# Unit 1, Lesson 11: Polarity of Molecules

In Lesson 7 we reviewed the polarity of ***bonds*****.**

If the difference in electronegativity between the bonding atoms is:

| | |
|---|---|
| < 0.50 | pure covalent bond: the electrons are essentially equally shared between the bonding atoms |
| 0.50 to 1.70 | polar covalent bond: the electrons are shifted toward the more electronegative atom, creating small or partial charges, $\delta-$ and $\delta+$ |
| > 1.70 | ionic bond: the electrons are essentially removed from the metal and shifted to the non-metal, creating full ionic charges |

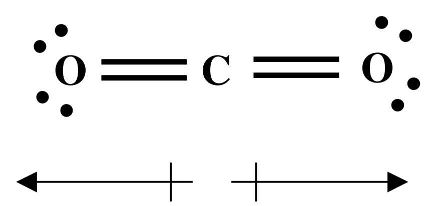
Polar bonds are said to have a "dipole moment" or a **<u>dipole</u>** because the poles (ends) of the bond have different partial charges. The electrons are shifted toward the element with higher electronegativity.

If a molecule is diatomic (contains only two atoms), such as HCℓ or NO, then the polarity of the bond is also the polarity of the molecule.

**A molecule can contain polar bonds, but the overall molecule may be non-polar. That is, the polarity of *bonds* can be quite different than the polarity of the *molecule* that contains those bonds.**

The polarity of ***molecules*** depends on both:
1. the polarity of the bonds in the molecule, and
2. the shape and symmetry of the molecule

eg. carbon dioxide ($CO_2$)



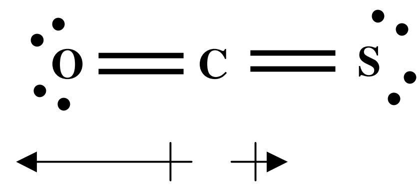

- the Lewis structure for $CO_2$ is:
- the VSEPR notation for $CO_2$ is $AX_2E_0$, it is a linear molecule
- the ΔEN for the C – O bond is 0.89, so the C – O ***bond*** is polar
- however, the $CO_2$ molecule is perfectly symmetrical
- the polar bonds are pulling equally on the bonded electrons, but in opposite directions as shown by the vector diagrams
- the dipoles cancel each other out
- carbon dioxide contains two polar ***bonds***, but the overall ***molecule*** is non-polar because the polarity of the bonds cancel out
- $CO_2$ is a non-polar molecule which explains why it is a gas at SATP

If a molecule is symmetrical in ALL planes, the dipole moments of any polar bonds will cancel out and the overall molecule will be non-polar.

If a molecule is asymmetrical, the dipole moments of any polar bonds will not cancel out, and the overall molecule will be polar.

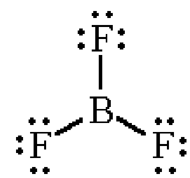
eg. CSO

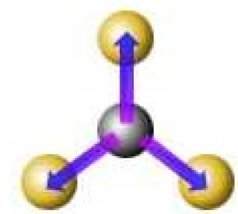
- the Lewis structure for CSO is shown to the right
- the VSEPR notation for CSO is $AX_2E_0$, it is a linear molecule
- the ΔEN for the C – O bond is 0.89, so the C – O ***bond*** is polar
- the ΔEN for the C – S bond is 0.03, so the C – S ***bond*** is non-polar
- the CSO molecule is not symmetrical
- the polar C – O bond pulls the bonding pairs toward the O, while the non-polar C – S bond can not offset this (as shown in the vector diagram)
- this gives the molecule a net dipole moment and it is a polar molecule
- because this is a polar molecule, we would predict that its melting and boiling points will be higher than for $CO_2$ (and they are)




eg. $BF_3$ (boron trifluoride)

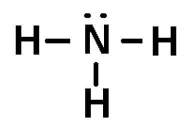
- Lewis structure is shown to the right
- the VSEPR notation for $BF_3$ is $AX_3E_0$
- the shape of the molecule is trigonal planar
- the ΔEN for the B – F bond is 1.94, so the B – F bond is very polar
- however, the molecule is symmetrical, so the dipole moments are pulling equally in opposite directions
- the dipoles cancel out and $BF_3$ is a non-polar molecule

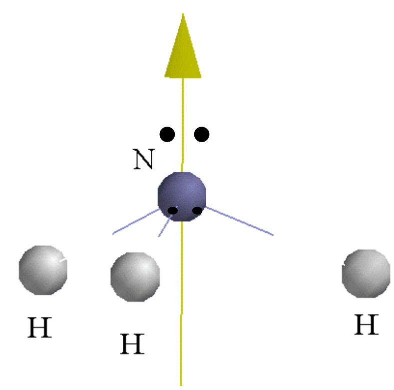


eg. $NH_3$ (ammonia)

- Lewis structure is shown to the right:
- the VSEPR notation for $NH_3$ is $AX_3E_1$
- the shape of the molecule is trigonal pyramidal
- the ΔEN for the N – H bond is 0.86, so the N – H bond is polar
- the ammonia molecule is asymmetrical, it has a lone pair at one point of the tetrahedral
- because the molecule is asymmetrical, the dipole moments will not cancel out and $NH_3$ is a polar molecule




eg. $H_2O$ (water)

- Lewis structure is shown to the right:
- the VSEPR notation for $H_2O$ is $AX_2E_2$
- the shape of the molecule is bent
- the ΔEN for the O – H bond is 1.24, so the O – H bond is polar
- the $H_2O$ molecule is not symmetrical in ALL planes
- because the molecule is asymmetrical, the dipole moments will not cancel out and $H_2O$ is a polar molecule