

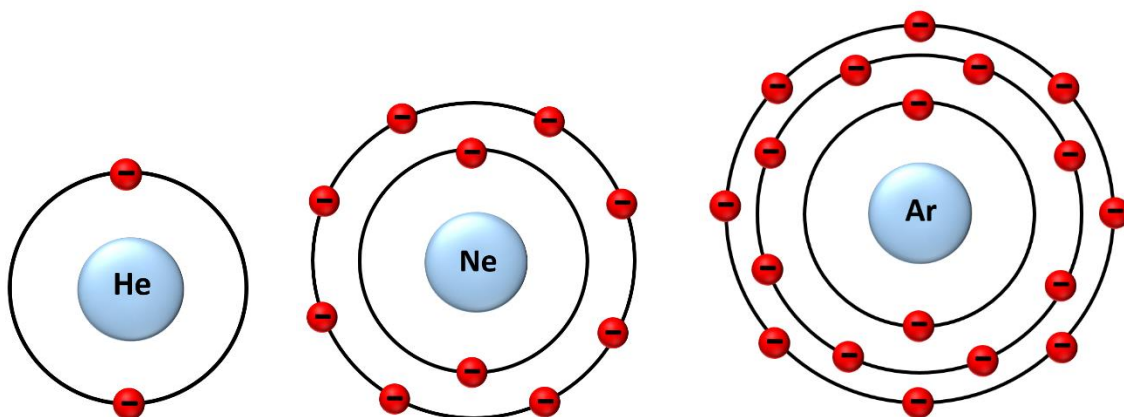
Bonding SL (answers)

IB CHEMISTRY SL

<div>25</div> <div>Mn</div> <div>Manganese 54.938045</div>	<div>16</div> <div>S</div> <div>Sulfur 32.065</div>	<div></div> <div>J</div> <div></div>	<div>6</div> <div>C</div> <div>Carbon 12.0107</div>	<div>2</div> <div>He</div> <div>Helium 4.002602</div>	<div>25</div> <div>Mn</div> <div>Manganese 54.938045</div>
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The octet rule

- The octet rule states that atoms lose or gain electrons to achieve the electron configurations of noble gases.
- Electron configurations of noble gases He, Ne and Ar are shown below.



- Noble gases are stable because they have full outer main energy levels.
- 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 outer energy level.
- Beryllium is stable with 4 electrons in its outer energy level.
- Boron is stable with 6 electrons in its outer energy level.
- Elements in period 3 onwards can hold more than 8 electrons in their outer shells (expanded octets).

Electronegativity and bonding

- Electronegativity is a measure of the attraction of an atom for a bonding pair of electrons.
- It increases across a period due to increasing nuclear charge, and decreases down a group as atomic radius increases.

Difference in electronegativity	Type of bonding
0–0.4	non-polar covalent
0.5–1.7	polar covalent
≥ 1.8	ionic

4.1 Ionic bonding

Understandings:

- Positive ions (cations) form by metals losing valence electrons.
- Negative ions (anions) form by non-metals gaining electrons.
- The number of electrons lost or gained is determined by the electron configuration of the atom.
- The ionic bond is due to the electrostatic attraction between oppositely charged ions.
- Under normal conditions, ionic compounds are usually solid with lattice structures.

Applications and skills:

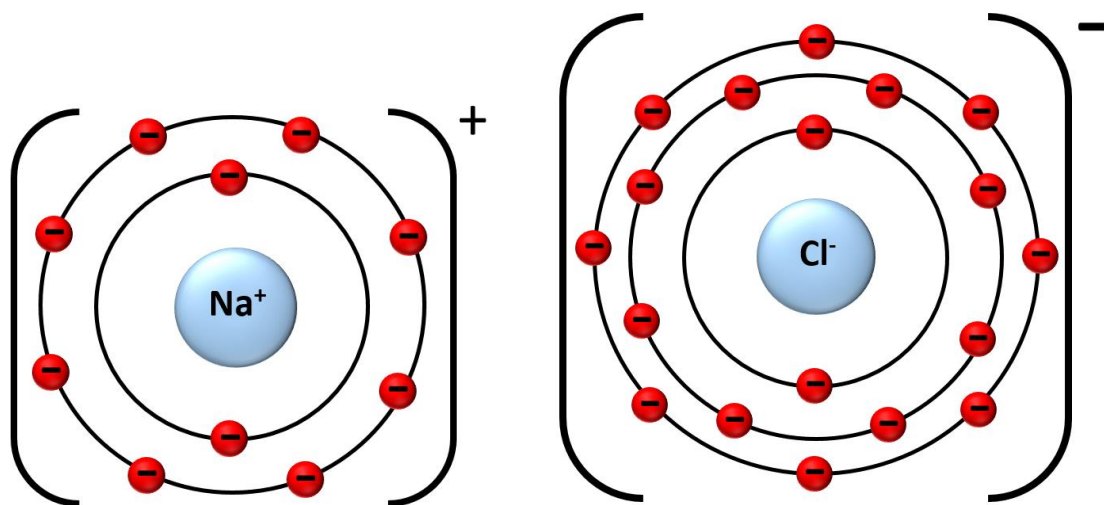
- Deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions.
- Explanation of the physical properties of ionic compounds (volatility, electrical conductivity and solubility) in terms of their structure.

Guidance:

- Students should be familiar with the following polyatomic ions: NH_4^+ OH^- HCO_3^- CO_3^{2-} SO_4^{2-} PO_4^{3-} NO_3^-

Ion formation

- Neutral atoms have equal numbers of protons and electrons.
- Positive ions are formed when atoms lose electrons.
- Negative ions are formed when atoms gain electrons.



Exercises:

1) Explain why the sodium atom has a larger radius than the sodium ion.

The sodium atom has 3 occupied energy levels and the sodium ion has 2 occupied energy levels.

The sodium atom has equal numbers of protons and electrons.

The sodium ion has more protons than electrons (11:10), therefore the remaining electrons experience a stronger electrostatic attraction from the protons, making the ion smaller.

2) Explain why the chloride ion has a larger radius than the chlorine atom.

The chloride ion has gained an electron, therefore it has more electrons than protons.

The extra electron causes extra repulsion between electrons making the ion bigger.

The electrons in the ion feel a weaker electrostatic attraction from the same number of protons, therefore the electrons are held less tightly (making the ion bigger).

3) Explain why metals tend to lose electrons and form positive ions and non-metals tend to gain electrons and form negative ions.

Metals have low electronegativity values, therefore they can lose electrons easily to form positive ions.

Non-metals have higher electronegativity values therefore they gain electrons to form negative ions.

4) Determine the ions formed by the following atoms:

a) Cs

b) Be

c) Cl

d) N

e) S

a) Cs^+

b) Be^{2+}

c) Cl^-

d) N^{3-}

e) S^{2-}

f) I

g) Ba

h) Al

i) O

j) K

f) I^-

g) Ba^{2+}

h) Al^{3+}

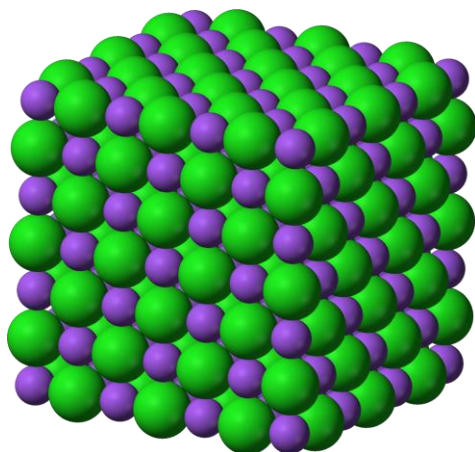
i) O^{2-}

j) K^+

Ionic bonding

- The ionic bond is the electrostatic attraction between oppositely charged ions.
- An ionic bond forms between two elements with a difference in electronegativity of equal to or greater than 1.8 units.
- Ionic bonds form between metal and non-metal elements.

Structure of ionic compounds



Ionic compounds have a lattice structure. The lattice is held together by the electrostatic attraction between the oppositely charged ions. Each Na^+ ion is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions. Ionic compounds are solids under standard conditions.

Exercises:

1) Define an ionic bond.

An ionic bond is the electrostatic attraction between oppositely charged ions.

2) An element on the far left and an element on the far right of the periodic table form a chemical bond. Determine the type of bond they would form and explain your reasoning.

An element on the far left is likely to be a metal, an element on the far right is likely to be a non-metal.

Metals and non-metals on opposite ends of the periodic table have a big difference in electronegativity (1.8 units or greater), therefore they form an ionic bond.

3) Describe the lattice structure of NaCl .

The sodium and chloride ions are held together by their opposite charges (+ and -).

Each sodium ion is surrounded by six chloride ions.

Each chloride ion is surrounded by six sodium ions.

Writing ionic formulae

List of common ions:

Positive ions		Negative ions	
Name	Symbol	Name	Symbol
Hydrogen	H ⁺	Fluoride	F ⁻
Sodium	Na ⁺	Chloride	Cl ⁻
Silver	Ag ⁺	Bromide	Br ⁻
Potassium	K ⁺	Iodide	I ⁻
Lithium	Li ⁺	Hydrogencarbonate	HCO ₃ ⁻
Ammonium	NH ₄ ⁺	Hydroxide	OH ⁻
Barium	Ba ²⁺	Nitrate	NO ₃ ⁻
Calcium	Ca ²⁺	Oxide	O ²⁻
Copper(II)	Cu ²⁺	Sulfate	SO ₄ ²⁻
Magnesium	Mg ²⁺	Carbonate	CO ₃ ²⁻
Zinc	Zn ²⁺	Phosphate	PO ₄ ³⁻
Mercury(I)	Hg ⁺	Nitride	N ³⁻
Lead	Pb ²⁺	Sulfide	S ²⁻
Iron(II)	Fe ²⁺	Phosphide	P ³⁻
Iron(III)	Fe ³⁺	Nitrite	NO ₂ ⁻
Aluminium	Al ³⁺	Sulfite	SO ₃ ²⁻

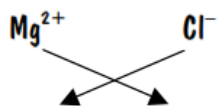
Exercise: Write the formulae for the following ionic compounds:

- Potassium bromide
- Calcium fluoride
- Beryllium sulfide
- Strontium iodide
- Magnesium nitride
- Ammonium sulfate
- Iron(III) sulfite
- Copper(II) nitrite
- Potassium hydrogencarbonate
- Aluminium sulfate
- Aluminium oxide
- Sodium carbonate
- Copper(II) phosphide
- Zinc phosphate
- Ammonium nitrate
- Mercury(I) nitride
- Iron(II) nitrite
- Barium nitrate
- Iron(II) phosphide
- Calcium hydrogencarbonate

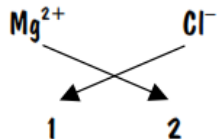
Step 1 Write the ions side by side:



Step 2 Draw arrows that cross each other:



Step 3 Write the **charges** at the arrow ends:



Step 4 Write the formula as follows:

a) write the positive ion without its charge:



b) write the number as a subscript unless it is 1:



c) write the negative ion without its charge:



d) write the number as a subscript unless it is 1:



Answers:

1) KBr

2) CaF₂

3) BeS

4) SrI₂

5) Mg₃N₂

6) Al₂O₃

7) Na₂CO₃

8) Cu₃P₂

9) Zn₃(PO₄)₂

10) NH₄NO₃

11) (NH₄)₂SO₄

12) Fe₂(SO₃)₃

13) Cu(NO₂)₂

14) KHCO₃

15) Al₂(SO₄)₃

16) Hg₃N

17) Fe(NO₂)₂

18) Ba(NO₃)₂

19) Fe₃P₂

20) Ca(HCO₃)₂

Structure and properties of ionic compounds

Electrical conductivity

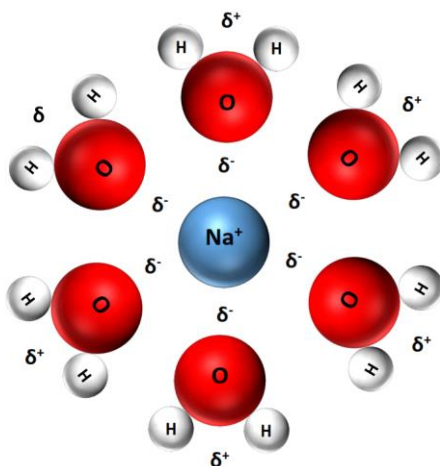
- Ionic compounds do not conduct electricity when solid, because the ions are held in fixed positions by the electrostatic attractions between the ions.
- They only conduct electricity when molten (melted) or dissolved in water (aqueous).
- When molten or dissolved, the ions are free to move and carry the electric current.

Melting and boiling point

- Ionic compounds have high melting and boiling points due to the strong electrostatic attraction between the oppositely charged ions (NaCl has a melting point of 800 °C).
- The greater the charge and smaller the ionic radius of the ions, the stronger the electrostatic attraction and the higher the melting point.

Solubility

- Ionic compounds are soluble in polar solvents.
- The ions are separated from the lattice structure by the polar water molecules.
- The ions become surrounded by water molecules (hydration) as shown below.



Exercises:

1. Explain the conductivity of ionic compounds when molten or dissolved and when solid.

Solid ionic compounds do not conduct electricity because the ions are held in fixed positions in the lattice structure by strong electrostatic attractions.

When molten or dissolved, the ions are free to move and can conduct electricity.

2. Explain the high melting point of ionic compounds.

The high melting point of ionic compounds is due to the strong electrostatic attractions between the oppositely charged ions.

3. Explain why NaF has a higher melting point than KF.

Sodium has a smaller ionic radius than potassium. The smaller ionic radius means that the electrostatic attraction between the ions in sodium chloride is stronger than the electrostatic attraction in potassium fluoride.

4.2 Covalent bonding

Understandings:

- A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
- Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively.
- Bond length decreases and bond strength increases as the number of shared electrons increases.
- Bond polarity results from the difference in electronegativities of the bonded atoms.

Applications and skills:

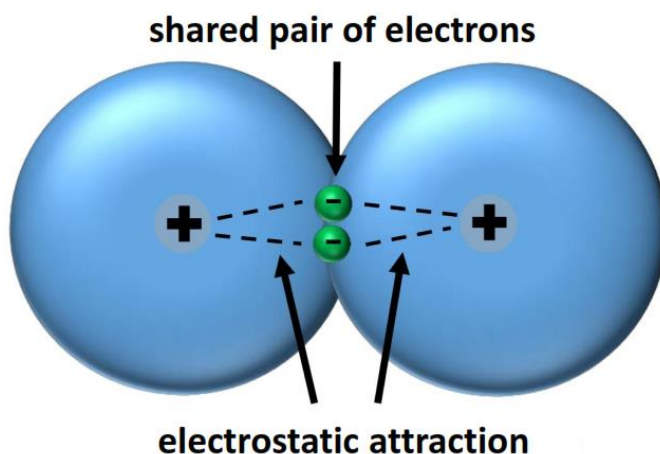
- Deduction of the polar nature of a covalent bond from electronegativity values.

Guidance:

- Bond polarity can be shown either with partial charges, dipoles or vectors.
- Electronegativity values are given in the data booklet in section 8.

Covalent bonding

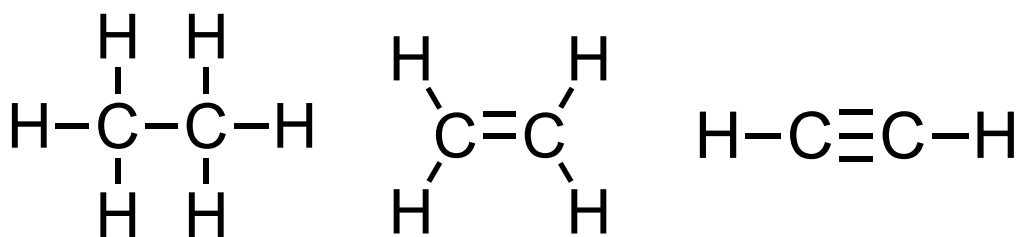
- A covalent bond is the electrostatic attraction between positive nuclei and a shared pair of bonding electrons.
- Covalent bonding occurs between non-metal elements only.
- The electronegativity difference between atoms is between 0.0 and 1.7 units.



- Both nuclei are attracted to the shared pair of bonding electrons.
- The electrostatic attraction between the nuclei and shared pair of bonding electrons forms 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

Polar covalent bonds

- Covalent bonds between atoms that have a difference in electronegativity of between 0.5 and 1.7 units form a polar covalent bond.
- Bond polarity can be shown by a vector arrow or by δ^+ and δ^- as shown below.



Exercises:

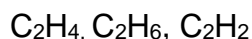
1. Define a covalent bond.

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

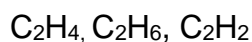
2. Where on the periodic table would you find elements that bond by covalent bonding?

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).



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



5. State the relationship between the number of electrons in a covalent bond and the length and strength of the bond.

As the number of electrons in the bond increases (2 in a single bond, 4 in a double bond and 6 in a triple bond), the strength of the bond increases and the length of the bond decreases.

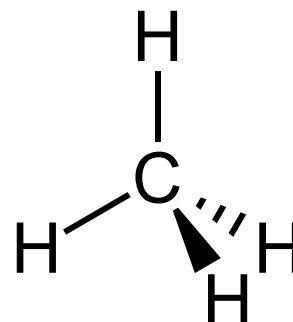
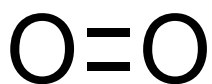
6. Draw a diagram showing the bonding polarity in HF (hydrogen fluoride).

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

Polar and non-polar molecules

The polarity of a molecule depends on two factors:

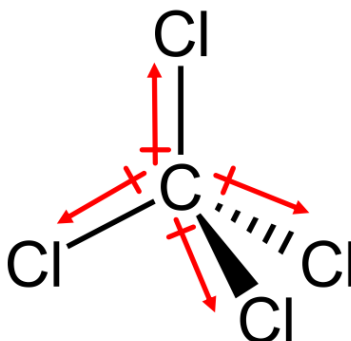
- The presence of polar bonds within the molecule.
- The shape (geometry) of the molecule.



- 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 (C-H)

Non-polar molecule with polar bonds

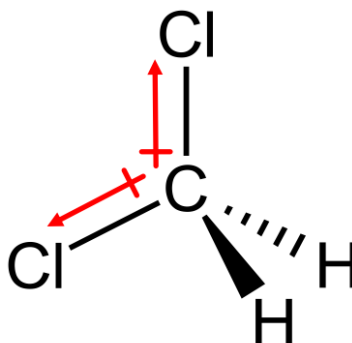
- CCl_4 is a non-polar molecule that has polar bonds.



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

Polar molecules

- CH_2Cl_2 is a polar molecule.



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

Exercises

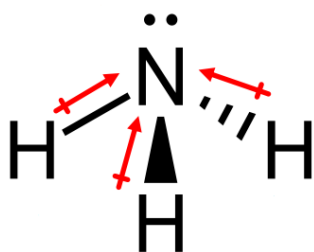
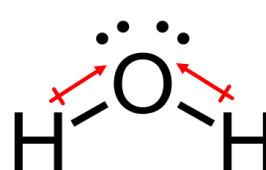
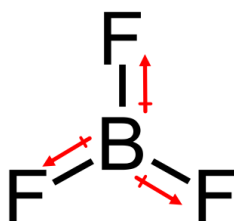
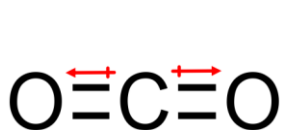
1. Outline the two factors that determine the polarity of a molecule.

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 in a bond due to a difference in electronegativity of 0.5 units or greater 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 – non-polar as the bond dipoles cancel out because of linear molecular geometry.

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

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

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

Coordinate covalent bonds

- In a coordinate covalent bond, both shared electrons come from the same atom.
- Draw the Lewis structure of CO, H₃O⁺ and NH₄⁺ in the boxes below.

CO	H ₃ O ⁺	NH ₄ ⁺
$\text{:C}\equiv\text{O:}$	$\left[\text{H}-\ddot{\text{O}}-\text{H} \right]^+$ H	$\left[\begin{array}{c} \text{H} \\ \\ \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array} \right]^+$

4.3 Covalent structures

Understandings:

- Lewis (electron dot) structures show all the valence electrons in a covalently bonded species.
- The “octet rule” refers to the tendency of atoms to gain a valence shell with a total of 8 electrons.
- Some atoms, like Be and B, might form stable compounds with incomplete octets of electrons.
- Resonance structures occur when there is more than one possible position for a double bond in a molecule.
- Shapes of species are determined by the repulsion of electron pairs according to VSEPR theory.
- Carbon and silicon form giant covalent/network covalent structures.

Applications and skills:

- Deduction of Lewis (electron dot) structure of molecules and ions showing all valence electrons for up to four electron pairs on each atom.
- The use of VSEPR theory to predict the electron domain geometry and the molecular geometry for species with two, three and four electron domains.
- Prediction of bond angles from molecular geometry and presence of non-bonding pairs of electrons.
- Prediction of molecular polarity from bond polarity and molecular geometry.
- Deduction of resonance structures, examples include but are not limited to CO₃²⁻, NO₃⁻
- Explanation of the properties of giant covalent compounds in terms of their structures

Guidance:

- The term “electron domain” should be used in place of “negative charge centre”.
- Electron pairs in a Lewis (electron dot) structure can be shown as dots, crosses, a dash or any combination.
- Allotropes of carbon (diamond, graphite, graphene, C₆₀ buckminsterfullerene) and SiO₂ should be covered.
- Coordinate covalent bonds should be covered.

Lewis structures

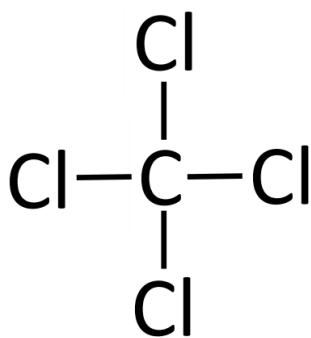
- Lewis structures show all the valence electrons in a molecule; the bonding electrons and the lone pairs of electrons (non-bonding electrons).

How to determine Lewis structures:

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

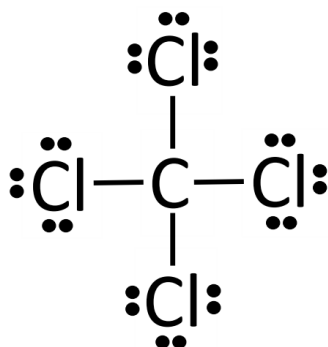
Example – CCl₄ (carbon tetrachloride)

- 1) Total number of valence electrons = $4 + (4 \times 7) = 32$
- 2) Number of valence electrons needed for each atom to complete its octet $(5 \times 8) = 40$
- 3) $40 - 32 = 8$ bonding electrons
- 4) Skeletal formula:



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

5) 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).

Exercise:

1) Draw Lewis structures for the following molecules:

a) Cl₂

b) HCl

c) CH₃Cl

d) CF₄

e) PF₃

f) BF₃

g) BeCl₂

h) N₂H₄

i) H₂O

j) CH₃OH

k) CO

l) C₂H₄

j) HCN

2) Draw Lewis structures for the following ions:

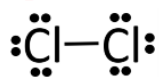
a) Cl⁻

b) NH₄⁺

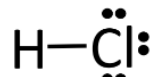
c) OH⁻

Answers

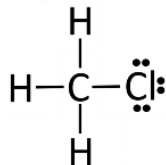
Cl_2



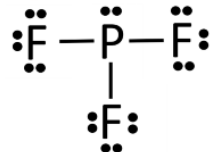
HCl



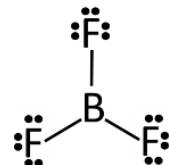
CH_3Cl



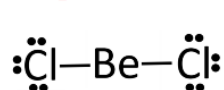
PF_3



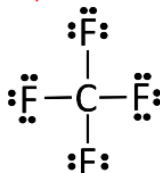
BF_3^*



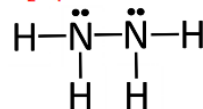
BeCl_2^*



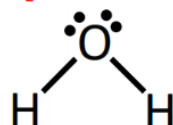
CF_4



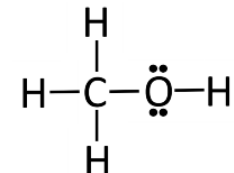
N_2H_4



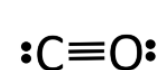
H_2O



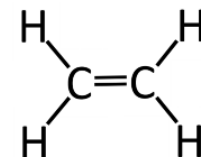
CH_3OH



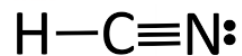
CO



C_2H_4



HCN

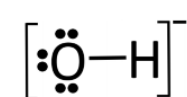


2)

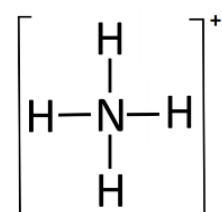
Cl^-



OH^-

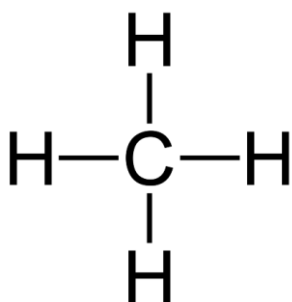


NH_4^+



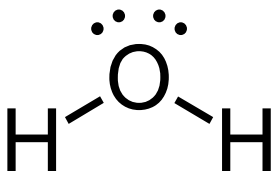
VSEPR theory

- 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?

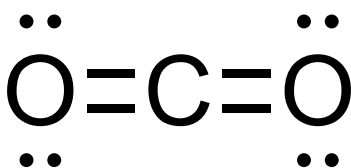
4 bonding domains



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:

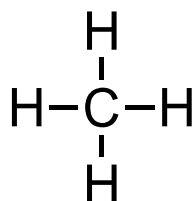
lone pair – lone pair > lone pair – bonding domain > bonding domain – bonding domain

most repulsion

least repulsion

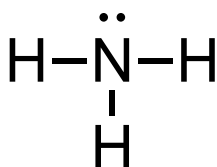
Molecules with four electron domains

- Four bonding domains, zero non-bonding domains



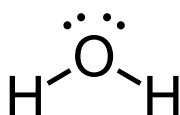
electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
4	4	0	tetrahedral	tetrahedral	109.5°

- 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°

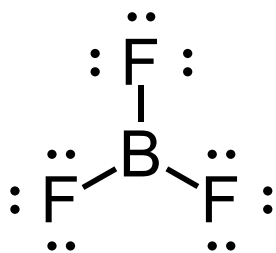
- Two bonding domains, two non-bonding domains



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

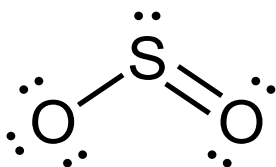
Molecules with three electron domains

- 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°

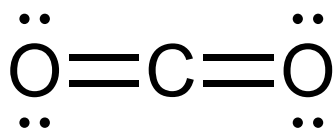
- Two bonding domains, one non-bonding domain



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

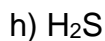
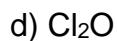
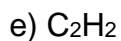
- Two bonding domains, zero non-bonding domains



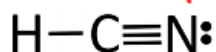
electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
2	2	0	linear	linear	180°

Exercise:

Determine the shape and bond angles in the following molecules:

**Answers:**

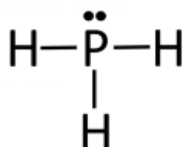
HCN – linear (180°)



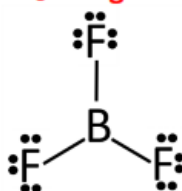
C₂H₂ – linear (180°)



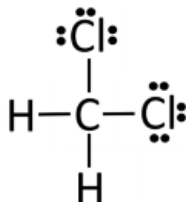
PH₃ – trigonal pyramidal
(<107.8°)*



BF₃ – trigonal planar (120°)



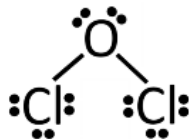
CH₂Cl₂ – tetrahedral (109.5°)



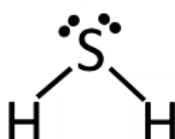
BeH₂ – linear (180°)



Cl₂O – bent/v-shaped (111°)*



H₂S – bent/v-shaped (<104.5°)*

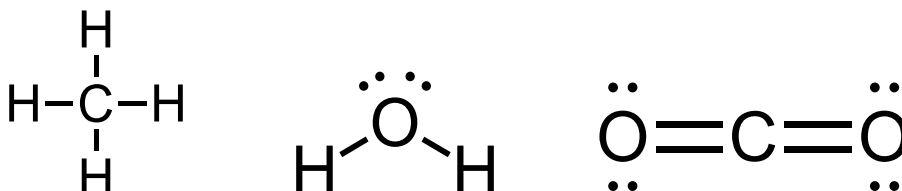


* The reasons for these bond angles are complex and are beyond the scope of the standard level IB course. It is more important that you can correctly predict the shape of the molecule in these cases.

Structure and properties of covalent compounds

- Covalent compounds have two types of structure; simple molecular and giant covalent (network covalent).
- Each type of structure has different physical properties, such as melting and boiling point, electrical conductivity, solubility and hardness.

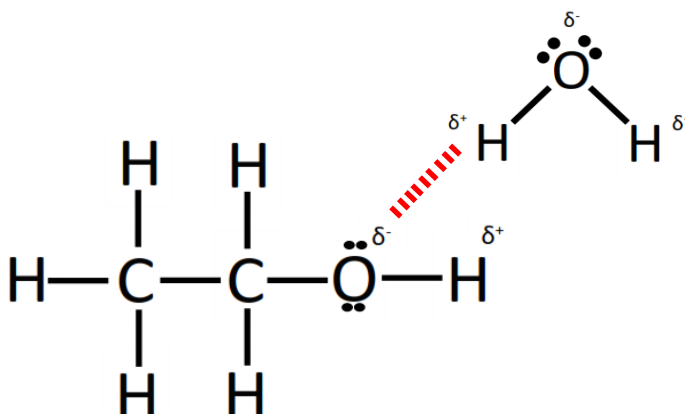
Simple molecular structures



- Simple molecular compounds are mostly liquids and gases due to the weak intermolecular forces between molecules.

Solubility

- Non-polar molecules are soluble in non-polar solvents (CH₄ is soluble in C₃H₈)
- Polar molecules are soluble in polar solvents (C₂H₅OH is soluble in H₂O)
- The phrase '**like dissolves like**' is useful to remember the solubility.
- Hydrogen bonding between a molecule of C₂H₅OH (ethanol) and a molecule of H₂O is shown below.



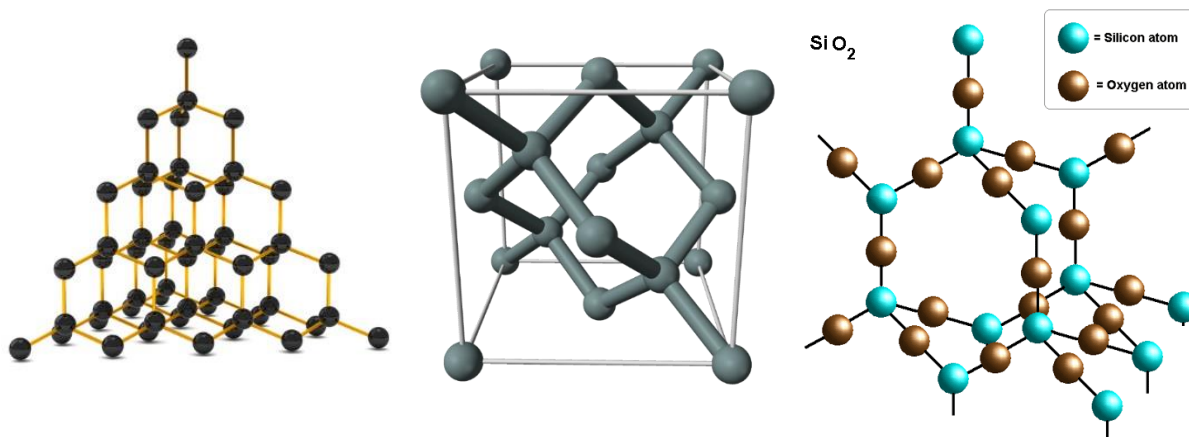
- Ethanol is soluble in water as it is able to form hydrogen bonds between itself and water molecules.

Electrical conductivity

- Simple molecular substances are poor conductors of electricity as they have no free moving (mobile) ions in solution.

Giant covalent structures (network covalent)

- 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 also very hard substances.

Exercise:

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 (C_5H_{12}) is insoluble in water but soluble in hexane (C_6H_{14}).

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. CO_2 is a gas under standard conditions.

The molecules of CO_2 have weak London dispersion forces between them.

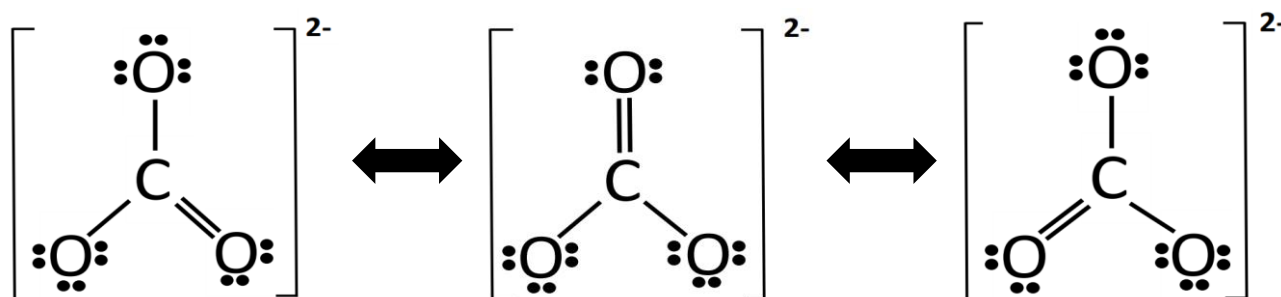
- d. SiO_2 has a melting point of $1600^\circ C$

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

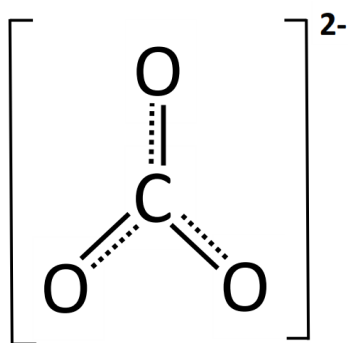
Resonance structures

- Resonance structures occur when there is more than one position for a multiple bond in a molecule.

The three resonance structures for the carbonate ion (CO_3^{2-}) are shown below.



- The actual structure is a resonance hybrid structure, shown below.



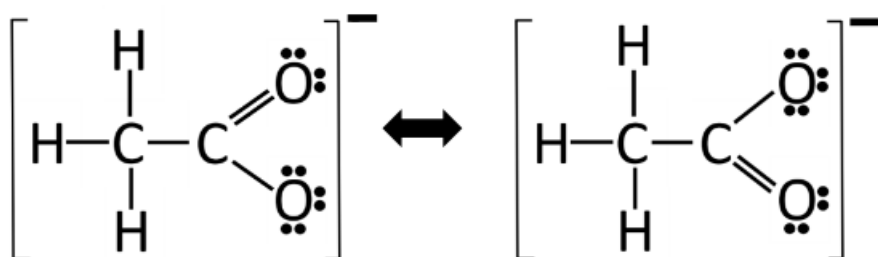
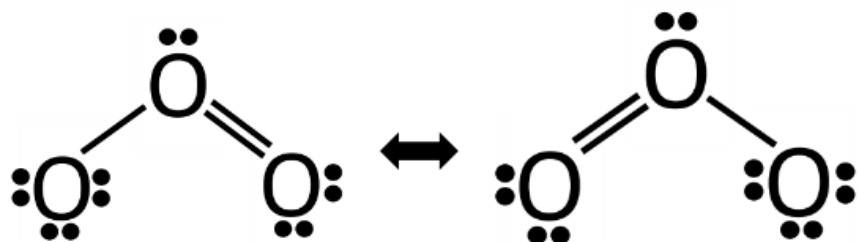
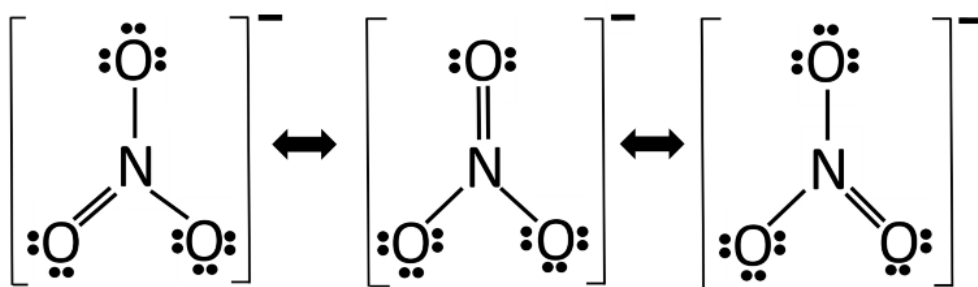
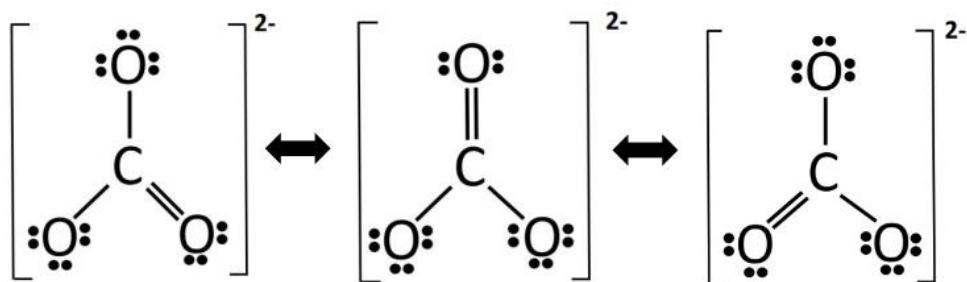
The bond lengths and bond strengths are intermediate between a single and a double bond.

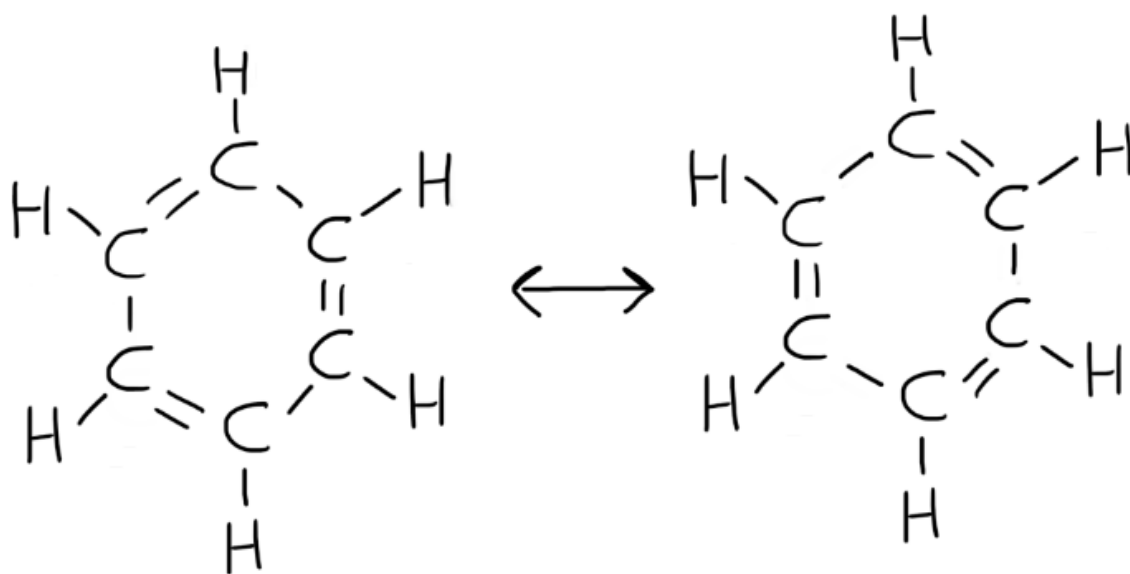
Exercise:

1) Draw the resonance structures for the following molecules and ions:

- | | |
|---------------------------------------|--|
| a) Carbonate ion – CO_3^{2-} | d) Ethanoate ion - CH_3COO^- |
| b) Nitrate ion - NO_3^- | e) Benzene – C_6H_6 |
| c) Ozone - O_3 | |

1)

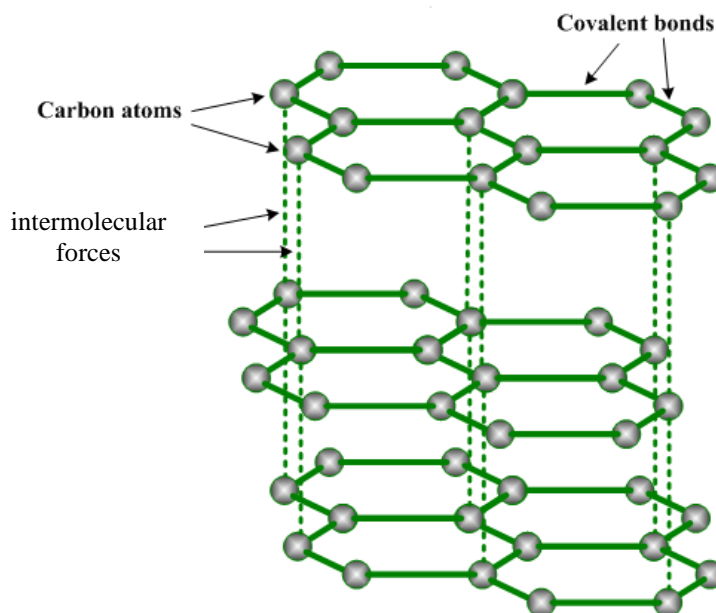




Allotropes of carbon

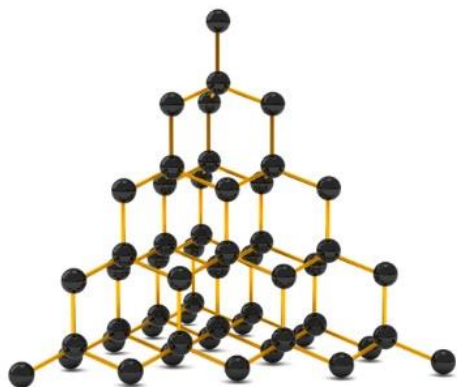
- Allotropes are different forms of the same element in the same physical state.
- Carbon has 4 allotropes – graphite, diamond, Fullerene C₆₀ and graphene.
- Different bonding within the structures gives the allotropes different properties.

Graphite



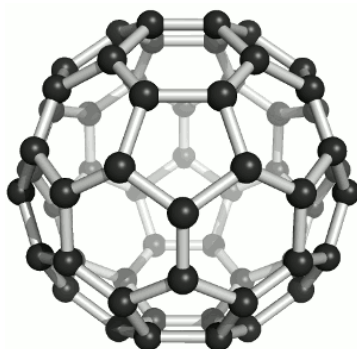
- Graphite has a layered structure.
- The layers are held together by weak intermolecular forces.
- The layers can slide over one another (because of weak intermolecular forces).
- Each carbon atom is bonded to 3 other carbon atoms.
- The bond angle between carbon atoms is 120°, trigonal planar.
- Graphite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

Diamond



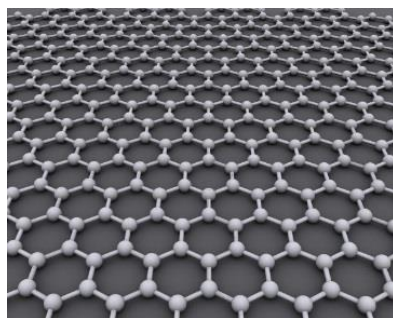
- Giant covalent structure.
- High melting and boiling point.
- Very hard (strong covalent bonds between atoms).
- Each carbon is bonded to 4 other carbon atoms.
- Bond angle is 109.5° , tetrahedral.
- Does not conduct electricity (no delocalised electrons).
- Used in jewellery and for cutting glass.

Fullerene C₆₀



- Each carbon atom is bonded to 3 other carbon atoms.
- Structure consists of 12 pentagons and 20 hexagons.
- Shows some electrical conductivity.

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 also very strong.

Allotropes of carbons 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/phonics 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 used for making electrodes.
The presence of delocalised electrons within the structure of graphite make it a good electrical conductor.
5. 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.

4.4 Intermolecular forces

Understandings:

- Intermolecular forces include London (dispersion) forces, dipole-dipole forces and hydrogen bonding.
- The relative strengths of these interactions are London (dispersion) forces < dipole-dipole forces < hydrogen bonds.

Applications and skills

- Deduction of the types of intermolecular force present in substances, based on their structure and chemical formula.
- Explanation of the physical properties of covalent compounds (volatility, electrical conductivity and solubility) in terms of their structure and intermolecular forces.

Guidance:

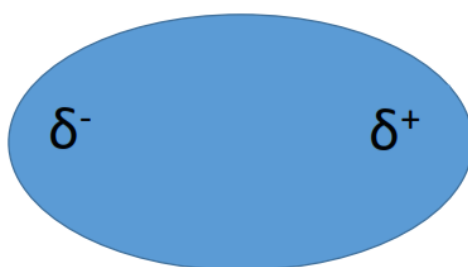
- The term “London (dispersion) forces” refers to instantaneous induced dipole-induced dipole forces that exist between any atoms or groups of atoms and should be used for non-polar entities. The term “van der Waals” is an inclusive term, which includes dipole–dipole, dipole-induced dipole and London (dispersion) forces.

Intermolecular forces

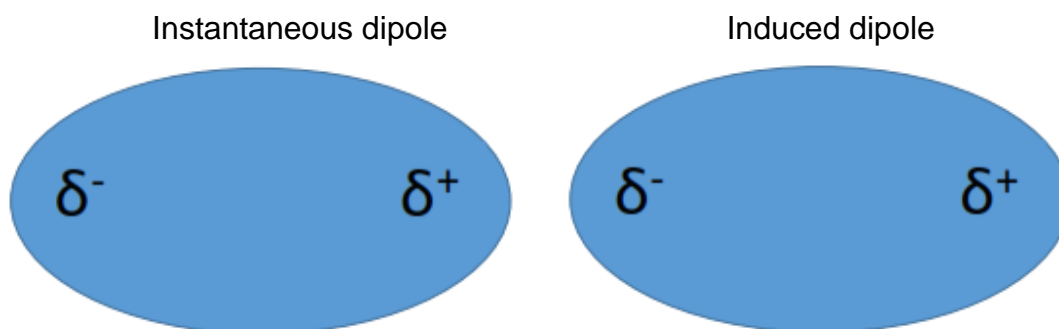
- Intermolecular forces are forces of attraction between molecules.
- The three types of intermolecular force are:
London dispersion forces, dipole-dipole forces and hydrogen bonding.
- London dispersion forces and dipole-dipole forces are collectively known as van der Waals forces.

London dispersion forces

- London dispersion forces are the weakest type of intermolecular force.
- 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.



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.

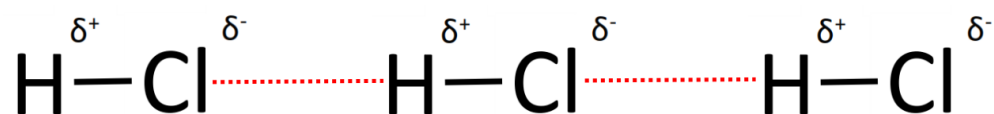
Boiling point of the halogens

	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.
- London dispersion forces increase 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).

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 in another.
- The dipole-dipole attractions between HCl molecules are shown below.



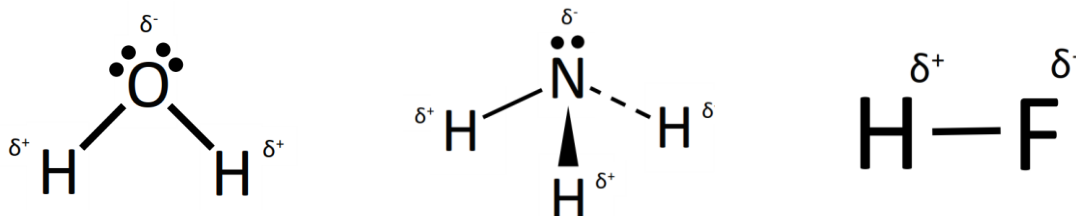
Exercises:

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.

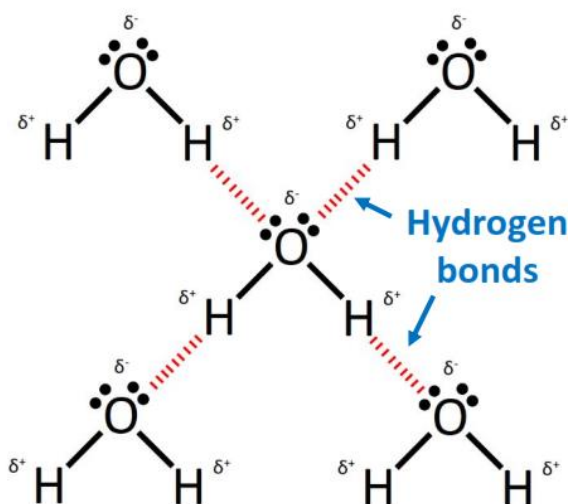
Hydrogen bonding

- Hydrogen bonding occurs when hydrogen is bonded to either a nitrogen, oxygen or fluorine atom.
- It is the strongest type of intermolecular force.

Molecules that have hydrogen bonding between molecules (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:

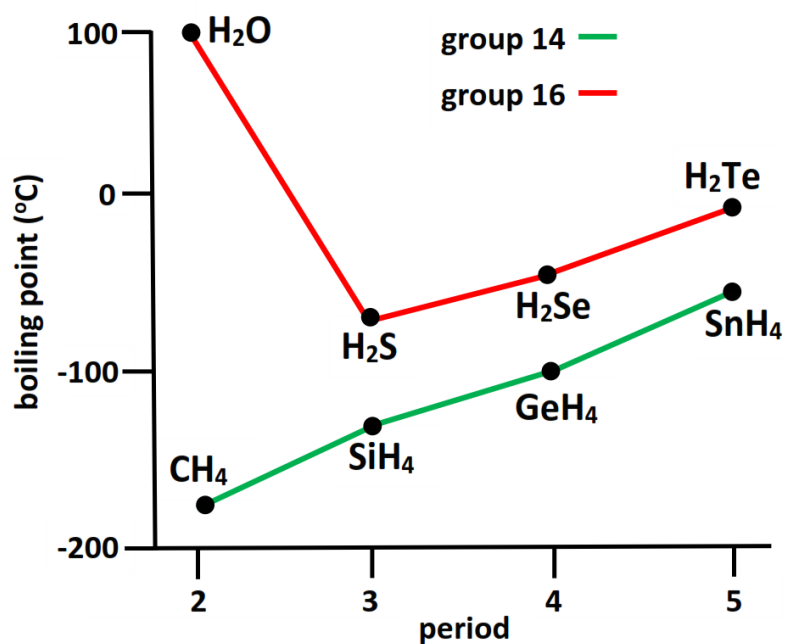
- Which groups of atoms are necessary for the formation of a hydrogen bond?

O-H, N-H, F-H

- 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 molecules also increases, which results in a higher boiling point.

Exercises:

1. Arrange the following in terms of increasing strength:

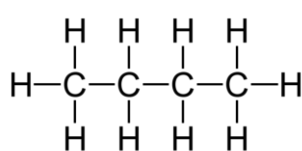
Hydrogen bonding London dispersion forces Dipole-dipole forces

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

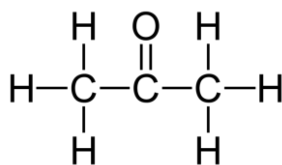
- | | |
|-------------------|---------------------------|
| a) Cl_2 | f) CH_3Cl |
| b) HCl | g) H_2O |
| c) HF | h) CH_3OH |
| d) CH_4 | i) C_2H_6 |
| e) CCl_4 | j) NH_3 |

3. Explain why at room temperature, F_2 and Cl_2 are gases, Br_2 is a liquid and I_2 is a solid.

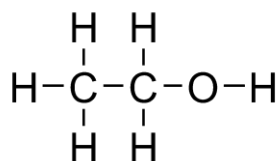
4. The structures of butane, propanone, and ethanol can be seen below. They have similar molar masses, but different boiling points. Explain the difference in boiling point between the three molecules.



butane B.P. -1°C



propanone B.P 56°C



ethanol B.P 78°C

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.

4.5 Metallic bonding

Understandings:

- A metallic bond is the electrostatic attraction between a lattice of positive ions and delocalized electrons.
- The strength of a metallic bond depends on the charge of the ions and the radius of the metal ion.
- Alloys usually contain more than one metal and have enhanced properties.

Applications and skills:

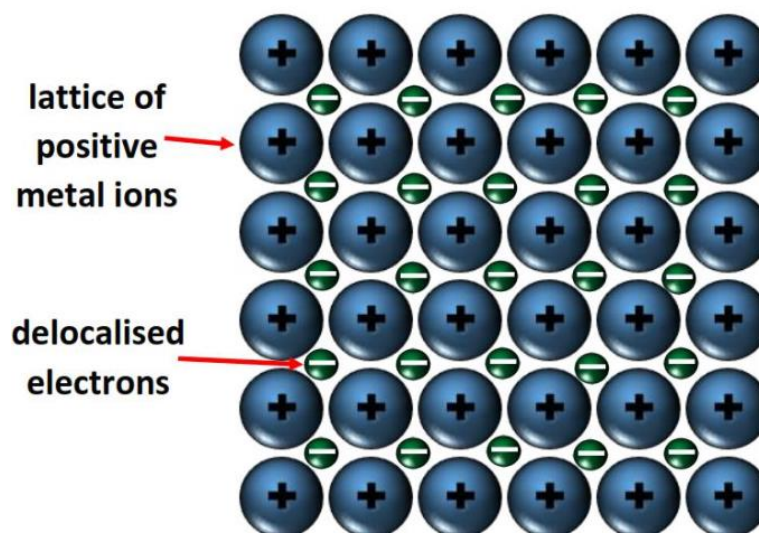
- Explanation of electrical conductivity and malleability in metals.
- Explanation of trends in melting points of metals.
- Explanation of the properties of alloys in terms of non-directional bonding.

Guidance:

- Trends should be limited to s- and p-block elements.
- Examples of various alloys should be covered.

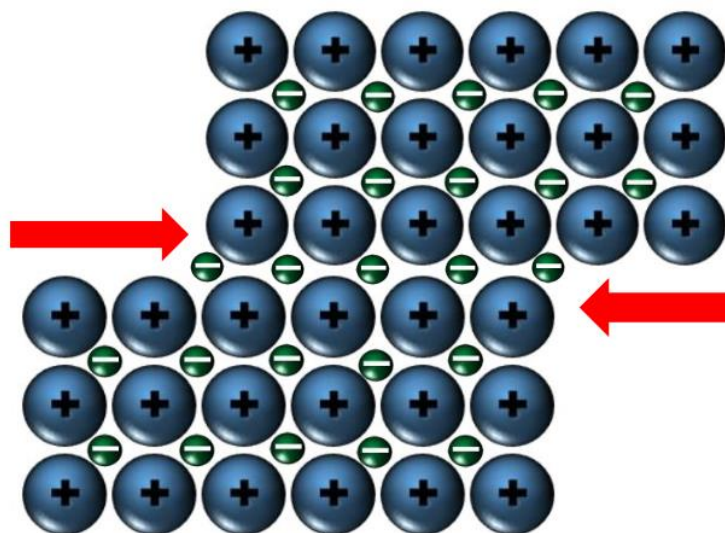
Metallic bonding

- The metallic bond is the electrostatic attraction between a lattice of positively charged metal ions and delocalized electrons.



- Metals are good conductors of heat and electricity because of the presence of delocalized electrons.
- Metals are malleable (can be bent into shape) and ductile (can be drawn into wires) because the metallic bond remains intact even if the structure is distorted.
- Metals are reflective – the delocalized electrons in the metallic structure reflect light.

- Metals are malleable and ductile – the layers can slide over each other when metals are bent, hammered, or stretched, without breaking the metallic bond.



Strength of the metallic bond

- The strength of the metallic bond is determined by the charge on the metal ion and the ionic radius of the metal ion.

Ion	charge on ion	ionic radius ($\times 10^{-12}$ m)	melting point ($^{\circ}\text{C}$)
Na^{+}	1+	102	98
Mg^{2+}	2+	72	650

- Mg has a higher melting point than Na due to its greater ionic charge and smaller ionic radius.
- The higher the charge on the ion, the more delocalised electrons exist in the metallic structure.
- This results in a stronger electrostatic attraction between the lattice of positive metal ions and the delocalized electrons, therefore a stronger metallic bond.
- The strength of the metallic bond decreases down a group as the size of the metal cation increases.

Exercises:

1. Describe the metallic bond.

The metallic bond is the electrostatic attraction between a lattice of positive metal ions and delocalised electrons.

2. Explain the following properties of metals:

- a. Metals are good conductors of heat and electricity

The delocalised electrons in the lattice structure allow the conduction of heat and electricity.

- b. Metals are malleable and ductile

The layers of positive ions in the lattice can slide over each other without breaking the metallic bond.

- c. Metals are shiny

The delocalised electrons reflect light.

3. Explain why aluminium has a higher melting point than sodium.

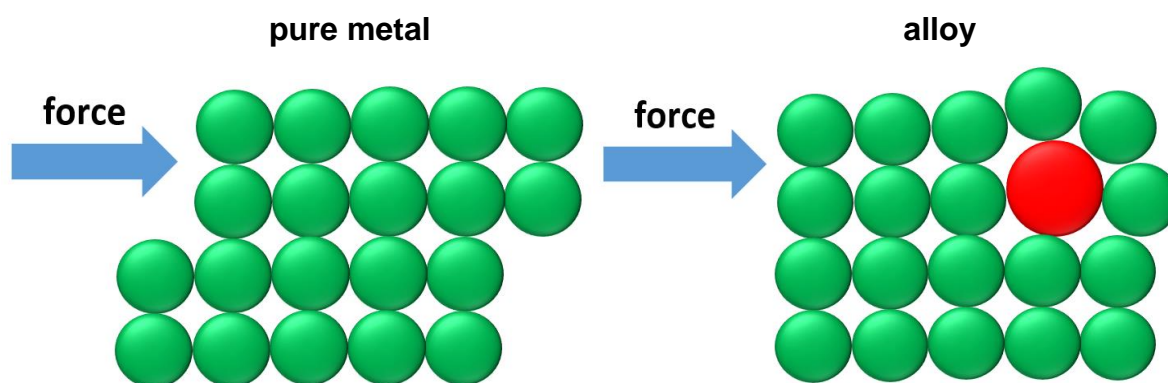
The aluminium ion has a higher charge and a smaller ionic radius than the sodium ion. This results in a stronger metallic bond and a higher melting point.

4. Explain why the melting point of the group 1 metals decreases down the group.

The strength of the metallic bond depends on the ionic radius of the ions. The ionic radius increases down the group, therefore, the metallic bond gets weaker which results in a decreasing melting point.

Alloys

- Alloys are materials that are composed of two or more metals or a metal and a non-metal.
- The bonding in metals is non-directional; the force of attraction between the positive metal ions and the delocalised electrons acts in all directions around the fixed metal ions.
- Alloys have enhanced properties (increased tensile strength and increased resistance to corrosion); they have different properties to the metals that they are made from.
- They tend to be harder (less malleable) and have greater tensile strength (stronger).
- The added metal atoms can distort the lattice structure.
- The distortion of the lattice structure makes it more difficult for the layers to slide over each other.



- In a pure metal the layers can slide over each other.
- The presence of different metal atoms means the layers cannot slide over each other as easily.

Exercise: Explain why alloys are harder than pure metals.

The addition of different sized metals atoms (or ions) in the alloy means that the layers in the lattice structure cannot move over each other as easily as in the pure metal.

Alloys and their uses

Alloy	Component metals	Properties and uses
Steel	iron, carbon	high tensile strength; used in construction
Stainless steel	iron, nickel, chromium	resistant to corrosion; used in cooking implements
Brass	copper and zinc	pipes
Bronze	copper and tin	coins, medals, tools
Pewter	tin, copper, antimony	decorative ornaments
Solder	lead and tin	low melting point; used to join metals in electrical circuits
Nichrome	nickel and chromium	heating elements

Exercises: Suggest an alloy for the following uses with a reason for your choice.

1. To use in electrical circuits.

Solder – low melting point

2. To make water pipes.

Brass – resistant to corrosion

3. To use in an ornament.

Pewter – resistant to corrosion

4. To construct a bridge.

Steel – high tensile strength

5. To make a saucepan.

Stainless steel – resistant to corrosion