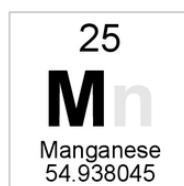
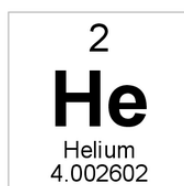
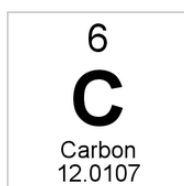
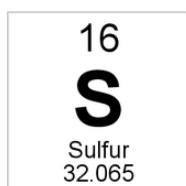
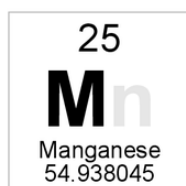


Structure 2.2

Answers

IB CHEMISTRY SL



Structure 2.2.1 and 2.2.2

Understandings:

- A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei (2.2.1).
- The octet rule refers to the tendency of atoms to gain a valence shell with a total of 8 electrons (2.2.1).
- Single, double and triple bonds involve one, two and three shared pairs of electrons respectively (2.2.2).

Learning outcomes:

- Deduce the Lewis formula of molecules and ions for up to four electron pairs on each atom (2.2.1).
- Explain the relationship between the number of bonds, bond length and bond strength (2.2.2).

Additional notes:

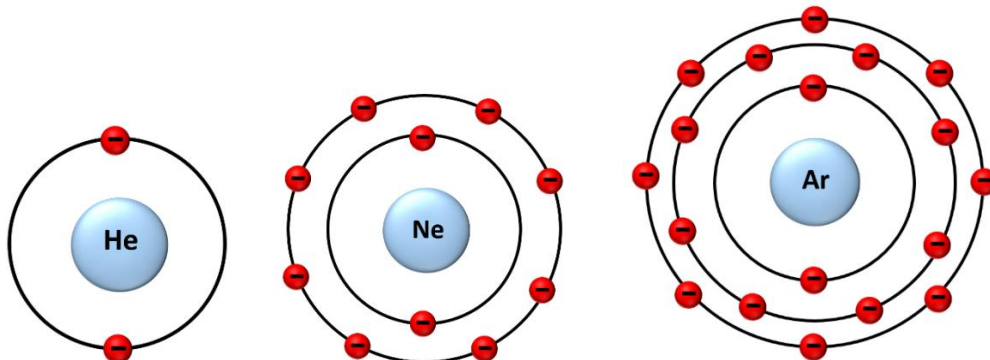
- Lewis formulas (also known as electron dot or Lewis structures) show all the valence electrons (bonding and non-bonding pairs) in a covalently bonded species.
- Electron pairs in a Lewis formula can be shown as dots, crosses or dashes.
- Molecules containing atoms with fewer than an octet of electrons should be covered.
- Organic and inorganic examples should be used.

Linking questions:

- Structure 1.3 Why do noble gases form covalent bonds less readily than other elements?
- Structure 2.1 Why do ionic bonds only form between different elements while covalent bonds can form between atoms of the same element?
- Reactivity 2.2 How does the presence of double and triple bonds in molecules influence their reactivity?

The octet rule

- The octet rule states that atoms are stable with eight electrons in their valence shells with the electron configuration of a noble gas.
- The electron configurations of noble gases He, Ne and Ar are shown below.



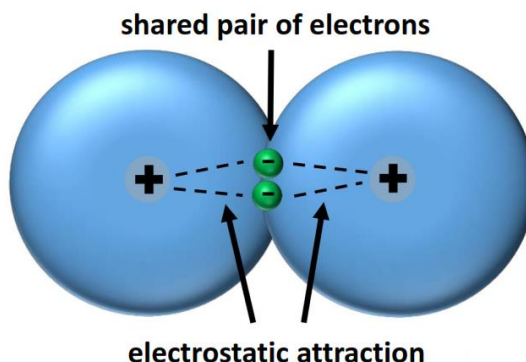
- Noble gases are stable because they have full valence shells.
- Atoms can achieve the electron configuration of a noble gas by either sharing electrons (covalent bonding) or by losing or gaining electrons (ionic bonding).

Exceptions to the octet rule

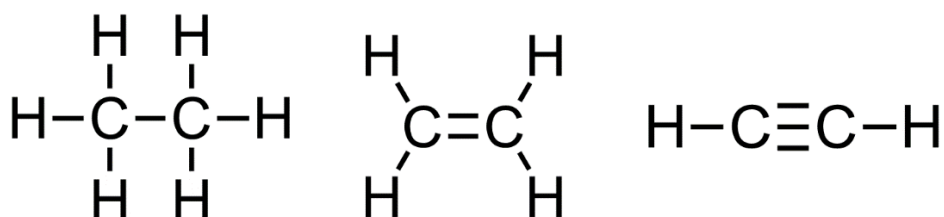
- Hydrogen and helium are stable with 2 electrons in their valence shells.
- Beryllium is stable with 4 electrons in its valence shell.
- Boron is stable with 6 electrons in its valence shell.
- Elements in period 3 onwards can hold more than 8 electrons in their outer shells (expanded octets).

Covalent bonding

- A covalent bond is the electrostatic attraction between positive nuclei and shared pairs of bonding electrons.
- A covalent bond is an intramolecular bond (bond within a molecule).
- Covalent bonding occurs between non-metal elements only.
- The electronegativity difference between atoms is between 0.0 and 1.7 units.



- Both nuclei are attracted to the shared pair of bonding electrons.
- The electrostatic attraction between the nuclei and shared pair of bonding electrons is the covalent bond.
- **Single, double and triple bonds**
- Single bonds are longer and weaker, triple bonds are shorter and stronger.



- The carbon-to-carbon bonds in C_2H_6 are longer and weaker than those in C_2H_2 .

Bond	Number of shared electrons	C to C bond strength (kJ mol^{-1})	C to C bond length (10^{-12} m)
Single	2	347	153
Double	4	614	134
Triple	6	839	120

Exercises:

1. Describe a covalent bond.

A covalent bond is the electrostatic attraction between positive nuclei and a shared pair of bonding electrons.

2. Where on the periodic table would you find elements that bond by covalent bonding?

Covalent bonding occurs between non-metal elements which are found on the right side of the periodic table.

3. Arrange the following in order of increasing carbon to carbon bond strength (weakest first).

C₂H₄, C₂H₆, C₂H₂

C₂H₆, C₂H₄ C₂H₂

4. Arrange the following in order of decreasing carbon to carbon bond length (longest first),

C₂H₄, C₂H₆, C₂H₂

C₂H₆, C₂H₄ C₂H₂

Bond order

- Bond order is the number of bonds between a pair of atoms.
- Single bonds have a bond order of 1, double bonds have a bond order of 2 and triple bonds have a bond order of 3.
- The higher the bond order, the stronger (and shorter) the bond.
- Polyatomic ions such as CO₃²⁻ and NO₃⁻ and molecules such as benzene, C₆H₆, can have fractional bond orders.
- Bond order can be calculated by dividing the sum of the individual bond orders by the number of bonding groups in the molecule or ion.

Example: Calculate the bond order of the bonds in the nitrate ion, NO₃⁻.

The nitrate ion has two single bonds and one double bond. The sum of the bond orders is 4. There are three bonding groups, so the bond order is $4/3 = 1.3$

Lewis formulas (Lewis structures)

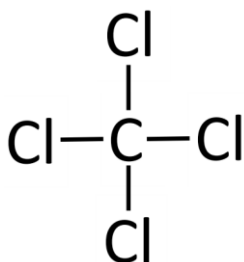
- Lewis formulas show all the valence electrons in a molecule; the bonding electrons and the lone pairs of electrons (non-bonding electrons).

How to determine a Lewis formula

- Calculate the number of valence electrons in the molecule.
- Calculate the number of electrons each atom needs to complete its octet.
- Subtract 1 from 2 – this will give you the number of bonding electrons in the molecule.
- Draw the skeletal structure of the molecule with the least electronegative atom at the center.
- Complete the octets of the atoms in the molecule.

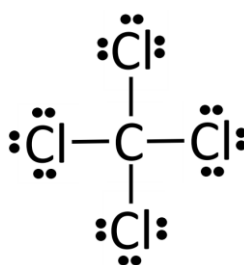
Example – CCl₄ (carbon tetrachloride)

- Total number of valence electrons = $4 + (4 \times 7) = 32$
- Number of valence electrons needed for each atom to complete its octet (5×8) = 40
- $40 - 32 = 8$ bonding electrons
- Skeletal structure:



One single bond = 2 electrons (total 8 bonding electrons).

- Complete the octets of all the atoms in the molecule.



Do a final check that all electrons are accounted for. It should match with the number in part 1 (32).

Exercises:

1. Draw Lewis formulas for the following molecules.

a) Cl_2

b) HCl

c) CH_3Cl

d) CF_4

e) PF_3

f) BF_3

g) BeCl_2

h) N_2H_4

i) H_2O

j) CH_3OH

k) CO

l) C_2H_4

j) HCN

2. Draw Lewis formulas for the following ions.

a) Cl^-

b) NH_4^+

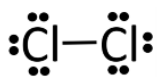
c) OH^-

3. Which molecules contain an incomplete octet of electrons?

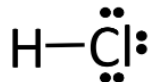
Answers

1.

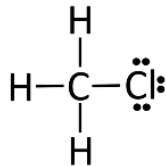
Cl_2



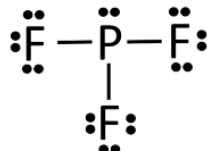
HCl



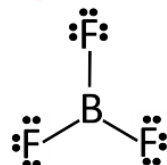
CH_3Cl



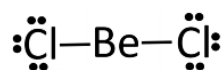
PF_3



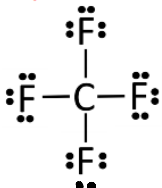
BF_3^+



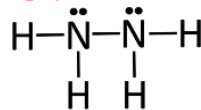
BeCl_2^+



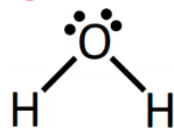
CF_4



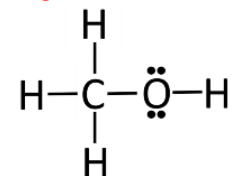
N_2H_4



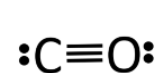
H_2O



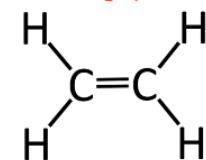
CH_3OH



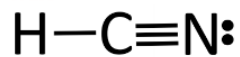
CO



C_2H_4



HCN

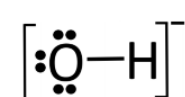


2)

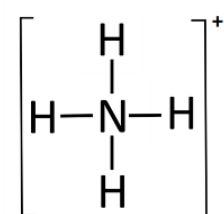
Cl^-



OH^-



NH_4^+



3. BeCl_2 and BF_3 both contain an incomplete octet of electrons.

Structure 2.2.3

Understandings:

- A coordination bond is a covalent bond in which both the electrons of the shared pair originate from the same atom.

Learning outcomes:

- Identify coordination bonds in compounds.

Additional notes:

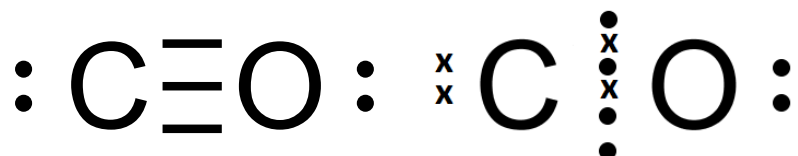
- Include coverage of transition element complexes (HL).

Linking questions:

- Reactivity 3.4 (HL) Why do Lewis acid–base reactions lead to the formation of coordination bonds?

Coordination bonds

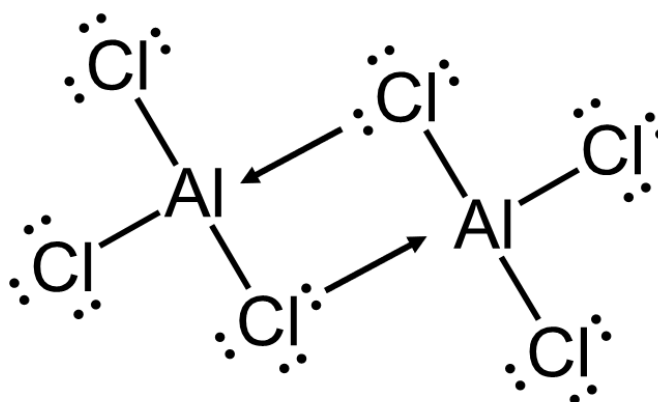
- A coordination bond is a covalent bond in which one atom contributes both the bonding electrons to the bond.
- Carbon monoxide (CO) has a triple covalent bond between the carbon and oxygen atoms – one of these bonds is a coordination bond where both the bonding electrons come from the oxygen atom.



- Other molecules and ions that have coordination bonds are H_3O^+ , NH_4^+ and Al_2Cl_6 .
- Draw the Lewis structures of CO, H_3O^+ and NH_4^+ in the boxes below.

CO	H_3O^+	NH_4^+
$:\text{C}\equiv\text{O}:$	$\left[\begin{array}{c} \cdot\cdot \\ \text{H}-\text{O}-\text{H} \\ \\ \text{H} \end{array} \right]^+$	$\left[\begin{array}{c} \text{H} \\ \\ \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array} \right]^+$

- Al_2Cl_6 is a dimer formed between two molecules of AlCl_3 .



Structure 2.2.4

Understandings:

- The valence shell electron pair repulsion (VSEPR) model enables the shapes of molecules to be predicted from the repulsion of electron domains around a central atom.

Learning outcomes:

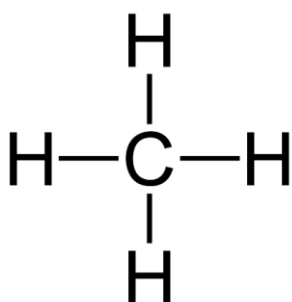
- Predict the electron domain geometry and the molecular geometry for species with up to four electron domains.

Additional notes:

- Include predicting how non-bonding pairs and multiple bonds affect bond angles.

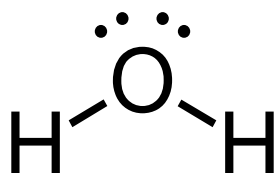
VSEPR theory

- Valence shell electron pair repulsion theory is used to predict the geometry (shape) of molecules.
- Electron pairs (bonds or lone pairs) repel each other and spread apart as far as possible.
- The term electron domain is used to refer to bonds or lone pairs of electrons (non-bonding electrons) around an atom in a molecule.
- Single bonds, double bonds, triple bonds and lone pairs of electrons (non-bonding electrons) count as one electron domain.



How many electron domains are there around the carbon atom?

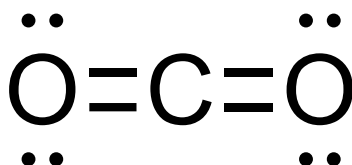
4 bonding domains



How many electron domains are there around the oxygen atom?

2 bonding domains

2 non-bonding domains



How many electron domains are there around the carbon atom?

2 bonding domains (a double bond counts as one electron domain)

The order of repulsion between electron domains is as follows:

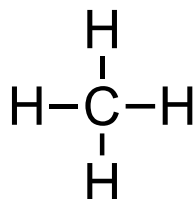
lone pair – lone pair > lone pair – bonding domain > bonding domain – bonding domain

most repulsion

least repulsion

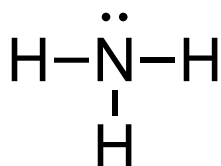
Molecules with four electron domains

- Four bonding domains, zero non-bonding domains



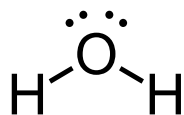
electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
4	4	0	tetrahedral	tetrahedral	109.5°

- Three bonding domains, one non-bonding domain



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
4	3	1	tetrahedral	trigonal pyramidal	107.8°

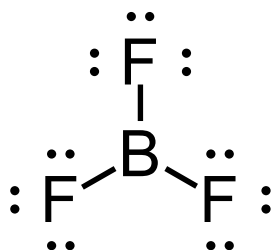
- Two bonding domains, two non-bonding domains



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
4	2	2	tetrahedral	bent	104.5°

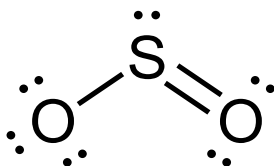
Molecules with three electron domains

- Three bonding domains, zero non-bonding domains



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
3	3	0	trigonal planar	trigonal planar	120°

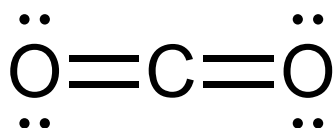
- Two bonding domains, one non-bonding domain



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
3	2	1	trigonal planar	bent	<120°

Molecules with two electron domains

- Two bonding domains, zero non-bonding domains



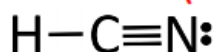
electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
2	2	0	linear	linear	180°

Exercise: Determine the molecular geometry and bond angles in the following molecules:

- | | |
|------------------------------------|----------------------------------|
| a) HCN | e) C ₂ H ₂ |
| b) PH ₃ | f) BF ₃ |
| c) CH ₂ Cl ₂ | g) BeH ₂ |
| d) Cl ₂ O | h) H ₂ S |

Answers:

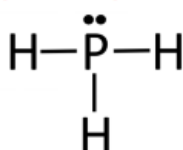
HCN – linear (180°)



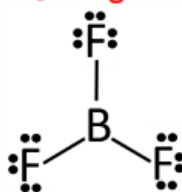
C₂H₂ – linear (180°)



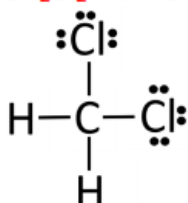
**PH₃ – trigonal pyramidal
(<107.8°)***



BF₃ – trigonal planar (120°)



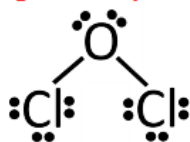
CH₂Cl₂ – tetrahedral (109.5°)



BeH₂ – linear (180°)



Cl₂O – bent/v-shaped (111°)*



H₂S – bent/v-shaped (<104.5°)*



* The reasons for these bond angles are complex and are beyond the scope of the standard level IB course. It is more important that you can correctly predict the shape of the molecule in these cases.

Structure 2.2.5

Understandings:

- Bond polarity results from the difference in electronegativities of the bonded atoms.

Learning outcomes:

- Deduce the polar nature of a covalent bond from electronegativity values.

Additional notes:

- Bond dipoles can be shown either with partial charges or vectors.
- Electronegativity values are given in the data booklet.

Linking questions:

- Structure 2.1 What properties of ionic compounds might be expected in compounds with polar covalent bonding?

Polar covalent bonds

- Polar covalent bonds occur between atoms that have a difference in electronegativity.

Difference in electronegativity	Polar or non-polar covalent bond	Example
0	non-polar (pure) covalent bond	Cl-Cl
0.1– 0.4	non-polar (weakly polar) covalent bond	C-H
0.5 –1.7	polar covalent bond	C-F

- Bond polarity can be shown by a vector arrow or by δ^+ and δ^- as shown below.



Exercises:

- Draw a diagram showing the bonding polarity in HF (hydrogen fluoride).

H-F with the partial negative charge on the F atom and the partial positive charge on the H atom.

2. For each bond, find the difference in electronegativity and classify as either pure covalent, non-polar covalent or polar covalent. Draw bond dipoles or partial charges on each atom.

Bond	Pure covalent, non-polar covalent or polar covalent	Bond dipoles
Br-Br	Pure covalent Difference in EN = 0	No bond dipole
C-Cl	Polar covalent Difference in EN = 0.6	Partial negative charge on Cl and partial positive charge on C
C-I	Weakly polar covalent Difference in EN = 0.1	Partial negative charge on I and partial positive charge on C
C-O	Polar covalent Difference in EN = 0.8	Partial negative charge on O and partial positive charge on C
N-F	Polar covalent Difference in EN = 1.0	Partial negative charge on F and partial positive charge on N
H-F	Polar covalent Difference in EN = 1.8	Partial negative charge on F and partial positive charge on H

Structure 2.2.6

Understandings:

- Molecular polarity depends on both bond polarity and molecular geometry.

Learning outcomes:

- Deduce the net dipole of a molecule or ion by considering bond polarity and geometry.

Additional notes:

- Examples should include species in which bond dipoles do and do not cancel each other.

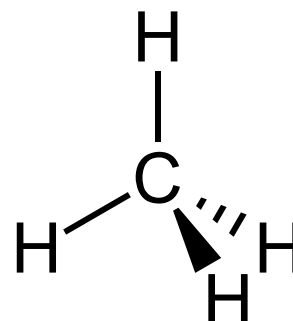
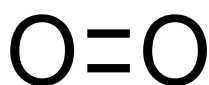
Linking questions:

- Structure 3.2 (HL) What features of a molecule make it “infrared (IR) active”?

Polar and non-polar molecules

The polarity of a molecule depends on two factors:

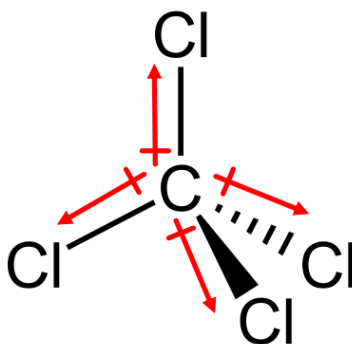
- The presence of polar bonds within the molecule.
- The shape (geometry) of the molecule.



- O_2 , N_2 and CH_4 are non-polar molecules.
- O_2 and N_2 are diatomic; both atoms have the same electronegativity.
- CH_4 has non-polar bonds.

Non-polar molecule with polar bonds

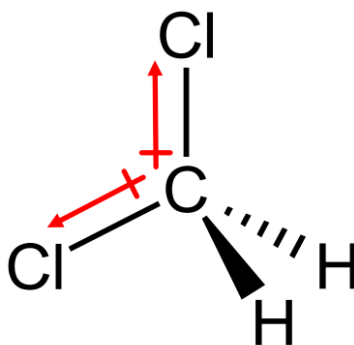
- CCl_4 is a non-polar molecule that has polar bonds.



- The bond polarities cancel out, therefore, the molecule has no net-dipole moment.

Polar molecules

- CH_2Cl_2 is a polar molecule.



- The bond polarities do not cancel out, therefore, it has a net-dipole moment.

Exercises

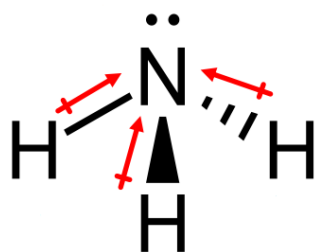
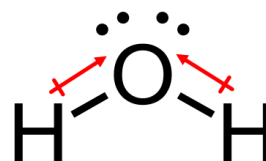
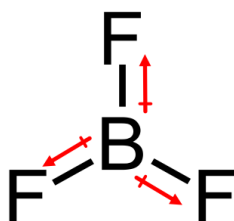
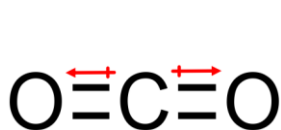
1. Outline the two factors that determine the polarity of a molecule.

The two factors that determine the polarity of a molecule are the presence of polar bonds within the molecule and the geometry (shape) of the molecule. Polar bonds are caused by unequal sharing of electrons due to a difference in electronegativity between the atoms in the bond.

2. Explain how a molecule can have polar bonds but overall have no net-dipole moment (be a non-polar molecule).

A molecule can have polar bonds but be non-polar overall if the bond polarities cancel out. This is due to the shape of the molecule, for example, CO_2 has two polar bonds and a linear shape – the bond dipoles cancel out leaving the molecule with no net-dipole moment.

3. Determine if the molecules below are polar or non-polar, giving a reason.



CO_2 is non-polar as the bond dipoles cancel out because of linear molecular geometry.

BF_3 is non-polar as the bond dipoles cancel out because of trigonal planar molecular geometry.

H_2O is polar as bond dipoles do not cancel out because of bent molecular geometry.

NH_3 is polar as bond dipoles do not cancel out because of trigonal pyramidal molecular geometry.

Structure 2.2.7

Understandings:

- Carbon and silicon form covalent network structures.

Learning outcomes:

- Describe the structures and explanation of the properties of silicon, silicon dioxide and carbon's allotropes: diamond, graphite, fullerenes and graphene.

Additional notes:

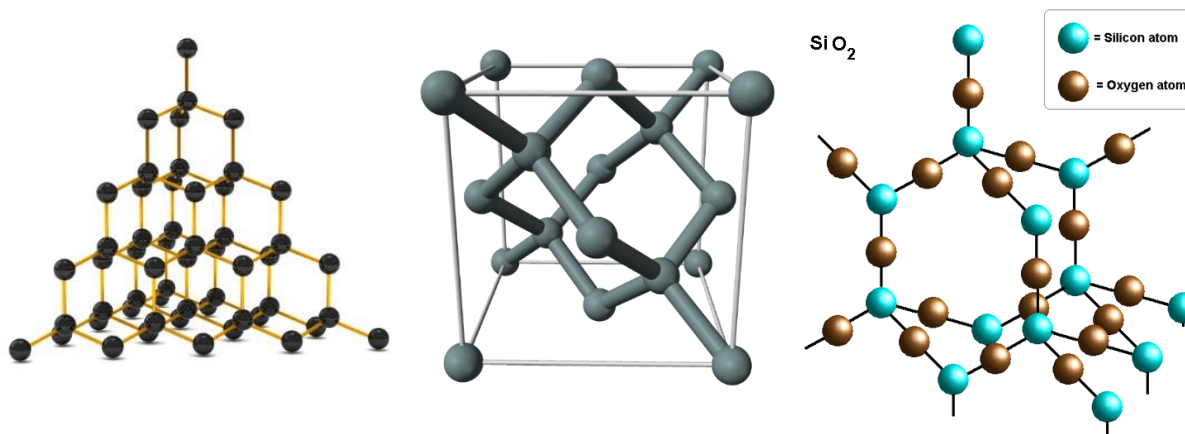
- Allotropes of the same element have different bonding and structural patterns, and so have different chemical and physical properties.

Linking questions:

- Structure 3.1 Why are silicon–silicon bonds generally weaker than carbon–carbon bonds?

Giant covalent structures

- Giant covalent structures do not form discrete molecules.
- Examples include diamond, Si and SiO₂



- Giant covalent structures do not conduct electricity (no free moving electrons).
- They are insoluble in polar and non-polar solvents.
- They have high melting and boiling points due to the strong covalent bonds between atoms.
- They are very hard substances.

Exercises:

1. Classify the following covalent compounds as simple molecular or giant covalent based on their properties:

Property	Simple molecular	Giant covalent
high melting and boiling point		x
insoluble in polar and non-polar solvents		x
liquids and gases under standard conditions	x	
do not form discrete molecules		x
poor electrical conductors	x	x
soluble in polar or non-polar solvents	x	
form discrete molecules	x	
very hard substances		x

2. Explain the following properties of the covalent compounds.

- a. Pentane (C₅H₁₂) is insoluble in water but soluble in hexane (C₆H₁₄).

Pentane is a non-polar molecule and is therefore soluble in a non-polar solvent such as hexane.

- b. Diamond does not conduct electricity.

Diamond does not have any free moving electrons in its structure.

- c. CO₂ is a gas under standard conditions.

The molecules of CO₂ have weak London dispersion forces between them.

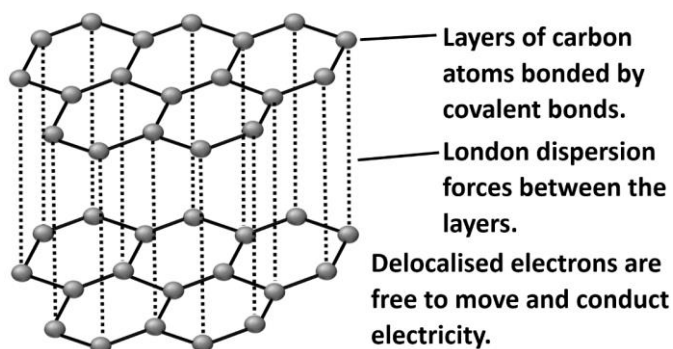
- d. SiO₂ has a melting point of 1600 °C.

There are strong covalent bonds between the silicon and oxygen atoms.

Allotropes of carbon

- Allotropes are different forms of the same element in the same physical state.
- Carbon has 4 allotropes – graphite, diamond, Fullerene C₆₀ and graphene.
- Different bonding within the structures gives the allotropes different properties.

Graphite



- Graphite has a layered structure.
- The layers are held together by weak intermolecular forces.
- The layers can slide over one another (because of weak intermolecular forces).
- Each carbon atom is bonded to 3 other carbon atoms.
- The bond angle between carbon atoms is 120°, trigonal planar.
- Graphite is a good conductor of electricity because of the delocalised electrons that are free to move within the structure.

Diamond



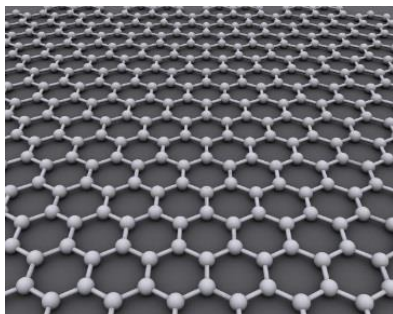
- Giant covalent structure.
- High melting and boiling point.
- Very hard (strong covalent bonds between atoms).
- Each carbon is bonded to 4 other carbon atoms.
- Bond angle is 109.5°, tetrahedral.
- Does not conduct electricity (no delocalised electrons).
- Used in jewellery and for cutting glass.

Fullerene C₆₀

- Each carbon atom is bonded to 3 other carbon atoms.
- Structure consists of 12 pentagons and 20 hexagons.
- Poor electrical conductor (better than diamond but worse than graphite).



Graphene



- Each carbon atom is bonded to 3 other carbon atoms.
- Bond angle between carbon atoms is 120° .
- Very good heat and electrical conductivity.
- Very thin (one layer thick) but very strong.

Allotropes of carbon and their uses

Allotrope	Uses
Graphite	dry lubricant, electrode rods, pencils
Diamond	jewellery, tools for cutting glass
Fullerene C ₆₀	lubricant, nanotubes
Graphene	lightweight, thin, flexible, yet durable display screens, electric/photonic circuits, solar cells

Exercises:

1. Outline why graphite is a good conductor of electricity but diamond is not.

Graphite has delocalised electrons within its structure whereas diamond does not.

2. Give one reason for the high melting and boiling point of diamond.

The carbon atoms in diamond are bonded by strong covalent bonds.

3. State and explain the bond angles of graphite and diamond.

Diamond – tetrahedral (109.5°). Each carbon is bonded to four other carbon atoms with four bonding domains around each carbon atom.

Graphite – trigonal planar (120°). Each carbon is bonded to three other carbon atoms with three electron domains around each carbon atom.

4. Explain why graphite is a very soft substance.

The layers in graphite are held together by weak intermolecular forces (London dispersion forces). Therefore, the layers can slide over each other relatively easily.

Structure 2.2.8 and 2.2.9

Understandings:

- The nature of the force that exists between molecules is determined by the size and polarity of the molecules.
- Intermolecular forces include London (dispersion), dipole-induced dipole, dipole–dipole and hydrogen bonding.
- Given comparable molar mass, the relative strengths of intermolecular forces are generally: London (dispersion) forces < dipole–dipole forces < hydrogen bonding.

Learning outcomes:

- Deduce the types of intermolecular force present from the structural features of covalent molecules.
- Explain the physical properties of covalent substances to include volatility, electrical conductivity and solubility in terms of their structure.

Additional notes:

- The term “van der Waals forces” should be used as an inclusive term to include dipole–dipole, dipole- induced dipole, and London (dispersion) forces.
- Hydrogen bonds occur when hydrogen, being covalently bonded to an electronegative atom, has an attractive interaction on a neighbouring electronegative atom.

Linking questions:

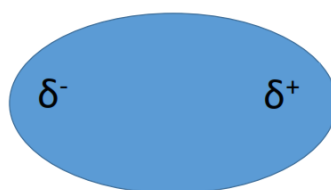
- Structure 1.5 To what extent can intermolecular forces explain the deviation of real gases from ideal behaviour?
- Structure 3.2 To what extent does a functional group determine the nature of the intermolecular forces?

Intermolecular forces

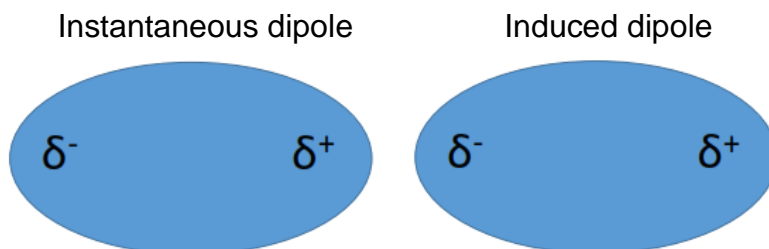
- Intermolecular forces are forces of attraction between molecules.
- The three types of intermolecular force are:
London dispersion forces, dipole-dipole forces and hydrogen bonding.
- London dispersion forces and dipole-dipole forces are collectively known as van der Waals forces.

London dispersion forces

- London dispersion forces are the weakest type of intermolecular force.
- They exist between all atoms or molecules (both polar and non-polar).
- Due to the constant motion of electrons, an atom or molecule can develop a temporary (instantaneous) dipole.



- An instantaneous dipole in one molecule can cause an induced dipole in a nearby molecule.



- London dispersion forces have a significant effect on the physical properties of molecules, such as boiling point.

Exercises:

1. State the types of intermolecular forces that are van der Waals forces.
London dispersion and dipole-dipole attractions.
2. Outline the formation of London dispersion forces.
London dispersion forces are made up of instantaneous and induced dipoles.

Boiling point of the halogens

	Molar mass (g mol ⁻¹)	Boiling point (°C)
F ₂	30.8	-188
Cl ₂	70.9	-34.0
Br ₂	160	58.0
I ₂	254	193

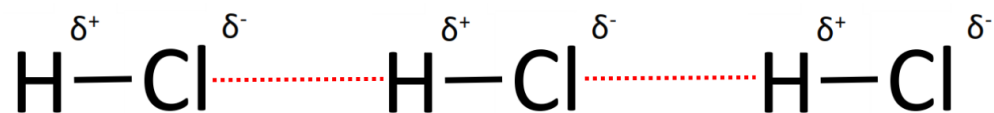
- The boiling points of the halogens increase as their molar masses increase.
- The strength of London dispersion forces increases with increasing molar mass.
- Stronger London dispersion forces result in a higher boiling point (more energy is required to overcome the attractive forces between molecules).

Exercise: Explain why the boiling points of the halogens increase down the group.

The molar mass of the halogens increases down the group which results in stronger London dispersion forces between the molecules and an increase in the boiling point.

Dipole-dipole forces

- Dipole-dipole forces occur between polar molecules that have a permanent dipole.
- They are the second strongest type of intermolecular force.
- There is an electrostatic attraction between the partial positive charge in one molecule and the partial negative charge on another.
- The dipole-dipole attractions between HCl molecules are shown below.



Exercises:

1. State the type of molecules that dipole-dipole forces occur between.

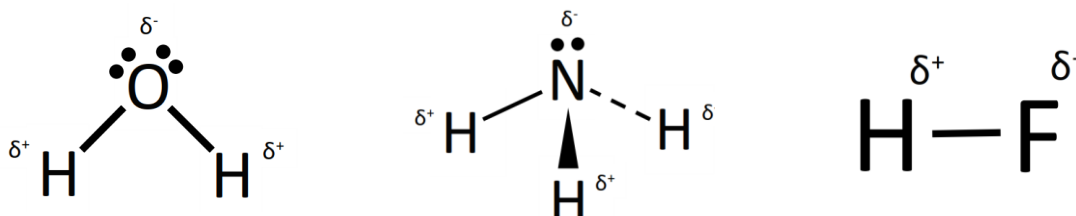
Polar molecules.

2. Outline the formation of the dipole-dipole forces in HCl

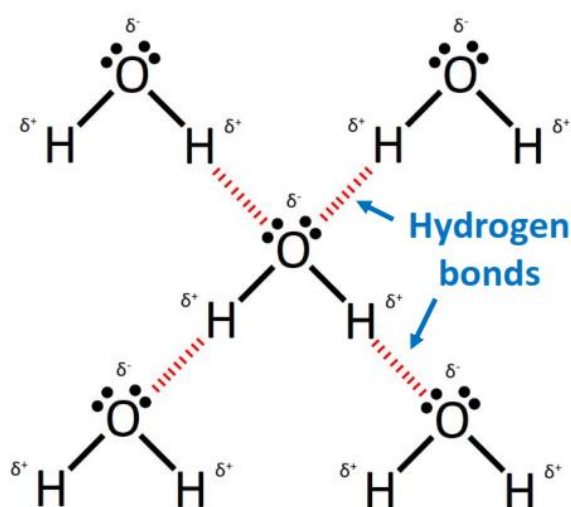
The difference in electronegativity between chlorine and hydrogen produces a partial negative charge on the chlorine atom and a partial positive charge on the hydrogen atom. The opposite charges on different molecules attract each other.

Hydrogen bonding

- Hydrogen bonding occurs when a hydrogen atom is bonded to either a nitrogen, oxygen or fluorine atom.
- It is the strongest type of intermolecular force.
- Example of compounds that have hydrogen bonding between their molecules are H_2O , NH_3 and HF .



- Hydrogen bonds between water molecules are shown below. These are responsible for the high boiling point of water.

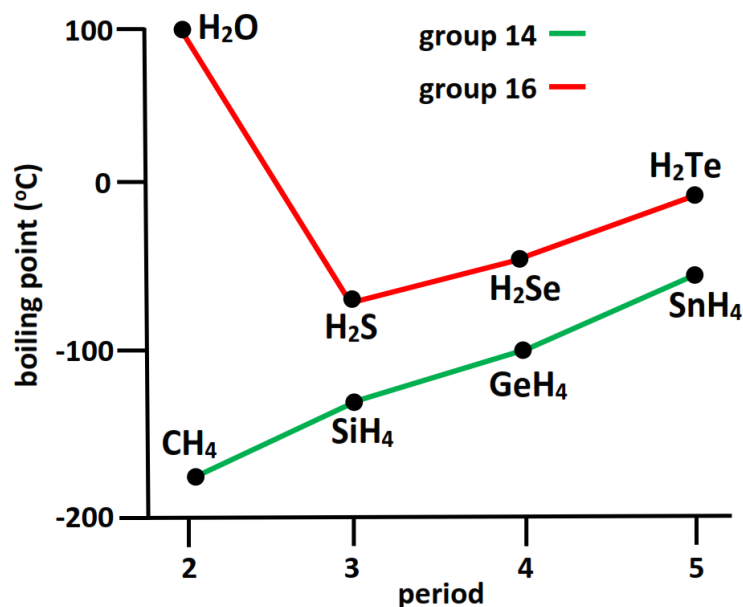


- The hydrogen bond is between the partial positive charge on the hydrogen atom and the partial negative charge on the oxygen atom
- Water has a much higher boiling point compared to other molecules with similar molar masses because of the effect of hydrogen bonding.

Exercises:

1. Which groups of atoms are necessary for the formation of a hydrogen bond?
O-H, N-H, F-H
2. Outline the formation of a hydrogen bond between water molecules.
The hydrogen bond is formed between the lone pair of electrons on the oxygen atom of one molecule and the partial positive charge on the hydrogen atom of another molecule.

Boiling points of group 14 and group 16 hydrides



Exercise: From the above graph, describe and explain the trend in boiling point of the group 14 and group 16 hydrides.

H₂O has the highest boiling point of the group 16 hydrides due to hydrogen bonding between its molecules. The boiling point decreases to H₂S and then starts to increase again as the molar mass of the remaining hydrides increases which results in stronger London dispersion forces between the molecules and a higher boiling point.

Of the group 14 hydrides, CH₄ has the lowest molar mass and therefore the lowest boiling point. As the molar mass increases, the strength of the London dispersion forces between the molecules also increases, which results in a higher boiling point.

Exercises:

1. Arrange the following in terms of increasing strength.

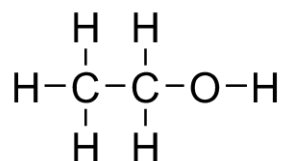
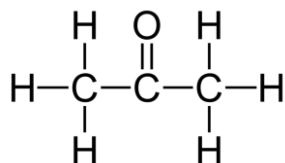
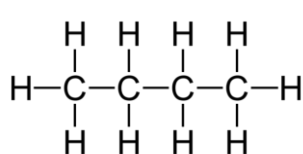
Hydrogen bonding London dispersion forces Dipole-dipole forces

2. Identify the **strongest** type of intermolecular forces in the following molecules.

- | | |
|---------------------|----------------------------------|
| a) Cl ₂ | f) CH ₃ Cl |
| b) HCl | g) H ₂ O |
| c) HF | h) CH ₃ OH |
| d) CH ₄ | i) C ₂ H ₆ |
| e) CCl ₄ | j) NH ₃ |

3. Explain why, at room temperature, F₂ and Cl₂ are gases, Br₂ is a liquid and I₂ is a solid.

4. The structures of butane, propanone, and ethanol can be seen below. They have similar molar masses, but different boiling points. Explain the difference in boiling point between the three molecules.



Answers:

1) London dispersion forces < Dipole-dipole forces < Hydrogen bonding

2) Non-polar molecules have London dispersion forces, polar molecules have dipole-dipole forces and molecules with H-O, H-N or H-F have hydrogen bonding as the **strongest** intermolecular force between molecules.

- | | |
|--|---|
| a) Cl ₂ London dispersion forces | f) CH ₃ Cl Dipole-dipole forces |
| b) HCl Dipole-dipole forces | g) H ₂ O Hydrogen bonding |
| c) HF Hydrogen bonding | h) CH ₃ OH Hydrogen bonding |
| d) CH ₄ London dispersion forces | i) C ₂ H ₆ London dispersion forces |
| e) CCl ₄ London dispersion forces | j) NH ₃ Hydrogen bonding |

3) F₂, Cl₂, Br₂ and I₂ are non-polar molecules, therefore they have London dispersion forces between molecules. The molar mass increases from F₂ to I₂, therefore the strength of the London dispersion forces also increases. Stronger London dispersion forces result in a higher boiling as more energy is required to overcome the forces between the molecules.

4) Butane is a non-polar molecule therefore it has London dispersion forces between molecules. These are the weakest type of intermolecular force, therefore it has the lowest boiling point.

Propanone is a polar molecule (due to the polar C=O bond) therefore it has dipole-dipole forces between molecules. They are stronger than London dispersion forces, therefore it has a higher boiling point than butane.

Ethanol has an OH group (O bonded to H) which means that it can form hydrogen bonds between molecules. Hydrogen bonding is the strongest type of intermolecular force, therefore ethanol has the highest boiling point.

Structure 2.2.10

Understandings:

- Chromatography is a technique used to separate the components of a mixture based on their relative attractions involving intermolecular forces to mobile and stationary phases.

Learning outcomes:

- Explain, calculate and interpret the retardation factor values, RF.

Additional notes:

- The use of locating agents is not required.
- The operational details of a gas chromatograph or high-performance liquid chromatograph will not be assessed.