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:

<table>
<thead>
<tr>
<th>Difference</th>
<th>Bond Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>&lt; 0.50</td>
<td>pure covalent bond: the electrons are essentially equally shared between the bonding atoms</td>
</tr>
<tr>
<td>0.50 to 1.70</td>
<td>polar covalent bond: the electrons are shifted toward the more electronegative atom, creating small or partial charges, $\delta^-$ and $\delta^+$</td>
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<tr>
<td>&gt; 1.70</td>
<td>ionic bond: the electrons are essentially removed from the metal and shifted to the non-metal, creating full ionic charges</td>
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</tbody>
</table>

Polar bonds are said to have a “dipole moment” or a dipole 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$_2$E$_0$, it is a linear molecule
- the $\Delta$EN 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 AX₂E₀, it is a linear molecule
- the ∆EN for the C – O bond is 0.89, so the C – O bond is polar
- the ∆EN 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 CO₂ (and they are)

eg. BF₃ (boron trifluoride)
- Lewis structure is shown to the right
- the VSEPR notation for BF₃ is AX₃E₀
- the shape of the molecule is trigonal planar
- the ∆EN 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 BF₃ is a non-polar molecule

eg. NH₃ (ammonia)
- Lewis structure is shown to the right
- the VSEPR notation for NH₃ is AX₃E₁
- the shape of the molecule is trigonal pyramidal
- the ∆EN 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 NH₃ is a polar molecule

eg. H₂O (water)
- Lewis structure is shown to the right
- the VSEPR notation for H₂O is AX₂E₂
- the shape of the molecule is bent
- the ∆EN for the O – H bond is 1.24, so the O – H bond is polar
- the H₂O molecule is not symmetrical in ALL planes
- because the molecule is asymmetrical, the dipole moments will not cancel out and H₂O is a polar molecule