

Theories are nets cast to catch what we call “the world”: to rationalize, to explain, and to master it. We endeavor to make the mesh ever finer and finer.

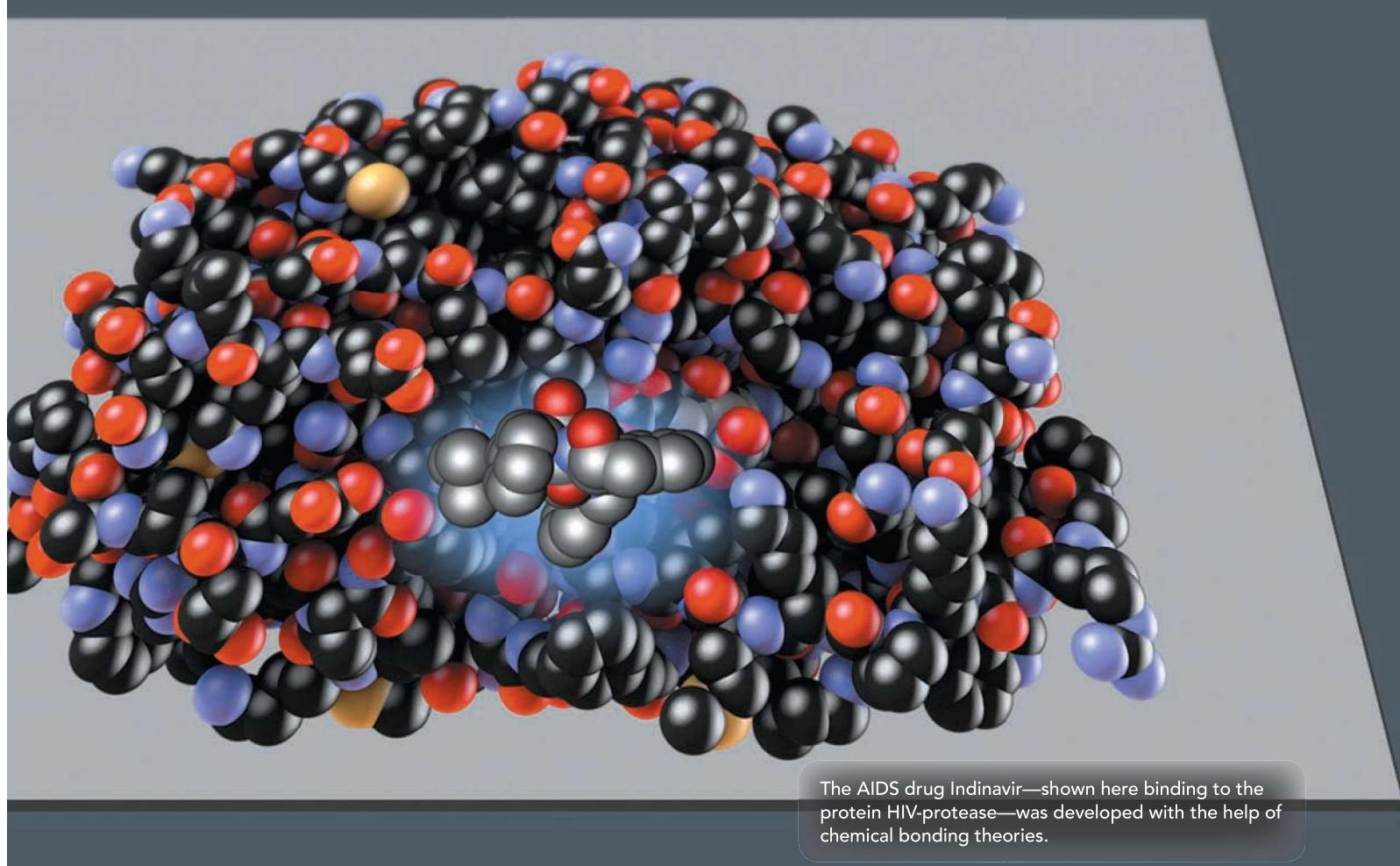
—KARL POPPER (1902–1994)

CHAPTER

10

Chemical Bonding I: The Lewis Model

Chemical bonding is at the heart of chemistry. The bonding theories that we are about to examine are—as Karl Popper, a philosopher of science, eloquently states in the chapter-opening quote—nets cast to understand the world. In the next two chapters, we will examine three theories with successively finer “meshes.” The first is the Lewis model, a simple model of chemical bonding, which can be sketched out on the back of an envelope. With just a few dots, dashes, and chemical symbols, the Lewis model can help us to understand and predict many chemical observations. The second is valence bond theory, which treats electrons in a more quantum-mechanical manner but stops short of viewing them as belonging to the entire molecule. The third is molecular orbital theory, essentially a full quantum-mechanical treatment of the molecule and its electrons as a whole. Molecular orbital theory has great predictive power, but at the expense of great complexity and intensive computational requirements. Which theory is “correct”? Remember that theories are models that help us understand and predict behavior. All three of these theories are extremely useful, depending on exactly what aspect of chemical bonding we want to predict or understand.



The AIDS drug Indinavir—shown here binding to the protein HIV-protease—was developed with the help of chemical bonding theories.

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10.1 Bonding Models and AIDS Drugs

In 1989, researchers using X-ray crystallography—a technique in which X-rays are scattered from crystals of the molecule of interest—determined the structure of a molecule called HIV-protease. HIV-protease is a protein (a class of large biological

We discuss X-ray crystallography in more detail in Section 13.2.

molecules) synthesized by the human immunodeficiency virus (HIV). This particular protein is crucial to the virus's ability to multiply and cause acquired immune deficiency syndrome, or AIDS. Without HIV-protease, HIV cannot spread in the human body because the virus cannot replicate. In other words, without HIV-protease, AIDS can't develop.

We discuss proteins in more detail in Chapter 23.



▲ G. N. Lewis

With knowledge of the HIV-protease structure, pharmaceutical companies set out to create a molecule that would disable HIV-protease by attaching to the working part of the molecule, called the active site. To design such a molecule, researchers used *bonding theories*—models that predict how atoms bond together to form molecules—to simulate the shape of potential drug molecules and determine how they would interact with the protease molecule. By the early 1990s, pharmaceutical companies had developed several drug molecules that seemed to work. Since these molecules inhibit the action of HIV-protease, they were named *protease inhibitors*. Protease inhibitors, when given in combination with other drugs, decrease the viral count in HIV-infected individuals to undetectable levels. Although protease inhibitors do not cure AIDS, many AIDS patients are still alive today because of these drugs.

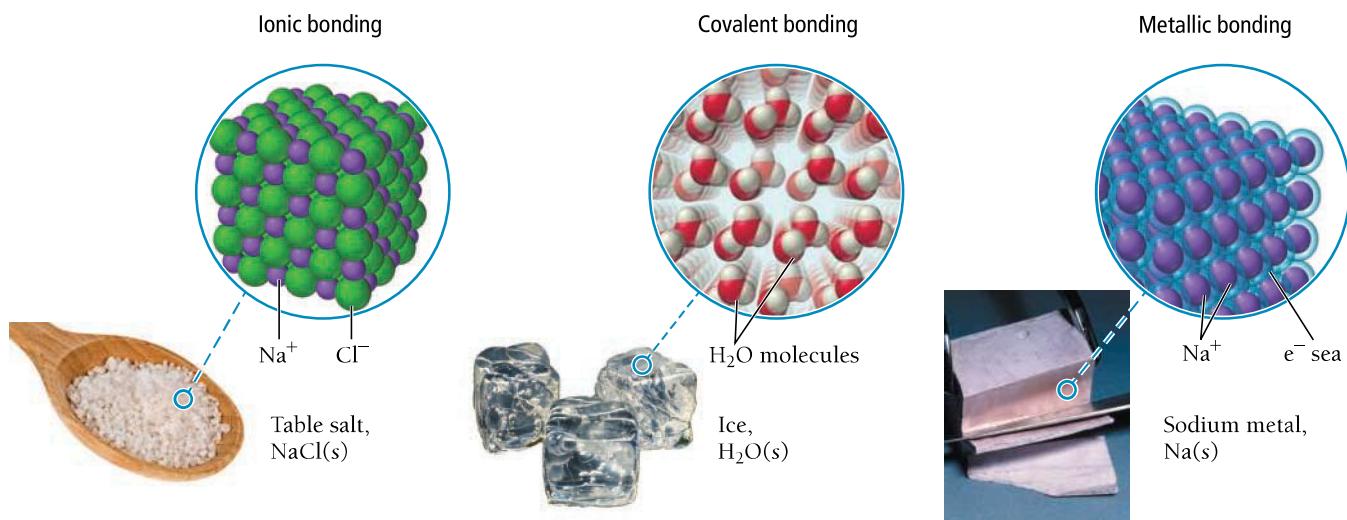
Bonding theories are central to chemistry because they explain how atoms bond together to form molecules. They explain why some combinations of atoms are stable and others are not. For example, bonding theories explain why table salt is NaCl and not NaCl₂ and why water is H₂O and not H₃O. Bonding theories also predict the shapes of molecules—a topic in the next chapter—which in turn determine many of the physical and chemical properties of compounds.

The bonding model we examine in this chapter is the **Lewis model**, named after the American chemist G. N. Lewis (1875–1946). In the Lewis model, valence electrons are represented as dots, and we draw **Lewis electron-dot structures** (or simply **Lewis structures**) to depict molecules. These structures, which are fairly simple to draw, have tremendous predictive power. With minimal computation, the Lewis model can be used to predict whether a particular set of atoms will form a stable molecule and what that molecule might look like. Although we will also examine more advanced theories in the following chapter, the Lewis model remains the simplest model for making quick, everyday predictions about most molecules.

10.2 Types of Chemical Bonds

We begin our discussion of chemical bonding by asking why bonds form in the first place. This seemingly simple question is vitally important. Imagine our universe without chemical bonding. Such a universe would contain only 91 different kinds of substances (the 91 naturally occurring elements). With such a poor diversity of substances, life would be impossible, and we would not be around to wonder why. The *answer* to the question of why bonds form, however, is not simple and involves not only quantum mechanics but also some thermodynamics that we do not introduce until Chapter 19. Nonetheless, we can address an important *aspect* of the answer now: *Chemical bonds form because they lower the potential energy between the charged particles that compose atoms.*

As you already know, atoms are composed of particles with positive charges (the protons in the nucleus) and negative charges (the electrons). When two atoms approach each other, the electrons of one atom are attracted to the nucleus of the other according to Coulomb's law (see Section 9.3) and vice versa. However, at the same time, the electrons of each atom repel the electrons of the other, and the nucleus of each atom repels the nucleus of the other. The result is a complex set of interactions among a potentially large number of charged particles. If these interactions lead to an overall net reduction of energy between the charged particles, a chemical bond forms. Bonding theories help us to predict the circumstances under which bonds form and also the properties of the resultant molecules.



▲ FIGURE 10.1 Ionic, Covalent, and Metallic Bonding

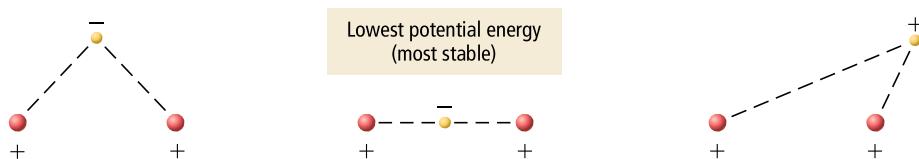
We broadly classify chemical bonds into three types, depending on the kind of atoms involved in the bonding (Figure 10.1▲):

Types of Atoms	Type of Bond	Characteristic of Bond
Metal and nonmetal	Ionic	Electrons transferred
Nonmetal and nonmetal	Covalent	Electrons shared
Metal and metal	Metallic	Electrons pooled

Recall from Chapter 9 that metals tend to have low ionization energies (their electrons are relatively easy to remove) and that nonmetals tend to have negative electron affinities (they readily gain electrons). When a metal atom bonds with a nonmetal atom, it transfers one or more electrons to the nonmetal. The metal atom becomes a cation and the nonmetal atom an anion. These oppositely charged ions attract one another, lowering their overall potential energy as described by Coulomb's law. The resulting bond is an **ionic bond**.

We also discussed in Chapter 9 that nonmetals tend to have high ionization energies (their electrons are relatively difficult to remove). Therefore, when a nonmetal bonds with another nonmetal, neither atom transfers electrons to the other. Instead, the two atoms *share* some electrons. The shared electrons interact with the nuclei of both of the bonding atoms, lowering their potential energy in accordance with Coulomb's law. The resulting bond is a **covalent bond**.

Recall from Section 3.2 that we can understand the stability of a covalent bond by considering the most stable arrangement (the one with the lowest potential energy) of two positively charged particles separated by a small distance and a negatively charged particle. As we can see in Figure 10.2▼, the arrangement in which the negatively charged particle lies *between* the two positively charged ones has the lowest potential energy because in this arrangement, the negatively charged particle interacts most strongly with *both of the positively charged ones*. In a sense, the negatively charged particle holds the two positively charged ones together. Similarly, shared electrons in a covalent chemical bond *hold* the bonding atoms together by attracting the positive charges of their nuclei.



◀ FIGURE 10.2 Possible Configurations of One Negatively Charged Particle and Two Positively Charged Ones

A third type of bonding, **metallic bonding**, occurs in metals. Since metals have low ionization energies, they tend to lose electrons easily. In the simplest model for metallic bonding—called the *electron sea* model—all of the atoms in a metal lattice pool their valence electrons. These pooled electrons are no longer localized on a single atom, but delocalized over the entire metal. The positively charged metal atoms are then attracted to the sea of electrons, holding the metal together. We discuss metallic bonding in more detail in Section 10.11.

ANSWER NOW!



10.1 Cc

Conceptual Connection

BOND TYPES Which compound is most likely to contain ionic bonds?

- (a) CH₄
- (b) N₂O
- (c) MgF₂

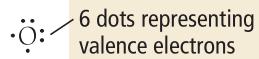
Remember, the number of valence electrons for any main group element is equal to the group number of the element (except for helium, which is in group 8A but has only two valence electrons).

10.3 Representing Valence Electrons with Dots

In Chapter 9, we saw that, for main-group elements, valence electrons are those electrons in the outermost principal energy level. Valence electrons are most important in bonding because they are held most loosely and chemical bonding involves the transfer or sharing of electrons between two or more atoms. Lewis models focus on valence electrons. In a **Lewis symbol**, we represent the valence electrons of main-group elements as dots surrounding the element symbol. For example, the electron configuration of O is:

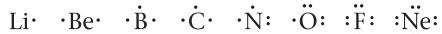


The Lewis symbol for oxygen is:



Although the exact location of dots is not critical, in this book we first place dots singly before pairing (except for helium, which always has two paired dots signifying its duet).

Each dot in a Lewis symbol represents a valence electron. The dots are placed around the element's symbol with a maximum of two dots per side. We draw the Lewis symbols for all of the period 2 elements similarly.



Lewis symbols provide a simple way to visualize the number of valence electrons in a main-group atom. Notice that atoms with eight valence electrons—which are particularly stable because they have full s and p sublevels—are easily identified because they have eight dots, an **octet**.

Helium is somewhat of an exception. Its electron configuration and Lewis symbol are:



The Lewis symbol of helium contains only two dots (a **duet**). For helium, a duet represents a stable electron configuration because the $n = 1$ quantum level fills with only two electrons.

In the Lewis model, a **chemical bond** is the sharing or transfer of electrons to attain stable electron configurations for the bonding atoms. If electrons are transferred, as occurs between a metal and a nonmetal, the bond is an *ionic bond*. If the electrons are shared, as occurs between two nonmetals, the bond is a *covalent bond*. In either case, the bonding atoms attain stable electron configurations; since the stable configuration is usually eight electrons in the outermost shell, this is known as the **octet rule**.

When applying the Lewis model, we do not try to calculate the energies associated with the attractions and repulsions between electrons and nuclei on neighboring atoms. The energy changes that occur because of these interactions are central to chemical

bonding (as we saw in Section 10.2), yet the Lewis model ignores them because calculating these energy changes is extremely complicated. Instead the Lewis model uses the simple octet rule, a practical approach that accurately predicts what we see in nature for a large number of compounds—hence the success and longevity of the Lewis model.

LEWIS SYMBOLS

What is the Lewis symbol for silicon?

- (a) $\cdot\text{Si}\cdot$
- (b) $:\ddot{\text{Si}}:$
- (c) $:\ddot{\text{Si}}:$
- (d) $\cdot\dot{\text{Si}}\cdot$

10.2
Cc

Conceptual Connection

ANSWER NOW!



10.4

Ionic Bonding: Lewis Symbols and Lattice Energies

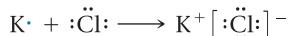
Although the Lewis model's strength is in modeling covalent bonding, it can also be applied to ionic bonding. To represent ionic bonding, we move electron dots from the Lewis symbol of the metal to the Lewis symbol of the nonmetal and allow the resultant ions to form a crystalline lattice composed of alternating cations and anions.

Ionic Bonding and Electron Transfer

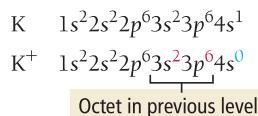
Consider potassium and chlorine, which have the following Lewis symbols:



When these atoms bond, potassium transfers its valence electron to chlorine:



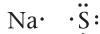
The transfer of the electron gives chlorine an octet (shown as eight dots around chlorine) and leaves potassium without any valence electrons but with an octet in the *previous* principal energy level (which is now the outermost level):



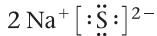
The potassium, having lost an electron, becomes positively charged (a cation), while the chlorine, which has gained an electron, becomes negatively charged (an anion). We typically write the Lewis symbol of an anion within brackets with the charge in the upper right-hand corner, outside the brackets. The positive and negative charges attract one another, resulting in the compound KCl.

This example shows how we can use the Lewis model to predict the correct chemical formulas for ionic compounds. For the compound that forms between K and Cl, for example, the Lewis model predicts a ratio of one potassium cation to every one chloride anion, KCl. In nature, when we examine the compound formed between potassium and chlorine, we indeed find one potassium ion to every chloride ion.

As another example, consider the ionic compound formed between sodium and sulfur. The Lewis symbols for sodium and sulfur are:



Sodium must lose its one valence electron in order to have an octet (in the previous principal shell), while sulfur must gain two electrons to attain an octet. Consequently, the compound that forms between sodium and sulfur requires a ratio of two sodium atoms to every one sulfur atom:



The two sodium atoms each lose their one valence electron, while the sulfur atom gains two electrons and attains an octet. The Lewis model predicts that the correct chemical formula is Na_2S , which is exactly what we see in nature.

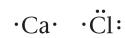
Recall that solid ionic compounds do not contain distinct molecules; they are composed of alternating positive and negative ions in a three-dimensional crystalline array.

Using Lewis Symbols to Predict the Chemical Formula of an Ionic Compound

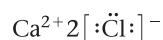
EXAMPLE 10.1 Use Lewis symbols to predict the formula for the compound that forms between calcium and chlorine.

SOLUTION

Draw Lewis symbols for calcium and chlorine based on their respective numbers of valence electrons, which you can determine from their group number in the periodic table.



Calcium must lose its two valence electrons (to end up with an octet in its previous principal shell), while chlorine only needs to gain one electron to attain an octet. Draw two chloride anions, each with an octet and a 1– charge, and one calcium cation with a 2+ charge. Place brackets around the chloride anions and indicate the charges on each ion.



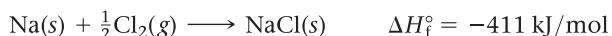
Finally, write the formula with subscripts to indicate the number of each ion.



FOR PRACTICE 10.1 Use Lewis symbols to predict the formula for the compound that forms between magnesium and nitrogen.

Lattice Energy: The Rest of the Story

The formation of an ionic compound from its constituent elements is usually quite exothermic. For example, when sodium chloride (table salt) forms from elemental sodium and chlorine, 411 kJ of heat evolves in the following violent reaction:



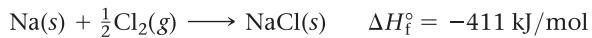
Where does this energy come from? It might seem that it comes solely from the tendency of metals to lose electrons and nonmetals to gain electrons—but it does not. In fact, the transfer of an electron from sodium to chlorine—by itself—actually *absorbs* energy. The first ionization energy of sodium is +496 kJ/mol, and the electron affinity of Cl is only –349 kJ/mol. Based only on these energies, the reaction should be *endothermic* by +147 kJ/mol. So why is the reaction so *exothermic*?

The answer lies in the **lattice energy**—the energy associated with the formation of a crystalline lattice of alternating cations and anions from the gaseous ions. Because the sodium ions are positively charged and the chlorine ions are negatively charged, the potential energy decreases—as prescribed by Coulomb's law—when these ions come together to form a lattice. That energy is emitted as heat when the lattice forms, as shown in Figure 10.3▶. The exact value of the lattice energy, however, is not simple to determine because it involves a large number of interactions among many charged particles in a lattice. The most straightforward way to calculate lattice energy is with the *Born-Haber cycle*.

The Born-Haber Cycle

The **Born-Haber cycle** is a hypothetical series of steps that represents the formation of an ionic compound from its constituent elements. The steps are chosen so that the change in enthalpy of each step is known except for the last one, which is the lattice energy. The change in enthalpy for the overall process is also known. Using Hess's law (see Section 7.8), we can therefore determine the enthalpy change for the unknown last step, the lattice energy.

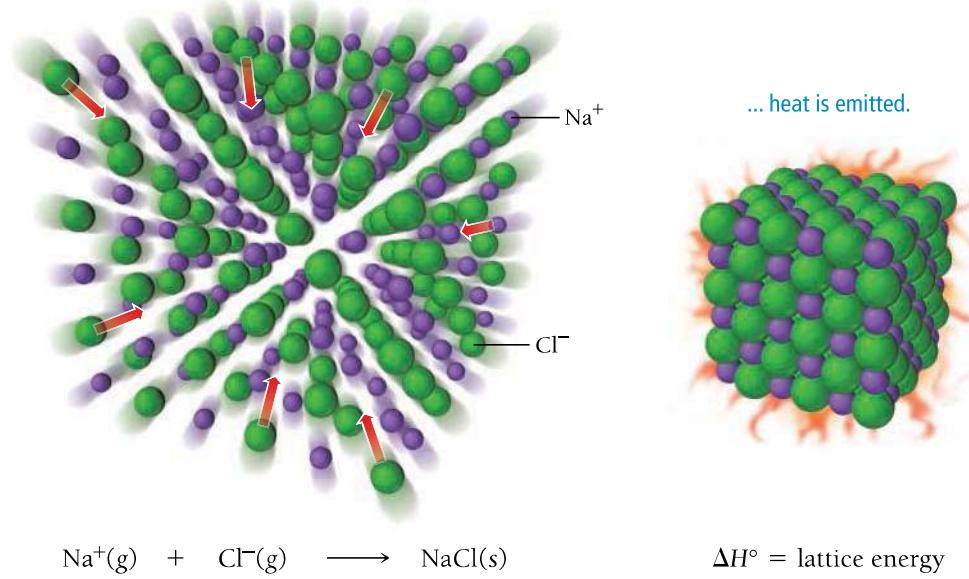
Consider the formation of NaCl from its constituent elements in their standard states. The enthalpy change for the overall reaction is simply the standard enthalpy of formation of NaCl(s):



Recall that Hess's law states that the change in the overall enthalpy of a stepwise process is the sum of the enthalpy changes of the steps.

Lattice Energy of an Ionic Compound

As gaseous ions coalesce...



▲ FIGURE 10.3 Lattice Energy The lattice energy of an ionic compound is the energy associated with the formation of a crystalline lattice of the compound from the gaseous ions.

Now consider the following set of steps—the Born–Haber cycle—from which $\text{NaCl}(\text{s})$ can also be made from $\text{Na}(\text{s})$ and $\text{Cl}_2(\text{g})$:

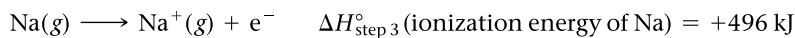
- The first step is the formation of gaseous sodium from solid sodium.



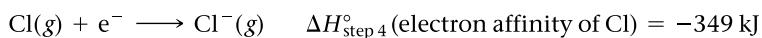
- The second step is the formation of a chlorine atom from a chlorine molecule.



- The third step is the ionization of gaseous sodium. The enthalpy change for this step is the ionization energy of sodium.



- The fourth step is the addition of an electron to gaseous chlorine. The enthalpy change for this step is the electron affinity of chlorine.



- The fifth and final step is the formation of the crystalline solid from the gaseous ions. The enthalpy change for this step is the lattice energy, the unknown quantity.

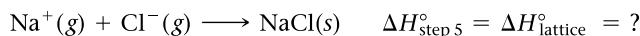
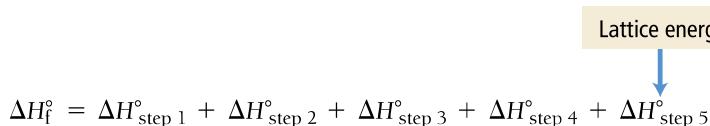
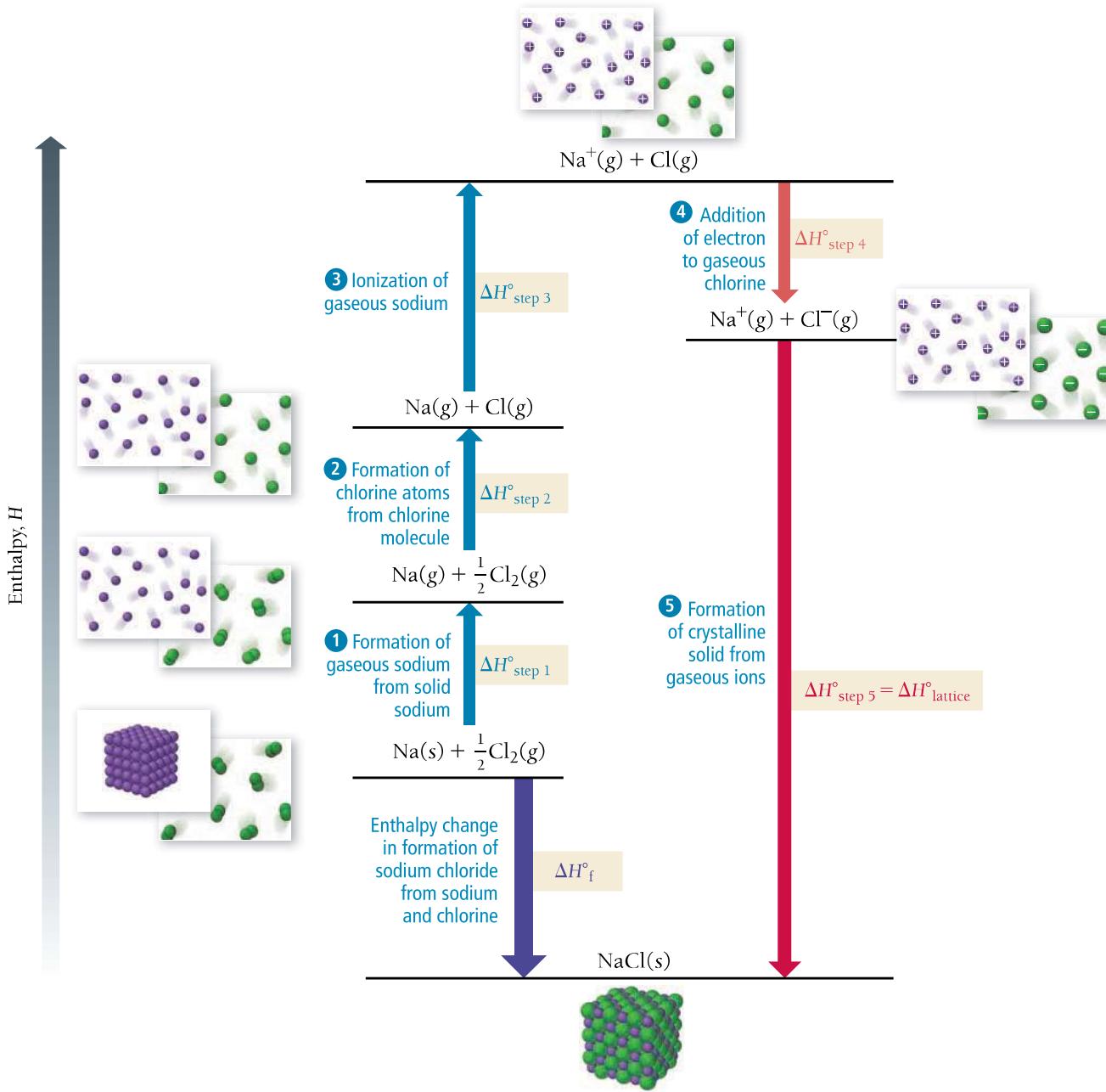


Figure 10.4► illustrates the entire Born–Haber cycle for NaCl .

Since the overall reaction obtained by summing the steps in the Born–Haber cycle is equivalent to the formation of NaCl from its constituent elements, we use Hess's law to set the overall enthalpy of formation for $\text{NaCl}(\text{s})$ equal to the sum of the steps in the Born–Haber cycle:



Born–Haber Cycle for Production of NaCl from Na(s) and Cl₂(g)



▲ FIGURE 10.4 Born–Haber Cycle for Sodium Chloride The sum of the steps is the formation of NaCl from elemental Na and Cl₂. The enthalpy change of the last step (step 5) is the lattice energy.

We then solve this equation for $\Delta H_{\text{step 5}}^{\circ}$, which is $\Delta H_{\text{lattice}}^{\circ}$, and substitute the appropriate values to calculate the lattice energy:

$$\begin{aligned}\Delta H_{\text{lattice}}^{\circ} &= \Delta H_{\text{step 5}}^{\circ} = \Delta H_f^{\circ} - (\Delta H_{\text{step 1}}^{\circ} + \Delta H_{\text{step 2}}^{\circ} + \Delta H_{\text{step 3}}^{\circ} + \Delta H_{\text{step 4}}^{\circ}) \\ &= -411 \text{ kJ} - (+108 \text{ kJ} + 122 \text{ kJ} + 496 \text{ kJ} - 349 \text{ kJ}) \\ &= -788 \text{ kJ}\end{aligned}$$

The value of the lattice energy is a large negative number. The formation of the crystalline NaCl lattice from sodium cations and chloride anions is highly exothermic and more than compensates for the endothermicity of the electron transfer process. In other

Some books define the lattice energy as the energy associated with separating one mole of an ionic lattice into its gaseous constituents, which makes the sign of the lattice energy positive instead of negative.

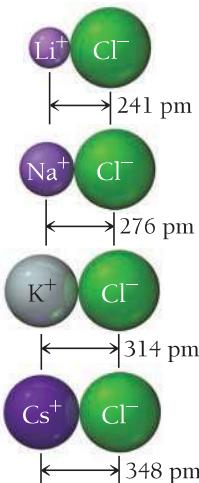
words, the formation of ionic compounds is not exothermic because sodium “wants” to lose electrons and chlorine “wants” to gain them; rather, it is exothermic because of the large amount of heat released when sodium and chloride ions coalesce to form a crystalline lattice.

Trends in Lattice Energies: Ion Size

Consider the lattice energies of the following alkali metal chlorides:

Metal Chloride	Lattice Energy (kJ/mol)
LiCl	-834
NaCl	-788
KCl	-701
CsCl	-657

Why does the magnitude of the lattice energy decrease as we move down the column? We know from the periodic trends discussed in Chapter 9 that ionic radius increases as we move down a column in the periodic table (see Section 9.7). We also know, from our discussion of Coulomb’s law in Section 9.3, that the potential energy of oppositely charged ions becomes less negative (or more positive) as the distance between the ions increases. As the size of the alkali metal ions increases down the column, so does the distance between the metal cations and the chloride anions. The magnitude of each successive lattice energy of the chlorides decreases accordingly, making the formation of the chlorides less exothermic as you move down the column. In other words, *as the ionic radii increase as we move down the column, the ions cannot get as close to each other and therefore do not release as much energy when the lattice forms.*

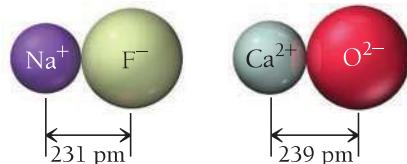


▲ Bond lengths of the group 1A metal chlorides.

Trends in Lattice Energies: Ion Charge

Consider the lattice energies of the following two compounds:

Compound	Lattice Energy (kJ/mol)
NaF	-910
CaO	-3414



Why is the magnitude of the lattice energy of CaO so much greater than the lattice energy of NaF? Na^+ has a radius of 95 pm and F^- has a radius of 136 pm, resulting in a distance between ions of 231 pm. Ca^{2+} has a radius of 99 pm and O^{2-} has a radius of 140 pm, resulting in a distance between ions of 239 pm. Even though the separation between the calcium and oxygen is slightly greater (which would tend to lower the lattice energy), the lattice energy for CaO is almost four times *greater*.

The explanation lies in the charges of the ions. Recall from Coulomb’s law that the magnitude of the potential energy of two interacting charges depends not only on the distance between the charges, but also on the product of the charges:

$$E = \frac{1}{4\pi\epsilon_0} \frac{q_1 q_2}{r}$$

For NaF, E is proportional to $(1+)(1-) = 1-$, while for CaO, E is proportional to $(2+)(2-) = 4-$, so the relative stabilization for CaO relative to NaF is roughly four times greater, as observed in the lattice energy.

Summarizing Trends in Lattice Energies:

- Lattice energies become less exothermic (less negative) with increasing ionic radius.
- Lattice energies become more exothermic (more negative) with increasing magnitude of ionic charge.

EXAMPLE 10.2 Predicting Relative Lattice Energies

Arrange these ionic compounds in order of increasing *magnitude* of lattice energy: CaO, KBr, KCl, SrO.

SOLUTION

KBr and KCl have lattice energies of smaller magnitude than CaO and SrO because of their lower ionic charges (1+, 1– compared to 2+, 2–). When you compare KBr and KCl, you expect KBr to have a lattice energy of lower magnitude due to the larger ionic radius of the bromide ion relative to the chloride ion. Between CaO and SrO, you expect SrO to have a lower magnitude lattice energy due to the larger ionic radius of the strontium ion relative to the calcium ion.

Order of increasing *magnitude* of lattice energy:



Actual lattice energy values:

Compound	Lattice Energy (kJ/mol)
KBr	−671
KCl	−701
SrO	−3217
CaO	−3414

FOR PRACTICE 10.2 Arrange the following in order of increasing magnitude of lattice energy: LiBr, KI, and CaO.

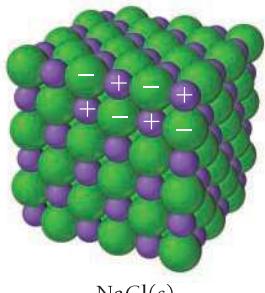
FOR MORE PRACTICE 10.2 Which compound has a higher magnitude lattice energy: NaCl or MgCl₂?

Ionic Bonding: Models and Reality

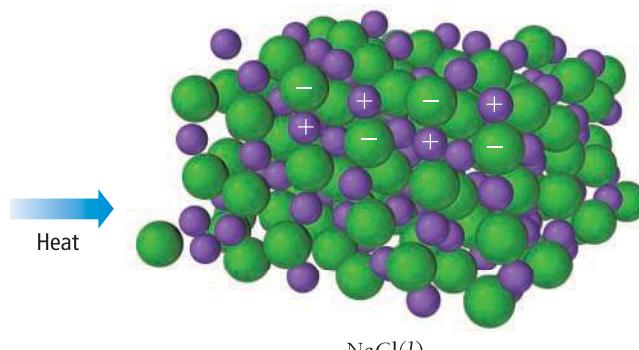
In this section, we have developed a model for ionic bonding. The value of a model is in how well it accounts for what we see in nature (through experiments). Does our ionic bonding model explain the properties of ionic compounds, including their high melting and boiling points, their tendency *not to conduct* electricity as solids, and their tendency *to conduct* electricity when dissolved in water?

We modeled an ionic solid as a lattice of individual ions held together by coulombic forces that are *nondirectional* (which means that, as we move away from the center of an ion, the forces are equally strong in all directions). To melt the solid, these forces must be overcome, which requires a significant amount of heat. Therefore, our model accounts for the high melting points of ionic solids. In the model, electrons transfer from the metal to the nonmetal, but the transferred electrons remain localized on one atom. In other words, our model does not include any free electrons that might conduct electricity (the movement or flow of electrons or other charged particles in response to an electric potential, or voltage, is electrical current). In addition, the ions themselves are fixed in place; therefore, our model accounts for the nonconductivity of ionic solids. When our idealized ionic solid dissolves in water, however, the cations and anions dissociate, forming free ions in solution. These ions can move in response to electrical forces, creating an electrical current. Thus, our model predicts that solutions of ionic compounds conduct electricity (which in fact they do).

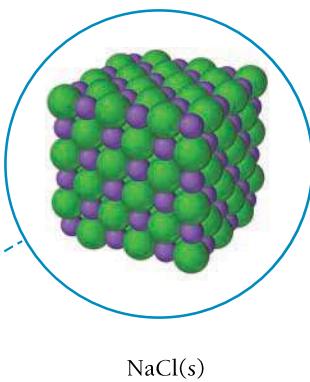
Ions are fixed in place.



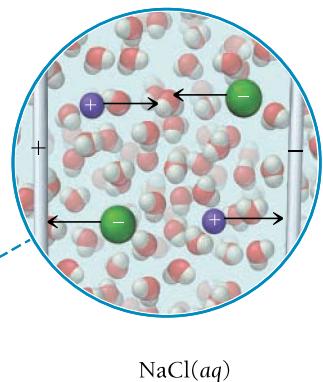
Ions can move.



► The melting of solid ionic compounds such as sodium chloride requires enough heat to overcome the electrical forces holding the anions and cations together in a lattice. Thus, the melting points of ionic compounds are relatively high.



▲ Solid sodium chloride does not conduct electricity.



▲ Aqueous sodium chloride conducts electricity.

MELTING POINTS OF IONIC SOLIDS Use the ionic bonding model to determine which has the higher melting point: NaCl or MgO. Explain the relative ordering.

- (a) MgO has a higher melting point because the ions have larger charges.
- (b) MgO has a higher melting point because the ions have smaller charges.
- (c) NaCl has a higher melting point because the ions have larger charges.
- (d) NaCl has a higher melting point because the ions have smaller charges.

10.3

Cc

Conceptual Connection

ANSWER NOW!



CHEMISTRY AND MEDICINE

Ionic Compounds in Medicine

Although most drugs are molecular compounds, a number of ionic compounds have medical uses. Consider the following partial list of ionic compounds used in medicine. Notice that many of these compounds contain polyatomic ions. The bonding between a metal and a polyatomic ion is ionic. However, the bonding within a polyatomic ion is covalent, the topic of our next section.



Formula	Name	Medical Use
AgNO_3	Silver nitrate	Topical anti-infective agent; in solution, used to treat and prevent eye infection, especially in newborn infants
BaSO_4	Barium sulfate	Given as a contrast medium—or image enhancer—in X-rays
CaSO_4	Calcium sulfate	Used to make plaster casts
KMnO_4	Potassium permanganate	Topical anti-infective agent; often used to treat fungal infections on the feet
KI	Potassium iodide	Antiseptic and disinfectant; given orally to prevent radiation sickness
Li_2CO_3	Lithium carbonate	Used to treat bipolar (manic-depressive) disorders
MgSO_4	Magnesium sulfate	Used to treat eclampsia (a condition that can occur during pregnancy in which elevated blood pressure leads to convulsions)
Mg(OH)_2	Magnesium hydroxide	Antacid and mild laxative
NaHCO_3	Sodium bicarbonate	Oral antacid used to treat heartburn and acid stomach; injected into blood to treat severe acidosis (acidification of the blood)
NaF	Sodium fluoride	Used to strengthen teeth
ZnO	Zinc oxide	Used as protection from ultraviolet light in sun



WATCH NOW!**KEY CONCEPT VIDEO 10.5**

 The Lewis Model for Chemical Bonding
10.5**Covalent Bonding: Lewis Structures**

The Lewis model provides us with a simple and useful model for covalent bonding. In this model, we represent molecular compounds with *Lewis structures*, which depict neighboring atoms as sharing some (or all) of their valence electrons in order to attain octets (or duets for hydrogen).

Single Covalent Bonds

To see how covalent bonding is conceived in terms of the Lewis model, consider hydrogen and oxygen, which have the following Lewis symbols:



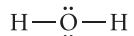
In water, these atoms share their unpaired valence electrons so that each hydrogen atom gets a duet and the oxygen atom gets an octet as represented with this Lewis structure:



The shared electrons—those that appear in the space between the two atoms—count toward the octets (or duets) of *both of the atoms*.

A shared pair of electrons is called a **bonding pair**, whereas a pair that is associated with only one atom—and therefore not involved in bonding—is a **lone pair** (shown at left). Lone pair electrons are also called **nonbonding electrons**.

We often represent a bonding pair of electrons by a dash to emphasize that it constitutes a chemical bond:



The Lewis model also shows why the halogens form diatomic molecules. Consider the Lewis symbol for chlorine:



If two Cl atoms pair together, they can each have an octet:

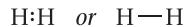


Elemental chlorine does indeed exist as a diatomic molecule in nature, just as the Lewis model predicts. The same is true for the other halogens.

Similarly, the Lewis model predicts that hydrogen, which has the Lewis symbol:



should exist as H₂. When two hydrogen atoms share their valence electrons, each gets a duet, a stable configuration for hydrogen:



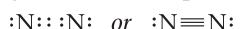
Again, the Lewis model predicts correctly. In nature, elemental hydrogen exists as H₂ molecules.

Double and Triple Covalent Bonds

In the Lewis model, two atoms may share more than one electron pair to get octets. For example, two oxygen atoms share two electron pairs in order for each oxygen atom to have an octet:



Each oxygen atom has an octet because *the additional bonding pair counts toward the octet of both oxygen atoms*. When two atoms share two electron pairs, the resulting bond is a **double bond**. In general, double bonds are shorter and stronger than single bonds. Atoms can also share three electron pairs. Consider the Lewis structure of N₂. Since each N atom has five valence electrons, the Lewis structure for N₂ has 10 electrons. Both nitrogen atoms attain octets by sharing three electron pairs:



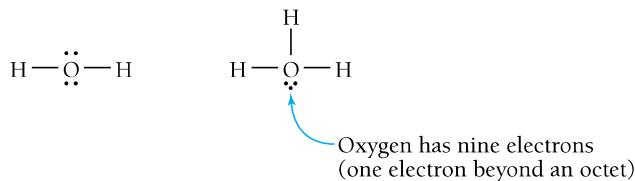
Keep in mind that each dash always stands for two electrons (a single bonding pair).

The bond is a **triple bond**. Triple bonds are even shorter and stronger than double bonds. When we examine nitrogen in nature, we find that it exists as a diatomic molecule with a very strong bond between the two nitrogen atoms. The bond is so strong that it is difficult to break, making N_2 a relatively unreactive molecule.

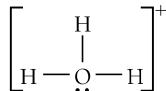
We explore the characteristics of multiple bonds more fully in Section 10.10.

Covalent Bonding: Models and Reality

The Lewis model predicts the properties of molecular compounds in many ways. First, it accounts for why particular combinations of atoms form molecules and others do not. For example, why is water H_2O and not H_3O ? We can write a good Lewis structure for H_2O , but not for H_3O .

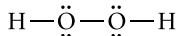


In this way, the Lewis model predicts that H_2O should be stable, while H_3O should not be, and that is in fact the case. However, if we remove an electron from H_3O , we get H_3O^+ , which should be stable (according to the Lewis model) because, when we remove the extra electron, oxygen attains an octet:



This ion, called the hydronium ion, is in fact stable in aqueous solutions (see Section 5.7).

The Lewis model predicts other possible combinations for hydrogen and oxygen as well. For example, we can write a Lewis structure for H_2O_2 as follows:



Indeed, H_2O_2 , or hydrogen peroxide, exists and is often used as a disinfectant and a bleach.

The Lewis model also accounts for why covalent bonds are highly *directional*. The attraction between two covalently bonded atoms is due to the sharing of one or more electron pairs in the space between them. Thus, each bond links just one specific pair of atoms—in contrast to ionic bonds, which are nondirectional and hold together an entire array of ions. As a result, the fundamental units of covalently bonded compounds are individual molecules. These molecules can interact with one another in a number of different ways that we cover in Chapter 12. However, in covalently bonded molecular compounds the interactions *between* molecules (intermolecular forces) are generally much weaker than the bonding interactions within a molecule (intramolecular forces), as shown in Figure 10.5▼.

Molecular Compound

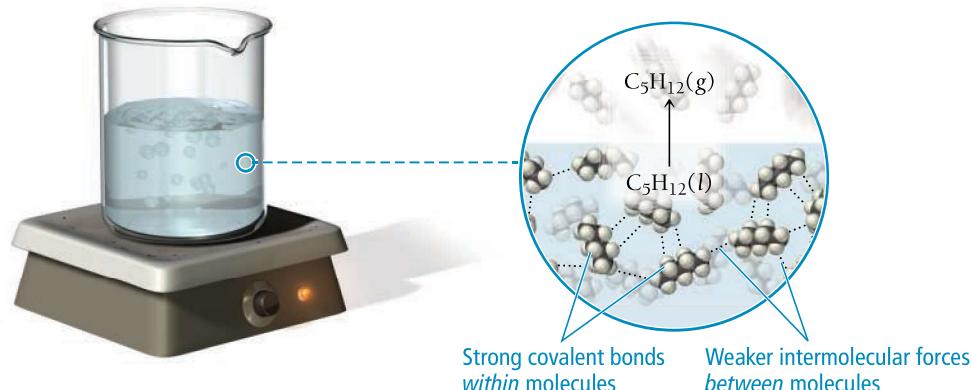


FIGURE 10.5 Intermolecular and Intramolecular Forces The covalent bonds between atoms of a molecule are much stronger than the interactions between molecules. To boil a molecular substance, we simply have to overcome the relatively weak intermolecular forces, so molecular compounds generally have low boiling points.

ANSWER NOW!



10.4 Cc

Conceptual Connection

WATCH NOW!

KEY CONCEPT VIDEO 10.6

Electronegativity and Bond Polarity

When a molecular compound melts or boils, the molecules themselves remain intact—only the relatively weak interactions between molecules must be overcome. Consequently, molecular compounds tend to have lower melting and boiling points than ionic compounds.

ENERGY AND THE OCTET RULE What is wrong with the following statement? *Atoms form bonds in order to satisfy the octet rule.*

- Nothing. The statement is correct.
- Atoms form bonds for complex reasons that include the lowering of potential energy. The octet rule helps us predict in which cases bonds form.
- Atoms form bonds for complex reasons that include the lowering of potential energy. The octet rule is not really useful.

10.6

Electronegativity and Bond Polarity

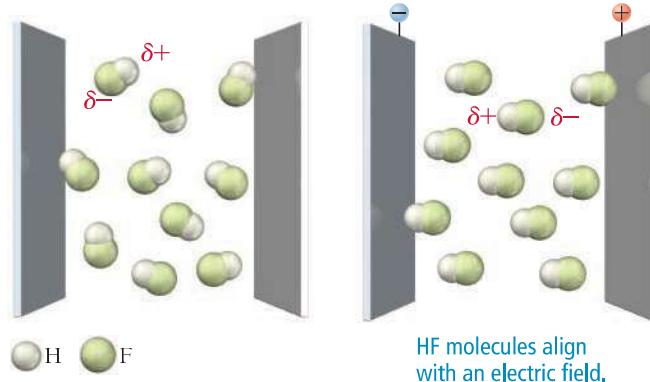
We know from Chapter 8 that representing electrons with dots, as we do in the Lewis model, is a drastic oversimplification. As we have already discussed, this does not invalidate the Lewis model—which is an extremely useful theory—but we must recognize and compensate for its inherent limitations. One limitation of representing electrons as dots, and covalent bonds as two dots shared between two atoms, is that the shared electrons always appear to be *equally* shared. Such is not the case.

For example, consider the Lewis structure of hydrogen fluoride:

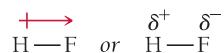


The two shared electron dots sitting between the H and the F atoms appear to be equally shared between hydrogen and fluorine. However, based on laboratory measurements, we know they are not. When HF is put in an electric field, the molecules orient as shown in Figure 10.6▼. From this observation, we know that the hydrogen side of the molecule

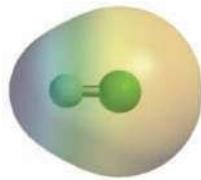
► **FIGURE 10.6** Orientation of Gaseous Hydrogen Fluoride in an Electric Field Because one side of the HF molecule has a slight positive charge and the other side a slight negative charge, the molecules align themselves with an external electric field.



must have a slight positive charge and the fluorine side of the molecule must have a slight negative charge. We represent this partial separation of charge as follows:



The red arrow on the left, with a positive sign on the tail, indicates that the left side of the molecule has a partial positive charge and that the right side of the molecule (the side the arrow is pointing toward) has a partial negative charge. Similarly, the δ^+ (delta plus) represents a partial positive charge and the δ^- (delta minus) represents a partial negative charge. Does this make the bond ionic? No. In an ionic bond, the electron is essentially transferred from one atom to another. In HF, the electron is *unequally shared*. In other words, even though the Lewis structure of HF portrays the bonding electrons as residing *between* the two atoms, in reality the electron density is greater on the fluorine atom than on the hydrogen atom (Figure 10.7◀). The bond is said to be *polar*—having a positive pole



▲ **FIGURE 10.7** Electrostatic Potential Map for the HF Molecule The F end of the molecule, with its partial negative charge, is yellow/pink; the H end, with its partial positive charge, is blue.

and a negative pole. A **polar covalent bond** is intermediate in nature between a pure covalent bond and an ionic bond. In fact, the categories of pure covalent and ionic are really two extremes within a broad continuum. Most covalent bonds between dissimilar atoms are actually *polar covalent*—somewhere between the two extremes.

Electronegativity

The ability of an atom to attract electrons to itself in a chemical bond (which results in polar and ionic bonds) is called **electronegativity**. We say that fluorine is more *electronegative* than hydrogen because it takes a greater share of the electron density in HF.

The American chemist Linus Pauling (1901–1994) quantified electronegativity in his classic book, *The Nature of the Chemical Bond*. Pauling compared the bond energy—the energy required to break a bond—of a heteronuclear diatomic molecule such as HF with the bond energies of its homonuclear counterparts, in this case H₂ and F₂. The bond energies of H₂ and F₂ are 436 kJ/mol and 155 kJ/mol, respectively. Pauling reasoned that if the HF bond were purely covalent—that is, if the electrons were shared exactly equally—the bond energy of HF should simply be an average of the bond energies of H₂ and F₂, which would be 296 kJ/mol. However, the bond energy of HF is experimentally measured to be 565 kJ/mol. Pauling suggested that the additional bond energy was due to the *ionic character* of the bond. Based on many such comparisons of bond energies, and by arbitrarily assigning an electronegativity of 4.0 to fluorine (the most electronegative element on the periodic table), Pauling developed the electronegativity values shown in Figure 10.8▼.

For main-group elements, notice the following periodic trends in electronegativity from Figure 10.8:

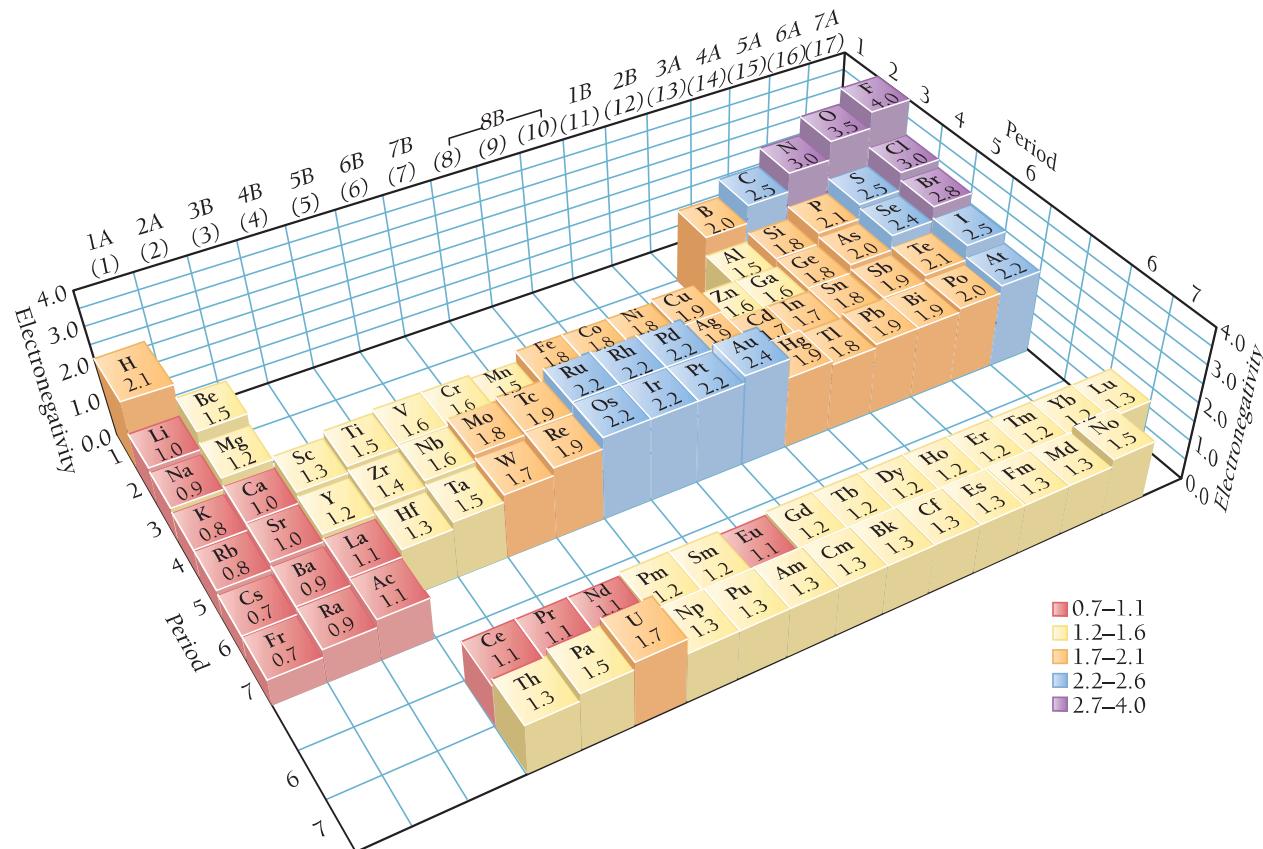
- Electronegativity generally increases across a period in the periodic table.
- Electronegativity generally decreases down a column in the periodic table.
- Fluorine is the most electronegative element.
- Francium is the least electronegative element (sometimes called the most *electropositive*).

We cover the concept of bond energy in more detail in Section 10.10.

Pauling's “average” bond energy was actually calculated a little bit differently than the normal average shown here. He took the square root of the product of the bond energies of the homonuclear counterparts as the “average.”

▼ FIGURE 10.8 Electronegativities of the Elements Electronegativity generally increases as we move across a row in the periodic table and decreases as we move down a column.

Trends in Electronegativity



The periodic trends in electronegativity are consistent with other periodic trends we have seen. In general, electronegativity is inversely related to atomic size—the larger the atom, the less ability it has to attract electrons to itself in a chemical bond.

ANSWER NOW!



10.5 Cc

Conceptual Connection

PERIODIC TRENDS IN ELECTRONEGATIVITY

Arrange these elements in order of decreasing electronegativity: P, Na, N, Al.

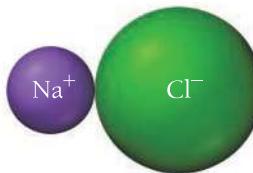
- (a) N > P > Al > Na (b) P > Na > N > Al
 (c) Na > Al > P > N (d) Al > Na > N > P

Bond Polarity, Dipole Moment, and Percent Ionic Character

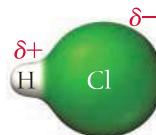
The degree of polarity in a chemical bond depends on the electronegativity difference (sometimes abbreviated ΔEN) between the two bonding atoms. The greater the electronegativity difference, the more polar the bond. If two atoms with identical electronegativities form a covalent bond, they share the electrons equally, and the bond is purely covalent or *nonpolar*. For example, the chlorine molecule, composed of two chlorine atoms (which necessarily have identical electronegativities), has a covalent bond in which electrons are evenly shared:



If there is a large electronegativity difference between the two atoms in a bond, such as normally occurs between a metal and a nonmetal, the electron from the metal is almost completely transferred to the nonmetal, and the bond is ionic. For example, sodium and chlorine form an ionic bond:



If there is an intermediate electronegativity difference between the two atoms, such as between two different nonmetals, then the bond is polar covalent. For example, HCl has a polar covalent bond:



Although all attempts to divide the bond polarity continuum into specific regions are necessarily arbitrary, it is helpful to classify bonds as covalent, polar covalent, and ionic, based on the electronegativity difference between the bonding atoms, as shown in Table 10.1 and Figure 10.9►.

TABLE 10.1 ■ The Effect of Electronegativity Difference on Bond Type

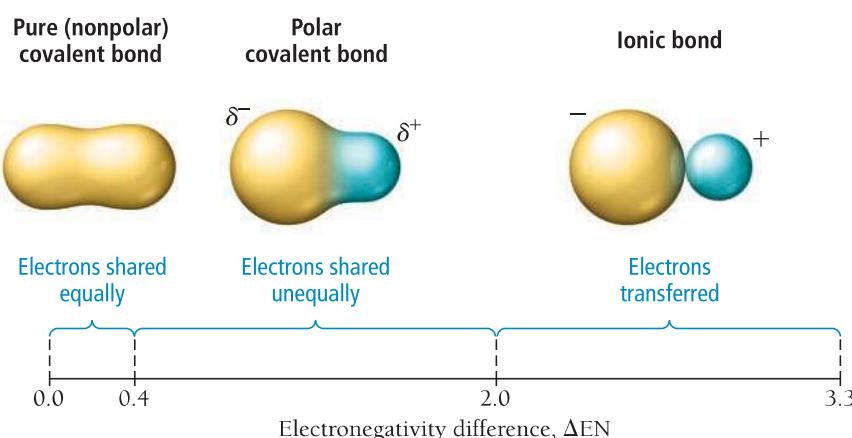
Electronegativity Difference (ΔEN)	Bond Type	Example
Small (0–0.4)	Covalent	Cl_2
Intermediate (0.4–2.0)	Polar covalent	HCl
Large (2.0+)	Ionic	NaCl

We quantify the polarity of a bond by the size of its dipole moment. A **dipole moment** (μ) occurs any time there is a separation of positive and negative charge. The magnitude of a dipole moment created by separating two particles of equal but opposite charges of magnitude q by a distance r is given by the equation:

$$\mu = qr$$

[10.1]

The Continuum of Bond Types



◀ FIGURE 10.9 Electronegativity Difference (ΔEN) and Bond Type

We can get a sense for the dipole moment of a completely ionic bond by calculating the dipole moment that results from separating a proton and an electron ($q = 1.6 \times 10^{-19} \text{ C}$) by a distance of $r = 130 \text{ pm}$ (the approximate length of a short chemical bond):

$$\begin{aligned}\mu &= qr \\ &= (1.6 \times 10^{-19} \text{ C})(130 \times 10^{-12} \text{ m}) \\ &= 2.1 \times 10^{-29} \text{ C} \cdot \text{m} \\ &= 6.2 \text{ D}\end{aligned}$$

The debye (D) is the unit commonly used for reporting dipole moments ($1 \text{ D} = 3.34 \times 10^{-30} \text{ C} \cdot \text{m}$). Based on this calculation, we would expect the dipole moment of completely ionic bonds with bond lengths close to 130 pm to be about 6 D. The smaller the magnitude of the charge separation, and the smaller the distance between the charges, the smaller the dipole moment. Table 10.2 lists the dipole moments of several molecules along with the electronegativity differences of their atoms.

By comparing the *actual* dipole moment of a bond to what the dipole moment would be if the electron were completely transferred from one atom to the other, we can get a sense of the degree to which the electron is transferred (or the degree to which the bond is ionic). The **percent ionic character** is the ratio of a bond's actual dipole moment to the dipole moment it would have if the electron were completely transferred from one atom to the other, multiplied by 100%:

$$\text{Percent ionic character} = \frac{\text{measured dipole moment of bond}}{\text{dipole moment if electron were completely transferred}} \times 100\%$$

For example, suppose a diatomic molecule with a bond length of 130 pm has a dipole moment of 3.5 D. We previously calculated that separating a proton and an electron by 130 pm results in a dipole moment of 6.2 D. Therefore, the percent ionic character of the bond is 56%:

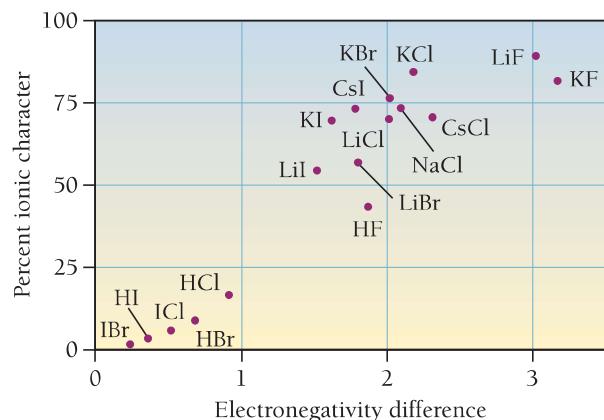
$$\begin{aligned}\text{Percent ionic character} &= \frac{3.5 \text{ D}}{6.2 \text{ D}} \times 100\% \\ &= 56\%\end{aligned}$$

A bond in which an electron is completely transferred from one atom to another would have 100% ionic character (although even the most ionic bonds do not reach this ideal). Figure 10.10▶ shows the percent ionic character of a number of diatomic gas-phase molecules plotted against the electronegativity difference between the bonding atoms. As expected, the percent ionic character generally increases as the electronegativity difference increases. However, as we can see, no bond is 100% ionic. In general, bonds with greater than 50% ionic character are referred to as ionic bonds.

TABLE 10.2 ▀ Dipole Moments of Several Molecules in the Gas Phase

Molecule	ΔEN	Dipole Moment (D)
Cl_2	0	0
ClF	1.0	0.88
HF	1.9	1.82
LiF	3.0	6.33

▼ FIGURE 10.10 Percent Ionic Character versus Electronegativity Difference for Some Compounds



EXAMPLE 10.3 Classifying Bonds as Pure Covalent, Polar Covalent, or Ionic

Classify the bond formed between each pair of atoms as covalent, polar covalent, or ionic.

- (a) Sr and F (b) N and Cl (c) N and O

SOLUTION

- (a) In Figure 10.8, find the electronegativity of Sr (1.0) and of F (4.0). The electronegativity difference (ΔEN) is $\Delta EN = 4.0 - 1.0 = 3.0$. Using Table 10.1, classify this bond as ionic.
- (b) In Figure 10.8, find the electronegativity of N (3.0) and of Cl (3.0). The electronegativity difference (ΔEN) is $\Delta EN = 3.0 - 3.0 = 0$. Using Table 10.1, classify this bond as covalent.
- (c) In Figure 10.8, find the electronegativity of N (3.0) and of O (3.5). The electronegativity difference (ΔEN) is $\Delta EN = 3.5 - 3.0 = 0.5$. Using Table 10.1, classify this bond as polar covalent.

FOR PRACTICE 10.3 Determine if the bond formed between each pair of atoms is pure covalent, polar covalent, or ionic.

- (a) I and I (b) Cs and Br (c) P and O

ANSWER NOW!



10.6 Cc Conceptual Connection

PERCENT IONIC CHARACTER The HCl(*g*) molecule has a bond length of 127 pm and a dipole moment of 1.08 D. Without doing detailed calculations, identify the best estimate for its percent ionic character.

- (a) 5% (b) 15% (c) 50% (d) 80%

WATCH NOW!

KEY CONCEPT VIDEO 10.7

Writing Lewis Structures for Molecular Compounds

10.7 Lewis Structures of Molecular Compounds and Polyatomic Ions

We now turn to the basic sequence of steps involved in actually writing Lewis structures for given combinations of atoms.

Writing Lewis Structures for Molecular Compounds

To write a Lewis structure for a molecular compound, follow these steps:

- 1. Write the correct skeletal structure for the molecule.** The Lewis structure of a molecule must have the atoms in the correct positions. For example, you could not write a Lewis structure for water if you started with the hydrogen atoms next to each other and the oxygen atom at the end (H H O). In nature, oxygen is the central atom and the hydrogen atoms are *terminal* (at the ends). The correct skeletal structure is H O H. The only way to determine the skeletal structure of a molecule with absolute certainty is to examine its structure experimentally. However, you can write likely skeletal structures by remembering two guidelines. First, *hydrogen atoms are always terminal*. Hydrogen does not ordinarily occur as a central atom because central atoms must form at least two bonds, and hydrogen, which has only a single valence electron to share and requires only a duet, can form just one bond. Second, *put the more electronegative elements in terminal positions* and the less electronegative elements (other than hydrogen) in the central position. Later in this section, you will learn how to distinguish between competing skeletal structures by applying the concept of formal charge.

- 2. Calculate the total number of electrons for the Lewis structure by summing the valence electrons of each atom in the molecule.** Remember that the number of valence electrons for any main-group element is equal to its group number in the periodic table. *If you are writing a Lewis structure for a polyatomic ion, you must consider the charge of the ion when calculating the total number of electrons.* Add one electron for each negative charge and subtract one electron for each positive charge. Don't worry about which electron comes from which atom—only the total number is important.

Often, chemical formulas provide clues to how the atoms are bonded together. For example, CH₃OH indicates that three hydrogen atoms and the oxygen atom are bonded to the carbon atom, but the fourth hydrogen atom is bonded to the oxygen atom.

There are a few exceptions to this rule, such as diborane (B₂H₆), which contains *bridging hydrogens*, but these are rare and cannot be adequately addressed by the Lewis model.

3. Distribute the electrons among the atoms, giving octets (or duets in the case of hydrogen) to as many atoms as possible. Begin by placing two electrons between every two atoms. These represent the minimum number of bonding electrons. Then distribute the remaining electrons as lone pairs, first to terminal atoms and then to the central atom, giving octets (or duets for hydrogen) to as many atoms as possible.

4. If any atoms lack an octet, form double or triple bonds as necessary to give them octets. Do this by moving lone electron pairs from terminal atoms into the bonding region with the central atom.

In the left column that follows, there is an abbreviated version of the procedure for writing Lewis structures; the center and right columns feature two examples of applying the procedure.

Sometimes distributing all the remaining electrons to the central atom results in more than an octet. This is called an expanded octet, which we discuss in Section 10.9.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE VIDEO 10.4

HOW TO: Write Lewis Structures for Molecular Compounds	EXAMPLE 10.4 Writing Lewis Structures	EXAMPLE 10.5 Writing Lewis Structures
<p>1. Write the correct skeletal structure for the molecule.</p>	<p>SOLUTION Because carbon is the less electronegative atom, put it in the central position.</p> <p style="text-align: center;">O C O</p>	<p>SOLUTION Since hydrogen is always terminal, put nitrogen in the central position.</p> <p style="text-align: center;">H N H H</p>
<p>2. Calculate the total number of electrons for the Lewis structure by summing the valence electrons of each atom in the molecule.</p>	<p>Total number of electrons for Lewis structure = $\left(\begin{array}{l} \text{number of} \\ \text{valence} \\ e^- \text{ for C} \end{array} \right) + 2 \left(\begin{array}{l} \text{number of} \\ \text{valence} \\ e^- \text{ for O} \end{array} \right)$ $= 4 + 2(6) = 16$</p>	<p>Total number of electrons for Lewis structure = $\left(\begin{array}{l} \text{number of} \\ \text{valence} \\ e^- \text{ for N} \end{array} \right) + 3 \left(\begin{array}{l} \text{number of} \\ \text{valence} \\ e^- \text{ for H} \end{array} \right)$ $= 5 + 3(1) = 8$</p>
<p>3. Distribute the electrons among the atoms, giving octets (or duets for hydrogen) to as many atoms as possible. Begin with the bonding electrons, and then proceed to lone pairs on terminal atoms, and finally to lone pairs on the central atom.</p>	<p>Bonding electrons are first. $O:C:O$ (4 of 16 electrons used)</p> <p>Lone pairs on terminal atoms are next. $\ddot{\bullet}:C:\ddot{\bullet}$ (16 of 16 electrons used)</p>	<p>Bonding electrons are first. $H:\ddot{\bullet}N\ddot{\bullet}:H$ (6 of 8 electrons used)</p> <p>Lone pairs on terminal atoms are next, but none is needed on hydrogen. Lone pairs on the central atom are last. $H-\overset{\ddot{\bullet}}{N}-H$ (8 of 8 electrons used)</p>
<p>4. If any atom lacks an octet, form double or triple bonds as necessary until all atoms have octets.</p>	<p>Since carbon lacks an octet, move lone pairs from the oxygen atoms to bonding regions to form double bonds.</p> <p style="text-align: center;"></p>	<p>Since all of the atoms have octets (or duets for hydrogen), the Lewis structure for NH_3 is complete as shown in the previous step.</p>
	<p>FOR PRACTICE 10.4 Write the Lewis structure for CO.</p>	<p>FOR PRACTICE 10.5 Write the Lewis structure for H_2CO.</p>

Writing Lewis Structures for Polyatomic Ions

Write Lewis structures for polyatomic ions by following the same procedure, but pay special attention to the charge of the ion when calculating the number of electrons for the Lewis structure. Add one electron for each negative charge and subtract one electron for each positive charge. Write the Lewis structure for a polyatomic ion within brackets with the charge of the ion in the upper right-hand corner, outside the bracket.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 10.6

EXAMPLE 10.6

Writing Lewis Structures for Polyatomic Ions

Write the Lewis structure for the NH_4^+ ion.

SOLUTION

Begin by writing the skeletal structure. Since hydrogen is always terminal, put the nitrogen atom in the central position.	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array}$
Calculate the total number of electrons for the Lewis structure by summing the number of valence electrons for each atom and subtracting 1 for the 1^+ charge.	<p>Total number of electrons for Lewis structure $= (\text{number of valence } e^- \text{ in N}) + 4 (\text{number of valence } e^- \text{ in H}) - 1$ $= 5 + 4(1) - 1$ $= 8$</p> <p style="margin-left: 200px;">Subtract $1 e^-$ to account for 1^+ charge of ion.</p>
Place two bonding electrons between every two atoms. Since all of the atoms have complete octets, no double bonds are necessary.	$\begin{array}{c} \text{H} \\ \\ \text{H}:\ddot{\text{N}}:\text{H} \\ \\ \text{H} \end{array}$ <p>(8 of 8 electrons used)</p>
Lastly, write the Lewis structure in brackets with the charge of the ion in the upper right-hand corner.	$\left[\begin{array}{c} \text{H} \\ \\ \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array} \right]^+$

FOR PRACTICE 10.6 Write the Lewis structure for the hypochlorite ion, ClO^- .

WATCH NOW!

KEY CONCEPT VIDEO 10.8

 Resonance and Formal Charge

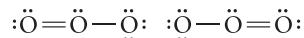
10.8

Resonance and Formal Charge

We need two additional concepts to write the best possible Lewis structures for a large number of compounds. The concepts are *resonance*, which is relevant when two or more valid Lewis structures can be drawn for the same compound, and *formal charge*, an electron bookkeeping system that allows us to discriminate between alternative Lewis structures.

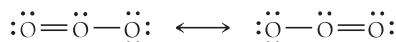
Resonance

For some molecules, we can write more than one valid Lewis structure. For example, consider writing a Lewis structure for O_3 . The following two Lewis structures, with the double bond on alternate sides, are equally correct:

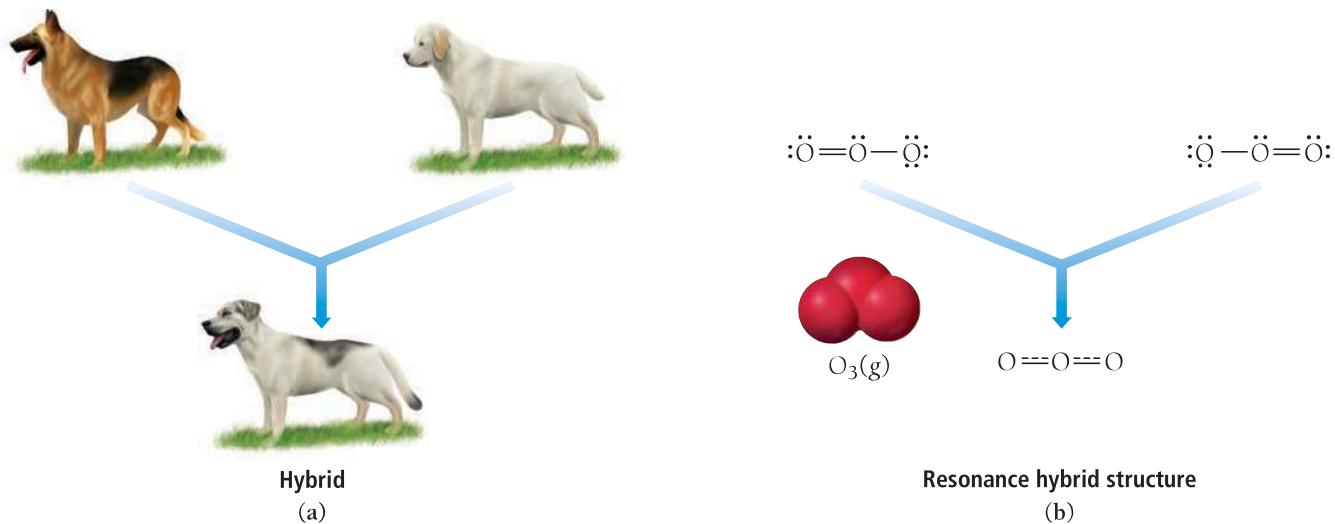


In cases such as this—where there are two or more valid Lewis structures for the same molecule—we find that, in nature, the molecule is an *average* of the two Lewis structures. Both of the two Lewis structures for O_3 predict that O_3 contains two different bonds

(one double bond and one single bond). However, when we experimentally examine the structure of O_3 , we find that the bonds in the O_3 molecule are equivalent and each is intermediate in strength and length between a double bond and single bond. We account for this by representing the molecule with both structures, called *resonance structures*, with a double-headed arrow between them:



A **resonance structure** is one of two or more Lewis structures that have the same skeletal formula (the atoms are in the same locations), but different electron arrangements. The actual structure of the molecule is intermediate between the two (or more) resonance structures and is called a **resonance hybrid**. The term *hybrid* comes from breeding and means the offspring of two animals or plants of different varieties or breeds. If we breed a Labrador retriever with a German shepherd, we get a *hybrid* that is intermediate between the two breeds (Figure 10.11(a)▼). Similarly, the actual structure of a resonance hybrid is intermediate between the two (or more) resonance structures (Figure 10.11(b)▼). The only structure that actually exists is the hybrid structure—the individual resonance structures do not exist and are merely a way to describe the actual structure. Notice that the actual structure of ozone has two equivalent bonds and a bent geometry (we discuss molecular geometries in Chapter 11).



▲ FIGURE 10.11 Hybridization Just as the offspring of two different dog breeds is a hybrid that is intermediate between the two breeds (a), the structure of a resonance hybrid is intermediate between that of the contributing resonance structures (b).

The concept of resonance is an adaptation of the Lewis model that helps account for the complexity of actual molecules. In the Lewis model, electrons are *localized* either on one atom (lone pair) or between atoms (bonding pair). However, in nature, the electrons in molecules are often *delocalized* over several atoms or bonds. The delocalization of electrons lowers their energy; it stabilizes them (for reasons that are beyond the scope of this book). Resonance depicts two or more structures with the electrons in different places in an attempt to more accurately reflect the delocalization of electrons. In the real hybrid structure, an average between the resonance structures, the electrons are more spread out (or delocalized) than in any of the resonance structures. The resulting stabilization of the electrons (that is, the lowering of their potential energy due to delocalization) is sometimes called *resonance stabilization*. Resonance stabilization makes an important contribution to the stability of many molecules.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 10.7

EXAMPLE 10.7 Writing Resonance StructuresWrite a Lewis structure for the NO_3^- ion. Include resonance structures.**SOLUTION**

Begin by writing the skeletal structure. Since nitrogen is the least electronegative atom, put it in the central position.

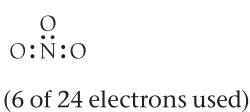


Calculate the total number of electrons for the Lewis structure by summing the number of valence electrons for each atom and adding 1 for the 1⁻ charge.

$$\begin{aligned} &\text{Total number of electrons for Lewis structure} \\ &= (\text{number of valence } e^- \text{ in N}) + 3 (\text{number of valence } e^- \text{ in O}) + 1 \\ &= 5 + 3(6) + 1 \\ &= 24 \end{aligned}$$

Add 1 e^- to account for 1⁻ charge of ion.

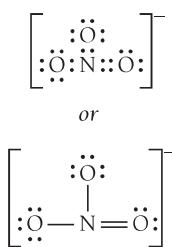
Place two bonding electrons between each pair of atoms.



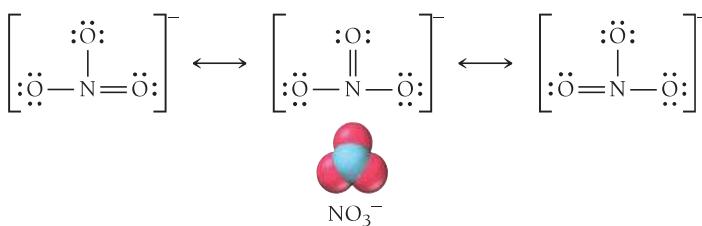
Distribute the remaining electrons, first to terminal atoms. There are not enough electrons to complete the octet on the central atom.



Form a double bond by moving a lone pair from one of the oxygen atoms into the bonding region with nitrogen. Enclose the structure in brackets and include the charge.



Since the double bond can form equally well with any of the three oxygen atoms, write all three structures as resonance structures. (The actual space-filling model of NO_3^- is shown here for comparison. Note that all three bonds are equal in length.)



FOR PRACTICE 10.7 Write a Lewis structure for the NO_2^- ion. Include resonance structures.

In the examples of resonance hybrids that we have examined so far, the contributing structures have been equivalent (or equally valid) Lewis structures. In these cases, the true structure is an equally weighted average of the resonance structures. In some cases, however, we can write resonance structures that are not equivalent. For reasons that we cover in the material that follows—such as formal charge—one resonance structure may be somewhat better than another. In such cases, the true structure is still an average of the resonance structures, but the better resonance structure contributes more to the true structure. In other words, multiple nonequivalent resonance structures may be weighted differently in their contributions to the true overall structure of a molecule (see Example 10.8).

Formal Charge

Formal charge is a fictitious charge assigned to each atom in a Lewis structure that helps us to distinguish among competing Lewis structures. The formal charge of an atom in a Lewis structure is *the charge it would have if all bonding electrons were shared equally*

RESONANCE STRUCTURES

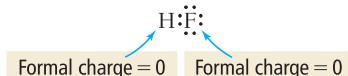
Which structure is NOT a resonance structure of the ion shown here?

- (a) $[\ddot{\text{S}}=\text{C}=\text{N}]^-$ (b) $[\ddot{\text{S}}-\text{N}\equiv\text{C}]^-$ (c) $[\text{S}\equiv\text{C}-\ddot{\text{N}}]^-$



ANSWER NOW!

the bonded atoms. In other words, formal charge is the calculated charge for an atom if we completely ignore the effects of electronegativity. For example, we know that because fluorine is more electronegative than hydrogen, HF has a dipole moment—the hydrogen atom has a slight positive charge and the fluorine atom has a slight negative charge. However, the *formal charges* of hydrogen and fluorine in HF (the calculated charges if we ignore their differences in electronegativity) are both zero:



We can calculate the formal charge on any atom as the difference between the number of valence electrons in the atom and the number of electrons that it “owns” in a Lewis structure. An atom in a Lewis structure can be thought of as “owning” all of its nonbonding electrons and one-half of its bonding electrons:

$$\text{Formal charge} = \text{number of valence electrons} - (\text{number of nonbonding electrons} + \frac{1}{2} \text{ number of bonding electrons})$$

So the formal charge of hydrogen in HF is 0:

$$\text{Formal charge} = 1 - [0 + \frac{1}{2}(2)] = 0$$

Number of valence electrons for H Number of electrons that H “owns” in the Lewis structure

Similarly, we calculate the formal charge of fluorine in HF as 0:

$$\text{Formal charge} = 7 - [6 + \frac{1}{2}(2)] = 0$$

Number of valence electrons for F Number of electrons that F “owns” in the Lewis structure

The concept of formal charge is useful because it can help us distinguish between competing skeletal structures or competing resonance structures. In general, when we calculate formal charges, these four rules apply:

1. The sum of all formal charges in a neutral molecule must be zero.
2. The sum of all formal charges in an ion must equal the charge of the ion.
3. Small (or zero) formal charges on individual atoms are better than large ones.
4. When formal charge cannot be avoided, negative formal charge should reside on the most electronegative atom.

We can use formal charge to determine which of the competing skeletal structures for hydrogen cyanide shown here is best (or most stable). Notice that both skeletal structures satisfy the octet rule equally. The formal charge of each atom in the structure is calculated below it.

	Structure A			Structure B				
	H	—	C ≡	N:	H	—	N ≡	C:
number of valence e ⁻	1	4		5	1	5	4	
—number of nonbonding e ⁻	-0	-0		-2	-0	-0	-2	
$-\frac{1}{2}(\text{number of bonding e}^-)$	$-\frac{1}{2}(2)$	$-\frac{1}{2}(8)$		$-\frac{1}{2}(6)$	$-\frac{1}{2}(2)$	$-\frac{1}{2}(8)$	$-\frac{1}{2}(6)$	
Formal charge	0	0		0	0	+1	-1	

Both HCN and HNC exist, but—as predicted by formal charge—HCN is more stable than HNC.

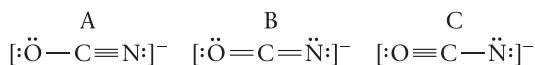
As required, the sum of the formal charges for each of these structures is zero (as it always must be for neutral molecules). However, structure B has formal charges on both the N atom and the C atom, while structure A has no formal charges on any atom. Furthermore, in structure B, the negative formal charge is not on the most electronegative element (nitrogen is more electronegative than carbon). Consequently, structure A is the better skeletal structure. Since atoms in the middle of a molecule tend to have more bonding electrons and fewer nonbonding electrons, they also tend to have more positive formal charges. Consequently, the best skeletal structure usually has the least electronegative atom in the central position, as we learned in step 1 of our procedure for writing Lewis structures.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 10.8

EXAMPLE 10.8 Assigning Formal Charges

Assign formal charges to each atom in the resonance structures of the cyanate ion (OCN^-). Which resonance structure is likely to contribute most to the correct structure of (OCN^-)?



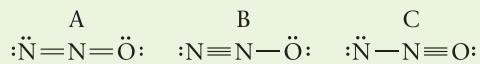
SOLUTION

Calculate the formal charge on each atom by finding the number of valence electrons and subtracting the number of nonbonding electrons and one-half the number of bonding electrons.

	A $[\ddot{\text{O}}-\text{C}\equiv\text{N}:]^-$			B $[\ddot{\text{O}}=\text{C}=\ddot{\text{N}}:]^-$			C $[\text{O}\equiv\text{C}-\ddot{\text{N}}:]^-$		
Number of valence e^-	6	4	5	6	4	5	6	4	5
-number of nonbonding e^-	-6	-0	-2	-4	-0	-4	-2	-0	-6
$-\frac{1}{2}$ (number of bonding e^-)	-1	-4	-3	-2	-4	-2	-3	-4	-1
Formal charge	-1	0	0	0	0	-1	+1	0	-2

The sum of all formal charges for each structure is -1 , as it should be for a $1-$ ion. Structures A and B have the fewest formal charges and are preferred over structure C. Structure A is preferable to B because it has the negative formal charge on the more electronegative atom. You therefore expect structure A to make the biggest contribution to the resonance forms of the cyanate ion.

FOR PRACTICE 10.8 Assign formal charges to each atom in the resonance structures of N_2O . Which resonance structure is likely to contribute most to the correct structure of N_2O ?



FOR MORE PRACTICE 10.8 Assign formal charges to each of the atoms in the nitrate ion (NO_3^-). The Lewis structure for the nitrate ion is shown in Example 10.7.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 10.9

EXAMPLE 10.9

Drawing Resonance Structures and Assigning Formal Charge for Organic Compounds

Draw the Lewis structure (including resonance structures) for nitromethane (CH_3NO_2). For each resonance structure, assign formal charges to all atoms that have formal charge.

SOLUTION

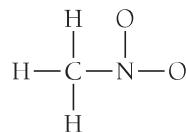
Begin by writing the skeletal structure. For organic compounds, the condensed structural formula (in this case CH_3NO_2) indicates how the atoms are connected.



Calculate the total number of electrons for the Lewis structure by summing the number of valence electrons for each atom.

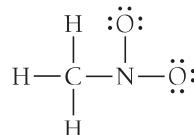
$$\begin{aligned} \text{Total number of electrons for Lewis structure} \\ &= (\# \text{ valence } e^- \text{ in C}) + 3(\# \text{ valence } e^- \text{ in H}) \\ &\quad + (\# \text{ valence } e^- \text{ in N}) + 2(\# \text{ valence } e^- \text{ in O}) \\ &= 4 + 3(1) + 5 + 2(6) \\ &= 24 \end{aligned}$$

Place a dash between each pair of atoms to indicate a bond. Each dash represents two electrons.



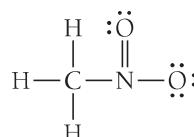
(12 of 24 electrons used)

Distribute the remaining electrons, first to terminal atoms, then to interior atoms.

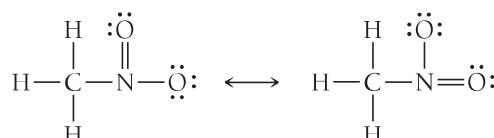


(24 of 24 electrons used)

If there are not enough electrons to complete the octets on the interior atoms, form double bonds by moving lone pair electrons from terminal atoms into the bonding region with interior atoms.

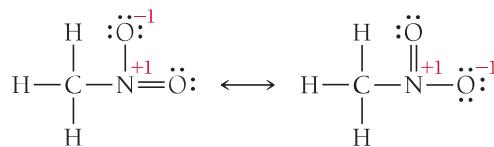


Draw any necessary resonance structures by moving only electron dots. (In this case, you can form a double bond between the nitrogen atom and the other oxygen atom.)



Assign formal charges (FC) to each atom.

$$\begin{aligned} \text{FC} &= \# \text{ valence } e^- \\ &\quad - (\# \text{ nonbonding } e^- + \frac{1}{2} \# \text{ bonding } e^-) \end{aligned}$$



Carbon, hydrogen, and the double-bonded oxygen atoms have formal charges of zero. Nitrogen has a +1 formal charge $[5 - \frac{1}{2}(8)]$ and the singly bonded oxygen atom in each resonance structure has a -1 formal charge $[6 - (6 + \frac{1}{2}(2))]$.

FOR PRACTICE 10.9 Draw the Lewis structure (including resonance structures) for diazomethane (CH_2N_2). For each resonance structure, assign formal charges to all atoms that have formal charge.

10.9

Exceptions to the Octet Rule: Odd-Electron Species, Incomplete Octets, and Expanded Octets

The octet rule in the Lewis model has some exceptions, which we examine in this section of the chapter. They include (1) *odd-electron species*, molecules or ions with an odd number of electrons; (2) *incomplete octets*, molecules or ions with *fewer than eight electrons* around an atom; and (3) *expanded octets*, molecules or ions with *more than eight electrons* around an atom.

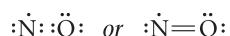
WATCH NOW!

KEY CONCEPT VIDEO 10.9

Exceptions to the Octet Rule and Expanded Octets

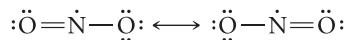
Odd-Electron Species

Molecules and ions with an odd number of electrons in their Lewis structures are **free radicals** (or simply *radicals*). For example, nitrogen monoxide—a pollutant found in motor vehicle exhaust—has 11 electrons. If we try to write a Lewis structure for nitrogen monoxide, we can't achieve octets for both atoms:



The unpaired electron in nitrogen monoxide is put on the nitrogen rather than the oxygen in order to minimize formal charges.

The nitrogen atom in nitrogen monoxide does not have an octet, so this Lewis structure does not satisfy the octet rule. Yet, nitrogen monoxide exists, especially in polluted air. Why? As with any simple theory, the Lewis model is not sophisticated enough to handle every single case. It is impossible to write good Lewis structures for free radicals; nevertheless, some of these molecules exist in nature. Perhaps it is a testament to the Lewis model, however, that *relatively few* such molecules exist and that, in general, they tend to be somewhat unstable and reactive. NO, for example, reacts with oxygen in the air to form NO₂, another odd-electron molecule represented with the following 17-electron resonance structures:



In turn, NO₂ reacts with water to form nitric acid (a component of acid rain) and also reacts with other atmospheric pollutants to form peroxyacetyl nitrate (PAN), an active component of photochemical smog. For free radicals, such as NO and NO₂, we simply write the best Lewis structure that we can.

ANSWER NOW!



10.8 Cc Conceptual Connection

ODD-ELECTRON SPECIES Which molecule would you expect to be a free radical?

- (a) CO (b) CO₂ (c) N₂O (d) ClO

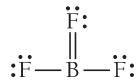
Incomplete Octets

Another significant exception to the octet rule involves elements that tend to form *incomplete octets*. The most important of these is boron, which forms compounds with only six electrons around B, rather than eight. For example, BF₃ and BH₃ lack an octet for B.



Beryllium compounds, such as BeH₂, also have incomplete octets.

You might be wondering why we don't just form double bonds to increase the number of electrons around B. For BH₃, we can't because there are no additional electrons to move into the bonding region. For BF₃, however, we could attempt to give B an octet by moving a lone pair from an F atom into the bonding region with B:

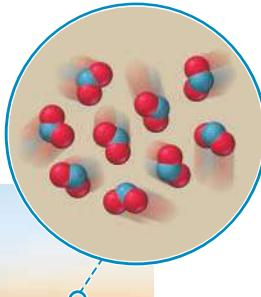




CHEMISTRY IN THE ENVIRONMENT

Free Radicals and the Atmospheric Vacuum Cleaner

Free radicals play a key role in much of the chemistry of the atmosphere. The free radical that is most important to atmospheric reactions is the hydroxyl radical:

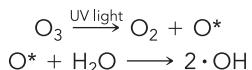


▲ $\text{NO}_2(g)$ is a pollutant found in urban air.

Many free radical structures are abbreviated by writing a single dot with the formula. Thus, the hydroxyl radical is often abbreviated as:

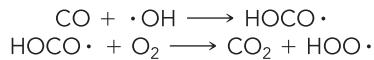


In the atmosphere, the hydroxyl radical forms when excited oxygen atoms—formed from the photodecomposition of ozone—react with water vapor:



The * next to the O indicates that the oxygen atom has excess energy.

The resulting hydroxyl radical reacts with a wide variety of molecules from both natural sources and from air pollution that are present in the atmosphere. For example, the hydroxyl radical reacts with carbon monoxide, an atmospheric pollutant that we first encountered in Chapter 1, in the following two-step process:

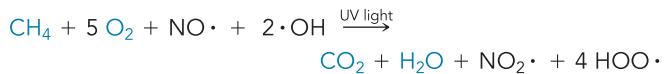


We can see from this reaction that the hydroxyl radical converts toxic CO into relatively nontoxic CO_2 .

The $\text{HO}_2\cdot$ free radical generated by the second reaction is converted back into the hydroxyl radical when it reacts with other atmospheric substances, and the process repeats itself. Therefore, a single hydroxyl radical can convert a lot of CO into CO_2 .

Do you ever wonder what happens to the hydrocarbons you accidentally spill when filling a car's gas tank or to the natural gas that is released into the atmosphere as you light a kitchen stove?

Hydrocarbons released into the atmosphere are converted to CO_2 and H_2O in a series of steps initiated by the hydroxyl free radical. Consider the following representative reaction of methane, the main hydrocarbon in natural gas:



Notice the similarity between this reaction and the direct combustion (or burning) of methane:



As you can see, the free radical reaction initiates a slow “burning” of CH_4 in a series of steps that produces carbon dioxide and water and some additional free radicals. The hydroxyl radical initiates similar reactions with other pollutants as well as undesirable naturally occurring atmospheric gases. Without the hydroxyl free radical—sometimes called the *atmospheric vacuum cleaner*—our atmosphere would be a much dirtier place.

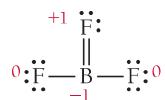


▲ Hydrocarbons such as octane evaporate into the atmosphere when a motor vehicle is fueled. What happens to them?

QUESTION Draw the best possible Lewis structures for the free radicals important in atmospheric chemistry: NO , NO_2 , $\text{HOO}\cdot$, OH , CH_3 .

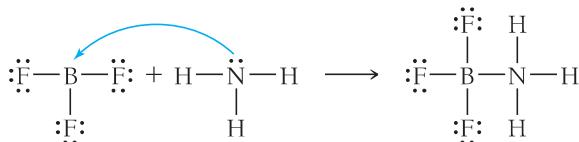


This Lewis structure has octets for all atoms, including boron. However, when we assign formal charges to this structure, we get a negative formal charge on B and a positive formal charge on F:



The positive formal charge on fluorine—the most electronegative element in the periodic table—makes this an unfavorable structure. This leaves us with some questions. Do we complete the octet on B at the expense of giving fluorine a positive formal charge? Or do we leave B without an octet in order to avoid the positive formal charge on fluorine? The answers to these kinds of questions are not always clear because we are pushing the limits of the Lewis model. In the case of boron, we usually accept the incomplete octet as the better Lewis structure. However, doing so does not rule out the possibility that the double-bonded Lewis structure might be a minor contributing resonance structure. The ultimate answers to these kinds of questions must be determined from experiments. Experimental measurements of the B—F bond length in BF_3 suggest that the bond may be slightly shorter than expected for a single B—F bond, indicating that it may indeed have a small amount of double-bond character.

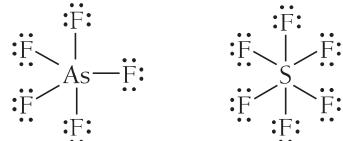
BF_3 can complete its octet in another way—via a chemical reaction. The Lewis model predicts that BF_3 might react in ways that would complete its octet, and indeed it does. For example, BF_3 reacts with NH_3 as follows:



The product has complete octets for all atoms in the structure.

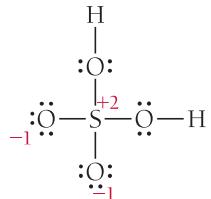
Expanded Octets

Elements in the third row of the periodic table and beyond often exhibit *expanded octets* of up to 12 (and occasionally 14) electrons. Consider the Lewis structures of arsenic pentafluoride and sulfur hexafluoride:



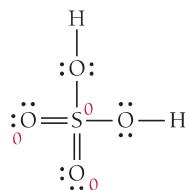
In AsF_5 arsenic has an expanded octet of 10 electrons, and in SF_6 sulfur has an expanded octet of 12 electrons. Both of these compounds exist and are stable. In third-period elements and beyond, 10- and 12-electron expanded octets are common because the *d* orbitals in these elements are energetically accessible (they are not much higher in energy than the orbitals occupied by the valence electrons) and can accommodate the extra electrons (see Section 9.3). Expanded octets *never* occur in second-period elements.

In some Lewis structures, we must decide whether or not to expand an octet in order to lower formal charge. For example, consider the Lewis structure of H_2SO_4 :



Notice that two of the oxygen atoms have a -1 formal charge and that sulfur has a $+2$ formal charge. While this amount of formal charge is acceptable, especially since the

negative formal charge resides on the more electronegative atom, it is possible to eliminate the formal charge by expanding the octet on sulfur:



Which of these two Lewis structures for H_2SO_4 is better? Again, the answer is not straightforward. Experiments show that the sulfur–oxygen bond lengths in the two sulfur–oxygen bonds without the hydrogen atoms are shorter than expected for sulfur–oxygen single bonds, indicating that the double-bonded Lewis structure plays an important role in describing the bonding in H_2SO_4 . In general, we expand octets in third-row (or beyond) elements in order to lower formal charge. However, we should *never* expand the octets of second-row elements. Second-row elements do not have energetically accessible *d* orbitals and therefore never exhibit expanded octets.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 10.10

EXAMPLE 10.10 Writing Lewis Structures for Compounds Having Expanded Octets


Write the Lewis structure for XeF_2 .

SOLUTION

Begin by writing the skeletal structure. Since xenon is the less electronegative atom, put it in the central position.	F Xe F
Calculate the total number of electrons for the Lewis structure by summing the number of valence electrons for each atom.	Total number of electrons for Lewis structure = (number of valence e^- in Xe) + 2(number of valence e^- in F) = 8 + 2(7) = 22
Place two bonding electrons between the atoms of each pair of atoms.	F:Xe:F (4 of 22 electrons used)
Distribute the remaining electrons to give octets to as many atoms as possible, beginning with terminal atoms and finishing with the central atom. Arrange additional electrons around the central atom, giving it an expanded octet of up to 12 electrons.	$\text{:}\ddot{\text{F}}\text{: Xe :}\ddot{\text{F}}\text{:}$ (16 of 22 electrons used) $\text{:}\ddot{\text{F}}\text{:}\ddot{\text{Xe}}\text{:}\ddot{\text{F}}\text{:}$ or $\text{:}\ddot{\text{F}}\text{---}\ddot{\text{Xe}}\text{---}\ddot{\text{F}}\text{:}$ (22 of 22 electrons used)

FOR PRACTICE 10.10 Write the Lewis structure for XeF_4 .

FOR MORE PRACTICE 10.10 Write the Lewis structure for H_3PO_4 . If necessary, expand the octet on any appropriate atoms to lower formal charge.

EXPANDED OCTETS Which molecule could have an expanded octet?

- (a) H_2CO_3
- (b) H_3PO_4
- (c) HNO_2



ANSWER NOW!



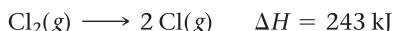
10.10 Bond Energies and Bond Lengths

In Chapter 7, we discussed how to calculate the standard enthalpy change for a chemical reaction ($\Delta H^\circ_{\text{rxn}}$) from tabulated standard enthalpies of formation. However, sometimes we may not easily find standard enthalpies of formation for all of the reactants and products of a reaction. In such cases, we can use individual *bond energies* to estimate enthalpy changes of reaction. In this section, we examine the concept of bond energy and how we can use bond energies to calculate enthalpy changes of reaction. We also look at average bond lengths for a number of commonly encountered bonds.

Bond Energy

Bond energy is also called bond enthalpy or bond dissociation energy.

The **bond energy** of a chemical bond is the energy required to break 1 mole of the bonds in the gas phase. For example, the bond energy of the Cl—Cl bond in Cl_2 is 243 kJ/mol.



The bond energy of HCl is 431 kJ/mol.

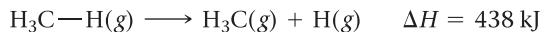


Bond energies are always positive because it always takes energy to break a bond. We say that the HCl bond is *stronger* than the Cl_2 bond because it requires more energy to break it. In general, compounds with stronger bonds tend to be more chemically stable, and therefore less chemically reactive, than compounds with weaker bonds. The triple bond in N_2 has a bond energy of 946 kJ/mol.



It is a very strong and stable bond, which explains nitrogen's relative inertness.

The bond energy of a particular bond in a polyatomic molecule is a little more difficult to determine because a particular type of bond can have different bond energies in different molecules. For example, consider the C—H bond. In CH_4 , the energy required to break one C—H bond is 438 kJ/mol.



However, the energy required to break a C—H bond in other molecules varies slightly, as shown here:

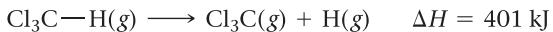
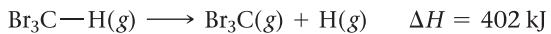
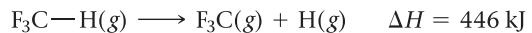


TABLE 10.3 ■ Average Bond Energies

Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)
H—H	436	N—N	163	Br—F	237
H—C	414	N=N	418	Br—Cl	218
H—N	389	N≡N	946	Br—Br	193
H—O	464	N—O	222	I—Cl	208
H—S	368	N=O	590	I—Br	175
H—F	565	N—F	272	I—I	151
H—Cl	431	N—Cl	200	Si—H	323
H—Br	364	N—Br	243	Si—Si	226
H—I	297	N—I	159	Si—C	301
C—C	347	O—O	142	S—O	265
C=C	611	O=O	498	Si=O	368
C≡C	837	O—F	190	S=O	523
C—N	305	O—Cl	203	Si—Cl	464
C=N	615	O—I	234	S=S	418
C≡N	891	F—F	159	S—F	327
C—O	360	Cl—F	253	S—Cl	253
C=O	736*	Cl—Cl	243	S—Br	218
C≡O	1072			S—S	266
C—Cl	339				

*799 in CO_2

We can calculate an *average bond energy* for a chemical bond, which is an average of the bond energies for that bond in a large number of compounds. For the four compounds just discussed, we calculate an average C—H bond energy of 422 kJ/mol. Table 10.3 lists average bond energies for a number of common chemical bonds averaged over a large number of compounds. Notice that the C—H bond energy listed is 414 kJ/mol, which is not too different from the value we calculated from our limited number of compounds. Notice also that bond energies depend, not only on the kind of atoms involved in the bond, but also on the type of bond: single, double, or triple. In general, for a given pair of atoms, triple bonds are stronger than double bonds, which are, in turn, stronger than single bonds. For example, consider the bond energies of carbon–carbon triple, double, and single bonds listed in Table 10.3.

Using Average Bond Energies to Estimate Enthalpy Changes for Reactions

We can use average bond energies to *estimate* the enthalpy change of a reaction. For example, consider the reaction between methane and chlorine:



We can imagine this reaction occurring by the breaking of a C—H bond and a Cl—Cl bond and the forming of a C—Cl bond and an H—Cl bond. We know that when bonds break, the process is endothermic (positive bond energy) and when bonds form, the process is exothermic (negative bond energy). So we can calculate the overall enthalpy change as a sum of the enthalpy changes associated with breaking the required bonds in the reactants and forming the required bonds in the products, as shown in Figure 10.12▼.



Bonds Broken		Bonds Formed	
C—H break	+414 kJ	C—Cl form	-339 kJ
Cl—Cl break	+243 kJ	H—Cl form	-431 kJ
<i>Sum (Σ)ΔH's bonds broken: +657 kJ</i>		<i>Sum (Σ)ΔH's bonds formed: -770 kJ</i>	

$$\begin{aligned}\Delta H_{\text{rxn}} &= \Sigma(\Delta H's \text{ bonds broken}) + \Sigma(\Delta H's \text{ bonds formed}) \\ &= +657 \text{ kJ} - 770 \text{ kJ} \\ &= -113 \text{ kJ}\end{aligned}$$

We find that $\Delta H_{\text{rxn}} = -113 \text{ kJ}$. Calculating $\Delta H_{\text{rxn}}^{\circ}$ from tabulated enthalpies of formation—the method we learned in Chapter 7—gives $\Delta H_{\text{rxn}}^{\circ} = -101 \text{ kJ}$, fairly close to the value we obtained from average bond energies.

In general, we can calculate ΔH_{rxn} from average bond energies by summing the changes in enthalpy for all of the bonds that are broken and adding the sum of the enthalpy changes for all of the bonds that are formed. Remember that ΔH is positive for breaking bonds and negative for forming them:

$$\Delta H_{\text{rxn}} = \underbrace{\Sigma(\Delta H's \text{ bonds broken})}_{\text{Positive}} + \underbrace{\Sigma(\Delta H's \text{ bonds formed})}_{\text{Negative}}$$

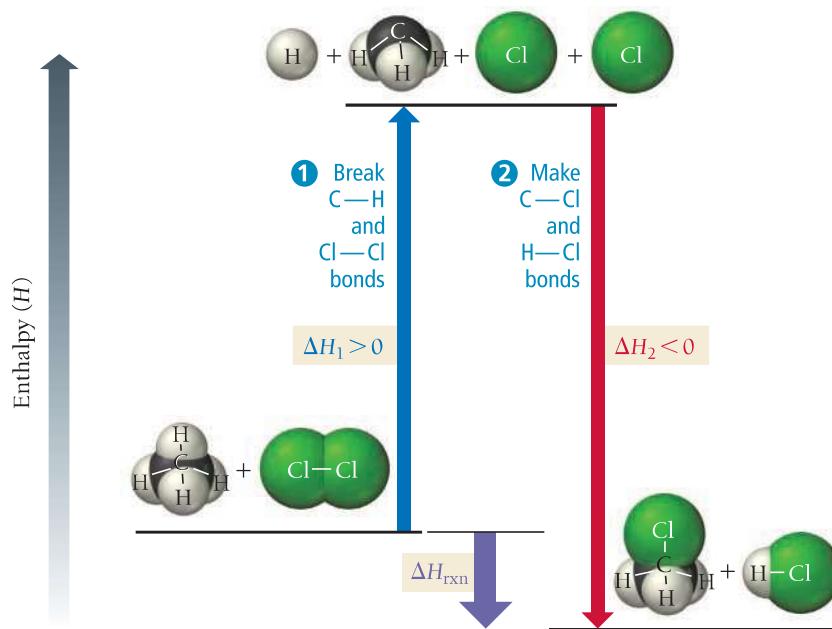
As we can see from the equation:

- A reaction is *exothermic* when weak bonds break and strong bonds form.
- A reaction is *endothermic* when strong bonds break and weak bonds form.

Scientists sometimes say that “energy is stored in chemical bonds or in a chemical compound,” which may make it sound as if breaking the bonds in the compound releases energy. For example, in biology we often hear that energy is stored in glucose or in ATP. However, *breaking a chemical bond always requires energy*. When we say that energy is stored in a compound, or that a compound is energy rich, we mean that the compound can undergo a reaction in which weak bonds break and strong bonds form, thereby releasing energy in the overall process. However, *it is always the forming of chemical bonds that releases energy*.

▼ FIGURE 10.12 Estimating ΔH_{rxn} from Bond Energies We can approximate the enthalpy change of a reaction by summing up the enthalpy changes involved in breaking old bonds and forming new ones.

Estimating the Enthalpy Change of a Reaction from Bond Energies



ANSWER NOW!



10.10 Cc

Conceptual Connection

BOND ENERGIES AND ΔH_{rxn} The reaction between hydrogen and oxygen to form water is highly exothermic. Which statement is true of the energies of the bonds that break and form during the reaction?

- (a) The energy needed to break the required bonds is greater than the energy released when the new bonds form.
- (b) The energy needed to break the required bonds is less than the energy released when the new bonds form.
- (c) The energy needed to break the required bonds is about the same as the energy released when the new bonds form.

WATCH NOW!

INTERACTIVE WORKED EXAMPLE 10.11

EXAMPLE 10.11 Calculating ΔH_{rxn} from Bond Energies

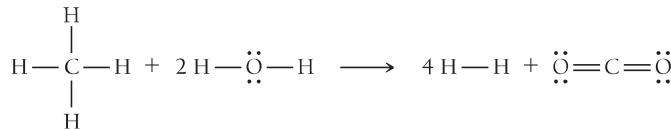
Hydrogen gas, a potential fuel, can be made by the reaction of methane gas and steam.



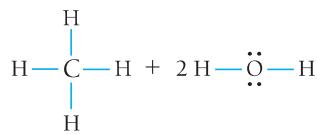
Use bond energies to calculate ΔH_{rxn} for this reaction.

SOLUTION

Begin by rewriting the reaction using the Lewis structures of the molecules involved.

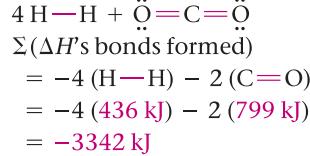


Determine which bonds are broken in the reaction and sum the bond energies of these.



$$\begin{aligned} & \Sigma(\Delta H^{\circ}\text{'s bonds broken}) \\ & = 4(\text{C—H}) + 4(\text{O—H}) \\ & = 4(414 \text{ kJ}) + 4(464 \text{ kJ}) \\ & = 3512 \text{ kJ} \end{aligned}$$

Determine which bonds are formed in the reaction and add the negatives of their bond energies.



Find ΔH_{rxn} by summing the results of the previous two steps.

$$\begin{aligned} \Delta H_{rxn} & = \Sigma(\Delta H^{\circ}\text{'s bonds broken}) + \Sigma(\Delta H^{\circ}\text{'s bonds formed}) \\ & = 3512 \text{ kJ} - 3342 \text{ kJ} \\ & = 1.70 \times 10^2 \text{ kJ} \end{aligned}$$

FOR PRACTICE 10.11 Another potential future fuel is methanol (CH_3OH). Write a balanced equation for the combustion of gaseous methanol and use bond energies to calculate the enthalpy of combustion of methanol in kJ/mol .

FOR MORE PRACTICE 10.11 Use bond energies to calculate ΔH_{rxn} for this reaction: $\text{N}_2(g) + 3 \text{H}_2(g) \longrightarrow 2 \text{NH}_3(g)$.

Bond Lengths

Just as we can tabulate average bond energies, which represent the average energy of a bond between two particular atoms in a large number of compounds, we can tabulate average bond lengths (Table 10.4). The average **bond length** represents the average

TABLE 10.4 ■ Average Bond Lengths

Bond	Bond Length (pm)	Bond	Bond Length (pm)	Bond	Bond Length (pm)
H—H	74	C—C	154	N—N	145
H—C	110	C=C	134	N=N	123
H—N	100	C≡C	120	N≡N	110
H—O	97	C—N	147	N—O	136
H—S	132	C=N	128	N=O	120
H—F	92	C≡N	116	O—O	145
H—Cl	127	C—O	143	O=O	121
H—Br	141	C=O	120	F—F	143
H—I	161	C—Cl	178	Cl—Cl	199
				Br—Br	228
				I—I	266

length of a bond between two particular atoms in a large number of compounds. Like bond energies, bond lengths depend not only on the kind of atoms involved in the bond, but also on the type of bond: single, double, or triple. In general, for a particular pair of atoms, triple bonds are shorter than double bonds, which are in turn shorter than single bonds. For example, consider the bond lengths (shown here with bond strengths) of carbon–carbon triple, double, and single bonds:

Bond	Bond Length (pm)	Bond Strength (kJ/mol)
C≡C	120	837
C=C	134	611
C—C	154	347

Notice that, as the bond gets longer, it also becomes weaker. This relationship between the length of a bond and the strength of a bond does not necessarily hold true for all bonds. Consider the following series of nitrogen–halogen single bonds:

Bond	Bond Length (pm)	Bond Strength (kJ/mol)
N—F	139	272
N—Cl	191	200
N—Br	214	243
N—I	222	159

Although the bonds generally get weaker as they get longer, the trend is not a smooth one.

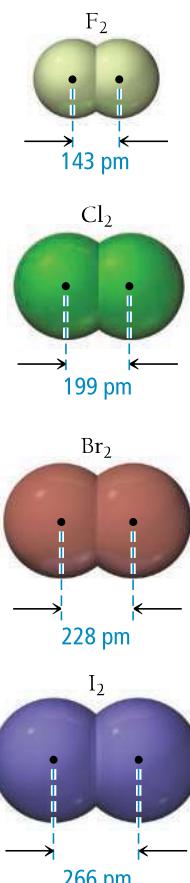
10.11

Bonding in Metals: The Electron Sea Model

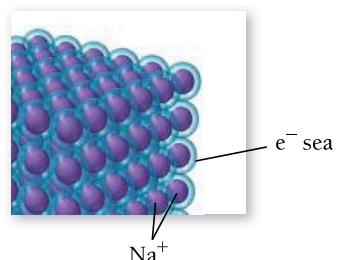
So far, we have explored simple models for bonding between a metal and a nonmetal (ionic bonding) and for bonding between two nonmetals (covalent bonding). We have seen how these models account for and predict the properties of ionic and molecular compounds. The last type of bonding that we examine in this chapter is metallic bonding, which occurs between metals.

As we have already discussed, metals have a tendency to lose electrons, which means that they have relatively low ionization energies. When metal atoms bond together to form a solid lattice, each metal atom donates one or more electrons to an *electron sea*. For example, we can think of sodium metal as an array of positively charged Na^+ ions immersed in a sea of negatively charged electrons (e^-), as shown in Figure 10.13►.

Bond Lengths



▲ Bond lengths in the diatomic halogen molecules.



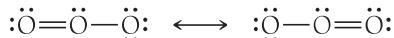
▲ **FIGURE 10.13** The Electron Sea Model for Sodium In this model of metallic bonding, Na^+ ions are immersed in a “sea” of electrons.



CHEMISTRY IN THE ENVIRONMENT

The Lewis Structure of Ozone

Ozone is a form of oxygen in which three oxygen atoms bond together. Its Lewis structure consists of the following resonance structures:



Compare the Lewis structure of ozone to the Lewis structure of O_2 :



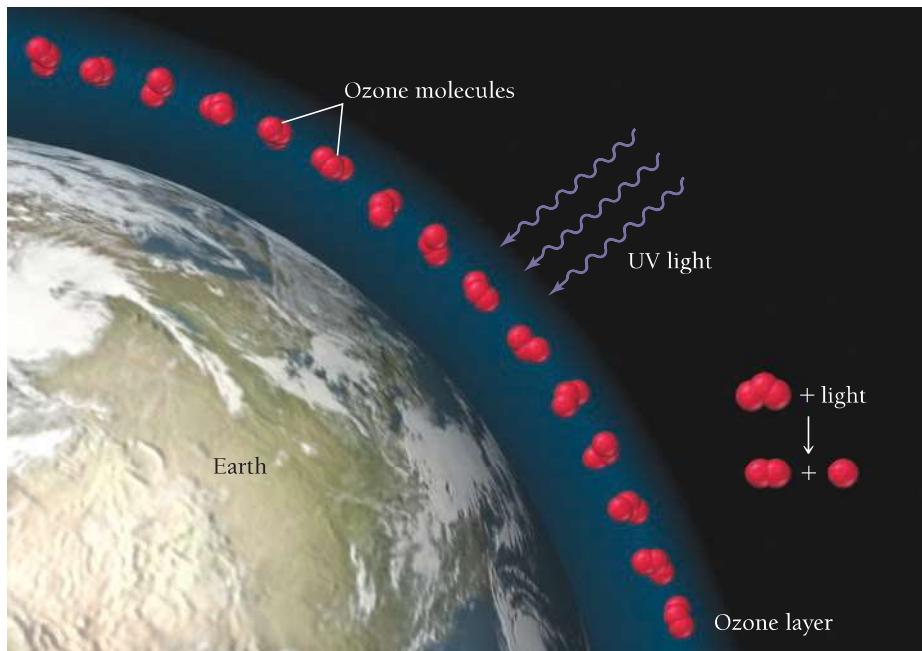
Since double bonds are stronger and shorter than single bonds, O_2 must have a stronger bond because it is a double bond. O_3 , however, has bonds that are intermediate between single and double, which are weaker bonds. The effects of this are significant. As we discussed in Section 7.10, O_3 absorbs harmful ultraviolet light entering Earth's atmosphere. Ozone is ideally suited to do this because photons at wavelengths of

280–320 nm (the most harmful components of sunlight) are just strong enough to break the bonds in the O_3 molecule:

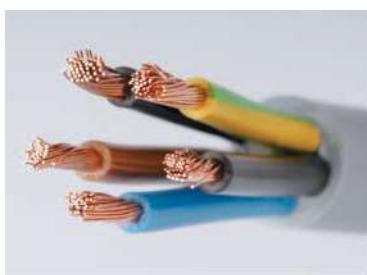


In this process, the photon is absorbed. O_2 and O then recombine to re-form O_3 , which can in turn absorb more UV light. The same wavelengths of UV light, however, do not have sufficient energy to break the stronger double bond of O_2 . No other molecules in our atmosphere can do the job that ozone does. Consequently, we should continue, and even strengthen, the ban on ozone-depleting compounds.

QUESTION Calculate the average bond energy of one O_3 bond. What wavelength of light has just the right amount of energy to break this bond?



▲ Ozone protects life on Earth from harmful ultraviolet light.



▲ Copper can easily be drawn into fine strands like those used in household electrical cords.

Each sodium atom donates its one valence electron to the “sea” and becomes a sodium ion. The sodium cations are held together by their attraction to the sea of electrons.

Although the electron sea model is simple, it accounts for many properties of metals. For example, metals conduct electricity because—in contrast to ionic solids, in which electrons are localized on an ion—the electrons in a metal are free to move. The movement or flow of electrons in response to an electric potential (or voltage) is an electric current. Metals are also excellent conductors of heat, again because of the highly mobile electrons, which help to disperse thermal energy throughout the metal.

The electron sea model also accounts for the *malleability* of metals (their capacity to be pounded into sheets) and the *ductility* of metals (their capacity to be drawn into wires). Since there are no localized or specific “bonds” in a metal, we can deform it relatively easily by forcing the metal ions to slide past one another. The electron sea easily accommodates deformations by flowing into the new shape.

Self-Assessment Quiz



Q1. Which pair of elements is most likely to form an ionic bond?

MISSED THIS? Read Section 10.2

- a) nitrogen and oxygen
- b) carbon and hydrogen
- c) sulfur and oxygen
- d) calcium and oxygen

Q2. Which set of elements is arranged in order of increasing electronegativity?

MISSED THIS? Read Section 10.6; Watch KCV 10.6

- a) O < S < As < Ge
- b) Ge < As < S < O
- c) S < O < As < Ge
- d) As < O < Ge < S

Q3. Which is the correct Lewis structure for magnesium bromide?

MISSED THIS? Read Section 10.4

- a) $2\text{Mg}^{2+}\left[\ddot{\text{Br}}:\right]^-$
- b) $\text{Mg}\ddot{\text{Br}}:$
- c) $\text{Mg}^{2+}2\left[\ddot{\text{Br}}:\right]^-$
- d) $\ddot{\text{Mg}}:\ddot{\text{Br}}:$

Q4. Which compound is likely to have an incomplete octet?

MISSED THIS? Read Section 10.9; Watch KCV 10.9

- a) NH_3
- b) SO_3
- c) N_2O
- d) BH_3

Q5. Which compound has the lattice energy with the greatest magnitude? **MISSED THIS? Read Section 10.4**

- a) MgS
- b) CaS
- c) SrS
- d) BaS

Q6. Which set of compounds is arranged in order of increasing magnitude of lattice energy?

MISSED THIS? Read Section 10.4

- a) CsI < NaCl < MgS
- b) NaCl < CsI < MgS
- c) MgS < NaCl < CsI
- d) CsI < MgS < NaCl

Q7. Which pair of atoms forms the most polar bond?

MISSED THIS? Read Section 10.6; Watch KCV 10.6

- a) N and O
- b) C and O
- c) C and F
- d) N and F

Q8. Which pair of atoms forms a nonpolar covalent bond?

MISSED THIS? Read Section 10.6; Watch KCV 10.6

- a) C and S
- b) C and O
- c) B and O
- d) Na and Cl

Q9. Which is the correct Lewis structure for nitrogen trifluoride?

MISSED THIS? Read Section 10.7; Watch KCV 10.7, IWE 10.4

- a) $\ddot{\text{F}}=\ddot{\text{N}}-\ddot{\text{F}}:$
- b) $\ddot{\text{F}}-\ddot{\text{N}}-\ddot{\text{F}}:$
- c) $\ddot{\text{F}}-\ddot{\text{N}}-\ddot{\text{F}}:$
- d) $\ddot{\text{F}}-\ddot{\text{N}}-\ddot{\text{F}}-\ddot{\text{F}}:$

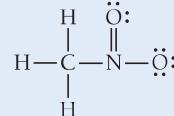
Q10. Which is the correct Lewis structure for CO_3^{2-} ?

MISSED THIS? Read Section 10.7; Watch KCV 10.7, IWE 10.6

- a) $\left[\ddot{\text{O}}=\text{C}=\ddot{\text{O}}:\right]^{2-}$
- b) $\left[\ddot{\text{O}}-\text{C}-\ddot{\text{O}}:\right]^{2-} \longleftrightarrow \left[\ddot{\text{O}}=\text{C}-\ddot{\text{O}}:\right]^{2-} \longleftrightarrow \left[\ddot{\text{O}}-\text{C}=\ddot{\text{O}}:\right]^{2-}$
- c) $\left[\ddot{\text{O}}-\text{C}-\ddot{\text{O}}:\right]^{2-}$
- d) $\left[\ddot{\text{O}}-\ddot{\text{C}}=\ddot{\text{O}}-\ddot{\text{O}}:\right]^{2-}$

Q11. Determine the formal charge of nitrogen in the structure shown here:

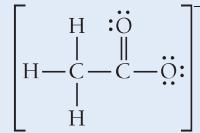
MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.8



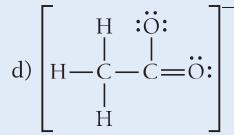
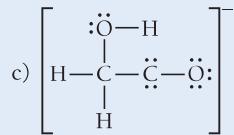
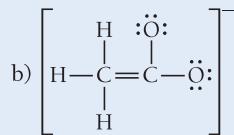
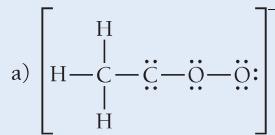
- a) +1
- b) +2
- c) -1
- d) -2

Q12. A Lewis structure for the acetate ion is shown here:

MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.7

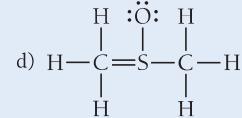
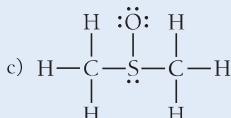
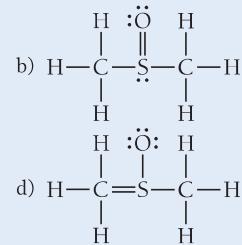
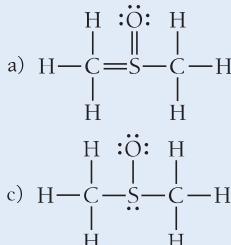


Which structure is the best resonance structure for the acetate ion?



Q13. Use formal charge to choose the best Lewis structure for CH_3SOCH_3 .

MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.8



—Continued on the next page

Continued—

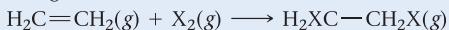
- Q14.** Use bond energies from Table 10.3 to determine ΔH_{rxn} for the reaction between ethanol and hydrogen chloride.

MISSED THIS? Read Section 10.10; Watch IWE 10.11



- a) -1549 kJ b) 1549 kJ c) -12 kJ d) 12 kJ

- Q15.** Consider the halogenation of ethene, where X is a generic halogen:



Use bond energies to determine which halogen produces the most exothermic halogenation reaction with ethene. The C—F, C—Br, and C—I bond energies are 552 kJ/mol, 280 kJ/mol, and 209 kJ/mol, respectively. Look up all other necessary bond energies in Table 10.3.

MISSED THIS? Read Section 10.10; Watch IWE 10.11

- | | |
|-------------|-------------|
| a) fluorine | b) chlorine |
| c) bromine | d) iodine |

Answers: 1. (d) 2. (b) 3. (c) 4. (d) 5. (a) 6. (a) 7. (c) 8. (a) 9. (c) 10. (b) 11. (a) 12. (d) 13. (b) 14. (c) 15. (a)

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CONCEPTS

Bonding Models and AIDS Drugs (10.1)

- Theories that predict how and why atoms bond together are central to chemistry because they explain compound stability and molecule shape.
- Bonding theories have been useful in combating HIV because they help in the design of molecules that bind to the active site of a protein crucial for the development of AIDS.

Types of Chemical Bonds (10.2)

- We divide chemical bonds into three general types: ionic bonds, which occur between a metal and a nonmetal; covalent bonds, which occur between two nonmetals; and metallic bonds, which occur within metals.
- In an ionic bond, an electron transfers from the metal to the nonmetal, and the resultant ions attract each other by coulombic forces.
- In a covalent bond, nonmetals share electrons that interact with the nuclei of both atoms via coulombic forces, holding the atoms together.
- In a metallic bond, the atoms form a lattice in which each metal loses electrons to an “electron sea.” The attraction of the positively charged metal ions to the electron sea holds the metal together.

The Lewis Model and Electron Dots (10.3)

- In the Lewis model, chemical bonds form when atoms transfer (ionic bonding) or share (covalent bonding) valence electrons to attain noble gas electron configurations.
- The Lewis model represents valence electrons as dots surrounding the symbol for an element. When two or more elements bond together, the dots are transferred or shared so that every atom has eight dots, an octet (or two dots, a duet, in the case of hydrogen).

Ionic Lewis Structures and Lattice Energy (10.4)

- In an ionic Lewis structure involving main-group metals, the metal transfers its valence electrons (dots) to the nonmetal.
- The formation of most ionic compounds is exothermic because of lattice energy, the energy released when metal cations and nonmetal anions coalesce to form the solid. The smaller the radius of the ions and the greater their charge, the more exothermic the lattice energy.

Covalent Lewis Structures, Electronegativity, and Polarity (10.5, 10.6, 10.7)

- In a covalent Lewis structure, neighboring atoms share valence electrons to attain octets (or duets).

- A single shared electron pair constitutes a single bond, while two or three shared pairs constitute double or triple bonds, respectively.
- The shared electrons in a covalent bond are not always *equally* shared; when two dissimilar nonmetals form a covalent bond, the electron density is greater on the more electronegative element. The result is a polar bond, in which one element has a partial positive charge and the other a partial negative charge.
- Electronegativity—the ability of an atom to attract electrons to itself in chemical bonding—increases as we move to the right across a period in the periodic table and decreases as we move down a column.
- Elements with very dissimilar electronegativities form ionic bonds, those with very similar electronegativities form nonpolar covalent bonds, and those with intermediate electronegativity differences form polar covalent bonds.

Resonance and Formal Charge (10.8)

- We can best represent some molecules not by a single Lewis structure, but by two or more resonance structures. The actual structure of these molecules is a resonance hybrid: a combination or average of the contributing structures.
- The formal charge of an atom in a Lewis structure is the charge the atom would have if all bonding electrons were shared equally between bonding atoms.
- In general, the best Lewis structures have the fewest atoms with formal charge, and any negative formal charges are on the most electronegative atom.

Exceptions to the Octet Rule (10.9)

- Although the octet rule normally applies when we draw Lewis structures, some exceptions occur.
- These exceptions include odd-electron species, which necessarily have Lewis structures with only seven electrons around an atom. Such molecules, called free radicals, tend to be unstable and chemically reactive.
- Other exceptions to the octet rule include molecules with incomplete octets—usually totaling six electrons (especially important in compounds containing boron)—and molecules with expanded octets—usually 10 or 12 electrons (which can occur in compounds containing elements from the third row of the periodic table and below). Expanded octets never occur in second-period elements.

Bond Energies and Bond Lengths (10.10)

- The bond energy of a chemical bond is the energy required to break one mole of the bonds in the gas phase.
- Average bond energies for a number of different bonds are tabulated, and we can use them to calculate enthalpies of reaction.
- Average bond lengths are also tabulated.
- In general, triple bonds are shorter and stronger than double bonds, which are in turn shorter and stronger than single bonds.

Bonding in Metals (10.11)

- When metal atoms bond together to form a solid, each metal atom donates one or more electrons to an *electron sea*. The metal cations are held together by their attraction to the sea of electrons.
- The *electron sea* model accounts for the electrical conductivity, thermal conductivity, malleability, and ductility of metals.

EQUATIONS AND RELATIONSHIPS

Dipole Moment (μ): Separation of Two Particles of Equal but Opposite Charge of Magnitude q by a Distance r (10.6)

$$\mu = qr$$

Percent Ionic Character (10.6)

Percent ionic character = $\frac{\text{measured dipole moment of bond}}{\text{dipole moment if electron were completely transferred}} \times 100\%$

Formal Charge (10.8)

Formal charge = number of valence electrons –

(number of nonbonding electrons + $\frac{1}{2}$ number of shared electrons)

Enthalpy Change of a Reaction (ΔH_{rxn}): Relationship of Bond Energies (10.10)

$$\Delta H_{rxn} = \Sigma(\Delta H^* \text{ bonds broken}) + \Sigma(\Delta H^* \text{ bonds formed})$$

LEARNING OUTCOMES

Chapter Objectives	Assessment
Write Lewis symbols for main group elements (10.3)	Exercises 35–38
Write and use Lewis symbols to describe and analyze ionic compounds (10.4)	Example 10.1 For Practice 10.1 Exercise 39–42
Compare the relative lattice energies of different compounds (10.4)	Example 10.2 For Practice 10.2 For More Practice 10.2 Exercises 43–46
Calculate lattice energies of ionic solids (10.4)	Exercises 47–48
Classify bonds as pure covalent, polar covalent, or ionic (10.5–10.6)	Example 10.3 For Practice 10.3 Exercises 55–58
Write and use Lewis structures to describe and analyze molecular compounds and polyatomic ions (10.7)	Examples 10.4–10.7 For Practice 10.4–10.7 Exercises 49–54, 59–62

—Continued on the next page

Continued—

Write the best possible Lewis structure using formal charge (10.8)	Examples 10.8, 10.9 For Practice 10.8, 10.9 For More Practice 10.8 Exercises 63–72
Write Lewis structures for molecules and ions that are exceptions to the octet rule (10.9)	Example 10.10 For Practice 10.10 For More Practice 10.10 Exercises 73–78
Analyze bonds in terms of bond energy and bond length (10.10)	Example 10.11 For Practice 10.11 For More Practice 10.11 Exercises 79–84

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

- Why are bonding theories important? Provide some examples of what bonding theories can predict.
- Why do chemical bonds form? What basic forces are involved in bonding?
- What are the three basic types of chemical bonds? What happens to electrons in the bonding atoms in each type?
- How do you determine how many dots to put around the Lewis symbol of an element?
- Describe the octet rule in the Lewis model.
- According to the Lewis model, what is a chemical bond?
- How do you draw an ionic Lewis structure?
- How can Lewis structures be used to determine the formula of ionic compounds? Give an example.
- What is lattice energy?
- Why is the formation of solid sodium chloride from solid sodium and gaseous chlorine exothermic, even though it takes more energy to form the Na^+ ion than the amount of energy released upon formation of Cl^- ?
- What is the Born–Haber cycle? List each step in the cycle and show how the cycle is used to calculate lattice energy.
- How does lattice energy relate to ionic radii? To ion charge?
- How does the ionic bonding model explain the relatively high melting points of ionic compounds?
- How does the ionic bonding model explain the nonconductivity of ionic solids, and at the same time the conductivity of ionic solutions?
- In a covalent Lewis structure, what is the difference between lone pair and bonding pair electrons?
- In what ways are double and triple covalent bonds different from single covalent bonds?
- How does the Lewis model for covalent bonding account for why certain combinations of atoms are stable while others are not?
- How does the Lewis model for covalent bonding account for the relatively low melting and boiling points of molecular compounds (compared to ionic compounds)?
- What is electronegativity? What are the periodic trends in electronegativity?
- How do a pure covalent bond, a polar covalent bond, and an ionic bond differ?
- Explain percent ionic character of a bond. Do any bonds have 100% ionic character?
- What is a dipole moment?
- What is the magnitude of the dipole moment formed by separating a proton and an electron by 100 pm? 200 pm?
- What is the basic procedure for writing a covalent Lewis structure?
- How do you determine the number of electrons in the Lewis structure of a molecule? A polyatomic ion?
- What are resonance structures? What is a resonance hybrid?
- Do resonance structures always contribute equally to the overall structure of a molecule? Explain.
- What is formal charge? How is formal charge calculated? How is it helpful?
- Why does the octet rule have exceptions? List the three major categories of exceptions and an example of each.
- Which elements can have expanded octets? Which elements should never have expanded octets?
- What is bond energy? How can you use average bond energies to calculate enthalpies of reaction?
- Explain the difference between endothermic reactions and exothermic reactions with respect to the bond energies of the bonds broken and formed.
- What is the electron sea model for bonding in metals?
- How does the electron sea model explain the conductivity of metals? The malleability and ductility of metals?

PROBLEMS BY TOPIC

Valence Electrons and Dot Structures

- Write the electron configuration for N. Then write the Lewis symbol for N and show which electrons from the electron configuration are included in the Lewis symbol.
MISSED THIS? Read Section 10.3
- Write the electron configuration for Ne. Then write the Lewis symbol for Ne and show which electrons from the electron configuration are included in the Lewis symbol.

- Write the Lewis symbol for each atom or ion.

MISSED THIS? Read Section 10.3

- Al
- Na^+
- Cl
- Cl^-

- Write the Lewis symbol for each atom or ion.

- S^{2-}
- Mg
- Mg^{2+}
- P

Ionic Lewis Symbols and Lattice Energy

39. Write the Lewis symbols for the ions in each ionic compound.
MISSED THIS? Read Section 10.4
- NaF
 - CaO
 - SrBr₂
 - K₂O
40. Write the Lewis symbols for the ions in each ionic compound.
- SrO
 - Li₂S
 - CaI₂
 - RbF
41. Use Lewis symbols to determine the formula for the compound that forms between each pair of elements.
MISSED THIS? Read Section 10.4
- Sr and Se
 - Ba and Cl
 - Na and S
 - Al and O
42. Use Lewis symbols to determine the formula for the compound that forms between each pair of elements:
- Ca and N
 - Mg and I
 - Ca and S
 - Cs and F
43. Explain the trend in the lattice energies of the alkaline earth metal oxides. **MISSED THIS? Read Section 10.4**

Metal Oxide	Lattice Energy (kJ/mol)
MgO	-3795
CaO	-3414
SrO	-3217
BaO	-3029

44. Rubidium iodide has a lattice energy of -617 kJ/mol, while potassium bromide has a lattice energy of -671 kJ/mol. Why is the lattice energy of potassium bromide more exothermic than the lattice energy of rubidium iodide?
45. The lattice energy of CsF is -744 kJ/mol, whereas that of BaO is -3029 kJ/mol. Explain this large difference in lattice energy.
MISSED THIS? Read Section 10.4
46. Arrange these compounds in order of increasing magnitude of lattice energy: KCl, SrO, RbBr, CaO.
47. Use the Born–Haber cycle and data from Appendix IIB, Chapter 9, and this chapter to calculate the lattice energy of KCl. (ΔH_{sub} for potassium is 89.0 kJ/mol.) **MISSED THIS? Read Section 10.4**
48. Use the Born–Haber cycle and data from Appendix IIB and Table 10.3 to calculate the lattice energy of CaO. (ΔH_{sub} for calcium is 178 kJ/mol; IE₁ and IE₂ for calcium are 590 kJ/mol and 1145 kJ/mol, respectively; EA₁ and EA₂ for O are -141 kJ/mol and 744 kJ/mol, respectively.)

Simple Covalent Lewis Structures, Electronegativity, and Bond Polarity

49. Use covalent Lewis structures to explain why each element (or family of elements) occurs as diatomic molecules.
MISSED THIS? Read Section 10.5; Watch KCV 10.5
- hydrogen
 - the halogens
 - oxygen
 - nitrogen
50. Use covalent Lewis structures to explain why the compound that forms between nitrogen and hydrogen has the formula NH₃. Show why NH₂ and NH₄ are not stable.
51. Write the Lewis structure for each molecule.
MISSED THIS? Read Section 10.7; Watch KCV 10.7, IWE 10.4
- PH₃
 - SCl₂
 - HI
 - CH₄
52. Write the Lewis structure for each molecule.
- NF₃
 - HBr
 - SBr₂
 - CCl₄

53. Write the Lewis structure for each molecule.

MISSED THIS? Read Section 10.7; Watch KCV 10.7, IWE 10.4

- SF₂
- SiH₄
- HCOOH (both O bonded to C)
- CH₃SH (C and S central)

54. Write the Lewis structure for each molecule.

- CH₂O
- C₂Cl₄
- CH₃NH₂
- CFCl₃ (C central)

55. Determine if a bond between each pair of atoms would be pure covalent, polar covalent, or ionic.

MISSED THIS? Read Section 10.6; Watch KCV 10.6

- Br and Br
- C and Cl
- C and S
- Sr and O

56. Determine if a bond between each pair of atoms would be pure covalent, polar covalent, or ionic.

- C and N
- N and S
- K and F
- N and N

57. Draw the Lewis structure for CO with an arrow representing the dipole moment. Refer to Figure 10.10 to estimate the percent ionic character of the CO bond.

MISSED THIS? Read Section 10.6; Watch KCV 10.6

58. Draw the Lewis structure for BrF with an arrow representing the dipole moment. Refer to Figure 10.10 to estimate the percent ionic character of the BrF bond.

Covalent Lewis Structures, Resonance, and Formal Charge

59. Write the Lewis structure for each molecule or ion.

MISSED THIS? Read Section 10.7; Watch KCV 10.7, IWE 10.4

- Cl₄
- N₂O
- SiH₄
- Cl₂CO

60. Write the Lewis structure for each molecule or ion.

- H₃COH
- OH⁻
- BrO⁻
- O₂²⁻

61. Write the Lewis structure for each molecule or ion.

MISSED THIS? Read Section 10.7; Watch KCV 10.7, IWE 10.4

- N₂H₂
- N₂H₄
- C₂H₂
- C₂H₄

62. Write the Lewis structure for each molecule or ion.

- H₃COCH₃
- CN⁻
- NO₂⁻
- ClO⁻

63. Write a Lewis structure that obeys the octet rule for each molecule or ion. Include resonance structures if necessary and assign formal charges to each atom.

MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.7

- SeO₂
- CO₃²⁻
- ClO⁻
- NO₂⁻

64. Write a Lewis structure that obeys the octet rule for each ion. Include resonance structures if necessary and assign formal charges to each atom.

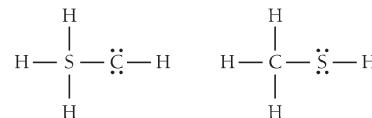
- ClO₃⁻
- ClO₄⁻
- NO₃⁻
- NH₄⁺

65. Use formal charge to identify the better Lewis structure.

MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.8



66. Use formal charges to identify the better Lewis structure.



67. How important is the resonance structure shown here to the overall structure of carbon dioxide? Explain.

MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.8



68. In N_2O , nitrogen is the central atom, and the oxygen atom is terminal. In OF_2 , however, oxygen is the central atom. Use formal charges to explain why.

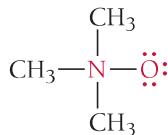
69. Draw the Lewis structure (including resonance structures) for the acetate ion (CH_3COO^-). For each resonance structure, assign formal charges to all atoms that have formal charge.

MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.9

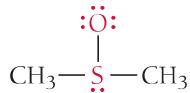
70. Draw the Lewis structure (including resonance structures) for methyl azide (CH_3N_3). For each resonance structure, assign formal charges to all atoms that have formal charge.

71. What are the formal charges of the atoms shown in red?

MISSED THIS? Read Section 10.8; Watch KCV 10.8, IWE 10.9



72. What are the formal charges of the atoms shown in red?



Odd-Electron Species, Incomplete Octets, and Expanded Octets

73. Write the Lewis structure for each molecule (octet rule not followed). **MISSED THIS?** Read Section 10.9; Watch KCV 10.9

- a. BCl_3 b. NO_2 c. BH_3

74. Write the Lewis structure for each molecule (octet rule not followed).

- a. BBr_3
b. NO
c. ClO_2

75. Write the Lewis structure for each ion. Include resonance structures if necessary and assign formal charges to all atoms. If necessary, expand the octet on the central atom to lower formal charge.

MISSED THIS? Read Section 10.9; Watch KCV 10.9, IWE 10.10

- a. PO_4^{3-} b. CN^- c. SO_3^{2-} d. ClO_2^-

76. Write Lewis structures for each molecule or ion. Include resonance structures if necessary and assign formal charges to all atoms. If necessary, expand the octet on the central atom to lower formal charge.

- a. SO_4^{2-} b. HSO_4^- c. SO_3 d. BrO_2^-

77. Write Lewis structures for each molecule or ion. Use expanded octets as necessary.

MISSED THIS? Read Section 10.9; Watch KCV 10.9, IWE 10.10

- a. PF_5 b. I_3^-
c. SF_4 d. GeF_4

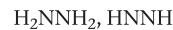
78. Write Lewis structures for each molecule or ion. Use expanded octets as necessary.

- a. ClF_5 b. AsF_6^-
c. Cl_3PO d. IF_5

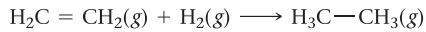
Bond Energies and Bond Lengths

79. List these compounds in order of increasing carbon–carbon bond strength and in order of decreasing carbon–carbon bond length: HCCH , H_2CCH_2 , H_3CCH_3 . **MISSED THIS?** Read Section 10.10

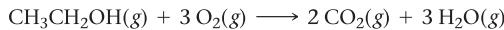
80. Which compound shown here has the stronger nitrogen–nitrogen bond? The shorter nitrogen–nitrogen bond?



81. Hydrogenation reactions are used to add hydrogen across double bonds in hydrocarbons and other organic compounds. Use average bond energies to calculate ΔH_{rxn} for the hydrogenation reaction. **MISSED THIS?** Read Section 10.10; Watch IWE 10.11

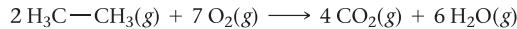


82. Ethanol is a possible fuel. Use average bond energies to calculate ΔH_{rxn} for the combustion of ethanol.



83. Ethane burns in air to form carbon dioxide and water vapor.

MISSED THIS? Read Section 10.10; Watch IWE 10.11



Use average bond energies to calculate ΔH_{rxn} for the reaction.

84. In the *Chemistry and the Environment* box on free radicals in this chapter, we discussed the importance of the hydroxyl radical in reacting with and eliminating many atmospheric pollutants. However, the hydroxyl radical does not clean up everything. For example, chlorofluorocarbons—which destroy stratospheric ozone—are not attacked by the hydroxyl radical. Consider the hypothetical reaction by which the hydroxyl radical might react with a chlorofluorocarbon:



Use bond energies to explain why this reaction is improbable. (The C—F bond energy is 552 kJ/mol.)

CUMULATIVE PROBLEMS

85. Write an appropriate Lewis structure for each compound. Make certain to distinguish between ionic and molecular compounds.

- a. BI_3 b. K_2S c. HCFO d. PBr_3

86. Write an appropriate Lewis structure for each compound. Make certain to distinguish between ionic and molecular compounds.

- a. Al_2O_3 b. ClF_5 c. MgI_2 d. XeO_4

87. Each compound contains both ionic and covalent bonds. Write ionic Lewis structures for each of them, including the covalent structure for the ion in brackets. Write resonance structures if necessary.

- a. BaCO_3 b. $\text{Ca}(\text{OH})_2$
c. KNO_3 d. LiIO

88. Each compound contains both ionic and covalent bonds. Write ionic Lewis structures for each of them, including the covalent structure for the ion in brackets. Write resonance structures if necessary.

- a. RbIO_2 b. NH_4Cl c. KOH d. $\text{Sr}(\text{CN})_2$

89. Carbon ring structures are common in organic chemistry. Draw a Lewis structure for each carbon ring structure, including any necessary resonance structures.

- a. C_4H_8 b. C_4H_4 c. C_6H_{12} d. C_6H_6

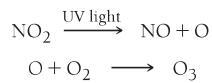
90. Amino acids are the building blocks of proteins. The simplest amino acid is glycine ($\text{H}_2\text{NCH}_2\text{COOH}$). Draw a Lewis structure for glycine. (Hint: The central atoms in the skeletal structure are nitrogen and the two carbon atoms. Each oxygen atom is bonded directly to the right-most carbon atom.)

- 91.** Formic acid is responsible for the sting of ant bites. By mass, formic acid is 26.10% C, 4.38% H, and 69.52% O. The molar mass of formic acid is 46.02 g/mol. Determine the molecular formula of formic acid and draw its Lewis structure.
- 92.** Diazomethane is a highly poisonous, explosive compound because it readily evolves N₂. Diazomethane has the following composition by mass: 28.57% C; 4.80% H; and 66.64% N. The molar mass of diazomethane is 42.04 g/mol. Find the molecular formula of diazomethane, draw its Lewis structure, and assign formal charges to each atom. Why is diazomethane not very stable? Explain.
- 93.** The reaction of Fe₂O₃(s) with Al(s) to form Al₂O₃(s) and Fe(s) is called the thermite reaction and is highly exothermic. What role does lattice energy play in the exothermicity of the reaction?
- 94.** NaCl has a lattice energy of -787 kJ/mol . Consider a hypothetical salt XY. X³⁺ has the same radius as Na⁺ and Y³⁻ has the same radius as Cl⁻. Estimate the lattice energy of XY.
- 95.** Draw the Lewis structure for nitric acid (the hydrogen atom is attached to one of the oxygen atoms). Include all three resonance structures by alternating the double bond among the three oxygen atoms. Use formal charge to determine which of the resonance structures is most important to the structure of nitric acid.
- 96.** Phosgene (Cl₂CO) is a poisonous gas used as a chemical weapon during World War I. It is a potential agent for chemical terrorism today. Draw the Lewis structure of phosgene. Include all three resonance forms by alternating the double bond among the three terminal atoms. Which resonance structure is the best?
- 97.** The cyanate ion (OCN⁻) and the fulminate ion (CNO⁻) share the same three atoms but have vastly different properties. The cyanate ion is stable, while the fulminate ion is unstable and forms explosive compounds. The resonance structures of the cyanate ion are explored in Example 10.8. Draw Lewis structures for the fulminate ion—including possible resonance forms—and use formal charge to explain why the fulminate ion is less stable (and therefore more reactive) than the cyanate ion.
- 98.** Draw the Lewis structure for each organic compound from its condensed structural formula.
- a. C₃H₈
 - b. CH₃OCH₃
 - c. CH₃COCH₃
 - d. CH₃COOH
 - e. CH₃CHO
- 99.** Draw the Lewis structure for each organic compound from its condensed structural formula.
- a. C₂H₄
 - b. CH₃NH₂
 - c. HCHO
 - d. CH₃CH₂OH
 - e. HCOOH
- 100.** Use Lewis structures to explain why Br₃⁻ and I₃⁻ are stable, while F₃⁻ is not.
- 101.** Draw the Lewis structure for HCSNH₂. (The carbon and nitrogen atoms are bonded together, and the sulfur atom is bonded to the carbon atom.) Label each bond in the molecule as polar or nonpolar.
- 102.** Draw the Lewis structure for urea, H₂NCONH₂, one of the compounds responsible for the smell of urine. (The central carbon atom is bonded to both nitrogen atoms and to the oxygen atom.) Does urea contain polar bonds? Which bond in urea is most polar?
- 103.** Some theories of aging suggest that free radicals cause certain diseases and perhaps aging in general. As you know from the Lewis model, such molecules are not chemically stable and will quickly react with other molecules. According to certain theories, free radicals may attack molecules within the cell, such as DNA, changing them and causing cancer or other diseases.

Free radicals may also attack molecules on the surfaces of cells, making them appear foreign to the body's immune system. The immune system then attacks the cells and destroys them, weakening the body. Draw Lewis structures for each free radical implicated in this theory of aging.

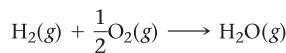
- a. O₂⁻
- b. O⁻
- c. OH
- d. CH₃OO (unpaired electron on terminal oxygen)

- 104.** Free radicals are important in many environmentally significant reactions (see the *Chemistry in the Environment* box on free radicals in this chapter). For example, photochemical smog—smog that results from the action of sunlight on air pollutants—forms in part by these two steps:



The product of this reaction, ozone, is a pollutant in the lower atmosphere. (Upper atmospheric ozone is a natural part of the atmosphere that protects life on Earth from ultraviolet light.) Ozone is an eye and lung irritant and also accelerates the weathering of rubber products. Rewrite the given reactions using the Lewis structure of each reactant and product. Identify the free radicals.

- 105.** If hydrogen were used as a fuel, it could be burned according to this reaction:



Use average bond energies to calculate ΔH_{rxn} for this reaction and also for the combustion of methane (CH₄). Which fuel yields more energy per mole? Per gram?

- 106.** Calculate ΔH_{rxn} for the combustion of octane (C₈H₁₈), a component of gasoline, by using average bond energies and then calculate it using enthalpies of formation from Appendix IIB. What is the percent difference between your results? Which result would you expect to be more accurate?

- 107.** Draw the Lewis structure for each compound.

- a. Cl₂O₇ (no Cl—Cl bond)
- b. H₃PO₃ (two OH bonds)
- c. H₃AsO₄

- 108.** The azide ion, N₃⁻, is a symmetrical ion, all of whose contributing resonance structures have formal charges. Draw three important contributing structures for this ion.

- 109.** List the following gas-phase ion pairs in order of the quantity of energy released when they form from separated gas-phase ions. List the pair that releases the least energy first. Na⁺F⁻, Mg²⁺F⁻, Na⁺O²⁻, Mg²⁺O²⁻, Al³⁺O²⁻.

- 110.** Calculate ΔH° for the reaction $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \longrightarrow 2 \text{HBr}(\text{g})$ using the bond energy values. The ΔH_f° of HBr(g) is not equal to one-half of the value calculated. Account for the difference.

- 111.** The heat of atomization is the heat required to convert a molecule in the gas phase into its constituent atoms in the gas phase. The heat of atomization is used to calculate average bond energies. Without using any tabulated bond energies, calculate the average C—Cl bond energy from the following data: the heat of atomization of CH₄ is 1660 kJ/mol, and the heat of atomization of CH₂Cl₂ is 1495 kJ/mol.

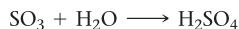
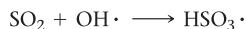
- 112.** Calculate the heat of atomization (see previous problem) of C₂H₃Cl, using the average bond energies in Table 10.3.

- 113.** A compound composed of only carbon and hydrogen is 7.743% hydrogen by mass. Propose a Lewis structure for the compound.

- 114.** A compound composed of only carbon and chlorine is 85.5% chlorine by mass. Propose a Lewis structure for the compound.

CHALLENGE PROBLEMS

- 115.** The main component of acid rain (H_2SO_4) forms from the SO_2 pollutant in the atmosphere via these steps:



Draw the Lewis structure for each of the species in these steps and use bond energies and Hess's law to estimate ΔH_{rxn} for the overall process. (Use 265 kJ/mol for the S—O single bond energy.)

- 116.** A 0.167-g sample of an unknown acid requires 27.8 mL of 0.100 M NaOH to titrate to the equivalence point. Elemental analysis of the acid gives the following percentages by mass: 40.00% C, 6.71% H, 53.29% O. Determine the molecular formula, molar mass, and Lewis structure of the unknown acid.

- 117.** Use the dipole moments of HF and HCl (given at the end of the problem) together with the percent ionic character of each bond (Figure 10.10) to estimate the bond length in each molecule. How well does your estimated bond length agree with the bond length in Table 10.4?

$$\text{HCl } \mu = 1.08 \text{ D}$$

$$\text{HF } \mu = 1.82 \text{ D}$$

- 118.** Use average bond energies together with the standard enthalpy of formation of $\text{C}(g)$ (718.4 kJ/mol) to estimate the standard enthalpy of formation of gaseous benzene, $\text{C}_6\text{H}_6(g)$. (Remember that average bond energies apply to the gas phase only.) Compare the value you obtain using average bond energies to the actual standard enthalpy of formation of gaseous benzene, 82.9 kJ/mol. What does the difference between these two values tell you about the stability of benzene?

- 119.** The standard state of phosphorus at 25 °C is P_4 . This molecule has four equivalent P atoms, no double or triple bonds, and no expanded octets. Draw its Lewis structure.

- 120.** The standard heat of formation of CaBr_2 is −675 kJ/mol. The first ionization energy of Ca is 590 kJ/mol, and its second ionization energy is 1145 kJ/mol. The heat of sublimation of $\text{Ca}[\text{Ca}(s) \longrightarrow \text{Ca}(g)]$ is 178 kJ/mol. The bond energy of Br_2 is 193 kJ/mol, the heat of vaporization of $\text{Br}_2(l)$ is 31 kJ/mol, and the electron affinity of Br is −325 kJ/mol. Calculate the lattice energy of CaBr_2 .

- 121.** The standard heat of formation of $\text{PI}_3(s)$ is −24.7 kJ/mol, and the PI bond energy in this molecule is 184 kJ/mol. The standard heat of formation of $\text{P}(g)$ is 334 kJ/mol, and that of $\text{I}_2(g)$ is 62 kJ/mol. The I_2 bond energy is 151 kJ/mol. Calculate the heat of sublimation of $\text{PI}_3[\text{PI}_3(s) \longrightarrow \text{PI}_3(g)]$.

- 122.** A compound has the formula C_8H_8 and does not contain any double or triple bonds. All the carbon atoms are chemically identical, and all the hydrogen atoms are chemically identical. Draw the Lewis structure for this molecule.

- 123.** Find the oxidation number of each sulfur in the molecule H_2S_4 , which has a linear arrangement of its atoms.

- 124.** Ionic solids of the O^- and O^{3-} anions do not exist, while ionic solids of the O^{2-} anion are common. Explain.

- 125.** The standard state of sulfur is solid rhombic sulfur. Use the appropriate standard heats of formation given in Appendix II to find the average bond energy of the $\text{S}=\text{O}$ in SO_2 .

CONCEPTUAL PROBLEMS

- 126.** Which statement is true of an endothermic reaction?

- Strong bonds break and weak bonds form.
- Weak bonds break and strong bonds form.
- The bonds that break and those that form are of approximately the same strength.

- 127.** When a firecracker explodes, energy is obviously released. The compounds in the firecracker can be viewed as being “energy rich.” What does this mean? Explain the source of the energy in terms of chemical bonds.

- 128.** A fundamental difference between compounds containing ionic bonds and those containing covalent bonds is the existence of molecules. Explain why molecules exist in solid covalent compounds but not in solid ionic compounds.

- 129.** In the very first chapter of this book, we described the scientific approach and put a special emphasis on scientific models or theories. In this chapter, we looked carefully at a model for chemical bonding (the Lewis model). Why is this theory successful? What are some of the limitations of the theory?

QUESTIONS FOR GROUP WORK

Discuss these questions with the group and record your consensus answer.

- 130.** Have each member of your group represent an atom of a metal or an atom of a nonmetal. Each group member holds a coin to represent an electron. Which group members are most reluctant to give up their electrons? Which group members are most willing to give up their electrons? Determine which kind of bond could form between each pair of group members. Tabulate your results.

- 131.** Spend a few minutes reviewing the Lewis dot symbols for the atoms H through Ne. Form a circle and have each group member ask the group member on his or her right to draw the Lewis symbol for a specific atom. Keep going around until each group

Active Classroom Learning

member can write all the Lewis dot symbols for the atoms H through Ne. Determine the formal charge for each symbol. In a complete sentence or two, describe why they are all the same.

- 132.** Draw the Lewis dot symbols for the atoms Al and O. Use the Lewis model to determine the formula for the compound formed from these two atoms.

- 133.** Draft a list of step-by-step instructions for writing the correct Lewis dot structure for any molecule or polyatomic ion.

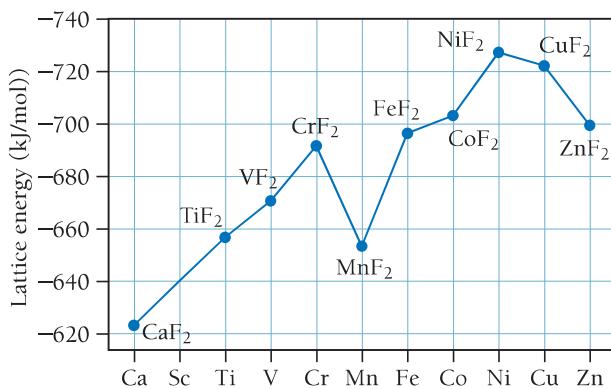
- 134.** Pass a piece of paper around the group and ask each group member in turn to perform the next step in the process of determining a correct Lewis structure (including formal charges on all atoms and resonance structures, if appropriate) for the following molecules and ions: N_2H_4 , CCl_4 , CO_3^{2-} , and NH_4^+ .



DATA INTERPRETATION AND ANALYSIS

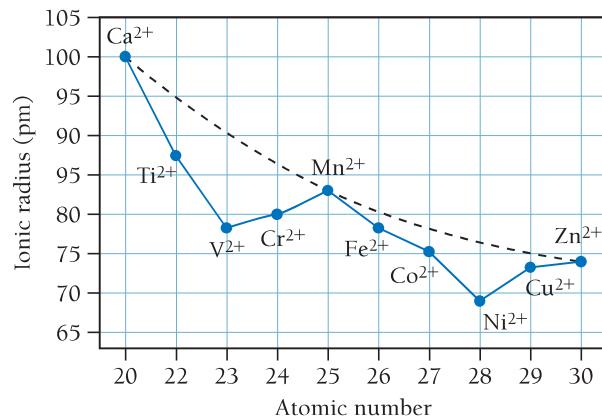
Lattice Energy

135. Evidence for the additional stabilization of certain electron configurations comes from the experimental lattice energies of the metal fluorides, MF_2 . The first figure below plots lattice energy for the 2+ metal cations for period 4 elements. The figure that follows plots the ionic radii of the 2+ metal cations of period 4 elements versus atomic number.



▲ Lattice Energy of Period 4 Metal Fluorides

Source: After C.S.G. Phillips, R. J. P Williams, Inorganic Chemistry, Volume 1, p. 179.



▲ Ionic Radius Versus Atomic Number of 2+ Metal Cations for Period 4 Elements

Use the information provided in the figures to answer the following questions:

- Explain the general trend in lattice energy.
- Is there a correlation between ionic radius and lattice energy? Explain.
- What could account for the decrease in lattice energy between CrF_2 and MnF_2 ?
- Which has the higher lattice energy: VF_2 or VCl_2 ? Explain.



ANSWERS TO CONCEPTUAL CONNECTIONS

Bond Types

10.1 (c) MgF_2 contains a metal (Mg) bonded to a nonmetal (F), so the bonding is ionic.

Lewis Symbols

10.2 (d) This Lewis symbol has four dots, corresponding to silicon's four valence electrons.

Melting Points of Ionic Solids

10.3 (a) You would expect MgO to have the higher melting point because, in our bonding model, the magnesium and oxygen ions are held together in a crystalline lattice by charges of 2+ for magnesium and 2- for oxygen. In contrast, the $NaCl$ lattice is held together by charges of 1+ for sodium and 1- for chlorine. The experimentally measured melting points of these compounds are 801 °C for $NaCl$ and 2852 °C for MgO , in accordance with our model.

Energy and the Octet Rule

10.4 (b) The reasons that atoms form bonds are complex. One contributing factor is the lowering of their potential energy. The octet rule is just a handy way to predict the combinations of atoms that have a lower potential energy when they bond.

Periodic Trends in Electronegativity

10.5 (a) N > P > Al > Na

Percent Ionic Character

10.6 (b) You are given that the dipole moment of the HCl bond is about 1 D and that the bond length is 127 pm. Previously, we calculated the dipole moment for a 130-pm bond that is 100% ionic to be about 6.2 D. You can therefore estimate the bond's ionic character as $1/6 \times 100$, which is closest to 15%.

Resonance Structures

10.7 (b) This structure is not a resonance structure of the ion because it has a different skeletal structure.

Odd-Electron Species

10.8 (d) ClO because the sum of the valence electrons of its atoms is an odd number

Expanded Octets

10.9 (b) The only molecule in this group that could have an expanded octet is H_3PO_4 because phosphorus is a third-period element. Expanded octets *never* occur in second-period elements such as carbon and nitrogen.

Bond Energies and ΔH_{rxn}

10.10 (b) In a highly exothermic reaction, the energy needed to break bonds is less than the energy released when the new bonds form, resulting in a net release of energy.