CHEMICAL BONDING

Question 1

AlF_3 has higher melting point than $AlCl_3$. Explain

Solution

AlF_3 is more ionic in characters as result of its lower degree of polarisation brought about by smaller polarisability of smaller sized F^- thus making its melting point higher

Question 2

Between $AlCl_3$ and NaCl which one sublimes on heating? Give reason for your choice.

Solution

$AlCl_3$ sublimes on heating.

Reason

$AlCl_3$ is more covalent in characters as result of its higher degree of polarisation brought about by smaller in size and higher charged Al^{3+}

Question 3

Molten AlF_3 conducts electricity while molten $AlCl_3$ does not. Explain

Solution

AlF_3 is more ionic in characters as result of its lower degree of polarisation brought about by smaller polarisability of smaller sized F^- thus in molten state it has enough concentration of free ions to do electrolytic conduction while $AlCl_3$ has not

Question 4

SO_2 is polar while CO_2 is non – polar although both have the similar empirical formula. Explain.

Solution

CO_2 Being symmetrical (linear) in structure ($O = C = O$) has zero resultant of dipole moment thus becoming non – polar while SO_2 being bent in structure (as result of repulsion exerted by lone pair in sulphur to the bonds), is unsymmetrical so SO_2 has non – zero resultant of dipole moment and hence become polar.

Question 5

Why non-metals do not form metallic bond although they have large number of valence electrons accompanied with smaller atomic radii compared to metals?

Solution

Atomic radii of non-metals are so small that their ionisation energies are very high thus there are no delocalised valence electrons in non metals.

Question 6

If oxygen is highly electronegative, why is a covalent bond between oxygen atoms considered non polar?

Solution

The two bonded oxygen atoms having equal electronegativity, the nucleus of each oxygen exert an equal pull on the shared electrons and hence the bond become non polar.

Question 7

a) Give the type of bonding present in $BeCl_2$
b) Give the type of bonding present in $BaCl_2$
c) Explain why the type of bonding is different in these two compounds.

Solution

a) Covalent
b) Ionic
c) Be being smaller in size than Ba, has greater polarising power and hence there is a greater degree of polarisation in $BeCl_2$ than in $BaCl_2$.

Question 8

Classify the dipole moments of the following molecules as either "Zero" or "nonzero"

i) Hydrogen iodide
ii) Carbon dioxide
iii) Ethanol
iv) 1-chloro-2-butyne$(ClCH_2C \equiv CCH_3)$
v) Nitrogen dioxide.

Solution

i) Non-zero
ii) Zero
iii) Non-zero
iv) Non-zero
v) Non-zero

Question 9

Copper is ductile while copper (II) sulphate is brittle. Explain

Solution

- In copper, stress applied on a metallic lattice causes sliding of layers of cations without breaking the metallic structure because the sea of electrons(delocalised electrons) are still holding the cations together and hence copper become ductile.
- In $CuSO_4$, stress applied on the ionic lattice with regular pattern causes sliding of layers resulting in ions of similar charge coming together leading to the repulsion that shatters the ionic structure and hence $CuSO_4$ become brittle.

Question 10

Compare the electrical conductivity of solid sodium metal with that NaCl

Solution

Na Metal conducts electricity by movement of its delocalised valence electrons in the solid state while NaCl conducts electicity by movement of ions which are Na^+ and Cl^- in molten state or aqueous

Question 11

Do compounds with covalent bonds conduct electricity? Explain why or why not.

Solution

Covalently bonded compounds do not conduct electricity, regardless of the state they are in. This is because electrons are being shared during the bond, not donated or accepted. Because there is no transfer of electrons, there are no ions. Ions are what conduct electricity and thus, without ions, the compounds cannot conduct electricity. Also, covalent compounds are uncharged molecules with the electrons held tightly together. This also means there are no free electrons, or charged ions, to conduct electricity.

The exception to this rule is when a covalent compound reacts with water. When this occurs, ions are formed and the compound is able to conduct electricity

Question 12

Are ionic compounds known as being brittle or malleable? Explain discussing the bonds.

Solution

Ionic compounds are said to be brittle. This is because of the arrangement of the ionic lattice. This arrangement is specific to alternate positive and negative ions and keep similar charges far apart. When it is altered, the like charges will become close to each other and will repel each other. This repulsion shatters the lattice causing the compound to break.

Question 13

Why is it that when you hit a metal such as silver with a hammer it will only deform it, whereas it would shatter NaCl?

Solution

Presence of delocalised electrons in a metal makes it to be malleable and ductile. While in the case of NaCl electrons are held tightly in respective ions in ionic bond thus making the ionic compound (NaCl) brittle.

Question 14

Would you expect SiO_2 to be ductile? Explain.

Solution

No. Molecules of SiO_2 are held together by network of giant covalent bonds and hence SiO_2 will be brittle rather than ductile.

Question 15

Why is aluminum able to conduct heat better than quartz glass?

Solution

The delocalised electrons of aluminum metal allow it to conduct heat while quartz which is actually SiO_2 has network of giant covalent bonds in which electrons are tightly held in the bonds and hence they cannot transfer heat well.

Question 16

Explain why third fluorine(F) cannot be added to F_2 to form F_3 despite the fact that fluorine is strongest electronegative element in the periodic table.

Solution

When two F atoms are covalent bonded to F_2, all of their atomic orbitals are full occupied with electrons and hence there is no possibility of forming third covalent bond so as to form F_3.

Question 17

Give differences in properties of ionic and covalent compounds.

Solution

Ionic Compounds	Covalent Compounds
Usually solid at room temperature	Usually liquid or gas at room temperature
High melting and boiling points	Low melting points
Soluble in water	Insoluble in water
Conducts electricity	Do not conduct electricity
Undergo fast reactions	Undergo slow reactions

Question 18

Which molecule should have the larger dipole moment HBr or HI?

Solution

HBr (Because Br is more electronegative than I, the $H-Br$ is more polar that the $H-I$ bond).

Question 19

Does SO_2 have a dipole moment? If so, in which direction does the net dipole point?

Solution

Yes. The dipole points towards the side with the two oxygen atoms because oxygen is more electronegative than sulphur.

Question 20

Explain the condition which must be met by a molecule with polar bonds so as to be non polar.

Solution

The molecule must be symmetrical. This means the polar bonds must be oriented in equal and opposite directions. In this way, there will be no net dipole overall. For example in CO_2, the $C = O$ bond are each polar, but they point towards each end of the linear arrangement and therefore the overall molecule is non polar.

Question 21

Which of the following molecules are polar?

BF_3, CO, CF_4, NCl_3,SF_2

Solution

CO , NCl_3, and SF_2

Question 22

Why are the bonds between H and Cl in HCl molecule never hundred electrovalent or covalent?

Solution

- It is no hundred percent electrovalent because electronegativity of H and Cl do not differ much and thus the covalent bond formation become more favourable.
- However Cl being more electronegative has greater ability of attracting covalent bonded electrons leading to the formation of partial positive charge and partial negative charge on hydrogen and chlorine respectively and hence the compound become partially ionic.

Question 23

For each of the entities listed, give the hybrid state of the selected atom(s) shown in the parenthesis beside it.

a) Iodine pentafluoride molecule(I)
b) Nitrous acid molecule (N)
c) Acetic (ethanoic) acid molecule(C_1 and C_2)
d) Hydrogen peroxide molecule(O)
e) Formic (methanoic)acid molecule(C)
f) Mercury(II) bromide molecule(Hg)

Solution

a) sp^3d^2 hybridisation

(Hint: there are five sigma bonds and one lone pair making a total of 6 pairs of electron in 6 hybrids orbital of about I).

b) sp^2 hybrisation

(Hint: with structure,$HO - \ddot{N} = O$ there are 2 sigma bond and 1 lone pair about and therefore making a total of three sp^2 hybrid orbitals in N)

c) C_1: sp^2 hybridisation

C_2: sp^3 hybridisation

(Hint: first carbon in acetic acid $CH_3 - C - OH$ is the carbon in $COOH$ group and the carbon has 3 sigma bonds and hence three sp^2 hybrid orbitals in it. The second carbon has four sigma bonds and hence it is sp^3 hybridised).
d) sp^3 hybridisation
(Hint: Each oxygen atom in hydrogen peroxide,$H - O - O - H$, has two sigma bonds and two lone pair making a total of four sp^3 hybrid orbitals in O)
e) sp^2 hybridisation
f) sp hybridisation

Question 24

Explain why PCl_5 exist while NCl_5 is not known.

Solution

With outermost energy level of $n = 3$, phosphorus may undergo sp^3d hydridisation so as to form covalency of 5 and therefore forming PCl_5 while in nitrogen there are no d-orbitals in its outermost energy level of $n = 2$ to enable it to form five sp^3d hybrid orbitals.

Question 25

Why PCl_5 is more reactive than PCl_3?

Solution

PCl_5 has higher energy electronic structure as it is formed through sp^3d hybrid orbitals (after excitation of electrons) compared to sp^3 hybrid orbitals used for formation of PCl_3 which do not involve excitation of electrons. This makes PCl_5 unstable and therefore more reactive.

Question 26

Indicate the type of hybridisation of orbitals by the central atom in each of the following:

a) SO_3^{2-}
b) SO_3
c) PCl_4^-
d) ClF_2^-
e) SCl_2

Solution

a) sp^3 hybridisation
b) sp^2 hybridisation
c) sp^3d hybridisation
d) sp^3d hybridisation
e) sp^3 hybridisation

Question 27

What types of hybridised orbitals can be formed by elements of the third period that cannot be formed by elements in the second period?

Solution

sp^3d and sp^3d^2 (Because elements in the second period have outermost energy level of n=2 which has no d sublevel to enable such hybridisation).

Question 28

Arrange the following molecules in order of increasing of their bond angles:

$$CH_4, H_2O, NH_3$$

Solution

CH_4 has no lone pair

H_2O has two lone pairs

NH_3 has one lone pair

Thus H_2O, NH_3, CH_4

$\xrightarrow{\text{Increase in bond angle}}$

Question 29

Arrange the following entities in order of increasing bond angle between a non-central atom, the central atom and a non –central atom: carbonate ion, carbon tetrachloride molecule, water molecule and hydrogen cyanide molecule.

Solution

$$H_2O\ (104.5°) < CCl_4(109.5°) < CO_3^{2-}(120°) < HCN\ (180°)$$

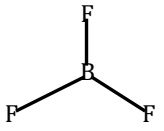
- H_2O and CCl_4 are sp^3-hybridised so their electron pair geometric shape is tetrahedral
  - CCl_4 having lone pair its bond angle is the tetrahedral angle which is $109.5°$
  - H_2O having 2 lone pairs its angle is decreased to $104.5°(109.5° - (2 \times 2.5°)$, keeping in mind that each lone pair decrease the bond angle by $2.5°$)
- The bond angle that would be observed if all electron pairs would be bonded electron pairs).
- CO_3^{2-} is sp^2 hybridised with no lone pair therefore its shape is trigonal planar with the bond angle of $120°$
- HCN is sp-hybridised and therefore its shape is linear with bond angle of 180°

Question 30

On the basis of Sidgwick -Powell theory, describe and explain the shape of BF_3.

Solution

- There are 3 **bonded electron pair without any lone** pair around Boron (B).
  - To minimise repulsion, the three electron pairs are directed to the **corners of an equilateral triangle** in three dimension and hence BF_3 is **trigonal planar in shape** with an angle of $120°$.



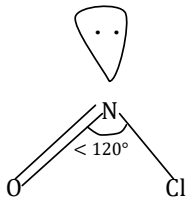
Shape of BF_3

Question 31

On the basis of VSEPR theory, describe and explain the shape of nitrosyl chloride$(NOCl)$ molecule.

Solution

- There are 2 bonded electron pair(double bond is counted as one electron pair) and 1 lone pair around oxygen (O)
- To minimise repulsion, the three electron pairs are directed to the corner of triangle in three dimensions.



Shape of NOCl

- This makes the electron pair geometric shape of the molecule to be trigonal planar. However because lone pair exert greater repulsion than bonding pair of electrons the bond angle is slightly less than 120°.

Question 32

Using VSEPR model, predict molecular geometry (molecular shape) and electron geometry of the following:

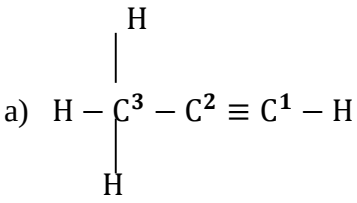
a. CO_3^{2-}
b. ICl_2^-
c. NF_3
d. $SiCl_3^-$

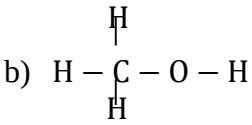
Solution

a) Molecular geometry: Trigonal planar
Electron geometry: Trigonal planar
b) Molecular geometry: Linear
Electron geometry: Trigonal (triangular) pyramidal
c) Molecular geometry : Trigonal (triangular) pyramidal
Electron geometry: Tetrahedral
d) Molecular geometry : Trigonal planar
Electron geometry: Tetrahedral

Question 33

Predict the approximate value for the bond angles for each carbon indicated in the following compounds:

a) 

b) 

Solution

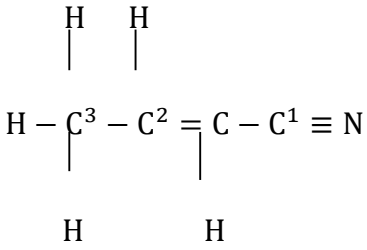
a) $C^1: 180°$: $C^2: 180°$: $C^3 : 109.5°$

Hint: C^1 is sp hybridised, C^2 is sp hybridised, C^3 is sp^3 hybridised

b) $109.5°$ (Hint: The carbon is sp^3 hybridised)

Question 34

Consider the following molecule



For each of the marked carbons

i. Give type of hybridisation
ii. Predict the molecular geometry and electron geometry around each marked carbon.
iii. Predict the bond angles

Solution

i. C^1: sp hybridisation

C^2 : sp^2 hybridisation

C^3 : sp^3hybridisation

ii. C^1 : molecular geometry is linear, electron geometry is linear

C^2 : Molecular geometry is trigonal planar, electron geometry is trigonal planar

C^3 : Molecular geometry is tetrahedral, electron geometry is tetrahedral

iii. $C^1: 180°$

$C^2: 120°$

$C^3: 109.5°$

Question 35

Use VSEPR theory to predict the shape of the following molecules:

a) BrF_5 b) HCN c) SCl_2

Solution

a) Square pyramid b) Linear c) Bent

Question 36

Predict the approximate bond angle in the following:

a) $SiCl_2$ b) IF_4^- c) SCl_4

Solution

a) $120°$ b) $90°$ c) $90°$ and $120°$

Note:

In this question we were asked to give just an approximate value of the angle. So don't worry about small deviation due to presence of lone pair in (a) and (c).

Question 37

CCl_4 is a perfect tetrahedron, but PCl_4^- is distorted tetrahedron. Explain.

Solution

CCl_4 has four electrons pairs about the central C all of which are bonded and therefore they exert equal repulsion and whence tetrahedral shape results. On another hand presence of one lone pair which exerts greater repulsion than the four bonded electron pairs about P of PCl_4^- prevents the formation of tetrahedron for PCl_4^-.

Question 38

Arrange the following species in order of decreasing bond angle:

NH_2^-, NH_3, NH_4^+

Give reason(s) for your arrangement.

Solution

$$NH_4^+ > NH_3 > NH_2^-$$

Decrease in bond angle

Reason

NH_4^+ has tetrahedral shape with 109.5° bond angles. NH_3 has a trigonal pyramidal shape and the lone pair pushes the bonds together making the bond angle less than the tetrahedral angle (bond angle in NH_3 is 107°).NH_2^- has two lone pairs and these push the two bond pairs even closer together than in NH_3 molecule. (With two lone pairs, bond angle in NH_2^- is 104.5°).

Question 39

Magnesium is the hard metal while sodium is soft. Explain

Solution

Magnesium has stronger metallic bond as result of its smaller metallic radius and greater number of valence electrons (Mg has two valence electrons while Na has one), and hence the metal has higher melting point.

Question 40

Between Lithium and potassium which one has higher melting point? Give reasons(s) for your choice.

Solution

Lithium has higher melting point.

Reason

Li being on the top of group I has smaller metallic (atomic) radius leading to its stronger metallic bond and hence the metal has higher melting point.

Question 41

H_2O is liquid while H_2S is gas at room temperature although the molecular weight of H_2S is greater than that of H_2O. Explain

Solution

There is strong hydrogen bonding existing between water molecules as result of higher electronegativity of oxygen than that of sulphur thus making boiling point of water which is above the room temperature higher than that of H_2O which is below.

Question 42

HF is liquid at room temperature while other hydrogen halides are gases. Explain

Solution

F being on the top of halogen group has highest electronegativity so there is strong hydrogen bonding existing between HF molecules while in other hydrogen halides there is a weaker intermolecular force of attraction which is Van der Waals force of attraction and hence boiling point of HF become higher (above the room temperature).

Question 43

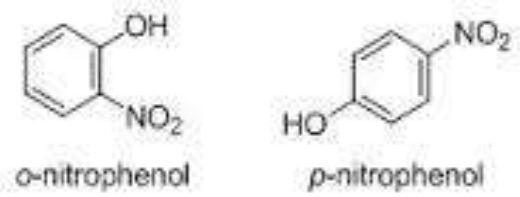
F_2 is gas while HF liquid. Explain

Solution

In HF, there is stronger intermolecular force of attraction which is hydrogen bonding while F_2 being non-polar covalent molecule has van-der-waals force of attraction and hence boiling point of F_2 become lower

Question 44

p-nitrophenol has higher boiling point than o-nitrophenol, explain.



Solution

The first compound has higher boiling point because possess intermolecular hydrogen bonding between its molecules which allows more association than intramolecular hydrogen bonding which exists in another given compound.

Question 45

Can non polar molecules such as CH_4 participate in hydrogen bond? Why or why not?

Solution

Non polar molecules cannot form hydrogen bonds because there is no electronegativity difference between the atoms and hence no partial charges in the atoms.

Question 46

Using your knowledge of the properties of water, explain the quote *"Hydrogen bonds sank the Titanic"*

Solution

Hydrogen bonds sank the titanic in two ways:-

i) Hydrogen bonds caused ice to float in water and therefore the iceberg that was hitted by titanic was formed in the cold water.
ii) Hydrogen bonds caused water molecule to stick together and therefore were able to flow into the ship and whence sinking the titanic.

Question 47

Explain why sodium has a lower melting point than magnesium

Solution

Sodium has weaker metallic bond due to:

- Its fewer valence electron (Na has only one valence electron while Mg has two)
- Its larger metallic radius.

Thus it is easier to weaken metallic bonds between Na atoms and hence it's lower melting point.

Question 48

By referring to their structure and bonding, explain why magnesium oxide has higher melting point than sulfur dioxide.

Solution

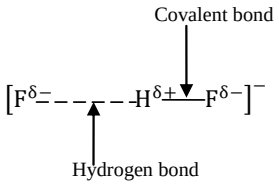
MgO being ionic has stronger intermolecular forces of ionic bonds in ion-ion interaction than Van der Waals forces existing between molecules of SO_2 which is covalent.

Question 49

Why KHF_2 exists while $KHCl_2$ does not?

Solution

F atom being smaller in size accompanied with higher electronegativity than Cl atom, can form very strong hydrogen bonding leading to the formation of HF_2^- in KHF_2 while Cl cannot form strong hydrogen bond so as to form HCl_2^-.



Question 50

Explain why the bond angles in ice are 109° and not 104° like in liquid water?

Solution

In ice, each water molecule forms **two hydrogen bonds** with adjacent water molecule such that **each oxygen** atom is surrounded by **four hydrogen** bonded atoms. Thus in ice there are **four bonded electron pairs** without **any lone pair** around oxygen atom making its shape tetrahedral with bond angle of 109° unlike in liquid water where there are two lone pairs (unbounded electron pairs) decreasing the bond angle to 104°.

Question 51

Carbon dioxide is a gas whereas silicon dioxide is a solid of high melting point. Explain.

Solution

CO_2 is a simple covalent molecule (where π-bond exists) whose molecules are held together by weak Van der Waals dispersion forces while SiO_2 is a giant molecule where single sigma bonds exists whose molecules are held together by very strong giant covalent bonds.

Question 52

Butane, $CH_3CH_2CH_2CH_3$, has higher boiling point than $2-$ methylpropane, $CH(CH_3)_3$. Explain.

Solution

Butane being unbranched has a stronger Van der Waals dispersion forces because:

- It has larger surface area than the branched $2-$ methyl propane

- It chains have more compact packing than the branched $2-$ methylpropane

The two factors make butane to have stronger Van der Waals dispersion forces and hence higher boiling point than $2-$ methyl propane although the two compounds have the same molecular weight.

Question 53

Explain why CO_2 has lower solubility in water?

Solution

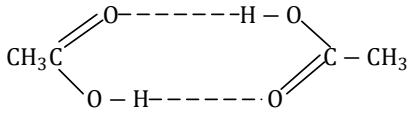
Weaker intermolecular Van der Waals dispersion forces in CO_2 are not able to displace the stronger intermolecular hydrogen bonds in water for hydration and hence the CO_2 become almost insoluble in water.

Question 54

Ethanoic acid,$C_2H_4O_2$, in the gas phase just above its boiling point has an apparent molar mass of $120g/mol$. Explain

Solution

In the vapour phase two molecules of ethanoic acid associate through intermolecular hydrogen bonding to form a dimer.



Dimer of ethanoic acid with Mr. of $120g/mol$

Question 55

State and explain with reasoning the relative solubility of $PbCl_2$ and $PbCl_4$ in water.

Solution

- $PbCl_2$ being ionic (Pb^{2+} in $PbCl_2$ has lower charge and therefore has smaller polarising power making $PbCl_2$ to have lower degree of polarisation) can form strong **ion-dipole forces** with water molecules that result in the **release of energy** to break the giant ionic lattice structure of $PbCl_2$ for hydration to occur. This make$PbCl_2$ more soluble in water.
- $PbCl_4$ being covalent (Pb^{4+} in $PbCl_4$ has higher charge and greater polarising power and hence higher degree of polarisation in $PbCl_4$) has no favourable interaction with water molecules because the weak intermolecular Van der Waals forces in it are not able to displace the stronger intermolecular hydrogen bonds in water. This makes $PbCl_4$ insoluble in water.

Question 56

Arrange the following compounds in order of their boiling point.

HF, HCl, HBr, HF

Solution

$$\text{HF} > HI > HBr > HCl$$

Decrease in their boiling point

Question 57

Which of the following pairs of substance is likely to have higher boiling point? Give reason for your choice.

i. Na and K
ii. C and Si
iii. Ne and Ar
iv. NH_3 and PH_3
v. NaCl and HCl
vi. F_2 and Cl_2

Solution

i. Na
Reason: Na has stronger metallic bond because has small metallic radius

ii. C
Reason: C form stronger giant covalent bond because has small atomic size

iii. Ar
Reason: Ar has a stronger Van der Waals dispersion forces because has greater atomic weight (Greater number of electrons and surface area)

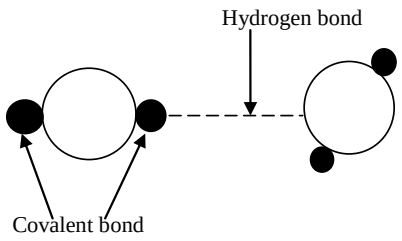
iv. NH_3
Reason: NH_3 has stronger intermolecular hydrogen bond than intermolecular Van der Waals dispersion forces of PH_3.

v. NaCl
Reason: has stronger ion-ion forces (ion bond) than intermolecular Van der Waals dispersion force of HCl.

vi. Cl_2
Reason: Cl_2 has a stronger intermolecular Van der Waals dispersion forces because has greater molecular weight (which means greater surface area and greater number of electrons in Cl_2).

Question 58

Consider the following representation of two water molecules in the liquid state



a. What bonds are broken when water boils? Are these intermolecular or intramolecular bonds?
b. What bonds are broken when water is electrolyzed? Are these intermolecular or intramolecular bonds?

Solution

a) Bond broken: Hydrogen bond
   Bond kind: Intermolecular bond
b) Bond broken :Covalent bond
   Bond kind: Intramolecular bond
   (Remember: Boiling is **physical change** while electrolysis is the **chemical change)**

Question 59

Arrange the following compounds according to increasing solubility in water, least soluble first

a. CH_3COOH
b. CH_3CH_3
c. $CH_3CH_2OCH_2CH_3$

Solution

Both Ar and F_2 being non-polar, their boiling point is determined by strength of Van der Waals dispersion forces. Since the two have the same electronic structure (Each has 18 electrons), they have the same strength of Van der Waals dispersion forces and hence similar boiling point.

Question 60

Explain why argon and fluorine have similar boiling point?

Solution

a) Giant covalent bond
b) London dispersion forces
c) Ion-ion forces (or ionic bonds)
d) Hydrogen bonds
e) London dispersion forces
f) Metallic bonds

Question 61

What kind of intermolecular forces must be broken or overcome in order to:

a) Melt diamond
b) Sublime solid iodine
c) Melt table salt
d) Boil liquid ammonia
e) Melt solid neon
f) Melt iron

Solution

a) Giant covalent bond
b) London dispersion forces
c) Ion-ion forces (or ionic bonds)
d) Hydrogen bonds
e) London dispersion forces

f) Metallic bonds

Question 62

Although CH_3CH_2Cl and $CH_3CH_2CH_2CH_3$ do not differ much in their molecular weights, the boiling point of the former is much higher than that of the later. Explain

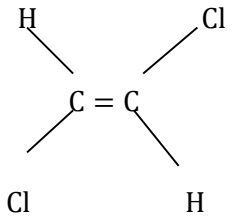
Solution

CH_3CH_2Cl is the polar molecule as the result of large electronegativity difference between carbon and chlorine thus there is electrostatic force of attraction between partial positive charge in carbon and partial negative charge in chlorine ($C^{\sigma+} - Cl^{\sigma-}$) unlike in $CH_3CH_2CH_3$ which is non-polar.

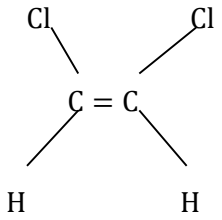
Question 63

Predict whether the following molecules are polar or non-polar

a)



b)



c)SO_2 d) NF_3 e) PF_5

Solution

a) Non polar
b) Polar
c) Polar
d) Polar
e) Non polar

Question 64

Which substance would most likely dissolve in water; CO_2 or CHF_3 why?

Solution

CHF_3 being polar is more likely to dissolve in water which is also polar solvent than CO_2 which is non-polar

Question 65

Which intermolecular forces exist in each of the following?

i. Water
ii. Carbon tetrachloride
iii. Ammonia

Solution

i. London dispersion forces and hydrogen bonds
ii. London dispersion forces

iii. London dispersion forces and hydrogen bonds
iv. London dispersion forces

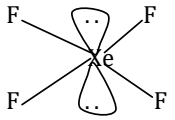
Question 66

a) The shape of the molecule BCl_3 and that of unstable molecule CCl_2 are shown below:

i) Why is each bond angle exactly 120° in BCl_3?
ii) Predict the bond angle in CCl_2 and explain why this angle is different from that in BCl_3.

b) Give the name which describes the shape of molecules having bond angles of $109°\ 28'(109.5°)$. Give an example of one such molecule.

c) The shape of the XeF_4 molecule is shown below:



i) State the bond angle in XeF_4
ii) Suggest why lone pairs of electrons are opposite to each other in the molecule.
iii) Name the shape of this molecule, given that the shape describes the positions of the Xe and F atoms only

Solution

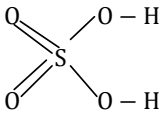
a)
i. Electrons pairs around the central atom (B) are all bonded and therefore they repel equally.
ii. 117.5°
Reason: Lone pair exerts greater repulsion than bonded pairs of electrons.

b) Name of the shape: Tetrahedral, Example:CH_4

c)
i. 90°
ii. Lone pairs repel more strongly than bonded pairs of electrons.
iii. Square planar

Question 67

a) Sulphuric acid is a liquid that can be represented by formula drawn below.

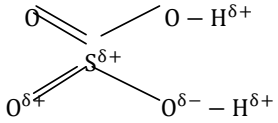


Given that electronegativity value for hydrogen, sulphur and oxygen are 2.1, 2.5 and 3.5 respectively; clearly indicated the polarity of each bond present in the formula given.

a) Name type of hybridisation and molecular shape of the sulphuric acid molecule.
b) Suggest the strongest type of intermolecular force present in pure sulphuric acid. Briefly explain how this type of intermolecular force arises.

Solution

a)



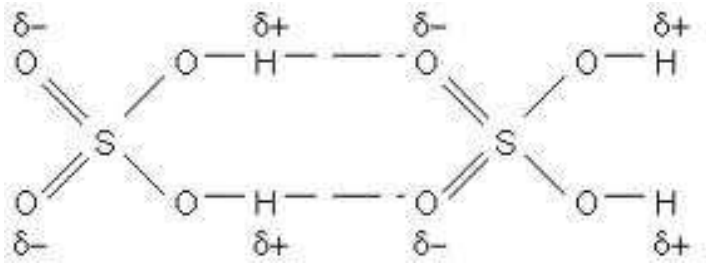
b) Type of hybridisation: sp^3 hybridisation

Molecular shape: Tetrahedral.

a) Hydrogen bond

Explanation:

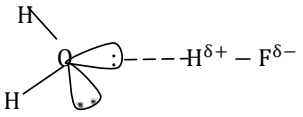
It occurs as result of electrostatic attraction between partial positively charge hydrogen of one molecule and lone pair in partial negatively charge oxygen atom of the adjacent sulphuric acid molecule.



Where ----- stands for hydrogen bonding.

Question 68

The diagram below show how a water molecule interacts with hydrogen fluoride molecule.



a) What is the value of the bond angle in a single molecule of water?
b) Explain your answer to part (a) by using the concept of electron pair repulsion

c) Name the type of interaction between a water molecule and hydrogen fluoride molecule shown in the diagram above.
d) Explain the origin of the δ^+charge shown on the hydrogen atom in the diagram.
e) When water interacts with hydrogen fluoride, the value of bond angle in water change slightly. Predict how the angle is different from that in a single molecule of water and explain your answer.

Solution

a) $104.5°$
b) Water molecule has two lone pairs (in its oxygen atom) which exert stronger repulsion than bonded electron pairs and therefore the bond angle become less than the expected tetrahedral angle of $109.5°$
c) Intermolecular hydrogen bonding.
d) F atom being stronger electronegative, pull the shared electrons from H atom which is weaker electronegative.
   (**Warning!** Don't talks about O atom because there is no δ^+ in H atom of H_2O in the given diagram and the question mention that the δ^+ charge shown on the hydrogen atom in the diagram).
e) Prediction: slight greater than $104.5°$
   Explanation: In presence of HF molecule, lone pairs in O of H_2O interact with $H^{\delta+}$ of HF through hydrogen bonding and therefore **they act like bonded electron pair** which always exert smaller repulsion.

Question 69

Ethene, C_2H_4 and hydrazine, N_2H_4, are hydrides of elements which are adjacent in the periodic table. Data about ethene and hydrazine are given in the table below.

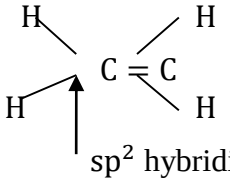
	C_2H_4	N_2H_4
Melting point/℃	-169	+2
Boiling point/℃	-104	+114
Solubility in water	Insoluble	High
Solubility in ethanol	High	High

a) Ethene and hydrazine have a similar arrangement of atoms but differently shape molecules.
   i. What is the $H-C-H$ bond angle in ethene?
   ii. What is the $H-N-H$ bond angle in hydrazine?
   iii. State and explain whether hydrazine is polar or non-polar
b) The melting and boiling points of hydrazine are much higher than those of ethene. Suggest reasons for these differences in term of the intermolecular forces each compound possesses.

c) Explain, with the aid of diagram showing lone pair of electrons and dipoles, why hydrazine is very soluble in ethanol.

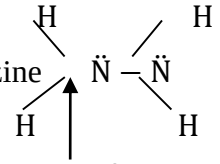
Solution

a) Structure of ethene



sp^2 hybridised without lone pair

Structure of hydrazine



$sp^3 -$ hybridised with 1 lone pair

Therefore:

i) Bond angle of $H - C - H$ is 120° (trigonal planar angle).

ii) Bond angle of $H - N - H$ is 107° (Due to presence of one lone pair the tetrahedral angle of 109.5° is reduced to 107°)

b) Hydrazine is **polar**

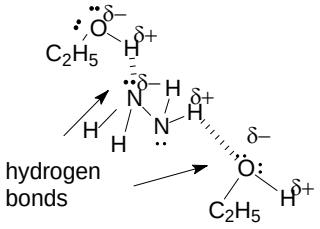
Explanation:

$N - H$ bond is polar and there is 1 lone pair around each sp^3-hybridised N making N_2H_4(hydrazine) trigonal bipyramidal about each N such that there is non-zero resultant of dipole moments and hence the hydrazine becomes polar.

c) More energy is required to overcome the stronger intermolecular hydrogen bonds in hydrazine than the weaker intermolecular Van der Waals dispersion forces in ethene and hence higher melting and boiling point of hydrazine.

d) Hydrazine molecules can form intermolecular hydrogen bonds with ethanol molecules and hence more it become very soluble

- The intermolecular hydrogen bonds between hydrazine and ethanol is illustrated in the diagram below:



Question 70

a.

i) How ions are held together in the solid sodium metal

ii) How ions are held together in solid sodium chloride

b. The melting point of sodium chloride is much higher than that of sodium metal. What can be deduced from this information?

Solution

a.

i) By metallic bond which exist between sodium ions and delocalised valence electrons

ii) By ionic bond which act as the electrostatic force of attraction between sodium cations (Na^+) and chloride anions(Cl^-)

b. Ionic bonds (ion-ion forces) are stronger forces than metallic bonds

Question 71

Name type of hybridisation and molecular shape of

i) NH_4^+

ii) CH_4

What is the relationship between the two?

Solution

i) sp^3 hybridisation, tetrahedral

ii) sp^3 hybridisation, tetrahedral

The two are i**soelectronic.**

Question 72

The water molecule is polar, how does this offer experimental proof that the molecule cannot be linear?

Solution

If H_2O were linear, then there would be no net dipole, and the molecule would be therefore non polar. Since it is polar, the two lone pairs must be on adjacent sides of the tetrahedral electron geometry, leaving the two bonds on adjacent sides. This results in a net dipole with the negative end near the oxygen atom, and the positive end near the two hydrogen atoms.

Question 73

Give reasons for the following observations;

i. The boiling point of water, ethanol and ethoxyethane (diethyl ether) are in the reverse order of their relative molecular masses unlike those of their analogous sulphur compounds H_2S, C_2H_5SH and $C_2H_5SC_2H_5$.

ii. Aluminium fluoride has much higher melting point than aluminium chloride.

Solution

i. For given first group of compounds boiling point is determined by the strength of hydrogen bonding which increases from ethoxyethane to water, where water has strongest hydrogen bonding and therefore highest boiling point. For the second group of compounds, boiling point is determined by Van-der Waals dispersion forces of attraction which increase with an increase in molecular weight; so H_2S having lowest molecular weight has lowest boiling while $C_2H_5SC_2H_5$ having highest molecular weight has highest boiling point.

ii. AlF_3 is more ionic in characters as result of its lower degree of polarisation brought by smaller polarisability of smaller sized F^-.

Question 74

Arrange the following bonds according to polarity.

$$H-H, O-H, Cl-H, S-H, F-H, N-H$$

Solution

$$H-H < S-H < Cl-H < N-H < O-H < F-H$$

Increase in bonds polarity

Question 75

Explain the following observation:

i. PCl_5 is more reactive than PCl_3.
ii. $AlCl_3$ is covalent while AlF_3 is ionic.

Solution

(i) PCl_5 has higher energy electronic structure and therefore less stable as it is formed by sp^3d hybridisation which has higher energy than sp^3 hybridisation which is used in forming PCl_3.

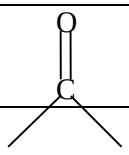
(ii) $AlCl_3$ has higher degree of polarisation as result of greater polarisability power of Cl^- which has larger size than F^- of AlF_3 and hence according to Fajan's rule, $AlCl_3$ becomes more covalent in characters.

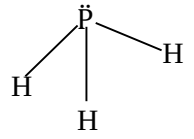
Question 76

By using modern molecular theory complete the table on the following molecules.

Molecule	**Geometrical structure**	**Name of the structure**	**hybridisation**
CO_2			
CH_2O			
PH_3			

Solution

Molecule	**Geometrical structure**	**Name of the structure**	**hybridisation**
CO_2	$O=C=O$	Linear	sp hybridisation
CH_2O			

	H H	Trigonal planar	sp^2 hybridisation
PH_3		Triangular pyramid	sp^3 hybridisation

Question 77

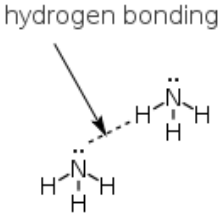
a. State the bond angle in NH_3 and BF_3 and state the name of the shape of each molecule.
b. Explain why NH_3 has polar molecules where as molecules of BF_3 are non-polar.
c. Polar molecules of NH_3 form hydrogen bonds. Draw diagram to show this hydrogen bonding.
d. Explain why the $H-N-H$ bond angle in NH_3 is less than that in NH_4^+

Solution

a) Bond angles:
107° in NH_3
120° in BF_3
Shape of molecules:
Triangular pyramid for NH_3
Trigonal planar for BF_3

b) Due to presence of lone pair in N, NH_3 is triangular pyramid shaped which is unsymmetrical structure while BF_3 having no lone pair has symmetrical structure in its trigonal planar shape and hence NF_3 is polar and BF_3 is non – polar.

c)



d) Due to presence of lone pair in N of NH_3 which exert stronger repulsion than bonding pair of electrons, $H-N-H$ bond angle is less than tetrahedral angle of NH_4^+ which has no lone pair.

Question 78

The boiling points of water, hydrogen chloride and argon are shown in the table below.

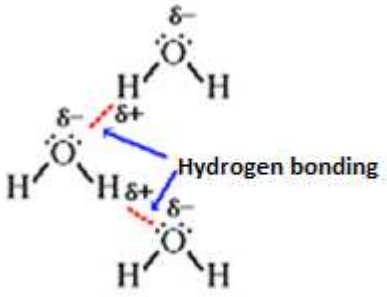
Substance	H_2O	HCl	Ar

Boiling point /℃	100	−85	−186
Total number of electrons	10	18	18

a. H_2O , HCl, Ar all have Van der Waals forces. Outline how Van-der-Waals force arise between molecules.
b. Liquid H_2O has additional intermolecular forces
   i) What are these forces?
   ii) Explain with the aid of diagram, how these forces arise between molecules of H_2O.
c. Liquid HCl also has additional intermolecular forces. What are these forces.
d. Explain the variation in boiling point shown in the above table.

Solution

a) Hint: explain how London dispersion forces occur (Remember when you are asked just about "Van der Waals forces", the implication is always the London dispersion forces)
b)
   i. Hydrogen bond
   ii. Due to electronegativity difference between H and O (oxygen is more electronegative than H), the $O-H$ bond is polar with partial positive charge in the hydrogen atom. Thus the partial positively charged H of one molecule of H_2O attracts the lone pair of partial negatively charged oxygen atom of neighbour H_2O molecule and whence hydrogen bond is formed.



c) Dipole –dipole forces
d) On additional to Van der Waals dispersion forces, water molecules are held together by hydrogen bonds which are stronger intermolecular forces than dipole-dipole forces which present in HCl in the addition to the Van der Waals dispersion forces. That is why water has highest boiling point among the three. Argon has lowest boiling point because has Van der Waals dispersion forces only.

Question 79

a. The table below shows boiling points of fluorine, fluoromethane (CH_3F) and hydrogen fluoride.

Compound	F_2	CH_3F	HF
Boiling point/K	85	194	293

Name the strongest type of intermolecular force present in;-

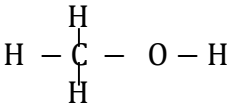
i. Liquid F_2
ii. Liquid CH_3F
iii. Liquid HF
iv. Explain how the strongest type of intermolecular force in liquid HF arises

b. The table below shows the boiling points of some other hydrogen halides.

Compound	HCl	HBr	HI
Boiling point/K	188	206	238

i. Explain the trend in the boiling points of the hydrogen halides from HCl to HI.
ii. Give one reason why the boiling point of HF is higher than that of all the other hydrogen halides.

c.
i. Methanol has the structure

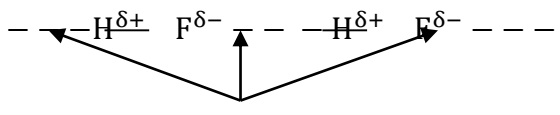


Explain why $O-H$ bond in a methanol molecule is polar?

ii. The boiling point of methanol is$+65℃$, the boiling point of oxygen is $-183℃$. Methanol and oxygen each has M_r value of 32. Explain why the boiling point of methanol is much higher than that oxygen.

Solution

a)
i. London dispersion forces
ii. London dispersion forces
iii. Hydrogen bonds
iv. H being highly less electronegative than F is partially positively charged. This hydrogen positive pole of one HF molecule attracts lone pair of partially negatively charged F of the neighbour HF molecule resulting to the hydrogen bonding.



Hydrogen bonding

b)

i. The strongest type of intermolecular forces in the given hydrogen halides is Van der Waals (London) dispersion forces whose strength increases with an increase in the molecular weight

- HI having greatest molecular weight has strongest of Van der Waals dispersion forces and therefore highest boiling point while HCl having lowest molecular weight has weakest forces.

ii. There are very strong hydrogen bonds holding molecules of HF together.

c)

i. O being more electronegative than H has partial negative charge while H has partial positive charge making the bond polar.

There are hydrogen bonds holding molecules of methanol together, those hydrogen bonds are stronger intermolecular forces than Van der Waals dispersion forces present in the oxygen

Question 80

a. Below are elements of group VI A their atomic numbers and symbols.

Element	Atomic number	Symbol
Oxygen	8	O
Sulphur	16	S
Selenium	34	Se
Tellurium	52	Te

i. Give full electronic distribution in selenium

ii. Can sulphur form S=S? Give reason(s) to support your answer.

iii. The hydrides of group VIA elements are H_2O, H_2S, H_2Se and H_2Te. Arrange the hydrides in order of increasing their boiling point starting with the least volatile giving clear reasons for your arrangement.

b.

i. Suggest why the strength of the $C-H$ bond in CH_4 is greater than that of the $Si-H$ bond in SiH_4. State the relationship if any, between the strength of covalent bond in CH_4 and the boiling temperature of CH_4 and hence state which one (CH_4 or SiH_4) has greater boiling point.

c. State the type of hybridisation shown by the nitrogen atoms in:

i. N_2

ii. N_2H_2

iii. N_2H_4

Solution

a)

i. It is more than easier for you (Don't forget to use boxes and arrow because you asked to give full electronic distribution)

ii. No

Reason:

Atomic size of sulphur (or size of p-orbitals in sulphur) is too large to enable side-way overlapping of atomic orbitals so as to form pi- bond in the double bond.

iii. H_2O, H_2Te, H_2Se, H_2S

Reason

- H_2O has hydrogen bonds which are stronger intermolecular forces than Van der Waals dispersion forces of other molecules.
- For the other molecules, H_2Te has greatest molecular weight and therefore strongest Van der Waals dispersion forces while H_2S has the least one.

b)

i.

- C **being smeller is size than Si** can form stronger covalent bond in $C-H$ than Si in $Si-H$.
- There is no relation between strength of covalent bond of CH_4 and its boiling point (because the covalent bond is the interatomic forces which has no effect on the physical properties like boiling).The boiling point of CH_4 and SiH_4 depends on the strength of Van der Waals dispersion intermolecular forces which increases with an increase in their molecular weight and hence SiH_4 has greater boiling point (because SiH_4 has greater molecular weight than CH_4).

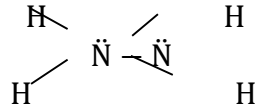
c)

i) $\ddot{\ddot{N}} \equiv \ddot{N}$

Type of hybridisation: sp hybridisation

ii) $H-\ddot{N}=\ddot{N}-H$

Type of hybridisation: sp^2 hybridisation



Types of hybridisation: sp^3 hybridisation

Question 81

a. Consider shape of the molecule of BCl_3 and that of unstable molecule CCl_2
   i. Why is each bond angle exactly $120°$ in BCl_3?
   ii. Predict the bond angle in CCl_2 and explain why this angle is different from that in BCl_3?

b. When considering electron pair repulsions in molecules, why does a lone pair of electrons repel more strongly than a bonding pair?

c. Co-ordinate bonding can be described as **dative covalency**.

i. In the above context, what is the meaning of each of terms **covalency** and **dative**?
ii. Write an equation for a reaction in which a coordinate bond is formed

d. Arrange the following elements in order of increasing their melting points giving clear reason(s) for your trend.

Neon, Sodium, Magnesium, Aluminium and silicon.

Solution

a) (i) In BCl_3, all three pairs of electrons are bonded electrons which repel equally so they get as far as possible in trigonal planar arrangement with bond angle of $120°$ so as to minimise repulsion.
(ii) Predicted bond angle is about $118°$.
Explanation:
This is because in CCl_2 there is a lone pair in C which repels more than bonding pair of electrons making the bonds relatively closer than in BCl_3 where all pair of electrons are bonded.

b) Lone pair are more compact (very close) than bonding pair

c)

i) Covalency meaning: shared electron pair
Dative meaning: both electrons comes from one atom

Neon $< sodium < Magnesium < Aluminium < Silicon$

Reason:

- Neon has Van der Waals dispersion forces which are very weak intermolecular forces.
- Sodium, magnesium and aluminium have metallic bonds (which are stronger than Van der Waals dispersion forces of Neon).
  - With only one valence electron accompanied with large atomic size, sodium has weakest metallic bond while aluminium has strongest one.
- Silicon has giant covalent bonds which are stronger intermolecular forces than metallic bonds.

Question 82

For each of the following species, identify type of hybridisation of an element in the bracket, **name** electron pair geometric shape, molecular geometric shape and bond angle.

i. X_eF_4 (Xe)
ii. H_2SO_4 (S)
iii. I_3^- (I)

Solution

i. Type of hybridisation: sp^3d^2 hybridisation
Electron pair geometric shape: Octahedral
Molecular geometric shape: Square planar
Bond angle: $90°$

ii. Type of hybridisation: sp^3 hybridisation
Electron pair geometric shape: Tetrahedral
Molecular geometric shape: Tetrahedral
Bond angle: $109.5°$

iii. Types of hybridisation : sp^3d hybridisation
Electron pair geometric shape: Trigonal bipyramidal
Molecular geometric shape: Linear
Bond angle: $180°$

Question 83

i. Predict and explain the polarity of the bonds within BF_3 and NF_3.
ii. State whether BF_3 and NF_3 are polar molecules. Explain your answer.

Solution

i.
B – F is a polar bond, because F, is more electronegative than B;
N – F is a polar bond, because F is more electronegative than N

ii.

- BF_3 is a non-polar molecule because of its symmetrical distribution of electron pairs in its trigonal planar shape makes it to have zero resultant of dipole moment.
- NF_3 is a polar molecule due to its unsymmetrical distribution of electron pair in its trigonal pyramidal shape as result of presence of lone pair. This makes NF_3 to have non-zero resultant of the dipole moment.

Question 84

What type of intermolecular force is found between:

(i) F_2 and Cl_2
(ii) Two water molecules

Solution

(i) London dispersion forces (or temporary dipoles-induced dipole forces)
(ii) hydrogen bonds.

Question 85

Although sodium and potassium are metals, they can never be used in construction of bridges; explain.

Solution

The given metals have very weak metallic bond as result of their large metallic radii accompanied with only one valence electron in each thus making them soft accompanied with low density and high solubility in water and hence the metals become unsuitable for making bridges.

Question 86

Starting with the shortest one, arrange the following bonds in order of their bond length

(1) (2) (3)

i. $CH_2 = CH - CH_2 - CH_3$

(1) (2) (3)

ii. $CH \equiv C - CH = CH_2$

(1) (2) (3)

iii. $CH_2 = C = CH - CH_3$

(1) (2) (3) (4) (5)

iv. $CH_2 = C = C = CH - C \equiv CH$

Solution

The bond which involves hybridised orbitals with greater percentage of s-characters will be shorter than the bond with hybridised orbitals of smaller percentage of s-characters.

i.

1. Is the bond with $sp^2 - sp^2$ carbon
2. Is the bond with $sp^2 - sp^3$ carbons
3. Is the bond with $sp^3 - sp^3$ carbons

Hence the order becomes:

$$(1) < (2) < (3)$$

ii.

1. Is the bond with $sp - sp$ hybridised carbons
2. Is the bond with $sp - sp^2$ hybridised carbons
3. Is the bond with $sp^2 - sp^2$ hybridised carbons

Hence the order becomes:

$$(1) < (2) < (3)$$

iii.

1. Is the bond with $sp^2 - sp$ hybridised carbons
2. Is the bond with $sp - sp^2$ hybridised carbons
3. Is the bond with $sp^2 - sp^3$ hybridised carbons

Hence the order becomes:

$$(1) = (2) < (3)$$

iv.

1. Is the bond with $sp^2 - sp$ hybridised carbons
2. Is the bond with $sp - sp$ hybridised carbons

3. Is the bond with $sp - sp^2$ hybridised carbons
4. Is the bond with $sp^2 - sp$ hybridised carbons
5. Is the bond with $sp - sp$ hybridised carbons

Hence the order becomes:

$$(2) = (5) < (1) = (3) = (4)$$