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CHEMISTRY

PHYSICAL CHEMISTRY

Level & Board	EDEXCEL (A-LEVEL)
TOPIC:	BONDING & STRUCTURE
PAPER TYPE:	SOLUTION - 4
TOTAL QUESTIONS	10
TOTAL MARKS	34

Bonding & Structure - 4

1.

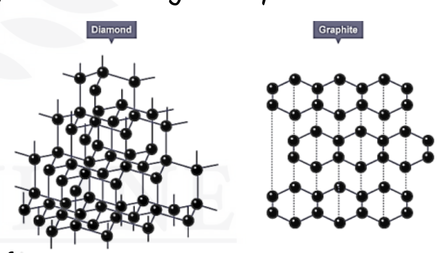
Diamond and graphite are both allotropes of carbon, meaning they are different structural forms of the same element. The significant difference in their properties, including melting temperatures, arises from their distinct crystal structures.

Graphite:

- *In graphite, carbon atoms are arranged in layers of hexagonal arrays. Each carbon atom is bonded to three others in a flat, two-dimensional plane.*
- *The layers in graphite are held together by weak van der Waals forces, allowing them to easily slide over each other. This gives graphite its lubricating properties and softness.*
- *The weak forces between layers make it easier to break the bonds and transition from a solid to a liquid, resulting in a relatively low melting temperature.*

Diamond:

- *In diamond, each carbon atom forms strong covalent bonds with four other carbon atoms, creating a three-dimensional tetrahedral structure.*
- *The strong, covalent bonds in diamond make the crystal structure exceptionally stable and rigid.*
- *The requirement to break strong covalent bonds to transition from a solid to a liquid gives diamond a much higher melting temperature compared to graphite.*



So, Diamond's three-dimensional network of strong covalent bonds requires much more energy to break compared to the weaker forces between the layers in graphite, leading to the significantly higher melting temperature of diamond.

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2.

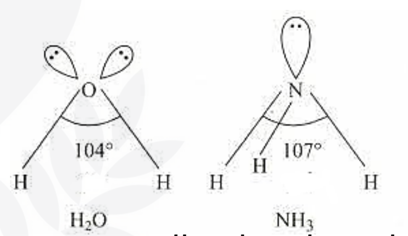
The difference in bond angles between water (H_2O) and ammonia (NH_3) is due to the structural differences in their molecular geometries and the repulsion between electron pairs. As:

Water ($H_2\,O$):

- *In a water molecule, oxygen is bonded to two hydrogen atoms. Oxygen has two lone pairs of electrons in addition to the two sigma bonds formed with hydrogen.*
- *The presence of lone pairs creates electron-electron repulsion, pushing the hydrogen atoms slightly closer together.*
- *This lone pair repulsion causes a distortion in the bond angles, leading to a reduction in the angle between the hydrogen atoms. The bond angle in water is approximately 104.5 degrees.*

2. **Ammonia (NH_3):**

- *In an ammonia molecule, nitrogen is bonded to three hydrogen atoms. Nitrogen has one lone pair of electrons in addition to the three sigma bonds formed with hydrogen.*
- *Similar to water, the lone pair in ammonia introduces electron-electron repulsion. However, since ammonia has three hydrogen atoms (as opposed to water's two), the repulsion is distributed over a larger area.*
- *The larger number of atoms around the central nitrogen atom in ammonia mitigates the impact of the lone pair repulsion, resulting in a larger bond angle compared to water. The bond angle in ammonia is approximately 107 degrees.*



So, the lone pairs in water create more significant electron-electron repulsion, leading to a smaller bond angle, while the larger number of atoms in ammonia helps distribute the repulsion, resulting in a larger bond angle.

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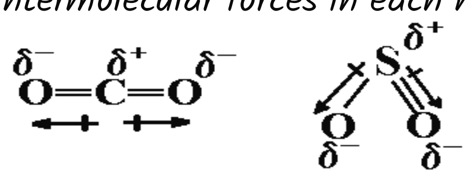
3.

The difference in boiling temperatures between carbon dioxide (CO_2) and sulfur dioxide (SO_2) can be explained by the nature and strength of the intermolecular forces in each molecule.



Carbon Dioxide (CO_2):

- *CO_2 is a linear molecule with a carbon atom double-bonded to two oxygen atoms.*

- The primary intermolecular force in CO_2 is London dispersion forces (Van der Waals forces).
- Although there are polar covalent bonds within the molecule due to the electronegativity difference between carbon and oxygen, the linear geometry results in a symmetric distribution of charge, leading to no permanent dipole moment.
- The London dispersion forces in CO_2 are relatively weak compared to other intermolecular forces.

Sulfur Dioxide (SO_2):

- SO_2 is a bent or V-shaped molecule with a sulfur atom double-bonded to two oxygen atoms.
- Like CO_2 , SO_2 also has London dispersion forces.
- Additionally, SO_2 has polar covalent bonds due to the electronegativity difference between sulfur and oxygen.
- The bent geometry of SO_2 results in an asymmetric distribution of charge, creating a permanent dipole moment.
- The presence of both London dispersion forces and dipole-dipole interactions makes the intermolecular forces in SO_2 stronger than those in CO_2 .

So, the distinction in boiling temperatures between CO_2 and SO_2 is explained by the differences in intermolecular forces. SO_2 has stronger intermolecular forces due to the presence of both London dispersion forces and dipole-dipole interactions, leading to a higher boiling temperature compared to CO_2 .

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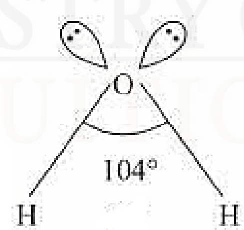
4.

The diagram of the Water molecule, showing its shape lone pairs of electrons and the value of the bond angle.



H_2O

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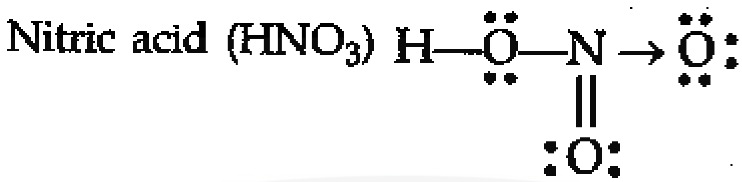
5.

The dot-and-cross diagram of a molecule of Nitric acid HNO_3 is as:



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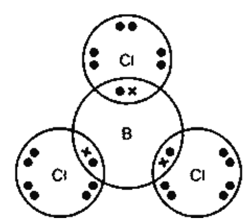
6.

Lewis Structure of BCl_3 :



- Boron has three valence electrons, and chlorine has seven valence electrons each. The Lewis structure for BCl_3 is formed by bonding one boron atom with three chlorine atoms, utilizing a single bond for each pair of shared electrons.

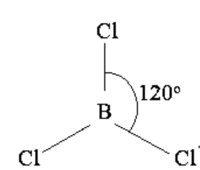
VSEPR Theory:

- According to the VSEPR (Valence Shell Electron Pair Repulsion) theory, the electron pairs around the central atom arrange themselves in a way that minimizes repulsion and maximizes the distance between them.
- Boron in BCl_3 has an incomplete octet and forms three sigma bonds with three chlorine atoms.

Trigonal Planar Geometry:

- With no lone pairs on the central boron atom, the molecular geometry of BCl_3 is trigonal planar.
- The three chlorine atoms are arranged symmetrically around the boron atom, forming a flat, triangular shape.

Bond Angles:



- The bond angles in BCl_3 are approximately 120 degrees. This is because the trigonal planar arrangement maximizes the separation between the bonding pairs, minimizing electron-electron repulsion.

So, the trigonal planar shape of BCl_3 with bond angles of 120 degrees is a result of the three sigma bonds formed between boron and chlorine atoms, following the principles of VSEPR theory to minimize electron pair repulsion and achieve a stable molecular geometry.

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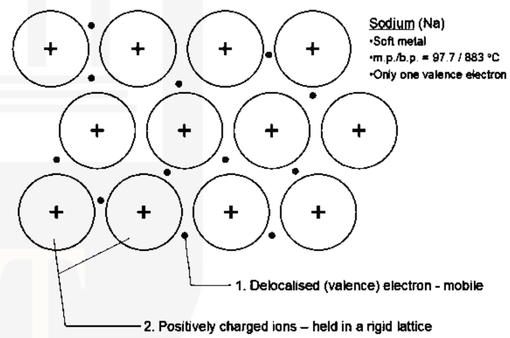
7.

Metallic bonding is a type of chemical bonding that occurs between metal atoms. It is a unique bonding mechanism that imparts several distinctive properties to metals. Here's an explanation of metallic bonding:

Electron Delocalization:



- *In metallic bonding, metal atoms contribute their valence electrons to form a "sea of electrons" that is delocalized and can move freely throughout the entire metallic structure.*
- *Unlike in ionic or covalent bonds where electrons are localized between specific pairs of atoms, metallic bonding involves a shared pool of electrons that is not associated with any individual metal atom.*

Positive Metal Ions (Cations):

- *As a result of contributing their valence electrons, metal atoms become positively charged ions or cations.*
- *The delocalized electrons move around these cations, creating a kind of electron cloud that surrounds the positively charged metal ions.*

Attraction Between Electrons and Cations:

- *The attraction between the negatively charged electrons and the positively charged metal ions is what holds the metallic lattice together.*
- *This attraction is often referred to as the metallic bond, and it is a strong force that gives metals their characteristic properties.*

Properties of Metallic Bonding:

- ***Conductivity***
- ***Malleability and Ductility***
- ***Luster:***

Examples of Metals with Metallic Bonding:

- *Common examples of metals with metallic bonding include copper, iron, aluminum, and gold.*

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8.

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons in a chemical compound. It is a relative scale, with elements assigned electronegativity values based on their ability to attract electrons. The Pauling scale is commonly used for this purpose.
In HCl (hydrochloric acid) molecule:
Chlorine (Cl) is more electronegative than hydrogen (H).
The electronegativity difference between chlorine and hydrogen in HCl results in a polar covalent bond.
Polarity of the H-Cl bond:

Chlorine attracts the shared electron pair more strongly than hydrogen, causing the electron cloud to be skewed towards chlorine.
This creates a partial negative charge (δ-) on the chlorine atom and a partial positive charge (δ+) on the hydrogen atom.
The H-Cl bond in hydrochloric acid is polar, with the chlorine end being more negative and the hydrogen end being more positive.

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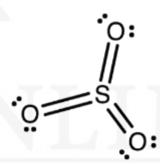
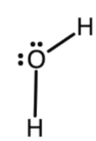
9.

Both water (H_2O) and sulfur trioxide (SO_3) have polar bonds, but only water is a polar molecule because:





Water (H_2O):
- In a water molecule, oxygen is more electronegative than hydrogen, leading to polar covalent bonds.
- Water has a bent or V-shaped molecular geometry.
- The two O-H bonds create a dipole moment, and the molecule as a whole has a net dipole moment.
- The dipole moments do not cancel each other out due to the asymmetry of the molecule, resulting in a polar molecule.

Sulfur Trioxide (SO_3):
- In sulfur trioxide, sulfur is more electronegative than oxygen, creating polar covalent bonds between sulfur and oxygen atoms.

- SO_3 has a trigonal planar molecular geometry.
- Although each S-O bond has a dipole moment, the overall molecule is symmetrical.
- The dipole moments in sulfur trioxide are arranged such that they cancel each other out, resulting in a molecule with no net dipole moment.
- Due to the symmetry and cancellation of dipole moments, sulfur trioxide is a nonpolar molecule.

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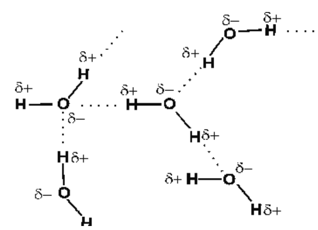
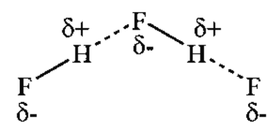
10.

The differences in boiling temperatures between HF (hydrogen fluoride, boiling point 19.0°C) and H_2O (water, boiling point 100.0°C) are due to the strength of intermolecular forces.



HF has dipole-dipole interactions and less hydrogen bonding per molecule, while water, with its ability to form extensive hydrogen bonds i.e two per molecule, has significantly stronger intermolecular forces, leading to a much higher boiling temperature.



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