Chapter 7 Molecular Structure: Solids and Liquids

Electron Configuration of Ionic Compounds
Review
Metals Form Positive Ions

Metals form

- Octets by *losing* all of their valence electrons.
- **Positive ions** with the electron configuration of the nearest noble gas.
- **Positive ions** with fewer electrons than protons.

- Group 1A(1) metals $\rightarrow$ ion $^{1+}$
- Group 2A(2) metals $\rightarrow$ ion $^{2+}$
- Group 3A(13) metals $\rightarrow$ ion $^{3+}$
Formation of Mg$^{2+}$

- Magnesium achieves an octet by losing its two valence electrons.
Formation of Negative Ions

In ionic compounds, nonmetals

- Achieve an octet arrangement.
- Gain electrons.
- Form negatively charged ions with 3-, 2-, or 1- charges.
Formation of a Chloride, Cl⁻

- Chlorine achieves an octet by adding an electron to its valence electrons.
Ions of Transition Metals

Transition metals

- Form positively charged ions.
- Lose s valence electrons to form 2+ ions.
- Lose d electrons to form ions with higher positive charges.
- Do not form octets like representative metals.
Ions of Transition Metals

Example: Fe forms Fe$^{3+}$ and Fe$^{2+}$

Fe $\quad 1s^22s^22p^63s^23p^64s^23d^6$

Loss of two 4s electrons

Fe$^{2+} \quad 1s^22s^22p^63s^23p^64s^03d^6$

Loss of two 4s electrons and one d electron

Fe$^{3+} \quad 1s22s22p63s23p64s03d5$

Note: The 3d subshell is half-filled and stable.
Electron-Dot Formulas

\[ \overset{\cdot}{O} = C = \overset{\cdot}{O} \quad \text{CO}_2 \]

Molecule of carbon dioxide

Why look at Electron-Dot Formulas?

Ionic Compounds: Helps to determine formulas

Covalent Compounds: Help us to understand molecular structures or molecular geometries.
Molecular Geometries or Shapes

- Linear: 180°
- Bent: 120°
- Trigonal planar: 120°
- Trigonal pyramidal shape
- Tetrahedral shape: 109.5°
Multiple Covalent Bonds

In a single bond
• One pair of electrons is shared.

In a double bond,
• Two pairs of electrons are shared.

In a triple bond.
• Three pairs of electrons are shared.
Electron-dot formulas show

- The order of bonded atoms in a covalent compound.
- The bonding pairs of electrons between atoms.
- The unshared (lone, non-bonding) valence electrons.
- A central atom with an octet.

Electron-dot formula

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{N} & \quad \text{H} \\
\text{H} &
\end{align*}
\]

Lone pair
Number of Covalent Bonds

The number of covalent bonds can be determined from the number of electrons needed to complete an octet.

Table 7.1

<table>
<thead>
<tr>
<th>1A (1)</th>
<th>3A (13)</th>
<th>4A (14)</th>
<th>5A (15)</th>
<th>6A (16)</th>
<th>7A (17)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 bond</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
<td></td>
</tr>
<tr>
<td>3 bonds</td>
<td>4 bonds</td>
<td>3 bonds</td>
<td>2 bonds</td>
<td>1 bond</td>
<td></td>
</tr>
<tr>
<td>Si</td>
<td>P</td>
<td>S</td>
<td>Cl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4 bonds</td>
<td>3 bonds</td>
<td>2 bonds</td>
<td>1 bond</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Br</td>
<td>I</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 bond</td>
<td>1 bond</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

*aH and B do not form eight-electron octets. H atoms share one electron pair; B atoms share three electron pairs for a set of 6 electrons.
Guide to Writing Electron-Dot Formulas

STEP 1  Determine the arrangement of atoms.
STEP 2  Add the valence electrons from all the atoms.
STEP 3  Attach the central atom to each bonded atom using one pair of electrons.
STEP 4  Add remaining electrons as lone pairs to complete octets (2 for H atoms).
STEP 5  If octets are not complete, form one or more multiple bonds.
Electron-Dot Formula of $\text{SF}_2$

Write the electron-dot formula for $\text{SF}_2$.

**STEP 1** Determine the atom arrangement.

S is the central atom.

F S F

**STEP 2** Total the valence electrons for 1S and 2F.

$$1\text{S}(6e^-) + 2\text{F}(7e^-) = 20e^-$$
**Electron-Dot Formula SF$_2$**

**STEP 3** Attach F atoms to S with one electron pair.

F : S : F

Calculate the remaining electrons.

20e$^-$ - 4 e$^-$ = 16e$^-$ left

**STEP 4** Complete the octets of all atoms by placing remaining e$^-$ as 8 lone pairs to complete octets.

:: F : S : F : or F─S─F :
Write the electron-dot formula for $\text{ClO}_3^-$.

**STEP 1** Determine atom arrangement.
Cl is the central atom. 

\[
\begin{array}{c}
\text{Cl} \\
\text{O} \\
\text{O} \\
\text{Cl} \\
\text{O}
\end{array}
\] 

**STEP 2** Add all the valence electrons for 1Cl and 3O plus 1e\(^-\) for negative charge on the ion.

\[
1\text{Cl}(7\text{e}^-) + 3 \text{O}(6\text{e}^-) + 1\text{e}^- = 26\text{e}^-
\]
Electron-Dot Formula ClO$_3^-$

**STEP 3** Attach each O atom to Cl with one electron pair.

\[
\begin{align*}
\text{O} : \\
\text{O} : \text{Cl} : \text{O} \\
\end{align*}
\]

Calculate the remaining electrons.

\[
26e^- - 6e^- = 20e^- \text{ left}
\]
STEP 4 Complete the octets of all atoms by placing the remaining 20 e⁻ as 10 lone pairs to complete octets.

Note: Bonding electrons can be shown by a \(-\).
Multiple Bonds

In a single bond
• One pair of electrons is shared.

In a double bond,
• Two pairs of electrons are shared.

In a triple bond.
• Three pairs of electrons are shared.
Multiple Bonds in N\textsubscript{2}

In nitrogen N\textsubscript{2},

- Octets are achieved by sharing three pairs of electrons, which is a **triple bond**.

\[ \text{N} \overset{\equiv}{\text{N}} \]
Electron-Dot Formula of $\text{CS}_2$

Write the electron-dot formula for $\text{CS}_2$.

**STEP 1** Determine the atom arrangement. The C atom is the central atom.

\[ \text{S} \quad \text{C} \quad \text{S} \]

**STEP 2** Total the valence electrons for 1C and 2S.

\[ 1\text{C}(4e^-) + 2\text{S}(6e^-) = 16e^- \]
STEP 3  Attach each S atom to C with electron pairs.

S : C : S

Calculate the remaining electrons.

\[ 16e^- - 4e^- = 12e^- \text{ left} \]
Electron-Dot Formula $\text{CS}_2$

**STEP 4** Attach 12 remaining electrons as 6 lone pairs to complete octets.

\[
: \cdot \cdot \cdot C : \cdot \cdot \cdot S :
\]

**STEP 5** To complete octets, move two lone pairs between C and S atoms to give two double bonds.

\[
\begin{align*}
: & \cdot \cdot \cdot S : \cdot \cdot \cdot C : \cdot \cdot \cdot S : \\
\text{or} & \quad : & \cdot \cdot \cdot S = C = S :
\end{align*}
\]
### Some Electron-Dot Formulas

#### Table 7.2

**Using Valence Electrons to Write Electron-Dot Formulas**

<table>
<thead>
<tr>
<th>Molecule or Polyatomic Ion</th>
<th>Valence Electrons</th>
<th>Bonds to Attach Atoms</th>
<th>Electrons Left</th>
<th>Completed Octets (or H₂)</th>
<th>Electron Check</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl₂</td>
<td>2(7) = 14</td>
<td>Cl—Cl</td>
<td>14 – 2 = 12</td>
<td>:Cl—Cl—Cl:</td>
<td>14</td>
</tr>
<tr>
<td>HCl</td>
<td>1 + 7 = 8</td>
<td>H—Cl</td>
<td>8 – 2 = 6</td>
<td>H—Cl:</td>
<td>8</td>
</tr>
<tr>
<td>H₂O</td>
<td>2(1) + 5 = 8</td>
<td>H—O—H</td>
<td>8 – 4 = 4</td>
<td>H—O—H</td>
<td>8</td>
</tr>
<tr>
<td>PCl₃</td>
<td>5 + 3(7) = 26</td>
<td>Cl—P—Cl</td>
<td>26 – 6 = 20</td>
<td>:Cl—P—Cl:</td>
<td>26</td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td>7 + 3(6) + 1 = 26</td>
<td>[O—Cl—O]</td>
<td>26 – 6 = 20</td>
<td>:O—Cl—O:</td>
<td>26</td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>5 + 2(6) + 1 = 18</td>
<td>[O—N—O]</td>
<td>18 – 4 – 14</td>
<td>:O=O=N=O:</td>
<td>18</td>
</tr>
</tbody>
</table>
Resonance Structures

Resonance structures are

- Two or more electron-dot formulas for the same arrangement of atoms.
- Related by a double-headed arrow.
- Written by changing location of a double bond from the central atom to a different attached atom.
- Sometimes written as a hybrid resonance structure.
Resonance structures for $\text{NO}_3^-$ are

\[
\begin{align*}
\overset{\text{O}:}{\text{O}} & \quad \overset{\text{O}:}{\text{O}} \\
\text{N} & \quad \overset{\text{O}:}{\text{O}} \\
\quad & \quad \overset{\text{O}:}{\text{O}} \\
\quad & \quad \overset{\text{O}:}{\text{O}}
\end{align*}
\]
Learning Check

Carbonate has three resonance structures. If the following is one, what are the other two?

\[
\begin{align*}
\text{2-} & \\
\text{O} & \\
\text{C} & \\
\text{O} & \\
\text{O} & \\
\cdot & \\
\cdot & \\
\cdot & \\
\end{align*}
\]
Solution

Carbonate has three resonance structures. If the following is one, what are the other two?

\[
\begin{align*}
\text{O} & \quad \text{O} \\
\text{\quad C} & \\
\text{O} & \quad \text{O}
\end{align*}
\]

\[
\begin{align*}
\text{O} & \quad \text{O} \\
\text{\quad C} & \\
\text{O} & \quad \text{O}
\end{align*}
\]

\[
\begin{align*}
\text{O} & \quad \text{O} \\
\text{\quad C} & \\
\text{O} & \quad \text{O}
\end{align*}
\]
Writing Resonance Structures

Cyanate ion $\text{NCO}^-$ has three resonance structures.

**STEP 1** Write the arrangement of atoms.

$$\text{N} \quad \text{C} \quad \text{O} \quad -$$

**STEP 2** Count the valence electrons.

$$1\text{N}(5\text{e}^-) + 1\text{C}(4\text{e}^-) + 1\text{O}(6\text{e}^-) + \text{charge (1e}^-) = 16\text{e}^-$$

**STEP 3** Connect bonded atoms by single electron pairs.

$$\text{N} : \text{C} : \text{O} \quad -$$

4e$^-$ used

Determine the remaining electrons.

$$16\text{e}^- - 4\text{e}^- = 12\text{e}^-$$
Writing Resonance Structures

STEP 4  Add 12 remaining electrons as 6 lone pairs

\[
\begin{align*}
\text{N—C—O} & \\
\cdot & \cdot & \cdot
\end{align*}
\]

STEP 5  Form double or triple bonds to make octets

\[
\begin{align*}
\text{N=C=O} & \\
\cdot & \cdot & \cdot
\end{align*} \quad \begin{align*}
\text{N—C≡O} & \\
\cdot & \cdot & \cdot
\end{align*} \quad \begin{align*}
\text{N≡C—O} & \\
\cdot & \cdot & \cdot
\end{align*}
\]

Chapter 7 Slide 30 of 80
Learning Check

Write two resonance structures for nitrite $\text{NO}_2^-$. 

$\text{NO}_2^- = 18 \text{ e}^-$

Electron-dot formula

Using a double bond to complete octets gives two resonance structures
Shapes of Molecules and Ions (VSEPR Theory)

Trigonal planar

120°
VSEPR Theory

In the **valence-shell electron-pair repulsion theory (VSEPR)**, the electron groups around a central atom [shell = another name for energy level]

- Are arranged as far apart from each other as possible.
- Have the least amount of repulsion of the negatively charged electrons.
- Have a geometry around the central atom that determines molecular shape.
Shapes of Molecules

The **three-dimensional shape** of a molecule

- Is the result of bonded groups and lone pairs of electrons around the central atom.
- Is predicted using the VSEPR theory (valence-shell-electron-pair repulsion).
Guide to Predicting Molecular Shape (VSEPR Theory)

STEP 1  Draw the electron-dot structure.

STEP 2  Arrange all electron groups around the central atom to minimize repulsion. Bonds as well as lone pairs.

STEP 3  Count the number of atoms bonded to the central atom to predict the shape of the molecule.
Two Electron Groups

In BeCl₂

- There are two electron groups bonded to the central atom Be (exception to the octet rule).

\[
\begin{array}{c}
\vdots \\
\text{Cl—Be—Cl} \\
\vdots 
\end{array}
\]

To minimize repulsion, the arrangement of two electron groups is 180° or opposite each other.

- The shape of the molecule is linear.
Two Electron groups with Double Bonds

In CO$_2$

**STEP 1** Two electron groups bond to C (electrons in a double bond count as one group).

**STEP 2** Minimal repulsion occurs when two electron groups are opposite each other (180°).

**STEP 3** CO$_2$ has a **linear** shape.
Three Electron Groups

In BF$_3$

**STEP 1** Three electron groups surround the central atom B. (B is an exception to the octet rule).

\[ \begin{array}{c}
\cdot \\
: F:
\end{array} \begin{array}{c}
\cdot \\
| \\
: F-B-F:
\end{array} \begin{array}{c}
\cdot \\
\cdot
\end{array} \]

**STEP 2** Minimal repulsion occurs when 3 electron groups are at angles of 120°

**STEP 3** 3 bonded atom give a **trigonal planar** shape.
Two Electron Groups and a Lone Pair

In SO$_2$

- S has 3 electron groups, (2 electron groups atoms and one lone pair).
  \[
  \begin{array}{c}
  \text{O} \\
  \text{S} \\
  \text{O}
  \end{array}
  \]
- Repulsion is minimized with the electron groups in a plane at angles of 120°, a trigonal planar arrangement.
- With two O atoms bonded to S, the shape is bent (120°).
Learning Check

The shape of a molecule of N₂O (N N O) is 1) linear
   In the electron-dot structure with 16 e⁻, octets are acquired using two double bonds to the central N atom. The shape of a molecule with **two electron groups and two bonded atoms** (no lone pairs on the central N) is linear.

\[
:\ddots N : : N : : \ddots\
\]

\[
:\ddots = N = \ddots : \quad \text{linear, 180°}
\]
Four Electron Groups

In a molecule of CH₄

- There are four electron groups bonded to C.
- Repulsion is minimized by placing four electron groups at angles of 109°, which is a tetrahedral arrangement.
- The shape with four bonded atoms is tetrahedral.
Three Bonding Atoms and One Lone Pair

In a molecule of NH$_3$

- Three electron groups bond to H atoms and the fourth one is a lone (nonbonding) pair.
- Repulsion is minimized with 4 electron groups at angles of 109°, which is a tetrahedral arrangement.
- With three bonded atoms, the shape is trigonal pyramidal.
Two Bonding Atoms and Two Lone Pairs

In a molecule of H₂O.

- Two electrons groups are bonded to H atoms and two are lone pairs (4 electron groups).
- Four electron groups minimize repulsion in a tetrahedral arrangement.
- The shape with two bonded atoms is bent(109°).
# Shapes with 2 and 3 Electron Groups

## Table 7.3

<table>
<thead>
<tr>
<th>Electron Groups</th>
<th>Bonded Atoms</th>
<th>Lone Pairs</th>
<th>Bond Angle</th>
<th>Molecular Shape</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>2</td>
<td>0</td>
<td>180°</td>
<td>linear</td>
<td>BeCl$_2$</td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>0</td>
<td>120°</td>
<td>trigonal planar</td>
<td>BF$_3$</td>
</tr>
<tr>
<td>3</td>
<td>2</td>
<td>1</td>
<td>$\sim$120°</td>
<td>bent</td>
<td>SO$_2$</td>
</tr>
</tbody>
</table>
Shapes with 4 Electron Groups

<table>
<thead>
<tr>
<th>4</th>
<th>4</th>
<th>0</th>
<th>109.5°</th>
<th>tetrahedral</th>
<th>CH₄</th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td>3</td>
<td>1</td>
<td>~109.5°</td>
<td>trigonal pyramidal</td>
<td>NH₃</td>
</tr>
<tr>
<td>4</td>
<td>2</td>
<td>2</td>
<td>~109.5°</td>
<td>bent</td>
<td>H₂O</td>
</tr>
</tbody>
</table>
Learning Check

State the number of electron groups, lone pairs, and use VSEPR theory to determine the shape of the following molecules or ions.

1) tetrahedral  2) pyramidal  3) bent

A. PF₃
   4 electron groups, 1 lone pair, (2) pyramidal

B. H₂S
   4 electron groups, 2 lone pairs, (3) bent

C. CCl₄
   4 electron groups, 0 lone pairs, (1) tetrahedral

D. PO₄³⁻
   4 electron groups, 0 lone pairs, (1) tetrahedral
Electronegativity and Polarity

Unequal sharing of electrons in a polar covalent bond

H⁺—Cl⁻
Electronegativity

Electronegativity values

• Indicate the attraction of an atom for shared electrons.
• Increases from left to right going across a period on the periodic table.
• Is high for the nonmetals with fluorine as the highest.
• Is low for the metals.
Some Electronegativity Values for Group A Elements

Electronegativity increases

- High values
  - H: 2.1
  - Group 18: 2.8

- Low values
  - Cs: 0.7
  - Group 1: 1.8

Copyright © 2008 by Pearson Education, Inc.
Publishing as Benjamin Cummings
Nonpolar Covalent Bonds

A nonpolar covalent bond
• Occurs between nonmetals.
• Is an equal or almost equal sharing of electrons.
• Has almost no electronegativity difference (0.0 to 0.4).

Examples:

<table>
<thead>
<tr>
<th>Atoms</th>
<th>Electronegativity Difference</th>
<th>Type of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>N-N</td>
<td>3.0 - 3.0 = 0.0</td>
<td>Nonpolar covalent</td>
</tr>
<tr>
<td>Cl-Br</td>
<td>3.0 - 2.8 = 0.2</td>
<td>Nonpolar covalent</td>
</tr>
<tr>
<td>H-Si</td>
<td>2.1 - 1.8 = 0.3</td>
<td>Nonpolar covalent</td>
</tr>
</tbody>
</table>
Polar Covalent Bonds

A polar covalent bond
- Occurs between nonmetal atoms.
- Is an unequal sharing of electrons.
- Has a moderate electronegativity difference (0.5 to 1.7).

Examples:

<table>
<thead>
<tr>
<th>Atoms</th>
<th>Electronegativity Difference</th>
<th>Type of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>O-Cl</td>
<td>3.5 - 3.0 = 0.5</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>Cl-C</td>
<td>3.0 - 2.5 = 0.5</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>O-S</td>
<td>3.5 - 2.5 = 1.0</td>
<td>Polar covalent</td>
</tr>
</tbody>
</table>
Comparing Nonpolar and Polar Covalent Bonds

Equal sharing of electrons in a nonpolar covalent bond

Unequal sharing of electrons in a polar covalent bond
Ionic Bonds

An ionic bond
- Occurs between metal and nonmetal ions.
- Is a result of complete electron transfer.
- Has a large electronegativity difference (1.8 or more).

Examples:

<table>
<thead>
<tr>
<th>Atoms</th>
<th>Electronegativity Difference</th>
<th>Type of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl-K</td>
<td>3.0 – 0.8 = 2.2</td>
<td>Ionic</td>
</tr>
<tr>
<td>N-Na</td>
<td>3.0 – 0.9 = 2.1</td>
<td>Ionic</td>
</tr>
<tr>
<td>S-Cs</td>
<td>2.5 – 0.7 = 1.8</td>
<td>Ionic</td>
</tr>
</tbody>
</table>
## Predicting Bond Types

### Table 7.4 Predicting Bond Type from Electronegativity Differences

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Type of Electron Sharing</th>
<th>Electronegativity Difference$^a$</th>
<th>Bond Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>H$_2$</td>
<td>H—H</td>
<td>Shared equally</td>
<td>2.1 − 2.1 = 0</td>
</tr>
<tr>
<td>Cl$_2$</td>
<td>Cl—Cl</td>
<td>Shared equally</td>
<td>3.0 − 3.0 = 0</td>
</tr>
<tr>
<td></td>
<td>δ$^+$ δ$^-$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HBr</td>
<td>H—Br</td>
<td>Shared unequally</td>
<td>2.8 − 2.1 = 0.7</td>
</tr>
<tr>
<td></td>
<td>δ$^+$ δ$^-$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>HCl</td>
<td>H—Cl</td>
<td>Shared unequally</td>
<td>3.0 − 2.1 = 0.9</td>
</tr>
<tr>
<td>NaCl</td>
<td>Na$^+$ Cl$^-$</td>
<td>Electron transfer</td>
<td>3.0 − 0.9 = 2.1</td>
</tr>
<tr>
<td>MgO</td>
<td>Mg$^{2+}$ O$^{2-}$</td>
<td>Electron transfer</td>
<td>3.5 − 1.2 = 2.3</td>
</tr>
</tbody>
</table>

$^a$Values are taken from Figure 6.2.
### Range of Bond Types

**Table 7.5**: Electronegativity Difference and Types of Bonds

<table>
<thead>
<tr>
<th>Electronegativity difference</th>
<th>0</th>
<th>0.4</th>
<th>1.8</th>
<th>3.3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bond type</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Nonpolar</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Polar</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ionic</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Electron Bonding</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equally</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Unequally</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Electron transfer</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>+</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>-</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Copyright © 2008 by Pearson Education, Inc. Publishing as Benjamin Cummings
Learning Check

Use the electronegativity difference to identify the type of bond between the following as: nonpolar covalent (NP), polar covalent (P), or ionic (I).

A. K-N 2.2 ionic (I)
B. N-O 0.5 polar covalent (P)
C. Cl-Cl 0.0 nonpolar covalent (NP)
D. H-Cl 0.9 polar covalent (P)
A polar molecule
- Contains polar bonds.
- Has a separation of positive and negative charge called a dipole indicated with $\delta^+$ and $\delta$.
- Has dipoles that do not cancel.

H–Cl

H–N–H
dipoles do not cancel
Nonpolar Molecules

A nonpolar molecule
- Contains nonpolar bonds.
  Cl–Cl H–H

- Or has a symmetrical arrangement of polar bonds.

O=C=O

Cl–C–Cl

\[ \text{dipoles cancel} \]
Determining Molecular Polarity

**STEP 1** Write the electron-dot formula.

**STEP 2** Determine the polarity of the bonds.

**STEP 3** Determine if any dipoles cancel or not.

Example: $\text{H}_2\text{O}$

$\text{H} - \text{O}$: $\text{H}_2\text{O}$ is polar
dipoles do not cancel
Learning Check

Identify each of the following molecules as 1) polar or 2) nonpolar. Explain.

A. \( \text{PBr}_3 \) 1) pyramidal; dipoles don’t cancel; **polar**
B. \( \text{HBr} \) 1) linear; one polar bond (dipole); **polar**
C. \( \text{Br}_2 \) 2) linear; nonpolar bond; **nonpolar**
D. \( \text{SiBr}_4 \) 2) tetrahedral; dipoles cancel; **nonpolar**
Attractive Forces Between Particles
Ionic Bonds

In ionic compounds, ionic bonds

- Are strong attractive forces.
- Hold positive and negative ions together.

### Table: Type of Force

<table>
<thead>
<tr>
<th>Type of Force</th>
<th>Particle Arrangement</th>
<th>Energy (kJ/mol)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Between atoms or ions</td>
<td>+ - + -</td>
<td>500–5000</td>
<td>Na⁺ ⋯ Cl⁻</td>
</tr>
</tbody>
</table>

Copyright © 2008 by Pearson Education, Inc.
Publishing as Benjamin Cummings
In covalent compounds, polar molecules

- Exert attractive forces called **dipole-dipole attractions**.
- Form strong dipole attractions called **hydrogen bonds** between hydrogen atoms bonded to F, O, or N.
Dipole-Dipole Attractions

<table>
<thead>
<tr>
<th>Type of Force</th>
<th>Particle Arrangement</th>
<th>Energy (kJ/mol)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Between molecules</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen bond (X = F, O, or N)</td>
<td>$\delta^+ \delta^- \cdot \delta^+ \delta^-$</td>
<td>10–40</td>
<td>$\delta^+ \delta^- \cdot \delta^+ \delta^-$</td>
</tr>
<tr>
<td>Dipole–dipole (X and Y = different nonmetals)</td>
<td>$\delta^+ \delta^- \cdot \delta^+ \delta^-$</td>
<td>5–20</td>
<td>$\delta^+ \delta^- \cdot \delta^+ \delta^-$</td>
</tr>
</tbody>
</table>

Copyright © 2008 by Pearson Education, Inc.
Publishing as Benjamin Cummings
Dispersion Forces

Dispersion forces are

- Weak attractions between nonpolar molecules.
- Caused by temporary dipoles that develop when electrons are not distributed equally.

<table>
<thead>
<tr>
<th>Type of Force</th>
<th>Particle Arrangement</th>
<th>Energy (kJ/mol)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Dispersion</td>
<td>$\delta^+ \delta^- \delta^+ \delta^-$ (temporary dipoles)</td>
<td>0.04 – 10</td>
<td>$\delta^+ \delta^- \delta^+ \delta^-$</td>
</tr>
</tbody>
</table>

Copyright © 2008 by Pearson Education, Inc.
Publishing as Benjamin Cummings
## Comparison of Bonding and Attractive Forces

<table>
<thead>
<tr>
<th>Type of Force</th>
<th>Particle Arrangement</th>
<th>Energy (kJ/mol)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Between atoms or ions</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ionic bond</td>
<td>![Ion Bond Image]</td>
<td>500 – 5000</td>
<td>Na⁺ (\cdots) Cl⁻</td>
</tr>
<tr>
<td>Covalent bond (X = nonmetal)</td>
<td>![Covalent Bond Image]</td>
<td>100 – 1000</td>
<td>Cl – Cl</td>
</tr>
<tr>
<td>Between molecules</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen bond (X = F, O, or N)</td>
<td>![Hydrogen Bond Image]</td>
<td>10 – 40</td>
<td>H – F (\cdots) H – F</td>
</tr>
<tr>
<td>Dipole–dipole (X and Y = different nonmetals)</td>
<td>![Dipole–dipole Image]</td>
<td>5 – 20</td>
<td>Br – Cl (\cdots) Br – Cl</td>
</tr>
<tr>
<td>Dispersion (Temporary shift of electrons in nonpolar bonds)</td>
<td>![Dispersion Image]</td>
<td>0.04 – 10</td>
<td>F – F (\cdots) F – F</td>
</tr>
</tbody>
</table>
Melting Points and Attractive Forces

- Ionic compounds require large amounts of energy to break apart ionic bonds. Thus, they have high melting points.
- Hydrogen bonds are the strongest type of dipole-dipole attractions. They require more energy to break than other dipole attractions.
- Dispersion forces are weak interactions and very little energy is needed to change state.
Learning Check

Identify the main type of attractive forces for each:
1) ionic 2) dipole-dipole
3) hydrogen bonds 4) dispersion

A. NCl₃
B. H₂O
C. Br-Br
D. KCl
E. NH₃
Learning Check

Identify the main type of attractive forces for each compound below as:

1) ionic  
2) dipole-dipole  
3) hydrogen bonds  
4) dispersion

A. NCl₃  
B. H₂O  
C. Br-Br  
D. KCl  
E. NH₃

A. NCl₃  2  dipole-dipole  
B. H₂O  3  hydrogen bonds  
C. Br-Br  4  dispersion  
D. KCl  1  ionic  
E. NH₃  3  hydrogen bonds
Matter and Changes of State
Melting and Freezing

A substance

- Is **melting** while it changes from a solid to a liquid.
- Is **freezing** while it changes from a liquid to a solid.
- Such as water has a freezing (melting) point of 0°C.
Heat of Fusion

The heat of fusion

• Is the amount of heat released when 1 gram liquid freezes (at its freezing point).
Sublimation

- Occurs when particles absorb heat to change directly from solid to a gas.
- Is typical of dry ice, which sublimes at -78°C.
- Takes place in frost-free refrigerators.
- Is used to prepare freeze-dried foods for long-term storage.
Evaporation and Condensation

Water

- **Evaporates** when molecules on the surface gain sufficient energy to form a gas.
- **Condenses** when gas molecules lose energy and form a liquid.
Boiling

At boiling,

- All the water molecules acquire enough energy to form a gas.
- Bubbles appear throughout the liquid.
Heat of Vaporization

The heat of vaporization is the amount of heat
- Absorbed to vaporize 1 g of a liquid to gas at the boiling point.
- Released when 1 g of a gas condenses to liquid at the boiling point.
Heats of Vaporization and Fusion of Polar and Nonpolar Compounds

<table>
<thead>
<tr>
<th>Nonpolar covalent</th>
<th>Polar covalent</th>
<th>Ionic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Propane C₃H₈</td>
<td>Benzene C₆H₆</td>
<td>Ammonia NH₃</td>
</tr>
<tr>
<td>18 J/g</td>
<td>128 J/g</td>
<td>351 J/g</td>
</tr>
<tr>
<td>336 J/g</td>
<td>395 J/g</td>
<td>390 J/g</td>
</tr>
<tr>
<td>192 J/g</td>
<td>Ethanol C₂H₅OH</td>
<td>Ammonia NH₃</td>
</tr>
<tr>
<td>109 J/g</td>
<td>351 J/g</td>
<td>334 J/g</td>
</tr>
<tr>
<td>334 J/g</td>
<td>Ethanol C₂H₅OH</td>
<td>Ammonia NH₃</td>
</tr>
<tr>
<td>109 J/g</td>
<td>351 J/g</td>
<td>334 J/g</td>
</tr>
<tr>
<td>880 J/g</td>
<td>Ethanol C₂H₅OH</td>
<td>Ammonia NH₃</td>
</tr>
<tr>
<td>390 J/g</td>
<td>351 J/g</td>
<td>334 J/g</td>
</tr>
<tr>
<td>336 J/g</td>
<td>Benzene C₆H₆</td>
<td>Ammonia NH₃</td>
</tr>
<tr>
<td>128 J/g</td>
<td>395 J/g</td>
<td>351 J/g</td>
</tr>
<tr>
<td>18 J/g</td>
<td>Propane C₃H₈</td>
<td>Ammonia NH₃</td>
</tr>
</tbody>
</table>

Heats of vaporization (J/g) and heats of fusion (J/g).
Summary of Changes of State
A heating curve

- Illustrates the changes of state as a solid is heated.
- Uses sloped lines to show an increase in temperature.
- Uses plateaus (flat lines) to indicate a change of state.
A cooling curve

- Illustrates the changes of state as a gas is cooled.
- Uses sloped lines to indicate a decrease in temperature.
- Uses plateaus (flat lines) to indicate a change of state.