5. Chemical Bonding: The Covalent Bond Model

5.1 The Covalent Bond Model

Almost all chemical substances are found as aggregates of atoms in the form of molecules and ions produced through the reactions of various atoms of elements except the noble-gas elements which are stable mono-atomic gases.

Chemical bond is a term that describes the attractive force that is holding the atoms of the same or different kind of atoms in forming a molecule or ionic solid that has more stability than the individual atoms. Depending on the kinds of atoms participating in the interaction there seem to be three types of bonding:

**Gaining or Losing Electrons:**

**Ionic bonding:** Formed between many ions formed by metal and nonmetallic elements.

**Sharing Electrons:**

**Covalent bonding:** sharing of electrons between two atoms of non-metals.

Metallic Bonding: sharing of electrons between many atoms of metals.

<table>
<thead>
<tr>
<th>Ionic Compounds</th>
<th>Covalent Compounds</th>
<th>Metallic Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Metal and non-meal element combinations.</td>
<td>Non-metal and non-meal elements combinations.</td>
<td>Metal of one type or, combinations of two or metal elements combinations.</td>
</tr>
<tr>
<td>2. High melting brittle crystalline solids.</td>
<td>Gases, liquids, or waxy, low melting soft solids.</td>
<td>Conducting, high melting, malleable, ductile crystalline solids.</td>
</tr>
<tr>
<td>3. Do not conduct as a solid but conducts electricity when molten.</td>
<td>Do not conduct electricity at any state.</td>
<td>Conduct electricity at solid and molten states.</td>
</tr>
<tr>
<td>4. Dissolved in water produce conducting solutions (electrolytes) and few are soluble in non-polar solvents.</td>
<td>Most are soluble in non-polar solvents and few in water. These solutions are non-conducting (non-electrolytes).</td>
<td>Insoluble in any type of solvents.</td>
</tr>
</tbody>
</table>

The differences in these three bonding types are mainly due to the number of valence electron of the interacting atoms compared to noble gas elements. Noble gases show no inherent tendency to form any type of bonding apparently due to their closed valence shell electron configurations. Non-metals need only few electrons to achieve a closed shell. We focus mainly on covalent bonding in this chapter.

**Covalent bond:** bond in which one or more pairs of valence electrons are shared by two atoms in achieving a noble gas in contrast to complete transfer of electrons in forming ionic bonds.

Covalent bonding can be visualized with the aid of the octet rule and the Lewis structure we learned previously in chapter 4.
A molecular or covalent compound is made up of two or more atoms which are held together by a covalent bond. Even though covalent bonds are mainly formed between two atoms of non-metals, there could be many metal and nonmetal combination that are held by covalent bonds.

5.2 Lewis Structures for Molecular Compounds

Two atoms in molecule or ion in which each atom shares one valence electron with the other atom to form a single bond that keeps the two atoms together in forming a covalent bond.

**Lewis Theory of Covalent Bonding:**

The idea that the stability of noble gas electron configurations and the realization of the connection of reactivity of elements to achieve octet of valance electrons through formation of electron-pair bonds with other atoms of main group elements except duet for hydrogen.

**E.g.** In fluorine molecule, F₂ each atom F atom has 7 valence electrons and shares one valence electrons to form a covalent bonding pair. Bonding pair is represented by a line.

\[
\begin{align*}
\text{F} & \quad \text{F} \\
\text{or} & \quad \vdash & \vdash
\end{align*}
\]

**Lewis Electron-Dot (Line) Formulas:** A simple way of writing out a formula that shows the disposition of the shared valance electron pairs between the different atoms in a molecule.

In chlorine molecule, Cl₂ each atom Cl atom has 7 valence electrons and shares one valence electrons to form a covalent bonding pair. Bonding pair is represented by a line.

\[
\begin{align*}
\text{Cl} & \quad \text{Cl} \\
\text{or} & \quad \vdash & \vdash
\end{align*}
\]

**Bond pairs:** An electron pair (●)shared by two atoms in a bond.

**Lone pair:** An electron pair (●) found solely on a single atom.

5.3 Single, Double, and Triple Covalent Bonds

**Single covalent bond:** An electron pair (●) shared by two atoms in a bond. Always there has to be a single covalent before forming multiple covalent bonds.
Multiple Covalent Bonds in Lewis Structures: For every pair of electrons shared between two atoms, a single covalent bond is formed. Some atoms can share multiple pairs of electrons, forming multiple covalent bonds.

Double covalent bond: Where two electron pairs are shared between two atoms. For example, oxygen, $O_2$ (which has six valence electrons) needs two electrons to complete its valence shell. In $O_2$, two pairs of electrons are shared, forming two covalent bonds called a double covalent bond.

\[ \text{Six valence electrons from each } O \]
\[ \text{Total of 12 electrons or 6 electron pairs.} \]

Triple covalent bond: Found in unsaturated hydrocarbons with formula

For example, nitrogen, $N_2$ (which has 5 valence electrons) needs three electrons to complete its valence shell. In $N_2$, three pairs of electrons are shared, forming three covalent bonds called a Triple double covalent bond.

\[ \text{Five valence electrons from each } N \]
\[ \text{Total of 10 electrons or 5 electron pairs.} \]

Bond Length and Bond Energy

Bond Energy order: single=1 < double=2 < triple=3
Bond length: single (1 pair) > double (2 pairs) > triple (3 pairs)

Bond lengths from periodic trends in atomic radii

Bond length is proportional to the sum of atomic radii forming the bond.

Atomic radii trend: Li > Be > B > C > N > O > F decrease across a period

Bond Length Order: Li-H > Be-H > B-H > C-H > N-H > O-H > F-H

5.4 Valence Electrons and Number of Covalent Bonds Formed

Predicting number of covalent bonds formed by representative elements:

Lewis dot symbols are useful in showing the ways in which non-noble gas electron configurations could be achieved by sharing electrons in forming covalent bonds. Number of covalent bonds formed is equal to the number of single electrons in the Lewis symbol of the representative element.

Group I A: alkali metals (Li, Na, K, Rb, Cs)

Common Lewis symbol of Group I A $X^-$ could only form a single covalent bond as found in NaH.

Group II A: alkali earth metals (Be, Mg, Ca, Ba)
Common Lewis symbol of Group II A \( X \): alkali earth metals (Be, Mg, Ca, Ba) could form two covalent bond as found in BeCl\(_2\).

**Group III A**: (B, Al, Ga, In)

Common Lewis symbol of Group III A \( X \): (B, Al, Ga, In) could form two covalent bond as found in NH\(_3\).

Hydrogen, H\(_2\) gas forms the simplest covalent bond in the **homo-nuclear diatomic molecule**, H\(_2\). The halogens such as chlorine also exist as **homo-nuclear diatomic gases** by forming covalent bonds. The nitrogen and oxygen which makes up the bulk of the atmosphere also exhibits covalent bonding in forming diatomic molecules. Hydrogen and chlorine forms the simplest covalent bond in the **hetero-nuclear diatomic molecule**, HCl.

\[
\text{H}^+ + \text{H}^- \rightarrow \text{H}_2
\]

\[
\text{Cl}^+ + \text{Cl}^- \rightarrow \text{Cl}_2
\]

<table>
<thead>
<tr>
<th>Periodic Table Group Number</th>
<th>IA</th>
<th>IIA</th>
<th>IIIA</th>
<th>IVA</th>
<th>VA</th>
<th>VIA</th>
<th>VIIA</th>
<th>VIIIA</th>
</tr>
</thead>
<tbody>
<tr>
<td>Common Lewis symbols</td>
<td>X</td>
<td>X</td>
<td>X</td>
<td>X</td>
<td>X</td>
<td>X</td>
<td>X</td>
<td>X</td>
</tr>
<tr>
<td>Lewis symbols</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
<td>Ne</td>
</tr>
<tr>
<td>2nd period elements</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
<td>Ne</td>
</tr>
<tr>
<td>Number of covalent bonds formed</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>3</td>
<td>2</td>
<td>1</td>
<td>0</td>
</tr>
</tbody>
</table>

### 5.5 Coordinate Covalent Bonds

**Coordinate Covalent bonds**: Two atoms in molecule in which one atom shares two valence electrons with the other atom to form a bond that keeps the two atoms together.

**E.g.** In ammonia ion fluorine molecule, NH\(_4^+\), N atom has a lone pair of electrons and shares with the H\(^+\) ion to form a single bond that keeps the two atoms together. Coordinate covalent bonding pair is represented by an arrow pointing to atom receiving the electron. Chemistry dealing with coordinative covalent bonds is called **Coordination Chemistry**.
Covalent Bonds in Lewis Structure of Molecules and Polyatomic Ions: Lewis electron dot structures are representations of the distribution of electrons in molecules and polyatomic ions. The procedure is described below.

5.6 Systematic Procedures for Drawing Lewis Structures

A Lewis structure can be drawn for any molecule or ion could be obtained by following six steps:

Steps for Writing Electron-Dot Lewis structures for Molecules and Polyatomic Ions

1. **Calculate the number of valence electron pairs** (including charges, if any). The total number of valence electrons for the molecule or ion is calculated by adding up the valence electrons for each of the atoms in the species (molecule or ion). The number of valence electrons contributed by an atom equals the group number of the element. If the species has a charge, you need to add (anions) or subtract (cations) electrons from the neutral atoms to give that charge. Add one electron for each negative charge and subtract one electron for each positive charge.

2. **Pick the central atom**: Lowest EN (electronegativity) atom, largest atom, and/or atom forming most bonds is usually the central atom.

3. **Connect central atom to all terminal atoms**. Connect atoms by single bonds, using either a pair of dots (●) or a dash (—). You may decide later that two or more of these bonds will add up to a double or triple bond.

4. **Fill octet to terminal atoms**. Distribute electron dots (●) or a dash (—) to the atoms surrounding the central atom to satisfy the octet rule in them. Calculate the number of remaining valence electrons pairs so far used.

5. **Fill octet to central atom**: Distribute the remaining valence electrons to the central atom. If there arc fewer than eight electrons around the central atom, this suggests that you need to make double triple bonds taking lone-electron pairs (●) from terminal atoms. When you have a choice of atoms from which to obtain such a lone-pair, note that C, N, O, and S atoms tend to form multiple bonds, so you take the lone pair from such an atom.

6. **Check that total number of valance electron pairs** in the structure match with the dots (●) or a dash (—) or lines.
### Draw Lewis Structure of NH₃

1) **Count Valence electrons:**
   
   \[ 5 + 3 \times 1 = 8 = 8/2 \]
   
   **4 electron pairs**

2) **N is the central atom:** Forms most covalent bonds
   
   ![Lewis Structure of NH₃](image)

3) **Connect central atom to terminal**
   
   ![Lewis Structure](image)

4) **Fill octet to terminal atoms.**
   
   Exception duet to H atoms.

5) **Fill octet to terminal atoms.**

6) **Check that total number in 5.**

   - 3 bond pairs = 3 pairs
   - 1 lone pairs = 1 pair
   - **4 electron pairs**

### Problem: Draw the Lewis structures for the following molecules: H₂O

1) **H₂O** sum of valence electrons: 2 + 6 = 8; 4 pairs

2) **Pick the central atom:** usually the biggest atom and the atom forming most bonds. **O is the central atom**

3) **Connect central atom to other terminal atoms** with a line representing a covalent bond (two electrons).

4) **Fill octet to central atom:**
   
   A dot represents one electron and a line represents two electrons.

   **Check to see whether duet on H**
   
   ![Lewis Structure](image)

   - Count electrons and check valence electrons on C.
     - 2 bond pairs = 2 x 2 = 4
     - 2 lone pairs = 2 x 2 = 4
     - Total = 8 = 4 electron pairs (an Octet)

   **Bond pairs:** an electron pair shared by two atoms in a bond. E.g. two bond pairs between O-H in water.

   **Lone pair:** an electron pair found solely on a single atom.
   
   ![Lewis Structure](image)

   - E.g. two lone pairs found on the O atom at the top and the bottom.

**H₂S has the same Lewis Structure as H₂O**

Since S is in the same **group VI A** as O. As a rule if an element from the period 2 is replaced by another element from the period 3 and same group in the periodic table they will have the same Lewis structure.

**Draw Lewis structure of Nitrogen (N₂)**

sum of valence electrons: 5 + 5 = 10; 5 electron pairs

**Draw Lewis structure of CO₂**
<table>
<thead>
<tr>
<th></th>
<th>Valence electrons: $4 + 2 \times 6 = 16$ (8 pairs)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2)</td>
<td>Central atom C;</td>
</tr>
<tr>
<td>3)</td>
<td>Give octet to oxygen $\overline{O}C\overline{O}$</td>
</tr>
<tr>
<td></td>
<td>Try to fill octet to C make double bonds</td>
</tr>
<tr>
<td></td>
<td>4) Count electrons:</td>
</tr>
<tr>
<td></td>
<td>4 bond pairs = 4 pairs</td>
</tr>
<tr>
<td></td>
<td>4 lone pairs = 4 pairs</td>
</tr>
<tr>
<td></td>
<td>8 electron pairs</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th></th>
<th>Draw Lewis structure of CH$_4$</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Sum of valence electrons:</td>
<td></td>
</tr>
<tr>
<td></td>
<td>$4$ (from one C) + $4$ (from four H) = 8 electrons = 4 electron pairs</td>
</tr>
<tr>
<td>2)</td>
<td>Central atom is C</td>
</tr>
<tr>
<td>3)</td>
<td>Octet on C atom and duet on H are already complete.</td>
</tr>
<tr>
<td>4)</td>
<td>Count valence electrons on the Lewis structure:</td>
</tr>
<tr>
<td></td>
<td>bond pairs = $2 \times 4 = 8$ = 4 electron pairs</td>
</tr>
<tr>
<td></td>
<td>lone pairs = 0</td>
</tr>
</tbody>
</table>
|   | $\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}$ or $\begin{array}{c}
\text{H} \\
\text{C} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}$ |

<table>
<thead>
<tr>
<th></th>
<th>Draw Lewis structure of Ethane( CH$_3$CH$_3$ )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Draw two carbons and attach to them six hydrogens.</td>
<td></td>
</tr>
<tr>
<td>Count valence electrons:</td>
<td></td>
</tr>
<tr>
<td></td>
<td>$2C = 8$</td>
</tr>
<tr>
<td></td>
<td>$6H = 6$</td>
</tr>
<tr>
<td></td>
<td>14 (7 electron pairs)</td>
</tr>
<tr>
<td>Octet on each carbon and duet on each hydrogen</td>
<td></td>
</tr>
</tbody>
</table>
|   | $\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}$ or $\begin{array}{c}
\text{H} \\
\text{C} \\
\text{C} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}$ |

<table>
<thead>
<tr>
<th></th>
<th>Draw Lewis structure of nitric acid( HNO$_3$ )</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>= 1 valence electron</td>
</tr>
<tr>
<td>N</td>
<td>= 5 valence electrons</td>
</tr>
<tr>
<td>O x 3 = 6 x 3 = 18 valence electrons</td>
<td></td>
</tr>
<tr>
<td>Total of 24 valence electrons</td>
<td></td>
</tr>
</tbody>
</table>
### 5.7 Bonding in Compounds with Polyatomic Ions Present

#### Draw Lewis structure of \( \text{(CO}_3^{2-}\text{)} \) 2 electrons are added to the total valence electrons

1) Valence electrons: \(4 + (3 \times 6) + 2 = 24\) (12 electron pairs)

2) Central atom C;

3) Give octet to 3 oxygens. Try to fill octet to C make double bonds

![Lewis structure of \( \text{(CO}_3^{2-}\text{)} \)](image)

4) Count electrons:
   - 4 bond pairs = 4 pairs
   - 8 lone pairs = 8 pairs
   - 12 electron pairs

#### Draw Lewis structure of \( \text{(SO}_4^{2-}\text{)} \) 2 electrons are added to the total valence electrons

2) Valence electrons: \(6 + (4 \times 6) + 2 = 32\) (16 electron pairs)

2) Central atom S;

3) Give octet to 4 oxygens. Try to fill octet to C make double bonds

![Lewis structure of \( \text{(SO}_4^{2-}\text{)} \)](image)

4) Count electrons:
   - 4 bond pairs = 4 pairs
   - 12 lone pairs = 12 pairs
   - 16 electron pairs

### 5.8 Molecular Geometry

**Lewis Structures and Molecular Geometry**

Lewis structures are only helpful explaining the formation of covalent bonds and chemical stability of molecules. It does not provide information about the molecular shape or geometry as it is written and only account for octet rule. The valence shell electron pair repulsion (VSEPR) theory is an extension to Lewis theory of covalent bonding and helps to explain (or predict) molecular geometry which including **linear**, **trigonal planar**, and **tetrahedral** arrangements of terminal atoms to the central atom of a Lewis structure.
Valence-Shell Electron-Pair Repulsion Theory (VSEPR Theory)

In the theory, valence shell electron pairs are assumed to repel each other, assuming orientations to minimize repulsions and establish certain groups of molecular shapes.

Basic Geometries created by molecules with 2, 3, and 4 electron around the central atom pairs.

<table>
<thead>
<tr>
<th>four things</th>
<th>three things</th>
<th>two things</th>
</tr>
</thead>
<tbody>
<tr>
<td>Linear</td>
<td>Trigonal planar</td>
<td>Tetrahedral</td>
</tr>
</tbody>
</table>

Molecules with no lone-pair electrons on central atom

<table>
<thead>
<tr>
<th>Example</th>
<th>BeF₂</th>
<th>BF₃</th>
<th>CH₄</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lewis Structure</td>
<td>![Lewis Structure for BeF₂]</td>
<td>![Lewis Structure for BF₃]</td>
<td>![Lewis Structure for CH₄]</td>
</tr>
<tr>
<td>Electron pairs around central atom</td>
<td>two</td>
<td>three</td>
<td>four</td>
</tr>
<tr>
<td>Type of Geometry only looking at bonding electron pairs</td>
<td>Linear</td>
<td>Trigonal planar</td>
<td>Tetrahedral</td>
</tr>
</tbody>
</table>

Molecules with lone-pair electrons on central atom

The molecular configuration is not exactly the same as the electron pair geometry of the central atom. The reason for this is that lone pairs of electrons not seen when you look at the bond connectivity which is the molecular geometry, although their effects are still seen (such as pushing bonding pair electrons closer together). This gives rise to more geometries (in addition to the 3 basic geometries: linear, trigonal planar).

The most common geometry with lone pairs is angular structure has four electron pairs surrounding the central atom. The water molecule (H₂O:H) has two lone pairs and two bond pairs (4 total pairs on O). Another common geometry is triangular pyramid found in
NH₃ with only one lone pair of electrons.

**Geometries with lone pair of electrons**

<table>
<thead>
<tr>
<th>Example</th>
<th>H₂O</th>
<th>NH₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lewis Structure</td>
<td><img src="image" alt="Lewis Structure H₂O" /></td>
<td><img src="image" alt="Lewis Structure NH₃" /></td>
</tr>
<tr>
<td>Electron pairs around central atom</td>
<td>four</td>
<td>four</td>
</tr>
<tr>
<td>Lone pair electron pairs around central atom</td>
<td>two</td>
<td>one</td>
</tr>
<tr>
<td>Type of Geometry only looking at bonding electron pairs</td>
<td>Angular or Bent</td>
<td>Triangular pyramid</td>
</tr>
</tbody>
</table>

A = central atom; X= terminal atoms  E= lone pair of electrons

**Chemistry at a Glance: The Geometry of Molecules**

**Summary of geometries without/with lone pairs of electrons**

A = central atom
X = attached (bonded) atom (or group of atoms) or lone pair

5.9 Electronegativity

Electronegativity is useful in describing the electron distribution in a covalent bond. It is a measure of the tendency of an atom to attract a bonding pair of electrons. The electronegativity scale created by **Linus Pauling** is commonly used in predicting bond polarity. Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to cesium and francium which are the least electronegative at 0.7.
Periodic trend in electronegativity
As you go down a group, electronegativity decreases and crossing a period it increases.

<table>
<thead>
<tr>
<th></th>
<th>H</th>
<th>Li</th>
<th>Be</th>
<th>B</th>
<th>C</th>
<th>N</th>
<th>O</th>
<th>F</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.1</td>
<td></td>
<td>1.0</td>
<td>1.5</td>
<td>2.0</td>
<td>2.5</td>
<td>3.0</td>
<td>3.5</td>
<td>4.0</td>
</tr>
</tbody>
</table>

5.10 Bond Polarity
The electron pairs shared between two atoms in a covalent bond are not necessarily shared equally except in homonuclear diatomic molecules. The degree of electron pair sharing in a covalent bond leads to three types of covalent bonds.

a) **Non-polar covalent bond**

b) **Polar covalent bond**

c) **Ionic bond** (considered as an extreme case of covalent bonding)

**Pure covalent bond**
Equally shared electron pair in a covalent bond is called a pure covalent bond as found in homonuclear diatomic molecule, Cl₂ the shared electron pairs is shared equally.

**Polar covalent bond** is one in which one atom has a greater attraction for the electrons than the other atom as found in heteronuclear diatomic molecule, HCl the shared electron pairs is shared is attracted more towards the Cl because of its greater electronegativity.

**Bond Properties: Bond Polarity and Electronegativity**

**Bond polarity** is a useful concept for describing the sharing of electrons between atoms

a) **Nonpolar covalent bond** is one in which the electrons are shared equally between two atoms
b) **Polar covalent bond** is one in which one atom has a greater attraction for the electrons than the other atom.

c) **Ionic bond**: If this relative attraction is great enough, then the bond is an ionic bond

<table>
<thead>
<tr>
<th>Electronegativity Difference</th>
<th>Non-polar covalent</th>
<th>Polar covalent bond</th>
<th>Ionic</th>
</tr>
</thead>
<tbody>
<tr>
<td>0-0.5</td>
<td>C-H</td>
<td>O-H</td>
<td>Na-Cl</td>
</tr>
<tr>
<td>0.6-1.6</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1.7, or greater</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Polar covalent bonds**

In the O-H bond O has an electronegativity of 3.5, H has an electronegativity of 2.1. The difference in electronegativity is $3.5 - 2.1 = 1.4$

1.4 is less than 1.7, so the resulting bond is **polar covalent**.

Oxygen is the more electronegative so it will have a greater share of the bonding electrons and therefore a partial negative charge, O$^{\delta-}$.

Hydrogen is less electronegative so it will have a lesser share of the bonding electrons and therefore a partial positive charge, H$^{\delta+}$.

Since the bond has two 'poles' or 'ends' it is sometimes referred to as a dipole.

The polar covalent bond can be represented as:

\[
\text{O}^{\delta-} - \text{H}^{\delta+} \quad \text{or} \quad \text{O}^{\delta-} - \text{H}^{\delta+} <------+
\]

Where the arrow head points towards the most electronegative atom.

**Ionic bond** (considered as an extreme case of covalent bonding) In **heteronuclear compounds** such as NaCl the 3s electron is stripped from the Na atom and is incorporated into the electronic structure of the Cl atom - and the compound is most accurately described as consisting of individual Na$^{+}$ and Cl$^{-}$ ions

**For most covalent substances, their bond character falls between these two extremes**

5.11 Molecular Polarity

**Lewis Structures and Polarity**

A molecule containing only non-polar bonds is always non-polar molecule. Having a polar bonds do not always lead to molecular polarity. In a molecule with polar covalent bonds there is a possibility that they might add up to give a overall stronger molecular
polarity or they might cancel each other to produce a non-polar molecule. Concepts of molecular polarity are better understood when viewed from the perspective of a Lewis structure. A polar molecule has at least one polar covalent bond. The electronegativity difference helps to assess the degree of polarity of a bond. A molecule containing polar bonds may be polar or non-polar, depending on the relative position of the bonds.

**Molecular Polarity**

Molecules composed of covalently bonded atoms may also be polar or non-polar. For a molecule to have polarity, it must, of course, have polar bonds. But the key factor for determining the polarity of a molecule is its shape. If the polar bonds (dipoles) are **symmetrical** around the central atom, they offset each other and the resulting molecule is non-polar. However, if the dipoles are **not symmetrical** (asymmetric) around the central atom, the electrons will be pulled to one end of the molecule. The resulting molecule is polar. Ball and stick models are often used to demonstrate molecular shape. In this exercise you will build several covalent molecules and predict each molecule's polarity on the basis of its molecular shape.

The geometry of the molecule affects its **dipole moment** or **polarity** - the asymmetric distribution of positive and negative charges.

Polar Molecules will orient in a magnetic field When two elements with different electronegativities are bonded together, one obtains a polar bond.

**Example**: HF

Depending on the molecular shape, polar bonds can give rise to a polar molecule, one with an overall uneven distribution of electron charge. We can see this charge imbalance when the molecules are placed in an electric field. Polarity is a key feature of a molecule because it can influence physical, chemical, and even biological properties.

For a **hetero-diatomic molecule**, a polar bond **must** lead to a polar molecule.

Consider **hydrogen fluoride**, $\text{H}^{-} \text{F}^{+}$. Note the polar arrows with the negative end pointing to F atom. You can see that the F end of the molecule is much more negative than the H end, and thus HF is highly polar.

If a molecule has more than two atoms, its shape can affect the polarity in a crucial way. For example, in carbon dioxide, $\text{C}=\text{O}$ since oxygen is more electronegative than carbon, each bond is highly polar. But the linear molecular shape (**symmetrical**) makes the bond polarities cancel each other, so the CO$_2$ molecule is **non-polar**. Notice that the orientation of the molecules is random, whether the field is off or on.
In water, \( \text{H}_2\text{O} \), the V-shape of the molecule allows the bond polarities to reinforce each other, so water is highly polar, as the large polar arrow shows.

![Water molecule](image)

The case of boron trifluoride is non-polar similar to that of \( \text{CO}_2 \). Because the three highly polar bonds point to the corners of an equilateral triangle, the bond polarities cancel each other, and \( \text{BF}_3 \) is nonpolar.

Like \( \text{BF}_3 \), ammonia has four atoms and three polar bonds, but the trigonal pyramidal shape means that the bond polarities reinforce each other. Thus, ammonia is highly polar for the same reason that water is. Substituting a hydrogen for one of the chlorines gives chloroform, \( \text{CHCl}_3 \) another tetrahedral molecule, but now the polar bonds reinforce each other. As a result, chloroform is highly polar.

**Using VSEPR molecular structures predict the polarity (net dipole moment or zero dipole moment) of the following molecules: \( \text{H}_2\text{O} \), \( \text{NH}_3 \), \( \text{CO}_2 \), and \( \text{SO}_3 \)**

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Shape</th>
<th>Electron pair distribution</th>
<th>Polarity</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{H}_2\text{O} )</td>
<td>bent</td>
<td>asymmetric</td>
<td>polar</td>
</tr>
<tr>
<td>( \text{NH}_3 )</td>
<td>Trigonal pyramid</td>
<td>asymmetric</td>
<td>polar</td>
</tr>
<tr>
<td>( \text{CO}_2 )</td>
<td>liner</td>
<td>symmetric</td>
<td>Non-polar</td>
</tr>
<tr>
<td>( \text{SO}_3 )</td>
<td>Trigonal planer</td>
<td>symmetric</td>
<td>Non-polar</td>
</tr>
</tbody>
</table>

**Chemistry at a Glance: Covalent Bonds and Molecular Compounds**

### 5.12 Naming Binary Molecular Compounds

Most covalent compounds are formed by the reactions of non-metals with another non-metal. Covalent compounds exist as molecules and are named using prefixes that denote the number of each element present in the compound. A covalent compound is given systematic name similar to an ionic compound following certain set of certain rules. Before the rules are made common names was given without following systematic rules: E.g. water for \( \text{H}_2\text{O} \). The "shorthand" symbol for a covalent compound is its formula. Formula gives types atoms and numbers each atom in the chemical compound. Many familiar covalent compounds have common names. It is useful to correlate both systematic and common names with the corresponding molecular formula.
A **binary covalent compound** (a molecule) is consisting of only **atoms** of two elements in which one or more electron pairs are shared to form covalent bonds. When naming these compounds, its composition must be considered. In naming covalent compounds, some of the features used in naming ionic compounds are flowed but no the part that deal with ionic charge. The most electronegative element is normally named similar to the anion part of the ionic compound by adding a prefix –ide to end of the element name. E.g CO₂, is named carbon dioxide.

Binary molecular compounds are composed of only two elements. Examples are H₂O, NO, SF₆ etc. Sometimes these compounds have generic or common names (e.g., H₂O is "water") and they also have systematic names (e.g., H₂O, dihydrogen monoxide). The common name must be memorized. The systematic name is more complicated but it has the advantage that the formula of the compound can be deduced from the name. Since there are no charges prefixes has to be used in front of the element name to indicate number of atoms of each element in the compound. If there are four chlorine it would be tetra-chloride.

### Prefixes used:

| 1 | mono | 6 | hexa |
| 2 | di   | 7 | hepta |
| 3 | tri  | 8 | octa |
| 4 | tetra| 9 | nona |
| 5 | penta| 10| deca |

<table>
<thead>
<tr>
<th>Compound</th>
<th>Systematic name</th>
<th>Common name (if it has one)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NF₃</td>
<td>nitrogen trifluoride</td>
<td></td>
</tr>
<tr>
<td>NO</td>
<td>nitrogen monoxide</td>
<td>nitric oxide</td>
</tr>
<tr>
<td></td>
<td><strong>note: for first element we don’t use mono- prefix</strong></td>
<td></td>
</tr>
<tr>
<td>NO₂</td>
<td>nitrogen dioxide</td>
<td></td>
</tr>
<tr>
<td>N₂O</td>
<td>dinitrogen monoxide</td>
<td>laughing gas nitrous oxide</td>
</tr>
<tr>
<td>N₂O₄</td>
<td>dinitrogen tetraoxide</td>
<td></td>
</tr>
<tr>
<td>PCl₅</td>
<td>phosphorous pentachloride</td>
<td></td>
</tr>
<tr>
<td>SF₆</td>
<td>sulfur hexafluoride</td>
<td></td>
</tr>
</tbody>
</table>
Problem: Give the names of following formulas from the names of following covalent compounds names:

<table>
<thead>
<tr>
<th>Compound</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>H2S</td>
<td>hydrogen sulfide</td>
</tr>
<tr>
<td>CS2</td>
<td>carbon disulfide</td>
</tr>
<tr>
<td>PCl5</td>
<td>phosphorus pentachloride</td>
</tr>
<tr>
<td>P2O5</td>
<td>diphosphorus pentoxide</td>
</tr>
</tbody>
</table>

Answer:

- a. hydrogen sulfide
- b. carbon disulfide
- c. phosphorus pentachloride
- d. diphosphorus pentoxide

Names of Acids and Bases

Binary acids: made up of only two elements - hydrogen and one other element.

Naming binary acids:
Begin with the prefix hydro.
Determine the "stem" - part of the name of the element that combines with hydrogen.
Add the suffix ic.

Examples:
- HF - hydro fluor ic - hydrofluoric acid
- HCl - hydro chlor ic - hydrochloric acid
- HBr - hydro brom ic - hydrobromic acid
- HI - hydro iod ic - hydroiodic acid

Ternary acids: made up of three elements - hydrogen, oxygen, and another element.

Naming ternary acids:
Acids made up of three elements including hydrogen
Determine the "stem" - part of the name of the third element.
The most common acid is given the suffix ic.
Add the prefix per for the acid with one more oxygen.
The suffix ous is given to the acid with one less oxygen.
Add the prefix hypo for the acid with two less oxygen atoms.

Examples:
- HClO4 - per chlor ic - perchloric acid - one more oxygen atom.
- HClO3 - chlor ic - chloric acid - the most common form of the acid.
- HClO2 - chlor ous - chlorous acid - one less oxygen atom.
- HClO - hypo chlor ous - hypochlorous acid - two less oxygen atoms.
- HNO3 - nitric acid
- HNO2 - nitrous acid
\(H_2SO_4\) - sulfuric acid
\(H_2SO_3\) -sulfurous acid
\(H_3PO_4\) -phosphoric acid
\(H_3PO_3\) -phosphorous acid
\(H_3BO_3\) -boric acid

**Examples:**
\(P_2O_5\) - this is named diphosphorus pentoxide, because there are two phosphorus atoms and five oxygen atoms.
\(CO\) - this is carbon monoxide (you need the "mono-" because there's only one oxygen atom).
\(CF_4\) - this is carbon tetrafluoride, because there's one carbon atom and four fluorine atoms.

Chemical Connections: Nitric Oxide: A Molecule Whose Bonding Does Not Follow "The Rules"; Molecular Geometry and Odor