Atoms bond to other atoms
- Atoms bonded together make up molecules (intramolecular)

Molecules can also “bond” to other molecules
- Intermolecular Forces attract to each other

Polar and Non-polar Covalent Bonds Review
1. Non-polar Covalent Bonds
   - Electrons are shared equally
   - Electronegativity differences are less than 0.5
2. Polar Covalent Bonds
   - Electrons are shared unequally
   - Electronegativity differences are ≥ 0.5
   - Polar – Two ends that are oppositely charged of one another.

Polar Molecules are called Dipolar
- Dipole means a pair of charges (+ and -)
- Di means two...so 2 opposite poles
- The larger the EN difference, the stronger the polarity
- Partial charges are represented by a lower-case delta symbol
  - \( \delta^+ \)
  - \( \delta^- \)

Dipolar Molecules
- Example = Water (Oxygen EN = 3.4, Hydrogen EN = 2.2)
- The atom with the greatest electronegativity has the negative dipole(s) (\( \delta^- \))
- The atom with the smallest electronegativity has the positive dipole(s) (\( \delta^+ \))
**Why does it matter if a molecule is polar?**

- Chemical and physical properties of a molecule depend on its polarity.
- Why are $\text{N}_2$ & $\text{O}_2$ gases?
- Chemical and physical properties are dependent on:
  - The polarity of each bond AND
  - The orientation of the bond.
  - $\text{CO}_2$ vs. $\text{H}_2\text{O}$

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**Intermolecular Forces**

- Attractive forces between molecules
- Weaker than ionic or covalent bonds
- 1. Dipole - Dipole ($\delta^+ - \delta^-$)
- 2. London / Dispersion Forces

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**Dipole-Dipole**

- Attraction of polar molecules of another molecule
- Stronger polarity = Stronger Dipole Force
  - **Dipole Force** (Allows for solids & liquids)

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**Why do polar molecules have higher boiling and melting points?**

- Breaks
- Temperature is a measure of kinetic energy
  - Energy in motion
  - Relatively Strong Bond $\rightarrow$ Higher KE $\rightarrow$ $\uparrow$ Temp
Hydrogen Bonds

• Specific kind of Dipole-Dipole
• Have strong dipole-dipole force
• Two reasons
  1. Large EN difference
  2. Partially exposed proton
• Not a bond between atoms
  • Hydrogen end of one molecule "bonds" to the high EN atom of another molecule.
  • High EN Atoms = O, N, F, Cl
• Common Examples
  • H₂O & DNA

Hydrogen bonds explain the unique properties of water:
• Surface Tension
• Adhesion
• Cohesion
• Less density as a solid
• High Boiling / Melting Points

London / Dispersion Forces

• ALWAYS exist but are usually covered up by other forces
• Only force that exists for non-polar molecules
• Instantaneous attractions between molecules.
  • Temporary Dipoles
  • Electrons are moving within their orbitals
  • Random movements will cause instantaneous moments of polarity
• Weaker than Dipole-Dipole
• Examples = O₂, N₂, H₂

London dispersion forces - Very weak attraction

• Very low temperatures are required to allow the bonds to hold.
• Nitrogen liquefies @ -321°F
• Why so low?
  • Too much Kinetic Energy @ Room Temperature
• Weak Bond → Temperatures
Intermolecular Forces Review

- Dipole-Dipole forces
  - Only occurs between POLAR molecules
  - Hydrogen bonds
    - Between Hydrogen and atom on another molecule
- London dispersion forces
  - Always exist (so both polar and non-polar)
  - The only force between NON-POLAR molecules

Shape can influence polarity!

- Even if you conclude that a molecule should be polar based on the difference in Electronegativities, the molecular shape might negate these charges.
  - Therefore, a molecule that appears to be polar may actually non-polar
  - For example, CO$_2$
  - So, a non-polar molecule could actually have polar bonds


\[ \text{CO}_2 \]

Electronegativity Difference = 3.4 - 2.6 = 0.8
- Conclusion = Polar

BUT - CO$_2$ is NON-POLAR!
- Why?

So, what about CO$_2$?

If the polar forces are treated as vectors, what’s the consequence?

\[ \text{CO}_2 \]

- Nothing!
- They Cancel!
- Therefore, It’s non-polar
**As a result of vectors**

- **Bent and Trigonal Pyramidal** molecules are **ALWAYS polar**.
  - Even if the individual bonds are non-polar
  - Why? Symmetry isn’t possible with the unshared pairs

![](image1)

- **Linear, Trigonal Planar, and Tetrahedral** can be **polar or non-polar**
  - **To be non-polar**, they can
    - Either have small EN differences (less than 0.5), OR...
    - Polar bonds that CANCEL out. This can happen when...
      1. All of the elements on the outside of the molecule are the same.
      2. Equal & opposite attraction
  - **To be polar**
    1. They must have polar bonds that do NOT cancel each other out.

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**Quick rules to determine non-polar vs polar molecules**

- Determine shape
- Bent or trigonal pyramidal always polar
- If there is symmetry in the molecule so that the polarity of the molecule cancels out, then the molecule is non-polar
  - Examples - CO₂ and CCl₄
  - Everything around central atom is the same
- If there are polar bonds but there is no symmetry such that they do not cancel each other out, the overall charge is polar.
  - Examples – CH₃Cl
  - Everything around the outside is not the same

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**Practice Problems**

- **CCl₄**
  - Polar or non-polar bonds?
  - Polar or non-polar molecule?
  - Shape?
- **O₂**
  - Polar or non-polar?
  - Shape?
  - What type of intermolecular force?
- **HCl**
  - Polar or non-polar molecule?
  - What type of intermolecular force?
- **H₂O vs CO₂**
  - Polar or non-polar bonds?
  - Two differences
    - Intermolecular forces and shape
**Vectors**

Vector - A quantity that has both magnitude and direction
- An arrow

1. **If I traveled on Vector 1 (1049.5 miles @ 169.5°) where would I be?**

2. **What if I get bored and travel on Vector 2 (1049.5 miles @ 349.5°)? Where will I be?**

**Vectors can be broken down...**

\[ \text{Vector} = V_y + V_x + V_z (3D) \]

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**Vectors can be added**

If an individual traveled on the following two vectors, where would they end up?

(Plot one vector at the end of the other vector)