

# Chapter 6

## Preview

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- Objectives
- Chemical Bond

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# Chapter 6

## Section 1 Introduction to Chemical Bonding

### Lesson Starter ▼

- Imagine getting onto a crowded elevator. As people squeeze into the confined space, they come in contact with each other. Many people will experience a sense of being too close together. ▼
- When atoms get close enough, their outer electrons repel each other. At the same time, however, each atom's outer electrons are strongly attracted to the nuclei of the surrounding atoms. ▼
- The degree to which these outer electrons are attracted to other atoms determines the kind of chemical bonding that occurs between the atoms.



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# Chapter 6

## Section 1 Introduction to Chemical Bonding

### Objectives ▼

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



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### Chemical Bond

- *Mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together.*

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## Section 1 Introduction to Chemical Bonding

### Ionic Bonding

- *Chemical bonding that results from the electrical attraction between cations and anions*

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# Chapter 6

## Section 1 Introduction to Chemical Bonding

### Covalent Bonding

- *Chemical bonding that results from the sharing of electron pairs between two atoms*

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### Comparing Polar and Nonpolar Covalent Bonds

***Nonpolar-covalent*** – A covalent bond in which the bonding electrons are shared equally by the bonded atoms, resulting in a balanced distribution of electrical charge

- Electronegativity difference between 0-0.3

***Polar-covalent*** – A covalent bond in which the bonding electrons have an uneven distribution of charge

- Electronegativity difference between 0.3-1.7

# Chapter 6

## Section 1 Introduction to Chemical Bonding

### Using Electronegativity Difference to Classify Bonding

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### Chemical Bonding, *continued*

#### Sample Problem A ▾

Use electronegativity values listed in **Figure 20** from the previous chapter in your book, on page 161, and **Figure 2** in your book, on page 176, to classify bonding between sulfur, S, and the following elements: hydrogen, H; cesium, Cs; and chlorine, Cl. In each pair, which atom will be more negative?



# Chapter 6

## Section 1 Introduction to Chemical Bonding

### Chemical Bonding, *continued*

#### Sample Problem A Solution ▼

The electronegativity of sulfur is 2.5. The electronegativities of hydrogen, cesium, and chlorine are 2.1, 0.7, and 3.0, respectively. In each pair, the atom with the larger electronegativity will be the more-negative atom. ▼

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



# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Preview

- Objectives
- Molecular Compounds
- Formation of a Covalent Bond
- Characteristics of the Covalent Bond
- The Octet Rule
- Electron-Dot Notation
- Lewis Structures
- Multiple Covalent Bonds

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# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Objectives ▼

- **Define** *molecule* and *molecular formula*. ▼
- **Explain** the relationships among potential energy, distance between approaching atoms, bond length, and bond energy. ▼
- **State** the octet rule.



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# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Objectives, *continued* ▼

- **List** the six basic steps used in writing Lewis structures. ▼
- **Explain** how to determine Lewis structures for molecules containing single bonds, multiple bonds, or both. ▼
- **Explain** why scientists use resonance structures to represent some molecules.



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### Molecular Compounds ▼

- A **molecule** is a neutral group of atoms that are held together by covalent bonds. ▼
- A chemical compound whose simplest units are molecules is called a **molecular compound**. ▼



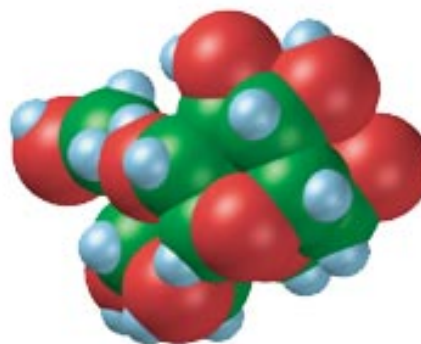
### Molecule



Water molecule  
 $\text{H}_2\text{O}$



Oxygen molecule  
 $\text{O}_2$



Sucrose molecule  
 $\text{C}_{12}\text{H}_{22}\text{O}_{11}$



### Molecular Compounds ▼

- The composition of a compound is given by its chemical formula. ▼
- A **chemical formula** indicates the relative numbers of atoms of each kind in a chemical compound by using atomic symbols and numerical subscripts. ▼
- A **molecular formula** shows the types and numbers of atoms combined in a single molecule of a molecular compound.





# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Chemical Formula

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## Section 2 Covalent Bonding and Molecular Compounds

### Structure of a Water Molecule

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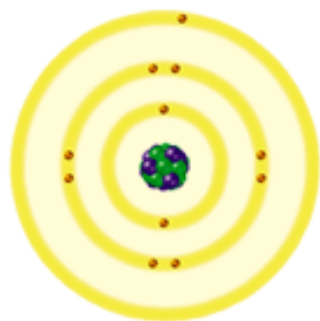
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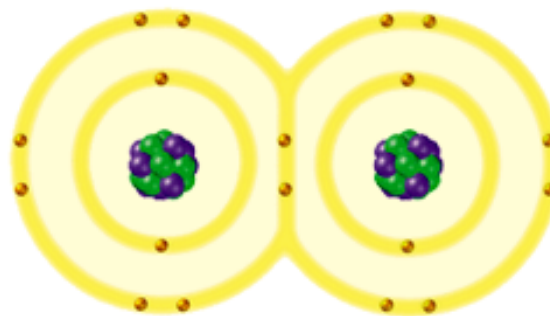
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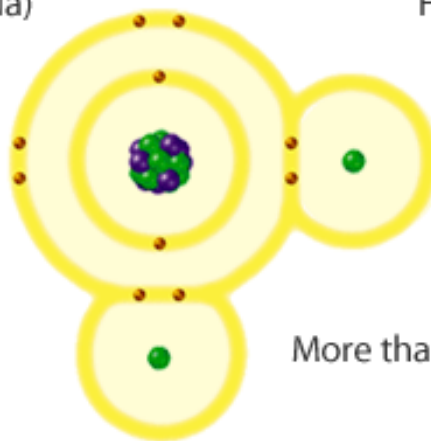
### Comparing Monatomic, Diatomic, and Polyatomic Molecules



1 atom = monatomic  
Sodium (Na)



2 atoms = diatomic  
Fluorine (F<sub>2</sub>)



More than 2 atoms = polyatomic  
Water (H<sub>2</sub>O)

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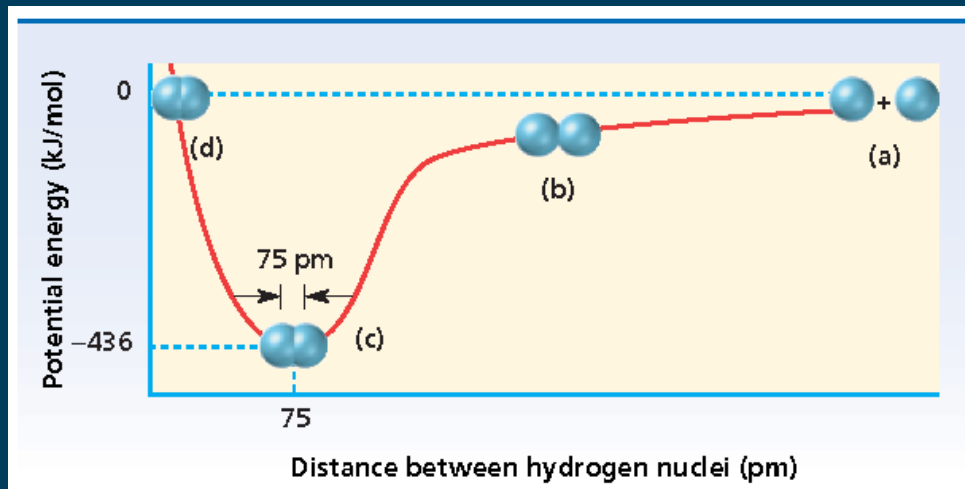
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# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Formation of a Covalent Bond ▾

- Most atoms have lower potential energy when they are bonded to other atoms than they have as they are independent particles. ▾
- The figure below shows potential energy changes during the formation of a hydrogen-hydrogen bond.



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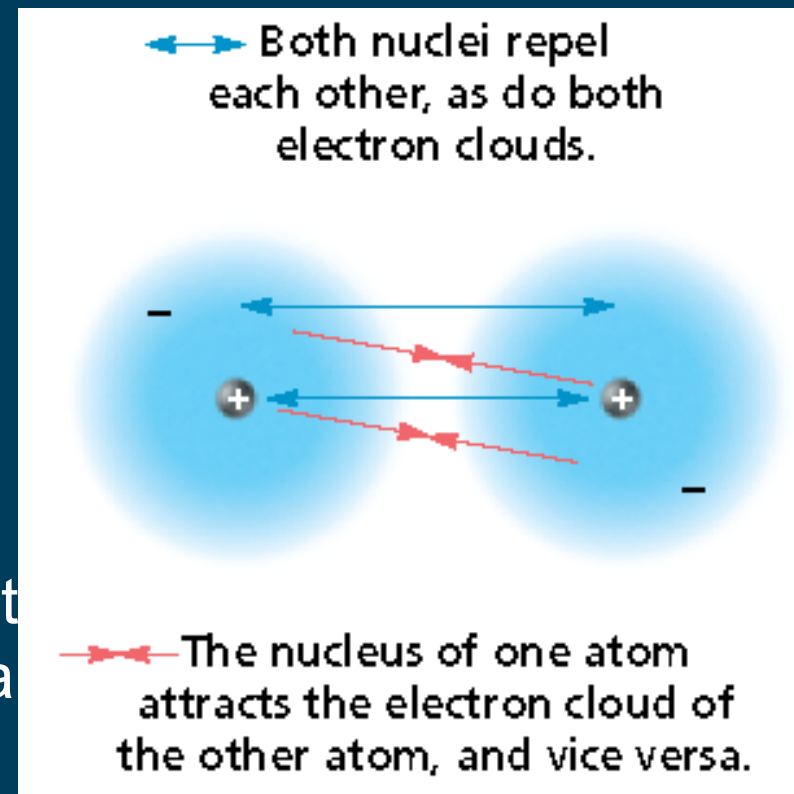
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## Section 2 Covalent Bonding and Molecular Compounds

### Formation of a Covalent Bond ▾

- The electron of one atom and proton of the other atom *attract* one another. ▾
- The two nuclei and two electrons *repel* each other. ▾
- These two forces cancel out to form a covalent bond at a length where the potential energy is at a minimum.



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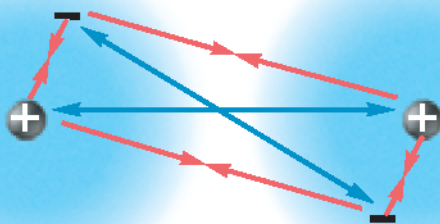
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### Formation of a Covalent Bond

Both nuclei repel each other,  
as do both electron clouds.

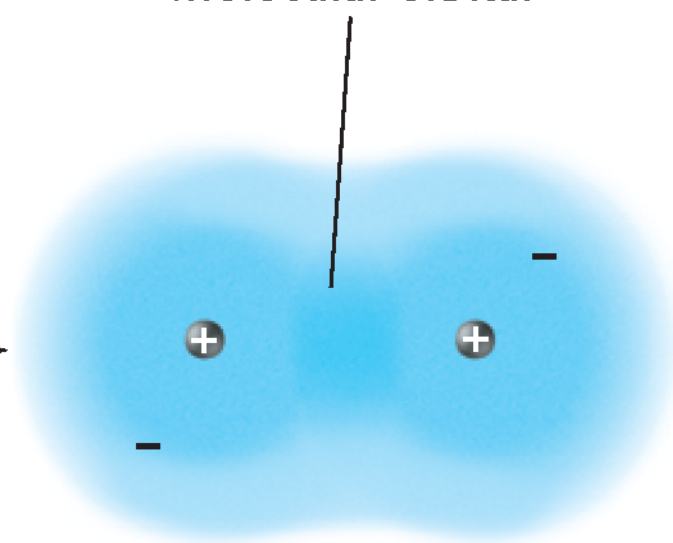
( repulsion)



The nucleus of each atom  
attracts both electron clouds.

( attraction)

Molecular orbital



A covalent bond is formed.

### Characteristics of the Covalent Bond ▼

- The distance between two bonded atoms at their minimum potential energy (the average distance between two bonded atoms) is the *bond length*. ▼
- In forming a covalent bond, the hydrogen atoms release energy. The same amount of energy must be added to separate the bonded atoms. ▼
- **Bond energy** is the energy required to break a chemical bond and form neutral isolated atoms.



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## Section 2 Covalent Bonding and Molecular Compounds

### Bond Length

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## Section 2 Covalent Bonding and Molecular Compounds

### Bond Energy

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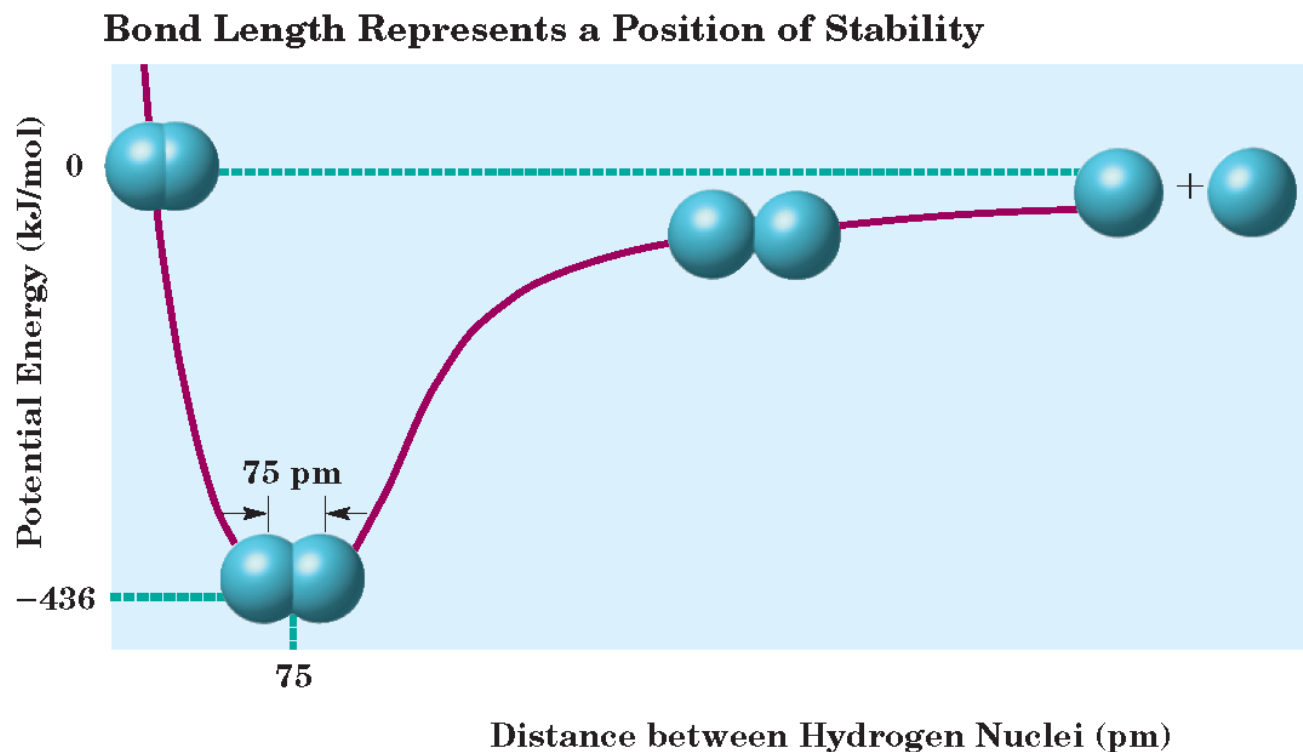
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## Section 2 Covalent Bonding and Molecular Compounds

### Bond Length and Stability

A bond forms when atoms are a certain distance from each other. At this distance, the atoms are in a low energy state. If they are closer together or farther apart, they will be in an unstable situation.



# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Bond Energies and Bond Lengths for Single Bonds

	Bond energy (kJ/mol)	Bond length (pm)		Bond energy (kJ/mol)	Bond length (pm)
<b>H—F</b>	570	92	<b>H—I</b>	299	161
<b>C—F</b>	552	138	<b>C—Br</b>	280	194
<b>O—O</b>	498	121	<b>Cl—Cl</b>	243	199
<b>H—H</b>	436	75	<b>C—I</b>	209	214
<b>H—Cl</b>	432	127	<b>Br—Br</b>	193	229
<b>C—Cl</b>	397	177	<b>F—F</b>	159	142
<b>H—Br</b>	366	141	<b>I—I</b>	151	266

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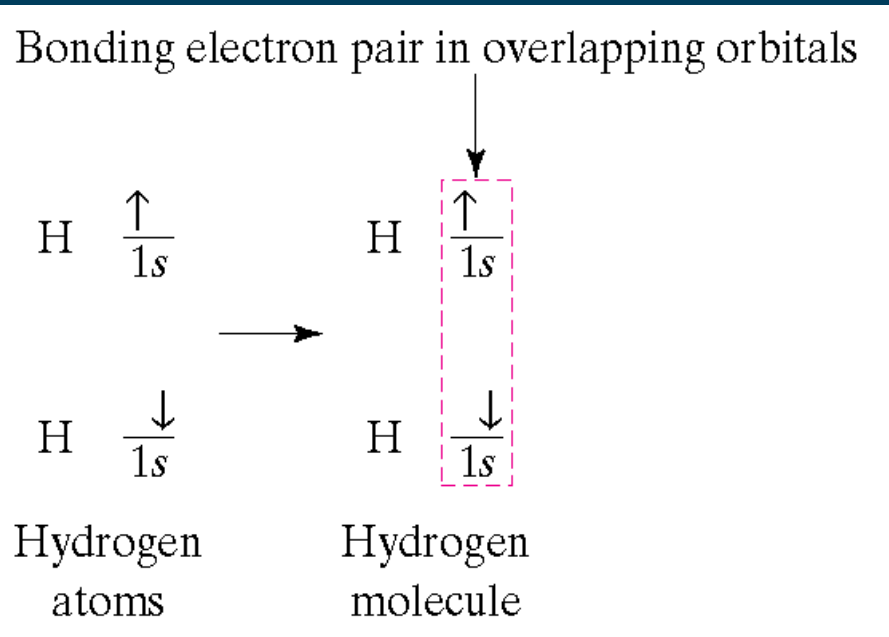
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# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Characteristics of the Covalent Bond ▾

- When two atoms form a covalent bond, their shared electrons form overlapping orbitals. ▾
- This achieves a noble-gas configuration. ▾
- The bonding of two hydrogen atoms allows each atom to have the stable electron configuration of helium,  $1s^2$ .



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### The Octet Rule ▼

- Noble gas atoms are unreactive because their electron configurations are especially stable. ▼
  - This stability results from the fact that the noble-gas atoms' outer  $s$  and  $p$  orbitals are completely filled by a total of eight electrons. ▼
- Other atoms can fill their outermost  $s$  and  $p$  orbitals by sharing electrons through covalent bonding. ▼
- Such bond formation follows the *octet rule*: Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an octet of electrons in its highest energy level.



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## Section 2 Covalent Bonding and Molecular Compounds

### The Octet Rule

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### The Octet Rule, *continued* Exceptions to the Octet Rule ▼

- Exceptions to the octet rule include those for atoms that cannot fit eight electrons, and for those that can fit more than eight electrons, into their outermost orbital. ▼
  - **Hydrogen** forms bonds in which it is surrounded by only two electrons. ▼
  - **Boron** has just three valence electrons, so it tends to form bonds in which it is surrounded by six electrons. ▼
  - Main-group elements in **Periods 3** and up can form bonds with *expanded valence*, involving *more* than eight electrons.



# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Electron-Dot Notation ▼

- To keep track of valence electrons, it is helpful to use electron-dot notation. ▼
- Electron-dot notation** is an electron-configuration notation in which only the valence electrons of an atom of a particular element are shown, indicated by dots placed around the element's symbol. The inner-shell electrons are not shown.

Number of valence electrons	Electron-dot notation	Example
1	X·	Na·
2	·X·	·Mg·
3	·X·	·B·
4	·X·	·C·
5	·X·	·N·
6	·X·	·O·
7	·X·	·F·
8	·X·	·Ne·



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# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Electron-Dot Notation

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### Electron-Dot Notation, *continued*

#### Sample Problem B ▼

- Write the electron-dot notation for hydrogen. ▼
- Write the electron-dot notation for nitrogen.



### Electron-Dot Notation, *continued*

#### Sample Problem B Solution ▼

- a. A hydrogen atom has only one occupied energy level, the  $n = 1$  level, which contains a single electron. ▼



- b. The group notation for nitrogen's family of elements is  $ns^2np^3$ . ▼

Nitrogen has five valence electrons. ▼



### Lewis Structures ▼

- Electron-dot notation can also be used to represent molecules. ▼



- The pair of dots between the two symbols represents the shared electron pair of the hydrogen-hydrogen covalent bond. ▼
  - For a molecule of fluorine,  $\text{F}_2$ , the electron-dot notations of two fluorine atoms are combined. ▼



### Lewis Structures ▼

- The pair of dots between the two symbols represents the shared pair of a covalent bond.



- In addition, each fluorine atom is surrounded by three pairs of electrons that are not shared in bonds.



- An *unshared pair*, also called a *lone pair*, is a pair of electrons that is not involved in bonding and that belongs exclusively to one atom.



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## Section 2 Covalent Bonding and Molecular Compounds

### Lewis Structures

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## Section 2 Covalent Bonding and Molecular Compounds

### Lone Pair of Electrons

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### Lewis Structures ▼

- The pair of dots representing a shared pair of electrons in a covalent bond is often replaced by a long dash. ▼



- A **structural formula** indicates the kind, number, and arrangement, and bonds but not the unshared pairs of the atoms in a molecule. ▼





# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Structural Formula

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### Lewis Structures ▼

- The Lewis structures and the structural formulas for many molecules can be drawn if one knows the composition of the molecule and which atoms are bonded to each other. ▼
- A single covalent bond, or **single bond**, is a covalent bond in which one pair of electrons is shared between two atoms.



# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Lewis Structures, *continued*

#### Sample Problem C ▼

Draw the Lewis structure of iodomethane,  $\text{CH}_3\text{I}$ .



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### Lewis Structures, *continued*

#### Sample Problem C Solution ▼

1. *Determine the type and number of atoms in the molecule.* ▼

The formula shows one carbon atom, one iodine atom, and three hydrogen atoms. ▼

2. *Write the electron-dot notation for each type of atom in the molecule.* ▼

Carbon is from Group 14 and has four valence electrons. ▼

Iodine is from Group 17 and has seven valence electrons. ▼

Hydrogen has one valence electron. ▼



# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Lewis Structures, *continued*

#### Sample Problem C Solution, continued ▼

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

$$\text{C} \quad 1 \times 4e^{-} \quad = \quad 4e^{-}$$

$$\text{I} \quad 1 \times 7e^{-} \quad = \quad 7e^{-}$$

$$3\text{H} \quad 3 \times 1e^{-} \quad = \quad 3e^{-}$$

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$$14e^{-}$$



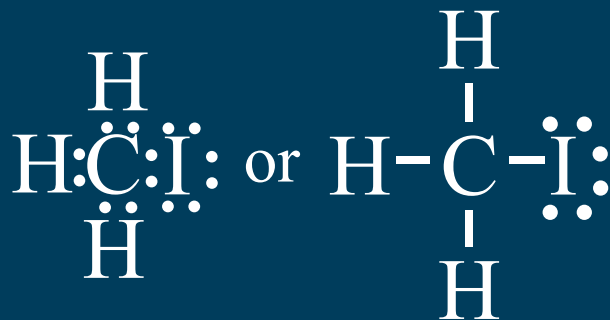
### Lewis Structures, *continued*

#### Sample Problem C Solution, *continued* ▼

4. *If carbon is present, it is the central atom. Otherwise, the least-electronegative atom is central. Hydrogen, is never central.* ▼

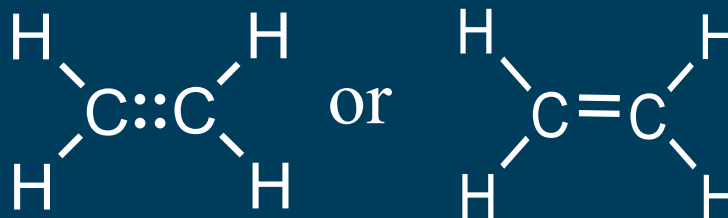


5. *Add unshared pairs of electrons to each nonmetal atom (except hydrogen) such that each is surrounded by eight electrons.* ▼



### Multiple Covalent Bonds ▾

- A double covalent bond, or simply a *double bond*, is a covalent bond in which two pairs of electrons are shared between two atoms. ▾
  - Double bonds are often found in molecules containing carbon, nitrogen, and oxygen. ▾
- A double bond is shown either by two side-by-side pairs of dots or by two parallel dashes. ▾



### Multiple Covalent Bonds ▼

- A triple covalent bond, or simply a *triple bond*, is a covalent bond in which three pairs of electrons are shared between two atoms. ▼

- example 1—diatomic nitrogen:



- example 2—ethyne,  $\text{C}_2\text{H}_2$ :





# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Comparing Single, Double, and Triple Bonds

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### Multiple Covalent Bonds ▼

- Double and triple bonds are referred to as **multiple bonds**, or multiple covalent bonds. ▼
  - In general, double bonds have greater bond energies and are shorter than single bonds. ▼
  - Triple bonds are even stronger and shorter than double bonds. ▼
- When writing Lewis structures for molecules that contain carbon, nitrogen, or oxygen, remember that multiple bonds between pairs of these atoms are possible.



### Drawing Lewis Structures with Many Atoms

#### 1. Gather information.

- Draw a Lewis structure for each atom in the compound. When placing valence electrons around an atom, place one electron on each side before pairing any electrons.
- Determine the total number of valence electrons in the compound.

#### 2. Arrange the atoms.

- Arrange the Lewis structure to show how the atoms bond in the molecule.
- Halogen and hydrogen atoms *often* bind to only one other atom and are *usually* at an end of the molecule.
- Carbon is often placed in the center of the molecule.
- You will find that, with the exception of carbon, the atom with the lowest electronegativity is often the central atom.

### Drawing Lewis Structures with Many Atoms

#### 3. Distribute the dots.

- Distribute the electron dots so that each atom, except for hydrogen, beryllium, and boron, satisfies the octet rule.

#### 4. Draw the bonds.

- Change each pair of dots that represents a shared pair of electrons to a long dash.

#### 5. Verify the structure.

- Count the number of electrons surrounding each atom. Except for hydrogen, beryllium, and boron, all atoms must satisfy the octet rule. Check that the number of valence electrons is still the same number you determined in step 1.

# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Multiple Covalent Bonds, *continued*

#### Sample Problem D ▼

Draw the Lewis structure for methanal,  $\text{CH}_2\text{O}$ , which is also known as formaldehyde.



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### Multiple Covalent Bonds, *continued*

#### Sample Problem D Solution ▼

1. *Determine the number of atoms of each element present in the molecule.* ▼

The formula shows one carbon atom, two hydrogen atoms, and one oxygen atom. ▼

2. *Write the electron-dot notation for each type of atom.* ▼

Carbon is from Group 14 and has four valence electrons. ▼

Oxygen, which is in Group 16, has six valence electrons. ▼

Hydrogen has only one valence electron. ▼



# Chapter 6

## Section 2 Covalent Bonding and Molecular Compounds

### Multiple Covalent Bonds, *continued*

#### Sample Problem D Solution, continued ▼

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

$$\text{C} \quad 1 \times 4e^{-} \quad = \quad 4e^{-}$$

$$\text{O} \quad 1 \times 6e^{-} \quad = \quad 6e^{-}$$

$$2\text{H} \quad 2 \times 1e^{-} \quad = \quad 2e^{-}$$

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$$12e^{-}$$



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### Multiple Covalent Bonds, *continued*

#### Sample Problem D Solution, *continued* ▼

4. *Arrange the atoms to form a skeleton structure for the molecule. Connect the atoms by electron-pair bonds.* ▼



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





### Multiple Covalent Bonds, *continued*

#### Sample Problem D Solution, *continued* ▼

6a. *Count the electrons in the Lewis structure to be sure that the number of valence electrons used equals the number available.* ▼

The structure has 14 electrons. The structure has two valence electrons too many. ▼

6b. *Subtract one or more lone pairs until the total number of valence electrons is correct.* ▼

*Move one or more lone electron pairs to existing bonds until the outer shells of all atoms are completely filled.*



## Multiple Covalent Bonds, *continued*

### Sample Problem D Solution, *continued* ▼

Subtract the lone pair of electrons from the carbon atom. Move one lone pair of electrons from the oxygen to the bond between carbon and oxygen to form a double bond.



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## Section 2 Covalent Bonding and Molecular Compounds

### Atomic Resonance

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# Chapter 6

## Section 3 Ionic Bonding and Ionic Compounds

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- Ionic Compounds
- Formation of Ionic Compounds
- A Comparison of Ionic and Molecular Compounds
- Polyatomic Ions

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# Chapter 6

## Section 3 Ionic Bonding and Ionic Compounds

### Objectives ▼

- **Compare** a chemical formula for a molecular compounds with one for an ionic compound. ▼
- **Discuss** the arrangements of ions in crystals. ▼
- **Define** *lattice energy* and explain its significance. ▼
- **List** and compare the distinctive properties of ionic and molecular compounds. ▼
- **Write** the Lewis structure for a polyatomic ion given the identity of the atoms combined and other appropriate information.



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### Ionic Compounds ▼

- Most of the rocks and minerals that make up Earth's crust consist of positive and negative ions held together by ionic bonding. ▼
  - **example:** table salt, NaCl, consists of sodium and chloride ions combined in a one-to-one ratio— $\text{Na}^+\text{Cl}^-$ —so that each positive charge is balanced by a negative charge. ▼
- An **ionic compound** is composed of positive and negative ions that are combined so that the numbers of positive and negative charges are equal.



### Ionic Compounds ▼

- Most ionic compounds exist as crystalline solids. ▼
- A crystal of any ionic compound is a three-dimensional network of positive and negative ions mutually attracted to each other. ▼
- In contrast to a molecular compound, an ionic compound is not composed of independent, neutral units that can be isolated.



### Ionic Compounds, *continued* ▼

- The chemical formula of an ionic compound represents not molecules, but the simplest ratio of the compound's ions. ▼
- A **formula unit** is the simplest collection of atoms from which an ionic compound's formula can be established.



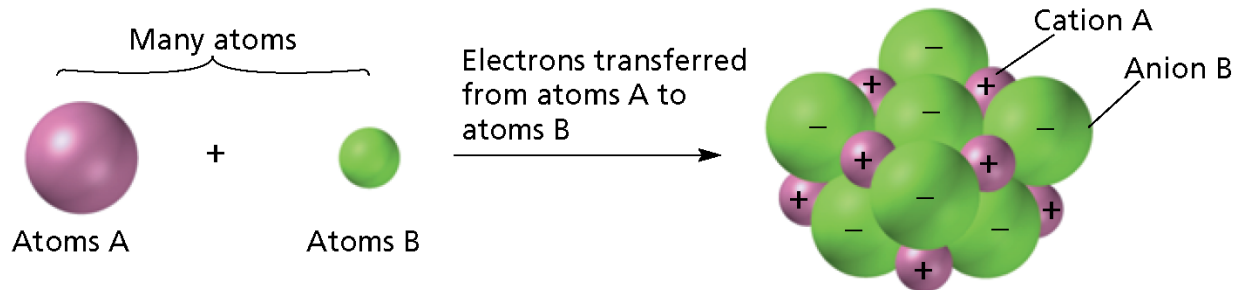


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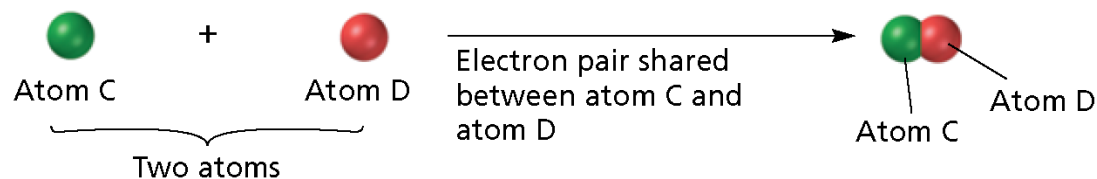
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### Ionic Vs. Covalent Bonding

**Ionic bonding**



**Covalent bonding**



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### Formation of Ionic Compounds ▼

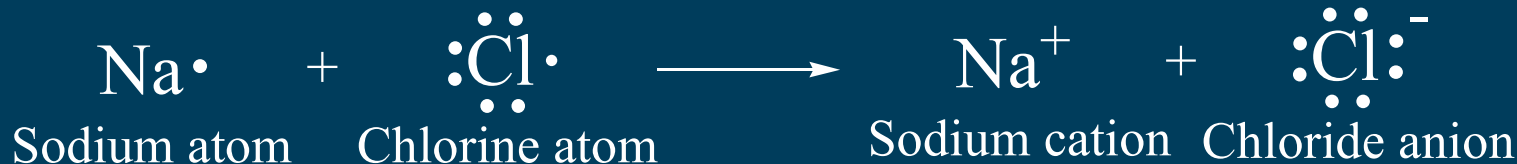


Sodium atom



Chlorine atom ▼

- The sodium atom has two valence electrons and the chlorine atom has seven valence electrons. ▼
- Atoms of sodium and other alkali metals easily lose one electron to form cations. ▼
- Atoms of chlorine and other halogens easily gain one electron to form anions. ▼



### Formation of Ionic Compounds, *continued* ▼

- In an ionic crystal, ions minimize their potential energy by combining in an orderly arrangement known as a *crystal lattice*. ▼
  - Attractive forces exist between oppositely charged ions within the lattice. ▼
  - Repulsive forces exist between like-charged ions within the lattice. ▼
- The combined attractive and repulsive forces within a crystal lattice determine: ▼
  - the distances between ions ▼
  - the pattern of the ions' arrangement in the crystal



# Chapter 6

## Section 3 Ionic Bonding and Ionic Compounds

### Characteristics of Ion Bonding in a Crystal Lattice

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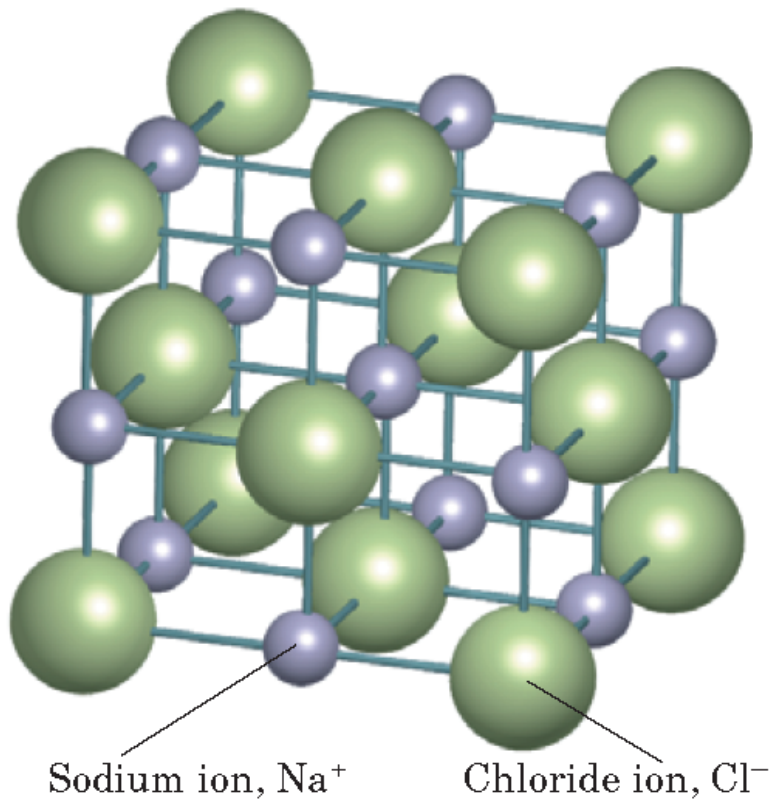
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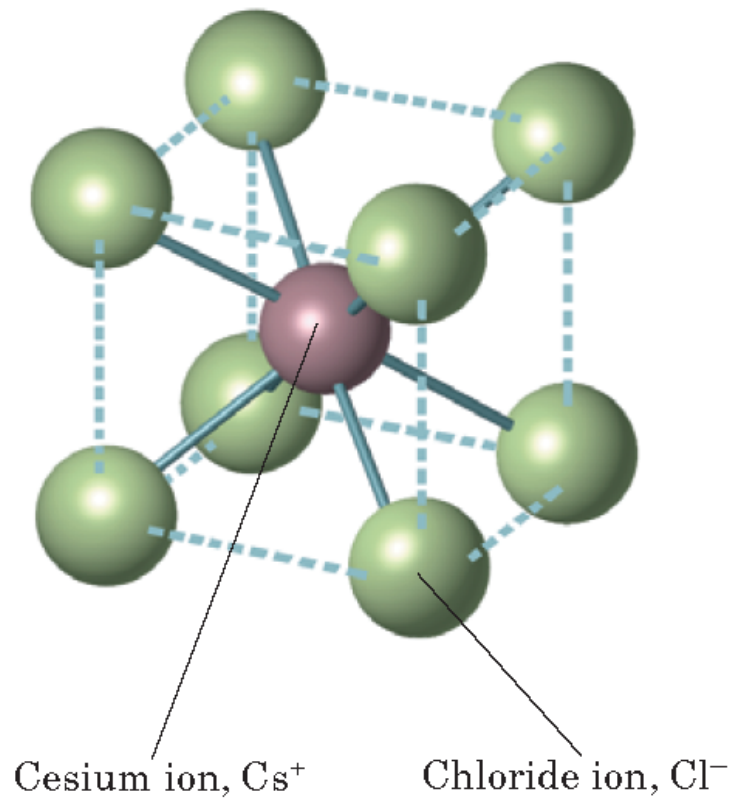
## Section 3 Ionic Bonding and Ionic Compounds

### NaCl and CsCl Crystal Lattices

NaCl lattice



CsCl lattice



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# Chapter 6

## Section 3 Ionic Bonding and Ionic Compounds

### Lattice Energy

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### A Comparison of Ionic and Molecular Compounds ▼

- The force that holds ions together in an ionic compound is a very strong electrostatic attraction. ▼
- In contrast, the forces of attraction *between* molecules of a covalent compound are much weaker. ▼
- This difference in the strength of attraction between the basic units of molecular and ionic compounds gives rise to different properties between the two types of compounds. ▼



### A Comparison of Ionic and Molecular Compounds, *continued* ▼

- Molecular compounds have relatively weak forces between individual molecules. ▼
  - They melt at low temperatures. ▼
- The strong attraction between ions in an ionic compound gives ionic compounds some characteristic properties, listed below. ▼
  - very high melting points ▼
  - hard but brittle ▼
  - not electrical conductors in the solid state, because the ions cannot move





# Chapter 6

## Section 3 Ionic Bonding and Ionic Compounds

### Melting and Boiling Points of Compounds

Compound name	Formula	Type of compound	Melting point		Boiling point	
			°C	K	°C	K
Magnesium fluoride	MgF <sub>2</sub>	ionic	1261	1534	2239	2512
Sodium chloride	NaCl	ionic	801	1074	1413	1686
Calcium iodide	CaI <sub>2</sub>	ionic	784	1057	1100	1373
Iodine monochloride	ICl	covalent	27	300	97	370
Carbon tetrachloride	CCl <sub>4</sub>	covalent	-23	250	77	350
Hydrogen fluoride	HF	covalent	-83	190	20	293
Hydrogen sulfide	H <sub>2</sub> S	covalent	-86	187	-61	212
Methane	CH <sub>4</sub>	covalent	-182	91	-164	109

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### How to Identify a Compound as Ionic

You can carry out the following procedures in a laboratory to determine if a substance is an ionic compound.

- Examine the substance. All ionic compounds are solid at room temperature. If the substance is a liquid or gas, then it is not an ionic compound. However, if it is a solid, then it *may* or *may not* be an ionic compound.
- Tap the substance gently. Ionic compounds are hard and brittle. If it is an ionic compound, then it should not break apart easily. If it does break apart, the substance should fracture into tinier crystals and not crumble into a powder.

### How to Identify a Compound as Ionic

- Heat a sample of the substance. Ionic compounds generally have high melting and boiling points.
- If the substance melts, use a conductivity apparatus to determine if the melted substance conducts electric current. Ionic compounds are good conductors of electric current in the liquid state.
- Dissolve a sample of the substance in water. Use a conductivity apparatus to see if it conducts electric current. Ionic compounds conduct electric current when dissolved in water.

# Chapter 6

## Section 3 Ionic Bonding and Ionic Compounds

### Comparing Ionic and Molecular Compounds

Click below to watch the Visual Concept.

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### Polyatomic Ions ▼

- Certain atoms bond covalently with each other to form a group of atoms that has both molecular and ionic characteristics. ▼
- A charged group of covalently bonded atoms is known as a **polyatomic ion**. ▼
- Like other ions, polyatomic ions have a charge that results from either a shortage or excess of electrons.



### Polyatomic Ions ▼

- An example of a polyatomic ion is the ammonium ion:  $\text{NH}_4^+$ . It is sometimes written as  $[\text{NH}_4]^+$  to show that the group of atoms *as a whole* has a charge of 1+. ▼
- The charge of the ammonium ion is determined as follows: ▼
  - The seven protons in the nitrogen atom plus the four protons in the four hydrogen atoms give the ammonium ion a total positive charge of 11+.



### Polyatomic Ions, *continued* ▼

- The charge of the ammonium ion is determined as follows, *continued*: ▼
  - When nitrogen and hydrogen atoms combine to form an ammonium ion, one of their electrons is lost, giving the polyatomic ion a total negative charge of  $10^-$ . ▼
  - The total charge is therefore  $(11^+) + (10^-) = 1^+$ .

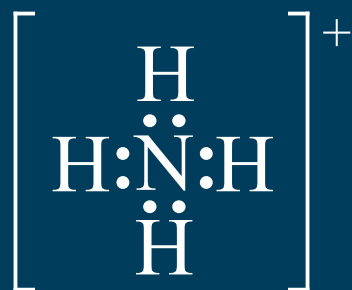


# Chapter 6

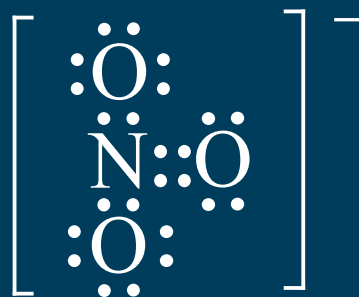
## Section 3 Ionic Bonding and Ionic Compounds

### Polyatomic Ions, *continued* ▼

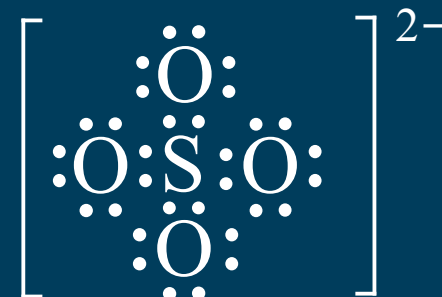
- Some examples of Lewis structures of polyatomic ions are shown below. ▼



Ammonium ion



Nitrate ion



Sulfate ion



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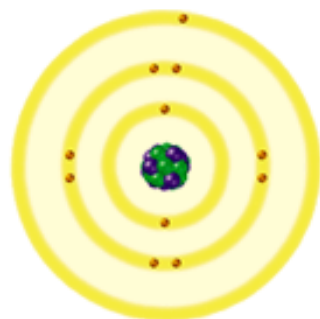
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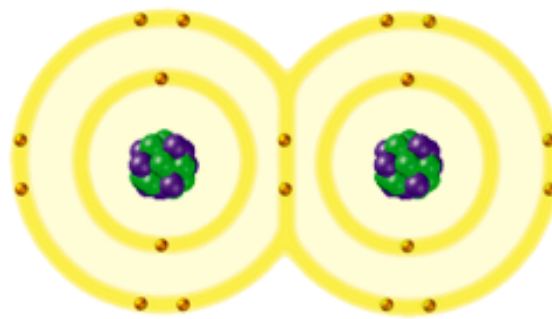
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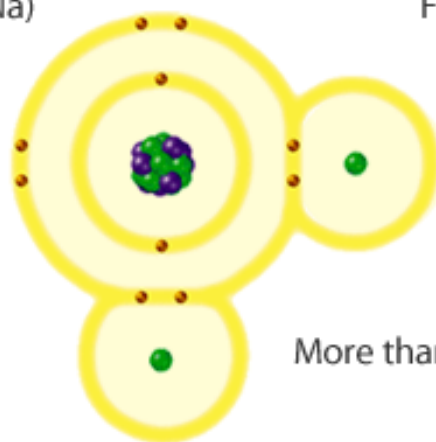
### Comparing Monatomic, Polyatomic, and Diatomic Structures



1 atom = monatomic  
Sodium (Na)



2 atoms = diatomic  
Fluorine (F<sub>2</sub>)



More than 2 atoms = polyatomic  
Water (H<sub>2</sub>O)



### Preview

- Objectives
- Metallic Bonding
- The Metallic-Bond Model

### Objectives ▼

- **Describe** the electron-sea model of metallic bonding, and explain why metals are good electrical conductors. ▼
- **Explain** why metal surfaces are shiny. ▼
- **Explain** why metals are malleable and ductile but ionic-crystalline compound are not.



### Metallic Bonding ▼

- Chemical bonding is different in metals than it is in ionic, molecular, or covalent-network compounds. ▼
- The unique characteristics of metallic bonding gives metals their characteristic properties, listed below. ▼
  - electrical conductivity ▼
  - thermal conductivity ▼
  - malleability ▼
  - ductility ▼
  - shiny appearance



### Metallic Bonding, *continued* ▼

- **Malleability** is the ability of a substance to be hammered or beaten into thin sheets. ▼
- **Ductility** is the ability of a substance to be drawn, pulled, or extruded through a small opening to produce a wire.

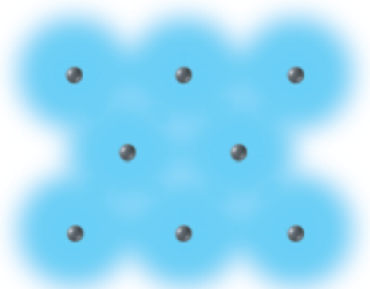


# Chapter 6

## Section 4 Metallic Bonding

### Properties of Substances with Metallic, Ionic, and Covalent Bonds

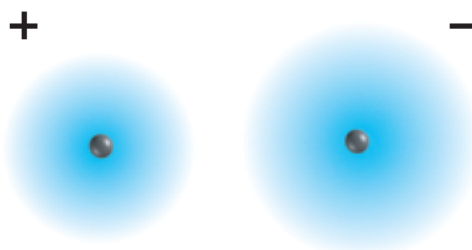
**Metallic**



Potassium



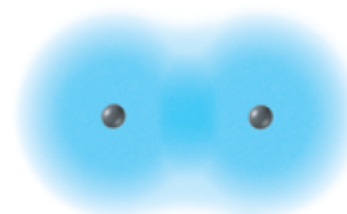
**Ionic**



Potassium chloride



**Covalent**



Chlorine



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### The Metallic-Bond Model ▼

- In a metal, the vacant orbitals in the atoms' outer energy levels overlap. ▼
- This overlapping of orbitals allows the outer electrons of the atoms to roam freely throughout the entire metal. ▼
- The electrons are *delocalized*, which means that they do not belong to any one atom but move freely about the metal's network of empty atomic orbitals.▼
- These mobile electrons form a *sea of electrons* around the metal atoms, which are packed together in a crystal lattice.



### The Metallic-Bond Model, *continued* ▾

- The chemical bonding that results from the attraction between metal atoms and the surrounding sea of electrons is called **metallic bonding**.





### Metallic Bonding

Click below to watch the Visual Concept.

[Visual Concept](#)

### Properties of Metals: Surface Appearance

Click below to watch the Visual Concept.

[Visual Concept](#)

### Properties of Metals: Malleability and Ductility

Click below to watch the Visual Concept.

[Visual Concept](#)

# Chapter 6

## Section 4 Metallic Bonding

### Properties of Metals: Electrical and Thermal Conductivity

Click below to watch the Visual Concept.

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### Preview

- Objectives
- Molecular Geometry
- VSEPR Theory
- Hybridization
- Intermolecular Forces

### Objectives ▼

- **Explain** VSEPR theory. ▼
- **Predict** the shapes of molecules or polyatomic ions using VSEPR theory. ▼
- **Explain** how the shapes of molecules are accounted for by hybridization theory.



### Objectives, *continued* ▼

- **Describe** dipole-dipole forces, hydrogen bonding, induced dipoles, and London dispersion forces and their effects on properties such as boiling and melting points. ▼
- **Explain** the shapes of molecules or polyatomic ions using VSEPR theory.



### Molecular Geometry ▼

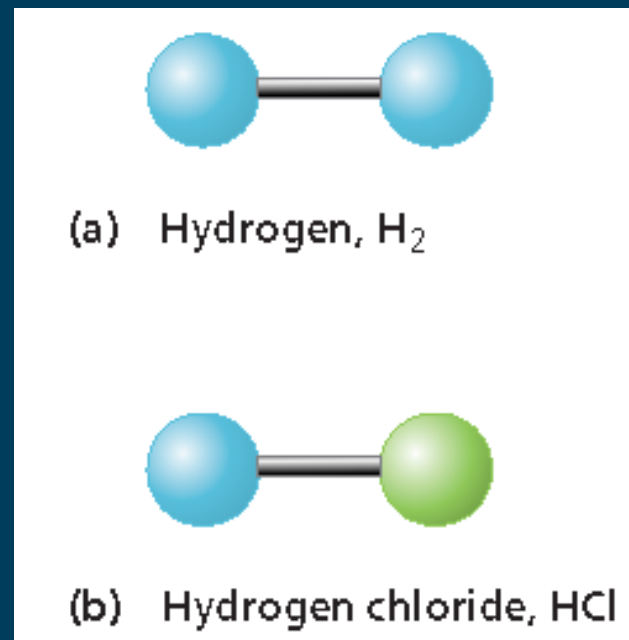
- The properties of molecules depend not only on the bonding of atoms but also on *molecular geometry*: the three-dimensional arrangement of a molecule's atoms. ▼
- The polarity of each bond, along with the geometry of the molecule, determines *molecular polarity*, or the uneven distribution of molecular shape. ▼
- Molecular polarity strongly influences the forces that act *between* molecules in liquids and solids. ▼
- A chemical formula, by itself, reveals little information about a molecule's geometry.





### VSEPR Theory ▼

- As shown at right, diatomic molecules, like those of **(a)** hydrogen,  $\text{H}_2$ , and **(b)** hydrogen chloride,  $\text{HCl}$ , can only be linear because they consist of only two atoms. ▼
- To predict the geometries of more-complicated molecules, one must consider the locations of all electron pairs surrounding the bonding atoms. This is the basis of VSEPR theory.



## VSEPR Theory ▼

- The abbreviation VSEPR (say it “VES-pur”) stands for “valence-shell electron-pair repulsion.” ▼
- **VSEPR theory** states that repulsion between the sets of valence-level electrons surrounding an atom causes these sets to be oriented as far apart as possible. ▼
  - **example:**  $\text{BeF}_2$  ▼
  - The central beryllium atom is surrounded by only the two electron pairs it shares with the fluorine atoms. ▼
  - According to VSEPR, the shared pairs will be as far away from each other as possible, so the bonds to fluorine will be  $180^\circ$  apart from each other. ▼
  - The molecule will therefore be linear:



### VSEPR Theory ▼

- Representing the central atom in a molecule by  $A$  and the atoms bonded to the central atom by  $B$ , then according to VSEPR theory,  $\text{BeF}_2$  is an example of an  $\text{AB}_2$  molecule, which is linear. ▼
- In an  $\text{AB}_3$  molecule, the three  $A\text{--}B$  bonds stay farthest apart by pointing to the corners of an equilateral triangle, giving  $120^\circ$  angles between the bonds. ▼
- In an  $\text{AB}_4$  molecule, the distance between electron pairs is maximized if each  $A\text{--}B$  bond points to one of four corners of a tetrahedron.



### VSEPR and Basic Molecular Shapes

Click below to watch the Visual Concept.

[Visual Concept](#)

### VSEPR Theory, *continued*

#### Sample Problem E ▼

Use VSEPR theory to predict the molecular geometry of boron trichloride,  $\text{BCl}_3$ .



### VSEPR Theory, *continued*

#### Sample Problem E Solution ▼

First write the Lewis structure for  $\text{BCl}_3$ . ▼

Boron is in Group 13 and has three valence electrons. ▼



Chlorine is in Group 17, so each chlorine atom has seven valence electrons. ▼



## VSEPR Theory, *continued*

### Sample Problem E Solution, *continued* ▼

The total number of electrons is calculated as shown below. ▼

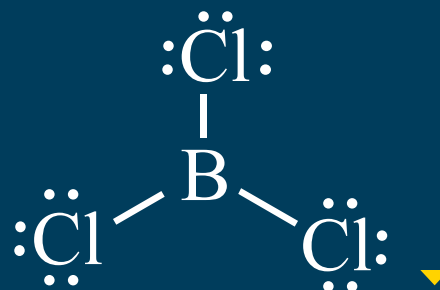
$$\begin{array}{rcl}
 \mathbf{B} & 1 \times 3e^{-} & = 3e^{-} \\
 \mathbf{3Cl} & 3 \times 7e^{-} & = 21e^{-} \\
 & \hline
 & & 24e^{-} \quad \blacktriangledown
 \end{array}$$

The following Lewis structure uses all 24 electrons. ▼



## VSEPR Theory, *continued*

### Sample Problem E Solution, *continued* ▼



Boron trichloride is an  $\text{AB}_3$  type of molecule. ▼

Its geometry should therefore be **trigonal-planar**.





## VSEPR Theory, *continued* ▼

- VSEPR theory can also account for the geometries of molecules with unshared electron pairs. ▼
  - **examples:** ammonia,  $\text{NH}_3$ , and water,  $\text{H}_2\text{O}$ . ▼
- The Lewis structure of ammonia shows that the central nitrogen atom has an unshared electron pair: ▼



- VSEPR theory postulates that the lone pair occupies space around the nitrogen atom just as the bonding pairs do.

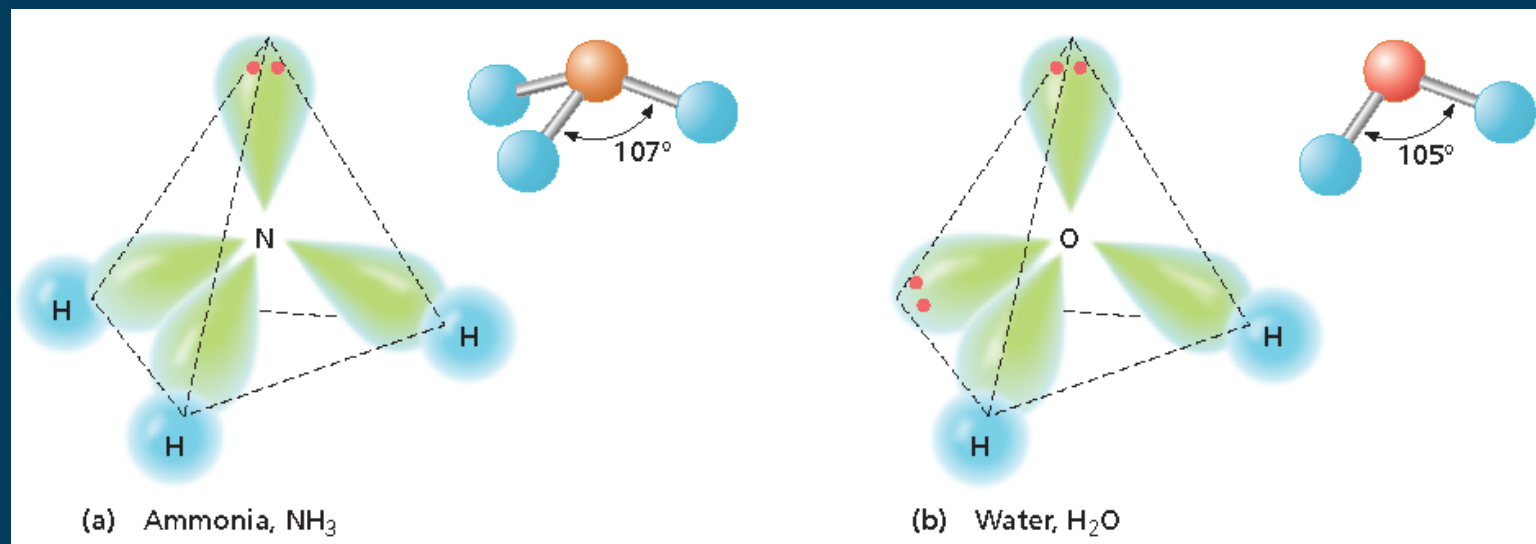


### VSEPR Theory, *continued* ▼

- Taking into account its unshared electron pair,  $\text{NH}_3$  takes a tetrahedral shape, as in an  $\text{AB}_4$  molecule. ▼
- The shape of a molecule refers to the *positions of atoms only*. ▼
- The geometry of an ammonia molecule is that of a pyramid with a triangular base. ▼
- $\text{H}_2\text{O}$  has two unshared pairs, and its molecular geometry takes the shape of a “bent,” or angular, molecule.



### VSEPR Theory, *continued* ▼



- Unshared electron pairs repel other electron pairs more strongly than bonding pairs do. ▼
  - This is why the bond angles in ammonia and water are somewhat less than the  $109.5^\circ$  bond angles of a perfectly tetrahedral molecule.

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### VSEPR and Lone Electron Pairs

Click below to watch the Visual Concept.

[Visual Concept](#)

### VSEPR Theory, *continued* ▼



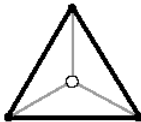
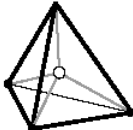
- The same basic principles of VSEPR theory that have been described can be used to determine the geometry of several additional types of molecules, such as  $AB_2E$ ,  $AB_2E_2$ ,  $AB_5$ , and  $AB_6$ . ▼
  - Treat double and triple bonds the same way as single bonds. ▼
  - Treat polyatomic ions similarly to molecules. ▼
  - The next slide shows several more examples of molecular geometries determined by VSEPR theory.



# Chapter 6

## Section 5 Molecular Geometry

### VSEPR and Molecular Geometry

	Molecular shape	Atoms bonded to central atom	Lone pairs of electrons	Formula example	Lewis structure
Linear		2	0	BeF <sub>2</sub>	$\text{:}\ddot{\text{F}}\text{--Be--}\ddot{\text{F}}\text{:}$
Bent		2	1	SnCl <sub>2</sub>	$\begin{array}{c} \ddot{\text{Sn}} \\ \diagdown \quad \diagup \\ \text{:}\ddot{\text{Cl}} \quad \ddot{\text{Cl}}\text{:} \end{array}$
Trigonal-planar		3	0	BF <sub>3</sub>	$\begin{array}{c} \ddot{\text{F}} \quad \ddot{\text{F}} \\ \diagdown \quad \diagup \\ \text{B} \\ \diagup \\ \text{:}\ddot{\text{F}}\text{:} \end{array}$
Tetrahedral		4	0	CH <sub>4</sub>	$\begin{array}{c} \text{H} \\   \\ \text{H--C--H} \\   \\ \text{H} \end{array}$

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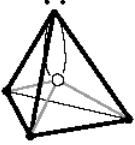
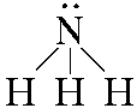
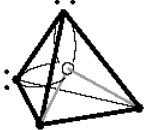
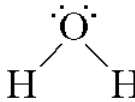
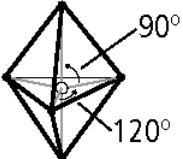
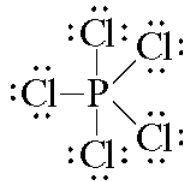
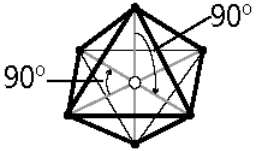
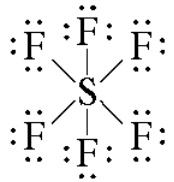
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# Chapter 6

## Section 5 Molecular Geometry

### VSEPR and Molecular Geometry

	Molecular shape	Atoms bonded to central atom	Lone pairs of electrons	Formula example	Lewis structure
Trigonal-pyramidal		3	1	NH <sub>3</sub>	
Bent		2	2	H <sub>2</sub> O	
Trigonal-bipyramidal		5	0	PCl <sub>5</sub>	
Octahedral		6	0	SF <sub>6</sub>	

### VSEPR Theory, *continued*

#### Sample Problem F ▼

- Use VSEPR theory to predict the shape of a molecule of carbon dioxide,  $\text{CO}_2$ . ▼
- Use VSEPR theory to predict the shape of a chlorate ion,  $\text{ClO}_3^-$ .





### VSEPR Theory, *continued*

#### Sample Problem F Solution ▼

a. Draw the Lewis structure of carbon dioxide. ▼



There are two carbon-oxygen double bonds and no unshared electron pairs on the carbon atom. ▼

This is an AB<sub>2</sub> molecule, which is ▼

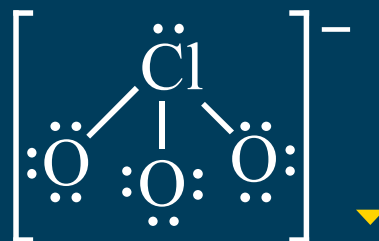
linear.



## VSEPR Theory, *continued*

### Sample Problem F Solution, *continued* ▼

b. Draw the Lewis structure of the chlorate ion. ▼



There are three oxygen atoms bonded to the central chlorine atom, which has an unshared electron pair. ▼

This is an AB<sub>3</sub>E molecule, which is ▼  
trigonal-pyramidal.



### Hybridization ▼

- VSEPR theory is useful for predicting and explaining the shapes of molecules. ▼
- A step further must be taken to explain how the orbitals of an atom are rearranged when the atom forms covalent bonds. ▼
- For this purpose, we use the model of **hybridization**, which is the mixing of two or more atomic orbitals of similar energies on the same atom to produce new hybrid atomic orbitals of equal energies.



### Hybridization ▼

- Take the simple example of methane,  $\text{CH}_4$ . The carbon atom has four valence electrons, two in the  $2s$  orbital and two in  $2p$  orbitals. ▼
- Experiments have determined that a methane molecule is tetrahedral. How does carbon form four equivalent, tetrahedrally arranged, covalent bonds? ▼
  - Recall that  $s$  and  $p$  orbitals have different shapes. To achieve four equivalent bonds, carbon's  $2s$  and three  $2p$  orbitals *hybridize* to form four new, identical orbitals called  $sp^3$  orbitals. ▼
  - The superscript 3 on the  $p$  indicates that there are three  $p$  orbitals included in the hybridization. The superscript 1 on the  $s$  is left out, like in a chemical formula.



### Hybridization, *continued* ▼

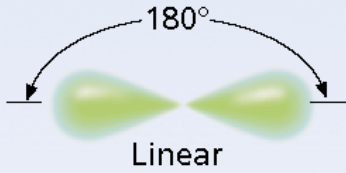
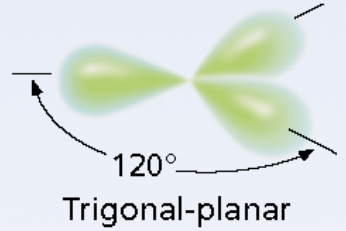
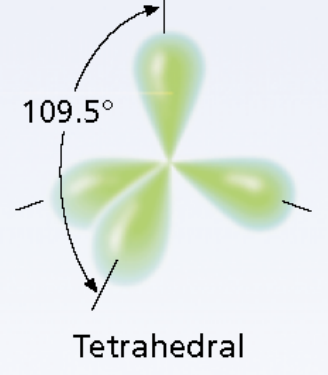
- The four ( $s + p + p + p$ ) hybrid orbitals in the  $sp^3$ -hybridized methane molecule are equivalent: they all have the same energy, which is greater than that of the  $2s$  orbital but less than that of the  $2p$  orbitals. ▼
- **Hybrid orbitals** are orbitals of equal energy produced by the combination of two or more orbitals on the same atom. ▼
- Hybridization explains the bonding and geometry of many molecules.



# Chapter 6

## Section 5 Molecular Geometry

### Geometry of Hybrid Orbitals

Atomic orbitals	Type of hybridization	Number of hybrid orbitals	Geometry
$s, p$	$sp$	2	 Linear
$s, p, p$	$sp^2$	3	 Trigonal-planar
$s, p, p, p$	$sp^3$	4	 Tetrahedral

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### Hybrid Orbitals

Click below to watch the Visual Concept.

[Visual Concept](#)

### Intermolecular Forces ▼

- The forces of attraction between molecules are known as *intermolecular forces*. ▼
  - The boiling point of a liquid is a good measure of the intermolecular forces between its molecules: the higher the boiling point, the stronger the forces between the molecules. ▼
  - Intermolecular forces vary in strength but are generally weaker than bonds between atoms within molecules, ions in ionic compounds, or metal atoms in solid metals. ▼
- Boiling points for ionic compounds and metals tend to be much higher than those for molecular substances: forces between molecules are weaker than those between metal atoms or ions.





## Comparing Ionic and Molecular Substances

Type of substance	Common use	State at room temperature	Melting point (°C)	Boiling point (°C)
<b>Ionic substances</b>				
Potassium chloride, KCl	salt substitute	solid	770	sublimes at 1500
Sodium chloride, NaCl	table salt	solid	801	1413
Calcium fluoride, CaF <sub>2</sub>	water fluoridation	solid	1423	2500
<b>Covalent substances</b>				
Methane, CH <sub>4</sub>	natural gas	gas	-182	-164
Ethyl acetate, CH <sub>3</sub> COOCH <sub>2</sub> CH <sub>3</sub>	fingernail polish	liquid	-84	77
Water, H <sub>2</sub> O	(many)	liquid	0	100
Heptadecane, C <sub>17</sub> H <sub>36</sub>	wax candles	solid	22	302

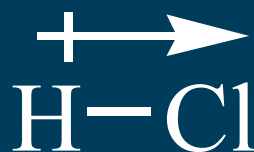
### Intermolecular Forces, *continued* ▼

- The strongest intermolecular forces exist between polar molecules. ▼
- Because of their uneven charge distribution, polar molecules have *dipoles*. A **dipole** is created by equal but opposite charges that are separated by a short distance. ▼
- The direction of a dipole is from the dipole's positive pole to its negative pole.



### Intermolecular Forces, *continued* ▾

- A dipole is represented by an arrow with its head pointing toward the negative pole and a crossed tail at the positive pole. The dipole created by a hydrogen chloride molecule is indicated as follows: ▾



### Intermolecular Forces, *continued* ▼

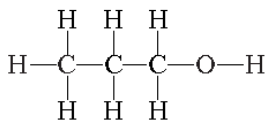
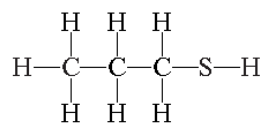
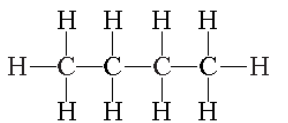
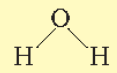
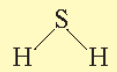
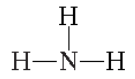
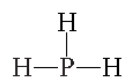
- The negative region in one polar molecule attracts the positive region in adjacent molecules. So the molecules all attract each other from opposite sides. ▼
- Such forces of attraction between polar molecules are known as *dipole-dipole forces*. ▼
- Dipole-dipole forces act at short range, only between nearby molecules. ▼
- Dipole-dipole forces explain, for example the difference between the boiling points of iodine chloride, I–Cl (97°C), and bromine, Br–Br (59°C).



# Chapter 6

## Section 5 Molecular Geometry

### Comparing Dipole-Dipole Forces

Substance	Boiling point (°C)	Polarity	State at room temperature	Structure
1-propanol, C <sub>3</sub> H <sub>7</sub> OH	97.4	polar	liquid	
1-propanethiol, C <sub>3</sub> H <sub>7</sub> SH	67.8	less polar	liquid	
Butane, C <sub>4</sub> H <sub>10</sub>	-0.5	nonpolar	gas	
Water, H <sub>2</sub> O	100.0	polar	liquid	
Hydrogen sulfide, H <sub>2</sub> S	-60.7	less polar	gas	
Ammonia, NH <sub>3</sub>	-33.35	polar	gas	
Phosphine, PH <sub>3</sub>	-87.7	less polar	gas	

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### Dipole-Dipole Forces

Click below to watch the Visual Concept.

[Visual Concept](#)

### Intermolecular Forces, *continued* ▼

- A polar molecule can *induce* a dipole in a nonpolar molecule by temporarily attracting its electrons. ▼
- The result is a short-range intermolecular force that is somewhat weaker than the dipole-dipole force. ▼
  - Induced dipoles account for the fact that a nonpolar molecule, oxygen,  $O_2$ , is able to dissolve in water, a polar molecule.



### Dipole-Induced Dipole Interaction

Click below to watch the Visual Concept.

[Visual Concept](#)



### Intermolecular Forces, *continued* ▼

- Some hydrogen-containing compounds have unusually high boiling points. This is explained by a particularly strong type of dipole-dipole force. ▼
- In compounds containing H–F, H–O, or H–N bonds, the large electronegativity differences between hydrogen atoms and the atoms they are bonded to make their bonds highly polar. ▼
- This gives the hydrogen atom a positive charge that is almost half as large as that of a bare proton.



### Intermolecular Forces, *continued* ▼

- The small size of the hydrogen atom allows the atom to come very close to an unshared pair of electrons in an adjacent molecule. ▼
- The intermolecular force in which a hydrogen atom that is bonded to a highly electronegative atom is attracted to an unshared pair of electrons of an electronegative atom in a nearby molecule is known as **hydrogen bonding**.

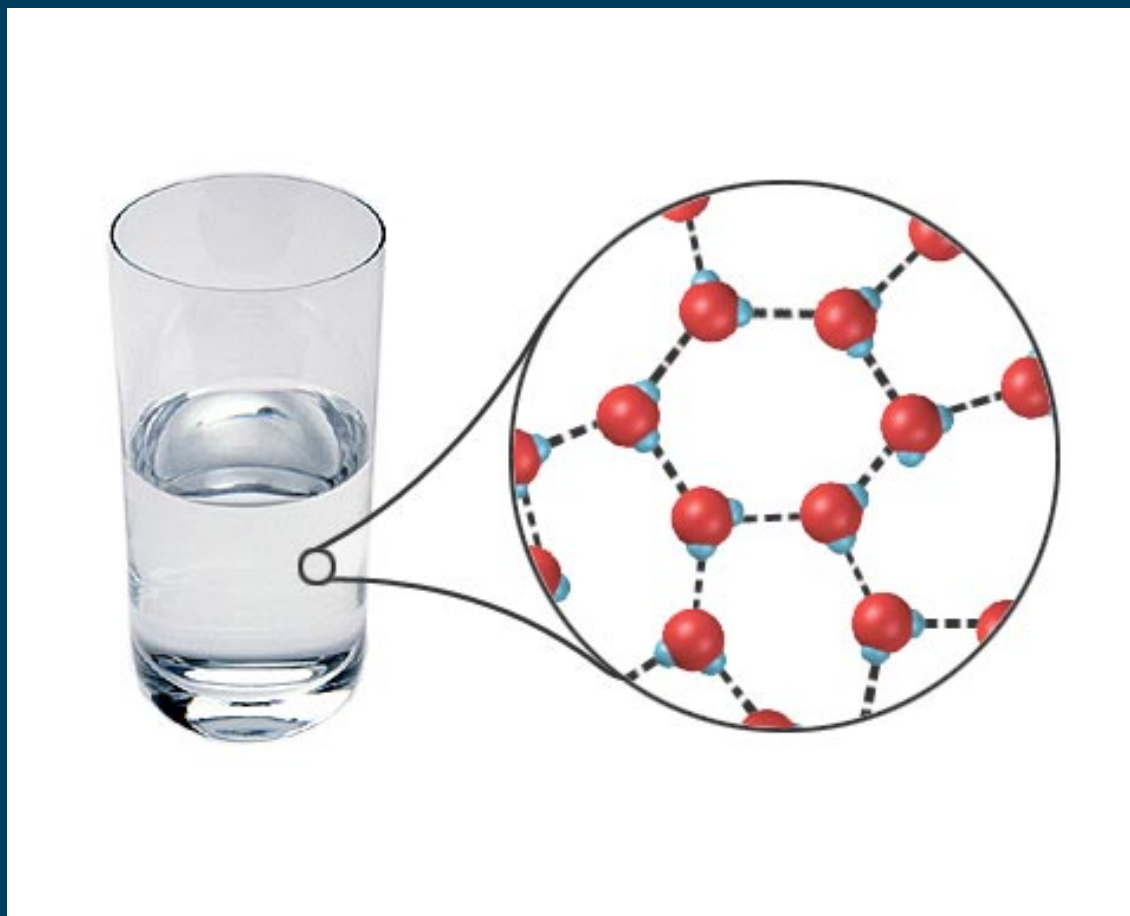


### Intermolecular Forces ▼

- Hydrogen bonds are usually represented by dotted lines connecting the hydrogen-bonded hydrogen to the unshared electron pair of the electronegative atom to which it is attracted. ▼
- An excellent example of hydrogen bonding is that which occurs between water molecules. The strong hydrogen bonding between water molecules accounts for many of water's characteristic properties.



### Hydrogen Bonding



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### Intermolecular Forces, *continued* London Dispersion Forces ▼

- Even noble gas atoms and nonpolar molecules can experience weak intermolecular attraction. ▼
- In any atom or molecule—polar or nonpolar—the electrons are in continuous motion. ▼
- As a result, at any instant the electron distribution may be uneven. A momentary uneven charge can create a positive pole at one end of an atom of molecule and a negative pole at the other.



### Intermolecular Forces, *continued*

#### London Dispersion Forces, *continued* ▼

- This temporary dipole can then induce a dipole in an adjacent atom or molecule. The two are held together for an instant by the weak attraction between temporary dipoles. ▼
- The intermolecular attractions resulting from the constant motion of electrons and the creation of instantaneous dipoles are called **London dispersion forces**. ▼
- Fritz London first proposed their existence in 1930.



### London Dispersion Force

Click below to watch the Visual Concept.

[Visual Concept](#)

# End of Chapter 6 Show

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