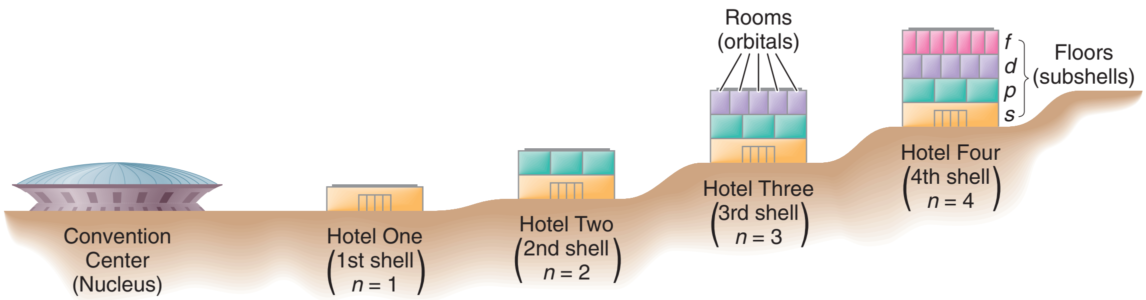
Study Skills 3.1 The Convention Hotels Analogy

A new concept is often made easier to understand by relating it to something familiar. The concept of electronic configurations is very likely new to you, but you are probably familiar with hotels. The way electrons fill up orbitals, subshells, and shells around a nucleus can be compared to the way rooms, floors, and hotels located near a convention center will fill with convention delegates. To make our analogy work, imagine that the hotels are located on a street that runs uphill from the convention center, as shown. Further imagine that none of the hotels has elevators, so the only way to get to upper floors is to climb the stairs.

In this analogy, the convention center is equivalent to the nucleus of an atom, and each hotel represents a shell, each floor represents a subshell, and each room represents an orbital. If you were a delegate who wanted to describe to a friend where you were staying, you would indicate the hotel (shell), floor (subshell), and room (orbital) assigned to you. Electronic configurations such as $1s^22s^22p^1$ give similar information for each electron: The numbers preceding each letter indicate the shells, the letters indicate the subshells, and the superscripts coupled with Hund's rule indicate which orbitals are occupied.

Three more assumptions allow us to extend the analogy: (1) No more than two delegates can be assigned to a room, (2) delegates prefer not to have roommates if an empty room on the same floor is available, and (3) all delegates want to use as little energy as possible when they walk from the convention center to their rooms (remember, there are no elevators).

With these assumptions in mind, it is obvious that the first delegate to check in will choose to stay in the single room of the *s* floor of Hotel One (a small but very exclusive hotel). The second delegate will also choose this same floor and room of Hotel One and will fill the hotel to capacity. The third delegate will have to go to Hotel Two and will choose the one room on the *s* floor. The fourth delegate will also choose the one room of the *s* floor, and fill that floor. The fifth delegate will choose a room on the *p* floor of Hotel Two because the *s* floor is full. The sixth delegate will also choose a room on the *p* floor of Hotel Two but will choose one that is not occupied. The seventh delegate will also choose an empty room on the *p* floor of Hotel Two. Delegates eight, nine, and ten can either pair up with the delegates already in the rooms of the *p* floor of Hotel Two or go uphill to Hotel Three. They choose to save energy by walking up the stairs and staying in rooms with roommates on the *p* floor of Hotel Two. Additional arriving delegates will occupy the floors of Hotels Three and Four in the order dictated by these same energy and pairing considerations. Thus, we see that the delegates in their rooms are analogous to the electrons in their orbitals.



Notice that each of these elements has a single electron in the valence shell. Further, each of these valence-shell electrons is located in an *s* subshell. Elements belonging to other groups also have valence-shell electronic configurations that are similar for all members of the group, and all members have similar chemical properties. It has been determined that the similar chemical properties of elements in the same group result from similar valence-shell electronic configurations.

The periodic table becomes more useful when we interpret it in terms of the electronic configurations of the elements in various areas. One relationship is shown in Figure 3.9, where the periodic table is divided into four areas on the basis of the type of subshell occupied by the highest-energy electron in the atom. This last electron added to an atom is called the **distinguishing electron.** Note that the *s* area is 2 columns (elements) wide, the *p* area is 6 columns wide, the *d* area is 10 columns wide, and the *f* area is 14 columns wide—exactly the number of electrons required to fill the *s, p, d,* and *f* subshells, respectively.

distinguishing electron The last and highest-energy electron found in an element.

Electronic configurations are also used to classify elements of the periodic table, as shown in Figure 3.10.

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 –	113 –	114 –	115 –	*p* Area		118 –

s Area (H, He; groups 1–2) — *d* Area — *p* Area

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

f Area

Figure 3.9 The periodic table divided by distinguishing electrons of the elements.

Representative elements — Transition elements — Representative elements — Noble gases

Period	(1) I A	(2) II A	(3) III B	(4) IV B	(5) V B	(6) VI B	(7) VII B	(8)	(9) VIII B	(10)	(11) I B	(12) II B	(13) III A	(14) IV A	(15) V A	(16) VI A	(17) VII A	(18) VIII A
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 –	113 –	114 –	115 –			118 –

Inner-transition elements														
	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Figure 3.10 Elemental classification on the basis of electronic configurations.

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Figure 3.11 Transition elements, such as gold, silver, copper, nickel, platinum, and zinc, are often used in coins and medals. List some chemical and physical properties that would be desirable in metals used for such purposes.

The *noble gases* make up the group of elements found on the extreme right of the periodic table. They are all gases at room temperature and are unreactive with most other substances (hence, the group name). With the exception of helium, the first member of the group, noble gases are characterized by filled *s* and *p* subshells of the highest occupied shell.

Representative elements are those found in the *s* and *p* areas of the periodic table, not including the noble gases. The distinguishing electrons of representative elements partially or completely fill an *s* subshell—groups IA(1) and IIA(2)—or partially fill a *p* subshell—groups IIIA(13), IVA(14), VA(15), VIA(16), and VIIA(17). Most of the common elements are representative elements.

representative element An element in which the distinguishing electron is found in an *s* or a *p* subshell.

The *d* area of the periodic table contains the **transition elements** (Figure 3.10) in which the distinguishing electron is found in a *d* subshell. Some transition elements are used for everyday applications (Figure 3.11). **Inner-transition elements** are those in the *f* area of the periodic table, and the distinguishing electron is found in an *f* subshell.

transition element An element in which the distinguishing electron is found in a *d* subshell.

inner-transition element An element in which the distinguishing electron is found in an *f* subshell.

Example 3.8

Use the periodic table and Figures 3.9 and 3.10 to determine the following for Ca, Fe, S, and Kr:

a. The type of distinguishing electron
b. The classification based on Figure 3.10

Solution

a. On the basis of the location of each element in Figure 3.9, the distinguishing electrons are of the following types:

Ca: *s* Fe: *d* S: *p* Kr: *p*

b. The classifications based on Figure 3.10 are: Ca, representative element; Fe, transition element; S, representative element; Kr, noble gas.

Learning Check 3.8 Determine the following for element numbers 38, 47, 50, and 86:

a. The type of distinguishing electron
b. The classification according to Figure 3.10

The elements can also be classified into the categories of metals, nonmetals, and metalloids. This approach, used in Figure 3.12, shows that most elements are classified as metals. It is also apparent from Figure 3.10 that all nonmetals and all metalloids are representative elements.

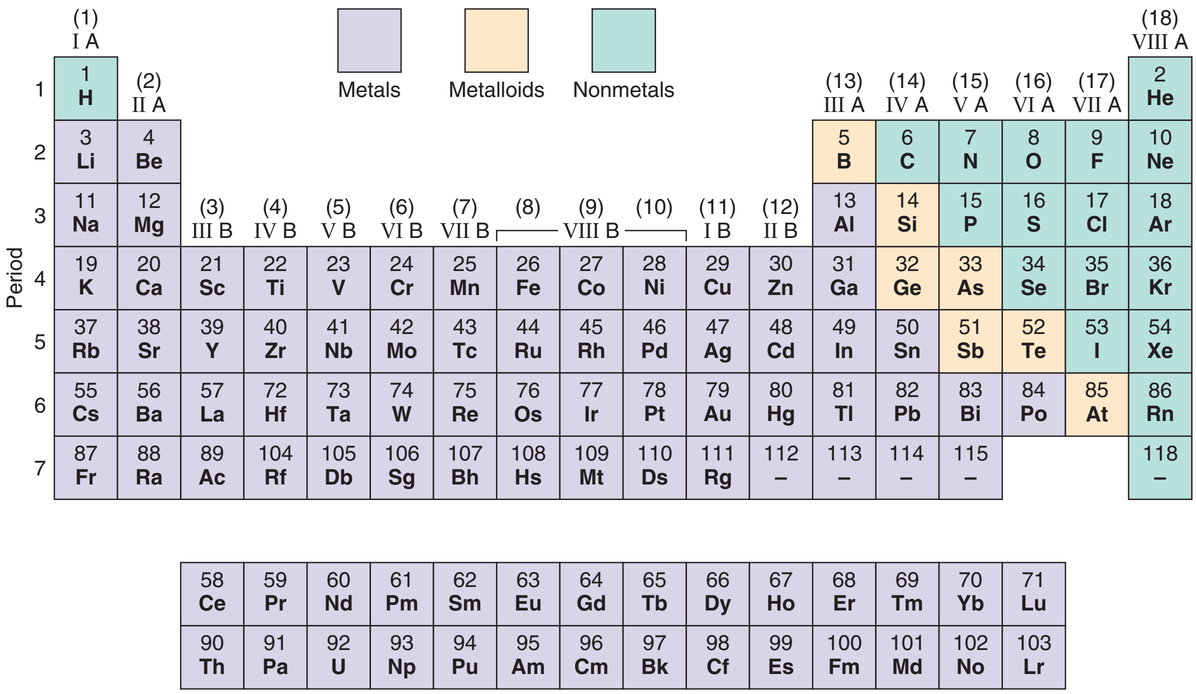

Figure 3.12 Locations of metals, nonmetals, and metalloids in the periodic table of the elements.

metals Elements found in the left two-thirds of the periodic table. Most have the following properties: high thermal and electrical conductivities, high malleability and ductility, and a metallic luster.

nonmetals Elements found in the right one-third of the periodic table. They often occur as brittle, powdery solids or gases and have properties generally opposite those of metals.

metalloids Elements that form a narrow diagonal band in the periodic table between metals and nonmetals. They have properties somewhat between those of metals and nonmetals.

Most **metals** have the following properties. (Are these properties physical or chemical?)

High thermal conductivity—they transmit heat readily.
High electrical conductivity—they transmit electricity readily.
Ductility—they can be drawn into wires.
Malleability—they can be hammered into thin sheets.
Metallic luster—they have a characteristic "metallic" appearance.

Nonmetals, the elements found in the right one-third of the periodic table, generally have chemical and physical properties opposite those of metals. Under normal conditions, they often occur as brittle, powdery solids or as gases.

Metalloids, such as boron (B) and silicon (Si), are the elements that form a diagonal separation zone between metals and nonmetals in the periodic table. Metalloids have properties somewhat between those of metals and nonmetals, and they often exhibit some of the characteristic properties of each type.

Learning Check 3.9 Classify each of the following elements as metal, nonmetal, or metalloid:

a. Xe **b.** As **c.** Hg **d.** Ba **e.** Th

3.6 Property Trends within the Periodic Table

Learning Objective

6. Recognize property trends of elements within the periodic table, and use the trends to predict selected properties of the elements.

In Figure 3.12, the elements are classified into the categories of metal, metalloid, or nonmetal according to their positions in the table and properties such as thermal conductivity

Chemistry and Your Health 3.1

Protecting Children from Iron Poisoning

The element iron plays a vital role in a number of body processes. Perhaps the most well-known of these functions is the role of iron as a component of hemoglobin, the oxygen-transporting protein of blood. When blood is iron poor, body tissues do not receive enough oxygen, and *anemia,* a general weakening of the body, results. Recommended Dietary Allowances (RDA) have been established for iron because of this and other important contributions to good health. The general RDA is 18 mg per day depending on age and sex, and 27 mg per day for pregnant women.

A well-balanced diet that includes meat, whole grains, and dark green vegetables will generally provide enough iron to satisfy the RDA for most individuals other than pregnant women. In an attempt to enhance the effectiveness of the diet in providing iron, many foods are enriched or fortified with iron—primarily breads, other flour products, and cereals. About 25% of all dietary iron consumed in the United States comes from such foods.

The emphasis on the health benefits of iron and the attempts to get enough iron into the diet make it somewhat surprising to be told that iron is also a serious poisoning threat for children. In fact, iron is the leading cause of poisoning deaths in children under the age of six. Since 1986, 110 thousand iron poisoning incidents in children have been reported, including 33 deaths. The iron-containing products involved range from innocent-appearing nonprescription daily multivitamin/mineral supplements for children to high-potency prescription iron supplements for pregnant women. Children showed symptoms of poisoning from consuming as few as 5 to as many as 98 iron-containing tablets. Immediate symptoms include nausea, vomiting, and diarrhea. Deaths occurred from ingesting as little as 200 mg to as much as 5850 mg of iron.

In an attempt to address this problem, the U.S. Food and Drug Administration (FDA) published regulations in 1997 that require warnings on all iron-containing drugs and dietary supplements about the risk of iron poisoning in children and the need to keep such products out of children's reach. In addition, the regulations require that most products containing 30 mg or more of iron per dosage unit have to be packaged as individual doses (such as in blister packages). This rule is designed to limit the number of pills or capsules a child might consume because of the difficulty encountered by a child in opening small individual packets. These 1997 FDA regulations added to rules already in place, including a U.S. Consumer Product Safety Commission regulation given in 1987. According to this regulation, most drugs and food supplements containing more than 250 mg of iron per container have to be packaged in child-resistant containers.

The primary responsibility for protecting young children from iron poisoning still rests with parents, older siblings, and other caregivers. These individuals must first recognize the hazards presented by iron-containing products and must then be very diligent in keeping such products out of the reach of young children.

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The warning at the bottom of this label is designed to help prevent accidental iron poisoning of children.

and metallic luster. It is generally true that within a period of the periodic table, the elements become less metallic as we move from left to right. Within a group, the metallic character increases from top to bottom.

Consider group VA(15) as an example. We see that the elements of the group consist of the two nonmetals nitrogen (N) and phosphorus (P), the two metalloids arsenic (As) and antimony (Sb), and the metal bismuth (Bi). We see the trend toward more metallic character from top to bottom that was mentioned above. According to what has been discussed concerning the periodic law and periodic table, these elements should have some similarities in chemical properties because they belong to the same group of the periodic table. Studies of their chemical properties show that they are similar but not identical. Instead, they follow trends according to the location of the elements within the group.

Certain physical properties also follow such trends, and some of these are easily observed, as shown in Figure 3.13. At room temperature and ordinary atmospheric pressure, nitrogen is a colorless gas, phosphorus is a white nonmetallic solid, arsenic is a brittle gray solid with a slight metallic luster, antimony is a brittle silver-white solid with a metallic luster, and bismuth is a lustrous silver-white metal.

We see that these physical properties are certainly not identical, but we do see that they change in a somewhat regular way (they follow a trend) as we move from element to element down the group. Nitrogen is a gas, but the others are solids. Nitrogen is colorless, phosphorus is white, then the other three become more and more metallic-looking. Also,

Figure 3.13 The elements of group VA(15) of the periodic table. Phosphorus must be stored under water because it will ignite when exposed to the oxygen in air. It has a slight color because of reactions with air.

as mentioned previously, we note a change from nonmetal to metalloid to metal as we come down the group.

The trends in these properties occur in a regular way that would allow us to predict some of them for one element if they were known for the other elements in the group. For example, if the properties of bismuth were unknown, we would have predicted that it would have a silvery white color and a metallic luster based on the appearance of arsenic and antimony. The three elements from group VIIA(17) shown in Figure 3.14 also demonstrate some obvious predictable trends in physical properties.

Trends in properties occur in elements that form periods across the periodic table as well as among those that form vertical groups. We will discuss two of these properties and their trends for representative elements. Our focus will be on the general trends, recognizing that some elements show deviations from these general behaviors. We will also propose explanations for the trends based on the electronic structure of atoms discussed in the chapter.

The first property we will consider is the size of the atoms of the representative elements. The size of an atom is considered to be the radius of a sphere extending from the center of the nucleus of the atom to the location of the outermost electrons around the nucleus. The behavior of this property across a period and down a group is shown in Figure 3.15. We see that the size increases from the top to the bottom of each group, and decreases from left to right across a period.

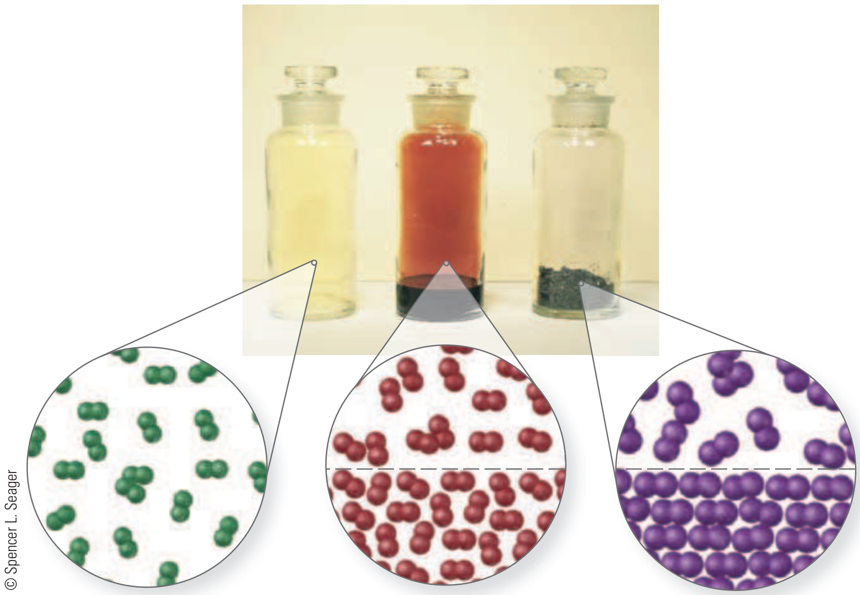

Figure 3.14 Chlorine, bromine, and iodine (left to right) all belong to group VIIA(17) of the periodic table. Their atoms all have the same number of electrons in the valence shell and therefore similar chemical properties. However, they do not have similar appearances. At room temperature and under normal atmospheric pressure, chlorine is a pale yellow gas, bromine is a dark red liquid that readily changes into a gas, and iodine is a gray-black solid that changes into a purple gas when heated slightly. Astatine is the next member of the group after iodine. Try to think like Mendeleev and predict its color and whether it will be a liquid, solid, or gas under normal conditions.

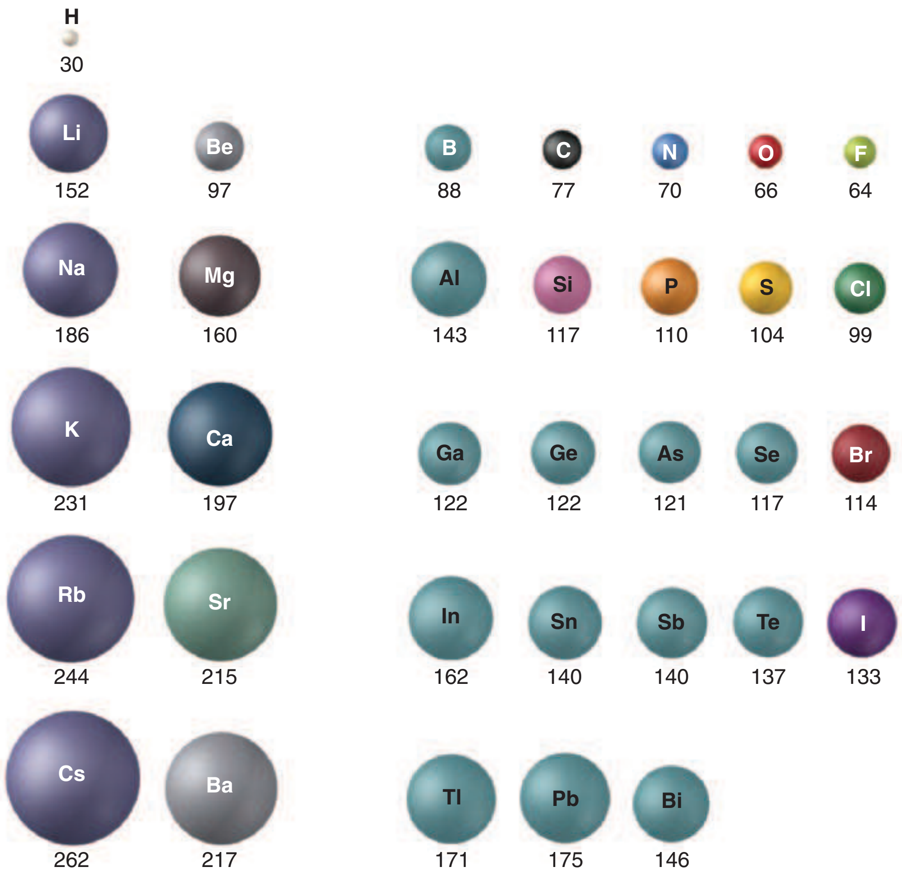

Figure 3.15 Scale drawings of the atoms of some representative elements enlarged about 60 million times. The numbers are the atomic radii in picometers (10^{-12} m).

Remember that when we go from element to element down a group, a new electronic shell is being filled. Consider group IA(1). In lithium atoms (Li), the second shell ($n = 2$) is beginning to fill, in sodium atoms (Na), the third shell ($n = 3$) is beginning to fill, and so on for each element in the group until the sixth shell ($n = 6$) begins to fill in cesium atoms (Cs). As n increases, the distance of the electrons from the nucleus in the shell designated by n increases, and the atomic radius as defined above also increases.

The decrease in atomic radius across a period can also be understood in terms of the outermost electrons around the nucleus. Consider period 3, for example. The abbreviated electronic structure for sodium atoms (Na) is [Ne]$3s^1$, and the outermost electron is in the third shell ($n = 3$). In magnesium atoms (Mg), the electronic structure is [Ne]$3s^2$, and we see that the outermost electron is still in the third shell. In aluminum atoms (Al), the electronic structure is [Ne]$3s^2 3p^1$, and we see that, once again, the outermost electron is in the third shell. In fact, the outermost electrons in all the atoms across period 3 are in the third shell.

Because all these electrons are in the same third shell, they should all be the same distance from the nucleus, and the atoms should all be the same size. However, each time another electron is added to the third shell of these elements, another positively charged proton is added to the nucleus. Thus, in sodium atoms there are 11 positive nuclear charges attracting the electrons, but in aluminum atoms there are 13. The effect of the increasing nuclear charge attracting electrons in the same shell is to pull all the electrons of the shell closer to the nucleus and cause the atomic radii to decrease as the nuclear charge increases.

The chemical reactivity of elements is dependent on the behavior of the electrons of the atoms of the elements, especially the valence electrons. One property that is related to the behavior of the electrons of atoms is the ionization energy.

Figure 3.16 First ionization energies for selected representative elements. The values are given in kJ/mole. (Source: Data from Bard, A. J.; Parsons, R.; Jordan, J. *Standard Potentials in Aqueous Solution*. New York: Dekker, 1985, pp. 24–27.)

IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
H 1311							**He** 2370
Li 521	**Be** 899	**B** 799	**C** 1087	**N** 1404	**O** 1314	**F** 1682	**Ne** 2080
Na 496	**Mg** 737	**Al** 576	**Si** 786	**P** 1052	**S** 1000	**Cl** 1245	**Ar** 1521
K 419	**Ca** 590	**Ga** 576	**Ge** 784	**As** 1013	**Se** 939	**Br** 1135	**Kr** 1351
Rb 402	**Sr** 549	**In** 559	**Sn** 704	**Sb** 834	**Te** 865	**I** 1007	**Xe** 1170
Cs 375	**Ba** 503	**Tl** 590	**Pb** 716	**Bi** 849	**Po** 791	**At** 926	**Rn** 1037

The *ionization energy* of an element is the energy required to remove an electron from an atom of the element in the gaseous state. The removal of one electron from an atom leaves the atom with a net 1+ charge because the nucleus of the resulting atom contains one more proton than the number of remaining electrons. The resulting charged atom is called an *ion*. We will discuss ions and the ionization process in more detail in Sections 4.2 and 4.3. A reaction for the removal of one electron from an atom of sodium is

$$Na(g) \rightarrow Na^{+}(g) + e^{-}$$

first ionization energy The energy required to remove the first electron from a neutral atom.

Because this process represents the removal of the first electron from a neutral sodium atom, the energy necessary to accomplish the process is called the **first ionization energy.** If a second electron were removed, the energy required would be called the second ionization energy, and so forth. We will focus on only the first ionization energy for representative elements. Figure 3.16 contains values for the first ionization energy of a number of representative elements.

The general trend is seen to be a decrease from the top to the bottom of a group, and an increase from left to right across a period. The higher the value of the ionization energy, the more difficult it is to remove an electron from the atoms of an element. Thus, we see that in general, the electrons of metals are more easily removed than are the electrons of nonmetals. Also, the farther down a group a metal is located, the easier it is to remove an electron.

The metals of group IA(1) all react with ethyl alcohol, C_2H_5OH, to produce hydrogen gas, H_2, as follows, where M is a general representation of the metals of the group:

$$2M + 2C_2H_5OH \rightarrow 2C_2H_5OM + H_2$$

In this reaction, electrons are removed from the metal and transferred to the hydrogen. The reaction is shown in progress for the first three members of the group in Figure 3.17. The rate (speed) of the reaction is indicated by the amount of hydrogen gas being released.

Learning Check 3.10 Refer to Figure 3.17, and do the following:

a. Arrange the three metals of the group vertically in order of the rate of the reaction with ethyl alcohol. Put the slowest reaction at the top and the fastest at the bottom.
b. Compare the trend in reaction rate coming down the group with the trend in ionization energy coming down the group.
c. Compare the trend in reaction rate coming down the group with the trend in ease of removing an electron from an atom of the metal coming down the group.

The trends in ionization energy and ease of removing an electron can be explained using arguments similar to those for explaining the sizes of atoms. As we come down a group,

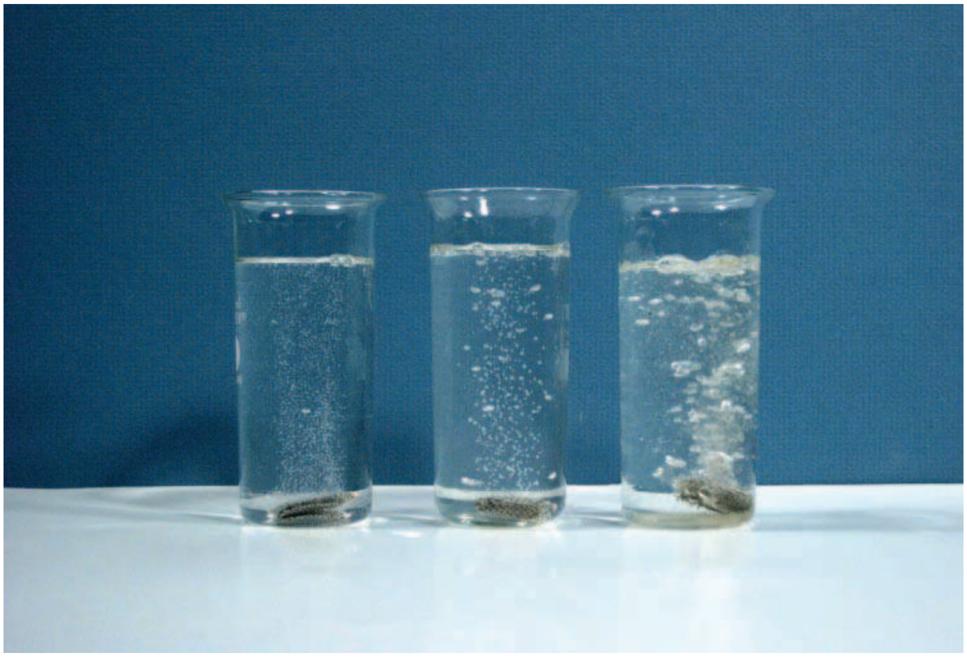

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Figure 3.17 The reaction of group IA(1) elements with ethyl alcohol. Left to right: Lithium (Li), sodium (Na), potassium (K). Each metal sample was wrapped in a wire screen to keep it from floating.

the valence electrons are located farther and farther away from the nucleus because they are located in higher-energy shells. The farther the electrons are away from the nucleus, the weaker is the attraction of the positively charged nucleus for the negatively charged electrons, and the easier it is to pull the electrons away.

Similarly, as we go from left to right in a period, the valence electrons are going into the same shell and therefore should be about the same distance from the nucleus. But as we saw before, the nuclear charge increases as well as the number of electrons. As a result, there is a greater nuclear charge attracting the electrons the farther we move to the right. Thus, a valence electron of an atom farther to the right is more difficult to remove than a valence electron of an atom farther to the left.

The general trends we have discussed are summarized in ◗ Figure 3.18. It is easiest to remember the trends by always going in the same direction in the table and noting how the properties change in that direction. We have chosen to summarize by always going from top to bottom for groups, and left to right for periods.

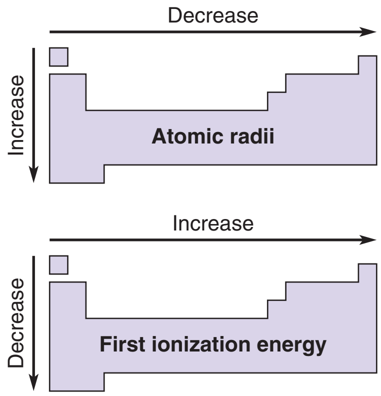

Figure 3.18 General trends for atomic size and ionization energy of representative elements.

Concept Summary

The Periodic Law and Table. The chemical properties of the elements tend to repeat in a regular (periodic) way when the elements are arranged in order of increasing atomic numbers. This periodic law is the basis for the arrangement of the elements called the periodic table. In this table, each element belongs to a vertical grouping, called a group or family, and a horizontal grouping, called a period. All elements in a group or family have similar chemical properties.

Objective 1, Exercise 3.4

Electronic Arrangements in Atoms. Niels Bohr proposed a theory for the electronic structure of hydrogen based on the idea that the electrons of atoms move around atomic nuclei in fixed circular orbits. Electrons change orbits only when they absorb or release energy. The Bohr model was modified as a result of continued research. It was found that precise Bohr orbits for electrons could not be determined. Instead, the energy and location of electrons could be specified in terms of shells, subshells, and orbitals, which are indicated by a notation system of numbers and letters.

Objective 2, Exercise 3.12

The Shell Model and Chemical Properties. The modified Bohr model, or shell model, of electronic structure provides an explanation for the periodic law. The rules governing electron occupancy in shells, subshells, and orbitals result in a repeating pattern of valence-shell electron arrangements. Elements with similar chemical properties turn out to be elements with identical numbers and types of electrons in their valence shells.

Objective 3, Exercises 3.18 and 3.22

Electronic Configurations. The arrangements of electrons in orbitals, subshells, and shells are called electronic configurations. Rules and patterns have been found that allow these configurations to be represented in a concise way. Written electronic configurations allow details of individual orbital, subshell, and shell electron occupancies to be seen readily. Also, the number of unpaired electrons in elements is easily determined. Electronic configurations can be represented in an abbreviated form by using noble gas symbols to represent some of the inner electrons.

Objective 4, Exercises 3.24 and 3.28

Another Look at the Periodic Table. Correlations between electronic configurations for the elements and the periodic table arrangement of elements make it possible to determine a number of details of electronic structure for an element simply on the basis of the location of the element in the periodic table. Special attention is paid to the last or distinguishing electron in an element. Elements are classified according to the type of subshell (*s, p, d, f*) occupied by this electron. The elements are also classified on the basis of other properties as metals, nonmetals, or metalloids.

Objective 5, Exercises 3.34 and 3.36

Property Trends within the Periodic Table. Chemical and physical properties of elements follow trends within the periodic table. These trends are described in terms of changes in properties of elements from the top to the bottom of groups, and from the left to the right of periods. The sizes of atoms and first ionization energies are two properties that show distinct trends.

Objective 6, Exercises 3.40 and 3.42

Key Terms and Concepts

Atomic orbital (3.2)
Distinguishing electron (3.5)
Electronic configuration (3.4)
First ionization energy (3.6)
Group or family of the periodic table (3.1)
Hund's rule (3.4)
Inner-transition element (3.5)
Metalloid (3.5)
Metal (3.5)
Noble gas configuration (3.4)
Nonmetal (3.5)
Pauli exclusion principle (3.4)
Periodic law (3.1)
Period of the periodic table (3.1)
Representative element (3.5)
Shell (3.2)
Subshell (3.2)
Transition element (3.5)
Valence shell (3.3)

Exercises

OWL Interactive versions of these problems are assignable in OWL.

Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

The Periodic Law and Table (Section 3.1)

3.1 Identify the group and period to which each of the following elements belongs:

a. Ca

b. element number 22

c. nickel

d. tin

3.2 Identify the group and period to which each of the following elements belongs:

a. element number 27

b. Pb

c. arsenic

d. Ba

3.3 Write the symbol and name for the elements located in the periodic table as follows:

a. Belongs to group VIA(16) and period 3

b. The first element (reading down) in group VIB(6)

c. The fourth element (reading left to right) in period 3

d. Belongs to group IB(11) and period 5

3.4 Write the symbol and name for the elements located in the periodic table as follows:

a. The noble gas belonging to period 4

b. The fourth element (reading down) in group IVA(14)

c. Belongs to group VIB(6) and period 5

d. The sixth element (reading left to right) in period 6

3.5 **a.** How many elements are located in group VIIIB(8, 9, 10) of the periodic table?

b. How many elements are found in period 2 of the periodic table?

c. How many total elements are in groups IIA(2) and VIA(16) of the periodic table?

3.6 **a.** How many elements are located in group VIIA(17) of the periodic table?

b. How many total elements are found in periods 2 and 3 of the periodic table?

c. How many elements are found in period 6 to the periodic table?

3.7 The following statements either define or are closely related to the terms *periodic law, period,* and *group.* Match the terms to the appropriate statements.

a. This is a vertical arrangement of elements in the periodic table

b. The chemical properties of the elements repeat in a regular way as the atomic numbers increase

c. The chemical properties of elements 11, 19, and 37 demonstrate this principle

d. Elements 4 and 12 belong to this arrangement

3.8 The following statements either define or are closely related to the terms *periodic law, period,* and *group.* Match the terms to the appropriate statements.

a. This is a horizontal arrangement of elements in the periodic table

b. Element 11 begins this arrangement in the periodic table

c The element nitrogen is the first member of this arrangement

d. Elements 9, 17, 35, and 53 belong to this arrangement

Electronic Arrangements in Atoms (Section 3.2)

3.9 According to the Bohr theory, which of the following would have the higher energy?

a. An electron in an orbit close to the nucleus

b. An electron in an orbit located farther from the nucleus

3.10 What particles in the nucleus cause the nucleus to have a positive charge?

3.11 What is the maximum number of electrons that can be contained in each of the following?

a. A 2*s* orbital

b. A 2*s* subshell

c. The first shell

3.12 What is the maximum number of electrons that can be contained in each of the following?

a. A 2*p* orbital

b. A 2*p* subshell

c. The second shell

3.13 How many orbitals are found in the third shell? Write designations for the orbitals.

3.14 How many orbitals are found in the fourth shell? Write designations for the orbitals.

3.15 How many orbitals are found in a 3*d* subshell? What is the maximum number of electrons that can be located in this subshell?

3.16 How many orbitals are found in a 4*f* subshell? What is the maximum number of electrons that can be located in this subshell?

3.17 Identify the subshells found in the fourth shell; indicate the maximum number of electrons that can occupy each subshell and the total number of electrons that can occupy the shell.

The Shell Model and Chemical Properties (Section 3.3)

3.18 Look at the periodic table and tell how many electrons are in the valence shell of the following elements:

a. element number 54

b. The first element (reading down) in group VA(15)

c. Sn

d. The fourth element (reading left to right) in period 3

3.19 Look at the periodic table and tell how many electrons are in the valence shell of the following elements:

a. element number 35

b. Zn

c. strontium

d. The second element in group VA(15)

3.20 What period 6 element has chemical properties most like sodium? How many valence-shell electrons does this element have? How many valence-shell electrons does sodium have?

3.21 What period 5 element has chemical properties most like silicon? How many valence-shell electrons does this element have? How many valence-shell electrons does silicon have?

3.22 If you discovered an ore deposit containing copper, what other two elements might you also expect to find in the ore? Explain your reasoning completely.

3.23 Radioactive isotopes of strontium were produced by the explosion of nuclear weapons. They were considered serious health hazards because they were incorporated into the bones of animals that ingested them. Explain why strontium would be likely to be deposited in bones.

Electronic Configurations (Section 3.4)

3.24 Write an electronic configuration for each of the following elements, using the form $1s^22s^22p^6$, and so on. Indicate how many electrons are unpaired in each case.

a. element number 37

b. Si

c. titanium

d. Ar

3.25 Write an electronic configuration for each of the following elements, using the form $1s^22s^22p^6$, and so on. Indicate how many of the electrons are unpaired in each case.

a. Br

b. element number 36

c. cadmium

d. Sb

3.26 Write electronic configurations and answer the following:

a. How many total *s* electrons are found in magnesium?

b. How many unpaired electrons are in nitrogen?

c. How many subshells are complerely filled in Al?

3.27 Write electronic configurations and answer the following:

a. How many total electrons in Ge have a number designation (before the letters) of 4?

b. How many unpaired *p* electrons are found in sulfur? What is the number designation of these unpaired electrons?

c. How many 3*d* electrons are found in tin?

3.28 Write the symbol and name for each of the elements described. More than one element will fit some descriptions.

a. Contains only two 2*p* electrons

b. Contains an unpaired 3*s* electron

c. Contains two unpaired 3*p* electrons

d. Contains three 4*d* electrons

e. Contains three unpaired 3*d* electrons

3.29 Write the symbol and name for each of the elements described. More than one element will fit some descriptions.

a. Contains one unpaired $5p$ electron

b. Contains a half-filled $5s$ subshell

c. Contains a half-filled $6p$ subshell

d. The last electron completes the $4d$ subshell

e. The last electron half fills the $4f$ subshell

3.30 Write abbreviated electronic configurations for the following:

a. selenium

b. element number 23

c. Ca

d. carbon

3.31 Write abbreviated electronic configurations for the following:

a. lead

b. element number 53

c. An element that contains 24 electrons

d. silicon

3.32 Refer to the periodic table and write abbreviated electronic configurations for all elements in which the noble gas symbol used will be [Ne].

3.33 Refer to the periodic table and determine how many elements have the symbol [Kr] in their abbreviated electronic configurations.

Another Look at the Periodic Table (Section 3.5)

3.34 Classify each of the following elements into the *s, p, d*, or *f* area of the periodic table on the basis of the distinguishing electron:

a. nickel

b. Rb

c. element 51

d. Cm

3.35 Classify each of the following elements into the *s, p, d*, or *f* area of the periodic table on the basis of the distinguishing electron:

a. Kr

b. tin

c. Pu

d. element 40

3.36 Classify the following elements as representative, transition, inner-transition, or noble gases:

a. iron

b. element 15

c. U

d. xenon

e. tin

3.37 Classify the following elements as representative, transition, inner-transition, or noble gases:

a. W

b. Cm

c. element 10

d. helium

e. barium

3.38 Classify the following as metals, nonmetals, or metalloids:

a. element 51

b. iodine

c. Al

d. radon

e. Pt

3.39 Classify the following as metals, nonmetals, or metalloids:

a. rubidium

b. arsenic

c. element 50

d. S

e. Br

Property Trends within the Periodic Table (Section 3.6)

3.40 Use trends within the periodic table to predict which member of each of the following pairs is more metallic:

a. Na or Mg

b. Pb or Ge

c. Mg or Ba

d. Cs or Li

3.41 Use trends within the periodic table to predict which member of each of the following pairs is more metallic:

a. C or Sn

b. Sb or In

c. Ca or As

d. Al or Mg

3.42 Use trends within the periodic table and indicate which member of each of the following pairs has the larger atomic radius:

a. Ga or Se

b. N or Sb

c. O or C

d. Te or S

3.43 Use trends within the periodic table and indicate which member of each of the following pairs has the larger atomic radius:

a. Mg or Sr

b. Rb or Ca

c. S or Te

d. I or Sn

3.44 Use trends within the periodic table and indicate which member of each of the following pairs gives up one electron more easily:

a. Li or K

b. C or Sn

c. Mg or S

d. Li or N

3.45 Use trends within the periodic table and indicate which member of each of the following pairs gives up one electron more easily:

a. Mg or Al

b. Ca or Be

c. S or Al

d. Te or O

Additional Exercises

3.46 How would you expect the chemical properties of isotopes of the same element to compare to each other? Explain your answer.

3.47 Bromine (Br) and mercury (Hg) are the only elements that are liquids at room temperature. All other elements in the periodic group that contains mercury are solids. Explain why mercury and bromine are not in the same group.

3.48 What would be the mass in mg of 3.0×10^{20} atoms that all have the same electronic configuration of $1s^2\ 2s^2\ 2p^4$?

3.49 Refer to Figure 3.15 and predict what would happen to the density of the metallic elements (purple color) as you go from left to right across a period of the periodic table. Explain your reasoning.

3.50 A 10.02-g sample of an element contains 0.250 mol of the element. Classify the element into the correct category of representative, transition, inner-transition, or noble gas. Will the element conduct electricity?

Allied Health Exam Connection

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3.51 The arrangement of the modern periodic table is based on atomic:

a. mass

b. number

c. radius

d. electronegativity

3.52 The horizontal rows of the periodic table are called:

a. families

b. groups

c. representative elements

d. periods

3.53 Which of the following is an example of a transition element?

a. aluminum

b. astatine

c. nickel

d. rubidium

3.54 Which two elements have chemical properties that are similar?

a. H and He

b. Fe and W

c. Li and Be

d. Mg and Ca

3.55 Which statement below is false?

a. Hydrogen is a nonmetal

b. Aluminum is a semimetal

c. Calcium is a metal

d. Argon is a gas at room temperature

3.56 Where on the periodic table are the nonmetals located?

a. upper right

b. upper left

c. lower right

d. lower left

3.57 What does the number 36 represent on the periodic table entry for krypton?

a. atomic number

b. relative atomic mass

c. group number

d. electron configuration

3.58 Which of the following is an alkali metal (group IA)?

a. calcium

b. sodium

c. aluminum

d. alkanium

3.59 Which of the following is an alkaline earth metal?

a. Na

b. Mg

c. Sc

d. Ti

3.60 What is the maximum number of electrons that each *p* orbital can hold?

a. 8

b. 2

c. 6

d. 4

3.61 From the periodic table, which of K and Br is larger?

a. K is larger

b. Br is larger

c. They are the same size

d. We cannot know which one is larger

3.62 The element with the smallest atomic radius of the following is:

a. Sr

b. Mg

c. Ba

d. Ra

3.63 Ionization energy is:

a. The energy required to completely remove an electron from an atom or ion

b. The energy created by an ion

c. The same as kinetic energy

d. The attraction between a proton and neutron

3.64 Which of the following has the largest first ionization energy?

a. Cs

b. Rb

c. Ba

d. Sr

3.65 Which elements conduct electricity?

a. metals

b. nonmetals

c. metalloids

d. ions

3.66 What term describes the electrons in the outermost principal energy level of an atom?

a. vector

b. core

c. kernel

d. valence

3.67 If the electron configuration of an element is written $1s^2\ 2s^2\ 2p_x^{\ 2}\ 2p_y^{\ 2}\ 2p_z^{\ 2}\ 3s^1$, the element's atomic:

a. number is 11

b. number is 12

c. weight is 11

d. weight is 12

3.68 Identify the 2 atoms with the same number of electrons in their outermost energy level.

a. Na/K

b. K/Ca

c. Na/Mg

d. Ca/Na

3.69 The number of unpaired electrons in the outer subshell of a phosphorus atom (atomic number: 15) is:

a. 2

b. 0

c. 3

d. 1

3.70 How many valence electrons are needed to complete the outer valence shell of sulfur?

a. 1

b. 2

c. 3

d. 4

3.71 An atom that has five 3 *p* electrons in its ground state is:

a. Si

b. P

c. Cl

d. O

Chemistry for Thought

3.72 Samples of three metals that belong to the same group of the periodic table are shown in Figure 3.5. When magnesium reacts with bromine, a compound with the formula $MgBr_2$ results. What would be the formulas of the compounds formed by reactions of bromine with each of the other metals shown? Explain your reasoning.

3.73 Answer the problem posed in Figure 3.14, then predict the same things for fluorine, the first member of the group. Explain the reasoning that led you to your answers.

3.74 Answer the problem posed in Figure 3.11. What property that makes gold suitable for coins and medals also makes it useful in electrical connectors for critical electronic parts such as computers in spacecraft?

3.75 Calcium metal reacts with cold water as follows:

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

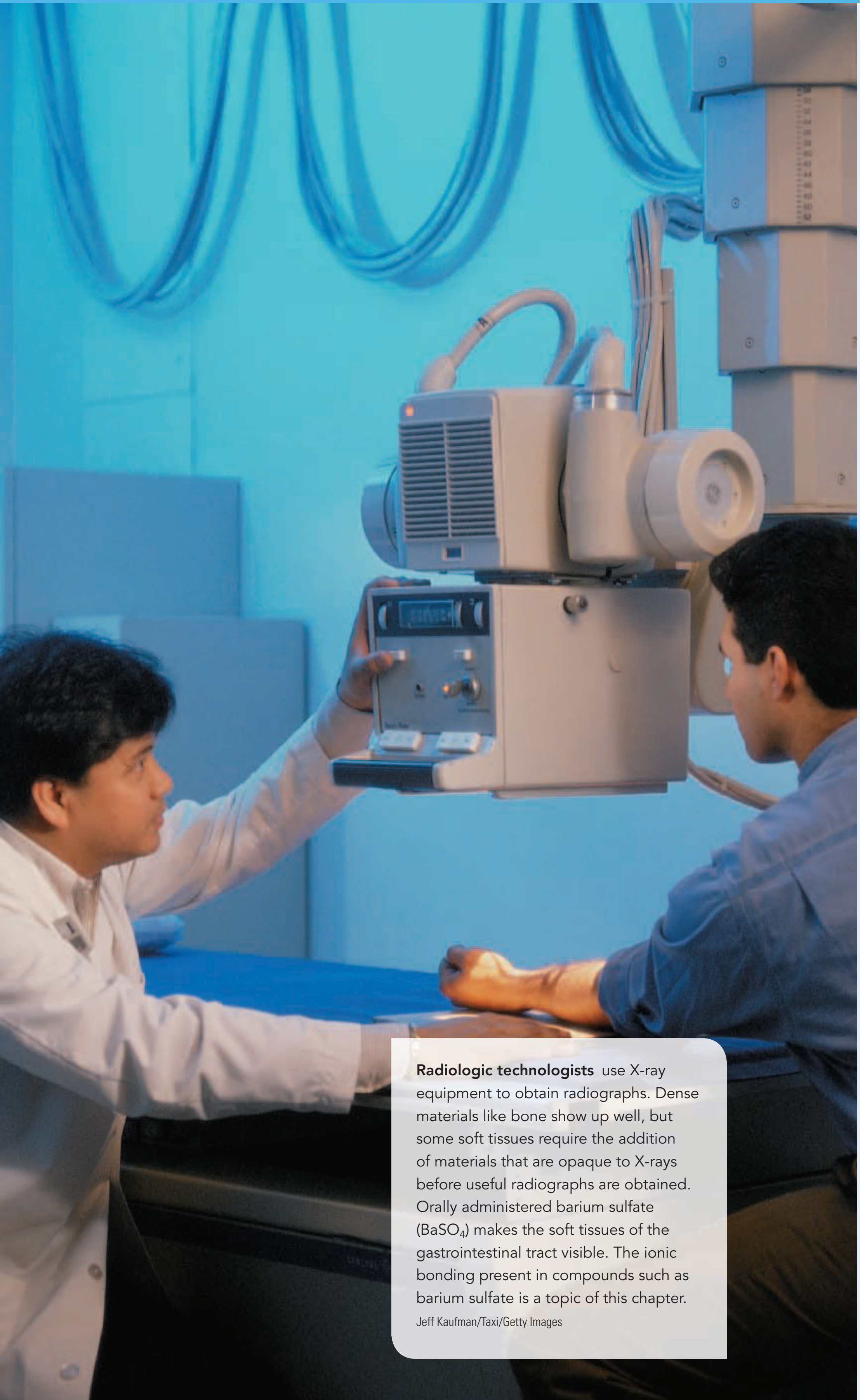
Magnesium metal does not react with cold water. What behavior toward cold water would you predict for strontium and barium? Write equations to represent any predicted reactions.

3.76 Refer to the hotels analogy in Study Skills 3.1 and determine the number of floors and the number of rooms on the top floor of Hotel Five.

3.77 A special sand is used by a company as a raw material. The company produces zirconium metal that is used to contain the fuel in nuclear reactors. What other metals are likely to be produced from the same raw material? Explain your answer.

Forces Between Particles

4



Radiologic technologists use X-ray equipment to obtain radiographs. Dense materials like bone show up well, but some soft tissues require the addition of materials that are opaque to X-rays before useful radiographs are obtained. Orally administered barium sulfate ($BaSO_4$) makes the soft tissues of the gastrointestinal tract visible. The ionic bonding present in compounds such as barium sulfate is a topic of this chapter.
Jeff Kaufman/Taxi/Getty Images

Learning Objectives

When you have completed your study of this chapter, you should be able to:

1 Draw correct Lewis structures for atoms of representative elements. **(Section 4.1)**

2 Use electronic configurations to determine the number of electrons gained or lost by atoms as they achieve noble gas electronic configurations. **(Section 4.2)**

3 Use the octet rule to correctly predict the ions formed during the formation of ionic compounds, and write correct formulas for binary ionic compounds containing a representative metal and a representative nonmetal. **(Section 4.3)**

4 Correctly name binary ionic compounds. **(Section 4.4)**

5 Determine formula weights for ionic compounds. **(Section 4.5)**

6 Draw correct Lewis structures for covalent molecules. **(Section 4.6)**

7 Draw correct Lewis structures for polyatomic ions. **(Section 4.7)**

8 Use VSEPR theory to predict the shapes of molecules and polyatomic ions. **(Section 4.8)**

9 Use electronegativities to classify covalent bonds of molecules, and determine whether covalent molecules are polar or nonpolar. **(Section 4.9)**

10 Write correct formulas for ionic compounds containing representative metals and polyatomic ions, and correctly name binary covalent compounds and compounds containing polyatomic ions. **(Section 4.10)**

11 Relate melting and boiling points of pure substances to the strength and type of interparticle forces present in the substances. **(Section 4.11)**

OWL Online homework for this chapter may be assigned in OWL.

In the discussion to this point, we have emphasized that matter is composed of tiny particles. However, we have not yet discussed the forces that hold these particles together to form the matter familiar to us. It is these forces that produce many of the properties associated with various types of matter—properties such as the electrical conductivity of copper wire, the melting point of butter, and the boiling point of water.

4.1 Noble Gas Configurations

Learning Objectives

1. Draw correct Lewis structures for atoms of representative elements.

The octet rule, proposed in 1916 by G. N. Lewis (an American) and Walter Kossel (a German), was the first widely accepted theory to describe bonding between atoms. The basis for the rule was the observed chemical stability of the noble gases. Because the noble gases did not react readily with other substances and because chemical reactivity depends on electronic structure, chemists concluded that the electronic structure of the noble gases represented a stable (low-energy) configuration. This noble gas configuration (see Section 3.4) is characterized by two electrons in the valence shell of helium and eight valence-shell electrons for the other members of the group (Ne, Ar, Kr, Xe, and Rn).

Example 4.1

Use the techniques described in Section 3.4 to write the electronic configurations for the noble gases. Identify the valence-shell electrons, and verify the noble gas configurations just mentioned.

Solution

The subshell-filling order in Figure 3.7 gives the electronic configurations shown below. Valence-shell electrons are in darker type, and the noble gas configurations are verified.

		Number of Valence-shell Electrons
He:	$\mathbf{1s^2}$	2
Ne:	$1s^2\mathbf{2s^2 2p^6}$	8
Ar:	$1s^2 2s^2 2p^6 \mathbf{3s^2 3p^6}$	8
Kr:	$1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^2} 3d^{10} \mathbf{4p^6}$	8
Xe:	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 \mathbf{5s^2} 4d^{10} \mathbf{5p^6}$	8
Rn:	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 \mathbf{6s^2} 4f^{14} 5d^{10} \mathbf{6p^6}$	8

Learning Check 4.1 Write the electronic configurations for F and K. What would have to be done to change these configurations to noble gas configurations (how many electrons would have to be added or removed)?

Lewis structure A representation of an atom or ion in which the elemental symbol represents the atomic nucleus and all but the valence-shell electrons. The valence-shell electrons are represented by dots arranged around the elemental symbol.

A simplified way to represent the valence-shell electrons of atoms was invented by G. N. Lewis. In these representations, called electron-dot formulas or **Lewis structures,** the symbol for an element represents the nucleus and all electrons around the nucleus except those in the valence shell. Valence-shell electrons are shown as dots around the symbol. Thus, rubidium is represented as Rb·.

When a Lewis structure is to be written for an element, it is necessary to determine the number of valence-shell electrons in the element. This can be done by writing the electronic configuration using the methods described in Chapter 3 and identifying the valence-shell electrons as those with the highest *n* value. A simpler alternative for representative elements is to refer to the periodic table and note that the number of valence-shell electrons in the atoms of an element is the same as the number of the group in the periodic table to which the element belongs. The group number used must be the one that precedes the letter A, not the one in parentheses as given in the periodic table inside the front cover of this book. Using this method, we see that rubidium belongs to group IA(1) and therefore has one valence-shell electron.

Example 4.2

Draw Lewis structures for atoms of the following:

a. element number 4
b. cesium (Cs)
c. aluminum (Al)
d. selenium (Se)

The accepted procedure is to write the element's symbol and put a dot for each valence electron in one of four equally spaced locations around the symbol. Imagine a square around the symbol. Each side of the square represents one of the four locations. An element with four valence electrons would have one dot in each of the four locations. A fifth electron would be represented by one additional dot in one of the locations. Each location can have a maximum of two dots.

Solution

a. Element number 4 is beryllium; it belongs to group IIA(2) and thus has two valence-shell electrons. The Lewis structure is $\underset{\cdot}{\mathrm{Be}}\cdot$.
b. Cesium is in group IA(1), has one valence-shell electron, and has the Lewis structure $\mathrm{Cs}\cdot$.
c. Aluminum is in group IIIA(13). It has three valence-shell electrons and is represented by the following Lewis structure: $\underset{\cdot}{\dot{\mathrm{Al}}}\cdot$.
d. Selenium is in group VIA(16) and thus has six valence-shell electrons. The Lewis structure is $:\underset{\cdot}{\ddot{\mathrm{Se}}}\cdot$.

Learning Check 4.2 Draw Lewis structures for atoms of the following:

a. element number 9
b. magnesium (Mg)
c. sulfur (S)
d. krypton (Kr)

It is important to clearly understand the relationship between Lewis structures and electronic configurations for atoms. The following example and learning check will help you review this relationship.

Example 4.3

Represent the following using abbreviated electronic configurations and Lewis structures:

a. F **b.** K **c.** Mg **d.** Si

Solution

a. Fluorine (F) contains 9 electrons, with 7 of them classified as valence electrons (those in the 2*s* and 2*p* subshells). The configuration and Lewis structure are

$$[\mathrm{He}]2s^22p^5 \qquad \text{and} \qquad :\underset{\cdot\cdot}{\ddot{\mathrm{F}}}\cdot$$

b. Potassium (K) contains 19 electrons, with 1 of them classified as a valence electron (the 1 in the 4*s* subshell). The configuration and Lewis structure are

$$[Ar]4s^1 \qquad \text{and} \qquad K\cdot$$

c. Magnesium (Mg) contains 12 electrons, with 2 of them classified as valence electrons (the 2 in the 3*s* subshell). The configuration and Lewis structure are

$$[Ne]3s^2 \qquad \text{and} \qquad \dot{Mg}\cdot$$

d. Silicon (Si) contains 14 electrons, with 4 of them classified as valence electrons (those in the 3*s* and 3*p* subshells). The configuration and Lewis structure are

$$[Ne]3s^23p^2 \qquad \text{and} \qquad \cdot\dot{\underset{\cdot}{Si}}\cdot$$

Learning Check 4.3 Using the two methods illustrated in Example 4.3, write electronic configurations and Lewis structures for the following atoms:

a. Li **b.** Br **c.** Sr **d.** S

4.2 Ionic Bonding

Learning Objective

2. Use electronic configurations to determine the number of electrons gained or lost by atoms as they achieve noble gas electronic configurations.

octet rule A rule for predicting electron behavior in reacting atoms. It says that atoms will gain or lose sufficient electrons to achieve an outer electron arrangement identical to that of a noble gas. This arrangement usually consists of eight electrons in the valence shell.

simple ion An atom that has acquired a net positive or negative charge by losing or gaining electrons.

ionic bond The attractive force that holds together ions of opposite charge.

According to the **octet rule** of Lewis and Kossel, atoms tend to interact through electronic rearrangements that produce a noble gas electronic configuration for each atom involved in the interaction. Except for the lowest-energy shell, this means that each atom ends up with eight electrons in the valence shell. There are exceptions to this octet rule, but it is still used because of the amount of information it provides. It is especially effective in describing reactions between the representative elements of the periodic table (see Figure 3.10).

During some chemical interactions, the octet rule is satisfied when electrons are transferred from one atom to another. As a result of the transfers, neutral atoms acquire net positive or negative electrical charges and become attracted to one another. These charged atoms are called **simple ions,** and the attractive force between oppositely charged atoms constitutes an **ionic bond.** A second type of interaction that also satisfies the octet rule is discussed in Section 4.6.

Example 4.4

Show how the following atoms can achieve a noble gas configuration and become ions by gaining or losing electrons:

a. Na **b.** Cl

Solution

a. The electronic structure of sodium (Na) is represented below using an abbreviated configuration and a Lewis structure:

$$[Ne]3s^1 \qquad \text{and} \qquad Na\cdot$$

The first representation makes it obvious that the Ne configuration with 8 electrons in the valence shell would result if the Na atom lost the single electron located in the 3*s* subshell. The loss is represented by the following equation:

$$Na \rightarrow Na^+ + 1e^-$$

Notice that the removal of a single negative electron from a neutral Na atom leaves the atom with 11 positive protons in the nucleus and 10 negative electrons. This gives the atom a net positive charge. The atom has become a positive ion.

b. The electronic structure for chlorine (Cl) is shown below using an abbreviated configuration and a Lewis structure:

$$[\text{Ne}]3s^23p^5 \qquad \text{and} \qquad :\ddot{\underset{..}{\text{Cl}}}\cdot$$

The Cl atom can achieve an Ne configuration by losing the 7 valence electrons. However, it is energetically much more favorable to achieve the configuration of argon (Ar) by adding 1 electron to the valence shell. This electron would complete an octet and change a Cl atom into a negative chloride ion as represented by the following equation, where Cl^- represents a Cl atom with 17 protons and 18 electrons, giving the atom a net negative charge:

$$\text{Cl} + 1e^- \rightarrow \text{Cl}^-$$

Learning Check 4.4 In Learning Check 4.1, you described what had to be done to change the electronic configurations of F and K to noble gas configurations. Write equations to illustrate these changes.

As a general rule, metals lose electrons and nonmetals gain electrons during ionic bond formation. The number of electrons lost or gained by a single atom rarely exceeds three, and this number can be predicted accurately for representative elements by using the periodic table. The number of electrons easily lost by a representative metal atom is the same as the group number (represented by roman numerals preceding the A in the periodic table). The number of electrons that tend to be gained by a representative nonmetal atom is equal to eight minus the group number. However, the nonmetal hydrogen is an exception. In most reactions, it loses one electron, consistent with its placement in group IA(1). In other, less common, reactions it gains one electron just as group VIIA(17) elements do.

Example 4.5

Use the periodic table to predict the number of electrons lost or gained by atoms of the following elements during ionic bond formation. Write an equation to represent the process in each case.

a. Li
b. Any group IIA(2) element, represented by the symbol M
c. element number 15
d. carbon

Solution

a. Lithium (Li), a metal, is in group IA(1); therefore it will lose 1 electron per atom:

$$\text{Li} \rightarrow \text{Li}^+ + 1e^-$$

b. Any group IIA(2) metal will lose 2 electrons; therefore,

$$\text{M} \rightarrow \text{M}^{2+} + 2e^-$$

c. Element number 15 is phosphorus (P), a nonmetal of group VA(15). It will gain $8 - 5 = 3$ electrons. The equation is

$$\text{P} + 3e^- \rightarrow \text{P}^{3-}$$

d. Carbon (C) is in group IVA(14) and is a nonmetal. It should, therefore, gain $8 - 4 = 4$ electrons. However, no more than 3 electrons generally become involved in ionic bond formation, so we conclude that carbon will not react readily to form ionic bonds.

Table 4.1 A Metal Atom, a Metal Ion, and a Noble Gas Atom Compared

Particle	Symbol	Number of Protons in Nucleus	Number of Electrons Around Nucleus	Net Charge on Particle
Magnesium atom	Mg	12	12	0
Magnesium ion	Mg^{2+}	12	10	+2
Neon atom	Ne	10	10	0

Learning Check 4.5 Predict the number of electrons that would be gained or lost during ionic bond formation for each of the following elements. Write an equation to represent each predicted change.

a. element number 34 **b.** Rb **c.** element number 18 **d.** In

isoelectronic A term that literally means "same electronic," used to describe atoms or ions that have identical electronic configurations.

While your attention is focused on noble gas configurations, it is appropriate to emphasize the following point: The attainment of a noble gas electronic configuration by an atom does not mean that the atom is converted into a noble gas. Instead, as seen above, the atoms are converted into simple ions (charged atoms). This point is emphasized by the comparison given in Table 4.1. Notice that the electronic configuration of Mg^{2+} is the same as that of Ne, but the number of protons (the atomic number) of Mg^{2+} is still the same as that of Mg. Atoms and simple ions that have identical electronic configurations are said to be **isoelectronic.**

4.3 Ionic Compounds

Learning Objective

3. Use the octet rule to correctly predict the ions formed during the formation of ionic compounds, and write correct formulas for binary ionic compounds containing a representative metal and a representative nonmetal.

So far the discussion has focused on the electron-transfer process as it occurs for isolated atoms. In reality, the electrons lost by a metal are the same ones gained by the nonmetal with which it is reacting. Substances formed from such reactions are called *ionic compounds*. No atom can lose electrons unless another atom is available to accept them, as shown in the formulas used to represent ionic compounds. These formulas represent the combining ratio of the positive and negative ions found in the compounds. This ratio is determined by the charges on the ions, which are determined by the number of electrons transferred.

Example 4.6

Represent the electron-transfer process that takes place when the following pairs of elements react ionically. Determine the formula for each resulting ionic compound.

a. Na and Cl **b.** Mg and F

Solution

a. Sodium (Na) is in group IA(1), and we can write

$$Na \rightarrow Na^{+} + 1e^{-}$$

Chemistry and Your Health 4.1

Fight Hypertension With Potassium

In the United States, hypertension (high blood pressure) is the primary reason people visit doctor's offices, and more prescriptions are written for its treatment than any other health problem. In addition to the use of prescription drugs, hypertension is also usually treated by reducing or eliminating the dietary intake of sodium in the form of table salt (sodium chloride). Recently-released research results indicate that combining an increase in dietary potassium intake with a reduction in sodium intake is probably the most important dietary decision (after excess weight loss) people can make to reduce cardiovascular diseases, including hypertension.

Studies reveal that in societies with diets rich in fruits and vegetables, only 1% of the population suffers from hypertension. By contrast, 33% of adults have hypertension in industrialized societies where the diet contains larger amounts of processed foods which often contain added salt. The typical diet in the United States contains about twice the sodium and only half the 4700 milligrams per day of potassium currently recommended by the American Heart Association.

It might seem that taking a daily supplement is the only way to insure a daily potassium intake as high as 4.7 grams, but that is not the case. Nature provides many potassium-rich foods such as squash, potatoes, tomatoes, carrots, spinach, beans, bananas, apricots, prunes, melons, peaches, halibut, tuna, trout, and low-fat dairy products. Specific examples of the potassium content of a few dietary potassium sources are given in the following table:

banana, 1 medium—422 mg 

© iStockphoto.com/ZoneCreative

sweet potato, 1 baked—694 mg 

© iStockphoto.com/EddWestmacott

white potato, 1 baked—610 mg 

© iStockphoto.com/Viktor Kitaykin

white beans, canned, ½ cup—595 mg 

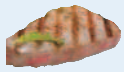
© iStockphoto.com/Stephan Hoerold

yogurt, plain, nonfat, 8 oz—595 mg 

Foodcollection/PhotoLibrary

halibut, cooked, 3 oz—490 mg 

© iStockphoto.com/Stacie Peterson

tuna, yellowfin, cooked, 3 oz—484 mg 

© iStockphoto.com/Eugene Bochkarev

lima beans, cooked, ½ cup—484 mg 

© iStockphoto.com/Joan Kimball

winter squash, cooked, ½ cup—448 mg 

© iStockphoto.com/velvetsnow

In addition to potassium, some studies have shown that the minerals magnesium and calcium may also have a positive influence in maintaining healthy blood pressure. The fruits and vegetables that provide potassium in the diet are also good sources of these two minerals. So, the parental directions traditionally given to children to "eat your fruits and vegetables and drink your milk" have been given scientific validity for all of us as a way to help maintain healthy levels of blood pressure.

Chlorine (Cl), in group VIIA(17), reacts as follows:

$$Cl + 1e^- \rightarrow Cl^-$$

The resulting ions, Na^+ and Cl^-, will combine in a 1:1 ratio because the total positive and total negative charges in the final formula must add up to zero. The formula then is NaCl. Note that the metal is written first in the formula, a generally followed practice. Also, the formula is written using the smallest number of each ion possible—in this case, one of each.

The actual electron-transfer process and the achievement of octets can more easily be visualized as follows:

$$\underset{Na}{[Ne]3s^1} + \underset{Cl}{[Ne]3s^23p^5} \rightarrow \underset{Na^+}{[Ne]} + \underset{Cl^-}{[Ne]3s^23p^6}$$

In this representation, it must be remembered that the $[Ne]3s^23p^6$ electronic configuration of Cl^- is the same configuration as argon, the noble gas that follows chlorine in the periodic table.

b. Magnesium (Mg) of group IIA(2) will lose two electrons per atom, whereas fluorine (F) of group VIIA(17) will gain one electron per atom. Therefore, we can write

$$Mg \rightarrow Mg^{2+} + 2e^-$$
$$F + 1e^- \rightarrow F^-$$

It is apparent that two fluorine atoms will be required to accept the electrons from one magnesium atom. From another point of view, two F^- ions will be needed to balance

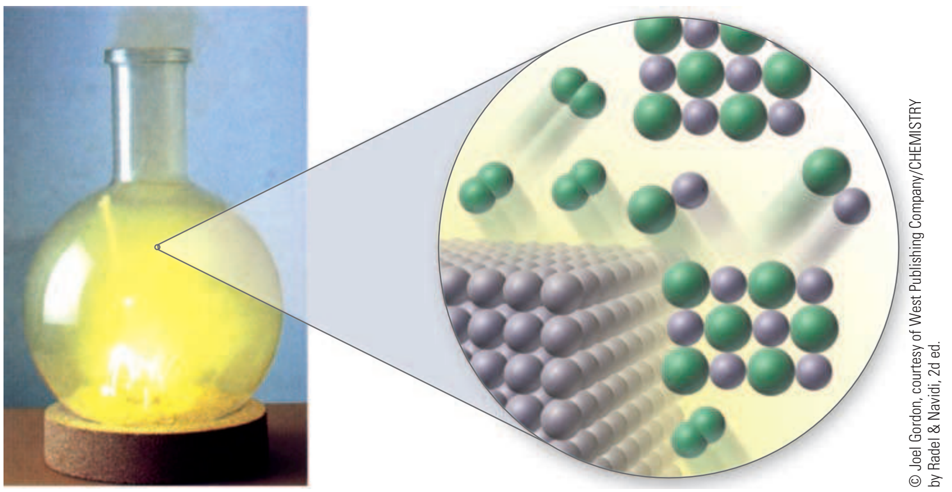


Figure 4.1 The reaction of sodium metal and chlorine gas. The vigorous reaction releases energy that heats the flask contents to a high temperature. Would you expect a similar reaction to occur between potassium metal and fluorine gas?

the charge of a single Mg^{2+} ion. Both observations lead to the formula MgF_2 for the compound. Note the use of subscripts to indicate the number of ions involved in the formula. The subscript 1, on Mg, is understood and never written.

Learning Check 4.6 Write equations to represent the formation of ions for each of the following pairs of elements. Write a formula for the ionic compound that would form in each case.

a. Mg and O **b.** K and S **c.** Ca and Br

binary compound A compound made up of two different elements.

Ionic compounds of the types used in Example 4.6 and Learning Check 4.6 are called **binary compounds** because each contains only two different kinds of atoms, Na and Cl in one case (see Figure 4.1) and Mg and F in the other.

4.4 Naming Binary Ionic Compounds

Learning Objective

4. Correctly name binary ionic compounds.

Names for binary ionic compounds are easily assigned when the names of the two elements involved are known. The name of the metallic element is given first, followed by the stem of the nonmetallic elemental name to which the suffix *-ide* has been added.

name = metal + nonmetal stem + *-ide*

The stem of the name of a nonmetal is the name of the nonmetal with the ending dropped. Table 4.2 gives the stem of the names of some common nonmetallic elements.

Example 4.7

Name the following binary ionic compounds:

a. KCl **b.** SrO **c.** Ca_3N_2

Solution

a. The metal is potassium (K), and the nonmetal is chlorine (Cl). Thus, the compound name is potassium chloride (the stem of the nonmetallic name is underlined).

Table 4.2 Stem Names and Ion Formulas of Common Nonmetallic Elements

Element	Stem	Formula of Ion
Bromine	brom-	Br^-
Chlorine	chlor-	Cl^-
Fluorine	fluor-	F^-
Iodine	iod-	I^-
Nitrogen	nitr-	N^{3-}
Oxygen	ox-	O^{2-}
Phosphorus	phosph-	P^{3-}
Sulfur	sulf-	S^{2-}

b. Similarly, strontium (Sr) and oxygen (O) give the compound name strontium <u>ox</u>ide.
c. The elements are calcium (Ca) and nitrogen (N), hence the name calcium <u>nitr</u>ide.

Learning Check 4.7 Assign names to the binary compounds whose formulas you wrote in Learning Check 4.6.

Names for the constituent ions of a binary compound are obtained in the same way as the compound name. Thus, K^+ is a potassium ion, whereas Cl^- is a chloride ion.

Some metal atoms, especially those of transition and inner-transition elements, form more than one type of charged ion. Copper, for example, forms both Cu^+ and Cu^{2+}, and iron forms Fe^{2+} and Fe^{3+}. The names of ionic compounds containing such elements must indicate which ion is present in the compound. A nomenclature system that does this well indicates the ionic charge of the metal ion by a roman numeral in parentheses following the name of the metal. Thus, CuCl is copper(I) chloride and $CuCl_2$ is copper(II) chloride. These names are expressed verbally as "copper one chloride" and "copper two chloride."

An older system is still in use but works only for naming compounds of metals that can form only two different charged ions. In this method, the endings *-ous* and *-ic* are attached to the stem of the metal name. For metals with elemental symbols derived from non-English names, the stem of the non-English name is used. The *-ous* ending is always used with the ion of lower charge and the *-ic* ending with the ion of higher charge. Thus, CuCl is cuprous chloride, and $CuCl_2$ is cupric chloride. In the case of iron, $FeCl_2$ is ferrous chloride, and $FeCl_3$ is ferric chloride (see Figure 4.2). Notice that the Cu^{2+} ion was called cupric, whereas the Fe^{2+} was called ferrous. The designations do not tell the actual ionic charge, but only which of the two is higher (Cu^{2+} and Fe^{3+}) or lower (Cu^+ and Fe^{2+}) for the metal in question.

Figure 4.2 Chloride compounds of copper and iron. Top: Cuprous and cupric chloride. Bottom: Ferrous and ferric chloride.

Example 4.8

Write formulas for ionic compounds that would form between the following simple ions. Note that the metal forms two different simple ions, and name each compound two ways.

a. Cr^{2+} and S^{2-} **b.** Cr^{3+} and S^{2-}

Solution

In each case, the metal ion is from the metal chromium (Cr).

a. Because the ionic charges are equal in magnitude but opposite in sign, the ions will combine in a 1:1 ratio. The formula is CrS. Chromium forms simple ions with 2+ and 3+ charges. This one is 2+, the lower of the two; thus, the names are chromium(II) sulfide and chromous sulfide.

b. In this case, the charges on the combining ions are 3+ and 2−. The smallest combining ratio that balances the charges is two Cr^{3+} (a total of 6+ charge) and three S^{2-} (a total of 6− charge). The formula is Cr_2S_3. The names are chromium(III) sulfide and chromic sulfide.

Learning Check 4.8 Write formulas for the ionic compounds that would form between the following simple ions. The metal is cobalt, and it forms only the two simple ions shown. Name each compound, using two methods.

a. Co^{2+} and Br^- **b.** Co^{3+} and Br^-

4.5 The Smallest Unit of Ionic Compounds

Learning Objective

5. Determine formula weights for ionic compounds.

In Section 1.3, a molecule was defined as the smallest unit of a pure substance capable of a stable, independent existence. As you will see later, some compound formulas are used to represent single molecules. True molecular formulas represent the precise numbers of atoms of each element that are found in a molecule. However, as we have seen, formulas for ionic compounds represent only the simplest combining ratio of the ions in the compounds.

The stable form of an ionic compound is not a molecule, but a crystal in which many ions of opposite charge occupy **lattice sites** in a rigid three-dimensional arrangement called a **crystal lattice.** For example, the lattice of sodium chloride (ordinary table salt) represented in Figure 4.3 is the stable form of pure sodium chloride.

lattice site The individual location occupied by a particle in a crystal lattice.

crystal lattice A rigid three-dimensional arrangement of particles.

Even though the formulas of ionic compounds do not represent true molecular formulas, they are often used as if they did. This is especially true when equations representing chemical reactions are written, or when the mole concept is applied to chemical formulas (see Section 2.7). When the atomic weights of the atoms making up a true molecular formula are added together, the result is called the molecular weight of the compound (Section 2.4). A similar quantity obtained by adding up the atomic weights of the atoms shown in the formula of an ionic compound is called a **formula weight.** The

formula weight The sum of the atomic weights of the atoms shown in the formula of an ionic compound.

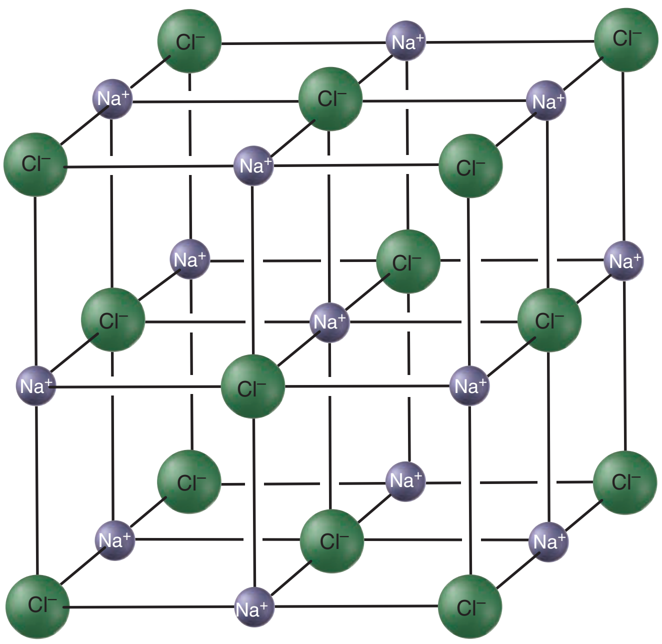


Figure 4.3 Crystal lattice for sodium chloride (table salt).

mole concept is applied to formula weights in a manner similar to the way it is applied to molecular weights.

Example 4.9

Carbon dioxide, CO_2, is a molecular compound, whereas magnesium chloride, $MgCl_2$, is an ionic compound.

a. Determine the molecular weight for CO_2 and the formula weight for $MgCl_2$ in atomic mass units.
b. Determine the mass in grams of 1.00 mol of each compound.
c. Determine the number of CO_2 molecules in 1.00 mol and the number of Mg^{2+} and Cl^- ions in 1.00 mol of $MgCl_2$.

Solution

a. For CO_2, the molecular weight is the sum of the atomic weights of the atoms in the formula:

$$\text{MW} = (1)(\text{at. wt. C}) + (2)(\text{at. wt. O}) = (1)(12.0\ \text{u}) + (2)(16.0\ \text{u})$$
$$\text{MW} = 44.0\ \text{u}$$

For $MgCl_2$, the formula weight is also equal to the sum of the atomic weights of the atoms in the formula:

$$\text{FW} = (1)(\text{at. wt. Mg}) + (2)(\text{at. wt. Cl}) = (1)(24.3\ \text{u}) + (2)(35.5\ \text{u})$$
$$\text{FW} = 95.3\ \text{u}$$

b. According to Section 2.6, 1.00 mol of a molecular compound has a mass in grams equal to the molecular weight of the compound. Thus, 1.00 mol CO_2 = 44.0 g CO_2.

For ionic compounds, 1.00 mol of compound has a mass in grams equal to the formula weight of the compound. Thus, 1.00 mol $MgCl_2$ = 95.3 g $MgCl_2$.

c. In Section 2.6 we learned that 1.00 mol of a molecular compound contains Avogadro's number of molecules. Thus, 1.00 mol $CO_2 = 6.02 \times 10^{23}$ molecules of CO_2.

In the case of ionic compounds, 1.00 mol of compound contains Avogadro's number of formula units. That is, 1.00 mol of magnesium chloride contains Avogadro's number of $MgCl_2$ units, where each unit represents one Mg^{2+} ion and two Cl^- ions. Thus,

$$1.00\ \text{mol}\ MgCl_2 = 6.02 \times 10^{23}\ MgCl_2\ \text{units}$$

or

$$1.00\ \text{mol}\ MgCl_2 = 6.02 \times 10^{23}\ Mg^{2+}\ \text{ions} + 1.20 \times 10^{24}\ Cl^-\ \text{ions}$$

Note that the number of Cl^- ions is simply twice Avogadro's number.

Learning Check 4.9 Determine the same quantities that were determined in Example 4.9, but use the molecular compound hydrogen sulfide, H_2S, and the ionic compound calcium oxide, CaO.

4.6 Covalent Bonding

Learning Objective

6. Draw correct Lewis structures for covalent molecules.

In Example 4.5, we found that carbon, a nonmetal, would have to gain four electrons in order to form ionic bonds. We concluded that carbon does not generally form ionic bonds. Yet carbon is known to form many compounds with other elements. Also,

Chemistry Around Us 4.1
Water: One of Earth's Special Compounds

An adequate supply of clean water is essential to our health and well-being. We can live without food for many days, but life would end in only a few days without water. Our circulatory system is an aqueous stream that distributes an amazing variety of substances throughout the body. Our cells are filled with water solutions in which the chemical reactions of life take place. Without water and its unique properties, life would not be possible on Earth.

Three-fourths of Earth's surface is covered with water, which gives the planet a blue color when it is viewed from space. The oceans represent the world's largest liquid solution. Water participates in most of the chemical reactions that occur in nature.

Most water used for human consumption comes from reservoirs, lakes, rivers, and wells. Much of this water is used and reused by numerous cities as it travels downstream. Such reused water may become seriously polluted with waste from previous users and with pathogenic microorganisms. As a safety precaution, most of the water we use undergoes chemical and physical treatment to purify it. This treatment process includes settling, whereby specific materials are added to bring down suspended solids. This is followed by filtration through sand and gravel to remove still more suspended matter. The filtered water may then be aerated by spraying it into the air. This part of the treatment process removes some odors and improves the taste of the water. Charcoal filtration, a treatment process that is gaining in use, also removes odors along with colored materials. In a final step, chlorine or another disinfectant is added to kill any remaining bacteria.

The treatment of water before use is an important safety precaution.

electron transfers would not be expected to take place between two atoms of the same element because such an exchange would not change the electronic configuration of the atoms involved. Yet molecules of elements containing two atoms, such as chlorine (Cl_2), oxygen (O_2), and nitrogen (N_2), are known to exist and, in fact, represent the stable form in which these elements occur in nature. What is the nature of the bonding in these molecules?

Again, G. N. Lewis provided an answer by suggesting that the valence-shell electrons of the atoms in such molecules are shared in a way that satisfies the octet rule for each of the atoms. This process is symbolized below for fluorine, (F_2), using Lewis structures:

$$:\ddot{\underset{..}{F}}\cdot + \cdot\ddot{\underset{..}{F}}: \longrightarrow :\ddot{\underset{..}{F}}:\ddot{\underset{..}{F}}:$$

shared pair (counted in octet of each atom)

Sometimes the shared electron pairs are shown as straight lines, and sometimes the nonshared pairs are not shown, as illustrated in Example 4.10.

Example 4.10

Represent the following reactions using Lewis structures:

a. $Cl + Cl \rightarrow Cl_2$ **b.** $H + H \rightarrow H_2$

Solution

a. Each chlorine (Cl) atom has seven valence-shell electrons. Because one electron from each atom is shared, the octet rule is satisfied for each atom

$$:\ddot{\underset{..}{Cl}}\cdot + \cdot\ddot{\underset{..}{Cl}}: \longrightarrow :\ddot{\underset{..}{Cl}}:\ddot{\underset{..}{Cl}}:$$

shared pair

The structure of Cl_2 can also be represented as

$$:\ddot{\underset{..}{Cl}}-\ddot{\underset{..}{Cl}}: \quad \text{or} \quad Cl-Cl$$

b. Each hydrogen (H) atom has one valence-shell electron. By sharing the two electrons, each H atom achieves the helium (He) structure:

$$H\cdot + \cdot H \longrightarrow H:H \quad \text{or} \quad H-H$$

Learning Check 4.10 Represent the following reactions using Lewis structures:

a. $N + N \rightarrow N_2$ (more than two electrons will be shared) **b.** $Br + Br \rightarrow Br_2$

To understand the origin of the attractive force between atoms that results from electron sharing, consider the H_2 molecule and the concept of atomic orbital overlap. Suppose two H atoms are moving toward each other. Each atom has a single electron in a spherical 1*s* orbital. While the atoms are separated, the orbitals are independent of each other; but as the atoms get closer together, the orbitals overlap and blend to create an orbital common to both atoms called a *molecular orbital.* The two shared electrons then move throughout the overlap region but have a high probability of being found somewhere between the two nuclei. As a result, both of the positive nuclei are attracted toward the negative pair of electrons and hence toward each other. This process, represented in Figure 4.4, produces a net attractive force between the nuclei. This attractive force is the **covalent bond** that holds the atoms together. Because forces cannot be seen, we represent only the presence of the shared electron pair between the atoms (by a pair of dots or a line) and remember that an attractive force is also there.

covalent bond The attractive force that results between two atoms that are both attracted to a shared pair of electrons.

Covalent bonding between like atoms has been described, but it also occurs between unlike atoms. Examples of both types are given in Table 4.3. Lewis structures, like those in Table 4.3, are most easily drawn by using a systematic approach such as the one represented by the following steps:

1. Use the molecular formula to determine how many atoms of each kind are in the molecule.
2. Use the given connecting pattern of the atoms to draw an initial structure for the molecule, with the atoms arranged properly.
3. Determine the total number of valence-shell electrons contained in the atoms of the molecule.
4. Put one pair of electrons between each bonded pair of atoms in the initial structure drawn in Step 2. Subtract the number of electrons used in this step from the total number determined in Step 3. Use the remaining electrons to complete the octets of all atoms in the structure, beginning with the atoms that are present in greatest number in the molecule. Remember, hydrogen atoms require only one pair to achieve the electronic configuration of helium.
5. If all octets cannot be satisfied with the available electrons, move nonbonding pairs (those that are not between bonded atoms) to positions between bonded atoms to complete octets. This will create **double** or **triple bonds** between some atoms.

double and triple bonds The bonds resulting from the sharing of two and three pairs of electrons, respectively.

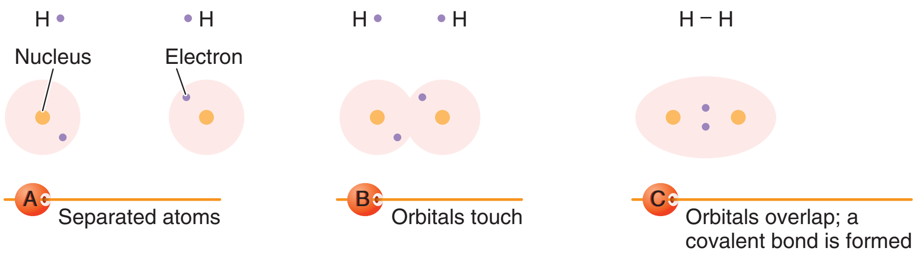
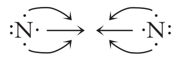
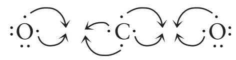
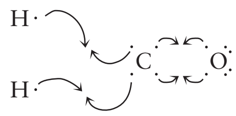
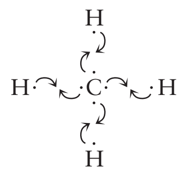
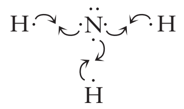
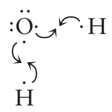
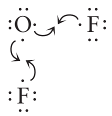


Figure 4.4 Orbital overlap during covalent bond formation.

Table 4.3 Examples of Covalent Bonding

Molecule	Atomic Lewis structure	Sharing pattern	Molecular Lewis structure
Nitrogen gas (N_2)	·N· $(1s^22s^22p^3)$		:N⋮⋮N:
Carbon dioxide (CO_2) (each O is bonded to the C)	·C· $(1s^22s^22p^2)$:O· $(1s^22s^22p^4)$		:O::C::O:
Formaldehyde (H_2CO) (each H and the O are bonded to the C)	·C· $(1s^22s^22p^2)$:O· $(1s^22s^22p^4)$ H· $(1s^1)$		H:C::O: H
Methane (CH_4) (each H is bonded to the C)	·C· $(1s^22s^22p^2)$ H· $(1s^1)$		H H:C:H H
Ammonia (NH_3) (each H is bonded to the N)	·N· $(1s^22s^22p^3)$ H· $(1s^1)$		H:N:H H
Water (H_2O) (each H is bonded to the O)	:O· $(1s^22s^22p^4)$ H· $(1s^1)$		:O:H H
Oxygen diflouride (OF_2) (each F is bonded to the O)	:O· $(1s^22s^22p^4)$:F: $(1s^22s^22p^5)$		:O:F: :F:

Example 4.11

Draw Lewis structures for the following molecules:

a. NH_3 **b.** SO_3 **c.** C_2H_2

Solution

a. Step 1. The formula indicates that the molecule contains 1 nitrogen (N) atom and 3 hydrogen (H) atoms.

Step 2. The connecting pattern will have to be given to you in some form. Table 4.3 shows that each H atom is connected to the N atom; thus, we draw

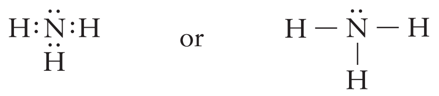
H N H
H

Step 3. Nitrogen is in group VA(15) of the periodic table and so has 5 valence electrons; hydrogen is in group IA(1) and has 1 valence electron. The total is 8 (5 from one N atom and 3 from the three H atoms).

Step 4. We put one pair of electrons between each H atom and the N atom in the initial structure drawn in Step 2:

H:N:H
H

This required 6 of the 8 available electrons. The remaining pair is used to complete the octet of nitrogen. Remember, hydrogen achieves the noble gas configuration of helium with only 2 electrons:



Step 5. All octets are satisfied by Step 4, so nothing more needs to be done.

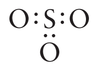
b. Step 1. The formula indicates that 1 sulfur (S) and 3 oxygen (O) atoms are present in a molecule.

Step 2. Each O atom is bonded only to the S atom; thus, we draw

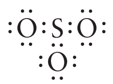
O S O
O

Step 3. Sulfur and oxygen are both in group VIA(16), and so each atom has 6 valence electrons. The total is 24 (6 from the one S atom and 18 from the three O atoms).

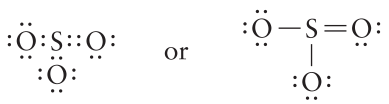
Step 4. We put one pair of electrons between each O atom and the S atom:



This required 6 of the 24 available electrons. The remaining 18 are used to complete the octets, beginning with the O atoms:



Step 5. We see that the octet of sulfur is not completed, even though all available electrons have been used. If one nonbonding pair of any one of the three O atoms is moved to a bonding position between the oxygen and sulfur, it will help satisfy the octet of both atoms. The final Lewis structure is given below. Note that it contains a double bond between one of the oxygens and the sulfur:



If a nonbonding pair of electrons from either of the other two oxygens is used, the following structures result. These are just as acceptable as the preceding structures:

c. Step 1. The formula indicates 2 carbon (C) atoms and 2 hydrogen (H) atoms per C_2H_2 molecule.

Step 2. The C atoms are bonded to each other, and 1 H atom is bonded to each C atom.

H C C H

Step 3. Carbon is in group IVA(14) and has 4 valence electrons; hydrogen is in group IA(1) and has 1 valence electron. The total electrons available is 10 (8 from the two C atoms and 2 from the two H atoms).

Step 4. We put one pair of electrons between the two C atoms and between each C atom and H atom.

H:C:C:H

This required 6 of the available 10 electrons. The remaining 4 electrons are used to complete the C octets (remember, hydrogen needs only 2 electrons):

$$\mathrm{H:C:\underset{\cdot\cdot}{\ddot{C}}:H}$$

Step 5. All electrons are used in Step 4, but the octet of 1 C atom is still incomplete. It needs two more pairs. If both nonbonding pairs on the second C atom were shared with the first C atom, both carbons would have complete octets. The result is

$$\mathrm{H:C:::C:H} \qquad \text{or} \qquad \mathrm{H{-}C{\equiv}C{-}H}$$

This molecule contains a triple bond (three shared pairs) between the C atoms.

Learning Check 4.11 Draw Lewis structures for the following molecules:

a. CH_4 (each H atom is bonded to the C atom)
b. H_2CO (the O atom and two H atoms are each bonded to the C atom)
c. HNO_3 (the O atoms are each bonded to the N atom, and the H atom is bonded to one of the O atoms)

4.7 Polyatomic Ions

Learning Objective

7. Draw correct Lewis structures for polyatomic ions.

polyatomic ions Covalently bonded groups of atoms that carry a net electrical charge.

An interesting combination of ionic and covalent bonding is found in compounds that contain **polyatomic ions.** These ions are covalently bonded groups of atoms that carry a net electrical charge. With the exception of the ammonium ion, NH_4^+, the common polyatomic ions are negatively charged.

Lewis structures can be drawn for polyatomic ions with just a slight modification in the steps given earlier for covalently bonded molecules. In Step 3, the total number of electrons available is obtained by first determining the total number of valence-shell electrons contained in the atoms of the ion. To this number are added electrons representing the number required to give the negative charge to the ion.

For example, in the SO_4^{2-} ion, the total number of valence-shell electrons is 30, with 6 coming from the sulfur atom and 6 from each of the four oxygen atoms. To this number we add 2 because two additional electrons are required to give the group of neutral atoms the charge of 2− found on the ion. This gives 32 total available electrons.

The only exception to this procedure for the common polyatomic ions is for the positively charged ammonium ion, NH_4^+. In this case, the number of electrons available is the total number of valence electrons minus 1, because one electron would have to be removed from the neutral group of atoms to produce the + charge on the ion.

Example 4.12

Draw Lewis structures for the following polyatomic ions. The connecting patterns of the atoms to each other are indicated.

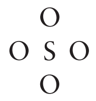
a. SO_4^{2-} (each O atom is bonded to the S atom)
b. NO_3^- (each O atom is bonded to the N atom)
c. $H_2PO_4^-$ (each O atom is bonded to the P atom, and each H atom is bonded to an O atom)

Solution

a. The only difference between drawing Lewis structures for molecules and polyatomic ions comes in Step 3, where the total number of valence-shell electrons is determined.

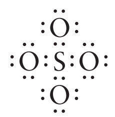
Step 1. The formula indicates the ion contains one sulfur (S) atom and four oxygen (O) atoms.

Step 2. The given bonding relationships lead to the following initial structure:

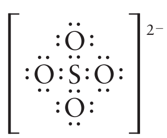


Step 3. Both the S and O atoms are in group VIA(16), so each atom contributes 6 valence electrons. The total number of valence-shell electrons from the atoms is therefore $5 \times 6 = 30$. However, the ion has a $2-$ charge. This charge comes from the presence of 2 electrons in the ion in addition to the electrons from the constituent atoms. These two additional electrons must be included in the total valence-shell electrons. Thus, the total number of electrons is $30 + 2 = 32$.

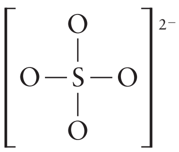
Step 4. The following is obtained when the 32 electrons are distributed among the atoms of the initial structure of Step 2:



Step 5. All octets are satisfied by Step 4. However, it is conventional to enclose Lewis structures of polyatomic ions in brackets and to indicate the net charge on the ion. Thus, the Lewis structures are



or



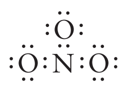
where the second structure does not show nonbonding (unshared) electron pairs, and the bonding pairs are represented by lines.

b. Step 1. Three oxygen (O) atoms and one nitrogen (N) atom are found in the ion.

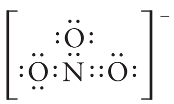
Step 2.

Step 3. Electrons from oxygen $= 3 \times 6 = 18$.
Electrons from nitrogen $= 1 \times 5 = 5$.
One electron comes from the ionic charge $= 1$.
Total electrons $= 18 + 5 + 1 = 24$.

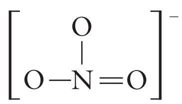
Step 4.



Step 5. The octet of nitrogen is not satisfied by Step 4, so an unshared pair of electrons on an O atom must be shared with the N atom. This will form a double bond in the ion. The resulting Lewis structures are



or



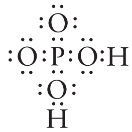
c. Step 1. Two hydrogen (H) atoms, one phosphorus (P) atom, and four oxygen (O) atoms are found in the ion.

Step 2.

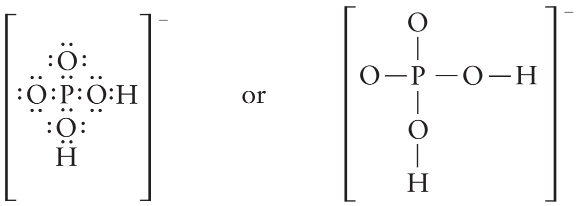
O
O P O H
O
H

Step 3. Electrons from hydrogen = $2 \times 1 = 2$.
Electrons from oxygen = $4 \times 6 = 24$.
Electrons from phosphorus = $1 \times 5 = 5$.
Electron from ionic charge = 1.
Total electrons = $2 + 24 + 5 + 1 = 32$.

Step 4.



Step 5. All octets are satisfied by Step 4 (remember, H atoms are satisfied by a single pair of electrons). The resulting Lewis structures are



Learning Check 4.12 Draw Lewis structures for the following polyatomic ions:

a. PO_4^{3-} (each O atom is bonded to the P atom)
b. SO_3^{2-} (each O atom is bonded to the S atom)
c. NH_4^{+} (each H atom is bonded to the N atom)

4.8 Shapes of Molecules and Polyatomic Ions

Learning Objective

8. Use VSEPR theory to predict the shapes of molecules and polyatomic ions.

Most molecules and polyatomic ions do not have flat, two-dimensional shapes like those implied by the molecular Lewis structures of Table 4.3. In fact, the atoms of most molecules and polyatomic ions form distinct three-dimensional shapes. Being able to predict the shape is important because the shape contributes to the properties of the molecule or ion. This can be done quite readily for molecules composed of representative elements.

To predict molecular or ionic shapes, first draw Lewis structures for the molecules or ions using the methods discussed in Sections 4.6 and 4.7. Once you have drawn the Lewis structure for a molecule or ion, you can predict its shape by applying a simple theory to the structure. The theory is called the *valence-shell electron-pair repulsion theory,* or **VSEPR theory** (sometimes pronounced "vesper" theory). According to VSEPR theory, electron pairs in the valence shell of an atom are repelled by other electron pairs and get as far away from one another as possible. Any atom in a molecule or ion that is bonded to two or more other atoms is called a *central atom.* When the VSEPR theory is applied to

VSEPR theory A theory based on the mutual repulsion of electron pairs. It is used to predict molecular shapes.

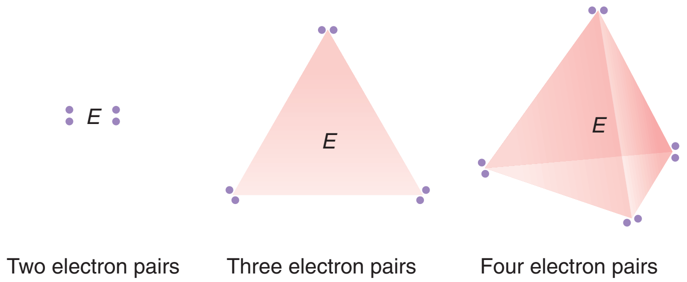


Figure 4.5 Arrangements of electron pairs around a central atom (E).

the valence-shell electrons of central atoms, the shape of the molecule or ion containing the atoms can be predicted. Two rules are followed:

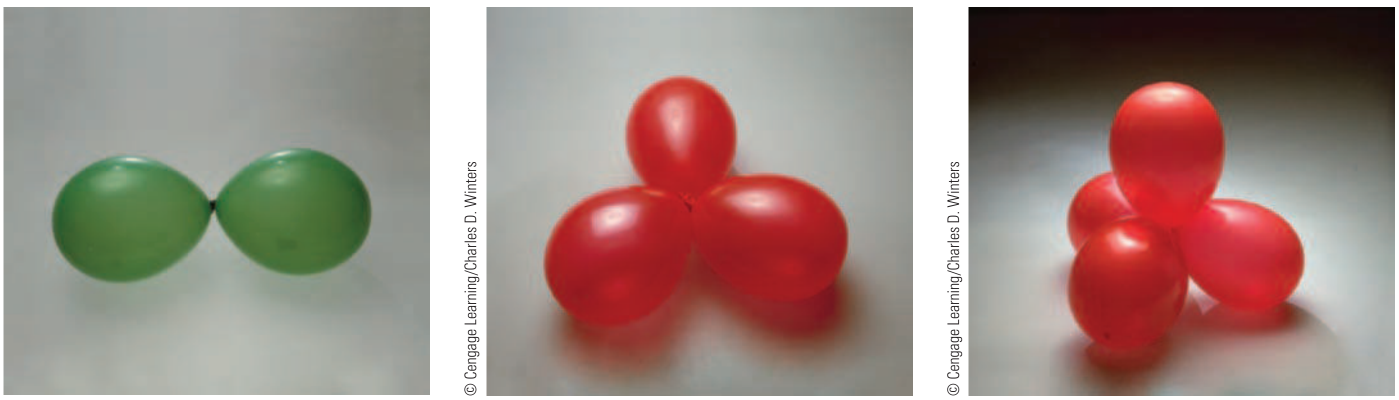
1. All valence-shell electron pairs around the central atom are counted equally, regardless of whether they are bonding or nonbonding pairs.
2. Double or triple bonds between atoms are treated like a single pair of electrons when predicting shapes.

The electron pairs around a central atom will become oriented in space to get as far away from one another as possible. Thus, two pairs will be oriented with one pair on each opposite side of the central atom. Three pairs will form a triangle around the central atom, and four pairs will be located at the corners of a regular tetrahedron with the central atom in the center (see Figure 4.5 and Active Figure 4.6). The VSEPR theory can be used to predict shapes of molecules and ions with five or more pairs on the central atom, but we will not go beyond four pairs in this book.

Example 4.13

Draw Lewis structures for the following molecules, apply the VSEPR theory, and predict the shape of each molecule:

a. CO_2 (each O atom is bonded to the C atom)
b. C_2Cl_4 (the C atoms are bonded together, and two Cl atoms are bonded to each C atom)
c. NH_3 (each H atom is bonded to the N atom)
d. H_2O (each H atom is bonded to the O atom)



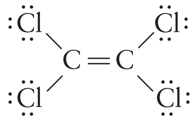
A Two balloons; linear
B Three balloons; triangular
C Four balloons; tetrahedral

Active Figure 4.6 When balloons of the same size and shape are tied together, they will assume positions like those taken by pairs of valence electrons around a central atom. Where would the central atom be located in these balloon models? **Go to www.cengage.com/chemistry/seager or OWL to explore an interactive version of this figure.**

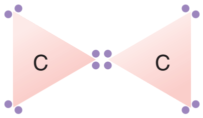
Solution

a. The Lewis structure is $:\ddot{O}=C=\ddot{O}:$. The central atom is carbon (C) and it has two pairs of electrons in the valence shell surrounding it (remember, each double bond is treated like a single pair of electrons in the VSEPR theory). The two pairs will be located on opposite sides of the C atom, so the molecule has the shape drawn above with the O, C, and O atoms in a line; it is a linear molecule.

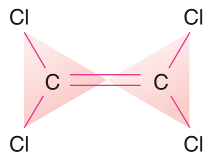
b. The Lewis structure is



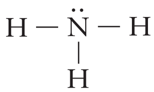
Both C atoms are central atoms in this molecule, and each one has three electron pairs in the valence shell (again, the double bond is treated as only one pair). The three pairs around each C atom will be arranged in the shape of a triangle:



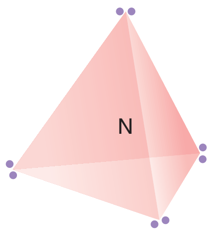
Thus, the molecule has the shape of two flat triangles in the same plane connected together at one point:



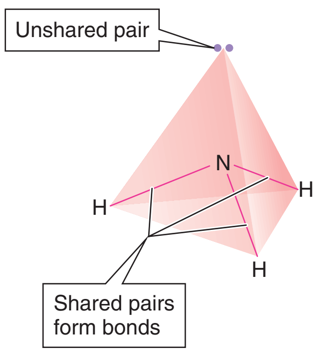
c. The Lewis structure is



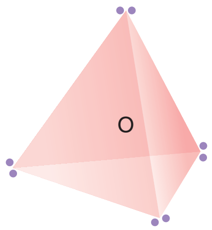
The central atom is nitrogen (N), and it has four electron pairs in the valence shell. The four pairs will be located at the corners of a tetrahedron with the N atom in the middle:



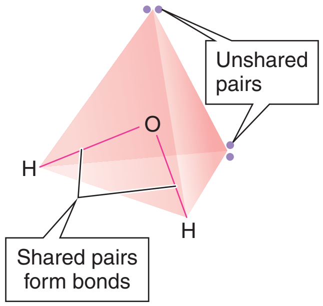
The shape of the molecule is determined only by the positions of the atoms, not by the position of the unshared pair of electrons. Thus, the NH_3 molecule has the shape of a pyramid with a triangular base. The N atom is at the peak of the pyramid, and an H atom is at each corner of the base:



d. The Lewis structure is H:Ö:H. The central atom is oxygen (O), and it is seen to have four pairs of electrons in its valence shell. The four pairs will be located at the corners of a tetrahedron with the O atom in the middle:

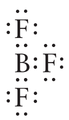


Once again, the shape of the molecule is determined by the locations of the atoms. Thus, H_2O is seen as having a bent or angular shape:



Learning Check 4.13 Predict the shapes of the following molecules by applying the VSEPR theory:

a. BF_3 (the Lewis structure is



for this molecule, which does not obey the octet rule).

b. CCl_4 (each Cl atom is bonded to the C atom)

c. SO_2 (each O atom is bonded to the S atom)

d. CS_2 (each S atom is bonded to the C atom)

Example 4.14

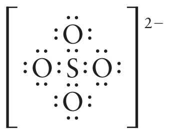
Use the VSEPR theory to predict the shapes of the following polyatomic ions. The Lewis structures were drawn in Example 4.12.

a. SO_4^{2-}

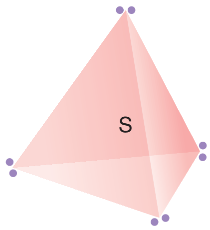
b. NO_3^-

Solution

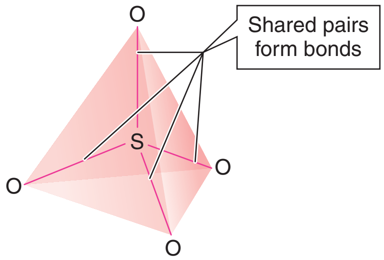
a. The Lewis structure from Example 4.12 is



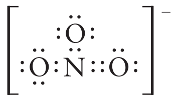
We see that sulfur is the central atom, and it has four electron pairs around it. The four pairs will be located at the corners of a tetrahedron with the S in the middle.



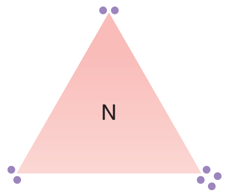
The shape of the ion is determined by the location of the oxygen atoms. Thus, the ion has a tetrahedral shape:



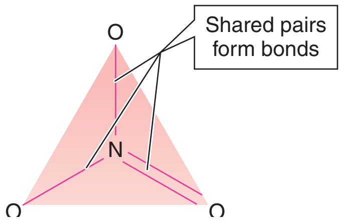
b. The Lewis structure from Example 4.12 is



We see that nitrogen is the central atom, and it has three electron pairs around it. Remember, the double bond formed by the two shared electron pairs between the nitrogen and one of the oxygens counts as only one pair in the VSEPR theory. The three pairs will form a triangle around the N.



An oxygen is located at each corner of the triangle, so the ion has a flat triangular shape:



Learning Check 4.14 Use the VSEPR theory to predict the shapes of the following polyatomic ions. The Lewis structures were drawn in Learning Check 4.12.

a. PO_4^{3-} **b.** SO_3^{2-} **c.** NH_4^+

At the Counter 4.1
Versatile Zinc Oxide

While the jury is still out on the usefulness of zinc compounds in treating the symptoms of the common cold (see At the Counter 3.1), the usefulness of zinc oxide as a topical astringent, antiseptic, and skin protectant is well established. Zinc oxide, with the simple formula ZnO, is a coarse white or grayish powder that is insoluble in water or in alcohol. Because of this insolubility, it is often blended with other ingredients to produce a suspension of the solid in such forms as lotions, creams, or salves. The solid compound in a finely divided form is sometimes added to products such as baby powders.

In these forms, zinc oxide is used topically on the skin as a sunscreen to block harmful ultraviolet rays, and as a shield from various chemical irritants. In addition, products containing the compound are used to relieve a number of common skin conditions, including diaper rash; poison ivy, oak, and sumac rashes eczema; external hemorrhoids; and insect bites. In these applications its antiseptic property helps prevent infections, and it is thought to promote healing by attracting protein to the affected areas of the skin, thereby encouraging new tissue growth. In these topical applications, the insolubility and consistency of zinc oxide prevent it from being absorbed into the skin. As a result, side effects or allergic reactions to the compound are rare, even with long-term use.

In addition to its valuable medical uses, this versatile compound is also found in a wide variety of commercial applications. It functions as a pigment and reinforcing agent in some rubber products; as a pigment and mold-growth inhibitor in paints; as a pigment in ceramics, floor tile, and glass; as a thickener in cosmetics; as a feed additive for cattle; and as a dietary supplement for humans.

Zinc oxide in the form of an ointment is convenient and easy to use.

4.9 The Polarity of Covalent Molecules

Learning Objective

9. Use electronegativities to classify covalent bonds of molecules, and determine whether covalent molecules are polar or nonpolar.

Molecules of chlorine (Cl—Cl) and hydrogen chloride (H—Cl) have a number of similarities. For example, they contain two atoms each and are therefore diatomic molecules, and the atoms are held together by a single covalent bond. Some differences also exist, the most obvious being the homoatomic nature of the chlorine molecule (an element) and the heteroatomic nature of the hydrogen chloride molecule (a compound). This difference influences the distribution of the shared electrons in the two molecules.

An electron pair shared by two identical atoms is attracted equally to each of the atoms, and the probability of finding the electrons close to an atom is equal for the two atoms. Thus, on average, the electrons spend exactly the same amount of time associated with each atom. Covalent bonds of this type are called **nonpolar covalent bonds.**

Different atoms generally have different tendencies to attract the shared electrons of a covalent bond. A measurement of this tendency is called **electronegativity.** The electronegativity of the elements is a property that follows trends in the periodic table. It increases from left to right across a period and decreases from top to bottom of a group. These trends are shown in ◗ Table 4.4, which contains electronegativity values for the common representative elements. As a result of electronegativity differences, the bonding electrons shared by two different atoms are shared unequally. The electrons spend more of their time near the atom with the higher electronegativity. The resulting shift in average location of bonding electrons is called **bond polarization** and produces a **polar covalent bond.**

nonpolar covalent bond A covalent bond in which the bonding pair of electrons is shared equally by the bonded atoms.

electronegativity The tendency of an atom to attract shared electrons of a covalent bond.

bond polarization A result of shared electrons being attracted to the more electronegative atom of a bonded pair of atoms.

polar covalent bond A covalent bond that shows bond polarization; that is, the bonding electrons are shared unequally.

Table 4.4 Electronegativities for the Common Representative Elements

Increasing electronegativity →

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

↓ *Decreasing electronegativity*

As a result of the unequal electron sharing, the more electronegative atom acquires a partial negative charge (δ−), and the less electronegative atom has a partial positive charge (δ+). The molecule as a whole has no net charge, just an uneven charge distribution.

Example 4.15

Use only the periodic table to determine the following for the diatomic covalent molecules listed below: (1) the more electronegative element, (2) the direction of bond polarization, and (3) the charge distribution resulting from the polarization.

a. I—Cl **b.** Br—Br **c.** C≡O

Solution

a. Both chlorine (Cl) and iodine (I) belong to group VIIA(17); chlorine is higher in the group and therefore is the more electronegative. The bond polarization (shift in average bonding-electron location) will be toward chlorine, as indicated by the arrow shown below the bond. The result is a partial negative charge on Cl and a partial positive charge on I:

$$\overset{\delta+}{\mathrm{I}}\underset{\rightarrow}{-}\overset{\delta-}{\mathrm{Cl}}$$

b. Since both bromine (Br) atoms have the same electronegativity, no bond polarization or unequal charge distribution results. The molecule is nonpolar covalent and is represented as Br—Br.

c. Oxygen (O) and carbon (C) both belong to the second period. Oxygen, being located farther to the right in the period, is the more electronegative. Thus, the molecule is represented as

$$\overset{\delta+}{\mathrm{C}}\underset{\rightarrow}{\equiv}\overset{\delta-}{\mathrm{O}}$$

Learning Check 4.15 Use only the periodic table and show the direction of any bond polarization and resulting charge distribution in the following molecules:

a. N≡N **b.** I—Br **c.** H—Br

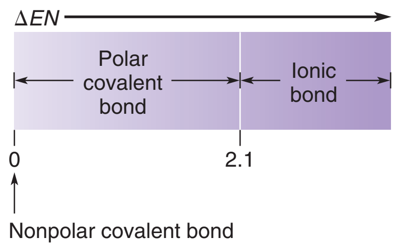


Figure 4.7 The darker the shading, the larger ΔEN is for the bonded atoms. Larger ΔEN values correspond to greater inequality in the sharing of the bonding electrons.

The extent of bond polarization depends on electronegativity differences between the bonded atoms and forms one basis for a classification of bonds. When the electronegativity difference (ΔEN) between the bonded atoms is 0.0, the bond is classified as nonpolar covalent (the electrons are shared equally). When ΔEN is 2.1 or greater, the bond is classified as ionic (the electrons are transferred). When ΔEN is between 0.0 and 2.1, the bond is classified as polar covalent (the electrons are unequally shared, and the bond is polarized). Note, however, that the change from purely covalent to completely ionic compounds is gradual and continuous as ΔEN values change. Thus, a bond between hydrogen (H) and boron (B) ($\Delta EN = 0.1$) is classified as polar, but the electron sharing is only slightly unequal and the bond is only slightly polar (see Figure 4.7). As the values in Table 4.4 indicate, compounds formed by a reaction of a metal with a nonmetal are generally ionic, while those between nonmetals are either nonpolar or polar covalent.

Example 4.16

Using Table 4.4, classify the bonds in the following compounds as nonpolar covalent, ionic, or polar covalent:

a. ClF **b.** MgO **c.** PI_3

Solution

a. The electronegativities of chlorine (Cl) and fluorine (F) are 3.0 and 4.0, respectively. Obtain the ΔEN value by subtracting the smaller from the larger, regardless of the order of the elements in the formula. Thus,

$$\Delta EN = 4.0 - 3.0 = 1.0$$

and the bond is polar covalent.

b. Similarly,

$$\Delta EN = 3.5 - 1.2 = 2.3$$

and the bond is ionic.

c. The ΔEN is determined for each phosphorus–iodine bond:

$$\Delta EN = 2.5 - 2.1 = 0.4$$

Thus, each bond is classified as polar covalent.

Learning Check 4.16 Using Table 4.4, classify the bonds in the following compounds as nonpolar covalent, ionic, or polar covalent:

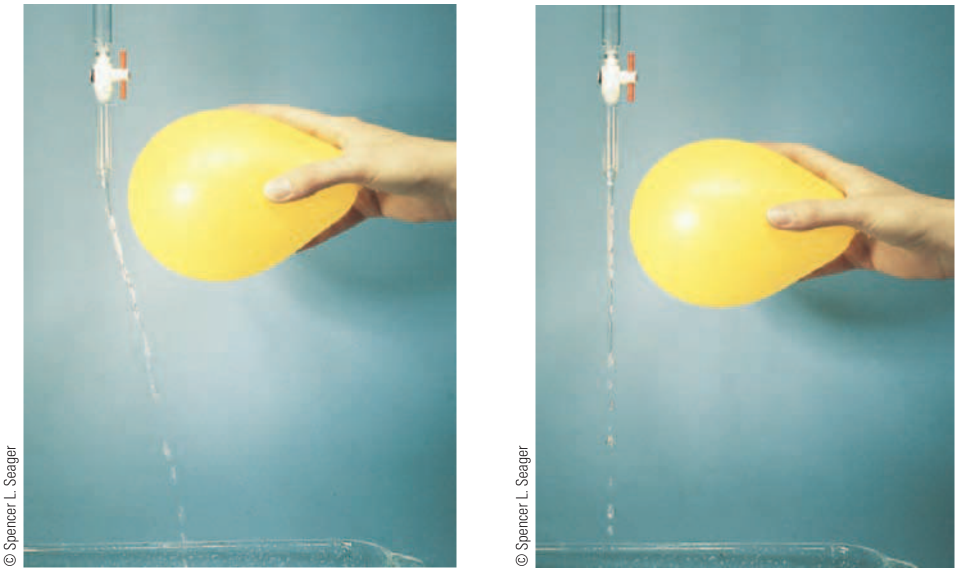
a. KF **b.** NO **c.** AlN **d.** K_2O

You can now predict the polarity of bonds within molecules, but it is also important to be able to predict the polar nature of the molecules themselves. The terms *polar* and *nonpolar,* when used to describe molecules, indicate their electrical charge symmetry (see Figure 4.8). In **polar molecules** (often called dipoles), the charge distribution resulting from bond polarizations is nonsymmetric. In **nonpolar molecules,** any charges caused by bond polarization are symmetrically distributed through the molecule. Table 4.5 illustrates polar and nonpolar molecules. The molecular shapes were determined by using the VSEPR theory.

polar molecule A molecule that contains polarized bonds and in which the resulting charges are distributed nonsymmetrically throughout the molecule.

nonpolar molecule A molecule that contains no polarized bonds, or a molecule containing polarized bonds in which the resulting charges are distributed symmetrically throughout the molecule.

Figure 4.8 Polar and nonpolar molecules are affected differently by electrical charges. The Δ*EN* between C and Cl is 0.5. Propose a reason why CCl_4 molecules are nonpolar.



A A stream of polar water molecules is attracted by a charged balloon.

B A stream of nonpolar carbon tetrachloride (CCl_4) molecules is not affected by a charged balloon.

Table 4.5 The Polarity of Molecules

Molecular formula	Geometric structure	Bond polarization and molecular charge distribution	Geometry of charge distribution	Classification of charge distribution	Classification of molecule
H_2	H—H	H—H	No charges	No charges	Nonpolar
HCl	H—Cl	$\overset{\delta+}{H} \rightrightarrows \overset{\delta-}{Cl}$	+ −	Nonsymmetric	Polar
CO_2	O=C=O	$\overset{\delta-}{O} \Lleftarrow \overset{\delta 2+}{C} \Rrightarrow \overset{\delta-}{O}$	− 2+ −	Symmetric	Nonpolar
BF_3	F–B(–F)–F (trigonal: F above B, two F below)	$\overset{\delta-}{F}$ ⇈ $\overset{\delta 3+}{B}$, $\overset{\delta-}{F}$ ⇙ B ⇘ $\overset{\delta-}{F}$	− 3+ − −	Symmetric	Nonpolar
N_2O	N≡N—O	$N \equiv \overset{\delta+}{N} \rightrightarrows \overset{\delta-}{O}$	+ −	Nonsymmetric	Polar
H_2O	H—O╲H	$\overset{\delta+}{H} \rightrightarrows \overset{\delta 2-}{O} \Lleftarrow \overset{\delta+}{H}$	+ 2− +	Nonsymmetric	Polar

4.10 More about Naming Compounds

Learning Objective

10. Write correct formulas for ionic compounds containing representative metals and polyatomic ions, and correctly name binary covalent compounds and compounds containing polyatomic ions.

In Section 4.4, we discussed the rules for naming binary ionic compounds. In this section, we expand the rules to include the naming of binary covalent compounds and ionic compounds that contain polyatomic ions.

It is apparent from Table 4.4 that covalent bonds (including polar covalent) form most often between representative elements classified as nonmetals. It is not possible to predict formulas for these compounds in a simple way, as was done for ionic compounds in Section 4.4. However, simple rules do exist for naming binary covalent compounds on the basis of their formulas. The rules are similar to those used to name binary ionic compounds: (1) Give the name of the less electronegative element first (the element given first in the formula), (2) give the stem of the name of the more electronegative element next and add the suffix *-ide,* and (3) indicate the number of each type of atom in the molecule by means of the Greek prefixes listed in Table 4.6.

Table 4.6 Greek Numerical Prefixes

Number	Prefix
1	*mono-*
2	*di-*
3	*tri-*
4	*tetra-*
5	*penta-*
6	*hexa-*
7	*hepta-*
8	*octa-*
9	*nona-*
10	*deca-*

Example 4.17

Name the following binary covalent compounds:

a. CO_2 **b.** CO **c.** NO_2 **d.** N_2O_5 **e.** CS_2

Solution

a. The elements are carbon (C) and oxygen (O), and carbon is the less electronegative. Because the stem of oxygen is *ox-,* the two portions of the name will be carbon and oxide. Only one C atom is found in the molecule, but the prefix *mono-* is not used when it appears at the beginning of a name. The two O atoms in the molecule are indicated by the prefix *di-*. The name therefore is carbon dioxide.
b. Similarly, we arrive at the name carbon monoxide for CO.
c. NO_2 is assigned the name nitrogen dioxide.
d. N_2O_5 is assigned the name dinitrogen pentoxide. Note: The a is dropped from the penta prefix for oxygen. This is done to avoid the pronunciation problem created by having two vowels next to each other in a name (dinitrogen pentaoxide).
e. CS_2 is named carbon disulfide.

Learning Check 4.17 Name the following binary covalent compounds:

a. SO_3 **b.** BF_3 **c.** S_2O_7 **d.** CCl_4

The formulas and names of some common polyatomic ions are given in Table 4.7. From this information, the formulas and names for compounds containing polyatomic ions can be written. The rules are essentially the same as those used earlier for binary ionic

Table 4.7 Some Common Polyatomic Ions

Very Common		Common	
NH_4^+	ammonium	CrO_4^{2-}	chromate
$C_2H_3O_2^-$	acetate	$Cr_2O_7^{2-}$	dichromate
CO_3^{2-}	carbonate	NO_2^-	nitrite
ClO_3^-	chlorate	MnO_4^-	permanganate
CN^-	cyanide	SO_3^{2-}	sulfite
HCO_3^-	hydrogen carbonate (bicarbonate)	ClO^-	hypochlorite
OH^-	hydroxide	HPO_4^{2-}	hydrogen phosphate
NO_3^-	nitrate	$H_2PO_4^-$	dihydrogen phosphate
PO_4^{3-}	phosphate	HSO_4^-	hydrogen sulfate (bisulfate)
SO_4^{2-}	sulfate	HSO_3^-	hydrogen sulfite (bisulfite)

Study Skills 4.1 Help with Polar and Nonpolar Molecules

When concepts are closely related, it is very easy to become confused. In this chapter, the terms *polar* and *nonpolar* are used to describe both bonds and molecules. What is the correct way to apply the terms?

It may be helpful to remember the numbers 3 and 2: three types of bonds, but only two types of molecules. The bonds are nonpolar covalent ($\Delta EN = 0$), polar covalent ($\Delta EN = 0.1$–2.0), and ionic ($\Delta EN = 2.1$ or greater). Remember that compounds bonded by ionic bonds do not form molecules, so there are only two types of molecules: nonpolar and polar.

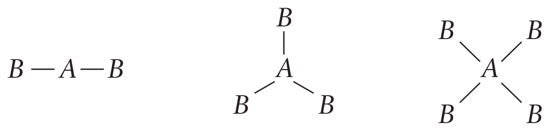
The polarity of a molecule depends on two factors: (1) the polarity of the bonds between the atoms of the molecule, and (2) the geometric arrangement of the atoms in space. First, determine whether any of the bonds between atoms are polar. If only nonpolar bonds are present, the molecule will be nonpolar regardless of the geometric arrangement of the atoms. However, if one or more polar bonds are found, the arrangement of the atoms must be considered as a second step.

For diatomic (two-atom) molecules, a polar bond between the atoms results in a polar molecule:

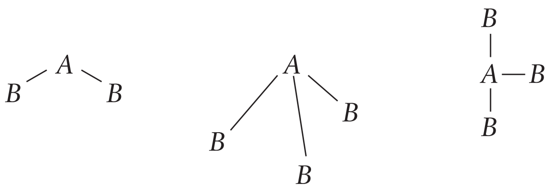
A—A	$\Delta EN = 0$ because atoms are identical; nonpolar
A—B	If $\Delta EN = 0$; nonpolar
A—B	If $\Delta EN = 0.1$–2.0; polar

Thus, we see that for diatomic molecules, the type of bond between atoms and the resulting type of molecule is the same.

For a molecule that contains polar bonds and three or more atoms, the molecular geometry must be known and taken into account before the polar nature of the molecule can be determined. Remember, the unequal sharing of electrons between two atoms bonded by a polar bond will cause one of the bonded atoms to have a partial positive charge and one a partial negative charge. If these atoms with their charges are symmetrically arranged in space to form the molecule, the molecule will be nonpolar. Thus, each of the following planar (flat) molecules will be nonpolar even if the bonds between the atoms are polar:



On the other hand, each of the following molecules will be polar if they contain polar bonds because the resulting charges on the atoms are not distributed symmetrically in space:



compounds. In the formulas, the metal (or ammonium ion) is written first, the positive and negative charges must add up to zero, and parentheses are used around the polyatomic ions if more than one is used. In names, the positive metal (or ammonium) ion is given first, followed by the name of the negative polyatomic ion (see Figure 4.9). No numerical

Figure 4.9 Examples of compounds that contain polyatomic ions. The samples are shown as solids and as solids dissolved in water. Referring to Table 4.7, write formulas for these compounds (clockwise from top): potassium carbonate, potassium chromate, potassium phosphate, and potassium permanganate.

prefixes are used except where they are a part of the polyatomic ion name. Names and formulas of acids (compounds in which hydrogen is bound to polyatomic ions) will be given in Chapter 9.

Example 4.18

Write formulas and names for compounds composed of ions of the following metals and the polyatomic ions indicated:

a. Na and NO_3^- **c.** K and HPO_4^{2-}
b. Ca and ClO_3^- **d.** NH_4^+ and NO_3^-

Solution

a. Sodium (Na) is a group IA(1) metal and forms Na^+ ions. Electrical neutrality requires a combining ratio of one Na^+ for one NO_3^-. The formula is $NaNO_3$. The name is given by the metal name plus the polyatomic ion name: sodium nitrate.
b. Calcium (Ca), a group IIA(2) metal, forms Ca^{2+} ions. Electrical neutrality therefore requires a combining ratio of one Ca^{2+} for two ClO_3^- ions. The formula is $Ca(ClO_3)_2$. (NOTE: The use of parentheses around the polyatomic ion prevents the confusion resulting from writing $CaClO_{32}$, which implies that there are 32 oxygen atoms in the formula. Parentheses are always used when multiples of a specific polyatomic ion are used in a formula and are indicated by a subscript.) The name is calcium chlorate.
c. Potassium (K), a group IA(1) metal, forms K^+ ions. The required combining ratio of 2:1 gives the formula K_2HPO_4. The name is potassium hydrogen phosphate.
d. The NH_4^+ is not a metallic ion, but it behaves like one in numerous compounds. The 1:1 combining ratio gives the formula NH_4NO_3. (NOTE: The polyatomic ions are written separately, and the nitrogen atoms are not grouped to give a formula such as $N_2H_4O_3$.) The name is ammonium nitrate.

Learning Check 4.18 Write formulas and names for compounds containing ions of the following metals and the polyatomic ions indicated:

a. Ca and HPO_4^{2-} **c.** K and MnO_4^-
b. Mg and PO_4^{3-} **d.** NH_4^+ and $Cr_2O_7^{2-}$

4.11 Other Interparticle Forces

Learning Objective

11. Relate melting and boiling points of pure substances to the strength and type of interparticle forces present in the substances.

Ionic and covalent bonding can account for certain properties of many substances. However, some experimental observations can be explained only by proposing the existence of other types of forces between particles.

Earlier in this chapter, a crystal lattice was described in connection with ionic bonding (see Figure 4.3). Most pure substances (elements or compounds) in the solid state also exist in the form of a crystal lattice. However, in some of these solids, neutral atoms or molecules occupy the lattice sites instead of ions. When solids are melted, the forces holding the lattice particles in place are overcome, and the particles move about more freely in what is called the liquid state. The addition of more heat overcomes the attractive interparticle forces to a still greater extent, and the liquid is converted into a gas or vapor; the liquid boils. Particles in the vapor state move about very freely and are influenced only slightly by interparticle attractions. Thus, the temperatures at which melting and boiling take place give an indication of the strength of the interparticle forces that are

Table 4.8 Some Characteristics of Selected Pure Substances

Substance	Formula or Symbol	Classification	Particles Occupying Lattice Sites
Sodium chloride	NaCl	Compound	Na^+ and Cl^- ions
Water	H_2O	Compound	H_2O molecules
Carbon monoxide	CO	Compound	CO molecules
Quartz (pure sand)	SiO_2	Compound	Si and O atoms
Copper metal	Cu	Element	Cu atoms
Oxygen	O_2	Element	O_2 molecules

being overcome. (These states of matter—solid, liquid, and gas—are discussed in more detail in Chapter 6.)

Suppose an experiment is carried out with several pure substances. Some are familiar to you, but others you probably have never seen in the solid state. The substances are listed in Table 4.8, together with some pertinent information.

The experiment is started at the very low temperature of –220°C to have all substances in the solid state. The temperature is then increased slowly and uniformly to 2600°C, at which point all the substances will be in the gaseous state. See Table 4.9, which gives only those temperatures corresponding to a specific change in one of the substances.

The melting points of the substances used in this experiment show that the weakest interparticle forces are found in solid oxygen. The forces then increase in the order CO, H_2O, NaCl, Cu, and SiO_2. With two exceptions, the boiling points follow the same order. How do these melting and boiling points relate to the lattice particles of these substances?

Table 4.9 The Behavior of Selected Pure Substances in Response to Heating

	Behavior or State of Substance					
Temperature (°C)	Oxygen (O_2)	Carbon Monoxide (CO)	Water (H_2O)	Salt (NaCl)	Copper (Cu)	Quartz (SiO_2)
−220	Solid	Solid	Solid	Solid	Solid	Solid
−218	**Melts**	Solid	Solid	Solid	Solid	Solid
−199	Liquid	**Melts**	Solid	Solid	Solid	Solid
−192	Liquid	**Boils**	Solid	Solid	Solid	Solid
−183	**Boils**	Gas	Solid	Solid	Solid	Solid
0	Gas	Gas	**Melts**	Solid	Solid	Solid
100	Gas	Gas	**Boils**	Solid	Solid	Solid
801	Gas	Gas	Gas	**Melts**	Solid	Solid
1083	Gas	Gas	Gas	Liquid	**Melts**	Solid
1413	Gas	Gas	Gas	**Boils**	Liquid	Solid
1610	Gas	Gas	Gas	Gas	Liquid	**Melts**
2230	Gas	Gas	Gas	Gas	Liquid	**Boils**
2595	Gas	Gas	Gas	Gas	**Boils**	Gas
2600	Gas	Gas	Gas	Gas	Gas	Gas

Chemistry Around Us 4.2

Nitric Oxide: A Simple but Vital Biological Molecule

Nitric oxide (NO) is a covalently bonded compound that is a toxic gas under ordinary conditions of temperature and pressure. The diatomic molecules of NO are only slightly polar, as indicated by an electronegativity difference (ΔEN) of only 0.5.

Until 1987, NO was regarded only as an environmental pollutant involved in numerous environmental problems, including the production of smog and acid precipitation. In 1987, researchers discovered that nitric oxide was produced by blood vessels. When NO was produced on the inside of blood vessels, it relaxed nearby muscles of the vessels, thereby reducing blood pressure. This discovery explained how a group of drugs, including amyl nitrite and nitroglycerine, worked to stop painful attacks of *angina.* During an angina attack, blood vessels to the heart constrict and reduce the supply of blood and oxygen to this vital organ. Drugs such as amyl nitrite and nitroglycerine produce NO inside the vessels, cause the vessels to relax, and restore the blood supply.

A second function of nitric oxide is protecting the body against unwanted foreign particles such as bacteria. Blood cells called macrophages seek out and destroy foreign particles by injecting them with a fatal dose of toxic NO.

In the early 1990s, researchers discovered that NO functions as a neurotransmitter, a chemical that carries messages from one nerve cell to another. This discovery was surprising because other known neurotransmitters are larger, more complicated molecules, and none are gases. The small size and low polarity of NO molecules allows them to diffuse quickly through cell membranes, a characteristic that enhances the role of a neurotransmitter. In this role, NO is known to be involved in long-term memory functions of the brain, the maintenance of blood pressure, central nervous system functions, and the immune system's response to infections caused by some types of viruses.

The lattice particles in solid silicon dioxide are individual atoms of silicon and oxygen. They are held together in the lattice by covalent bonds. Solids of this type are called **network solids,** and when such solids are melted or vaporized, strong covalent bonds must be broken.

network solid A solid in which the lattice sites are occupied by atoms that are covalently bonded to each other.

The individual atoms that are the lattice particles of copper metal are held together in the lattice by what is called a **metallic bond.** As described in Section 4.2, the atoms of metals lose valence-shell electrons readily. Imagine a large number of metal atoms occupying lattice sites. Now imagine that the valence-shell electrons of each of the atoms move readily throughout the lattice. The attraction of the positive kernels (the metal atom nuclei plus low-level electrons) to the mobile electrons, and hence to one another, constitutes a metallic bond. The mobile electrons of the metallic bond are responsible for a number of the observed properties of metals, including high thermal conductivity, high electrical conductivity, and the characteristic metallic luster.

metallic bond An attractive force responsible for holding solid metals together. It originates from the attraction between positively charged atomic kernels that occupy the lattice sites and mobile electrons that move freely through the lattice.

The bonds broken when sodium chloride (NaCl) melts or boils are ionic bonds resulting from attractions between positive (Na^+) and negative (Cl^-) ions, the lattice particles in Figure 4.3. As shown by Table 4.9, these bonds are generally quite strong.

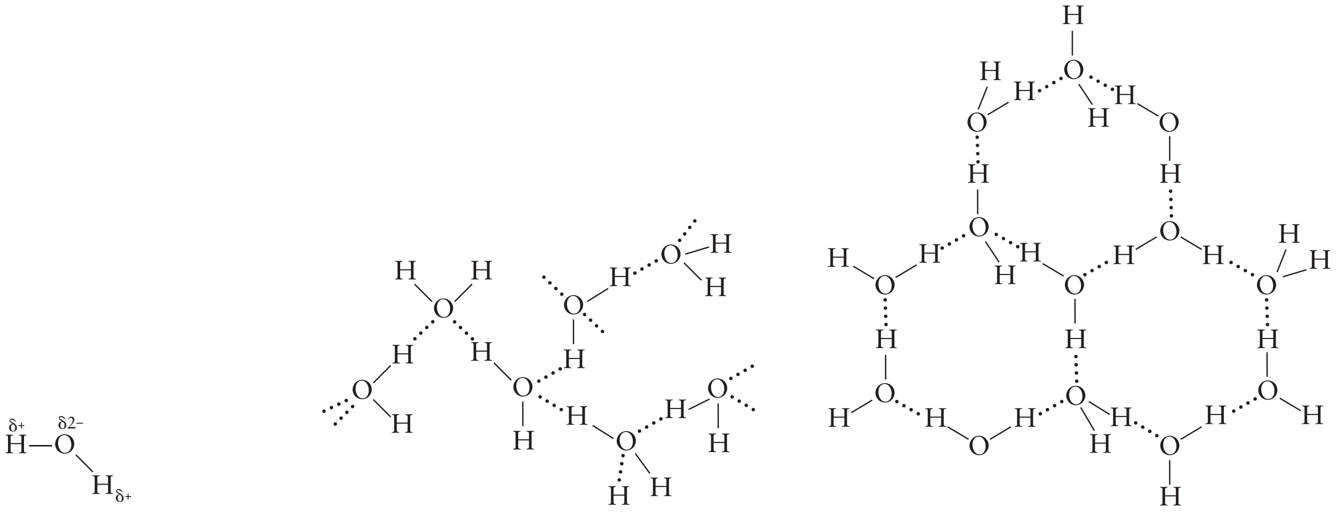
The lattice particles of solid carbon monoxide (CO) are the nonsymmetric CO molecules. These molecules are polar because of the nonsymmetric distribution of charges between the carbon and oxygen atoms. They are held in the solid lattice by **dipolar forces** resulting from the attraction of the positive end of one polar molecule to the negative end of another polar molecule. These forces are usually weak, and melting and vaporization take place at very low temperatures. Such substances are usually thought of as gases because that is their normal state at room temperature.

dipolar force The attractive force that exists between the positive end of one polar molecule and the negative end of another.

Water molecules, the lattice particles of ice, are also held in place by dipolar forces. However, these forces are stronger than those of the dipoles for CO. In water molecules, the hydrogens carry a partial positive charge, and the oxygen has a partial negative charge (see Table 4.5). Thus, the hydrogens of one molecule are attracted to the oxygens of other molecules. This attraction, called **hydrogen bonding,** is stronger than most other dipolar attractions because of the small size of the hydrogen atom and the high electronegativity of oxygen (see ◗ Figure 4.10). Hydrogen bonding occurs in gases, liquids, and solids composed of polar molecules in which hydrogen atoms are covalently bonded to highly electronegative elements (generally O, N, or F).

hydrogen bonding The result of attractive dipolar forces between molecules in which hydrogen atoms are covalently bonded to very electronegative elements (O, N, or F).

Water is the most well-known substance in which hydrogen bonding dramatically influences the properties. These properties make water useful in many processes, including those characteristic of living organisms. Because of this widespread use, water is not often



A Polar water molecule

B Hydrogen bonding in liquid water

C Hydrogen bonding in solid water

Figure 4.10 Hydrogen bonding in water. Hydrogen bonds are dotted, and covalent bonds are solid.

thought of as being a peculiar substance. However, it is often the peculiarities that make it so useful.

As we saw in Example 4.13d, water molecules are angular. Because of this shape and the large difference in electronegativity between hydrogen and oxygen, water molecules are polar, with partial positive charges on the hydrogens and a partial negative charge on the oxygen, as shown in Figure 4.10a. The attractions between these oppositely charged parts of the molecules result in hydrogen bonding. This is represented for the liquid and solid states of water in Figures 4.10b and c.

Let's look at some of the peculiar properties of water. As shown in Table 4.9, water has a normal boiling point of 100°C. This is much higher than would be predicted from the measured boiling points of compounds containing hydrogen and the other members of group VIA(16) of the periodic table. The boiling point of H_2Te is –2.2°C, of H_2Se is –41.3°C, and of H_2S is –60.3°C. Thus, it appears that the boiling point decreases with decreasing compound molecular weight. On the basis of this trend, water should boil at approximately –64°C. The 164-degree difference between predicted and measured boiling points is caused by strong hydrogen bonds between water molecules in the liquid. The other three compounds do not have strong hydrogen bonds between molecules because the electronegativites of the elements to which hydrogen is covalently bonded are too low. This deviation from prediction makes water a liquid at normal temperatures and allows it to be used in many of the ways familiar to us.

It is also common knowledge that solid water (ice) floats on liquid water (see Figure 4.11). Nothing peculiar here—or is there? In fact, there is. Water, like most other liquids, increases in density as it is cooled. This means that a specific mass of liquid decreases in volume as its temperature is lowered. Unlike most liquids, water behaves this way only until it reaches a temperature of about 4°C, as shown by the data in Table 4.10. Then, its density decreases as it is cooled further. When its temperature reaches 0°C, water freezes and, in the process, undergoes a dramatic decrease in density. Put another way, a quantity of liquid water at 0°C expands significantly when it freezes to solid at 0°C. Thus, a specific volume of ice has a lower mass than the same volume of liquid water, so the ice floats on the water. Hydrogen bonds between water molecules are, once again, responsible. The strong hydrogen bonds orient water molecules into a very open three-dimensional crystal lattice when it freezes (see Figure 4.10c). This open lattice of the solid occupies more space than is occupied by the molecules in the liquid state.

Table 4.10 Water Density as a Function of Temperature

Temperature (°C)	Water Density (g/mL)
100	0.9586
80	0.9719
60	0.9833
40	0.9923
20	0.9982
10	0.9997
5	0.9999
4 (actually 3.98)	1.0000
2	0.9999
0 (liquid H_2O)	0.9998
0 (solid H_2O)	0.9170

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Figure 4.11 Icebergs move as a result of wind and ocean currents. They can be very dangerous to ships, so their positions are reported and their probable courses estimated by an International Ice Patrol. The patrol was established in 1914 following the sinking of the *Titanic*.

Even though containers or engines might be ruptured or cracked if water is allowed to freeze in them, the overall effect of this characteristic of water is beneficial. If ice did not float, it would form on natural waters in the winter and sink to the bottom. Gradually, ponds, lakes, and other bodies of water would fill with ice. In the spring, the ice would melt from the top down because the heavier solid ice would stay under the water. As a result, most natural waters in areas that experience freezing winters would contain significant amounts of ice year-round. It would be a strange (and possibly hostile) environment for us and for wildlife.

The amount of heat required to melt a quantity of solid and the amount of heat required to vaporize a quantity of liquid also depend on the attractive forces between molecules. As expected, these are high for water, compared with the hydrogen compounds of the three other elements in group VIA(16). The beneficial results of these high values for water will be discussed in Chapter 6.

The forces between O_2 molecules, the lattice particles of solid oxygen, do not fit into any of the classifications we have discussed to this point. The O_2 molecule is not polar and contains no ionic or metallic bonds, and solid oxygen melts at much too low a temperature to fit into the category of a network solid. The forces between O_2 molecules are called **dispersion forces** and result from momentary nonsymmetric electron distributions in the molecules.

dispersion forces Very weak attractive forces acting between the particles of all matter. They result from momentary nonsymmetric electron distributions in molecules or atoms.

The nonbonding electrons in an O_2 molecule can be visualized as being uniformly distributed. However, there is a small statistical chance that, during their normal movement, more of the electrons will momentarily be on one side of a molecule than on the other. This condition causes the molecule to become dipolar for an instant. The resulting negative side of the dipole will tend to repel electrons of adjoining molecules and induce them also to become dipolar (they become induced dipoles). The original (statistical) dipole and all induced dipoles are then attracted to one another. This happens many times per second throughout the solid or liquid oxygen. The net effect is a weak dispersion force of attraction between the molecules. The net force is weak because it represents the average result of many weak, short-lived attractions per second between molecules.

Dispersion forces exist in all matter, but because they are very weak, their contribution is negligible when other, stronger forces are also present. Thus, solid water is held together by dispersion forces and hydrogen bonds, but the properties are almost entirely the result of the hydrogen bonds. The ease with which a dipole can be induced increases with the size of the particle (atoms or molecules). The larger the particle, the stronger the resulting dipolar attraction (dispersion force), and the harder to separate the particles by melting or boiling. Thus, we expect to find the melting and boiling points of the elements increase as we move down a group of the periodic table.

Table 4.11 The Normal Melting and Boiling Points of Group VIIA(17) Elements

Substance	Melting Point (°C)	Boiling Point (°C)
F_2	–223	–188
Cl_2	–103	–34.6
Br_2	–7.2	58.8
I_2	113.9	184.3

Example 4.19

Illustrate the behavior of dispersion forces by using the information in Table 4.11.

Solution

The molecules increase in size in the order F_2, Cl_2, Br_2, I_2. The strength of dispersion forces increases in the same order, as shown in Table 4.11.

Learning Check 4.19 Using only the periodic table, predict which member of each of the following pairs of elements would have the higher melting and boiling points:

a. O and Se **b.** Sb and P **c.** He and Ne

The melting and boiling points in Example 4.19 indicate the variation in magnitude of dispersion forces. Similar variations are found for the other forces described in this section; sometimes their strengths overlap. For example, some metallic bonds are weaker than ionic bonds, whereas others are stronger. This is illustrated in Figure 4.12, which summarizes the relative strengths of interparticle forces.

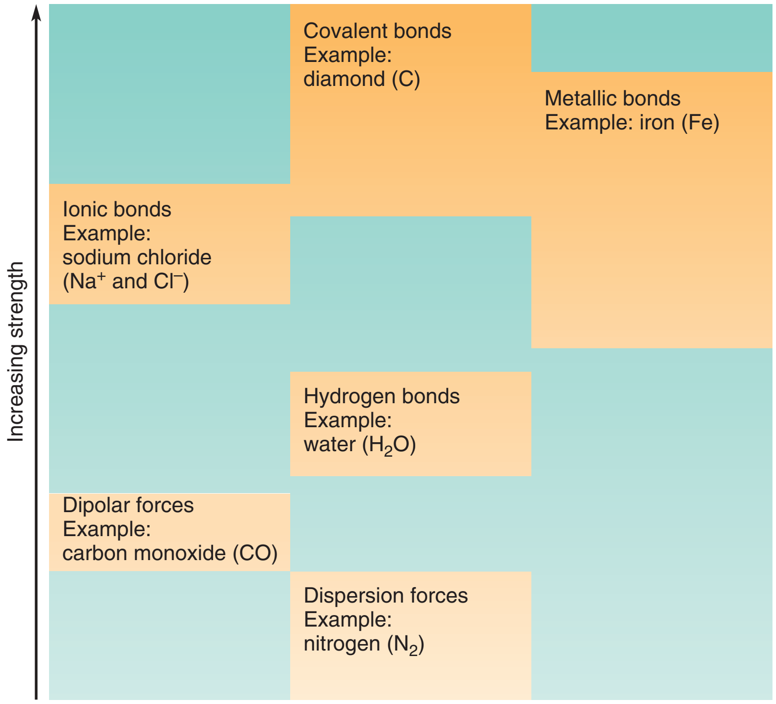


Figure 4.12 The relative strengths of interparticle forces.

Concept Summary

Noble Gas Configurations. The lack of reactivity for noble gases led to the proposal that the electronic configurations of the noble gases represented stable configurations. These configurations, usually consisting of eight electrons in the valence shell, can be represented in several useful ways.

Objective 1, Exercise 4.2

Ionic Bonding. Ionic compounds are formed when reacting atoms gain or lose electrons to achieve a noble gas configuration of eight electrons in the valence shell. This octet rule predicts that atoms will be changed into charged particles called simple ions. Ions of opposite charge are attracted to each other; the attractive force is called an ionic bond.

Objective 2, Exercise 4.12

Ionic Compounds. Oppositely charged ions group together to form compounds in ratios determined by the positive and negative charges of the ions. The formulas representing these ratios contain the lowest number of each ion possible in a proportion such that the total positive charges and total negative charges used are equal.

Objective 3, Exercises 4.20 and 4.22

Naming Binary Ionic Compounds. Binary ionic compounds contain a metal and a nonmetal. They are named by naming the metal, then adding the suffix *-ide* to the stem of the nonmetal.

Objective 4, Exercise 4.30

The Smallest Unit of Ionic Compounds. Ionic compounds do not exist in the form of molecules but as three-dimensional arrangements of oppositely charged ions. The sum of the atomic weights of the elements in the formula of ionic compounds is called the formula weight and is used in calculations somewhat like the molecular weight of molecular compounds.

Objective 5, Exercise 4.38

Covalent Bonding. Elements with little or no tendency to gain or lose electrons often react and achieve noble gas electronic configurations by sharing electrons. Lewis structures are useful in representing electron sharing. Shared pairs of electrons exert an attractive force on both atoms that share them. The atoms are held together by this attraction to form a covalent bond.

Objective 6, Exercise 4.48

Polyatomic Ions. Polyatomic ions are groups of two or more covalently bonded atoms that carry a net electrical charge. They are conveniently represented using Lewis structures.

Objective 7, Exercise 4.50

Shapes of Molecules and Polyatomic Ions. The shapes of many molecules and polyatomic ions can be predicted by using the valence-shell electron-pair repulsion theory (VSEPR). According to the VSEPR theory, electron pairs in the valence shell of the central atom of a molecule or ion repel one another and become arranged so as to maximize their separation distances. The resulting arrangement determines the molecular or ionic shape when one or all of the electron pairs involved form bonds between the central atom and other atoms.

Objective 8, Exercises 4.52 and 4.54

The Polarity of Covalent Molecules. Shared electrons may be shared equally, or they may be attracted more strongly to one of the atoms they bond together. The tendency of a covalently bonded atom to attract shared electrons is called the electronegativity of the atom. Unequally shared bonding-electron pairs form polar covalent bonds. The extent of bond polarization can be estimated from the electronegativity differences between the bonded atoms. The higher the electronegativity difference, the more polar (or ionic) is the bond. Polar covalent bonds cause partial positive and negative charges to form within molecules. When these charges are symmetrically distributed in the molecule, it is said to be nonpolar. An unsymmetric distribution gives rise to a polar molecule.

Objective 9, Exercises 4.58 and 4.64

More about Naming Compounds. Binary covalent compounds are named using the name of the less electronegative element first, followed by the stem plus *-ide* of the more electronegative element. Greek prefixes are used to represent the number of each type of atom in molecules of the compounds. Ionic compounds that contain a metal ion (or ammonium ion) plus a polyatomic ion are named by first naming the metal (or ammonium ion) followed by the name of the polyatomic ion.

Objective 10, Assessment Exercises 4.66, 4.70, and 4.72

Other Interparticle Forces. Forces other than ionic and covalent bonds are also known to hold the particles of some pure substances together in the solid and liquid states. These forces include metallic bonds, dipolar attractions, hydrogen bonds, and dispersion forces. The strength of the predominant force acting in a substance is indicated by the melting and boiling points of the substance.

Objective 11, Exercises 4.78 and 4.80

Key Terms and Concepts

Binary compound (4.3)
Bond polarization (4.9)
Covalent bond (4.6)
Crystal lattice (4.5)
Dipolar force (4.11)
Dispersion force (4.11)
Double bond (4.6)
Electronegativity (4.9)
Formula weight (4.5)
Hydrogen bonding (4.11)
Ionic bond (4.2)
Isoelectronic (4.2)
Lattice site (4.5)
Lewis structure (4.1)
Metallic bond (4.11)
Network solid (4.11)
Nonpolar covalent bond (4.9)
Nonpolar molecule (4.9)
Octet rule (4.2)
Polar covalent bond (4.9)
Polar molecule (4.9)
Polyatomic ion (4.7)
Simple ion (4.2)
Triple bond (4.6)
VSEPR theory (4.8)

Exercises

OWL Interactive versions of these problems are assignable in OWL.

Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

Noble Gas Configurations (Section 4.1)

4.1 Refer to the group numbers of the periodic table and draw Lewis structures for atoms of the following:

a. lithium

b. sodium

c. chlorine

d. boron

4.2 Refer to the group numbers of the periodic table and draw Lewis structures for atoms of the following:

a. arsenic

b. silicon

c. lead

d. barium

4.3 Write abbreviated electronic configurations for the following:

a. iodine

b. element number 38

c. As

d. phosphorus

4.4 Write abbreviated electronic configurations for the following:

a. element number 50

b. Se

c. cesium

d. iodine

4.5 Draw Lewis structures for the elements given in Exercise 4.3.

4.6 Draw Lewis structures for the elements given in Exercise 4.4.

4.7 Use the symbol E to represent an element in a general way and draw Lewis structures for atoms of the following:

a. Any group IA(1)element

b. Any group IVA(14)element

4.8 Use the symbol E to represent an element in a general way and draw Lewis structures for atoms of the following:

a. Any group IIIA(13) element

b. Any group VIA(16) element

Ionic Bonding (Section 4.2)

4.9 Indicate both the minimum number of electrons that would have to be added and the minimum number that would have to be removed to change the electronic configuration of each element listed in Exercise 4.3 to a noble gas configuration.

4.10 Indicate both the minimum number of electrons that would have to be added and the minimum number that would have to be removed to change the electronic configuration of each element listed in Exercise 4.4 to a noble gas configuration.

4.11 Use the periodic table and predict the number of electrons that will be lost or gained by the following elements as they change into simple ions. Write an equation using elemental symbols, ionic symbols, and electrons to represent each change.

a. Ca

b. aluminum

c. fluorine

d. element number 34

4.12 Use the periodic table and predict the number of electrons that will be lost or gained by the following elements as they change into simple ions. Write an equation using elemental symbols, ionic symbols, and electrons to represent each change.

a. Cs

b. oxygen

c. element number 7

d. iodine

4.13 Write a symbol for each of the following ions:

a. A bromine atom that has gained one electron

b. A sodium atom that has lost one electron

c. A sulfur atom that has gained two electrons

4.14 Write a symbol for each of the following ions:

a. A selenium atom that has gained two electrons

b. A rubidium atom that has lost one electron

c. An aluminium atom that has lost three electrons

4.15 Identify the element in period 2 that would form each of the following ions. E is used as a general symbol for an element.

a. E^-

b. E^{2+}

c. E^{3-}

d. E^+

4.16 Identify the element in period 3 that would form each of the following ions. E is used as a general symbol for an element.

a. E^{2-}

b. E^{3+}

c. E^+

d. E^-

4.17 Identify the noble gas that is isoelectronic with each of the following ions:

a. Mg^{2+}

b. Te^{2-}

c. N^{3-}

d. Be^{2+}

4.18 Identify the noble gas that is isoelectronic with each of the following ions:

a. Li^+

b. I^-

c. S^{2-}

d. Sr^{2+}

Ionic Compounds (Section 4.3)

4.19 Write equations to represent positive and negative ion formation for the following pairs of elements. Then write a formula for the ionic compound that results when the ions combine.

a. Mg and S

b. strontium and nitrogen

c. elements number 3 and 34

4.20 Write equations to represent positive and negative ion formation for the following pairs of elements. Then write a formula for the ionic compound that results when the ions combine.

a. Ba and F

b. potassium and bromine

c. elements number 13 and 35

4.21 Write the formula for the ionic compound formed from Sr^{2+} and each of the following ions:

a. S^{2-}

b. Br^-

c. N^{3-}

d. Cl^-

4.22 Write the formula for the ionic compound formed from Ba^{2+} and each of the following ions:

a. Te^{2-}

b. N^{3-}

c. F^-

d. P^{3-}

4.23 Classify each of the following as a binary compound or not a binary compound:

a. HF

b. OF_2

c. H_2SO_4

d. H_2S

e. $MgBr_2$

4.24 Classify each of the following as a binary compound or not a binary compound:

a. PbO_2

b. $CuCl_2$

c. KNO_3

d. Be_3N_2

e. $CaCO_3$

Naming Binary Ionic Compounds (Section 4.4)

4.25 Name the following metal ions:

a. Ca^{2+}

b. K^+

c. Al^{3+}

d. Rb^+

4.26 Name the following metal ions:

a. Li^+

b. Mg^{2+}

c. Ba^{2+}

d. Cs^+

4.27 Name the following nonmetal ions:

a. Cl^-

b. N^{3-}

c. S^{2-}

d. Se^{2-}

4.28 Name the following nonmetal ions:

a. Br^-

b. O^{2-}

c. P^{3-}

d. Te^{2-}

4.29 Name the following binary ionic compounds:

a. K_2O

b. $SrCl_2$

c. Al_2O_3

d. LiBr

e. CaS

4.30 Name the following binary ionic compounds:

a. MgO

d. CaS

c. ZnO

d. $AlCl_3$

e. Na_3N

4.31 Name the following binary ionic compounds, using a roman numeral to indicate the charge on the metal ion:

a. $CrCl_2$ and $CrCl_3$

b. CoS and Co_2S_3

c. FeO and Fe_2O_3

d. $PbCl_2$ and $PbCl_4$

4.32 Name the following binary ionic compounds using a roman numeral to indicate the charge on the metal ion:

a. SnS and SnS_2

b. $FeCl_2$ and $FeCl_3$

c. Cu_2O and CuO

d. AuCl and $AuCl_3$

4.33 Name the binary compounds of Exercise 4.31 by adding the endings *-ous* and *-ic* to indicate the lower and higher ionic charges of the metal ion in each pair of compounds. The non-English root for lead (Pb) is *plumb-*.

4.34 Name the binary compounds of Exercise 4.32 by adding the endings *-ous* and *-ic* to indicate the lower and higher ionic charges of the metal ion in each pair of compounds. The non-English root for gold (Au) is *aur-*, and that of tin (Sn) is *stann-*.

4.35 Write formulas for the following binary ionic compounds:

a. manganese(II) chloride

b. iron(III) sulfide

c. chromium(II) oxide

d. iron(II) bromide

e. tin(II) chloride

4.36 Write formulas for the following binary ionic compounds:

a. mercury(I) oxide

b. lead(II) oxide

c. platinum (IV) iodine

d. copper(I) nitride

e. cobalt(II) sulfide

The Smallest Unit of Ionic Compounds (Section 4.5)

4.37 Determine the formula weight in atomic mass units for each of the following binary ionic compounds:

a. Na_2O

b. FeO

c. PbS_2

d. $AlCl_3$

4.38 Determine the formula weight in atomic mass units for each of the following binary ionic compounds:

a. KF

b. Be_3N_2

c. Li_3P

d. Cu_2O

4.39 Identify the ions that would occupy lattice sites in a solid sample of each compound given in Exercise 4.37.

4.40 Identify the ions that would occupy lattice sites in a solid sample of each compound given in Exercise 4.38.

4.41 Calculate the mass in grams of positive ions and negative ions contained in 1 mol of each compound given in Exercise 4.37.

4.42 Calculate the mass in grams of positive ions and negative ions contained in 1 mol of each compound given in Exercise 4.38.

4.43 Calculate the number of positive ions and negative ions contained in 1.00 mol of each compound given in Exercise 4.37.

4.44 Calculate the number of positive ions and negative ions contained in 1.00 mol of each compound given in Exercise 4.38.

Covalent Bonding (Section 4.6)

4.45 Represent the following reaction using Lewis structures:

$$I + I \rightarrow I_2$$

4.46 Represent the following reaction using Lewis structures:

$$8S \rightarrow S_8 \text{ (the atoms form a ring)}$$

4.47 Represent the following molecules by Lewis structures:

a. HF

b. IBr

c. PH_3 (each H atom is bonded to the P atom)

d. $HClO_2$ (the O atoms are each bonded to the Cl, and the H is bonded to one of the O atoms)

4.48 Represent the following molecules by Lewis structures:

a. H_2S (each H atom is bonded to the S atom)

b. CIF

c. HBr

d. HClO (the H and Cl are each bonded to O)

Polyatomic Ions (Section 4.7)

4.49 Draw Lewis structures for the following polyatomic ions:

a. ClO_3^- (each O atom is bonded to the Cl atom)

b. CN^-

c. CO_3^{2-} (each O atom is bonded to the C atom)

4.50 Draw Lewis strucrures for the following polyatomic ions:

a. PH_4^+(each H atom is bonded to the P atom)

b. HPO_4^{2-} (each O atom is bonded to the P atom, and the H atom is bonded to an O atom)

c. HSO_4^-(each O atom is bonded to the S atom, and the H atom is bonded to an O atom)

Shapes of Molecules and Polyatomic Ions (Section 4.8)

4.51 Draw Lewis structures for the following molecules:

a. O_3 (the O atoms are bonded together, like beads on a string)

b. CS_2 (each S atom is bonded to the C atom)

c. SeO_2 (each O atom is bonded to the Se atom)

d. H_2SO_3 (each O atom is bonded to the S atom, and one H atom is bonded to each of the two O atoms)

4.52 Predict the shape of each of the following molecules by first drawing a Lewis structure, then applying the VSEPR theory:

a. H_2S (each H atom is bonded to the S atom)

b. PCl_3 (each Cl atom is bonded to the P atom)

c. OF_2 (each F atom is bonded to the O atom)

d. SnF_4 (each F atom is bonded to the Sn atom)

4.53 Predict the shape of each of the following molecules by first drawing a Lewis structure, then applying the VSEPR theory:

a. O_3 (see Exercise 4.51 for Lewis structure)

b. SeO_2 (see Exercise 4.51 for Lewis structure)

c. PH_3 (each H atom is bonded to the P atom)

d. SO_3 (each O atom is bonded to the S atom)

4.54 Predict the shape of each of the following polyatomic ions by first drawing a Lewis structure, then applying the VSEPR theory:

a. NO_2^- (each O is bonded to N)

b. ClO_3^- (each O is bonded to Cl)

c. CO_3^{2-} (each O is bonded to C)

d. H_3O^+ (each H is bonded to O)

Note the positive charge; compare with NH_4^+.

4.55 Predict the shape of each of the following polyatomic ions by first drawing a Lewis structure, then applying the VSEPR theory:

a. NH_2^- (each H is bonded to N)

b. PO_3^{3-} (each O is bonded to P)

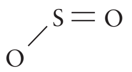
c. $BeCl_4^{2-}$ (each Cl is bonded to Be)

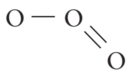
d. ClO_4^- (each O is bonded to Cl)

The Polarity of Covalent Molecules (Section 4.9)

4.56 Use the periodic table and Table 4.4 to determine which of the following bonds will be polarized. Show the resulting charge distribution in those molecules that contain polarized bonds.

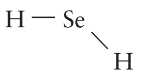
a. H—I

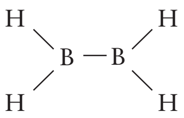
b. 

c. 

4.57 Use the periodic table and Table 4.4 to determine which of the following bonds will be polarized. Show the resulting charge distribution in those molecules that contain polarized bonds.

a. Cl—F

b. 

c. 

4.58 Use Table 4.4 and classify the bonds in the following compounds as nonpolar covalent, polar covalent, or ionic:

a. KI

b. NH_3

c. CO

d. CaO

e. NO

4.59 Use Table 4.4 and classify the bonds in the following compounds as nonpolar covalent, polar covalent, or ionic:

a. PCl_3

b. H_2Se

c. $MgCl_2$

d. BeI_2

e. NCl_3

4.60 On the basis of the charge distributions you drew for the molecules of Exercise 4.56, classify each of the molecules as polar or nonpolar.

4.61 On the basis of the charge distributions you drew for the molecules of Exercise 4.57, classify each of the molecules as polar or nonpolar.

4.62 Use Table 4.4 and predict the type of bond you would expect to find in compounds formed from the following elements:

a. nitrogen and oxygen

b. magnesium and oxygen

c. N and H

4.63 Use Table 4.4 and predict the type of bond you would expect to find in compounds formed from the following elements:

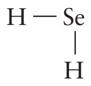
a. sulfur and oxygen

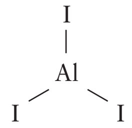
b. aluminum and bromine

c. C and Cl

4.64 Show the charge distribution in the following molecules, and predict which are polar molecules:

a. C≡O

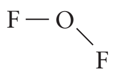
b. 

c. 

4.65 Show the charge distribution in the following molecules, and predict which are polar molecules:

a. S═C═S

b. H—C≡N

c. 

More about Naming Compounds (Section 4.10)

4.66 Name the following binary covalent compounds:

a. NCl_3

b. P_4O_6

c. BrCl

d. SF_4

e. ClO_2

4.67 Name the following binary covalent compounds:

a. SiO_2

b. SiF_4

c. P_2O_5

d. $AlBr_3$

e. CBr_4

4.68 Write formulas for the following binary covalent compounds:

a. selenium tetrafluoride

b. oxygen difluoride

c. dinitrogen monoxide

d. phosphorus trichloride

4.69 Write formulas for the following binary covalent compounds:

a. disulfur monoxide

b. sulfur hexafluoride

c. silicon tetrachloride

d. carbon diselenide

4.70 Write the formulas and names for compounds composed of ions of the following metals and the indicated polyatomic ions:

a. calcium and the hypochloritc ion

b. cesium and the nitrite ion

c. Mg and SO_3^{2-}

d. K and $Cr_2O_7^{2-}$

4.71 Write the formulas and names for compounds composed of ions of the following metals and the indicated polyatomic ions:

a. calcium and the phosphate ion

b. sodium and the dichromate ion

c. Li and CO_3^{2-}

d. Na and PO_4^{3-}

4.72 Write formulas for the following compounds:

a. magnesium hydroxide

b. calcium sulfite

c. ammonium phosphate

d. sodium hydrogen carbonate

e. barium sulfate

4.73 Write formulas for the following compounds:

a. potassium permanganate

b. calcium hydroxide

c. calcium phosphate

d. ammonium dihydrogen phosphate

e. calcium hypochlorite

4.74 Write a formula for the following compounds, using *M* with appropriate charges to represent the metal ion:

a. Any group IA(1) element and SO_3^{2-}

b. Any group IA(1) element and $C_2H_3O_2^-$

c. Any metal that forms M^{2+} ions and $Cr_2O_7^{2-}$

d. Any metal that forms M^{3+} ions and PO_4^{3-}

e. Any metal that forms M^{3+} ions and NO_3^-

4.75 Write a formula for the following compounds, using *M* with appropriate charges to represent the metal ion:

a. Any group IIA(2) element and HSO_3^-

b. Any group IIA(2) element and HPO_4^{2-}

c. Any metal that forms M^+ ions and NO_2^-

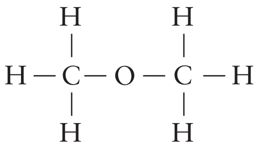
d. Any metal that forms M^{3+} ions and CO_3^{2-}

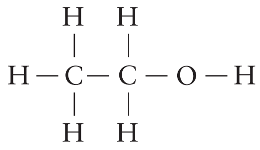
e. Any metal that forms M^{2+} ions and HPO_4^{2-}

Other Interparticle Forces (Section 4.11)

4.76 The covalent compounds ethyl alcohol and dimethyl ether both have the formula C_2H_6O. However, the alcohol melts at −117.3°C and boils at 78.5°C, whereas the ether melts at −138.5°C and boils at −23.7°C. How could differences in forces between molecules be used to explain these observations?

4.77 The following structural formulas represent molecules of ethyl alcohol and dimethyl ether. Assign the correct name to each formula and explain how your choice is consistent with your answer to Exercise 4.76.





4.78 Describe the predominant forces that exist between molecules of the noble gases. Arrange the noble gases in a predicted order of increasing boiling point (lowest first) and explain the reason for the order.

4.79 Use the concept of interparticle forces to propose an explanation for the fact that CO_2 is a soft, low-melting solid (dry ice), whereas SiO_2 is a hard solid (sand). Focus on the nature of the particles that occupy lattice sites in the solid.

4.80 Table sugar, sucrose, melts at about 185°C. Which interparticle forces do you think are unlikely to be the predominant ones in the lattice of solid sucrose?

4.81 The formula for sucrose is $C_{12}H_{22}O_{11}$, where many of the hydrogens and oxygens are combined to form OH groups that are bonded to carbon atoms. What type of predominant interparticle bonding would you now propose for solid sucrose (see Exercise 4.80)?

Additional Exercises

4.82 Three atoms with an electronic configuration of $1s^1$ are covalently bonded to one atom with an electronic configuration of $1s^2\,2s^2\,2p^3$ to form a molecule. What is the formula of this molecule? Make a drawing to show how two of these molecules would hydrogen bond with each other.

4.83 Suppose an element from group IIA(2), and period 3 of the periodic table forms an ionic compound with the element that has an electronic configuration of $1s^2\,2s^2\,2p^5$. What would be the formula of the compound, and what would be the name of the compound?

4.84 What would be the mass in grams of 0.200 moles of the ionic compound formed when magnesium metal reacts with oxygen?

4.85 The ampere unit is used to describe the flow of electricity in an electrical circuit. One ampere is an amount of electricity corresponding to the flow of 6.2×10^{18} electrons past a point in a circuit in 1 second. In a hydrogen fuel cell, hydrogen atoms are dissociated into H^+ ions and electrons ($H \rightarrow H^+ + 1e^-$). How many grams of hydrogen must be dissociated each second in a fuel cell to produce 1 ampere of electricity?

4.86 If one atom of oxygen reacted with two atoms of nitrogen to form a molecule, what would be the formula of the molecule? Use Table 4.4 and Figure 4.7 to determine if the bond between the atoms is ionic or covalent. Also, name the compound that is formed.

Allied Health Exam Connection

The following questions are from these sources:

1. *Nursing School Entrance Exam* © 2005, Learning Express, LLC.
2. *McGraw-Hill's Nursing School Entrance Exams* by Thomas A. Evangelist, Tamara B. Orr and Judy Unrein © 2009, The McGraw-Hill Companies, Inc.
3. *NSEE Nursing School Entrance Exams,* 3rd Edition © 2009, Kaplan Publishing.
4. *Cliffs Test Prep: Nursing School Entrance Exams* by Fred N. Grayson © 2004, Wiley Publishing, Inc.
5. *Peterson's Master the Nursing School and Allied Health Entrance Exams,* 18th Edition by Marion F. Gooding © 2008, Peterson's, a Nelnet Company.

4.87 Nobel gases:

a. have low boiling points
b. are all gases at room temperature
c. are also called inert gases
d. all of the above

4.88 The elements in group Zero of the periodic table are considered inert gases because each has how many electrons in its outermost energy level?

a. 8
b. 7
c. 4
d. 2

4.89 Name the type of bond that is formed when *electrons are shared* between two atoms.

a. shared bond
b. ionic bond
c. covalent bond
d. multiple bond

4.90 Which of the following is the correct name for Li_2SO_3?

a. lithium sulfite
b. lithium sulfide
c. lithium sulfate
d. lithium disulfate

4.91 An atom becomes an ion that possesses a negative charge. The atom must have:

a. gained protons
b. lost protons
b. lost electrons
d. gained electrons

4.92 When calcium reacts with chlorine to form calcium chloride, it:

a. shares two electrons
b. gains two electrons
c. loses two electrons
d. gains one electron

4.93 Which of the following is the probable charge for an ion formed from Ca?

a. $+1$
b. $+2$
c. -1
d. -2

4.94 A covalent bond is believed to be caused by a:

a. transfer of electrons
b. sharing of electrons
c. release of energy
d. none of the above

4.95 Which molecule below has a nonpolar bond in which the electrons are being shared equally?

a. H_2O
b. NH_3
c. Cl_2
d. CH_4

4.96 What is the formula for bismuth (III) hydroxide?

a. Bi_3OH
b. $Bi(OH)_3$
c. $Bi(OH)_2$
d. BiOH

4.97 Which of the following species will combine with a chloride ion to produce ammonium chloride?

a. NH_3
b. K^+
c. $NH_4{}^+$
d. Al^{+++}

4.98 What type of bond is created when bromine and magnesium are reacted to form $MgBr_2$?

a. polar covalent
b. metallic
c. ionic
d. nonpolar covalent

4.99 The parts of an atom directly involved in ionic bonding are the:

a. protons in the nucleus
b. neutrons in the nucleus
c. electrons in the outer energy shell
d. electrons in the innermost energy shell

4.100 In forming an ionic bond with an atom of chlorine, a sodium atom will:

a. receive 1 electron from the chlorine atom

b. receive 2 electrons from the chlorine atom

c. give up 1 electron to the chlorine atom

d. give up 2 electrons to the chlorine atom

4.101 In bonding, what would happen between the electrons of K and Br?

a. they would be transferred

b. they would be shared

c. none of the above

d. they would be both transferred and shared

4.102 Which compound contains a bond with **no** ionic character?

a. CO

b. CaO

c. K_2O

d. Na_2O

4.103 Which of the following molecules is nonpolar?

a. NH_3

b. H_2O, NO_2

c. PCl_3

d. N_2

4.104 Which molecule is nonpolar and contains a nonpolar covalent bond?

a. F_2

b. HI

c. KCl

d. NH_3

4.105 Which of the following is a nonpolar covalent bond?

a. the bond between two carbons

b. the bond between sodium and chloride

c. the bond between two water molecules

d. the bond between nitrogen and hydrogen

4.106 The type of bond that forms between two molecules of water is a:

a. polar covalent bond

b. hydrogen bond

c. nonpolar covalent bond

d. peptide bond

4.107 Which of the following is an example of hydrogen bonding?

a. The bond between O and H in a single molecule of water

b. The bond between the O of one water molecule and the H of a second water molecule

c. The bond between the O of one water molecule and the O of a second water molecule

d. The bond between the H of one water molecule and the H of a second water molecule

Chemistry for Thought

4.108 Refer to Figure 4.1, and answer the question in the caption. What other metals and nonmetals would you predict might react in a similar way?

4.109 The colors of some compounds, such as those shown in Figure 4.2, result from the presence of water in the compounds. Propose an experiment you could perform to see if this was true for the compounds shown in Figure 4.2.

4.110 Refer to Figure 4.6, and answer the question in the caption. Propose the shapes that would be assumed by a group of five balloons and by a group of six balloons.

4.111 Refer to Figure 4.9. Two of the compounds are highly colored (other than white). All the compounds consist of potassium and a polyatomic ion. If you have not yet done so, write formulas for the compounds and see if you can find a characteristic of the polyatomic ions of the colored compounds that is not found in the white compounds. Then refer to Table 4.7 and predict which of the other polyatomic anions would form colored compounds of potassium.

4.112 Recall how a metal atom changes to form a positively charged metal ion. How do you think the sizes of a metal atom and a positive ion of the same metal will compare?

4.113 Recall how a nonmetal atom such as chlorine changes to form a negatively charged ion. How do you think the size of a nonmetal atom and a negatively charged ion of the same nonmetal will compare?

4.114 Neon atoms do *not* combine to form Ne_2 molecules. Explain.

4.115 Refer to Figure 4.8, and answer the question in the caption. The balloon carries a negative charge. What would happen if a positively charged object was used in place of the balloon?

4.116 In Chemistry Around Us 4.2, NO was described as a vital biological molecule. Explain how NO forms when a fuel such as natural gas, CH_4, is burned in air at a high temperature.

Chemical Reactions

5

Medical technologists provide data to help physicians diagnose and treat patients. Much of these data come from analyses of body fluids. These analyses are performed by mixing body fluid samples with reagents that react with specific materials such as glucose. In this chapter, several different types of chemical reactions are presented, some of which are used in body fluid analysis.
Jim West/PhotoLibrary

Learning Objectives

When you have completed your study of this chapter, you should be able to:

1. Identify the reactants and products in written reaction equations, and balance the equations by inspection. **(Section 5.1)**
2. Assign oxidation numbers to elements in chemical formulas, and identify the oxidizing and reducing agents in redox reactions. **(Section 5.3)**
3. Classify reactions into the categories of redox or nonredox, then into the categories of decomposition, combination, single replacement, or double replacement. **(Sections 5.4–5.6)**
4. Write molecular equations in total ionic and net ionic forms. **(Section 5.7)**
5. Classify reactions as exothermic or endothermic. **(Section 5.8)**
6. Use the mole concept to do calculations based on chemical reaction equations. **(Section 5.9)**
7. Use the mole concept to do calculations based on the limiting-reactant principle. **(Section 5.10)**
8. Use the mole concept to do percentage-yield calculations. **(Section 5.11)**

OWL Online homework for this chapter may be assigned in OWL.

In previous chapters, we introduced the terms *molecule, element, compound,* and *chemical change.* In Chapter 1, you learned that chemical changes result in the transformation of one or more substances into one or more new substances. The processes involved in such changes are called *chemical reactions.* In this chapter, you will learn to write and read chemical equations that represent chemical reactions, to classify reactions, and to do calculations based on the application of the mole concept to chemical equations.

5.1 Chemical Equations

Learning Objective

1. Identify the reactants and products in written reaction equations, and balance the equations by inspection.

A simple chemical reaction between elemental hydrogen and oxygen has been used to power the engines of a number of spacecraft, including the space shuttle. The products are water and much heat. For the moment, we will focus only on the substances involved; we will deal with the heat later. The reaction can be represented by a word equation:

$$\text{hydrogen} + \text{oxygen} \rightarrow \text{water} \tag{5.1}$$

The word equation gives useful information, including the reactants and products of the reaction. The **reactants** are the substances that undergo the chemical change; by convention, these are written on the left side of the equation. The **products,** or substances produced by the chemical change, are written on the right side of the equation. The reactants and products are separated by an arrow that points to the products. A plus sign (+) is used to separate individual reactants and products.

reactants of a reaction The substances that undergo chemical change during the reaction. They are written on the left side of the equation representing the reaction.

products of a reaction The substances produced as a result of the reaction taking place. They are written on the right side of the equation representing the reaction.

Word equations convey useful information, but the chemical equations used by chemists convey much more:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) \tag{5.2}$$

In this equation, the reactants and products are represented by molecular formulas that tell much more than the names used in the word equation. For example, it is apparent that both hydrogen and oxygen molecules are diatomic. The equation is also consistent with a fundamental law of nature called the **law of conservation of matter.** According to this law, atoms are neither created nor destroyed in chemical reactions but are rearranged to form new substances. Thus, atoms are conserved in chemical reactions, but molecules are not. The numbers written as coefficients to the left of the molecular formulas make the equation consistent with this law by making the total number of each kind of atom equal in the reactants and products. Note that coefficients of 1 are never written but are understood. Equations written this way are said to be **balanced.** In addition, the symbol in parentheses to the right of each formula indicates the state or form in which the substance exists. Thus, in Equation 5.2, the reactants hydrogen and oxygen are both in the form of gases (g), and the product water is in the form of a liquid (ℓ). Other common symbols you will encounter are (s) to designate a solid and (aq) to designate a substance dissolved in water. The symbol (aq) comes from the first two letters of *aqua,* the Latin word for water.

law of conservation of matter Atoms are neither created nor destroyed in chemical reactions.

balanced equation An equation in which the number of atoms of each element in the reactants is the same as the number of atoms of that same element in the products.

Example 5.1

Determine the number of atoms of each type on each side of Equation 5.2.

Solution

The coefficient 2 written to the left of H_2 means that two hydrogen molecules with the formula H_2 are reacted. Since each molecule contains two hydrogen atoms, a total of four hydrogen atoms is represented. The coefficient to the left of O_2 is 1, even though it is not written in the equation. The oxygen molecule contains two atoms, so a total of two oxygen atoms is represented. The coefficient 2 written to the left of H_2O means that two molecules of H_2O are produced. Each molecule contains two hydrogen atoms and one oxygen atom. Therefore, the total number of hydrogen atoms represented is four, and the number of oxygen atoms is two. These are the same as the number of hydrogen and oxygen atoms represented in the reactants.

Learning Check 5.1 Determine the number of atoms of each type on each side of the following balanced equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

When the identity and formulas of the reactants and products of a reaction are known, the reaction can be balanced by applying the law of conservation of matter.

Example 5.2

Nitrogen dioxide (NO_2) is an air pollutant that is produced in part when nitric oxide (NO) reacts with oxygen gas (O_2). Write a balanced equation for the production of NO_2 by this reaction.

Solution

The reactants written to the left of the arrow will be NO and O_2. The product is NO_2:

$$NO(g) + O_2(g) \rightarrow NO_2(g) \tag{5.3}$$

A quick inspection shows that the reactants contain three oxygen atoms, whereas the product has just two. The number of nitrogen atoms is the same in both the reactants and the products. A practice sometimes resorted to by beginning chemistry students is to change the formula of oxygen gas to O by changing the subscript and to write

$$NO(g) + O(g) \rightarrow NO_2(g) \tag{5.4}$$

This is not allowed. The natural molecular formulas of any compounds or elements cannot be adjusted; they are fixed by the principles of chemical bond formation described in Chapter 4. All that can be done to balance a chemical equation is to change the coefficients of the reactants and products. In this case, inspection reveals that the following is a balanced form of the equation:

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g) \tag{5.5}$$

Notice that the following forms of the equation are also balanced:

$$4NO(g) + 2O_2(g) \rightarrow 4NO_2(g) \tag{5.6}$$

and

$$NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g) \tag{5.7}$$

In Equation 5.6, the coefficients are double those used in Equation 5.5; in Equation 5.7, one coefficient is a fraction. The lowest possible whole-number coefficients are used in balanced equations. Thus, Equation 5.5 is the correct form.

Learning Check 5.2 Write and balance an equation that represents the reaction of sulfur dioxide (SO_2) with oxygen gas (O_2) to give sulfur trioxide (SO_3).

5.2 Types of Reactions

A large number of chemical reactions are known to occur; only a relatively small number of them will be studied in this book. This study is made easier by classifying the reactions according to certain characteristics. Such a classification scheme could be developed in a

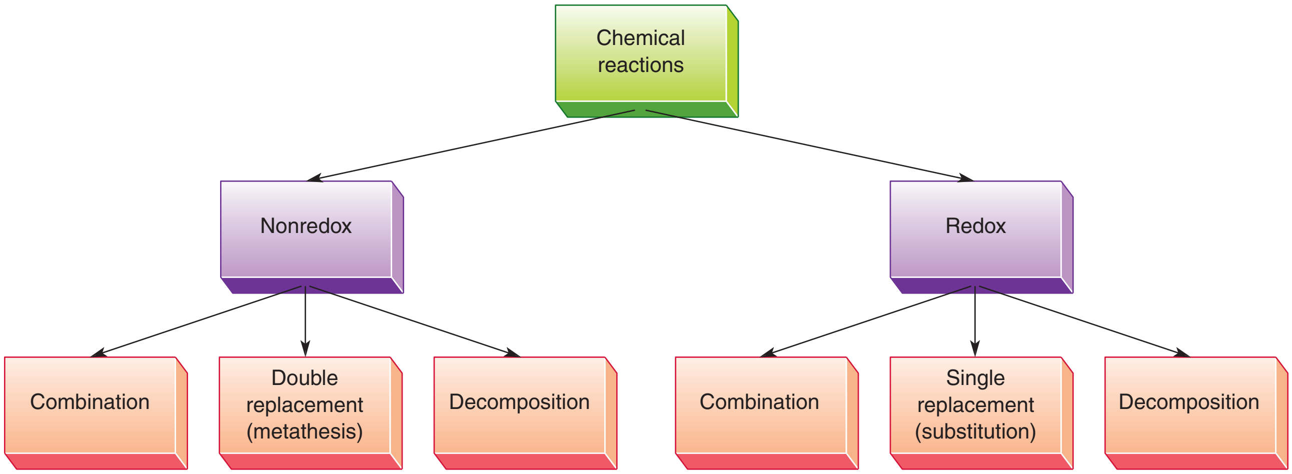


Figure 5.1 A classification of chemical reactions.

number of ways, but we have chosen to first classify reactions as being either oxidation–reduction (redox) reactions or nonredox reactions. As you will see, redox reactions are very important in numerous areas of study, including metabolism. Once reactions are classified as redox or nonredox, many can be further classified into one of several other categories, as shown in Figure 5.1. Notice that according to Figure 5.1, single-replacement, or substitution, reactions are redox reactions, whereas double-replacement, or metathesis, reactions are nonredox. Combination and decomposition reactions can be either redox or nonredox.

5.3 Redox Reactions

Learning Objective

2. Assign oxidation numbers to elements in chemical formulas, and identify the oxidizing and reducing agents in redox reactions.

Almost all elements react with oxygen to form oxides. The process is so common that the word *oxidation* was coined to describe it. Some examples are the rusting of iron,

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) \tag{5.8}$$

and the burning of hydrogen,

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) \tag{5.9}$$

The reverse process, reduction, originally referred to the technique of removing oxygen from metal oxide ores to produce the free metal. Some examples are

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(\ell) \tag{5.10}$$

and

$$2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g) \tag{5.11}$$

Today, the words **oxidation** and **reduction** are used in a rather broad sense. Table 5.1 contains most of the common meanings. To understand oxidation and reduction in terms of electron transfer and oxidation number (O.N.) change, you must become familiar with the concept of oxidation numbers.

oxidation Originally, a process involving a reaction with oxygen. Today it means a number of things, including a process in which electrons are given up, hydrogen is lost, or an oxidation number increases.

reduction Originally, a process in which oxygen was lost. Today it means a number of things, including a process in which electrons are gained, hydrogen is accepted, or an oxidation number decreases.

Oxidation numbers, sometimes called **oxidation states,** are positive or negative numbers assigned to the elements in chemical formulas according to a set of rules. The following rules will be used; be sure to note that Rule 1 applies only to uncombined elements—that is, elements in their free state. Rules 2 through 7 apply to elements combined to form compounds or ions.

oxidation numbers or oxidation states Positive or negative numbers assigned to the elements in chemical formulas according to a specific set of rules.

Rule 1. The oxidation number (O.N.) of any uncombined element is 0.
Examples: Al(0), O_2(0), Br_2(0), and Na(0)

Rule 2. The O.N. of a simple ion is equal to the charge on the ion.
Examples: $Na^+(+1)$, $Mg^{2+}(+2)$, $S^{2-}(-2)$, and $Br^-(-1)$

Rule 3. The O.N.s of group IA(1) and IIA(2) elements are +1 and +2, respectively.
Examples: Na_2CO_3(Na = +1), $Sr(NO_3)_2$ (Sr = +2), and $CaCl_2$ (Ca = +2)

Rule 4. The O.N. of hydrogen is +1.
Example: HCl (H = +1) and H_3PO_4 (H = +1)

Rule 5. The O.N. of oxygen is −2 except in peroxides, where it is −1.
Examples: CaO (O = −2), H_2SO_4 (O = −2), H_2O (O = −2), and H_2O_2 (O = −1)

Rule 6. The algebraic sum of the O.N.s of all atoms in a complete compound formula equals zero.

$$\begin{array}{llllll} \text{Examples: } K_2CO_3: & 2(\text{O.N. of K}) & + (\text{O.N. of C}) & + 3(\text{O.N. of O}) & = 0 \\ & 2(+1) & +4 & +3(-2) & = 0 \\ & +2 & +4 & +(-6) & = 0 \end{array}$$

$$\begin{array}{llllll} HNO_2: & (\text{O.N. of H}) & + (\text{O.N. of N}) & + 2(\text{O.N. of O}) & = 0 \\ & +1 & +3 & +2(-2) & = 0 \\ & +1 & +3 & +(-4) & = 0 \end{array}$$

Rule 7. The algebraic sum of the O.N.s of all atoms in a polyatomic ion is equal to the charge on the ion.

$$\begin{array}{lllll} \text{Examples: } MnO_4^-: & (\text{O.N. of Mn}) & + 4(\text{O.N. of O}) & = -1 \\ & +7 & +4(-2) & = -1 \\ & +7 & +(-8) & = -1 \end{array}$$

$$\begin{array}{llllll} HPO_4^{2-}: & (\text{O.N. of H}) & + (\text{O.N. of P}) & + 4(\text{O.N. of O}) & = -2 \\ & +1 & +5 & +4(-2) & = -2 \\ & +1 & +5 & +(-8) & = -2 \end{array}$$

Table 5.1 Common Uses of the Terms Oxidation and Reduction

Term	Meaning
Oxidation	To combine with oxygen To lose hydrogen To lose electrons To increase in oxidation number
Reduction	To lose oxygen To combine with hydrogen To gain electrons To decrease in oxidation number

Example 5.3

a. Assign O.N.s to the blue element in each of the following:

CO_2, NO_2^-, H_2O, N_2, K^+

b. Assign O.N.s to each element in the following:

CO_2, $Mg(NO_3)_2$, NO_3^-, CH_2O

Solution

a. CO_2, O = −2 (Rule 5); NO_2^-, O = −2 (Rule 5); H_2O, H = +1 (Rule 4); N_2, N = 0 (Rule 1); K^+, K = +1 (Rule 2).

b. CO_2: The O.N. of O is -2 (Rule 5), and the O.N. of C can be calculated by using Rule 6 as follows:

$$\begin{aligned}(\text{O.N. of C}) + 2(\text{O.N. of O}) &= 0\\ (\text{O.N. of C}) + 2(-2) &= 0\\ (\text{O.N. of C}) + (-4) &= 0\end{aligned}$$

Therefore,

$$\text{O.N. of C} = +4$$

$Mg(NO_3)_2$: The O.N. of Mg is $+2$ (Rule 3), the O.N. of O is -2 (Rule 5), and the O.N. of N can be calculated by using Rule 6 as follows:

$$\begin{aligned}(\text{O.N. of Mg}) + 2(\text{O.N. of N}) + 6(\text{O.N. of O}) &= 0\\ (+2) + 2(\text{O.N. of N}) + 6(-2) &= 0\\ 2(\text{O.N. of N}) + 2 + (-12) &= 0\\ 2(\text{O.N. of N}) - 10 &= 0\end{aligned}$$

Therefore,

$$\text{O.N. of N} = +5$$

NO_3^-: The O.N. of O is -2 (Rule 5), and the O.N. of N can be calculated by using Rule 7 as follows:

$$\begin{aligned}(\text{O.N. of N}) + 3(\text{O.N. of O}) &= -1\\ (\text{O.N. of N}) + 3(-2) &= -1\\ (\text{O.N. of N}) + (-6) &= -1\end{aligned}$$

Therefore,

$$\text{O.N. of N} = +5$$

Note that N has the same O.N. in NO_3^- and in $Mg(NO_3)_2$. This is expected because the polyatomic NO_3^- ion is present in $Mg(NO_3)_2$.

CH_2O: The O.N. of H is $+1$ (Rule 4), the O.N. of O is -2 (Rule 5), and the O.N. of C can be calculated by using Rule 6 as follows:

$$\begin{aligned}(\text{O.N. of C}) + 2(\text{O.N. of H}) + (\text{O.N. of O}) &= 0\\ (\text{O.N. of C}) + 2(+1) + (-2) &= 0\\ (\text{O.N. of C}) + 2 - 2 &= 0\end{aligned}$$

Therefore,

$$\text{O.N. of C} = 0$$

This example shows that an O.N. of 0 may be found for some combined elements as well as those in an uncombined state (Rule 1).

Learning Check 5.3 Assign O.N.s to each element in the following:

a. SO_3 **b.** $Ca(ClO_3)_2$ **c.** ClO_4^-

Now that you can assign O.N.s to the elements involved in reactions, you are ready to look at redox processes in terms of O.N.s and electron transfers. The reaction when sulfur is burned in oxygen is represented by the equation

$$S(s) + O_2(g) \rightarrow SO_2(g) \tag{5.12}$$

You can see that this reaction represents an oxidation because sulfur has combined with oxygen (remember Table 5.1). When oxidation numbers are assigned to sulfur, the reactant S has an oxidation number of 0, while the S in the SO_2 product has an oxidation number of $+4$.

Thus, the oxidation that has taken place resulted in an increase in the oxidation number of sulfur. Oxidation always corresponds to an increase in oxidation number.

The same conclusion is arrived at by thinking of oxidation numbers as charges. Thus, a sulfur atom with 0 charge is oxidized as it acquires a $+4$ charge. This change in charge results when the sulfur atom releases four electrons that have a -1 charge each. This leads to another definition of oxidation: Oxidation takes place when electrons are lost.

The oxygen of Equation 5.12 undergoes an oxidation number change from 0 in O_2 to -2 in SO_2. This decrease in oxidation number corresponds to a reduction of the oxygen. In terms of electrons, this same change is accomplished when each uncharged oxygen atom in the O_2 molecule accepts two electrons. Thus, the four electrons lost by sulfur as it was oxidized is the same total number accepted by the oxygen as it was reduced.

An oxidation process must always be accompanied by a reduction process, and all the electrons released during oxidation must be accepted during reduction. Thus, oxidation and reduction processes always take place simultaneously, hence the term redox. In redox reactions, the substance that is oxidized (and releases electrons) is called the **reducing agent** because it is responsible for reducing another substance. Similarly, the substance that is reduced (accepts electrons) is called the **oxidizing agent** because it is responsible for oxidizing another material (see Figure 5.2). These characteristics are summarized in Table 5.2.

A word of caution is appropriate here. Although electron transfers are a useful concept for understanding redox reactions involving covalent substances, it must be remembered that such transfers actually take place only during the formation of ionic compounds (see Section 4.3). In covalent substances, the electrons are actually shared (see Section 4.6). The oxidation number assignment rules that were given, as well as the electron-transfer idea, are based on the arbitrary practice of assigning shared electrons to the more electronegative element sharing them. However, it must be remembered that none of the atoms in covalent molecules actually acquire a net charge.

Figure 5.2 Combustion, the first reaction known to be carried out by humans, is a rapid redox reaction. Identify the oxidizing and reducing agents of the reaction.

reducing agent The substance that contains an element that is oxidized during a chemical reaction.

oxidizing agent The substance that contains an element that is reduced during a chemical reaction.

Example 5.4

Determine oxidation numbers for each atom represented in the following equations and identify the oxidizing and reducing agents:

a. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
b. $CO(g) + 3H_2(g) \rightarrow H_2O(g) + CH_4(g)$
c. $S_2O_8^{2-}(aq) + 2I^-(aq) \rightarrow I_2(aq) + 2SO_4^{2-}(aq)$

Solution

The O.N. under each elemental symbol was calculated by using the methods demonstrated in Example 5.3.

a. $\underset{0}{4Al} + \underset{0}{3O_2} \rightarrow \underset{+3}{2Al_2}\,\underset{-2}{O_3}$

The O.N. of Al has changed from 0 to $+3$. Therefore, Al has been oxidized and is the reducing agent. The O.N. of O has decreased from 0 to -2. The oxygen has been reduced and is the oxidizing agent.

Table 5.2 Properties of Oxidizing and Reducing Agents

Oxidizing Agent	Reducing Agent
Gains electrons	Loses electrons
Oxidation number decreases	Oxidation number increases
Becomes reduced	Becomes oxidized

At the Counter 5.1
Antiseptics and Disinfectants

Antiseptics and disinfectants are both used for the same purpose: to kill bacteria. The difference between these two categories of bacteria killers is where they are used. Antiseptics are used to kill bacteria on living tissue, such as wounds. Disinfectants are used to kill bacteria on inanimate objects. Some antiseptics, such as iodine and hydrogen peroxide, operate by oxidizing and thus destroying compounds essential to the normal functioning of the bacteria. A solution containing 3% hydrogen peroxide dissolved in water is an antiseptic found in most pharmacies, and it is often used to treat minor cuts and abrasions. A 2% solution of iodine dissolved in alcohol, called tincture of iodine, is also generally available, and it is used in a way similar to hydrogen peroxide. One disadvantage of the iodine solution is that it stains the skin a yellow-brown color.

Oxidizing antiseptics are often regarded as being too harsh. They may damage skin and other normal tissue, as well as kill the bacteria. For this reason, they have been replaced in many products by antiseptics derived from phenol. Water solutions of phenol, called carbolic acid, were first introduced as hospital antiseptics in 1867 by the English surgeon Joseph Lister. Before that time, antiseptics had not been used, and very few patients survived even minor surgery because of postoperative infections. These phenolic derivatives can often be recognized on ingredient labels by the characteristic *-ol* ending of their names. Some examples are thymol, eucalyptol, and eugenol.

Because disinfectants are used on inanimate objects, there is much less concern about the damage they might do to living tissue, and many of them are oxidizing agents. Sodium hypochlorite is one of the most widely used disinfectant compounds. In 5% solutions, it is marketed as liquid laundry bleach. This solution is an effective disinfectant for sinks, toilets, and similar fixtures. A chemically similar compound called calcium hypochlorite is the active ingredient in bleaching powder, and it is also used in hospitals as a disinfectant for clothing and bedding. Chlorine gas and ozone gas are two widely used, strong, oxidizing disinfectants. Their most well-known use is water treatment; they are added in small quantities to municipal water supplies to kill any harmful bacteria that may be present.

Antiseptics and disinfectants are both used to kill bacteria.

b. $\underset{+2}{C}\,\underset{-2}{O} + 3\underset{0}{H_2} \rightarrow \underset{+1}{H_2}\,\underset{-2}{O} + \underset{-4}{C}\,\underset{+1}{H_4}$

The O.N. of H_2 increased from 0 to +1. H_2 has been oxidized and is the reducing agent. The O.N. of C has decreased from +2 to −4. Carbon has been reduced and could be called the oxidizing agent. However, when one element in a molecule or ion is the oxidizing or reducing agent, the convention is to refer to the entire molecule or ion by the appropriate term. Thus, carbon monoxide (CO) is the oxidizing agent.

c. $\underset{+7}{S_2}\,\underset{-2}{O_8}^{2-} + 2\underset{-1}{I^-} \rightarrow \underset{0}{I_2} + 2\underset{+6}{S}\,\underset{-2}{O_4}^{2-}$

The O.N. of I has increased from −1 to 0, and the I^- ion is the reducing agent. The O.N. of S in $S_2O_8^{2-}$ has decreased from +7 to +6. Thus, S has been reduced, and the $S_2O_8^{2-}$ is the oxidizing agent.

Learning Check 5.4 Assign oxidation numbers to each atom represented in the following equations and identify the oxidizing and reducing agents:

a. $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
b. $2KI(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + I_2(aq)$
c. $IO_3^-(aq) + 3HSO_3^-(aq) \rightarrow I^-(aq) + 3HSO_4^-(aq)$

5.4 Decomposition Reactions

Learning Objective

3. Classify reactions into the categories of redox or nonredox, then into the categories of decomposition, combination, single replacement, or double replacement.

In **decomposition reactions,** a single substance is broken down to form two or more simpler substances, as shown in Figure 5.3. In this and other "box" representations of reactions, the number of molecules in the boxes will not match the coefficients of the reaction equation. However, they will be in the correct proportions. In Figure 5.3, for example, the box on the left contains the same number of H_2O_2 molecules as the number of H_2O molecules on the right. Also, the number of H_2O molecules on the right is twice the number of O_2 molecules. The general form of the equation for a decomposition reaction is

$$A \rightarrow B + C \tag{5.13}$$

Some decomposition reactions are also redox reactions, whereas others are not. Examples of decomposition reactions are given by Equations 5.14 and 5.15:

$$2HgO(s) \rightarrow 2Hg(\ell) + O_2(g) \tag{5.14}$$

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g) \tag{5.15}$$

Equation 5.14 represents the redox reaction that takes place when mercury(II) oxide (HgO) is heated. Mercury metal (Hg) and oxygen gas (O_2) are the products. This reaction was used by Joseph Priestley in 1774 when he discovered oxygen. Equation 5.15 represents a nonredox reaction used commercially to produce lime(CaO) by heating limestone ($CaCO_3$) to a high temperature. The decomposition of H_2O_2 represented by Figure 5.3 is shown in Figure 5.4.

decomposition reaction A chemical reaction in which a single substance reacts to form two or more simpler substances.

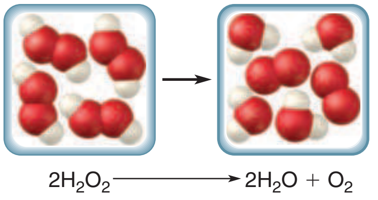


Figure 5.3 A decomposition reaction.

5.5 Combination Reactions

Learning Objective

3. Classify reactions into the categories of redox or nonredox, then into the categories of decomposition, combination, single replacement, or double replacement.

Combination reactions are sometimes called *addition* or *synthesis reactions.* Their characteristic is that two or more substances react to form a single substance (see

combination reaction A chemical reaction in which two or more substances react to form a single substance.

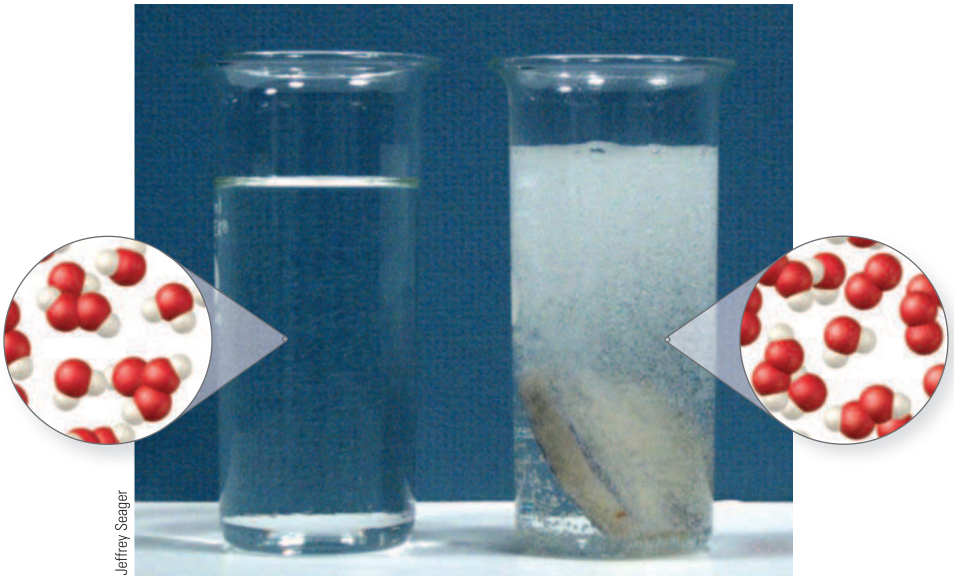


Figure 5.4 The decomposition of hydrogen peroxide. A solution of hydrogen peroxide (H_2O_2) in water (left) does not decompose rapidly until an enzyme catalyst from a piece of freshly-cut potato is added (right). How can you tell that one of the products of the reaction is a gas?

Figure 5.5 A combination reaction.

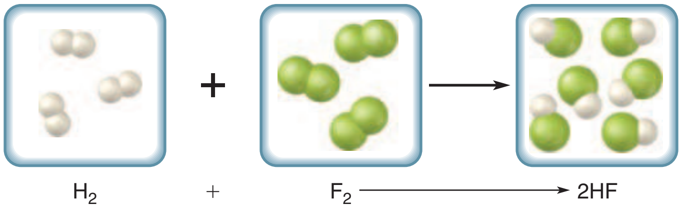


Figure 5.6 Magnesium metal burns in air to form magnesium oxide.

Figure 5.5). The reactants can be any combination of elements or compounds, but the product is always a compound. The general form of the equation is

$$A + B \rightarrow C \tag{5.16}$$

At high temperatures, a number of metals will burn and give off very bright light. This burning is a redox combination reaction, represented for magnesium metal by Equation 5.17 and shown in Figure 5.6.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s) \tag{5.17}$$

A nonredox combination reaction that takes place in the atmosphere contributes to the acid rain problem. An air pollutant called sulfur trioxide (SO_3) reacts with water vapor and forms sulfuric acid. The reaction is represented by Equation 5.18:

$$SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(aq) \tag{5.18}$$

5.6 Replacement Reactions

Learning Objective

3. Classify reactions into the categories of redox or nonredox, then into the categories of decomposition, combination, single replacement, or double replacement.

single-replacement reaction A chemical reaction in which an element reacts with a compound and displaces another element from the compound.

Single-replacement reactions, also called *substitution reactions,* are always redox reactions and take place when one element reacts with a compound and displaces another element from the compound. A single-replacement reaction is represented in Figure 5.7 and demonstrated in Figure 5.8. The general equation for the reaction is shown in Equation 5.19.

$$A + BX \rightarrow B + AX \tag{5.19}$$

This type of reaction is useful in a number of processes used to obtain metals from their oxide ores. Iron, for example, can be obtained by reacting iron(III) oxide ore (Fe_2O_3) with carbon. The carbon displaces the iron from the oxide, and carbon dioxide is formed. The equation for the reaction is

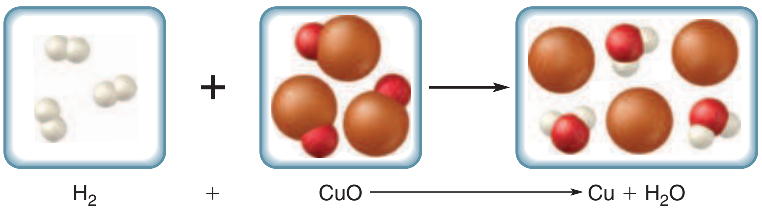
$$3C(s) + 2Fe_2O_3(s) \rightarrow 4Fe(s) + 3CO_2(g) \tag{5.20}$$

double-replacement reaction A chemical reaction in which two compounds react and exchange partners to form two new compounds.

Double-replacement reactions, also called *metathesis reactions,* are never redox reactions. These reactions, represented in Figure 5.9, often take place between substances dissolved in water. The general Equation 5.21 shows the partner-swapping characteristic of these reactions:

$$AX + BY \rightarrow BX + AY \tag{5.21}$$

Figure 5.7 A single-replacement reaction.



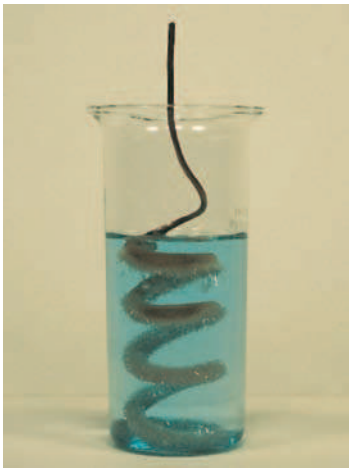


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Figure 5.8 When a piece of copper wire (Cu) is placed in a solution of silver nitrate ($AgNO_3$) in water, crystals of silver metal (Ag) form on the wire, and the liquid solution that was originally colorless turns blue as copper nitrate [$Cu(NO_3)_2$] forms in solution. Write a balanced equation for this single-replacement reaction.

The reaction that takes place when a base is used to neutralize an acid is a good example of a double-replacement reaction:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(\ell) \tag{5.22}$$

Example 5.5

Classify each of the reactions represented by the following equations as redox or nonredox. Further classify them as decomposition, combination, single-replacement, or double-replacement reactions.

a. $SO_2(g) + H_2O(\ell) \rightarrow H_2SO_3(aq)$
b. $2K(s) + 2H_2O(\ell) \rightarrow 2KOH(aq) + H_2(g)$
c. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
d. $BaCl_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + 2NaCl(aq)$

Solution

The O.N. under each elemental symbol was calculated by using the methods demonstrated in Example 5.3.

a. $\underset{+4}{S}\ \underset{-2}{O_2} + \underset{+1}{H_2}\ \underset{-2}{O} \rightarrow \underset{+1}{H_2}\ \underset{+4}{S}\ \underset{-2}{O_3}$

No O.N. changes take place; therefore, the reaction is nonredox. Because two substances combine to form a third, the reaction is a combination reaction.

b. $2\underset{0}{K} + 2\underset{+1}{H_2}\ \underset{-2}{O} \rightarrow 2\underset{+1}{K}\ \underset{-2}{O}\ \underset{+1}{H} + \underset{0}{H_2}$

The O.N. of K increases from 0 to +1 and that of H decreases from +1 to 0. The reaction is a redox reaction. Because K displaces H, it is a single-replacement reaction.

c. $\underset{0}{N_2} + 3\underset{0}{H_2} \rightarrow 2\underset{-3}{N}\ \underset{+1}{H_3}$

The O.N.s of both N and H change; therefore, the reaction is a redox reaction. Two substances combine to form a third, so it is a combination reaction.

d. $\underset{+2}{Ba}\ \underset{-1}{Cl_2} + \underset{+1}{Na_2}\ \underset{+4}{C}\ \underset{-2}{O_3} \rightarrow \underset{+2}{Ba}\ \underset{+4}{C}\ \underset{-2}{O_3} + 2\underset{+1}{Na}\ \underset{-1}{Cl}$

No changes in O.N. occur. The reaction is nonredox and is an example of a double-replacement, or partner-swapping, reaction.

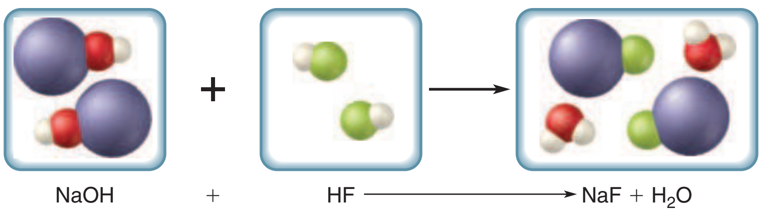

Figure 5.9 A double-replacement, or metathesis, reaction.

Chemistry and Your Health 5.1

The Importance of Color in Your Diet

Scientific evidence accumulated during the 1990s suggested that diets rich in fruits and vegetables had a protective effect against a number of different types of cancer. Studies showed that simply increasing the levels of vitamins and minerals in the diet did not provide the increased protection. This led to research into the nature of other substances found in fruits and vegetables that are important for good health. As a result of this research, a number of chemical compounds found in plants and called *phytonutrients* have been shown to be involved in the maintenance of healthy tissues and organs. The mechanism for their beneficial action in the body is not understood for all phytonutrients, but a significant number are known to work as antioxidants that stop harmful oxidation reactions from occurring.

The colors of fruits and vegetables help identify those containing beneficial compounds. The table below contains a list of some of the more well-known phytonutrients together with sources, colors, and beneficial actions. The amount of evidence supporting the existence of benefits from phytonutrients is not the same for all those listed in the table. In some cases, the experimental evidence is extensive (e.g., the cancer-blocking behavior of isothiocyanates), while in other cases the listed benefits are based on a limited amount of research and more studies are being done (e.g., the contribution to eye health by anthocyanins).

Fruit/Vegetable Color	Fruit/Vegetable Examples	Phytonutrients	Possible Benefits
Red  © iStockphoto.com/DNY59	Tomatoes, watermelon, pink grapefruit	Lycopene (a carotenoid)	Protect against prostate, cervical, and pancreatic cancer and heart and lung disease
Red/purple  © iStockphoto.com/Michael Hill	Red and blue grapes, blueberries, strawberries, beets, eggplant, red cabbage, red peppers, plums, red apples	Anthocyanins (flavonoids)	Antioxidants; block formation of blood clots and help maintain good eye health
Orange  © iStockphoto.com/kgfoto	Carrots, mangoes, sweet potatoes, cantaloupe, winter squash	Alpha- and beta-carotenes	Cancer fighters; protect skin against free radicals, promote repair of damaged DNA
Orange/yellow  © iStockphoto.com/Doug Cannell	Oranges, peaches, papaya, nectarines	Beta-cryptoxanthin	May help prevent heart disease
Yellow/green  © iStockphoto.com/eyewave	Spinach, collards, corn, green peas, avocado, honeydew	Lutein and zeaxanthin (both are carotenoids)	Reduce risk of cataracts and age-related macular degeneration
Green  © iStockphoto.com/Dmitry Margolin	Broccoli, brussels sprouts, cabbage, kale, bok choy	Sulforaphane, isothiocyanates, and indoles	Cancer blocking
White/green  © iStockphoto.com/Ashok Rodrigues	Onions, leeks, garlic, celery, asparagus, pears, green grapes, cauliflower, mushrooms	Allicin (in onions) and flavanoids	Antitumor agent; antioxidants

Learning Check 5.5 Classify each of the reactions represented by the following equations as redox or nonredox. Further classify them as decomposition, combination, single-replacement, or double-replacement reactions.

a. $2HI(g) \rightarrow H_2(g) + I_2(g)$
b. $2H_2O_2(aq) \rightarrow 2H_2O(\ell) + O_2(g)$ (NOTE: H_2O_2 is a peroxide)
c. $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
d. $4P(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$
e. $2NaI(aq) + Cl_2(aq) \rightarrow 2NaCl(aq) + I_2(aq)$

5.7 Ionic Equations

Learning Objective

4. Write molecular equations in total ionic and net ionic forms.

Many of the reactions of interest in the course you are taking occur between compounds or elements dissolved in water. Ionic compounds and some polar covalent compounds break apart (dissociate) into ions when they are dissolved in water. Thus, a water solution of sodium hydroxide (NaOH), an ionic compound, does not contain molecules of NaOH but, rather, contains equal numbers of sodium ions (Na^+) and hydroxide ions (OH^-). Covalently bonded hydrogen chloride, HCl, dissolves readily in water to form H^+ and Cl^- ions. Equations for reactions between substances that form ions in solution can be written in several ways. For example, Equation 5.22 contains three substances that form ions, HCl, NaOH, and NaCl. Equation 5.22 is written in the form of a **molecular equation** in which each compound is represented by its formula. This same reaction, when represented by a **total ionic equation,** becomes

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(\ell) \quad (5.23)$$

In this equation each ionic compound is shown dissociated into ions, the form it takes when it is dissolved in water. Some of the ions appear as both reactants and products. These so-called **spectator ions** do not actually undergo any changes in the reaction. Because of this, they are dropped from the equation when it is written as a **net ionic equation:**

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell) \quad (5.24)$$

The net ionic equation makes the partner-swapping characteristics of this double-replacement reaction less obvious, but it does emphasize the actual chemical changes that take place. Figure 5.10 shows an experiment in which an ionic reaction occurs.

molecular equation An equation written with each compound represented by its formula.

total ionic equation An equation written with all soluble ionic substances represented by the ions they form in solution.

spectator ions The ions in a total ionic reaction that are not changed as the reaction proceeds. They appear in identical forms on the left and right sides of the equation.

net ionic equation An equation that contains only un-ionized or insoluble materials and ions that undergo changes as the reaction proceeds. All spectator ions are eliminated.

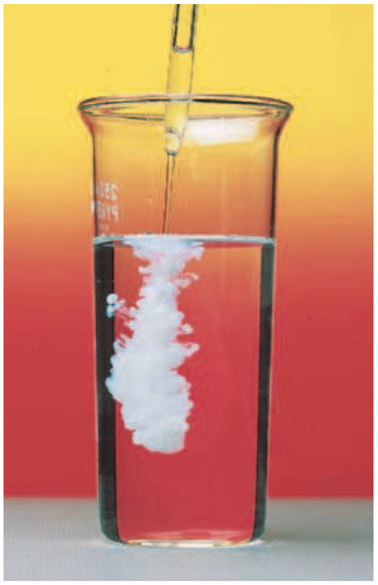

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Figure 5.10 The liquid in the large container is a solution of solid sodium chloride (NaCl) in water. The liquid being added is a solution of solid silver nitrate ($AgNO_3$) in water. When the two liquids are mixed, an insoluble white solid forms. The solid is silver chloride (AgCl). Write the molecular, total ionic, and net ionic equations for the reaction.

Example 5.6

Write equations for the following double-replacement reaction in total ionic and net ionic forms. Note that barium sulfate ($BaSO_4$) does not dissolve in water and should not be written in dissociated form. All other compounds are ionic and soluble in water.

$$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

Solution

Total ionic:

$$2Na^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2Cl^-(aq) \rightarrow BaSO_4(s) + 2Na^+(aq) + 2Cl^-(aq)$$

Net ionic:

$$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$$

The Na^+ and Cl^- ions appear in equal numbers as reactants and products. Thus, they are spectator ions and are not shown in the net ionic equation.

Learning Check 5.6 Write equations for the following reactions in total ionic and net ionic forms. Consider all ionic compounds to be soluble except $CaCO_3$ and $BaSO_4$, and remember that covalent molecules do not form ions when they dissolve.

a. $2NaI(aq) + Cl_2(aq) \rightarrow 2NaCl(aq) + I_2(aq)$
b. $CaCl_2(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CaCO_3(s)$
c. $Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow 2H_2O(\ell) + BaSO_4(s)$

5.8 Energy and Reactions

Learning Objective

5. Classify reactions as exothermic or endothermic.

Besides changes in composition, energy changes accompany all chemical reactions. The equation for the reaction between hydrogen and oxygen given at the beginning of this chapter (Equation 5.2) can also be written

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g) + \text{energy} \qquad (5.25)$$

Most of the energy of this reaction appears as heat, but the energy released or absorbed during chemical changes can take many forms including sound, electricity, light, high-energy chemical bonds, and motion. Some of these forms are discussed in detail in later chapters, but here the focus will be on energy in general, and it will be expressed in calories or joules, as if it all took the form of heat.

On the basis of heat, Equation 5.25 can be written

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g) + 115.6 \text{ kcal } (483.7 \text{ kJ}) \qquad (5.26)$$

This equation assumes that the hydrogen and oxygen are gases when they react and that the water produced is also a gas (vapor). According to Equation 5.26, whenever 2 mol of water vapor is formed from 2 mol of hydrogen gas and 1 mol of oxygen gas, 115.6 kcal, or 483.7 kJ, of energy is also released.

exothermic reaction A reaction that liberates heat.

endothermic reaction A reaction that absorbs heat.

The reaction represented by Equation 5.26 is an example of an **exothermic** (heat out) **reaction,** in which heat is released as the reaction takes place. In **endothermic** (heat in) **reactions,** heat is absorbed. A reaction of this type is utilized in emergency cold packs, which consist of a small plastic pouch of liquid sealed inside a larger pouch that also contains a solid substance. When the large pouch is squeezed firmly, the small pouch of liquid breaks, the liquid and solid react, heat is absorbed, and the liquid contents of the larger pouch become cold. This reaction is described in more detail in Chapter 7.

Chemistry Around Us 5.1

Ozone: The Good and The Bad

Ozone, a colorless gas with the characteristic odor of an electrical discharge or a burned-out electrical motor, has the simple molecular formula of O_3. As the title indicates, ozone has a dual personality. In one setting, it is a necessary protector of life on Earth. In another setting, it is a serious health hazard. Location is the key.

A part of the earth's upper atmosphere called the stratosphere extends from about 15 km (10 miles) to 50 km (30 miles) above the earth's surface. At this altitude, high-energy ultraviolet light from the sun (UV_1) causes the bond between oxygen atoms in O_2 to break and converts ordinary oxygen molecules into oxygen atoms (Equation 1). The oxygen atoms are very reactive

$$O_2(g) \xrightarrow{UV_1} 2O(g) \qquad \text{(Eq. 1)}$$

and combine with other oxygen molecules to form ozone (Equation 2).

$$O(g) + O_2(g) \longrightarrow O_3(g) \qquad \text{(Eq. 2)}$$

The ozone produced by this reaction is changed back into oxygen atoms and oxygen gas (Equation 3) by absorbing high energy ultraviolet light from the sun (UV_2) that is of a different energy than the light involved in Equation 1.

$$O_3(g) \xrightarrow{UV_2} O(g) + O_2(g) \qquad \text{(Eq. 3)}$$

As a result of these processes, most of the harmful high-energy ultraviolet light from the sun that reaches the upper atmosphere is absorbed and never reaches the Earth's surface where it would cause serious injury to most plant and animal life.

Ozone also forms at the Earth's surface, but it is produced by reactions involving oxygen gas, nitrogen dioxide gas (NO_2) from the exhaust of internal combustion engines, and ultraviolet light from the sun that was not absorbed by the reactions in the upper atmosphere. The absorption of this ultraviolet light (designated UV_3) causes NO_2 to break apart into NO gas and oxygen atoms (Equation 4). The oxygen atoms produced by this reaction are very reactive and combine with

$$NO_2(g) \xrightarrow{UV_3} NO(g) + O(g) \qquad \text{(Eq. 4)}$$

oxygen molecules to form ozone, just as they did in the upper atmosphere (Equation 5).

$$O(g) + O_2(g) \longrightarrow O_3(g) \qquad \text{(Eq. 5)}$$

As we have seen, ozone produced in this way in the upper atmosphere provides a protective blanket that keeps harmful ultraviolet light from reaching the Earth's surface. However, ozone formed at ground level causes numerous problems. Ozone is a very strong oxidizing agent that reacts aggressively with other molecules. As a result, it causes deterioration, discoloration, and other physical and chemical changes in materials. When ozone reacts with the molecules of living organisms, serious damage and health problems often result. The most serious health problem for humans is damage to lung tissue. Research shows that prolonged exposure to ozone diminishes lung function and increases the incidence of negative respiratory symptoms. Young children whose lungs are still developing and individuals suffering from asthma are especially vulnerable to damage from exposure to ozone.

With ozone, location is everything. The good ozone is formed in the upper atmosphere and protects life on the Earth's surface from harmful high-energy ultraviolet radiation from the sun. The bad ozone is formed at ground level and behaves as an oxidizing agent that damages materials and causes serious health problems for susceptible individuals.

Ozone, a common component of smog, presents serious health risks to susceptible individuals.

5.9 The Mole and Chemical Equations

Learning Objective

6. Use the mole concept to do calculations based on chemical reaction equations.

The mole concept introduced in Section 2.6 and applied to chemical formulas in Section 2.7 can also be used to calculate mass relationships in chemical reactions. The study of such mass relationships is called **stoichiometry,** a word derived from the Greek *stoicheion* (element) and *metron* (measure).

stoichiometry The study of mass relationships in chemical reactions.

Stoichiometry calculations require that balanced equations be used for the reactions being studied. Consider the following equation for the redox reaction that takes place when methane (CH_4), a major constituent of natural gas, is burned:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) \qquad (5.27)$$

Remember, coefficients in balanced equations refer to the formula that follows the coefficient. With this in mind, note that the following statement is consistent with this balanced equation:

1. 1 CH_4 molecule + 2 O_2 molecules → 1 CO_2 molecule + 2 H_2O molecules

This statement indicates that one molecule of CH_4 gas reacts with two molecules of O_2 gas and produces one molecule of CO_2 gas plus two molecules of H_2O liquid. Put another way, according to this reaction equation, the number of O_2 molecules that react with CH_4 will be equal to twice the number of CH_4 molecules that react. We remember that one mole of any kind of molecule is equal to 6.02×10^{23} molecules. Thus, one mole of CH_4 is equal to 6.02×10^{23} CH_4 molecules. As shown earlier, the number of O_2 molecules that will react with CH_4 will be twice the number of CH_4 molecules that react. Thus, the number of O_2 molecules that will react with 6.02×10^{23} CH_4 molecules is $2(6.02 \times 10^{23})$. Remembering that 6.02×10^{23} molecules is 1 mole of molecules leads to the conclusion that 1 mole of CH_4 molecules will react with 2 moles of O_2 molecules. We note that the numbers 1 and 2 are identical to the coefficients of the CH_4 and O_2 in the balanced equation. This leads to the conclusion that the coefficients in reaction equations can be interpreted as the number of moles of reactants and products involved in the reaction, which leads to the following statement for the reaction of CH_4 with O_2:

2. 1 mol $CH_4(g)$ + 2 mol $O_2(g)$ → 1 mol $CO_2(g)$ + 2 mol $H_2O(\ell)$

In general, any balanced reaction equation can be interpreted this way, where the coefficients of the equation become the number of moles of reactants and products involved in the reaction.

The fact used earlier that 1 mole of any substance is equal to 6.02×10^{23} particles of the substance leads to statement 3 given below for the reaction of CH_4 with O_2:

3. 6.02×10^{23} CH_4 molecules + $2(6.02 \times 10^{23})$ O_2 molecules →
6.02×10^{23} CO_2 molecules + $2(6.02 \times 10^{23})$ H_2O molecules

Application of the fact that one mole of a compound has a mass in grams equal to the molecular weight of the compound leads to statement 4 given below for the reaction:

4. 16.0 g CH_4 + 2(32.0) g O_2 → 44.0 g CO_2 + 2(18.0) g H_2O

Statements 2, 3 and 4 are all based on the mole definition, and are very useful in solving numerical problems involving balanced reaction equations and the factor-unit method described earlier in Section 1.9. Any two quantities from statements 2, 3 and 4 can be used to form factors that can be used to solve problems. For example, the following factors are just four of the many that can be obtained from statements 2, 3 and 4 by combining various quantities from the statements:

$$\frac{2\text{ mol }O_2}{1\text{ mol }CH_4} \qquad \frac{16.0\text{ g }CH_4}{44.0\text{ g }CO_2} \qquad \frac{2(18.0)\text{ g }H_2O}{6.02 \times 10^{23}\ CH_4\text{ molecules}} \qquad \frac{1\text{ mol }CH_4}{44.0\text{ g }CO_2}$$

Suppose you were asked this question: How many moles of CO_2 could be formed by reacting together 2 mol of CH_4 and 4 mol of O_2? Statement 2 can be used to solve this problem quickly. We see that the amounts reacted correspond to twice the amounts represented by Statement 2. Thus, 2 mol of CO_2 would be formed, which is also twice the amount represented by Statement 2. However, many stoichiometric problems cannot be solved quite as readily, so it is helpful to learn a general approach that works well for many problems of this type. This approach is based on the factor-unit method described earlier in Section 1.9. The needed factors are obtained from Statements 2, 3 and 4.

Example 5.7

The following questions are based on Equation 5.27.

a. How many moles of oxygen gas (O_2) will be required to react with 1.72 mol of methane (CH_4)?

b. How many grams of H_2O will be produced when 1.09 mol of CH_4 reacts with an excess of O_2?

c. How many grams of O_2 must react with excess CH_4 to produce 8.42 g of carbon dioxide (CO_2)?

Solution

a. The steps to follow in the factor-unit method are as follows (see Section 1.9):

1. Write down the known or given quantity. It will include both a number and a unit.
2. Leave working space, and set the known quantity equal to the unit of the unknown quantity.
3. Multiply the known quantity by a factor such that the unit of the known quantity is canceled and the unit of the unknown quantity is generated.
4. After the units on each side of the equation match, do the necessary arithmetic to get the final answer.

In this problem, the known quantity is 1.72 mol of CH_4, and the unknown quantity has the unit of mol O_2.

Step 1. 1.72 mol CH_4

Step 2. $1.72 \text{ mol CH}_4 \qquad = \text{mol O}_2$

Step 3. $1.72 \cancel{\text{mol CH}_4} \times \dfrac{2 \text{ mol O}_2}{1 \cancel{\text{mol CH}_4}} = \text{mol O}_2$

The factor

$$\frac{2 \text{ mol O}_2}{1 \text{ mol CH}_4}$$

was used because it properly canceled the unit of the known and generated the unit of the unknown.

Step 4. $(1.72)\left(\dfrac{2 \text{ mol O}_2}{1}\right) = 3.44 \text{ mol O}_2$

The answer contains the proper number of significant figures because the 2 and 1 of the factor are exact numbers (see Section 1.8).

b.

Step 1. 1.09 mol CH_4

Step 2. $1.09 \text{ mol CH}_4 \qquad = \text{g H}_2\text{O}$

Step 3. $1.09 \cancel{\text{mol CH}_4} \times \dfrac{2(18.0) \text{ g H}_2\text{O}}{1 \cancel{\text{mol CH}_4}} = \text{g H}_2\text{O}$

Step 4. $1.09 \times \dfrac{2(18.0) \text{ g H}_2\text{O}}{1} = 39.24 \text{ g H}_2\text{O} = 39.2 \text{ g H}_2\text{O}$

The answer was rounded to 3 significant figures to match the three in 1.09 and 18.0 g of H_2O. The 1 and 2 are exact counting numbers.

c.

Step 1. 8.42 g CO_2

Step 2. $8.42 \text{ g CO}_2 \qquad = \text{g O}_2$

Step 3. $8.42 \cancel{\text{g CO}_2} \times \dfrac{2(32.0) \text{ g O}_2}{44.0 \cancel{\text{g CO}_2}} = \text{g O}_2$

Step 4. $8.42 \times \dfrac{2(32.0) \text{ g O}_2}{44.0} = 12.2472 \text{ g O}_2 = 12.2 \text{ g O}_2$

The answer was rounded to 3 significant figures to match the three in 8.42, 32.0 g of O_2 and the 44.0. The 2 is an exact counting number.

Learning Check 5.7 The following reaction equation is balanced:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

a. Write statements equivalent to statements 2, 3, and 4 based on this equation.
b. Calculate the number of mol of NH_3 that would be produced if 2.11 mol of H_2 was reacted with excess N_2.
c. Calculate the number of g of N_2 required to react with 9.47 g H_2.

5.10 The Limiting Reactant

Learning Objective

7. Use the mole concept to do calculations based on the limiting-reactant principle.

Nearly everyone is familiar with two characteristics of automobiles. They stop running when they run out of gasoline, and they stop running when the air intake to the engine becomes clogged. Most people also know that gasoline and air are mixed, and the mixture is burned inside the automobile engine. However, the combustion reaction occurs only as long as both reactants, gasoline and air, are present. The reaction (and engine) stops when gasoline is absent, despite the presence of ample amounts of air. On the other hand, the reaction stops when air is absent, even though plenty of gasoline is available.

The behavior of the auto engine illustrates the **limiting-reactant principle** that a chemical reaction will take place only as long as all necessary reactants are present. The reaction will stop when one of the reactants, the **limiting reactant,** is used up. According to this principle, the amount of product formed depends on the amount of limiting reactant present because the reaction will stop (and no more product will form) when the limiting reactant is used up. This is represented in Figure 5.11 for the reaction $F_2(g) + H_2(g) \rightarrow 2HF(g)$.

limiting-reactant principle The maximum amount of product possible from a reaction is determined by the amount of reactant present in the least amount, based on its reaction coefficient and molecular weight.

limiting reactant The reactant present in a reaction in the least amount, based on its reaction coefficients and molecular weight. It is the reactant that determines the maximum amount of product that can be formed.

Example 5.8

A mixture containing 20.0 g of CH_4 and 100 g of O_2 is ignited and burns according to Equation 5.27, which is repeated here: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$. What substances will be found in the mixture after the reaction stops?

Solution

The task is to identify the limiting reactant. It will be completely used up, and the other reactant will be found mixed with the products (CO_2 and H_2O) after the reaction stops. A simple way to identify the limiting reactant is to find out which reactant will give the least amount of products. To do this, two problems must be solved, one in which the known quantity is 20.0 g of CH_4 and one in which the known quantity is 100.0 g of O_2. In each case, the theoretical amount of CO_2 produced is calculated, although the same information would be obtained by solving for the amount of water produced.

$$20.0\ \cancel{g\ CH_4} \times \frac{44.0\ g\ CO_2}{16.0\ \cancel{g\ CH_4}} = 55.0\ g\ CO_2$$

$$100\ \cancel{g\ O_2} \times \frac{44.0\ g\ CO_2}{2(32.0)\ \cancel{g\ O_2}} = 68.75\ g\ CO_2 = 68.8\ g\ CO_2$$

Since the 20.0 g of CH_4 produces the smaller amount of CO_2, CH_4 is the limiting reactant. Therefore, the final mixture will contain the products CO_2 and H_2O and the leftover O_2.

Chemistry Around Us 5.2
Air Bag Chemistry

Air bags installed in automobiles have proven to be effective, yet somewhat controversial, safety devices. According to the National Transportation Safety Administration, air bags have saved an estimated 1100 lives and prevented many more serious injuries from automobile accidents. This effectiveness is tempered by the fact that air bags have caused the deaths of a number of small children and infants. Short adults have also died as a result of being hit in the face rather than the chest by an expanding air bag. For this reason, it is now generally recommended that infants, small children, and short adults not occupy the front passenger seat. Also, most newer model automobiles are equipped with a switch that allows the air bag to be deactivated.

An air bag is essentially a nylon fabric bag that fills very rapidly with a gas when a collision occurs. Inflated air bags cushion and protect the driver and front-seat passenger. The gas that inflates an air bag is nitrogen, N_2, that is produced in a gas generator by a redox reaction of a toxic, explosive material called sodium azide, NaN_3. When a collision occurs, an electrical impulse triggers the rapid decomposition of the sodium azide according to the following equation:

$$2NaN_3(s) \rightarrow 2Na(g) + 3N_2(g)$$

The sodium vapor produced by this reaction would be very hazardous to the occupants of the automobile, but potassium nitrate, KNO_3, also included in the gas-generating mixture, reacts with the sodium vapor to produce potassium oxide, K_2O, sodium oxide, Na_2O, and some additional nitrogen gas. The equation for the reaction is

$$10Na(g) + 2KNO_3(s) \rightarrow K_2O(s) + 5Na_2O(s) + N_2(g)$$

The potassium oxide and sodium oxide are also hazardous materials; they are both basic oxides that react with water to form very caustic potassium hydroxide and sodium hydroxide, respectively. However, silicon dioxide, SiO_2, a third substance included in the gas-generating mixture of chemicals, is an acidic oxide that reacts with the basic potassium and sodium oxides, neutralizes their caustic characteristics, and converts them into a safe silicate-glass powder.

Thus, we see that from a chemical standpoint, the products of an air bag inflation are rendered safe by a series of carefully designed reactions. However, most air bags are never inflated, which means that old cars sent to scrap yards still contain a small amount of very toxic and explosive sodium azide. This is a recycling or disposal problem that is yet to be resolved.

Don Johnston/Tony Stone/Getty Images

Air bags: Protection provided by chemical reactions.

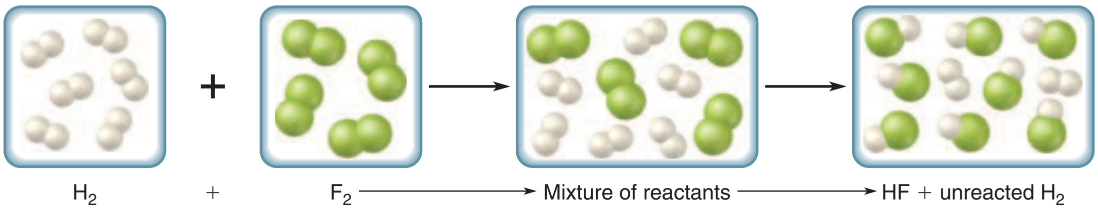

Figure 5.11 The limiting reactant is used up in a reaction.

Example 5.9

The reaction shown in Figure 5.8 was carried out by a student who wanted to produce some silver metal. The equation for the reaction is

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

The student dissolved 17.0 g of $AgNO_3$ in distilled water, then added 8.50 g of copper metal, Cu, to the resulting solution. What mass of silver metal, Ag, was produced?

Solution

The limiting reactant must be identified because it will determine the amount of silver produced. To find the limiting reactant, we calculate the theoretical amount of silver produced by each reactant. The reactant that will produce the smallest mass of silver is the limiting reactant, and the mass of silver calculated for it is the amount produced by the reaction.

$$17.0 \text{ } \cancel{\text{g AgNO}_3} \times \frac{2(108) \text{ g Ag}}{2(170) \text{ } \cancel{\text{g AgNO}_3}} = 10.8 \text{ g Ag}$$

$$8.50 \text{ } \cancel{\text{g Cu}} \times \frac{2(108) \text{ g Ag}}{63.6 \text{ } \cancel{\text{g Cu}}} = 28.8679 \text{ g Ag} = 28.9 \text{ g Ag}$$

We see that the $AgNO_3$ is the limiting reactant, and it will produce 10.8 g of Ag.

Learning Check 5.8 Refer to the reaction used for Learning Check 5.7. Assume that 2.00 mol of H_2 and 15.5 g of N_2 are reacted. What is the maximum mass of ammonia (NH_3) that can be produced? Which reactant is the limiting reactant?

5.11 Reaction Yields

Learning Objective

8. Use the mole concept to do percentage-yield calculations.

side reactions Reactions that do not give the desired product of a reaction.

In Sections 5.9 and 5.10, calculations were done to determine how much product could be obtained from the reaction of specified amounts of reactants. In many instances, the amount determined by such calculations would be greater than the amount produced by reacting the specified reactants in a laboratory. Does this mean matter is destroyed when reactions are actually done in the laboratory? No, this would violate the law of conservation of matter, a fundamental law of nature. Less product might be obtained because some of the reactants form compounds other than the desired product. Such reactions are called **side reactions.** Carbon dioxide gas, CO_2, is produced when carbon-containing fuels burn in an ample supply of oxygen. However, a side reaction produces toxic carbon monoxide gas, CO, when the oxygen supply is limited.

Another cause of less product than expected might be poor laboratory technique resulting in such things as losses when materials are transferred from one container to another. Whatever the cause, the mass of product obtained in an experiment is called the *actual yield,* and the mass calculated according to the methods of Sections 5.9 and 5.10 is called

Study Skills 5.1 Help with Oxidation Numbers

Assigning oxidation numbers is a task that seems to be straightforward, but mistakes caused by carelessness often show up when it is done on exams. You can minimize such mistakes by using a systematic approach to the task. First, have the seven rules for assigning oxidation numbers well in mind. Then, use an organized method such as the boxlike approach (shown below) to assign an oxidation number to the Mn in Mn_2O_3. In this approach, we begin with an oxidation number that is given by one of the rules (O is –2 according to Rule 5):

Each Mn must be +3 because two Mn give +6 total.

$$\begin{array}{cc} Mn_2 & O_3 \\ +3 & \downarrow \\ \uparrow & -2 \\ \uparrow & \downarrow \\ +6 & \leftarrow -6 \end{array}$$

Rule 5: O.N. of O = −2.

Rule 6: Sum of total O.N. of Mn and O = 0.

Total O.N. for 3 O's = −6

Follow the arrows in a clockwise direction to solve for the O.N. of each atom of Mn (+3 is correct). Remember that oxidation numbers are reported for individual atoms. For example, the total oxidation number due to Mn atoms is +6, but because there are two atoms of Mn in the formula, the oxidation number for each atom is just +3.

the *theoretical yield*. The **percentage yield** is the actual yield divided by the theoretical yield multiplied by 100 (see Section 1.10):

percentage yield The percentage of the theoretical amount of a product actually produced by a reaction.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \qquad (5.28)$$

Example 5.10

A chemist wants to produce urea (N_2CH_4O) by reacting ammonia (NH_3) and carbon dioxide (CO_2). The balanced equation for the reaction is

$$2NH_3(g) + CO_2(g) \rightarrow N_2CH_4O(s) + H_2O(\ell)$$

The chemist reacts 5.11 g of NH_3 with excess CO_2 and isolates 3.12 g of solid N_2CH_4O. Calculate the percentage yield of the experiment.

Solution

First, the theoretical yield must be calculated. Because this involves only masses of reactants and products, statements 2 and 4 based on the balanced equation will be used:

$$2NH_3(g) + CO_2(g) \rightarrow N_2CH_4O(s) + H_2O(\ell)$$

2. $2 \text{ mol } NH_3(g) + 1 \text{ mol } CO_2(g) \rightarrow 1 \text{ mol } N_2CH_4O(s) + 1 \text{ mol } H_2O(\ell)$
4. $2(17.0) \text{ g } NH_3 + 44.0 \text{ g } CO_2 \rightarrow 60.0 \text{ g } N_2CH_4O + 18.0 \text{ g } H_2O$

The known quantity is 5.11 g of NH_3, and the unit of the unknown quantity is g N_2CH_4O. The needed factor will come from statement 4.

Step 1. $5.11 \text{ g } NH_3$

Step 2. $5.11 \text{ g } NH_3 \qquad = \text{g } N_2CH_4O$

Step 3. $5.11 \cancel{\text{g } NH_3} \times \frac{60.0 \text{ g } N_2CH_4O}{2(17.0) \cancel{\text{g } NH_3}} \qquad = \text{g } N_2CH_4O$

Step 4. $5.11 \times \frac{60.0 \text{ g } N_2CH_4O}{2(17.0)} = 9.0176 \text{ g } N_2CH_4O = 9.02 \text{ g } N_2CH_4O$

The actual yield was 3.12 g, so the percentage yield is

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{3.12 \text{ g}}{9.02 \text{ g}} \times 100 = 34.5898\% = 34.6\%$$

Learning Check 5.9

a. A chemist isolates 17.43 g of product from a reaction that has a calculated theoretical yield of 21.34 g. What is the percentage yield?

b. Lime (CaO) is produced by heating calcium carbonate. The equation for the reaction is $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$. Suppose 510 g of $CaCO_3$ is heated, and 235 g of CaO is isolated after the reaction mixture cools. What is the percentage yield for the reaction?

Concept Summary

Chemical Equations. Chemical reactions are conveniently represented by equations in which reacting substances, called reactants, and the substances produced, called products, are written in terms of formulas. Coefficients are placed before reactant and product formulas to balance the equation. A balanced equation satisfies the law of conservation of matter.

Objective 1, Exercises 5.2 and 5.6

Types of Reactions. To facilitate their study, reactions are classified into the general category redox or nonredox, and further classified as decomposition, combination, single-replacement, or double-replacement reactions.

Redox Reactions. Reactions in which reactants undergo oxidation or reduction are called redox reactions. Oxidation and reduction are indicated conveniently by the oxidation number changes that occur.

Oxidation numbers are assigned according to a specific set of rules. A substance is oxidized when the oxidation number of a constituent element increases, and it is reduced when the oxidation number of a constituent element decreases.
Objective 2, Exercises 5.10 and 5.15

Decomposition Reactions. Decomposition reactions are characterized by one substance reacting to give two or more products. Decomposition reactions can be redox or nonredox.
Objective 3, Exercise 5.20

Combination Reactions. Combination reactions, also called addition or synthesis reactions, are characterized by two or more reactants that form a single compound as a product. Combination reactions can be redox or nonredox.
Objective 3, Exercise 5.20

Replacement Reactions. Single-replacement reactions, sometimes called substitution reactions, are always redox reactions. One element reacts with a compound and displaces another element from the compound. Double-replacement reactions, also called metathesis reactions, are always nonredox reactions. They can be recognized by their partner-swapping characteristics.
Objective 3, Exercise 5.20

Ionic Equations. Many water-soluble compounds separate (dissociate) into ions when dissolved in water. Reactions of such materials can be represented by molecular equations in which no ions are shown, total ionic equations in which all ions are shown, or net ionic equations in which only ions actually undergoing a change are shown.
Objective 4, Exercise 5.30a, b, c

Energy and Reactions. Energy changes accompany all chemical reactions. The energy can appear in a variety of forms, but heat is a common one. Exothermic reactions liberate heat, and endothermic reactions absorb it.
Objective 5, Exercise 5.34

The Mole and Chemical Equations. The mole concept, when applied to chemical equations, yields relationships that can be used to obtain factors for doing factor-unit calculations.
Objective 6, Exercise 5.42

The Limiting Reactant. The limiting reactant is the reactant present in a reaction in an amount that determines the maximum amount of product that can be made. Factor-unit calculations can be used to determine which reactant is limiting.
Objective 7, Exercise 5.52

Reaction Yields. The mass of product isolated after a reaction is often less than the mass theoretically possible. The ratio of the actual isolated mass (actual yield) to the calculated theoretical yield multiplied by 100 is the percentage yield.
Objective 8, Exercise 5.56

Key Terms and Concepts

Balanced equation (5.1)
Combination reaction (5.5)
Decompostion reaction (5.4)
Double-replacement reaction (5.6)
Endothermic reaction (5.8)
Exothermic reaction (5.8)
Law of conservation of matter (5.1)
Limiting reactant (5.10)
Limiting-reactant principle (5.10)
Molecular equation (5.7)
Net ionic equation (5.7)
Oxidation (5.3)
Oxidation number (oxidation state) (5.3)
Oxidizing agent (5.3)
Percentage yield (5.11)
Product of a reaction (5.1)
Reactant of a reaction (5.1)
Reducing agent(5.3)
Reduction (5.3)
Side reaction (5.11)
Single-replacement reaction (5.6)
Spectator ion (5.7)
Stoichiometry (5.9)
Total ionic equation (5.7)

Key Equations

1. Decomposition reaction—one substance changes to two or more new substances (Section 5.4):	$A \rightarrow B + C$	Equation 5.13
2. Combination reaction—two or more substances react to produce one new substance (Section 5.5):	$A + B \rightarrow C$	Equation 5.16
3. Single-replacement reaction—one element reacts with a compound to produce a new compound and new element (Section 5.6):	$A + BX \rightarrow B + AX$	Equation 5.19
4. Double-replacement reaction—partner-swapping reaction (Section 5.6):	$AX + BY \rightarrow BX + AY$	Equation 5.21
5. Percentage yield (Section 5.11):	$\%\ \text{yield} = \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100$	Equation 5.28

Exercises

OWL Interactive versions of these problems are assignable in OWL.

Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

Chemical Equations (Section 5.1)

5.1 Identify the reactants and products in each of the following reaction equations:

a. $BaO_2(s) + H_2SO_4(\ell) \rightarrow BaSO_4(s) + H_2O_2(\ell)$

b. $2H_2O_2(aq) \rightarrow 2H_2O(\ell) + O_2(g)$

c. methane + water → carbon monoxide + hydrogen

d. copper(II) oxide + hydrogen → copper + water

5.2 Identify the reactants and products in each of the following reaction equations:

a. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

b. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

c. magnesium oxide + carbon → magnesium + carbon monoxide

d. ethane + oxygen → carbon dioxide + water

5.3 Identify which of the following are consistent with the law of conservation of matter. For those that are not, explain why they are not.

a. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

b. $P_4(s) + O_2(g) \rightarrow P_4O_{10}(s)$

c. 3.20 g oxygen + 3.21 g sulfur → 6.41 g sulfur dioxide

d. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

5.4 Identify which of the following are consistent with the law of conservation of matter. For those that are not, explain why they are not.

a. $ZnS(s) + O_2(g) \rightarrow ZnO(s) + SO_2(g)$

b. $Cl_2(aq) + 2I^-(aq) \rightarrow I_2(aq) + 2Cl^-(aq)$

c. 1.50 g oxygen + 1.50 g carbon → 2.80 g carbon monoxide

d. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

5.5 Determine the number of atoms of each element on each side of the following equations and decide which equations are balanced:

a. $H_2S(aq) + I_2(aq) \rightarrow 2HI(aq) + S(s)$

b. $KClO_3(s) \rightarrow KCl(s) + O_2(g)$

c. $SO_2(g) + H_2O(\ell) \rightarrow H_2SO_3(aq)$

d. $Ba(ClO_3)_2(aq) + H_2SO_4(aq) \rightarrow 2HClO_3(aq) + BaSO_4(s)$

5.6 Determine the number of atoms of each element on each side of the following equations and decide which equations are balanced:

a. $2Ag_2O(s) \rightarrow 2Ag(s) + O_2(g)$

b. $Al(s) + O_2(g) \rightarrow Al_2O_3(s)$

c. $2AgNO_3(aq) + K_2SO_4(aq) \rightarrow Ag_2SO_4(s) + 2KNO_3(aq)$

d. $SO_2(g) + O_2(g) \rightarrow SO_3(g)$

5.7 Balance the following equations:

a. $Cl_2(aq) + NaBr(aq) \rightarrow NaCl(aq) + Br_2(aq)$

b. $CaF_2(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + HF(g)$

c. $Cl_2(g) + NaOH(aq) \rightarrow NaOCl(aq) + NaCl(aq) + H_2O(\ell)$

d. $KClO_3(s) \rightarrow KClO_4(s) + KCl(s)$

e. dinitrogen monoxide → nitrogen + oxygen

f. dinitrogen pentoxide → nitrogen dioxide + oxygen

g. $P_4O_{10}(s) + H_2O(\ell) \rightarrow H_4P_2O_7(aq)$

h. $CaCO_3(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$

5.8 Balance the following equations:

a. $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(\ell)$

b. hydrogen + chlorine → hydrogen chloride

c. $H_2S(g) + O_2(g) \rightarrow SO_2(g) + H_2O(\ell)$

d. sulfur + oxygen → sulfur dioxide

e. $Na_2CO_3(aq) + Ca(NO_3)_2(aq) \rightarrow NaNO_3(aq) + CaCO_3(s)$

f. $NaBr(aq) + Cl_2(aq) \rightarrow NaCl(aq) + Br_2(aq)$

g. $Ag_2CO_3(s) \rightarrow Ag(s) + CO_2(g) + O_2(g)$

h. $H_2O_2(aq) + H_2S(aq) \rightarrow H_2O(\ell) + S(s)$

Redox Reactions (Section 5.3)

5.9 Assign oxidation numbers to the blue element in each of the following formulas:

a. Cl_2O_5

b. $KClO_4$

c. Ba^{2+}

d. F_2

e. $H_4P_2O_7$

f. H_2S

5.10 Assign oxidation numbers to the blue element in each of the following formulas:

a. ClO_3^-

b. H_2SO_4

c. $NaNO_3$

d. N_2O

e. $KMnO_4$

f. $HClO_2$

5.11 Find the element with the highest oxidation number in each of the following formulas:

a. N_2O_5

b. $KHCO_3$

c. $NaOCl$

d. $NaNO_3$

e. $HClO_4$

f. $Ca(NO_3)_2$

5.12 Find the element with the highest oxidation number in each of the following formulas:

a. $Na_2Cr_2O_7$

b. $K_2S_2O_3$

c. HNO_3

d. P_2O_5

e. $Mg(ClO_4)_2$

f. $HClO_2$

5.13 For each of the following equations, indicate whether the blue element has been oxidized, reduced, or neither oxidized nor reduced.

a. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

b. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$

c. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

d. $BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

e. $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

5.14 For each of the following equations, indicate whether the blue element has been oxidized, reduced or neither oxidized nor reduced.

a. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

b. $SO_2(g) + H_2O(\ell) \rightarrow H_2SO_3(aq)$

c. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

d. $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

e. $2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH(aq) + H_2(g)$

5.15 Assign oxidation numbers to each element in the following equations and identify the oxidizing and reducing agents:

a. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

b. $H_2O(g) + CH_4(g) \rightarrow CO(g) + 3H_2(g)$

c. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$

d. $B_2O_3(s) + 3Mg(s) \rightarrow 2B(s) + 3MgO(s)$

e. $Fe_2O_3(s) + CO(g) \rightarrow 2FeO(s) + CO_2(g)$

f. $Cr_2O_7^{2-}(aq) + 2H^+(aq) + 3Mn^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3MnO_2(s) + H_2O(\ell)$

5.16 Assign oxidation numbers to each element in the following equations and identify the oxidizing and reducing agents:

a. $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

b. $Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$

c. $3MnO_2(s) + 4Al(s) \rightarrow 2Al_2O_3(s) + 3Mn(s)$

d. $2H^+(aq) + 3SO_3^{2-}(aq) + 2NO_3^-(aq) \rightarrow 2NO(g) + H_2O(\ell) + 3SO_4^{2-}(aq)$

e. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

f. $4NO_2(g) + O_2(g) \rightarrow 2N_2O_5(g)$

5.17 The tarnish of silver objects is a coating of silver sulfide (Ag_2S), which can be removed by putting the silver in contact with aluminum metal in a dilute solution of baking soda or salt. The equation for the cleaning reaction is

$$3Ag_2S(s) + 2Al(s) \rightarrow 6Ag(s) + Al_2S_3(s)$$

The sulfur in these compounds has a -2 oxidation number. What are the oxidizing and reducing agents in the cleaning reaction?

5.18 Aluminum metal reacts rapidly with highly basic solutions to liberate hydrogen gas and a large amount of heat. This reaction is utilized in a popular solid drain cleaner that is composed primarily of lye (sodium hydroxide) and aluminum granules. When wet, the mixture reacts as follows:

$$6NaOH(aq) + 2Al(s) \rightarrow 3H_2(g) + 2Na_3AlO_3(aq) + \text{heat}$$

The liberated H_2 provides agitation that, together with the heat, breaks the drain stoppage loose. What are the oxidizing and reducing agents in the reaction?

5.19 Identify the oxidizing and reducing agents in the Haber process for producing ammonia from elemental nitrogen and hydrogen. The equation for the reaction is

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Decomposition, Combination, and Replacement Reactions (Sections 5.4–5.6)

5.20 Classify each of the reactions represented by the following equations, first as a redox or nonredox reaction. Then further classify each redox reaction as a decomposition, single-replacement, or combination reaction, and each nonredox reaction as a decomposition, double-replacement, or combination reaction.

a. $K_2CO_3(s) \rightarrow K_2O(s) + CO_2(g)$

b. $Ca(s) + 2H_2O(\ell) \rightarrow Ca(OH)_2(s) + H_2(g)$

c. $BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

d. $SO_2(g) + H_2O(\ell) \rightarrow H_2SO_3(aq)$

e. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

f. $2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$

5.21 Classify each of the reactions represented by the following equations, first as a redox or nonredox reaction. Then further classify each redox reaction as a decomposition, single-replacement, or combination reaction, and each nonredox reaction as a decomposition, double-replacement, or combination reaction:

a. $N_2O_5(g) + H_2O(\ell) \rightarrow 2HNO_3(aq)$

b. $Cr_2O_3(s) + 2Al(s) \rightarrow 2Cr(s) + Al_2O_3(s)$

c. $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(s)$

d. $H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(\ell)$

e. $PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$

f. $Zn(s) + Cl_2(g) \rightarrow ZnCl_2(s)$

5.22 Baking soda ($NaHCO_3$) can serve as an emergency fire extinguisher for grease fires in the kitchen. When heated, it liberates CO_2, which smothers the fire. The equation for the reaction is

$$2NaHCO_3(s) \xrightarrow{\text{Heat}} Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

Classify the reaction into the categories used in Exercises 5.20 and 5.21.

5.23 Baking soda may serve as a source of CO_2 in bread dough. It causes the dough to rise. The CO_2 is released when $NaHCO_3$ reacts with an acidic substance:

$$NaHCO_3(aq) + H^+(aq) \rightarrow Na^+(aq) + H_2O(\ell) + CO_2(g)$$

Classify the reaction as redox or nonredox.

5.24 Many homes are heated by the energy released when natural gas (represented by CH_4) reacts with oxygen. The equation for the reaction is

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Classify the reaction as redox or nonredox.

5.25 Hydrogen peroxide will react and liberate oxygen gas. In commercial solutions, the reaction is prevented to a large degree by the addition of an inhibitor. The equation for the oxygen-liberating reaction is

$$2H_2O_2(aq) \rightarrow 2H_2O(\ell) + O_2(g)$$

Classify the reaction as redox or nonredox.

5.26 Chlorine, used to treat drinking water, undergoes the reaction in water represented by the following equation:

$$Cl_2(aq) + H_2O(\ell) \rightarrow HOCl(aq) + HCl(aq)$$

Classify the reaction as redox or nonredox.

5.27 Triple superphosphate, an ingredient of some fertilizers, is prepared by reacting rock phosphate (calcium phosphate) and phosphoric acid. The equation for the reaction is

$$Ca_3(PO_4)_2(s) + 4H_3PO_4(aq) \rightarrow 3Ca(H_2PO_4)_2(s)$$

Classify the reaction into the categories used in Exercises 5.20 and 5.21.

Ionic Equations (Section 5.7)

5.28 Consider all of the following ionic compounds to be water soluble, and write the formulas of the ions that would be formed if the compounds were dissolved in water. Table 4.7 will be helpful.

a. $LiNO_3$

b. Na_2HPO_4

c. $Ca(ClO_3)_2$

d. KOH

e. $MgBr_2$

f. $(NH_4)_2SO_4$

5.29 Consider all the following ionic compounds to be water soluble, and write the formulas of the ions that would be formed if the compounds were dissolved in water. Table 4.7 will be helpful.

a. $K_2Cr_2O_7$

b. H_2SO_4

c. NaH_2PO_4

d. Na_3PO_4

e. NH_4Cl

f. $KMnO_4$

5.30 Reactions represented by the following equations take place in water solutions. Write each molecular equation in total ionic form, then identify spectator ions and write the equations in net ionic form. Solids that do not dissolve are designated by (s), gases that do not dissolve are designated by (g), and substances that dissolve but do not dissociate appear in blue.

a. $Cl_2(aq) + 2NaI(aq) \rightarrow 2NaCl(aq) + I_2(aq)$

b. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

c. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

d. $BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

e. $SO_3(aq) + H_2O(\ell) \rightarrow H_2SO_4(aq)$

f. $2NaI(aq) + 2H_2SO_4(aq) \rightarrow I_2(aq) + SO_2(aq) + Na_2SO_4(aq) + H_2O(\ell)$

5.31 Reactions represented by the following equations take place in water solutions. Write each molecular equation in total ionic form, then identify spectator ions and write the equations in net ionic form. Solids that do not dissolve are designated by (s), gases that do not dissolve are designated by (g), and substances that dissolve but do not dissociate appear in blue.

a. $H_2O(\ell) + Na_2SO_3(aq) + SO_2(aq) \rightarrow 2NaHSO_3(aq)$

b. $3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$

c. $2HCl(aq) + CaO(s) \rightarrow CaCl_2(aq) + H_2O(\ell)$

d. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(aq) + H_2O(\ell)$

e. $MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + Cl_2(aq) + 2H_2O(\ell)$

f. $2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

5.32 The following molecular equations all represent neutralization reactions of acids and bases. These reactions will be discussed further in Chapter 9. Write each equation in total ionic form, identify the spectator ions, then write the net ionic equation. Water is the only substance that does not dissociate. What do you notice about all the net ionic equations?

a. $HNO_3(aq)\ KOH(aq) \rightarrow KNO_3(aq) + H_2O(\ell)$

b. $H_3PO_4(aq) + 3NH_4OH(aq) \rightarrow (NH_4)_3PO_4(aq) + 3H_2O(\ell)$

c. $HI(aq) + NaOH(aq) \rightarrow NaI(aq) + H_2O(\ell)$

5.33 The following molecular equations all represent neutralization reactions of acids and bases. These reactions will be discussed further in Chapter 9. Write each equation in total ionic form, identify the spectator ions, then write the net ionic equation. Water is the only substance that does not dissociate. What do you notice about all the net ionic equations?

a. $HBr(aq) + RbOH(aq) \rightarrow RbBr(aq) + H_2O(\ell)$

b. $H_2SO_4(aq) + 2LiOH(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(\ell)$

c. $HCl(aq) + CsOH(aq) \rightarrow CsCl(aq) + H_2O(\ell)$

Energy and Reactions (Section 5.8)

5.34 In addition to emergency cold packs, emergency hot packs are available that heat up when water is mixed with a solid. Is the process that takes place in such packs exothermic or endothermic? Explain.

5.35 In refrigeration systems, the area to be cooled has pipes running through it. Inside the pipes, a liquid evaporates and becomes a gas. Is the evaporation process exothermic or endothermic? Explain.

5.36 An individual wants to keep some food cold in a portable picnic cooler. A piece of ice is put into the cooler, but it is wrapped in a thick insulating blanket to slow its melting. Comment on the effectiveness of the cooler in terms of the direction of heat movement inside the cooler.

5.37 The human body cools itself by the evaporation of perspiration. Is the evaporation process endothermic or exothermic? Explain.

The Mole and Chemical Equations (Section 5.9)

5.38 For the reactions represented by the following equations, write statements equivalent to statements 2, 3, and 4 given in Section 5.9.

a. $Ca(s) + 2H_2O(\ell) \rightarrow H_2(g) + Ca(OH)_2(s)$

b. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

c. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(\ell)$

d. $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$

e. $2HCl(aq) + Mg(OH)_2(s) \rightarrow MgCl_2(aq) + 2H_2O(\ell)$

5.39 For the reactions represented by the following equations, write statements equivalent to Statements 2, 3 and 4 given in Section 5.9.

a. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

b. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

c. $BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

d. $2H_2O_2(aq) \rightarrow 2H_2O(\ell) + O_2(g)$

e. $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$

5.40 For the following equation, write statements equivalent to Statements 2, 3, and 4 given in Section 5.9. Then write at least six factors (including numbers and units) that could be used to solve problems by the factor-unit method.

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

5.41 Calculate the number of grams of SO_2 that must react to produce 350 g of SO_3. Use the statements written in Exercise 5.40 and express your answer using the correct number of significant figures.

5.42 Calculate the mass of limestone ($CaCO_3$) that must be decomposed to produce 500 g of lime (CaO). The equation for the reaction is

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

5.43 Calculate the number of moles of CO_2 generated by the reaction of Exercise 5.42 when 500 g of CaO is produced.

5.44 Calculate the number of grams of bromine (Br_2) needed to react exactly with 50.1 g of aluminum (Al). The equation for the reaction is

$$2Al(s) + 3Br_2(\ell) \rightarrow 2AlBr_3(s)$$

5.45 Calculate the number of moles of $AlBr_3$ produced by the process of Exercise 5.44.

5.46 How many grams of $AlBr_3$ are produced by the process in Exercise 5.44?

5.47 In Exercise 5.17 you were given the following equation for the reaction used to clean tarnish from silver:

$$3Ag_2S(s) + 2Al(s) \rightarrow 6Ag(s) + Al_2S_3(s)$$

a. How many grams of aluminum would need to react to remove 0.250 g of Ag_2S tarnish?

b. How many moles of Al_2S_3 would be produced by the reaction described in (a)?

5.48 Pure titanium metal is produced by reacting titanium(IV) chloride with magnesium metal. The equation for the reaction is

$$TiCl_4(s) + 2Mg(s) \rightarrow Ti(s) + 2MgCl_2(s)$$

How many grams of Mg would be needed to produce 1.00 kg of pure titanium?

5.49 An important metabolic process of the body is the oxidation of glucose to water and carbon dioxide. The equation for the reaction is

$$C_6H_{12}O_6(aq) + 6O_2(aq) \rightarrow 6CO_2(aq) + 6H_2O(\ell)$$

a. What mass of water in grams is produced when the body oxidizes 1.00 mol of glucose?

b. How many grams of oxygen are needed to oxidize 1.00 mol of glucose?

5.50 Caproic acid is oxidized in the body as follows:

$$C_6H_{12}O_2(aq) + 8O_2(aq) \rightarrow 6CO_2(aq) + 6H_2O(\ell)$$

How many grams of oxygen are needed to oxidize 1.00 mol of caproic acid?

The Limiting Reactant (Section 5.10)

5.51 A sample of 4.00 g of methane (CH_4) is mixed with 15.0 g of chlorine (Cl_2).

a. Determine which is the limiting reactant according to the following equation:

$$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(\ell) + 4HCl(g)$$

b. What is the maximum mass of CCl_4 that can be formed?

5.52 Nitrogen and oxygen react as follows:

$$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$$

Suppose 1.25 mol of N_2 and 50.0 g of O_2 are mixed together.

a. Which one is the limiting reactant?

b. What is the maximum mass in grams of NO_2 that can be produced from the mixture?

5.53 Suppose you want to use acetylene (C_2H_2) as a fuel. You have a cylinder that contains 500 g of C_2H_2 and a cylinder that contains 2000 g of oxygen (O_2). Do you have enough oxygen to burn all the acetylene? The equation for the reaction is

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g)$$

5.54 Ammonia, carbon dioxide, and water vapor react to form ammonium bicarbonate as follows:

$$NH_3(aq) + CO_2(aq) + H_2O(\ell) \rightarrow NH_4HCO_3(aq)$$

Suppose 50.0 g of NH_3, 80.0 g of CO_2, and 2.00 mol of H_2O are reacted. What is the maximum number of grams of NH_4HCO_3 that can be produced?

5.55 Chromium metal (Cr) can be prepared by reacting the oxide with aluminum. The equation for the reaction is

$$Cr_2O_3(s) + 2Al(s) \rightarrow 2Cr(s) + Al_2O_3(s)$$

What three substances will be found in the final mixture if 150 g of Cr_2O_3 and 150 g of Al are reacted?

Reaction Yields (Section 5.11)

5.56 The actual yield of a reaction was 12.18 g of product, while the calculated theoretical yield was 15.93 g. What was the percentage yield?

5.57 A product weighing 14.37 g was isolated from a reaction. The amount of product possible according to a calculation was 17.55 g. What was the percentage yield?

5.58 For a combination reaction, it was calculated that 7.59 g of *A* would exactly react with 4.88 g of *B*. These amounts were reacted, and 9.04 g of product was isolated. What was the percentage yield of the reaction?

5.59 A sample of calcium metal with a mass of 2.00 g was reacted with excess oxygen. The following equation represents the reaction that took place:

$$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$$

The isolated product (CaO) weighed 2.26 g. What was the percentage yield of the reaction?

5.60 Upon heating, mercury(II) oxide undergoes a decomposition reaction:

$$2HgO(s) \rightarrow 2Hg(\ell) + O_2(g)$$

A sample of HgO weighing 7.22 g was heated. The collected mercury weighed 5.95 g. What was the percentage yield of the reaction?

Additional Exercises

5.61 Would you expect argon, Ar, to be involved in a redox reaction? Explain your reasoning.

5.62 Rewrite the following word equation using chemical formulas. Balance the equation and write the balanced equation in net ionic form:

barium chloride (aq) + sodium sulfate (aq) →
sodium chloride (aq) + barium sulfate (s)

5.63 The element with an electron configuration of $1s^22s^22p^63s^1$ undergoes a combination reaction with the element that has 15 protons in its nucleus. Assume the product of the reaction consists of simple ions, and write a balanced chemical equation for the reaction.

5.64 Assuming a 100% reaction yield, it was calculated that 6.983 g of naturally occurring elemental iron would be needed to react with another element to form 18.149 g of product. If iron composed only of the iron-60 isotope was used to form the product, how many grams of iron-60 would be required?

5.65 The decomposition of a sample of a compound produced 1.20×10^{24} atoms of nitrogen and 80.0 grams of oxygen atoms. What was the formula of the sample that was decomposed? What is the correct name of the decomposed sample?

Allied Health Exam Connection

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5.66 Balance the following redox reaction:

$Mg(s) + H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$

a. $Mg(s) + H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
b. $Mg(s) + 4H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
c. $Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$
d. $Mg(s) + H_2O(g) \rightarrow Mg(OH)_2(s) + \frac{1}{2}H_2(g)$

5.67 Which one of the following equations is balanced?

a. $H_2O \rightarrow H_2 + O_2$
b. $Al + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2$
c. $S + O_2 \rightarrow SO_3$
d. $2HgO \rightarrow 2Hg + O_2$

5.68 Which of the following equations is balanced?

a. $Mg + N_2 \rightarrow Mg_3N_2$
b. $Fe + O_2 \rightarrow Fe_2O_3$
c. $C_{12}H_{22}O_{11} \rightarrow 12C + 11H_2O$
d. $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$

5.69 The coefficient of O_2 after the following equation is balanced is: $\_CH_4 + \_O_2 \rightarrow \_CO_2 + \_H_2O$

a. 1
b. 2
c. 3
d. 4

5.70 Ammonia can be produced from nitrogen and hydrogen, according to the unbalanced equation

$$N_2(g) + H_2(g) = NH_3(g)$$

After balancing the equation, the coefficient before ammonia should be:

a. 1
b. 2
c. 3
d. 4

5.71 What is the oxidation number of sodium in the following reaction? $Pb(NO_3)_{2(aq)} + 2NaI_{(aq)} \rightarrow PbI_{2(s)} + 2NaNO_{3(aq)}$

a. +1
b. +2
c. −1
d. −2

5.72 What is the oxidation number for nitrogen in HNO_3?

a. −2
b. +5
c. −1
d. −5

5.73 The oxidation number of sulfur in the ion SO_4^{-2} is:

a. −2
b. +2
c. +6
d. +10

5.74 Which of the following is the oxidation number of sulfur in the compound sodium thiosulfate, $Na_2S_2O_3$?

a. +1
b. −1
c. +2
d. −2

5.75 Which best describes the following redox reaction: $Br^-_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Br_{2(1)} + Mn^{+2}_{(aq)}$?

a. Br and Mn are both reduced
b. Br is oxidized and Mn is reduced
c. Br is oxidized and O is reduced
d. Br is reduced and Mn is oxidized

5.76 Which reactant is oxidized and which is reduced in the following reaction?
$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

a. oxidized: $C_2H_4(g)$, reduced: $3O_2(g)$

b. oxidized: $C_2H_4(g)$, reduced: $2H_2O(g)$

c. oxidized: $C_2H_4(g)$, reduced: $2CO_2(g)$

d. oxidized: $2CO_2(g)$, reduced: $C_2H_4(g)$

5.77 Which of the following species is being oxidized in this redox reaction?
$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$

a. $Zn_{(s)}$

b. $Cu^{2+}{}_{(aq)}$

c. $Zn^{2+}{}_{(aq)}$

d. $Cu_{(s)}$

5.78 Identify the oxidizing agent and the reducing agent in the following reaction:
$8H^+(aq) + 6Cl^-(aq) + Sn(s) + 4NO_3^-(aq) \rightarrow$
$SnCl_6^{2-}(aq) + 4NO_2(g) + 4H_2O(\ell)$

a. oxidizing agent: $8H^+(aq)$, reducing agent: $Sn(s)$

b. oxidizing agent: $4NO_3^-(aq)$, reducing agent: $Sn(s)(g)$

c. oxidizing agent: $4NO_3^-(aq)$, reducing agent: $4NO_2(g)$

d. oxidizing agent: $4NO_3^-(aq)$, reducing agent: $8H^+(aq)$

5.79 The chemical reaction: $2Zn + 2HCl \rightarrow 2ZnCl + H_2$ is an example of:

a. double displacement

b. synthesis

c. analysis

d. single displacement

5.80 Identify the following as an oxidation, a reduction, a decomposition, or a dismutation reaction:
$Cl_2 + 2e^- \rightarrow 2Cl^-$

a. a reduction

b. an oxidation

c. a decomposition

d. a dismutation

5.81 The reaction $2NaI + Cl_2 \rightarrow 2NaCl + I_2$ demonstrates a:

a. decomposition reaction

b. synthesis reaction

c. single replacement reaction

d. double replacement reaction

5.82 Classify the reaction as to the reaction type:
$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

a. synthesis

b. decomposition

c. single replacement

d. double replacement

5.83 Identify the statement which is NOT characteristic of exergonic reactions:

a. They are downhill reactions

b. They have a negative energy change ($-H$)

c. They are uphill reactions

d. The products have less energy than the reactants

5.84 What is the net ionic equation of the following transformations?
$Ca(NO_3)_2(aq) + 2KCl(aq) \rightarrow CaCl_2(s) + 2KNO_3(aq)$

a. $2NO_3^-(aq) + 2K^+(aq) \rightarrow 2KNO_3(aq)$

b. $Ca^{2+}(aq) + 2Cl^-(aq) \rightarrow CaCl_2(s)$

c. $Ca^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2Cl^-(aq) \rightarrow CaCl_2(s) + 2K^+(aq) + 2NO_3^-(aq)$

d. $Ca^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + 2K^+(aq) + 2NO_3^-(aq)$

5.85 In the Haber process, ammonia is produced according to the following equation:

$$N_2(g) + 3H_2(g) = 2NH_3(g)$$

How many moles of hydrogen gas are needed to react with one of nitrogen?

a. 1

b. 3

d. 6

d. 22.4

5.86 What mass of product would you expect given that you started with 17 g of NH_3 and 36.5 g of HCl?

$$NH_3 + HCl \rightarrow NH_4Cl$$

a. 17 g

b. 36.5 g

c. 53.5 g

d. 19.5 g

5.87 The number of grams of hydrogen formed by the action of 6 grams of magnesium (atomic weight = 24) on an appropriate quantity of acid is:

a. 0.5

b. 8

c. 22.4

d. 72

5.88 In the reaction $CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2NaCl$, if 0.5 mole of NaCl is to be formed, then:

a. 1 mole of Na_2CO_3 is needed

b. 0.5 mole of $CaCO_3$ is also formed

c. 0.5 rnole of Na_2CO_3 is needed

d. 0.25 mole of $CaCl_2$ is needed

5.89 In the reaction $4Al + 3O_2 \rightarrow 2Al_2O_3$, how many grams of O_2 are needed to completely react with 1.5 moles of Al?

a. 24 g

b. 36 g

c. 48 g

d. 60 g

Chemistry for Thought

5.90 In experiments where students prepare compounds by precipitation from water solutions, they often report yields of dry product greater than 100%. Propose an explanation for this.

5.91 Refer to Figure 5.2 and follow the instructions. Then explain how the concept of limiting reactant is used to extinguish a fire.

5.92 Refer to Figure 5.4 and answer the question. Suggest another material that might provide the enzyme catalyst.

5.93 What do the observations of Figure 5.10 indicate to you about the abilities of the following solids to dissolve in water: NaCl, $AgNO_3$, and AgCl? What would you expect to observe if you put a few grams of solid AgCl into a test tube containing 3 mL of water and shook the mixture? What would you expect to observe if you repeated this experiment but used a few grams of $NaNO_3$ instead of AgCl?

5.94 The concentration of alcohol (CH_3CH_2OH) in the breath of an individual who has been drinking is measured by an instrument called a Breathalyzer. The breath sample is passed through a solution that contains the orange-colored $Cr_2O_7^{2-}$ ion. The equation for the net ionic redox reaction that occurs is:

$$2Cr_2O_7^{2-}(aq) + 3CH_3CH_2OH(aq) + 16H^+(aq) \rightarrow 4Cr^{3+}(aq) + 3CH_3COOH(aq) + 11H_2O(\ell)$$

The Cr^{3+} ion has a pale violet color in solution. Explain how color changes in the solution could be used to indicate the amount of alcohol in the breath sample.

5.95 Certain vegetables and fruits, such as potatoes and apples, darken quickly when sliced. Submerging the slices in water slows this process. Explain.

5.96 In an ordinary flashlight battery, an oxidation reaction and a reduction reaction take place at different locations to produce an electrical current that consists of electrons. In one of the reactions, the zinc container of the battery slowly dissolves as it is converted into zinc ions. Is this the oxidation or the reduction reaction? Is this reaction the source of electrons, or are electrons used to carry out the reaction? Explain your answers with a reaction equation.

6 The States of Matter

Learning Objectives

When you have completed your study of this chapter, you should be able to:

1 Do calculations based on the property of density. **(Section 6.1)**

2 Demonstrate an understanding of the kinetic molecular theory of matter. **(Sections 6.2–6.4)**

3 Use the kinetic molecular theory to explain and compare the properties of matter in different states. **(Section 6.5)**

4 Do calculations to convert pressure and temperature values into various units. **(Section 6.6)**

5 Do calculations based on Boyle's law, Charles's law, and the combined gas law. **(Section 6.7)**

6 Do calculations based on the ideal gas law. **(Section 6.8)**

7 Do calculations based on Dalton's law. **(Section 6.9)**

8 Do calculations based on Graham's law. **(Section 6.10)**

9 Classify changes of state as exothermic or endothermic. **(Section 6.11)**

10 Demonstrate an understanding of the concepts of vapor pressure and evaporation. **(Section 6.12)**

11 Demonstrate an understanding of the process of boiling and the concept of boiling point. **(Section 6.13)**

12 Demonstrate an understanding of the processes of sublimation and melting. **(Section 6.14)**

13 Do calculations based on energy changes that accompany heating, cooling, or changing the state of a substance. **(Section 6.15)**

OWL Online homework for this chapter may be assigned in OWL.

Respiratory therapists assist in both the treatment and diagnostic testing of pulmonary function. They dispense gases, vapors, and drug-containing therapeutic aerosols to patients. They also use devices such as a spirometer to measure lung capacity. Gaseous behavior, as represented by the gas laws of this chapter, is an important part of their training.

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If you live in an area that has cold winters, you have probably seen water in the three different forms used to categorize the states in which matter occurs. On a cold day, you can usually find solid water (ice) floating in a pool of cold liquid water, and at the same time you can see a small cloud of tiny water droplets that forms when gaseous water condenses as you exhale into the cold air.

Most matter is not as easily observed in all three states as the water in the preceding example. In fact, most matter is classified as a solid, a liquid, or a gas on the basis of the form in which it is commonly observed. However, according to Section 4.11, the state of a substance depends on temperature. You will see in this chapter that the state also depends on pressure. Therefore, when a substance is classified as a solid, a liquid, or a gas, we are usually simply stating its form under normal atmospheric pressure and at a temperature near 25°C. ◗ Figure 6.1 gives the states of the elements under those conditions.

In this chapter, we will study the characteristics of the common states of matter, along with a theory that relates these states to molecular behavior. We will also investigate the energy relationships that accompany changes of state.

6.1 Observed Properties of Matter

Learning Objective

1. Do calculations based on the property of density.

Solids, liquids, and gases can easily be recognized and distinguished by differences in physical properties. Four properties that can be used are density, shape, compressibility, and thermal expansion.

Density was defined in Section 1.11 as the mass of a sample of matter divided by the volume of the same sample. A failure to remember this is the basis for the wrong answers given to the children's riddle, "Which is heavier, a pound of lead or a pound of feathers?" Of course, they weigh the same, but many people say lead is heavier because they think in terms of samples having the same volume. Of course, if you have lead and feather samples of the same volume, the lead sample will be much heavier because its density is much greater.

The **shape** of matter is sometimes independent of a container (for solids), or it may be related to the shape of the container (for liquids and gases). In the case of liquids, the shape depends on the extent to which the container is filled. Gases always fill the container completely (see ◗ Figure 6.2).

shape Shape depends on the physical state of matter.

Compressibility, the change in volume resulting from a pressure change, is quite high for gases. Compressibility allows a lot of gas to be squeezed into a small volume if the gas

compressibility The change in volume of a sample resulting from a pressure change acting on the sample.

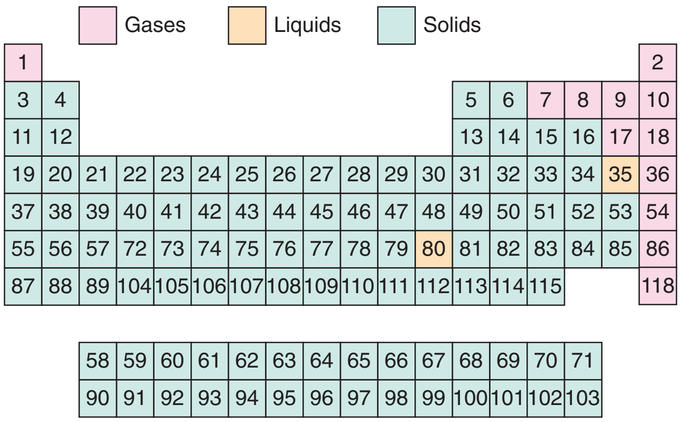


Figure 6.1 The common states of the elements at normal atmospheric pressure and 25°C.

1 Solids have a shape and volume that does not depend on the container.

2 Liquids take the shape of the part of the container they fill. Each sample above has the same volume.

3 Gases completely fill and take the shape of their container. When the valve separating the two parts of the container is opened, the gas fills the entire container volume (bottom photo).

Figure 6.2 Characteristics of the states of matter.

is put under sufficient pressure—think of automobile tires, or a tank of compressed helium used to fill toy balloons.

thermal expansion The change in volume of a sample resulting from a change in temperature of the sample.

Thermal expansion, the change in volume resulting from temperature changes, is a property used in thermometers. As the temperature of the liquid increases, the liquid expands and fills more of the fine capillary tube on which the temperature scale is engraved. You see the liquid move up the tube and know the temperature is increasing.

These four properties are compared for the three states of matter in Table 6.1.

Example 6.1

Samples of plumber's solder, rubbing alcohol, and air are collected. The volume and mass of each sample are determined at 20°C as follows: solder—volume = 28.6 mL, mass = 268.8 g; alcohol—volume = 100.0 mL, mass = 78.5 g; air—volume = 500.0 mL, mass = 0.602 g. Calculate the density of each substance in grams per milliliter.

Solution

The density (d) is obtained by dividing the sample mass (in grams) by the sample volume (in milliliters):

Table 6.1 Physical Properties of Solids, Liquids, and Gases

	State		
Property	**Solid**	**Liquid**	**Gas**
Density	High	High—usually lower than that of corresponding solid	Low
Shape	Definite	Indefinite—takes shape of container to the extent it is filled	Indefinite—takes shape of container it fills
Compressibility	Small	Small—usually greater than that of corresponding solid	Large
Thermal expansion	Very small	Small	Moderate

solder: $d = \dfrac{268.8 \text{ g}}{28.6 \text{ mL}} = 9.40 \text{ g/mL}$

alcohol: $d = \dfrac{78.5 \text{ g}}{100.0 \text{ mL}} = 0.785 \text{ g/mL}$

air: $d = \dfrac{0.602 \text{ g}}{500.0 \text{ mL}} = 0.00120 \text{ g/mL}$

The calculated densities follow the pattern given in Table 6.1.

Learning Check 6.1 Samples of copper, glycerin, and helium are collected and weighed at 20°C. The results are as follows: copper—volume = 12.8 mL, mass = 114.2 g; glycerin—volume = 50.0 mL, mass = 63.0 g; helium—volume = 1500 mL, mass = 0.286 g.

a. Calculate the density of each sample at 20°C.
b. Describe qualitatively (increase, decrease, little change, no change, etc.) what happens to the following properties for each sample when the pressure on the sample is doubled: sample volume, sample mass, sample density (see Table 6.1).

The density calculations done in Example 6.1 used Equation 1.9, which is repeated here:

$$d = \frac{m}{V}$$

Remember, in this equation d is density, m is mass, and V is volume. This useful equation can be rearranged so that any one quantity in it can be calculated if the other two are known.

Example 6.2

The sample of rubbing alcohol used in Example 6.1 is heated to 50°C. The density of the sample is found to decrease to a value of 0.762 g/mL. What is the volume of the sample at 50°C?

Solution

It is assumed that no alcohol is lost by evaporation, so the sample mass has the same value of 78.5 g it had in Example 6.1. Equation 1.9 is rearranged and solved for volume:

$$V = \frac{m}{d} \tag{6.1}$$

The new volume is

$$V = \frac{m}{d} = \frac{78.5 \cancel{\text{g}}}{0.762 \cancel{\text{g}}/\text{mL}} = 103 \text{ mL}$$

Thus, the heating causes the liquid to expand from a volume of 100 mL to 103 mL.

Learning Check 6.2 Calculate the mass in grams of a 1200-mL air sample at a temperature where the air has a density of 1.18×10^{-3} g/mL. NOTE: Equation 6.1 can be rearranged to $m = d \times V$.

6.2 The Kinetic Molecular Theory of Matter

Learning Objective

2. Demonstrate an understanding of the kinetic molecular theory of matter.

The long title of this section is the name scientists have given to a model or theory used to explain the behavior of matter in its various states. Some theories, including this one, are made up of a group of generalizations or postulates. This useful practice makes it possible to study and understand each postulate individually, instead of the entire theory.

The postulates of the kinetic molecular theory are:

1. Matter is composed of tiny particles called molecules.
2. The particles are in constant motion and therefore possess kinetic energy.
3. The particles possess potential energy as a result of attracting or repelling each other.
4. The average particle speed increases as the temperature increases.
5. The particles transfer energy from one to another during collisions in which no net energy is lost from the system.

kinetic energy The energy a particle has as a result of its motion. Mathematically, it is $KE = \frac{1}{2}mv^2$.

These postulates contain two new terms, kinetic energy and potential energy. **Kinetic energy** is the energy a particle has as a result of its motion. Mathematically, kinetic energy is calculated as

$$KE = \frac{1}{2}mv^2 \tag{6.2}$$

where m is the particle mass and v is its velocity. Thus, if two particles of different mass are moving at the same velocity, the heavier particle will possess more kinetic energy than the other particle. Similarly, the faster moving of two particles with equal masses will have more kinetic energy.

Example 6.3

Calculate the kinetic energy of two particles with masses of 2.00 g and 3.00 g if they are both moving with a velocity of 15.0 cm/s.

Solution

The kinetic energy of the 2.00-g particle is

$$KE = \frac{1}{2}(2.00\ \text{g})\left(15.0\ \frac{\text{cm}}{\text{s}}\right)^2 = 225\ \frac{\text{g cm}^2}{\text{s}^2}$$

The kinetic energy of the 3.00-g particle is

$$KE = \frac{1}{2}(3.00\ \text{g})\left(15.0\ \frac{\text{cm}}{\text{s}}\right)^2 = 337.5\ \frac{\text{g cm}^2}{\text{s}^2}$$

Rounding gives 338 g cm^2/s^2; thus, the more massive particle has more kinetic energy.

Learning Check 6.3 Calculate the kinetic energy of two 3.00-g particles if one has a velocity of 10.0 cm/s and the other has a velocity of 20.0 cm/s.

potential energy The energy a particle has as a result of attractive or repulsive forces acting on it.

Potential energy results from attractions or repulsions of particles. A number of these interactions are familiar, such as the gravitational attraction of Earth and the behavior of the poles of two magnets brought near each other. In each of these examples, the size of the force and the potential energy depend on the separation distance. The same behavior is found for the potential energy of atomic-sized particles. The potential energy of attraction increases as separation increases, whereas the potential energy of repulsion decreases with increasing separation.

cohesive force The attractive force between particles; it is associated with potential energy.

disruptive force The force resulting from particle motion; it is associated with kinetic energy.

The kinetic molecular theory provides reasonable explanations for many of the observed properties of matter. An important factor in these explanations is the relative influence of cohesive forces and disruptive forces. **Cohesive forces** are the attractive forces associated with potential energy, and **disruptive forces** result from particle motion (kinetic energy). Disruptive forces tend to scatter particles and make them independent of each other; cohesive forces have the opposite effect. Thus, the state of a substance depends on the relative strengths of the cohesive forces that hold the particles together and the disruptive forces tending to separate them. Cohesive forces are essentially temperature-independent because they involve interparticle attractions of the type described in Chapter 4. Disruptive forces increase with temperature because they arise from particle motion, which increases with temperature (Postulate 4). This explains why temperature plays such an important role in determining the state in which matter is found.

6.3 The Solid State

Learning Objective

3. Use the kinetic molecular theory to explain and compare the properties of matter in different states.

In the solid state, the cohesive forces are stronger than the disruptive forces (see Figure 6.3A). Each particle of a crystalline solid occupies a fixed position in the crystal lattice (see Section 4.11). Disruptive kinetic energy causes the particles to vibrate about their fixed positions, but the strong cohesive forces prevent the lattice from breaking down. The properties of solids in Table 6.1 are explained by the kinetic theory as follows:

High density. The particles of solids are located as closely together as possible. Therefore, large numbers of particles are contained in a small volume, resulting in a high density.
Definite shape. The strong cohesive forces hold the particles of solids in essentially fixed positions, resulting in a definite shape.
Small compressibility. Because there is very little space between particles of solids, increased pressure cannot push them closer together, and it will have little effect on the volume.
Very small thermal expansion. Increased temperature increases the vibrational motion of the particles and the disruptive forces acting on them. Each particle vibrates with an increased amplitude and "occupies" a slightly larger volume. However, there is only a slight expansion of the solid because the strong cohesive forces prevent this effect from becoming very large.

6.4 The Liquid State

Learning Objective

3. Use the kinetic molecular theory to explain and compare the properties of matter in different states.

Particles in the liquid state are randomly packed and relatively close to each other (see Figure 6.3B). They are in constant, random motion, sliding freely over one another, but

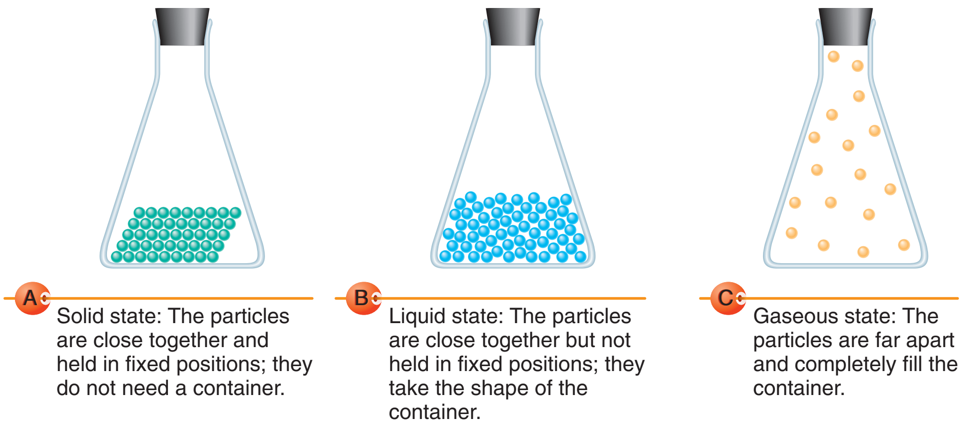


Figure 6.3 A kinetic molecular view of solids, liquids, and gases.

without sufficient kinetic energy to separate completely from each other. The liquid state is a situation in which cohesive forces dominate slightly. The characteristic properties of liquids are also explained by the kinetic theory:

High density. The particles of liquids are not widely separated; they essentially touch each other. There will, therefore, be a large number of particles per unit volume and a high density.

Indefinite shape. Although not completely independent of each other, the particles in a liquid are free to move over and around each other in a random manner, limited only by the container walls and the extent to which the container is filled.

Small compressibility. Because the particles in a liquid essentially touch each other, there is very little space between them. Therefore, increased pressure cannot squeeze the particles much more closely together.

Small thermal expansion. Most of the particle movement in a liquid is vibrational because the particles can move only a short distance before colliding with a neighbor. Therefore, the increased particle velocity that accompanies a temperature increase results only in increased vibration. The net effect is that the particles push away from each other a little more, thereby causing a slight volume increase in the liquid.

6.5 The Gaseous State

Learning Objective

3. Use the kinetic molecular theory to explain and compare the properties of matter in different states.

Disruptive forces completely overcome cohesive forces between particles in the gaseous or vapor state. As a result, the particles of a gas move essentially independently of one another in a totally random way (see Figure 6.3C). Under ordinary pressure, the particles are relatively far apart except when they collide with each other. Between collisions with each other or with the container walls, gas particles travel in straight lines. The particle velocities and resultant collision frequencies are quite high for gases, as shown in Table 6.2.

The kinetic theory explanation of gaseous-state properties follows the same pattern seen earlier for solids and liquids:

Low density. The particles of a gas are widely separated. There are relatively few of them in a given volume, which means there is little mass per unit volume.

Indefinite shape. The forces of attraction between particles have been overcome by kinetic energy, and the particles are free to travel in all directions. The particles, therefore, completely fill the container and assume its inner shape.

Table 6.2 Some Numerical Data Related to the Gaseous State

Gas	Average Speed at 0°C	Average Distance Traveled Between Collisions at 1 atm of Pressure	Number of Collisions of 1 Molecule in 1 s at 1 atm of Pressure
Hydrogen (H_2)	169,000 cm/s (3700 mi/h)	1.12×10^{-5} cm	1.6×10^{10}
Nitrogen (N_2)	45,400 cm/s (1015 mi/h)	0.60×10^{-5} cm	0.80×10^{10}
Carbon dioxide (CO_2)	36,300 cm/s (811 mi/h)	0.40×10^{-5} cm	0.95×10^{10}

Large compressibility. The gas particles are widely separated, so that a gas sample is mostly empty space. When pressure is applied, the particles are easily pushed closer together, decreasing the amount of empty space and the gas volume, as shown by ◗ Figure 6.4.

Moderate thermal expansion. Gas particles move in straight lines except when they collide with each other or with container walls. An increase in temperature causes the particles to collide with more energy. Thus, they push each other away more strongly, and at constant pressure this causes the gas itself to occupy a significantly larger volume.

It must be understood that the size of the particles is not changed during expansion or compression of gases, liquids, or solids. The particles are merely moving farther apart or closer together, and the space between them is changed.

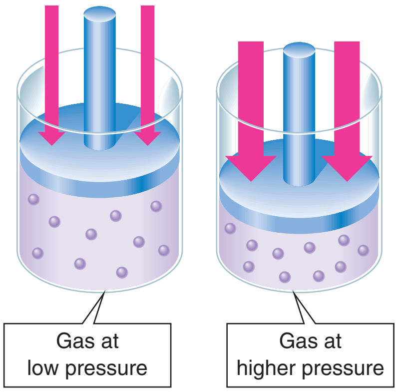


Figure 6.4 The compression of a gas.

6.6 The Gas Laws

Learning Objective

4. Do calculations to convert pressure and temperature values into various units.

The states of matter have not yet been discussed in quantitative terms. We have pointed out that solids, liquids, and gases expand when heated, but the amounts by which they expand have not been calculated. Such calculations for liquids and solids are beyond the intended scope of this text, but gases obey relatively simple quantitative relationships that were discovered during the 17th, 18th, and 19th centuries. These relationships, called **gas laws,** describe in mathematical terms the behavior of gases as they are mixed, subjected to pressure or temperature changes, or allowed to diffuse.

gas law A mathematical relationship that describes the behavior of gases as they are mixed, subjected to pressure or temperature changes, or allowed to diffuse.

To use the gas laws, you must clearly understand the units used to express pressure, and you must utilize the Kelvin temperature scale introduced in Section 1.6. **Pressure** is defined as force per unit area. However, most of the units commonly used to express pressure reflect a relationship to barometric measurements of atmospheric pressure. The mercury barometer was invented by Italian physicist Evangelista Torricelli (1608–1647). Its essential components are shown in ◗ Figure 6.5. A glass tube, sealed at one end, is filled with mercury, stoppered, and inverted so the stoppered end is under the surface of a pool of mercury. When the stopper is removed, the mercury in the tube falls until its weight is just balanced by the weight of air pressing on the mercury pool. The pressure of the atmosphere is then expressed in terms of the height of the supported mercury column.

pressure A force per unit area of surface on which the force acts. In measurements and calculations involving gases, it is often expressed in units related to measurements of atmospheric pressure.

Despite attempts to standardize measurement units, you will probably encounter a number of different units of pressure in your future studies and employment. Some of the more common are standard atmosphere (atm), torr, millimeters of mercury (mmHg), inches of mercury (in. Hg), pounds per square inch (psi), bar, and kilopascals (kPa). One **standard atmosphere** is the pressure needed to support a 760-mm column of mercury in a barometer tube, and 1 **torr** (named in honor of Torricelli) is the pressure needed to support a 1-mm column of mercury in a barometer tube. The relationships of the various units to the standard atmosphere are given in ◗ Table 6.3. Note that the values of 1 atm and 760 torr (or 760 mmHg) are exact numbers based on definitions and do not limit the number of significant figures in calculated numbers.

standard atmosphere The pressure needed to support a 760-mm column of mercury in a barometer tube.

torr The pressure needed to support a 1-mm column of mercury in a barometer tube.

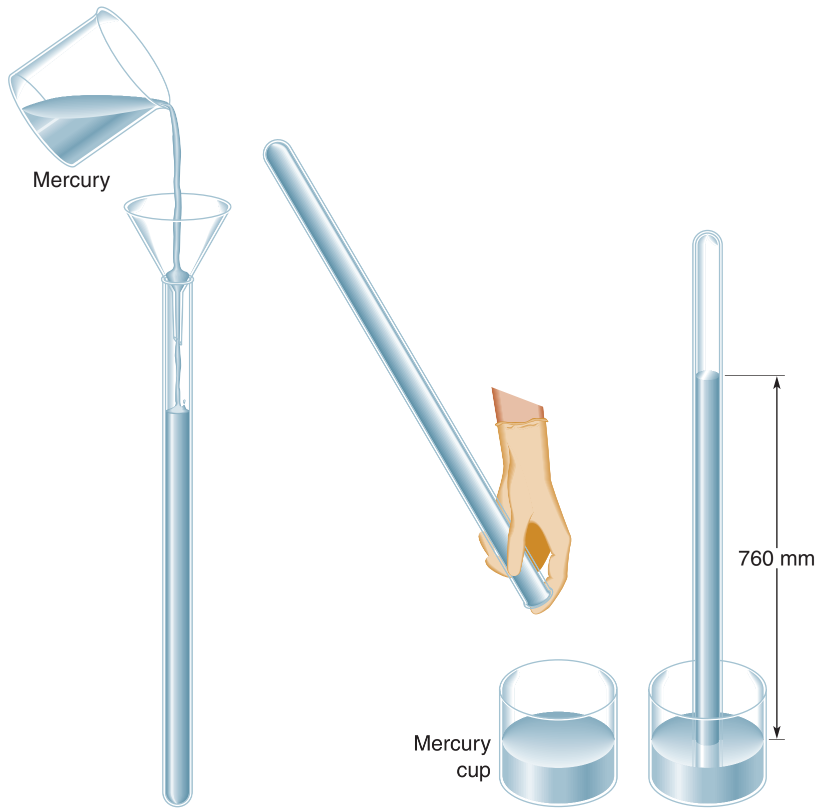
◗ Example 6.4

The gauge on a cylinder of compressed oxygen gas reads 1500 psi. Express this pressure in terms of (a) atm, (b) torr, and (c) mmHg.

Solution

We will use the factor-unit method of calculation from Section 1.9, and the necessary factors from Table 6.3.

Figure 6.5 Setting up a simple mercury barometer. Standard atmospheric pressure is 760 mm of mercury.



a. It is seen from Table 6.3 that 14.7 psi = 1 atm. Therefore, the known quantity, 1500 psi, is multiplied by the factor atm/psi in order to generate units of atm. The result is

$$1500\ \cancel{\text{psi}}\left(\frac{1\ \text{atm}}{14.7\ \cancel{\text{psi}}}\right) = 102\ \text{atm}$$

b. Table 6.3 does not give a direct relationship between psi and torr. However, both psi and torr are related to atm. Therefore, the atm is used somewhat like a "bridge" between psi and torr:

$$(1500\ \cancel{\text{psi}})\left(\frac{1\ \cancel{\text{atm}}}{14.7\ \cancel{\text{psi}}}\right)\left(\frac{760\ \text{torr}}{1\ \cancel{\text{atm}}}\right) = 77{,}551\ \text{torr} = 7.76 \times 10^4\ \text{torr}$$

Note how the "bridge" term canceled out.

c. Because the torr and mmHg are identical, the problem is worked as it was in part b:

$$(1500\ \cancel{\text{psi}})\left(\frac{1\ \cancel{\text{atm}}}{14.7\ \cancel{\text{psi}}}\right)\left(\frac{760\ \text{mmHg}}{1\ \cancel{\text{atm}}}\right) = 7.76 \times 10^4\ \text{mmHg}$$

Learning Check 6.4 A barometer has a pressure reading of 670 torr. Convert this reading into (a) atm and (b) psi.

According to the kinetic molecular theory, a gas expands when it is heated at constant pressure because the gaseous particles move faster at the higher temperature. It makes no difference to the particles what temperature scale is used to describe the heating. However, gas law calculations are based on the Kelvin temperature scale (Section 1.6) rather than the Celsius or Fahrenheit scales. The only apparent difference between the Kelvin and Celsius scales is the location of the zero reading (see Figure 1.11). It should be noted that a 0 reading on the Kelvin scale has a great deal of significance. It is the lowest possible temperature and is called **absolute zero.** It represents the temperature at which particles have no

absolute zero The temperature at which all motion stops; a value of 0 on the Kelvin scale.

Chemistry and Your Health 6.1

Huffing: A Potential Introduction of Children to Drug Abuse

"Huffing" is a general term used to describe various types of inhalant abuse. The inhalants involved are usually liquid chemicals with relatively high vapor pressures. The chemicals are often found in ordinary household products. For example, the propellants or solvents found in such products as hair spray, deodorant spray, spray paint, or cooking oil spray are often used for huffing. Some relatively pure liquids such as paint thinner, cleaning fluid, or fingernail polish remover are also commonly used for huffing.

Parents should be on the lookout for the use of certain words by their children which could indicate that inhalant abuse is taking place. While "huffing" is a general word for inhalant abuse, it also describes a specific technique in which a rag is soaked in a liquid inhalant and the rag is then pressed against the mouth and nose area of the face while the fumes are inhaled. "Sniffing" refers to the practice of sniffing or snorting fumes sprayed directly from an aerosol container. In some cases the contents of an aerosol container might by sprayed directly into the nostrils or mouth. The term "bagging" is used to describe the practice of inhaling fumes from a product that has been sprayed or poured into a plastic or paper bag.

The first effect of huffing, sniffing or bagging is generally a sense of euphoria. However, if an inhalant is abused for a longer time, the initial euphoria is replaced by dizziness, slurred speech and loss of coordination, inhibitions and control. Some individuals become irritable or agitated, and some have delusions or hallucinations. Other potential risks associated with huffing include a rapid, irregular heartbeat that might trigger lethal heart failure. Chronic huffing can result in weakness and fatigue, and serious liver and kidney damage. Permanent brain damage and hearing loss are also possible.

Some of the chemicals frequently used in huffing along with their common sources given in parentheses are: acetone (solvent—paint or hardware store), butane (disposable cigar and cigarette lighters), propane (outdoor grill fuel, spray paint propellent), toluene (solvent—paint or hardware store), methylene chloride (paint stripper—paint or hardware store), Freon (propellent—compressed gas duster) and xylene (solvent—paint or hardware store).

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Propane is often used as a fuel for outdoor grills and space heaters.

Charles D. Winters/Photo Researchers, Inc.

Numerous common products contain solvents used for huffing.

kinetic energy because all motion stops. The Kelvin and Celsius temperature scales have the same size degree, but the 0 of the Kelvin scale is 273 degrees below the freezing point of water, which is the 0 of the Celsius scale. Thus, a Celsius reading can be converted to a Kelvin reading simply by adding 273.

Table 6.3 Units of Pressure

Unit	Relationship to One Standard Atmosphere	Typical Application
Atmosphere	—	Gas laws
Torr	760 torr = 1 atm	Gas laws
Millimeters of mercury	760 mmHg = 1 atm	Gas laws
Pounds per square inch	14.7 psi = 1 atm	Compressed gases
Bar	1.01 bar = 29.9 in. Hg = 1 atm	Meteorology
Kilopascal	101 kPa = 1 atm	Gas laws

Example 6.5

Refer to Figure 1.11 for the necessary conversion factors, and make the following temperature conversions (remember, Kelvin temperatures are expressed in kelvins, K):

a. 37°C (body temperature) to K **b.** −50°C to K **c.** 400 K to °C

Solution

a. Celsius is converted to Kelvin by adding 273, K = °C + 273. Therefore,

$$K = 37°C + 273 = 310\ K\ (\text{body temperature})$$

b. Again, 273 is added to the Celsius reading. However, in this case the Celsius reading is negative, so the addition must be done with the signs in mind. Therefore,

$$K = -50°C + 273 = 223\ K$$

c. Because K = °C + 273, an algebraic rearrangement shows that °C = K − 273. Therefore,

$$°C = 400\ K - 273 = 127°C$$

Learning Check 6.5 Convert the following Celsius temperatures into kelvins, and the Kelvin (K) temperatures into degrees Celsius:

a. 27°C **b.** 0°C **c.** 0 K **d.** 100 K

6.7 Pressure, Temperature, and Volume Relationships

Learning Objective

5. Do calculations based on Boyle's law, Charles's law, and the combined gas law.

Experimental investigations into the behavior of gases as they were subjected to changes in temperature and pressure led to several gas laws that could be expressed by simple mathematical equations. In 1662 Robert Boyle, an Irish chemist, reported his discovery of a relationship between the pressure and volume of a gas sample kept at constant temperature. This relationship, known as **Boyle's law,** is

Boyle's law A gas law that describes the pressure and volume behavior of a gas sample kept at constant temperature. Mathematically, it is $PV = k$.

$$P = \frac{k}{V} \tag{6.3}$$

or

$$PV = k \tag{6.4}$$

where k is an experimentally determined constant, and (remember) the measurements are made without changing the temperature of the gas. Mathematically, the pressure and volume are said to be related by an inverse relationship because the change in one is in a direction opposite (inverse) to the change in the other. That is, as the pressure on a gas sample is increased, the volume of the gas sample decreases.

Another gas law was discovered in 1787 by Jacques Charles, a French scientist. He studied the volume behavior of gas samples kept at constant pressure as they were heated. He found that at constant pressure, the volume of a gas sample was directly proportional to its temperature expressed in kelvins. In other words, if the temperature was doubled, the sample volume doubled as long as the pressure was

Figure 6.6 A balloon collapses when the gas it contains is cooled in liquid nitrogen (temperature = −196°C, or 77 K). Assume the inflated yellow balloon had a volume of 4.0 L at room temperature (25°C) and the prevailing pressure. Assume the pressure remained constant and calculate the volume of the gas in the balloon at the temperature of liquid nitrogen.

kept constant (see Figure 6.6). This behavior, known as **Charles's law,** is represented mathematically as

Charles's law A gas law that describes the temperature and volume behavior of a gas sample kept at constant pressure. Mathematically, it is $V/T = k'$.

$$V = k'T \tag{6.5}$$

or

$$\frac{V}{T} = k' \tag{6.6}$$

where k' is another experimentally determined constant.

Boyle's law and Charles's law can be combined to give a single law called the combined gas law, which provides a relationship between the pressure, volume, and temperature of gases. The **combined gas law** is written as

combined gas law A gas law that describes the pressure, volume, and temperature behavior of a gas sample. Mathematically, it is $PV/T = k''$.

$$\frac{PV}{T} = k'' \tag{6.7}$$

where k'' is another experimentally determined constant.

Equation 6.7 can be put into a useful form by the following line of reasoning. Suppose a gas sample is initially at a pressure and temperature of P_i and T_i and has a volume of V_i. Now suppose that the pressure and temperature are changed to some new (final) values represented by P_f and T_f and that the volume changes to a new value of V_f. According to Equation 6.7,

$$\frac{P_iV_i}{T_i} = k''$$

and

$$\frac{P_fV_f}{T_f} = k''$$

Since both quotients on the left side are equal to the same constant, they can be set equal to each other, and we get

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f} \tag{6.8}$$

Example 6.6

a. A sample of helium gas has a volume of 5.00 L at 25°C and a pressure of 0.951 atm. What volume will the sample have if the temperature and pressure are changed to 50°C and 1.41 atm?

b. A sample of gas has a volume of 3.75 L at a temperature of 25°C and a pressure of 1.15 atm. What will the volume be at a temperature of 35°C and a pressure of 620 torr?

Solution

a. Equation 6.8 will be used to solve this problem. An important step is to identify the quantities that are related and will thus have the same subscripts in Equation 6.8. In this problem, the 5.00-L volume, 25°C temperature, and 0.951 atm pressure are all related and will be used as the initial conditions. The final conditions are the 50°C temperature, the 1.41 atm pressure, and the final volume V_f, which is to be calculated. Substitution of these quantities into Equation 6.8 gives

$$\frac{(0.951 \text{ atm})(5.00 \text{ L})}{(298 \text{ K})} = \frac{(1.41 \text{ atm})V_f}{(323 \text{ K})}$$

Note that the Celsius temperatures have been changed to kelvins. The desired quantity, V_f, can be isolated by multiplying both sides of the equation by 323 K and dividing both sides by 1.41 atm:

$$\frac{(323 \cancel{\text{K}})(0.951 \cancel{\text{atm}})(5.00 \text{ L})}{(298 \cancel{\text{K}})(1.41 \cancel{\text{atm}})} = \frac{(\cancel{323 \text{ K}})(\cancel{1.41 \text{ atm}})V_f}{(\cancel{323 \text{ K}})(\cancel{1.41 \text{ atm}})}$$

or

$$V_f = \frac{(323 \cancel{\text{K}})(0.951 \cancel{\text{atm}})(5.00 \text{ L})}{(298 \cancel{\text{K}})(1.41 \cancel{\text{atm}})} = 3.66 \text{ L}$$

b. This is the same as the problem in part a, except that the initial and final pressures are given in different units. As we saw in part a, the pressure units must cancel in the final calculation, so the initial and final units must be the same. The units of pressure used make no difference as long as they are the same, so we will convert the initial pressure into torr. The necessary conversion factor was obtained from Table 6.3:

$$P_i = (1.15 \cancel{\text{atm}})\left(\frac{760 \text{ torr}}{1 \cancel{\text{atm}}}\right) = 874 \text{ torr}$$

The quantities are substituted into Equation 6.8 to give

$$\frac{(874 \text{ torr})(3.75 \text{ L})}{(298 \text{ K})} = \frac{(620 \text{ torr})(V_f)}{(308 \text{ K})}$$

or

$$V_f = \frac{(308 \cancel{\text{K}})(874 \cancel{\text{torr}})(3.75 \text{ L})}{(298 \cancel{\text{K}})(620 \cancel{\text{torr}})} = 5.46 \text{ L}$$

Example 6.7

a. A steel cylinder has a bursting point of 10,000 psi. It is filled with gas at a pressure of 2500 psi when the temperature is 20°C. During transport, the cylinder is allowed to sit in the hot sun, and its temperature reaches 100°C. Will the cylinder burst?

b. A sample of gas has a volume of 3.00 ft^3 at a temperature of 30°C. If the pressure on the sample remains constant, at what Celsius temperature will the volume be half the volume at 30°C?

Solution

a. The volume of the gas is constant (unless the cylinder bursts), so $V_i = V_f$. With this in mind, Equation 6.8 is used with the following values:
$P_i = 2500$ psi, $V_i = V_f$, $T_i = 20°C + 273$ K, $P_f =$ unknown pressure, and $T_f = 100°C + 273$ K. Substitution into Equation 6.8 gives

$$\frac{(2500 \text{ psi})(V_i)}{293 \text{ K}} = \frac{(P_f)(V_f)}{373 \text{ K}}$$

or

$$P_f = \frac{(2500 \text{ psi})(\cancel{V_f})(373 \cancel{\text{K}})}{(\cancel{V_f})(293 \cancel{\text{K}})} = 3.18 \times 10^3 \text{ psi}$$

The cylinder will not burst.

b. In this case, $P_i = P_f$ because the pressure remains constant. Also, $V_i = 3.00 \text{ ft}^3$; $V_f = \frac{1}{2}V_i$, or 1.50 ft^3; $T_i = 30°C + 273 = 303$ K; and T_f is to be determined. Substitution into Equation 6.8 gives

$$\frac{(P_i)(3.00 \text{ ft}^3)}{(303 \text{ K})} = \frac{(P_f)(1.50 \text{ ft}^3)}{T_f}$$

The quantity to be calculated, T_f, is in the denominator of the right side. We put it into the numerator by inverting both sides of the equation:

$$\frac{(303 \text{ K})}{(P_i)(3.00 \text{ ft}^3)} = \frac{T_f}{(P_f)(1.50 \text{ ft}^3)}$$

The desired quantity, T_f, is now isolated by multiplying both sides of the equation by $(P_f)(1.50 \text{ ft}^3)$:

$$\frac{(\cancel{P_f})(1.50 \cancel{\text{ft}^3})(303 \text{ K})}{(\cancel{P_i})(3.00 \cancel{\text{ft}^3})} = \frac{(\cancel{P_f})(\cancel{1.50 \text{ ft}^3})(T_f)}{(\cancel{P_f})(\cancel{1.50 \text{ ft}^3})}$$

or

$$T_f = \frac{(\cancel{P_f})(1.50 \cancel{\text{ft}^3})(303 \text{ K})}{(\cancel{P_i})(3.00 \cancel{\text{ft}^3})} = 152 \text{ K}$$

Remember, P_f and P_i canceled because they were equal (the pressure was kept constant). The answer is given in kelvins, but we want it in degrees Celsius:

$$\text{K} = °\text{C} + 273$$

or

$$°\text{C} = \text{K} - 273 = 152 - 273 = -121°\text{C}$$

Learning Check 6.6

a. A sample of argon gas is confined in a 10.0-L container at a pressure of 1.90 atm and a temperature of 30°C. What volume would the sample have at 1.00 atm and −10.2°C?

b. A sample of gas has a volume of 500 mL at a temperature and pressure of 300 K and 800 torr. It is desired to compress the sample to a volume of 250 mL at a pressure of 900 torr. What temperature in both kelvins and Celsius degrees will be required?