

- Assumptions
(need confirmation)
- Raj bhaiya imp notes
- Stuff I added
(confirmed)

Chemical Bonding

The basic difference between an ionic and covalent bond is that, the distribution of e^- in a covalent bond is symmetrical but in ionic bond, it is assymetrical.

Ionic bond is the result of strong electrostatic forces of attraction between ~~one~~ positively charged metal ion and negatively charged non-metal ion.

Covalent bond is the result of ~~weak~~ weak electrostatic force of attraction between the nuclei of the bonded atoms and the ~~for~~ bonding pair of e^- .

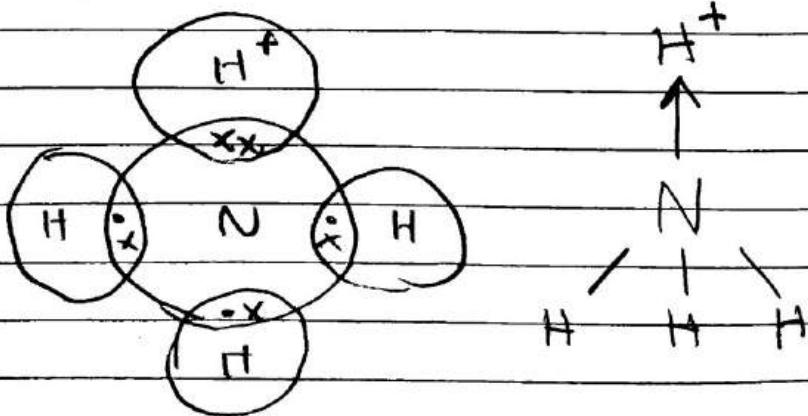
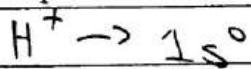
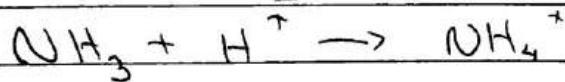
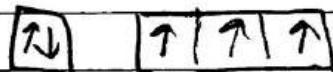
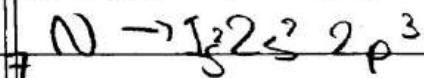
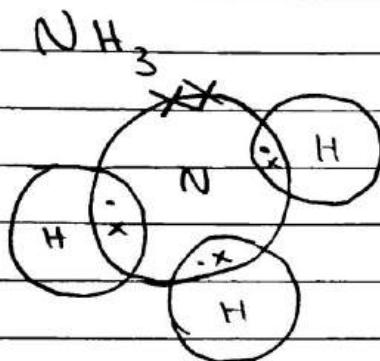
Define dative covalent bond or coordinate bond

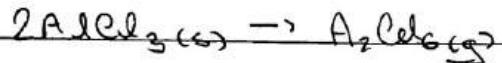
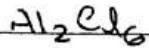
It is a special type of covalent bond where ~~one~~ of the bonded atoms ~~shares~~ both the bonding e^- .

* Covalent bonds are formed by the over-lapping of orbitals.

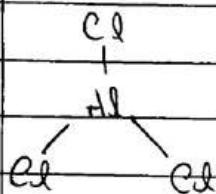
* The number of covalent bonds that an atom can form depends on the number of singly filled orbitals.

For the formation of Dative Covalent bonds, the atom that will share both the e⁻ must have a fully-filled orbital and the other bonded atom must have a vacant orbital.

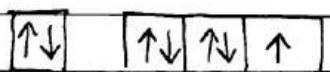
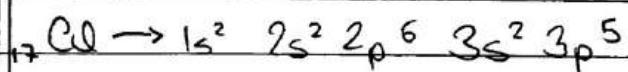
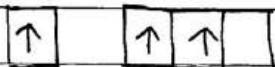
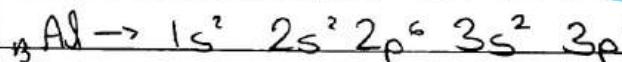




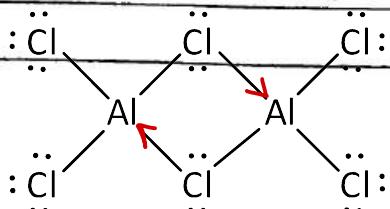
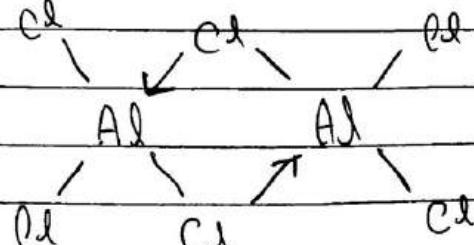
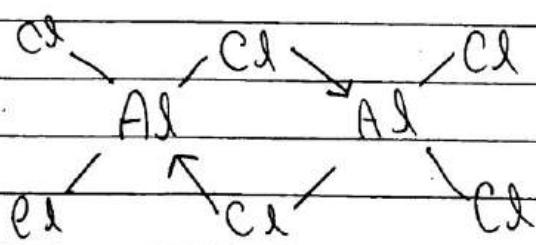
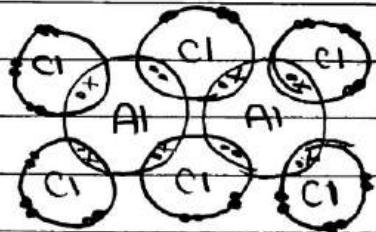
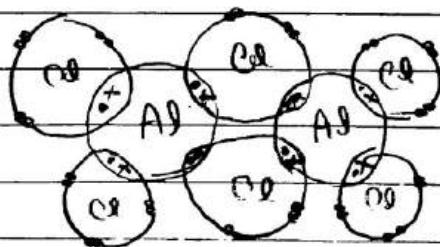
Dimer

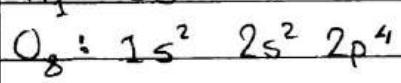
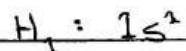
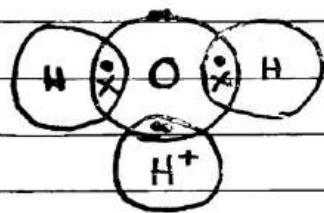
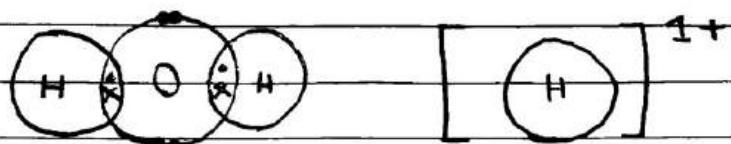
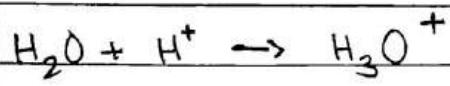


A dimer is a polymer with 2 monomers in its chain (not to be confused with having 2 different types of monomers in its chain) (Since there are only 2 monomers in the chain, there must only be 1 type of monomer/repeated unit in the chain)

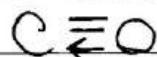
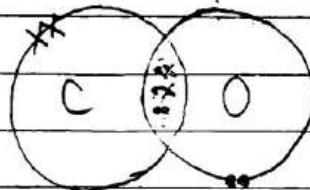
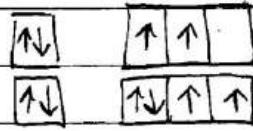
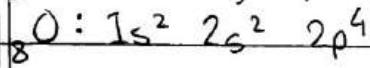
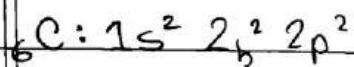


Cl can form up to 7 bonds when energized





Q. Carbon Monoxide Molecule has 2 covalent bonds and 1
Dative Covalent Bond



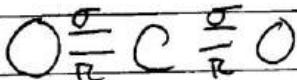
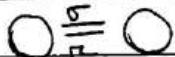
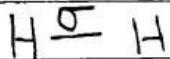
Types of covalent bonds based on e^- sharing

- ↳ Regular covalent bond
- ↳ Dative covalent / Coordinate bond

Type of Covalent Bonds; Based on the type of orbital overlap or overlapping:

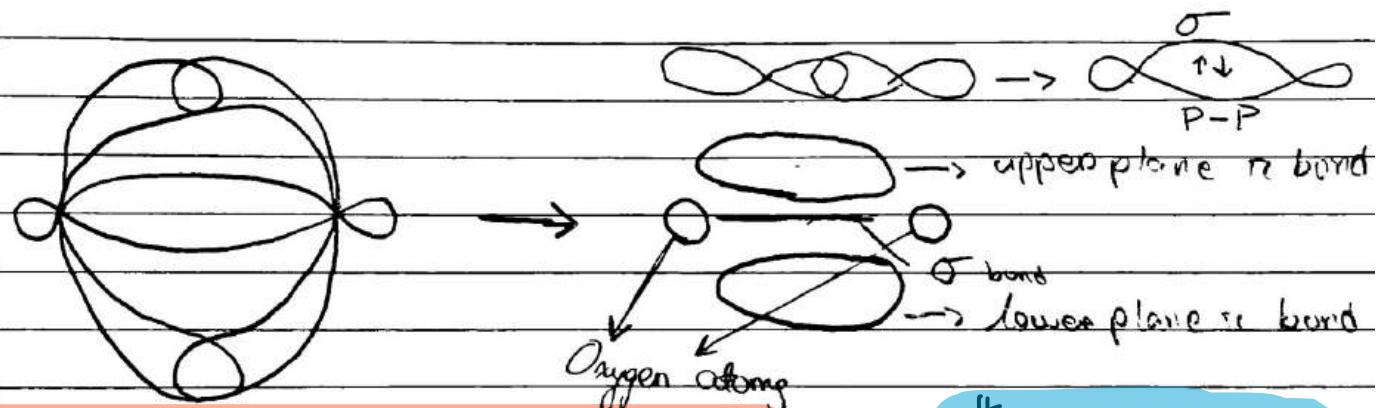
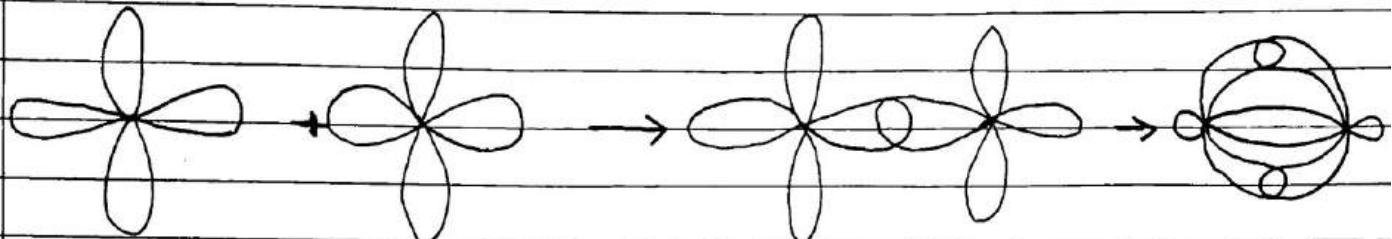
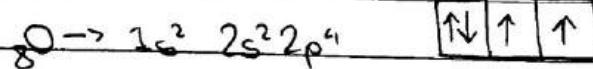
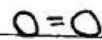
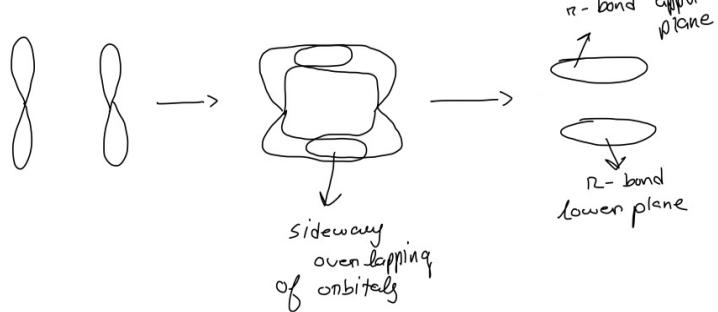
- 1) **Sigma Bond**: They are formed by head to head overlapping of orbitals
- 2) **Pi Bond & Pi Bond**: They are formed by sideways overlapping of orbitals

* π bond never exist alone, the first covalent bond between any 2 atoms is a Sigma Bond, and the other bonds are always π bonds



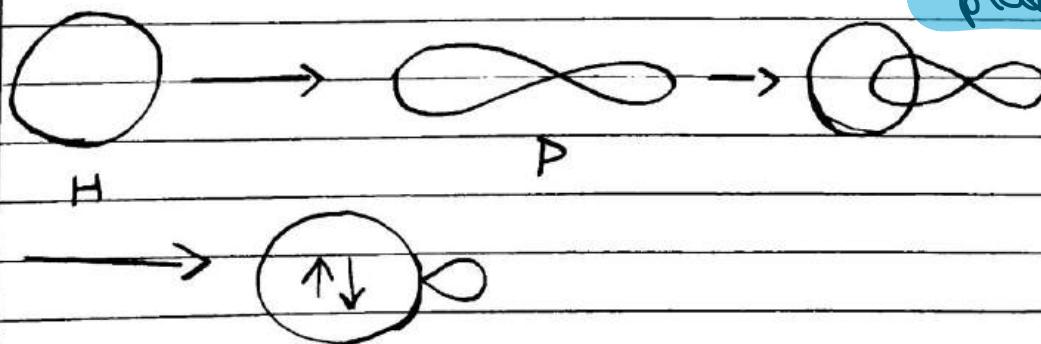
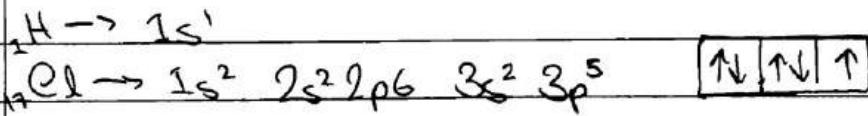
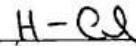
* π bonds are always weaker than Sigma Bonds

Draw a labelled diagram to show the formation of one pi (π) bond.

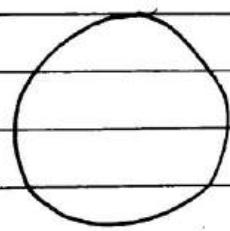
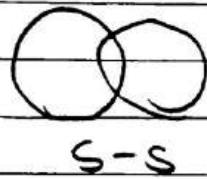
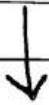
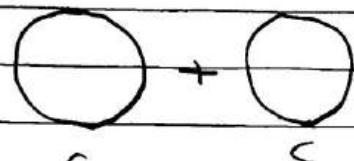


* π bonds are called e⁻ charged clouds.

e⁻ in π bonds
are ~~delocalised~~,
they are outside
the molecule and
move between the
upper and lower
planes of the
 π -bond



H - H

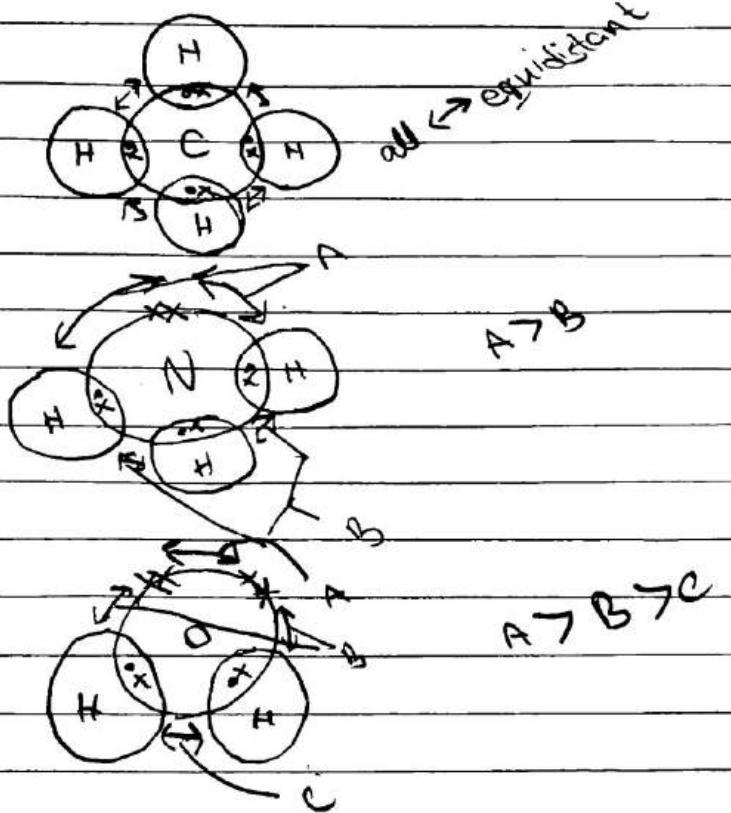


~~Effx~~

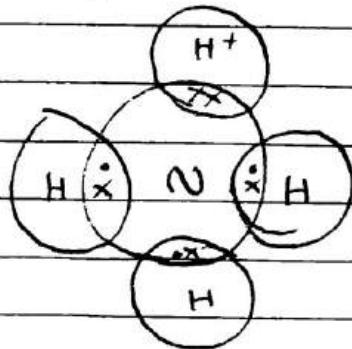
Shape of Molecules

Valence Shell Electron Pair Repulsion Theory (VSEPR Theory)

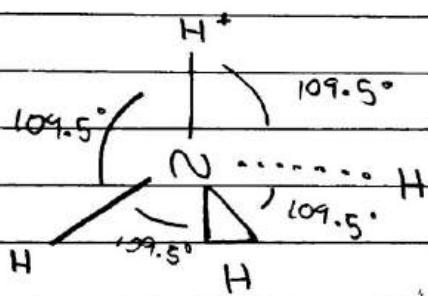
1. In a covalent molecule, e^- pairs arrange themselves as far apart from each other to minimise the repulsion force.
2. Lone pair - Lone pair ($lp-lp$) repulsion force is greater than Lone pair - Bond pair ($lp-bp$) repulsion force which is greater than Bond pair - Bond pair ($bp-bp$) repulsion force.
 $lp-lp > lp-bp > bp-bp$
3. Only the e^- pairs that participate in Sigma Bonds is responsible for the shape of the molecule, e^- in π -bond are neglected. # As the e^- in the π -bonds are delocalised and lie outside the molecule. Refer ↑
4. While determining the sp₃ shape of the molecule only the e^- pairs in the central atom of the molecule is considered.



1) no. of e.p. = 4
 no. of b.p. = 4
 no. of l.p. = 0

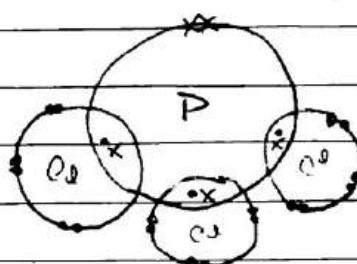


Shape = Tetrahedral



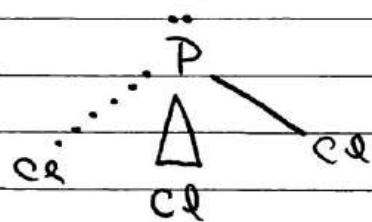
Bond angle = 109.5°

2) no. of e.p. = 4
 no. of b.p. = 3
 no. of l.p. = 1



Shape = (Triangular) Trigonal pyramidal

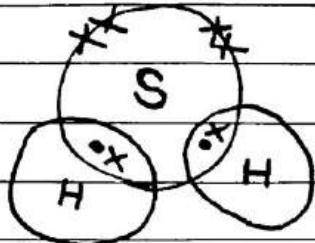
Bond Angle = 107°



For molecules with 4 electron pairs, with each lone pair gained, the bond angle decreases by 2.5°

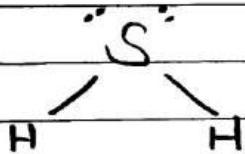
(.....) Into the page (Δ) Out of the page (—) Same plane of the page
 Central atom

- (3) no. of e⁻.p. = 4
 no. of b.p. = 2
 no. of l.p. = 2



Shape = V-shape OR Bent-Shape

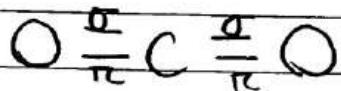
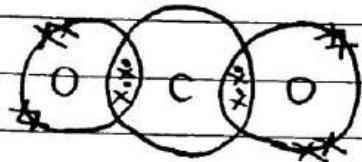
Bond angle = 104.5°



there cannot be a bond angle and/or bond shape for a molecule, as for there to be a bond angle, there must be 2 b.p. or bonds

\rightarrow there cannot be any "shape" of there is no reference point/perspective when it is just 2 atoms

4) CO_2



no. of c.p. = 2

no. of b.p. = 2

no. of l.p. = 0

π -bonds are neglected when determining the shape of molecules

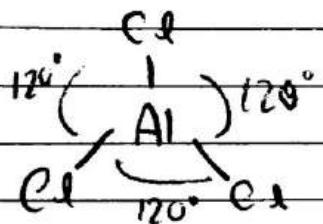
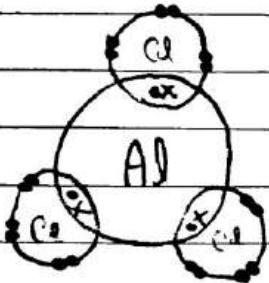
(Shape = linear)

Bond Angle = 180°

5) no. of e.p. = 3

no. of b.p. = 3

no. of l.p. = 0

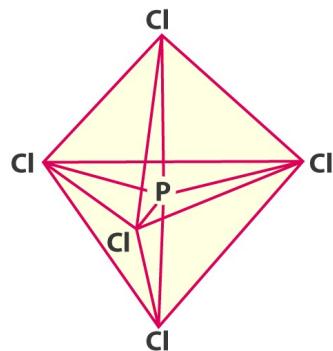
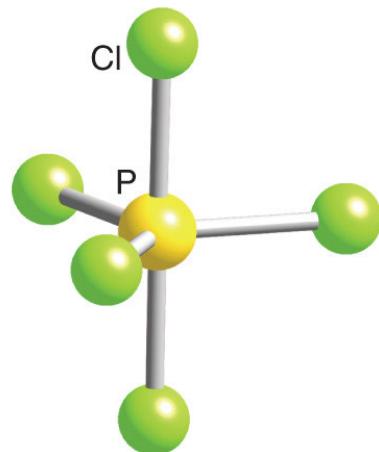
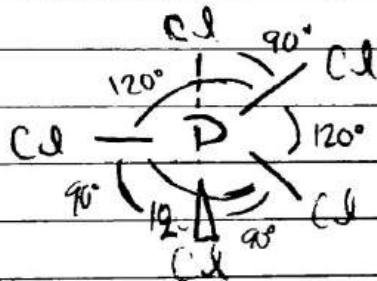
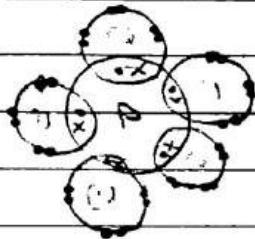


Shape = Trigonal Planar

Bond angle = 120°

6) no. of e.p. = 5
 no. of b.p. = 5
 no. of l.p. = 0

PCl_5



Shape = Trigonal bipyramidal

Bond angle = 120° , 90° , 180° We will not consider 180° when answering Qs

${}^{15}\text{P}$: $1s^2 2s^2 2p^6 3s^2 3p^3$



When energy is applied an e^- from 3s can jump to 4s due to the low proximity in energy level between 3s and 4s

${}^{15}\text{P}$: $1s^2 2s^2 2p^6 3s^1 3p^3 4s^1$



Initially P only has 3 singly filled orbitals and hence can only form 3 covalent bonds, but when energized P has 5 singly filled orbitals and hence can form 5 covalent bonds as seen in PCl_5 .

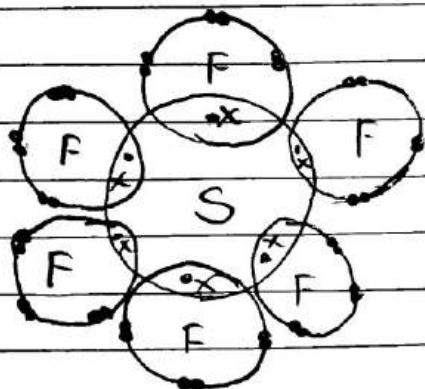
7.)

no. of e.p. = 6

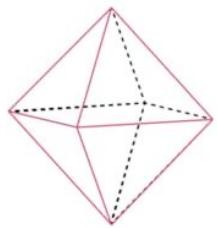
no. of b.p. = 6

no. of l.p. = 0

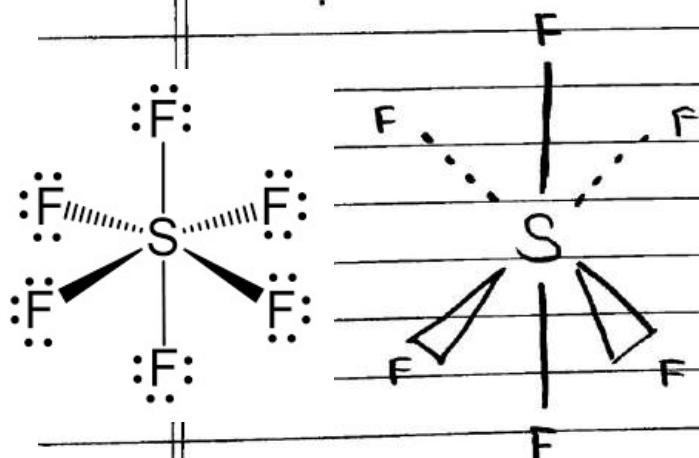
SF_6



Sulfur Hexafluoride, SF_6



Octahedron



The plane is perpendicular to the paper. The base square is parallel to the plane.
-paper plane

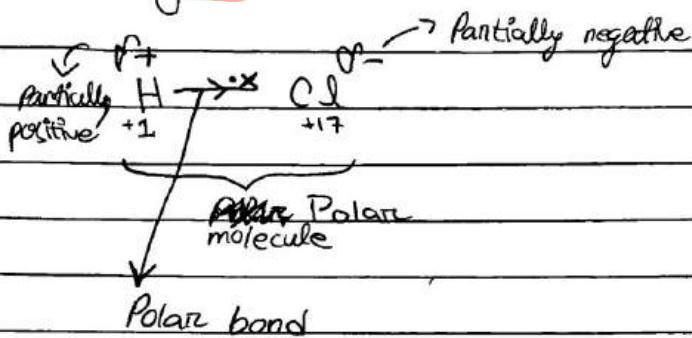
Shape = Square planar, ~~or~~ bi-pyramidal / Octahedral

Bond Angle = 90° , 180° We will not consider 180° when answering Qs

Bond Polarity

Define Electronegativity

→ It is the ability of a bonded atom to attract bonding pair of e^- towards it.



Define Bond Polarity

→ It is the uneven distribution of bonded e^- in a covalent bond due to difference in the electronegativity of the bonded atoms.

Factors Affecting Electronegativity

1. Nuclear Charge : Atom with

2. Atomic Radius :

Bonded

Atoms with greater nuclear charge and lesser atomic radius than atoms with smaller nuclear charge and larger atomic radius are more electronegative.

Fluorine is the most electronegative atom in the periodic table.

Bond Polarity is a vector quantity

If two polar bonds of equal magnitude (dipole moment) are oppositely directed to each other, they will get cancelled out, resulting overall the molecule to be non-polar.

Atom Electronegativity

H 2.1

C 2.5

N 3.0

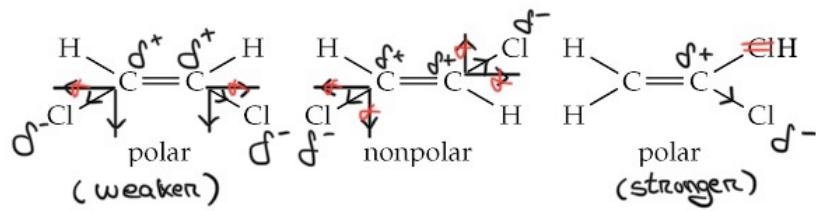
O 3.5

F 4.0

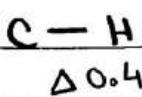
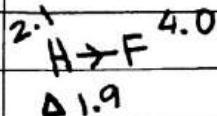
Cl 3.0

Br 2.8

S 2.5

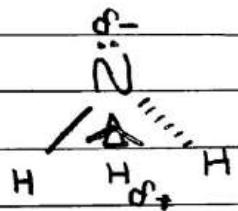


//↓ and the end with l.p. is slightly more negative relative to the rest of the atom. Since the overall charge in an atom is constant, the end of the atom directly opposite to the l.p. is slightly more positive relative to the rest of the atom. Hence the b.p. e⁻ at that end are pulled faster than the rest of the b.p.e⁻. Those e⁻ arrive first, a polar bond is formed, the central atom has a partial -ve charge. The bonding e⁻ are no longer attracted to the central atom. No other polar bonds are formed.



anything below 0.4 is non-polar

- # If the difference in electronegativity is greater than 1.9, then it will be a pure ionic bond / formula unit
- # If the difference in electronegativity is less than 1.9, then it will be a covalent bond / molecule

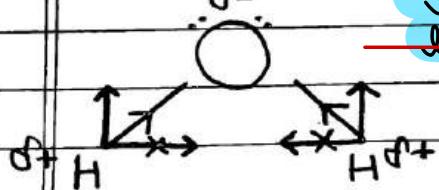


> In an atom with 0 l.p., the e⁻ are evenly distributed. Hence it can form 4 polar bonds as all the e⁻ are pulled at the same rate. However, in an atom with a l.p., there is uneven distribution of e⁻ as l.p.-b.p. repulsion > b.p.-b.p. repulsion, //↑

no. of l.p. in central atom = no. of polar bonds in the molecule

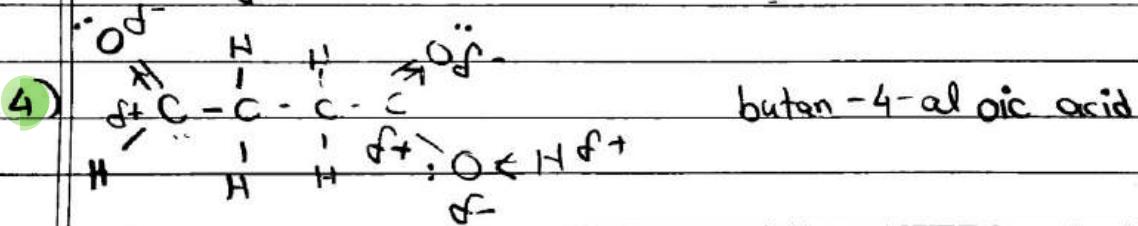
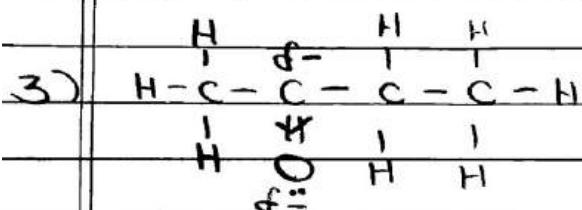
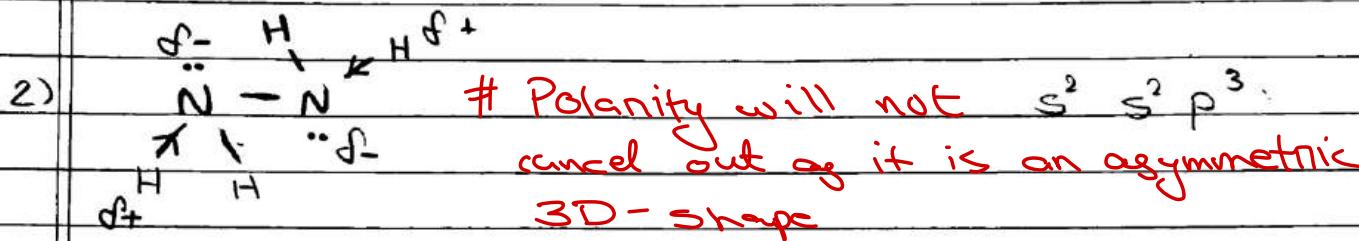
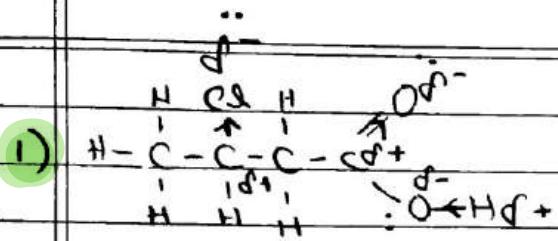
(the bonds that are opposite to the l.p. is/are the polar bonds)

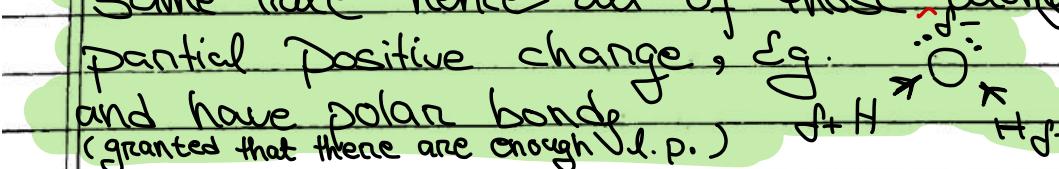
↳ 1st point of VSEPR Theory. The e⁻ pair will arrange themselves as far apart as possible. As polarity ↑, l.p.-b.p. Repulsion ↑



When cancelling out vectors in bond polarity, the two or more bond polarities must be with the same atom (max gap of 2 atoms)
 $C=C$ They must also be on the same plane

Example Molecules:



When a highly electronegative atom is bonded with 2 or more atoms of a different element, it will pull the bonded e⁻ from all those atoms at the same rate hence all of those atoms will gain a partial positive charge, e.g.  and have polar bonds (granted that there are enough V.L.P.)

However when it is bonded to 2 or more atoms of different elements, it will pull the bonded e⁻ from the atom with the least electronegativity faster than from the other atoms. This will cause the highly electronegative atom to become partially negatively charged. Thus it will no longer be able to attract the bonded e⁻ from the other atoms and they only have 1 polar bond. E.g. Example molecule 1 and 4

Intermolecular Forces

There are 3 types of intermolecular forces (imf)

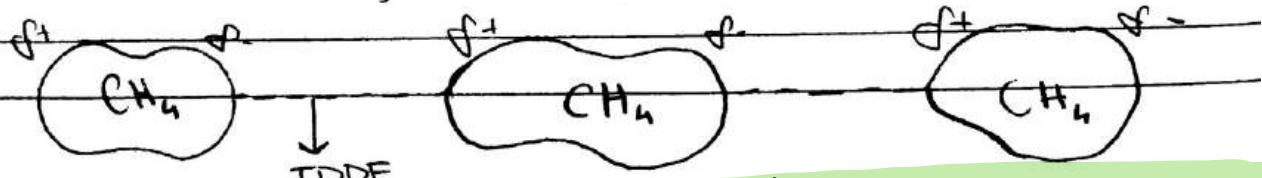
- ↳ Temporary dipole-dipole force (TDDF)
/ Instantaneous dipole-dipole forces / Induced dipole-dipole force
- ↳ Permanent dipole-dipole force (PDDF)
- ↳ Hydrogen Bonding

TDDF is the weakest imf and hydrogen bonding is the strongest imf.

TDDF

All non-polar molecules will only contain TDDF

↳ Due to random movement of e⁻s in a non-polar molecule, at any instant one end of the molecule will contain more e⁻s than the other end. This causes 1 end to have partially negative charge and the other end to have partially positive charge. This causes the formation of a dipole in a non-polar molecule. However this dipole is temporary / not permanent. At the very next instant, the higher number of e⁻s can shift if end, resulting in the change in the direction of the dipole.



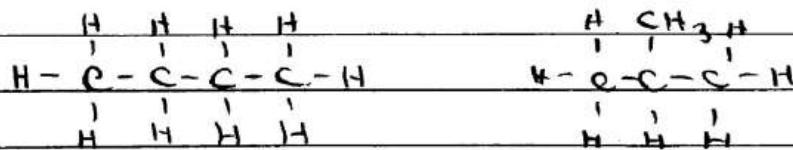
A temporary dipole induces temporary polarity on surrounding non-polar molecules, resulting in TDDF between them.

Factors affecting Strength of T.D.D.F.

1 Number of e^- , $e^- \uparrow$ T.D.D.F. \uparrow , with increase in no. of e^- , strength of T.D.D.F. increases

Q. Explain why down the group, b.p. of Halogens increases

Ans. As we go down the group the no. of e^- in each halogen molecule increases, causing strength of T.D.D.F. to increase. Thus more energy is required to break the force.



Surface area of straight chain isomer \rightarrow S.A. of

In case of 2 isomers, the straight chain isomer has more surface area than the branched isomer. The straight chain isomer can interact more with each other using T.D.D.F. than branched isomer. This causes straight chain isomer to form have greater b.p. am and m.p.

L, contact area \uparrow

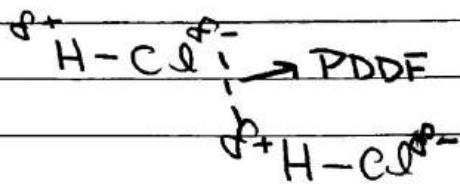
effective strength of the force \uparrow

in chemistry, force = bond

straight chain molecules have larger surface area and thus can interact more with each other, straight chain molecules can pack more closely therefore stronger van der Waals' forces

Permanent dipole-dipole Force (PDDF)

It only exists in Polar molecules



* The greater the bond polarity in a molecule the stronger will be the PDDF

Van der Waals force

(Vanderwaals' Force)

* Even a polar molecule has a non-polar part, they also act between polar molecules. # b.p. of halogenoalkanes increases down the group

The TDDF acting between polar molecules is called van der Waals' force in MCQs and Pgjt papers.

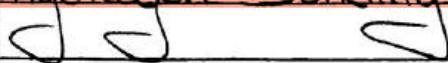
Do polar molecules have stronger TDDF than non-polar molecules? NO

TDDF in polar molecules only concerns the non-polar part of the molecule

The partial charge is confined to the involved atom and NOT the entire molecule. Hence the effect of the partial charges on the rest of the molecule (the non-polar part of the molecule) is minimal and thus the TDDF in these parts is (mostly) unaffected.

Hydrogen bonding is a special type of PDDF that occurs between the δ -p. of a highly electronegative atom (N, O or F) and the δ^+ hydrogen atom in a N-H, O-H or F-H bond.

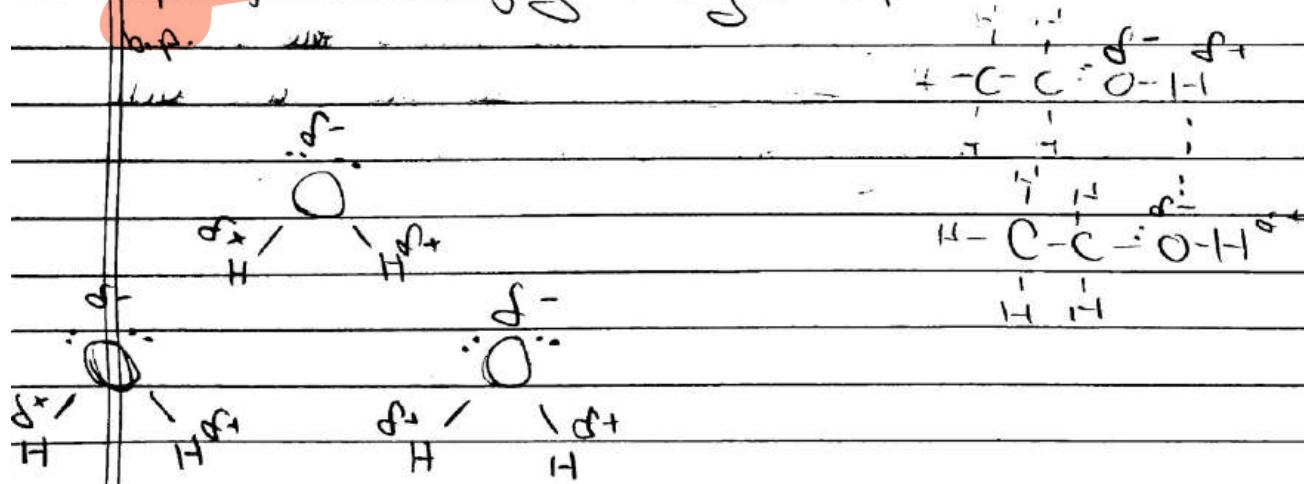
Hydrogen Bonding



It's a special type of PDDF where the Polar bond must have δ^+ Hydrogen atom bonded to highly electronegative Oxygen, Nitrogen or Fluorine atom.

The molecule to which the H atom belongs to is the molecule whose hydrogen bond it is.

Q. Ethanol has a boiling point of 78°C and water has a b.p. of 100°C . Using a diagram explain their difference in b.p.



Per water molecule can form 2 hydrogen bond whereas each Ethanol molecules can only form 1 hydrogen bond. The more amount of Energy is required to break the larger number of hydrogen bond in water.

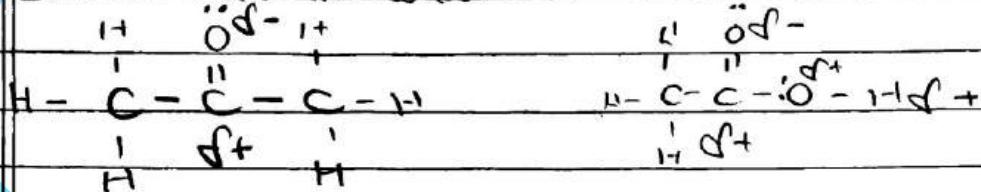
Why is water less dense than ice?

The effect of hydrogen bonding on water

most significant intermolecular force = m simf

- ★ Polar molecules have an affinity towards each other
- ★ Non-Polar " " " " " "

Q. * Explain why Propanone is less soluble than Ethanoic Acid in water



Same as why Br_2 more soluble in hexane than Water

Aps.

If the most significant force in the solute matches the most significant force in the ^{intermolecular} solvent, then the solute will dissolve more readily in the solvent.

The relative strengths of the intermolecular forces are:

hydrogen bonding $>$ PDDF $>$ TDDF

Intermolecular forces/bonds of the same type can be considered to have similar relative strengths.

When a solute is added to a solvent, intermolecular forces/bonds act to form between unbonded solute molecules and unbonded solvent molecules. When these bonds form, energy is released. This energy is used to overcome the solute-solute and solvent-solvent intermolecular forces so that they are free/available to form solute-solvent intermolecular forces/bonds.

In compounds with the same most significant intermolecular force the ratio of bond forming to bond breaking is almost 1:1. The formation of 2 solute-solvent bonds releases enough energy to break 1 solute-solute bond and 1 solvent-solvent bond which

then go on to form 2 other solute-solvent bonds. Hence they dissolve readily in each other.

For compounds with different msimf, the solute-solvent bond is the weaker msimf. They multiple bonds need to be formed in order to release enough energy to break 1 of the stronger msimf bond. The ratio is no longer 1:1 and they less solubility

Polarization of Ionic Compound

Define Polarization

→ The distortion of the negative charge cloud of an anion by the cation in an ionic compound is called Polarization.



Asymmetric distribution of e⁻ decreases.

Distortion of anion charge increases

Polarization of anion increases

Ionic character decreases.

Covalent " increases.

Nothing to do with electronegativity

Factors affecting the Polarizing Ability of Cations

- Cations with greater effective charge and smaller size have greater polarizing ability.
- In other words, cations with greater charge density have greater polarizing ability.

Factors affecting the Polarization of Anions

- Anions with larger size and smaller charge are easier to distort than anions with smaller size and greater charge.

Electrostatic Force of Charge $1 \times \text{Charge } 2$

Q) Why certain atoms form cations in preference to anions and vice versa?

- Removing all the e^- from the atom requires very high ΔH_i , more so than gaining e^- , and results in the formation of very unstable ions (cations). Hence anions are formed in preference to cations.
- In Chemistry, the more energy that is required to form an ion or compound, the less stable that ion or compound will be.
- The less stable an ion or compound is, the more readily it will revert back to its original form or base component.

Bond length and Bond energy

Bond length: The shortest possible distance between the nuclei of 2 covalently bonded atoms

Bond Energy: It is the amount of energy required to break 1 mol of a covalent bond in gaseous state.

Bond length or Bond Energy

Covalent bond with smaller bond length have greater orbital overlapping.

$B.L.C-C > B.L.C=C > B.L.C\equiv C$

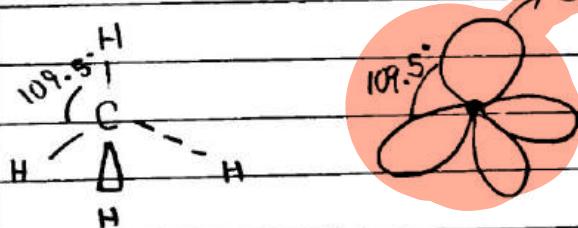
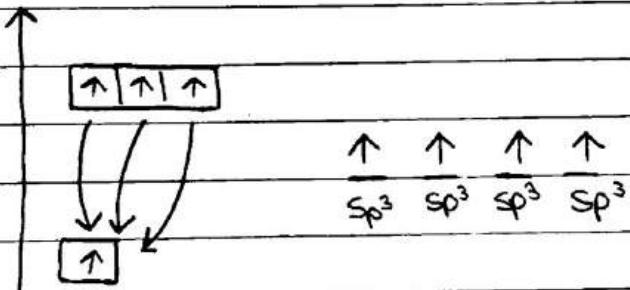
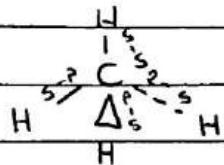
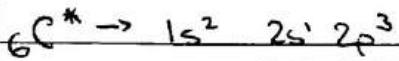
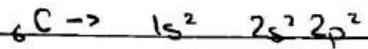
$B.E.C-C < B.E.C=C < B.E.C\equiv C$

HI is decomposed using a heated glass rod
↳ The purpose of a heated glass rod in any system is to provide E_a for the reaction to occur

↳ Hydrogen Halides are misty / Steamy fumes

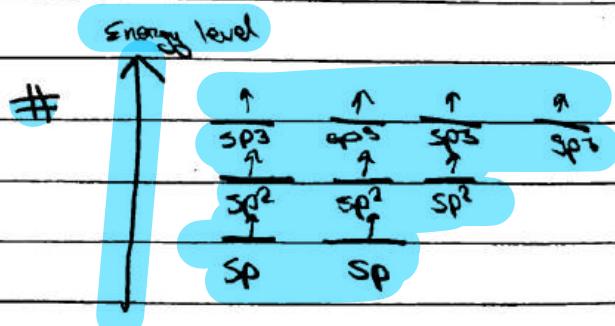
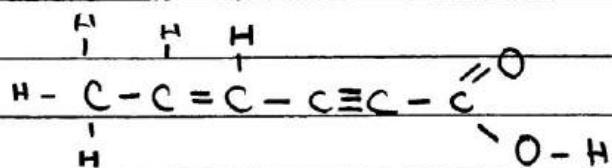
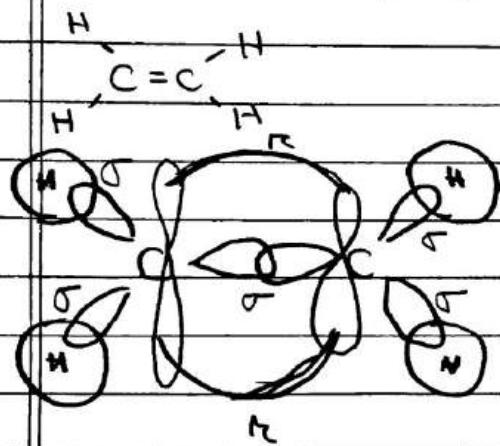
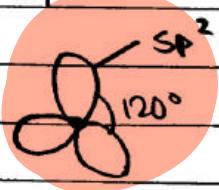
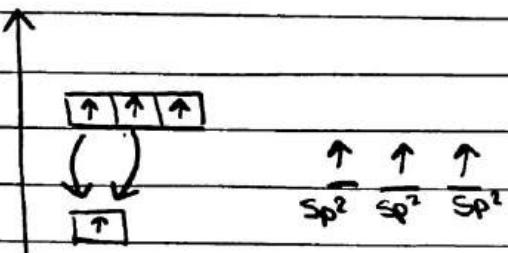
Bond length and Bond Energy have no effect on the b.p. or m.p. of the compound. They only affect the energy required to decompose the compound. b.p. and m.p. are only affected by the strength of intermolecular forces. (for covalent compound)

Hybridization



Hybridised orbitals will never overlap sideways, they will only overlap head to head.

No. of hybridised orbitals = no. of σ (sigma) bonds + no. of π



double pairing counts towards hybridization the same way
 σ -bond pairs do. Only π -bond pairs are ignored. e.g. H_2O

The Periodic Table of Elements

1		2		Group												18																	
				1						2																							
Key												Group																					
atomic number atomic symbol name relative atomic mass												Group																					
3 Li lithium 6.9	4 Be beryllium 9.0	5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2	11 Na sodium 23.0	12 Mg magnesium 24.3	13 Al aluminum 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	16 S sulfur 32.1	17 Cl chlorine 35.5	18 Ar argon 39.9	19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Kr krypton 83.8	
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Nb niobium 92.9	41 Mo molybdenum 95.9	42 Tc technetium –	43 Ru ruthenium 101.1	44 Rh rhodium 102.9	45 Pd palladium 106.4	46 Ag silver 107.9	47 Cd cadmium 112.4	48 In indium 114.8	49 Sn tin 118.7	50 Sb antimony 121.8	51 Te tellurium 127.6	52 I iodine 126.9	53 Xe xenon 131.3	55 Cs caesium 132.9	56 Ba barium 137.3	57–71 lanthanoids 137.3	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 208.0	84 Po polonium –	85 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids –	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –	113 Nh nihonium –	114 Fl ferovium –	115 Mc moscovium –	116 Lv livmorium –	117 Ts tennessine –	118 Og oganesson –																
Lanthanoids												63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0													
Actinoids												93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –											
57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.4	61 Pm promethium –	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0	72 Ac actinium –	73 Th thorium 232.0	74 Pa protactinium 231.0	75 U uranium 238.0	76 Np neptunium –	77 Pu plutonium –	78 Am americium –	79 Cm curium –	80 Bk berkelium –	81 Cf californium –	82 Es einsteinium –	83 Fm fermium –	84 Md mendelevium –	85 No nobelium –	86 Lr lawrencium –				

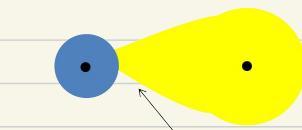
Misc. Notes from Past Paper Questions

- # b.p. of halogenoalkanes increases down the group
- # Effects of hydrogen bonding on water - It causes water to have a higher than expected:
 - i) m.p.
 - ii) viscosity
 - iii) surface tension

Ice also floats on water due to hydrogen bonding between water molecules

- # A solute is ^{very / readily} soluble in water because it forms hydrogen bonds with water
- # Ethylene glycol ($\text{HOCH}_2\text{CH}_2\text{OH}$) is the most abundantly produced diol. It works as anti-freeze.
It lowers the m.p. of ice as it interferes with the hydrogen bonding network of pure water.
- # Weaken i.m.f. will make it more likely that a gas will approach ideal behaviour
- # ~~A dative bond is always polar~~ Ignore as it is not included in the syllabus

- # AlCl_3 is a covalent compound because the high degree of polarization of the Cl^- ion creates a region where the Al and Cl orbitals overlap and e⁻ sharing occurs.



Region where electrons are "shared"

The difference in electronegativity between Al and Cl is 1.5 which is less than 1.9 (on p. 7) and hence it is a covalent bond.

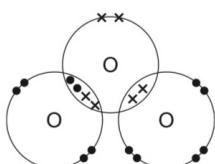
It sublimes due to the weak i.m.f. between the AlCl_3 molecules

- # The standard enthalpy change of atomisation of a diatomic ~~gaseous~~ molecule is half the bond energy / standard enthalpy
- # In hydrogen bonding, I.P. function of H-bond acceptor and the H on the bonding orbital function of H-bond donor.

- # Molecules in which all the atoms surrounding the central atom are the same (tend to be) ^{are} non-polar if there are no l.p. in the central atom (eg such molecules tend to be ^{are} symmetrical)
- # A symmetric molecule is always a non-polar molecule
- # SiCl_4 reacts with water whereas CCl_4 does not. This is because C does not have any d-orbitals to accept l.p. of e^- from the O atom in water, whereas Si does have vacant d-orbitals to accept l.p. from the O in water. // common sense
- # General eq. for the decomposition of G-II Nitrates
 $2X(\text{NO}_3)_2 \rightarrow 2X\text{O} + 4\text{NO}_2 + \text{O}_2$
- # CH_4 has a higher b.p. than Ne, despite having the same no. of e^- , because CH_4 has a larger surface area.
- # Geometric isomers do not have identical b.p.
 Optical isomers have identical b.p.
- # A solute does not fully dissolve in water because:
 M1) Hydrogen bonds exist between H_2O molecules
 M2) Hydrogen bonds do not exist between the solute molecules
- # How ionic covalent (co-ordinate) bonds are formed?
 ↳ X (atom(s)) donates a pair of e^- to the X-Y covalent bond(s) - mg

13 VSEPR theory should be used in answering this question.

The dot-and-cross diagram for an ozone, O_3 , molecule is shown.



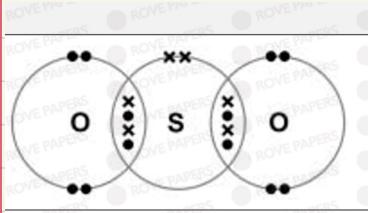
What is the predicted bond angle in this molecule?

- A 107° B 109.5° C 117° D 120°

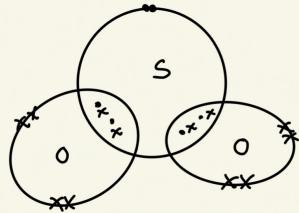
3 bp is 120° so 2bp and 1lp will be slightly lower than that

#

The σ -bond of a compound n-bond will always be formed by the overlapping of 2 P-orbitals and NOT S-S or S-P overlap - Sp^2 orbital



(f)



Shape and of SO_2
Bonding

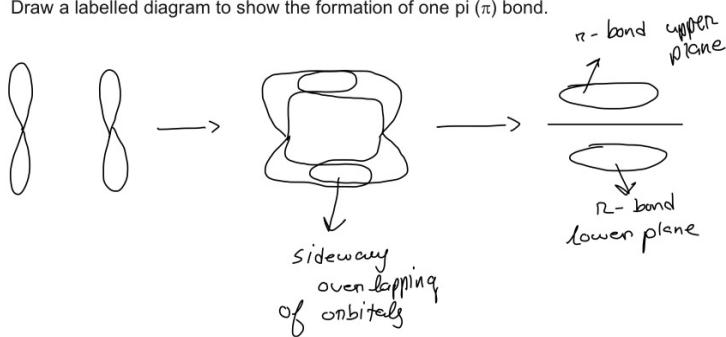
#

- ★ 6 Which statement about the $\text{Cl}-\text{N}=\text{O}$ molecule is correct?
- A Each molecule contains one σ and two π bonds.
 - B It is a non-polar molecule.
 - C It is a linear molecule.
 - D The nitrogen atom is sp^2 hybridised.

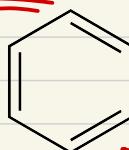
Lone pair DEFINITELY
count towards
hybridization of a
 σ -bond

N_2 is inert because : needs a lot of energy to break...
: .. triple bond
: non-polar molecule

Draw a labelled diagram to show the formation of one pi (π) bond.



According to our syllabus, the e^- in π -bonds are NOT delocalised. The ONLY exception being in Benzene rings // phenyl groups.

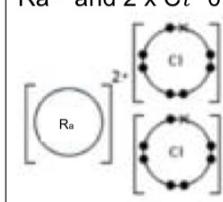


Cyclohexene is NOT Benzene Ring



The 3 imfs : Instantaneous dipole - Induced dipole
: Permanent Dipole - Permanent Dipole
: Hydrogen bonding

Because 1 instantaneous dipole induces a dipole in the neighbouring molecule resulting in electrostatic forces between them

- # Suggest why the molecule of CO contains a dipole moment
 \hookrightarrow O is more ELECTRONEGATIVE than C - mg
- # although aldehydes and ketones can't hydrogen bond with themselves, they can hydrogen bond with water molecules
- # 2(b) Ra²⁺ and 2 x Cl⁻ 0 electrons surrounding Ra AND 8 electrons surrounding Cl
- 
- # outer e⁻
meas empty shell
for acting

- # High mp due to MANY strong metallic/ionic/covalent bonds
- # Describe how sp hybridised orbitals are formed:
 \hookrightarrow Mixing/overlap/combination of one s and one p orbital - mg

Sketch a diagram to show how two sp hybrid orbitals can form a σ-bond (sigma)



m₁

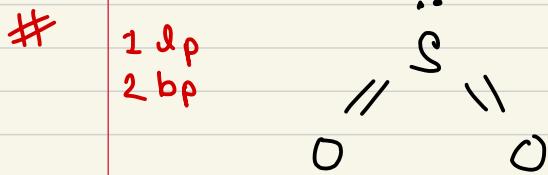


m₂



- mg

- # Why Branched molecule has lower bp: straight chain molecules have larger surface area and thus can interact more with each other, straight chain molecules can pack more closely therefore stronger van der Waals' forces



Shape: non-linear or bent
Bond Angle: 115° - 120°

Intermolecular Forces+ Electronegativity & Bond Properties

Intermolecular Forces+ Electronegativity & Bond Properties

8

Ethylene glycol, HOCH₂CH₂OH, is used as a de-icer. It allows ice to melt at temperatures below 0 °C.

Which statements are correct?

- 1 Ethylene glycol changes the extensive network of hydrogen bonds in ice.
- 2 Ethylene glycol molecules form hydrogen bonds with other ethylene glycol molecules.
- 3 Ethylene glycol molecules will dissolve in the water formed from the ice.

Paper 1 Variant 3 Summer 2017 | Q33

Answer:

A

Intermolecular Forces+ Electronegativity & Bond Properties

12

Halogenoalkanes show trends in their physical and chemical properties.

Which properties steadily increase from C₂H₅Cl to C₂H₅Br to C₂H₅I ?

- 1 the polarity of the carbon-halogen bond
- 2 the boiling point of the halogenoalkane
- 3 the rate of reaction of the halogenoalkane with nucleophiles

Paper 1 Variant 3 Summer 2016 | Q37

Answer:

C

Water has some unusual physical properties compared to other hydrides of Group 16 elements. Some of these properties are due to hydrogen bonds. These intermolecular forces are much stronger in water than they are in H₂S, for example.

Which statements are correct?

- 1 Hydrogen bonds cause the melting point of ice to be higher than expected.
- 2 Hydrogen bonds cause the surface tension of water to be higher than expected.
- 3 Hydrogen bonds cause the viscosity of water to be higher than expected.

Paper 1 Variant 3 Summer 2016 I Q32**Answer:**

A

What will make it more likely that a gas will approach ideal behaviour?

- A higher pressure
- B lower temperature
- C more polar molecules
- D weaker intermolecular forces

Paper 1 Variant 3 Summer 2016 I Q2**Answer:**

D

Which molecules have an overall dipole moment?

- 1 carbon monoxide, CO
- 2 phosphine, PH₃
- 3 carbon dioxide, CO₂

Paper 1 Variant 2 March 2016 I Q31**Answer:**

B

Why does aluminium chloride, Al_2Cl_6 , sublime at the relatively low temperature of 180°C ?

- 1 The intermolecular forces between the Al_2Cl_6 molecules are weak.
- 2 The co-ordinate bonds between aluminium and chlorine are weak.
- 3 The covalent bonds between aluminium and chlorine are weak.

Paper 1 Variant 2 Summer 2015 | Q32**Answer:**

D

Which physical properties are due to hydrogen bonding between water molecules?

- 1 Water has a higher boiling point than H_2S .
- 2 Ice floats on water.
- 3 The H–O–H bond angle in water is approximately 104° .

Paper 1 Variant 3 Summer 2015 | Q34**Answer:**

B

Which energy change corresponds to the enthalpy change of atomisation of hydrogen at 298 K ?

- A the bond energy of a H–H bond
- B half the bond energy of a H–H bond
- C minus half the bond energy of a H–H bond
- D minus the bond energy of a H–H bond

Paper 1 Variant 2 Summer 2013 | Q11**Answer:**

B

What is involved when a hydrogen bond is formed between two molecules?

- 1 a hydrogen atom bonded to an atom less electronegative than itself
- 2 a lone pair of electrons
- 3 an electrostatic attraction between opposite charges

Paper 1 Variant 2 Summer 2011 | Q34

Answer:

C

Covalent Bonding & Shapes of Molecules

2

Ammonia, NH_3 , and hydrazine, NH_2NH_2 , are two compounds of nitrogen, N_2 .

- Which statement is correct?
- A The N-N bond in NH_2NH_2 is polar.
 - B NH_3 and NH_2NH_2 have lone pairs of electrons but N_2 does not.
 - C The oxidation number of each nitrogen in NH_2NH_2 is +2.
 - D The reaction of nitrogen with hydrogen has a high activation energy.

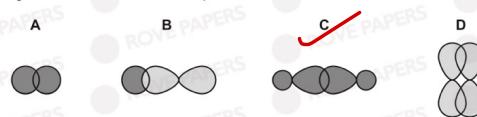
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[Turn over]

A σ bond is made between two carbon atoms in a molecule of ethene.

Which diagram shows the orbital overlap that occurs to form this bond?



Paper 1 Variant 2 March 2019 | Q19

Answer:

D

Paper 1 Variant 2 Summer 2019 | Q4

Answer: # The carbons

C have sp^2 orbitals

has to
be 2 p orbitals
because the
C atoms will
also form a
 π -bond with
each other
which cannot
be done by
 s -orbitals