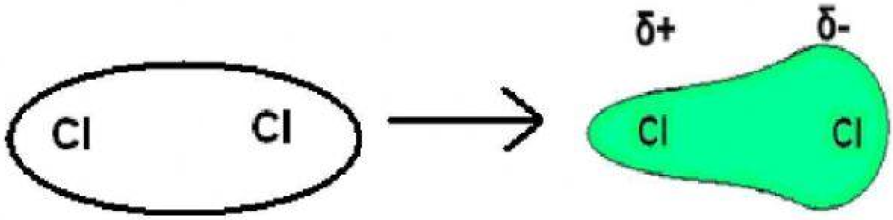
Induced dipole forces (London, Dispersion or Momentary Dipole forces)

Occurs when two non-polar molecules come into contact.
Electrons are constantly moving around the nucleus. At any moment it's possible that more electrons lie to the one side of the molecule, creating a dipole for a split second. When two non-polar molecules have a temporary dipole they will attract each other.

Examples: Halogens

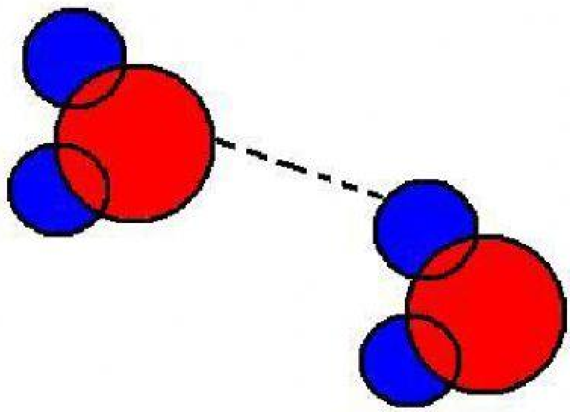


Other examples:
Other non-polar molecules like CO_2 and CCl_4
Condensed noble gases

Hydrogen bonds

Involve a hydrogen atom covalently bonded to a **highly electronegative atom**.

It is a **special strong type of dipole-dipole force**. It is seen when hydrogen bonds to **N, O or F**. These forces are relatively strong intermolecular forces and are stronger than other dipole- dipole forces. Hydrogen bonds are weaker than ionic bonds.



Other examples: NH_3 and HF

Watch this video explaining hydrogen bonds

Intermolecular forces will determine the following:

Which substances are likely to dissolve in which other substances

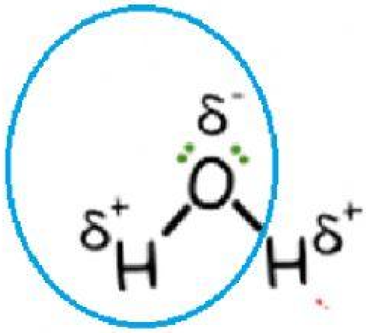
Melting point, boiling point, viscosity and vapour pressure of substances

Difference between molecular and bond polarity

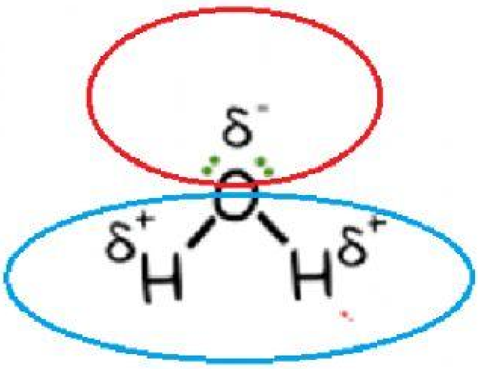
Bond polarity looks at one bond at a time and asks if the bond itself is polar

Eg 1) In the water molecule below the **bond between the bonded H and O is polar**, since they have an EN (electronegativity) difference.

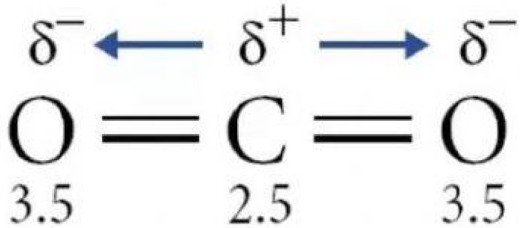
{Hydrogen's EN = 2,1 AND Oxygen's EN is 3,5}



The molecular polarity: the molecule is polar (due to it's shape and the fact that it has a slightly positive and negative side on opposite ends of the molecule)



Eg 2) Compare that to the CO_2 molecule



The bond between C and O is polar, since there is an EN difference between the 2 atoms. (notice that you never add the electronegativities of the 2 oxygens together. You always consider one at a time)

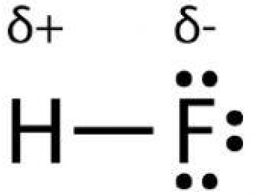
Eg 3) H-H

The bond polarity: is non-polar, since there is no EN difference between H and the other H.

Similarly, any diatomic molecule will always have non-polar bonds.

The molecular polarity is also non-polar since the molecule does not have positive and negative opposite ends.

Eg 4)



The bond polarity: is polar, since there is an EN difference between H and the other F.

The molecular polarity is: polar since the molecule has two oppositely charged ends.

- With an electronegativity difference $\Delta EN > 2.1$ electron transfer will take place and the bond would be ionic
- With an electronegativity difference $\Delta EN > 1$ the bond will be covalent and polar
- With an electronegativity difference $\Delta EN < 1$ the bond will be covalent and very weakly polar
- With an electronegativity difference $\Delta EN = 0$ the bond will be covalent and nonpolar

Watch this video on molecular polarity before continuing

State the

- ✓ Molecular polarity of the following and
- ✓ The type of **intermolecular forces** in the following

Molecule	Polar or non-polar bond	Polar or non-polar molecule	Type of intermolecular force Choose between London Dipole-dipole Hydrogen
H_2			
N_2			
HCl			
HBr			
HF			
H_2O			
CO_2			
CH_4			

VSEPR theory

DEFINITION: *Valence Shell Electron Pair Repulsion Theory*

Valence shell electron pair repulsion (VSEPR) theory is a model in chemistry, which is used to predict the shape of individual molecules. VSEPR is based upon minimising the extent of the electron-pair repulsion around the central atom being considered.

Molecular shapes: (VSEPR theory)

- Molecular shape is determined by the repulsion between electron pairs present in the valance shell of the central atom
- The number of electron pairs around the central atom can be determined by writing the Lewis structure for the molecule
- The shape of the molecule depends on the number of bonding electron groups (or atoms bonded to the central atom) and the number of lone pairs on the central atom
- A is used represent the central atom and X is used to represent terminal atoms.
- There are five ideal shaped found when there are NO lone pairs on the central atom (thus only bonding pairs)

Molecular shape

✓ Remember to draw the lewis structures before you try to determine the shape of a molecule

✓ Then look at the whether there are lone pairs on the terminal (central atom)

✓ Whether or not it has lone pairs has a huge impact on the shape

There are 5 basic shapes if the central atom has zero lone pairs

Number bonding pairs on central atom	Number of lone **pairs on central atom**	Geometry	Examples
1,2	**0**	linear	CO_2, $BeCl_2$
3	**0**	trigonal planar	BF_3
4	**0**	tetrahedral	CH_4
5		Trigonal bipyramidal	PCl_4
6	**0**	octahedral	SF_6

If the central atom does contain lone pairs

If the central atom does contain one or more lone pairs of electrons then the molecule will not have one of the above ideal shapes of VSPER theory. Then you must adapt the shape and use the following rule:

lone pair - lone pair repulsion

lone pair - bonding pair repulsion

Bonding pair repulsion-bonding pair

↓

Repulsion force decreases

Number bonding pairs on central atom	Number of lone pairs on central atom	Geometry	General formula	Examples	Reason
2	1	bent or angular	AX_2E	SO_2	The lone-pair vs. bonding pair repulsion is greater than the bonding pair vs. bonding pair repulsion
3	1	trigonal pyramidal	AX_3E	NH_3	The lone-pair vs bonding pair repulsion is greater than the bonding-pair vs bonding pair repulsion
2	2	bent or angular	AX_2E_2	H_2O	The lone-pair vs lone-pair repulsion is greater than the lone-pair vs bonding pair repulsion