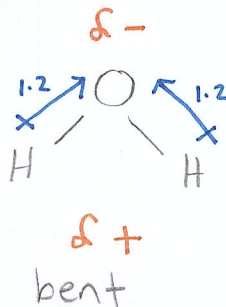


Polar Bonds vs. Polar Molecules

- We already learned covalent bonds can be polar or nonpolar based on the electronegativity of the atoms in the molecule
- We now need to look at all the individual covalent bonds in a molecule to determine if the entire molecule will be polar or not
- To determine if a molecule is polar or nonpolar follow these steps
 1. Draw the VSEPR shape of the molecule
 2. Use the electronegativities of the atoms in the molecule to determine if any bond dipoles exist. Draw the bond dipole arrows onto the VSEPR drawing
 3. "Add" up the bond dipole arrows to see if they all cancel out or not to determine if the entire molecule is polar or nonpolar.
 - If the bond dipoles arrows generally point in the same direction, the entire molecule has a side that is overall slightly positive or negative and is called a polar molecule
 - If the bond dipoles arrows point in opposite directions and are equal in magnitude/strength, the molecule is called nonpolar. In order for polar bonds to cancel each other out and result in a nonpolar molecule, the molecule needs to be symmetrical.

- EXAMPLE:** Show that H_2O is a polar molecule.

$$\begin{array}{l} \text{H} = 1e^- \times 2 \\ \text{O} = 6e^- \\ \hline = 8e^- \end{array}$$

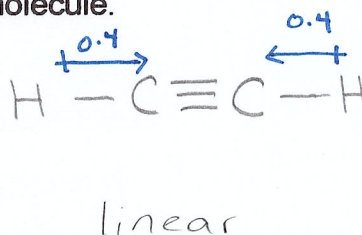


$$\begin{array}{c} \xleftarrow{1.2} + \\ \text{O} - \text{H} \\ \Delta \text{EN} = 3.4 - 2.2 \\ \Delta \text{EN} = 1.2 \end{array}$$

$\therefore$ polar

- EXAMPLE:** Show that C_2H_2 is a nonpolar molecule.

$$\begin{array}{l} \text{C} = 4e^- \times 2 \\ \text{H} = 1e^- \times 2 \\ \hline = 10e^- \end{array}$$



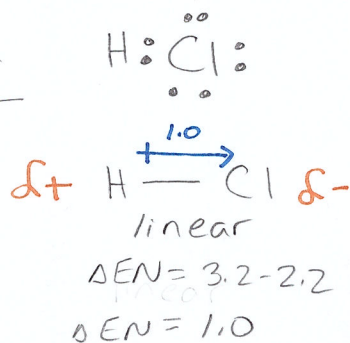
$$\begin{array}{c} \xleftarrow{0.4} + \\ \text{C} - \text{H} \\ \Delta \text{EN} = 2.6 - 2.2 \\ \Delta \text{EN} = 0.4 \end{array}$$

$\therefore$ non polar

- EXAMPLES:** Draw the Lewis structure and the VSEPR shape for each molecule and then determine if the following molecules are polar or nonpolar.

1. HCl = polar

$$\begin{array}{l} \text{H} = 1e^- \\ \text{Cl} = 7e^- \\ \hline = 8e^- \end{array}$$



4. N_2 = non polar

$$\begin{array}{l} \text{N} = 5e^- \times 2 \\ \hline = 10e^- \end{array}$$



linear

$$\Delta \text{EN} = 3.0 - 3.0$$

$$\Delta \text{EN} = 0$$

$\therefore$ no bond dipole

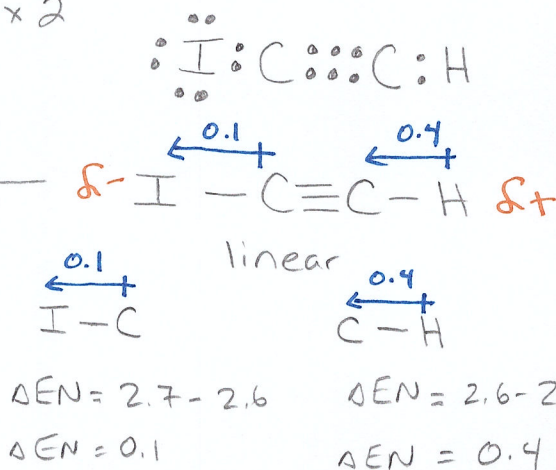
2. C_2HI = polar

$$\text{C} = 4e^- \times 2$$

$$\text{H} = 1e^-$$

$$\text{I} = 7e^-$$

$$\hline = 16e^-$$

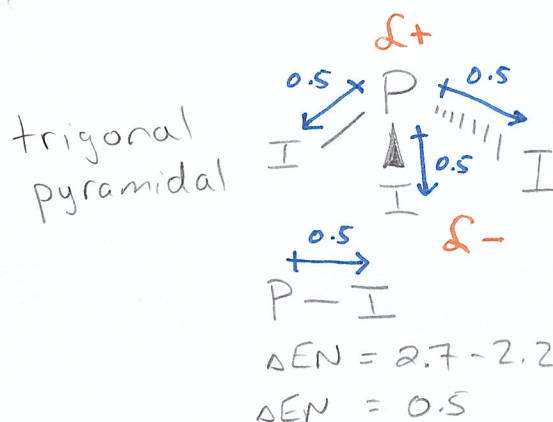
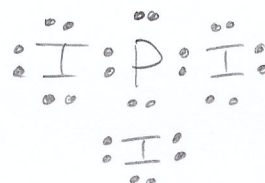


5. PI_3 = polar

$$\text{P} = 5e^-$$

$$\text{I} = 7e^- \times 3$$

$$\hline = 26e^-$$

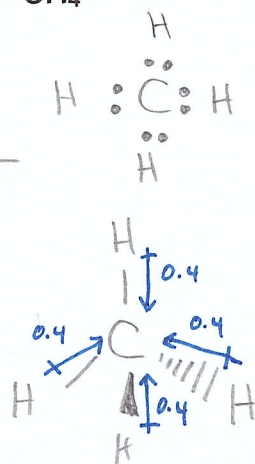


3. CH_4 = non polar

$$\text{C} = 4e^-$$

$$\text{H} = 1e^- \times 4$$

$$\hline = 8e^-$$



Now Try Practice Problems

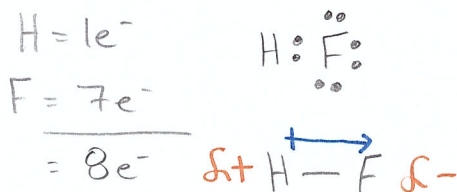
- Summary Chart on Polar vs. Nonpolar Molecules

Shape	VSEPR Diagram	Polar (P) or Non-polar (NP)
Tetrahedral		NP if all atoms same P if atoms attached to central atom are different
Trigonal planar		NP if all atoms attached to central atom are the same P if atoms attached to central atom are different
Pyramidal		Always P
Bent		Always P
Linear		NP if symmetrical P if non symmetrical

Practice Problems

Draw the VSEPR shape for each molecule and determine if the molecule will be polar or nonpolar.

1. HF = polar

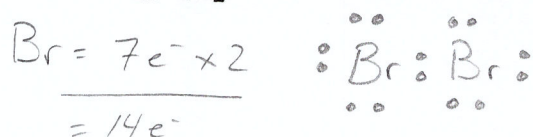


linear

$$\Delta EN = 4.0 - 2.2$$

$$\Delta EN = 1.8$$

2. Br₂ = non polar



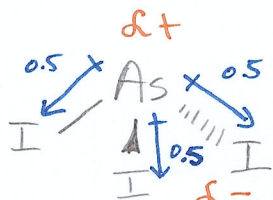
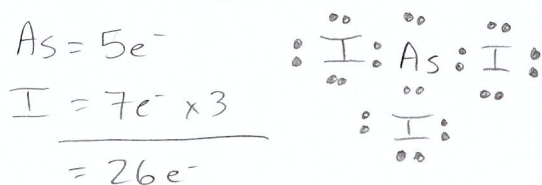
linear

$$\Delta EN = 3.0 - 3.0$$

$$\Delta EN = 0$$

$\therefore$ no bond dipole

3. AsI₃ = polar

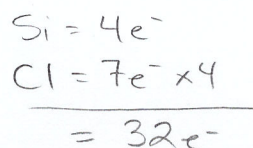


trigonal pyramidal

$$\Delta EN = 2.7 - 2.2$$

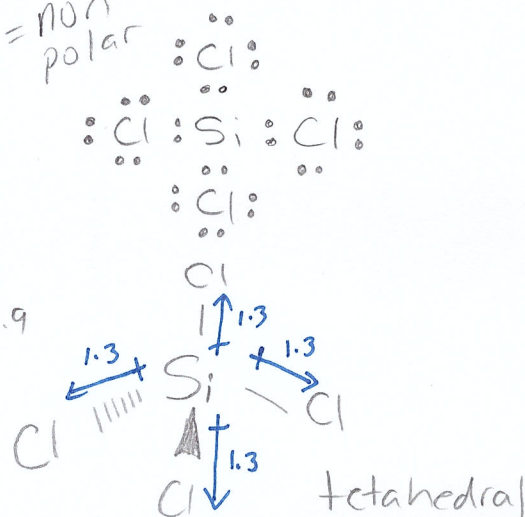
$$\Delta EN = 0.5$$

4. SiCl₄ = non polar

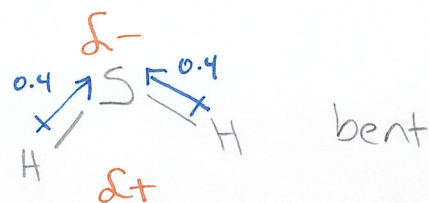
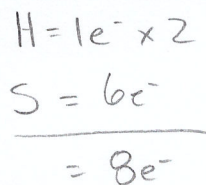


$$\Delta EN = 3.2 - 1.9$$

$$\Delta EN = 1.3$$



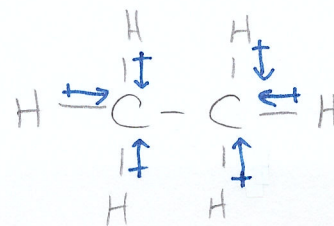
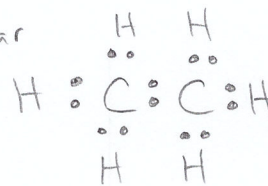
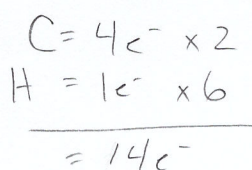
5. H₂S = polar



$$\Delta EN = 2.6 - 2.2$$

$$\Delta EN = 0.4$$

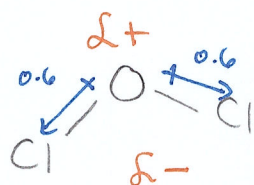
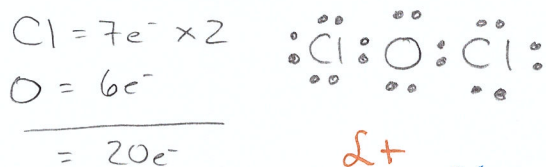
6. C₂H₆ = non polar



$$\Delta EN = 2.6 - 2.2$$

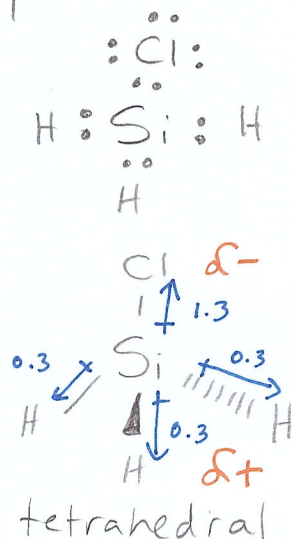
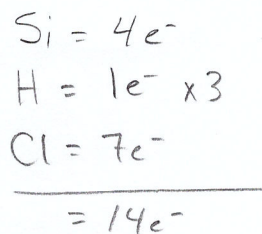
$$\Delta EN = 0.4$$

7. Cl_2O = polar



bent
 $\Delta\text{EN} = 3.2 - 2.6$
 $\Delta\text{EN} = 0.6$

10. SiH_3Cl = polar

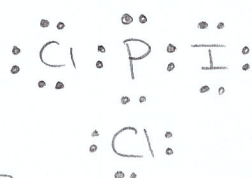
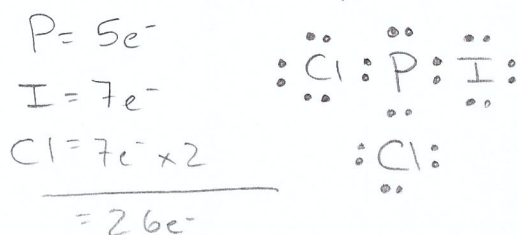


$\text{Si} - \text{Cl}$
 $\Delta\text{EN} = 3.2 - 1.9$
 $\Delta\text{EN} = 1.3$

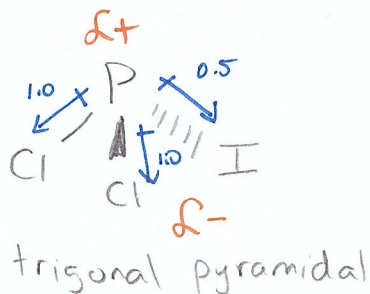
$\text{Si} - \text{H}$
 $\Delta\text{EN} = 2.2 - 1.9$
 $\Delta\text{EN} = 0.3$

tetrahedral

8. PCl_2I = polar



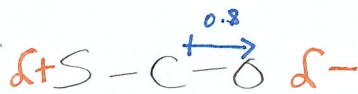
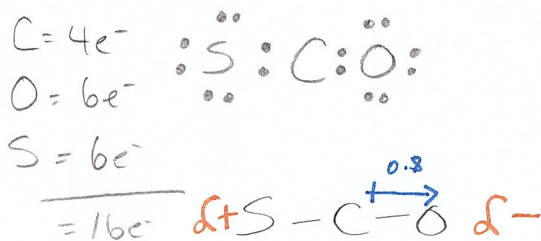
$\text{P} - \text{I}$
 $\Delta\text{EN} = 2.7 - 2.2$
 $\Delta\text{EN} = 0.5$



trigonal pyramidal

$\text{P} - \text{Cl}$
 $\Delta\text{EN} = 3.2 - 2.2$
 $\Delta\text{EN} = 1.0$

9. COS = polar

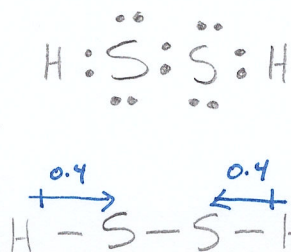
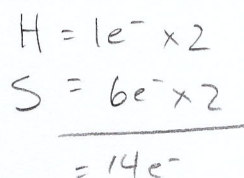


linear

$\text{S} - \text{C}$
 $\Delta\text{EN} = 2.6 - 2.6$
 $\Delta\text{EN} = 0$

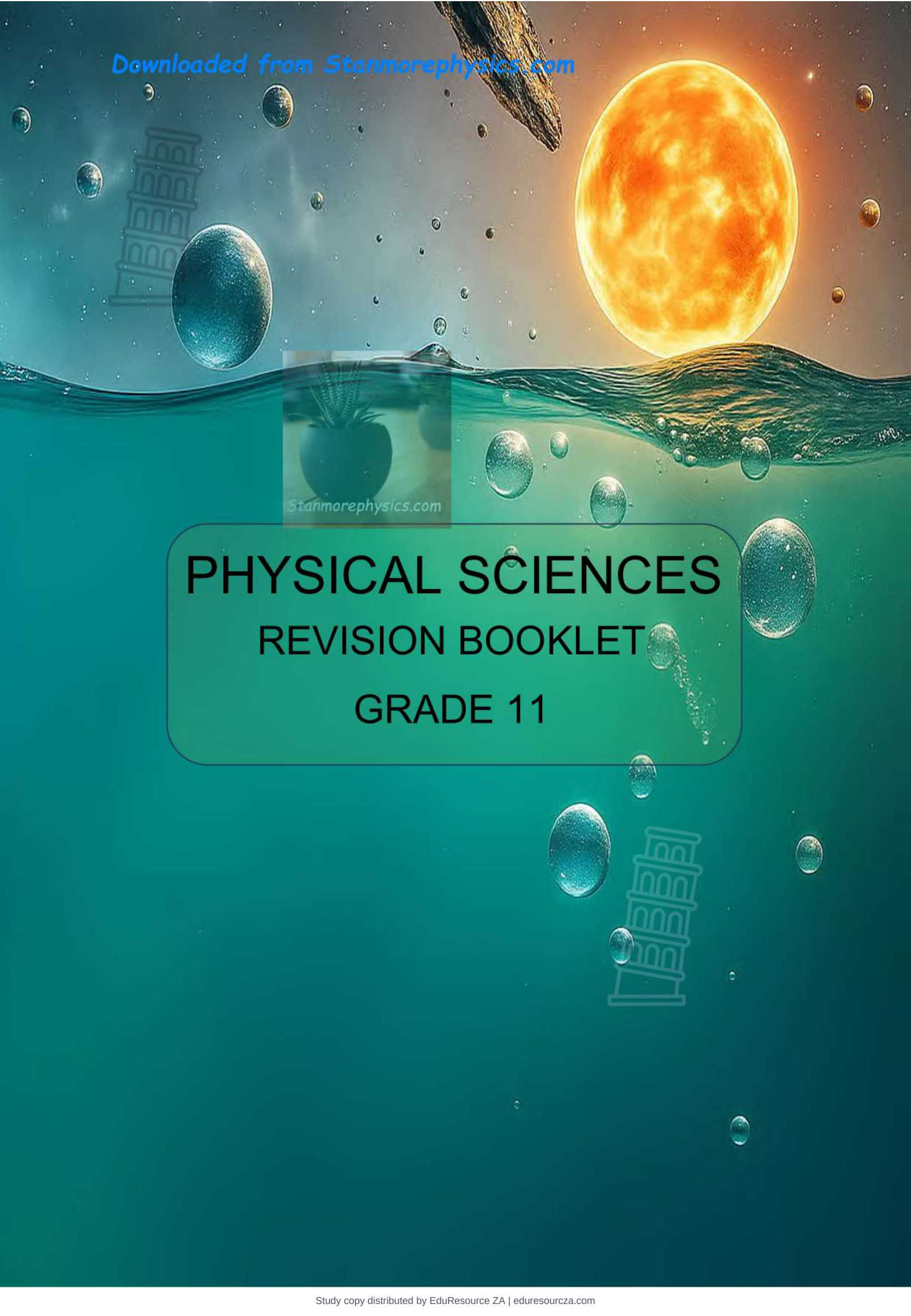
$\text{C} - \text{O}$
 $\Delta\text{EN} = 3.4 - 2.6$
 $\Delta\text{EN} = 0.8$

11. H_2S_2 = non polar



linear

$\Delta\text{EN} = 2.6 - 2.2$
 $\Delta\text{EN} = 0.4$



PHYSICAL SCIENCES

REVISION BOOKLET

GRADE 11

Secondary Schools
Directorate



education

Department of
Education
FREE STATE PROVINCE

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CREDITS

The following question papers were used to compile this book:
Department of Basic Education, *National Senior Certificate
Physical Sciences Question Papers*, 2014 – 2022, Pretoria

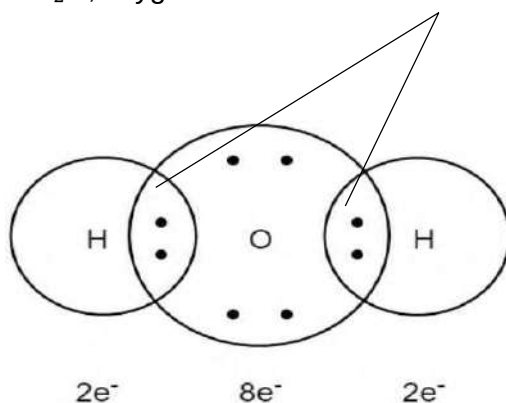
ATOMIC COMBINATIONS

- **Covalent Compounds:** Lower melting and boiling points, poor electrical conductors, may or may not dissolve in water.
- **Metallic Compounds:** Good electrical and thermal conductors, malleable, ductile, generally high melting points.

Lewis Structures (Electron Dot Diagrams)

Lewis structures represent molecules showing the valence electrons of atoms. Covalent bonds are depicted as shared pairs of dots or lines between atoms.

Example: For a molecule like H_2O , oxygen forms TWO bonds with two hydrogen atoms.



Valency

Valency is the combining power of an atom, indicating how many electrons an atom needs to lose, gain, or share to achieve a stable electron configuration.

For example, oxygen has a valency of 2 (it needs to gain 2 electrons to fill its outer shell).

RULES FOR BOND FORMATION.

IONIC BOND

Ionic bonds form when electrons are transferred from one atom to another.

Involves: Metals and non-metals.

- **Metals** (usually from Groups 1, 2, or 3 of the periodic table) lose electrons to form positive ions (cations).
- **Non-metals** (usually from Groups 15, 16, or 17) gain electrons to form negative ions (anions).
- **Bond Strength:** Strong electrostatic attraction between oppositely charged ions.
- **Example:** Sodium (Na) donates an electron to Chlorine (Cl), forming Na^+ and Cl^- , which combine to form NaCl (table salt).

Atomic combinations:

DEFINITIONS

- Define a **chemical bond** as a **mutual attraction between two atoms resulting from the simultaneous attraction between their nuclei and the outer electrons**. (NOTE: The energy of the combined atoms is lower than that of the individual atoms resulting in higher stability.)
- Define a **covalent bond** as the sharing of electrons between two atoms to form a molecule.

Molecule: A group of two or more atoms covalently bonded and that function as a unit.

Describe rules for bond formation:

- Different atoms, each with an unpaired valence electron can share these electrons to form a chemical bond e.g. two H atoms form a H_2 molecule by sharing an electron pair.
- Different atoms with **paired valence electrons**, called **lone pairs**, cannot share these four electrons and cannot form a chemical bond e.g. no bond forms between two He atoms.
- Different atoms, with unpaired valence electrons, can share these electrons and form a chemical bond for each electron pair shared. The two atoms can form multiple bonds between them. If two pairs of electrons are shared, a double bond is formed e.g. between two O atoms to form O_2 .
- Atoms with an empty valence shell can share a lone pair of electrons from another atom to form a coordinate covalent or dative covalent bond e.g. in NH_4^+ the lone pair of nitrogen is shared with H^+ and in H_3O^+ the lone pair of oxygen is shared with H^+ .
- Define a **bonding pair** as a **pair of electrons that is shared between two atoms in a covalent bond**.
- Define a **lone pair** as a **pair of electrons in the valence shell of an atom that is not shared with another atom**.
- Describe the formation of the dative covalent (or coordinate covalent) bond by means of electron diagrams using NH_4^+ and H_3O^+ as examples.
- Define **electronegativity** as a **measure of the tendency of an atom in a molecule to attract bonding electrons**.
- Describe a **non-polar covalent bond** as a **bond in which the electron density is shared equally between the two atoms**. An example is the bond between two H atoms.
- Describe a **polar covalent bond** as a **bond in which the electron density is shared unequally between the two atoms**. An example is the bond between an H atom and a Cl atom.

State the major principles used in the VSEPR:

- Molecular shape is determined by the repulsions between electron pairs present in the valence shell of the central atom.
- The number of electron pairs around the central atom can be determined by writing the Lewis structure for the molecule.
- The shape of the molecule depends on the number of bonding electron groups (or atoms bonded to the central atom) and the number of lone pairs on the central atom.
- A** is used to represent the central atom and **X** is used to represent terminal atoms.
- There are five ideal shapes found when there are NO lone pairs on the central atom, ONLY bond pairs.
- Define **electronegativity** as a **measure of the tendency of an atom in a molecule to attract bonding electrons**.
- Describe a **non-polar covalent bond** as a **bond in which the electron density is shared equally between the two atoms**. An example is the bond between two H atoms.

- Describe a **polar covalent bond** as **a bond in which the electron density is shared unequally between the two atoms**. An example is the bond between an H atom and a Cl atom.
- Define **bond energy** of a compound as **the energy needed to break one mole of its molecules into separate atoms**.
- Define **bond length** as **the average distance between the nuclei of two bonded atoms**.

COVALENT BONDS

Covalent bonds form when atoms share electrons.

Involves: Non-metals only.

- Each atom contributes one or more electrons to be shared in the bond.
- The shared electrons allow each atom to achieve a stable electron configuration (usually achieving a full outer shell, like noble gases).

Types:

- **Single bond:** One pair of electrons is shared (e.g., H_2).
- **Double bond:** Two pairs of electrons are shared (e.g., O_2).
- **Triple bond:** Three pairs of electrons are shared (e.g., N_2).

Example: Two oxygen atoms share two pairs of electrons to form an O_2 molecule.

METALLIC BONDS

Metallic bonds occur when atoms in a metal release some of their electrons into a "sea of electrons."

Involves: Metals only.

- The electrons are free to move throughout the entire metal structure.
- The positively charged metal ions are held together by their attraction to the sea of delocalized electrons.

Properties: This gives metals their characteristic properties such as conductivity, malleability, and ductility.

Example: Copper (Cu) atoms are held together in a metallic bond.

STRUCTURED QUESTIONS

QUESTION 1 (November Exemplar)

Ammonia (NH_3) is an important gas used in the preparation of fertilisers. An ammonia molecule is formed when electrons are shared between three hydrogen atoms and a nitrogen atom.

- 1.1 Name the type of chemical bond formed between a hydrogen and a nitrogen atom. (1)
- 1.2 How many valence electrons does a nitrogen atom have? (1)
- 1.3 Write down a Lewis structure for the ammonia molecule. (2)
- 1.4 For the ammonia molecule, write down the:
 - 1.4.1 Number of electron pairs surrounding the central atom (1)
 - 1.4.2 Number of atoms surrounding the central atom (1)
 - 1.4.3 Name used to describe the molecular shape (1)

Ammonia dissolves readily in water to form ammonium ions, $\text{NH}_4^+(\text{aq})$. An ammonium ion is formed when an ammonia molecule shares a lone pair of electrons with a hydrogen ion.

- 1.5 Name the type of bond formed between an ammonia molecule and a hydrogen ion. (1)
- 1.6 Represent the formation of an ammonium ion with the aid of Lewis structures. (4)
- 1.7 For the ammonium ion, write down the:
 - 1.7.1 Number of atoms surrounding the central atom (1)
 - 1.7.2 Name used to describe the molecular shape (1)

The nitrogen atom can also bond with itself to form the nitrogen molecule.

- 1.8 Which ONE of the following bonds will be the strongest?

I: Bond between a nitrogen atom and a hydrogen atom

OR

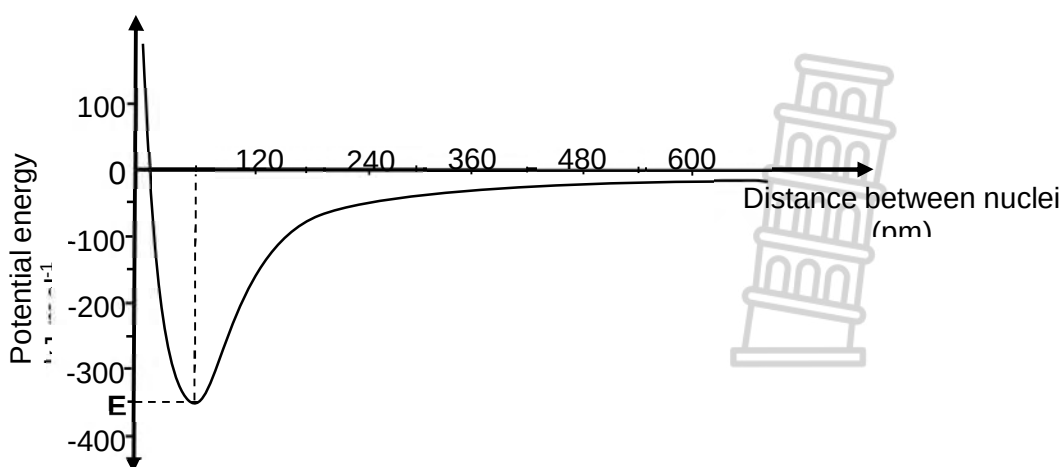
II: Bond between a nitrogen atom and a nitrogen atom

Write down **I** or **II**. Give a reason for the answer.

(2)
[16]

QUESTION 2 (November 2014)

The graph below shows the change in energy that takes place when a hydrogen (H) atom approaches a bromine (Br) atom.



- 2.1 Define the term *bond length*. (2)
- 2.2 From the graph, write down the:
 - 2.2.1 Bond length, in pm, of the H-Br bond (2)
 - 2.2.2 Energy, in $\text{kJ}\cdot\text{mol}^{-1}$, needed to break the H-Br bond (2)

- 2.2.3 Name of the potential energy represented by E (1)
- 2.3 How will the bond length of an H-F bond compare to that of the H-Br bond? (2)
- Write down EQUAL TO, SHORTER THAN or LONGER THAN. Give a reason for the answer. [9]

QUESTION 3 (November 2014)

Both aluminium fluoride (AlF_3) and phosphorous trifluoride (PF_3) contain fluorine. Aluminium fluoride is a colourless solid used in the production of aluminium, whilst phosphorous trifluoride is a poisonous, colourless gas.

- 3.1 Explain the difference between a *covalent bond* and an *ionic bond*. (2)
- 3.2 Name the type of chemical bond between particles in: (1)
- 3.2.1 AlF_3 (1)
- 3.2.2 PF_3 (1)
- 3.3 Draw the Lewis structures for: (3)
- 3.3.1 AlF_3 (2)
- 3.3.2 PF_3 (1)
- 3.4 Write down the molecular shape of PF_3 . (1)
- 3.5 The melting point of AlF_3 is $1\,291\text{ }^\circ\text{C}$ and that of PF_3 is $-151\text{ }^\circ\text{C}$. (4)
- Fully explain this difference in melting point. [14]

QUESTION 4 (November 2015)

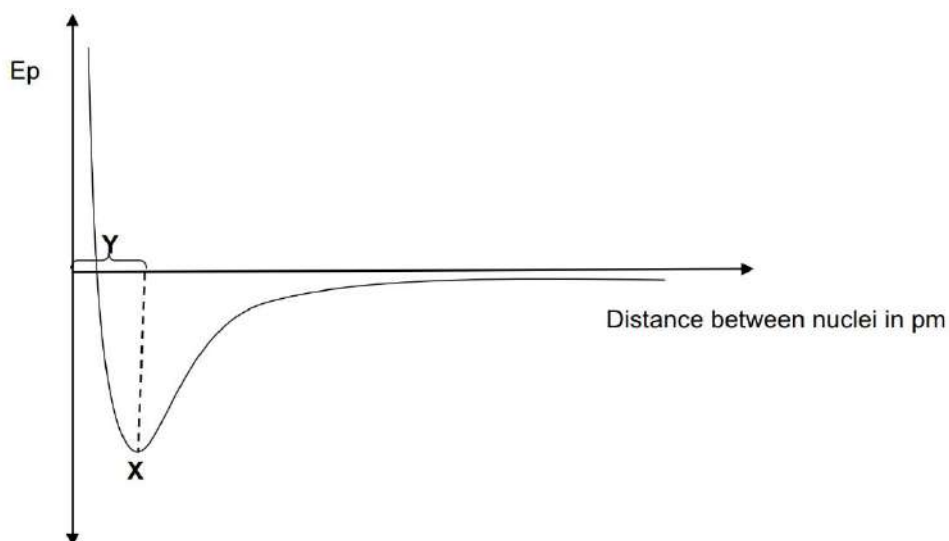
Molecules such as CO_2 and H_2O are formed through covalent bonding.

- 4.1 Define the term *covalent bonding*. (2)
- 4.2 ONE of the above molecules has lone pairs of electrons on the central atom. Draw the Lewis diagram for this molecule. (2)
- 4.3 H_3O^+ is formed when H_2O forms a dative covalent bond with an H^+ ion. (1)
- 4.3.1 Draw the Lewis diagram for H_3O^+ . (2)
- 4.3.2 State TWO conditions for the formation of such a bond. (2)
- 4.4 The polarity of molecules depends on the DIFFERENCE IN ELECTRONEGATIVITY and the MOLECULAR SHAPE. (2)
- 4.4.1 Define the term *electronegativity*. (1)
- 4.4.2 Calculate the difference in electronegativity between: (1)
- (a) C and O in CO_2 (1)
- (b) H and O in H_2O (6)
- 4.4.3 Explain the difference in polarity between CO_2 and H_2O by referring to the polarity of the bonds and the shape of the molecules. [17]

QUESTION 5 (November 2016)

Electronegativity of atoms may be used to explain the polarity of bonds.

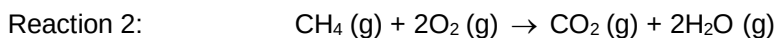
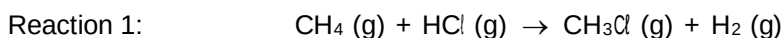
- 5.1 Define the term *electronegativity*. (2)
- 5.2 Draw the Lewis diagram of an oxygen difluoride molecule. (2)
- 5.3 Calculate the electronegativity difference between O and F in oxygen difluoride and predict the polarity of the bond. (2)
- 5.4 A polar bond does not always lead to a polar molecule. Explain the statement by referring to OF_2 and CO_2 molecules. In your explanation, include the polarity of the bonds and the shape of the molecules. (4)
- 5.5 The diagram below shows the energy change that takes place when two atoms move towards each other.



- 5.5.1 What does **X** and **Y** represent? (2)
- 5.5.2 Define the concept represented by **X**. (2)
- 5.5.3 Explain the relationship between bond order, bond length and bond energy. (3)
- [17]**

QUESTION 6 (November 2017)

Consider the following two reactions of methane (CH_4):



- 6.1 Define the term *covalent bond*. (2)
- 6.2 Draw Lewis structures for: (2)
- 6.2.1 CH_3Cl (2)
- 6.2.2 CO_2 (2)
- 6.3 How many lone-pair electrons are on the central atom in the CO_2 molecule? (1)
- 6.4 Identify ONE of the substances in Reaction 2 that can form a dative covalent bond when reacting with an acid. (1)
- 6.5 Write down the shape of the: (1)
- 6.5.1 H_2O molecule (1)

6.5.2 CO₂ molecule (1)

- 6.6 Although the molecules of CH₄ and CH₃Cl have the same shape, CH₄ is non-polar, while CH₃Cl is polar. Give a reason for the difference in molecular polarity. (1)
[11]

QUESTION 7 (November 2018)

Hydrogen cyanide (HCN) is a very poisonous compound used in the manufacturing of plastics, mining of gold and as a poison.

- 7.1 Define the term *chemical bond*. (2)
7.2 Draw Lewis structures for: (2)
7.2.1 HCN (2)
7.2.2 H₂O (2)
7.3 What is the shape of the HCN molecule? (1)
7.4 Calculate the electronegativity difference for the CN bond. (1)
7.5 What is polarity of the HCN molecule? Write only POLAR or NON-POLAR. (1)

The table below indicates the values of the bond length and bond energy of the different bonds in HCN.

BOND	BOND LENGTH (nm)	BOND ENERGY (kJ·mol ⁻¹)
CH	0,109	413
CN	0,116	890

- 7.6 Explain why the bond energy of the CN bond is more than the bond energy of the CH bond. (2)
7.7 Explain the difference between the *bond length of the CH bond* and the *bond length of the CN bond*. (2)
7.8 Will HCN be soluble in water? Write only YES or NO. (1)
7.9 Explain the answer to QUESTION 2.8 by referring to the polarity and intermolecular forces of the compounds. (3)
[17]

QUESTION 8 (November 2019)

- 8.1 Ammonia NH₃(g) and hypochlorous acid HOCl(l) are both examples of covalent compounds. (2)
8.1.1 Define the term *bonding pair*. (2)
8.1.2 Draw Lewis structures for the following molecules: (2)
(a) NH₃ (2)
(b) HOCl (2)
8.1.3 Write down the: (1)
(a) Number of bonding pairs in NH₃ (1)
(b) Number of lone pairs on the oxygen atom in HOCl (1)
(c) Shape of an ammonia molecule (1)
8.1.4 Which bond, N-H or O-H, is more polar? Give a reason for the answer. (2)
8.1.5 Write down the type of intermolecular forces present in BOTH ammonia and hypochlorous acid. (1)

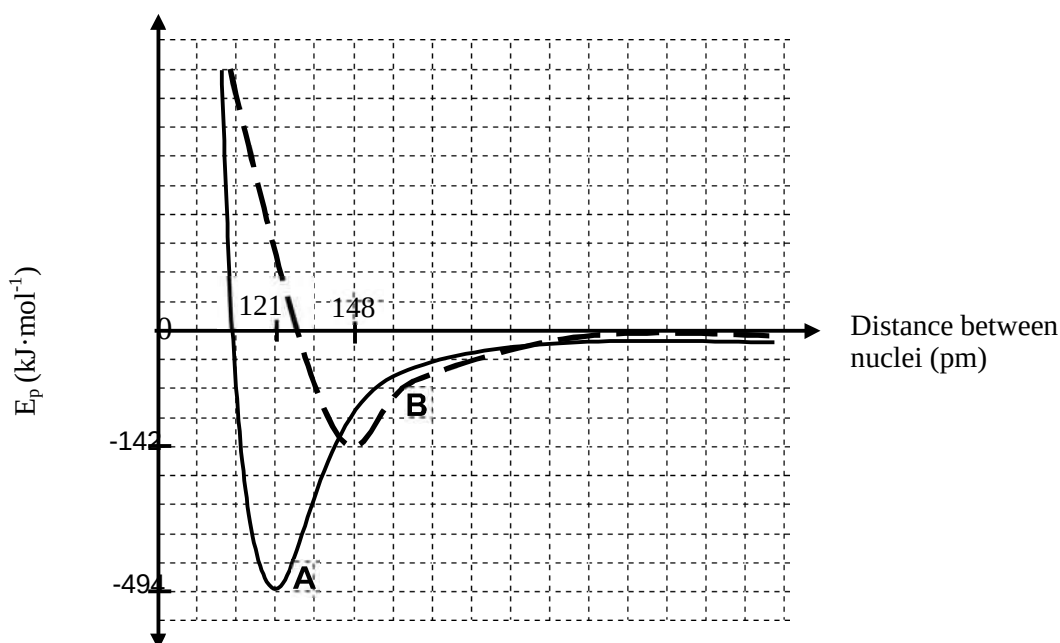
8.1.6 When ammonia dissolves in water, the ammonium ion (NH_4^+) is formed.

What type of bond forms between the ammonia molecule and the hydrogen ion?

(1)

8.2 The graph of potential energy versus distance between the nuclei of two Oxygen atoms during bond formation is shown below.

Graph of potential energy versus distance between nuclei



8.2.1 Define the term *bond energy*.

(2)

8.2.2 Which curve, **A** or **B**, represents the formation of the double bond ($\text{O}=\text{O}$) between oxygen atoms? Briefly explain the answer.

(3)

8.2.3 Write down the bond length of the bond represented by curve **B**.

(1)

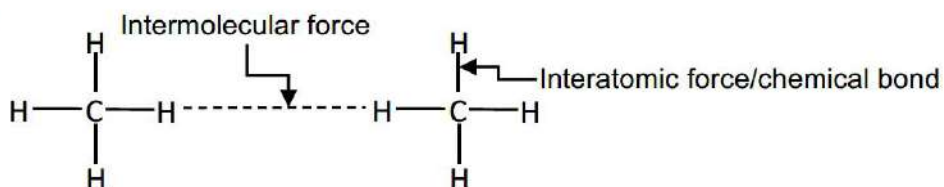
[19]

INTERMOLECULAR FORCES

TERMS AND DEFINITIONS	
Boiling point	The temperature at which the vapour pressure of a substance equals atmospheric pressure.
Melting point	The temperature at which the solid and liquid phases of a substance are at equilibrium.
Vapour pressure	The pressure exerted by a vapour at equilibrium with its liquid in a closed system.
Solubility	The property of a solid, liquid, or gaseous chemical substance (solute) to dissolve in a solid, liquid, or gaseous solvent to form a homogeneous solution

Understanding **bonds** and **intermolecular forces**:

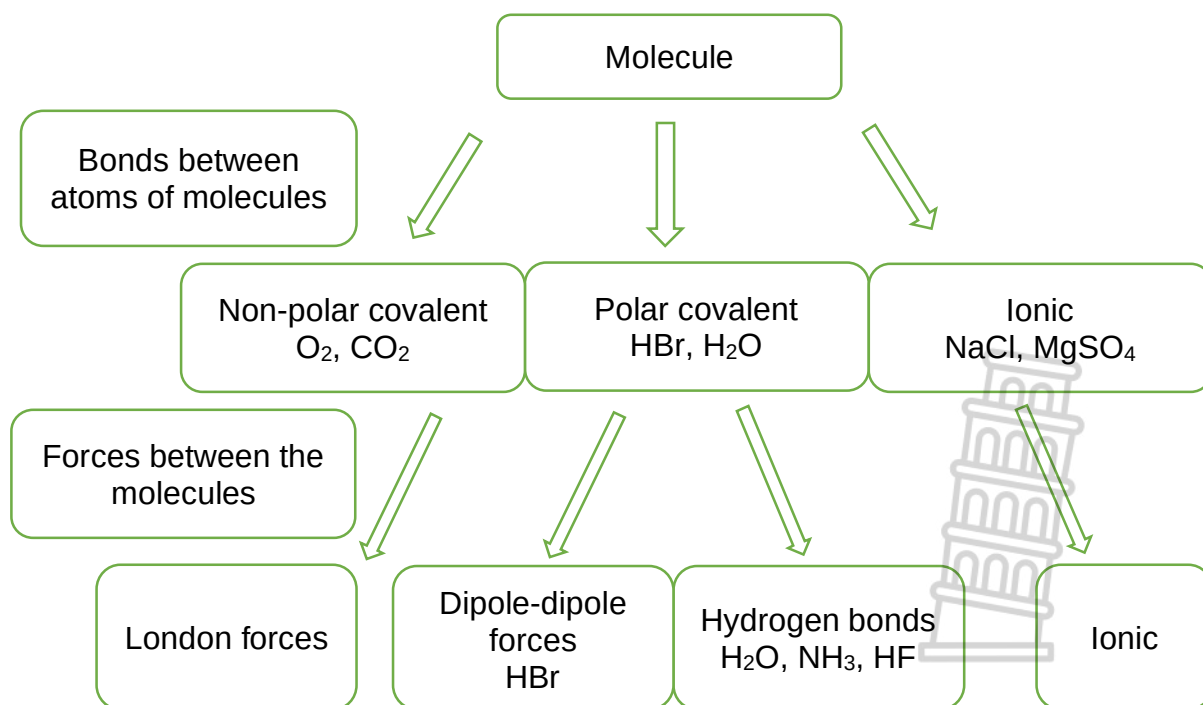
Example:



Bonds: between atoms. This is the force

that holds atoms together within a molecule - intramolecular force. Covalent bonds are such an example.

Intermolecular forces: This is the force that holds molecules together. The strength of intermolecular forces will influence properties like boiling-, melting points and vapour pressure.



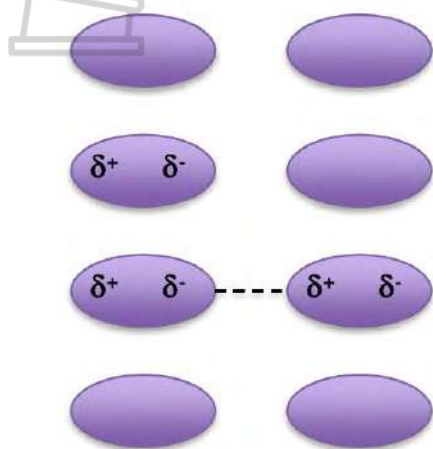
INTERMOLECULAR FORCES

London forces

Sometimes called induced dipole forces or just dispersion forces. Temporary dipole attractions between non-polar molecules that form due to shifting electrons.

This is the only type of force between non-polar molecules.

Bigger molecules or atoms usually have stronger dispersion forces. (More electrons)



- 1) Two nonpolar molecules with no attractive forces between them.
- 2) As electrons shift within one of the molecules, a temporary dipole may appear.
- 3) An adjacent molecule will be attracted to the molecule with the temporary dipole and a new dipole within the second molecule will be induced. This creates the London dispersion force.
- 4) The electrons move back, and the temporary dipoles disappear. This makes the London forces a weak force, it is only temporary.

Dipole-Dipole Forces

A permanent dipole force exists between polar molecules. Attractions form between the partially positive and partially negative ends of adjacent polar molecules. The opposite dipoles align themselves to create attractive forces.

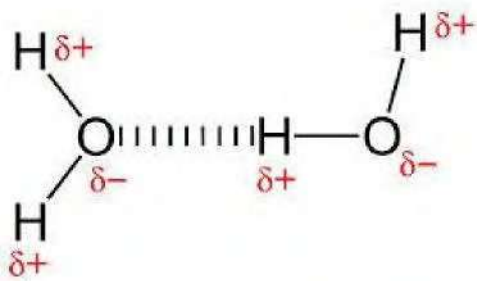


Dipole forces are usually stronger than London forces since the dipoles are permanent.

Only polar molecules can form dipole-dipole forces.

Hydrogen Bonds

An especially strong dipole force exists between molecules containing H-F, H-O or H-N bonds. (These bonds are highly polar due to the large electronegativity difference.)

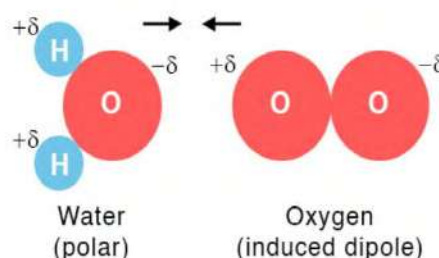


A very strong type of intermolecular force between specific polar molecules.

Take note: Hydrogen bond also occur between organic molecules: CH_3OH , $\text{CH}_3\text{CH}_2\text{OH}$ etc. because of the H-O strong dipole.

Dipole-induced dipole forces

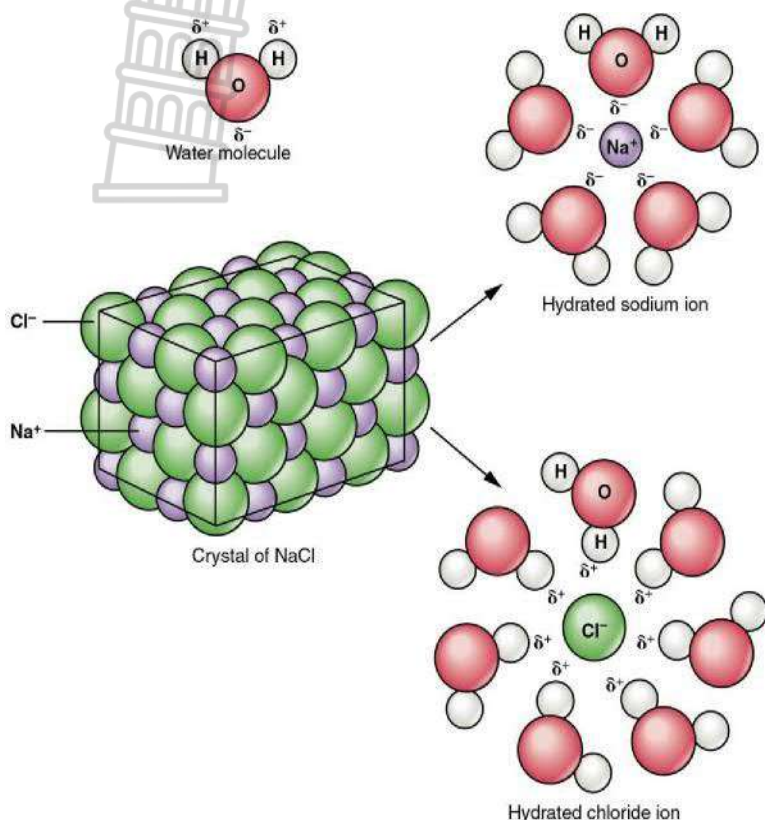
Forces between polar and non-polar molecules



Terms, definitions, questions & answers

Ion-Dipole Forces

When an ionic compound such as NaCl dissolves in water, the water molecules arrange their oppositely charged dipole to be attracted to the fully charged ion, creating a very strong attractive force called an ion-dipole force.



The partial negative charge on the water molecule is attracted to the fully charged positive sodium ion (Na^+).

The partial positive charge on the water molecule is attracted to the fully charged negative chloride ion (Cl^-).

Boiling point

The stronger the intermolecular forces, the higher the boiling point.

Melting point:

The stronger the intermolecular forces, the higher the melting point.

Vapour pressure

The stronger the intermolecular forces, the lower the vapour pressure.



STRUCTURED QUESTIONS

QUESTION 1 (November Exemplar 2013)

The table below shows the boiling points of the hydrides of group IV (compounds in which hydrogen is bonded to elements from group IV in the periodic table).

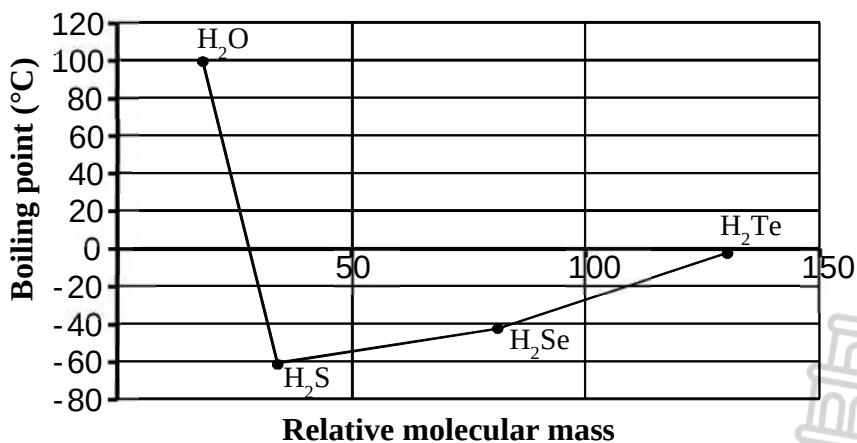
BOILING POINTS OF HYDRIDES FROM GROUP IV

HYDRIDES OF GROUP IV	RELATIVE MOLECULAR MASS	BOILING POINT (°C)
CH ₄	16	-164
SiH ₄	32	-112
GeH ₄	77	-89
SnH ₄	123	-52

- 1.1 What is the phase (solid, liquid or gas) of the hydrides above at 25 °C? (1)
- 1.2 Name the type of Van der Waals forces between molecules of the hydrides in the table above. (1)
- 1.3 Explain the trend in boiling points observed for the hydrides in the above table. In your explanation, refer to molecular size, intermolecular forces and the energy needed. (3)

The graph below shows the boiling points of the hydrides of group VI in the periodic table versus their relative molecular masses.

GRAPH OF BOILING POINT VERSUS RELATIVE MOLECULAR MASS



- 1.4 From the graph above, deduce and then write down the NAME of the hydride:
 - 1.4.1 With the weakest intermolecular forces (1)
 - 1.4.2 With hydrogen bonds between molecules (1)
 - 1.4.3 That requires the most energy to undergo a phase change (1)
 - 1.5 Refer to intermolecular forces and energy and give a reason why one of the hydrides of group VI deviates from the trend in boiling point observed for the others. (2)
- [10]**

QUESTION 2 (November 2014)

The boiling points of four compounds of hydrogen, represented by the letters **P**, **Q**, **R** and **S**, are given in the table below.

	Formula	Boiling point (°C)
P	CH ₄	-164
Q	NH ₃	-33
R	H ₂ O	100
S	SiH ₄	-112

- 2.1 Define the term *boiling point*. (2)
- 2.2 Fully explain the difference in boiling points between compound **P** and: (2)
- 2.2.1 Compound **Q** (3)
- 2.2.2 Compound **S** (3)
- 2.3 Explain why the boiling points of compounds **Q** and **R** differ by referring to ELECTRONEGATIVITY and DEGREE OF POLARITY. (2)
- 2.4 Write down the letter from the table that represents the following: (1)
- 2.4.1 ONE polar compound (1)
- 2.4.2 ONE non-polar compound (1)
- [12]

QUESTION 3 (November 2015)

- 3.1 The boiling point of compounds **A** to **E** are given in the table below.

COMPOUND	FORMULA	BOILING POINT (°C)
A	CH ₄	-164
B	C ₂ H ₆	-89
C	C ₅ H ₁₂	36
D	C ₆ H ₁₄	69
E	C ₂₀ H ₄₂	343

- 3.1.1 Define the term *boiling point*. (2)
- 3.1.2 Calculate the molecular mass of compound **D**. (1)
- 3.1.3 In what phase is compound **B** at 25 °C? (1)
- 3.1.4 Name the type of intermolecular force present in compound **A**. (1)
- 3.1.5 Explain why the boiling point increases from compound **A** to **E**. (3)
- 3.1.6 How does the vapour pressure of compound **B** compare to the vapour pressure of compound **C**? Write down only HIGHER THAN, LOWER THAN or EQUAL TO. (1)
- 3.2 Consider the boiling points of the compounds in the table below.



SUBSTANCE	BOILING POINT (°C)
H ₂ S	-60
NH ₃	-33
H ₂ O	100

- 3.2.1 Which ONE of the substances in the table above has the weakest forces between its molecules? (1)
- 3.2.2 Name the type of intermolecular force found between NH₃ molecules. (1)
- 3.2.3 Explain the following statement:

Although the shape of the molecules of H₂S and H₂O is similar, there is a remarkable difference in their boiling points. (4)

[15]

QUESTION 4 (November 2016)

Learners conduct an experiment to investigate the effects of intermolecular forces on boiling points. They use 20 ml of each of the following compounds in their investigation: water, sunflower oil, nail polish remover, glycerine and methylated spirits. The results are shown in the table below:

NAME OF COMPOUND	BOILING POINT (°C)
water	93
sunflower oil	230
nail polish remover	56
glycerine	290
methylated spirits	62

- 4.1 Define the term *boiling point*. (2)
- 4.2 Formulate an investigative question for this experiment. (2)
- 4.3 In which compound in the table above will the strongest intermolecular forces occur? Give a reason for the answer. (2)
- 4.4 The learners now use 40 ml of each of the compounds above in the experiment. Will it affect the boiling points? Choose YES or NO. Give a reason for the answer. (2)
- 4.5 Methylated spirits is highly flammable. State TWO safety precautions that should be taken when using methylated spirits in the laboratory. (2)
- 4.6 Which compound in the table above will have the highest rate of evaporation? Give reason for the answer. (2)
- 4.7 Sunflower oil is a non-polar compound with induced dipole forces between the molecules, while water is a polar molecule with hydrogen bonds between its molecules. Explain why the boiling point of sunflower oil is higher than the boiling point of water. (2)

[14]

QUESTION 5 (November 2017)

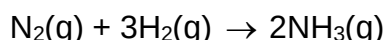
Consider the list of six substances with their formulae and boiling points in the table below.

NAME OF SUBSTANCE	FORMULA	BOILING POINT (°C)
Water	H ₂ O	100
Ethanol	CH ₃ CH ₂ OH	78
Bromine	Br ₂	58,8
Iodine	I ₂	184,3
Ammonia	NH ₃	-33,3
Phosphine	PH ₃	-87,7

- 5.1 Explain why ethanol is soluble in water. Refer to the relative strength of the intermolecular forces in ethanol and water. (3)
- 5.2 Explain why the boiling point of iodine is higher than that of bromine. Refer to the intermolecular forces present in EACH substance in the explanation. (3)
- 5.3 Explain why phosphine will evaporate faster than ammonia by referring to the types of intermolecular forces present in EACH substance. (4)
- 5.4 Water, ethanol and bromine are all liquids at room temperature. Which ONE will have the highest vapour pressure? (1)
- 5.5 Give a reason for the answer to QUESTION 5.4 by referring to the relative strength of the intermolecular forces and boiling points. (2)
- [13]

QUESTION 6 (November 2018)

The reaction below is used in the Haber process to manufacture ammonia.



The boiling points of the substances in the reaction are as follows:

SUBSTANCE	BOILING POINT (°C)
H ₂	-252,9
N ₂	-195,8
NH ₃	-33,3

- 6.1 Refer to the intermolecular forces and explain the difference in boiling point between NH₃ and N₂. (3)
- 6.2 Write down the FORMULA of the substance in the table that will have the lowest melting point. (1)
- 6.3 Explain why H₂ will evaporate faster than N₂. Refer to the type and relative strength of the intermolecular forces. (3)
- 6.4 Write down the FORMULA of the substance in the table that will have the highest vapour pressure. Explain your answer. (3)
- [10]

QUESTION 7 (November 2019)

The melting points and boiling points of four substances (**A**, **B**, **C** and **D**) are shown in the table below.

	SUBSTANCES	MELTING POINT (°C)	BOILING POINT (°C)
A	HF	- 83,11	19,54
B	HCl	- 114,2	- 81,7
C	CS ₂	- 111	46,0
D	CO ₂	- 56,6	- 78,5

- 7.1 Define the term *melting point*. (2)
- 7.2 Explain the difference in melting points of HF and HCl by referring to the TYPE of intermolecular forces. (4)
- 7.3 Which ONE of the substances (**A**, **B**, **C** or **D**) above is a liquid at 25 °C? (1)
- 7.4 Explain why CS₂ has a higher boiling point than CO₂. (3)
- 7.5 Which ONE of the substances (**A**, **B**, **C** or **D**) above has the highest vapour pressure? Give a reason for the answer by referring to the data in the table. (2)

[12]



PHYSICAL SCIENCES Grade 11 TERM 1 RESOURCE PACK

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TOPIC 3: Atomic Combinations (Molecular Structure)

WORKSHEET

MULTIPLE CHOICE

- The type of particle which results from the covalent bonding of atoms is called:
A an ion.
B an atom.
C a molecule.
D an isotope. (2)
- The bond in the H_2O molecule is described as polar covalent because :
A the shared electrons are closer to the oxygen atom than the hydrogen atom.
B the shared electrons are closer to the hydrogen atom than the oxygen atom.
C there is a large electronegativity difference between the two atoms.
D electrons are transferred from hydrogen atoms to oxygen atoms. (2)
- The CCl_4 molecule is considered to be non-polar because :
A the shape of the molecule is linear.
B the electronegativity difference is very small.
C the molecule is asymmetrical.
D the molecule is symmetrical. (2)
- Consider the methane molecule (CH_4). How much energy is required to break the chemical bonds in the molecule (in $\text{kJ}\cdot\text{mol}^{-1}$)?
A 412
B 348
C 1648
D 1996

BOND	BOND ENERGY
C – C	348
C = C	619
C $\equiv$ C	835
C - H	412
C = O	799
O - H	463
O = O	499

(2)

RESOURCE PACK

LONG QUESTIONS

1. Copy and complete the following table:

MOLECULE	CO ₂	CH ₄	SF ₆	NH ₃
What is the central atom in the molecule?				
Shape of molecule				
Are the molecular shapes IDEAL or NON-IDEAL?				
Does the molecule contain polar bonds?				
Is the molecule polar or non-polar?				

(20)

2. A water molecule is polar while a carbon dioxide molecule is non-polar.

2.1 What is the difference between a polar and a non-polar molecule? (4)

2.2 Use electronegativity difference to determine the type of bond present in the carbon dioxide molecule. (2)

2.3 State the shapes of water and carbon dioxide molecules. (2)

2.4 Explain why carbon dioxide, which contains polar bonds, is non-polar. (2)

3. What is the nature of the charges on the H and O atoms in the H₂O molecule? Explain your answer. (2)

4. H₂O can form a dative covalent bond to form H₃O⁺. Use the Lewis diagram for a water molecule to help explain the idea of a dative covalent bond. (6)

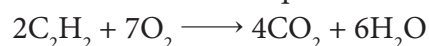
TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

5. Consider the following list of molecules:
 Cl_2 H_2O BF_3 HCl CO_2 NH_3
- 5.1 Which of the following molecules is polar covalent and which are non-polar covalent? (6)
- 5.2 CO_2 contains polar bonds, but is considered to be a non-polar molecule. Explain why this is so. (3)
- 5.3 Name the shape of each molecule in the list. (7)
6. Consider a molecule of hydrogen sulfide (H_2S).
- 6.1 Draw the Lewis diagram for hydrogen sulfide. (2)
- 6.2 Identify the central atom in H_2S . (1)
- 6.3 How many lone pairs of electrons are present in H_2S molecule? (1)
- 6.4 How many bonding pairs of electrons are present in the H_2S molecule. (1)
- 6.5 What is the molecular shape of this molecule? (1)
- 6.6 Is this an IDEAL or a NON-IDEAL molecular shape? Explain your answer. (3)
7. The table below lists the bond energies of certain covalent bonds.

BOND	O = O	O - O	C - C	C = O	C - H	H - H	H - O
BOND ENERGY (kJ·mol ⁻¹)	494	142	346	499	413	435	459

- 7.1 Use the information from the table to compare the length and strength of an O – O bond to the length and strength of an O = O bond. (3)

Consider the chemical equation shown below:



- 7.2 Calculate the total energy absorbed to break the bonds of the reactants. (5)
- 7.3 Calculate the total energy released on bond formation of the products. (4)
- 7.4 How much excess energy was released in the whole reaction? (3)

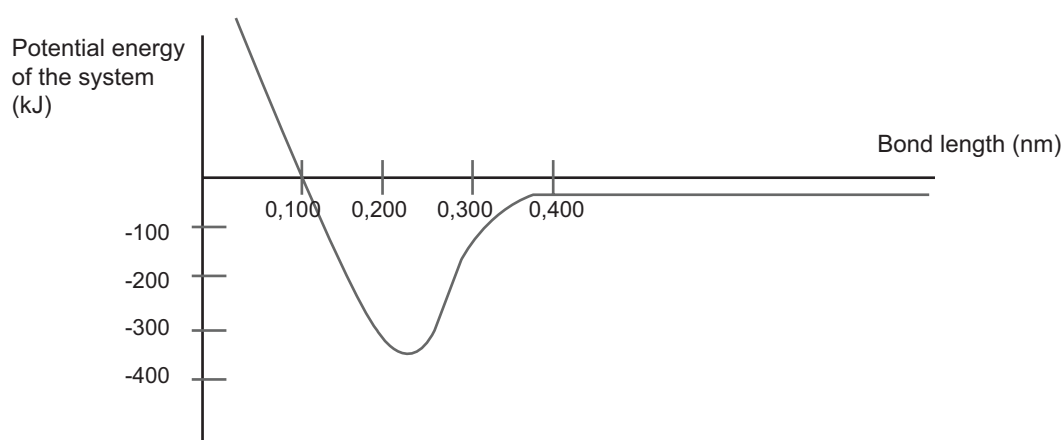
RESOURCE PACK

CONSOLIDATION EXERCISE

MULTIPLE CHOICE

1. A covalent bond will form between two atoms when...
- A there is repulsion between the nuclei of two approaching atoms.
 - B there is attraction between the electrons of two approaching atoms.
 - C one atom transfers an electron to another atom.
 - D the potential energy of the atoms is at its lowest. (2)
2. In VSEPR theory, the greatest amount of repulsion between orbitals will be between:
- A lone pair and lone pair.
 - B bonding pair and lone pair.
 - C bonding pair and bonding pair.
 - D bonding pair and central atom. (2)
3. The HBr molecule forms a dipole because the...
- (i) Br atom is more electronegative than the H atom.
 - (ii) H atom is more electronegative than the Br atom.
 - (iii) HBr molecule is polar.
- Which of the above statements is/are true?
- A Only (i) is true.
 - B Both (i) and (iii) are true.
 - C Both (ii) and (iii) are true.
 - D Only (ii) is true. (2)

Study the energy diagram below which represents the formation of a covalent bond between two atoms. Use this diagram to answer Questions 4 and 5.



TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

4. The bond energy of the of the chemical bond is ...

A 0 kJ
B 30 kJ
C 330 kJ
D 200 kJ

(2)

5. The bond length between the two atoms forming the chemical bond is ...

A 0,200 nm
B 0,370 nm
C 0,240 nm
D 0,050 nm

(2)

6. Using the table alongside, calculate the total energy required to break the bonds of the reactants when methane reacts with oxygen to form carbon dioxide and water. The reaction equation is written below.



BOND	BOND ENERGY (kJ·mol ⁻¹)
C - C	348
C = C	619
C ≡ C	835
C - H	412
C = O	799
O - H	463
O = O	499

A 911 kJ·mol⁻¹
B 1 410 kJ·mol⁻¹
C 2 646 kJ·mol⁻¹
D 3 644 kJ·mol⁻¹

(2)

RESOURCE PACK

LONG QUESTIONS

1. Consider the table below. Redraw this table in your answer books and complete it in the spaces provided.

CHEMICAL FORMULA OF COMPOUND	ELECTRO- NEGATIVITY DIFFERENCE	NON-POLAR COVALENT, POLAR COVALENT OR IONIC BOND	SHAPE OF MOLECULE	TYPE OF MOLECULE (POLAR OR NON-POLAR)
Cl ₂				
CF ₄				
CO ₂				
HB _r				
NH ₃				
H ₂ O				
BF ₃				

(28)

2. Draw Lewis structures for:

2.1 a water molecule.

2.2 a carbon dioxide molecule. (4)

3. 3.1 Define the term electronegativity. (2)

3.2 Consider the following molecules:



a) Assign electronegativity values to each atom. (6)

b) Calculate the electronegativity difference for the bonds in each molecule. (6)

c) Identify which of the molecules contain polar covalent or non-polar covalent bonds. (6)

d) Identify the shapes of each molecule using VSEPR theory. (6)

4. Consider the following list of molecules:



4.1 Draw Lewis diagrams for the molecules HF and NH₃ to show the bonding that takes place. (4)

4.2 Which of the molecules in the list are non-polar? Explain your answer. (6)

4.3 According to VSEPR theory, what are the shapes of molecules BCl₃, CF₄, NH₃ and SCl₆? (4)

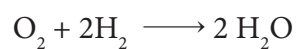
TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

- 4.4** NH_3 is said to have a non-ideal shape where SCl_6 is said to have an ideal shape. What is the difference between molecules with these different shape classifications? (2)
- 5.** For each of the following bonding pairs, say which bond is more polar. Show all calculations and indicate the partially positive (δ^+) and partially negative (δ^-) poles on each bond.
- 5.1** $\text{C} - \text{O}$ and $\text{C} - \text{N}$ (3)
- 5.2** $\text{P} - \text{Br}$ and $\text{P} - \text{Cl}$ (3)
- 5.3** $\text{C} - \text{O}$ and $\text{C} - \text{S}$ (3)
- 6.** Ammonia (NH_3) is able to make a dative covalent bond with a hydrogen ion (H^+) forming the ammonium cation (NH_4^+)
- 6.1** Draw the Lewis diagram for a molecule of ammonia. (2)
- 6.2** Explain why the hydrogen ion can form a dative covalent bond with ammonia. (2)
- 6.3** Draw a Lewis diagram of the resulting ammonium cation. (3)
- 7. 7.1** Define the term 'bond energy'. (2)
- 7.2** What is the relationship between bond energy and bond length in a stable molecule? (2)
- 7.3** Consider the table of bond energies and bond lengths below. Use the table to answer the following questions
- a)** Cl_2 and I_2 are both examples of diatomic halogen molecules (Group 17). Explain why I_2 has a smaller bond energy than Cl_2 . (2)
- b)** Explain why I_2 has a larger bond length compared to Cl_2 . (2)
- c)** Explain why the bond energies of the different carbon to carbon bonds differ from single to triple bonds. (3)

BOND	BOND ENERGY ($\text{kJ}\cdot\text{mol}^{-1}$)	BOND LENGTH (pm)
$\text{Cl} - \text{Cl}$	242	199
$\text{C} - \text{C}$	347	540
$\text{C} = \text{C}$	614	340
$\text{C} \equiv \text{C}$	839	120
$\text{I} - \text{I}$	151	266

RESOURCE PACK

8. Consider the reaction between oxygen and hydrogen which is used to form water:



BOND	BOND ENERGY (kJ·mol ⁻¹)
O=O	498
O - H	463
H - H	436

Using the table of bond energies supplied, calculate the following quantities.

- 8.1 The amount of energy required to break the bonds in the reactants. (3)
- 8.2 The amount of energy released on formation of bonds of products. (2)
- 8.3 The total amount of energy released from the reaction in the formation of H₂O. (2)

MARKING GUIDELINES

MULTIPLE CHOICE

- 1.1 C ✓✓ All molecules are made from non-metal atoms only in their structure.
- 1.2 A ✓✓ The electronegativity of O is greater than that of H, thus the shared electron pair will be closer to the O atom and result in the bond becoming polar. Water is an asymmetric molecule and will thus will have a definite δ^+ and δ^- end, hence polar covalent.
- 1.3 D ✓✓ The central C atom is surrounded by 4 Cl atoms to form a tetrahedral shape. Tetrahedrons are symmetrical arrangements hence the molecule will be non-polar covalent due to symmetry.
- 1.4 C ✓✓ There are $4 \times \text{C} - \text{H}$ bonds in the molecule, thus it is simply $4 \times 416 \text{ kJ}\cdot\text{mol}^{-1}$ to give a total amount of energy to break the bonds as $1\,648 \text{ kJ}\cdot\text{mol}^{-1}$ (8)

LONG QUESTIONS

1. Complete the following table

MOLECULE	CO ₂	CH ₄	SF ₆	NH ₃
WHAT IS THE CENTRAL ATOM IN THE MOLECULE?	C ✓	C ✓	S ✓	N ✓
SHAPE OF MOLECULE	Linear ✓	Tetrahedral ✓	Octahedral ✓	Pyramidal ✓
ARE THE MOLECULAR SHAPES IDEAL OR NON-IDEAL?	IDEAL ✓	IDEAL ✓	IDEAL ✓	NON-IDEAL ✓
DOES THE MOLECULE CONTAIN POLAR BONDS?	Yes ✓	Yes ✓	Yes ✓	Yes ✓
IS THE MOLECULE POLAR OR NON-POLAR?	Non-polar ✓	Non-polar ✓	Non-polar ✓	Polar ✓

(20)

2. 2.1 A polar molecule – A molecule with a distinct region of charge ✓ at either end of the molecule due to one atom having a larger electronegativity than the other. ✓
 A non-polar molecule – There is no distinct region of charge due to a zero electronegativity difference between the atoms ✓ or the molecule is perfectly symmetrical. ✓ (4)

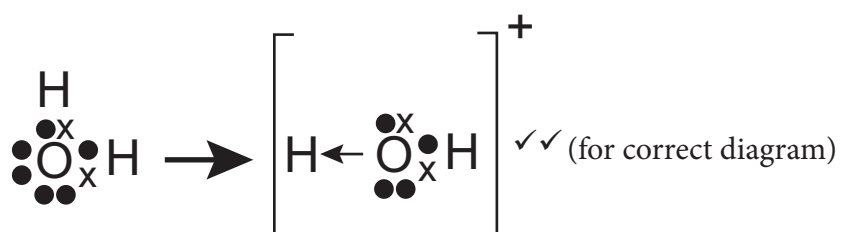
- 2.2 C : 2,5 } electronegativity difference is 1,0 ✓ thus bonds are
 O : 3,5 } polar covalent ✓ (2)

RESOURCE PACK

- 2.3 water – angular ✓
carbon dioxide – linear ✓ (2)

- 2.4 Carbon dioxide is non-polar due to the symmetry of the molecule. ✓ $\delta^- \quad \delta^+ \quad \delta^-$
There is no distinct region of δ^+ and δ^- , thus the molecule is considered non-polar
even though it has polar bonds. ✓ (2)

3. In the water molecule, H has a δ^+ charge and O has a δ^- charge. ✓ This is due to the
electronegativity of O being greater than the electronegativity of H. ✓ (2)
4. H_2O has two lone pairs of electrons. ✓ One of these lone pairs can be used to fill the
empty outermost energy level of the H^+ ion. ✓ This now forms a dative covalent bond
between H_2O and H^+ ✓ producing the H_3O^+ ion. ✓



(6)

5. 5.1 Polar covalent : H_2O ✓ HCl ✓ NH_3 ✓
Non-polar covalent : Cl_2 ✓ BF_3 ✓ CO_2 ✓ (6)

- 5.2 CO_2 is a symmetrical molecule ✓ meaning that it does not have two definite
regions of polarity ✓, that is, a δ^+ side and a δ^- side. Even though the bonds are
polar due to the electronegativity difference between C and O, there is no distinct
polarity within the molecule, hence it is non-polar. ✓



- 5.3 Cl_2 : linear ✓
 H_2O : angular ✓
 BF_3 : trigonal planar ✓
 HCl : linear ✓
 CO_2 : linear ✓
 NH_3 : trigonal pyramidal ✓ (6)

TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

6. 6.1



✓ correct atoms and shape

✓ correct electron arrangement

(2)

6.2 S ✓

(1)

6.3 two ✓

(1)

6.4 two ✓

(1)

6.5 angular ✓

(1)

6.6 NON-IDEAL ✓✓ - the central atom has lone pairs of electrons around it. ✓ (3)

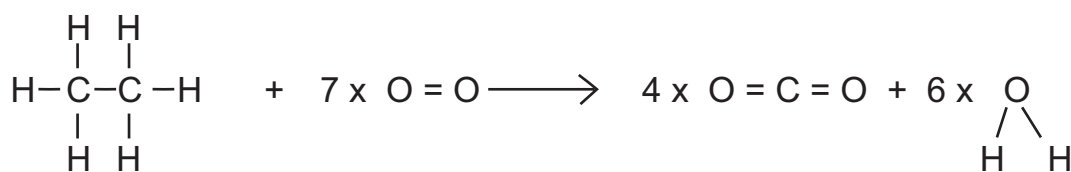
7. 7.1 The O = O bond requires more energy to break the bond compared to the O - O bond. ✓

The greater the amount of energy required, the shorter the bond length, ✓

thus the O = O bond will be shorter in length in comparison to the O - O bond. ✓

(3)

7.2



$$2 \times (\text{C}-\text{C}) = 692 \checkmark + 7 \times (\text{O}=\text{O}) = 3\,458 \checkmark$$

$$2 \times (6 \times \text{C}-\text{H}) = 4\,956 \checkmark$$

$$\text{Total energy to break bonds of reactants} = 5\,648 + 3\,458 \checkmark$$

$$\text{Total} = 9\,106 \text{ kJ}\cdot\text{mol}^{-1} \checkmark \quad (5)$$

$$7.3 \quad 4 \times (2 \times \text{C}=\text{O}) = 3\,992 \text{ kJ}\cdot\text{mol}^{-1} \checkmark + 6 \times (2 \times \text{O}-\text{H}) = 5\,508 \text{ kJ}\cdot\text{mol}^{-1} \checkmark$$

$$\text{Total energy released on bond formation of products} = 3\,992 + 5\,508 \checkmark$$

$$\text{Total} = 9\,500 \text{ kJ}\cdot\text{mol}^{-1} \checkmark \quad (4)$$

$$7.4 \quad \text{Excess energy} = (\text{energy to break bonds}) - (\text{energy released in new bond formation})$$

$$= 9\,106 - 9\,500 \checkmark$$

$$= -394 \text{ kJ}\cdot\text{mol}^{-1} \checkmark$$

$$\text{Thus, } 394 \text{ kJ}\cdot\text{mol}^{-1} \text{ of excess energy was released } \checkmark \quad (3)$$

RESOURCE PACK

MARKING GUIDELINES

CONSOLIDATION EXERCISE

TOTAL : 122 MARKS

MULTIPLE CHOICE

1. D ✓✓ When atoms approach each other, there are both attractive and repulsive forces between the atoms. Only when these forces are balanced will a covalent bond be formed and the potential energy of the system will be at its lowest. (2)
2. A ✓✓ Lone pair-lone pair electron repulsions are the largest of all the valence electron pair repulsions. (2)
3. B ✓✓ Br has an electronegativity of 2,8 and H has an electronegativity of 2,1. This gives an electronegativity difference of 0,7 which makes the bond polar. HBr is also a linear molecule which will form definite δ^+ and δ^- ends as a result of this difference, thus HBr will be polar. (2)
4. C ✓✓ The bond energy will be at the position where the potential energy in the system is the lowest and the bond is stable. This is represented by the lowest point of the graph and the value is read off the y-axis of the graph. (2)
5. C ✓✓ At the point where the potential energy in the system is the lowest is where the bond is most stable. This is where the atoms now establish a stable chemical bond and the bond length is fixed. This is now read off the x – axis of the graph. (2)
6. C ✓✓ This is the sum of the bond energies of four C – H bonds and two O = O bonds as there are 2 moles of O₂ in the reactants. This adds up to 2 646 kJ·mol⁻¹. (2)

TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

LONG QUESTIONS

1. Consider the table below. Redraw this table in your answer books and complete in the space provided.

CHEMICAL FORMULA OF COMPOUND	ELECTRO-NEGATIVITY DIFFERENCE BETWEEN ATOMS IN BOND	NON-POLAR COVALENT, POLAR COVALENT OR IONIC BOND	SHAPE OF MOLECULE	IDEAL OR NON IDEAL SHAPE
Cl_2	0 ✓	Non-polar covalent ✓	Linear ✓	IDEAL ✓
CF_4	1,5 ✓	Non-polar covalent ✓	Tetrahedral ✓	IDEAL ✓
CO_2	1,0 ✓	Non-polar covalent ✓	Linear ✓	IDEAL ✓
HBr	0,7 ✓	Polar covalent ✓	Linear ✓	IDEAL ✓
NH_3	0,9 ✓	Polar covalent ✓	Pyramidal ✓	NON-IDEAL ✓
H_2O	1,4 ✓	Polar covalent ✓	Angular ✓	NON-IDEAL ✓
BF_3	2,0 ✓	Non-polar covalent ✓	Trigonal planar ✓	IDEAL ✓

(28)

2. 2.1



- ✓ correct shape
- ✓ electrons arranged correctly

(2)

2.2



- ✓ correct shape
- ✓ electrons arranged correctly

(2)

3. 3.1 This is the measure of the amount of attraction ✓ that an atom exerts on a shared electron pair ✓ within a chemical bond. (2)

3.2 a) F_2 : F = 4,0 ✓ HCl : H = 2,1 ✓ CH_4 : C = 2,5 ✓

Cl = 3,0

H = 2,1

 H_2O : H = 2,1 ✓ NH_3 : N = 3,0 ✓ SF_6 : S = 2,5 ✓

(6)

O = 3,5

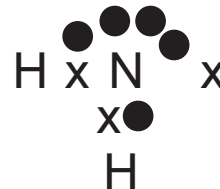
H = 2,1

F = 4,0

RESOURCE PACK

- b) F_2 : EN difference $4,0 - 4,0 = 0$ ✓
 HCl: EN difference $3,0 - 2,1 = 0,9$ ✓
 CH_4 : EN difference $2,5 - 2,1 = 0,4$ ✓
 H_2O : EN difference $3,5 - 2,1 = 1,4$ ✓
 NH_3 : EN difference $3,0 - 2,1 = 0,9$ ✓
 SF_6 : EN difference $4,0 - 2,5 = 1,5$ ✓ (6)
- c) F_2 : non-polar covalent ✓
 HCl: polar covalent ✓
 CH_4 : polar covalent ✓
 H_2O : polar covalent ✓
 NH_3 : polar covalent ✓
 SF_6 : polar covalent ✓ (6)
- d) F_2 = linear ✓
 HCl = linear ✓
 CH_4 = tetrahedral ✓
 H_2O = angular ✓
 NH_3 = trigonal pyramidal ✓
 SF_6 = octahedral ✓ (6)

4. 4.1



Two marks for each correct structure.

- 4.2 BCl_3 ✓ CF_4 ✓ CO_2 ✓ SCl_6 ✓
 symmetrical molecules hence non-polar ✓✓ (6)

- 4.3 BCl_3 : trigonal planar ✓
 CF_4 : tetrahedral ✓
 NH_3 : trigonal pyramidal ✓
 SCl_6 : octahedral ✓ (4)

- 4.4 NH_3 - has lone pair of electrons around central atom hence NON-IDEAL ✓
 SCl_6 - has no lone pairs of electrons around central atom, hence IDEAL ✓ (2)

5. 5.1 δ^+ δ^- ✓

- C - O EN difference (C - O) = $3,5 - 2,5 = 1,0$ ✓
 C - N EN difference (C - N) = $3,0 - 2,5 = 0,5$ ✓ (3)

5.2 δ^+ δ^- ✓

- P - Cl EN difference (P - Cl) = $3,0 - 2,1 = 0,9$ ✓
 P - Br EN difference (P - Br) = $2,8 - 2,1 = 0,7$ ✓ (3)

TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

5.3 $\delta^+ \delta^-$ ✓

C – O EN difference (C – O) = 3,5 – 2,5 = 1,0 ✓

C – S EN difference (C – S) = 2,5 – 2,5 = 0 ✓ (3)

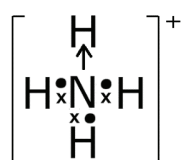
6.1



(2)

6.2 N has a lone pair of electrons in the ammonia molecule which it can share with the H^+ ion which has an empty valence energy level. ✓✓ (2)

6.3



✓✓ for the correct Lewis diagram
 ✓ There is now an arrow replacing the lone pair of electrons to indicate dative covalent bond (3)

7.1 The bond energy of a compound is the energy needed to break one mole of its molecules into separate atoms. ✓✓ (2)

7.2 As the bond length gets larger, so the amount of energy required to break the bond gets smaller. ✓✓ (2)

7.3 a) I_2 is made up of much larger atoms than in Cl_2 and thus has a much greater bond length. ✓ The larger the bond length, the less the bond energy. ✓ (2)

b) I_2 is made up of much larger atoms compared to Cl_2 . ✓ The larger the atoms, the larger the bond length between the atoms. ✓ (2)

c) The more bonds there are between two atoms, the shorter the bond length. ✓ This means that the amount of energy required to break the bonds will be more, thus bond energy will increase. ✓ The carbon triple bond required $839 \text{ kJ}\cdot\text{mol}^{-1}$ of energy to break the bonds while the carbon single bond only requires $347 \text{ kJ}\cdot\text{mol}^{-1}$ to break the bond. ✓ (3)

8.1 $O = O$ $498 \text{ kJ}\cdot\text{mol}^{-1}$ ✓ $2 \times H - H$ $872 \text{ kJ}\cdot\text{mol}^{-1}$ ✓Total $1\,370 \text{ kJ}\cdot\text{mol}^{-1}$ ✓ (3)8.2 $4 \times O - H$: $1\,852 \text{ kJ}\cdot\text{mol}^{-1}$ ✓✓ (2)

8.3 Energy released = $1\,370 - 1\,852$ ✓
 = $-482 \text{ kJ}\cdot\text{mol}^{-1}$ ✓ (2)

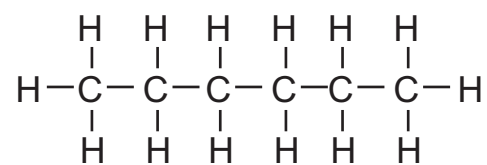
RESOURCE PACK

TOPIC 4: Intermolecular Forces and Chemistry of Water

WORKSHEET

MULTIPLE CHOICE

- Which one of the following statements represents the best explanation for the term 'electronegativity'?
A A measure of the tendency of an atom to attract a bonding pair of electrons.
B A measure of the tendency of an atom to attract an electron.
C A measure of the strength of a covalent bond.
D A measure of the strength of an ionic bond. (2)
- Sodium chloride (NaCl) is a solid which is soluble in water. Which one of the following describes the intermolecular forces that exist between sodium chloride and water in solution?
A Ion-dipole
B Dipole-dipole
C Ion-induced dipole
D Induced dipole-dipole (2)
- Hydrogen bonding is a type of intermolecular force that can exist between the molecules of certain compounds. Which one of the statements below best describes the conditions under which hydrogen bonding is most likely to occur? It occurs between:
A small molecules which contain hydrogen atoms.
B molecules in which hydrogen is bonded to small atoms with high electronegativity.
C large molecules which contain both hydrogen and oxygen atoms.
D molecules in which hydrogen is bonded to small atoms with low electronegativity. (2)
- Consider the structure of hexane.

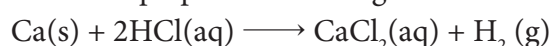


A molecule of hexane is considered to be non-polar. Which one of the following statements best describes the reason why hexane is non-polar?

- A Hexane contains only single bonds between atoms.
 - B The electronegativity difference between C and H atoms is so small as to be considered non-polar.
 - C Hexane is a linear molecule hence is symmetrical.
 - D The charge distribution of electrons within the hexane molecule is symmetrical.
- (2)

LONG QUESTIONS

1. Calcium chloride is prepared according to the following balanced chemical equation:



Hydrochloric acid is produced for this preparation by the ionisation of hydrogen chloride gas in water.

- 1.1 Name the type of bonding that is present in a molecule of hydrogen chloride. (1)
- 1.2 Define the term 'electronegativity'. (2)
- 1.3 Make use of the Pauling scale of electronegativities, as provided in the Periodic Table, to explain the type of bonding found in hydrogen chloride and crystalline calcium chloride. (4)
- 1.4 HCl molecules are described as 'dipoles'. Explain what is meant by this term. (2)
- 1.5 Calcium chloride is soluble in water. The structure of its crystal lattice is broken down by the water molecules to form aqueous ions in solution.
 - 1.5.1 Name the type of crystal lattice of which calcium chloride is an example. (1)
 - 1.5.2 Name the types of intermolecular forces present between:
 - a) calcium chloride ions in the crystal lattice. (1)
 - b) the water molecules of the solvent. (1)
 - c) the water and the crystal lattice. (1)
 - 1.5.3 Using diagrams to illustrate your answer, explain how the crystal lattice of calcium chloride is broken down during the dissolving process. (5)
- 2. Use only substances from the list below when answering Question 2.1 to 2.6 (Only write down the question number and the formula of the substance next to it.) The state symbols (phase indicators) represent the physical state of each of the substances at room temperature.

$\text{SiO}_2\text{(s)}$ HCl(g) $\text{H}_2\text{O(l)}$ $\text{PH}_3\text{(g)}$ Mg(s) $\text{Br}_2\text{(g)}$ KF(s)

Select one substance from the list that has:

 - 2.1. pure covalent intramolecular bonds. (1)
 - 2.2. a high melting point due to the strong electrostatic attraction between the cations and anions in the crystal lattice. (1)
 - 2.3. hydrogen bonding intermolecular forces. (1)

RESOURCE PACK

2.4 a very high melting point due to its giant covalent network structure. (1)

2.5 delocalised valence electrons. (1)

2.6 dipole – dipole intermolecular forces. (1)

3. Consider the Table below which shows the boiling points of the halogens.

Halogens	F ₂	Cl ₂	Br ₂	I ₂
Boiling point (°C)	-188	-35	59	184

3.1 Name the specific type of interatomic bond that occurs between atoms in the F₂ molecule. (2)

3.2 The table show an increase in the boiling points of the halogens from F₂ to I₂. Explain this trend by making reference to the relevant intermolecular force between the halogen molecules and the factor influencing its strength. (4)

4. Consider the following pairs of substances.

HCl and CO NaCl and CCl₄ KBr and H₂S CCl₄ and Br₂

Which of the pairs will have:

4.1 ion-dipole intermolecular forces? (2)

4.2 dipole – dipole intermolecular forces? (2)

4.3 ion – induced dipole intermolecular forces? (2)

4.4 induced dipole – induced dipole intermolecular forces? (2)

5. Water is considered to be a highly remarkable substance that has many properties that are beneficial to life forms on earth. One of the interesting factors with regards to this is the fact that ice is able to float on water.

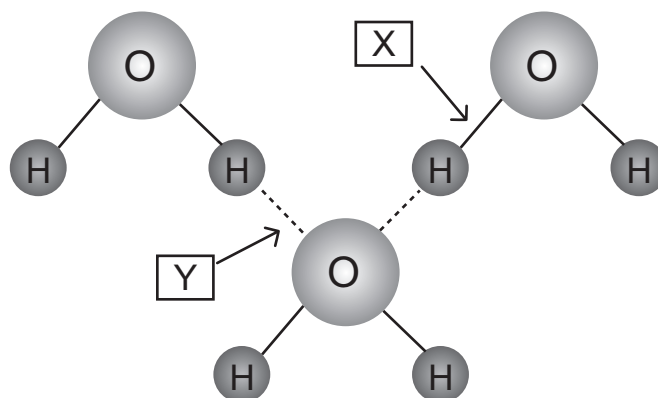
5.1 Explain why ice is able to float on water. (2)

5.2 Why is this beneficial to aquatic life? (1)

CONSOLIDATION EXERCISE

TOTAL: 47 MARKS

1. Hydrogen fluoride (HF) can be prepared by treating calcium fluoride (CaF) with an acid.
 - 1.1 Define the term 'electronegativity'. (2)
 - 1.2 Use electronegativity differences to explain the difference in the type of bonding found in HF and CaF₂. (4)
 - 1.3 Using a Lewis diagram, illustrate fully how bonding occurs in CaF₂. (4)
 - 1.4 HF is described as a 'polar molecule'. Explain what is meant by this term. (2)
 - 1.5 In fact, HF is an example of a group of three molecules which exhibit what is known as hydrogen bonding. What TWO identifying characteristics are crucial in a molecule being able to exhibit hydrogen bonding? (2)
2. Two crystalline solids, P and Q have melting points of 710°C and 723°C respectively. At 700°C, substance P conducts electricity but Q does not. At 750°C, both P and Q conduct electricity.
 - 2.1 What can be deduced from the melting points about the relative magnitude of the forces between the particles in P and Q? (2)
 - 2.2 Explain the difference observed in electrical conductivity between P and Q. (4)
 - 2.3 You are told that one of the solids is barium and the other is potassium iodide. Which is P and which is Q? (2)
3. Consider the diagram below, showing an arrangement of water molecules in the liquid phase.



- 3.1 Explain the term 'interatomic bond'. (1)
- 3.2 Name the specific type of interatomic bond represented by the letter X in the diagram. (2)
- 3.3 Define the term 'intermolecular force'. (2)

RESOURCE PACK

- 3.4 Name the specific type of intermolecular force represented by the letter Y in the diagram. (1)
- 3.5 State two properties of the oxygen atom that make this type of intermolecular force (Y) possible. (2)
- 3.6 What is the partial charge (δ^+ or δ^-) on the hydrogen atom in a water molecule? (1)
4. Diagram 1 below shows the crystal lattice structure of the giant ionic solid sodium chloride.

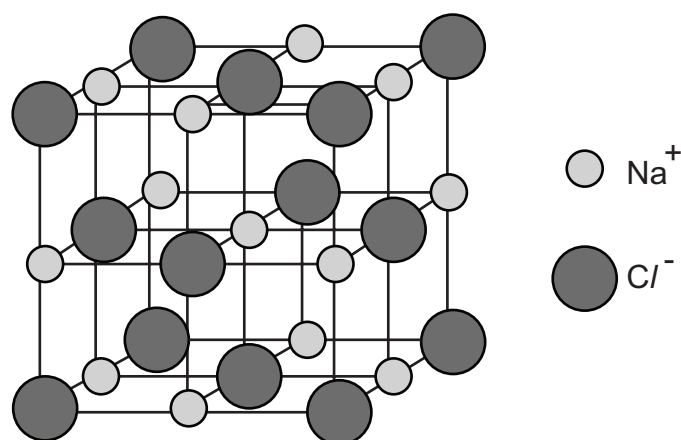


Diagram 1

- 4.1 Define the term 'ionic bond'. (2)
- 4.2 With reference to the crystal lattice structure, explain why sodium chloride has a very high melting point (801°C). (3)
- 4.3 Sodium chloride dissolves in water. The ions are surrounded by water molecules as shown in Diagram 2 below.

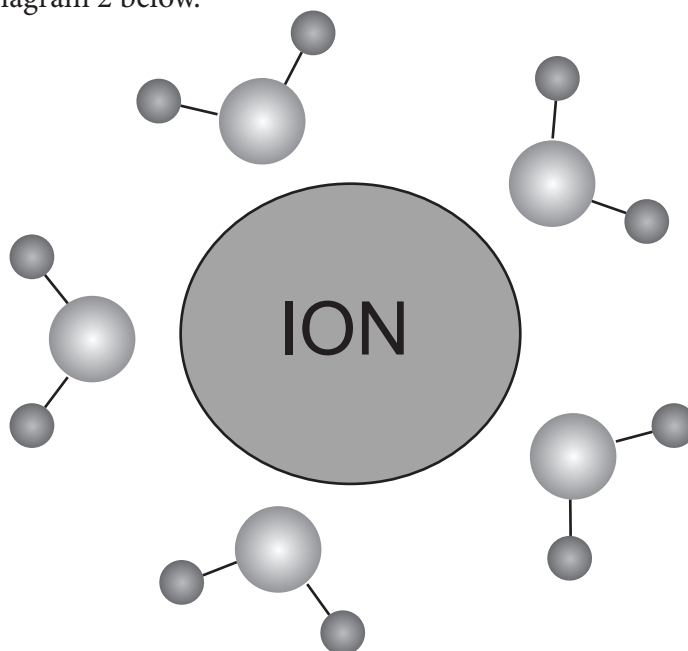


Diagram 2

TOPIC 4: INTERMOLECULAR FORCES AND CHEMISTRY OF WATER

WORKSHEETS

- a) Name the specific type of force between the ion and the water molecules. (1)
- b) Is the ion shown in the above diagram a sodium ion (Na^+) or a chloride ion (Cl^-)? Explain. (2)
- c) A relatively large amount of water is needed to dissolve a small amount of sodium chloride. Explain why this is so by referring to the structure of sodium chloride and the strengths of the forces involved. (4)
5. 5.1 Explain what is meant by the term 'Heat of vaporisation'. (2)
- 5.2 Explain why in the early morning just before dawn breaks, the temperature of the air suddenly drops. (2)
- (47)

RESOURCE PACK

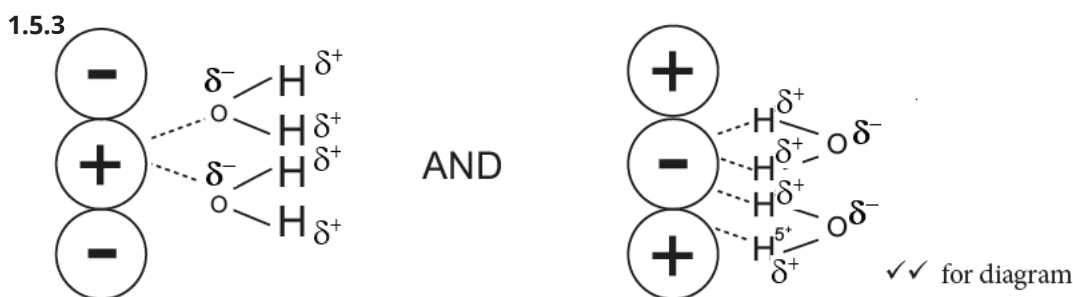
MARKING GUIDELINES

MULTIPLE CHOICE

1. A ✓✓ This question is a direct understanding of the definition so learners must learn this. (2)
2. A ✓✓ Sodium chloride is an ionic compound; water is a dipole. (2)
3. B ✓✓ The criteria for a molecule to be able to exert hydrogen bonding intermolecular forces is when a H atom is bonded to a small atom of high electronegativity such as N, O or F. (2)
4. D ✓✓ Hexane is a non-polar molecule as it is perfectly symmetrical and thus the distribution of electrons in the electron cloud is symmetrical hence no region of δ^+ or δ^- . (2)

LONG QUESTIONS

- 1.1 Polar covalent bonding ✓ (1)
- 1.2 The measure of the tendency of an atom to attract ✓ a shared pair of electrons ✓ within the chemical bond. (2)
- 1.3 HCl Electronegativity difference = $3,5 - 2,1 = 1,4$ ✓
 CaCl_2 Electronegativity difference = $3,5 - 1,0 = 2,5$ ✓
 Pauling scale states that electronegativity difference greater than 2,0 will make the substance ionic in character. CaCl_2 is greater than 2,0 hence will be ionic ✓ whereas HCl at 1,4 will be polar covalent in character as Pauling scale states that from 0 to 2,0, substance will be covalent. ✓ (4)
- 1.4 HCl molecules have two distinct regions of charge ✓ δ^+ and δ^- at either end of the molecule ✓ (2)
- 1.5 1.5.1 Ionic crystal lattice ✓ (1)
- 1.5.2 a) ion – ion (ionic) ✓
 b) hydrogen bonding (dipole) ✓
 c) ion – dipole ✓ (3)



Polar water molecules surround the ions in the crystal lattice creating an ion-dipole intermolecular force. ✓ A single ion-dipole intermolecular force is not

TOPIC 4: INTERMOLECULAR FORCES AND CHEMISTRY OF WATER

strong enough to overcome the very strong ion-ion forces within the lattice. ✓
More water molecules thus move to surround the ions to create a combined ion-dipole force that will overcome the forces within the crystal lattice. ✓ The ions are then removed from the lattice and go into solution. (5)

- 2.1** Br₂ or PH₃ Atoms in the molecule with equal electronegativities, hence electronegativity difference of 0, hence a pure covalent bond (non-polar) ✓ (1)
- 2.2** KF This is the only ionic substance in the list and contains cations and anions. These are very strong forces of attraction, thus a very high boiling point ✓ (1)
- 2.3** H₂O H is bonded to the small O atom which has a high electronegativity. ✓ (1)
- 2.4** SiO₂ This is solid made of covalent bonds only hence called a covalent network solid. These covalent bonds are very strong hence a very high melting point. ✓ (1)
- 2.5** Mg A pure metal with metallic bonding between atoms, Contains delocalised valence electrons. ✓ (1)
- 2.6** HCl There is an electronegativity difference of 1,4 between atoms, hence it is a dipole. ✓ (1)
- 3.1** Non-polar ✓ covalent ✓ (The electronegativity difference between two F atoms is 0, hence will be non-polar). (2)
- 3.2** All halogen molecules will have London/dispersion intermolecular forces between them ✓. The size of the halogen molecules increases as one moves down the Table. This means that the electron cloud density increases allowing for a greater surface area ✓ meaning that there will be more points of contact for intermolecular forces to occur. There will be stronger intermolecular forces ✓ between the molecules and more energy needed ✓ to overcome these forces. (4)
- 4.1** KBr and H₂S ✓✓ (2)
- 4.2** HCl and CO ✓✓ (2)
- 4.3** NaCl and CCl₄ ✓✓ (2)
- 4.4** CCl₄ and Br₂ ✓✓ (2)
- 5.1** Ice has a lower density compared to liquid water at around temperatures of 4 °C ✓. This is due to the fact that the hydrogen bonding intermolecular forces are further apart in the solid phase compared to water in the liquid phase at these very cold temperatures. ✓ (Thus due to the increase in the distance when the intermolecular forces act in the solid phase the density of ice decreases and it floats. In the liquid phase at very cold temperatures, the water molecules can approach each other more closely, hence the density of water increases at these temperatures). (2)
- 5.2** Ice will float, cold water will sink ✓ – thus aquatic life can survive at the bottom of the ocean/river. (1)

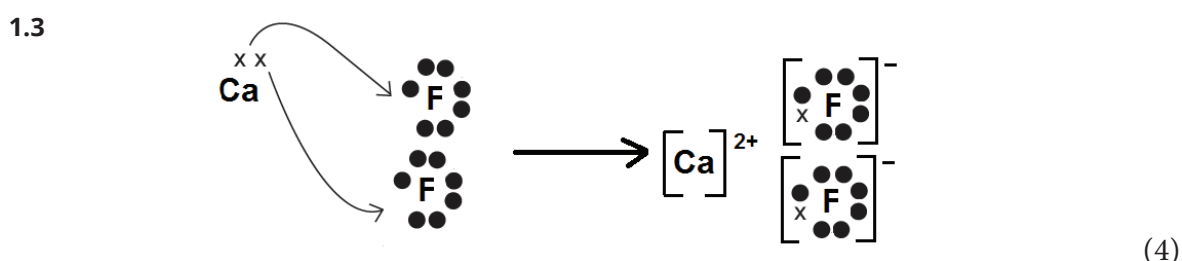
RESOURCE PACK

CONSOLIDATION EXERCISE

TOTAL: 47 MARKS

1.1 The measure of the amount of attraction ✓ an atom has on a shared pair of electrons ✓
(2)

1.2 HF : electronegativity difference = 1,9 ✓
CaF : electronegativity difference = 3,0 ✓
HF will be expected to be covalent ✓ although it does have some ionic character due to the relatively large electronegativity difference. CaF₂ will be strongly ionic ✓ due to a very high electronegativity difference. (4)



1.4 HF has two distinct regions of charge on the molecule ✓✓ $\delta^+ \quad \delta^-$
H – F (2)

1.5 The H atom in the molecule must be bonded to a small atom ✓ of very high electronegativity. ✓ (2)

2.1 Very high melting points indicate that there are very strong intermolecular forces between the particles in each solid. ✓✓ (2)

2.2 At 700°C, both P and Q are still in the solid phase. The fact that P is able to conduct electricity tells us it must have charges that are free to move in the solid phase ✓. This means that P must be a metal with delocalised electrons hence the conductivity ✓. Q has no free charges in the solid phase ✓ thus will not be able to conduct electricity, thus Q cannot be a metallic substance. Q conducts when molten therefore it is an ionic substance. ✓ (4)

2.3 P = Barium ✓ (metallic substance thus has delocalised electrons hence conductivity)
Q = Potassium iodide ✓ (ionic substance where the ions are held in the crystal lattice and not free to move) (2)

3.1 A bond that exists within a molecule between the atoms in that molecule. ✓ (1)

3.2 Polar ✓ covalent bond ✓ (It will be polar due to the electronegativity difference between H and O) (2)

3.3 This is the force of attraction between molecules in a system. ✓✓ (2)

3.4 Hydrogen bonding intermolecular force. ✓ (1)

3.5 O is a very small atom in terms of its atomic size. ✓
O has a very high electronegativity. ✓ (2)

3.6 δ^+ ✓

(H will not have any electrons near its nucleus as the high electronegativity of O will pull the shared electron pair away from the H atom and towards itself, leaving the H atom δ^+) (1)

**4.1 Force of attraction that occurs between cations and anions in a crystal lattice ✓✓
(Bond formed by transfer of electrons) ✓ (2)**

4.2 The Na^+ cations are surrounded by Cl^- anions and vice versa. This means that there are many strong electrostatic forces of attraction ✓ between the cations and anions leading to very strong inter-particle (inter-ionic) forces. ✓ Large amounts of energy will be needed to overcome these forces, ✓ hence a very high melting point. (3)

4.3 a) Ion-dipole intermolecular force ✓ (1)

4.3 b) Na^+ ✓ – the O end of the water molecule is attracted to the ion. The O end is δ^- thus it is attracted to a positive ion, hence the sodium cation. ✓ (2)

4.3 c) The ion-dipole force between the water molecule and the ion is a strong intermolecular force, yet is not stronger than the ion-ion electrostatic forces within the crystal lattice. ✓✓ Water thus needs to accumulate around an ion so that the combined ion-dipole intermolecular forces ✓ eventually will overcome the ion-ion forces ✓ and the ion is only then removed from the lattice and goes into solution. (4)

5.1 This is the amount of energy one mole of water will absorb before it is able to escape to the gaseous phase ✓✓ (2)

5.2 Energy is absorbed by the water molecules that sit on grass and plants (dew) just before the sun rises. ✓ This allows the water molecule to escape to the gas phase with a certain kinetic energy. This means that energy is taken away from the system hence it begins to feel much colder as there is less energy in the system's surroundings. ✓ (2)

Topic 3: Atomic Combinations (Molecular Structure)

QUESTIONS

MULTIPLE CHOICE

1. Which chlorine compound has bonding that can be described as ionic with some covalent character?

A NaCl

B MgCl₂

C AlCl₃

D SiCl₄

(2)
2. Which molecule has only six bonding electrons?

A C₂H₄

B C₂F₆

C H₂O

D NF₃

(2)
3. What set of data would hydrogen sulfide, H₂S, be expected to have?

	Number of bonding pairs	Shape
A	1	Trigonal planar
B	2	Pyramidal
C	2	Angular
D	3	Pyramidal

(2)

4. Which ONE of the following best describes the bond formed between an H⁺ ion and the NH₃ molecule to form an NH₄⁺ ion?

A covalent bond

B dative covalent bond

C ionic bond

D hydrogen bond

(2)

RESOURCE PACK

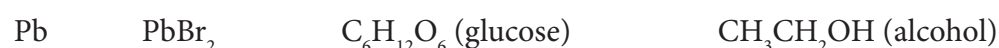
5. Refer to the following list of substances and the electronegativities provided on the Periodic Table.



Which of the following lists all the non-polar molecules from the above list?

- A N_2
 B N_2 and NH_3
 C N_2 and CO_2
 D N_2 , CO_2 and NH_3 (2)

6. Which of the following pure substances will conduct electric current when molten (in the liquid phase)?

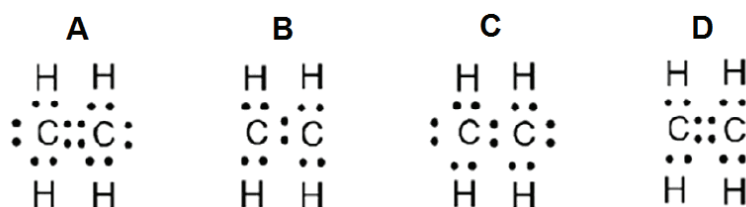


- A Pb and PbBr_2
 B Pb and $\text{CH}_3\text{CH}_2\text{OH}$ (alcohol)
 C Pb, PbBr_2 and $\text{CH}_3\text{CH}_2\text{OH}$ (alcohol)
 D All of these substances (2)

7. A dative covalent bond forms when...

- A water decomposes.
 B ionic compounds dissociate in water.
 C hydronium ions are produced.
 D hydroxide ions are produced. (2)

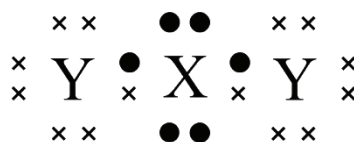
8. Which of one of the following Lewis structures correctly represents a molecule of C_2H_4 ?



(2)

TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

9. Consider the Lewis structure of a compound shown below:



Which ONE of the following is correct?

	Name of element X	Name of element Y	Molecular shape of compound
A	chlorine	sulfur	angular
B	oxygen	chlorine	angular
C	chlorine	oxygen	linear
D	sulfur	chlorine	linear

(2)

10. Which of the following substances has polar covalent bonds between its atoms, but its molecule is non-polar?

- A H_2O
 B H
 C HCl
 D CCl_4

(2)

LONG QUESTIONS

- 1.1 Study the following list of substances:

Hg I_2 SF_6 KI PCl_3 Ne C H_2O

For each statement below choose a substance that displays those properties. You can choose a substance once, more than once or not at all.

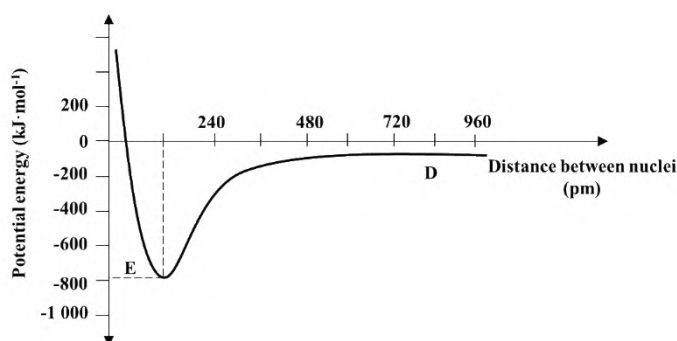
- 1.1.1 A substance that conducts electricity when molten but not when solid. (1)
 1.1.2 A substance that forms a covalent crystal lattice. (1)
 1.1.3 A substance that consists of non-polar molecules containing polar bonds. (1)
 1.1.4 Its molecules have two lone pairs on the central atom. (1)
 1.1.5 A substance which has delocalised electrons. (1)
 1.1.6 Its molecules have a non-ideal shape. (1)
 1.1.7 A substance which consists of monatomic molecules. (1)

- 1.2 Refer to the bond energies supplied below:

A	B
C–C	C=C
346 $\text{kJ}\cdot\text{mol}^{-1}$	602 $\text{kJ}\cdot\text{mol}^{-1}$

RESOURCE PACK

- 1.2.1 What information does the bond energy provide? (1)
- 1.2.2 Which of A or B will have the shorter bond length? (1)
- 1.2.3 Explain why the bond energy in B is not double the bond energy in A. (2)
- [11]
- 2.1 South Africa has comparable greenhouse gas emissions to many industrialised countries. The main air pollutants generated by South African industries are carbon dioxide (CO_2), methane (CH_4), carbon monoxide (CO), sulfur dioxide (SO_2), the oxides of nitrogen (NO_2 , NO and N_2O) and ammonia (NH_3).
- 2.1.1 Draw Lewis structures to represent:
- a) carbon dioxide. (2)
- b) methane. (2)
- 2.1.2 Define the term “electronegativity”. (2)
- 2.1.3 By means of a suitable calculation describe the type of bonding that occurs in a molecule of sulfur dioxide. (2)
- 2.1.4 The chemical bonds within the methane molecule are polar covalent and yet methane is known to be a non-polar molecule. Explain how this phenomenon comes about. Show the polarity of the molecule with the use of a Lewis diagram. (4)
- 2.1.5 Of the pollutants listed above, choose
- a) one trigonal pyramidal
- b) one tetrahedral and
- c) one linear molecule. (3)
- 2.2 The graph below shows the change in energy that takes place when an oxygen atom approaches a carbon atom during the formation of a $\text{C}=\text{O}$ bond.



- 2.2.1 Define the term ‘bond length’. (2)
- 2.2.2 From the graph, write down the:
- (a) bond length, in pm, of the $\text{C}=\text{O}$ bond. (1)
- (b) bond energy, in $\text{kJ}\cdot\text{mol}^{-1}$ needed to break the $\text{C}=\text{O}$ bond. (2)
- (c) Name the potential energy represented by E. (1)
- 2.2.3 How will the bond length of an $\text{S}=\text{O}$ bond compare to that of the $\text{C}=\text{O}$ bond? Write down EQUAL TO, SHORTER THAN or LONGER THAN. Give a reason for your answer. (2)
- 2.2.4 What can be said about the forces between the two atoms at point D? (2)
- [25]

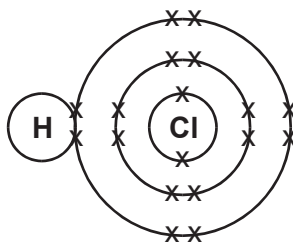
TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

3. 3.1 $\text{CO}_2(\text{g})$, SO_2 and $\text{CO}(\text{g})$ are some of the air pollutants in the South African atmosphere.
- 3.1.1 CO has a dative covalent bond. Draw the Lewis structure for a carbon monoxide molecule and use it to help explain the idea of a dative covalent bond. (3)
- 3.1.2 Explain why you would expect the CO molecule to be polar. (2)
- 3.1.3 What is the nature of the partial charges on (i.e. the polarity of) the C and O atoms in the CO molecule? Explain your answer. (2)
- 3.2 Consider the gases CO_2 and SO_2 .
- 3.2.1 Refer to Lewis Diagrams and the VSEPR theory. Compare the shapes of a CO_2 molecule and a SO_2 molecule, both of which contain two oxygen atoms bonded to one other atom. (4)
- 3.2.2 Comment on the polarity of both the CO_2 and SO_2 molecules. In each instance, explain your answer with reference to the molecular shape. (4)
- [15]
4. Sulfur forms many compounds by reacting with other elements.
- 4.1 Three of these compounds are sodium sulfide, SOCl_2 , SF_2 and two other fluorides.
- 4.1.1 Draw Lewis diagrams to show the bonding in:
- a) sodium sulfide (3)
- b) SF_2 (2)
- 4.1.2 There are two possible ways of drawing the Lewis diagram of SOCl_2 . Draw the two possible Lewis structures. Explain how you arrived at each of the structures. (5)
- 4.1.3 Use the valence shell electron pair repulsion (VSEPR) theory to deduce the shapes of the following molecules:
- a) SF_2 (2)
- b) BeCl_2 (2)
- 4.2 Sulfur is in group 16 and period 3 of the Periodic Table.
- 4.2.1 Define the terms:
- a) Electronegativity (2)
- b) Bond polarity (2)
- 4.2.2 Give reasons for your answers to the next two questions. Identify the type of bonds that form in
- a) SF_2 (3)
- b) H_2S (3)
- [24]

RESOURCE PACK

5. 5.1 The hydrogen halides (hydrogen fluoride, hydrogen chloride, hydrogen bromide and hydrogen iodide) are important chemicals.

The diagram below represents a molecule of hydrogen chloride.



- 5.1.1 What type of particles are represented by the crosses (X)? (1)
- 5.1.2 What type of chemical bond holds the atoms in this molecule together? (1)
- 5.2 The relative amount of energy (in $\text{kJ}\cdot\text{mol}^{-1}$) required to break the bond in each of the hydrogen halide molecules is shown below.
- | | |
|--------|-----|
| H – F | 569 |
| H – Cl | 432 |
| H – Br | 366 |
| H – I | 298 |
- One of the important properties of the hydrogen halides is that they dissolve in water to form acids. For example, hydrogen chloride reacts with water to form hydrochloric acid.
- To form an acid the bond between the hydrogen and the halogen atoms must be broken and ions are formed. The stronger the acid the more molecules that split up to form ions.
- 5.2.1 Which ion must be formed to make a solution acidic? (1)
- 5.2.2 Which of the hydrogen halides would you expect to react with water to form the strongest acid? Explain your answer. (3)
- [6]
6. A learner studies fluorine and carbon and the bonds that atoms of each can form with other atoms. Fluorine is a reactive element that can react with water and with many metals. Carbon is a very special element because it is one of the elements in the food we eat, the clothes we wear and the fuel in our cars.
- 6.1 Write down the number of electrons in the valence energy level of an atom of:
- 6.1.1 fluorine. (1)
- 6.1.2 carbon. (1)
- 6.2 Describe what is meant by the term 'covalent bond'. (2)
- 6.3 Write down the Lewis structure for:
- 6.3.1 F_2 (2)
- 6.3.2 HCN (2)
- 6.4 Write down the name of the type of bond between particles in molecules of:
- 6.4.1 F_2 (1)
- 6.4.2 HF (1)

TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

6.5 Write down the molecular shape as predicted by the VSEPR model of:

6.5.1 BF_3 (2)

6.5.2 CH_4 (2)

6.5.3 HF (2)

6.6 The table below shows bond lengths and bond energies for different bonds between two carbon atoms.

	Bond	Length (pm)	Energy ($\text{kJ}\cdot\text{mol}^{-1}$)
A	$\text{C} - \text{C}$	154	348
B	$\text{C} = \text{C}$	134	614
C	$\text{C} \equiv \text{C}$	120	839

6.6.1 Describe the relationship between bond length and bond energy as shown in the above table. (2)

6.6.2 Which one of the bonds (A, B or C) will be the weakest? Explain. (3)

[20]

RESOURCE PACK

MARKING GUIDELINES

MULTIPLE CHOICE

1. B ✓✓ Δ electronegativity = $3,0 - 1,2 = 1,8$ which implies that the bond is polar covalent. However, Mg is a metal bonding with the non-metal Cl, therefore the bond is ionic. This is the compound which is ionic with some covalent character. [CL3] (2)
2. D ✓✓ Nitrogen has three unpaired electrons which it is able to share with the three fluorine atoms. There are three bonds around each nitrogen atoms. Each bond consists of a shared pair of electrons. Therefore, this compound has 6 bonding electrons (six electrons involved in chemical bonds). [CL2] (2)
3. C ✓✓ There are two bonding pairs because two hydrogen atoms are attached to the sulfur atoms. The two lone pairs of sulfur repel each other strongly so the shape of H_2S is similar to that of water. It is angular (or bent). [CL2] (2)
4. B ✓✓ The hydrogen ion bonds with the ammonia molecule by attaching to the lone pair of electrons. This type of covalent bond where one atom donates both electrons to form the bond is called a dative covalent bond. [CL3] (2)
5. C ✓✓ Nitrogen is non-polar because it consists of three non-polar covalent bonds. CO_2 is non-polar because its molecule is linear, with the effect of the two polar covalent bonds directed in opposite directions, resulting in an overall non-polar molecule. [CL2] (2)
6. A ✓✓ Pb is a metal with delocalised electrons to carry charge through the substance. $PbBr_2$ in molten form dissociates into its ions (cations and anions). These ions are free to carry charge throughout the molten liquid. [CL2] (2)
7. C ✓✓ Hydronium ions are formed when a hydrogen ion (H^+ ion) attaches itself to one of the lone pairs of electrons of a water molecule. This kind of bond is called a dative covalent bond. [CL3] (2)
8. D ✓✓ Carbon has four electrons in the outermost energy level of its atoms. This Lewis structure is the only one which correctly shows carbon atoms with four valence electrons. [CL4] (2)
9. B ✓✓ The element X has six valence electrons therefore it is either a sulfur or an oxygen atom. The element Y must belong to the halogens because it has seven valence shell electrons. The only possible choices are narrowed to choices B and D. However, there are two lone pairs on atom X, so the shape of the molecule will be angular. [CL4] (2)

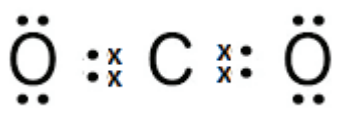
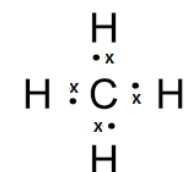
TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

10. D ✓✓ This molecule has a symmetrical shape (tetrahedral). Even though it has polar covalent bonds between the atoms, the overall electron distribution is symmetrical, therefore its molecule is non-polar. [CL3] (2)

LONG QUESTIONS

- 1.1.1 KI ✓ An ionic solid conducts electric current ONLY when molten or dissolve in water. (1)
- 1.1.2 C ✓ Carbon as diamond is a covalent crystal lattice. (1)
- 1.1.3 SF₆ ✓ The six fluorine atoms are arranged symmetrically around the sulfur atom therefore the overall charge distribution in the molecule is symmetrical – a non-polar molecule. (1)
- 1.1.4 H₂O ✓ The oxygen atom has two lone pairs on it in a water molecule. (1)
- 1.1.5 Hg ✓ Mercury is a metal therefore it has delocalised electrons. (1)
- 1.1.6 H₂O ✓ or PCl₃ ✓ Shapes are angular (bent) and trigonal pyramidal (each of these shapes has lone pairs around the central atom, therefore it has a non-ideal shape). (1)
- 1.1.7 Ne ✓ Neon is a noble (inert) gas. Its molecules consist of one atom only (therefore monatomic). (1)
- Because of the choices that learners must make from the list of substances, the overall cognitive level of this question is Level 3. It requires learners to have a good understanding and knowledge of molecular geometry and bonding. [CL3]

- 1.2.1 The bond energy is the amount of energy required to break all the chemical bonds in one mole of the substance. ✓ [CL1] (1)
- 1.2.2 B ✓ [CL2] (1)
- The greater the bond energy the shorter the bond length.
- 1.2.3 The first bond that forms in a double bond has the same bond energy as a single bond. It is a non-polar covalent bond between the two carbon atoms. ✓ The second bond that forms has slightly less bond energy as it is a less stable bond. ✓ OR It breaks more easily than the first single bond. (This is a result of the orientation of the electron orbitals when the bond is formed.) [CL4] (2)
- [11]

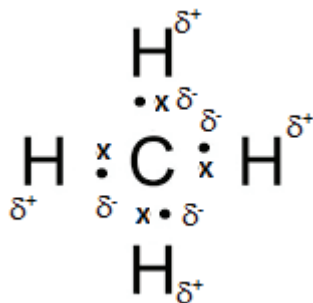
- 2.1.1 a)  b)  [CL2] (4)

- ✓ Correct number of electrons around each atom.
 ✓ Correct positioning of chemical bonds.
 (2 marks for each Lewis diagram)

- 2.1.2 The electronegativity is a measure of the tendency of an atom in a molecule to attract the bonding electrons. ✓✓ [CL1] (2)
- 2.1.3 S: EN = 2,5 O: EN = 3,5
 Δ electronegativity = 3,5 – 2,5 = 1,0 ✓
 Polar covalent bond ✓ [CL2] (2)

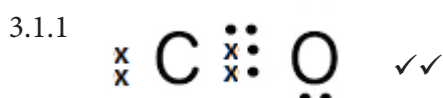
RESOURCE PACK

- 2.1.4 The methane molecule has a tetrahedral shape. ✓ Its polar covalent bonds are arranged symmetrically within the molecule ✓ therefore the overall electron distribution within the molecule is symmetrical ✓ – and the molecule is non-polar.
 ✓ Lewis diagram with dipoles shown



[CL3] (4)

- 2.1.5 a) NH_3 ✓
 b) CH_4 ✓
 c) CO_2 or CO or NO ✓ [CL3] (3)
- 2.2.1 The bond length is the average distance between the nuclei of two bonded atoms. ✓✓ [CL1] (2)
- 2.2.2 a) 120 pm ✓ [CL3] (1)
 b) $800 \text{ kJ} \cdot \text{mol}^{-1}$ ✓✓ [CL2] (2)
 c) Bond energy: the amount of energy per mole of the substance required to break the double bond between the carbon and oxygen atoms. ✓ [CL2] (1)
- 2.2.3 LONGER THAN ✓ The sulfur atom is larger than the carbon atom ✓ therefore the bond length will be longer (the attraction between the two atoms will be less). [CL4] (2)
- 2.2.4 At point D there is very little (or no) force ✓ between the two atoms because the atoms are fairly far apart from one another. ✓ [CL3] (2)
- [25]



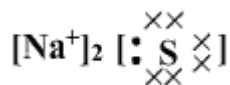
Dative covalent bond occurs when one atom contributes both electrons (from a lone pair of electrons) to the bond. ✓ In this case oxygen donates the electrons to the dative covalent bond. ✓ [CL4] (4)

- 3.1.2 $\Delta \text{ electronegativity} = 3,5 - 2,5 = 1,0$ ✓
 Polar covalent bond
 The molecule is linear therefore it is a dipole (polar molecule). ✓ [CL2] (2)
- 3.1.3 Oxygen is more electronegative than carbon therefore the shared pairs of electrons will be more strongly attracted to oxygen ✓ making it slight negative (δ^-). Carbon will be δ^+ . ✓ [CL3] (2)
- 3.2.1 The CO_2 molecule is linear. ✓ It's shape is ideal since there are no lone pairs in the molecule. ✓
 The SO_2 molecule is angular (bent) ✓ since there are two lone pairs on the central atom (S). ✓ [CL4] (4)

TOPIC 3: ATOMIC COMBINATIONS (MOLECULAR STRUCTURE)

3.2.2 The CO_2 molecule is non-polar✓ because although the bonds are polar covalent, they are arranged symmetrically in the molecule, so the overall electron distribution is also symmetrical.✓ The SO_2 molecule has a bent structure since the lone pairs repel each other strongly. Oxygen is more electronegative than sulfur,✓ therefore the oxygen atoms will be slightly negative (δ^-) and the sulfur atom will be slightly positive (δ^+)✓ (a polar molecule). [CL3] (4)

4.1.1 a)



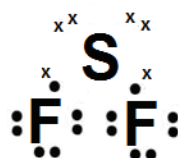
✓ Correct sodium ion

✓ Two sodium ions

✓ Correct sulfide ion

[CL3] (3)

b)

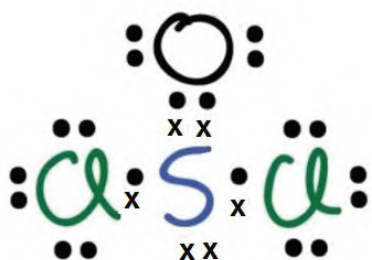


✓ Correct number of electrons sulfur atom

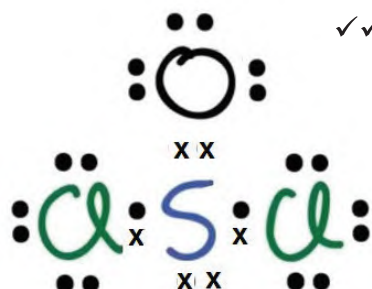
✓ Correct fluoride ions

[CL3] (2)

4.1.2 a)



b)



✓✓ One mark for each

Sulfur is the central atom with its six valence electrons.✓ Oxygen also has six valence electrons and chlorine has seven valence electrons.✓ In the first molecule sulfur and oxygen form a double bond between them; in the second molecule there is a dative covalent bond between sulfur and oxygen.✓ [CL4] (5)

4.1.3 a) SF_2 Angular ✓ Two lone pairs on sulfur atom repel each other causing the shape to be bent (angular). [CL3] (2)
b) BeCl_2 Linear ✓ An ideal shape as there are no lone pairs on Be.✓ [CL2] (2)

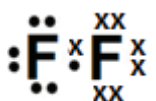
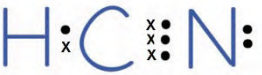
4.2.1 a) Electronegativity: a measure of the tendency of an atom in a molecule to attract the bonding electrons.✓✓ [CL1] (2)

b) Bond polarity: A non-polar covalent bond is one in which the electron density is shared equally between the bonding atoms, ✓ and a polar covalent bond is one in which the electron density is shared unequally between the bonding atoms.✓ [CL1] (2)

4.2.2 a) SF_2 Polar covalent bonds✓
 Δ electronegativity ✓ (method) = $4,0 - 2,5 = 1,5$ ✓ [CL2] (3)

b) H_2S Polar covalent bonds✓
 Δ electronegativity = $2,5 - 2,1 = 0,4$ ✓ [CL2] (2)
[21]

RESOURCE PACK

- 5.1.1 electrons✓ [CL2] (1)
- 5.1.2 polar covalent bonds✓ [CL2] (1)
- 5.2.1 H^+ ion (or H_3O^+ ion or hydrogen ion or hydronium ion) [CL3] (1)
- 5.2.2 HI ✓
It requires the lowest amount of energy to break the bond between H and I✓
therefore it will react most readily with water✓ to release H^+ ions. [CL4] (3)
[6]
- 6.1.1 7✓ [CL2] (1)
- 6.1.2 4✓ [CL2] (1)
- 6.2 A covalent bond is the sharing of electrons to form a chemical bond between two atoms. [CL!] (2)
- 6.3.1

 ✓ correct number of valence electrons in each atom
 ✓ covalent bond shown clearly [CL2] (2)
- 6.3.2

 ✓ correct number of valence electrons in each atom
 ✓ covalent bonds shown clearly [CL2] (2)
- 6.4.1 covalent bond ✓ (non-polar) [CL2] (1)
- 6.4.2 polar covalent bond✓ [CL2] (1)
- 6.5.1 trigonal planar✓ [CL2] (2)
- 6.5.2 tetrahedral✓ [CL2] (2)
- 6.5.3 linear✓ [CL2] (2)
- 6.6.1 The shorter the bond length the higher the bond energy.✓✓ [CL3] (2)
- 6.6.2 $\text{C} \checkmark \text{C}$ can break to form B losing $839 - 614 = 225 \text{ kJ} \cdot \text{mol}^{-1}$ ✓ which is less than the energy required to break a single covalent bond between the carbon atoms.✓
[CL4] (3)
[20]

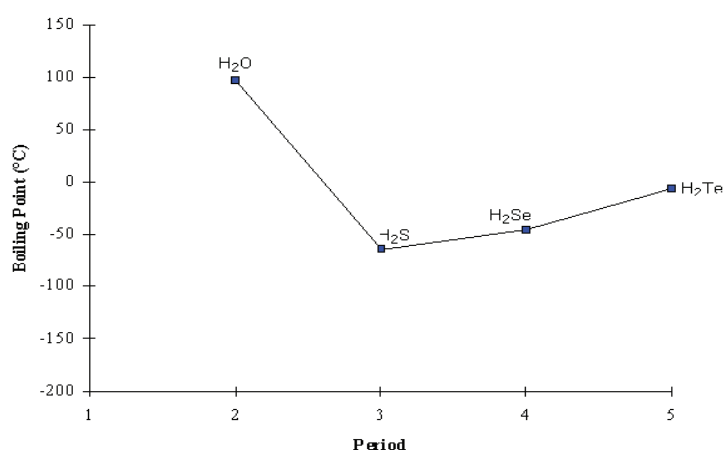
Topic 4: Intermolecular Forces

QUESTIONS

MULTIPLE-CHOICE QUESTIONS

- Between which particles can hydrogen bonds occur?
 - Small molecules that contain hydrogen.
 - Molecules in which hydrogen is bonded to small atoms with high electronegativity.
 - Small atoms with very high electronegativity.
 - Molecules in which hydrogen is bonded to small atoms with low electronegativity.(2)
- The gecko is a small lizard that can climb up a smooth glass window. The gecko has millions of microscopic hairs on its toes and each hair has thousands of pads on its tip. The result is that the molecules in the pads are extremely close to the glass surface on which the gecko is climbing.
What is the attraction between the gecko's toe pads and the glass surface?
 - dative covalent bonds
 - covalent bonds
 - ionic bonds
 - van der Waals forces(2)

Questions 3. and 4. refer to the graph below.



- Water has a higher boiling point than expected because it...
 - is a liquid.
 - has strong hydrogen bonding intermolecular forces.
 - is a polar covalent molecule.
 - has small polar molecules.

RESOURCE PACK

4. The tendency for boiling point to increase from H_2S to H_2Te is due to...
- A stronger London intermolecular forces.
 - B stronger polar covalent bonds.
 - C increasing molecular polarity.
 - D stronger dipole-dipole intermolecular forces.

5. Which one of the following combinations is correct?

	Compound	Strongest Intermolecular force
A	CH_3F	Hydrogen bonding
B	HCl	Hydrogen bonding
C	CH_3OH	Dipole-dipole forces
D	CCl_4	London forces

(2)

6. The fact that noble gases liquefy can be given as evidence for the existence of ...
- A covalent bonds.
 - B dipole-dipole forces.
 - C hydrogen bonds.
 - D London forces.

(2)

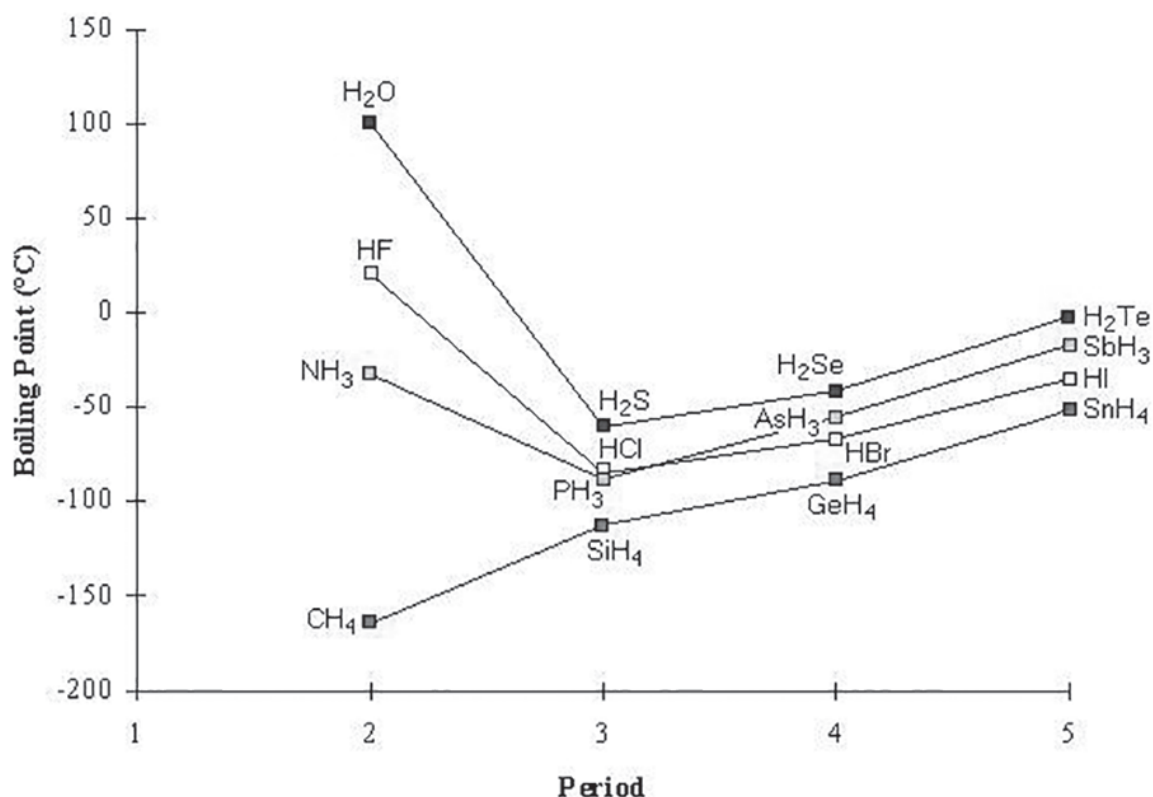
7. Crude oil consists of a mixture of different compounds called hydrocarbons. The molecular formula of some of these compounds is given below. Which one of these compounds will have the highest boiling point?

- A C_4H_{10}
- B C_6H_{14}
- C C_8H_{18}
- D $\text{C}_{10}\text{H}_{22}$

(2)

TOPIC 4: INTERMOLECULAR FORCES AND CHEMISTRY OF WATER

8. The graph shown below gives the boiling points of the hydrides of group 14, 15, 16 and 17 elements.



Water, hydrogen fluoride and ammonia have higher boiling points than expected, which suggests that these substances have stronger intermolecular forces between their molecules.

Which of the following does not contribute to the unexpectedly high boiling points of these three substances?

Nitrogen, oxygen and fluorine

- A have small atoms.
- B are non-metals.
- C have high electronegativities.
- D are bonded to hydrogen atoms.

(2)

RESOURCE PACK

9. Hydrogen chloride occurs naturally as a gas at room temperature. When the temperature is reduced, hydrogen chloride liquefies and eventually solidifies. Which of the following pairs correctly names the bonds and the intermolecular forces found in hydrogen chloride?

	Bond	Intermolecular force
A	Non-polar covalent	Hydrogen bond force
B	Polar covalent	Hydrogen bond force
C	Non-polar covalent	London force
D	Polar covalent	Dipole-dipole force

(2)

10. In which one of the following compounds will hydrogen bonds be present? (All these compounds are in the liquid phase).

- A HCl
B HF
C H₂
D H₂S

(2)

LONG QUESTIONS

1. Water exists as a liquid at room temperature, but methane exists as a gas at this temperature. The molar mass of water is 18 g·mol⁻¹ and that of methane is 16 g·mol⁻¹. Explain, by discussing their intra- and intermolecular forces, why these two substances of similar molar mass exist in different phases at room temperature.

[6]

2. H₂O and CCl₄ are both liquids that also act as solvents for other substances. CCl₄ has been used in the dry-cleaning business, but because it is toxic, it is no longer used for this commercially.

The table below gives information about these two substances.

Substance	Viscosity (kg·m ⁻¹ ·s)	Surface tension (J·m ⁻²)
H ₂ O	1 × 10 ⁻³	7,3 × 10 ⁻²
CCl ₄	9,7 × 10 ⁻⁴	2,7 × 10 ⁻²

- 2.1 Determine the types of bonds in each of these substances. Justify your answer. (3)
2.2 Draw Lewis dot diagrams for:
a) H₂O
b) CCl₄ (4)
2.3 Identify the specific type of intermolecular forces in each compound. (2)
2.4 Explain the differences in viscosity and the surface tension of these two substances as shown in the data in the table above. (5)
2.5 KCl and I₂ are both solids at room temperature. Which of these substances will dissolve most readily in CCl₄? Explain briefly. (3)

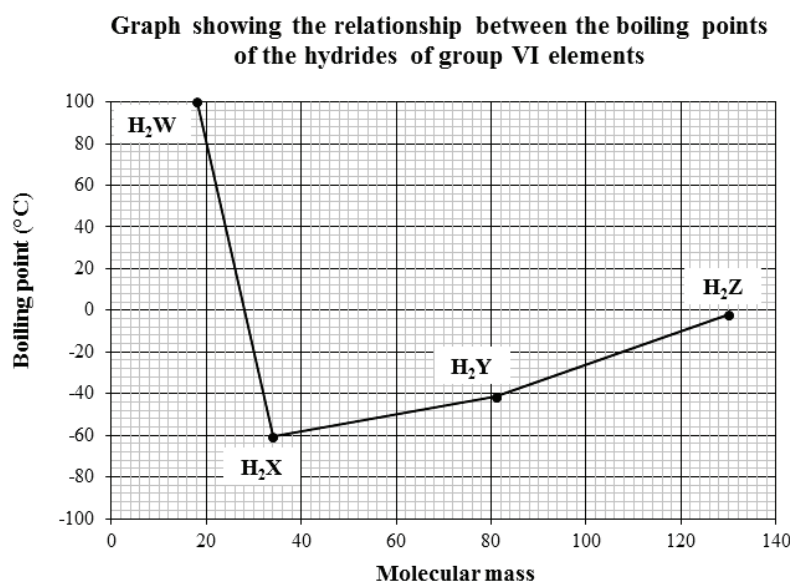
[17]

TOPIC 4: INTERMOLECULAR FORCES AND CHEMISTRY OF WATER

- 3.1 Consider the following information for the compounds CF_4 and NH_3 at standard pressure.

Compound	Melting Point ($^{\circ}\text{C}$)	Boiling Point ($^{\circ}\text{C}$)	Solubility in Water at 20°C
CF_4	-183,6	-127,8	insoluble
NH_3	-77,7	-33,3	soluble

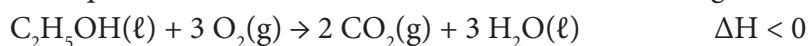
- 3.1.1 What evidence is there in the table that suggests that NH_3 molecules experience stronger intermolecular forces than CF_4 ? (1)
- 3.1.2 Explain fully why CF_4 is insoluble in water but NH_3 is soluble. (5)
- 3.1.3 Ammonia dissolves in water to form the ammonium anion.
 (a) Write down the ionic equation for the reaction. (1)
 (b) Explain the formation of a dative covalent bond in ammonia. (2)
- 3.2 The graph below shows the relationship between the boiling points of the hydrides of group VI elements, namely H_2W , H_2X , H_2Y and H_2Z , and molecular mass.



- 3.2.1 What unexpected result do you observe on the graph? (2)
- 3.2.2 Identify H_2W . (2)
- 3.2.3 Explain the phenomenon referred to in question 3.2.1. (4)
- [17]

RESOURCE PACK

4. The equation for the combustion reaction of ethanol is given below.



- 4.1 Draw the Lewis electron dot diagram for the ethanol molecule ($\text{CH}_3\text{CH}_2\text{OH}$). Use dots and crosses or different colours to show which electrons come from which atoms. (3)
- 4.2 Name and define the specific type of intramolecular bond that forms between the C atom and the O atom in the ethanol molecule. (3)
- 4.3 Name/state the molecular geometry (shape) around... (1)
- (a) each C atom in the ethanol molecule. (1)
- (b) the O atom in the ethanol molecule. (1)
- 4.4 Explain clearly why ethanol is a versatile (very useful) solvent. (4)
- [12]

5. The table below shows the boiling points of four alcohols.

Alcohol	Molecular formula	Boiling point (°C)
Methanol	CH_3OH	65
Ethanol	$\text{C}_2\text{H}_5\text{OH}$	79
Propanol	$\text{C}_3\text{H}_7\text{OH}$	97
Butanol	$\text{C}_4\text{H}_9\text{OH}$	117

- 5.1 Draw a graph of the boiling points (on the y-axis) against the number of carbon atoms in the molecule (on the x-axis). Start with the temperatures from 60°C. Draw the line of best fit for the data. (5)
- 5.2 Describe the trend in the boiling points of the alcohols. (2)
- 5.3 Name the strongest intermolecular forces in these substances. (1)
- 5.4 Explain the trend that this data shows about boiling points of alcohols, with reference to the intermolecular forces. (2)
- 5.5 Methanol (CH_3OH) has a boiling point of 65°C whereas ethane (C_2H_6) boils at -164°C. Explain why these two substances have such wide differences in their boiling points even though they have the same molecular mass. (4)
- [14]

6. The boiling points of four compounds of hydrogen at standard pressure are given in the table below.

Formula	Boiling point (°C)
CH_4	-164
NH_3	-33
H_2O	100
SiH_4	-112

- 6.1 Define the term 'boiling point'. (2)

TOPIC 4: INTERMOLECULAR FORCES AND CHEMISTRY OF WATER

- 6.2 Fully explain the difference in boiling points between CH_4 and NH_3 . (3)
- 6.2.1 NH_3 . (3)
- 6.2.2 SiH_4 . (3)
- 6.3 Explain why the boiling points of NH_3 and H_2O differ by referring to electronegativity, the molecular shapes and the intermolecular forces in these substances. (4)
- 6.4 Select a substance from the table above which has
- 6.4.1 polar molecules. (1)
- 6.4.2 non-polar molecules. (1)
- [14]
7. The properties of water play an important role for survival on Earth.
- 7.1 Explain why the density of water decreases when it freezes. (3)
- 7.2 Give ONE reason why this decrease in the density of water is important for life on Earth. (2)
- 7.3 Briefly explain, by referring to the unique property of water involved, how water regulates temperatures on Earth. (2)
- 7.4 Which ONE of intermolecular forces or intramolecular forces is stronger? Explain your answer using water as an example. (3)
- 7.5 Water evaporates slower than eucalyptus oil. Explain this statement in terms of intermolecular forces. (3)
- [13]

RESOURCE PACK

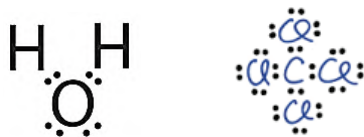
MARKING GUIDELINES**MULTIPLE CHOICE**

1. B ✓✓ It is the small atom with high electronegativity that is bonded to a hydrogen atom that allows hydrogen bonds to form between molecules of a substance. [CL2] (2)
2. D ✓✓ It is the force between the molecules of glass and the molecules in the pads of gecko's toes that allows it climb the smooth glass surface. The other choices here concern intramolecular forces (chemical bonds). The gecko does not bond permanently with the glass! [CL3] (2)
3. B ✓✓ The other substances do not have hydrogen bonds between their molecules. It must therefore be the strong hydrogen bonds between water molecules that accounts for its high boiling point. [CL2] (2)
4. D ✓✓ The molecules of H_2S , H_2Se and H_2Te are polar, therefore the intermolecular forces between them are dipole-dipole forces. The forces become stronger as the size of the molecules increases. [CL2] (2)
5. D ✓✓ CH_3F does not have hydrogen bonding because the fluorine atom is attached to the carbon atom (not to a hydrogen atom). HCl has dipole-dipole forces (not hydrogen bonds). CH_3OH has dipole-dipole forces, but its strongest intermolecular forces are hydrogen bonds. The only correct answer is therefore CCl_4 which does have London forces between its molecules. [CL3] (2)
6. D ✓✓ The noble gases are non-polar substances. The fact that they liquefy gives evidence that forces exist between their molecules in the liquid phase. These forces are London forces. [CL2] (2)
7. D ✓✓ These compounds are all the same type of substance; therefore, they will have the same type of intermolecular forces between their molecules. The intermolecular forces will increase in strength as the size of the molecules increases because there will be more points of contact available for the forces to hold the molecules closer together. [CL2] (2)
8. B ✓✓ It is the small atom with high electronegativity that is bonded to a hydrogen atom that allows hydrogen bonds to form between molecules of a substance. Therefore, the fact that these elements are non-metals does not contribute directly to the hydrogen bonding. [CL 4] (2)

TOPIC 4: INTERMOLECULAR FORCES AND CHEMISTRY OF WATER

9. D ✓✓ Hydrogen chloride has one polar covalent bond between its atoms. It has dipole-dipole intermolecular forces between its molecules. [CL 3] (2)
10. B ✓✓ F is a small atom with high electronegativity bonded to H. It is these types of molecules that can form hydrogen bonds between them. [CL2] (2)

LONG QUESTIONS

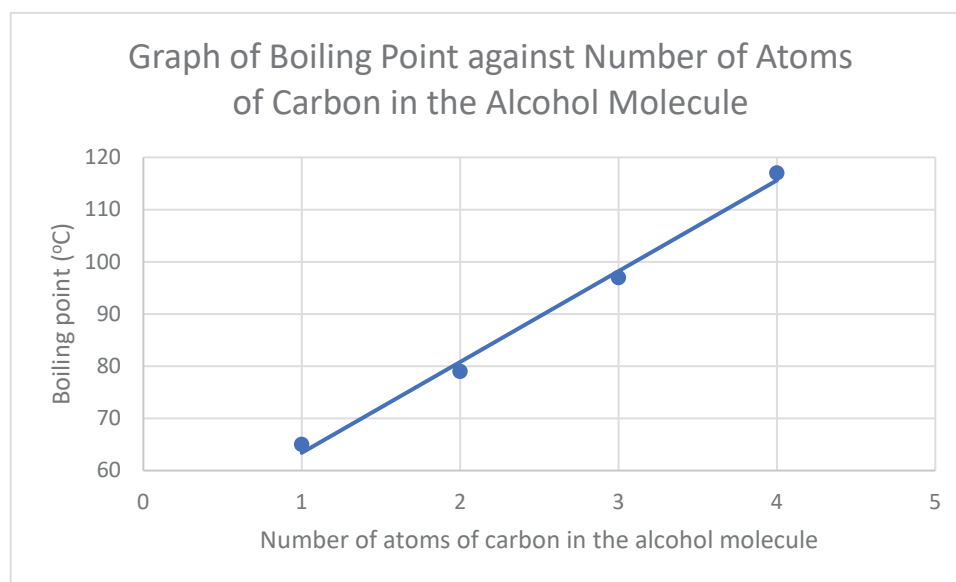
1. Both methane and water have polar covalent bonds between their atoms. ✓
 Δ electronegativity C-H = 2,5 – 2,1 = 0,4
 Δ electronegativity H-O = 3,5 – 2,1 = 1,4
 (for calculating Δ electronegativity correctly) ✓
 The bond between oxygen and hydrogen is more polar than that between carbon and hydrogen.
 Water has an angular structure, whereas methane is tetrahedral. The electron distribution in the water molecule is asymmetrical, but it is symmetrical in the methane molecule. ✓
 Water molecules are polar; methane molecules are non-polar. ✓
 The intermolecular forces between water molecules are hydrogen bond (dipole-dipole) forces which are much stronger ✓ than the London forces (induced dipole-induced dipole forces) that exist between methane molecules, ✓ therefore water is a liquid and methane is a gas at room temperature. [CL 4] [6]
- 2.1 Water: polar covalent bond ✓
 CCl₄: polar covalent bond ✓
 The differences in electronegativity in each of these bonds is greater than 0. [CL 2] (3)
- 2.2  2 marks for each correct diagram.
 -1 any mistake. [CL 3] (4)
- 2.3 Water: hydrogen bond ✓ (dipole-dipole) forces CCl₄: London forces ✓ [CL 2] (2)
- 2.4 The viscosity of water is about 10 times that of CCl₄, and its surface tension is just a little more than twice as high. ✓ The hydrogen bonds ✓ between molecules in water are much stronger ✓ than the dispersion (London) forces in CCl₄. ✓ Therefore, it makes sense that water molecules will cling to one another more tightly than CCl₄ molecules – and therefore have stronger forces in the surface of the liquid (surface tension) and greater viscosity than CCl₄. ✓ [CL 4] (4)
- 2.5 Iodine is a non-polar substance, whereas KCl is an ionic substance. ✓ Substances with similar strength of intermolecular forces are able to dissolve in one another. ✓ Therefore, iodine will dissolve in CCl₄ (but KCl will not). ✓ [CL 3] (3)
- [17]

RESOURCE PACK

- 3.1.1 The melting and boiling points of NH_3 are much higher than those of CF_4 therefore it takes more energy to overcome the intermolecular forces in NH_3 . [CL2] (2)
- 3.1.2 CF_4 molecules are non-polar molecules; they have London forces between them.
 ✓ Water molecules are polar molecules with stronger hydrogen bonds between them. ✓ If CF_4 is added to water, the water molecules are so strongly attracted to each other that they will not move apart to allow CF_4 molecules to mix in with them. ✓ Therefore CF_4 is insoluble in water.
 NH_3 molecules are polar molecules with hydrogen bonding force between them.
 ✓ Because these forces are of comparable strength to the forces between water molecules, ammonia molecules will mix with water molecules ✓ – ammonia will dissolve in water. [CL3] (5)
- 3.1.3 a) $\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$ ✓ [CL3] (1)
 b) The lone pair of electrons on the NH_3 molecule is donated to form the bond
 ✓ with the H^+ ion that is removed from the water molecule. ✓ [CL3] (2)
- 3.2.1 The boiling point of H_2W is much higher than expected. ✓✓ [CL2] (2)
- 3.2.2 Water (or H_2O) ✓✓ [CL2] (2)
- 3.2.3 Water molecules have much stronger intermolecular forces between their molecules than the other hydrides of group VI elements do. ✓ This occurs because hydrogen is bonded to oxygen which is a very small atom ✓ with high electronegativity ✓. Hydrogen bonds exist between water molecules; dipole-dipole forces exist between the molecules of the other substances. ✓ [CL3] (4)
- 4.1. [17]
- ✓ 4 electrons on each carbon atom
 ✓ 1 electron on each hydrogen atom
 ✓ 6 electrons on oxygen atoms (1 lone pair)
 [CL3] (3)
- 4.2 polar covalent bond ✓
 The bonded pair of shared electrons ✓ is more strongly attracted to the more electronegative atom (oxygen). ✓ [CL2] (3)
- 4.3 a) tetrahedral ✓ [CL2] (1)
 b) angular ✓ [CL3] (1)
- 4.4 Ethanol molecules have a non-polar side and a polar side to them ✓✓ therefore they are able to dissolve polar ✓ and non-polar substances. ✓ [CL4] (4)

TOPIC 4: INTERMOLECULAR FORCES AND CHEMISTRY OF WATER

5.1



- ✓ Appropriate title for the graph
- ✓ Axes labelled correctly with SI unit (y-axis)
- ✓ Appropriate scales on axes
- ✓ Points plotted correctly
- ✓ Line of best fit (straight line) [CL2] (5)

5.2 The boiling point increases at a steady rate ✓ as the number of carbon atoms in the alcohol molecule increases. ✓ [CL3] (2)

5.3 Hydrogen bonds ✓ [CL2] (1)

5.4 As the number of carbon atoms in the molecules increases, the size of the molecules also increases. ✓ The intermolecular forces between the molecules become stronger because there are more points of contact available on the bigger molecules. ✓ [CL3] (2)

5.5 Methanol has hydrogen bonding between its molecules whereas ethane which is a non-polar substance ✓ has weaker London forces between its molecules. ✓ Hydrogen bonding forces are the strongest type of intermolecular force; ✓ London forces are the weakest forces between molecules ✓ therefore there is a wide difference between their boiling points. [CL3] (4)

6.1 The boiling point is the temperature at which the vapour pressure of a liquid is equal to the atmospheric pressure. ✓ ✓ [CL1] (2)

6.2 NH_3 is a polar molecule which has strong hydrogen bonds between its molecules. ✓ CH_4 is a non-polar molecule with weaker London forces between its molecules. ✓ It takes much more energy to overcome the forces between the molecules of NH_3 ✓ therefore its boiling point is much higher than that of CH_4 . [CL3] (3)

RESOURCE PACK

- 6.3 NH_3 and H_2O both have polar molecules with hydrogen bonds between them.
 ✓ The electronegativity difference between N and H is $3,0 - 2,1 = 0,9$ whereas the electronegativity difference between H and O is $3,5 - 2,1 = 1,4$. ✓ The covalent bonds inside the water molecule are more polar than those in the NH_3 molecule. ✓ Water molecules have an angular shape, whereas NH_3 molecules are trigonal pyramidal. ✓ This allows each water molecule to form two hydrogen bonds with two other water molecules making the intermolecular forces in water much stronger than those in NH_3 . ✓ [CL4] (4)
- 6.4.1 NH_3 or H_2O ✓ [CL2] (1)
- 6.4.2 CH_4 or SiH_4 ✓ [CL2] (1)
- [14]
- 7.1 Hydrogen bonds between water molecules act in particular directions. ✓ When ice crystallises, this directional property of hydrogen causes the molecules to align next to each in a more open structure than they do when water is liquid. ✓ Water's volume increases as the molecules move into their positions in the crystal, therefore its density decreases. ✓ [CL3] (3)
- 7.2 Water freezes from the top down so, ice forms on the surface of water first while the bottom of the lake, dam, river or ocean remains liquid. ✓ The sheet of ice over the top of a frozen lake insulates the water below it from the colder air in the atmosphere. In this way plants and animals that live in water are able to do so under the sheet of ice. ✓ [CL2] (2)
- 7.3 Water has a very high specific heat capacity. (It takes a large amount of energy to be absorbed per kg of water to raise the temperature of water by 1°C). ✓ Water also releases a large amount of energy per kg when its temperature drops by 1°C . In this way water acts as a heat reservoir – and regulates the temperatures on Earth. ✓ [CL3] (2)
- 7.4 Intramolecular forces are stronger than intermolecular forces. ✓ In water the polar covalent bond between hydrogen atoms and oxygen atoms requires more energy to break ✓, that the energy required to overcome the hydrogen bonding forces that keep the molecules together in liquids and solids. ✓ [CL3] (3)
- 7.5 Eucalyptus oil molecules are non-polar molecules with weaker London forces between them. ✓ Water molecules have stronger hydrogen bonds between their molecules. ✓ It takes more energy to release a water molecule from the surface than it does to release molecules of eucalyptus oil. ✓ [CL3] (3)



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**PHYSICAL SCIENCES
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2025

DEFINITIONS

- Chemical bond – a mutual attraction between two atoms resulting from the simultaneous attraction between their nuclei and the outer electrons.
- Intramolecular bond – bond which occurs between atoms within molecules
- Lewis dot diagram – a structural formula in which valence electrons are represented by dots or crosses.
- Valence electrons – the electrons in the highest energy level of an atom in which there are electrons.
- Covalent bond – the sharing of electrons between two atoms to form a molecule.
- Ionic bond – a transfer of electrons and subsequent electrostatic attraction
- Metallic bond – metallic bonding as the bond between positive ions and delocalized valence electrons in a metal
- Molecule – a group of two or more atoms covalently bonded and that function as a unit.
- Bonding pair – a pair of electrons that is shared between two atoms in a covalent bond.
- Lone pair – a pair of electrons in the valence shell of an atom that is not shared with another atom.
- Electronegativity – a measure of the tendency of an atom in a molecule to attract bonding electrons.
- Non-polar covalent bond – a bond in which the electron density is shared equally between the two atoms.
- Polar covalent bond – a bond in which the electron density is shared unequally between the two atoms.
- Bond energy of a compound – the energy needed to break one mole of its molecules into separate atoms.
- Bond length – the average distance between the nuclei of two bonded atoms

MOLECULAR STRUCTURE

- The structure mainly depends on the type of chemical bond (force) that exists between the atoms that the molecule consists of.

RULES FOR BOND FORMATION

Different atoms:

- each with an unpaired valence electron, can share these electrons to form a chemical bond
example: two H atoms form a H_2 molecule by sharing an electron pair
- with paired valence electrons, called lone pairs, cannot share these 4 electrons and cannot form a chemical bond
example: no bond forms between 2 He atoms
- with unpaired valence electrons, can share these electrons and form a chemical bond for each electron pair shared (multiple bonds can be formed)
example: two O atoms form an O_2 molecule

Atoms with an empty valence shell can share a lone pair of electrons from another atom to form a coordinate bond or a dative covalent

- example: in NH_4^+ the lone pair of nitrogen from NH_3 is shared with H^+

CHEMICAL BONDS

A chemical bond is the net electrostatic force that two atoms sharing electrons exert on each other.

- Bonding occurs when valence electrons are shared between two atoms or transferred from one atom to another.
- Valence electrons correspond to the group number of an element in the Periodic table
- Valence electrons: Valence electrons or outer electrons are the electrons in the highest energy level of an atom in which there are electrons.
- The type of bond that forms is dependent on the electronegativity difference (ΔEN) between the atoms.
- Electronegativity is a measure of the tendency of an atom in a molecule to attract bonding electrons. If one atom has a greater electronegativity than another, the electrons will be pulled more towards one atom than another. Such a shift of electrons creates negative and positive charge distributions inside the molecule.
 - $\Delta EN = 0$: Non-polar (pure) covalent
 - $\Delta EN < 1$: Weakly polar covalent
 - $\Delta EN > 1$: polar covalent
 - $\Delta EN > 2,1$: Ionic (transfer of electrons)
 - $\Delta EN > 3$: purely ionic

Types of bonds and molecules

- Non-polar bond: a bond in which the electron density is shared equally between the two atoms
Example: $H - H$
- EN (H) = EN (H): Bonding electrons are evenly shared. Charge is evenly distributed and no dipole formed.
 $\Delta EN = 2,1 - 2,1 = 0$
- Polar bond: a bond in which the electron density is shared unequally between the two atoms.

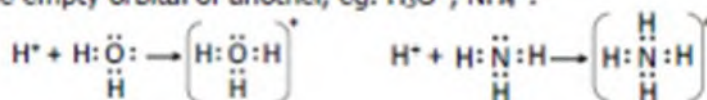
- Example: $\text{H} - \text{Cl}$
- $\text{EN}(\text{Cl}) > \text{EN}(\text{H})$: Electrons shift towards chlorine. Chlorine is slightly negative (δ^-) and hydrogen is slightly positive (δ^+).
 $\Delta\text{EN} = 3,0 - 2,1 = 0,9$
- Polar molecule: A molecule over which charge is distributed unevenly.
 Example: H_2O
- Non-polar molecule: A molecule over which the charge is evenly distributed.

Covalent Bonding (Between non-metals)

- Covalent bonding involves the sharing of electrons between two atoms to form a molecule.
- The covalent bond may be non-polar, weakly polar or polar.
- A non-polar (pure) covalent bond is a bond in which the electron density is shared equally between the two atoms.
 example:
 Two chlorine atoms are joined by a non-polar covalent bond ($\text{Cl} - \text{Cl}$)
 EN for chlorine atom is 3,0.
 $\Delta\text{EN} = 3,0 - 3,0 = 0$
- In a weakly polar covalent bond there is an unequal sharing of electrons
 example:
 A hydrogen atom and a bromine atom are joined by a weakly polar bond ($\text{H} - \text{Br}$)
 EN for hydrogen is 2,1 and EN for Br is 2,8.
 $\Delta\text{EN} = 2,8 - 2,1 = 0,7$
- A polar covalent bond has an unequal sharing of electrons leading to a dipole forming
- example:
 A hydrogen atom and an oxygen atom are joined by a polar covalent bond (H_2O)
 EN for oxygen atom is 3,5.
 $\Delta\text{EN} = 3,5 - 2,1 = 1,4$
- bonds between them. If two pairs of electrons are shared, a double bond is formed, e.g. between two O atoms to form H_2O_2 .
- Atoms with an empty valence shell can share a lone pair of electrons from another atom to form a coordinate covalent or dative covalent bond, e.g. in NH_4^+ the lone pair of nitrogen is shared with H^+ and in H_3O^+ the lone pair of oxygen is shared with H^+ .

DATIVE BONDING/ COORDINATE BONDING

A coordinate bond is formed when one atom donates both electrons to form a bonding pair. The filled electron orbital (lone pair) of one atom overlaps with the empty orbital of another, eg. H_3O^+ , NH_4^+ .



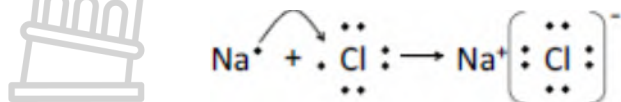
Ionic Bonding (between metals and non-metals)

- metal atom gives electron or electrons to non-metal atom
- metal forms a positive ion and positive ion is called a cation
- non-metal forms a negative ion and negative ion is called an anion
- electrostatic attraction of ions leads to formation of giant crystal lattice

Example:

Show how the ionic bond is formed between the sodium and the chlorine atoms.

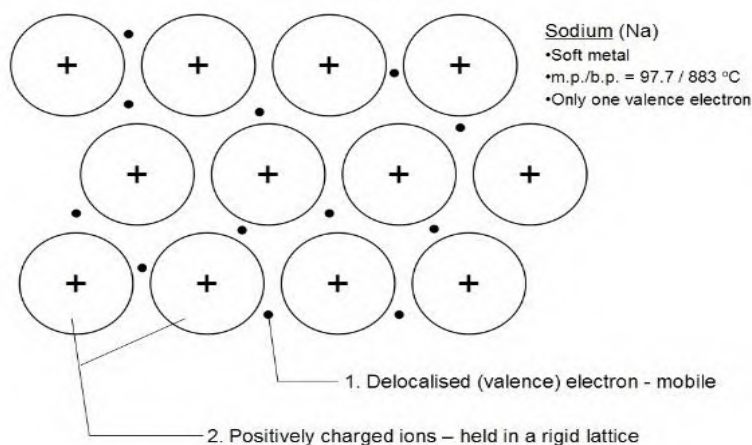
1. $\text{Na} - 1e^- \rightarrow \text{Na}^+$ (Na atom loses one electron)
2. $\text{Cl} + 1e^- \rightarrow \text{Cl}^-$ (Cl atom gains one electron)
3. $\text{Na}^+ + \text{Cl}^- \rightarrow \text{NaCl}$ (electrostatic attraction between the two ions)



Metallic Bonding (between metals)

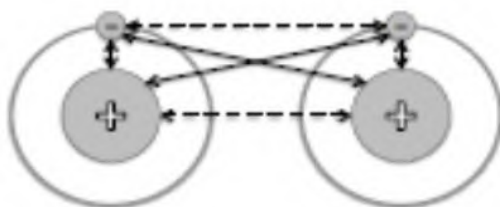
- Metallic bonding forms between the positive metal kernels and the sea of delocalized electrons.
- The metal atoms release their valence electrons to surround them.
- There is a strong but flexible bond between the positive metal kernels and a sea of delocalised electrons

Metallic Bonding



BOND ENERGY

- Bond energy is required to break one mole of its molecules into separate atoms.



- There are various attractive and repulsive forces at play between the two atoms during bonding.
- Attractive forces between the protons of one atom and the electrons of another.
- Attractive forces between the protons and electrons from the same atom.
- A repulsive force between the protons from each atom.
- A repulsive force between the electrons from each atom.
- The net electrostatic forces will determine bond strength, which can be quantified as the bond energy. This is the energy required to break the bond, or it is the energy released when bonds are formed.

FACTORS INFLUENCING BOND STRENGTH

*Bond length

The shorter the length of the bond, the stronger the bond.

*Atom size

The smaller the atoms, the stronger the bond.

***Bond order**

The more bonds (double, triple bonds etc.) between the atoms, the stronger the bond.

BOND LENGTH

Bond length: The average distance between the nuclei of two bonded atoms.

As the atoms get closer, the attractive forces get stronger until the minimum possible potential energy is reached (bond energy). The distance between the nuclei of the atoms at the minimum potential energy is the bond length. If the two atoms get closer than the bonding length, they will be forced apart by the repulsive forces, increasing the potential energy.

FACTORS INFLUENCING BOND LENGTH

*Atom size- The smaller the atoms, the shorter the bond.

*Bond order -The more bonds (double, triple bonds etc.) between the atoms, the shorter the bond.

*Difference in electronegativity (ΔEN). The greater the ΔEN , the shorter the bond.

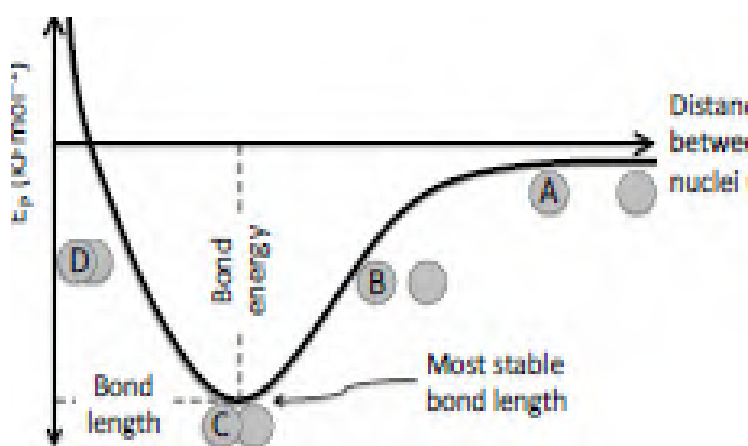
Process of bond formation

*Atoms are infinitely separated; potential energy is nearly zero.

*As the atoms move closer, the forces of attraction and repulsion increase until the forces of attraction dominate.

*The lowest, most stable energy state reached, chemical bond forms.

*Atoms move too close, forces of repulsion increase, potential energy increases



WORKED EXAMPLE 1

Both aluminium fluoride (AlF_3) and phosphorus trifluoride (PF_3) contain fluorine. Aluminium fluoride is a colorless solid in the production of aluminum, whilst phosphorus trifluoride is a poisonous, colorless gas.

- 1.1 Explain the difference between a covalent bond and ionic bond (2)
- 1.2 Name the type of chemical bond between particles in:
 - 1.2.1 AlF_3 (1)
 - 1.2.2 PF_3 (1)
- 1.3 Draw a lewis structure for:
 - 1.3.1 AlF_3 (3)
 - 1.3.2 PF_3 (2)
- 1.4 Write down the molecular shape of PF_3 . (1)
- 1.5 The melting point of AlF_3 is $1\ 291^\circ C$ and that of PF_3 is $-151^\circ C$. Fully explain this difference in melting point. (4)

[14]

SOLUTIONS

- 1.1 A covalent bond is a sharing of electrons between two atoms, whilst, ionic bond forms when there is a transfer of electrons from one atom to another (non-metal to metal) and oppositely charge atoms attract each other. $\checkmark\checkmark$ (2)

- 1.2 1.2.1 Ionic bond ✓ (1)
1.2.2 Covalent bond ✓ (1)

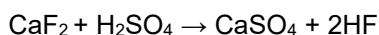


- 1.4 Pyramidal ✓ (1)
1.5 The ionic bonds between the aluminum fluoride are stronger than the dipole-dipole forces in between phosphorus fluoride molecules. ✓ ✓ (1)
More energy is needed to overcome ionic bond in AlF_3 than intermolecular forces in PF_3 ✓ ✓ (4)

[14]

WORKED EXAMPLE 2

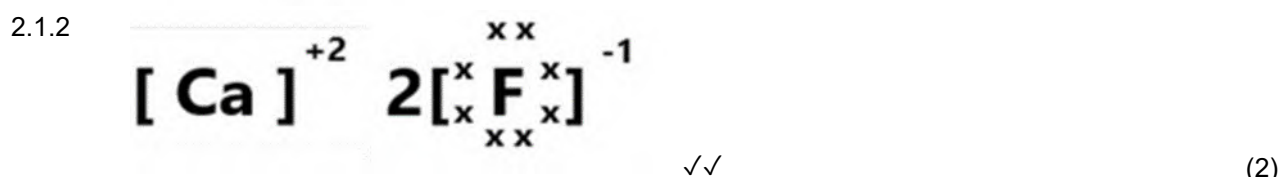
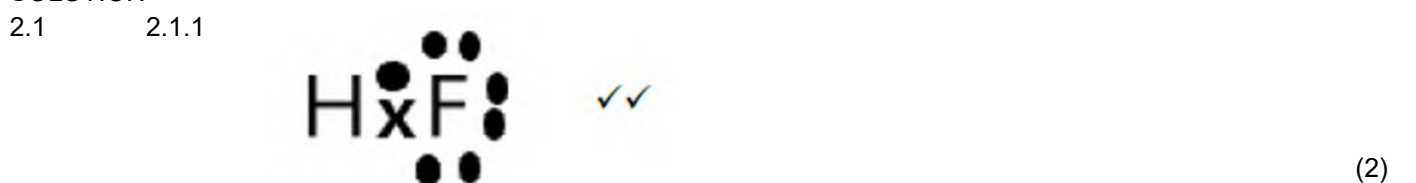
2. Hydrogen fluoride (HF) is used in the manufacturing of refrigerants, herbicides, pharmaceuticals and gasoline. Hydrogen fluoride (HF) can be prepared by treating calcium fluoride with sulphuric acid, according to the following balanced equation.



- 2.1 Draw a Lewis diagram for:
2.1.1 HF (2)
2.1.2 CaF_2 (2)
2.2 Define a polar covalent bond (2)
2.3 Which of the product will likely be:
2.3.1 A gas at room temperature (1)
2.3.2 A crystalline solid at room temperature (1)
2.4 Define the term electronegativity (2)
2.5 Use electronegativity to determine the types of bonds in CaF_2 (2)
2.6 Will HF be a polar or non-polar molecule? (1)
2.7 Explain your answer in question using calculation (2)

[15]

SOLUTION



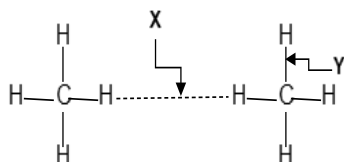
- 2.2 A polar covalent bond is a bond in which the electron density is shared unequally between the two atoms ✓✓ (2)
2.3 2.3.1 HF/Hydrogen fluoride ✓ (1)
2.3.2 $CaSO_4$ /Calcium sulfate ✓ (1)

- 2.4 Electronegativity is a measure of the tendency of an atom in a molecule to attract bonding electrons. ✓✓ (2)
- 2.5 $\Delta EN = 4 - 1 = 3$ ✓
Ionic bond ✓ (2)
- 2.6 Polar molecule ✓ (1)
- 2.7 $\Delta EN = 4 - 2,1 = 1,9$ ✓ polar bond ✓
Thus, the molecule is polar (2)

[15]

MULTIPLE CHOICE QUESTIONS

- 1.1 The diagram below is used to demonstrate the differences between interatomic bonds (chemical bonds) and intermolecular forces.



- A X is stronger than Y
B X is an interatomic bond
C X is an intermolecular force
D Y is an intermolecular force (2)

- 1.2 Which ONE of the following bonds will have the highest bond energy?

- A H – H
B C – H
C C = C
D C ≡ C

(2)

- 1.3 Consider the following compounds and their respective boiling points.

COMPOUNDS	BOILING POINT (°C)
He	-268,9
HBr	-66
HF	19,5

The correct arrangement for the decreasing strength of the intermolecular forces of the given compounds is ...

- A HF, He, HBr.
B HBr, He, HF.
C HF, HBr, He.
D He, HBr, HF.

(2)

- 1.4 Consider the substances **P**, **Q** and **R**. **P** dissolves in **Q** but not in **R**.

The MOST LIKELY bonds or intermolecular forces in substances **P**, **Q** and **R** are :

	P	Q	R
A	ionic bonds	Dipole-dipole forces	London forces
B	London forces	Dipole-dipole forces	ionic bonds
C	Dipole-dipole forces	London forces	ionic bonds

D	Ionic bonds	London forces	Dipole-dipole forces
---	-------------	---------------	----------------------

(2)

1.5 The intermolecular forces holding the molecules in ice together are ...

- A ion-dipole forces.
- B Van der Waals forces.
- C covalent bonds.
- D hydrogen bonds.

(2)

[10]

LONG QUESTIONS

QUESTION 2

Consider the following compounds.

O_2 ; H_2O ; H_3O^+

2.1 Define the term *molecule*.

(2)

2.2 Draw the Lewis structures for:

2.2.1 O_2

(2)

2.2.2 H_2O

(2)

2.3 The H_3O^+ ions forms when the oxygen atom in H_2O donates its lone pair of electrons into the vacant orbital of H^+ .

2.3.1 Write down the NAME of the type of bond described by the underlined phrase.

(2)

2.3.2 Draw the Lewis structure for the H_3O^+ -ion.

(2)

2.4 Which molecule is polar, H_2O or O_2 ? Explain your answer.

(2)

[12]

QUESTION 3

3.1 Draw a Lewis diagram of methane.

(2)

3.2 What is the name of the special kind of bond found inside the methane molecule?

(1)

3.3 Answer the following questions about magnesium chloride:

3.3.1 Make a Lewis diagram to indicates the transfer of electrons between magnesium and chlorine

(4)

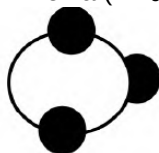
3.3.2 Magnesium chloride is ionic substance. Prove this statement with relevant calculation

(2)

[9]

QUESTION 4

4 Ammonia (NH_3) is represented in the diagram below.



4.1 Write down the number of valence electrons for nitrogen

(1)

4.2 Draw the Lewis structure for ammonia

(3)

4.3 Write down the molecular shape of the ammonia molecule

(1)

4.4 Explain why ammonia does not have a trigonal planar shape.

(2)

4.5 Is ammonia a polar or non-polar molecule?

(1)

4.6 Refer to the structure and difference in electronegativity to explain the answer to QUESTION 4.5.

(2)

When a hydrogen ion reacts with an ammonia molecule, it forms the ammonium ion.

4.7 By referring to the ammonium ion, explain what is a dative covalent bond

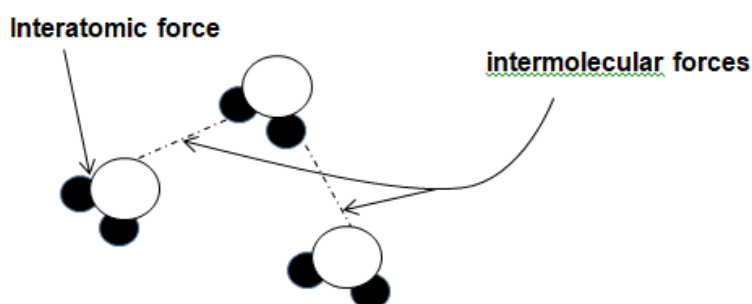
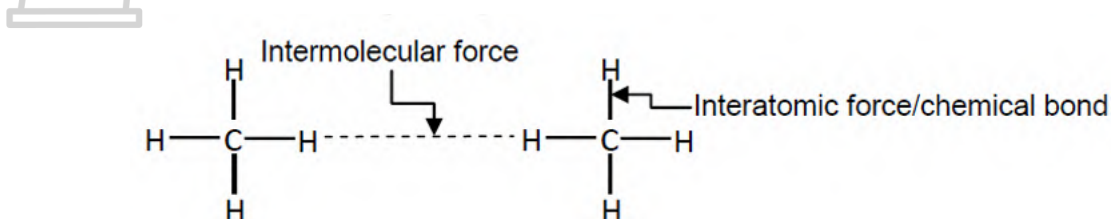
(2)

[12]

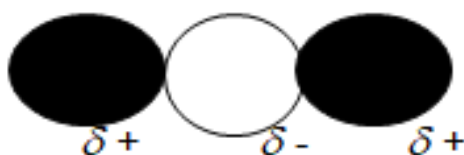
INTERMOLECULAR FORCES

- Intermolecular forces are forces of attraction between molecules in a substance.
- Intermolecular forces are weak forces of attraction between molecules or between atoms of noble gases
- Intermolecular forces are not the same as intramolecular bonds.
- Intramolecular (interatomic) bonds exist between atoms in a molecule.
- The intermolecular forces are **weaker** than interatomic forces.

Examples of intermolecular forces and intramolecular (interatomic) forces

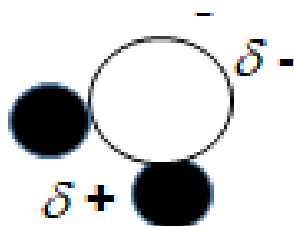


- Molecules can be **polar** or **non-polar**.
- The polarity of a molecule is determined by two factors:
 - The difference in electronegativity between the bonding atoms.
 - The geometry (shape) of a molecule
- **Non-polar molecules** have no dipoles (positive and negative ends) e.g., CO_2



Polar molecule

- In polar molecule there are δ^+ and δ^- ends e.g., H_2O and the geometry of the water molecule is angular (bent).



- H_2O molecule has H-end being δ^+ and O-end being δ^- , it has **dipoles**. The molecule is polar

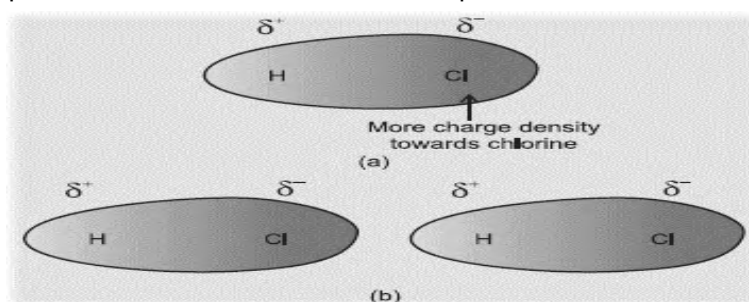
- **Non-polar bonds:**

- H_2 is non-polar because the two atoms are identical, there is even or symmetrical distribution of charge, this makes H_2 a non-polar molecule.

TYPES OF INTERMOLECULAR FORCES

- Different intermolecular forces (Van der Waals forces) ARE

- Mutually induced dipole forces or London forces: Forces between non-polar Molecules
- Dipole-dipole forces: Forces between two polar molecules
- Dipole-induced dipole forces: Forces between polar and non-polar molecules
- Hydrogen bonding: Forces between molecules in which hydrogen is covalently bonded to nitrogen, oxygen, or fluorine – a special case of dipole-dipole forces
- Ion-dipole forces: Forces between ions and polar molecules



INTERMOLECULAR FORCES AND PHYSICAL PROPERTIES (factors affecting intermolecular forces)

Boiling point and melting point

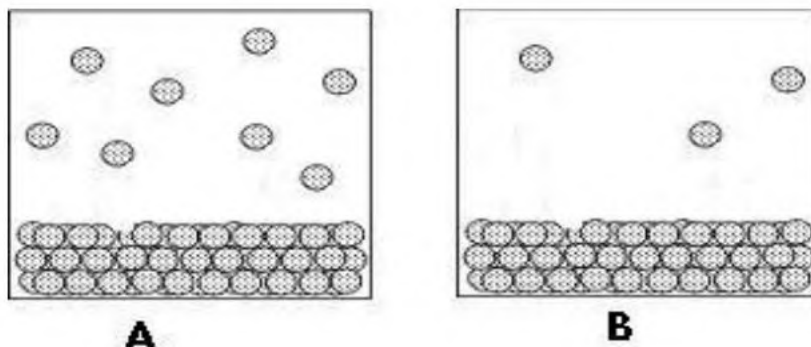
- Intermolecular forces dictate several properties, such as melting points, boiling points, solubility of substances, vapour pressure, viscosity etc.
- Boiling point is the temperature at which the vapour pressure of a substance equals atmospheric pressure. The stronger the intermolecular forces, the higher the boiling point.
- Melting point is the temperature at which the solid and liquid phases of a substance are at equilibrium. The stronger the intermolecular forces, the higher the melting point.
- A liquid with high boiling point, like water (H_2O , b.p. $100\text{ }^{\circ}C$), exhibits stronger intermolecular forces compared to a liquid with low boiling-point, like hexane (C_6H_{14} , b.p. $68,73\text{ }^{\circ}C$).

Solubility

- Solubility is the property of a solid, liquid, or gaseous chemical substance (solute) to dissolve in a solid, liquid, or gaseous solvent to form a homogeneous solution.
- Miscible liquids have similar polarities.
- For example, methanol and water are miscible and are both polar and capable of hydrogen bonding.
- When methanol and water are mixed, they interact through intermolecular hydrogen bonds of comparable strength to the methanol–methanol, and water–water interactions; thus, they are miscible.
- Likewise, nonpolar liquids like hexane and bromine are miscible with each other through dispersion forces.
- The chemical axiom “like dissolves like” is useful to predict the miscibility of compounds.
- For example, nonpolar hexane is immiscible in polar water. Relatively weak attractive forces between the hexane and water do not adequately overcome the stronger hydrogen bonding forces between water molecules.

Vapor Pressure

- Vapor pressure is the pressure exerted by a vapour at equilibrium with its liquid in a closed system. The stronger the intermolecular forces, the lower the vapour pressure.
- At the surface, some molecules of a liquid have enough kinetic energy to break their attractive forces with neighbouring molecules.
- These molecules escape from the liquid phase and form a gas above the surface of the liquid. If there is a lid, pressure develops.

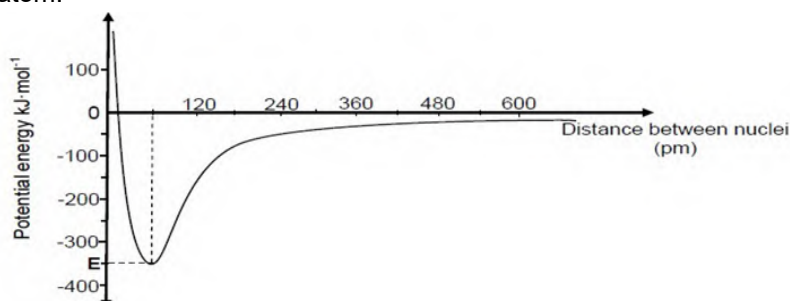


- In the picture above, A has more gas, meaning gas molecules are breaking away from the liquid state easier than molecules in B.
- The intermolecular forces must be weaker in A than in B.
- Molecules in A are less attracted to each other than molecules in B.
- Intermolecular forces in A are weaker than those in B

WORKED EXAMPLES

QUESTION 1

The graph below shows the change in energy that takes place when a hydrogen (H) atom approaches a bromine (Br) atom.



- 1.1 Define the term *bond length* (2)
- 1.2 From the graph, write down the:
 - 1.2.1 Bond length, in pm, of the H-Br bond (2)
 - 1.2.2 Energy, in kJ mol^{-1} (1)
 - 1.2.3 Name of the potential energy represented by E (2)
- 1.3 How will the bond length of H-F bond compare to that of H-Br bond? (2)
Write down EQUAL TO, SHORTER THAN, LONGER THAN. Give a reason for your answer [9]

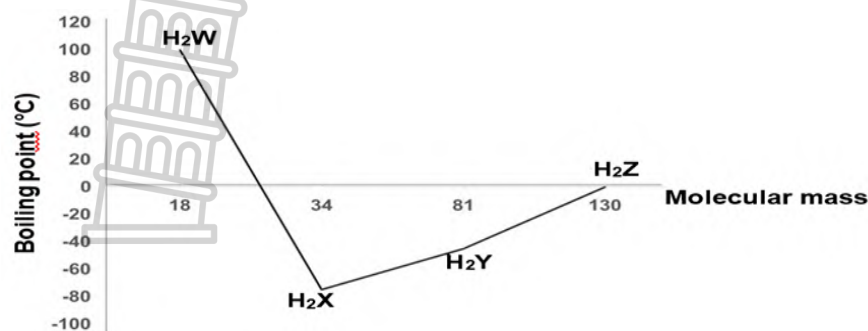
SOLUTION

- 1.1 The average distance between nuclei of two bonded atoms in a molecule. (2)
- 1.2
 - 1.2.1 60 (pm) ✓✓ (2)
 - 1.2.2 350 (kJ mol^{-1}) ✓✓ (2)
 - 1.2.3 Bond energy ✓ (1)
- 1.3 Shorter than, ✓ F atoms are smaller than Br atoms and come closer to H atom. ✓ (2)

[9]

QUESTION 2

The graph below shows the relationship between the boiling points of the hydrides of the Group 16 elements and molecular mass.



- 2.1 Define the term *boiling point*. (2)
- 2.2 Consider the boiling points of hydrides, excluding H₂W.
 - 2.2.1 What is the relationship between molecular mass and boiling points of compounds? (2)
 - 2.2.2 Explain the trend in boiling points. Refer to molecular mass, intermolecular forces and energy to explain your answer. (3)
- 2.3 H₂W does not follow the trend in boiling points for the 16 hydrides.
 - 2.3.1 Identify the compound H₂W (1)
 - 2.3.2 Name the types of intermolecular forces found in this compound (2)
 - 2.3.3 Explain why the boiling point is higher than expected (2)

[12]

SOLUTION

- 2.1 The temperature at which the vapour pressure of a substance equals atmospheric pressure (2)
- 2.2 2.2.1 As the molecular mass increases, the boiling point increases. ✓✓ (2)
- 2.2.2 As the molecular mass increases, ✓ the strength of the intermolecular forces increases. ✓
Therefore more energy is needed to overcome/ weaken the intermolecular forces. ✓(No mark if BROKEN is used instead of overcome or weaken.) Thus the boiling point increases. (3)
- 2.3 2.3.1 water✓ (1)
- 2.3.2 Hydrogen Bond✓ (2)
- 2.3.3 Hydrogen bonds are stronger than dipole-dipole forces ✓ therefore more energy is needed to overcome/weaken the forces. ✓ Thus the boiling point is higher than expected. (2)

[12]

ACTIVITIES: INTERMOLECULAR FORCES

QUESTION 1

Which intermolecular forces are found in:

- 1.1 hydrogen fluoride (HF) (1)
- 1.2 methane (CH₄) (1)
- 1.3 potassium chloride in ammonia (KCl in NH₃) (1)
- 1.4 krypton (Kr) (1)

[4]

QUESTION 2

Refer to the table below and answer questions that follow.

	distance	Melting Point	Boiling Point	Relative molecular/ atomic mass



Sodium	97,79	882,8	23
Bromine (Br ₂)		58,8	160
Hydrogen chloride (HCl)	-1 14,2	-85,1	36,5
Water (H ₂ O)		100	18

- 2.1 Define the term *boiling point*. (2)
- 2.2 By referring to the forces present in hydrogen chloride, explain the difference between intermolecular forces and interatomic forces. (2)
- 2.3 Which of these substances will be a liquid at +50 °C? (1)
- 2.4 Refer to intermolecular forces and energy and explain why the boiling point of HCl is lower than the boiling point of H₂O (4)
- 2.5 NaCl dissolved in 1-120. (4)
- 2.5.1 Write down the name the intermolecular forces between NaCl and 1-120. (1)
- 2.5.2 Are the intermolecular forces stated in QUESTION 2.5.1 weaker or stronger than those of H₂O molecules. (1)
- 2.6 Refer to TYPE of INTERMOLECULAR FORCES, MOLECULAR MASS, STRENGTH of INTERMOLECULAR FORCES and ENERGY to explain the big difference in boiling points of Br₂ and HCl. (4)

[15]

QUESTION 3

The table below shows the relationship between the melting points of three molecules:

Molecule	Melting point (°C)
CH ₄	-182,5
CF ₄	-150
CCl ₄	-23

- 3.1 Define the term *melting point*. (2)
- 3.2 Explain the trend in melting points of the molecules in the above table by referring to the intermolecular forces and energy involved. (4)
- 3.3 Which molecule in the table will have the highest vapour pressure at a given temperature? Explain the answer by referring to the data in the table. (2)
- 3.4 Water (H₂O) is a smaller molecule than CCl₄, but water has a higher melting point than CCl₄. Explain this observation by referring to the type of intermolecular forces involved. (2)
- 3.5 Write down the name of intermolecular force that will exist in a mixture of H₂O and CCl₄. (2)

[12]

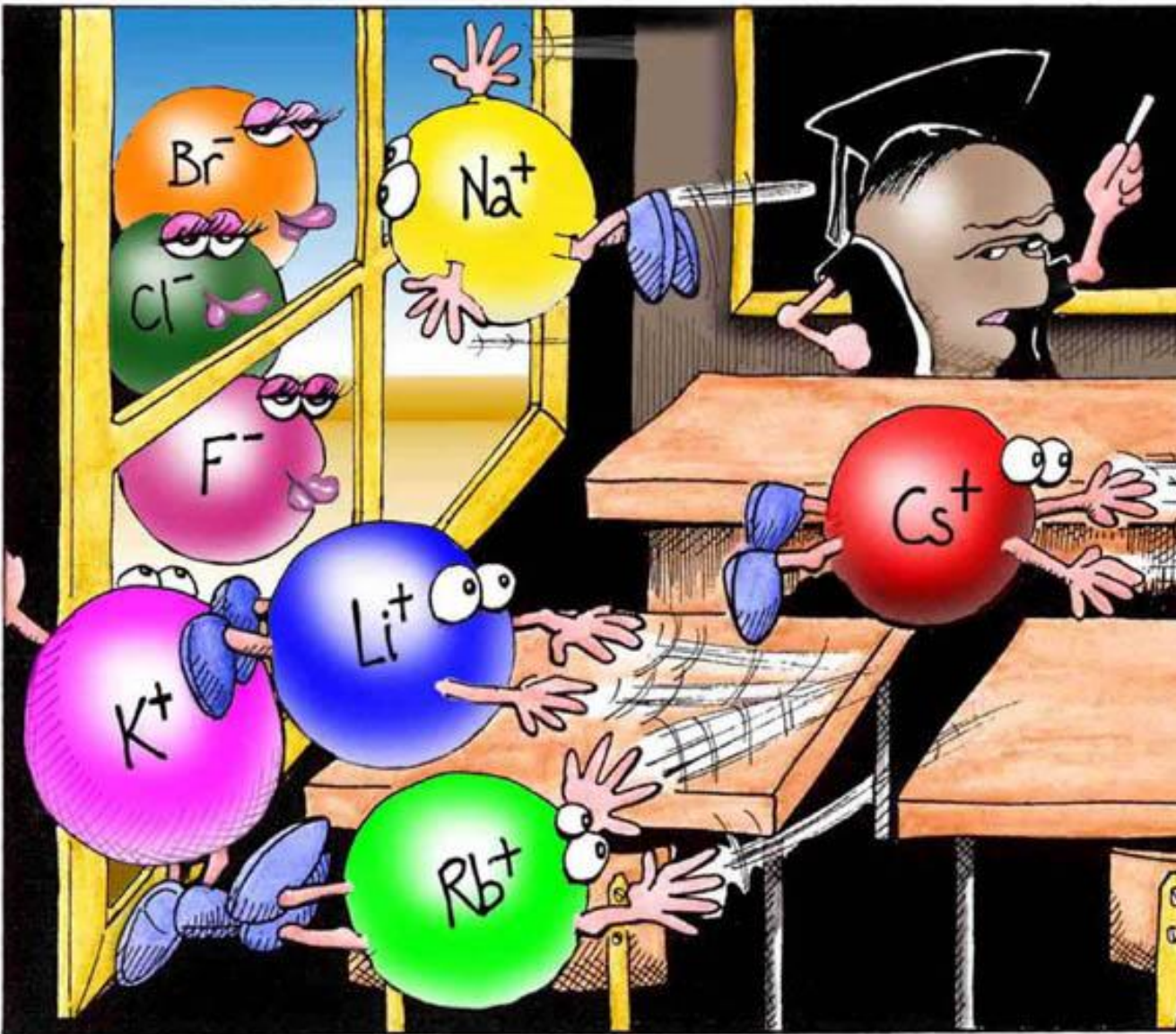
QUESTION 4

In the table below the boiling points of four hydrogen halides are given.

Hydrogen halides	A	B	C	D
	HF	HCl	HBr	HI
Boiling points (°C)	19,4	-85	-67	-35,5

- 4.1 Define the term *boiling point*. (2)
- 4.2 What is the relationship between strength of the intermolecular forces and boiling point? (1)
- 4.3 Which of the hydrogen halides (A, B, C or D) in the table above has the ...

Unit 4:
Chemical
Bonding &
Molecules
(Chapter 6 in
book)



"Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive..?"

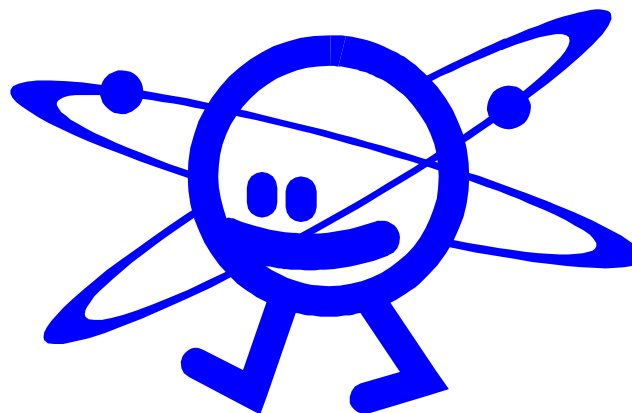
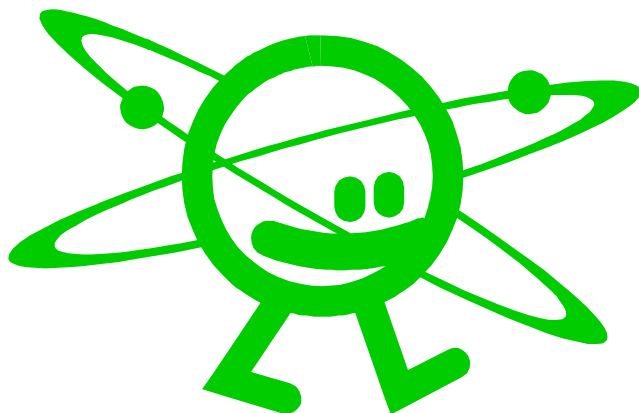
Cartoon courtesy of NearingZero.net

Chemical Bonding

pgs.161-182

Chemical Bonds

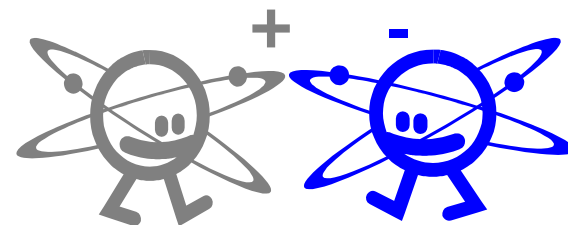
- ❑ Attraction between the *nuclei* and *valence electrons* of different atoms that "glues" the atoms together.
 - ❑ the difference between materials as diverse as diamonds and pencils is how they're glued together.
- ❑ Why?
 - ❑ Bonded atoms are more stable than solo atoms
- ❑ How?
 - ❑ Atoms will share or exchange valence electrons to achieve a full outer shell (usually octet).



3 Main Types of Bonds

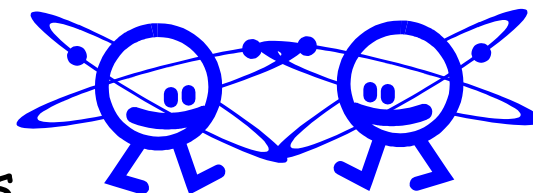
Ionic Bonds - Transfer of electrons
between atoms

- electrical attraction between *cations* & *anions*
- *Formed by: metals & non-metals*



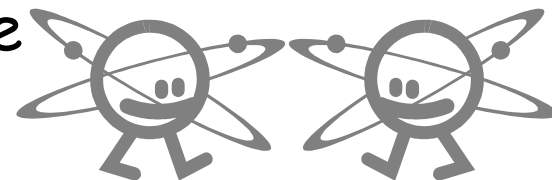
Covalent Bonds - sharing of electrons
between atoms

- "co" = sharing, "valent" = outer electrons
- *Formed by: non-metals & non-metals*



Metallic Bonds - Metal atoms that share
a "sea of electrons"

- *Formed by: metals & metals*



Predicting Bond Types

- Bonding is not usually purely ionic or covalent, but somewhere in between
- The difference in *electronegativity* strength of the atoms in a bond can help us estimate what percentage of the bond will be ionic
(see example on next slide)

Using the Periodic Table to Determine Bond Types

Electronegativity

Electronegativity

1											13	14	15	16	17	
H 2.1											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	
Li 1.0	Be 1.5											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
Na 0.9	Mg 1.2	3	4	5	6	7	8	9	10	11	12	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Cs 0.8	Ba 0.9	La* 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9					
Fr 0.7	Ra 0.9	Ac† 1.1	* Lanthanides: 1.1–1.3 † Actinides: 1.3–1.5													

below 1.0

1.0–1.4

1.5–1.9

2.0–2.4

2.5–2.9

3.0–4.0

Ionic bond=

- **metal** (weak) & **non-metal** (strong)
- huge difference in strength (1.7 or more)

Metallic Bond =

- **2 Metals** (both weak)

Covalent bond =

- **2 non-metals** (strong)
- close to same strength

Summary: Ionic Bonds vs. Covalent Bonds

<u>Ionic</u>	<u>Covalent</u>
Electrons are Transferred (become charged ions that are attracted)	Electrons are shared
Metal + non-metal (ex: Li + K)	2 non-metals (O + O or O + N)
One atom is a lot higher electronegativity than the other (1.7)	Close to equal electronegativities (less than 1.7)

Lewis Dot Structures

Octet Rule



- Most* atom wants to have 8 electrons in their valence shell (outermost shell)
- Chemical compounds form so that each atom can complete their octet by gaining, losing or sharing electrons
- *Exceptions =
 - H & He (they only want 2 electrons in their valence shell)
 - B (forms bonds so it will have only 6 electrons)
 - F, O & Cl (will sometimes be surrounded by more than 8 electrons because they are so electronegative)

Lewis Dot Structure

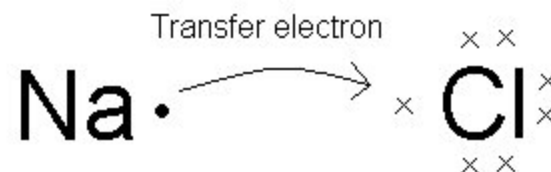
- Picture showing how many **valence** electrons an atom has (dots).
- Helps determine how atoms will bond.



Ex: Phosphorus (has 5 valence electrons)

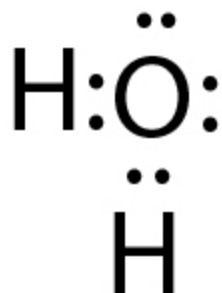
Lewis Dot Structures for Ionic Compounds

- A way to show how atoms achieve the octet with each other.
- Note:
 - the transfer of the electron
 - the charges ions that result

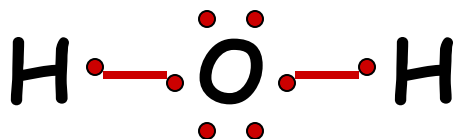


**This is how
we draw it**

Lewis Dot Structures for Covalent Molecules



Or



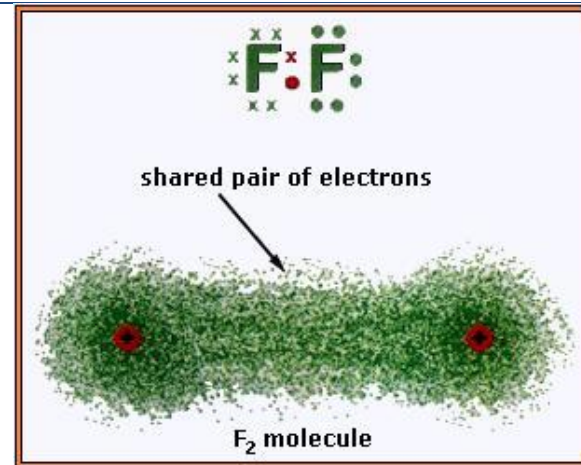
2 ways to show:

- With electrons being shared in between
- Line showing the sharing of pair of electrons

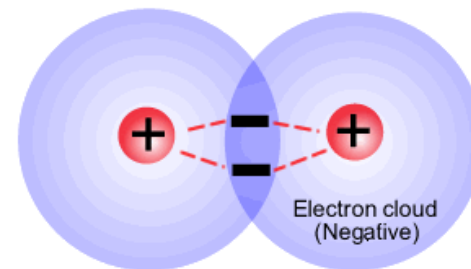
3 Bonds Types in More Depth

Covalent Bonds

- Result from the sharing of electron pairs between two atoms
- Molecule = termed used to describe atoms are held together by covalent bonds



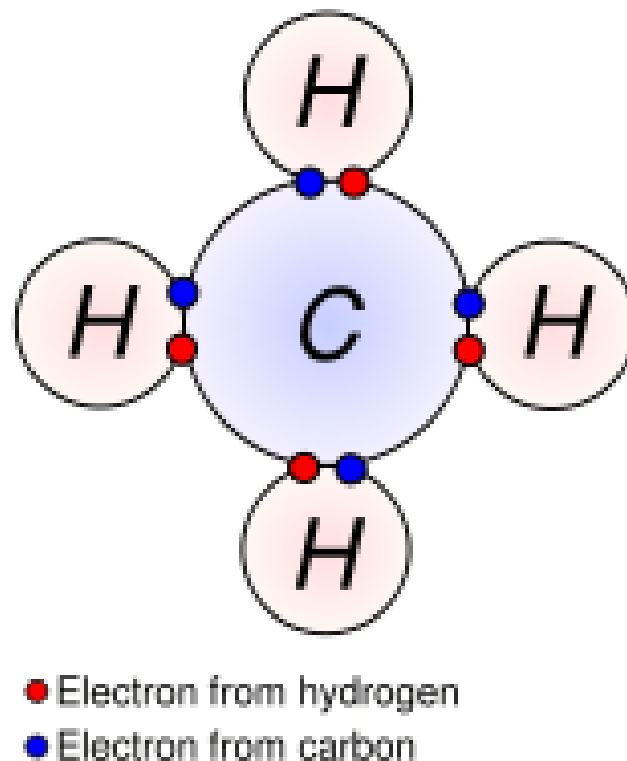
The electrons experience a force of attraction from both nuclei. This negative - positive - negative attraction holds the two particles together



This attraction is called a chemical bond
one pair of electrons constitutes ONE bond

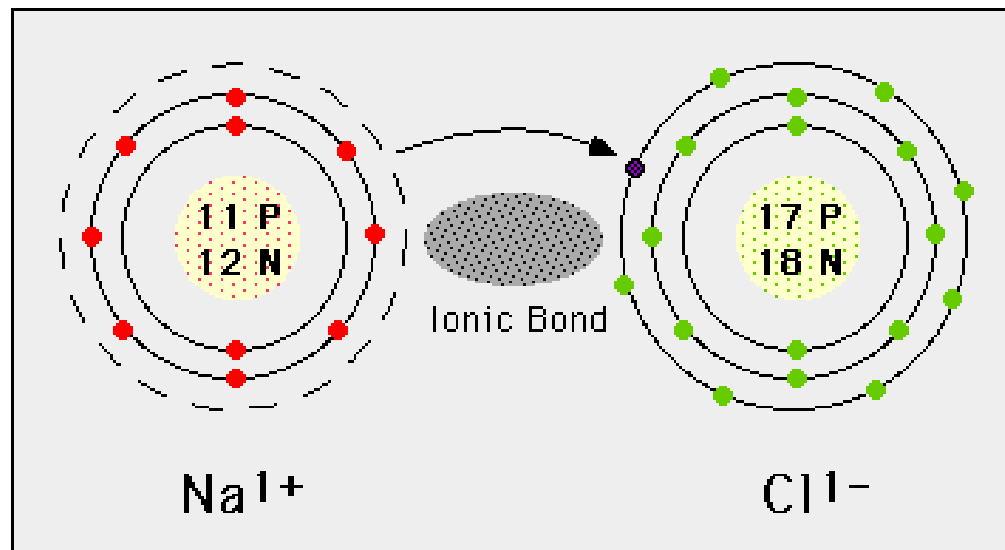
Covalent Bonds

- Occurs between 2 non-metals
- electrons are shared
- 2 types of covalent bonds: Polar and non-polar (to be discussed later)
- Ex: Water & most biological molecules (sugars, fats, proteins)
- Can form single, double, or triple bonds

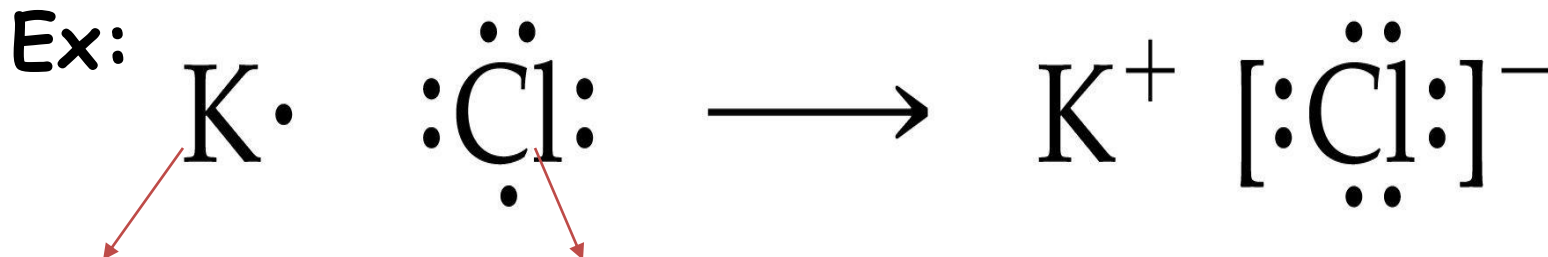


Ionic Bonds

- Forms between:
Metal + Non-Metal
- Electrons are transferred



Ionic Bonds (cont.)



Electroneg= .8

Electroneg= 3.0

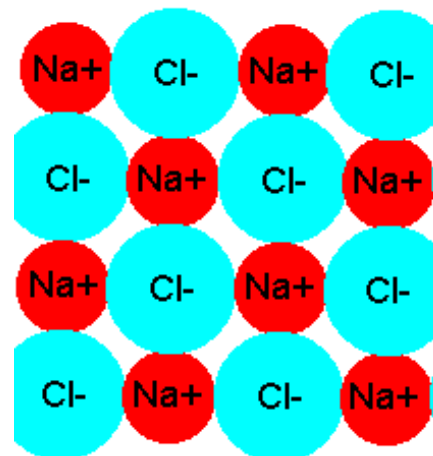
- Cl is so much stronger that it will “take” K’s electron

- The transfer of electron causes K to be a cation (+) and Cl to be an anion (-).

- Oppositely charged particles are highly attracted to each other... Ionic bond!

Characteristics of Ionic Compounds

- Shape- crystal lattice of alternating positive and negative ions
- Ex: NaCl and salts
- Ionic bonds are strong so they are:
 - hard
 - have a high melting point
 - high boiling point

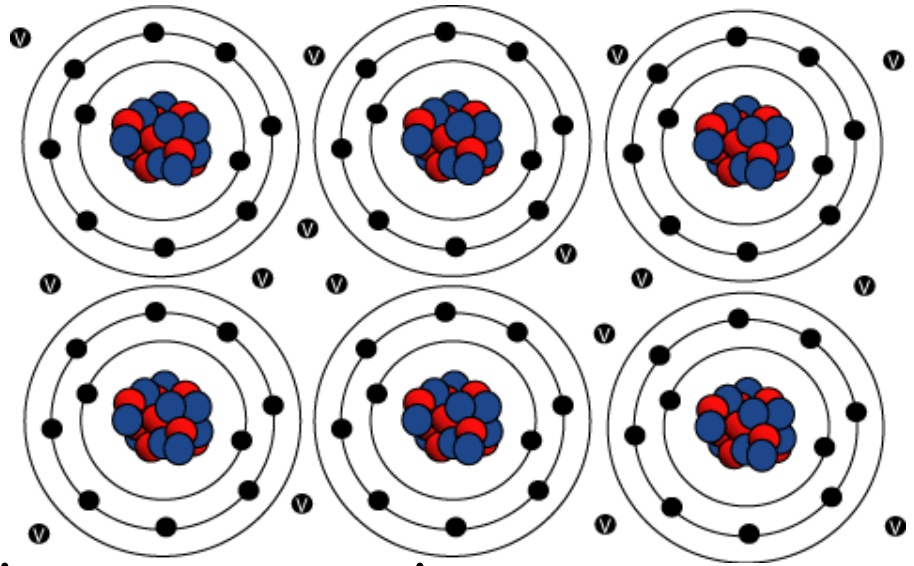


Crystal
Lattice



Metallic Bonds- "sea of electrons"

- Forms between 2 metals
- Metal atoms valence electrons overlap creating a "sea of electrons".



• Electrons do not belong to any one atom, but roam freely throughout the metal atoms

• Ex: Brass (alloy of Cu + Zn)

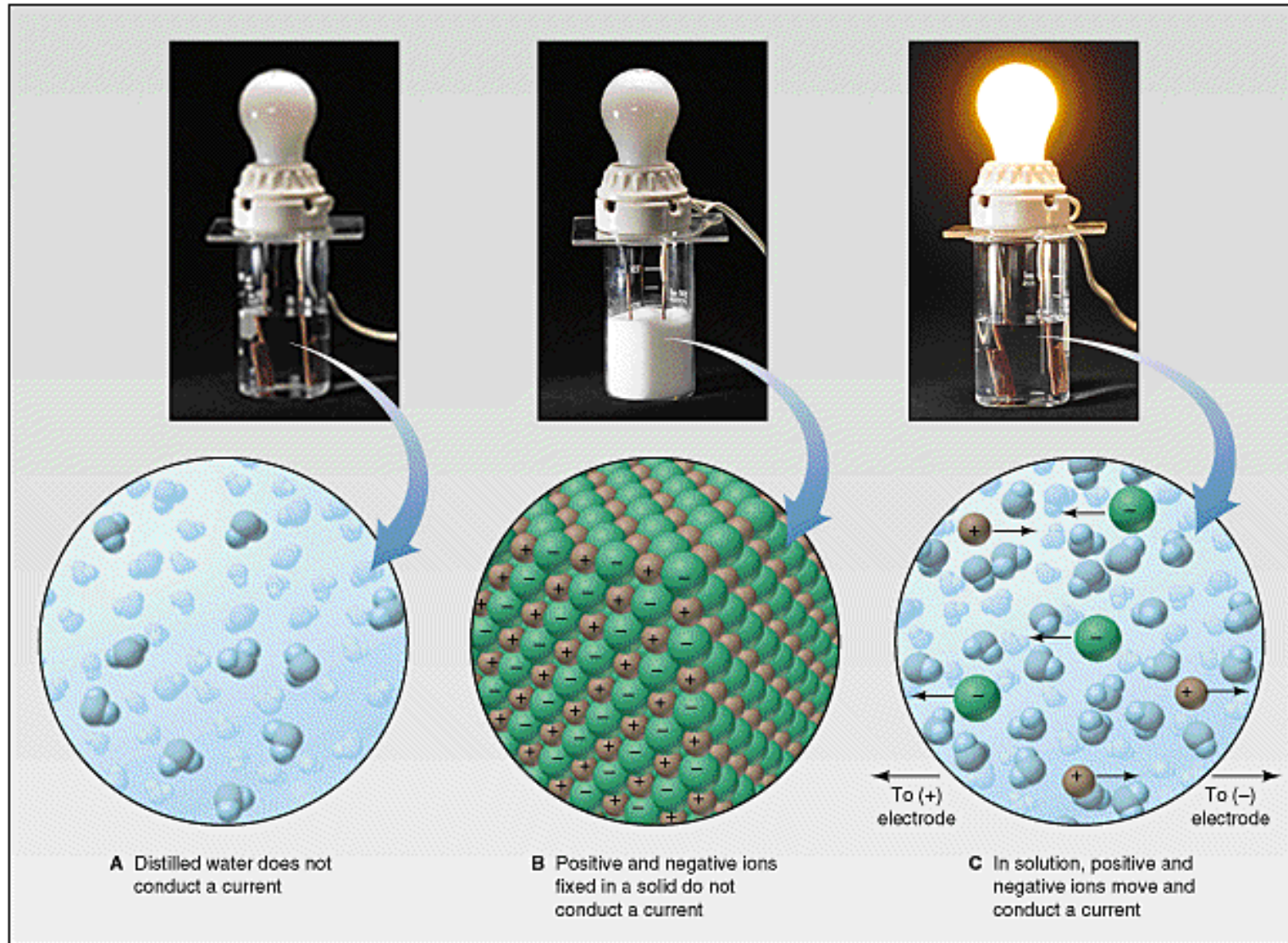


Metallic Characteristics

- Because of these roaming “sea” of electrons:
 - metals are great conductors of heat/electricity
 - they are ductile (can be made into wire)
 - they are malleable (can be hammered into sheets)



Electrical Conductivity



Properties & Bonding Type

pgs.161-182

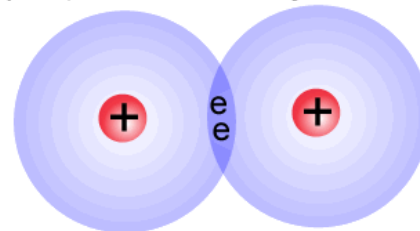
Comparison	Covalent	Ionic Bonds	Metallic Bonds
Formation			
<i>Types of Atoms</i>	Non-metal & non-metal	Non-metal & metal	Metal & metal
<i>Electron Distribution</i>	Shared	Transferred	Sea of electrons
Characteristics			
<i>Bond Strength</i>	Strong	Very strong	Varies
<i>Structure</i>	Neutral group	Crystal lattice	crystalline
Properties of Compounds			
<i>Type of Compound</i>	Molecular	Ionic	metallic
<i>Melting Point</i>	Low	Very high	n/a
<i>Boiling Point</i>	Low	High	Very high
<i>Malleability</i>	n/a	Not malleable, brittle	Very malleable
<i>Ductility</i>	n/a	Not ductile	Very ductile
<i>Conductivity</i>	Not conductive	Conductive	Highly conductive

Bond Energy & Bond Length

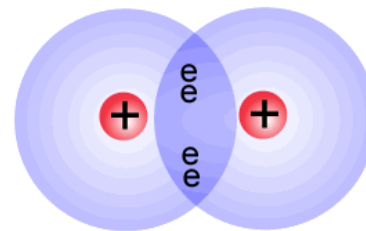
Bond Energy- energy required to break bond

	<u>Bond Length</u>	<u>Bond Energy</u>
Single Bond	————	Low
Double Bond	=====	
Triple Bond	≡≡≡	High

Only one pair of electrons holding the nuclei together



Two pair of electrons hold the nuclei tighter and closer



Bond Energy & Bond Lengths

<u>Bond</u>	<u>Length (picometers)</u>	<u>Energy (kJ/mol)</u>
H—Br	141	366
H—C	109	413
H—N	101	391
H—O	96	464
H—S	93	339
C—O	143	360
C=O	129	799
C—C	154	348
C=C	134	614
C C	120	839
O—O	148	145
O=O	121	498
N—N	145	170
N=N	125	418
N N	110	945

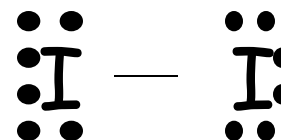
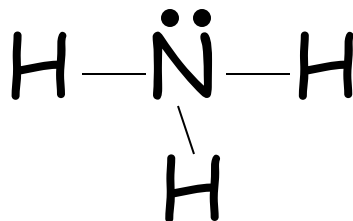
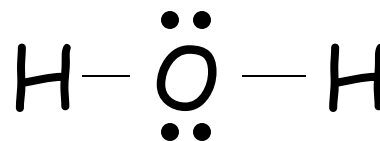
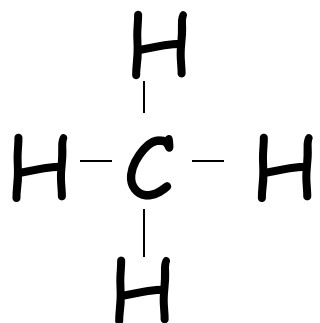
Lewis Structures in Covalently Bonded Molecules & HONC Rule

pgs. 183 - 186

Drawing Lewis Dot Structures for Molecules

- arrange atoms to form a skeleton
 - Carbon is center atom
 - Hydrogen is never a central atom
- Pair up all electrons
 - unpaired electrons can pair unpaired electron from another atom to form a bond
- Make sure each atom of the molecule obeys the octet rule & HONC rule
- Make sure you have correct # of valence electrons

Examples of Lewis Dot Structure



Multiple Covalent Bonds: Double bonds

Two pairs of shared electrons

$O_2 :$

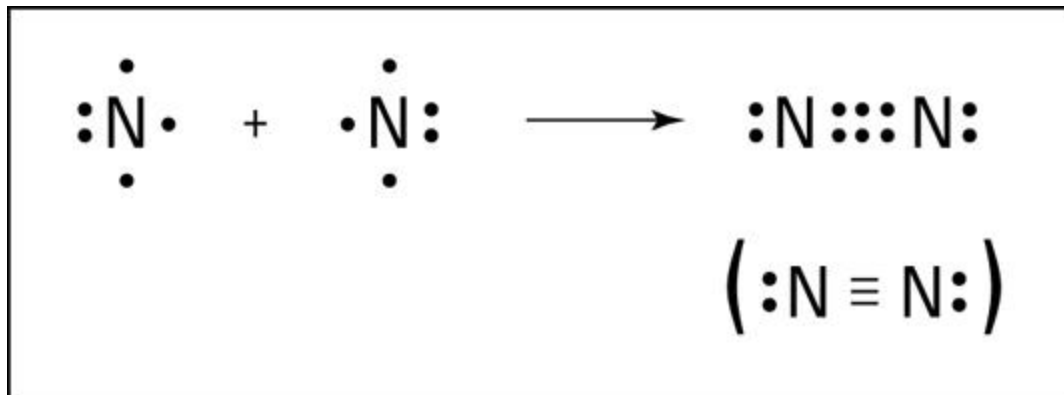


each oxygen
has 8 electrons
in the valence shell

$CO_2 :$



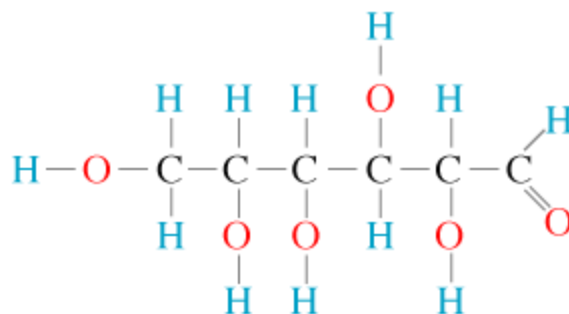
Multiple Covalent Bonds: Triple bonds



Three pairs of shared electrons

Molecular vs. Structural Formulas

- Molecular formulas - show how many atoms of each element are in the molecules
 - Ex: $C_6H_{12}O_6$ = 6 carbons, 12 hydrogens & 6 oxygens
- Structural formulas - show the 2-dimensional shape of the molecule
 - Ex:



HONC 1-2-3-4 Rule

- Hydrogen, oxygen, nitrogen & carbon are common elements found in biological molecules.
 - Hydrogen needs 1 electron to fill its “octet”
 - Oxygen needs 2 electrons to fill its octet
 - Nitrogen needs 3 electrons to fill its octet
 - Carbon needs 4 electrons to fill its octet
- “1-2-3-4” can be used to predict how these atoms will form bonds with other atoms to build molecules.

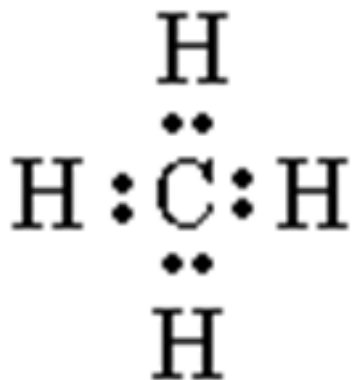
Molecular Geometry

Seeing Molecules in 3-D

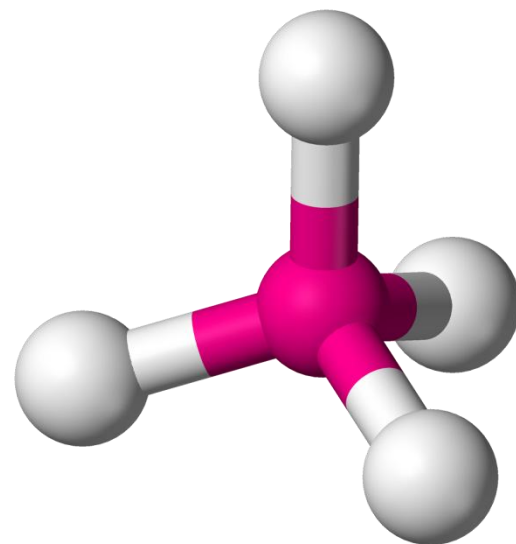
Molecular Geometry

molecules are really 3-D!

CH_4 in 2-D on a
sheet of paper



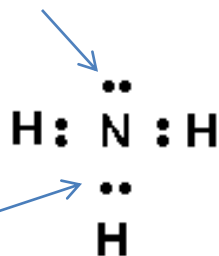
CH_4 looks like this
in 3-D



Valence Electrons determine Molecular "VSEPR" Shape

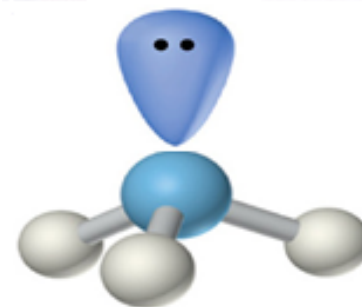
- VSEPR = "Valence-Shell Electron Pair Repulsion"
- Electron pairs (bonding or lone pairs) in a molecule repel each other and will try and get as far away from each other as possible... this determines the shape.

Lone pair electrons



bonding pair electrons

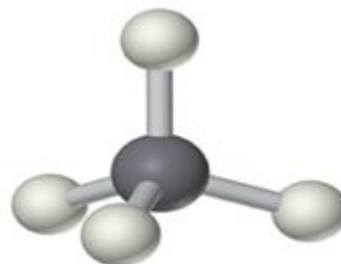
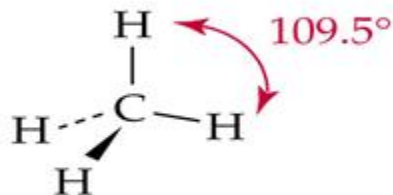
NH_3 in 2-D



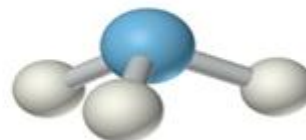
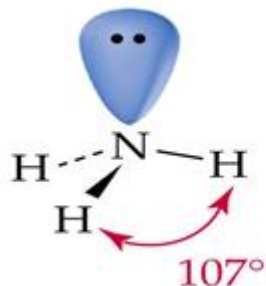
NH_3 VSEPR shape
in 3-D

4 Shapes to Know

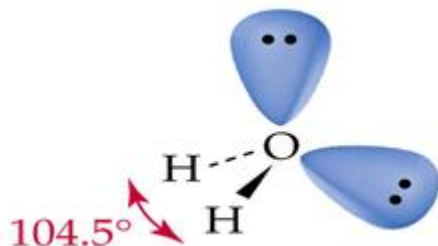
Tetrahedral



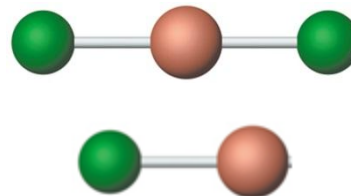
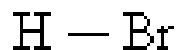
Pyramidal



Bent

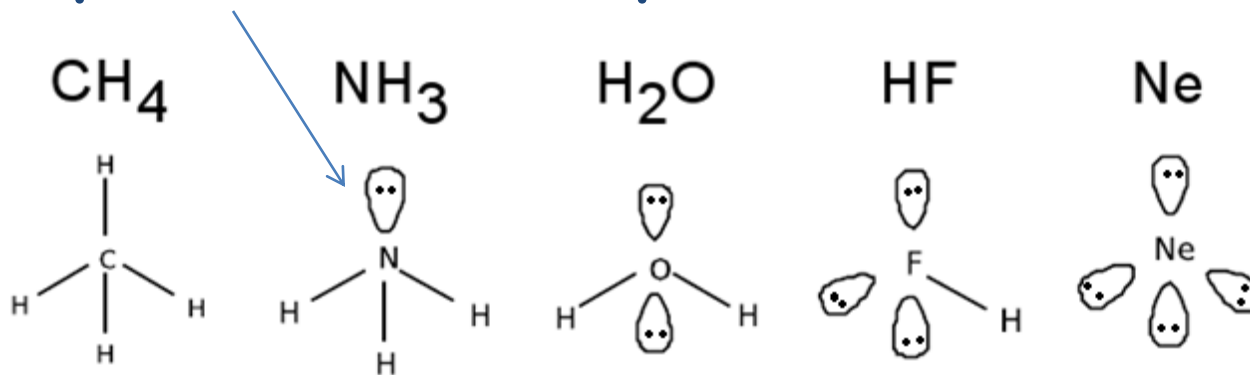


Linear

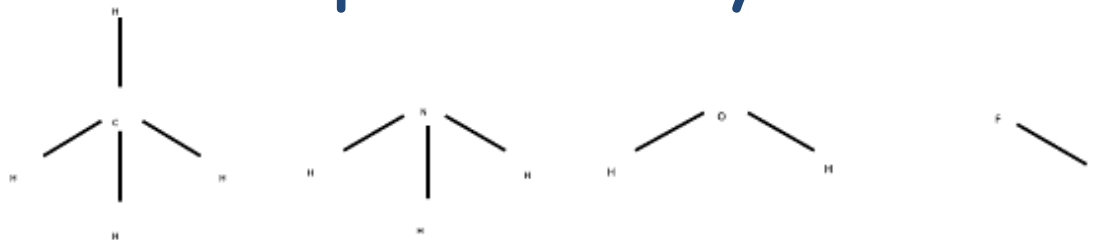


How Lone Pairs Affect Molecular Shape

“paddles” are lone pairs of electrons.



Remove the paddles and you can see the shapes.



tetrahedral

pyramidal

bent

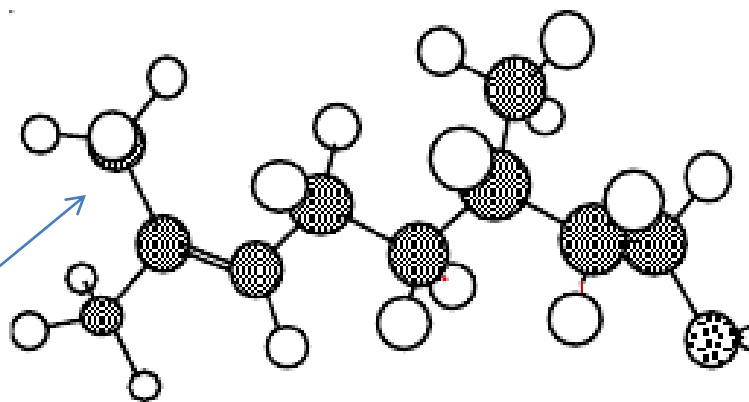
linear

Steps for Determining Molecular Geometry

1. Draw Lewis dot structure
2. Count number atoms bonded to the central atom
3. Count number of lone-pair electrons on the central atom
4. Look up the Geometry on the chart

Shapes in Large Molecules

Large molecules
are composed of
the small shapes
we've studied

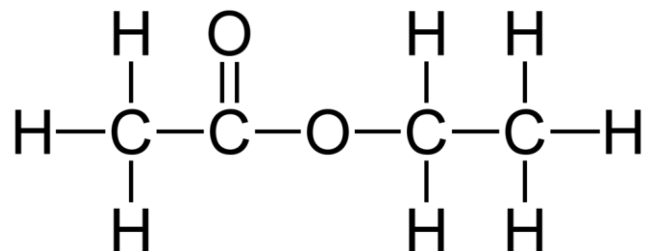


Ex: tetrahedral

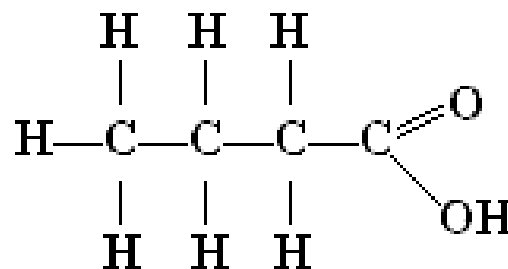
ball-and-stick
model of citronellol

Why Shape Matters

Ethyl Acetate ($C_4H_8O_2$)



Butyric Acid ($C_4H_8O_2$)



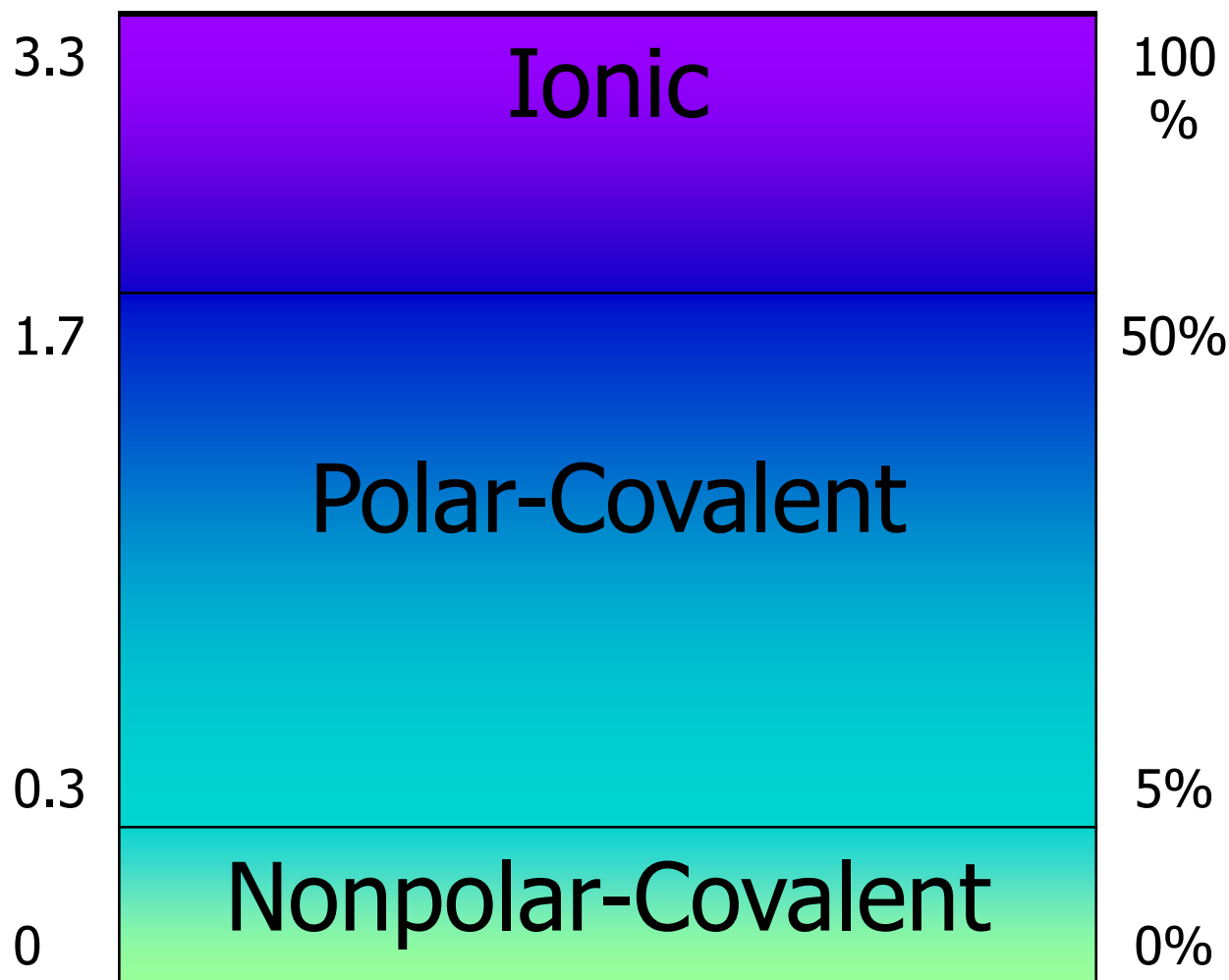
**Same formula, but different shapes
= very different smells**

Rum extract smell

Rancid butter smell

Polarity

Differences In Electronegativities

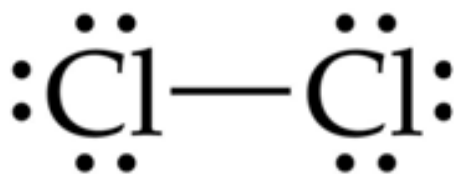


Practice Problems

Bonding Between:	Difference in Electronegativity	Bond Type
<i>Cl & Ca</i>	$3.0 - 1.0 = 2.0$	Ionic
<i>O & H</i>	$3.5 - 2.1 = 1.4$	Polar-covalent
<i>B & H</i>	$2.0 - 2.1 = 0.1$	Nonpolar-covalent

2 types of Covalent Bonds:

Non- Polar

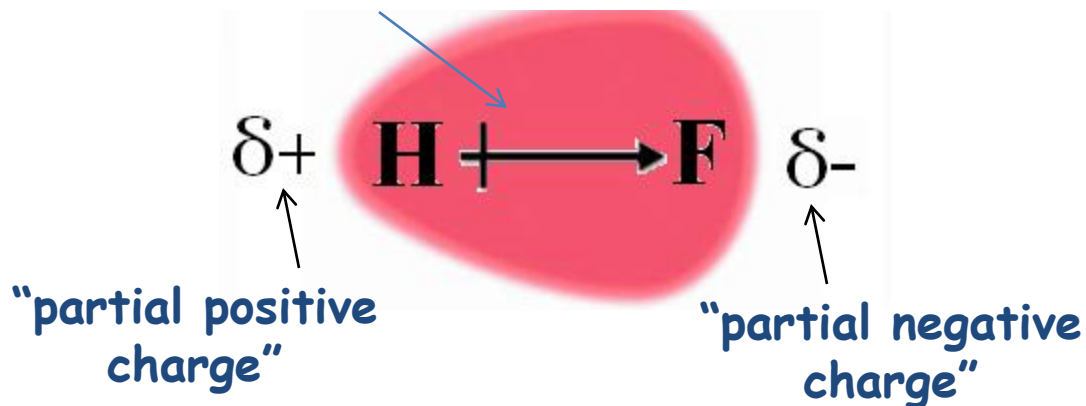


-Electrons are shared equally

- Usually the same element bonded to itself

Polar

(Arrow shows F is "pulling" electrons)

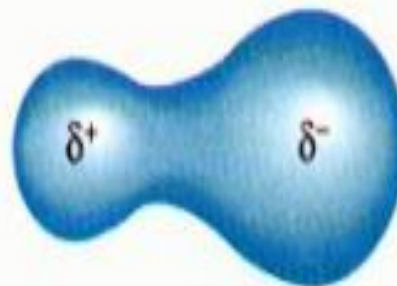


- Unequal sharing of electrons between atoms
- more electronegative atoms "hogs" electrons

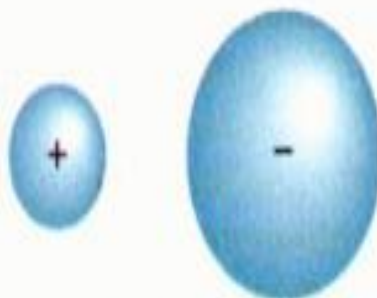
Visual Comparison of Bond Types



Nonpolar covalent bond



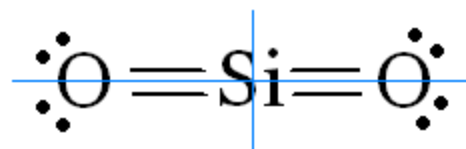
Polar covalent bond



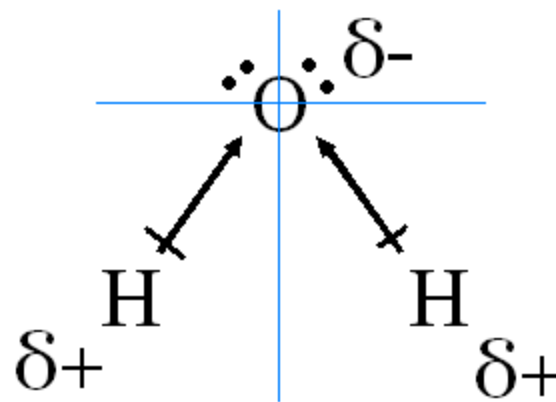
Ionic bond

Determining Polarity

1. Draw correct VSEPR Shape
2. Determine if molecule is symmetrical.
3. If the molecule is symmetrical = non-polar
 - no partial charges are needed!
4. If the molecule is NOT symmetrical = polar
 - you must show partial charges.
 - always bent or pyramidal shapes



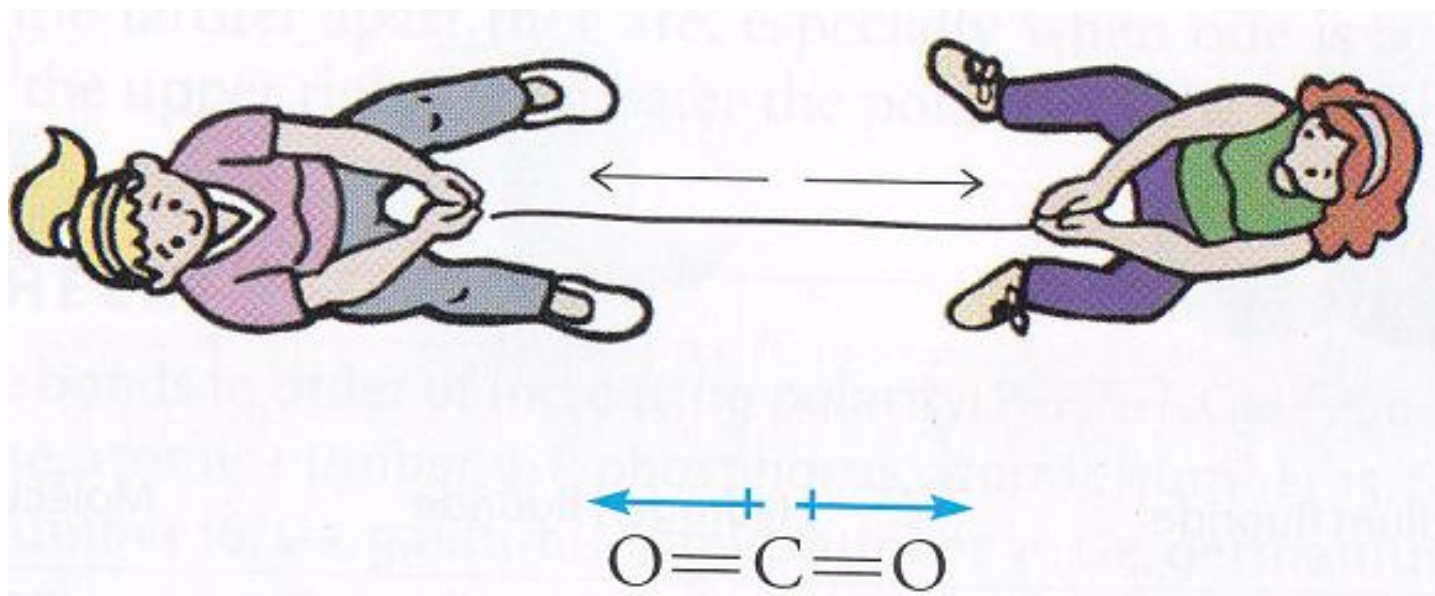
Linear shape (symmetrical)



Bent shape (assymetrical)

Ex: CO_2

- Carbon dioxide = nonpolar
- has polar bonds, but they cancel each other out.



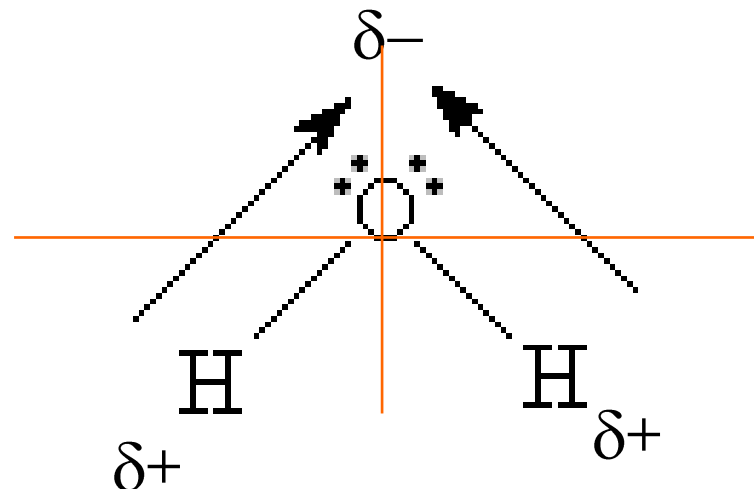
EX: Water= Polar Molecule

How we know:

1) Cut the molecule on 2 planes

- see how it's different above the horizontal line = non symmetrical

1) One atom is "pulling", look at periodic table to determine which one.



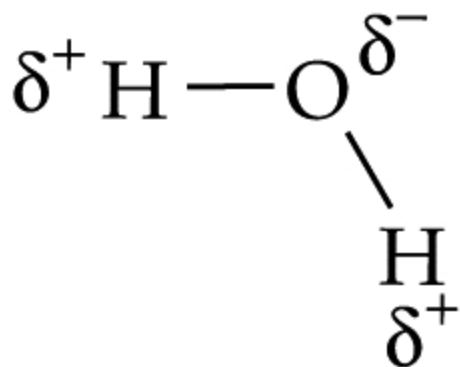
Indicates which atom "pulls" the electrons

$\delta-$

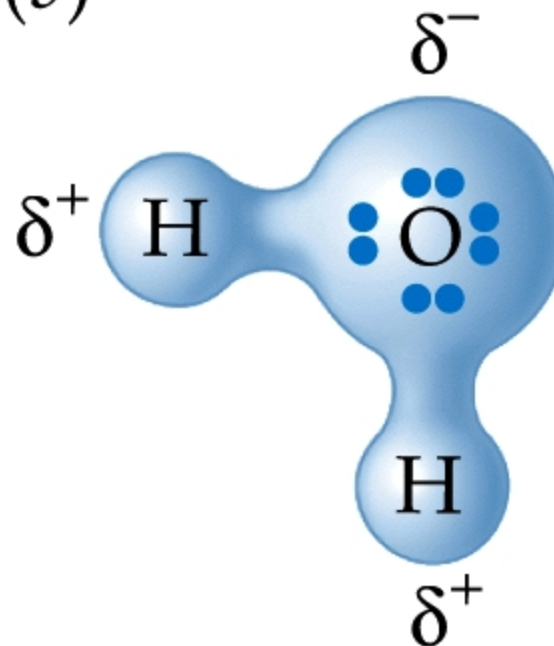
Means oxygen is slightly negative because it "hogs" electrons

2 views of Polar water

(a)



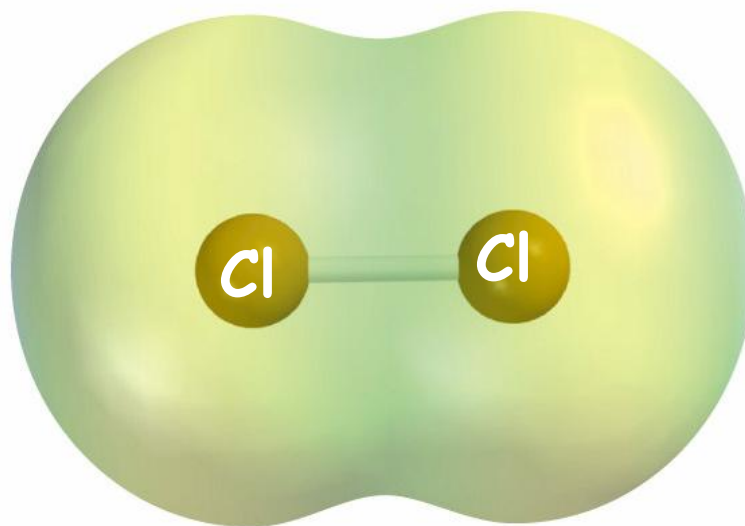
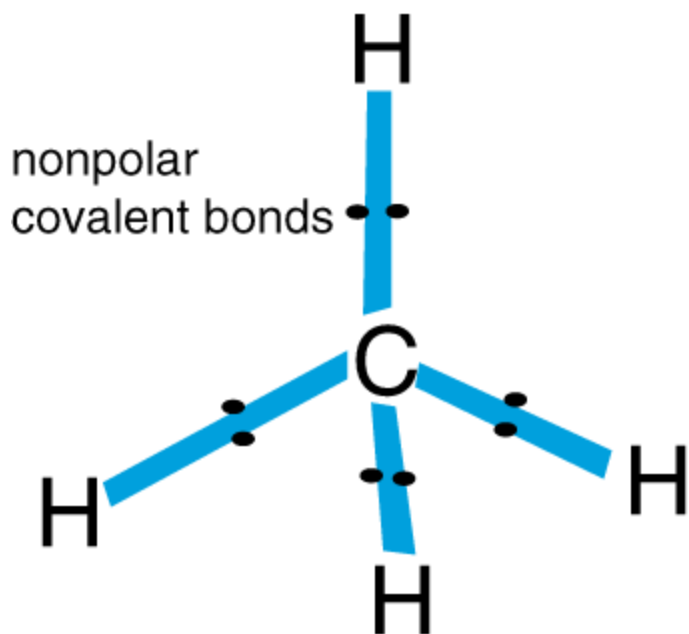
(b)



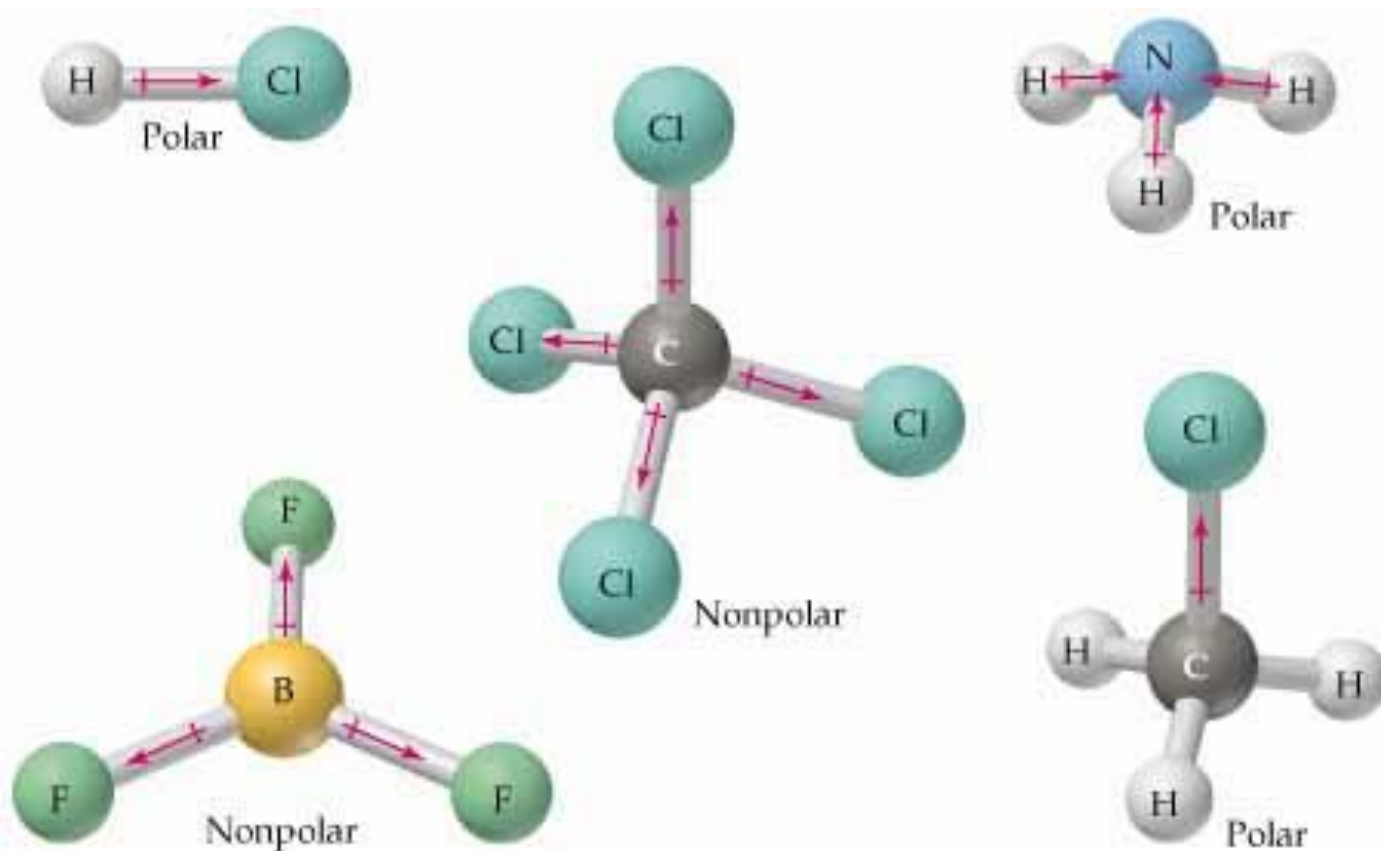
© 2001 Sinauer Associates, Inc.

Non Polar Molecules

- Non Polar molecule= "no pull"
 - equal sharing of electrons
 - No difference in electronegativity
 - symmetrical in shape



Examples of Polar & Nonpolar Molecules



Electronegativity



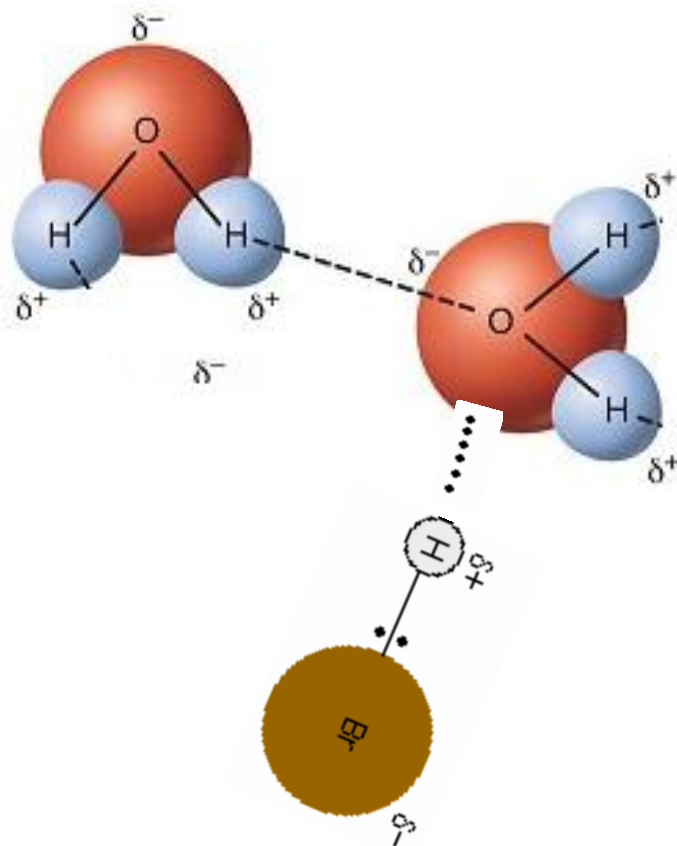
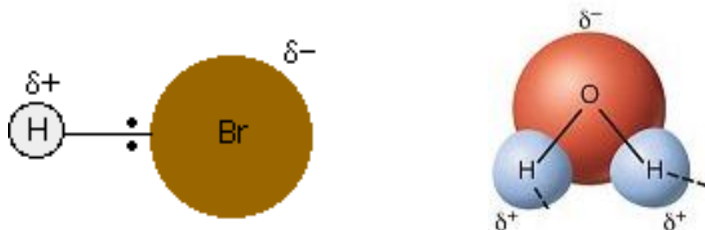
1	2											3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)
H 2.1																	He --
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne --
Na 0.9	Mg 1.3	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	Al 1.6	Si 1.9	P 2.2	S 2.5	Cl 3.0	Ar --
K 0.8	Ca 1.3	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.7	Ga 1.6	Ge 2.0	As 2.2	Se 2.6	Br 2.8	Kr --
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 2.1	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe 2.6
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 2.0	Pb 2.3	Bi 2.0	Po 2.0	At 2.2	Rn --
Fr 0.7	Ra 0.9	Ac 1.1	Rf --	Db --	Sg --	Bh --	Hs --	Mt --	Uun --	Uuu --	Uub --		Uuq				

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Inter vs. Intra molecular Forces

Why Polarity Matters: Molecular Attractions

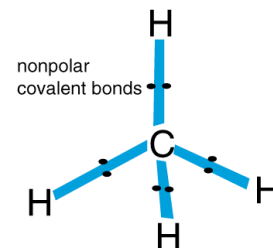
- 1 molecule can be attracted to another molecule
 - “inter”molecular force
- You can predict how one molecule might react with another: Ex: $\text{HBr} + \text{H}_2\text{O}$



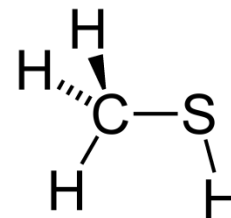
Intermolecular Attractions & Smell

- Besides shape, polarity also plays a role in your ability to smell.
 - Polar molecules = smell
 - Non-polar = don't smell
- Your smell receptors are polar and surrounded by mucous (a watery substance)

Ex: Methane gas is odorless



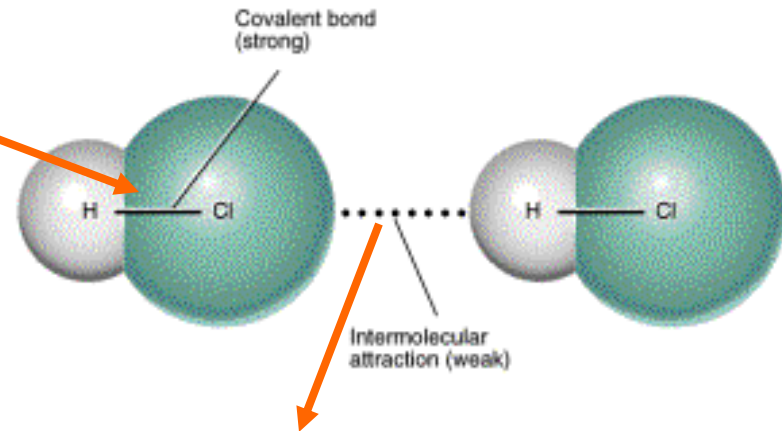
-They add a this stinky chemical to it so that you can smell it it:



Intermolecular Forces vs. Intramolecular Forces

Intramolecular Forces:
(within in a molecule)

Ex: -Covalent bond
- Ionic bond
-Metallic bond

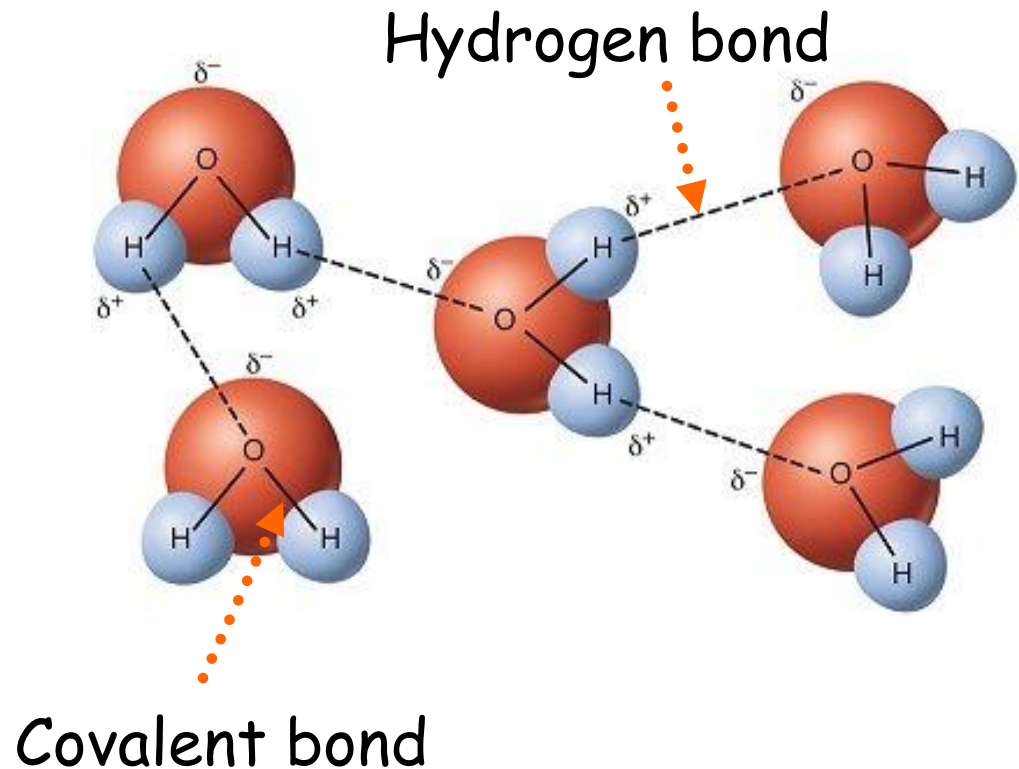


Intermolecular Forces:
(between molecules)

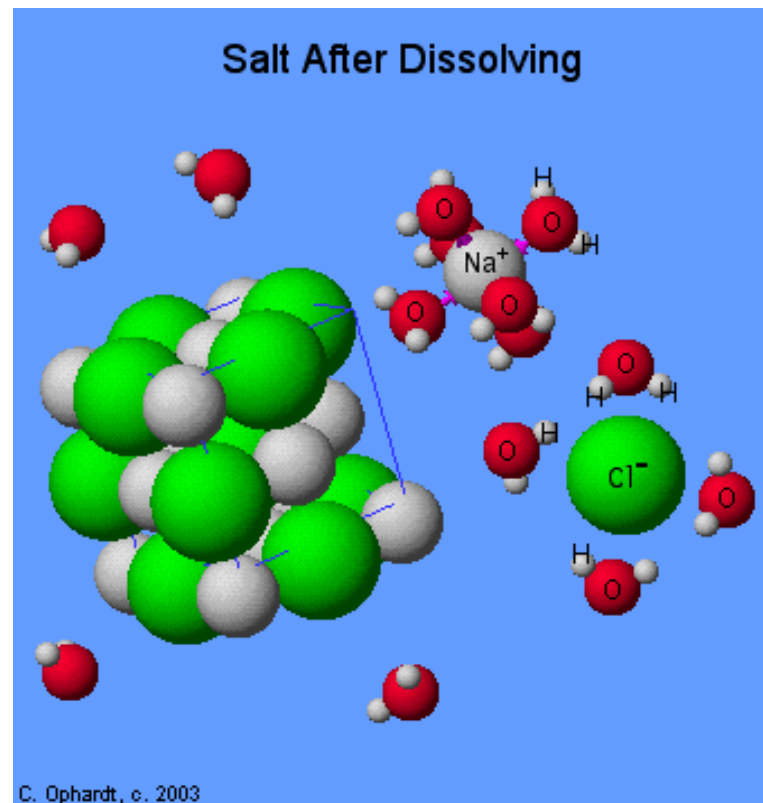
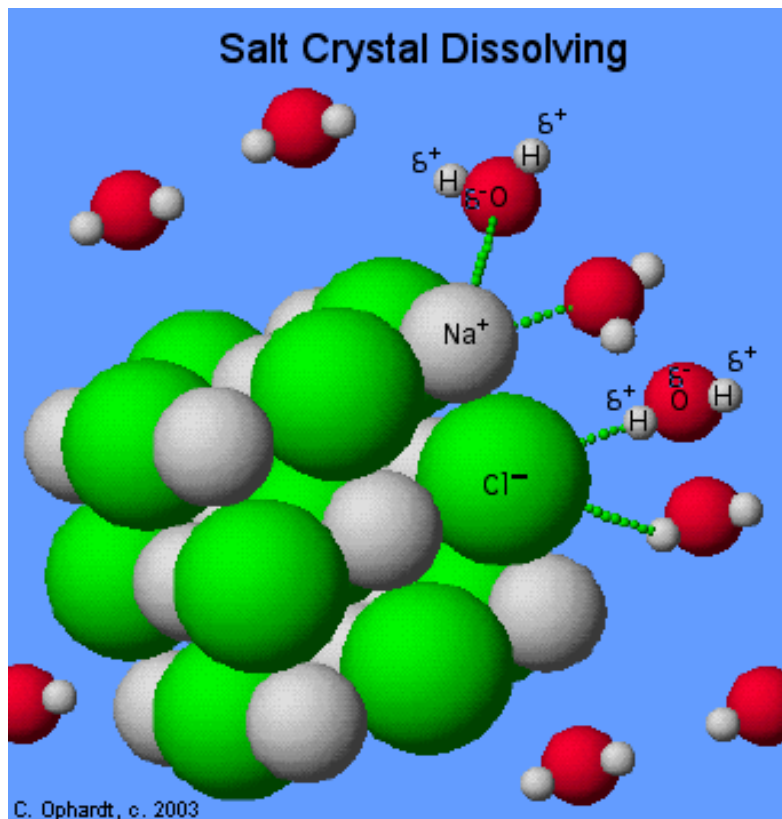
Ex: Hydrogen bonds
-Weaker than covalent,
ionic, & metallic bonds

Hydrogen Bonding- (an **inter**molecular force) in Water

- Water is polar
 - (has a + and - end)
 - It's "sticky"
- Will stick to any other thing that is:
 - polar (ex: other water molecules)
 - charged ionic substances (NaCl)



Water's Polarity leads to its ability to dissolve things so well



The slight charges on water attract the NaCl's ions and cause them to separate from each other

Unique Properties of Water due to polarity & hydrogen bonding

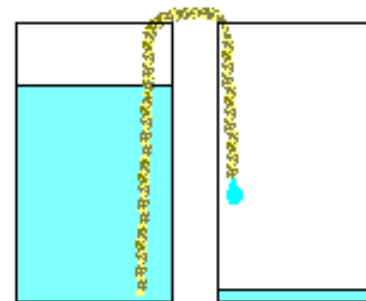
1) Surface tension (hydrogen bonds create surface on water)



3) Adhesion/ Cohesion (water is attracted to other water molecules)

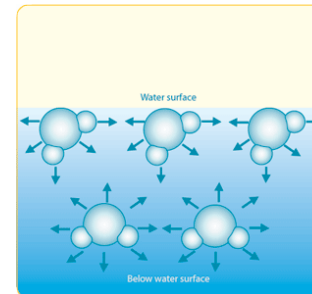


4) Capillary action water is attracted to other water molecules and will “rise”

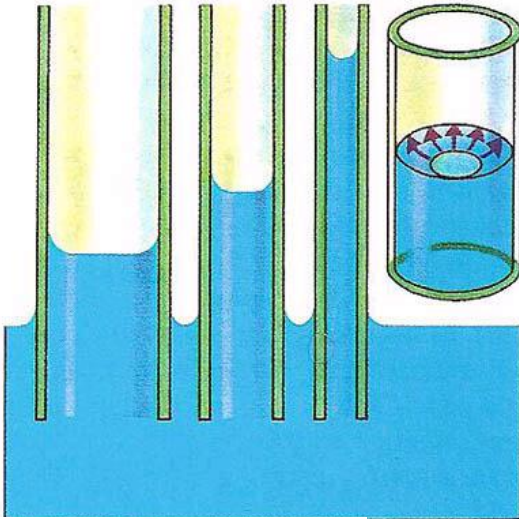


Properties of Water due to Hydrogen Bonding & Polarity

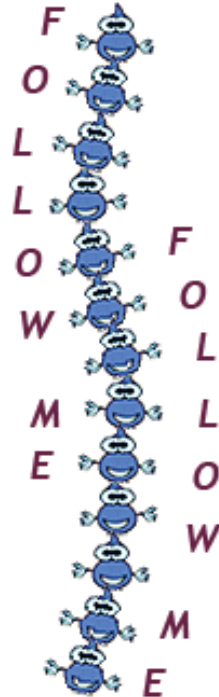
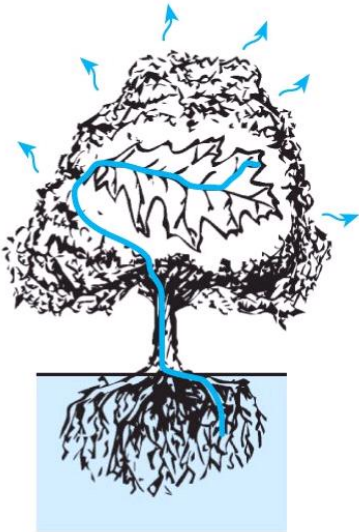
- **Cohesion** - water molecules are attracted to one another
 - Causes water to be "Sticky"
 - This is why water forms droplets
- **Adhesion** - water is attracted to other substances
 - Water will "stick" to containers & objects
- **Surface tension** - strong forces between molecules cause the surface of a liquid to contract



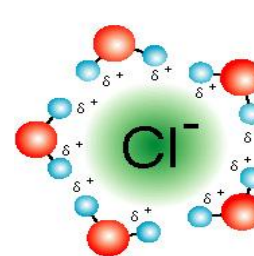
More properties...



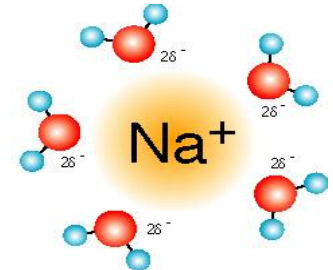
- **Capillary Action** - the movement of water within the spaces of a porous material due to the forces of adhesion, cohesion, and surface tension.



- **Universal Solvent** things dissolve in water- polarity)



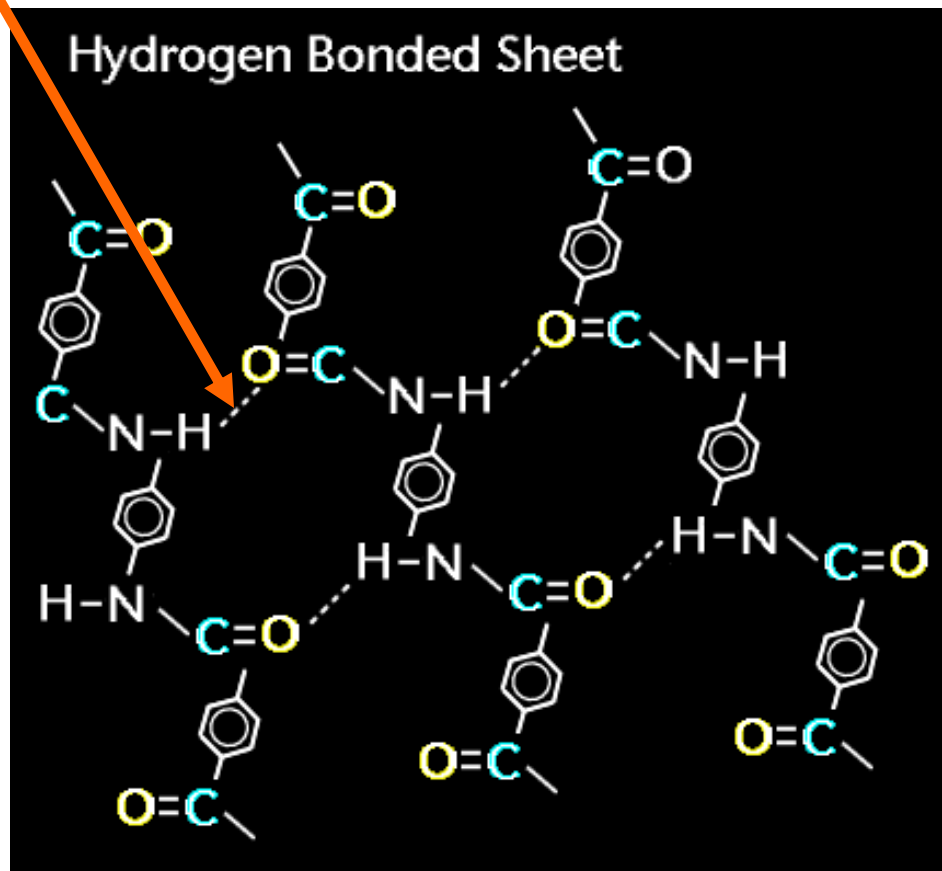
Slightly positive hydrogen are attracted to chlorine anions



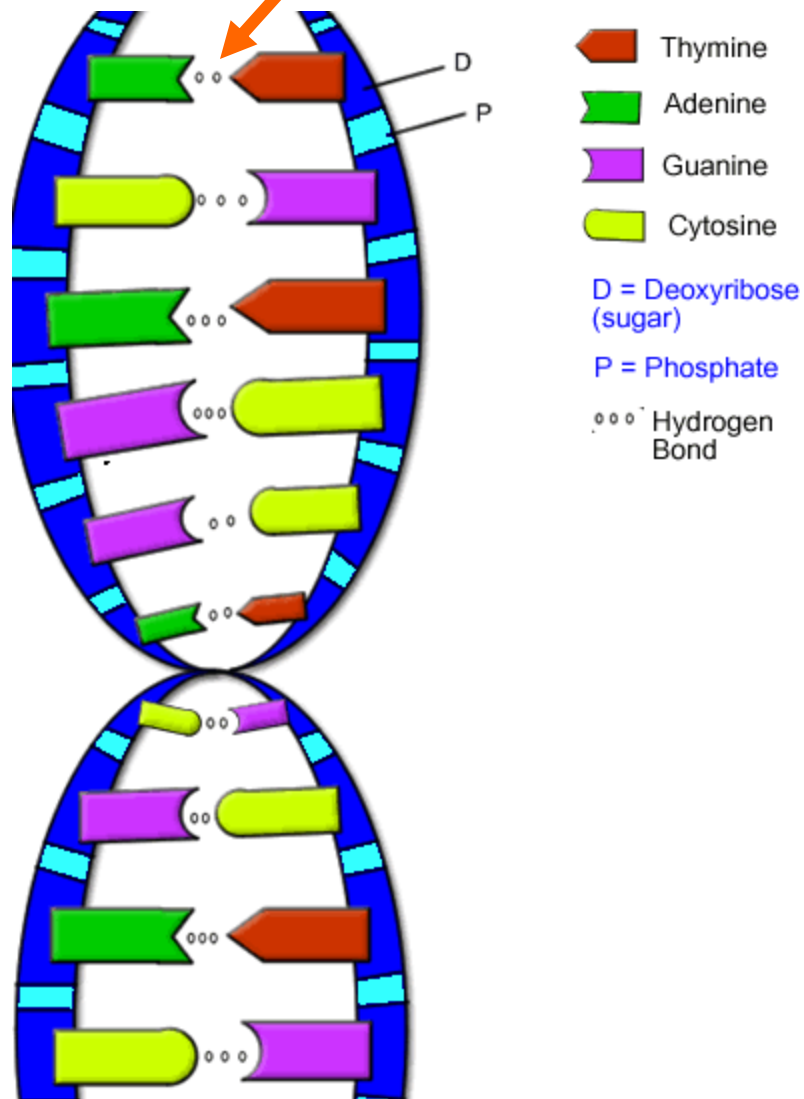
Slightly negative oxygen are attracted to sodium cations

Hydrogen Bonding in Kevlar

Hydrogen bonding in Kevlar, a strong polymer used in bullet-proof vests.

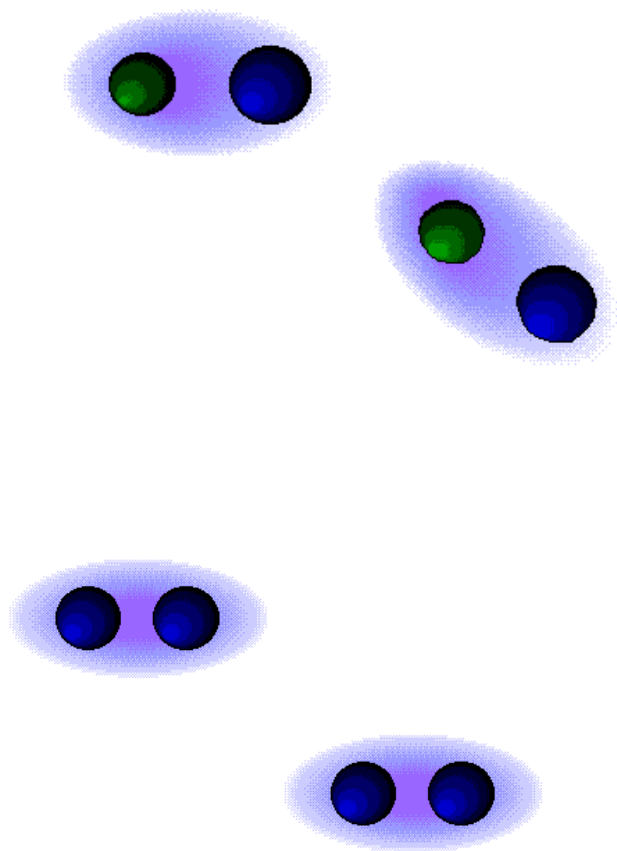


Hydrogen Bonding in DNA



Other Intermolecular Forces (FYI... not part of this class)

- Van der Waals Forces include:
 - **Dipole-Dipole forces** - results from the tendency of polar molecules to align themselves so that the positive end of one molecule is near the negative end of another molecule.
 - **London (Dispersion) forces** - results from the small, instantaneous dipoles that occur because of the varying positions of the electrons during their motion about nuclei



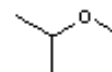
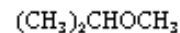
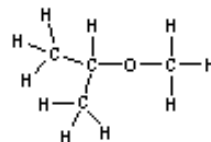
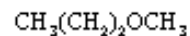
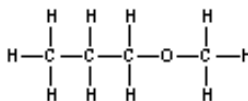
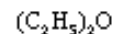
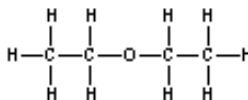
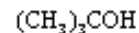
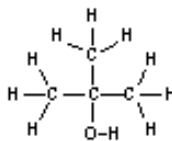
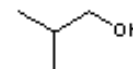
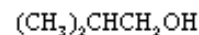
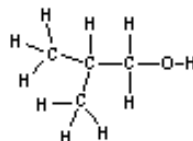
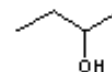
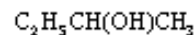
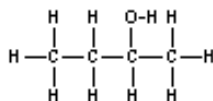
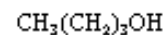
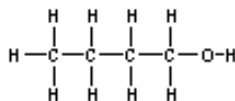
Organic Chemistry

Organic Chemistry- shows the versatility of carbon

- has 4 valence electrons = 4 bonding spaces available.

- Backbone to many large, complex biological molecules (Carbs, Lipids, Proteins, Nucleic Acids)

- Over 16 million carbon-containing compounds are known.

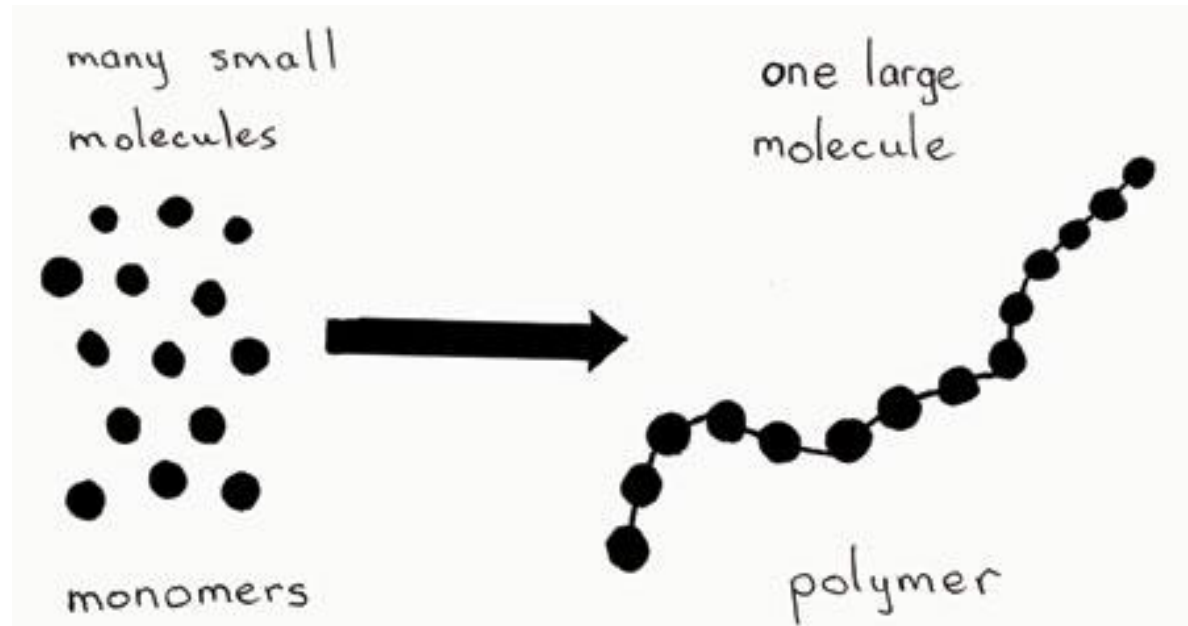


Monomers combine to make Polymers (small unit) (large)

Common Examples of Polymers:

- Carbohydrates
- Lipids
- Proteins
- Nucleic Acids

(CLPN)



Ex: Carbohydrates

Monomer

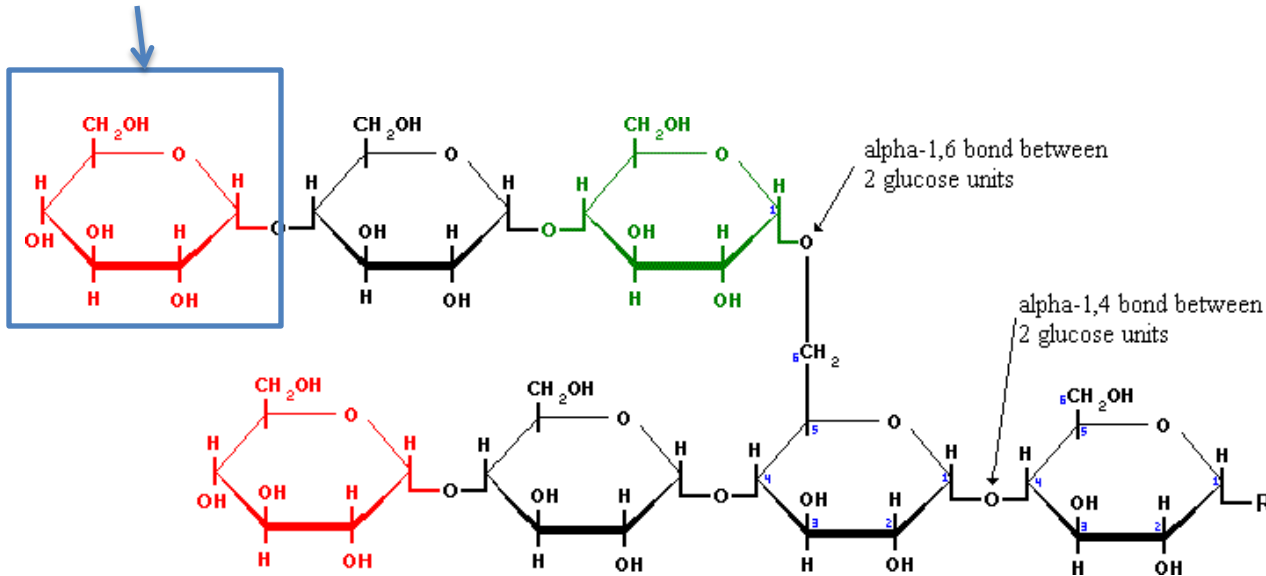
Polymer

Examples

Monosaccharide

Polysaccharide

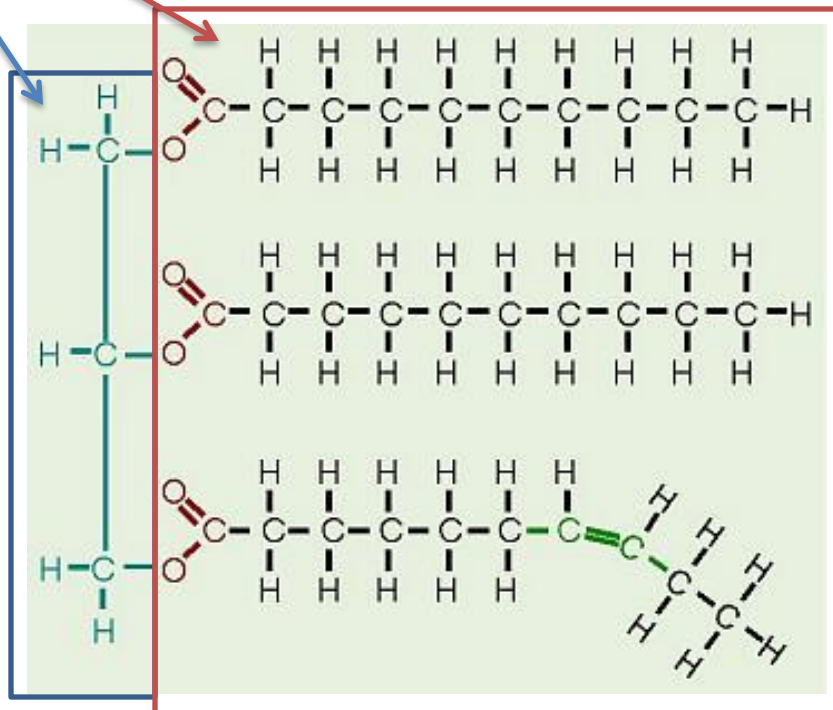
- Starch
- Fiber
- sucrose



Ex: Lipids

Monomers

Glycerol
& Fatty Acid
tails



Polymer

Tri-glyceride

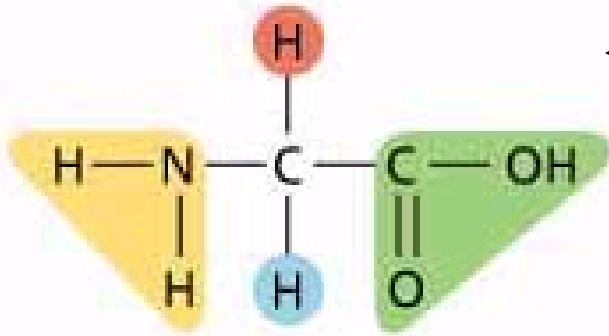
Examples

- Saturated Fats
- Unsaturated fats
- Steroids
- Cholesterol

Ex: Proteins

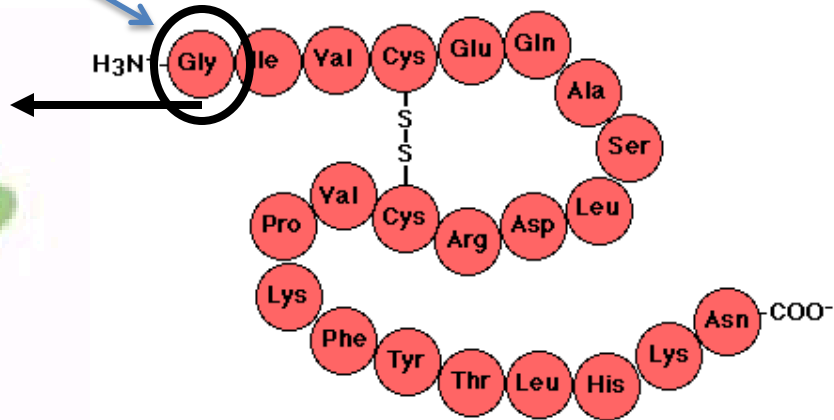
Monomer

Amino Acids



Polymer

Polypeptide



Examples

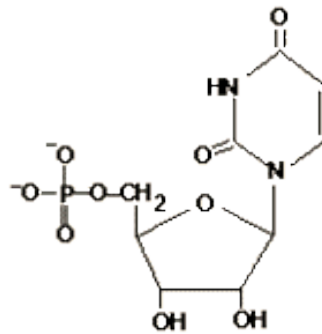
- enzymes
- pigments
- Meat/dairy



Ex: Nucleic Acids

Monomer

Nucleotide



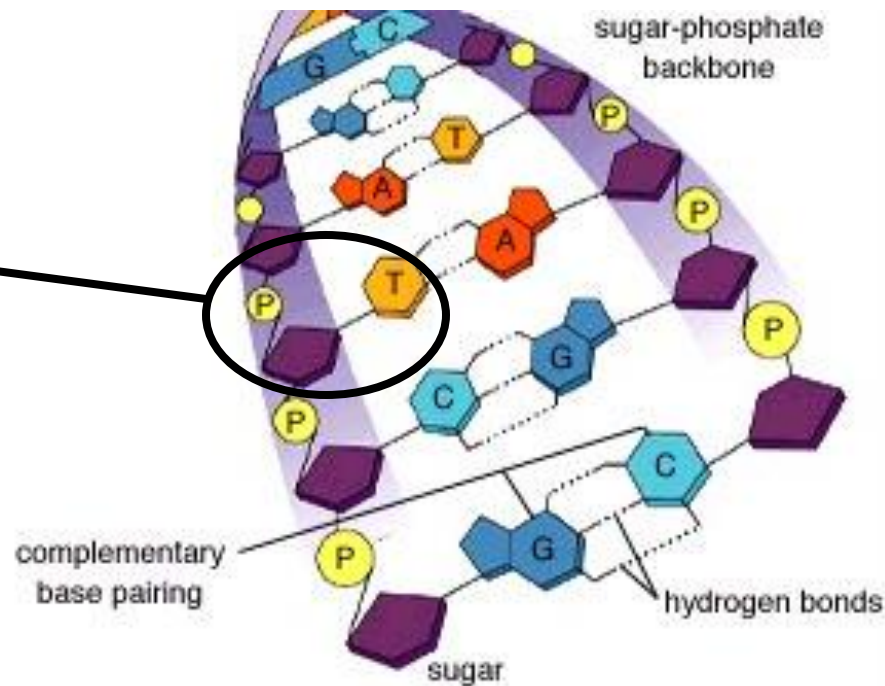
Polymer

Polynucleotide

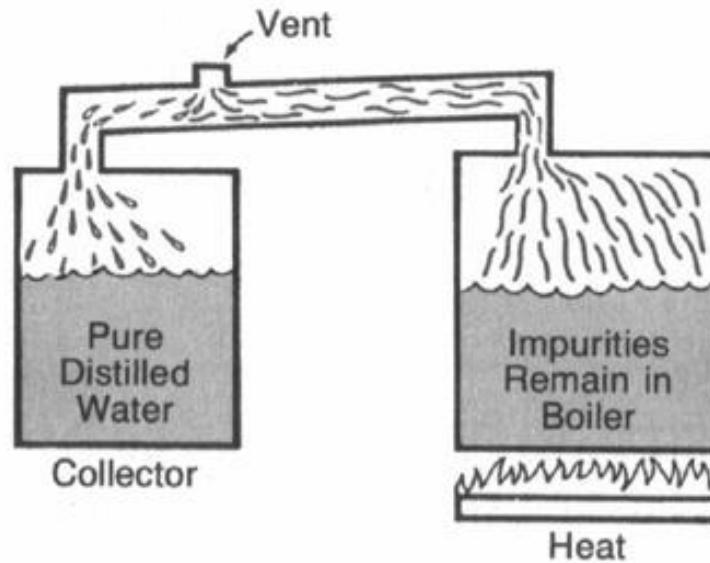
Examples

-DNA

-RNA



Distilled Water vs. Tap Water



Water Poisoning/ water Intoxication

Cause: excessive consumption of water during a short period of time.

Why: leads to a disruption in normal brain function due to the imbalance of electrolytes in the body's fluids.

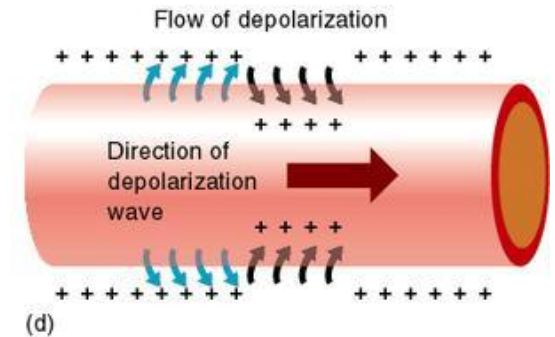
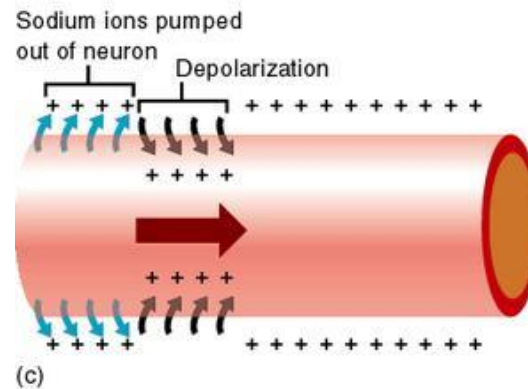
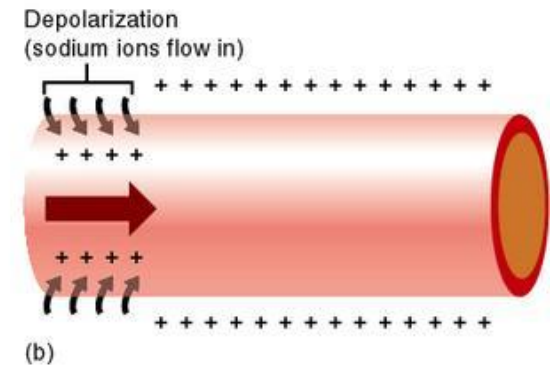
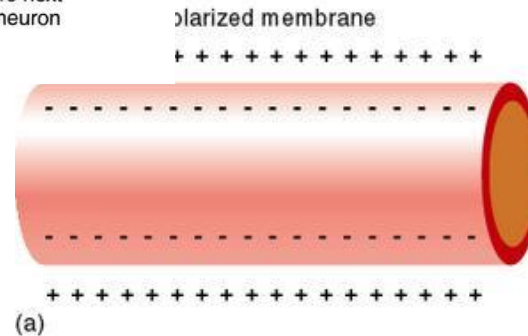
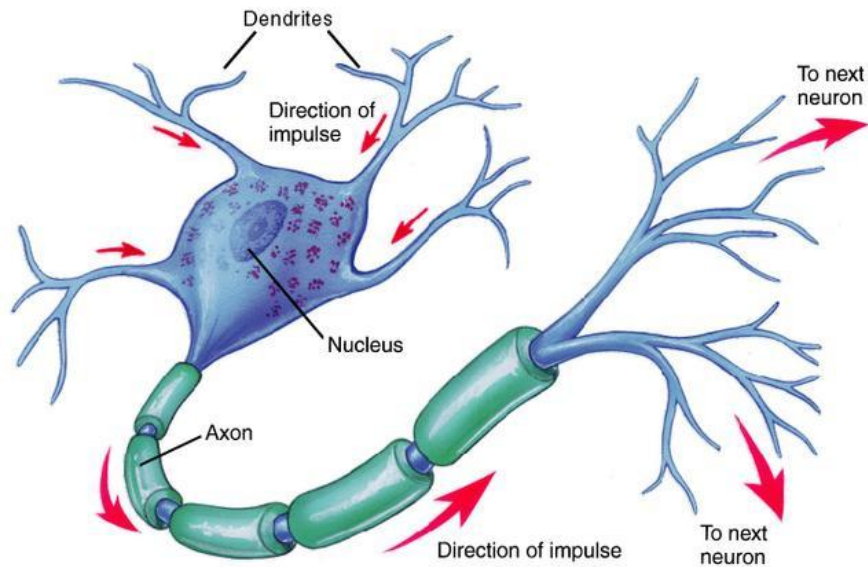
- can dilute the careful balance of sodium compounds in the body fluids

Who: individuals in water drinking contests...consume more than 10 liters (10.5 quarts) of water over the course of just a few minutes

- People doing endurance sports which electrolytes are not properly replenished, yet massive amounts of fluid are still consumed



Neural transmission



Electric Stimulation Machine- stimulates muscles for you



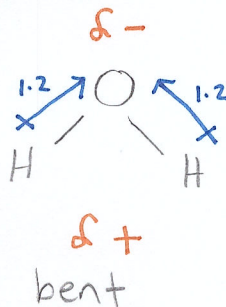
See video clips
on web links

Polar Bonds vs. Polar Molecules

- We already learned covalent bonds can be polar or nonpolar based on the electronegativity of the atoms in the molecule
- We now need to look at all the individual covalent bonds in a molecule to determine if the entire molecule will be polar or not
- To determine if a molecule is polar or nonpolar follow these steps
 1. Draw the VSEPR shape of the molecule
 2. Use the electronegativities of the atoms in the molecule to determine if any bond dipoles exist. Draw the bond dipole arrows onto the VSEPR drawing
 3. "Add" up the bond dipole arrows to see if they all cancel out or not to determine if the entire molecule is polar or nonpolar.
 - If the bond dipoles arrows generally point in the same direction, the entire molecule has a side that is overall slightly positive or negative and is called a polar molecule
 - If the bond dipoles arrows point in opposite directions and are equal in magnitude/strength, the molecule is called nonpolar. In order for polar bonds to cancel each other out and result in a nonpolar molecule, the molecule needs to be symmetrical.

- EXAMPLE:** Show that H_2O is a polar molecule.

$$\begin{array}{r} \text{H} = 1e^- \times 2 \\ \text{O} = 6e^- \\ \hline = 8e^- \end{array}$$

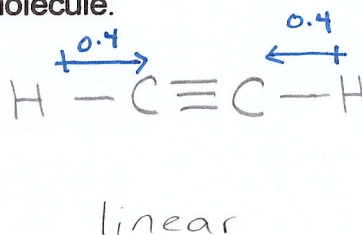


$$\begin{array}{c} \xleftarrow{1.2} + \\ \text{O} - \text{H} \\ \Delta \text{EN} = 3.4 - 2.2 \\ \Delta \text{EN} = 1.2 \end{array}$$

$\therefore$ polar

- EXAMPLE:** Show that C_2H_2 is a nonpolar molecule.

$$\begin{array}{r} \text{C} = 4e^- \times 2 \\ \text{H} = 1e^- \times 2 \\ \hline = 10e^- \end{array}$$



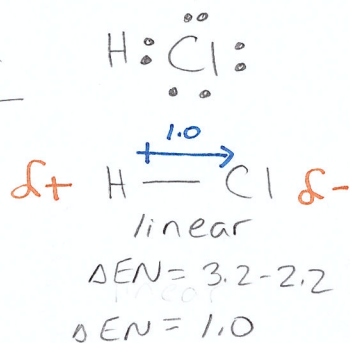
$$\begin{array}{c} \xleftarrow{0.4} + \\ \text{C} - \text{H} \\ \Delta \text{EN} = 2.6 - 2.2 \\ \Delta \text{EN} = 0.4 \end{array}$$

$\therefore$ non polar

- EXAMPLES:** Draw the Lewis structure and the VSEPR shape for each molecule and then determine if the following molecules are polar or nonpolar.

1. HCl = polar

$$\begin{array}{l} \text{H} = 1e^- \\ \text{Cl} = 7e^- \\ \hline = 8e^- \end{array}$$



∴ polar

4. N_2 = non polar

$$\begin{array}{l} \text{N} = 5e^- \times 2 \\ \hline = 10e^- \end{array}$$



linear

$$\Delta\text{EN} = 3.0 - 3.0$$

$$\Delta\text{EN} = 0$$

∴ no bond dipole

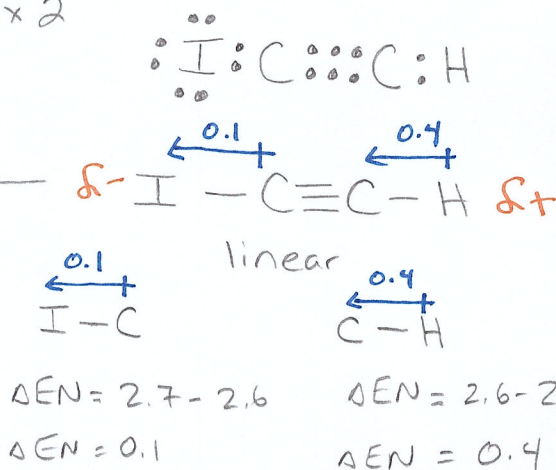
2. C_2HI = polar

$$\text{C} = 4e^- \times 2$$

$$\text{H} = 1e^-$$

$$\text{I} = 7e^-$$

$$\hline = 16e^-$$



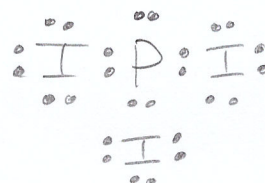
∴ polar

5. PI_3 = polar

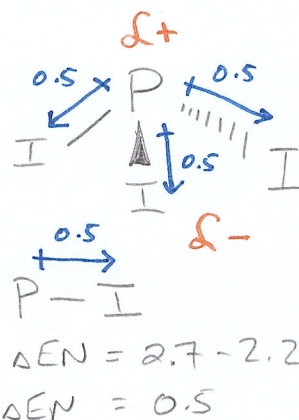
$$\text{P} = 5e^-$$

$$\text{I} = 7e^- \times 3$$

$$\hline = 26e^-$$



trigonal pyramidal

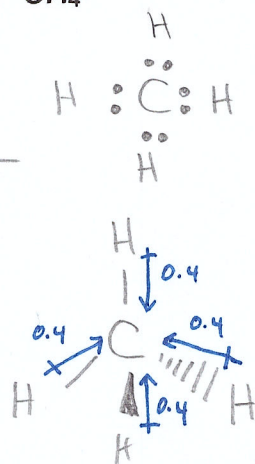


3. CH_4 = non polar

$$\text{C} = 4e^-$$

$$\text{H} = 1e^- \times 4$$

$$\hline = 8e^-$$



∴ non polar

tetrahedral

Now Try Practice Problems

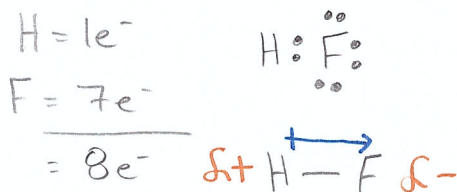
- Summary Chart on Polar vs. Nonpolar Molecules

Shape	VSEPR Diagram	Polar (P) or Non-polar (NP)
Tetrahedral		NP if all atoms same P if atoms attached to central atom are different
Trigonal planar		NP if all atoms attached to central atom are the same P if atoms attached to central atom are different
Pyramidal		Always P
Bent		Always P
Linear		NP if symmetrical P if non symmetrical

Practice Problems

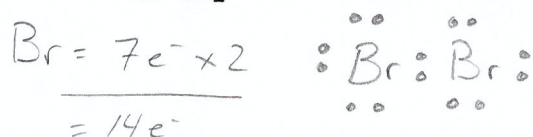
Draw the VSEPR shape for each molecule and determine if the molecule will be polar or nonpolar.

1. HF = polar



$$\begin{aligned} \Delta EN &= 4.0 - 2.2 \\ \Delta EN &= 1.8 \end{aligned}$$

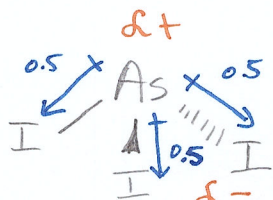
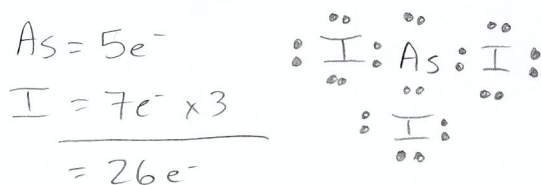
2. Br₂ = non polar



linear

$$\begin{aligned} \Delta EN &= 3.0 - 3.0 \\ \Delta EN &= 0 \\ \therefore \text{no bond dipole} \end{aligned}$$

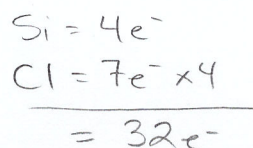
3. AsI₃ = polar



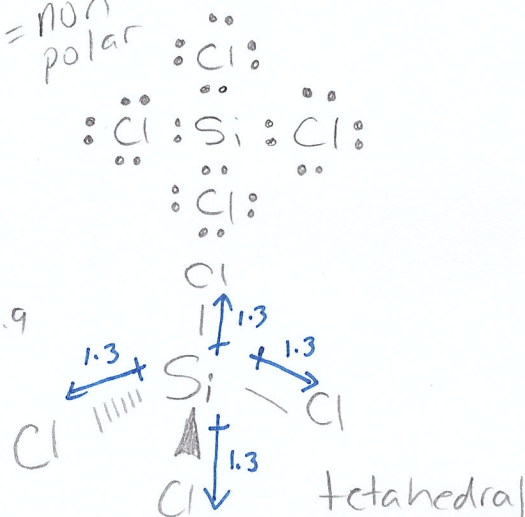
trigonal pyramidal

$$\begin{aligned} \Delta EN &= 2.7 - 2.2 \\ \Delta EN &= 0.5 \end{aligned}$$

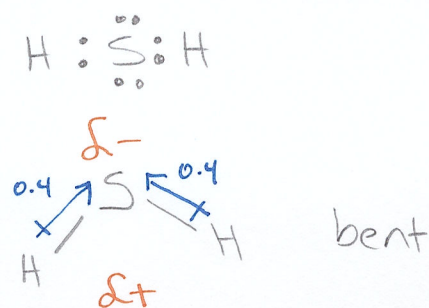
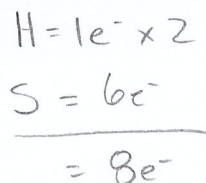
4. SiCl₄ = non polar



$$\begin{aligned} \Delta EN &= 3.2 - 1.9 \\ \Delta EN &= 1.3 \end{aligned}$$

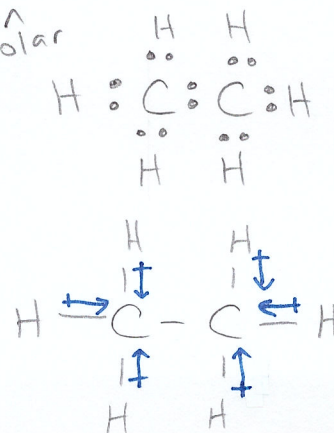
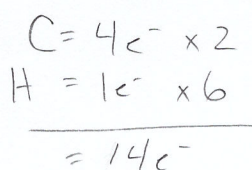


5. H₂S = polar



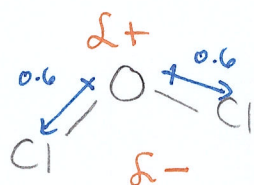
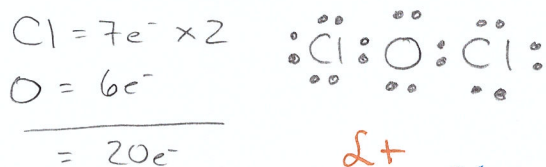
$$\begin{aligned} \Delta EN &= 2.6 - 2.2 \\ \Delta EN &= 0.4 \end{aligned}$$

6. C₂H₆ = non polar



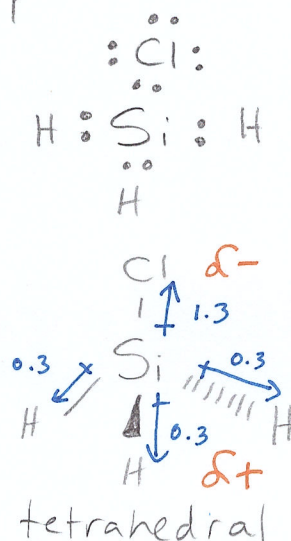
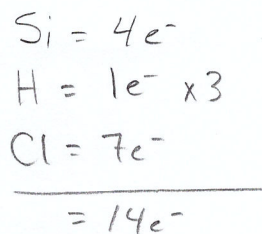
$$\begin{aligned} \Delta EN &= 2.6 - 2.2 \\ \Delta EN &= 0.4 \end{aligned}$$

7. Cl_2O = polar



bent
 $\Delta\text{EN} = 3.2 - 2.6$
 $\Delta\text{EN} = 0.6$

10. SiH_3Cl = polar

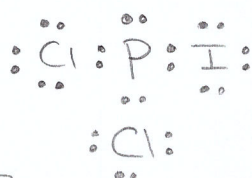
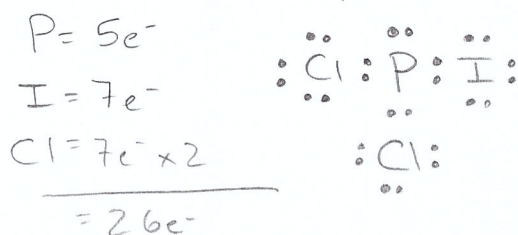


$\text{Si} - \text{Cl}$
 $\Delta\text{EN} = 3.2 - 1.9$
 $\Delta\text{EN} = 1.3$

$\text{Si} - \text{H}$
 $\Delta\text{EN} = 2.2 - 1.9$
 $\Delta\text{EN} = 0.3$

tetrahedral

8. PCl_2I = polar

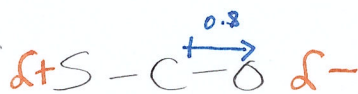
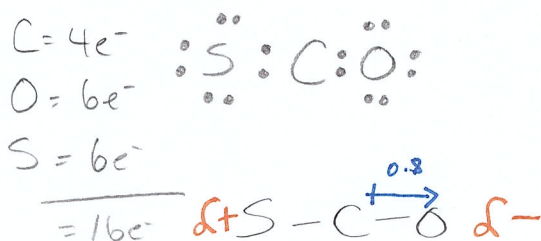


$\text{P} - \text{I}$
 $\Delta\text{EN} = 2.7 - 2.2$
 $\Delta\text{EN} = 0.5$

trigonal pyramidal

$\text{P} - \text{Cl}$
 $\Delta\text{EN} = 3.2 - 2.2$
 $\Delta\text{EN} = 1.0$

9. COS = polar

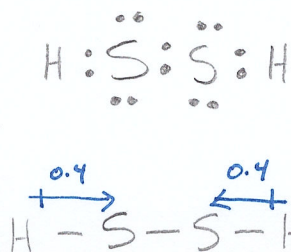
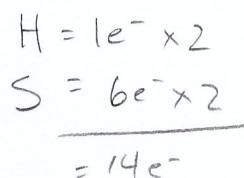


linear

$\text{S} - \text{C}$
 $\Delta\text{EN} = 2.6 - 2.6$
 $\Delta\text{EN} = 0$

$\text{C} - \text{O}$
 $\Delta\text{EN} = 3.4 - 2.6$
 $\Delta\text{EN} = 0.8$

11. H_2S_2 = non polar



linear

$\Delta\text{EN} = 2.6 - 2.2$
 $\Delta\text{EN} = 0.4$



KWAZULU-NATAL PROVINCE

EDUCATION
REPUBLIC OF SOUTH AFRICA

CURRICULUM GRADE 10 -12 DIRECTORATE

NCS (CAPS)

LEARNER SUPPORT DOCUMENT

GRADE 11

**PHYSICAL SCIENCES
REVISION DOCUMENT**

2025

PREFACE

This support document serves to assist Physical Sciences learners on how to deal with curriculum gaps and learning losses. It addresses all the topics in the Grade 11 curriculum.

Activities serve as a guide on how various topics are assessed at different cognitive levels and to prepare learners for informal and formal tasks in Physical Sciences. It covers the following topics:

No.	Topic	Page
1.	Vectors in 2 dimensions	3 – 10
2.	Newton's Laws	10 – 18
3.	Electrostatics	19 – 29
4.	Electromagnetism	30 – 36
5.	Electrodynamics	36 – 39
5.	Electric Circuits	39 – 45
6.	Atomic combination	46 – 52
7.	Intermolecular forces	53 – 58
8.	Quantitative Aspects of Chemical Change	58 – 70
9.	Energy and chemical change	70 – 75
10.	Acids and Bases	75 – 82
11.	Ideal Gases	82 – 89
12.	Redox Reactions	89 - 94

DEFINITIONS

- Chemical bond – a mutual attraction between two atoms resulting from the simultaneous attraction between their nuclei and the outer electrons.
- Intramolecular bond – bond which occurs between atoms within molecules
- Lewis dot diagram – a structural formula in which valence electrons are represented by dots or crosses.
- Valence electrons – the electrons in the highest energy level of an atom in which there are electrons.
- Covalent bond – the sharing of electrons between two atoms to form a molecule.
- Ionic bond – a transfer of electrons and subsequent electrostatic attraction
- Metallic bond – metallic bonding as the bond between positive ions and delocalized valence electrons in a metal
- Molecule – a group of two or more atoms covalently bonded and that function as a unit.
- Bonding pair – a pair of electrons that is shared between two atoms in a covalent bond.
- Lone pair – a pair of electrons in the valence shell of an atom that is not shared with another atom.
- Electronegativity – a measure of the tendency of an atom in a molecule to attract bonding electrons.
- Non-polar covalent bond – a bond in which the electron density is shared equally between the two atoms.
- Polar covalent bond – a bond in which the electron density is shared unequally between the two atoms.
- Bond energy of a compound – the energy needed to break one mole of its molecules into separate atoms.
- Bond length – the average distance between the nuclei of two bonded atoms

MOLECULAR STRUCTURE

- The structure mainly depends on the type of chemical bond (force) that exists between the atoms that the molecule consists of.

RULES FOR BOND FORMATION

Different atoms:

- each with an unpaired valence electron, can share these electrons to form a chemical bond
example: two H atoms form a H_2 molecule by sharing an electron pair
- with paired valence electrons, called lone pairs, cannot share these 4 electrons and cannot form a chemical bond
example: no bond forms between 2 He atoms
- with unpaired valence electrons, can share these electrons and form a chemical bond for each electron pair shared (multiple bonds can be formed)
example: two O atoms form an O_2 molecule

Atoms with an empty valence shell can share a lone pair of electrons from another atom to form a coordinate bond or a dative covalent

- example: in NH_4^+ the lone pair of nitrogen from NH_3 is shared with H^+

CHEMICAL BONDS

A chemical bond is the net electrostatic force that two atoms sharing electrons exert on each other.

- Bonding occurs when valence electrons are shared between two atoms or transferred from one atom to another.
- Valence electrons correspond to the group number of an element in the Periodic table
- Valence electrons: Valence electrons or outer electrons are the electrons in the highest energy level of an atom in which there are electrons.
- The type of bond that forms is dependent on the electronegativity difference (ΔEN) between the atoms.
- Electronegativity is a measure of the tendency of an atom in a molecule to attract bonding electrons. If one atom has a greater electronegativity than another, the electrons will be pulled more towards one atom than another. Such a shift of electrons creates negative and positive charge distributions inside the molecule.
 - $\Delta EN = 0$: Non-polar (pure) covalent
 - $\Delta EN < 1$: Weakly polar covalent
 - $\Delta EN > 1$: polar covalent
 - $\Delta EN > 2,1$: Ionic (transfer of electrons)
 - $\Delta EN > 3$: purely ionic

Types of bonds and molecules

- Non-polar bond: a bond in which the electron density is shared equally between the two atoms
Example: $H - H$
- EN (H) = EN (H): Bonding electrons are evenly shared. Charge is evenly distributed and no dipole formed.
 $\Delta EN = 2,1 - 2,1 = 0$
- Polar bond: a bond in which the electron density is shared unequally between the two atoms.

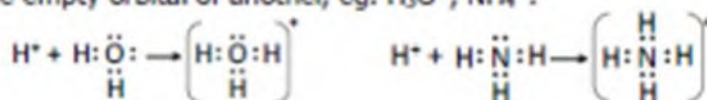
- Example: $\text{H} - \text{Cl}$
- $\text{EN}(\text{Cl}) > \text{EN}(\text{H})$: Electrons shift towards chlorine. Chlorine is slightly negative (δ^-) and hydrogen is slightly positive (δ^+).
 $\Delta\text{EN} = 3,0 - 2,1 = 0,9$
- Polar molecule: A molecule over which charge is distributed unevenly.
 Example: H_2O
- Non-polar molecule: A molecule over which the charge is evenly distributed.

Covalent Bonding (Between non-metals)

- Covalent bonding involves the sharing of electrons between two atoms to form a molecule.
- The covalent bond may be non-polar, weakly polar or polar.
- A non-polar (pure) covalent bond is a bond in which the electron density is shared equally between the two atoms.
 example:
 Two chlorine atoms are joined by a non-polar covalent bond ($\text{Cl} - \text{Cl}$)
 EN for chlorine atom is 3,0.
 $\Delta\text{EN} = 3,0 - 3,0 = 0$
- In a weakly polar covalent bond there is an unequal sharing of electrons
 example:
 A hydrogen atom and a bromine atom are joined by a weakly polar bond ($\text{H} - \text{Br}$)
 EN for hydrogen is 2,1 and EN for Br is 2,8.
 $\Delta\text{EN} = 2,8 - 2,1 = 0,7$
- A polar covalent bond has an unequal sharing of electrons leading to a dipole forming
- example:
 A hydrogen atom and an oxygen atom are joined by a polar covalent bond (H_2O)
 EN for oxygen atom is 3,5.
 $\Delta\text{EN} = 3,5 - 2,1 = 1,4$
- bonds between them. If two pairs of electrons are shared, a double bond is formed, e.g. between two O atoms to form H_2O_2 .
- Atoms with an empty valence shell can share a lone pair of electrons from another atom to form a coordinate covalent or dative covalent bond, e.g. in NH_4^+ the lone pair of nitrogen is shared with H^+ and in H_3O^+ the lone pair of oxygen is shared with H^+ .

DATIVE BONDING/ COORDINATE BONDING

A coordinate bond is formed when one atom donates both electrons to form a bonding pair. The filled electron orbital (lone pair) of one atom overlaps with the empty orbital of another, eg. H_3O^+ , NH_4^+ .



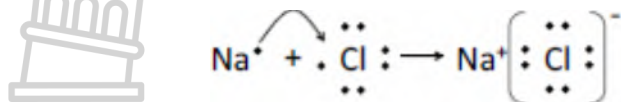
Ionic Bonding (between metals and non-metals)

- metal atom gives electron or electrons to non-metal atom
- metal forms a positive ion and positive ion is called a cation
- non-metal forms a negative ion and negative ion is called an anion
- electrostatic attraction of ions leads to formation of giant crystal lattice

Example:

Show how the ionic bond is formed between the sodium and the chlorine atoms.

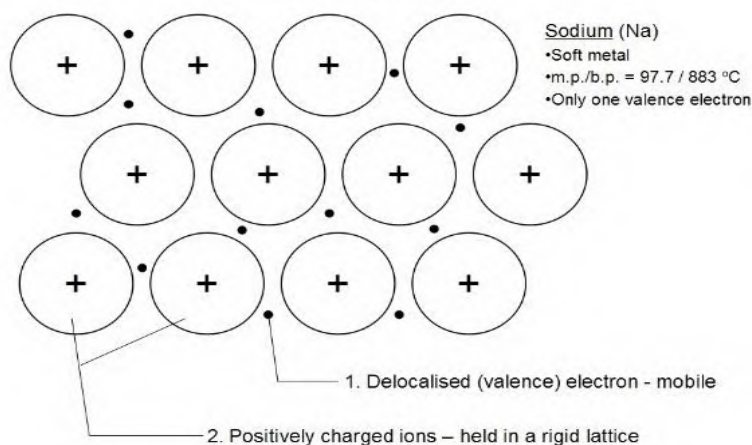
1. $\text{Na} - 1e^- \rightarrow \text{Na}^+$ (Na atom loses one electron)
2. $\text{Cl} + 1e^- \rightarrow \text{Cl}^-$ (Cl atom gains one electron)
3. $\text{Na}^+ + \text{Cl}^- \rightarrow \text{NaCl}$ (electrostatic attraction between the two ions)



Metallic Bonding (between metals)

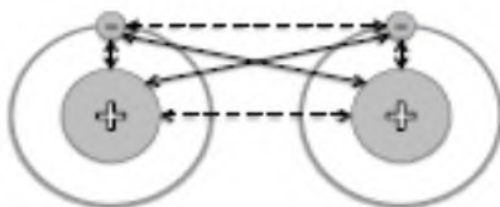
- Metallic bonding forms between the positive metal kernels and the sea of delocalized electrons.
- The metal atoms release their valence electrons to surround them.
- There is a strong but flexible bond between the positive metal kernels and a sea of delocalised electrons

Metallic Bonding



BOND ENERGY

- Bond energy is required to break one mole of its molecules into separate atoms.



- There are various attractive and repulsive forces at play between the two atoms during bonding.
- Attractive forces between the protons of one atom and the electrons of another.
- Attractive forces between the protons and electrons from the same atom.
- A repulsive force between the protons from each atom.
- A repulsive force between the electrons from each atom.
- The net electrostatic forces will determine bond strength, which can be quantified as the bond energy. This is the energy required to break the bond, or it is the energy released when bonds are formed.

FACTORS INFLUENCING BOND STRENGTH

*Bond length

The shorter the length of the bond, the stronger the bond.

*Atom size

The smaller the atoms, the stronger the bond.

***Bond order**

The more bonds (double, triple bonds etc.) between the atoms, the stronger the bond.

BOND LENGTH

Bond length: The average distance between the nuclei of two bonded atoms.

As the atoms get closer, the attractive forces get stronger until the minimum possible potential energy is reached (bond energy). The distance between the nuclei of the atoms at the minimum potential energy is the bond length. If the two atoms get closer than the bonding length, they will be forced apart by the repulsive forces, increasing the potential energy.

FACTORS INFLUENCING BOND LENGTH

*Atom size- The smaller the atoms, the shorter the bond.

*Bond order -The more bonds (double, triple bonds etc.) between the atoms, the shorter the bond.

*Difference in electronegativity (ΔEN). The greater the ΔEN , the shorter the bond.

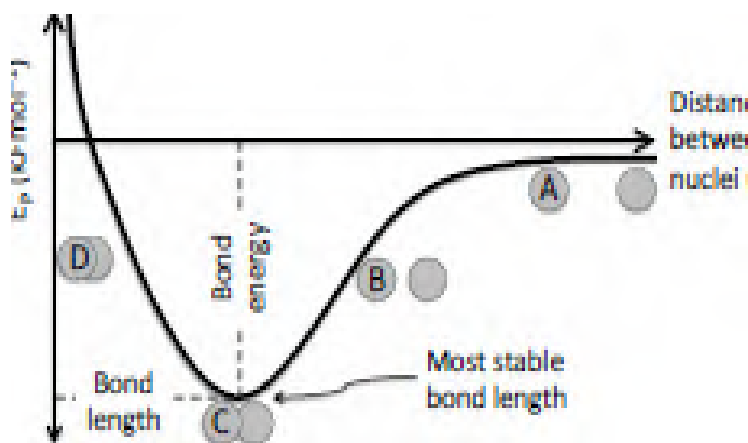
Process of bond formation

*Atoms are infinitely separated; potential energy is nearly zero.

*As the atoms move closer, the forces of attraction and repulsion increase until the forces of attraction dominate.

*The lowest, most stable energy state reached, chemical bond forms.

*Atoms move too close, forces of repulsion increase, potential energy increases



WORKED EXAMPLE 1

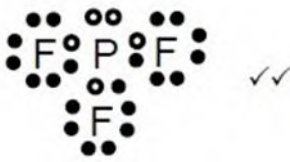
Both aluminium fluoride (AlF_3) and phosphorus trifluoride (PF_3) contain fluorine. Aluminium fluoride is a colorless solid in the production of aluminum, whilst phosphorus trifluoride is a poisonous, colorless gas.

- 1.1 Explain the difference between a covalent bond and ionic bond (2)
- 1.2 Name the type of chemical bond between particles in:
 - 1.2.1 AlF_3 (1)
 - 1.2.2 PF_3 (1)
- 1.3 Draw a lewis structure for:
 - 1.3.1 AlF_3 (3)
 - 1.3.2 PF_3 (2)
- 1.4 Write down the molecular shape of PF_3 . (1)
- 1.5 The melting point of AlF_3 is $1\ 291^\circ C$ and that of PF_3 is $-151^\circ C$. Fully explain this difference in melting point. (4)

[14]

SOLUTIONS

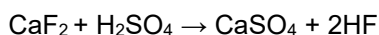
- 1.1 A covalent bond is a sharing of electrons between two atoms, whilst, ionic bond forms when there is a transfer of electrons from one atom to another (non-metal to metal) and oppositely charge atoms attract each other. $\checkmark\checkmark$ (2)

- 1.2 1.2.1 Ionic bond ✓ (1)
- 1.2.2 Covalent bond ✓ (1)
- 1.3 1.3.1  ✓✓ OR/OF $Al^{3+} 3 [F:]^{-}$ (3)
- 1.3.2  ✓✓ (2)
- 1.4 Pyramidal ✓ (1)
- 1.5 The ionic bonds between the aluminum fluoride are stronger than the dipole-dipole forces in between phosphorus fluoride molecules. ✓ ✓ (1)
- More energy is needed to overcome ionic bond in AlF_3 than intermolecular forces in PF_3 ✓ ✓ (4)

[14]

WORKED EXAMPLE 2

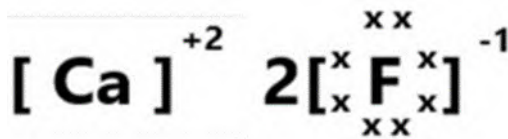
2. Hydrogen fluoride (HF) is used in the manufacturing of refrigerants, herbicides, pharmaceuticals and gasoline. Hydrogen fluoride (HF) can be prepared by treating calcium fluoride with sulphuric acid, according to the following balanced equation.



- 2.1 Draw a Lewis diagram for:
- 2.1.1 HF (2)
- 2.1.2 CaF_2 (2)
- 2.2 Define a polar covalent bond (2)
- 2.3 Which of the product will likely be:
- 2.3.1 A gas at room temperature (1)
- 2.3.2 A crystalline solid at room temperature (1)
- 2.4 Define the term electronegativity (2)
- 2.5 Use electronegativity to determine the types of bonds in CaF_2 (2)
- 2.6 Will HF be a polar or non-polar molecule? (1)
- 2.7 Explain your answer in question using calculation (2)

[15]

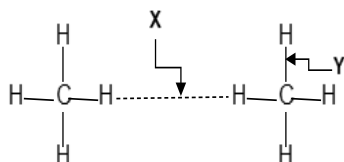
SOLUTION

- 2.1 2.1.1  ✓✓ (2)
- 2.1.2  ✓✓ (2)
- 2.2 A polar covalent bond is a bond in which the electron density is shared unequally between the two atoms ✓✓ (2)
- 2.3 2.3.1 HF/Hydrogen fluoride ✓ (1)
- 2.3.2 $CaSO_4$ /Calcium sulfate ✓ (1)

- 2.4 Electronegativity is a measure of the tendency of an atom in a molecule to attract bonding electrons. ✓✓ (2)
- 2.5 $\Delta EN = 4 - 1 = 3$ ✓
Ionic bond ✓ (2)
- 2.6 Polar molecule ✓ (1)
- 2.7 $\Delta EN = 4 - 2,1 = 1,9$ ✓ polar bond ✓
Thus, the molecule is polar (2)
- [15]**

MULTIPLE CHOICE QUESTIONS

- 1.1 The diagram below is used to demonstrate the differences between interatomic bonds (chemical bonds) and intermolecular forces.



- A X is stronger than Y
B X is an interatomic bond
C X is an intermolecular force
D Y is an intermolecular force (2)

- 1.2 Which ONE of the following bonds will have the highest bond energy?

- A H – H
B C – H
C C = C
D C ≡ C (2)

- 1.3 Consider the following compounds and their respective boiling points.

COMPOUNDS	BOILING POINT (°C)
He	-268,9
HBr	-66
HF	19,5

The correct arrangement for the decreasing strength of the intermolecular forces of the given compounds is ...

- A HF, He, HBr.
B HBr, He, HF.
C HF, HBr, He.
D He, HBr, HF. (2)

- 1.4 Consider the substances **P**, **Q** and **R**. **P** dissolves in **Q** but not in **R**.

The MOST LIKELY bonds or intermolecular forces in substances **P**, **Q** and **R** are :

	P	Q	R
A	ionic bonds	Dipole-dipole forces	London forces
B	London forces	Dipole-dipole forces	ionic bonds
C	Dipole-dipole forces	London forces	ionic bonds

D	Ionic bonds	London forces	Dipole-dipole forces
---	-------------	---------------	----------------------

(2)

1.5 The intermolecular forces holding the molecules in ice together are ...

- A ion-dipole forces.
- B Van der Waals forces.
- C covalent bonds.
- D hydrogen bonds.

(2)

[10]

LONG QUESTIONS

QUESTION 2

Consider the following compounds.

O_2 ; H_2O ; H_3O^+

2.1 Define the term *molecule*.

(2)

2.2 Draw the Lewis structures for:

2.2.1 O_2

(2)

2.2.2 H_2O

(2)

2.3 The H_3O^+ ions forms when the oxygen atom in H_2O donates its lone pair of electrons into the vacant orbital of H^+ .

2.3.1 Write down the NAME of the type of bond described by the underlined phrase.

(2)

2.3.2 Draw the Lewis structure for the H_3O^+ -ion.

(2)

2.4 Which molecule is polar, H_2O or O_2 ? Explain your answer.

(2)

[12]

QUESTION 3

3.1 Draw a Lewis diagram of methane.

(2)

3.2 What is the name of the special kind of bond found inside the methane molecule?

(1)

3.3 Answer the following questions about magnesium chloride:

3.3.1 Make a Lewis diagram to indicates the transfer of electrons between magnesium and chlorine

(4)

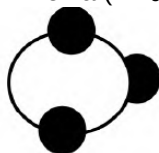
3.3.2 Magnesium chloride is ionic substance. Prove this statement with relevant calculation

(2)

[9]

QUESTION 4

4 Ammonia (NH_3) is represented in the diagram below.



4.1 Write down the number of valence electrons for nitrogen

(1)

4.2 Draw the Lewis structure for ammonia

(3)

4.3 Write down the molecular shape of the ammonia molecule

(1)

4.4 Explain why ammonia does not have a trigonal planar shape.

(2)

4.5 Is ammonia a polar or non-polar molecule?

(1)

4.6 Refer to the structure and difference in electronegativity to explain the answer to QUESTION 4.5.

(2)

When a hydrogen ion reacts with an ammonia molecule, it forms the ammonium ion.

4.7 By referring to the ammonium ion, explain what is a dative covalent bond

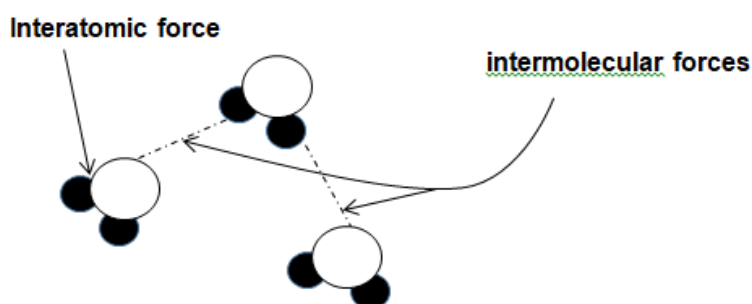
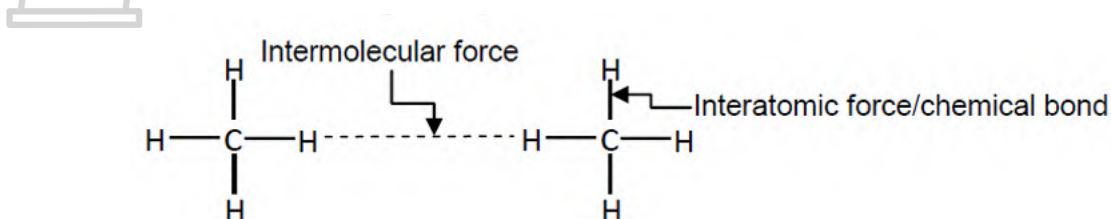
(2)

[12]

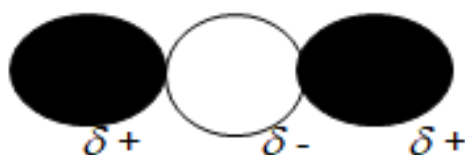
INTERMOLECULAR FORCES

- Intermolecular forces are forces of attraction between molecules in a substance.
- Intermolecular forces are weak forces of attraction between molecules or between atoms of noble gases
- Intermolecular forces are not the same as intramolecular bonds.
- Intramolecular (interatomic) bonds exist between atoms in a molecule.
- The intermolecular forces are **weaker** than interatomic forces.

Examples of intermolecular forces and intramolecular (interatomic) forces

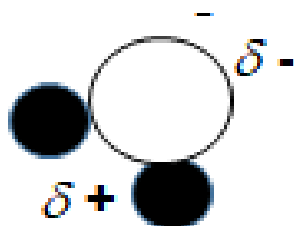


- Molecules can be **polar** or **non-polar**.
- The polarity of a molecule is determined by two factors:
 - The difference in electronegativity between the bonding atoms.
 - The geometry (shape) of a molecule
- **Non-polar molecules** have no dipoles (positive and negative ends) e.g., CO_2



Polar molecule

- In polar molecule there are δ^+ and δ^- ends e.g., H_2O and the geometry of the water molecule is angular (bent).



- H_2O molecule has H-end being δ^+ and O-end being δ^- , it has **dipoles**. The molecule is polar

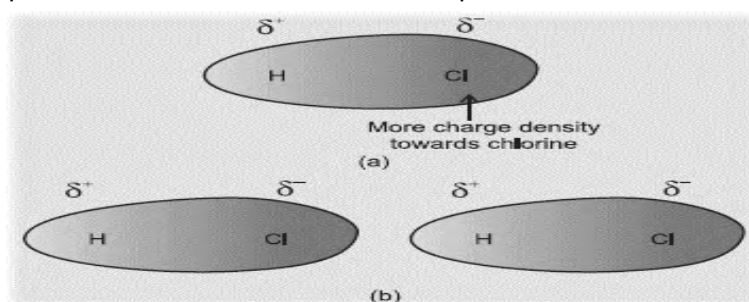
- **Non-polar bonds:**

- H_2 is non-polar because the two atoms are identical, there is even or symmetrical distribution of charge, this makes H_2 a non-polar molecule.

TYPES OF INTERMOLECULAR FORCES

- Different intermolecular forces (Van der Waals forces) ARE

- Mutually induced dipole forces or London forces: Forces between non-polar Molecules
- Dipole-dipole forces: Forces between two polar molecules
- Dipole-induced dipole forces: Forces between polar and non-polar molecules
- Hydrogen bonding: Forces between molecules in which hydrogen is covalently bonded to nitrogen, oxygen, or fluorine – a special case of dipole-dipole forces
- Ion-dipole forces: Forces between ions and polar molecules



INTERMOLECULAR FORCES AND PHYSICAL PROPERTIES (factors affecting intermolecular forces)

Boiling point and melting point

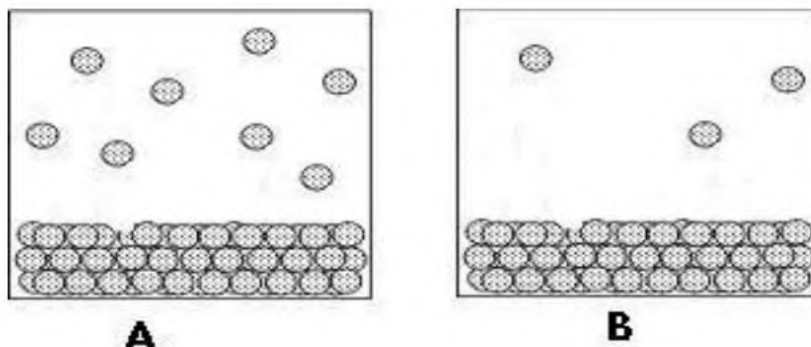
- Intermolecular forces dictate several properties, such as melting points, boiling points, solubility of substances, vapour pressure, viscosity etc.
- Boiling point is the temperature at which the vapour pressure of a substance equals atmospheric pressure. The stronger the intermolecular forces, the higher the boiling point.
- Melting point is the temperature at which the solid and liquid phases of a substance are at equilibrium. The stronger the intermolecular forces, the higher the melting point.
- A liquid with high boiling point, like water (H_2O , b.p. $100\text{ }^\circ\text{C}$), exhibits stronger intermolecular forces compared to a liquid with low boiling-point, like hexane (C_6H_{14} , b.p. $68,73\text{ }^\circ\text{C}$).

Solubility

- Solubility is the property of a solid, liquid, or gaseous chemical substance (solute) to dissolve in a solid, liquid, or gaseous solvent to form a homogeneous solution.
- Miscible liquids have similar polarities.
- For example, methanol and water are miscible and are both polar and capable of hydrogen bonding.
- When methanol and water are mixed, they interact through intermolecular hydrogen bonds of comparable strength to the methanol–methanol, and water–water interactions; thus, they are miscible.
- Likewise, nonpolar liquids like hexane and bromine are miscible with each other through dispersion forces.
- The chemical axiom “like dissolves like” is useful to predict the miscibility of compounds.
- For example, nonpolar hexane is immiscible in polar water. Relatively weak attractive forces between the hexane and water do not adequately overcome the stronger hydrogen bonding forces between water molecules.

Vapor Pressure

- Vapor pressure is the pressure exerted by a vapour at equilibrium with its liquid in a closed system. The stronger the intermolecular forces, the lower the vapour pressure.
- At the surface, some molecules of a liquid have enough kinetic energy to break their attractive forces with neighbouring molecules.
- These molecules escape from the liquid phase and form a gas above the surface of the liquid. If there is a lid, pressure develops.

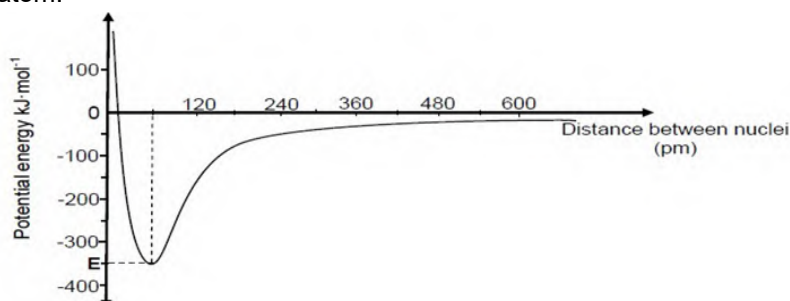


- In the picture above, A has more gas, meaning gas molecules are breaking away from the liquid state easier than molecules in B.
- The intermolecular forces must be weaker in A than in B.
- Molecules in A are less attracted to each other than molecules in B.
- Intermolecular forces in A are weaker than those in B

WORKED EXAMPLES

QUESTION 1

The graph below shows the change in energy that takes place when a hydrogen (H) atom approaches a bromine (Br) atom.



- 1.1 Define the term *bond length* (2)
- 1.2 From the graph, write down the:
 - 1.2.1 Bond length, in pm, of the H-Br bond (2)
 - 1.2.2 Energy, in kJ mol^{-1} (1)
 - 1.2.3 Name of the potential energy represented by E (2)
- 1.3 How will the bond length of H-F bond compare to that of H-Br bond? (2)
Write down EQUAL TO, SHORTER THAN, LONGER THAN. Give a reason for your answer [9]

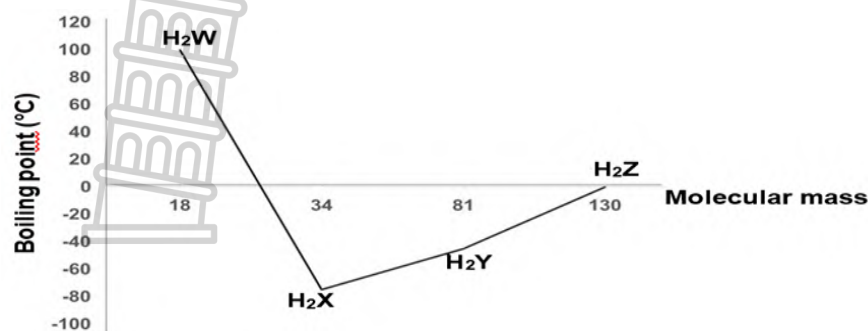
SOLUTION

- 1.1 The average distance between nuclei of two bonded atoms in a molecule. (2)
- 1.2
 - 1.2.1 60 (pm) ✓✓ (2)
 - 1.2.2 350 (kJ mol^{-1}) ✓✓ (2)
 - 1.2.3 Bond energy ✓ (1)
- 1.3 Shorter than, ✓ F atoms are smaller than Br atoms and come closer to H atom. ✓ (2)

[9]

QUESTION 2

The graph below shows the relationship between the boiling points of the hydrides of the Group 16 elements and molecular mass.



- 2.1 Define the term *boiling point*. (2)
- 2.2 Consider the boiling points of hydrides, excluding H₂W.
 - 2.2.1 What is the relationship between molecular mass and boiling points of compounds? (2)
 - 2.2.2 Explain the trend in boiling points. Refer to molecular mass, intermolecular forces and energy to explain your answer. (3)
- 2.3 H₂W does not follow the trend in boiling points for the 16 hydrides.
 - 2.3.1 Identify the compound H₂W (1)
 - 2.3.2 Name the types of intermolecular forces found in this compound (2)
 - 2.3.3 Explain why the boiling point is higher than expected (2)

[12]

SOLUTION

- 2.1 The temperature at which the vapour pressure of a substance equals atmospheric pressure (2)
- 2.2 2.2.1 As the molecular mass increases, the boiling point increases. ✓✓ (2)
- 2.2.2 As the molecular mass increases, ✓ the strength of the intermolecular forces increases. ✓
Therefore more energy is needed to overcome/ weaken the intermolecular forces. ✓(No mark if BROKEN is used instead of overcome or weaken.) Thus the boiling point increases. (3)
- 2.3 2.3.1 water✓ (1)
- 2.3.2 Hydrogen Bond✓ (2)
- 2.3.3 Hydrogen bonds are stronger than dipole-dipole forces ✓ therefore more energy is needed to overcome/weaken the forces. ✓ Thus the boiling point is higher than expected. (2)

[12]

ACTIVITIES: INTERMOLECULAR FORCES

QUESTION 1

Which intermolecular forces are found in:

- 1.1 hydrogen fluoride (HF) (1)
- 1.2 methane (CH₄) (1)
- 1.3 potassium chloride in ammonia (KCl in NH₃) (1)
- 1.4 krypton (Kr) (1)

[4]

QUESTION 2

Refer to the table below and answer questions that follow.

	distance	Melting Point	Boiling Point	Relative molecular/ atomic mass



Sodium	97,79	882,8	23
Bromine (Br ₂)		58,8	160
Hydrogen chloride (HCl)	-1 14,2	-85,1	36,5
Water (H ₂ O)		100	18

- 2.1 Define the term *boiling point*. (2)
- 2.2 By referring to the forces present in hydrogen chloride, explain the difference between intermolecular forces and interatomic forces. (2)
- 2.3 Which of these substances will be a liquid at +50 °C? (1)
- 2.4 Refer to intermolecular forces and energy and explain why the boiling point of HCl is lower than the boiling point of H₂O (4)
- 2.5 NaCl dissolved in 1-120. (4)
- 2.5.1 Write down the name the intermolecular forces between NaCl and 1-120. (1)
- 2.5.2 Are the intermolecular forces stated in QUESTION 2.5.1 weaker or stronger than those of H₂O molecules. (1)
- 2.6 Refer to TYPE of INTERMOLECULAR FORCES, MOLECULAR MASS, STRENGTH of INTERMOLECULAR FORCES and ENERGY to explain the big difference in boiling points of Br₂ and HCl. (4)

[15]

QUESTION 3

The table below shows the relationship between the melting points of three molecules:

Molecule	Melting point (°C)
CH ₄	-182,5
CF ₄	-150
CCl ₄	-23

- 3.1 Define the term *melting point*. (2)
- 3.2 Explain the trend in melting points of the molecules in the above table by referring to the intermolecular forces and energy involved. (4)
- 3.3 Which molecule in the table will have the highest vapour pressure at a given temperature? Explain the answer by referring to the data in the table. (2)
- 3.4 Water (H₂O) is a smaller molecule than CCl₄, but water has a higher melting point than CCl₄. Explain this observation by referring to the type of intermolecular forces involved. (2)
- 3.5 Write down the name of intermolecular force that will exist in a mixture of H₂O and CCl₄. (2)

[12]

QUESTION 4

In the table below the boiling points of four hydrogen halides are given.

Hydrogen halides	A	B	C	D
	HF	HCl	HBr	HI
Boiling points (°C)	19,4	-85	-67	-35,5

- 4.1 Define the term *boiling point*. (2)
- 4.2 What is the relationship between strength of the intermolecular forces and boiling point? (1)
- 4.3 Which of the hydrogen halides (A, B, C or D) in the table above has the ...