Polarity and Intermolecular Forces
Types of bonds

- **Ionic** – transfer of e- from one atom to another
- **Covalent** – sharing of e- between atoms
  - a) *nonpolar covalent* – equal sharing of e-
  - b) *polar covalent* – unequal sharing of e-
Polar bonds and Electronegativity

- Electronegativity is the ability of an atom to attract electrons in a chemical bond.
- Polar bonds result when a highly electronegative atom bonds to a less electronegative atom.
A covalent bond is polar if there is a significant difference between the electronegativities of the two atoms (see below):

<table>
<thead>
<tr>
<th>Electronegativity Difference</th>
<th>Type of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>0-0.3</td>
<td>Nonpolar covalent</td>
</tr>
<tr>
<td>0.4-1.9</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>2.0 or greater</td>
<td>Ionic</td>
</tr>
</tbody>
</table>
Fluorine has a stronger attraction for the electrons. They are still shared, but spend more time around the fluorine giving partial opposite charges to opposite ends of the bond (a dipole).
Nonpolar Bond (no dipole) vs. Polar Bond (dipole)

Electrons are evenly distributed

Electrons are polarized toward Cl.

$\chi = 2.1$  $\chi = 3.0$
Showing Polarity of a Bond

\[ \delta^+ \quad \delta^- \quad \text{or} \quad \text{H} - \text{F} \]
Give the electronegativity difference and determine the bond type in the following molecules

1) CH₄     1) polar
2) HCl     2) polar
3) NaF     3) ionic
4) MgCl₂   4) ionic
5) SO₂     5) polar
6) NH₃     6) polar
7) H₂O     7) polar
8) KCl     8) ionic
9) CsF     9) ionic
10) Cl₂    10) nonpolar
Determining Polarity of Molecules

- If one end of a molecule is slightly positive and another end is slightly negative, the molecule is polar.

- Polarity depends on the shape of the molecule.

- Ex. CO$_2$ (nonpolar) and H$_2$O (polar)
To determine polarity of a molecule you need the following:

- Lewis Structure

- ABE designation and molecular shape (using your chart)

- If surrounding atoms are identical in the following shapes, the molecule has no dipole (it’s nonpolar):

  linear, trigonal planar, tetrahedral, trigonal bipyramidal, octahedral, square planar
\( \mu = 0 \text{D} \)
No net dipole moment

\[ \text{CH}_4 \]

Net dipole \( \mu = 1.92 \text{D} \)

\[ \text{CH}_3\text{Cl} \]

Net dipole \( \mu = 1.60 \text{D} \)

\[ \text{CH}_2\text{Cl}_2 \]

Net dipole \( \mu = 1.04 \text{D} \)

\[ \text{CHCl}_3 \]

\( \mu = 0 \text{D} \)
No net dipole moment

\[ \text{CCl}_4 \]

---

\[ \text{Boron Trifluoride} \]

polar bonds
Determine the Polarity of the following molecules:

1) Water

2) Carbon tetrachloride

3) Carbon monoxide

4) Carbon dioxide

5) Ammonia (NH₃)

6) Methyl chloride (CH₃Cl)

7) Sulfur dioxide

8) Boron trichloride

9) ICl₄⁻
Intermolecular forces – the attractions between molecules

- Determine whether a compound is a solid, liquid or gas at a given temperature (determine melting and boiling points of substances)

- 3 Main Types:
  a) Hydrogen bonding
  b) Dipole-dipole interactions
  c) Dispersion forces
Hydrogen Bonding

- Attraction formed between the hydrogen atom of one molecule and an electronegative atom of an adjacent molecule (O, N, or F)
- A type of dipole interaction and the strongest intermolecular force
Dipole-dipole interactions

Dipoles interact by the positive end of one molecule being attracted to the negative end of another molecule (similar to but much weaker than ionic bonds).
Dispersion Forces

- Caused by electron motion. Electrons around one molecule momentarily repel electrons a nearby molecule creating a momentary charge difference.

- Can exist between nonpolar molecules as well as polar.

- Weakest intermolecular force but increases as the number of electrons increases.
Intermolecular forces and melting/boiling point

Stronger intermolecular forces
- ion-ion
- hydrogen bonding
- dipole-dipole
- dispersion

Weaker intermolecular forces

Higher melting and boiling points

Lower melting and boiling points