Various forces of attraction between molecules

1. Ionic bonds
2. Covalent bonds (also co-ordinate covalent bonds)
3. Metallic bonds
4. Van der Waals forces
5. Hydrogen bonds

Relationship between forces of attraction and state of matter

Force of attraction	**State of matter**
Ionic	Solid at room temperature and pressure
Covalent (between atoms)	Solid at room temperature and pressure
Metallic	Solid at room temperature and pressure
Van der Waals forces	Usually a liquid or gas at room temperature and pressure or a low melting point solid
Hydrogen bonds	Liquid at room temperature and pressure or a low melting point solid

Physical properties of matter in relation to force of attraction

The stronger the force of attraction, the higher the melting point, which would also determine the solubility of the substance as well.

For example, in group VII, the forces of attraction (van der Waals forces) increases from fluorine to iodine and the melting points increase as well. Fluorine is a gas at room temperature and pressure while bromine is a liquid and iodine is a solid.

In a simple molecular or non-polar substance like iodine, it does not dissolve in a polar substance like water, but its solubility is high in non- polar solvent (**remember the adage "like" dissolves "like"**)

However, giant molecular substances like silica or graphite have very high melting points and are insoluble in water because of the high strength of their forces of attraction.

Formation of:-

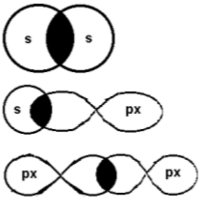
a) **ionic bonds**

Valence electrons from the metal atom are transferred to a non-metal atom resulting in the formation of positive and negative ions. The electrostatic attractions between the positive and negative ions hold the compound together.

b) **covalent bonds**

Covalent bonds are formed from the *overlap of orbitals*.
Two types of bonds are formed:- sigma σ and pi π bonds

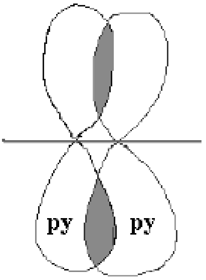
Sigma bonds can be formed via the overlap of s+s OR s + px OR px + px



Sigma bonds are **SINGLE** bonds.

The electrons in a σ bond occupy a region directly between the nuclei of the atoms involved.
In a π bond, the electrons are above and below the plane of the molecule. A double bond is one σ and one π. A triple bond is one σ and two π.

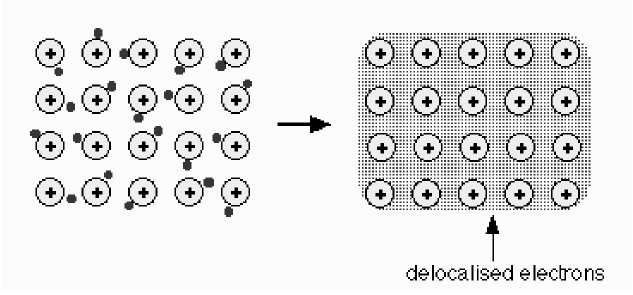
Formation of pi bonds



NB either the ***LATERAL*** *overall of* $p_y + p_y$ ***or*** $p_z + p_z$ *gives a pi bond*

A polar bond exists when there is a large difference in electronegativity between the atoms e.g. HCl, in such a case the more electronegative element would have δ^- ***and the less electronegative element would have*** δ^+ ***Non-polar bonds have little or no difference in electronegativity e.g. O=O, C-H***

c) **metallic bonds**



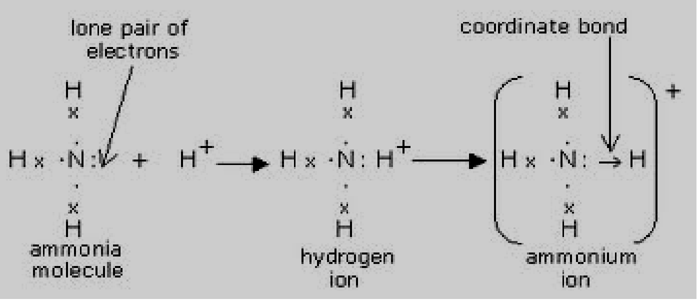
In metals, there is a regular, repeating array of metal ions surrounded by a "sea" of electrons. These electrons are mobile and are called **delocalised electrons**. The electrostatic force of attraction between the delocalised electrons and the positive metal ions forms the metallic bond. The delocalised electrons are responsible for the physical properties of metals.

Co-ordinate or dative bonding

In the formation of a simple covalent bond, each atom supplies one electron to the bond - but that doesn't have to be the case. A co-ordinate bond (also called a dative covalent bond) is a covalent bond (a shared pair of electrons) in which **both** electrons come from the **same** atom. **NB For a dative bond to form, one species MUST have at least ONE lone pair of electrons and the other species must be either a cation, partially positively charged or electron deficient.**

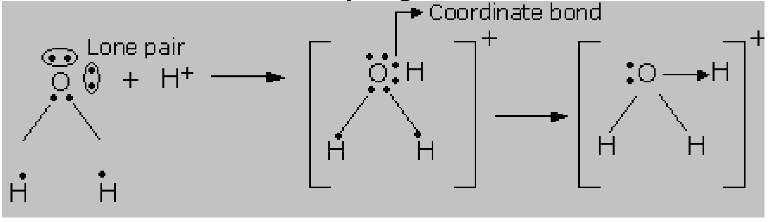
Examples of dative bonding

Reaction between ammonia and hydrogen ions



When the dative bond has been formed, it is indistinguish-able from the other covalent bonds

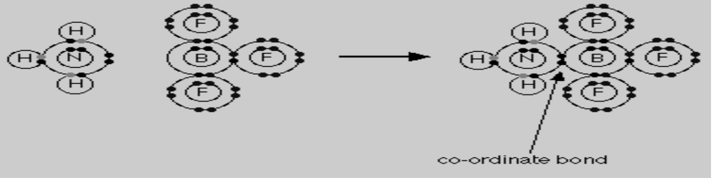
Reaction between water and hydrogen ions



Reaction between ammonia and boron trifluoride, BF_3

Boron only has 3 electrons in its valence shell and when BF_3 is formed, the boron only has 6 electrons in its valence shell as opposed to eight **(octet rule)**, therefore boron is **electron deficient**.

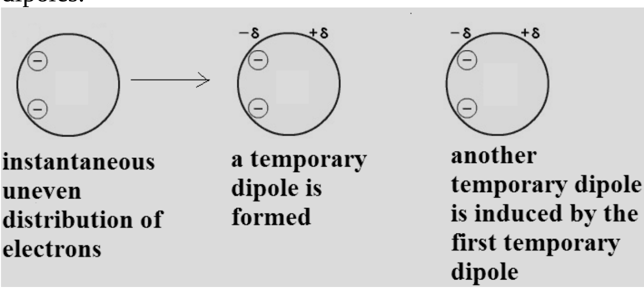
The lone pair on the nitrogen of an ammonia molecule can be used to overcome this deficiency, and a compound is formed involving a co-ordinate bond.



Van der Waals forces

a) Temporary dipole-induced dipole

Generally electrons are distributed evenly in a molecule thus the molecule is non-polar. Electrons are mobile and at any given instant can have more electrons on one side of the molecule than another resulting in a temporary dipole.
This temporary dipole in one molecule can **INDUCE** a dipole in an adjacent molecule which ultimately results in a large number of molecules now possessing dipoles.

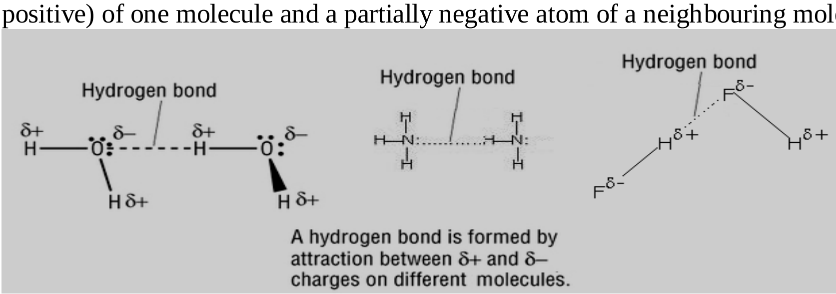


b) Permanent dipole-dipole interactions

A molecule like HCl has a permanent dipole because chlorine is more electronegative than hydrogen. These permanent, in-built dipoles will cause the molecules to have a **stronger attraction** to each other than just the temporary dipole-induced dipole interactions.

c) Hydrogen bonding

A hydrogen bond is formed between two neighbouring molecules when the molecules contain the element H and a **highly electronegative element such as nitrogen, oxygen or fluorine**. Hydrogen bonds exists in molecules as such as ammonia, water and hydrogen fluoride, as seen below. The hydrogen bond is the electrostatic force of attraction between the hydrogen atom (which is partially positive) of one molecule and a partially negative atom of a neighbouring molecule.

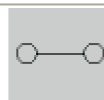
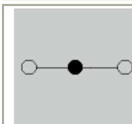
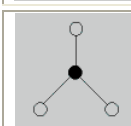
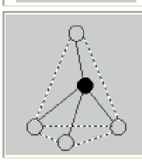
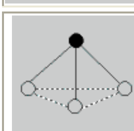
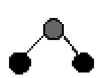


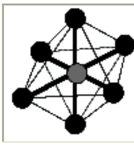
Hydrogen bonding occurs in ice as well. Ice is **less dense than water** because the molecules are further apart than in water resulting in a more open structure than liquid water. This is why ice of the same mass as liquid water occupies a **GREATER** volume.

Shapes and bond angles of simple molecules and ions

There is no direct relationship between the formula of a compound and the shape of its molecules. The shapes of these molecules can be predicted using a model developed about 30 years ago, known as the **valence-shell electron-pair repulsion (VSEPR) theory**. VSEPR theory assumes that each atom in a molecule will achieve a geometry that minimizes the repulsion between electrons in the valence shell of that atom. Just remember that the presence of lone pairs influences the geometry of the molecule, making the bond angle **SMALLER** than if lone pairs were **NOT** present.

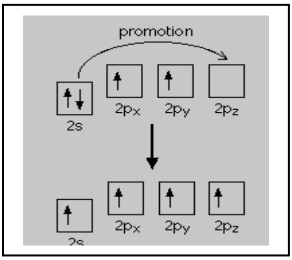
Study the table below

Total Number of electron pairs	***Arrangement of electron pairs***	***Number of bonding pairs of electrons***	***Number of lone pairs of electrons***	***Shape of Molecule***	***Name of Shape***	***Bond Angle***	
NA	Linear	1	NA		linear	180°	H_2, HCl
2	Linear	2	0		linear	180°	CO_2, HCN
3	trigonal planar	3	0		trigonal planar	120°	BCl_3, $AlCl_3$
4	Tetrahedral	4	0		tetrahedral	109.5°	CH_4, SiF_4
		3	1		trigonal pyramidal	107°	NH_3, PCl_3
		2	2		bent	104.5°	H_2O, SCl_2

5	trigonal bipyramidal	5	0		trigonal bipyramidal	120^o in the trigonal planar part of the molecule, 90^o for the others	PCl_5
6	octahedral	6	0		octahedral	90^o	SF_6

Hybridisation

Carbon is in group IV and we know that carbon can and does form 4 covalent bonds. However on examining its electronic configuration $1s^2 2s^2 2p^2$, there are only 2 unpaired electrons!! Unpaired electrons are needed to form covalent bonds. In order to have 4 unpaired electrons, a process called hybridization has to occur. **Why is the formation of 4 covalent bonds more desirable than 2? The more bonds formed, the more energy released and thus the more stable the compound.**

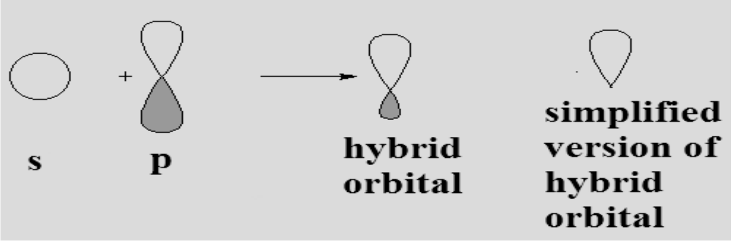


One of the 2s electrons is promoted to a vacant 2p orbital and now the E.C is $1s^2\ 2s^1\ 2p^3$. Now there are 4 unpaired electrons!

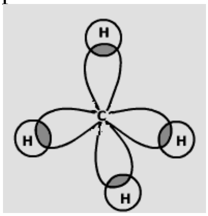
The # of sigma bonds attached from each carbon can be used to determine the type of hybridization present.

FOUR sigma bonds means the carbon atom is sp^3 hybridised ➔ this mean 4 hybrid orbitals arranged in a tetrahedral fashion. THREE sigma bonds means the carbon atom is sp^2 hybridised ➔ this means 3 hybrid orbitals arranged in a trigonal planar fashion. TWO sigma bonds means the carbon atom is sp hybridised ➔ this means 2 hybrid orbitals arranged in a linear fashion

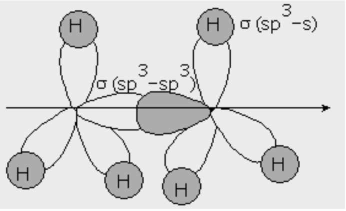
The formation of a hybrid orbital occurs when an s orbital and a p orbital mix.



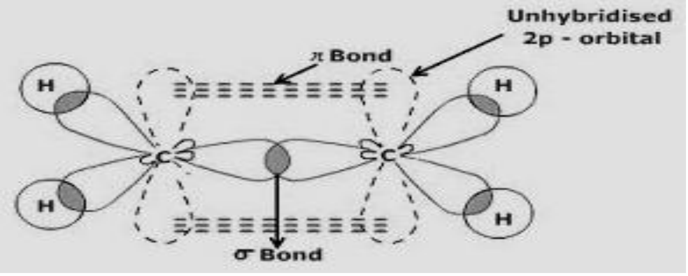
In **methane**, you now have to show to overlap of the 1s orbitals of the 4 hydrogen atoms. This is how the diagram would look.



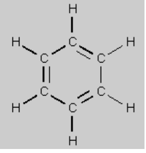
The bond angle in **ethane** would also be 109.5 ° since there are 4 single bonds attached to each carbon atom and thus the carbon atoms are sp^3 hybridised.



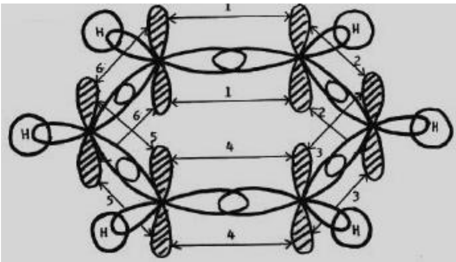
In ethene, carbon atoms have 3 sigma bonds attached to themselves ➔ thus carbon atoms are **sp^2 hybridised**. This means the bond angle is trigonal planar (120°). The **unhybridised** p orbital of each carbon atom then overlap laterally forming the **pi bond**. Remember the ONE pi bond counts as the overlap above **AND** below the plane of the molecule.



The molecule benzene (C_6H_6) can be thought of as



Each carbon has 3 sigma bonds attached which means **sp^2 hybridisation** again.



Each unhybridised p orbital of each carbon atoms would overlap with an adjacent p orbital forming a **pi system**. These electrons are delocalized electrons.

This π system results in the spreading of the negative charge around the molecule making it very stable. In fact, the reactions of aromatic compounds (compounds containing the benzene ring) occur so that the π system is not broken**. Please note the extra stability of benzene is called delocalization energy or resonance energy.**

<u>Relationship of lattice structure of crystalline solids and their physical properties</u>

Solid (type of solid and forces of attraction)	**Physical properties**
Iodine crystals (simple molecular solid, van der Waals forces)	Low melting and boiling point, sublimes, does not conduct electricity
Ice (simple molecular solid, hydrogen bonded)	Low melting and boiling point, although in comparison with other hydrides, it is much higher than expected because of the hydrogen bonds present. Does not conduct electricity
Silicon dioxide (giant molecular solid, strong covalent bonds between atoms)	Very high melting and boiling point, insoluble in water. Does not conduct electricity
Sodium chloride (ionic, strong ionic bonds)	High melting and boiling point, soluble in polar solvents. Conducts electricity when molten or in aqueous solution
Copper (metallic, strong metallic bonds)	High melting and boiling point, insoluble in solvents. Conducts electricity in solid or molten form
Graphite or diamond (giant atomic solid)	Extremely high melting and boiling point, insoluble in solvents. Diamond does not conduct electricity but graphite conducts electricity

Metals and graphite can conduct electricity in their solid state as they have mobile charge carriers.

(b) SiO_2 (quartz) melts at 1610°C whilst iodine crystals sublime on gentle heating. By considering the structure and bonding in these two compounds account for the difference in these physical properties. **[8 marks]**

(c) Correct the inaccuracy or inaccuracies in EACH of the following statements:

(i) Carbon dioxide and water are non-linear molecules.

(ii) Magnesium oxide has a giant molecular lattice.

(iii) Sulphur hexafluoride (SF_6) and phosphorus penta-chloride have the same number of electron pairs around the central atoms in each molecule.

(iv) Ice is denser than water because there is more extensive H-bonding in ice than in water. **[4 marks]**

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(b) Ammonia forms an addition compound with covalent beryllium chloride, $BeCl_2$.

(i) Suggest the formula of the compound formed.

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[1 mark]

(ii) Draw a dot-cross diagram to show the electron distribution in the addition compound.

(iii) Draw a diagram to illustrate the shape of the addition compound.

[1 mark]

(b) Based on the m.p. data obtained, the student concludes that A is a covalent compound. Another student, who has been given a sample of B, decides that B is not a covalent compound. Further analysis of A and B reveals that A is soluble in tetrachloro-methane but that B is not, and that neither A nor B dissolves in water.

(i) Suggest the type of forces of attraction that exist between particles of A and describe how they are formed.

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[3 marks]

(ii) Name and describe the forces of attraction present in B.

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[3 marks]