Chapter 10: Chemical Bonding

Chemical Bonds are the **attractive forces** that hold atoms together in more complex units.

- An understanding of **how** and **why** atoms attach together in the manner they do is central to chemistry

**Lewis Theory**

Emphasizes the importance of **valence electrons**

**VALENCE ELECTRONS** - electrons in the outermost shell - ARE THE ELECTRONS INVOLVED IN BONDING

Uses **dots** to represent valence electrons either **ON** or **SHARED** by atoms

**Bonding between atoms occurs by either** transfer or sharing of electrons to achieve outer shells with **8 electrons** (exceptions: Li, Be and He)

**Lewis Electron Dot Symbols**

- Uses symbol of element to represent the **nucleus and inner (core) electrons**
- Uses dots around the symbol to represent **valence electrons**
  (Put one electron on each side first, then pair)

Li:
Be:
B:
C:
N:
O:
F:
Ne:

Remember that elements in the **same group** have the **same number of valence electrons**; therefore, their Lewis dot symbols will **look alike**.
Noble gases are considered stable because they do not react with other elements. This stability is attributed to their **FULL VALENCE SHELLS** - they have a complete OCTET of ELECTRONS (exception: helium)

- He: \(1s^2\)
- Ne: \(1s^22s^22p^6\)
- Ar: \(1s^22s^22p^63s^23p^6\)
- Kr: \(1s^22s^22p^63s^23p^64s^23d^{10}4p^6\)
- Xe: \(1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6\)

### Formation of Ionic Compounds

Atoms of many elements that **LACK** a complete octet of electrons in their outer shells **react in such a way to attain it.**

- They may **lose** or **gain** electrons depending on the type of element the atom is (metal or nonmetal)

**Recall:** When a neutral atom loses or gains one or more electrons an ion is formed.

- **Ion formation occurs when atoms of two elements (a metal and nonmetal) are present.**

The metal will lose one or more electrons - forms a cation

The nonmetal will gain one or more electrons - forms an anion

\[
\text{Li}^+ + \text{F}^- \rightarrow \text{Li}^+ \left[\text{:F:}\right]^\text{-}
\]

**Lithium cation**  **fluoride anion**
Formation of Ionic Compounds

\[ \text{Li}^+ + \text{F}^- \rightarrow \text{Li}^+ [\text{F}^-]^- \]

**Example:**

- **Li** has a full valence shell - it has an electron configuration *just like He*
- **F** has a full valence shell (a complete octet) - it has an electron configuration *just like Ne*

Show how the following ionic compounds form using Lewis Electron Dot Symbols.

- **a)** Na\(_2\)S
- **b)** K\(_3\)P
- **c)** MgBr\(_2\)

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**Covalent Bonds**

Covalent bonding results from the **SHARING** of one or more electron pairs between atoms

- Most **nonmetal elements** try to achieve a **NOBLE GAS CONFIGURATION** by sharing electrons with other nonmetals

**Example:**

- **H** wants to be like He

\[ \text{H} \cdot \rightarrow \text{H}:\text{Cl}: \cdot \]

- **Cl** wants to be like Ar

\[ \cdot \text{Cl}: \rightarrow \cdot \text{Cl}: \cdot \]

By sharing e-s, **both** attain a noble gas configuration
Covalent Lewis Structures

\[ \cdot\cdot H\cdot:\cdot Cl\cdot:\cdot \]
\[ \cdot\cdot H\cdot:\cdot Cl\cdot:\cdot \]

The HCl molecule has 1 bonding pair of electrons and 3 nonbonding pairs of electrons (also called "lone pairs")

In Lewis theory:

- Nonmetal atoms share electrons to complete their octet, called:
  "OCTET RULE" (Note: there are exceptions to the octet rule)

- Completing octets may involve sharing electrons with multiple atoms or sharing multiple pairs of electrons with the same atom.

\[ \cdot\cdot O\equiv\cdot\cdot \]

Double covalent bond
Atoms share two e⁻ pairs

\[ \cdot\cdot H\cdot:\cdot C\cdot:\cdot H\cdot:\cdot \]

Single covalent bonds
Atoms share one pair of electrons
(carbon sharing e⁻s with multiple hydrogen atoms)

\[ \cdot\cdot N\equiv\cdot\cdot \]

Triple covalent bond
Atoms share three e⁻ pairs
Rules for Writing Lewis Structures for Covalent Compounds

1. Decide on the central atom (it will never be H or F). *(The central atom is usually the one that is by itself)*

2. Determine the **total number of valence electrons** in the structure.
   - If the structure is an ion:
     - Add 1e\(^-\) for a -1 charge, add 2e\(^-\) for a -2 charge, etc...
     - Subtract 1e\(^-\) for a +1 charge, subtract 2e\(^-\) for +2, etc...

3. Form covalent bonds between the central atom and the surrounding atoms - called the "skeletal structure". Count how many electrons have been used to form these bonds.

4. Subtract electrons used to form covalent bonds from total number of valence electrons in the molecule or ion to determine how many electrons remain (if any).

5. Any remaining electrons become **lone pairs**, FIRST ON THE OUTSIDE ATOMS to complete their octets, and then on the central atom.

6. If any atoms that need an octet of electrons do not have it, form double and triple bonds as necessary by bringing outer atom lone pair electrons down between two atoms so they can share them.
   - **Important:** C, N, O and F **always** follow the octet rule

Write the Lewis structure for NH\(_3\)

1. Central atom?
2. Valence electrons?
3. Form skeletal structure
4. Remaining electrons?
5. Lewis Structure?
C, N, O and F ALWAYS have 8 electrons surrounding them in their Lewis structures. They always follow the octet rule.

NOTE: Boron may have less than 8e⁻s (usually 6). Phosphorus and atoms of higher atomic number have AT LEAST 8e⁻s but sometimes can have more than 8e⁻s in Lewis structures. These are elements that can be exceptions to the octet rule. **We will not write Lewis structures of this type.**

Write the Lewis Structures for the following molecules and polyatomic ions.

a) CH₄  

b) CH₂O  

c) O₂  

d) CO₂
Resonance Structures

Write the Lewis Structures for the following molecules and polyatomic ions (continued).

e) HCN  
f) SO$_3^{2-}$  
g) NH$_2$Br$_2^+$

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**Resonance Structures**

Sometimes we can draw more than one valid Lewis structure for a molecule or polyatomic ion.

- In this situation, no **one** Lewis structure can adequately describe the actual structure of the molecule.
- The actual molecule or ion will have characteristics of **all the valid Lewis structures** that can be drawn. (It is a **hybrid** of these Lewis structures).

Write the Lewis structures of:

a) SO$_2$  
b) HCO$_2^-$
Shapes of Molecules

The most important factor in determining the shape of a molecule or polyatomic ion is the *relative repulsion between electron pairs.*

- A molecule or ion will be **most stable** when the electron pairs or groups are **as far apart as possible.**

**Number of Electron "Groups" around the Central Atom**

<table>
<thead>
<tr>
<th>Number of Electron &quot;Groups&quot;</th>
<th>Electron Geometry &amp; Molecular Geometry</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal Planar</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral</td>
</tr>
</tbody>
</table>

(Molecular Geometry = Shape of Molecule)
Determining Molecular Geometry Using VSEPR Theory

Valence Shell Electron Pair Repulsion (VSEPR) Theory: Repulsion between the negative charges on electron groups determines the molecular geometry (shape) of a molecule or polyatomic ion.

Determine the Electron and Molecular Geometry of:

a) NH₃

1ˢᵗ - Determine the Lewis structure of NH₃

2ⁿᵈ - Count electron "Groups" around the central atom → 4

3ʳᵈ - Determine Electron Geometry then Molecular Geometry

b) H₂O

Four electron groups are at the corners of a tetrahedron

Electron Geometry: Tetrahedral

Molecular Geometry: Trigonal Pyramidal
Determining Molecular Geometry

Lewis Structure of H₂O

\[
\text{H} \quad \hat{\text{O}} \quad \text{H}
\]

4 Groups of Electrons

Electron Pair Geometry = Tetrahedral
Molecular Geometry = Bent

**TABLE 10.1 Electron and Molecular Geometries**

<table>
<thead>
<tr>
<th>Electron Groups*</th>
<th>Bonding Groups</th>
<th>Lone Pairs</th>
<th>Electron Geometry</th>
<th>Angle between Electron Groups**</th>
<th>Molecular Geometry</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>2</td>
<td>0</td>
<td>linear</td>
<td>180°</td>
<td>linear</td>
<td>:\text{O}–\text{C}–\text{O}:</td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>0</td>
<td>trigonal planar</td>
<td>120°</td>
<td>trigonal planar</td>
<td>:\text{O}–\text{C}–\text{O}:</td>
</tr>
<tr>
<td>3</td>
<td>2</td>
<td>1</td>
<td>trigonal planar</td>
<td>120°</td>
<td>bent</td>
<td>:\text{O}–\text{S}–\text{O}:</td>
</tr>
<tr>
<td>4</td>
<td>4</td>
<td>0</td>
<td>tetrahedral</td>
<td>109.5°</td>
<td>tetrahedral</td>
<td>\text{H}–\text{C}–\text{H}–\text{H}</td>
</tr>
<tr>
<td>4</td>
<td>3</td>
<td>1</td>
<td>tetrahedral</td>
<td>109.5°</td>
<td>trigonal pyramidal</td>
<td>\text{H}–\text{N}–\text{H}–\text{H}</td>
</tr>
<tr>
<td>4</td>
<td>2</td>
<td>2</td>
<td>tetrahedral</td>
<td>109.5°</td>
<td>bent</td>
<td>\text{H}–\text{O}–\text{H}</td>
</tr>
</tbody>
</table>

*Count only electron groups around the central atom. Each of the following is considered one electron group: a lone pair, a single bond, a double bond, and a triple bond.

**Angles listed here are idealized. Actual angles in specific molecules may vary by several degrees.

Determine the Electron and Molecular Geometry of:

a) CCl₄

b) HCN
Bond Polarity

Determine the Electron and Molecular Geometry of:

c) CH$_2$S  
d) SO$_2$

e) H$_2$S  
f) PH$_3$

Practice for next class:

Bond Polarity

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 BOND**.

The end with the larger electron density gets a partial negative charge (δ⁻) and the end that is electron deficient gets a partial positive charge (δ⁺).

Example:  

HCl  

δ+  δ⁻
Electronegativity

How can we determine which atom is $\delta^+$ and which is $\delta^-$ in a polar bond?

➢ Use electronegativity values of the atoms

Electronegativity

Electronegativity is a measure of relative attraction that an atom has for the shared electrons in a covalent bond.

Electronegativity

- Increases across the period (left to right)
- Decreases down the group (top to bottom)

The larger the difference in electronegativity, the more polar the bond.
Main Classes of Chemical Bonds

**Nonpolar Covalent Bond**
- Electronegativity difference between zero and 0.4
- Many times between two identical atoms
  
  Example:
  
  \[ \text{H—H} \]

**Polar Covalent Bond**
- Electronegativity difference between 0.4 and 2
- Between two different NONMETAL atoms

**Ionic Bond**
- Electronegativity difference is greater than 2
- Primarily exists between METALS and NONMETALS
Polar Molecules

Can show the direction of bond polarity with $\delta^+$ and $\delta^-$ and/or a special arrow:

![Direction of Bond Polarity](image)

Show the direction of bond polarity for the bond in HCl.

In order for a molecule to be polar:
- It must have polar bonds
- AND
- It must have an unsymmetrical shape

If there are no polar bonds, then molecule is NONPOLAR

![F-F](image)

If there are polar bonds and the bond dipoles cancel out
- Molecule is NONPOLAR

![Nonpolar Molecule](image)

If there are polar bonds and the bond dipoles DO NOT cancel out,
- Molecule is POLAR

![Polar Molecule](image)
Molecular Polarity

TABLE 10.3 Common Cases of Adding Dipole Moments to Determine Whether a Molecule Is Polar

Determine whether the following molecules are POLAR or NONPOLAR.

a) CF₄
b) CH₂F₂

c) NH₃
d) NOCl

Lewis Structure

\[
\begin{align*}
\text{O} &= \text{N} - \text{Cl} \\
\end{align*}
\]
Molecular Polarity Affects Solubility

- Polar molecules are attracted to other polar molecules
- Since water is a polar molecule, other polar molecules dissolve well in water (Many ionic compounds dissolve in water as well).
- Nonpolar molecules are attracted to other nonpolar molecules and dissolve in each other

"LIKE DISSOLVES LIKE"