Lecture Presentation

Chapter 9

Chemical Bonding I: The Lewis Model
**Conductivity of NaCl**

In NaCl\((s)\), the ions are stuck in position and not allowed to move to the charged rods.

In NaCl\((aq)\), the ions are separated and allowed to move to the charged rods.
Lewis Theory of Covalent Bonding

• Lewis theory implies that another way atoms can achieve an octet of valence electrons is to share their valence electrons with other atoms.

• The shared electrons would then count toward each atom’s octet.

• The sharing of valence electrons is called covalent bonding.
Covalent Bonding: Bonding and Lone Pair Electrons

• Electrons that are shared by atoms are called **bonding pairs**.
• Electrons that are not shared by atoms but belong to a particular atom are called **lone pairs**.
  – Also known as **nonbonding pairs**
• End 10/31/16 Monday Lecture
• Start 11/2/16 Wednesday lecture

Show example of Lewis Dot structure illustrating ionic bond (selfish atoms) vs covalent bonding (friendly & cooperative atoms)
Single Covalent Bonds

- When two atoms share one pair of electrons, it is called a **single covalent bond**.
  - Two electrons
- One atom may use more than one single bond to fulfill its octet.
  - To different atoms
  - H only duet
Double Covalent Bond

• When two atoms share two pairs of electrons the result is called a **double covalent bond**.
  – Four electrons

\[
\begin{align*}
\cdot\overset{\cdot}{\O} : + \cdot\overset{\cdot}{\O} : \\
\downarrow \\
\cdot\overset{\cdot}{\O} : & : \overset{\cdot}{\O} : \text{ or } \cdot\overset{\cdot}{\O} = \overset{\cdot}{\O} : \\
\text{Octet} & \quad \text{Octet}
\end{align*}
\]
Triple Covalent Bond

• When two atoms share three pairs of electrons the result is called a **triple covalent bond**.
  – Six electrons

\[ \text{N::N: or N≡N:} \]
Predictions of Molecular Formulas by Lewis Theory

Oxygen is more stable when it is singly bonded to two other atoms.

\[
\begin{align*}
\text{H} & \overset{\cdot}{\text{O}} - \text{H} \\
\text{H} \overset{\cdot}{\text{O}} - \overset{\cdot}{\text{O}} - \text{H}
\end{align*}
\]

Oxygen has nine electrons (one electron beyond an octet)
Covalent Bonding: Model versus Reality

• Lewis theory of covalent bonding implies that the attractions between atoms are *directional*.
  – The shared electrons are most stable between the bonding atoms.

• Therefore, Lewis theory predicts covalently bonded compounds will be found as individual molecules.
  – Rather than an array like ionic compounds

• **Compounds of nonmetals are made of individual molecule units.**
Covalent Bonding: Model versus Reality

• Lewis theory predicts that the melting and boiling points of molecular compounds should be relatively low.
  – This involves breaking the attractions between the molecules but not the bonds between the atoms.
  – The covalent bonds are strong, but the attractions between the molecules are generally weak.

• Molecular compounds have low melting points and boiling points.
  – Melting points generally < 300 °C
  – Molecular compounds are found in all three states at room temperature.
Intermolecular Attractions versus Bonding

Molecular Compound

- Strong covalent bonds *within* molecules
- Weaker intermolecular forces *between* molecules

C₅H₁₂(g) → C₅H₁₂(l)
Covalent Bonding: Model versus Reality

• Lewis theory predicts that neither molecular solids nor liquids should conduct electricity.
  – There are no charged particles around to allow the material to conduct.

• Molecular compounds do not conduct electricity in the solid or liquid state.

• Molecular acids conduct electricity when dissolved in water but not in the solid or liquid state, due to them being ionized by the water.
Covalent Bonding: Model versus Reality

- Lewis theory predicts that the more electrons two atoms share, the stronger the bond should be.

- Bond strength is measured by how much energy must be added into the bond to break it in half.

- In general, triple bonds are stronger than double bonds, and double bonds are stronger than single bonds.
  
  - However, Lewis theory would predict that double bonds are twice as strong as single bonds; the reality is that they are less than twice as strong.
Covalent Bonding: Model versus Reality

• Lewis theory predicts that the more electrons two atoms share, the shorter the bond should be.
  – When comparing bonds to like atoms

• Bond length is determined by measuring the distance between the nuclei of bonded atoms.

• In general, triple bonds are shorter than double bonds, and double bonds are shorter than single bonds.
Polar Covalent Bonding

• Covalent bonding between unlike atoms results in unequal sharing of the electrons.
  – One atom pulls the electrons in the bond closer to its side.
  – One end of the bond has larger electron density than the other.

• The result is a polar covalent bond.
  – Bond polarity
  – The end with the larger electron density gets a partial negative charge.
  – The end that is electron deficient gets a partial positive charge.
Unequal Electron Sharing: HF
Electronegativity

• The ability of an atom to attract bonding electrons to itself is called **electronegativity**.

• Increases across period (left to right) and decreases down group (top to bottom)
  – Fluorine is the most electronegative element.
  – Francium is the least electronegative element.
  – Noble gas atoms are not assigned values.
  – Opposite of atomic size trend

• The larger the difference in electronegativity, the more polar the bond.
  – Negative end toward more electronegative atom
Electronegativity Scale

Trends in Electronegativity

Period 1
- H 2.1
- Li 1.0
- Be 1.5

Period 2
- B 2.0
- C 2.5
- N 3.0
- O 3.5
- F 4.0

Period 3
- Li 1.0
- Be 1.5
- B 2.0
- C 2.5
- N 3.0
- O 3.5
- F 4.0

Period 4
- Li 1.0
- Be 1.5
- B 2.0
- C 2.5
- N 3.0
- O 3.5
- F 4.0

Period 5
- Li 1.0
- Be 1.5
- B 2.0
- C 2.5
- N 3.0
- O 3.5
- F 4.0

Period 6
- Li 1.0
- Be 1.5
- B 2.0
- C 2.5
- N 3.0
- O 3.5
- F 4.0

Electronegativity Scale:
- 0.7–1.1
- 1.2–1.6
- 1.7–2.1
- 2.2–2.6
- 2.7–4.0
Electronegativity Difference and Bond Type

• If the difference in electronegativity between bonded atoms is 0, the bond is **pure covalent**.
  – Equal sharing
• If the difference in electronegativity between bonded atoms is 0.1 to 0.4, the bond is **nonpolar covalent**.
• If the difference in electronegativity between bonded atoms is 0.4 to 1.9, the bond is **polar covalent**.
• If the difference in electronegativity between bonded atoms is larger than or equal to 2.0, the bond is “100%” **ionic**.
• Most electronegative element is Florine.
Effect of Electronegativity Difference on Bond Type

<table>
<thead>
<tr>
<th>Electronegativity Difference ($\Delta EN$)</th>
<th>Bond Type</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Small (0–0.4)</td>
<td>Covalent</td>
<td>Cl$_2$</td>
</tr>
<tr>
<td>Intermediate (0.4–2.0)</td>
<td>Polar covalent</td>
<td>HCl</td>
</tr>
<tr>
<td>Large (2.0+)</td>
<td>Ionic</td>
<td>NaCl</td>
</tr>
</tbody>
</table>
The Continuum of Bond Types

- Pure (nonpolar) covalent bond: Electrons shared equally.
- Polar covalent bond: Electrons shared unequally.
- Ionic bond: Electrons transferred.

Electronegativity difference, $\Delta EN$:
- 0.0 to 0.4
- 0.4 to 2.0
- 2.0 to 3.3
Bond Dipole Moments

• Dipole moment, \( \mu \), is a measure of bond polarity.
  – A dipole is a material with a + and − end.
  – It is directly proportional to the size of the partial charges and directly proportional to the distance between them.
    • \( \mu = (q)(r) \)
    • Not Coulomb’s law
    • Measured in Debyes, D

• Generally, the more electrons two atoms share and the larger the atoms are, the larger the dipole moment.
### Dipole Moments

**TABLE 9.2 Dipole Moments of Several Molecules in the Gas Phase**

<table>
<thead>
<tr>
<th>Molecule</th>
<th>ΔEN</th>
<th>Dipole Moment (D)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl(_2)</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>ClF</td>
<td>1.0</td>
<td>0.88</td>
</tr>
<tr>
<td>HF</td>
<td>1.9</td>
<td>1.82</td>
</tr>
<tr>
<td>LiF</td>
<td>3.0</td>
<td>6.33</td>
</tr>
</tbody>
</table>
Writing Lewis Structures of Molecules

1. Write the correct skeletal structure for the molecule. **Examples on board**
   - Hydrogen atoms are always terminal.
   - The more electronegative atoms are placed in terminal positions.

2. Calculate the total number of electrons for the Lewis structure by summing the valence electrons of each atom in the molecule.

3. Distribute the electrons among the atoms, giving octets (or duets in the case of hydrogen) to as many atoms as possible.

4. If any atoms lack an octet, form double or triple bonds as necessary to give them octets.
End 11/2/16 lecture
Start 11/7/16 lecture