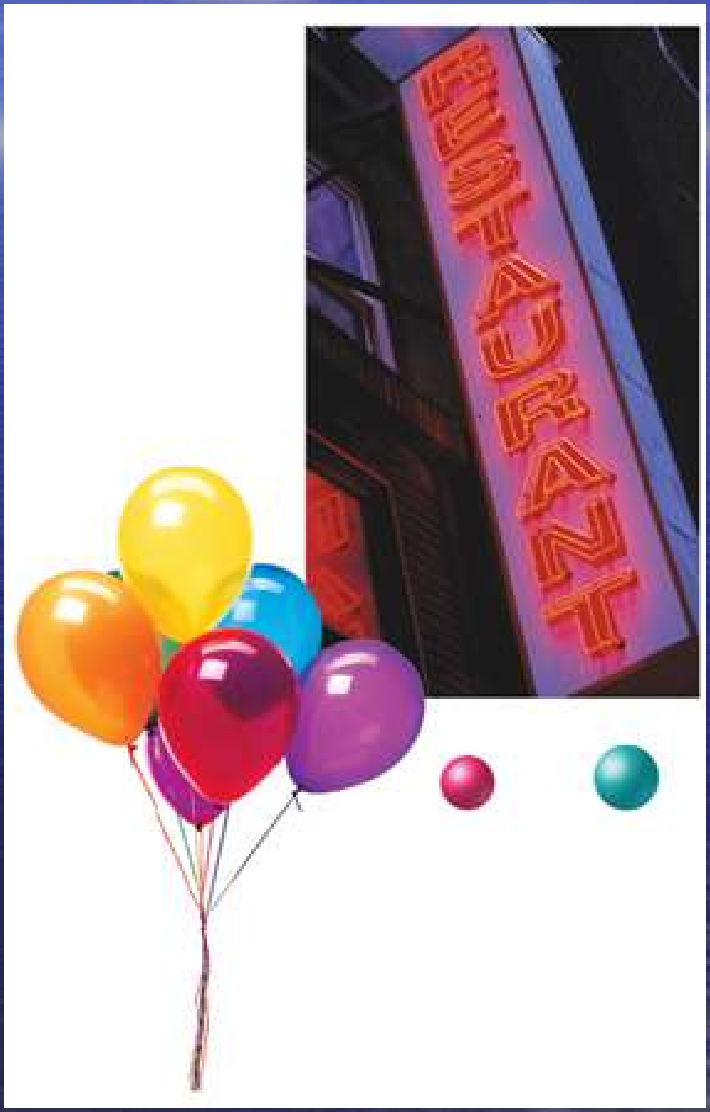
Chapter 8: Covalent Bonding

8.1 Molecular Compounds

8.1 Molecules and Molecular Compounds

- In nature, matter takes many forms. The noble gases, including helium and neon, are monatomic. That means they exist as single atoms.



8.1 Molecules and Molecular Compounds

- Some compounds are so different from ionic compounds that attractions between ions fail to explain their bonding.
- The atoms held together by sharing electrons are joined by a **covalent bond**.

8.1

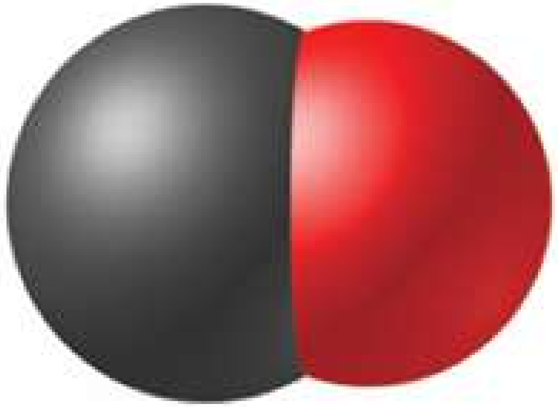
Molecules and Molecular Compounds

- A **molecule** is a neutral group of atoms joined together by covalent bonds. Air contains oxygen molecules.
- A **diatomic molecule** is a molecule consisting of two atoms. An oxygen molecule is a diatomic molecule.

8.1 Molecules and Molecular Compounds

- A compound composed of molecules is called a **molecular compound**. Water and carbon monoxide are molecular compounds.

Water (H_2O)



Carbon monoxide (CO)

8.1 Molecules and Molecular Changes

- Molecular compounds tend to have relatively lower melting and boiling points than ionic compounds.

Molecular Formulas

- A **molecular formula** is the chemical formula of a molecular compound.
  - **A molecular formula shows how many atoms of each element a molecule contains.**

8.1 Section Quiz.

- 1.Compared to ionic compounds, molecular compounds tend to have relatively
  a) low melting points and high boiling points.
  b) low melting points and low boiling points.
  c) high melting points and high boiling point.
  d) high melting points and low boiling points.

8.1 Section Quiz

- 2.A molecular compound usually consists of

  a) two metal atoms and a nonmetal atom.

  b) two nonmetal atoms and a metal atom.

  c) two or more metal atoms.

  d)two or more nonmetal atoms.

8.1 Section Quiz

- 3.A molecular formula shows

  a) how many atoms of each element a molecule contains.

  b) a molecule's structure.

  c) which atoms are bonded together.

  d) how atoms are arranged in space.

8.2 The Nature of Covalent Bonding

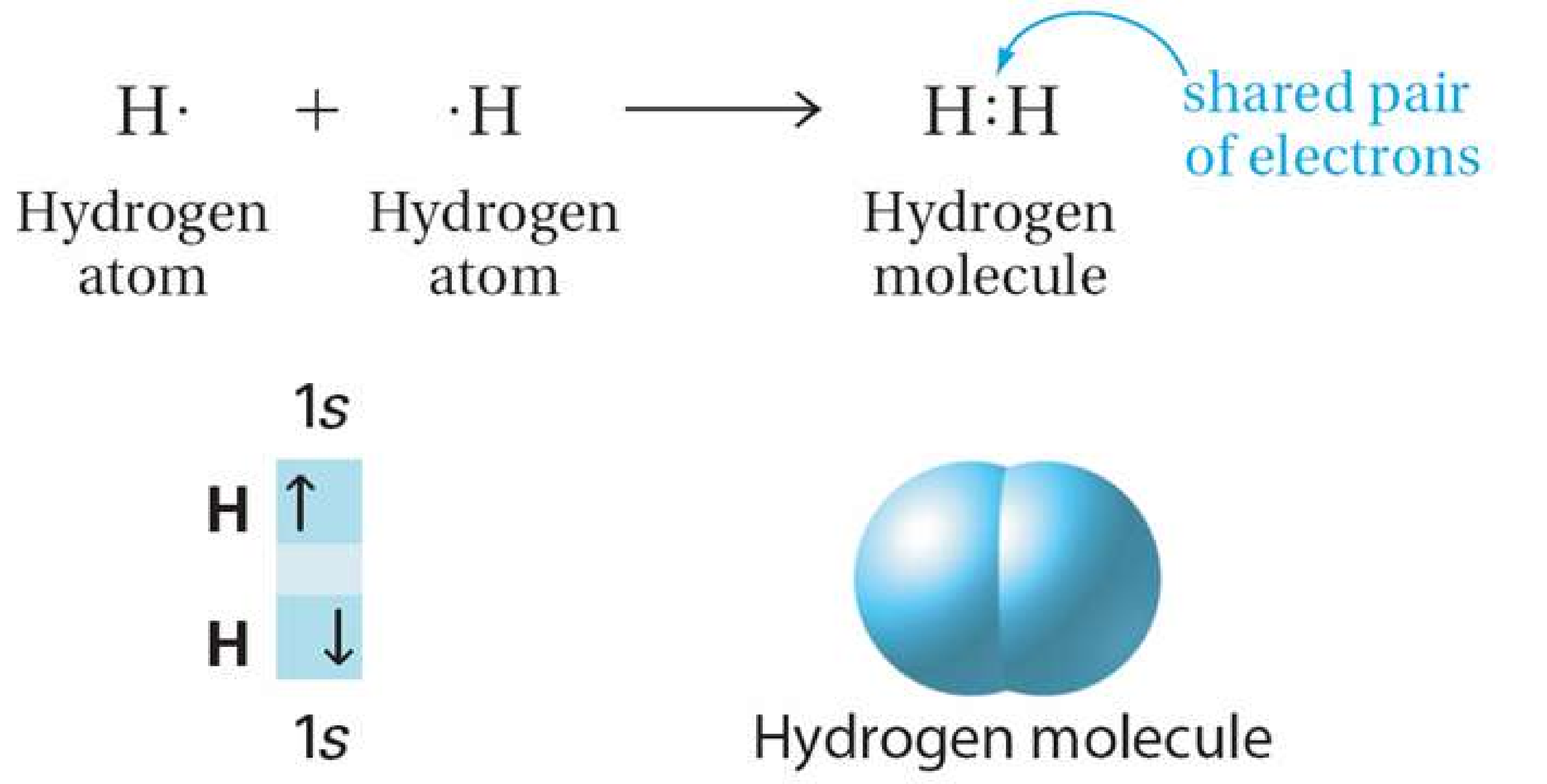
8.2

The Octet Rule in Covalent Bonding

- In covalent bonds, electron sharing usually occurs so that atoms attain the electron configurations of noble gases.

8.2 Single Covalent Bonds

- Two atoms held together by sharing a pair of electrons are joined by a single covalent bond.



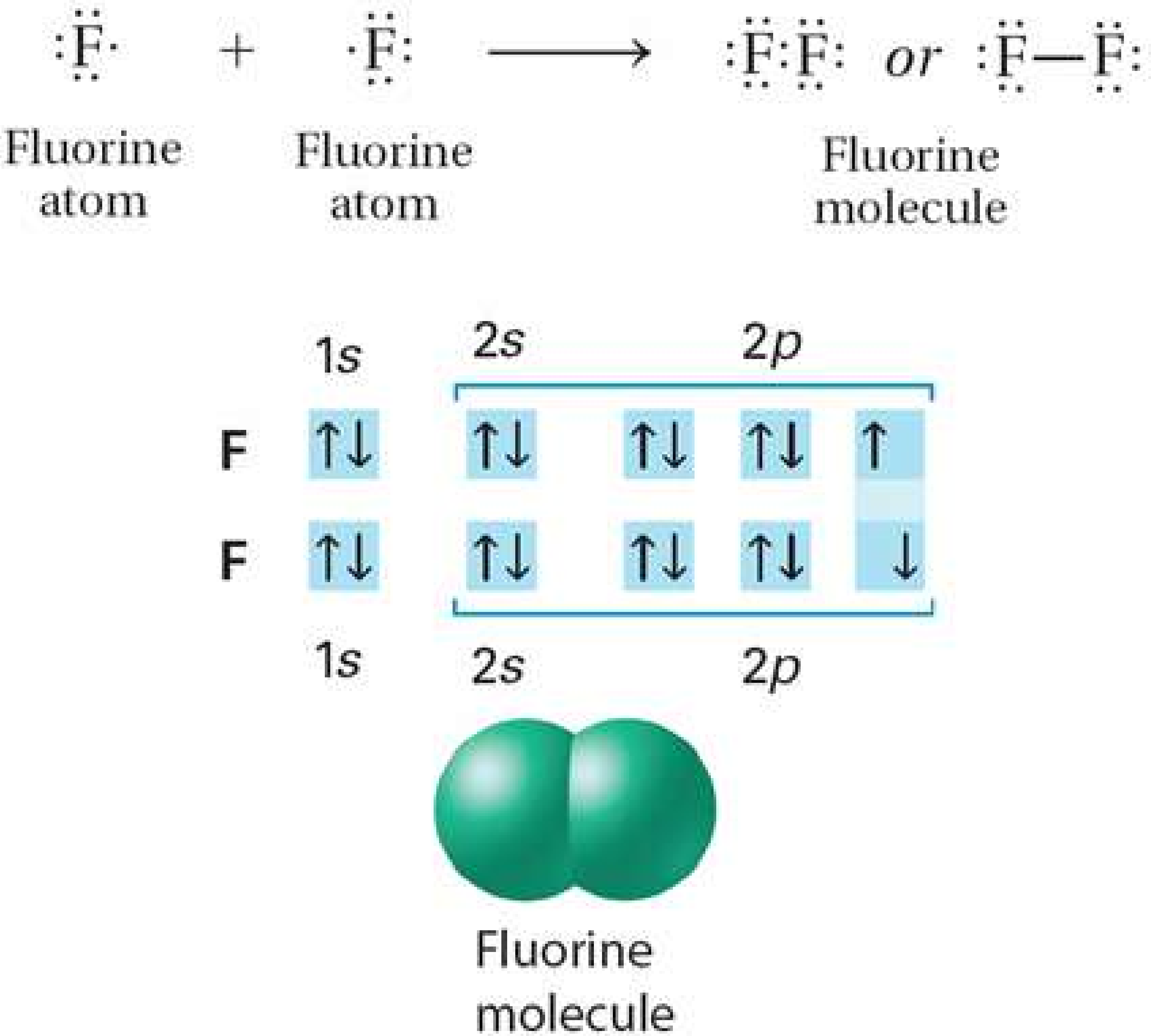
8.2

Single Covalent Bonds

- An electron dot structure such as H:H represents the shared pair of electrons of the covalent bond by two dots.
  - A structural formula represents the covalent bonds by dashes and shows the arrangement of covalently bonded atoms.

8.2

- The halogens form single covalent bonds in their diatomic molecules. Fluorine is one example.



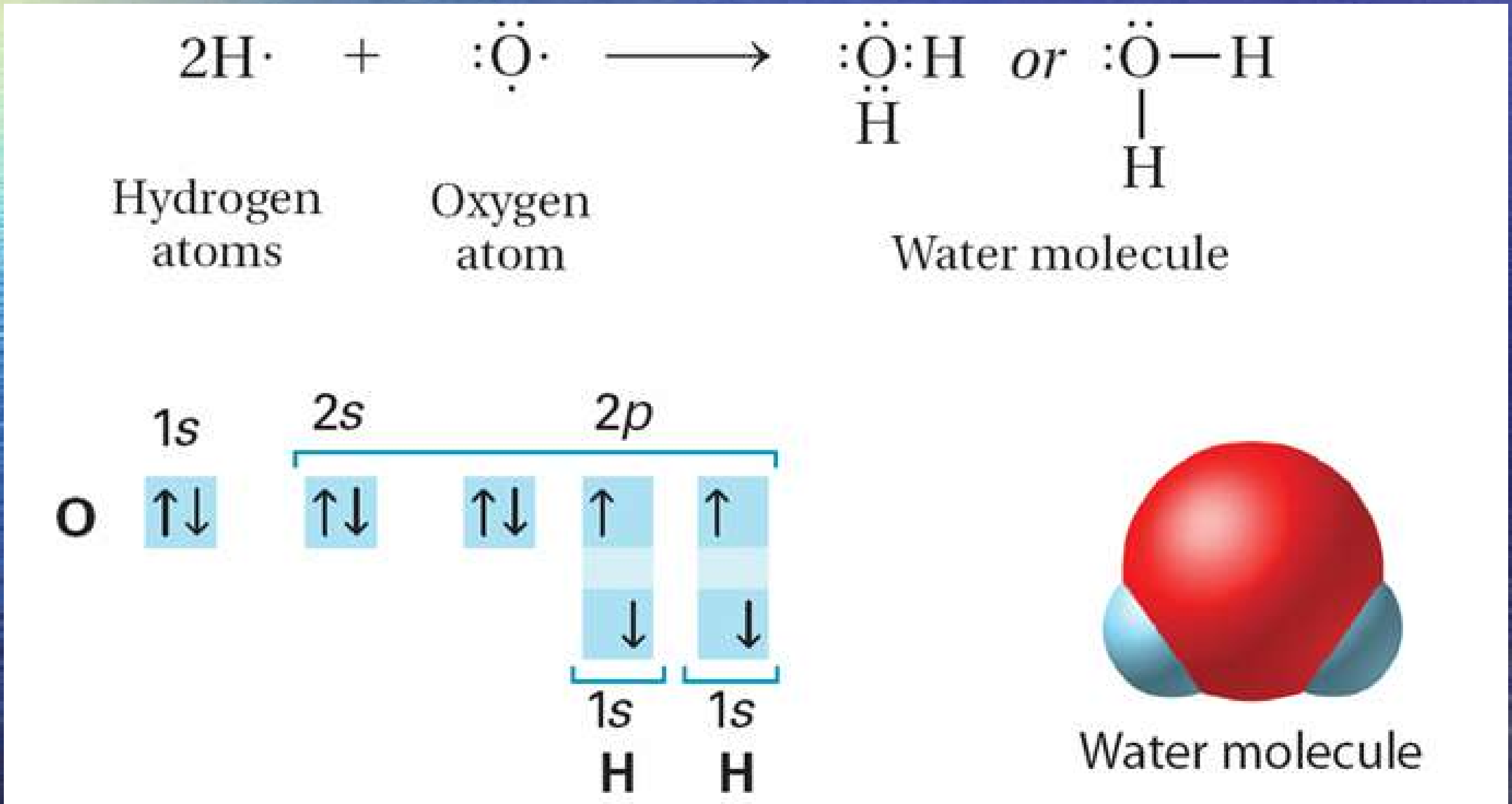
8.2

Single Covalent Bonds

- A pair of valence electrons that is not shared between atoms is called an **unshared pair**, also known as a lone pair or a nonbonding pair.

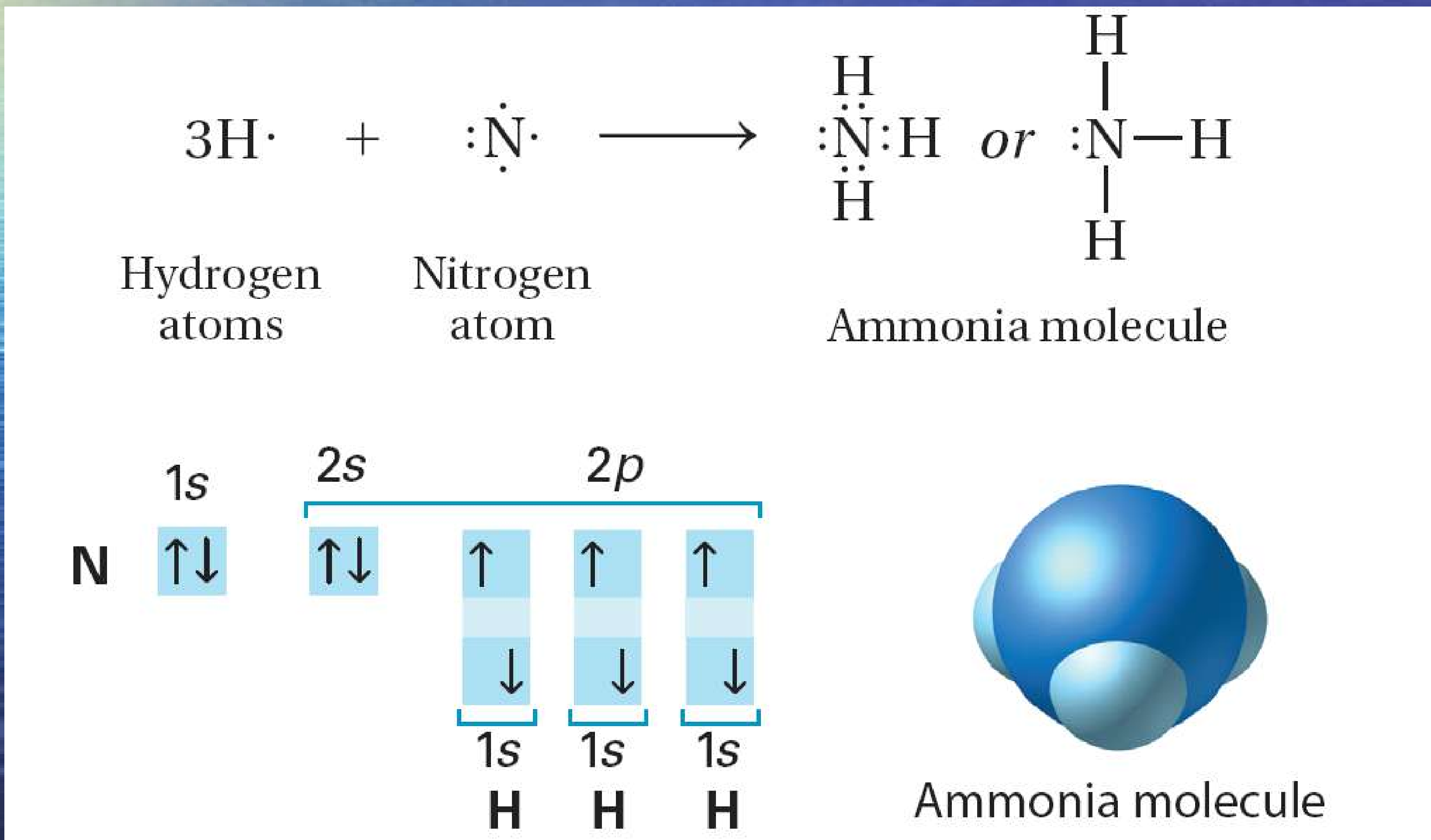
Single Covalent Bonds

- The hydrogen and oxygen atoms attain noble-gas configurations by sharing electrons.



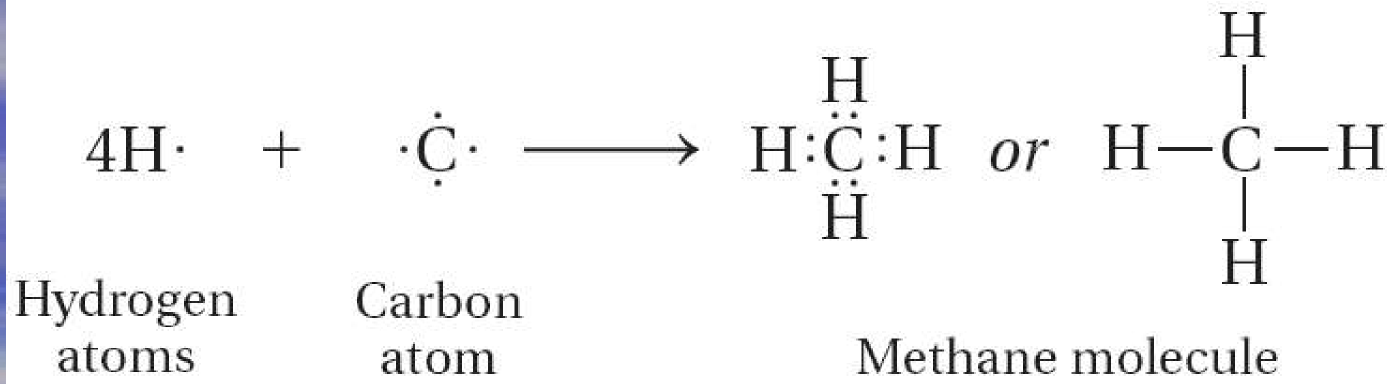
8.2 Single Covalent Bonds

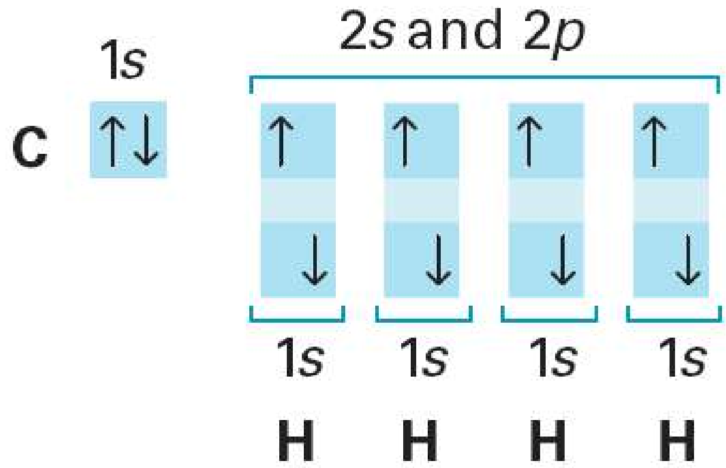
- The ammonia molecule has one unshared pair of electrons.

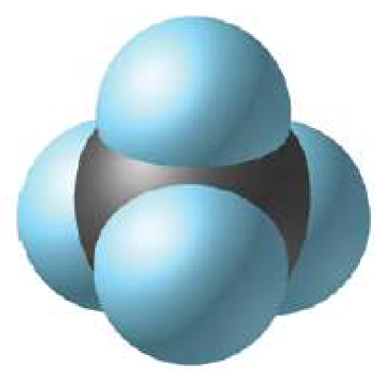


8.2

- Methane has no unshared pairs of electrons.







Methane molecule

Covalent Molecules

- Hydrogen shares 2 electrons (duet rule)
- Ex: H_2
- HH → H H
- Helium does not form bonds because its valence orbital is filled (He:)

Covalent Molecules

- Other nonmetals require 8 valence electrons (<u>octet rule</u>)
- Ex: F_2
- F F → F F

Steps for Lewis Structures

1) Add up all valence electrons for all atoms
2) Use one pair to form bonds between each pair of atoms (use a line to represent 2 dots)
3) Arrange remaining electrons to satisfy duet or octet rule

Drawing Lewis Diagrams

- **CF_4**

CONCEPTUAL PROBLEM 8.1

Drawing an Electron Dot Structure

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

Section Assessment

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

a. H_2O_2 **b.** PCl_3

Draw Lewis structures for:

a) H_2S

b) SiH_4

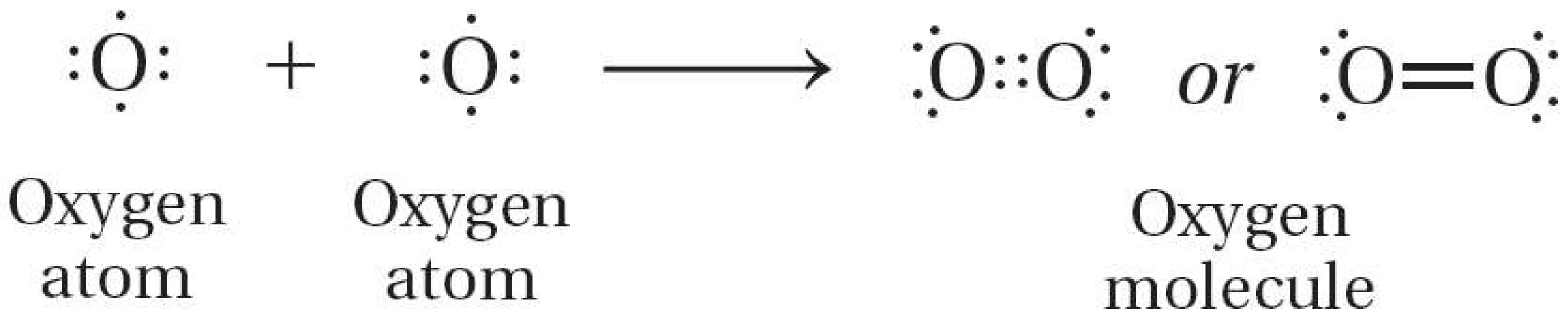
8.2 Double and Triple Covalent Bonds

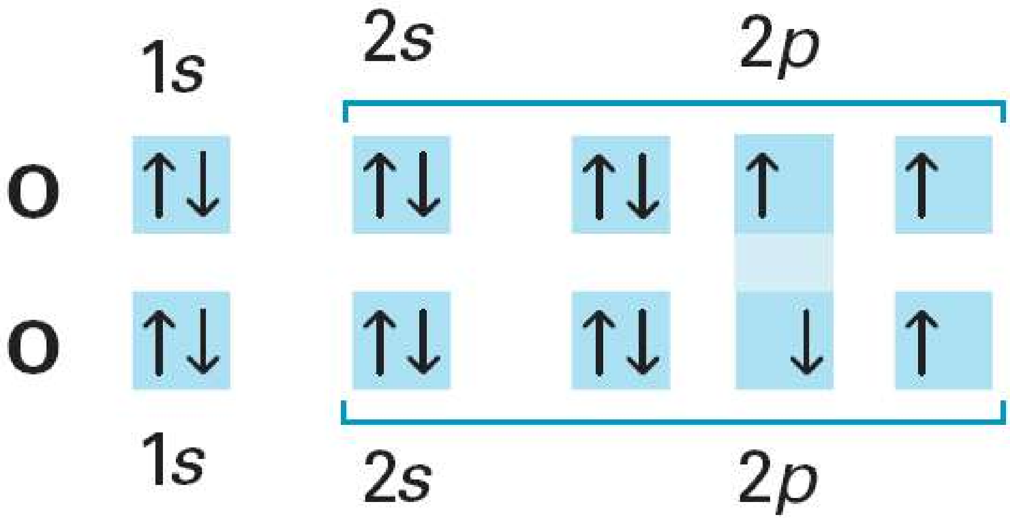
- Atoms form double or triple covalent bonds if they can attain a noble gas structure by sharing two pairs or three pairs of electrons.

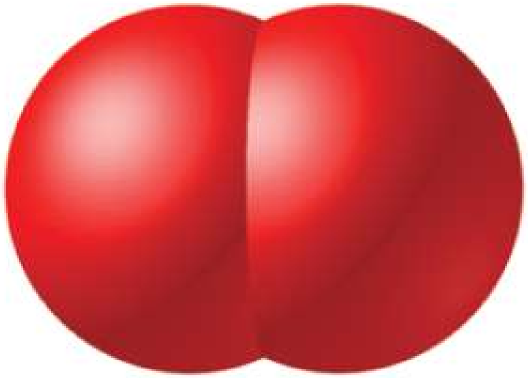
8.2 Double and Triple Covalent Bonds

- A bond that involves two shared pairs of electrons is a **double covalent bond**.
- A bond formed by sharing three pairs of electrons is a **triple covalent bond**.

8.2 Double and Triple Covalent Bonds



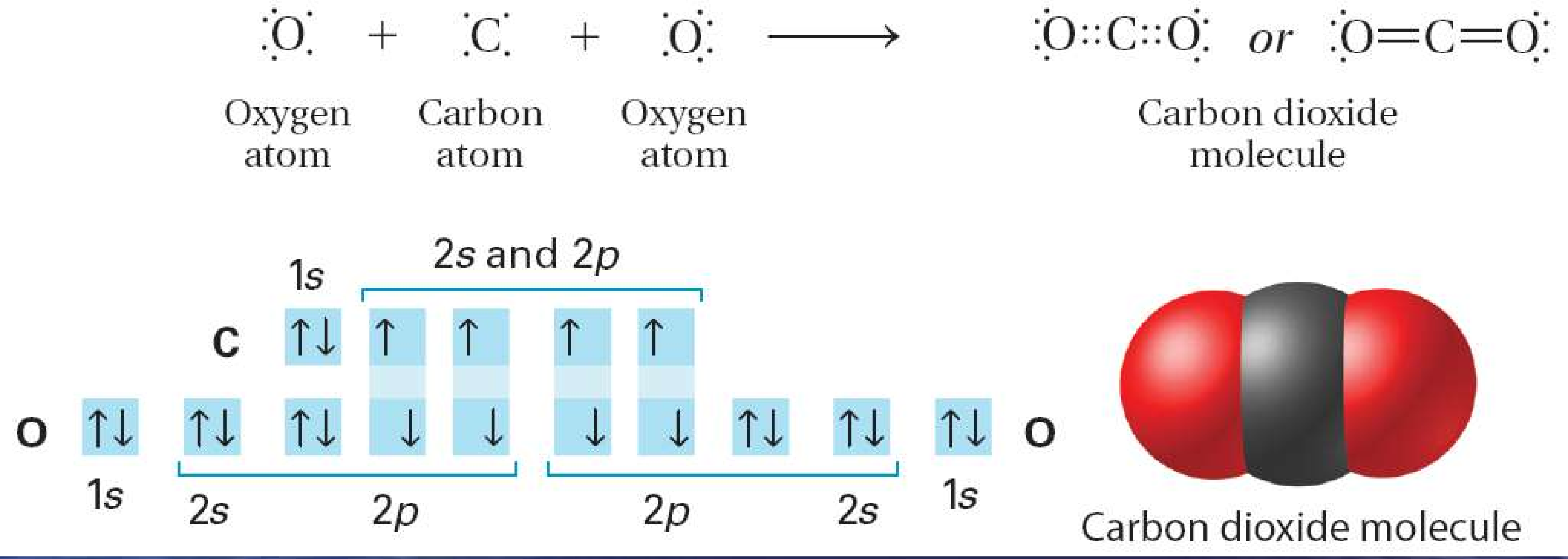




Oxygen molecule

8.2

- Carbon dioxide is an example of a triatomic molecule.



Example

- CH_2O
- formaldehyde

Practice

- C_2H_2
- Carbon disulfide

8.2 Polyatomic Ions

- A **polyatomic ion**, such as NH_4^+, is a tightly bound group of atoms that has a positive or negative charge and behaves as a unit.
  - Most plants need nitrogen that is already combined in a compound to grow.

Polyatomic Ions

- **ClO_4^-**

Polyatomic Ions

- **NH_4^+**

Example: CN^-

CONCEPTUAL PROBLEM 8.2

Drawing the Electron Dot Structure of a Polyatomic Ion

The polyatomic hydronium ion (H_3O^+), which is found in acidic mixtures such as lemon juice, contains a coordinate covalent bond. The H_3O^+ ion forms when a hydrogen ion is attracted to an unshared electron pair in a water molecule. Draw the electron dot structure of the hydronium ion.

Section Assessment

10. Draw the electron dot structure of the polyatomic boron tetrafluoride anion (BF_4^-).

Draw the Lewis structure:

1) ClO_3^-
2) ClO_4^-
3) HCO_3^-

8.2

Bond Dissociation Energies

- The energy required to break the bond between two covalently bonded atoms is known as the **bond dissociation energy**.
  - A large bond dissociation energy corresponds to a strong covalent bond.

8.2

Resonance

- A **resonance structure** is a structure that occurs when it is possible to draw two or more valid electron dot structures that have the same number of electron pairs for a molecule or ion.

Resonance Structures

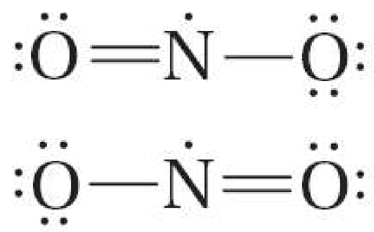
- **SO_3**

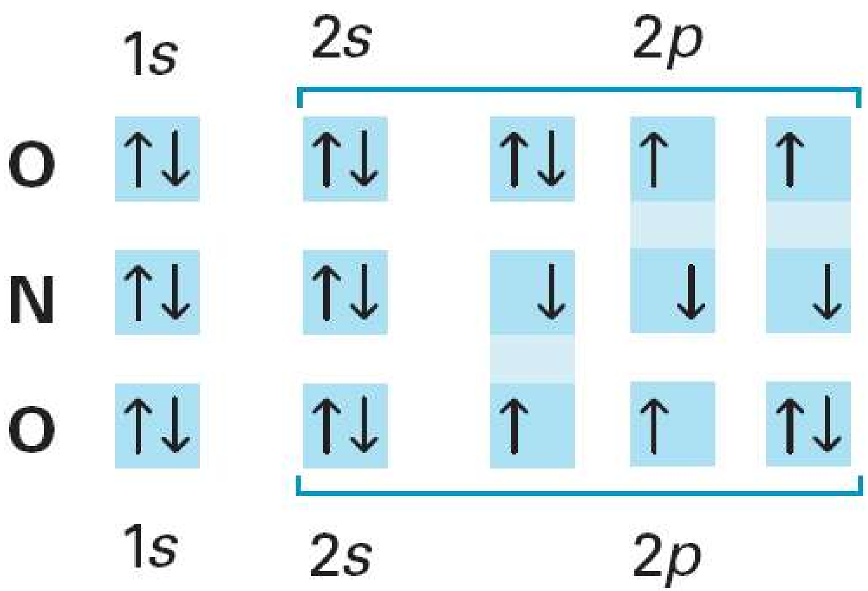
Exceptions to the Octet Rule

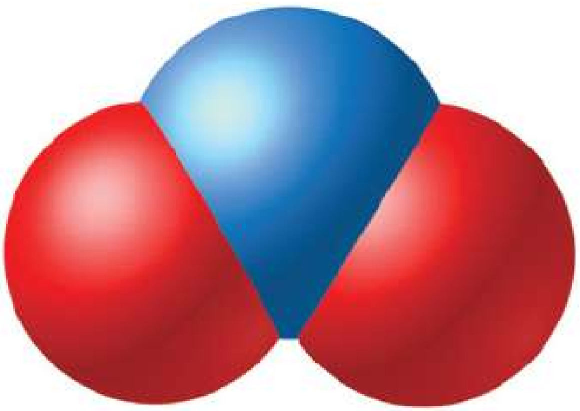
- The octet rule cannot be satisfied in molecules whose total number of valence electrons is an odd number. There are also molecules in which an atom has fewer, or more, than a complete octet of valence electrons.

8.2 Exceptions to the Octet Rule

- Two electron dot structures can be drawn for the NO_2 molecule.





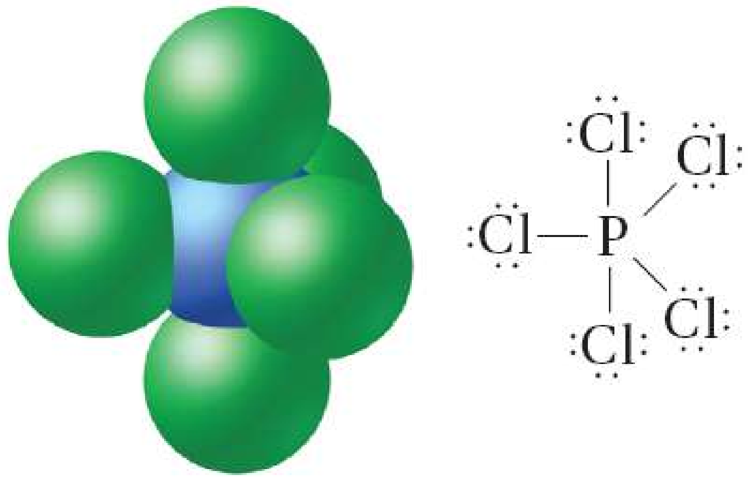


Nitrogen dioxide molecule

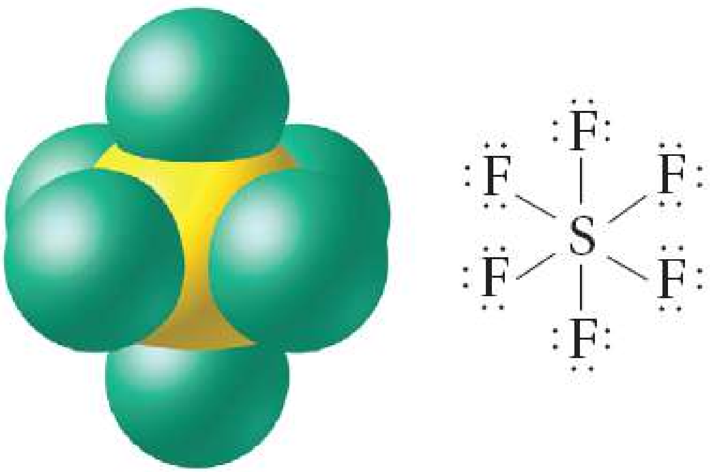
8.2

Exceptions to the Octet Rule

- The electron dot structure for PCl_5 can be written so that phosphorus has ten valence electrons.



Phosphorus pentachloride



Sulfur hexafluoride

Drawing Lewis Diagrams

- **$BeCl_2$**

Practice Problems

- 1) BH_3
- 2) PF_6^-
- 3) ClF_3

8.2 Section Quiz.

- 1. In covalent bonding, atoms attain the configuration of noble gases by
  - a) losing electrons.
  - b) gaining electrons.
  - c) transferring electrons.
  - d) sharing electrons.

8.2 Section Quiz

- 2. Electron dot diagrams are superior to molecular formulas in that they
  - a) show which electrons are shared.
  - b) indicate the number of each kind of atom in the molecule.
  - c) show the arrangement of atoms in the molecule.
  - d) are easier to write or draw.

8.2 Section Quiz

- 3. Which of the following molecules would contain a bond formed when atoms share three pairs of electrons?
  - a) Se_2
  - b) As_2
  - c) Br_2
  - d) Te_2

8.3 Bonding Theories

Molecular Orbitals

- When two atoms combine, the molecular orbital model assumes that their atomic orbitals overlap to produce **molecular orbitals**, or orbitals that apply to the entire molecule.

8.3

Molecular Orbitals

- Just as an atomic orbital belongs to a particular atom, a molecular orbital belongs to a molecule as a whole.

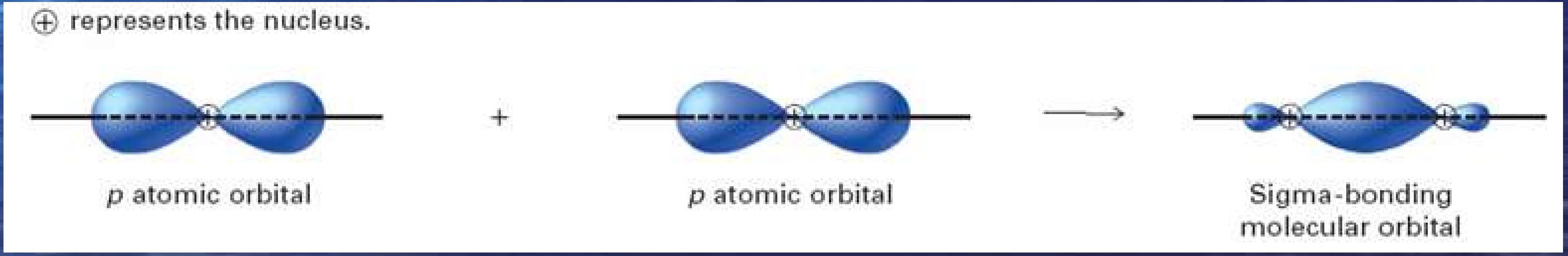
- A molecular orbital that can be occupied by two electrons of a covalent bond is called a **bonding orbital**.

Molecular Orbitals

- Sigma Bonds
  - When two atomic orbitals combine to form a molecular orbital that is symmetrical around the axis connecting two atomic nuclei, a **sigma bond** is formed.

Molecular Orbitals

- When two fluorine atoms combine, the *p* orbitals overlap to produce a bonding molecular orbital. The F—F bond is a sigma bond.



Molecular Orbitals

- Pi Bonds
  - In a **pi bond** (symbolized by the Greek letter π), the bonding electrons are most likely to be found in sausage-shaped regions above and below the bond axis of the bonded atoms.

VSEPR Theory

- The valence-shell electron-pair repulsion theory, or **VSEPR theory**, explains the three-dimensional shape of methane.

VSEPR Theory

- According to VSEPR theory, the repulsion between electron pairs causes molecular shapes to adjust so that the valence-electron pairs stay as far apart as possible.

Linear

- Total pairs: 2
- Shared pairs: 2
- Lone pairs: 0
- Bond Angle: 180°
- Hybridization: sp
- Ex: $BeCl_2$

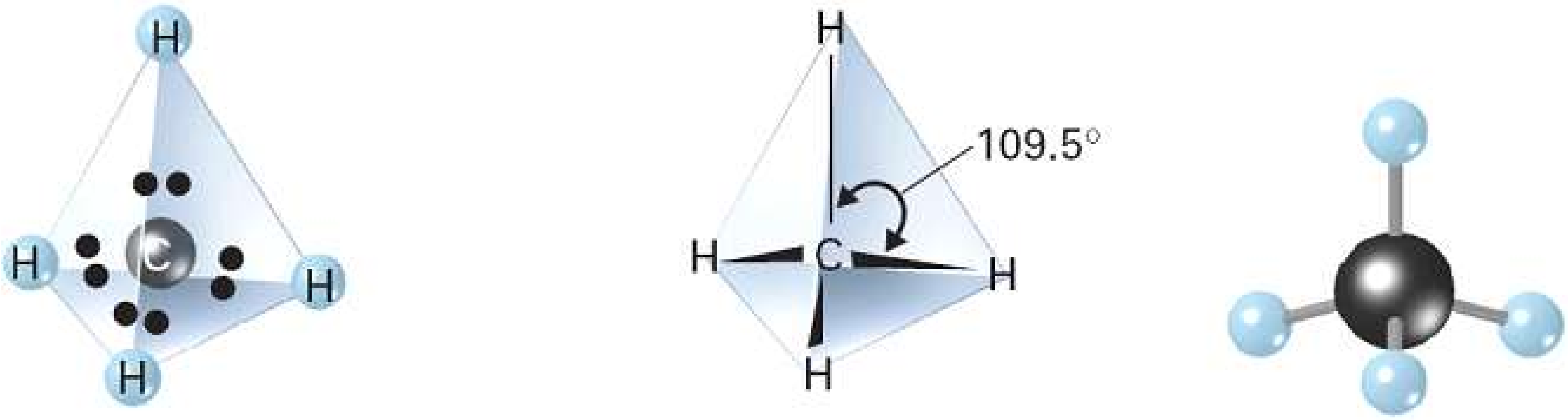
Trigonal Planar

- Total Pairs: 3
- Shared Pairs: 3
- Lone Pairs: 0
- Bond Angle: 120°
- Hybridization: sp^2
- Ex: $AlCl_3$

Tetrahedral

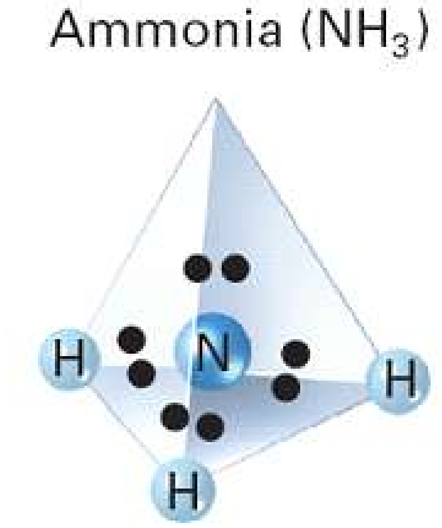
- Total Pairs: 4
- Shared Pairs: 4
- Lone Pairs: 0
- Hybridization: sp^3

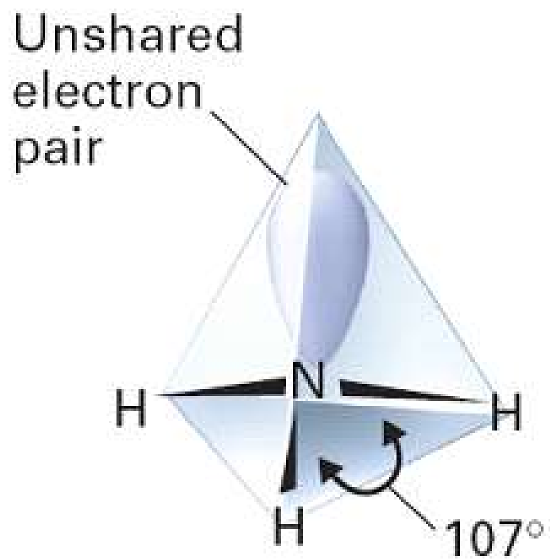
Methane (CH_4)

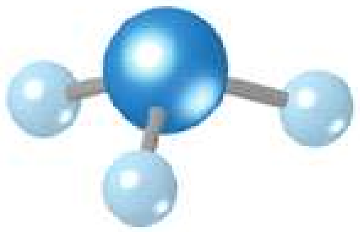


Trigonal Pyramidal

- Total Pairs: 4
- Shared Pairs: 3
- Lone Pairs: 1
- Bond Angle: 107.3°
- Hybridization: sp^3







Bent (V-shaped)

- Total Pairs: 4
- Shared Pairs: 2
- Lone Pairs: 2
- Hybridization: sp^3
- Bond Angle: 104.5°
- Ex: H_2O

Trigonal Bipyramidal

- Total pairs: 5
- Shared pairs: 5
- Lone pairs: 0
- Hybridization: sp^3d
- Example: $NbBr_5$
- Bond angles: equatorial is 120° and axial is 90°

Octahedral

- Total pairs: 6
- Bonding pairs: 6
- Lone pairs: 0
- Hybridization: sp^3d^2
- Example SF_6
- Bond angles: 90°

Examples

- PF_3

Examples

- CO_2

Draw Lewis Structure and predict Shape of:

- 1) NH_3
- 2) Cl_2O
- 3) N_2O
- 4) H_2Se
- 5) ClO_4^-
- 6) NH_4^+
- 7) SO_4^{2-}
- 8) NF_3
- 9) H_2S
- 10) ClO_3^-
- 11) BeF_2
- 12) NO_3^-

8.3

Hybrid Orbitals

- Orbital hybridization provides information about both molecular bonding and molecular shape.
  - In hybridization, several atomic orbitals mix to form the same total number of equivalent hybrid orbitals.
- http://mhhe.com/physsci/chemistry/essentialchemistry/flash/hybrv18.swf

8.3 Section Quiz.

- 1.VSEPR theory enables prediction of 3-dimensional molecular shape because the valence electron pairs

  a) are attracted to each other.

  b) form molecules with only four possible shapes.

  c) stay as far apart as possible.

  d) always form tetrahedral shapes.

8.3 Section Quiz.

- 2.Orbital hybridization provides information about
  - a) both molecular bonding and molecular shape.
  - b) both molecular bonding and bond energy.
  - c) neither molecular bonding nor molecular shape.
  - d) neither molecular bonding nor bond energy.

8.4 Polar Bonds and Molecules

8.4

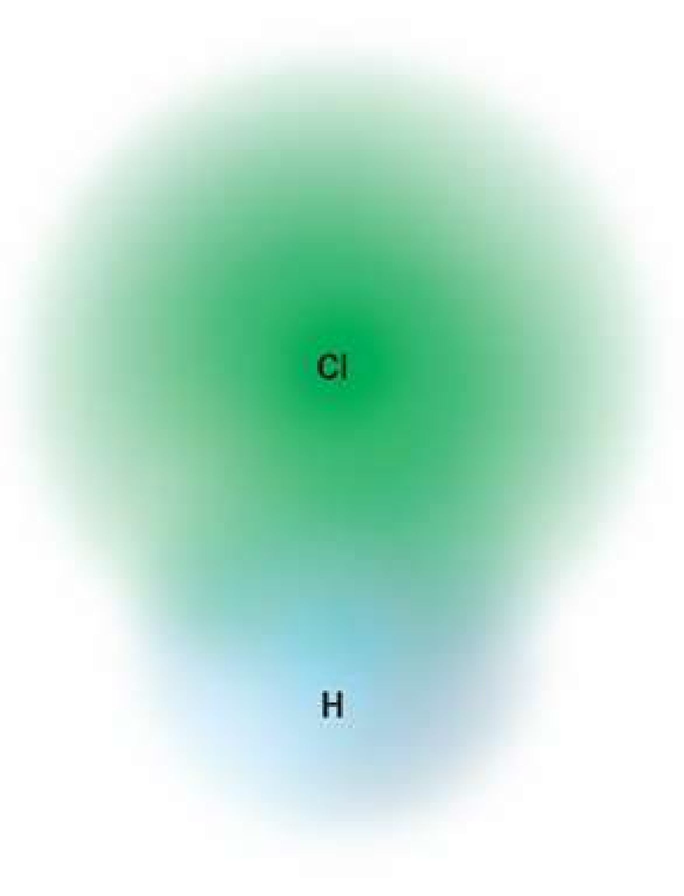
Bond Polarity

- When the atoms in a bond pull equally (as occurs when identical atoms are bonded), the bonding electrons are shared equally, and the bond is a **nonpolar covalent bond**.

8.4

Bond Polarity

- The chlorine atom attracts the electron cloud more than the hydrogen atom does.



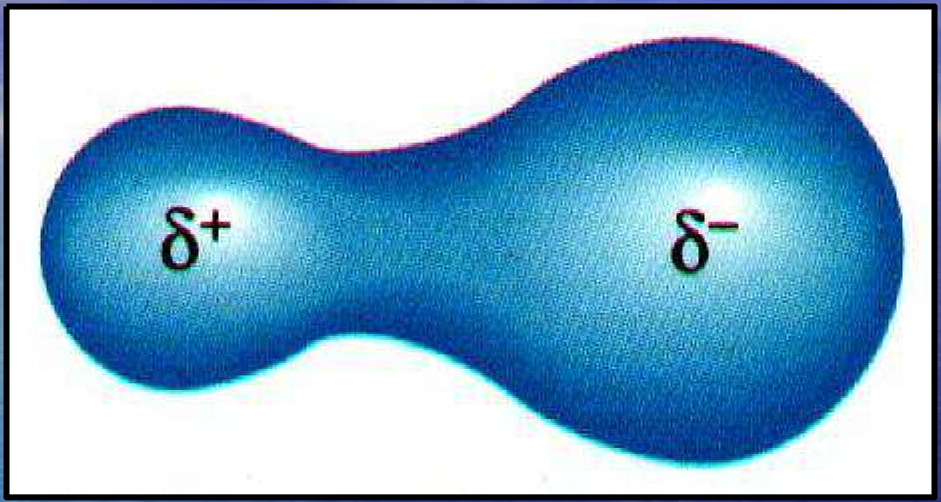
8.4

Bond Polarity

- A **polar covalent bond**, known also as a polar bond, is a covalent bond between atoms in which the electrons are shared unequally.
  - The more electronegative atom attracts electrons more strongly and gains a slightly negative charge. The less electronegative atom has a slightly positive charge.

- **Polar Covalent Bond**

δ^+



δ^-

8.4 Bond Polarity

Table 8.3

Electronegativity Differences and Bond Types		
Electronegativity difference range	**Most probable type of bond**	**Example**
0.0–0.4	Nonpolar covalent	H—H (0.0)
0.4–1.0	Moderately polar covalent	$\overset{\delta+}{H}$ — $\overset{\delta-}{Cl}$ (0.9)
1.0–2.0	Very polar covalent	$\overset{\delta+}{H}$ — $\overset{\delta-}{F}$ (1.9)
≥ 2.0	Ionic	Na^+Cl^- (2.1)

Electronegativity difference (ΔEN)

- Calculate the electronegativity difference and tell the type of bond formed between sulfur and each of the following: hydrogen, cesium, and chlorine. In each pair which atom will be more negative?
- Use electronegativity differences to classify bonding between chlorine and calcium, oxygen and bromine. In each pair which atom will be more negative?

Conceptual Problem 8.3

- Tell type of bond (nonpolar, polar or ionic) that will form between the following pairs of atoms:
  - N and H
  - F and F
  - Ca and Cl
  - Al and Cl

Practice
- tell type of bond

- H and Br
- Cl and F
- K and Cl
- Li and O
- C and O
- Br and Br

Dipole Moments in Polar Molecules

- H F

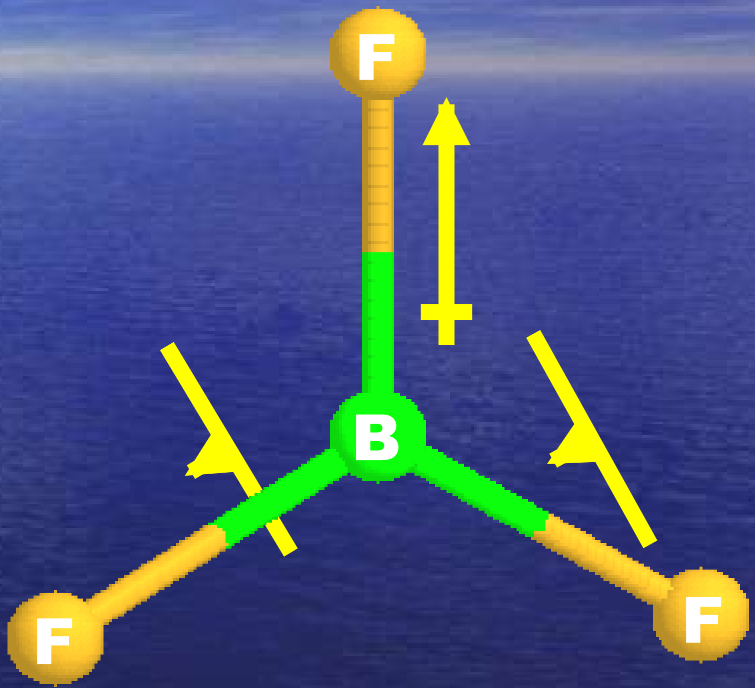
- δ^+ δ^-

Use an arrow to represent the dipole character that points toward the partial negative end

Determining Molecular Polarity

- **Nonpolar Molecules**
  - Dipole moments are symmetrical and cancel out.

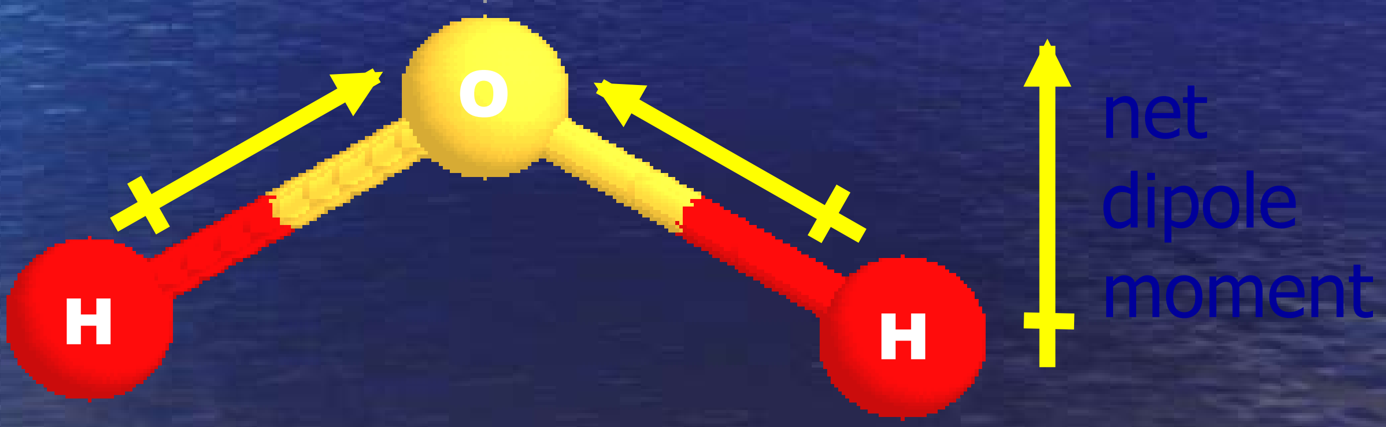
BF_3



Determining Molecular Polarity

- **Polar Molecules**
  - Dipole moments are asymmetrical and don't cancel .

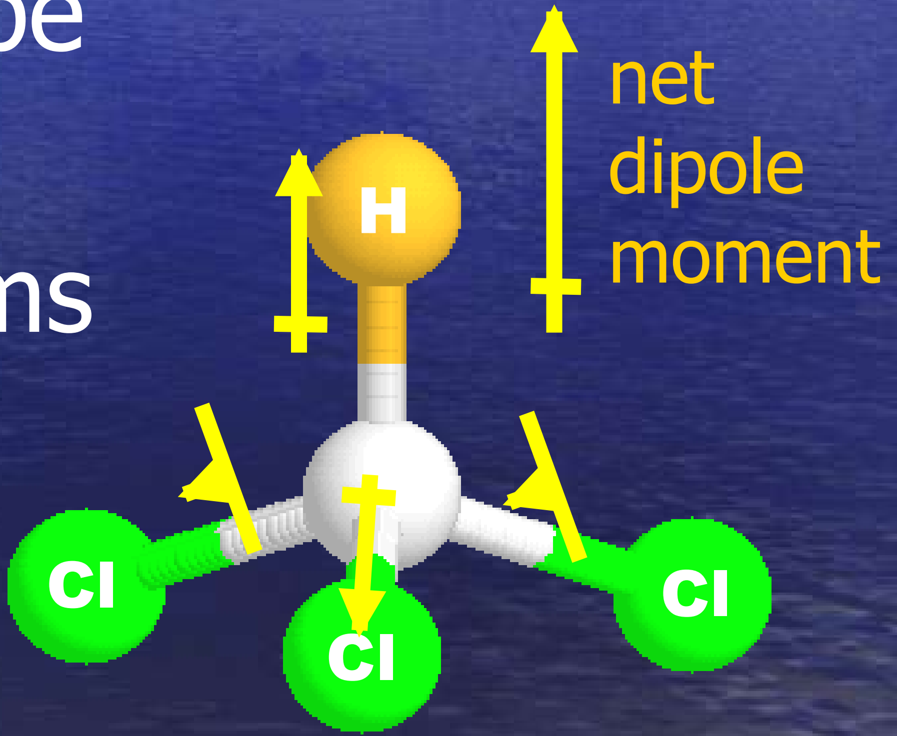
H_2O



Determining Molecular Polarity

- **Therefore, polar molecules have...**
  - asymmetrical shape (lone pairs) or
  - asymmetrical atoms

$CHCl_3$



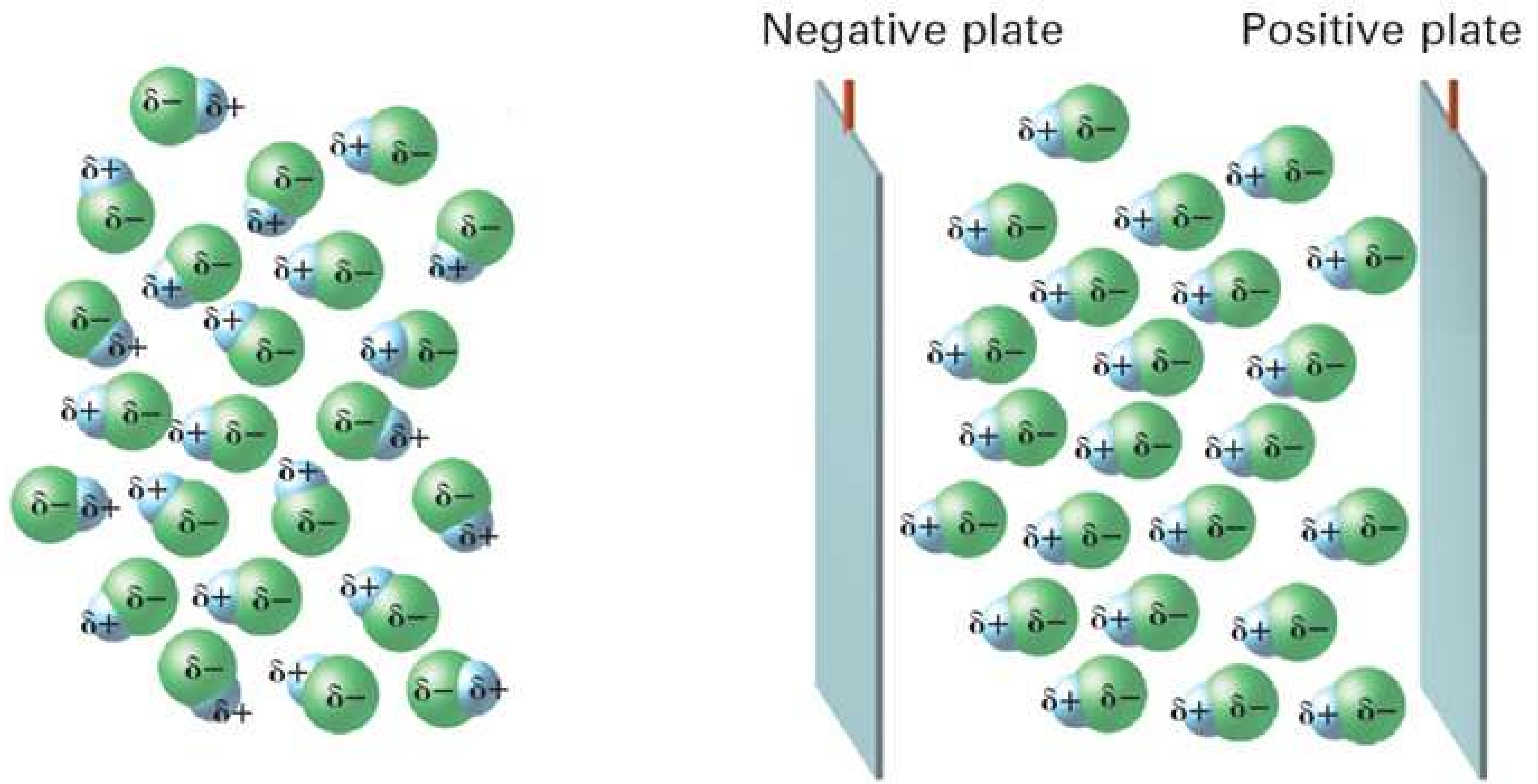
8.4 Polar Molecules

- In a **polar molecule**, one end of the molecule is slightly negative and the other end is slightly positive.
- A molecule that has two poles is called a dipolar molecule, or **dipole**.

8.4

Polar Molecules

- A hydrogen chloride molecule is a dipole.



Electric field absent.
Polar molecules orient randomly.

Electric field on.
Polar molecules line up.

8.4

Attractions Between Molecules

- Intermolecular attractions are weaker than either ionic or covalent bonds.
  - These attractions are responsible for determining whether a molecular compound is a gas, a liquid, or a solid at a given temperature.

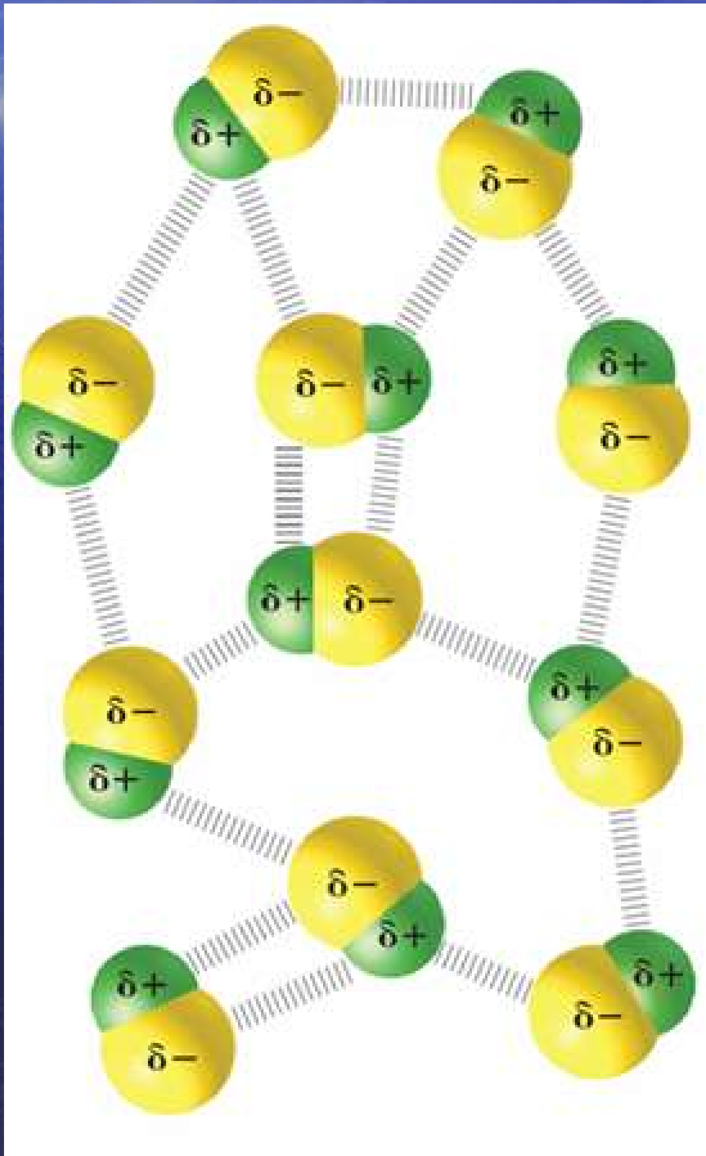
Attractions Between Molecules

- Van der Waals Forces
  - The two weakest attractions between molecules are collectively called **van der Waals forces,** named after the Dutch chemist Johannes van der Waals (1837–1923).

8.4

Attractions Between Molecules

- **Dipole interactions** occur when polar molecules are attracted to one another.



8.4

Attractions Between Molecules

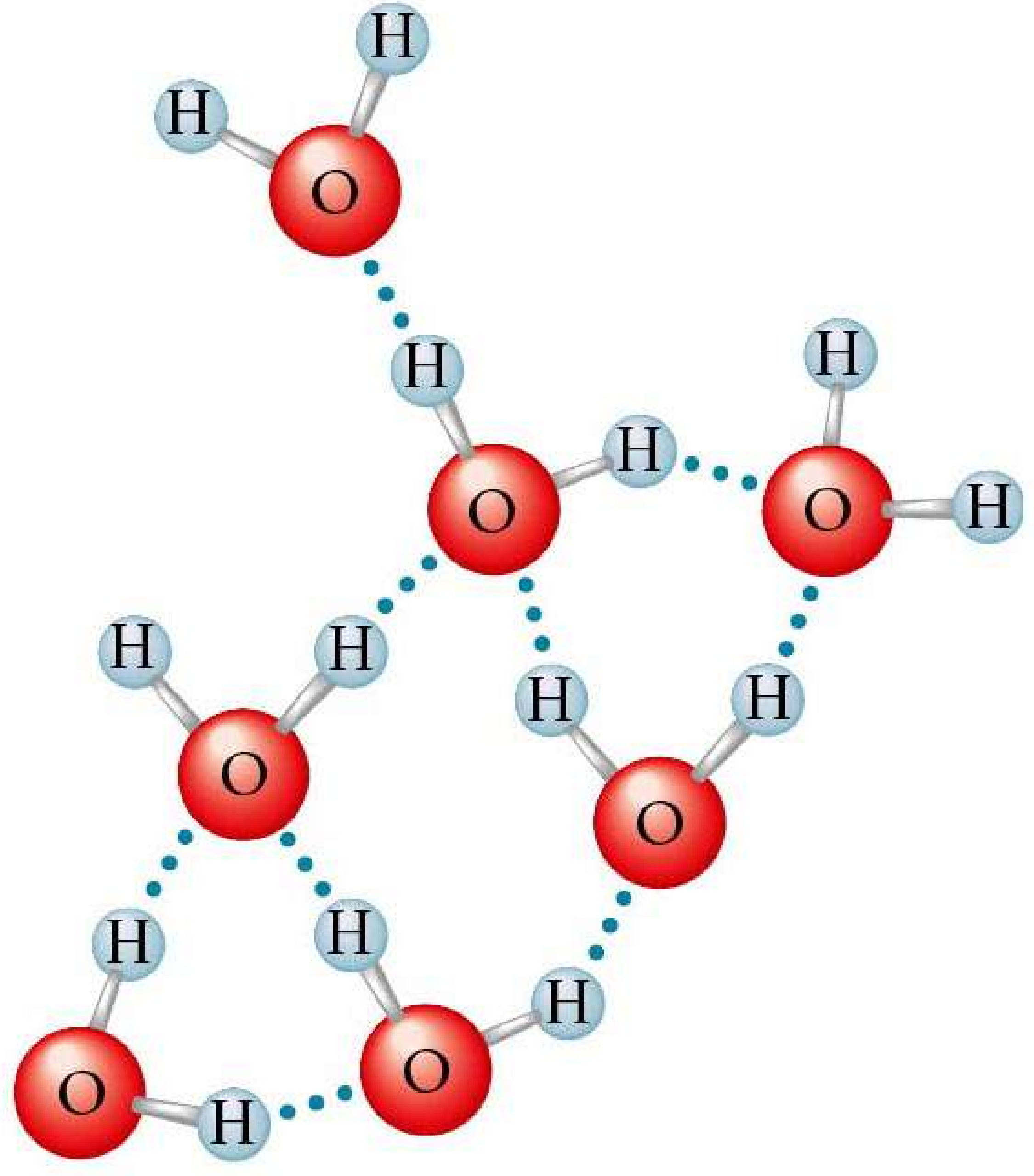
- **Dispersion forces**, the weakest of all molecular interactions, are caused by the motion of electrons.
  - The strength of dispersion forces generally increases as the number of electrons in a molecule increases.

8.4

Attractions Between Molecules

- Hydrogen Bonds
  - **Hydrogen bonds** are attractive forces in which a hydrogen covalently bonded to a very electronegative atom is also weakly bonded to an unshared electron pair of another electronegative atom.

- Hydrogen bonding among water molecules

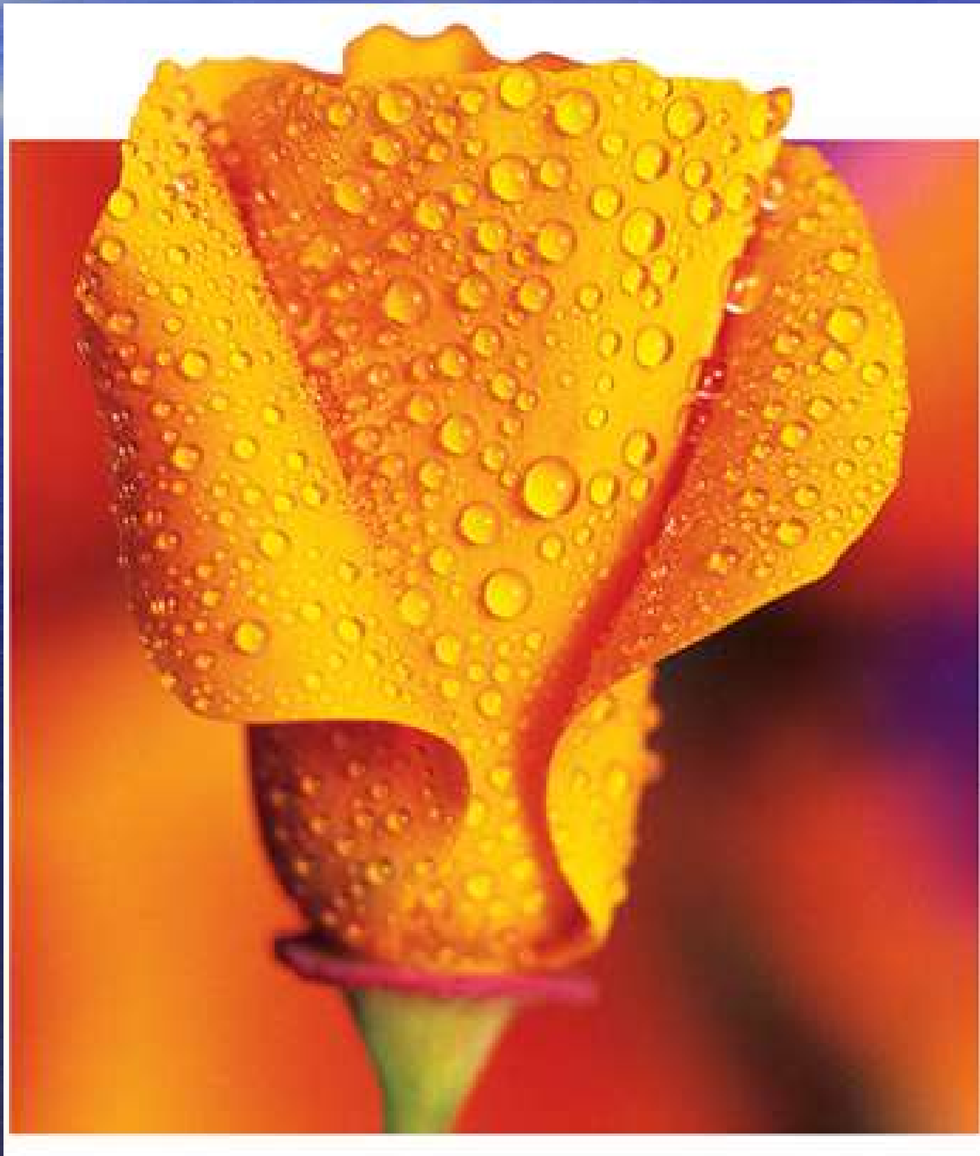


(b)

8.4

Attractions Between Molecules

- The relatively strong attractive forces between water molecules cause the water to form small drops on a waxy surface.



8.4

Intermolecular Attractions and Molecular Properties

- **Network solids** (or network crystals) are solids in which all of the atoms are covalently bonded to each other.
  - Network solids consist of molecules that do not melt until the temperature reaches 1000°C or higher, or they decompose without melting at all.

8.4

- Diamond is an example of a network solid. Diamond does not melt. It vaporizes to a gas at 3500°C or above.

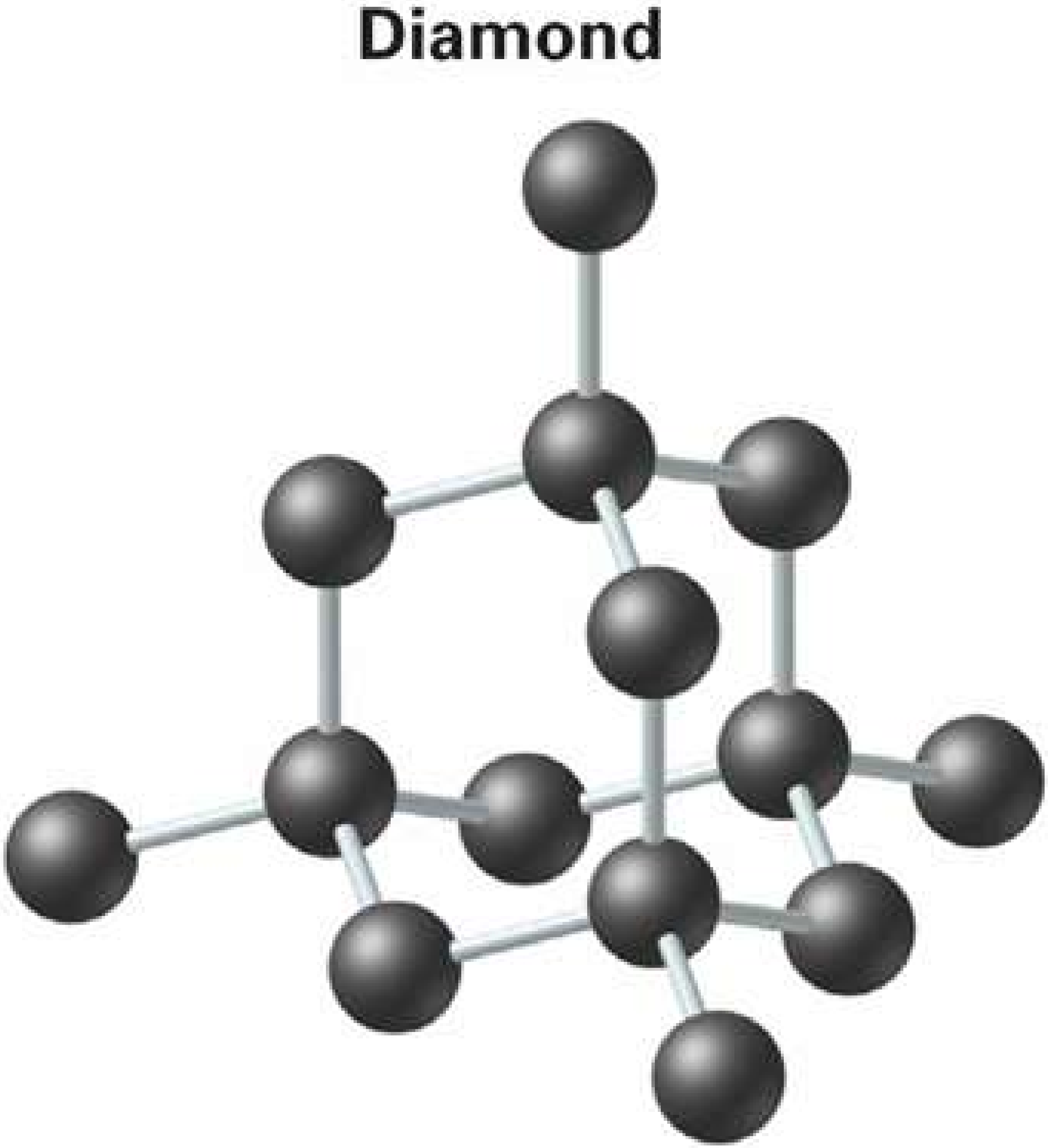


Table 8.4

Characteristics of Ionic and Covalent Compounds		
Characteristic	**Ionic compound**	**Covalent compound**
Representative unit	Formula unit	Molecule
Bond formation	Transfer of one or more electrons between atoms	Sharing of electron pairs between atoms
Type of elements	Metallic and nonmetallic	Nonmetallic
Physical state	Solid	Solid, liquid, or gas
Melting point	High (usually above 300°C)	Low (usually below 300°C)
Solubility in water	Usually high	High to low
Electrical conductivity of aqueous solution	Good conductor	Poor to nonconducting

8.4 Section Quiz.

- 1.In a molecule, the atom with the largest electronegativity value

  a) repels electrons more strongly and aquires a slightly negative charge.

  b) repels electrons more strongly and aquires a slightly positive charge.

  c) attracts electrons more strongly and aquires a slightly positive charge.

  d) attracts electrons more strongly and aquires a slightly negative charge.

8.4 Section Quiz.

- 2.When polar molecules are placed between oppositely charged plates, the negative
  - a) molecules stick to the positive plates.
  - b) molecules stick to the negative plates.
  - c) ends of the molecules turn toward the positive plates.
  - d) ends of the molecules turn toward the negative plates.

8.4 Section Quiz.

- 3.Which of the following bond types is the weakest?
  - a) ionic bond
  - b) Van der Waals force
  - c) covalent bond
  - d) hydrogen bond

9.3 Naming and Writing Formulas for Molecular Compounds

9.3 Naming Binary Molecular Compounds

- **A prefix in the name of a binary molecular compound tells how many atoms of an element are present in each molecule of the compound.**

Table 9.4

Prefixes Used in Naming Binary Molecular Compounds

Prefix	Mono-	Di-	Tri-	Tetra-	Penta-	Hexa-	Hepta-	Octa-	Nona-	Deca-
Number	1	2	3	4	5	6	7	8	9	10

9.3

Naming Binary Molecular Compounds

- Some guidelines for naming binary molecular compounds:
  - Name the elements in the order listed in the formula.
  - Use prefixes to indicate the number of each kind of atom.

Naming Binary Molecular Compounds

- Omit the prefix *mono-* when the formula contains only one atom of the first element in the name.
- The suffix of the name of the second element is -*ide*.

Practice Traditional Naming

- PF_5
- XeF_4
- CCl_4
- SF_4
- NO
- ClF_3

Naming Binary Molecular Compounds

- 1) BF_3
- 2) NO
- 3) N_2O_5
- 4) CCl_4
- 5) NO_2
- 6) I_2O_7
- 7) CO_2
- 8) CF_4
- 9) NH_3
- 10) PCl_3

9.3 Writing Formulas for Binary Molecular Compounds

- **Use the prefixes in the name to tell you the subscript of each element in the formula. Then write the correct symbols for the two elements with the appropriate subscripts.**

Writing Formulas

- dihydrogen monoxide
- phosphorus pentachloride
- sulfur dioxide
- carbon dioxide
- dinitrogen pentoxide
- sulfur hexafluoride

Write Formulas for:

- 1) sulfur trioxide
- 2) dinitrogen monoxide
- 3) phosphorus pentachloride
- 4) oxygen difluoride
- 5) chlorine monoxide

9.3 Section Quiz.

- 1.Which of the following compounds is named INCORRECTLY?
  - a) CS_2, carbon disulfide
  - b) BCl_3, boron trichloride
  - c) IF_7, iodine heptafluoride
  - d) PCl_5, phosphorus hexachloride

9.3 Section Quiz.

– 2.Which of the following molecular compounds is named INCORRECTLY?
  - a) $SbCl_3$, antimony trichloride
  - b) C_2O_5, dicarbon pentoxide
  - c) CF_4, carbon tetrafluoride
  - d) H_3As, hydrogen arsenide

9.3 Section Quiz.

- 3. The correct formula for tetraphosphorus trisulfide is
  - a) P_3S_4
  - b) S_3P_4
  - c) P_4S_3
  - d) S_4P_3

9.4 Naming and Writing Formulas for Acids and Bases

Naming Acids

- An **acid** is a compound that contains one or more hydrogen atoms and produces hydrogen ions (H^+) when dissolved in water. Acids have various uses.

Naming Acids

- Three rules can help you name an acid with the general formula H_nX.
  - When the name of the anion (X) ends in *-ide,* the acid name begins with the prefix *hydro-*. The stem of the anion has the suffix *-ic* and is followed by the word *acid.*

Naming Acids

- When the anion name ends in -*ite*, the acid name is the stem of the anion with the suffix -*ous*, followed by the word *acid*.

Naming Acids

- When the anion name ends in *-ate,* the acid name is the stem of the anion with the suffix *-ic* followed by the word *acid*.

Naming Acids

- **A Summary of the Three Rules for Naming Acids.**

Table 9.5

Naming Common Acids			
Anion ending	**Example**	**Acid name**	**Example**
-ide	chlor*ide*, Cl^-	*hydro*-(stem)-*ic acid*	*hydro*chlor*ic acid*
-ite	sulf*ite*, SO_3^{2-}	(stem)-*ous acid*	sulfur*ous acid*
-ate	nitr*ate*, NO_3^-	(stem)-*ic acid*	nitr*ic acid*

Name the acids:

- 1) HF7) HI
- 2) $HNO_3$8) $HClO_2$
- 3) $HC_2H_3O_2$9) H_2SO_4
- 4) HBr10) H_2S
- 5) $HClO_4$
- 6) $HClO_3$

Writing Formulas for Acids

- **Use the rules for writing the names of acids in reverse to write the formulas for acids.**
  - What is the formula for hydrobromic acid? Following Rule 1, hydrobromic acid (*hydro-* prefix and -*ic* suffix) must be a combination of hydrogen ion (H^+) and bromide ion (Br^-). The formula of hydrobromic acid is HBr.

Write formulas for:

- 1) sulfuric acid
- 2) sulfurous acid
- 3) hydrosulfuric acid
- 4) hydroiodic acid
- 5) nitric acid
- 6) carbonic acid

Names and Formulas for Bases

- **Bases are named in the same way as other ionic compounds—the name of the cation is followed by the name of the anion.**
  - For example, aluminum hydroxide consists of the aluminum cation (Al^{3+}) and the hydroxide anion (OH^{-}). The formula for aluminum hydroxide is $Al(OH)_3$.

9.4 Section Quiz

- 1. The name for $H_2S(aq)$ is
  - a) sulfuric acid.
  - b) hydrosulfuric acid.
  - c) sulfurous acid.
  - d) hydrosulfurous acid.

9.4 Section Quiz

- 2.The chemical formula for chlorous acid is
  - a) $HClO_2$.
  - b) $HClO_3$.
  - c) $HClO_4$.
  - d) HCl.

9.4 Section Quiz

- 3.The correct chemical name for NH_4OH is
  - nitrogen tetrahydrogen hydroxide.
  - nitrogen pentahydrogen oxide.
  - ammonium oxyhydride.
  - ammonium hydroxide.

9.5 The Laws Governing Formulas and Names

- The Law of Definite Proportions
  - The **law of definite proportions** states that in samples of any chemical compound, the masses of the elements are always in the same proportions.

- The Law of Multiple Proportions
  - The **law of multiple proportions**: Whenever the same two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.

9.1

Calculating Mass Ratios

Carbon reacts with oxygen to form two compounds. Compound A contains 2.41 g of carbon for each 3.22 g of oxygen. Compound B contains 6.71 g of carbon for each 17.9 g of oxygen. What is the lowest whole number mass ratio of carbon that combines with a given mass of oxygen?

34. Lead forms two compounds with oxygen. One contains 2.98 g of lead and 0.461 g of oxygen. The other contains 9.89 g of lead and 0.763 g of oxygen. For a given mass of oxygen, what is the lowest whole number mass ratio of lead in the two compounds?

Mixed Naming

- 1) $FeBr_3$
- 2) PCl_3
- 3) $NaHCO_3$
- 4) $BaSO_4$
- 5) BrF_5
- 6) NaBr
- 7) KClO
- 8) MgI_2
- 9) $Fe(OH)_2$
- 10) $Zn_3(PO_4)_2$

Formula Review

- 1) potassium hydroxide
- 2) sodium carbonate
- 3) nitric acid
- 4) cobalt (III) nitrate
- 5) calcium chloride

Write formulas for:

- 1) magnesium bromide
- 2) hydrochloric acid
- 3) carbon disulfide
- 4) nickel (II) perchlorate
- 5) cesium fluoride

Section Quiz 9.5.

– 1.The law of definite proportions states that in samples of any chemical compound, the elements are always in the same proportion by

a. mass.

b. volume.

c. group number.

d. period number.