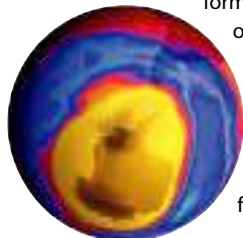


## 8.2 The Nature of Covalent Bonding

### Connecting to Your World

You know that without oxygen to breathe, you could not live. But did you know that oxygen plays another important role in your life? High in the atmosphere, a different form of oxygen, called ozone, forms a layer that filters out harmful radiation from the sun. The colors in this map indicate the concentrations of ozone in various parts of Earth's atmosphere. In this section, you will learn how oxygen atoms can join in pairs to form the oxygen you breathe and can also join in groups of three oxygen atoms to form ozone.

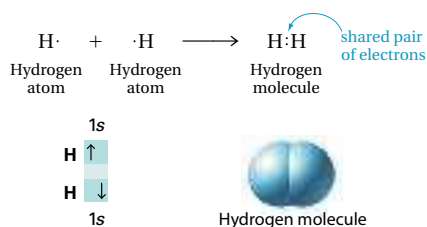


### The Octet Rule in Covalent Bonding

Recall that when ionic compounds form, electrons tend to be transferred so that each ion acquires a noble gas configuration. A similar rule applies for covalent bonds. **In covalent bonds, electron sharing usually occurs so that atoms attain the electron configurations of noble gases.** For example, each hydrogen atom has one electron. But a pair of hydrogen atoms share these two electrons when they form a covalent bond in a hydrogen molecule. Each hydrogen atom thus attains the electron configuration of helium, a noble gas with two electrons. Combinations of atoms of the nonmetals and metalloids in Groups 4A, 5A, 6A, and 7A of the periodic table are likely to form covalent bonds. In this case the atoms usually acquire a total of eight electrons, or an octet, by sharing electrons, so that the octet rule applies.

### Single Covalent Bonds

The hydrogen atoms in a hydrogen molecule are held together mainly by the attraction of the shared electrons to the positive nuclei. Two atoms held together by sharing a pair of electrons are joined by a **single covalent bond**. Hydrogen gas consists of diatomic molecules whose atoms share only one pair of electrons, forming a single covalent bond.



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### Guide for Reading

#### Key Concepts

- What is the result of electron sharing in covalent bonds?
- How do electron dot structures represent shared electrons?
- How do atoms form double or triple covalent bonds?
- How are coordinate covalent bonds different from other covalent bonds?
- How is the strength of a covalent bond related to its bond dissociation energy?
- How are oxygen atoms bonded in ozone?
- What are some exceptions to the octet rule?

#### Vocabulary

single covalent bond  
structural formula  
unshared pair  
double covalent bond  
triple covalent bond  
coordinate covalent bond  
polyatomic ion  
bond dissociation energy  
resonance structure

#### Reading Strategy

**Identifying Main Idea/Details**  
List the main idea in the paragraph following the heading *The Octet Rule in Covalent Bonding*. As you read, list examples of how this rule is followed when a single covalent bond, a double covalent bond, and a triple covalent bond form.

## 8.2

### 1 FOCUS

#### Objectives

- 8.2.1 Describe** how electrons are shared to form covalent bonds and identify exceptions to the octet rule.
- 8.2.2 Demonstrate** how electron dot structures represent shared electrons.
- 8.2.3 Describe** how atoms form double or triple covalent bonds.
- 8.2.4 Distinguish** between a covalent bond and a coordinate covalent bond and describe how the strength of a covalent bond is related to its bond dissociation energy.
- 8.2.5 Describe** how oxygen atoms are bonded in ozone.

### Guide for Reading

#### Build Vocabulary

L2

**Word Parts** The word *structure* comes from the Latin verb *struere*, which means “to build.” A structural formula is one that shows how a molecule is built, that is, how the atoms are joined together by chemical bonds. Ask students to define structural steel. (*Steel is used to form the skeleton of buildings.*)

#### Reading Strategy

L2

#### Directed Reading/Thinking Activity

Have students read the red headings throughout the section first and ask themselves questions about what they might learn when they read the text.

### 2 INSTRUCT

#### Connecting to Your World

Explain that scientists attribute the thinning of the ozone layer, in part, to the action of compounds called chlorofluorocarbons (CFCs), which have been released into the atmosphere. Ask, **Why is ozone important in the atmosphere?** (*It filters out radiation that could harm living beings on Earth.*)



### Section Resources

#### Print

- **Guided Reading and Study Workbook**, Section 8.2
- **Core Teaching Resources**, Section 8.2 Review
- **Probeware Laboratory Manual**, Quick Lab, p. 226
- **Transparencies**, T87–T89

#### Technology

- **Interactive Textbook with ChemASAP**, Problem-Solving 8.8, 8.10; Simulation 6, Assessment 8.2
- **Go Online**, Section 8.2

## The Octet Rule in Covalent Bonding

### Discuss

L2

Write electron configurations for carbon, nitrogen, oxygen, fluorine, and neon on the chalkboard. Ask, **How many electrons would carbon, nitrogen, oxygen, and fluorine need to share in order to achieve the same electron configuration as neon?** (4, 3, 2, and 1 respectively)

### CLASS Activity

### Representing Molecules

L2

**Purpose** Students practice different ways to represent molecules.

**Materials** paper and pencil

**Procedure** Divide students into groups of three or four. Have them practice drawing molecular diagrams, structural formulas, electron-dot structures, and orbital diagrams for molecules such as  $\text{OF}_2$ ,  $\text{SCl}_2$ ,  $\text{N}_2\text{H}_4$ ,  $\text{CCl}_4$ ,  $\text{CHCl}_3$ , and  $\text{C}_2\text{H}_6$ .

## Single Covalent Bonds

### Go Online

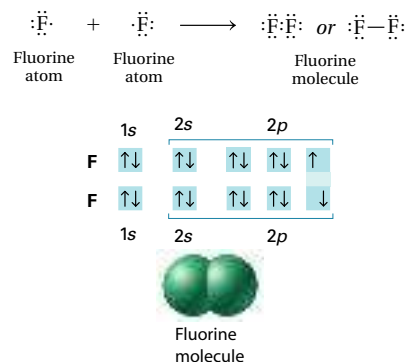
NSTA SciLINKS

Download a worksheet on **Valence Electrons** for students to complete, and find additional teacher support from NSTA SciLinks.

**Go Online**  
NSTA SciLINKS  
**For:** Links on Valence Electrons  
**Visit:** www.SciLinks.org  
**Web Code:** cdn-1082

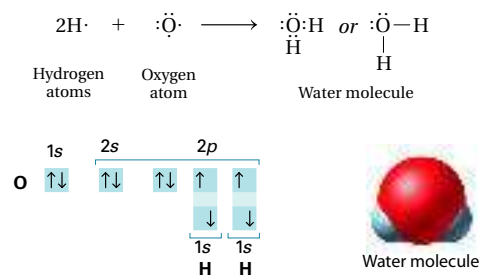
An electron dot structure such as  $\text{H:H}$  represents the shared pair of electrons of the covalent bond by two dots. The pair of shared electrons forming the covalent bond is also often represented as a dash, as in  $\text{H—H}$  for hydrogen. A **structural formula** represents the covalent bonds by dashes and shows the arrangement of covalently bonded atoms. In contrast, the molecular formula of hydrogen,  $\text{H}_2$ , indicates only the number of hydrogen atoms in each molecule.

The halogens also form single covalent bonds in their diatomic molecules. Fluorine is one example. Because a fluorine atom has seven valence electrons, it needs one more to attain the electron configuration of a noble gas. By sharing electrons and forming a single covalent bond, two fluorine atoms each achieve the electron configuration of neon.



In the  $\text{F}_2$  molecule, each fluorine atom contributes one electron to complete the octet. Notice that the two fluorine atoms share only one pair of valence electrons. A pair of valence electrons that is not shared between atoms is called an **unshared pair**, also known as a lone pair or a nonbonding pair.

You can draw electron dot structures for molecules of compounds in much the same way that you draw them for molecules of diatomic elements. Water ( $\text{H}_2\text{O}$ ) is a molecule containing three atoms with two single covalent bonds. Two hydrogen atoms share electrons with one oxygen atom. The hydrogen and oxygen atoms attain noble-gas configurations by sharing electrons. As you can see in the electron dot structures below, the oxygen atom in water has two unshared pairs of valence electrons.



**Checkpoint** What does a structural formula represent?

## Differentiated Instruction

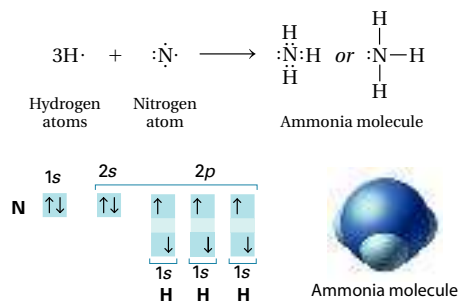
### Special Needs

L1

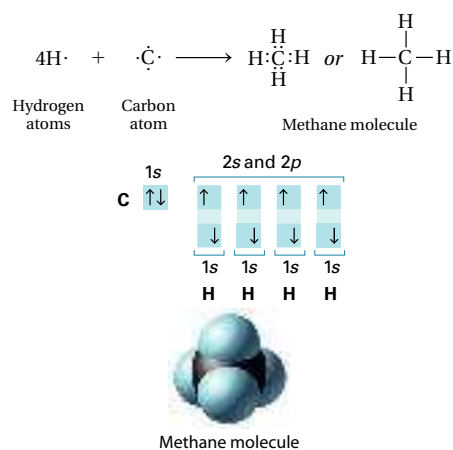
Pair each student with a study partner. Have them use the periodic table and quiz each other on writing electron dot structures for single atoms and bonded atoms. Make sure they understand that the Group number for

any atom, 1A to 7A, indicates the number of valence electrons that atom has, and that it is valence electrons that appear in the electron dot structures.

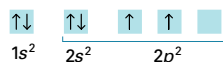
You can draw the electron dot structure for ammonia ( $\text{NH}_3$ ), a suffocating gas, in a similar way. The ammonia molecule has one unshared pair of electrons.



Methane ( $\text{CH}_4$ ) contains four single covalent bonds. The carbon atom has four valence electrons and needs four more valence electrons to attain a noble-gas configuration. Each of the four hydrogen atoms contributes one electron to share with the carbon atom, forming four identical carbon-hydrogen bonds. As you can see in the electron dot structure below, methane has no unshared pairs of electrons.



When carbon forms bonds with other atoms, it usually forms four bonds. You would not predict this based on carbon's electron configuration, shown below.



**Checkpoint** What is the electron dot structure of a methane molecule?

## Discuss

L2

First review the molecular and structural formulas, electron dot structures, and orbital diagrams for fluorine, water, and ammonia molecules. If possible, display physical models. Call attention to the fact that fluorine has one half-filled orbital and forms one bond, oxygen has two half-filled orbitals and forms two bonds, and nitrogen has three and forms three bonds. Tell students carbon has two. Ask, **How many covalent bonds do you think carbon forms?** (Students may logically say two.) Tell students that  $\text{CH}_2$  does not represent a stable molecule, but  $\text{CH}_4$  (methane) is a stable molecule. Explain the concept of electron promotion, which allows carbon to form four single covalent bonds. Point out that elements in groups 3A and 4A promote  $s$  electrons to  $p$  orbitals, increasing their bonding capacity. For example, boron's electron configuration is  $1s^2 2s^2 2p^1$ . Based on this configuration, students might infer that boron can form only one covalent bond. However, the chloride of boron is  $\text{BCl}_3$  rather than  $\text{BCl}$ . The promotion of one  $2s$  electron to the  $2p$  orbital allows for the formation of three bonds. Boron does not achieve a noble-gas configuration, but it does achieve added stability by forming three bonds rather than one.

## Facts and Figures

### Expanding the Octet

Nonmetals in the third row and beyond such as phosphorus, sulfur, and iodine can form more than four bonds because they have empty  $d$  orbitals. Phosphorus, for example, can unpair a  $3s$  electron and promote it to an empty  $3d$  orbital. The promotion allows

phosphorus to form five bonds. The amount of energy needed to promote an electron is less than the energy released with the formation of an extra bond. Sulfur can promote one  $3s$  electron and one  $3p$  electron and form two extra bonds.

### Answers to...

**Checkpoint** It shows the covalent bonds as dashes and shows the arrangement of covalently bonded atoms.



## CLASS Activity

Bonding for Second Row Elements L2

**Purpose** Students gain understanding of covalent bonding and distinguish between covalent and ionic bonding.

**Procedure** Have students draw electron dot structures for each element in the second row of the periodic table: Li, Be, B, C, N, O, and F. Then have them answer the following:

- Predict how many bonds each atom must form to attain a noble-gas configuration. (1, 2, 3, 4, 3, 2, 1)
- Can lithium form a covalent bond and reach stability? (no)
- Which elements can reach stability by forming covalent bonds? (C, N, O, F)
- Can fluorine form an ionic bond? (yes)
- Are the bonds in nitrogen molecules (N<sub>2</sub>) ionic or covalent? (covalent)

## CONCEPTUAL PROBLEM 8.1

## Answers

7. a.  $\text{Cl}:\text{Cl}:$     b.  $\text{Br}:\text{Br}:$     c.  $\text{I}:\text{I}:$   
 8. a.  $\text{H}:\text{O}:\text{O}:\text{H}$     b.  $\text{Cl}:\text{P}(\text{Cl})_3$

Practice Problems Plus L2

The following covalent molecules have only single covalent bonds. Draw an electron dot structure for each.

- a.  $\text{NF}_3$   $\text{F}:\text{N}(\text{F})_3$   
 b.  $\text{SBr}_2$   $\text{Br}:\text{S}(\text{Br})_2$

If you tried to form covalent C—H bonds for methane by combining the two 2*p* electrons of the carbon with two 1*s* electrons of hydrogen atoms, you would incorrectly predict a molecule with the formula CH<sub>2</sub> (instead of CH<sub>4</sub>). The formation of four bonds by carbon can be simply explained. One of carbon's 2*s* electrons is promoted to the vacant 2*p* orbital to form the following electron configuration.



This electron promotion requires only a small amount of energy. The promotion provides four electrons of carbon that are capable of forming covalent bonds with four hydrogen atoms. Methane, the carbon compound formed by electron sharing of carbon with four hydrogen atoms, is much more stable than CH<sub>2</sub>. The stability of the resulting methane more than compensates for the small energy cost of the electron promotion. Therefore, formation of methane (CH<sub>4</sub>) is more energetically favored than the formation of CH<sub>2</sub>.

## CONCEPTUAL PROBLEM 8.1

## Drawing an Electron Dot Structure

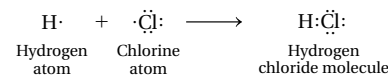
Hydrochloric acid (HCl (aq)) is prepared by dissolving gaseous hydrogen chloride (HCl (g)) in water. Hydrogen chloride is a diatomic molecule with a single covalent bond. Draw the electron dot structure for HCl.



## 1 Analyze Identify the relevant concepts.

In a single covalent bond, a hydrogen and a chlorine atom must share a pair of electrons. Each must contribute one electron to the bond. First, draw the electron dot structures for the two atoms. Then show the electron sharing in the compound they produce.

## 2 Solve Apply concepts to the situation.



In the electron dot structures, the hydrogen atom and the chlorine atom are each correctly shown to have an unpaired electron. Through electron sharing, the hydrogen and chlorine atoms are shown to attain the electron configurations of the noble gases helium and argon, respectively.

## Practice Problems

7. Draw electron dot structures for each molecule.  
 a. chlorine    b. bromine    c. iodine  
 8. The following molecules have single covalent bonds. Draw an electron dot structure for each.  
 a. H<sub>2</sub>O<sub>2</sub>    b. PCl<sub>3</sub>



**Problem Solving 8.8** Solve Problem 8 with the help of an interactive guided tutorial.

with ChemASAP

## Facts and Figures


## Inventing Electron Dot Structures

Gilbert Newton Lewis (1875–1946) was an American chemist who invented electron dot structures, which are often called Lewis structures or diagrams in his honor. These structures supported Lewis's theory of the electron pair in chemical bonding. As a professor of physical chemistry, he expanded

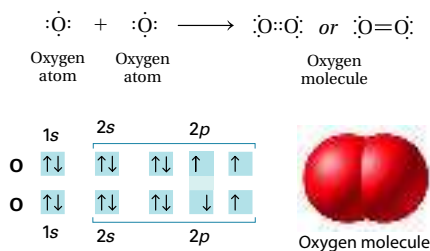
the theory of acids and bases by defining an acid as an electron pair acceptor and a base as an electron pair donor. The definitions encompass all Brønsted-Lowry acid-base reactions and include many others not previously categorized as acid-base reactions.



## Double and Triple Covalent Bonds

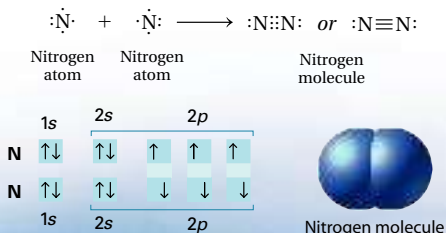
Sometimes atoms bond by sharing more than one pair of electrons.  **Atoms form double or triple covalent bonds if they can attain a noble gas structure by sharing two pairs or three pairs of electrons.** A bond that involves two shared pairs of electrons is a **double covalent bond**. A bond formed by sharing three pairs of electrons is a **triple covalent bond**.

You might think that an oxygen atom, with six valence electrons, would form a double bond by sharing two of its electrons with another oxygen atom.



In such an arrangement, all the electrons within the molecule would be paired. Experimental evidence, however, indicates that two of the electrons in  $\text{O}_2$  are still unpaired. Thus, the bonding in the oxygen molecule ( $\text{O}_2$ ) does not obey the octet rule. You cannot draw an electron dot structure that adequately describes the bonding in the oxygen molecule.

An element whose molecules contain triple bonds is nitrogen ( $\text{N}_2$ ), a major component of Earth's atmosphere illustrated in Figure 8.7. In the nitrogen molecule, each nitrogen atom has one unshared pair of electrons. A single nitrogen atom has five valence electrons. Each nitrogen atom in the nitrogen molecule must gain three electrons to have the electron configuration of neon.



### Interactive Textbook

**Simulation 6** Simulate the covalent bonding between molecules.

with ChemASAP

## Double and Triple Covalent Bonds

### Discuss

L2

To introduce the discussion of multiple covalent bonds, use the electron dot structure for the nitrogen molecule. Ask students to show how the structure of diatomic nitrogen can satisfy the octet rule. (*The nitrogen atoms can share six electrons.*) Have them compare the bonding in ammonia with the bonding in nitrogen gas. Then introduce the oxygen molecule. Ask students to draw a structure that obeys the octet rule. Ask, **Why doesn't oxygen form a triple bond?** (*Each oxygen needs to share only two electrons to achieve a stable electron configuration.*) Explain that although a double bond in the oxygen molecule fulfills the octet rule, it does not fit with experimental evidence that shows that the oxygen molecule contains two unpaired electrons. Thus the structure of  $\text{O}_2$  is an exception to the octet rule. Help students draw the electron dot structure and orbital diagram for carbon dioxide. Ask, **What type of bonds does carbon form with the two oxygen atoms in  $\text{CO}_2$ ?** (*double covalent bonds*) Note that carbon can form single, double, and triple bonds, but a quadruple bond is impossible because of geometric restrictions. Have students draw diagrams for hydrogen cyanide (HCN) and formaldehyde ( $\text{H}_2\text{CO}$ ). Ask, **What kind of bonds does carbon form in each of these molecules?** (*HCN: one single carbon-to-hydrogen bond and one triple carbon-to-nitrogen bond;  $\text{H}_2\text{CO}$ : two single carbon-to-hydrogen bonds and one double carbon-to-oxygen bond*) As the discussion proceeds, provide models for the molecules discussed.

**Figure 8.7** Oxygen and nitrogen are the main components of Earth's atmosphere. The oxygen molecule is an exception to the octet rule. It has two unpaired electrons. Three pairs of electrons are shared in a nitrogen molecule.

## Differentiated Instruction

### Less Proficient Reader

L1

Divide students into groups of three or four. Have them draw electron dot formulas for oxygen and fluorine, then combine the drawings into the electron dot formula for

$\text{OF}_2$ . Have them do the same for  $\text{SCl}_2$  and  $\text{CCl}_4$ . When students have mastered these, continue the exercise with  $\text{CHCl}_3$ ,  $\text{N}_2\text{H}_4$ , and  $\text{C}_2\text{H}_6$ .

## Relate

L1

Earth's atmosphere is approximately 80 percent nitrogen gas, but surprisingly few nitrogen compounds exist compared with the numerous compounds of oxygen, which constitutes only 20 percent of the atmosphere. Students may correctly surmise that the triple bond in  $N_2$  is harder to break than a double bond and considerably harder to break than a single bond. Thus,  $N_2$  is a stable molecule. Plant and animal life depends on nitrogen, but in order to be usable to living systems, the element must be converted to a compound, a process called nitrogen fixing. Nitrogen fixing occurs naturally when lightning provides the energy for atmospheric nitrogen to react with oxygen to form nitrogen oxides. The oxides dissolve in rain and fall to the ground where they can be utilized by plants. Nitrogen fixing bacteria in the soil are also able to convert atmospheric nitrogen to usable compounds.

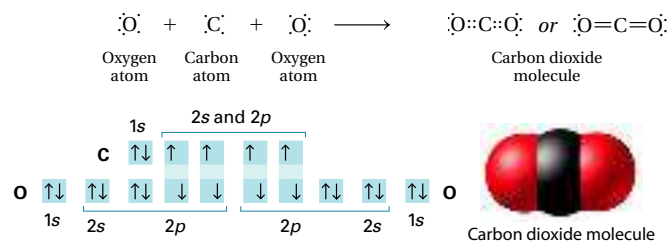
Table 8.1

| The Diatomic Elements |                  |                            |   |
|-----------------------|------------------|----------------------------|---|
| Name                  | Chemical formula | Electron dot structure     | Properties and uses   |
| Fluorine              | $F_2$            | $:\ddot{F}-\ddot{F}:$      | Greenish-yellow reactive toxic gas. Compounds of fluorine, a halogen, are added to drinking water and toothpaste to promote healthy teeth.    |
| Chlorine              | $Cl_2$           | $:\ddot{Cl}-\ddot{Cl}:$    | Greenish-yellow reactive toxic gas. Chlorine is a halogen used in household bleaching agents.   |
| Bromine               | $Br_2$           | $:\ddot{Br}-\ddot{Br}:$    | Dense red-brown liquid with pungent odor. Compounds of bromine, a halogen, are used in the preparation of photographic emulsions.             |
| Iodine                | $I_2$            | $:\ddot{I}-\ddot{I}:$      | Dense gray-black solid that produces purple vapors; a halogen. A solution of iodine in alcohol (tincture of iodine) is used as an antiseptic. |
| Hydrogen              | $H_2$            | $H-H$                      | Colorless, odorless, tasteless gas. Hydrogen is the lightest known element.   |
| Nitrogen              | $N_2$            | $:\text{N}\equiv\text{N}:$ | Colorless, odorless, tasteless gas. Air is almost 80% nitrogen by volume.   |
| Oxygen                | $O_2$            | Inadequate                 | Colorless, odorless, tasteless gas that is vital for life. Air is about 20% oxygen by volume.   |



**Figure 8.8** Carbon dioxide gas is soluble in water and is used to carbonate many beverages. A carbon dioxide molecule has two carbon–oxygen double bonds.

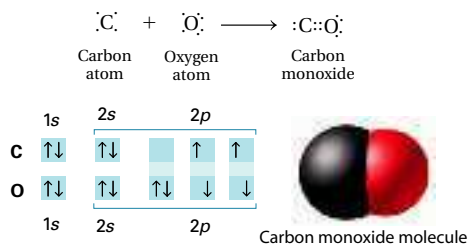
Up to this point in your textbook, the examples of single and triple covalent bonds have involved diatomic molecules. Table 8.1 lists the properties and uses of the elements that exist as diatomic molecules. Single, double, and triple covalent bonds can also exist between unlike atoms. For example, consider carbon dioxide ( $CO_2$ ), which is present in the atmosphere and is used to carbonate many soft drinks as shown in Figure 8.8. The carbon dioxide molecule contains two oxygens, each of which shares two electrons with carbon to form a total of two carbon–oxygen double bonds.



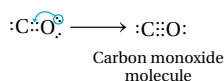
The two double bonds in the carbon dioxide molecule are identical to each other. Carbon dioxide is an example of a triatomic molecule, which is a molecule consisting of three atoms.

## Coordinate Covalent Bonds

Carbon monoxide (CO) is an example of a type of covalent bonding different from that seen in water, ammonia, methane, and carbon dioxide. A carbon atom needs to gain four electrons to attain the electron configuration of neon. An oxygen atom needs two electrons. Yet it is possible for both atoms to achieve noble-gas electron configurations by a type of bonding called coordinate covalent bonding. To see how, begin by looking at the double covalent bond between carbon and oxygen.

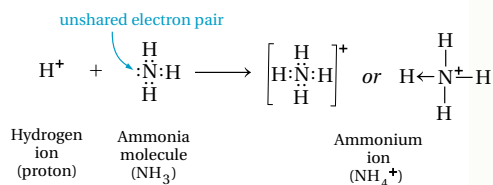


With the double bond in place, the oxygen has a stable configuration but the carbon does not. As shown below, the dilemma is solved if the oxygen also donates one of its unshared pairs of electrons for bonding.



A **coordinate covalent bond** is a covalent bond in which one atom contributes both bonding electrons. In a structural formula, you can show coordinate covalent bonds as arrows that point from the atom donating the pair of electrons to the atom receiving them. The structural formula of carbon monoxide, with two covalent bonds and one coordinate covalent bond, is  $\text{C}\equiv\text{O}$ . **In a coordinate covalent bond, the shared electron pair comes from one of the bonding atoms.** Once formed, a coordinate covalent bond is like any other covalent bond.

The ammonium ion ( $\text{NH}_4^+$ ) consists of atoms joined by covalent bonds, including a coordinate covalent bond. A **polyatomic ion**, such as  $\text{NH}_4^+$ , is a tightly bound group of atoms that has a positive or negative charge and behaves as a unit. The ammonium ion forms when a positively charged hydrogen ion ( $\text{H}^+$ ) attaches to the unshared electron pair of an ammonia molecule ( $\text{NH}_3$ ). Most plants need nitrogen that is already combined in a compound rather than molecular nitrogen ( $\text{N}_2$ ) to grow. The fertilizer shown in Figure 8.9 contains the nitrogen compound ammonium sulfate.



**Checkpoint** What is a polyatomic ion?

**Figure 8.9** The polyatomic ammonium ion ( $\text{NH}_4^+$ ), present in ammonium sulfate, is an important component of fertilizer for field crops, home gardens, and potted plants.



## Coordinate Covalent Bonds

### Discuss

L2

A coordinate covalent bond is an exception to the rule that covalent bonding occurs between two half-empty orbitals of two atoms. In a coordinate covalent bond, one atom has an empty orbital and the other an orbital filled with an electron pair that is not yet involved in a chemical bond. Note that the bonding still involves only one pair of electrons and one pair of orbitals but one atom provides both of the shared electrons.

### Discuss

L2

Ask students to compare the electron dot structure for ammonia ( $\text{NH}_3$ ) and for a hydrogen ion ( $\text{H}^+$ ). After forming the three nitrogen-to-hydrogen bonds, nitrogen has an unshared pair of electrons. The hydrogen ion has no electrons available for bonding. Ask, **How do you explain the existence of the ammonium ion,  $\text{NH}_4^+$ ?** (*The empty orbital of the  $\text{H}^+$  must overlap the filled orbital in ammonia. The electrons from nitrogen are attracted to both the nitrogen nucleus and the hydrogen nucleus. A bond forms when electrons are simultaneously attracted to two nuclei. This type of bond is called a coordinate covalent bond.*) Ask students to explain the bonding in  $\text{CO}_2$  and then explain the bonding in CO using coordinate covalent bonding.

### Answers to...

**Checkpoint** a group of atoms that has a positive or negative charge and behaves as a unit

## Section 8.2 (continued)

### Discuss

L2

Have students write the electron dot structure for  $\text{SO}_2$ . Emphasize that the structure should satisfy the bonding requirements of all three atoms. Students should find that, to satisfy the octet rule for all the atoms, they must write a structure in which one oxygen atom is double bonded to sulfur. The other oxygen is single bonded by a coordinate covalent bond in which the electrons are donated by sulfur. Point out that experimental evidence indicates that both sulfur-oxygen bonds are identical. Explain that this evidence indicates that the bonding in  $\text{SO}_2$  must be some intermediate between a single and double bond. Ask, **How does the formation of a coordinate covalent bond differ from that of a covalent bond?** (In a covalent bond, each atom provides one electron, in a coordinate covalent bond, both electrons are provided by the same atom.)

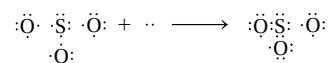
### Relate

L2

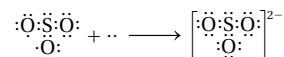
Have students think of everyday examples in which the shape of an object is as important as its composition. For example, several keys might be made of the same metal, but only one will fit into a particular lock.

Most polyatomic cations and anions contain both covalent and coordinate covalent bonds. Therefore, compounds containing polyatomic ions include both ionic and covalent bonding.

As another example, draw the electron dot structure of the polyatomic ion  $\text{SO}_3^{2-}$ . Start by drawing the electron dot structures for the oxygen and sulfur atoms, and the two extra electrons indicated by the charge. Then join two of the oxygens to sulfur by single covalent bonds.



Then join the remaining oxygen by a coordinate covalent bond, with sulfur donating one of its unshared pairs to oxygen, and add the two extra electrons. Put brackets about the structure and indicate the 2- charge, giving the result shown.



Each of the atoms of the completed structure has eight valence electrons, satisfying the octet rule. Without the two extra electrons, two of the oxygens would be electron-deficient.

Table 8.2 lists electron dot structures of some common compounds with covalent bonds.

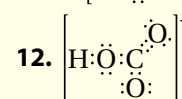
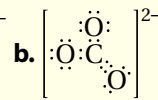
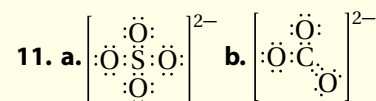
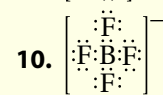
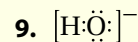
Table 8.2

Some Common Molecular Compounds

| Name              | Chemical formula       | Structure   | Properties and uses   |
|-------------------|------------------------|---|---|
| Hydrogen peroxide | $\text{H}_2\text{O}_2$ | $\begin{array}{c} \text{H} \\ \diagdown \\ \text{O} \\ \diagup \\ \text{H} \end{array}$   | Colorless, unstable liquid when pure. It is used as rocket fuel. A 3% solution is used as a bleach and antiseptic.  |
| Sulfur dioxide    | $\text{SO}_2$          | $\begin{array}{c} \cdot\ddot{\text{O}}=\text{S}\cdot \\ \diagdown \\ \cdot\ddot{\text{O}}\cdot \end{array}$   | Oxides of sulfur are produced in combustion of petroleum products and coal. They are major air pollutants in industrial areas. Oxides of sulfur can lead to respiratory problems.   |
| Sulfur trioxide   | $\text{SO}_3$          | $\begin{array}{c} \cdot\ddot{\text{O}}\cdot \\ \diagup \\ \cdot\ddot{\text{O}}=\text{S}\cdot \\ \diagdown \\ \cdot\ddot{\text{O}}\cdot \end{array}$ |   |
| Nitric oxide      | $\text{NO}$            | $\cdot\ddot{\text{O}}=\ddot{\text{N}}\cdot$   | Oxides of nitrogen are major air pollutants produced by the combustion of fossil fuels in automobile engines. They irritate the eyes, throat, and lungs.                            |
| Nitrogen dioxide  | $\text{NO}_2$          | $\begin{array}{c} \cdot\ddot{\text{O}}=\text{N}\cdot \\ \diagdown \\ \cdot\ddot{\text{O}}\cdot \end{array}$   |   |
| Nitrous oxide     | $\text{N}_2\text{O}$   | $\cdot\ddot{\text{O}}\leftarrow\text{N}\equiv\text{N}\cdot$   | Colorless, sweet-smelling gas. It is used as an anesthetic commonly called laughing gas.  |
| Hydrogen cyanide  | $\text{HCN}$           | $\text{H}-\text{C}\equiv\text{N}\cdot$  | Colorless, toxic gas with the smell of almonds.   |
| Hydrogen fluoride | $\text{HF}$            | $\text{H}-\ddot{\text{F}}\cdot$   | Two hydrogen halides, all extremely soluble in water. Hydrogen chloride, a colorless gas with pungent odor, readily dissolves in water to give a solution called hydrochloric acid. |
| Hydrogen chloride | $\text{HCl}$           | $\text{H}-\ddot{\text{Cl}}\cdot$  |   |

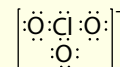


## Answers



## Practice Problems Plus

L2

 Draw the electron dot structure of the polyatomic chlorate anion,  $\text{ClO}_3^{-}$ .


## TEACHER Demo

## Bond Energies

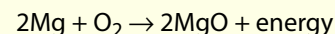
L2

**Purpose** Students observe two reactions in which bonds are formed and compare ionic and covalent bond energies.

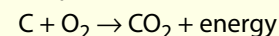
**Materials** 10- to 15- cm piece of clean magnesium ribbon, a small piece of charcoal, tongs, Bunsen burner, filters for viewing

**Safety** Wear safety goggles and lab apron.

**Procedure** Using tongs, hold the piece of magnesium in a Bunsen burner. **CAUTION!** Tie back long hair and loose clothing and wear safety goggles. Do not look directly at the flame. Observe through filters. Discuss with students the large amount of heat and light given off in the formation of  $\text{MgO}$ . Write the balanced equation on the board.



Using tongs, place a small piece of charcoal in the Bunsen burner flame and try to ignite it. Write the balanced chemical equation.



**Expected Outcomes** Students note that much less energy is given off in forming  $\text{CO}_2$  than in forming  $\text{MgO}$ . Ask, **What kind of bonds are in  $\text{MgO}$ ?** (ionic) **What kind of bonds are in  $\text{CO}_2$ ?** (covalent) Ionic bond energies are, in general, greater than covalent bond energies.

Remember, the electron dot structure for a neutral molecule contains the same number of electrons as the total number of valence electrons in the combining atoms. The negative charge of a polyatomic ion shows the number of electrons in addition to the valence electrons of the atoms present. Because a negatively charged polyatomic ion is part of an ionic compound, the positive charge of the cation of the compound balances these additional electrons.

## CONCEPTUAL PROBLEM 8.2

## Drawing the Electron Dot Structure of a Polyatomic Ion

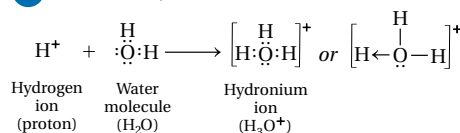
The polyatomic hydronium ion ( $\text{H}_3\text{O}^+$ ), which is found in acidic mixtures such as lemon juice, contains a coordinate covalent bond. The  $\text{H}_3\text{O}^+$  ion forms when a hydrogen ion is attracted to an unshared electron pair in a water molecule. Draw the electron dot structure of the hydronium ion.



## 1 Analyze Identify the relevant concepts.

$\text{H}_3\text{O}^+$  forms by the addition of a hydrogen ion to a water molecule. Draw the electron dot structure of the water molecule. Then, add the hydrogen ion. Oxygen must share a pair of electrons with the added hydrogen ion to form a coordinate covalent bond.

## 2 Solve Apply the concepts to this situation.



The oxygen in the hydronium ion has eight valence electrons, and each hydrogen shares two valence electrons. This satisfies the needs of both hydrogen and oxygen for valence electrons. The water molecule is electrically neutral, and the hydrogen ion has a positive charge. The combination of these two species must have a charge of  $1+$ , as is found in the hydronium ion.

## Practice Problems

- Draw the electron dot structure of the hydroxide ion ( $\text{OH}^-$ ).
- Draw the electron dot structure of the polyatomic boron tetrafluoride anion ( $\text{BF}_4^-$ ).
- Draw the electron dot structures for sulfate ( $\text{SO}_4^{2-}$ ) and carbonate ( $\text{CO}_3^{2-}$ ). Sulfur and carbon are the central atoms, respectively.
- Draw the electron dot structure for the hydrogen carbonate ion ( $\text{HCO}_3^-$ ). Carbon is the central atom, and hydrogen is attached to oxygen in this polyatomic anion.



**Problem-Solving 8.10** Solve Problem 10 with the help of an interactive guided tutorial.

with ChemASAP

## Bond Dissociation Energies

### Quick LAB

### Strengths of Covalent Bonds L2

**Objective** After completing this activity, students will understand that the dissociation energy of a covalent bond increases in order from single bond to double bond to triple bond.

**Skills Focus** Solving, interpreting

 **Prep Time** 10 minutes

**Advance Prep** Obtain the necessary materials in advance. You may wish to ask students to help with this task by bringing cans of food, coat hangers, and plastic grocery bags to class.

**Class Time** 30 minutes

### Teaching Tips

- Use tape to secure the hanger in place if necessary.
- Have groups of students record the data and then create a graph as a class using an overhead projector.

**Expected Outcome** As the mass of the load increases, the stretch of the rubber band or bands increases. For a given mass, a single rubber band stretches farther than a double and a double rubber band stretches farther than a triple.

### Analyze and Conclude

1. Triple covalent bonds are stronger than double covalent bonds, which are stronger than single covalent bonds.
2. The change in bond dissociation energies in going from a carbon-carbon single bond to a carbon-carbon double bond to a carbon-carbon triple bond is nearly constant. The change in length of one, two, and three rubber bands, as given by the slopes of the lines, is not constant. It is large going from one to two rubber bands and small going from two to three rubber bands.

### Strengths of Covalent Bonds


#### Purpose

To compare and contrast the stretching of rubber bands and the dissociation energy of covalent bonds.

#### Materials

- 1 170-g (6-oz) can of food
- 2 454-g (16-oz) cans of food
- 3 No. 25 rubber bands
- metric ruler
- coat hanger
- plastic grocery bag
- paper clip
- graph paper
- pencil
- motion detector (optional)

#### Procedure

 Sensor version available in the *Probeware Lab Manual*.

1. Bend the coat hanger to fit over the top of a door. The hook should hang down on one side of the door. Measure the length of the rubber bands (in cm). Hang a rubber band on the hook created by the coat hanger.
2. Place the 170-g can in the plastic bag. Use the paper clip to fasten the bag to the end of the rubber band. Lower the bag gently until it is suspended from the end of the rubber band. Measure and record the length of the stretched rubber band. Using different combinations of food cans, repeat this process three times with the following masses: 454 g, 624 g, and 908 g.
3. Repeat Step 2, first using two rubber bands to connect the hanger and the paper clip, and then using three.
4. Graph the length difference: (stretched rubber band) – (unstretched rubber band) on the y-axis versus mass (kg) on the x-axis for one, two, and three rubber bands. Draw the straight line that




you estimate best fits the points for each set of data. (Your graph should have three separate lines.) The x-axis and y-axis intercepts of the lines should pass through zero, and the lines should extend past 1 kg on the x-axis. Determine the slope of each line in cm/kg.

#### Analyze and Conclude

1. Assuming the rubber bands are models for covalent bonds, what can you conclude about the relative strengths of single, double, and triple bonds?
2. How does the behavior of the rubber bands differ from that of covalent bonds?

## Bond Dissociation Energies

A large quantity of heat is released when hydrogen atoms combine to form hydrogen molecules. This suggests that the product is more stable than the reactants. The covalent bond in the hydrogen molecule ( $H_2$ ) is so strong that it would take 435 kJ of energy to break apart all of the bonds in 1 mole (about 2 grams) of  $H_2$ . (You will study the mole, abbreviated mol, in Chapter 12.) The energy required to break the bond between two covalently bonded atoms is known as the **bond dissociation energy**. This is usually expressed as the energy needed to break one mole of bonds, or  $6.02 \times 10^{23}$  bonds. The bond dissociation energy for the  $H_2$  molecule is 435 kJ/mol.

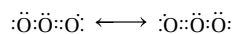
 **A large bond dissociation energy corresponds to a strong covalent bond.** A typical carbon-carbon single bond has a bond dissociation energy of 347 kJ/mol. Typical carbon-carbon double and triple bonds have bond dissociation energies of 657 kJ/mol and 908 kJ/mol, respectively. Strong carbon-carbon bonds help explain the stability of carbon compounds. Compounds with only C—C and C—H single covalent bonds, such as methane, tend to be quite unreactive. They are unreactive partly because the dissociation energy for each of these bonds is high.

### For Enrichment L3

Have students draw a graph of actual bond dissociation energies versus number of bonds for carbon-carbon single, double, and triple bonds. Have students write a statement that relates this graph to their conclusions in the activity.

## Resonance

Ozone in the upper atmosphere blocks harmful ultraviolet radiation from the sun. At the lower elevations shown in Figure 8.10, it contributes to smog. The ozone molecule has two possible electron dot structures.



Notice that the structure on the left can be converted to the one on the right by shifting electron pairs without changing the positions of the oxygen atoms.

As drawn, these electron dot structures suggest that the bonding in ozone consists of one single coordinate covalent bond and one double covalent bond. Because earlier chemists imagined that the electron pairs rapidly flip back and forth, or resonate, between the different electron dot structures, they used double-headed arrows to indicate that two or more structures are in resonance.

Double covalent bonds are usually shorter than single bonds, so it was believed that the bond lengths in ozone were unequal. Experimental measurements show, however, that this is not the case. The two bonds in ozone are the same length. This result can be explained if you assume that the actual bonding in the ozone molecule is the average of the two electron dot structures. The electron pairs do not actually resonate back and forth.

**Checkmark** **The actual bonding of oxygen atoms in ozone is a hybrid, or mixture, of the extremes represented by the resonance forms.**

The two electron dot structures for ozone are examples of what are still referred to as resonance structures. A **resonance structure** is a structure that occurs when it is possible to draw two or more valid electron dot structures that have the same number of electron pairs for a molecule or ion. Resonance structures are simply a way to envision the bonding in certain molecules. Although no back-and-forth changes occur, double-headed arrows are used to connect resonance structures.

**Checkmark** **Checkpoint** *What notation is used to show that the two covalent bonds in  $\text{O}_3$  are the same?*



**Figure 8.10** Although ozone high above the ground forms a protective layer that absorbs ultraviolet radiation from the sun, at lower elevations ozone is a pollutant that contributes to smog.

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## Resonance

### Discuss

L2

Although the electron dot structures of  $\text{O}_2$  on page 221 show that all electrons are paired, experimental evidence reveals that two electrons are unpaired. Ask, **Would you consider oxygen a reactive gas? (yes) What could account for oxygen's reactivity? (Unpaired electrons are reactive.)** After a brief discussion of the other exceptions to the octet rule that are presented in this section, remind students that many bonds do follow the octet rule.

### Relate

L2

Students may be interested in knowing that substances containing unpaired electrons can be identified through a phenomenon called *paramagnetism*. When molecules with unpaired electrons are placed in a magnetic field, they tend to be drawn into the field. These substances are paramagnetic. In contrast, molecules in which all electrons are paired tend to be pushed from a magnetic field. These substances are *diamagnetic*.

Paramagnetism should not be confused with ferromagnetism, which is the familiar attraction of metals such as iron, cobalt and nickel to a magnetic field. The property of paramagnetism is the evidence that shows that the oxygen molecule has unpaired and thus cannot be described exactly by application of the octet rule. However, the oxygen bond does have double bond character; its bond length and bond energy are similar to those of double bonds in other molecules that conform to the octet rule.

### Answers to...

**Checkmark** **Checkpoint** double-headed arrows

## Exceptions to the Octet Rule

### TEACHER Demo

#### A Resonance Hybrid

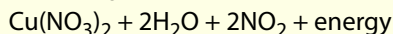
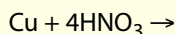
L2

**Purpose** Students observe the formation of nitrogen dioxide and write its resonance structures.

**Materials** small piece of copper metal, concentrated nitric acid, evaporating dish, fume hood

**Safety** Wear safety goggles, gloves, and lab apron. Perform the experiment under an efficient hood.


**Procedure** Place a small piece of copper in an evaporating dish and cover it with concentrated  $\text{HNO}_3$ . While students observe the reaction, write the balanced equation for the reaction on the board.

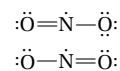


Tell students that  $\text{NO}_2$  is one of the pollutants in automobile exhaust. It gives smog its reddish-brown color and is very reactive and poisonous. Have students try to write the electron dot structure for  $\text{NO}_2$ . They will be unable to find a way to arrange the 17 electrons around the central nitrogen and the two oxygen atoms so that the octet rule is obeyed. Refer students to the two resonance structures on this page, each with an unpaired electron on the nitrogen atom. Ask them to draw another plausible resonance structure. (They could draw a structure with the unpaired electron on an oxygen atom.)

**Expected Outcome** A reddish-brown gas is produced.

## Exceptions to the Octet Rule

The octet rule provides guidance for drawing electron dot structures. For some molecules or ions, however, it is impossible to draw structures that satisfy the octet rule.  **The octet rule cannot be satisfied in molecules whose total number of valence electrons is an odd number. There are also molecules in which an atom has fewer, or more, than a complete octet of valence electrons.** The nitrogen dioxide ( $\text{NO}_2$ ) molecule, for example, contains a total of seventeen, an odd number, of valence electrons. Each oxygen contributes six electrons and the nitrogen contributes five. Two plausible electron dot structures can be drawn for the  $\text{NO}_2$  molecule.



|   | 1s | 2s | 2p |   |    |
|---|----|----|----|---|----|
| O | ↑↓ | ↑↓ | ↑↓ | ↑ | ↑  |
| N | ↑↓ | ↑↓ | ↓  | ↓ | ↓  |
| O | ↑↓ | ↑↓ | ↑  | ↑ | ↑↓ |
|   | 1s | 2s | 2p |   |    |

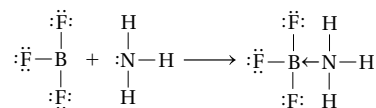


Nitrogen dioxide molecule

An unpaired electron is present in each of these structures, both of which fail to follow the octet rule. It is impossible to draw an electron dot structure for  $\text{NO}_2$  that satisfies the octet rule for all atoms. Yet,  $\text{NO}_2$  does exist as a stable molecule. In fact, it is produced naturally by lightning strikes of the sort shown in Figure 8.11.

A number of other molecules also have an odd number of electrons. In these molecules, as in  $\text{NO}_2$ , complete pairing of electrons is not possible. It is not possible to draw an electron dot structure that satisfies the octet rule. Examples of such molecules include chlorine dioxide ( $\text{ClO}_2$ ) and nitric oxide ( $\text{NO}$ ).

Several molecules with an even number of valence electrons, such as some compounds of boron, also fail to follow the octet rule. This may happen because an atom acquires less than an octet of eight electrons. The boron atom in boron trifluoride ( $\text{BF}_3$ ), for example, is deficient by two electrons, and therefore is an exception to the octet rule. Boron trifluoride readily reacts with ammonia to make the compound  $\text{BF}_3 \cdot \text{NH}_3$ . In doing so, the boron atom accepts the unshared electron pair from ammonia and completes the octet.



 **Checkpoint** Give two examples of exceptions to the octet rule.

**Figure 8.11** Lightning is one means by which nitrogen and oxygen in the atmosphere produce nitrogen dioxide.



## Facts and Figures

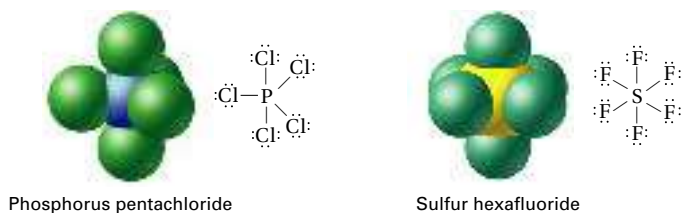
### Free Radicals

Nitrogen dioxide has an odd number of valence electrons (17), so one electron must be unpaired. Molecules with unpaired electrons are called free radicals and tend to be reactive. Two resonance structures for  $\text{NO}_2$  appear on this page. These have the unpaired electron on the nitrogen atom. Two other possible structures place the unpaired

electron on oxygen atoms. When free radicals interact, they share their unpaired electrons and create a dimer. Thus, dinitrogen tetroxide,  $\text{N}_2\text{O}_4$  consists of two nitrogen dioxide molecules joined by a N-N covalent bond. The bond is temperature dependent; at higher temperatures the bond breaks as shown in an equilibrium diagram on p. 557.



**Figure 8.12** Phosphorus pentachloride, used as a chlorinating and dehydrating agent, and sulfur hexafluoride, used as an insulator for electrical equipment, are exceptions to the octet rule. **Interpreting Diagrams** How many valence electrons does the sulfur in sulfur hexafluoride ( $SF_6$ ) have for the structure shown in the figure?



A few atoms, especially phosphorus and sulfur, sometimes expand the octet to include ten or twelve electrons. Consider phosphorus trichloride ( $PCl_3$ ) and phosphorus pentachloride ( $PCl_5$ ). Both are stable compounds in which all of the chlorines are bonded to a single phosphorus atom. Covalent bonding in  $PCl_3$  follows the octet rule because all the atoms have eight valence electrons. However, as shown in Figure 8.12, the electron dot structure for  $PCl_5$  can be written so that phosphorus has ten valence electrons.

## 8.2 Section Assessment

- Key Concept** What electron configurations do atoms usually achieve by sharing electrons to form covalent bonds?
  - Key Concept** How is an electron dot structure used to represent a covalent bond?
  - Key Concept** When are two atoms likely to form a double bond between them? A triple bond?
  - Key Concept** How is a coordinate covalent bond different from other covalent bonds?
  - Key Concept** How is the strength of a covalent bond related to its bond dissociation energy?
  - Key Concept** Draw the electron dot resonance structures for ozone and explain how they describe its bonding.
  - Key Concept** List three ways in which the octet rule can sometimes fail to be obeyed.
20. What kinds of information does a structural formula reveal about the compound it represents?
21. Draw electron dot structures for the following molecules, which have only single covalent bonds.
- $H_2S$
  - $PH_3$
  - $ClF$

22. Use the bond dissociation energies of  $H_2$  (435 kJ/mol) and of a typical carbon-carbon bond (347 kJ/mol) to decide which bond is stronger. Explain your reasoning.

### Elements Handbook

**Ozone** Read the feature on ozone on page R31 of the Elements Handbook. Describe the effect of CFCs on the ozone layer. Explain why the United States has banned the use of CFCs in aerosols.

### Interactive Textbook

**Assessment 8.2** Test yourself on the important concepts of Section 8.2.

with ChemASAP

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## ASSESS

### Evaluate Understanding L2

Ask student to write electron dot structures for  $SBr_2$  and  $C_3H_8$ . Have pairs of students decide whether their structures are correct and work together to correct any errors. Have students define a coordinate covalent bond in their own words. (A coordinate covalent bond is formed when one atom contributes both bonding electrons.) Ask students to explain resonance structures. (Resonance structures occur when two or more valid electron dot structures can be written for a molecule.)

### Elements Handbook

CFCs react with ozone, thinning the ozone layer. Banning the use of CFCs should help decrease the amount of CFCs reaching Earth's ozone layer.

### Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 8.2.

with ChemASAP

### Answers to...

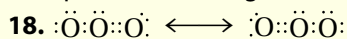
**Checkpoint**  $NO_2, BF_3$

**Figure 8.12** In sulfur hexafluoride, the sulfur atom must have twelve valence electrons.

## Section 8.2 Assessment

- the configurations of noble gases
- Two dots represent each covalent bond.
- when they can attain a noble-gas structure by sharing two pairs or three pairs of electrons
- The shared electron pair comes from one of the bonding atoms. In other covalent bonds each bonding atom provides an electron.
- A large bond dissociation energy corre-

sponds to a strong covalent bond.

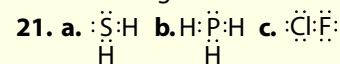


The actual bonding of oxygen atoms in ozone is a hybrid, or mixture, of the extremes represented by the resonance forms.

19. The octet rule cannot be satisfied in molecules whose total number of valence electrons is an odd number.

There are also molecules in which an atom has fewer, or more, than a complete octet of valence electrons.

20. the arrangement of atoms in a molecule



22. The H-H bond is stronger because it has a greater dissociation energy.