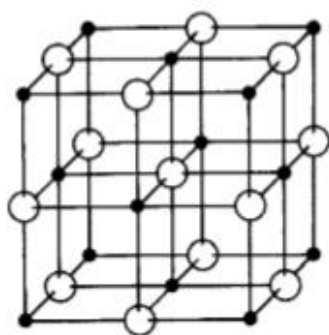
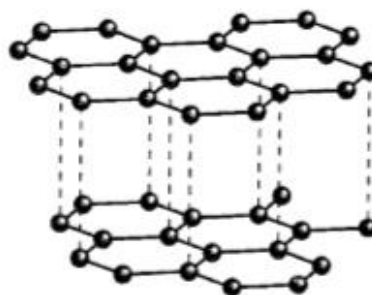
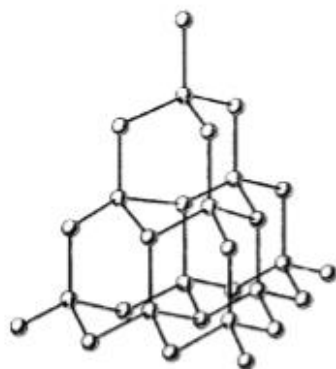


Chemical Bonding



● = Na⁺
○ = Cl⁻



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01. What is a Chemical bond?

A chemical bond is an electrostatic attractive force between positively and negatively charged species, which holds atoms or ions together.

02. Why do atoms form bonds?

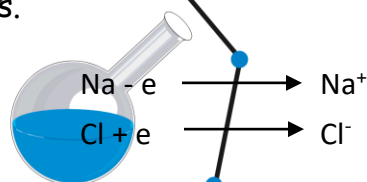
To acquire a stable electronic configuration by losing, gaining or by sharing electrons.

Ionic Bonding

03. Describe how ions are formed by the gain or loss of electrons?

Metals tend to lose electrons; hence form cations, non-metals will gain electrons to form anions.

e.g :



04. A. What is an ionic or electrovalent bond?

An ionic bond is an electrostatic attraction between positively charged and negatively charged ions.

B. Explain the term "electro static attraction"

Forces attracting oppositely charged ions

05. Explain how an ionic bond is formed?

Ionic bond is formed by the complete transfer of electrons from one atom to another resulting the formation of charged particles (cation & anion)

06. What groups of compounds are involved in the formation of an ionic bond?

Metals of group I & II reacting with non metals of group VI & VII & will result in ionic compounds. Exceptional: BeF₂, BeCl₂, AlCl₃ - covalent

07. Give some common examples for ionic compounds

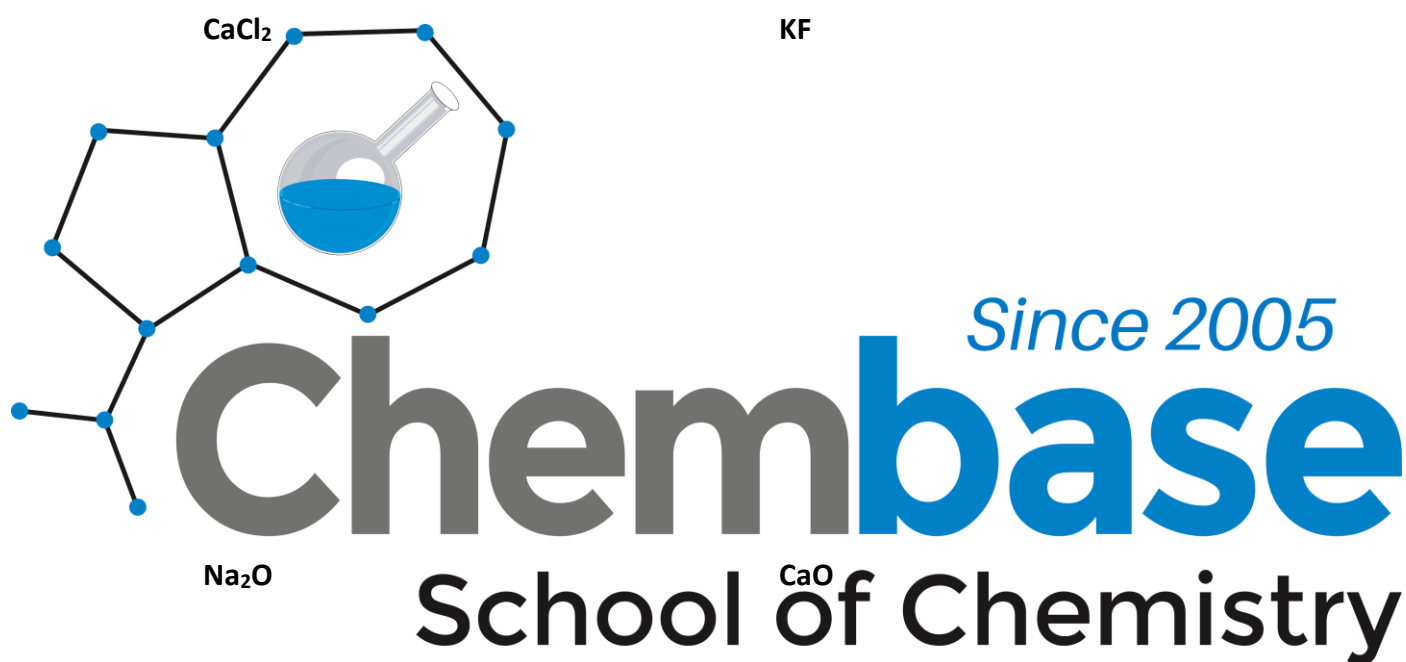
All salts are ionic, metal oxides and hydroxides are ionic.

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08. Draw dot & cross diagrams for the ionic compounds below.



CaCl₂

KF

Na₂O

CaO

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10. What are the conditions that favor the formation of an ionic compound?

- Both cation & anions should be present in ionic compound
- The net charge of the compound should be zero.

11. Describe how a **crystal lattice** is formed?

In a crystal lattice, each cation surrounds many anions around it. Likewise each anion will be surrounded by many cations in order to preserve the electrical neutrality. This type of arrangement of ions produces a giant assembly of ions held in a rigid crystal lattice.

12. Draw the lattice structure of NaCl

13. What are the main characteristics of the NaCl lattice ?

- Each Na^+ ion accommodates 6 Cl^- ion around it
- Similarly, each Cl^- ion is surrounded by 6 Na^+ ions to preserve electrical neutrality.
- The ions are packed closely, held by many electrostatic attractive forces.

14. State 3 physical properties of ionic compounds

- High melting and boiling points
- Ionic compounds are soluble in water
- They conduct electricity in aqueous or in molten state due to the presence of mobile ions.

15. Explain why ionic compounds does not conduct electricity when solid but conducts in molten and aqueous state?

The ions in an ionic compound in solid state are in fixed positions held together by strong electrostatic forces of attractions therefore unable to move thus does not conduct electricity. Whereas ions in aqueous or molten state are mobile and able to conduct.

16. Explain why ionic compounds possess high melting points?

There are many electrostatic attractions around each ion in an ionic lattice, therefore lot of energy is required to separate them.

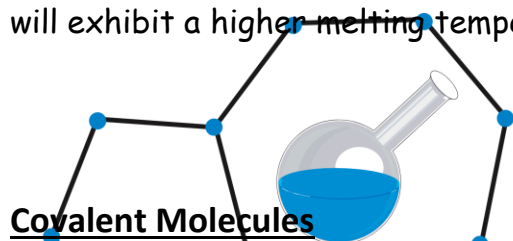
17. Explain the effect of ionic radius and ionic charge on the strength of ionic bonding.

Cations and anions with smaller ionic radii will possess greater attraction between them. As the sum of the ionic radii is smaller the cations and anions will be strongly attracted to each other resulting in stronger ionic bond.

Cations and anions with greater charges will attract each other strongly resulting stronger ionic bond.

18. Explain why MgO has a much higher melting point & boiling point than NaCl thus list Li_2O , CaO , Al_2O_3 & KF in the increasing order of boiling and melting points.

NaCl contains attraction between Na^+ & Cl^- ions, whereas in MgO the attraction is between Mg^{+2} and O^{-2} ions, the attraction between ions in MgO is greater than NaCl . Therefore MgO will exhibit a higher melting temperature.



01. Why do atoms tend to form covalent rather than ionic bonds?

To remove the outermost electrons of group III or group IV elements requires a very high energy and this process is extremely unlikely. Instead these atoms attain the noble gas configuration by sharing electrons as this requires much less energy.

02. What is a covalent bond?

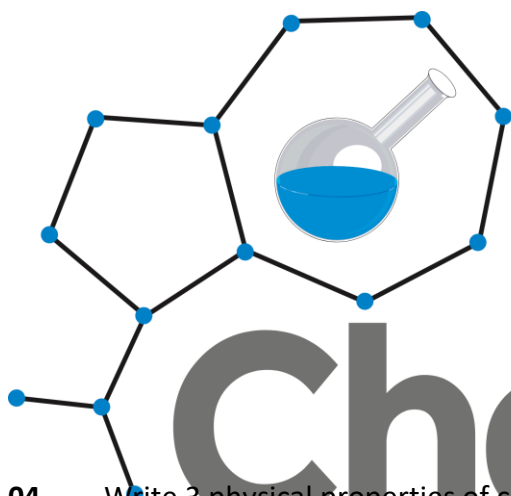
A covalent bond is an electrostatic attraction between the bonding electrons and the nuclei of the atoms involved in the bonding.

03. Draw the structure of the covalent compounds below, showing all outermost electrons in dots and crosses.



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04. Write 3 physical properties of covalent compounds.

- Low melting and boiling points due to weak intermolecular attraction between the molecules.
- They are insoluble in water (except for acids) but soluble in organic solvents (e.g.: ethanol, CCl_4)
- They do not conduct electricity in aqueous or molten state.

05. Explain why covalent compounds show low boiling & melting points

Covalent compounds possess weak intermolecular attractions between the molecules. Therefore; less energy is required to separate them.

06. Draw all the different possible atomic overlappings, which takes place during the formation of covalent bonds.

07. Explain how sigma & pi bonds are formed as a result of orbital overlap.

Sigma Bond (σ) :

Sigma bond is formed as a result of linear overlapping (head to head overlap) of 2 S/P sub orbitals. This type of bond is the strongest as the bonding pair of electrons lie closer to the nucleus, thus strongly attracted.

Pi Bond (π) :

Pi bond is formed as a result of lateral overlapping (side by side overlapping) of two 'p' orbitals. This type of bond is weaker compared to sigma bond as the shared electron pair is much away from the nucleus hence attracted weakly.

(Mark the pi & sigma bonds in the above atomic overlapping diagram)

08. Draw the valence shell orbitals to show the formation of the covalent bonds of the below molecules.





09. Explain the concept of orbital hybridization hence describe how sp^3 , sp^2 & sp hybridization occur taking methane, ethene & ethyne as examples.

▪ **SP^3 Hybridization:**

The four C-H bonds in CH_4 are identical, each has the same bond length & bond energy. But 'C' uses 's' & 'p' orbitals to form the four C-H bonds.

Therefore the bonds formed with the s-orbital would differ from the 3 bonds formed with the p-orbitals (energy). The formation of identical bonds are explained by the concept of orbital hybridization.

Hybrid orbitals are mixed orbitals. When one 's' orbital & three 'p' orbitals are mixed together, they form four sp^3 hybrid orbitals with equal energy.

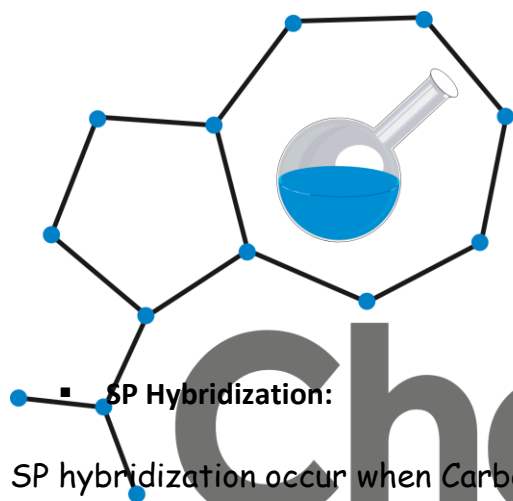
Each sp^3 orbital has 25% 's' character & 75% 'p' character. (The superscript 3 means that three p-orbitals are mixed with one s-orbital) The four sp^3 hybrid orbitals are degenerate (has equal energy)

The four sp^3 orbitals are arranged in tetrahedral shape in order to minimize the repulsions. Each C-H bond is formed by the overlap of the S-orbital of 'H' with a sp^3 orbital of 'C'

▪ **SP² Hybridization:**

SP² hybridization occurs when Carbon forms a C=C double bond. Two p-orbitals combine with one s-orbital to form 3 SP² hybrid orbitals with equal energy. And there will be one unhybridized p-orbital.

The SP² hybrid orbitals will overlap linearly to form 3 sigma bonds (2 with 'H' & one with 'C' in the case of C₂H₄) the unhybridized 'p' orbital will overlap laterally with another p-orbital of a 'C' to form a pi bond. The SP² hybrid orbitals lie 120° apart in order to minimize repulsions and the unhybridized p-orbital lies 90° to the sp² plane as shown below:



SP hybridization occurs when Carbon forms a triple bond. Single p-orbitals combine with one s-orbital to form 2 SP hybrid orbitals with equal energy. And there will be two unhybridized p-orbitals.

The SP hybrid orbitals will overlap linearly to form 2 sigma bonds (1 with 'H' & one with 'C' in the case of C₂H₂) the two unhybridized 'p' orbitals will overlap laterally with another p-orbital of a 'C' to form two pi bonds. The SP hybrid orbitals lie 180° apart in order to minimize repulsions and the unhybridized p-orbital lies 90° to the sp plane as shown below:

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GIANT COVALENT MOLECULES

01. What are "Allotropes"?

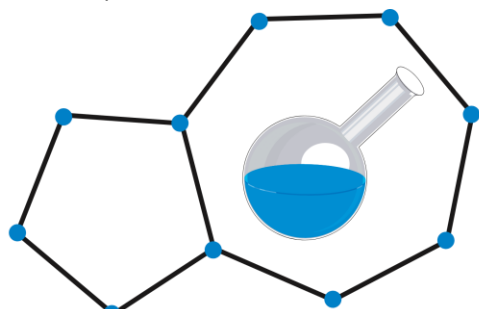
Allotropes are different structural forms of the same element existing in the same physical state.

02. State the natural and synthetic allotropes of Carbon.

Natural Allotropes : Diamond and Graphite.

Synthetic Allotropes : Bucky ball and Graphene

03. Draw the structure of Diamond (you should be able to draw at least 5 carbon atoms arranged tetrahedrally. The carbon atoms on the outside should indicate the tetra valency of 'C' by unattached bonds)



04. State the structural characteristics of Diamond.

- Each C atom tetrahedrally linked to 4 other C atoms by single co-valent bonds
- It is a continuous giant structure.

05. State 3 physical properties of Diamond.

- High melting and boiling points. When heated does not melt but sublimates at very high temperature.
- Does not conduct electricity, as there are no ions or free electrons to carry charge.
- Extremely strong and rigid. Therefore diamond is used as a cutting tool.

06. Why do allotropes of Carbon possess high sublimation points?

Atoms in diamond and graphite are held stronger by strong co-valent bonds in a giant molecular crystal.

07. Why does diamond not conduct electricity?

Because all electrons of 'C' in diamond are co-valently bonded. Therefore there are no free electrons to carry charge.

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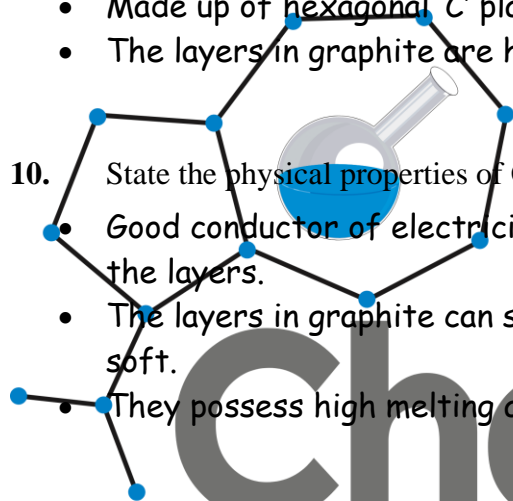
08. Draw the structure of Graphite. (Showing at least 2 layers & 3 hexagons in each layer)

09. State the structural characteristics of Graphite.

- Each 'C' atom is covalently bonded to 3 other 'C' atoms
- Made up of hexagonal 'C' plates or sheets of 'C' atoms.
- The layers in graphite are held together weak vander waal forces.

10. State the physical properties of Graphite

- Good conductor of electricity due to the presence of delocalized electrons between the layers.
- The layers in graphite can slide over each other. Therefore graphite is slippery and soft.
- They possess high melting and boiling points.



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Dative covalent bonds (Co-ordinate bonds)

1. Define a dative covalent bond.

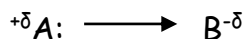
A dative covalent bond is formed between 2 atoms by sharing of 2 electrons between them both electrons are donated entirely by 1 atom.

2. Describe how a dative or coordinate bond is formed.

One atom donates (the donor atom) an electron pair while the other atom accepts the electron pair (the acceptor atom). The acceptor atom should possess an empty orbital to accept the electron pair. A co-ordinate bond is represented as $A \rightarrow B$, where A is the donor atom and B is the acceptor atom.

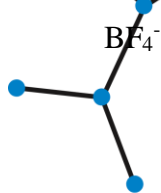
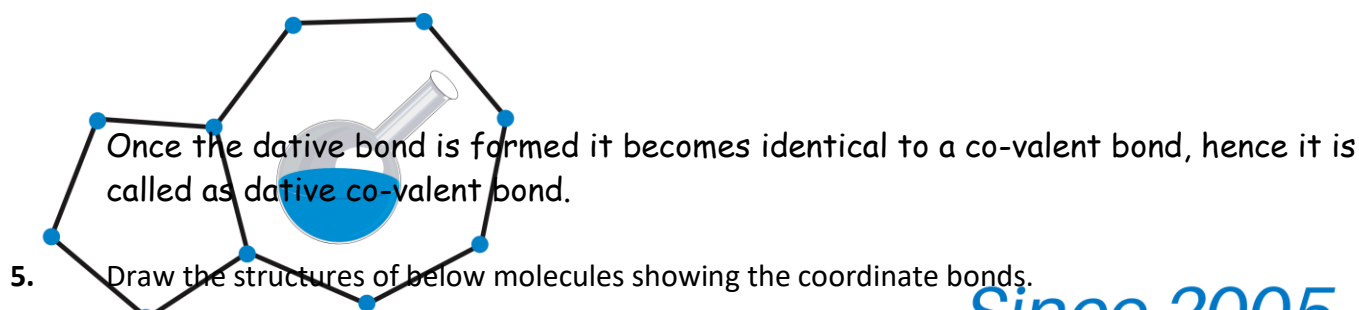
3. Explain why dative bonds are called as semi polar bonds.

In the formation of the co-ordinate bond, the partial positive charge is developed on the donor atom and an equal negative charge is developed on the acceptor atom.



Due to this co-ordinate bonds become slightly polar, hence called as semi polar bonds.

4. Show the dative bonds between BF_3 & NH_3 , name the donor & acceptor atom, hence comment why this bond is called dative **covalent** bond?



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6. Describe the co-ordinate bonding formation in Al_2Cl_6 & draw a dot & cross diagram to represent the bonding

7. Draw a dot & cross diagram to show the bonding when methyl amine CH_3NH_2 is reacted with a hydrogen ion.

Metallic Bonds

1. Explain what a metal is.

A metal is a giant structure where electrons are free to move throughout the whole structure.

2. Describe a "metallic bond"

A metallic bond is the electrostatic attraction between the metal cations and the delocalized sea of electrons.

3. Describe how a metallic bond is formed

The metal atoms lose their valence electrons and form cations. The electrons which have been lost, will act as the sea of delocalized electrons between the cations.

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4. What does the strength of a metallic bond depend on?

1. The number of delocalized electrons greater the number, stronger the bond.
2. Size of the cation, smaller the size greater the strength

5. Explain why across the period the strength and melting point of metals increase?

Across the period from left to right the strength of the metallic bond increases due to the decrease in the size of the metal cations and increase number of delocalized electrons. As the size of the metal cation decrease the attraction towards the electrons increase resulting in stronger metal.

6. The physical properties of metals are not only related to the strength of the metallic bond, but also the structure of the metallic lattice. State the properties of metals.

- Solids at room temperature (except for mercury Hg, Cs, Fr which are liquids)
- High melting and boiling points.
- Conducts electricity and heat
- Has a shiny, reflecting surface.
- Metals are malleable (can be stretched into sheets)
- Metals are ductile (can be stretched into wires)

7. A. Explain why metals conduct electricity?

Due to the presence of mobile sea of free delocalized electrons, metals carry electricity.

B. Name the non-metal, which conducts electricity?

Graphite.

8. Why are metals good conductors of heat?

The mobile sea of electrons takes in heat and they convert it into kinetic energy, as a result the electrons vibrate and move faster and transfer the heat energy.

9. Describe why metals are malleable and ductile?

When force is applied the metal cations slide over one another without breaking the metallic bond as a result metals are malleable and ductile.

10. Explain why across the period the strength and melting point of metals increase?

Across the period from left to right the strength of the metallic bond increases due to the decrease in the size of the metal cations and increase number of delocalized electrons. As the size of the metal cation decrease the attraction towards the electrons increase resulting in stronger metal.

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Polarization.

1. Define the term 'Electronegativity'

Electronegativity is the power of an atom in a covalent molecule to attract the bonding pair of electrons to itself.

2. State the factors affecting the magnitude of electronegativity

- Size of the atom - Smaller atoms have higher electronegativity than larger atoms.
- Nuclear attraction - Higher the nuclear attraction, greater the electronegativity.

3. Explain the variation of electronegativity values across a period & down a group

- Down the group the electronegativity values of elements decreases due to increase in atomic radius.
- Electronegativity values increase across the period from left to right with increase in atomic number due to the decrease in atomic radius.

4. Account why increase electronegativity difference between atoms in a molecule favours ionic character.

In an ionic compound, the nonmetal takes the electron/electrons from the metal to achieve the octet. Therefore in the formation of the ionic bond, the electron is taken in full control by one atom. This is possible only when the difference of electronegativity is very large. Therefore, the greater the electronegativity difference, the stronger the ionic bond.

5. Explain in terms of electronegativity difference the ionic & covalent bond

When the difference in electronegativity is large (more than 1.6 as to the Pauling scale) it results in an ionic bond. When the difference of electronegativity is less, the electrons will be more concentrated between the atoms, resulting in a polar covalent bond. (Refer the last page for detailed clarification)

The relationship between the difference in electronegativity and the percentage ionic character is shown below.

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6. Define the terms below:-

➤ Polarization : The overall distortion of charges in molecules which results in unequal sharing of electrons.

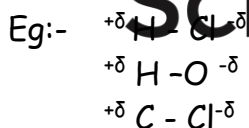
➤ Polar : The molecules where the polarization takes place.

➤ Dipole : The separation of charge in a molecule as a result of difference in electronegativity. Direction of the dipole is shown by the sign + → the arrow points to the partially negatively charged end of the dipole.

➤ Dipole moment: The measure of the degree of polarity of a molecule.

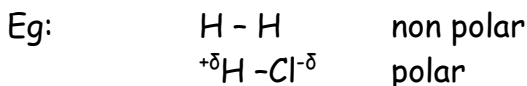
7. Explain how a polar bond results in terms of electronegativity.

When covalently bonded atoms have different electronegativity, a small positive charge (+ δ) will be developed on the lower electronegative atom, and a small negative charge (- δ) will be developed on the higher electronegative atom. When both + δ and - δ charges are present, the bond is said to be polar. The bonding electrons are more drawn towards the higher electronegative element, resulting in the other atom to be + δ .



8. Explain how electronegativity difference leads to polarity of bonds in a covalent molecule

When a covalent bond is formed between 2 atoms of the same electronegativity, the pair of electrons in the molecular orbital is equally attracted to both the nuclei, resulting in no charge separation. Such molecules are said to be non polar. If the atoms in the covalent bond are different elements, then the difference in electronegativity will result in charge separation. The higher electronegative atom will attract the electron more than the other, resulting in a polarized covalent bond.



9. State and explain the factors affecting the magnitude of polarizing power of a cation.

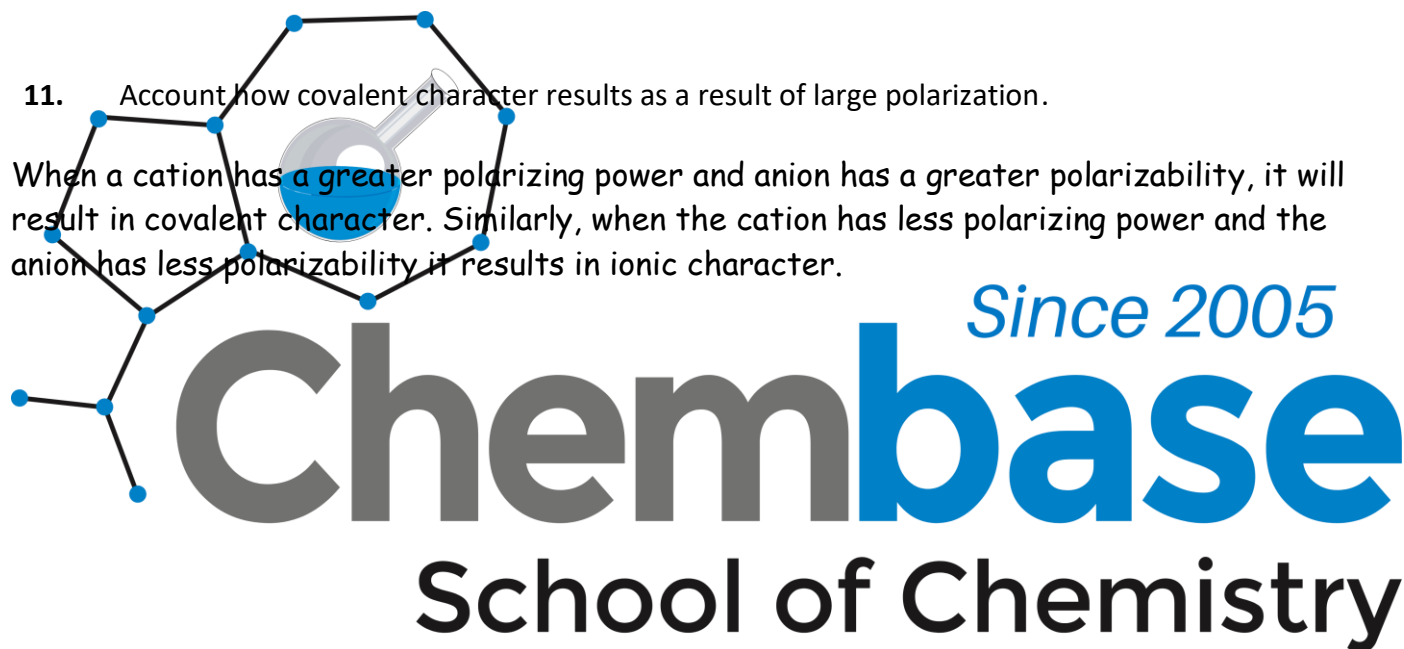
The polarizing power is the ability of a cation to polarize (cause charge separation) an anion. It depends on the charge & the size. Polarizing power is directly proportional to the charge and it is inversely proportional to the size. The charge/size ratio is referred to as charge density (also called as polarizing power). When the cation has a greater charge & smaller size, it will exhibit a greater charge density hence greater polarizing power.

10. Explain how the charge and the ionic radius affect the polarizability of an anion.

The polarizability is the ability of an anion to get polarized by a cation. The polarisability of an anion depends on the charge and the size. Anions with greater charge and greater size will exhibit greater polarizability.

11. Account how covalent character results as a result of large polarization.

When a cation has a greater polarizing power and anion has a greater polarizability, it will result in covalent character. Similarly, when the cation has less polarizing power and the anion has less polarizability it results in ionic character.



12. Account for the variation in melting points of below compounds based on their polarization.

NaCl, MgCl₂, AlCl₃

13. Account for the below observations based on their polarization

- i. Al_2O_3 is more ionic in contrast to AlCl_3
- ii. LiCl is insoluble in water whereas NaCl is soluble
- iii. Sulphides are less soluble in water than the oxides of the same metal

14. Account for below decomposition temperatures?



BeCO_3	100 °C	SrCO_3	1280 °C
MgCO_3	400 °C	BaCO_3	1360 °C
CaCO_3	900 °C		

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15. Explain how symmetry & asymmetry of molecules give rise to non polar & polar molecules.

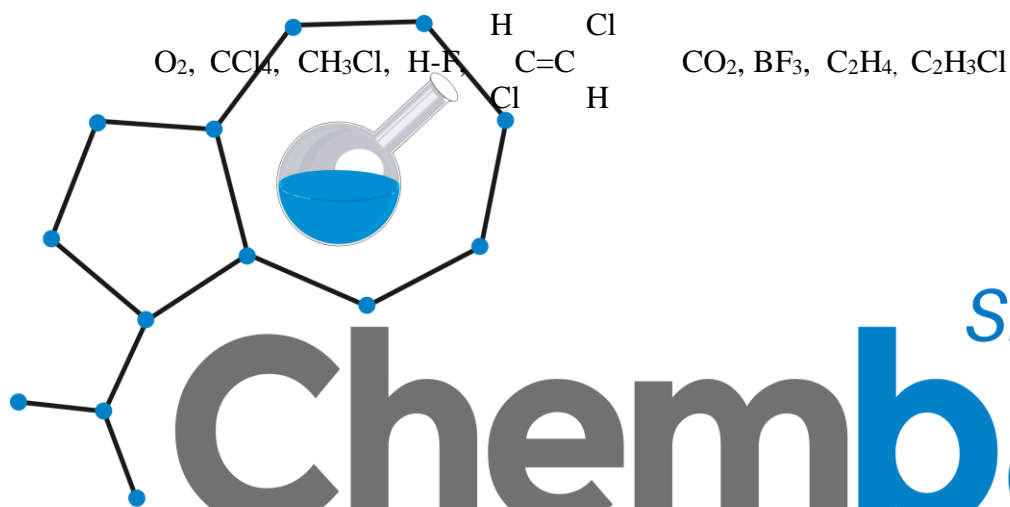
In symmetrical molecules, the center of positive charge exactly coincides with its center of negative charge, resulting in a no net dipole. Such molecules where equal dipoles cancel each other are said to be non polar. Therefore molecules with center of symmetry will not show permanent dipoles.

Eg:-

In asymmetrical molecules, the center of positive charge does not exactly coincide with its center of negative charge, resulting in a dipole. Such molecules are said to be polar.

Eg:-

16. State which of the below molecules are polar & non-polar



CO_2 , BF_3 , C_2H_4 , C_2H_3Cl

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17. Describe an experiment to determine the effect of electrostatic force on jets of liquid (water, ethanol & cyclohexane) hence use the results to determine the polarity of the molecules.

Intermolecular Forces.

(01) A. What is electrostatic attraction between opposite poles of molecules, which gives ionic character to co-valent compounds also responsible for phase change referred to as?

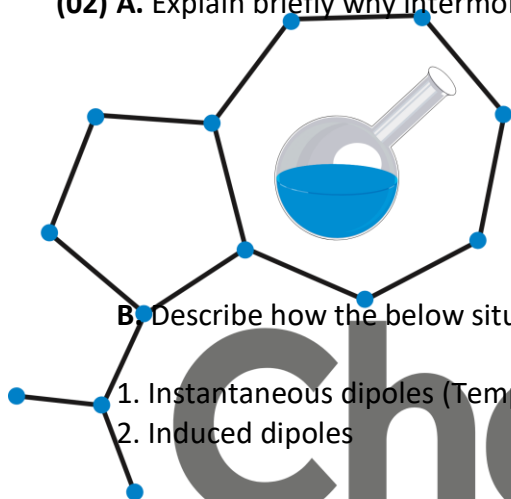
B. What are the main two categories of intermolecular forces?

C. What are the two types of **van der waals forces** present?

Permanent dipole - dipole forces

Temporary dipole-induced dipole forces

(02) A. Explain briefly why intermolecular forces are weak compared to intra molecular forces?



B. Describe how the below situations occur in “He” (helium) with the aid of diagrams

1. Instantaneous dipoles (Temporary dipoles)
2. Induced dipoles

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(03) State two alternative names, which are used for temporary dipole-induced dipole attraction?

(04) Explain how symmetrical, non-polar molecules show temporary fluctuating dipoles

The electrons found in the orbitals within the molecules are never stationary, but they oscillate around the nucleus. Also, the molecule itself is spinning and rotating. This random motion causes the electrons to be more concentrated in one direction than the other, resulting in a temporary fluctuating dipole. This is due to their positive and negative centers not coinciding at any moment of time. This molecule with temporary dipole will induce a charge separation on another non polar molecule, hence forming an intermolecular attraction.

(05) Describe the change in physical state of halogens from gases to liquids to solids down the group.

The total number of electrons within the molecule increases down the group (eg: Cl_2 - 34 electrons, Br_2 - 70 electrons). This will result in a larger $+\delta$ and $-\delta$ charge separation down the group. Therefore the strength of the dispersion forces between the molecules increases from Cl_2 to Br_2 to I_2 . The increase in strength of the intermolecular forces account for the change in the physical state from gas to liquid to solid.

(06) Explain why permanent dipole-dipole attractions are not necessarily stronger than the induced dipole- induced dipole attractions?

The strength of the permanent dipole attraction depends on the charge separation ($+\delta$, $-\delta$), which is the result of the electronegativity difference. Greater the difference in electronegativity greater will be the charge separation thus the strength of permanent dipole attraction.

The strength of the dispersion forces depends mainly upon the number of electrons. Molecule may have stronger dispersion force due to its greater number of electrons and weaker permanent dipole attraction due to lesser electronegativity difference.
Eg: H-I

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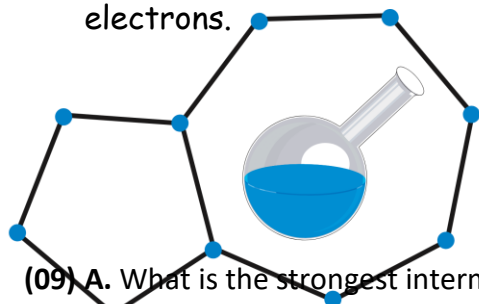
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(07) Using the knowledge of above forces, explain why the boiling & melting points increase in inert gases down the group?

Larger atoms can form instantaneous dipoles more easily than smaller atoms due to the higher number of electrons. The higher the number of electrons, the stronger the dispersion force. As a result the strength of the intermolecular forces increases down the inert gas group.

(08) Comment on the nature of Intermolecular forces in polar molecules

Polar molecules are molecules with permanent dipoles. The $+\delta$ atom of one molecule may attract the $-\delta$ atom of another molecule, resulting in an intermolecular attraction. They also do possess dispersion / London forces due to the presence of electrons.



(09) A. What is the strongest intermolecular force?

B. Define the term "Hydrogen Bond"?

Electrostatic attraction between a hydrogen atom of one molecule and highly electronegative atoms such as Nitrogen, Oxygen and Fluorine of another molecule. The Hydrogen itself, should be covalently bonded to one of the above electronegative elements. (N,O,F)

Eg:- HF, H₂O, NH₃

C. Account for the increase of the bond dissociation energy in the compounds below in their respective order as below; **NH₃ , H₂O , HF**

Greater electronegativity difference results in stronger bonds. Electronegativity increases in the order of N-H < O-H < H-F. Therefore the bond dissociation energy will increase in the order of NH₃ < H₂O < HF

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D. With the help of diagrams show the intermolecular & intra molecular formation of Hydrogen bonds in the above compounds.

□ **'H' bonding in HF:-**

The $+\delta$ H in one molecule of HF forms hydrogen bonds with $-\delta$ F atoms in the adjacent molecule.

□ **'H' bonding in H₂O:-**

All compounds which contains OH groups forms hydrogen bonds. In water the $+\delta$ 'H' atom form 'H' bonds with $-\delta$ 'O' atom of another molecule. There are 4 'H' bonds around a water molecule. This accounts for the unusual high boiling point of water (100 °C) even though the relative molecular mass of water is 18.

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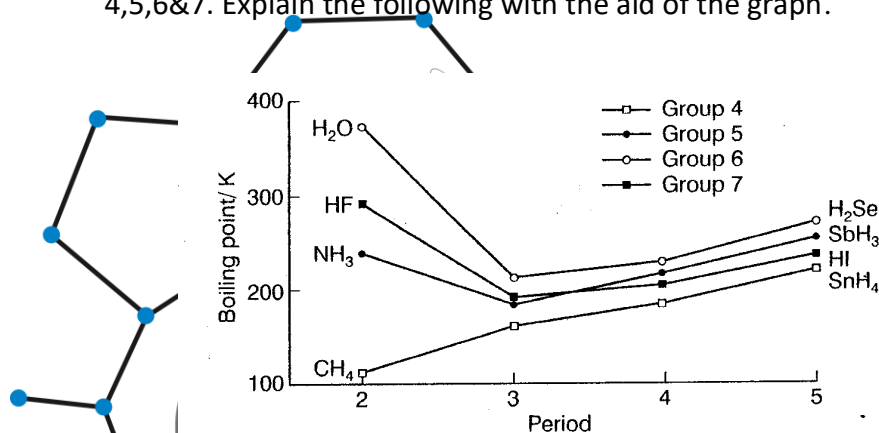
□ **'H' bonding in NH₃:-**

The lone pair in 'N' atom enables the formation of 'H' bonds with a $+\delta$ H of another NH₃ molecule. All nitrogen compounds with N-H group form Hydrogen bonds.

(10) Explain the factors which effects the melting point & boiling point of substances

The boiling and melting point depends on the strength of the vanderwaal forces between the molecules, not the strength of the covalent bond within the molecule, except in giant atomic solids or ionic compounds. When a giant atomic solid such as diamond or graphite is heated, the covalent bonds between the atoms are broken. But the covalent bonds are not broken when a molecular substance is boiled or melted. Therefore the strength of the covalent bonds within the molecules does not influence the boiling and melting points of a molecular substance.

(11) Below is a graph showing the variation in the Boiling points of the Group 4,5,6&7. Explain the following with the aid of the graph.



A. Water shows high boiling point than "HF"

Although the given hydrogen bond is stronger between H and F than H and O, there are more hydrogen bonds per water molecule than in HF, as a result lot of energy is required to separate the molecules hence exhibit greater boiling points.

B. Boiling point of NH₃ is lower than "HF" & higher than CH₄

Both NH₃ and HF contain the same number of hydrogen bonds around a molecule. The hydrogen bond between N and H are weaker than the hydrogen bond between H and F due to lesser electronegativity difference.

C. Account for the variation of Boiling points of the periods 3,4&5

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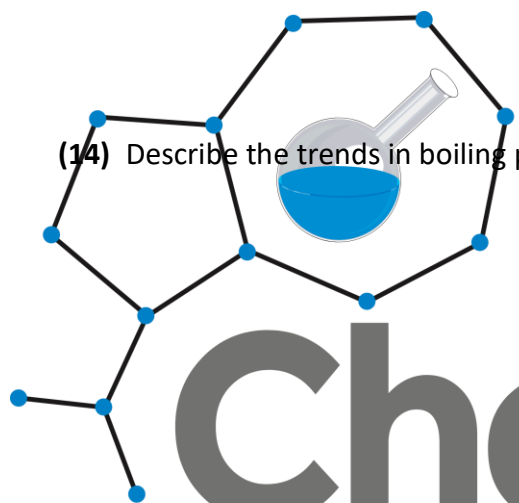
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(12) State why water (RMM 18) is a liquid while H_2S (RMM 34) is a gas even the molecular weight of H_2S is higher than water?

(13) Explain why alcohols have a low volatility & higher boiling point compared to alkanes with a similar no. of electrons.

(14) Describe the trends in boiling points of molecules from HF to HI



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(15) Explain why the melting point of Sulfur is higher than that of Phosphorous & why silicon has a much greater melting point than both 'S' & 'P'

Both sulfur and phosphorous are simple molecular structures with dispersion/London forces between the molecules. Sulfur exists as S_8 molecules, which has more electrons than phosphorous which exists as P_4 . This accounts for stronger dispersion forces between sulfur molecules. Therefore more energy is needed to separate them. Thus the melting point of sulfur is higher than that of phosphorous.

Silicon is a giant diamond like atomic structure with 4 strong covalent bonds on each atom, arranged in a tetrahedral shape. The energy required to overcome the strong covalent bonds between silicon atoms are much greater. Therefore silicon has a greater boiling point than both sulfur and phosphorous.

Shapes of molecules

(01) A. What is a **Lewis Structure**?

A Lewis structure is a representation of covalent bonding in a covalent molecule, using Lewis dot symbols, in which shared electron pairs are shown either as lines or pairs of dots, or dots and crosses.

B. Why only **valence electrons** are shown in a Lewis dot symbol?

Because only the valence electrons take part in a chemical bonding.

C. When drawing a Lewis structure for a complex polyatomic molecules & ions how to choose the central atom?

The central atom is the least electronegative atom.

(02) A. Write the strategy for drawing the Lewis structure (6 main points)

a. Draw the skeleton structure of the molecule by joining the atoms by a single covalent bond. The least electronegative atom is the central atom and H always occupies the terminal position.

b. Establish the total number of valence electrons in the structure. For polyatomic ions, add an extra electron for each negative charge and subtract an electron for each positive charge.

c. For each bond subtract 2 electrons from the total number of valence electrons. Thus convert this number to electron pairs by dividing the total number of electrons by 2.

d. Distribute the electron pairs first around the terminal atoms in such that they acquire the noble gas configuration and then the rest to the central atom.

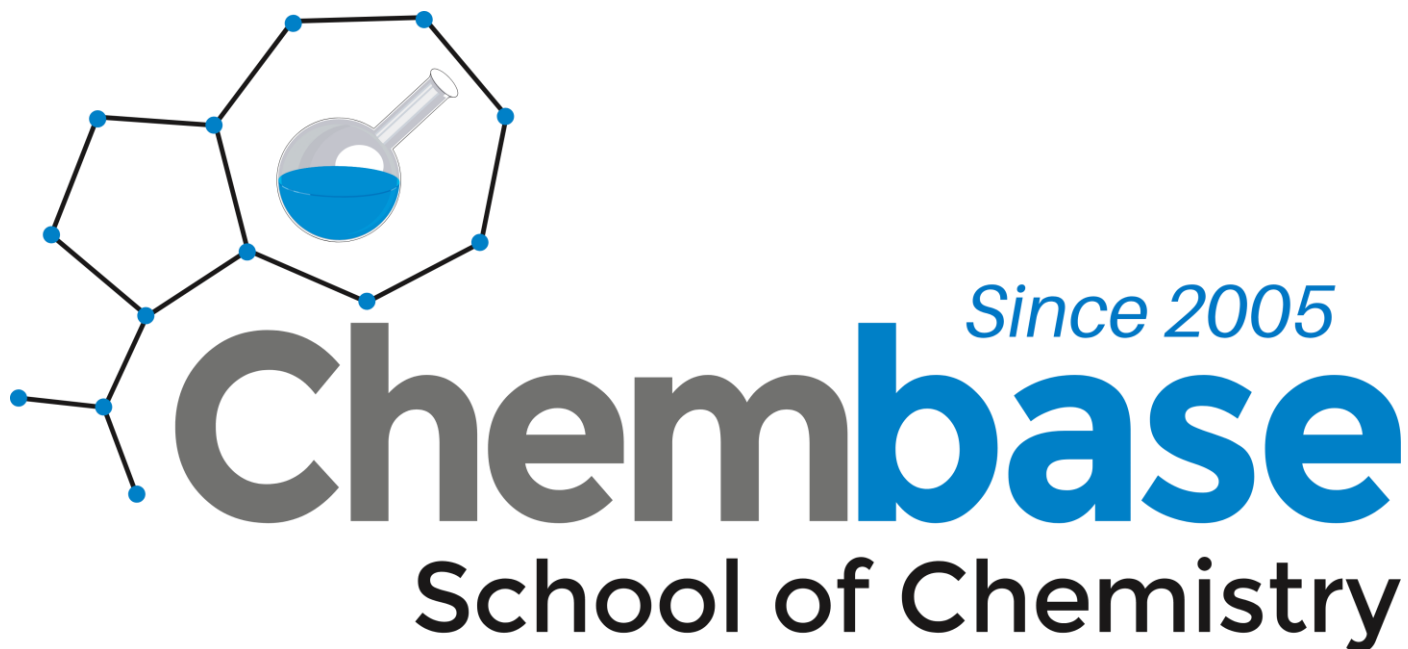
e. If the central atom lacks its valence electrons, then convert single bonds to double bonds by shifting lone pairs from the terminal atoms.

f. Assign the formal charges on the ion. The most possible Lewis structure is the one with the least amount of formal charges.

C. Draw the Lewis dot symbols for below molecules & ions.

HNO_2 , NOCl , H_2SO_4 , PCl_5 , SF_6 , BrF_5 , AlCl_3 ,

NH_4^+ , NO_2^- , SO_3^{2-} , PCl_4^+ , ICl_2^- , ClO_4^- , ClO_2^- , CO_3^{2-}





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(03) A. State the Rules for applying the Valence Shell Electron Pair Repulsion Theory (VSEPR) to predict molecular shapes.

- Draw the possible Lewis structure for the species.
- Determine the total number of lone pairs and sigma bonds around the central atom. These lone pairs and sigma bonds are called VSEPR electron pairs which determines the shape of the molecule.
- Establish the geometrical orientation of the VSEPR electron pairs around the central atom. If the central atom contains:

2 VSEPR electron pairs - linear

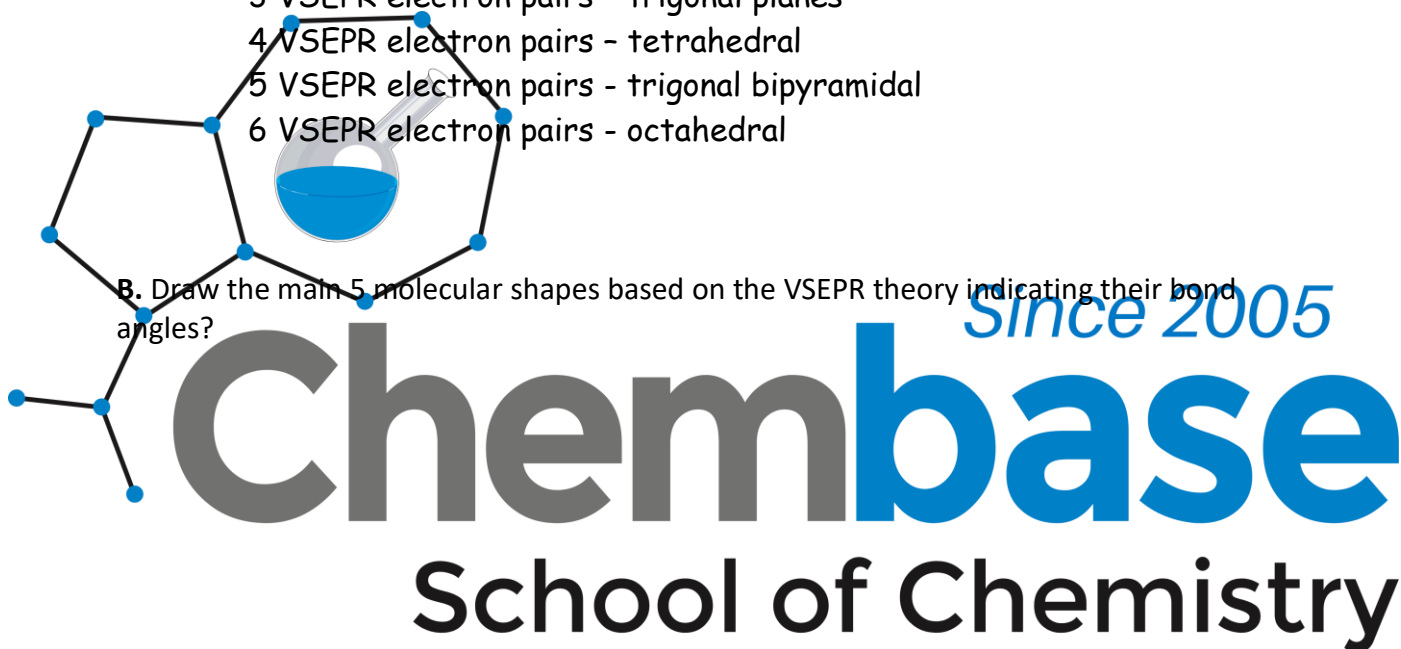
3 VSEPR electron pairs - trigonal planes

4 VSEPR electron pairs - tetrahedral

5 VSEPR electron pairs - trigonal bipyramidal

6 VSEPR electron pairs - octahedral

B. Draw the main 5 molecular shapes based on the VSEPR theory indicating their bond angles?



C. Explain why lone pairs always occupy equatorial plane & not axial position in Trigonal bipyramidal structure?

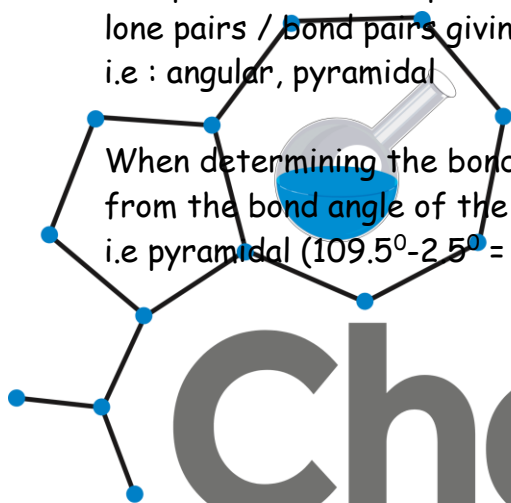
Axial position is unstable for the lone pair due to much lone pair - bond pair repulsion. But the equatorial position is stable as there is only 2 lone pair - bond pair repulsion.

(04) A. What is molecular distortion?

The presence of lone pair/s on the central atom causes repulsion between the lone pairs / bond pairs giving rise to a distorted shape from the basic shape.
i.e : angular, pyramidal

When determining the bond angle of a distorted shape deduct 2.5° per lone pair from the bond angle of the basic shape.

i.e pyramidal ($109.5^\circ - 2.5^\circ = 107^\circ$), Angular ($109.5^\circ - 5^\circ = 104.5^\circ$)



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B. State the repulsion of electron pairs in ascending order in terms of lone pair & bond pairs?

C. Considering the molecular distortion determine the shapes of the following molecules.

BeCl_2 , BCl_3 , NH_3 , H_2O , CO_2 , PCl_5 , SF_6 , SO_2 , PF_6 , NO_2 ,

SO_3^{2-} , CO_3^{2-} , NH_4^+ , ClF_3 , BrF_5 , ClO_3^- ,

