Introduction to Chemical Bonding

SECTION 1

OBJECTIVES

- Define *chemical bond.*
- Explain why most atoms form chemical bonds.
- Describe ionic and covalent bonding.
- Explain why most chemical bonding is neither purely ionic nor purely covalent.
- Classify bonding type according to electronegativity differences.

Atoms seldom exist as independent particles in nature. The oxygen you breathe, the water you drink, and nearly all other substances consists of combinations of atoms that are held together by chemical bonds. *A* **chemical bond** *is a mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together.*

Why are most atoms chemically bonded to each other? As independent particles, most atoms are at relatively high potential energy. Nature, however, favors arrangements in which potential energy is minimized. This means that most atoms are less stable existing by themselves than when they are combined. By bonding with each other, atoms decrease in potential energy, thereby creating more stable arrangements of matter.

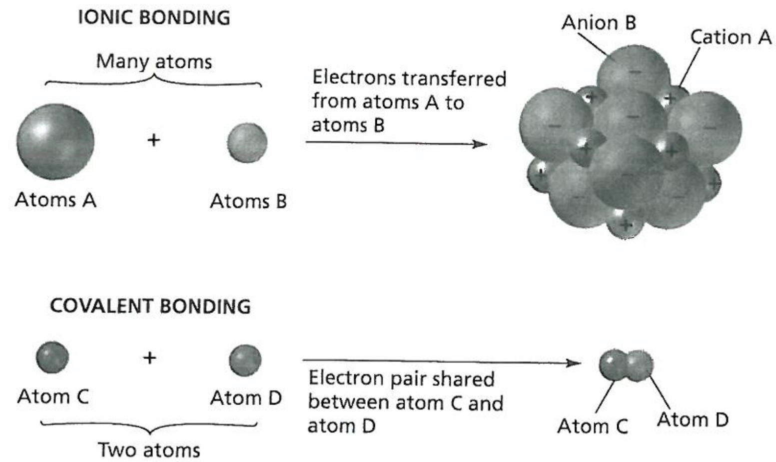
Types of Chemical Bonding

When atoms bond, their valence electrons are redistributed in ways that make the atoms more stable. The way in which the electrons are redistributed determines the type of bonding. In Chapter 5, you read that main-group metals tend to lose electrons to form positive ions, or cations, and nonmetals tend to gain electrons to form negative ions, or anions. *Chemical bonding that results from the electrical attraction between cations and anions is called* **ionic bonding.** In purely ionic bonding, atoms completely give up electrons to other atoms, as illustrated in **Figure 1** on the next page. In contrast to atoms joined by ionic bonding, atoms joined by covalent bonding share electrons. **Covalent bonding** *results from the sharing of electron pairs between two atoms* (see **Figure 1**). In a purely covalent bond, the shared electrons are "owned" equally by the two bonded atoms.

Ionic or Covalent?

Bonding between atoms of different elements is rarely purely ionic or purely covalent. It usually falls somewhere between these two extremes, depending on how strongly the atoms of each element attract electrons. Recall that electronegativity is a measure of an atom's ability to attract electrons. The degree to which bonding between atoms of two elements is ionic or covalent can be estimated by calculating the difference in

FIGURE 1 In ionic bonding, many atoms transfer electrons. The resulting positive and negative ions combine due to mutual electrical attraction. In covalent bonding, atoms share electron pairs to form independent molecules.



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the elements' electronegativities (see **Figure 2**). For example, the electronegativity difference between fluorine, F, and cesium, Cs, is 4.0 − 0.7 = 3.3. (See **Figure 20** on page 161 for a periodic table of electronegativity values.) So, according to **Figure 2,** cesium-fluorine bonding is ionic. Fluorine atoms, which are highly electronegative, gain valence electrons, causing the atoms to become anions. Cesium atoms, which are less electronegative, lose valence electrons, causing the atoms to become cations.

Bonding between atoms with an electronegativity difference of 1.7 or less has an ionic character of 50% or less. These compounds are typically classified as covalent. Bonding between two atoms of the same element is completely covalent. Hydrogen, for example, exists in nature not as isolated atoms, but as pairs of atoms held together by covalent bonds. The hydrogen-hydrogen bond is a **nonpolar-covalent bond,** *a covalent bond in which the bonding electrons are shared equally by the bonded atoms, resulting in a balanced distribution of electrical charge.* Bonds having 0% to 5% ionic character, corresponding to electronegativity differences of roughly 0 to 0.3, are generally considered nonpolar-covalent bonds. In bonds with significantly different electronegativities, the electrons are more strongly attracted by the more-electronegative atom. *Such bonds are* **polar,** *meaning that they have an uneven distribution of charge.* Covalent bonds having 5% to 50% ionic character, corresponding to electronegativity differences of 0.3 to 1.7, are classified as polar. *A* **polar-covalent bond** *is a covalent bond in which the bonded atoms have an unequal attraction for the shared electrons.*

Nonpolar- and polar-covalent bonds are compared in **Figure 3,** which illustrates the electron density distribution in hydrogen-hydrogen and hydrogen-chlorine bonds. The electronegativity difference between chlorine and hydrogen is 3.0 − 2.1 = 0.9, indicating a polar-covalent bond. The electrons in this bond are closer to the more-electronegative chlorine atom than to the hydrogen atom, as indicated in **Figure 3b.** Consequently, the chlorine end of the bond has a partial negative charge, indicated by the symbol $\delta-$. The hydrogen end of the bond then has an equal partial positive charge, $\delta+$.

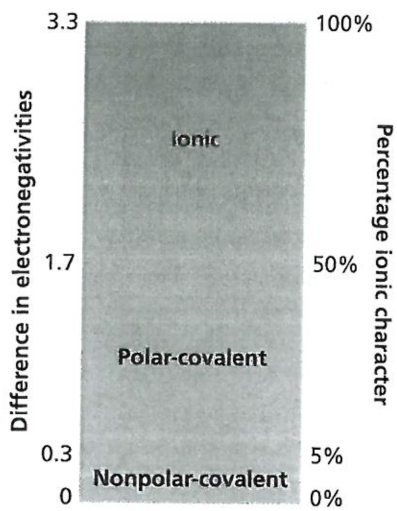

FIGURE 2 Differences in electronegativities reflect the character of bonding between elements. The electronegativity of the less-electronegative element is subtracted from that of the more-electronegative element. The greater the electronegativity difference, the more ionic is the bonding.

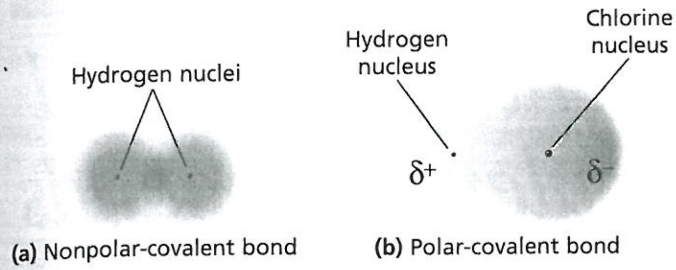

FIGURE 3 Comparison of the electron density in (a) a nonpolar, hydrogen-hydrogen bond and (b) a polar, hydrogen-chlorine bond. Because chlorine is more electronegative than hydrogen, the electron density in the hydrogen-chlorine bond is greater around the chlorine atom.

SAMPLE PROBLEM A

Use electronegativity differences and Figure 2 to classify bonding between sulfur, S, and the following elements: hydrogen, H; cesium, Cs; and chlorine, Cl. In each pair, which atom will be more negative?

SOLUTION From **Figure 20** on page 161, we know that the electronegativity of sulfur is 2.5. The electronegativities of hydrogen, cesium, and chlorine are 2.1, 0.7, and 3.0, respectively. In each pair, the atom with the larger electronegativity will be the more-negative atom.

Bonding between sulfur and	Electronegativity difference	Bond type	More-negative atom
hydrogen	$2.5 - 2.1 = 0.4$	polar-covalent	sulfur
cesium	$2.5 - 0.7 = 1.8$	ionic	sulfur
chlorine	$3.0 - 2.5 = 0.5$	polar-covalent	chlorine

PRACTICE *Answers in Appendix E*

Use electronegativity differences and **Figure 2** to classify bonding between chlorine, Cl, and the following elements: calcium, Ca; oxygen, O; and bromine, Br. Indicate the more-negative atom in each pair.

extension
Go to **go.hrw.com** for more practice problems that ask you to classify bonds.

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SECTION REVIEW

1. What is the main distinction between ionic and covalent bonding?
2. How is electronegativity used in determining the ionic or covalent character of the bonding between two elements?
3. What type of bonding would be expected between the following atoms?
   a. Li and F
   b. Cu and S
   c. I and Br
4. List the three pairs of atoms referred to in the previous question in order of increasing ionic character of the bonding between them.

Critical Thinking

5. **INTERPRETING CONCEPTS** Compare the following two pairs of atoms: Cu and Cl; I and Cl.
   a. Which pair would have a bond with a greater percent ionic character?
   b. In which pair would Cl have the greater negative charge?
6. **INFERRING RELATIONSHIPS** The isolated K atom is larger than the isolated Br atom.
   a. What type of bond is expected between K and Br?
   b. Which ion in the compound KBr is larger?

SECTION 2

Covalent Bonding and Molecular Compounds

OBJECTIVES

- Define *molecule* and *molecular formula.*
- Explain the relationships among potential energy, distance between approaching atoms, bond length, and bond energy.
- State the octet rule.
- List the six basic steps used in writing Lewis structures.
- Explain how to determine Lewis structures for molecules containing single bonds, multiple bonds, or both.
- Explain why scientists use resonance structures to represent some molecules.

Many chemical compounds, including most of the chemicals that are in living things and are produced by living things, are composed of molecules. *A* **molecule** *is a neutral group of atoms that are held together by covalent bonds.* A single molecule of a chemical compound is an individual unit capable of existing on its own. It may consist of two or more atoms of the same element, as in oxygen, or of two or more different atoms, as in water or sugar (see **Figure 4** below). *A chemical compound whose simplest units are molecules is called a* **molecular compound.**

The composition of a compound is given by its chemical formula. *A* **chemical formula** *indicates the relative numbers of atoms of each kind in a chemical compound by using atomic symbols and numerical subscripts.* The chemical formula of a molecular compound is referred to as a molecular formula. *A* **molecular formula** *shows the types and numbers of atoms combined in a single molecule of a molecular compound.* The molecular formula for water, for example, is H_2O, which reflects the fact that a single water molecule consists of one oxygen atom joined by separate covalent bonds to two hydrogen atoms. A molecule of oxygen, O_2, is an example of a diatomic molecule. A *diatomic molecule* is a molecule containing only two atoms.

(a) Water molecule, H_2O

(b) Oxygen molecule, O_2

(c) Sucrose molecule, $C_{12}H_{22}O_{11}$

FIGURE 4 The models for (a) water, (b) oxygen, and (c) sucrose, or table sugar, represent a few examples of the many molecular compounds in and around us. Atoms within molecules may form one or more covalent bonds.

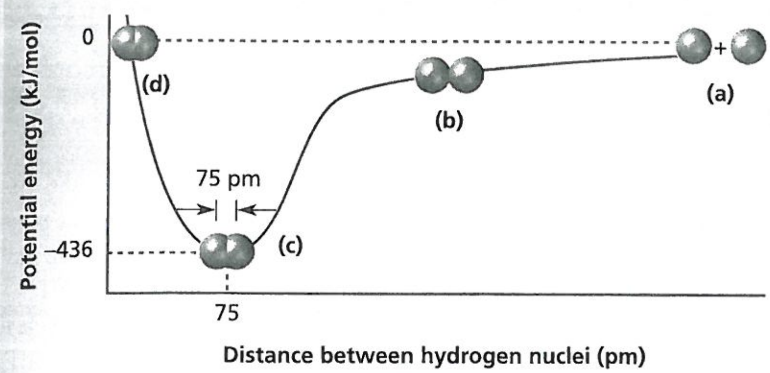

FIGURE 5 Potential energy changes during the formation of a hydrogen-hydrogen bond. (a) The separated hydrogen atoms do not affect each other. (b) Potential energy decreases as the atoms are drawn together by attractive forces. (c) Potential energy is at a minimum when attractive forces are balanced by repulsive forces. (d) Potential energy increases when repulsion between like charges outweighs attraction between opposite charges.

Formation of a Covalent Bond

As you read in Section 1, nature favors chemical bonding because most atoms have lower potential energy when they are bonded to other atoms than they have as they are independent particles. In the case of covalent bond formation, this idea is illustrated by a simple example, the formation of a hydrogen-hydrogen bond.

Picture two isolated hydrogen atoms separated by a distance large enough to prevent them from influencing each other. At this distance, the overall potential energy of the atoms is arbitrarily set at zero, as shown in part (a) of **Figure 5.**

Now consider what happens if the hydrogen atoms approach each other. Each atom has a nucleus containing a single positively charged proton. The nucleus of each atom is surrounded by a negatively charged electron in a spherical 1*s* orbital. As the atoms near each other, their charged particles begin to interact. As shown in **Figure 6,** the approaching nuclei and electrons are *attracted* to each other, which corresponds to a *decrease* in the total potential energy of the atoms. At the same time, the two nuclei *repel* each other and the two electrons *repel* each other, which results in an *increase* in potential energy.

The relative strength of attraction and repulsion between the charged particles depends on the distance separating the atoms. When the atoms first "sense" each other, the electron-proton attraction is stronger than the electron-electron and proton-proton repulsions. Thus, the atoms are drawn to each other and their potential energy is lowered, as shown in part (b) of **Figure 5.**

The attractive force continues to dominate and the total potential energy continues to decrease until, eventually, a distance is reached at which the repulsion between the like charges equals the attraction of the opposite charges. This is shown in part (c) of **Figure 5.** At this point, which is represented by the bottom of the valley in the curve, potential energy is at a minimum and a stable hydrogen molecule forms. A closer approach of the atoms, shown in part (d) of **Figure 5,** results in a sharp rise in potential energy as repulsion becomes increasingly greater than attraction.

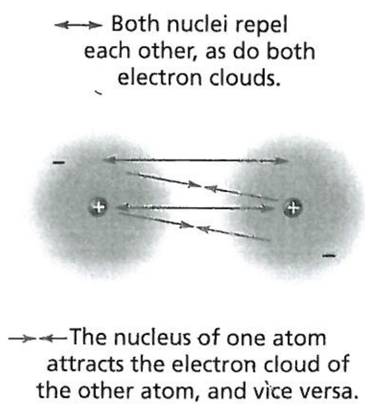

FIGURE 6 The arrows indicate the attractive and repulsive forces between the electrons (shown as electron clouds) and nuclei of two hydrogen atoms. Attraction between particles corresponds to a decrease in potential energy of the atoms, while repulsion corresponds to an increase.

Chemistry in Action

Ultrasonic Toxic-Waste Destroyer

Paints, pesticides, solvents, and sulfides are just a few components of the 3 million tons of toxic waste that flow out of U.S. factories every year. Some of this waste ends up in groundwater and contaminates our streams and drinking water.

Eliminating hazardous waste is a constant challenge. Unfortunately, today's disposal methods often damage the environment as much as they help it. Incinerators burning certain waste, for example, produce dioxins, one of the most dangerous class of toxins known to man.

Finding new methods to destroy toxic waste is a puzzle. Michael Hoffmann, a professor of environmental chemistry at the California Institute of Technology, thinks that part of the solution lies in sound-wave technology.

According to Hoffmann, cavitation is the key to eliminating certain chemical wastes from polluted water. Cavitation occurs when the pressure in water is made to fluctuate from slightly above to slightly below normal, causing bubbles. The bubbles are unstable and collapse, creating tiny areas of extremely high pressure and heat. The pressure inside a collapsing bubble can be 1000 times greater than normal, and the temperature reaches about 5000°C—just a bit cooler than the surface of the sun. These conditions are harsh enough to combust most toxic-waste compounds in the water, breaking them down into harmless components.

Hoffmann has employed a device that uses ultrasound—sound waves at frequencies just above the range of human hearing—to create cavitation in polluted water. As water flows between two panels that generate ultrasound at different frequencies, the ultrasonic waves generated by one panel form cavitation bubbles. An instant later, the ultrasound produced by the other panel collapses the bubbles. The intense pressure and heat generated break down toxic compounds into innocuous substances, such as carbon dioxide, chloride ions, and hydrogen ions.

"With ultrasound," says Hoffmann, "we can harness frequencies . . . of about 16 kilohertz up to 1 megahertz, and different . . . compounds are destroyed more readily at one frequency versus another . . . applying a particular frequency range, we can destroy a very broad range of chemical compounds."

The device destroys simple toxins in a few minutes and other toxins in several hours. To be destroyed completely, some compounds must form intermediate chemicals first and then be treated again. To be sure the waste is totally removed, scientists use sophisticated tracking methods to trace what happens to every single molecule of the toxin.

The ultrasound toxic-waste destroyer treats about 10% of all types of waste, eliminating both organic and inorganic compounds, such as hydrogen cyanide, TNT, and many pesticides. While the device cannot destroy complex mixtures of compounds, such as those found in raw sewage, it does have many advantages over current technologies. Aside from having no harmful environmental side effects, ultrasonic waste destruction is cheaper and simpler than the process of combustion.

Questions

1. How does Dr. Hoffmann's ultrasound device benefit society?
2. Briefly explain why the bulk temperature of the water remains low (at room temperature).

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Characteristics of the Covalent Bond

In **Figure 5,** the bottom of the valley in the curve represents the balance between attraction and repulsion in a stable covalent bond. At this point, the electrons of each hydrogen atom of the hydrogen molecule are shared between the nuclei. As shown below in **Figure 7,** the molecule's electrons can be pictured as occupying overlapping orbitals, moving about freely in either orbital.

The bonded atoms vibrate a bit, but as long as their potential energy remains close to the minimum, they are covalently bonded to each other. The distance between two bonded atoms at their minimum potential energy, that is, the average distance between two bonded atoms, is the *bond length.* The bond length of a hydrogen-hydrogen bond is 75 pm.

In forming a covalent bond, the hydrogen atoms release energy as they change from isolated individual atoms to parts of a molecule. The amount of energy released equals the difference between the potential energy at the zero level (separated atoms) and that at the bottom of the valley (bonded atoms) in **Figure 5.** The same amount of energy must be added to separate the bonded atoms. **Bond energy** *is the energy required to break a chemical bond and form neutral isolated atoms.* Scientists usually report bond energies in kilojoules per mole (kJ/mol), which indicates the energy required to break one mole of bonds in isolated molecules. For example, 436 kJ of energy is needed to break the hydrogen-hydrogen bonds in one mole of hydrogen molecules and form two moles of separated hydrogen atoms.

The energy relationships described here for the formation of a hydrogen-hydrogen bond apply generally to all covalent bonds. However, bond lengths and bond energies vary with the types of atoms that have combined. Even the energy of a bond between the same two types of atoms varies somewhat, depending on what other bonds the atoms have formed. These facts should be considered when examining the data in **Table 1** on the next page. The first three columns in the table list bonds, bond lengths, and bond energies of atoms in specific diatomic molecules. The last three columns give average values of specified bonds in many different compounds.

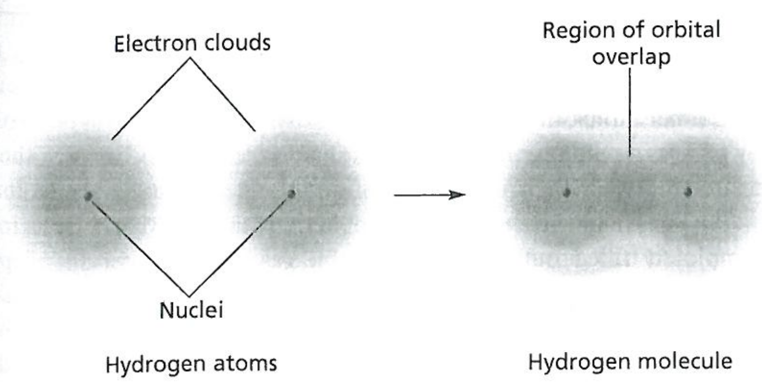


FIGURE 7 The orbitals of the hydrogen atoms in a hydrogen molecule overlap, allowing each electron to feel the attraction of both nuclei. The result is an increase in electron density between the nuclei.

TABLE 1 Bond Lengths and Bond Energies for Selected Covalent Bonds

Bond	Average bond length (pm)	Average bond energy (kJ/mol)	Bond	Average bond length (pm)	Average bond energy (kJ/mol)
H–H	75	436	C–C	154	346
F–F	142	159	C–N	147	305
Cl–Cl	199	243	C–O	143	358
Br–Br	229	193	C–H	109	418
I–I	266	151	C–Cl	177	327
H–F	92	569	C–Br	194	285
H–Cl	127	432	N–N	145	163
H–Br	141	366	N–H	101	386
H–I	161	299	O–H	96	459

All individual hydrogen atoms contain a single, unpaired electron in a 1*s* atomic orbital. When two hydrogen atoms form a molecule, they share electrons in a covalent bond. As **Figure 8** shows, sharing electrons allows each atom to have the stable electron configuration of helium, $1s^2$. This tendency for atoms to achieve noble-gas configurations by bonding covalently extends beyond the simple case of a hydrogen molecule.

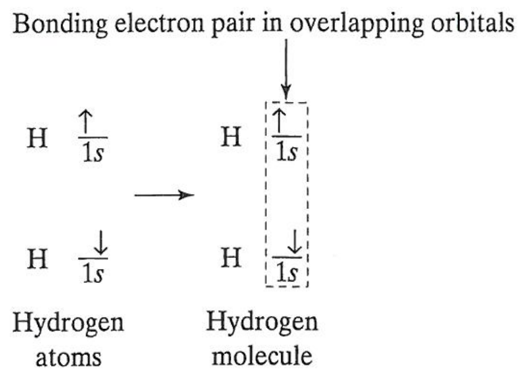


FIGURE 8 By sharing electrons in overlapping orbitals, each hydrogen atom in a hydrogen molecule experiences the effect of a stable $1s^2$ configuration.

The Octet Rule

Unlike other atoms, the noble-gas atoms exist independently in nature. They possess a minimum of energy existing on their own because of the special stability of their electron configurations. This stability results from the fact that, with the exception of helium and its two electrons in a completely filled outer shell, the noble-gas atoms' outer *s* and *p* orbitals are completely filled by a total of eight electrons. Other main-group atoms can effectively fill their outermost *s* and *p* orbitals with electrons by sharing electrons through covalent bonding. Such bond formation follows the *octet rule:* Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an octet of electrons in its highest occupied energy level.

Let's examine how the bonding in a fluorine molecule illustrates the octet rule. An independent fluorine atom has seven electrons in its highest energy level ([He]$2s^22p^5$). Like hydrogen atoms, fluorine atoms bond covalently with each other to form diatomic molecules, F_2. When two fluorine atoms bond, each atom shares one of its valence electrons with its partner. The shared electron pair effectively fills each atom's outermost energy level with an octet of electrons, as illustrated in **Figure 9a. Figure 9b** shows another example of the octet rule, in which the chlorine atom in a molecule of hydrogen chloride, HCl, achieves an outermost octet by sharing an electron pair with an atom of hydrogen.

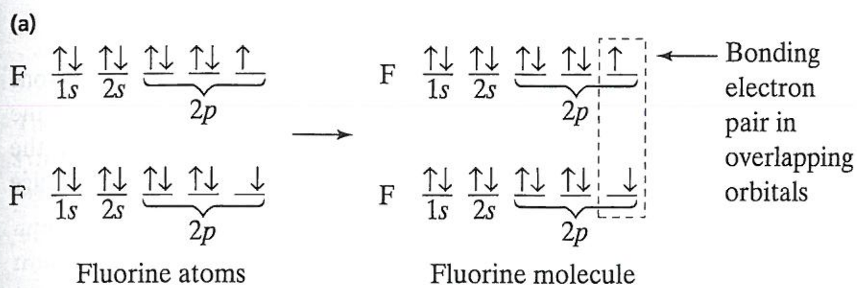
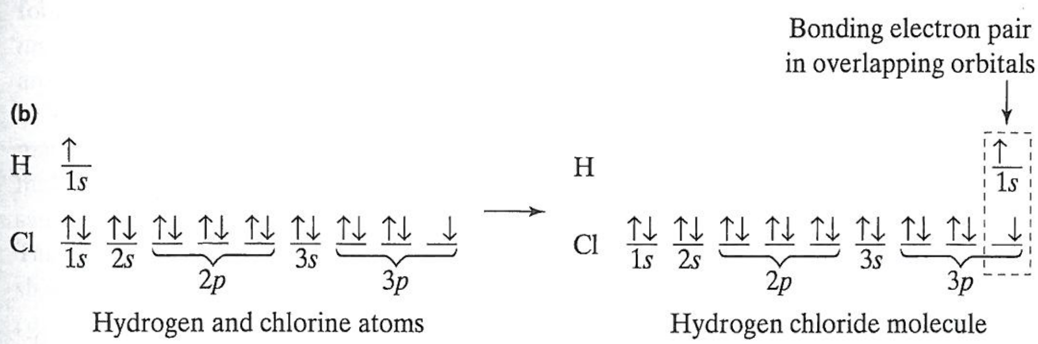


FIGURE 9 (a) By sharing valence electrons in overlapping orbitals, each atom in a fluorine molecule feels the effect of neon's stable configuration, [He]$2s^22p^6$. (b) In a hydrogen chloride molecule, the hydrogen atom effectively fills its 1*s* orbital with two electrons, while the chlorine atom experiences the stability of an outermost octet of electrons.



Exceptions to the Octet Rule

Most main-group elements tend to form covalent bonds according to the octet rule. However, there are exceptions. As you have seen, hydrogen forms bonds in which it is surrounded by only two electrons. Boron, B, has just three valence electrons ([He]$2s^22p^1$). Because electron pairs are shared in covalent bonds, boron tends to form bonds in which it is surrounded by six electrons. In boron trifluoride, BF_3, for example, the boron atom is surrounded by its own three valence electrons plus one from each of the three fluorine atoms bonded to it. Other elements can be surrounded by *more* than eight electrons when they combine with the highly electronegative elements fluorine, oxygen, and chlorine. In these cases of *expanded valence,* bonding involves electrons in *d* orbitals as well as in *s* and *p* orbitals. Examples of compounds that have an expanded valence include PF_5 and SF_6, as shown in **Table 5.**

Number of valence electrons	Electron-dot notation	Example
1	X·	Na·
2	·X·	·Mg·
3	·Ẋ·	·Ḃ·
4	·Ẋ̣·	·Ċ̣·
5	·Ẋ̣:	·Ṇ̇:
6	:Ẋ̣:	:Ọ̇:
7	:Ẍ̣:	:F̣̈:
8	:Ẍ̤:	:N̈ẹ̈:

FIGURE 10 To write an element's electron-dot notation, determine the element's number of valence electrons. Then place a corresponding number of dots around the element's symbol, as shown.

Electron-Dot Notation

Covalent bond formation usually involves only the electrons in an atom's outermost energy levels, or the atom's valence electrons. To keep track of these electrons, it is helpful to use electron-dot notation. **Electron-dot notation** *is an electron-configuration notation in which only the valence electrons of an atom of a particular element are shown, indicated by dots placed around the element's symbol.* The inner-shell electrons are not shown. For example, the electron-dot notation for a fluorine atom (electron configuration [He]$2s^22p^5$) may be written as follows.

:F̣̈:

In general, an element's number of valence electrons can be determined by adding the superscripts of the element's noble-gas notation. In this book, the electron-dot notations for elements with 1–8 valence electrons are written as shown in **Figure 10.**

SAMPLE PROBLEM B *For more help, go to the Math Tutor at the end of this chapter.*

a. **Write the electron-dot notation for hydrogen.**
b. **Write the electron-dot notation for nitrogen.**

SOLUTION a. A hydrogen atom has only one occupied energy level, the $n = 1$ level, which contains a single electron. Therefore, the electron-dot notation for hydrogen is written as follows.

H·

b. The group notation for nitrogen's family of elements is ns^2np^3, which indicates that nitrogen has five valence electrons. Therefore, the electron-dot notation for nitrogen is written as follows.

·Ṇ̇:

Lewis Structures

Electron-dot notation can also be used to represent molecules. For example, a hydrogen molecule, H_2, is represented by combining the notations of two individual hydrogen atoms, as follows.

H:H

The pair of dots represents the shared electron pair of the hydrogen-hydrogen covalent bond. For a molecule of fluorine, F_2, the electron-dot notations of two fluorine atoms are combined.

$$:\underset{\cdot\cdot}{\overset{\cdot\cdot}{F}}:\underset{\cdot\cdot}{\overset{\cdot\cdot}{F}}:$$

Here also the pair of dots between the two symbols represents the shared pair of a covalent bond. In addition, each fluorine atom is surrounded by three pairs of electrons that are not shared in bonds. An *unshared pair,* also called a *lone pair,* is a pair of electrons that is not involved in bonding and that belongs exclusively to one atom.

The pair of dots representing a shared pair of electrons in a covalent bond is often replaced by a long dash. According to this convention, hydrogen and fluorine molecules are represented as follows.

$$\mathrm{H{-}H} \qquad :\underset{\cdot\cdot}{\overset{\cdot\cdot}{F}}-\underset{\cdot\cdot}{\overset{\cdot\cdot}{F}}:$$

These representations are all **Lewis structures,** *formulas in which atomic symbols represent nuclei and inner-shell electrons, dot-pairs or dashes between two atomic symbols represent electron pairs in covalent bonds, and dots adjacent to only one atomic symbol represent unshared electrons.* It is common to write Lewis structures that show only the electrons that are shared, using dashes to represent the bonds. *A* **structural formula** *indicates the kind, number, arrangement, and bonds but not the unshared pairs of the atoms in a molecule.* For example, F−F and H−Cl are structural formulas.

The Lewis structures (and therefore the structural formulas) for many molecules can be drawn if one knows the composition of the molecule and which atoms are bonded to each other. The following sample problem illustrates the basic steps for writing Lewis structures. The molecule described in this problem contains bonds with single shared electron pairs. A single covalent bond, or a **single bond,** *is a covalent bond in which one pair of electrons is shared between two atoms.*

SAMPLE PROBLEM C *For more help, go to the Math Tutor at the end of this chapter.*

Draw the Lewis structure of iodomethane, CH_3I.

SOLUTION 1. *Determine the type and number of atoms in the molecule.*
The formula shows one carbon atom, one iodine atom, and three hydrogen atoms.

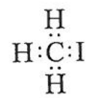
2. *Write the electron-dot notation for each type of atom in the molecule.*
Carbon is from Group 14 and has four valence electrons. Iodine is from Group 17 and has seven valence electrons. Hydrogen has one valence electron.

$$\cdot\underset{\cdot}{\overset{\cdot}{C}}\cdot \qquad :\underset{\cdot}{\overset{\cdot\cdot}{I}}: \qquad H\cdot$$

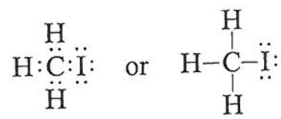
3. *Determine the total number of valence electrons available in the atoms to be combined.*

C	$1 \times 4e^- =$	$4e^-$
I	$1 \times 7e^- =$	$7e^-$
3H	$3 \times 1e^- =$	$3e^-$
		$14e^-$

4. *Arrange the atoms to form a skeleton structure for the molecule. If carbon is present, it is the central atom. Otherwise, the least-electronegative atom is central (except for hydrogen, which is never central). Then connect the atoms by electron-pair bonds.*



5. *Add unshared pairs of electrons to each nonmetal atom (except hydrogen) such that each is surrounded by eight electrons.*



6. *Count the electrons in the structure to be sure that the number of valence electrons used equals the number available. Be sure the central atom and other atoms besides hydrogen have an octet.*
There are eight electrons in the four covalent bonds and six electrons in the three unshared pairs, giving the correct total of 14 valence electrons.

PRACTICE *Answers in Appendix E*

1. Draw the Lewis structure of ammonia, NH_3.
2. Draw the Lewis structure for hydrogen sulfide, H_2S.
3. Draw the Lewis structure for silane, SiH_4.
4. Draw the Lewis structure for phosphorus trifluoride, PF_3.

extension
Go to **go.hrw.com** for more practice problems that ask you to draw Lewis structures.
Keyword: HC6BNDX

Multiple Covalent Bonds

Atoms of some elements, especially carbon, nitrogen, and oxygen, can share more than one electron pair. A double covalent bond, or simply a *double bond,* is a covalent bond in which two pairs of electrons are shared between two atoms. A double bond is shown either by two side-by-side pairs of dots or by two parallel dashes. All four electrons in a double bond "belong" to both atoms. In ethene, C_2H_4, for example, two electron pairs are simultaneously shared by two carbon atoms.