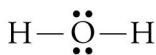


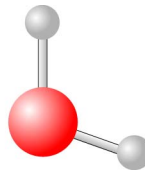
Chapter 5 Chemical Compounds



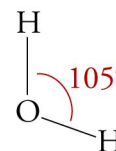
Lewis structure



Space-filling model



Ball-and-stick model



Geometric Sketch

◆ Review Skills

5.1 Classification of Matter

5.2 Compounds and Chemical Bonds

- Equal and Unequal Sharing of Electrons
- Transfer of Electrons
- Summary of Covalent and Ionic Bond Formation
- Predicting Bond Type
- Classifying Compounds

5.3 Ionic Compounds

- Cations and Anions
- Predicting Ion Charges
- Structure of Ionic Compounds
- Polyatomic Ions

5.4 Molecular Compounds

- The Strengths and Weaknesses of Models
- The Valence-Bond Model

5.5 Drawing Lewis Structures

- Simple Procedure
- General Procedure
- More Than One Possible Structure

5.6 Resonance

Internet: Resonance

5.7 Molecular Geometry from Lewis Structures

Special Topic 5.1: Molecular Shapes, Intoxicating Liquids, and the Brain

◆ Chapter Glossary

Internet: Glossary Quiz

◆ Chapter Objectives

Review Questions

Key Ideas

Chapter Problems

Section Goals and Introductions

The Review Skills section for this chapter is very important. Because students are so busy, it's always a battle to keep up, but if you don't keep up in chemistry, you're in real trouble. For this reason, the Review Skills section at the beginning of each chapter tells you what you really need to know (or be able to do) from earlier chapters to understand the chapter you are about to read. They are very important; don't just skip over them.

Section 5.1 Classification of Matter

Goal: To show how forms of matter can be classified as elements, compounds, and mixtures.

This section begins the process of teaching you how to classify matter into the categories of element, compound, and mixture. See Sample Study Sheet 5.1 for help classifying matter with respect to these categories. The distinctions among these categories will become increasingly clear as you study this chapter and Chapter 6. This section also describes how compounds are represented by chemical formulas.

Section 5.2 Compounds and Chemical Bonds

Goals

- *To show how atoms of different elements form links (chemical bonds) between them and to introduce three types of chemical bonding: nonpolar covalent, polar covalent, and ionic.*
- *To show how you can predict whether a pair of atoms will form a covalent bond or an ionic bond.*
- *To show how you can predict whether a chemical formula for a compound represents an ionic compound or a molecular compound.*

We can now take what you have learned about elements in Chapter 4 and expand on it to explain the formation of chemical bonds and chemical compounds. Be sure that you understand the similarities and differences among nonpolar covalent bonds, covalent bonds, and ionic bonds. The most important skills to develop from studying this section are (1) the ability to predict whether atoms of two elements would be expected to form an ionic bond or a covalent bond and (2) the ability to predict whether a chemical formula for a compound represents an ionic compound or a molecular compound. The ability to make these predictions is extremely important for many of the tasks in this chapter and in the chapters that follow.

Section 5.3 Ionic Compounds

Goals

- *To show why some atoms gain or lose electrons to form charged particles called ions.*
- *To show how the charges on ions can be predicted.*
- *To describe the structure of ionic solids.*
- *To describe polyatomic ions, which are charged collections of atoms held together by covalent bonds.*

This section is an important one. It provides more detailed information about ions than that found in Chapter 3, including how to predict their charges, how to convert between their names and symbols, and how they combine to form ionic compounds. Polyatomic ions (charged collections of atoms held together by covalent bonds) are also described. The section also includes a description of the structure of ionic solids. All of this information is used extensively in the rest of the book.

Section 5.4 Molecular Compounds

Goals

- *To describe the strengths and weaknesses of scientific models.*
- *To introduce a model, called the valence-bond model, which is very useful for describing the formation of covalent bonds.*
- *To explain the most common covalent bonding patterns for the nonmetallic atoms in terms of the valence-bond model.*

This section shows how the information learned in Chapter 4 can be combined with a model for covalent bonding called the valence-bond model to explain the common covalent bonding patterns of the nonmetallic atoms. It's important to recognize that although the valence bond model is only a model (and therefore a simplification of reality), it is extremely useful. You will find the information in Table 5.1: Covalent Bonding Patterns very helpful when drawing Lewis structures (the task described in Section 5.5).

Section 5.5 Drawing Lewis Structures

Goal: To show how Lewis structures can be drawn from chemical formulas.

First, you will learn to draw simple Lewis structures by arranging the atoms to yield the most common bonding pattern for each atom. This technique works very well for many molecules, but it is limited. For example, it does not work for polyatomic ions. This section also describes a procedure for drawing Lewis structures from chemical formulas that works for a broader range of molecules and polyatomic ions. See Sample Study Sheet 5.2 for help drawing Lewis structures.

Section 5.6 Resonance

Goal: To introduce a concept called resonance and show how it can be used to explain the characteristics of certain molecules and polyatomic ions.

The Lewis structures derived for some molecules and polyatomic ions by the technique described in Section 5.5 do not explain their characteristics adequately. One way that the valence-bond model has been expanded to better explain some of these molecules and polyatomic ions is by introducing the concept of resonance described in this section. See the related section on our Web site:

Internet: Resonance

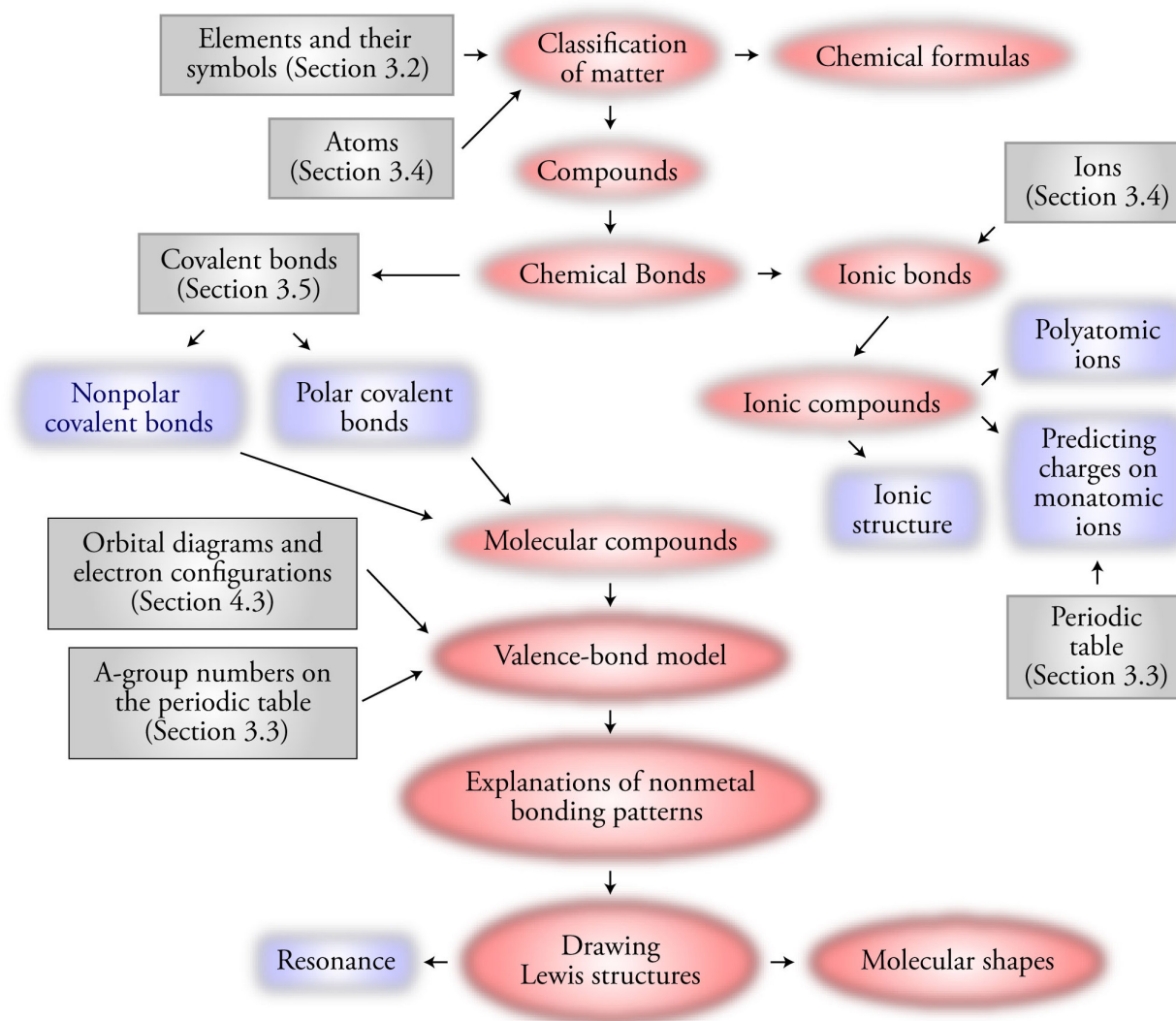
Section 5.7 Molecular Geometry from Lewis Structures

Goals

- *To show how you can predict the arrangement of atoms in molecules and polyatomic ions (called molecular geometry).*
- *To show how to make sketches of the molecular geometry of atoms in molecules and polyatomic ions.*

The arrangement of atoms in a molecule or polyatomic ion (that is, its molecular geometry) plays a significant role in determining its properties. This section explains why molecules and polyatomic ions have the geometry that they do, and Sample Study Sheet 5.3: Predicting Molecular Geometry and Table 5.2: Electron Group and Molecular Geometry show you how to predict and sketch these geometries.

Chapter 5 Map



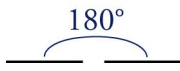
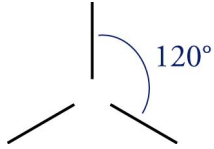
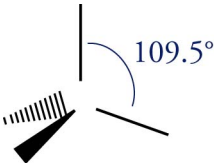
- The most common bonding patterns for the nonmetallic elements.

Elements	Number of covalent bonds	<i>Number of lone pairs</i>
C	4	0
N, P, & As	3	1
O, S, Se	2	2
F, Cl, Br, & I	1	3

- Although it's not absolutely necessary, it will help you to draw Lewis structures to know the expanded list of bonding patterns listed below.

Element	Frequency of pattern	Number of bonds	Number of lone pairs	Example
H	always	1	0	H-
B	most common	3	0	$\begin{array}{c} -\text{B}- \\ \end{array}$
C	most common	4	0	$\begin{array}{c} \\ -\text{C}- \\ \end{array} \quad \text{or} \quad \begin{array}{c} -\text{C}= \\ \end{array} \quad \text{or} \quad -\text{C}\equiv$
	rare	3	1	$\equiv\text{C}:$
N, P, & As	most common	3	1	$\begin{array}{c} \cdot\cdot \\ -\text{N}- \\ \end{array}$
	common	4	0	$\begin{array}{c} \\ -\text{N}- \\ \end{array}$
O, S, & Se	most common	2	2	$\begin{array}{c} \cdot\cdot \\ -\text{O}- \\ \cdot\cdot \end{array} \quad \text{or} \quad \begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \\ \end{array}$
	common	1	3	$\begin{array}{c} \cdot\cdot \\ -\text{O}: \\ \cdot\cdot \end{array}$
	rare	3	1	$\equiv\text{O}:$
F, Cl, Br, & I	most common	1	3	$\begin{array}{c} \cdot\cdot \\ -\text{X}: \\ \cdot\cdot \end{array}$

- The information found in the table below.

e ⁻ groups	e ⁻ group geometry	General Geometric Sketch	Bond angles	Bond groups	Lone pairs	molecular geometry
2	linear		180°	2	0	linear
3	trigonal planar		120°	3	0	trigonal planar
				2	1	bent
4	tetrahedral		109.5°	4	0	tetrahedral
				3	1	trigonal pyramid
				2	2	bent

- To get a review of the most important topics in the chapter, fill in the blanks in the Key Ideas section.
- Work all of the selected problems at the end of the chapter, and check your answers with the solutions provided in this chapter of the study guide.
- Ask for help if you need it.

Web Resources

Internet: Resonance

Internet: Glossary Quiz

Exercises Key

Exercise 5.1 - Classification of Matter: The label on a container of double-acting baking powder tells us that it contains cornstarch, bicarbonate of soda (also called sodium hydrogen carbonate, NaHCO_3), sodium aluminum sulfate, and acid phosphate of calcium (which chemists call calcium dihydrogen phosphate, $\text{Ca}(\text{H}_2\text{PO}_4)_2$). Classify each of the following as a pure substance or a mixture. If it is a pure substance, is it an element or a compound?

(Oblis 3 & 4)

- | | |
|---------------------------------|-----------------|
| a. calcium | element |
| b. calcium dihydrogen phosphate | compound |
| c. double-acting baking powder | mixture |

Exercise 5.2 - Classifying Compounds: Classify each of the following substances as either a molecular compound or an ionic compound. (*Obj 9*)

- a. formaldehyde, CH_2O (used in embalming fluids)

all nonmetal atoms - molecular

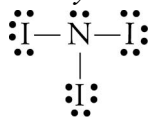
- b. magnesium chloride, MgCl_2 (used in fireproofing wood and in paper manufacturing)

metal-nonmetal - ionic

Exercise 5.3 - Drawing Lewis Structures from Formulas: Draw a Lewis structure for each of the following formulas: (*Obj 21*)

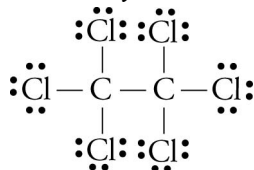
- a. nitrogen triiodide, NI_3 (explodes at the slightest touch)

Nitrogen atoms usually have three covalent bonds and one lone pair, and iodine atoms usually have one covalent bond and three lone pairs.



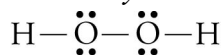
- b. hexachloroethane, C_2Cl_6 (used to make explosives)

Carbon atoms usually have four covalent bonds and no lone pairs, and chlorine atoms usually have one covalent bond and three lone pairs.



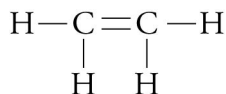
- c. hydrogen peroxide, H_2O_2 (a common antiseptic)

Hydrogen atoms always have one covalent bond and no lone pairs, and oxygen atoms usually have two covalent bonds and two lone pairs.

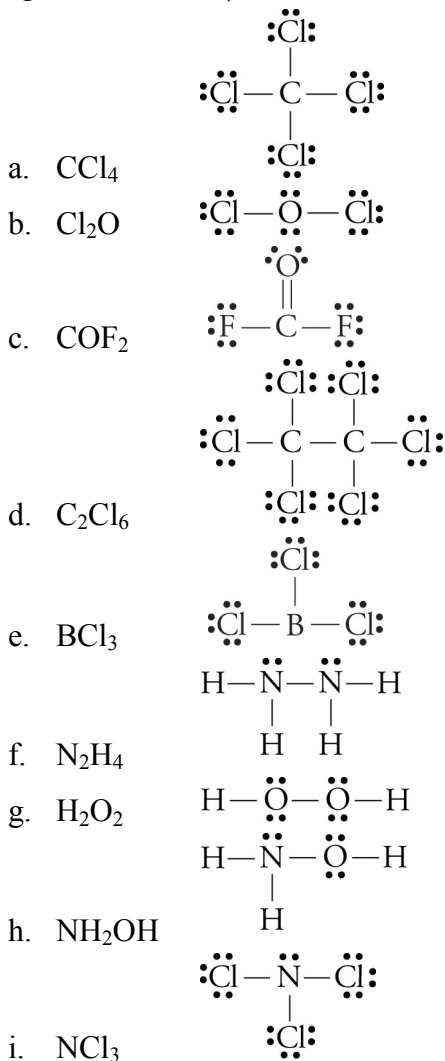


- d. ethylene (or ethene), C_2H_4 (used to make polyethylene)

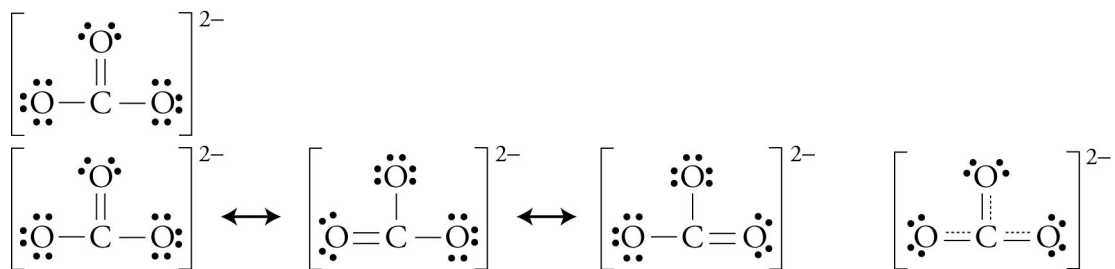
Carbon atoms form four bonds with no lone pairs, and hydrogen atoms form one bond with no lone pairs. To achieve these bonding patterns, there must be a double bond between the carbon atoms.



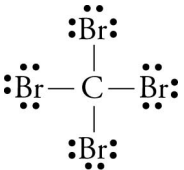
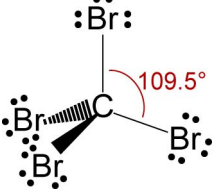
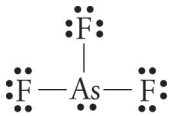
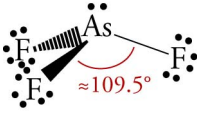
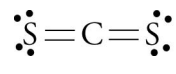
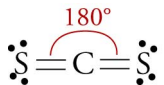
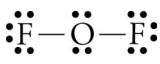
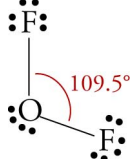
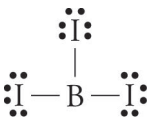
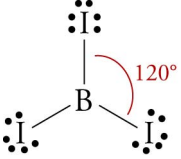
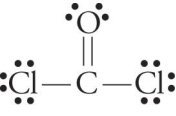
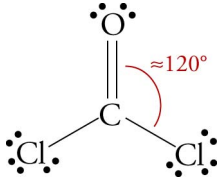
Exercise 5.4 - Lewis Structures: Draw a reasonable Lewis structure for each of the following formulas. (Obj 22)

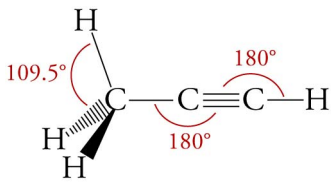
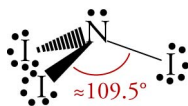


Exercise 5.5 - Resonance: Draw all of the reasonable resonance structures and the resonance hybrid for the carbonate ion, CO_3^{2-} . A reasonable Lewis structure for the carbonate ion is (Obj 25)



Exercise 5.6 - Molecular Geometry: For each of the Lewis structures that follow, (a) write the name of the electron group geometry around each atom that has two or more atoms attached to it, (b) draw the geometric sketch of the molecule, including bond angles, and (c) write the name of the molecular geometry around each atom that has two or more atoms attached to it. (*Obj 30*)

<p>a.</p> 	<p><i>Electron Group Geometry - tetrahedral</i></p>  <p><i>Molecular Geometry - tetrahedral</i></p>
<p>b.</p> 	<p><i>Electron Group Geometry - tetrahedral</i></p>  <p><i>Molecular Geometry - trigonal pyramid</i></p>
<p>c.</p> 	<p><i>Electron Group Geometry - linear</i></p>  <p><i>Molecular Geometry - linear</i></p>
<p>d.</p> 	<p><i>Electron Group Geometry - tetrahedral</i></p>  <p><i>Molecular Geometry - bent</i></p>
<p>e.</p> 	<p><i>Electron Group Geometry - trigonal planar</i></p>  <p><i>Molecular Geometry - trigonal planar</i></p>
<p>f.</p> 	<p><i>Electron Group Geometry - trigonal planar</i></p>  <p><i>Molecular Geometry - trigonal planar</i></p>

<p>g.</p> $\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{C}\equiv\text{C}-\text{H} \\ \\ \text{H} \end{array}$	<p><i>Electron group geometry for left carbon - tetrahedral</i> <i>Electron group geometry for middle and right carbons - linear</i></p>  <p><i>Molecular Geometry for left carbon - tetrahedral</i> <i>Molecular Geometry for middle and right carbon - linear</i></p>
<p>h.</p> $\begin{array}{c} \text{:}\ddot{\text{I}}-\ddot{\text{N}}-\ddot{\text{I}}\text{:} \\ \\ \text{:}\ddot{\text{I}}\text{:} \end{array}$	<p><i>Electron Group Geometry - tetrahedral</i></p>  <p><i>Molecular Geometry - trigonal pyramid</i></p>

Review Questions Key

- Using the A-group convention, what is the group number of the column in the periodic table that includes the element chlorine, Cl? (See Section 3.3.) **7A**

Write in each blank the word or words that best complete each sentence.

- An atom or group of atoms that has lost or gained one or more electrons to create a charged particle is called a(n) **ion**.
- An atom or collection of atoms with an overall positive charge is a(n) **cation**.
- An atom or collection of atoms with an overall negative charge is a(n) **anion**.
- A(n) **covalent** bond is a link between atoms that results from the sharing of two electrons.
- A(n) **molecule** is an uncharged collection of atoms held together with covalent bonds.
- A molecule such as H₂, which is composed of two atoms, is called **diatomic**.
- Describe the particle nature of solids, liquids, and gases. Your description should include the motion of the particles and the attractions between the particles.

See Figures 3.1, 3.2, and 3.4 in the textbook.

- Describe the nuclear model of the atom.

Protons and neutrons are in a tiny core of the atom called the nucleus, which has a diameter about 1/100,000 the diameter of the atom. The position and motion of the electrons are uncertain, but they generate a negative charge that is felt in the space that surrounds the nucleus.

- Describe the hydrogen molecule, H₂. Your description should include the nature of the link between the hydrogen atoms and a sketch that shows the two electrons in the molecule.

The hydrogen atoms are held together by a covalent bond formed due to the sharing of two electrons. See the image of H₂ in Figure 3.13 in the textbook.

11. Complete the following table.

Element name	Element symbol	Group number on periodic table	Metal, nonmetal, or metalloid?
lithium	Li	1 or 1A	metal
carbon	C	14 or 4A	nonmetal
chlorine	Cl	17 or 7A	nonmetal
oxygen	O	16 or 6A	nonmetal
copper	Cu	11 or 1B	metal
calcium	Ca	2 or 2A	metal
scandium	Sc	3 or 3B	metal

12. Write the name of the group to which each of the following belongs.

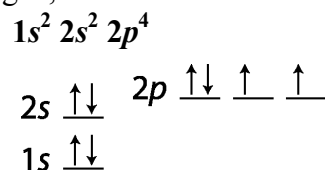
- a. chlorine **halogens** c. sodium **alkali metals**
 b. xenon **noble gases** d. magnesium **alkaline earth metals**

13. Define the term *orbital*. (See Section 4.2.)

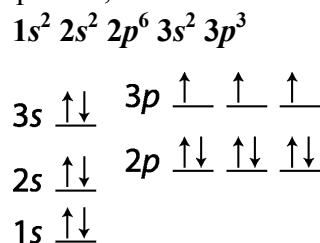
Orbital can be defined as the volume that contains a high percentage of the electron charge generated by an electron in an atom. It can also be defined as the volume within which an electron has a high probability of being found.

14. Write a complete electron configuration and an orbital diagram for each of the following. (See Section 4.3.)

a. oxygen, O



b. phosphorus, P



Key Ideas Answers

15. A compound is a substance that contains two or more elements, the atoms of those elements always combining in the same **whole-number ratio**.
17. A chemical formula is a concise written description of the components of a chemical compound. It identifies the elements in the compound by their **symbols** and indicates the relative number of atoms of each element with **subscripts**.
19. Mixtures are samples of matter that contain two or more pure substances and have **variable** composition.
21. Because particles with opposite charges attract each other, there is an attraction between **cations** and **anions**. This attraction is called an ionic bond.

-
23. The atom in a chemical bond that attracts electrons more strongly acquires a **negative** charge, and the other atom acquires a **positive** charge. If the electron transfer is significant but not enough to form ions, the atoms acquire **partial negative** and **partial positive** charges. The bond in this situation is called a polar covalent bond.
 25. When a metallic atom bonds to a nonmetallic atom, the bond is **usually ionic**.
 27. Metallic atoms hold some of their electrons relatively loosely, and as a result, they tend to **lose** electrons and become cations. In contrast, nonmetallic atoms attract electrons more strongly than metallic atoms, so nonmetals tend to **gain** electrons and form anions.
 29. When atoms gain electrons and form anions, they get **larger**. When atoms lose electrons and form cations, they get significantly **smaller**.
 31. When developing a model of physical reality, scientists take what they think is **true** and simplify it enough to make it **useful**.
 33. One characteristic of models is that they **change** with time.
 35. Valence electrons are the highest-energy **s and p** electrons in an atom.
 37. Paired valence electrons are called **lone pairs**.
 39. When atoms other than the noble-gas atoms form bonds, they often have **eight valence** electrons around them in total.
 41. When forming four bonds to hydrogen atoms in a methane molecule, each carbon atom behaves **as if** one electron is promoted from the $2s$ orbital to the $2p$ orbital.
 43. The shortcut for drawing Lewis structures for which we try to give each atom its **most common** bonding pattern works well for many simple uncharged molecules, but it does not work reliably for molecules that are more complex or for **polyatomic ions**.
 45. For polyatomic cations, the total number of valence electrons is the sum of the valence electrons for each atom **minus** the charge.
 47. Hydrogen and fluorine atoms are **never** in the center of a Lewis structure.
 49. The element with the **fewest** atoms in the formula is often in the center of a Lewis structure.
 51. Oxygen atoms rarely bond to other **oxygen** atoms.
 53. In a reasonable Lewis structure, hydrogen will always have a total of **two** electrons from its one bond.
 55. Substances that have the same molecular formula but different **structural** formulas are called isomers.
 57. Lewis structures are useful for showing how the atoms in a molecule are connected by covalent bonds, but they do not always give a clear description of how the atoms are **arranged in space**.
 59. The geometric arrangement in which four electron groups are as far apart as possible is a tetrahedral arrangement, with bond angles of 109.5° .
 61. The most stable arrangement for electron groups is one in which they are **as far apart** as possible.

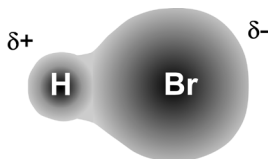
Problems Key

Section 5.1 Classification of Matter

63. Classify each of the following as a pure substance or a mixture. If it is a pure substance, is it an element or a compound? Explain your answer. (*Obj 3 & 4*)
- apple juice
mixture – *variable composition*
 - potassium (A serving of one brand of apple juice provides 6% of the recommended daily allowance of potassium.)
element and pure substance – *The symbol K is on the periodic table of the elements. All elements are pure substances.*
 - ascorbic acid (vitamin C), $C_6H_8O_6$, in apple juice
*The formula shows the constant composition, so ascorbic acid is a **pure substance**. The three element symbols in the formula indicate a **compound**.*
65. Write the chemical formula for each of the following compounds. List the symbols for the elements in the order that the elements are mentioned in the description. (*Obj 2*)
- a compound with molecules that consist of two nitrogen atoms and three oxygen atoms.
 N_2O_3
 - a compound with molecules that consist of one sulfur atom and four fluorine atoms.
 SF_4
 - a compound that contains one aluminum atom for every three chlorine atoms.
 $AlCl_3$
 - a compound that contains two lithium atoms and one carbon atom for every three oxygen atoms.
 Li_2CO_3

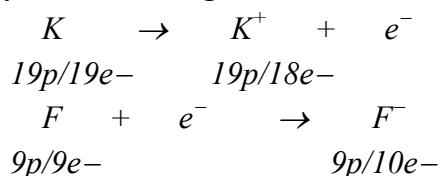
Section 5.2 Chemical Compounds and Chemical Bonds

67. Hydrogen bromide, HBr, is used to make pharmaceuticals that require bromine in their structure. Each hydrogen bromide molecule has one hydrogen atom bonded to one bromine atom by a polar covalent bond. The bromine atom attracts electrons more than does the hydrogen atom. Draw a rough sketch of the electron-cloud that represents the electrons involved in the bond. (*Obj 5*)



69. Atoms of potassium and fluorine form ions and ionic bonds in a very similar way to atoms of sodium and chlorine. Each atom of one of these elements loses one electron, and each atom of the other element gains one electron. Describe the process that leads to the formation of the ionic bond between potassium and fluorine atoms in potassium fluoride. Your answer should include mention of the charges that form on the atoms. (Obj 6)

The metallic potassium atoms lose one electron and form +1 cations, and the nonmetallic fluorine atoms gain one electron and form -1 anions.



The ionic bonds are the attractions between K^+ cations and F^- anions.

71. Explain how a nonpolar covalent bond, a polar covalent bond, and an ionic bond differ. Your description should include rough sketches of the electron clouds that represent the electrons involved in the formation of each bond. (Obj 7)

See Figure 5.6 in the textbook.

73. Would you expect the bonds between the following atoms to be ionic or covalent bonds? (Obj 8)

- a. N-O **covalent...nonmetal-nonmetal**
 b. Al-Cl **ionic...metal-nonmetal**

75. Classify each of the following as either a molecular compound or an ionic compound. (Obj 9)

- a. acetone, CH_3COCH_3 (a common paint solvent)
 *all nonmetallic atoms - **molecular***
 b. sodium sulfide, Na_2S (used in sheep dips) *metal-nonmetal - **ionic***

Section 5.3 Ionic Compounds

77. Explain why metals usually combine with nonmetals to form ionic bonds. (Obj 10)

Because metallic atoms hold some of their electrons relatively loosely, they tend to lose electrons and form cations. Because nonmetallic atoms attract electrons more strongly than metallic atoms, they tend to gain electrons and form anions. Thus, when a metallic atom and a nonmetallic atom combine, the nonmetallic atom often pulls one or more electrons far enough away from the metallic atom to form ions and an ionic bond.

78. How many protons and electrons do each of the following ions have?

- a. Be^{2+} **4 protons and 2 electrons**
 b. S^{2-} **16 protons and 18 electrons**

80. When atoms of the following elements form ions, what charge or charges would the ions have? (*Obj 11*)

- | | | | |
|-------------|------------------|-------------|--------------------------------------|
| a. hydrogen | H^- | e. iron | Fe^{2+} or Fe^{3+} |
| b. Ca | Ca^{2+} | f. Zn | Zn^{2+} |
| c. fluorine | F^- | g. nitrogen | N^{3-} |
| d. K | K^+ | h. Sc | Sc^{3+} |

82. Silver bromide, AgBr, is the compound on black and white film that causes the color change when the film is exposed to light. It has a structure similar structure to that of sodium chloride. What are the particles that form the basic structure of silver bromide? What type of attraction holds these particles together? Draw a rough sketch of the structure of solid silver bromide. (*Obj 12*)

The metallic silver atoms form cations, and the nonmetallic bromine atoms form anions. The anions and cations alternate in the ionic solid with each cation surrounded by six anions and each anion surrounded by six cations. See Figure 5.11 in the textbook, picturing Ag^+ ions in the place of the Na^+ ions and Br^- in the place of the Cl^- ions.

Section 5.4 Molecular Compounds

84. Describe the advantages and disadvantages of using models to describe physical reality. (*Obj 14*)

Our models come with advantages and disadvantages. They help us to visualize, explain, and predict chemical changes, but we need to remind ourselves now and then that they are only models, and as models, they have their limitations. For example, because a model is a simplified version of what we think is true, the processes it depicts are sometimes described using the phrase “as if.” When you read, “It is as if an electron were promoted from one orbital to another,” the phrase is a reminder that we do not necessarily think this is what really happens. We find it useful to talk about the process as if this is the way it happens.

One characteristic of models is that they change with time. Because our models are a simplification of what we think is real, we are not surprised when they sometimes fail to explain experimental observation. When this happens, the model is altered to fit the new observations.

86. How many valence electrons do the atoms of each of the following elements have? Write the electron configuration for these electrons. (For example, fluorine has 7 valence electrons, which can be described as $2s^2 2p^5$.) (*Obj 16*)

- | | |
|----------------|---------------------------------------------------------|
| a. nitrogen, N | five valence electrons – $2s^2 2p^3$ |
| b. sulfur, S | six valence electrons – $3s^2 3p^4$ |
| c. iodine, I | seven valence electrons – $5s^2 5p^5$ |
| a. argon, Ar | eight valence electrons – $3s^2 3p^6$ |

88. Draw electron-dot symbols for each of the following elements. (Obj 19)

- a. nitrogen, N $\cdot\ddot{\text{N}}\cdot$
 b. sulfur, S $\cdot\ddot{\text{S}}\cdot$
 c. iodine, I $\cdot\ddot{\text{I}}\cdot$
 d. argon, Ar $\cdot\ddot{\text{Ar}}\cdot$

90. To which group on the periodic table would atoms with the following electron-dot symbols belong? List the group numbers using the 1-18 convention and using the A-group convention.

- a. $\cdot\ddot{\text{X}}\cdot$ **Group 15 or 5A**
 b. $\cdot\ddot{\text{X}}\cdot$ **Group 18 or 8A**
 c. $\cdot\ddot{\text{X}}\cdot$ **Group 13 or 3A**

92. For each of the following elements, sketch *all* of the ways mentioned in Section 5.4 that their atoms could look in a Lewis structure. For example, fluorine has only one bonding pattern, and it looks like $\text{—}\ddot{\text{F}}\cdot$.

- a. nitrogen, N $\text{—}\ddot{\text{N}}\text{—}$ or $\text{—}\ddot{\text{N}}\text{—}$
 b. boron, B $\text{—}\ddot{\text{B}}\text{—}$
 c. carbon, C $\text{—}\ddot{\text{C}}\text{—}$ or $\text{—}\ddot{\text{C}}\text{=}$ or $\text{—}\ddot{\text{C}}\equiv$ or $\equiv\ddot{\text{C}}\text{:}$

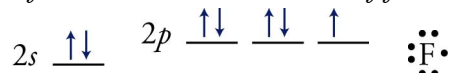
94. Use the valence-bond model to explain the following observations. (Obj 17)

The answers to each of these problems is based on the following assumptions of the valence bond model.

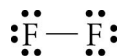
- Only the highest energy electrons participate in bonding.
- Covalent bonds usually form to pair unpaired electrons.

a. Fluorine atoms have one bond and three lone pairs in F_2 .

Fluorine is in group 7A, so it has seven valence electrons per atom. The orbital diagram for the valence electrons of fluorine is below.

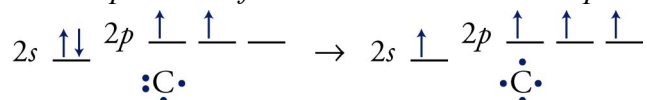


The one unpaired electron leads to one bond, and the three pairs of electrons give fluorine atoms three lone pairs.

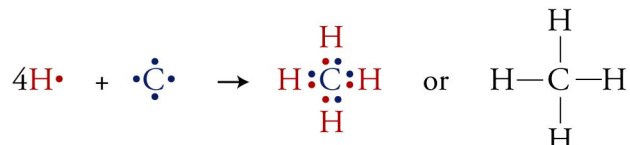


- b. Carbon atoms have four bonds and no lone pairs in CH₄.

Carbon is in group 4A, so it has four valence electrons per atom. It is as if one electron is promoted from the 2s orbital to the 2p orbital.

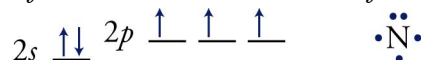


The four unpaired electrons lead to four covalent bonds. Because there are no pairs of electrons, carbon atoms have no lone pairs when they form four bonds.

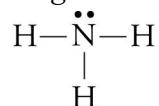


- c. Nitrogen atoms have three bonds and one lone pair in NH₃.

Nitrogen is in group 5A, so it has five valence electrons per atom. The orbital diagram for the valence electrons of nitrogen is below.

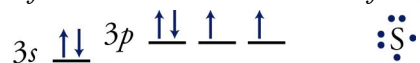


The three unpaired electrons lead to three bonds, and the one pair of electrons gives nitrogen atoms one lone pair.

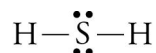


- d. Sulfur atoms have two bonds and two lone pairs in H₂S.

Sulfur is in group 6A, so it has six valence electrons per atom. The orbital diagram for the valence electrons of sulfur is below.

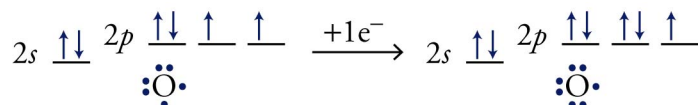


The two unpaired electrons lead to two bonds, and the two pairs of electrons give sulfur atoms two lone pairs.



- e. Oxygen atoms have one bond and three lone pairs in OH⁻.

Oxygen is in group 6A, so it has six valence electrons per atom. If it gains one electron, it will have a total of seven.



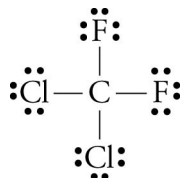
The one unpaired electron leads to one bond, and the three pairs of electrons give oxygen atoms with an extra electron three lone pairs.



96. Based on your knowledge of the most common bonding patterns for the nonmetallic elements, predict the formulas with the lowest subscripts for the compounds that would form from the following pairs of elements. (For example, hydrogen and oxygen can combine to form H_2O and H_2O_2 , but H_2O has lower subscripts.)

- C and H **CH_4**
- S and H **H_2S**
- B and F **BF_3**

98. The following Lewis structure is for CFC-12, which is one of the ozone-depleting chemicals that has been used as an aerosol can propellant and as a refrigerant. Describe the information given in this Lewis structure. (*Obj 18*)



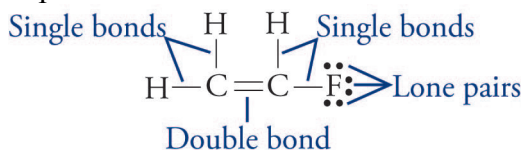
The molecule contains a carbon atom, two chlorine atoms, and two fluorine atoms. There are two covalent C–Cl bonds and two covalent C–F bonds. The Cl and F atoms have three lone pairs each.

100. Write the most common number of covalent bonds and lone pairs for atoms of each of the following nonmetallic elements. (*Obj 20*)

- H **1 bond, no lone pairs**
- iodine **1 bond, 3 lone pairs**
- sulfur **2 bonds, 2 lone pairs**
- N **3 bonds, 1 lone pair**

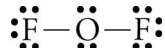
Section 5.5 Drawing Lewis Structures

102. Copy the following Lewis structure and identify the single bonds, the double bond, and the lone pairs.

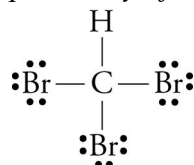


104. Draw a Lewis structure for each of the following formulas. (*Obj 21 & 22*)

- oxygen difluoride, OF_2 (an unstable, colorless gas)
Oxygen atoms usually have two covalent bonds and two lone pairs, and fluorine atoms have one covalent bond and three lone pairs.

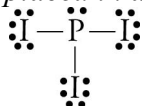


- bromoform, CHBr_3 (used as a sedative)
Carbon atoms usually have four covalent bonds and no lone pairs, hydrogen atoms always have one covalent bond and no lone pairs, and bromine atoms usually have one covalent bond and three lone pairs. The hydrogen atom can be put in any of the four positions.



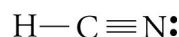
- c. phosphorus triiodide, PI_3 (used to make organic compounds)

Phosphorus atoms usually have three covalent bonds and one lone pair, and iodine atoms usually have one covalent bond and three lone pairs. The lone pair can be placed in any one of the four positions around the phosphorus atom.

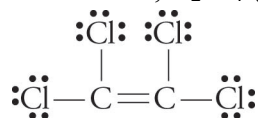


106. Draw Lewis structures for the following compounds by adding any necessary lines and dots to the skeletons given. (*Obj's 21 & 22*)

- a. hydrogen cyanide, HCN (used to manufacture dyes and pesticides)



- b. dichloroethene, C_2Cl_4 (used to make perfumes)



108. Write two different names for each of the following alcohols. (*Obj's 23 & 24*)

- a.
$$\begin{array}{c} \text{H} \\ | \\ \text{H} - \text{C} - \text{O} - \text{H} \\ | \\ \text{H} \end{array}$$
 methanol and methyl alcohol

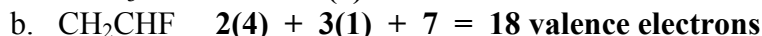
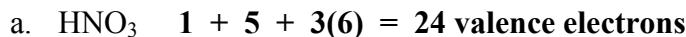
- b.
$$\begin{array}{c} \text{H} \quad \text{H} \\ | \quad | \\ \text{H} - \text{C} - \text{C} - \text{O} - \text{H} \\ | \quad | \\ \text{H} \quad \text{H} \end{array}$$
 ethanol and ethyl alcohol

- c.
$$\begin{array}{c} \text{:O:} - \text{H} \\ | \\ \text{H} \quad | \quad \text{H} \\ | \quad | \quad | \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{H} \\ | \quad | \quad | \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$$
 2-propanol and isopropyl alcohol

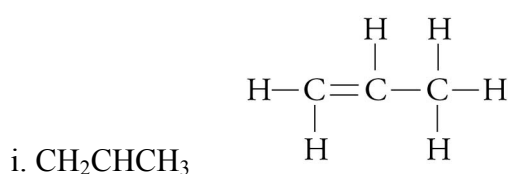
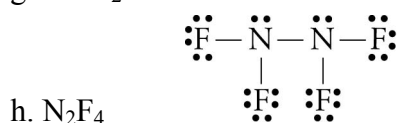
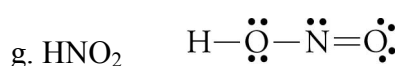
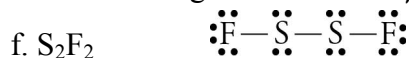
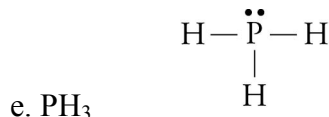
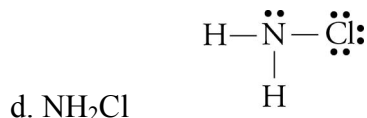
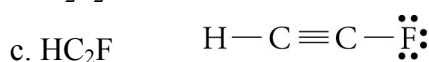
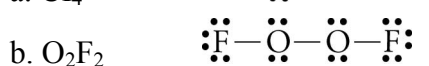
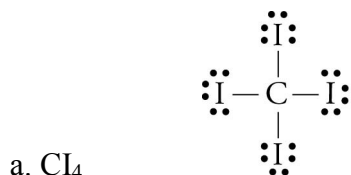
109. For each of the following molecular compounds, identify the atom that is most likely to be found in the center of its Lewis structure. Explain why.

- a. CBr_4 **Carbon** –*The element with the fewest atoms in the formula is often in the center. The atom that is capable of making the most bonds is often in the center. Carbon atoms usually form four bonds, and bromine atoms usually form one bond.*
- b. SO_2 **Sulfur** –*The element with the fewest atoms in the formula is often in the center. Oxygen atoms are rarely in the center. Oxygen atoms rarely bond to other oxygen atoms.*
- c. H_2S **Sulfur** –*The element with the fewest atoms in the formula is often in the center. The atom that is capable of making the most bonds is often in the center. Sulfur atoms usually form two bonds, and hydrogen atoms form one bond. Hydrogen atoms are never in the center.*
- d. NOF **Nitrogen** –*The atom that is capable of making the most bonds is often in the center. Nitrogen atoms usually form three bonds, oxygen atoms usually form two bonds, and fluorine atoms form one bond. Fluorine atoms are never in the center.*

111. Calculate the total number of valence electrons for each of the following formulas.



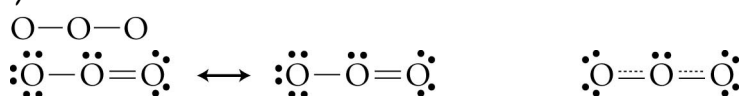
113. Draw a reasonable Lewis structure for each of the following formulas. (Obj's 21 & 22)



Section 5.6 Resonance

115. Draw a reasonable Lewis structure for the ozone molecule, O_3 , using the skeleton that follows. The structure is best described in terms of resonance, so draw all of its reasonable resonance structures and the resonance hybrid that summarizes these structures.

(Obj's 25)



Section 5.7 Molecular Geometry from Lewis Structures

118. Compare and contrast the information given in the Lewis structure, the space-filling model, the ball-and-stick model, and the geometric sketch of a methane molecule, CH_4 .

(Obj's 28 & 29)

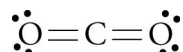
See Figure 5.19 in the textbook. The Lewis structure shows the four covalent bonds between the carbon atoms and the hydrogen atoms. The space-filling model provides the most accurate representation of the electron-charge clouds for the atoms in CH_4 . The ball-and-stick model emphasizes the molecule's correct molecular shape and shows the covalent bonds more clearly. Each ball represents an atom, and each stick represents a covalent bond between two atoms. The geometric sketch shows the three-dimensional tetrahedral structures with a two-dimensional drawing. Picture the hydrogen atoms connected to the central carbon atom with solid lines as being in the same plane as the carbon atom. The hydrogen atom connected to the central carbon with a solid wedge comes out of the plane toward you. The hydrogen atom connected to the carbon atom by a dashed wedge is located back behind the plane of the page.

120. Compare and contrast the information given in the Lewis structure, the space-filling model, the ball-and-stick model, and the geometric sketch of a water molecule, H_2O . (*Obj's 28 & 29*)

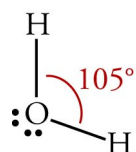
See Figure 5.23 in the textbook. The Lewis structure shows the two O–H covalent bonds and the two lone pairs on the oxygen atom. The space-filling model provides the most accurate representation of the electron-charge clouds for the atoms and the bonding electrons. The ball-and-stick model emphasizes the molecule's correct molecular shape and shows the covalent bonds more clearly. The geometric sketch shows the structure with a two-dimensional drawing.

121. Although both CO_2 molecules and H_2O molecules have three atoms, CO_2 molecules are linear, and H_2O molecules are bent. Why?

Atoms are arranged in molecules to keep the electron groups around the central atom as far apart as possible. An electron group is either (1) a single bond, (2) a multiple bond (double or triple), or (3) a lone pair. The Lewis structure for CO_2 shows that the carbon atom has two electron groups around it. The best way to get two things as far apart as possible is in a linear arrangement.



The Lewis structure for H_2O shows that the oxygen atom has four electron groups around it. The best way to get four things as far apart as possible is in a tetrahedral arrangement.



123. Using the symbol X for the central atom and Y for the outer atoms, draw the general geometric sketch for a three-atom molecule with linear geometry.



126. For each of the Lewis structures that follows, (Obj 30)

- Write the name of the electron group geometry around each atom that has two or more atoms attached to it.
- Draw the geometric sketch of the molecule, including bond angles (or approximate bond angles).
- Write the name of the molecular geometry around each atom that has two or more atoms attached to it.

