Bond polarity

- See electronegativity table on course website

If difference in electronegativity between atoms in a covalent bond is greater than 0.4, the bond is considered polar

- this means that one atom (the more electronegative one) attracts the e⁻ more strongly than the other atom – one side of the bond has more e⁻ density than the other

2.5 2.5 = 0
C – C Not polar

δ^+ δ^-

partially (+) partially (−)

electrons are closer to B so it has a partial (−)

2.5 3.5 = 0
C – O

3.0 3.5 = 0.5
N – O

2.5 2.5 = 0
C – S Not polar

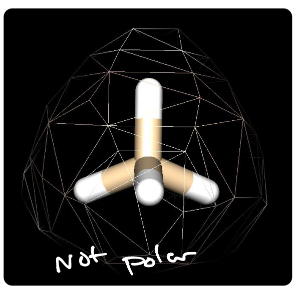
Today we are going to work toward understanding the forces that dictate the temperature that a substance melt (T_m) or boils (T_b) at.

For example, methane (CH_4) is a gas at room temperature, but water is a liquid. Why?

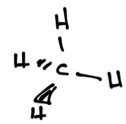
These forces are collectively known as Intermolecular Forces (IMF) and to predict them, you need to be able to predict the shape of a molecule AND whether it is polar

Molecular Polarity is a sum of individual bond polarities → a polar molecule MUST contain at least one polar bond, BUT nonpolar molecules can also contain polar bonds

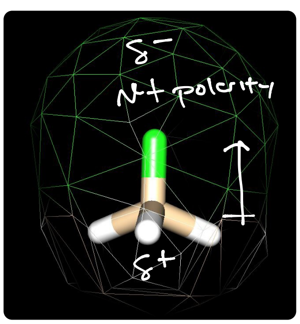
Let's look at a few examples: CH_4



Non-polar

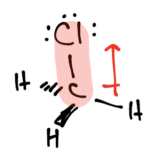


None of these bonds are polar, so this is not a polar molecule

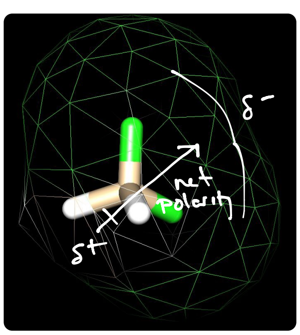


CH_3Cl → replacing one Hydrogen with a Cl

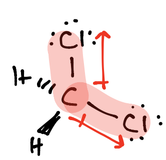
With only one polar bond, this molecule is polar



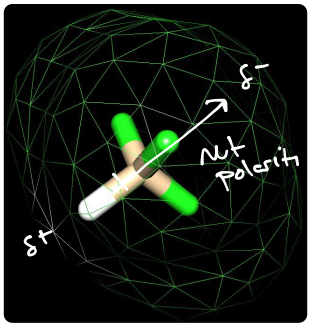
C−Cl bonds are polar
2.0 3.0



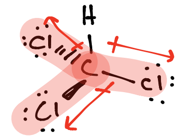
CH_2Cl_2



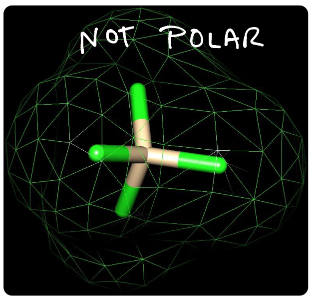
This molecule is also polar. One whole side of the molecule is δ−



$CHCl_3$

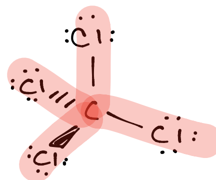


yes, this molecule is polar



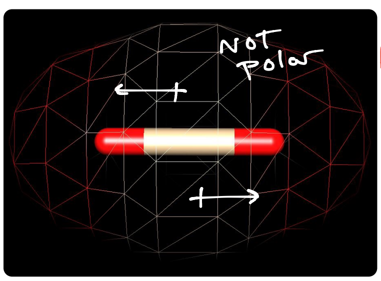
CCl_4

ALL BONDS are polar, but Not polar!

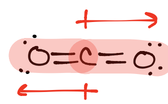


All bonds are polar
- since this molecule is perfectly symmetrical, the molecule is

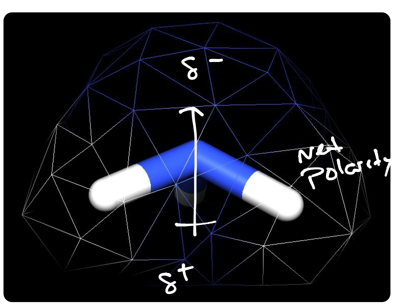
NOT POLAR



ANOTHER EXAMPLE of this: CO_2



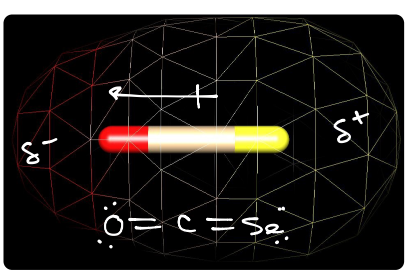
linear molecule
polar bonds cancel out

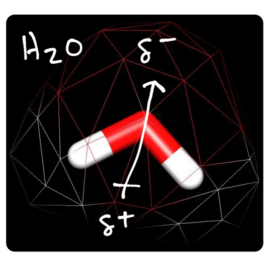


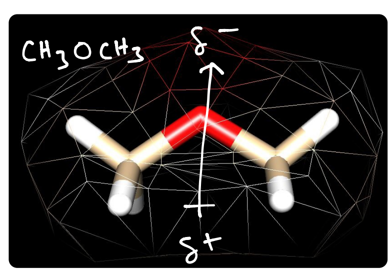
NH_3

Polar!

Trigonal pyramidal
these don't cancel out!







So why does this matter? Intermolecular Forces are dependent on molecule charge and Polarity!

Four types of IMF: listed from strongest to weakest

Ion-Ion > H-bond > dipole-dipole > London Dispersion Forces

$F \propto \frac{1}{r}$ $\quad$ $F \propto \frac{1}{r^3}$ $\quad$ $F \propto \frac{1}{r^6}$

stabilizing Force
(basically coulomb's law)

* IMF Need a (+) + (-) to exist! These can be fixed (Ion), partial (dipole), or induced (LDF)

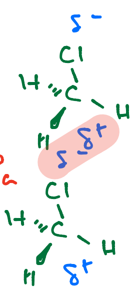
Ion-Ion

e.g. NaCl → two permanent charges interact

- these should be very easy to recognize → any ionic compound

Dipole-Dipole

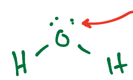
e.g. CH_3Cl → when two polar molecules interact



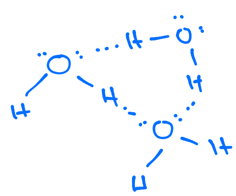
δ+ lines up with a δ-

H-bond: a special / stronger form of Dipole-dipole → shared Hydrogen

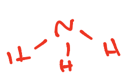
* ONLY occurs when Hydrogen is linked to N, O, or F! *



can accept an H-bond

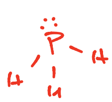


water forms a very beautiful network of H-bonds!



yes

H - F: yes

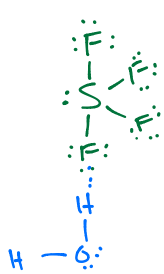


NO

In general, H-bonding pattern

X—H····Y

where X & Y are both N, O, or F



this molecule CAN be a part of H-bonds, but only the acceptor (Y in general pattern)

London Dispersion Forces → these exist in ALL molecules

- they occur when a dipole is induced in a nonpolar molecule
- stronger as molecular weight increases

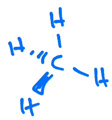
These forces are the only way that non-polar molecules form liquid and solids

- Which has stronger London Dispersion forces: CH_4 or CF_4

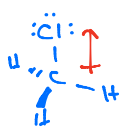
both are nonpolar

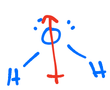
CF_4 wins because it's bigger

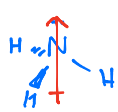
Determine all IMF: CH_4, CH_3Cl, H_2O, NH_3, NCl_3, XeF_4, IF_5

CH_4 →

Only LDF

nonpolar

CH_3Cl →

LDF + dipole/dipole

H_2O →

LDF, dipole/dipole + H-bond

NH_3 →

LDF, dipole/dipole + H-bond

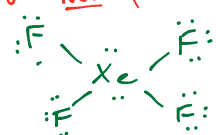
NCl_3 →

LDF

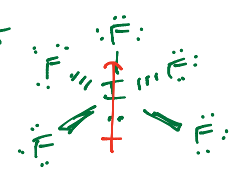
N-Cl bonds are NOT polar!

XeF_4

nonpolar


LDF

IF_5



IF another H-bond donor is around, H-bond is possible

LDF, dipole/dipole

but there is not, so we don't count it

Trends in melting and boiling temperatures can be predicted by understanding IMF

* the stronger the IMF, the higher the T_b & T_m (more energy needed to break the interaction) *

I_2 = solid Br_2 = liquid, Cl_2 = gas : Justify this

- each of these molecules experience LDF + nothing else. Since size determines the strength of LDF, bigger molecules have stronger IMF

$I_2 > Br_2 > Cl_2$ Cl_2 must have a T_b below room temp

Br_2 T_b > room temp > T_m → liquid @ RT

I_2 T_m > room temp → so solid @ RT

List these by increasing T_m: CH_2F_2, F_2, NaF, HF

IMF strength

3. CH_2F_2 → polar → LDF & dipole/dipole
4. F_2 → nonpolar → LDF only
1. NaF → ion/ion & LDF
2. HF → H-bond, dipole/dipole, LDF

$F_2 < CH_2F_2 < HF < NaF$