Chapter 12

Chemical Bonding
Chemical Bonding

• **Valence electrons** are the electrons in the outer shell (highest energy level) of an atom.

• A **chemical bond** is a mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together.

• During bonding, valence electrons are redistributed in ways that make the atoms more stable.
The Three Major Types of Chemical Bonding

• **Ionic Bonding** results from the electrical attraction between oppositely-charged ions.

• **Covalent Bonding** results from the sharing of electron pairs between two atoms.

• **Metallic Bonding** results from the attraction between metal atoms and the surrounding sea of electrons.
Ionic or Covalent?

- Bonding is usually somewhere between ionic and covalent, depending on the electronegativity difference between the two atoms.
- In **polar** covalent bonds, the bonded atoms have an unequal attraction for the shared electron.
Use electronegativity values (in table on pg 362) 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?

**Solution:**

<table>
<thead>
<tr>
<th>Bonding between sulfur and:</th>
<th>Electroneg. difference</th>
<th>Bond type</th>
<th>More negative atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen</td>
<td>2.5 − 2.1 = 0.4</td>
<td>polar-covalent</td>
<td>sulfur</td>
</tr>
<tr>
<td>cesium</td>
<td>2.5 − 0.7 = 1.8</td>
<td>ionic</td>
<td>sulfur</td>
</tr>
<tr>
<td>chlorine</td>
<td>3.0 − 2.5 = 0.5</td>
<td>polar-covalent</td>
<td>chlorine</td>
</tr>
</tbody>
</table>
Molecules

- A **covalent bond** is formed from *shared pairs* of electrons.
- A **molecule** is a neutral group of atoms held together by **covalent bonds**.
Why Do Covalent Bonds Form?

- When two atoms form a covalent bond, their shared electrons form overlapping orbitals.
- This gives both atoms a stable noble-gas configuration.
The Octet Rule

- Atoms are the most stable when they have completely full valence shells (like the noble gases).
- **The Octet Rule** — Compounds tend to form so that each atom has an octet (group of eight) electrons in its highest energy level.
- *Hydrogen* is an exception to the octet rule since it can only have two electrons in its valence shell.

Visual Concept
Electron-Dot Notation

- **Electron-dot notation** is indicated by dots placed around the element’s symbol. Only the valence electrons are shown. Inner-shell electrons are not shown.

<table>
<thead>
<tr>
<th>Number of valence electrons</th>
<th>Electron-dot notation</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>X·</td>
<td>Na·</td>
</tr>
<tr>
<td>2</td>
<td>·X·</td>
<td>·Mg·</td>
</tr>
<tr>
<td>3</td>
<td>··X·</td>
<td>··B·</td>
</tr>
<tr>
<td>4</td>
<td>···X·</td>
<td>···C·</td>
</tr>
<tr>
<td>5</td>
<td>····X·</td>
<td>····N·</td>
</tr>
<tr>
<td>6</td>
<td>·····X·</td>
<td>·····O·</td>
</tr>
<tr>
<td>7</td>
<td>······X·</td>
<td>······F·</td>
</tr>
<tr>
<td>8</td>
<td>·······X·</td>
<td>·······Ne·</td>
</tr>
</tbody>
</table>
Electron-Dot Notation

Sample Problem

a. Write the electron-dot notation for hydrogen.
b. Write the electron-dot notation for nitrogen.

Solution:

a. Hydrogen is in group 1. It has one valence electron.

\[ \text{H} : \]

a. Nitrogen is in group 15. It has 5 valence electrons.

\[ \text{N} : \]
Lewis Structures

- Electron-dot notations of two or more atoms can be combined to represent molecules.
- Unpaired electrons will pair up to form a shared pair or covalent bond.
• The pair of dots representing the shared pair of electrons in a covalent bond is often replaced by a long dash.

• 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.
How to Draw Lewis Structures

1. Draw the electron-dot notation for each type of atom, and count the valence electrons.
2. Put the least electronegative atom in the center (except H.)
3. Use electron pairs to form bonds between all atoms.
4. Make sure all atoms (except H) have octets.
5. Count the total electrons in your Lewis structure. Does it match the number you counted in step 1? If not, introduce multiple bonds.
Lewis Structures
Sample Problem A

Draw the Lewis structure of iodomethane, CH$_3$I.

Solution:

Step 1 - Draw the electron-dot notation for each type of atom, and count the valence electrons.

\[ \text{C} \quad 1 \times 4 \text{e}^- = 4 \text{e}^- \\
3\text{H} \quad 3 \times 1 \text{e}^- = 3 \text{e}^- \\
\text{I} \quad 1 \times 7 \text{e}^- = 7 \text{e}^- \\
\]

\[ \text{Total} = 14 \text{e}^- \]
**Step 2** – Put the least electronegative atom in the center (except H).

**Step 3** – Use electron pairs to form bonds between all atoms.

**Step 4** – Make sure all atoms (except H) have octets.

**Step 5** – Count the total electrons. Does it match your beginning total?

14 Total e⁻
Multiple Covalent Bonds

• In a single covalent bond, one pair of electrons is shared between two atoms.

• A double bond is a covalent bond in which two pairs of electrons are shared between two atoms.

• A triple bond is a covalent bond in which three pairs of electrons are shared between two atoms.

• Multiple bonds are often found in molecules containing carbon, nitrogen, and oxygen.
Draw the Lewis structure for methanal, CH₂O.

**Solution:**

**Step 1** - Draw the electron-dot notation for each type of atom, and count the valence electrons.

\[
\begin{align*}
\text{C} & : \quad 1 \times 4 \text{ e}^- = 4 \text{ e}^- \\
2\text{H} & : \quad 2 \times 1 \text{ e}^- = 2 \text{ e}^- \\
\text{O} & : \quad 1 \times 6 \text{ e}^- = 6 \text{ e}^- \\
\hline
\text{Total} & : \quad 12 \text{ e}^-
\end{align*}
\]
Lewis Structures

Sample Problem B (continued)

**Step 2** – Put the least electronegative atom in the center (except H).

**Step 3** – Use electron pairs to form bonds between all atoms.

**Step 4** – Make sure all atoms (except H) have octets.

**Step 5** – Count the total electrons. Does it match your beginning total? 14 Total e⁻  
If not, introduce multiple bonds (remove 2 lone pairs to make 1 shared pair.)

Now does it match? 12 Total e⁻
Formation of Ionic Compounds

- Sodium and other metals easily lose electrons to form positively-charged ions called **cations**.
- Chlorine and other non-metals easily gain electrons to form negatively-charged ions called **anions**.
Ionic Bonding

• Cations (+) and anions (-) are attracted to each other because of their opposite electrical charges.

• An ionic bond is a bond that forms between oppositely-charged ions because of their mutual electrical attraction.
Ionic Bonding and the Crystal Lattice

• In an ionic crystal, ions minimize their potential energy by combining in an orderly arrangement known as a **crystal lattice**.

• A **formula unit** is the smallest repeating unit of an ionic compound.

Sodium Chloride crystal lattice (many Na and Cl atoms)
Formula Unit = NaCl
Comparing Ionic and Covalent Compounds

- Covalent compounds have relatively weak forces of attraction between molecules, but ionic compounds have a strong attraction between ions. This causes some differences in their properties:

<table>
<thead>
<tr>
<th>Ionic</th>
<th>Covalent</th>
</tr>
</thead>
<tbody>
<tr>
<td>crystals</td>
<td>molecules</td>
</tr>
<tr>
<td>very high melting points</td>
<td>low melting points</td>
</tr>
<tr>
<td>hard, but brittle</td>
<td>usually gas or liquid</td>
</tr>
<tr>
<td>Ex: NaCl, CaF$_2$, KNO$_3$</td>
<td>Ex: H$_2$O, CO$_2$, O$_2$</td>
</tr>
</tbody>
</table>
Polyatomic Ions

• A charged group of covalently bonded atoms is known as a **polyatomic ion**.

• Draw a Lewis structure for a polyatomic ion with brackets around it and the charge in the upper right corner.

\[
\begin{align*}
\text{hydroxide ion, } &\text{OH}^- \\
\text{ammonium ion, } &\text{NH}_4^+ \\
\end{align*}
\]
VSEPR Theory

• The abbreviation VSEPR (say it “VES-pur”) stands for “valence-shell electron-pair repulsion.”

• **VSEPR theory** — repulsion between pairs of valence electrons around an atom causes the electron pairs to be oriented as far apart as possible.

• Treat double and triple bonds the same as single bonds.
VSEPR theory can also account for the geometries of molecules with unshared electron pairs.

VSEPR theory postulates that the lone pairs occupy space around the central atom just like bonding pairs, but they repel other electron pairs more strongly than bonding pairs do.

(a) Ammonia, NH$_3$

(b) Water, H$_2$O
VSEPR Theory (continued)

- 2 electron pairs around a central atom will be 180° apart, and the molecule’s shape will be **linear**.
- 3 bonding pairs around a central atom will be 120° apart, and the molecule’s shape will be **trigonal planar**. If one of the pairs is a lone pair, the shape will be **bent**.
VSEPR Theory (continued)

• 4 bonding pairs around a central atom will be 109.5° apart, and the molecule’s shape will be tetrahedral. If one of the pairs is a lone pair, the shape will be trigonal pyramidal. If two of the pairs are lone pairs, the shape will be bent.

• Unshared pairs repel electrons more strongly and will result in smaller bond angles.
Use VSEPR theory to predict the molecular geometry of water, H₂O.

Solution:
Draw the Lewis Structure for H₂O:
How many total electron pairs are surrounding the central atom?
4
How many are unshared pairs?
2
The shape is bent.
VSEPR Theory
Sample Problem B

Use VSEPR theory to predict the molecular geometry of carbon dioxide, CO₂.

Solution:
Draw the Lewis Structure for CO₂:
How many total electron pairs are surrounding the central atom?

2 (double or triple bonds count the same as single)

The shape is linear.
Molecular Polarity depends on both bond polarity and molecular geometry.

- If all bonds are non-polar, the molecule is always non-polar.
- If bonds are polar, but there is symmetry in the molecule so that the polarity of the bonds cancels out, then the molecule is non-polar. (Ex: CO$_2$, CCl$_4$)
- If bonds are polar but there is no symmetry such that they cancel each other out, the overall molecule is polar. (Ex: H$_2$O, CH$_3$Cl)