

CHEMICAL BONDING

Three main sections to this module

- **Intramolecular bonds**
 - (a) The covalent bond
 - (b) Electronegativity and molecular polarity
 - (c) Basics of ionic and metallic bonding
- **Intermolecular forces**
 - (a) Basic molecular geometry
 - (b) The concept of van der Waals forces
 - (i) Induced dipole (London Forces) - NON POLAR
 - (ii) Dipole-dipole (Keesom forces) - POLAR
 - (ii) Hydrogen bonding intermolecular forces
- **Physical Properties**
 - (a) Melting and boiling points
 - (b) Electrical conductivity

INTRAMOLECULAR BONDS

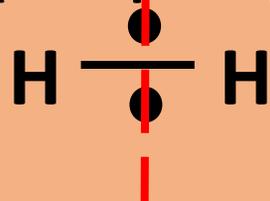
DEFINITION : *a bond that occurs between atoms within molecules*

Molecules are chemical compounds that are produced when a covalent bond is formed between two or more non-metal atoms

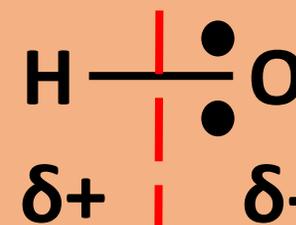
DEFINITION: *a covalent bond is a sharing of at least one pair of electrons by two non-metal atoms*

2 types of covalent bonds

Pure (non-polar) covalent



Polar covalent



DEFINITION: *Non-polar(pure) covalent bond is the equal sharing of electrons within that bond*

DEFINITION: *Polar covalent bond is the unequal sharing of electrons leading to the formation of a dipole*

Electronegativity

DEFINITION : the measure of the tendency of an atom to attract a bonding pair of electrons

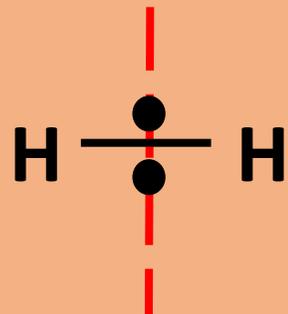


diagram 1

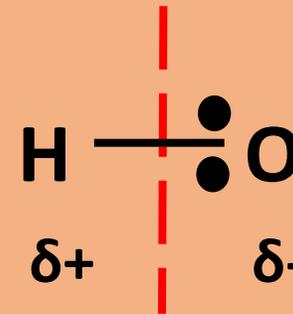


diagram 2

Consider the examples of covalent bonds shown above. In diagram 1, the two atoms involved with the bond are identical, hence the intramolecular forces between the two atoms will be the same allowing for the shared electron pair to be shared equally between the atoms. This is now known as a symmetrical distribution of charge within the chemical bond.

In diagram 2, it can be clearly seen that here we have different atoms involved with the bond and the O atom is exerting a stronger force of attraction than the H atom on the shared electron pair. The electrons are thus pulled closer to the O atom and thus we say that there is an asymmetrical distribution of charge within the chemical bond.

The electronegativity (Pauling) scale

This is a comparative scale of attractive forces developed by an American chemist by the name Linus Pauling, hence the name given to the scale. Pauling took the atom of the element with the greatest attractive force, namely fluorine and assigned it an arbitrary value of 4.0 called the 'electronegativity number' for fluorine. All other elements were assigned electronegativity numbers relative to fluorine and these numbers were placed in the Periodic Table for each element.

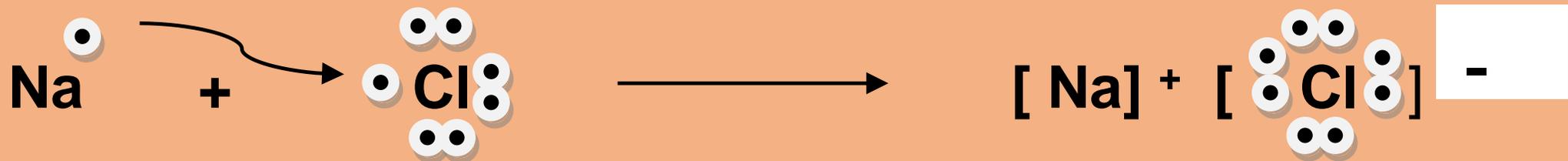
	7	3,0	8	3,5	9	4,0	10
	N		O		F		
	14		16		19		20
	15	2,1	16	2,5	17	3,0	18
	P		S		Cl		
	31		32		35,5		40
	33	2,0	34	2,4	35	2,0	36
	As		Se		Br		Kr

Ionic bonding

DEFINITION: the transfer of electrons and subsequent electrostatic attraction

Ionic bonding occurs between metal and non metal atoms which then form ions upon the transfer of electrons. Metal atoms form metal cations and non metal atoms form non-metal anions

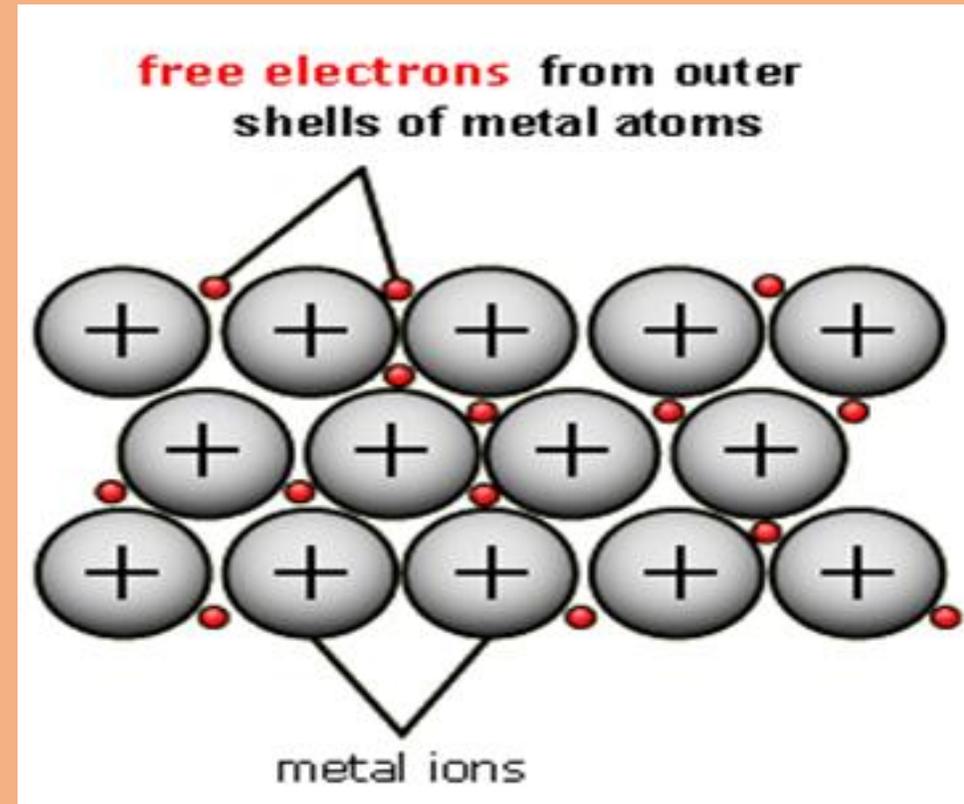
This is when two atoms, one metal and the other non metal are able to bond through the process of electron transfer from the metal atom to the non metal atom to form an ionic compound.



Metallic bonding

DEFINITION: *bonding between a positive atomic kernel and a sea of delocalised electrons*

This is when a crystal lattice of metal atoms are found where the atoms are embedded in a sea of delocalised electrons formed due to the overlapping of the outermost orbitals of the metal atoms which allow for the electrons to move easily through the lattice structure.



Practice example

A, B, C and D represents elements with consecutive atomic numbers from the Periodic Table. C is a noble gas.

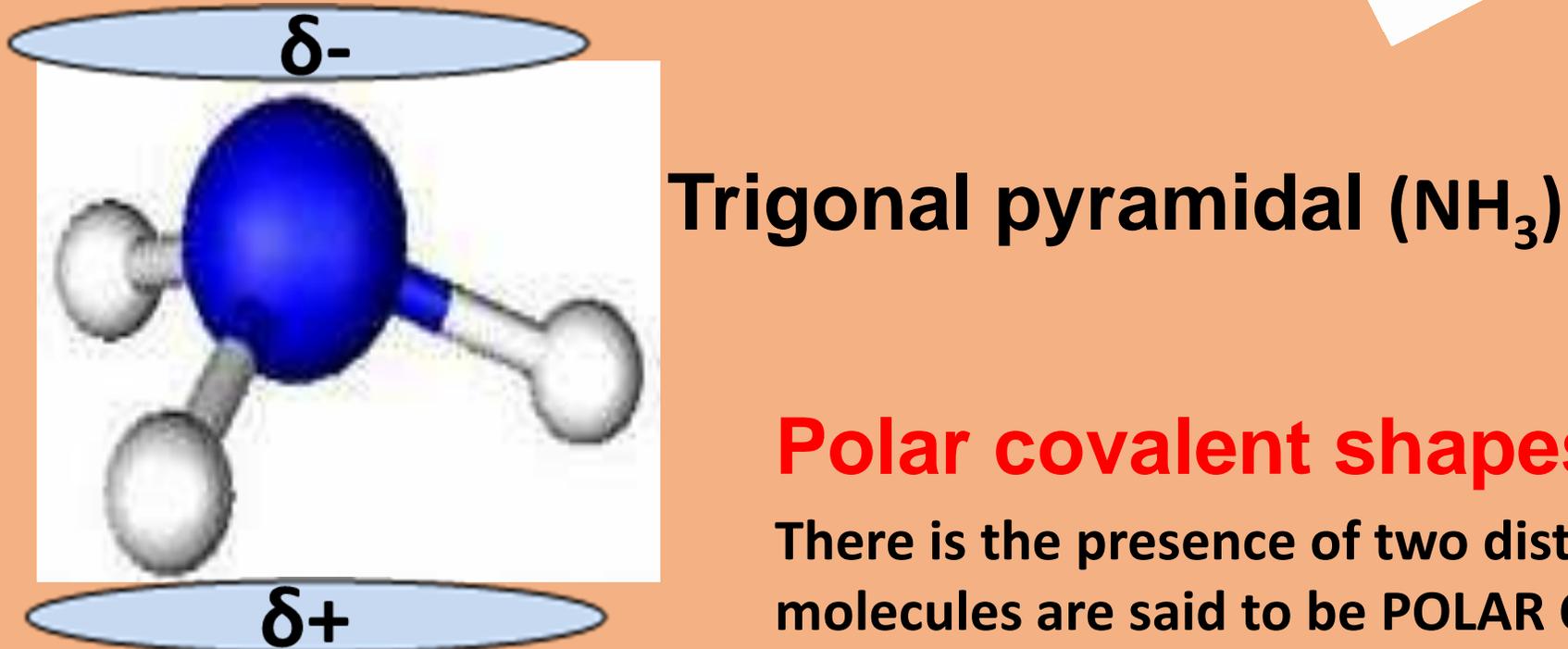
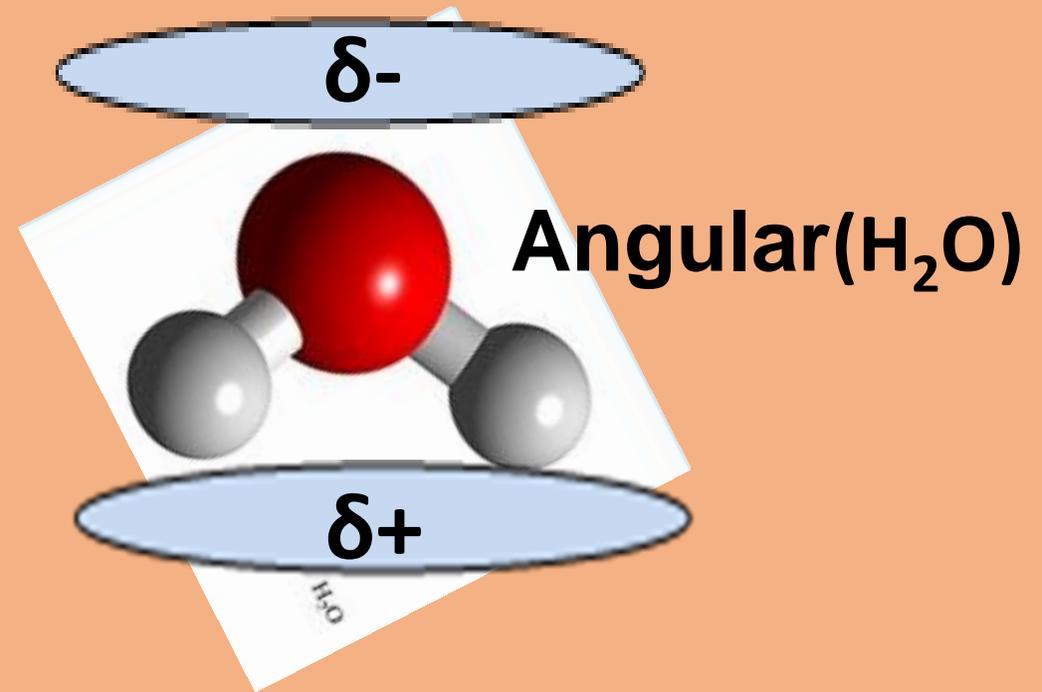
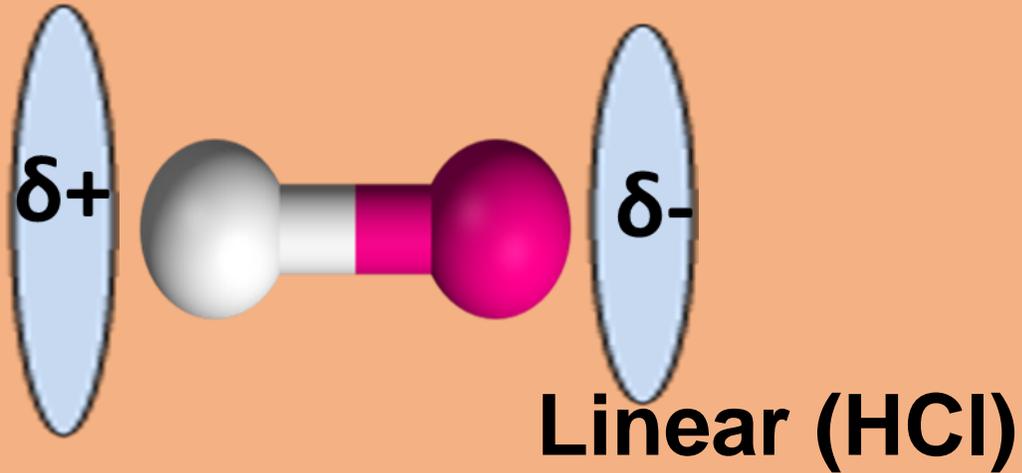
1. What does the electronegativity of an atom represent?
2. Which one, (A, B, C or D) would you expect to have the highest electronegativity value?
3. What type of bond could be expected when B and D form a compound?
4. Write down the formulae of the compounds formed, in terms of the letters given, when hydrogen combines with A and B respectively.
5. Which of the hydrides (hydrogen compounds) of A or B will exhibit greater polarity? Explain.

Memorandum

- 1. This is the measure of the amount of attraction an atom has on a shared pair of electrons within a chemical bond.**
- 2. Bthis element has to be a Group 17 halogen. These atoms exert the strongest attraction on the valence electrons, thus will have the highest electronegativity value .**
- 3. Ionic bond.....if B is a halogen, then D is a Group 1 alkali metal. This will lead to a metal bonding with a non metal which will form an ionic bond.**
- 4. H_2A and HB**
- 5. HB.....B is a halogen with a higher electronegativity, thus stronger force of attraction on the shared electron pair, hence greater polarity**

INTRAMOLECULAR FORCES

Molecular geometry

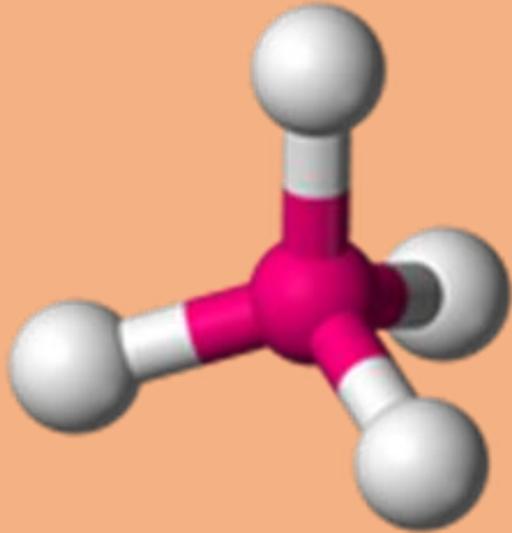


Polar covalent shapes

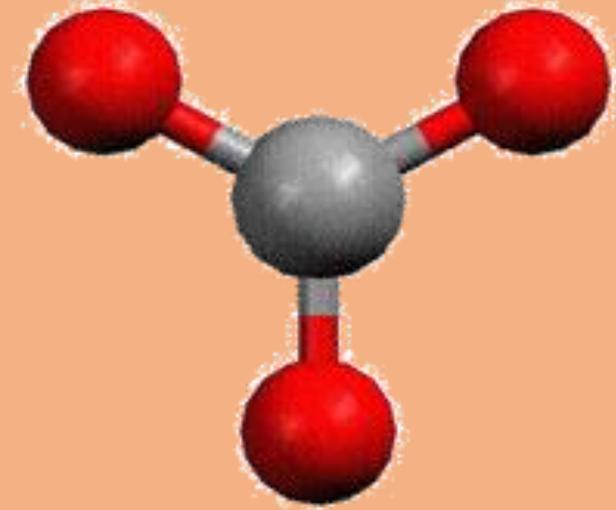
There is the presence of two distinct regions of polarity, thus molecules are said to be POLAR COVALENT

Molecular geometry

Non Polar covalent shapes



Tetrahedral(CH_4)



Trigonal planar(BF_3)

Symmetrical molecular structures - there is no distinct region of polarity, thus the molecules are non polar

Note : the intramolecular bonds may be polar, but the overall molecule is non-polar by symmetry

The concept of intermolecular forces

DEFINITION: *a weak force of attraction between molecules, ions or atoms of a noble gas*

(important to distinguish between intramolecular bonds and intermolecular forces)

- Also known as **van der Waals forces**
 - (i) Induced dipole (London Forces) - NON POLAR
 - (ii) Dipole-dipole (Keesom forces) - POLAR
 - (ii) Hydrogen bonding intermolecular forces



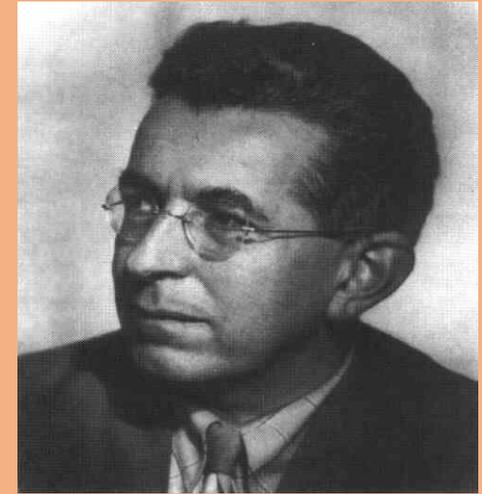
Johannes Jurgens van der Waals

The most important properties of van der Waals forces are :

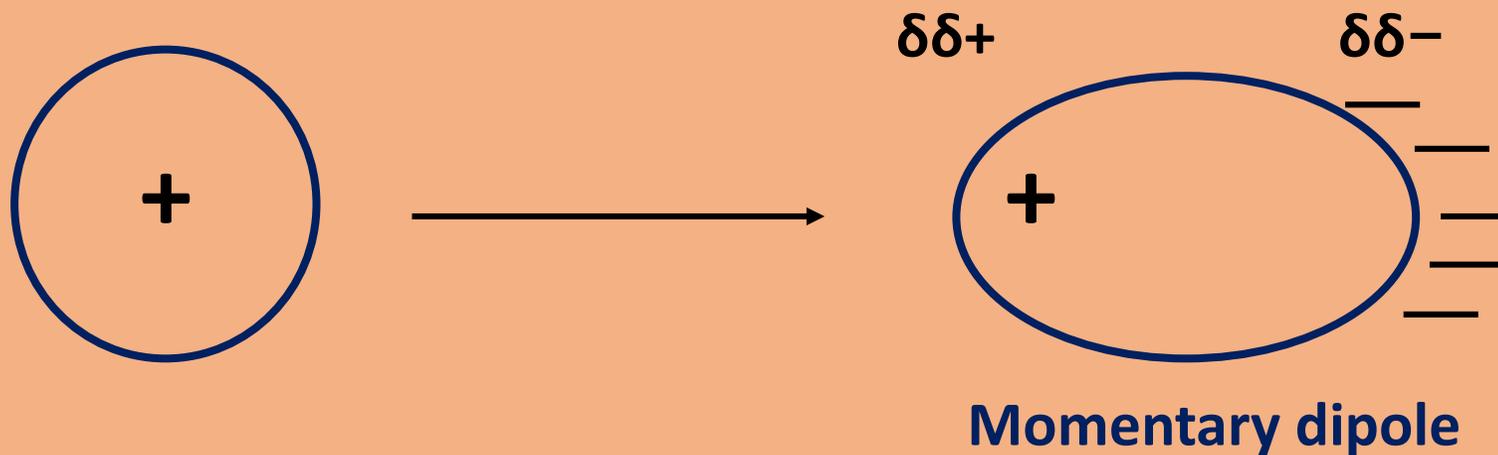
- They are not bonds, but forces of attraction
- They are relatively weak forces of attraction between molecules that are overall electrically neutral
- They arise due to the proximity of molecules
- These forces of attraction are exerted in all directions

Induced dipole (London Forces)

- These intermolecular forces were discovered by another Dutch physicist, Fritz London in 1929
- Intermolecular forces that arise between non polar molecules.

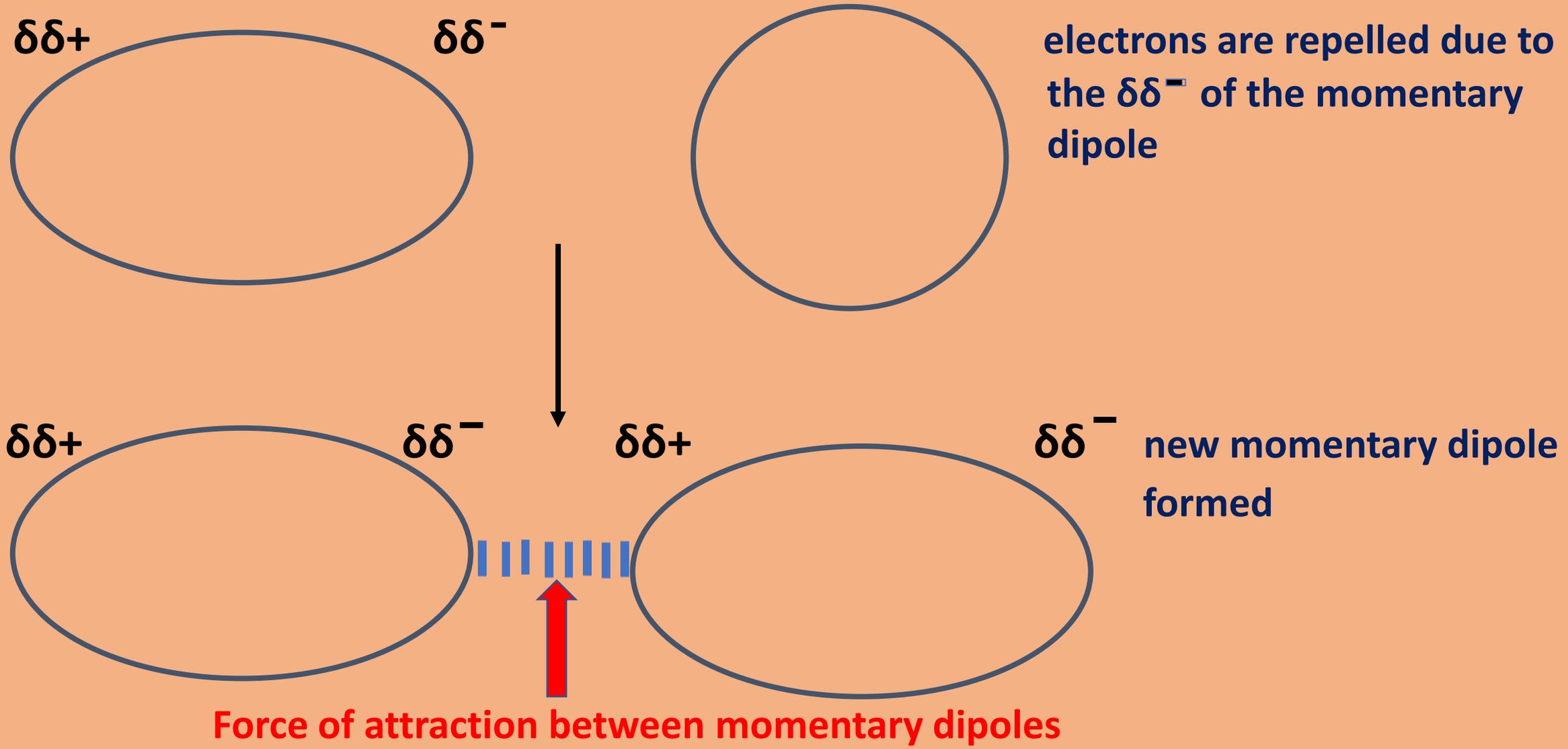


When non polar molecules approach each other, the negative electrons in the space around the positive nucleus of each atom will repel each other.



These electron spaces, also known as the electron cloud, become distorted so that the electron cloud becomes momentarily unsymmetrical. This is sometimes called an “electron wobble” and this wobble literally last only for a brief moment – hence the name **momentary dipole** (also called an instantaneous dipole)

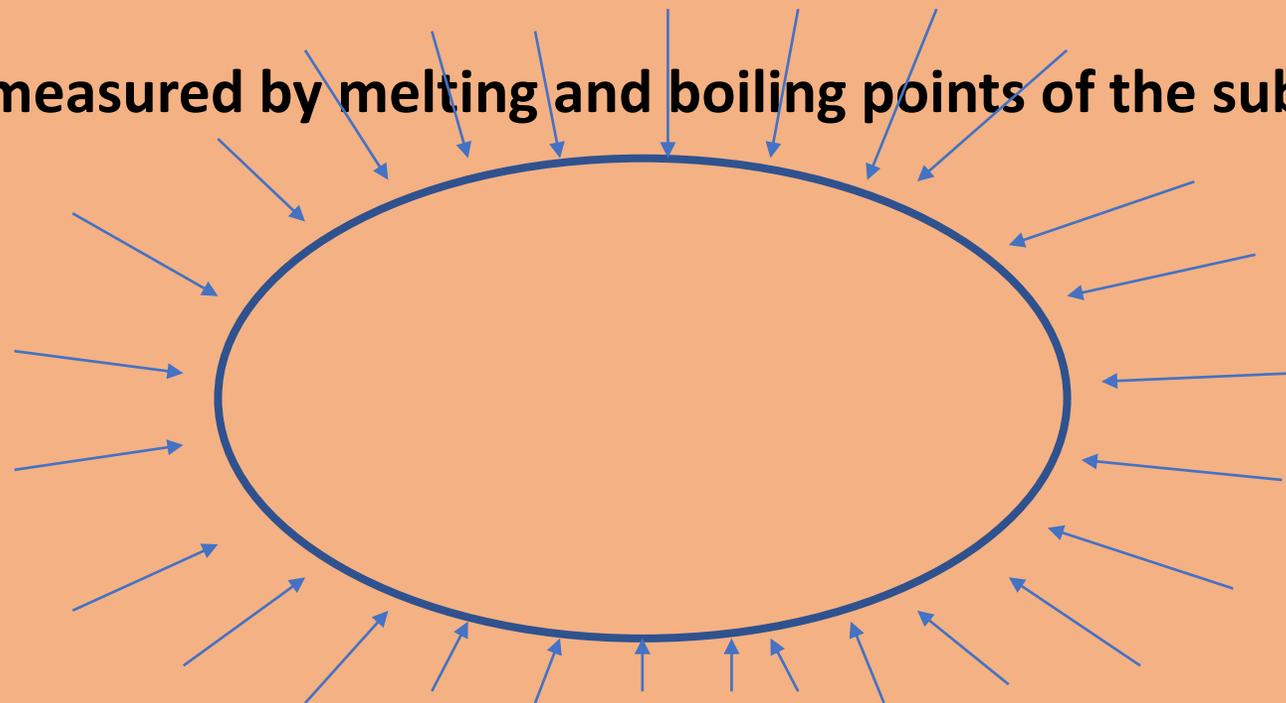
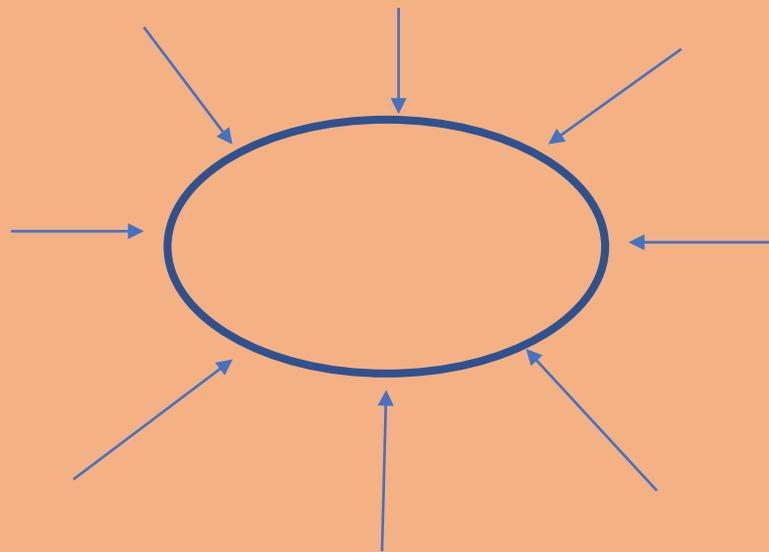
This momentary dipole can now cause a molecule close by to become polarised through the process of induction – thus we now find an instantaneous induced dipole.



The opposite charges of the momentary dipoles attract one another and hence give rise to a weak intermolecular force.

The different strengths of London forces is influenced by the size of the momentary dipoles. This is due to :

- **The size of the momentary dipole due to the electron cloud density**
- **The greater the electron cloud density, the larger the surface area of the momentary dipole**
- **The larger the surface area, the greater the number of positions of contact where attractions can take place.**
- **Strengths of London forces are measured by melting and boiling points of the substances**



Melting points and boiling points

- Melting and boiling points represent the amount of energy needed to transition between phases.
- This energy is equivalent to the strength of the intermolecular forces between the particles of the substance
- Low melting/boiling point = weak intermolecular forces
- High melting/boiling point = strong intermolecular forces
- Consider melting points of the following diatomic and Noble gases

N_2	O_2	H_2	Ne
- 196	- 183	-253	-246

- Consider the boiling points of the following halogens. (non polar molecules with London IMF between particles)

F_2	Cl_2	Br_2	I_2
- 188	-34	+ 59	+ 184

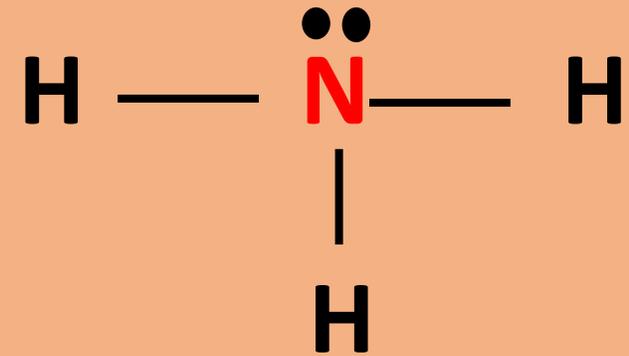
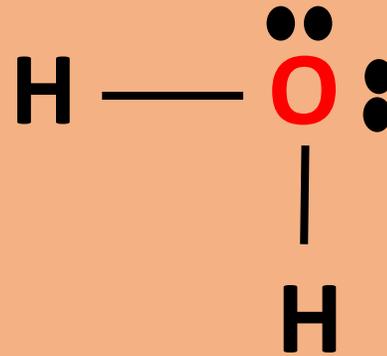
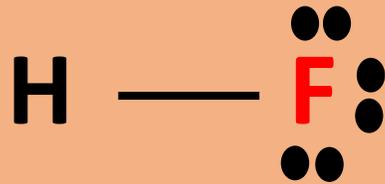
Dipole –Dipole (Keesom Forces)

- These intermolecular forces were discovered by Wilhelm Hendrik Keesom in 1921
- Are intermolecular forces that are formed between polar molecules.
- These attractive forces are somewhat stronger than the London forces as the molecule have permanent dipoles.



Hydrogen bonding intermolecular forces

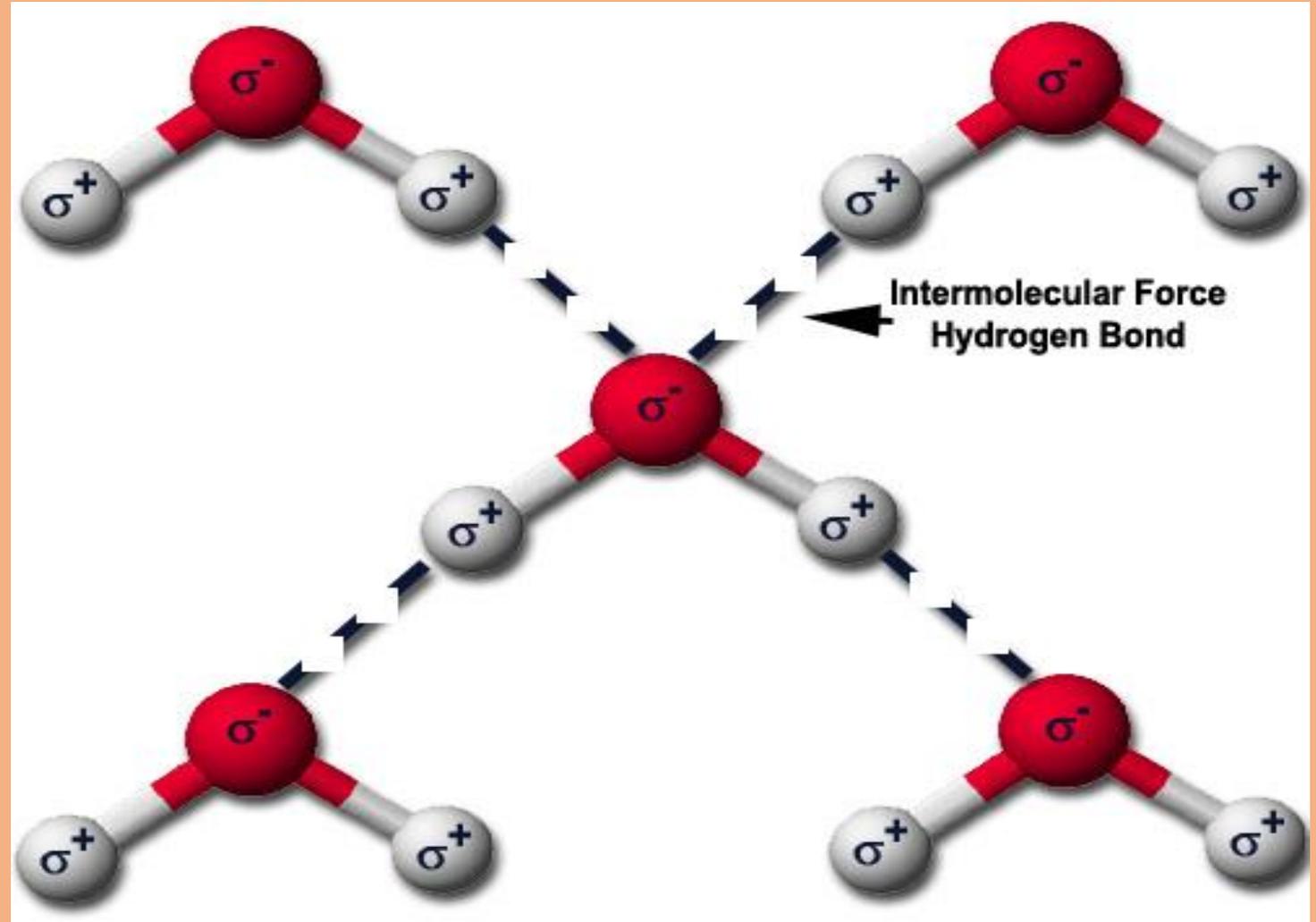
- A special type of dipole-dipole IMF
 1. Hydrogen atoms bonded covalently to a small atom of high electronegativity
 2. There must be at least one pair of lone pair electrons



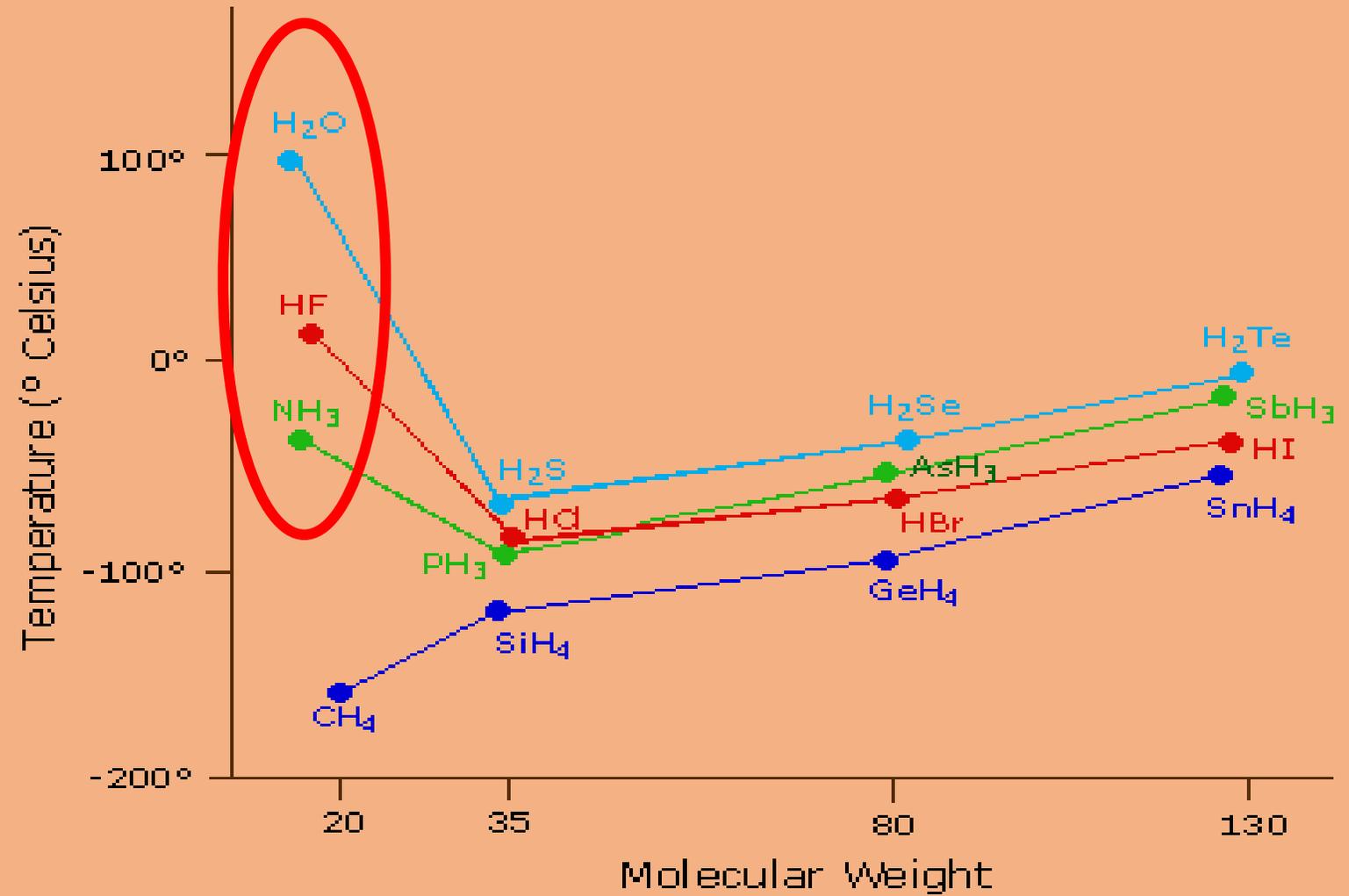
- Atoms that fit these criteria are **nitrogen, oxygen** and **fluorine** which all three, when bonded with hydrogen will exhibit hydrogen bonding forces between their molecules.
- This means that the molecules are highly polar with a large electron cloud density on the δ^- end of the molecule and a dense δ^+ on the other due to the very small hydrogen atom. The dipole-dipole forces between the molecules are much stronger than normal dipole-dipole forces

Here is an example of water which shows hydrogen bonding forces between each molecule. It is important to note that these forces are directional in that the intermolecular force is between H and O specifically

Note as well that these can be multiple H – bonding forces to each water molecule. The forces are specifically between the $\delta+$ H atoms and the lone pairs associated with each $\delta-$ O atom



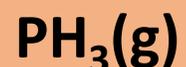
Graphical representation of boiling points of various hydrides



Practice examples

Question 1

- 1 Use only substances from the list below when answering Question 1.1 to 1.6. The state symbols (phase indicators) represent the physical state of each of the substances at room temperature.**



Select one substance from the list that has: (Only write down the question number and the formula of the substance next to it.)

1.1 Pure covalent intramolecular bonds.

1.2 A high melting point due to the strong electrostatic attraction between the cations and anions in the crystal lattice.

1.3 Hydrogen bonding intermolecular forces.

1.5 Delocalised valence electrons.

1.6 Dipole – dipole intermolecular forces.

Question 2

2. Consider the table below, which shows the boiling point of the halogens.

Halogens	F ₂	Cl ₂	Br ₂	I ₂
Boiling point (°C)	-188	-35	59	184

2.1 Name the specific type of intramolecular bond that occurs between the atoms in a fluorine molecule.

2.2 The table shows that the boiling point of the halogens increases from F₂ to I₂. Explain this trend by making reference to the relevant intermolecular force between the halogen molecules and the factor influencing its strength.

QUESTION 1 - MEMORANDUM

1.1 PH₃

1.2 KF

1.3 H₂O

1.4 Mg

1.5 HCl

2.1 pure (non-polar) covalent

2.2

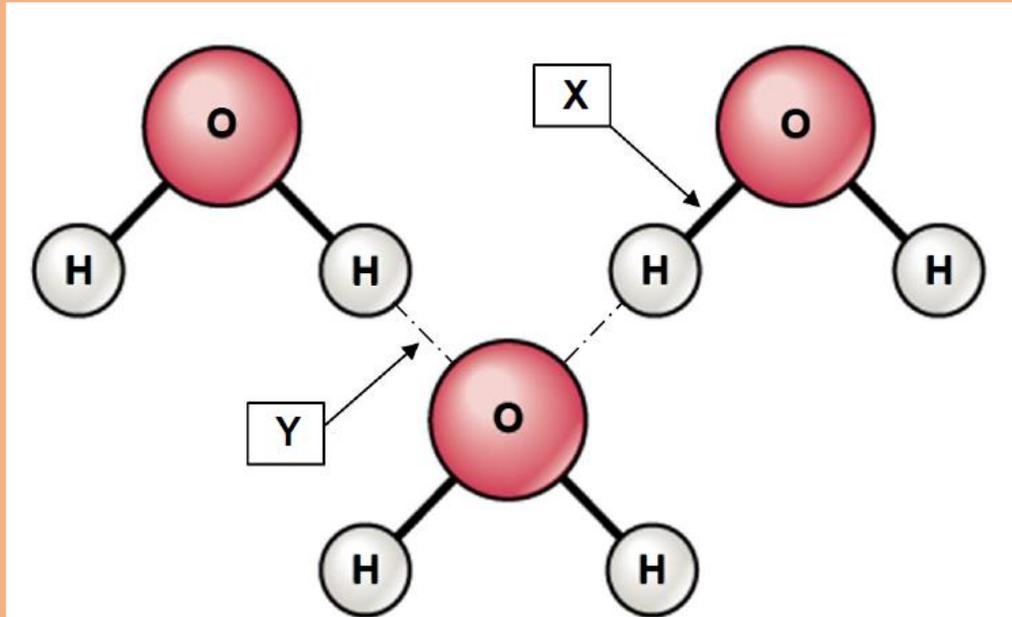
Intermolecular force = induced/temporary/momentary dipole forces or London forces or dispersion forces.

Bigger molecules have MORE ELECTRONS (greater electron density) therefore bigger (stronger) temporary dipoles are set up (or temporary dipoles are set up more easily), or greater electron density creates a greater surface area for more points where intermolecular forces can exist.

MORE ENERGY is needed to overcome the stronger intermolecular forces (NOT bonds) therefore a higher boiling point.

Question 2

2.1 Consider the diagram below, showing an arrangement of water molecules in the liquid phase.



- 2.1.1 What is an intramolecular bond?
- 2.1.2 Name the specific type of intramolecular bond represented by the letter X in the diagram.
- 2.1.3 Define the term *intermolecular force*.
- 2.1.4 Name the specific type of intermolecular force represented by the letter Y in the diagram.
- 2.1.5 State two properties of the oxygen atom that make this type of intermolecular force (Y) possible.
- 2.1.6 What is the partial charge ($\delta+$ or $\delta-$) on the hydrogen atom in a water molecule?

QUESTION 2 – MEMORANDUM

2.1.1 It is a bond occurring between atoms within molecules.

2.1.2 Polar/covalent/bond.

2.1.3 A (weak) force of attraction between molecules or between atoms of noble gases.

2.1.4 Hydrogen bond (Hydrogen bonding intermolecular force).

2.1.5 It is a (small) atom with a high electronegativity and it has at least one lone pair of electrons.

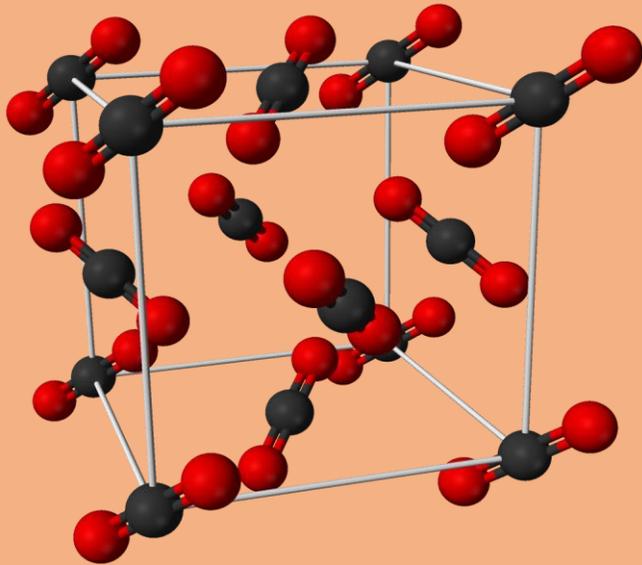
2.1.6 $\delta+$

Intermolecular forces in solids

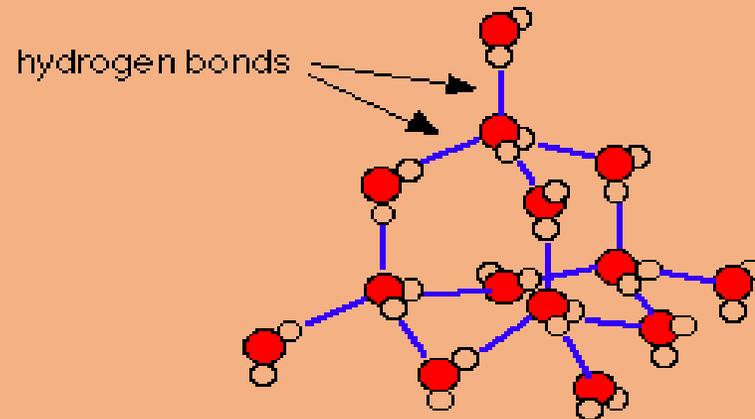
1. Molecular solids

The particles in crystals of this kind of solid consist of molecules. There are two types of intermolecular forces that can be found in molecular solids

a) London type van der Waal's intermolecular forces found in, for example - iodine (I_2), frozen CO_2 (dry ice). These forces are very weak and as a result these substances usually have low melting points



b) Hydrogen bonding intermolecular forces - found in frozen water (ice). The hydrogen bonding forces are stronger than London forces thus the melting points are somewhat higher in comparison, but still pretty low in comparison to the melting points in other solid substances.



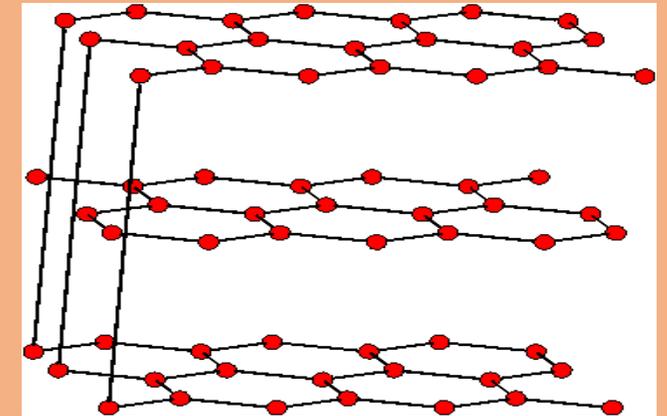
2. Network solids

The crystals of these substances consists of atoms which are strongly covalently bonded into a lattice forming a giant molecule. These giant molecules can be :

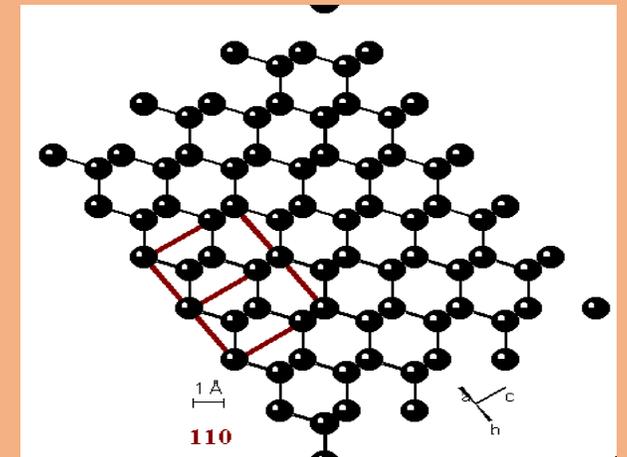
- i. Plate like such as graphite
- ii. Three dimensional as in the diamond

The strong covalent bonds between (intermolecular) the carbon atoms explain the hardness, chemical stability and very high melting point of these types of solids

In **graphite** , the carbon atoms are bonded strongly together by **three** covalent bonds in a **planar** hexagonal pattern. These hexagonal structures are repeated to build up the sheet like structure of graphite. The sheets are layered one on top of each other and held together by weak London forces. This is why graphite is able to leave a mark when pressure is placed on the structure, such as in a pencil.

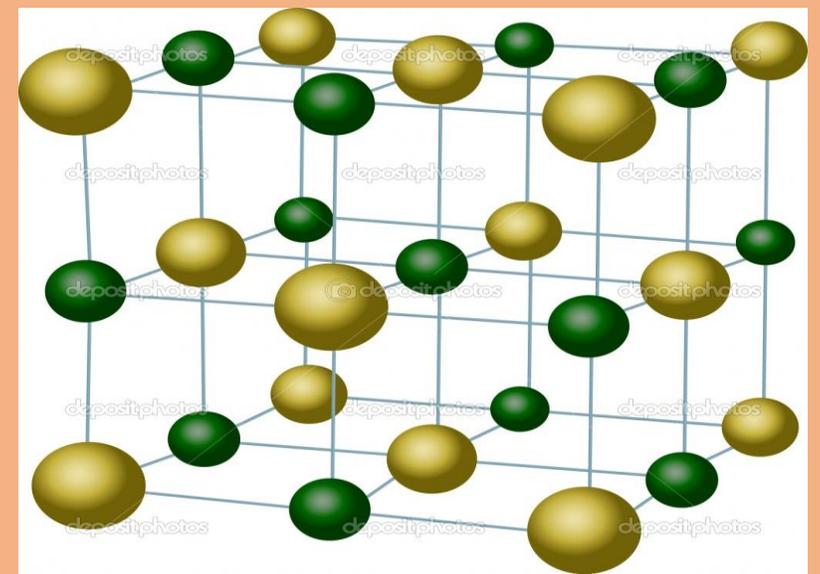


In the **diamond**, the carbon atoms are bonded strongly together by **four** covalent bonds in a **buckled** hexagonal pattern. This buckled pattern is repeated to form a crystal lattice which is now exceptionally strong. The forces holding the carbon atoms together are the forces in the covalent bond itself which is why the diamond has such a very high melting point



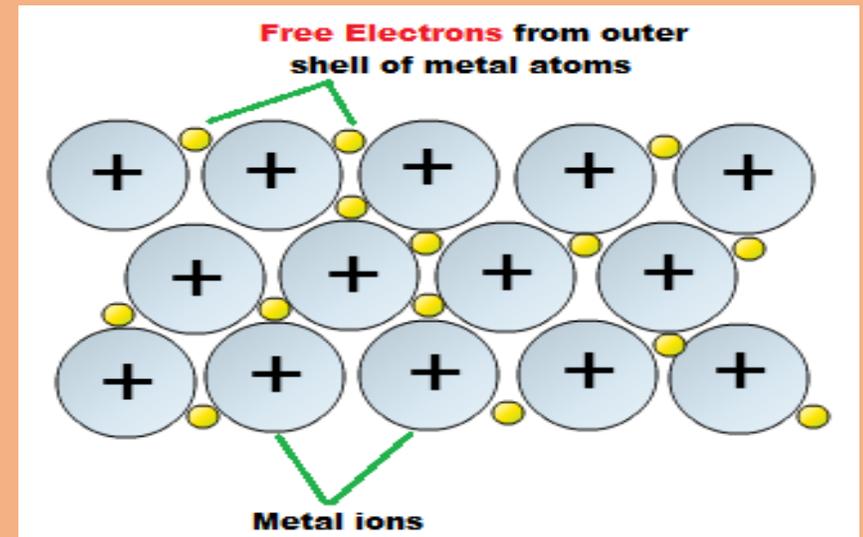
3. Ionic solids

These solids consist of ions with opposite charges (cations and anions) which are bonded into a regular crystal lattice by strong electrostatic forces of attraction (Coulombic forces). Because these electrostatic forces are strong forces, ionic crystals are usually hard and brittle as well as having high melting points



4. Metallic solids

The bonding model for metals yields a structure according to which atoms form the positive atomic kernels surrounded by a sea of delocalised electrons. The force that holds the metallic crystal lattice together is an electrostatic (Coulombic) force of attraction between the positive atomic kernels and the delocalised electrons. The atoms in a metallic crystal are arranged symmetrically as close together as possible. This is called the “close packing” of atoms. These forces of attraction are particularly strong and hence the very high melting points of the metals



SUMMARY

Type of solid	Smallest unit	Intermolecular force	Example	General properties
Molecular	Molecules	London type van der Waals or hydrogen bonding forces	Iodine Water (ice)	low melting points
Network	atoms	Covalent bonding	Diamond Graphite	Very high melting points
Ionic	cations and anions	Ionic bonding (electrostatic forces of attraction)	Sodium chloride Copper sulphate	High melting points
Metallic	Positive atomic kernels	Metallic bonding (electrostatic forces of attraction)	Copper Iron	High melting points