BONDING AND STRUCTURE

Competency:

The learner analyzes the types of chemical bonds and molecular structures and relates them to the properties and uses of substances in real-life contexts.

SUB-TOPIC 3.1: FORMATION OF IONIC AND METALLIC BONDS

Learning Outcomes: The learner should be able to:

- Justify the formation of ionic and metallic bonds based on electron transfer and electrostatic forces (u, s).
- Evaluate the properties of ionic and metallic compounds in relation to bond strength and structural composition (u).

INTRODUCTION TO CHEMICAL BONDING

What is Chemical Bonding?

Chemical bonding is the glue that holds atoms together, creating the substances we see every day.

Chemical bonding refers to the forces that hold similar or different atoms, ions and molecules together to form new chemical substances in which they are more stable.

Imagine lonely people (atoms) looking for friends to feel complete. They either share, donate, or accept "friends" (electrons) to form a stable group (molecule).

An atom chemically combines in order to attain:

(i) a stable configuration. The atoms in the compound formed have a stable electronic structure like that of a noble gas. These atoms are more stable than the individual atoms.

(ii) a state of minimum energy. The process must be energetically favourable leading to atoms in a combined state to attain minimum amount of energy.

During bond formation, the outermost energy level (valence shell) contains electrons known as the valence electrons that are involved in bonding. However, in some cases the electrons in the penultimate shell of the d-block elements (in the 3d-orbitals) are also involved in bonding.

N.B: The penultimate shell is the 2nd last energy level.

Atoms with stable electronic configuration are relatively inert/unreactive and they are also in a state of minimum energy. e.g. the noble gases.

A **chemical bond** is a lasting attraction between atoms, ions or molecules that enables the formation of new chemical substances through either transferring or sharing of electrons.

Types of Chemical Bonds:

- **Ionic Bonding** or **Electrovalent bonding**. This **occurs between a metal and a non-metal** in which electrons are transferred from a metal atom to a non-metal atom to achieve stability, forming positively and negatively charged ions that attract each other.

Think of table salt (**NaCl**). Sodium (Na) "gives away" an electron to Chlorine (Cl), just like a generous person giving away money to a friend. The two become strongly connected due to this transaction.

- **Covalent Bonding.** This **occurs between two non-metals** in which electrons are shared between the atoms to achieve stability.

e.g. Water ($\mathbf{H_2O}$) forms because hydrogen and oxygen non-metal atoms "share" electrons, just like two friends sharing an umbrella during a rainy day.

- **Metallic Bonding.** This occurs **in metals**. Electrons are free to move, creating a "sea of electrons" (delocalized electrons) that allows metals to conduct electricity.

Metallic bonding is like a group of people at a concert where everyone shares their energy by passing around a glowing bracelet (electrons), keeping the whole crowd connected.

- **Dative covalency** or **coordinate covalent bonding**. This is a type of covalent bond where **both electrons in the bond come from the same atom**. This is in contrast to a regular covalent bond, where each atom contributes one electron to the shared pair.

- **Intermolecular Forces.** These include; Van der Waal's forces and Hydrogen bonding.

Intermolecular forces are forces of attraction that occur between molecules. These forces are weaker than **intramolecular forces** (like covalent or ionic bonds), but they are essential for understanding the differences in the; **Boiling/melting points**, **Solubility**, **Viscosity**, **Surface tension**, **Physical state (solid/liquid/gas)** at room temperature, etc.

Conclusion: Chemical bonding is essential for forming the substances we use daily. Whether it's salt in food, water to drink, or metals in electronics, bonds hold everything together.

IONIC BONDING / ELECTROVALENT BONDING

General objectives; By the end of the subtopic, one should be able to:

- Define ionic bonding.
- Explain the formation of ionic bonds using examples.
- Describe the properties of ionic compounds.
- Relate ionic bonding to real-life scenarios.

Introduction to Ionic Bonding:

Chemistry is not just about composition of substances but it is also about **relationships between** the components of substances. Just like people form bonds to become stronger together, atoms also form bonds to achieve **stability and purpose**.

In the world around us, from the salt in our food to the minerals in our bodies, atoms are constantly interacting. But atoms are not complete on their own but they seek balance. They strive to achieve a stable state, similar to the noble gases. And to do this, they must **give, take, and share**.

One of the most powerful types of chemical relationships is **ionic bonding**.

Imagine one atom with something to give, and another in great need to receive. This transfer is not random but it is driven by the natural desire of atoms to become stable.

When one atom **loses electrons** and another **gains the electrons**, something remarkable happens: they become charged particles called **ions**. These oppositely charged ions are then held together by a strong force of attraction known as an **ionic bond**.

This is not just a simple interaction but it is a transformation. Atoms change their identities to form compounds that are essential to life and industry.

Take common salt, for example. Sodium, a highly reactive metal, and chlorine, a poisonous gas, come together to form sodium chloride which is a stable, safe, and essential compound. This is the power of ionic bonding; turning danger into usefulness, instability into balance.

💡 Key Idea to Emphasize

Ionic bonding is the electrostatic force of attraction between oppositely charged ions formed when electrons are transferred from one atom to another.

🔥 Motivation to Learners

As we study ionic bonding, ask yourself:

- Why do atoms prefer to lose or gain electrons instead of sharing?
- What makes ionic compounds strong and stable?
- How does this bonding explain properties like high melting points and conductivity?

Understanding ionic bonding will unlock the ability to explain **how matter is structured and why substances behave the way they do**.

Elements at the opposite ends of the Period Table attain a noble gas configuration by transferring electrons to one another.

In the electron transfer process, the elements of the left-hand side of the Period Table (i.e. metallic elements) give up their outer electrons to form positively charged ions (cations) while those at the right-hand side of the Periodic Table (i.e. non-metallic elements) gain electrons to form negatively charged ions (anions).

Since the ions formed have opposite charges, an electrostatic force of attraction occurs between these two oppositely charged ions and this constitutes an **ionic bond**.

Key Terms in Ionic Bonding:

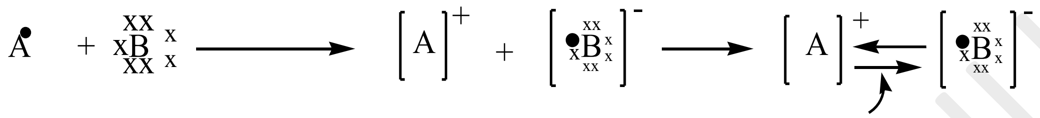
A cation: A positively charged ion, formed when an atom loses one or more electrons. (Usually formed by metals).

An anion: A negatively charged ion, formed when an atom gains one or more electrons. (Usually formed by non-metals).

Electron Transfer: The process of transferring electrons from one atom to another, resulting in the formation of ions.

Electrostatic Attraction: The force that holds the oppositely charged ions together in an ionic bond.

i.e.



Electrostatic force that forms the ionic bond

Examples of formation of compounds through ionic bonding

1. Formation of Sodium Chloride (NaCl)

Step 1: Identify the atoms involved.

Sodium (Na) is a metal with one electron in its outermost shell.

Chlorine (Cl) is a non-metal with seven electrons in its outermost shell.

Step 2: Electron Transfer.

Sodium (Na) needs to lose its one electron to achieve a stable electron configuration similar to the nearest noble gas (Neon, Ne), which has a full outer shell.

Chlorine (Cl) needs to gain one electron to complete its outer shell and achieve the stable configuration of Argon (Ar).

Step 3: Formation of Ions.

Sodium loses its electron, becoming the sodium ion, **Na**$^+$ (a cation).

Chlorine gains the electron, becoming the chloride ions, **Cl**$^-$ (an anion).

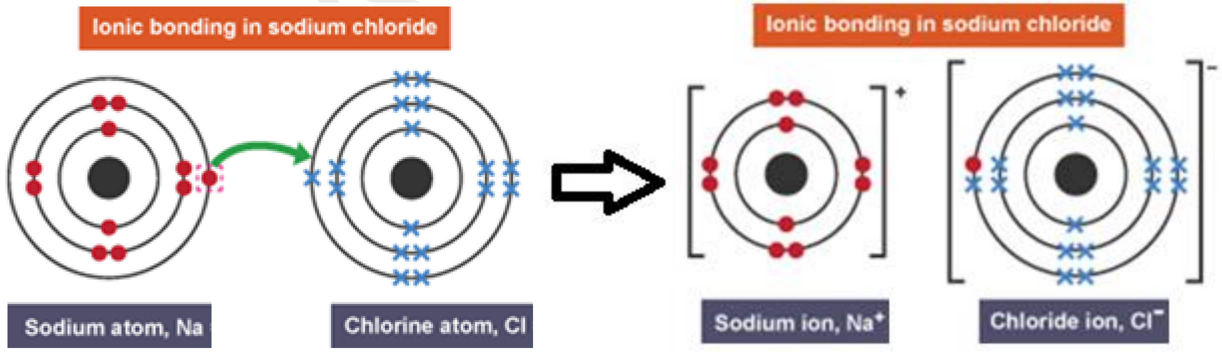
Step 4: Electrostatic Attraction.

The **Na**$^+$ and **Cl**$^-$ ions are attracted to each other due to their opposite charges, forming an ionic bond.

This results in the creation of **sodium chloride (NaCl)**, a common salt.

Conclusion:

Sodium chloride is an ionic compound formed from the transfer of an electron from sodium to chlorine.



2. Formation of Magnesium Chloride (MgCl$_2$)

Step 1: Identify the atoms involved.

Magnesium (Mg) is a metal with two electrons in its outer shell.

Chlorine (Cl) is a non-metal with seven electrons in its outer shell.

Step 2: Electron Transfer.

Magnesium (Mg) needs to lose two electrons to achieve the electron configuration of Neon (Ne), a noble gas. Chlorine (Cl) needs to gain one electron to complete its outer shell and achieve the stable configuration of Argon (Ar). Since one chlorine atom can only accept only one electron before it can achieve its noble gas configuration; therefore, 2 atoms of chlorine are required to accept the 2 electrons lost and donated by the magnesium.

Step 3: Formation of Ions.

Magnesium loses two electrons, becoming a magnesium ion, **Mg^{2+}** (a cation).

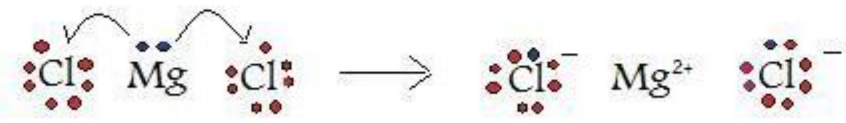
Each one of the two chlorine atoms gains the one electron, becoming two chloride ions, **$2Cl^{-}$** (anions).

Step 4: Electrostatic Attraction.

The **Mg^{2+}** and **$2Cl^{-}$** ions are attracted to each other due to their opposite charges, forming an ionic bond. This results in the creation of **magnesium chloride ($MgCl_2$)**.

Conclusion:

Magnesium chloride is an ionic compound formed by the transfer of two electrons from magnesium to two oxygen atoms.



3. Formation of Magnesium Oxide (MgO)

Step 1: Identify the atoms involved.

Magnesium (Mg) is a metal with two electrons in its outer shell.

Oxygen (O) is a non-metal with six electrons in its outer shell.

Step 2: Electron Transfer.

Magnesium (Mg) needs to lose two electrons to achieve the electron configuration of Neon (Ne), a noble gas. Oxygen (O) needs to gain two electrons to complete its outer shell and achieve the configuration of Neon (Ne).

Step 3: Formation of Ions.

Magnesium loses two electrons, becoming the magnesium ion, **Mg^{2+}** (a cation).

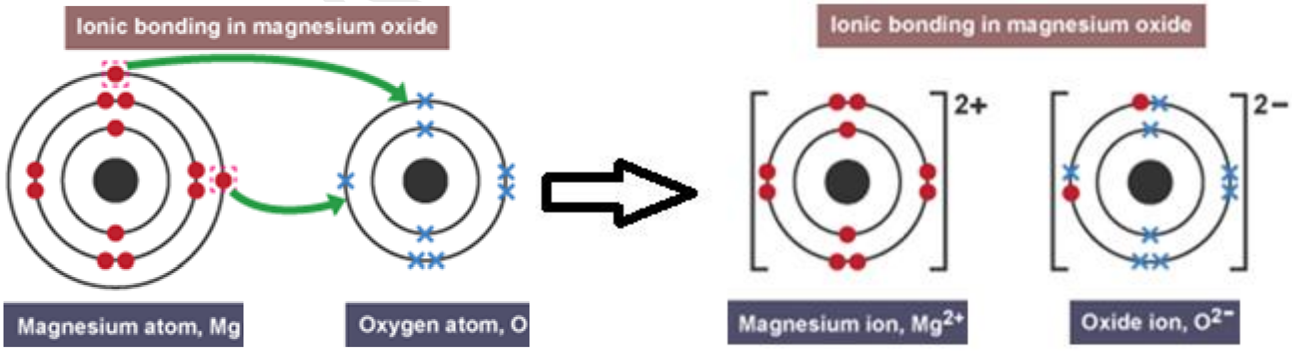
Oxygen gains the two electrons, becoming the oxide ion, **O^{2-}** (an anion).

Step 4: Electrostatic Attraction.

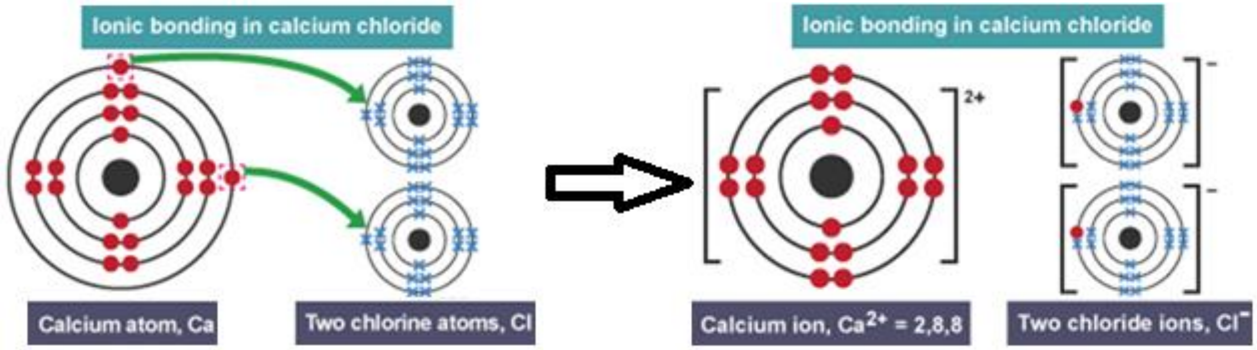
The **Mg^{2+}** and **O^{2-}** ions are attracted to each other due to their opposite charges, forming an ionic bond. This results in the creation of **magnesium oxide (MgO)**.

Conclusion:

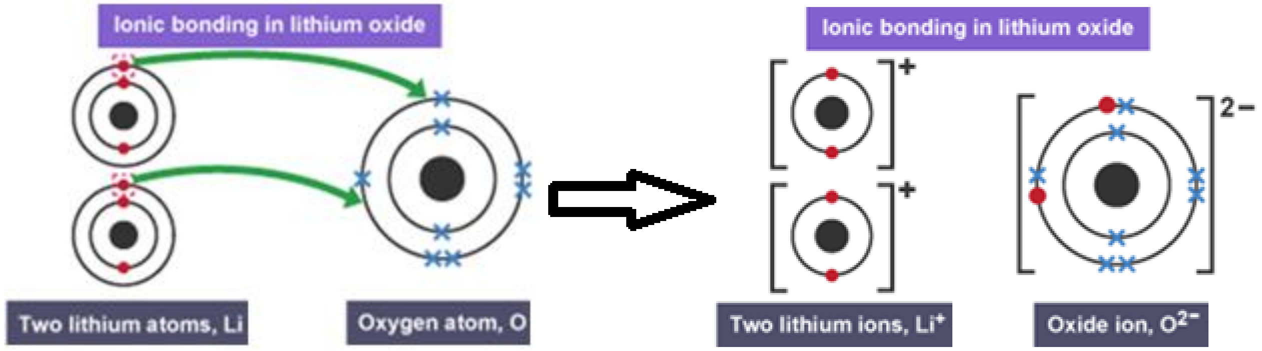
Magnesium oxide is an ionic compound formed by the transfer of two electrons from magnesium to oxygen.



4. Formation of Calcium Chloride ($CaCl_2$)



5. Formation of Lithium Oxide (Li_2O)



Ionic bonds extend continuously equally in all directions to form a giant ionic structure of a high melting or boiling point.

Factors that influence the strength of an ionic bond

The strength of an ionic bond depends on the following factors:

(a) Ionic charge. From Coulomb's Law (electrostatic force law), the force of attraction is directly proportional to the **product of the charges** on the ions.

The higher the ionic charge, the stronger the attraction for the oppositely charged ion (ionic bond).

(b) Ionic radius. According to Coulomb's Law, the force of attraction between two charges increases as the distance between them decreases.

The smaller the ionic radius, the stronger the attraction for the oppositely charged ion (ionic bond). This is because, smaller ions can get closer together. Thus, there is a shorter distance between charges leading to stronger electrostatic forces of attraction between the ions.

⚠ **Important note:** Ionic radius works together with the ionic charge and the lattice energy (Lattice Energy)

So, the strongest ionic bonds occur when **charges are high AND ions are small**.

(c) Covalent character (degree of covalency). When covalent character increases, the bond becomes **less purely ionic**, weakening the electrostatic attraction between ions.

This behavior is explained by Fajans' Rules which state that covalent character increases when; the **cation is small and highly charged** (strong polarizing power) and when the **anion is large** (strong polarizability)

i.e. Covalent character is generally introduced by:

(i) cations with a high polarizing power which include those usually with a:

- high ionic charge of 3+ and above.
- small ionic radius such as a cation formed by period 2 metals (Li^+, Be^{2+}), transition metals.

(ii) anions with high polarizability which include those usually with a large ionic radius. e.g. most anions formed by non-metals except those in period 1 and 2 together with hydroxide ions.

General characteristics of ionic compounds:

(a) They are made up of particles called ions of opposite charges.

Ionic compounds consist of **ions**, which are atoms or groups of atoms that have gained or lost electrons.

- Metals lose electrons to form **positive ions (cations).**
- Non-metals gain electrons to form **negative ions (anions).**

These oppositely charged ions attract each other through strong **electrostatic forces**, forming an ionic bond.
For example, sodium chloride (NaCl) is made of Na^+ and Cl^- ions.

(b) They conduct electricity in molten or aqueous state.

In solid form, ions are **fixed in position** and cannot move, so they cannot conduct electricity.
However, when **molten (melted)** or when **dissolved in water (aqueous solution),** the ions become **free and mobile (delocalized)**, allowing them to carry electric charge.
Because of this, ionic compounds are called **electrolytes**.

(c) They are hard solids under normal conditions.

At normal temperature and pressure, ionic compounds exist as **crystalline solids**.
This is because of the strong electrostatic forces between oppositely charged ions (ionic bonds) which hold them tightly together.
The ions are arranged in a **closely packed, giant ionic lattice structure** which makes them **hard and rigid**.

(d) They have high melting and boiling points.

A large amount of energy is needed to break the strong electrostatic forces of attraction between the oppositely charged ions (ionic bonds). As a result, ionic compounds melt and boil at **high temperatures**
The stronger the ionic bond, the higher the melting and boiling points.

(e) They are soluble in polar solvents but insoluble in nonpolar solvents.

Ionic compounds dissolve well in **polar solvents** like water.
This is because, the water molecules are polar hence have partial positive and negative ends which enable them to interact with (get attracted to) and **separate the ions** (a process called hydration).
Ionic compounds do not dissolve in **nonpolar solvents** like benzene because they are nonpolar solvents which cannot effectively attract or separate the charged ions.

Polarization in Ionic Bonds

Polarization is the process by which *a cation distorts an electron cloud of either an anion in an ionic compound or an atom in neighbouring neutral molecule*.
Polarization is as a result of a cation to attract either the outermost electrons of an anion in an ionic compound or lone pairs of electrons of an atom in a neighbouring neutral molecule.
e.g. - In solid aluminium chloride, $AlCl_3$ the aluminium ion, Al^{3+} will attract the outermost electrons of the chloride ions, Cl^- towards itself.
- In aqueous aluminium chloride, the aluminium ion, Al^{3+} will attract the lone pairs of electrons on the oxygen atom of the water molecule, H_2O towards itself.
Polarization in solid ionic compounds results into partial sharing of the distorted electrons between the cation and the anion leading to a partial covalent character (degree of covalency). Therefore, in these solid ionic compounds where polarization occurs, the ionic character reduces and hence the ionic bond becomes weaker.
Polarization in ionic compounds involves the following features:
1. **Polarizing power of the cation.** *This is the tendency of a cation to distort the electron cloud of either a neighbouring anion in an ionic compound or an atom in a neutral molecule.*
Polarizing power is a property of a cation. Thus, the polarizing power of a cation depends on:
(a) Ionic (cationic) radius.
Trend: The smaller the ionic radius, the higher the polarizing power. *The reverse is true*.
Explanation: When the radius of a cation is smaller, there is a stronger nuclear attracting power for the electrons on either a neighbouring anion or molecule and hence the higher the polarizing power.
e.g. - In the chlorides formed by group(II) metals, the polarizing power decreases in the order:

$Be^{2+} > Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$.

The ionic radius increases down the group(II) metal cations. This reduces the charge density and hence the polarizing power such that the covalent character of the chlorides reduces.

As a consequence, the ionic character increases from beryllium chloride to barium chloride as the ionic bond becomes stronger.

- In the chlorides formed by Period 3 metals, the polarizing power increases in the order:

$Na^+ < Mg^{2+} < Al^{3+}$.

Keeping other factors constant, the ionic radius decreases across the Period 3 metal cations. The charge density increases leading to increase in the polarizing power and hence increase in the covalent character. Thus, the ionic character decreases from sodium chloride to aluminium chloride as the ionic bond becomes weaker.

(b) Ionic (cationic) charge.

Trend: The higher the ionic charge, the higher the polarizing power. *The reverse is true*.

Explanation: When charge on the cation is higher, there is a stronger nuclear attracting power for the electrons on either a neighbouring anion or a molecule and hence the higher the polarizing power.

e.g. In the chlorides of Period 3 metals, the polarizing power increases in the order:

$Na^+ < Mg^{2+} < Al^{3+}$.

There is an increase in the ionic charge across Period 3 metals. The charge density increases leading to an increase in the polarizing power and hence an increase in the covalent character. Thus, the ionic character reduces from sodium chloride to aluminium chloride as the ionic bond becomes weaker.

On the other hand, there is also a decrease in the cationic radius across the Period 3 metals. This as well increases the charge density leading to further increase in the polarizing power and hence there is a greater increase in the covalent character. Thus, the ionic character further reduces from sodium chloride to aluminium chloride as the ionic bond becomes much weaker.

2. **Polarizability of the anion.** *This is the ease with which the electron cloud of either an anion in an ionic compound or an atom of a molecule is distorted by a neighbouring cation.*

Polarizability is a property of an anion or a molecule. Polarizability of an anion depends on:

(a) Ionic (anionic) radius.

Trend: The larger the ionic radius, the higher the polarizability. *The reverse is true*.

Explanation: When the radius of the anion is larger, the outermost electrons are further away from the nucleus and the nuclear attracting power for their outermost electrons reduces. Thus, there is a greater ease with which the outermost electrons of the anion are attracted (distorted) by a *neighbouring* cation and hence the higher the polarizability.

e.g. In group(VII), polarizability of the halide ions of sodium increases in the order;

$F^- < Cl^- < Br^- < I^-$.

The ionic radius increases down the group(VII) elements. The polarizability of the anions and the covalent character increases. Thus, the ionic character reduces from sodium fluoride to sodium iodide as the ionic bond becomes weaker. i.e. Fluoride ions having the smallest ionic radius are not easily polarized and most of its compounds are highly ionic.

(b) Ionic (anionic) charge.

Trend: The higher the ionic charge, the higher the polarizability. *The reverse is true*.

Explanation: When the charge on the anion is higher, there is a greater ease with which the electrons of the anion responsible for the negative charge are attracted (distorted) by a *neighbouring* cation and hence the higher the polarizability.

- In the aluminium compounds containing anions of Period 2 non-metals, the polarizability decreases in the order; $N^{3-} > O^{2-} > F^-$.

Keeping other factors constant, the ionic radius decreases across the Period 3 non-metal ions. Thus, the electrons of the anion responsible for the negative charge are less easily attracted (distorted) by a *neighbouring* cation and hence the polarizability reduces leading to a decrease in the covalent character. Thus, the ionic character increases from aluminium nitride to aluminium fluoride as the ionic bond becomes stronger.

✅ **How covalent character affects ionic bond strength:**

- **More covalent character → weaker ionic bond**

  - The electron cloud is distorted (polarized).
  - Attraction is no longer purely between full charges.
  - The bond becomes partly shared → reduces ionic strength.
- **Less covalent character → stronger ionic bond**
  - Ions remain more "pure" (full charges).
  - Strong electrostatic attraction is maintained.

Generally, in expected ionic compounds:

(a) Covalent character dominates the ionic character when the cation has a high polarizing power and on the other hand, the anion has a high polarizability. Hence the compound is said to be predominantly covalent.

e.g aluminium bromide, iron(III) bromide, lead(IV) chloride, lithium chloride, etc.

(b) Ionic character dominates the covalent character when either the cation has a high polarizing power while the anion has a low polarizability or the cation has a low polarizing power while the anion has a high polarizability. Hence the compound is said to be predominantly ionic.

e.g. sodium bromide, aluminium fluoride, etc.

(c) Purely ionic character is established when the cation has a low polarizing power and the anion has a low polarizability. e.g. sodium fluoride.

⚠ **Important note:**

Even though covalent character reduces *ionic* strength, it doesn't always mean the bond is "weak" overall—it just means it is **less purely ionic**.

The effects of polarization on properties of some ionic compounds include:

1. Low melting points and boiling points than expected in some ionic compounds.

e.g. The melting point of sodium chloride, NaCl is 801^0C while that of aluminium chloride, $AlCl_3$ is 193^0C.

Explanation: The aluminium ion has a *higher ionic charge* and *smaller ionic radius* unlike the sodium ion. This gives the aluminium ion a *higher charge density* and hence a *higher polarizing power*. Since the chloride ions are relatively large, they have a high polarizability and are consequently *easily polarized* by the aluminium ion forming a predominantly *covalent character*. i.e more of a covalent character than ionic. This reduces the strength of the ionic bond and makes aluminium chloride to have a lower melting point.

On the other hand, the chloride ion is *not polarized* by the sodium ion forming a highly *ionic character*. This makes sodium chloride to have a higher melting point.

2. Low solubility of some ionic compounds in water, a polar solvent and soluble in organic (non-polar) solvents such as benzene.

In terms of solubility of chemical substances, like dissolves like. i.e. compounds with an ionic character dissolve in polar (ionic) solvents such as water but those with a covalent character dissolve in non-polar organic solvents.

e.g. Aluminium chloride is moderately soluble in water and also soluble in benzene.

However, sodium chloride is very soluble in water and not soluble in benzene.

Explanation: The aluminium ion has a *higher ionic charge* and *smaller ionic radius* unlike the sodium ion. This gives the aluminium ions a *higher charge density* and hence a *higher polarizing power*. Since the chloride ions are relatively large, they have a high polarizability and are consequently *easily polarized* by the aluminium ion forming a predominant *covalent character*.

Thus, due to the *less ionic character*, aluminium chloride can moderately dissolve in water which is a *polar solvent*. However, due to the *dominant covalent character*, aluminium chloride is able to dissolve in organic solvents which are *non-polar*.

On the other hand, the sodium ion *cannot polarize* the chloride ion forming a *highly ionic character* and therefore can readily dissolve in water but does not dissolve in benzene.

3. Low thermal stability of some salts. e.g. Beryllium carbonate, $BeCO_3$ decomposes at room temperature while calcium carbonate, $CaCO_3$ only decomposes on strong heating.

Explanation: The beryllium ion being a *smaller cation* as compared to the calcium ion, it has a *higher charge density* and therefore a *higher polarizing power*.

The beryllium ion thus *strongly polarizes* the large carbonate ion establishing a *dominantly covalent character*. Consequently, the *ionic bond (electrostatic forces of attraction are) is weaker* and are easily overcome by heat such that beryllium carbonate is easily decomposed at room temperature.

On the other hand, the calcium ion *cannot easily polarize* the large carbonate ion leading to formation of a dominant *ionic bond*. The electrovalent bonds are held by *strong electrostatic forces of attraction* which are not easily overcome by heat and thus decomposes on strong heating.

4. Easy hydrolysis of some salts e.g. $AlBr_3$, $CrCl_3$, $ZnCl_2$, $FeCl_3$, etc.

Hydrolysis is a chemical reaction of any chemical species with water molecules.

Thus, aqueous solutions of the above salts contain hydrogen ions which make them to become acidic and hence having a pH less than 7.

e.g. An aqueous solution of aluminium bromide for illustration.

Explanation: The aluminium ions have a *high ionic charge* and *small ionic radius* that enables them to have a *high charge density*. Hence in an aqueous solution, the aluminium ions attract lone pairs of electrons on the oxygen atom of the water molecules to establish coordinate bonds and become *heavily hydrated*. Thus, the aluminium ions in aqueous solution exist as hexaaquaaluminium ions.

i.e. $Al^{3+}_{(aq)} + 6H_2O_{(l)} \longrightarrow [Al(H_2O)_6]^{3+}_{(aq)}$

Due to the *high polarizing power* of the aluminium ions, the oxygen to hydrogen bond in the water molecules is weakened while the aluminium to oxygen bond is strengthened and consequently *hydrolysis takes place* producing *hydrogen ions* that make the solution acidic.

i.e. $[Al(H_2O)_6]^{3+}_{(aq)} \longrightarrow [Al(OH)_3(H_2O)_3]_{(aq)} + 3H^{+}_{(aq)}$

or $[Al(H_2O)_6]^{3+}_{(aq)} + 3H_2O_{(l)} \longrightarrow [Al(OH)_3(H_2O)_3]_{(aq)} + 3H_3O^{+}_{(aq)}$

5. High hydration energies of small and / or highly charged cations in ionic compounds.

Hydration energy is the energy liberated when one mole of gaseous ions dissolves in water (is completely surrounded by water molecules) to form an infinitely dilute solution at standard conditions.

Explanation: Small and / or highly charged cations have higher charge densities as compared to large cations. Therefore, small and / or highly charged cations have a strong attraction for water molecules and become easily hydrated. This means that small cations attract water molecules more strongly than large cations. Hence more energy is released / evolved during the hydration of small cations leading to high *hydration* energy.

6. High lattice energies for small and / or highly charged ions in ionic compounds.

Lattice energy is the heat energy given out when one mole of solid ionic compounds is formed from its constituent gaseous ions at standard conditions. Its value is negative (exothermic).

Alternatively, this is the heat energy absorbed when one mole of solid ionic compounds is decomposed (broken down) into its constituent gaseous ions at standard conditions. Its value is positive (endothermic).

Explanation: Small and / or highly charged cations have higher charge densities than large cations. On the hand, small anions have low a polarizability. Therefore, the small and / or highly charged cations have a strong attraction for the oppositely charged negative ions. This means that small ions attract each other more strongly than large ions. Thus, in small oppositely charged ions, more energy is either liberated when one mole of ionic compounds is formed from the gaseous ions or absorbed when one mole of ionic compounds is broken down to form the gaseous ions.

Conclusion:

Ionic bonding is essential in forming stable compounds used in daily life. Understanding its principles helps explain why certain materials behave the way they do in different environments.

Sample items:

1. Two students are testing the electrical conductivity of different substances. They dissolve **table salt (NaCl)** in water and test its conductivity using a circuit. They also test solid sodium chloride using the same circuit.

Task: As a chemistry learner, help the two students to predict the results of the experiment and then explain why the substance behaves differently in solid and dissolved states.

Response:

Solid sodium chloride does **not** conduct electricity because the ions are locked or fixed as they are closely packed in a rigid lattice structure. This is as a result of the strong ionic bond (forces of attraction between the ions).

Dissolved sodium chloride conducts electricity because the strong ionic bond is broken by the water molecules. Therefore, the ions (Na^+ and Cl^-) are free to move in solution, allowing charge to flow.

2. A chef heats a pot of water containing dissolved table salt. The chef notices that while the water starts to evaporate, the salt remains.

Task: As a chemistry student, help the chef to understand why sodium chloride was unable to evaporate with water and what this tells us about the melting and boiling points of ionic compounds.

Response:

- Sodium chloride has a **high melting and boiling point** due to strong electrostatic forces between Na^+ and Cl^-.
- Water molecules evaporate easily because they have **weaker intermolecular forces** compared to the strong **ionic bonds** in salt.

Item 3: A student is given two elements: **Potassium (K) with one valence electron** and **Oxygen (O) with six valence electrons**.

- Predict and describe the type of bond that will form between these elements.
- Determine the formula of the resulting compound.

Response:

- Potassium will **lose one electron**, forming potassium ion, **K^+**.
- Oxygen needs **two electrons** to complete its octet, so it will gain **two electrons**, forming the oxide ion, **O^{2-}**.
- Two **K^+ ions** will combine with one **O^{2-} ion**, forming **K_2O (potassium oxide)**.

Item 4: A doctor prescribes potassium chloride (KCl) supplements to a patient with a deficiency.

- Explain why potassium ions are important for the human body.
- How does potassium chloride dissolve in the body to help maintain cell function?

Response:

- **K^+ ions** are essential for **nerve signal transmission** and **muscle contraction**.
- Potassium chloride dissolves in body fluids, breaking into **K^+ and Cl^- ions**, which are absorbed by cells and help maintain **electrolyte balance**.

Assignment Items:

1: Observing Chemical Properties in a School Laboratory.

During a chemistry practical lesson at a secondary school, the teacher gives learners samples of different compounds to investigate their physical properties. Two groups of learners heat small quantities of **aluminium chloride, calcium chloride, and aluminium oxide** using controlled laboratory equipment.

Group A observes that **aluminium chloride melts much more easily when heated**, while **calcium chloride requires a higher temperature before it melts**. This surprises the learners because aluminium and calcium are both metals and the compounds look similar in appearance.

Group B later heats **aluminium oxide** and **aluminium chloride** under the same conditions. They observe that **aluminium oxide does not melt easily and requires a very high temperature**, while **aluminium chloride melts relatively quickly**.

Task: As a student who understands chemical bonding and the structure of substances:

(a) Help the learners explain why **aluminium chloride has a lower melting point than calcium chloride**.

(b) Explain to the class why **aluminium oxide has a much higher melting point than aluminium chloride**, even though both compounds contain aluminium.

2: Understanding Ionic Bonding in Everyday Chemistry.

During a science club meeting at your school, students are discussing how different substances are formed through chemical bonding. The teacher brings examples of compounds such as **magnesium oxide and sodium chloride** and asks learners to relate the concept of bonding to substances they encounter in daily life.

Task: As a chemistry learner, you are asked to guide the discussion by explaining how **ionic compounds form and why they have certain properties**.

(a) In your own words, explain **what is meant by ionic bonding**.

(b) Using magnesium and oxygen as examples, explain **how an ionic bond is formed between these two elements**.

(c) The teacher asks the class to identify characteristics of ionic substances.
List **three properties of ionic compounds** and explain **why ionic compounds show those properties**.
(d) Ionic compounds are important in many aspects of daily life in Uganda.
Give **two real-life applications of ionic compounds** and explain how they are useful in those situations.

METALLIC BONDING

General objectives; By the end of the subtopic, one should be able to:

- Define metallic bonding.
- Explain the formation of metallic bonds using suitable examples.
- Describe the properties of metallic substances.
- Relate metallic bonding to real-life scenarios.

Introduction to Metallic Bonding

In the world around us, from the strength of bridges and buildings, to the shine of jewelry and to the wires that power our homes, metals play an essential role in shaping modern life. But what gives metals their unique properties such as conductivity, strength, malleability, and luster? The answer lies in a special type of chemical bonding known as **metallic bonding**.

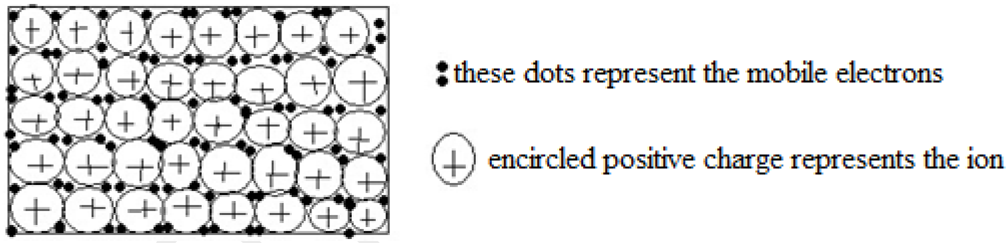
Metallic bonding is the type of bonding that exists in all metallic elements.

Metals generally have larger atomic radii than non-metals which increases the distance between the nuclei of metal atoms and their outermost (valence) electrons. Thus, the outermost (valence) electrons in atoms of metals are weakly attracted to their nuclei and as a result they separate from their parent atoms and begin to move randomly throughout the crystal structure which leads into formation of a pool of delocalized (mobile) electrons.

In metallic structures, each metallic atom loses its valence electrons to form cations which are attracted together by a pool (cloud) of the released delocalised electrons. These electrons are not tied to any one atom (or ion) but they move throughout the structure, creating a strong, cohesive force that binds the entire metal together.

A **metallic bond** is therefore formed as the electrostatic force of attraction between the negatively charged delocalized electrons and the positively charged nuclei of the adjacent atoms (ions) in the metallic structure.

Metallic bonds extend equally in all directions to form a giant metallic structure of a high melting point. The ions or atoms in the metal structure are assumed to be spherical and identical in size.



This remarkable arrangement explains why metals can conduct electricity, be hammered into sheets, or drawn into wires without breaking. Understanding metallic bonding is not just about learning chemistry—it is about unlocking the science behind the materials that build nations, drive industries, and transform ideas into reality. As we explore this topic, think of metallic bonding as the invisible force that turns simple atoms into powerful materials that support innovation, development, and progress in Uganda and beyond.

Factors that determine the strength of a metallic bond

The strength of the metallic bond is affected or determined by:

(a) The number of delocalized electrons contributed per atom to the formation of the metallic bond. The higher the number of delocalized electrons contributed per metallic atom, the stronger the metallic bond.

Metals with many electrons in their outermost shells tend to contribute more delocalised electrons per atom which results into a stronger nuclear attraction for the delocalised electrons. This leads to increase in the strength of the metallic bond.

Strong metallic bond means hardness, high density, high melting point and high boiling point.

This explains the trend in the melting or boiling points of metals in the same period.
e.g. In Period 3, magnesium ($1s^22s^22p^63s^2$) is harder with also a higher melting point and boiling point than sodium ($1s^22s^22p^63s^1$).

Element	Na	Mg	Al
Melting point/°C	98	650	660
Boiling point/°C	890	1120	2450

(b) The atomic (metallic) radius. i.e. radius of the metal.
The larger the atomic radius the weaker the metallic bond. This is because the distance between the delocalised electrons and the nucleus is longer. Thus, the nuclear attraction for the delocalised electrons becomes weaker and the strength of the metallic bond reduces which requires less heat energy to be weakened or broken.
This explains mainly the trend in melting or boiling points of metals in the same group.
e.g. In Group(I), the melting and boiling points of the metals decrease down the group.

Elements	Li	Na	K	Rb	Cs
Melting point (°C)	180.5	98	63.5	39.3	28.44
Boiling point (°C)	1330	882.8	758.8	688	670.8

General Characteristics of Metals

(a) Metals are made up of atoms (or cations) and delocalised electrons.
This is because, metals atoms **lose their outer (valence) electrons** easily and the atoms become **positive ions (cations)** arranged in a regular lattice.
The lost electrons don't belong to any one atom and so they move freely throughout the structure. These are called **delocalised electrons** often called a **"sea of electrons"** surrounding positive ions.

👉 So, a metal is basically **positive ions + a "sea" of mobile electrons held together by electrostatic force of attraction (metallic bonding).**

(b) They conduct electricity and heat (in solid or molten state).
This is because, the **delocalised electrons are free to move** thus, when a voltage is applied, these moving electrons carry with them an electric charge hence leading to flow of **electric current**.
Unlike ionic compounds, metals **conduct even when in the solid state** because electrons are already free.

👉 For heat; Electrons quickly transfer energy by moving and colliding with other particles.
This makes metals **excellent conductors of heat**.

(c) They are usually solids at room temperature.
This is because, the **metallic bonds are strong** (electrostatic forces of attraction between positive ions and delocalised electrons). These forces **hold particles tightly in fixed positions**, forming a solid structure.
The atoms are also **closely packed**, increasing the metallic strength.

👉 Exception is **Mercury** which is a liquid because its metallic bonds are relatively weak.

(d) They have high melting and boiling points.
This is because, to melt or boil a metal, the **strong metallic bonds** must be broken or weakened. Since these bonds involve many electrons and ions, they are strong and require **a lot of energy** to break.

👉 Stronger bonding = **higher melting/boiling point**.

(e) They are good conductors of heat.
This is because, when one part of a metal is heated, the delocalised electrons gain more energy and move faster.
These electrons transfer this energy quickly across the metal.
The metal atoms (ions) also vibrate faster helping pass heat along as well.

👉 Result: heat spreads **very fast** through metals.

(f) They are shiny (lustrous) appearance.
This is because, when light hits a metal; the **free electrons absorb energy** and become excited.
These electrons quickly fall back to lower energy levels and **re-emit the energy** in form of light energy which is what is seen as a **reflection**.

👉 Because electrons are free and mobile, metals; they reflect most visible light and appear **shiny or polished.**

(g) They have a high density.

This is because, metal atoms are; **Closely packed** and; **Arranged in compact structures** (like layers). This means a lot of mass is packed into a small volume.

👉 More particles per unit volume = **high density**.

(h) They are malleable (can be hammered into sheets**) and ductile (**can be drawn into wires**).**

This is because the layers of metal ions can **slide over each other**. However, the "sea of electrons" keeps holding the structure together even when layers move.

👉 Unlike ionic compounds; Metals don't shatter when layers shift and the metallic bonds don't break easily during movement.

Applications of Metallic Bonding:

1. Cooking Utensils. They are made of aluminum or stainless steel.
Good heat conductors due to free electrons transferring heat evenly.
2. Electric Wiring. Copper wires used in homes and electronics. This is due to the fact that copper is ductile and also contains delocalised electrons that allow smooth flow of electric current.
3. Car Bodies and Planes. Use of aluminum for light weight and strength.
Strong metallic bonds hold atoms firmly together.
4. Metal Coins and Jewelry. Use of metals like silver, gold, and copper. This is because they are malleable and shiny due to metallic bonding.
5. Steel Bridges and Skyscrapers. Iron or steel's strong metallic bonds provide strength and durability.
6. Iron bars, nails. Used in construction. Iron metal has strong metallic bonds that give it strength.
7. Roofing sheets. Iron metal is malleable.

Impact of Ionic and Metallic Compounds on Life and the Environment and Their Mitigations

Ionic and metallic compounds are widely used in industries, agriculture, medicine, and household applications. However, their presence in the environment and their interaction with living organisms can have severe negative consequences.

1. Impact of Ionic Compounds and The Possible Mitigation.

(a) Soil and Water Pollution.

(i) Ammonium nitrate (NH_4NO_3) used in fertilizers dissolves in rainwater and leaches into nearby rivers and lakes. High nitrate levels in drinking water cause "blue baby syndrome" (methemoglobinemia), which reduces oxygen transport in infants. It also causes eutrophication, leading to oxygen depletion, fish kills, and loss of aquatic biodiversity.

Mitigation: Use controlled-release fertilizers, apply fertilizers in the right amounts, and encourage organic farming to reduce nitrate leaching.

(ii) Sodium chloride (NaCl) from road salt accumulates in the soil. High soil salinity reduces crop productivity, affecting food security. This disrupts soil microbial activity, leading to decreased soil fertility and desertification.

Mitigation: Use alternative de-icing agents like calcium magnesium acetate and improve drainage systems to reduce salt accumulation.

(b) Toxicity and Health Hazards.

(i) Lead(II) nitrate, $Pb(NO_3)_2$ in old plumbing pipes dissolves into drinking water. This causes lead poisoning leading to brain damage, kidney failure and developmental problems in children. It also contaminates water bodies, making them unsafe for aquatic life.

Mitigation: Replace lead pipes with safer alternatives that do not corrode easily (e.g., PVC or copper pipes), and implement strict regulations on lead-containing materials.

(ii) Mercury chloride ($HgCl_2$) used in industrial processes is highly toxic. Mercury poisoning causes neurological disorders, birth defects, and immune system damage. It also causes bioaccumulates in fish, affecting food chains and causing long-term ecological damage. *Mitigation:* Ban mercury in non-essential applications, promote safe disposal of mercury containing waste, and use non-toxic alternatives.

(c) Acid Rain Formation.

(i) Sulphuric acid (H_2SO_4) is formed when sulphur dioxide (SO_2) from industrial emissions dissolves in rainwater. This causes respiratory problems, skin irritation, and damage to the eyes and acidifies soil, leaches nutrients from the ground, and damages plant life.

Mitigation: Reduce sulphurdioxide emissions by installing scrubbers in industries, and switch to cleaner energy sources like wind and solar.

(ii) Nitric acid (HNO_3) from vehicle exhausts contributes to acid rain. This aggravates asthma and other lung diseases.

It also corrodes buildings, statues, and bridges made of limestone and marble.

Mitigation: Use catalytic converters in vehicles and enforce stricter emission regulations.

(c) Corrosion and Structural Damage.

Sodium chloride (NaCl) used in de-icing roads accelerates rusting of metal structures. This weakens bridges and buildings, increasing the risk of accidents and leads to increased maintenance costs and loss of historical monuments.

Mitigation: Use alternative de-icing agents and apply protective coatings on metal structures.

2. Impact of Metallic Compounds and The Possible Mitigation.

(a) Heavy Metal Toxicity.

Cadmium (Cd) in batteries and pigments leaches into soil and water. Long term exposure leads to lung cancer, kidney disease, and osteoporosis and pollutes groundwater, making it unsafe for drinking.

Mitigation: Promote battery recycling programs and restrict the use of cadmium in consumer products.

(b) Metal Corrosion and Deterioration.

Iron (Fe) in bridges and pipelines rusts when exposed to moisture and oxygen. This weakens structures, increasing the risk of collapse and accidents and increases metal waste, requiring more energy for replacement and recycling.

Mitigation: Use galvanized or stainless steel, and apply protective coatings like paint or oil.

(c) Soil Contamination.

(i) Lead (Pb) from electronic waste (e-waste) accumulates in soil. Lead poisoning from contaminated food leads to nervous system disorders and reduced IQ in children and makes soil unsuitable for agriculture, reducing food production.

Mitigation: Encourage e-waste recycling, and use biodegradable alternatives for electronics.

(ii) Copper (Cu) from corroded pipes and wiring leaches into the ground. High copper intake causes liver and kidney damage and disrupts microbial activity in soil, reducing fertility.

Mitigation: Use PVC or composite materials for pipes and recycle copper materials efficiently.

(c) Greenhouse Gas Emissions from Metal Reactions.

Aluminum (Al) and magnesium (Mg) combustion releases fine metal particles into the air Causing respiratory diseases and lung inflammation and contributes to global warming and deteriorates air quality.

Mitigation: Improve industrial air filtration systems and develop alternative production methods that reduce emissions.

Note; Ionic and metallic compounds, while essential in various applications, pose significant risks to human health and the environment. Their negative effects, such as toxicity, pollution, corrosion, and greenhouse gas emissions, can be minimized through responsible usage, recycling, emission controls, and material substitutions. Sustainable management strategies are essential to ensure that their benefits outweigh their harmful impacts.

COVALENT BONDING

General objectives; By the end of the subtopic, one should be able to:

- Define the term covalent bonding.
- Explain the formation of covalent bonds using suitable examples.
- Describe the properties of covalent compounds.
- Relate covalent bonding to real-life scenarios.

Introduction to Covalent Bonding

Imagine a world where atoms don't just exist alone, but *choose* to come together by sharing what they have in order to become more stable, more complete, and more powerful. This is the story of **covalent bonding**.

In chemistry, not all atoms win by taking or losing electrons. Some succeed through cooperation.

These atoms **share electrons**, forming bonds that hold molecules together e.g. from the oxygen we breathe to the water that sustains life.

Covalent bonding is the foundation of life itself. Every living organism is built on compounds like proteins, carbohydrates, and DNA which are all formed through this special type of bonding.

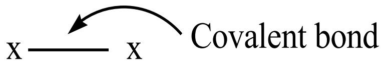
This teaches us a powerful lesson; sometimes, strength doesn't come from dominance, but from **sharing and balance**.

As we begin this topic, think of atoms not as tiny particles, but as partners in a system that is working together to achieve stability. By understanding covalent bonding, the secrets behind the structure of matter and the chemistry of life are unlocked.

A covalent bond *is a chemical bond formed when two electronegative atoms (usually nonmetallic atoms) combine together by sharing of electrons.*

Since the nonmetallic atoms have a tendency to gain electrons, the noble gas configuration is attained by sharing of their unpaired electrons.

i.e. Because both atoms have the affinity for electrons and neither has a tendency to donate them, they share electrons in order to achieve octet configuration and become more stable.

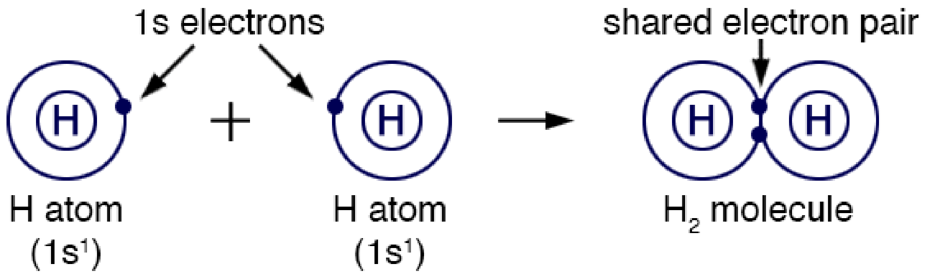
A covalent bond is represented by a short straight line. i.e. 

Covalent bonds are formed when non-metal atoms share valence electrons to achieve a stable, full outer shell. Common visual examples include dot-and-cross diagrams (showing shared electrons) and Lewis structures (using dots/sticks for bonds), typically involving molecules like

Key Examples of Covalent Bond Formation:

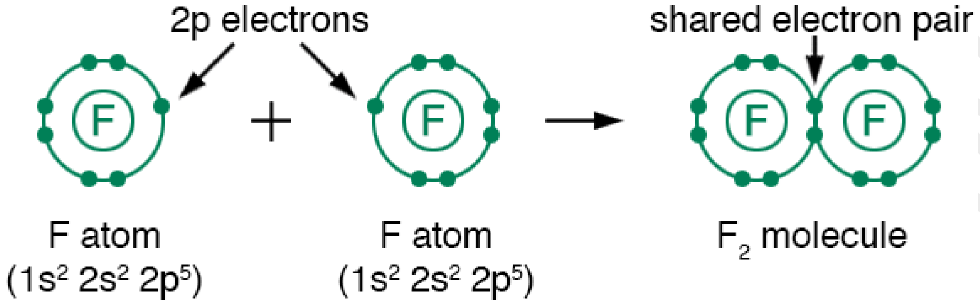
(a) Hydrogen molecule (H_2 with a single covalent bond): It is formed when two hydrogen atoms share one electron in their valence electron shell ($1s^1$) with each other. This allows each hydrogen atom to achieve a full electron shell with the electron configuration of a noble gas, He:$1s^2$.

i.e. Each hydrogen atom (with 1 electron) shares its electron with another hydrogen atom, forming a single covalent bond and achieving a stable pair.

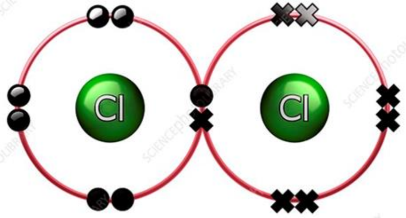


(b) Fluorine molecule (F_2 with a single covalent bond): It is formed when two fluorine atoms share a pair of electrons in their valence electron shell ($2s^22p^5$). This allows each fluorine atom to achieve a full electron shell with the electron configuration of a noble gas, Ne: $1s^22s^22p^6$.

i.e. Each fluorine atom (with 1 electron) shares its electron with another fluorine atom, forming a single covalent bond and achieving a stable pair.



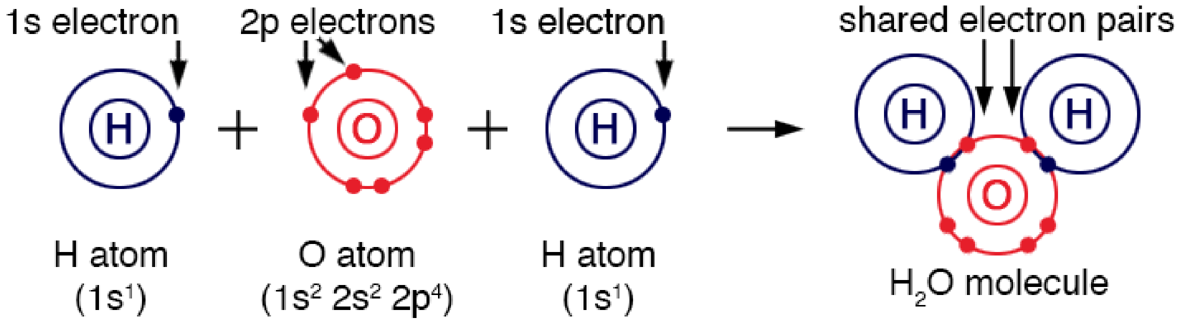
Similarly, each chlorine atom in the chlorine molecule **(Cl_2 with a single covalent bond)** achieves a full electron shell with the electron configuration of a noble gas, Ar: $1s^22s^22p^63s^23p^6$.



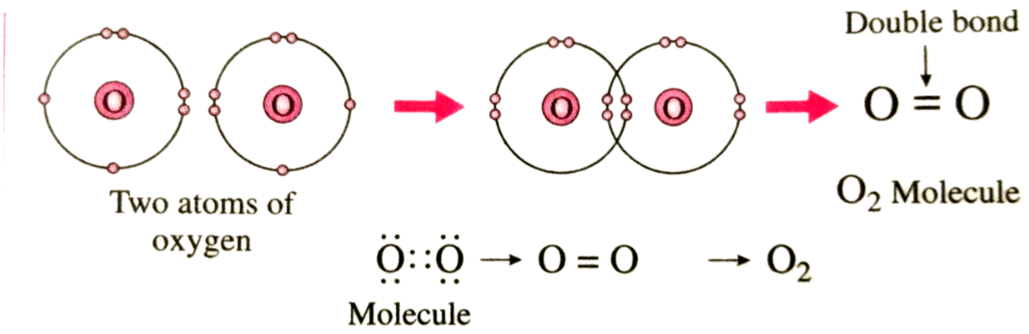
(c) Water Molecule (H_2O with polar single covalent bonds):

Water, H_2O, is formed when two hydrogen atoms share one electron in their valence electron shell ($1s^1$) with an oxygen atom. The oxygen atom shares one electron in their valence electron shell ($2s^22p^4$) with each hydrogen atom. This allows each hydrogen atom to achieve a full electron shell with the electron configuration of a noble gas, He:$1s^2$. It also allows the oxygen atom to achieve a full electron shell with the electron configuration of a noble gas, Ne: $1s^22s^22p^6$.

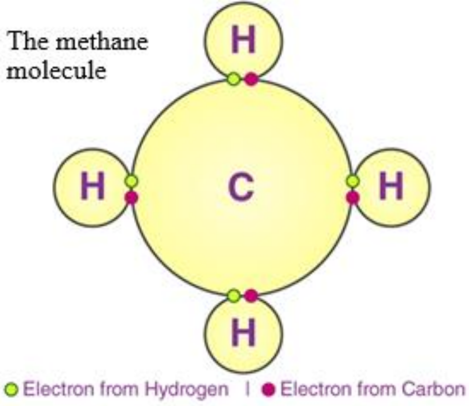
i.e. One oxygen atom shares two pairs of electrons with two hydrogen atoms. Oxygen's 6 valence electrons and two hydrogen atoms share 1 electron each, creating two single bonds.



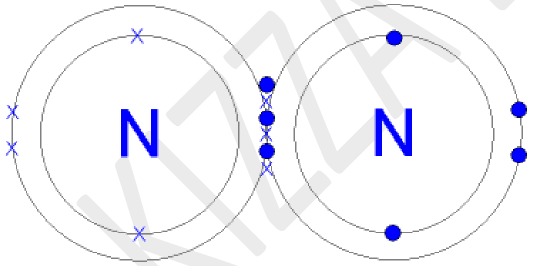
(d) Oxygen Molecule (O_2 with a double covalent bond): Two oxygen atoms, each needing two electrons, share two pairs (4 electrons total), creating a double covalent bond.



(e) Methane Molecule (CH_4 with single covalent bonds): A central carbon atom shares its 4 valence electrons with 4 hydrogen atoms, showing 4 distinct single covalent bonds.



(f) Nitrogen Molecule (N_2 with a triple covalent bond): Two nitrogen atoms share three pairs of electrons (6 electrons total), resulting in a strong triple covalent bond.
By sharing the six electrons where the shells touch each nitrogen atom can count 8 electrons in its **outer shell**. These full outer shells with their shared electrons are **now stable**.



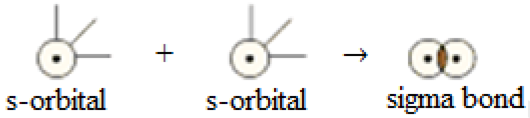
Sample item:
At Namanve Industrial Area, a group of Senior Two learners visit a factory during their science field study. While there, they observe different processes such as fuel burning, gas production, and chemical reactions in large tanks.
Their teacher explains that many of the substances they see are made from tiny particles called atoms, which combine to form molecules. The learners become curious about how these molecules are formed.
Back at school, they are given an assignment to investigate and illustrate how different molecules are formed from atoms.
Task: As a chemistry student, use diagrams for illustration to explain to the S.2 learners how the following molecules are formed:
(a) Phosphorus trichloride.
(b) Ammonia.
(c) Carbon tetrachloride.
(d) Hydrogen chloride.
(e) Carbon dioxide.
(f) Carbon monoxide.
(g) Hydrogen cyanide.
(h) Sulphur dioxide.
(i) Sulphur trioxide.

Note: In this case, each atom involved in bonding provides an electron to form the shared electron pair which constitutes or makes up the single covalent bond.
If two or three electron pairs are shared between two bonding atoms, a double or triple covalent bond is established respectively. i.e multiple covalent bonds are formed.

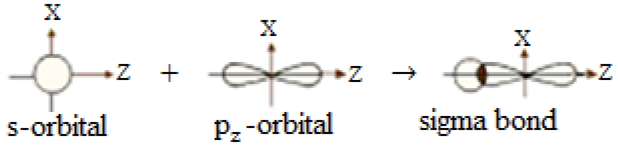
The Orbital Approach to the Types of Covalent Bonding

The **orbital approach to covalent bonding** explains how atoms form covalent bonds by overlapping their atomic orbitals and sharing electrons. This gives a more detailed picture than simple "electron dot" models.

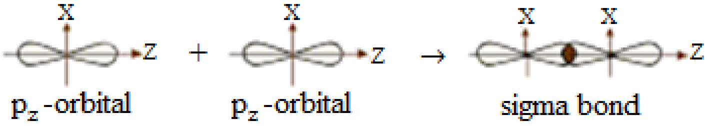
The Atomic Orbitals and Overlap: Atoms have regions in space where electrons are likely to be found, called **orbitals** (like s, p, d). When two atoms come close, their orbitals **overlap** which allows electrons to be **shared**, forming a covalent bond. The stronger the overlap, the stronger the covalent bond.
A **molecular orbital** *is an overlap of two atomic orbitals each containing an electron in which there is a high probability of finding the pair of bonding electrons of a covalent bond.*
Hybridization *is the fusing (mixing or overlapping) of atomic orbitals of different energies and shapes suitable for the pairing of valence electrons to form covalent bonds with new hybrid orbitals of equal energy.*
When each one of two different orbitals has an unpaired electron and the two orbitals combine, they form a molecular orbital with two electrons.
Because of the overlap, it is highly probable that a pair of electrons are found in the physical region or space where the orbitals overlap. The types of covalent bonding based on orbital approach are shown below:
(a) The **sigma (σ) bond**. This is formed by the head-to-head (linear) overlap of atomic orbitals (i.e. overlap between the nuclei of two atoms). These sigma bonds are formed in all single covalent bonds.
The sigma bond *is a covalent bond formed by the overlapping of two s-orbitals or two p-orbitals or an s-orbital and p-orbitals, whereby the electron density of the two bonding electrons is highly concentrated in the region between the two atoms.* This makes a sigma bond strong as illustrated by the:
(i) s - s overlap.



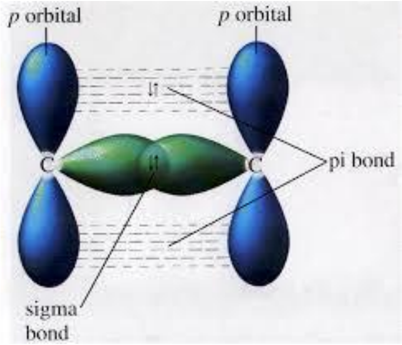
(ii) s - p (s - p_z) overlap.

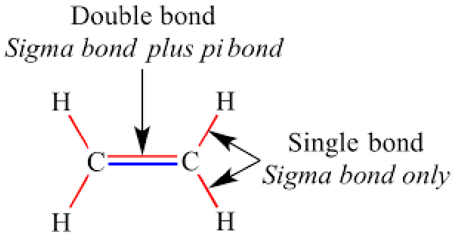


(iii) p - p (p_z - p_z) overlap.



(b) The **pi (π) bond.** This is formed by the side way overlap of atomic p-orbitals (i.e. two atomic orbitals overlap the space between the nuclei).
The pi bond *is a covalent bond formed by overlapping of two p-orbitals each containing one electron of opposite spin whereby the electron density is concentrated on two opposite sides of the inter nuclear axis perpendicular to the sigma bond.*
The sigma and pi bonds are illustrated in the diagram below:





Overlapping of two p-orbital electrons perpendicular to a sigma bond forms a pi-bond.
This implies that a pi bond is thus a second bond formed after formation of a sigma bond.
The pi bond has a lower electron density along the axis that joins the two nuclei than a sigma bond and this shows that a pi bond is weaker than a sigma bond.
However, the presence of pi bonds together with a sigma bond, results into formation of multiple (double and triple) covalent bonds. The pi bonds therefore exist in multiple covalent bonds.
These pi bonds further strengthen the force of attraction between atoms and this explains why multiple covalent bonds are shorter and stronger than single covalent bonds.

Types of Covalent Bonds

(a) A single covalent bond is one that takes place between two atoms by sharing of one pair of electrons forming one sigma bond. It is represented by a short straight line between the atoms.
When two electrons are shared between two atoms, a **single covalent bond** forms.
(b) A double covalent bond is one that occurs between two atoms by the sharing of two pairs of electrons forming one sigma bond and one pi bond.
Multiple covalent bonds are formed between two atoms when they share more than one pair of electrons to acquire a stable valence shell. For instance, in carbondioxide, CO_2 the carbon atom forms two covalent bonds (two pairs of electrons or four electrons) with each oxygen atom involved.

$O=C=O$

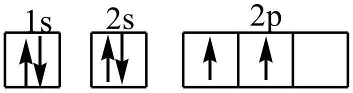
These are called **double covalent bonds**. Double bonds are stronger than single bonds.
(c) A triple covalent bond is one that occurs between two atoms by the sharing of three pairs of electrons forming one sigma bond and two pi bonds. Hence the pi bond electron cloud in a triple bond is larger than that in a double bond.
In the same way, when six electrons are shared between two atoms, a **triple covalent bond** forms.
For instance, in ethyne is an example of a molecular compound that forms a triple bond. One carbon atom shares six electrons with another carbon atom to complete the octet. These types of bonds are stronger than double bonds.

$HC\equiv CH$

Hybridization Concept:

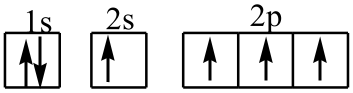
To explain molecular shapes, atomic orbitals mix to form **hybrid orbitals**.
The types of bonds can be studied using the **hybridization in carbon** as shown below**:**
Carbon has an electronic structure of $1s^2 2s^2 2p^2$ with two unpaired electrons in its *ground state*.

Carbon atom, in the ground state



However, when it combines with another atom, excitation takes place as the carbon atom acquires sufficient energy to cause unpairing of the 2s orbital electrons, $1s^2 2s^1 2p^3$. The carbon atom attains four unpaired electrons in its excited state.

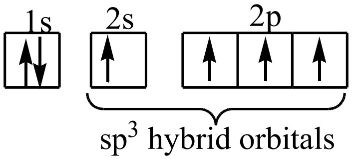
Carbon atom, in the excited state



Types of Hybridization:

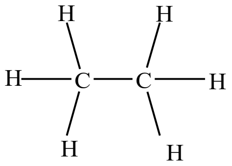
(a) The **sp^3 (tetrahedral) hybridization.** The combination of the one electron in the 2s orbital and three electrons in the 2p atomic orbitals results in formation of four coplanar hybridized atomic orbitals (sp^3 hybrid) which enclose angles of 109^{0}28'.

Carbon atom, in the excited state



This accounts for the tetrahedral arrangement of bonds. **Single covalent bonds** in carbon are formed.

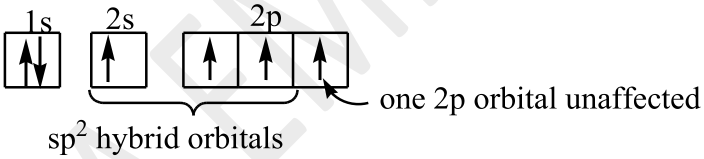
e.g. In the ethane molecule where both carbon atoms use the sp^3 hybridized atomic orbitals each in bonding. The carbon atoms bond to each other and to six hydrogen atoms to form seven sigma bonds.



All valence electrons in each carbon atom are utilized in sigma bond formation.

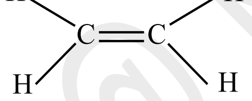
(b) The **sp^2 (trigonal) hybridization.** The combination of the 2s and two 2p atomic orbitals results in formation of three coplanar hybridized atomic orbitals (sp^2 hybrid) which enclose angles of 120^0. This accounts for the trigonal planar arrangement of bonds leaving a single occupied 2p orbital unaffected.

Carbon atom,in the excited state



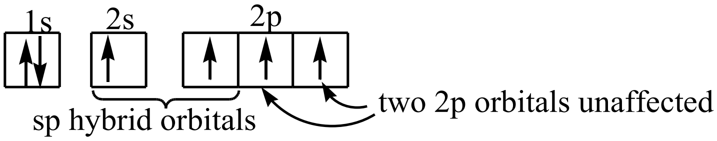
This leads to formation of **double covalent bonds** in carbon.

e.g. In ethene where each carbon atom uses the sp^2 hybridized atomic orbitals in bonding. The carbon atoms bond to each other and to four hydrogen atoms to form five sigma bonds. This leaves the 2p$_z$ atomic orbital (1 electron) on each carbon atom directed at right angles to the plane of the molecule which overlap laterally in a pair to form one pi bond. The pi bond and the sigma bond between the two carbon atoms means that the two carbon atoms are bound more tightly than in alkanes like ethane.



(c) The **sp (linear) hybridization.** The combination of the 2s and one 2p atomic orbitals results in formation of two hybridized atomic orbitals (sp hybrid) which are directed at an angle of 180^0 to each other. This accounts for the linear arrangement of bonds leaving two 2p orbitals unaffected.

Carbon atom, in the excited state



This leads to formation of **triple covalent bonds** in carbon.

e.g. In ethyne where each carbon atom uses sp hybridized atomic orbitals in bonding. The carbon atoms bond to one another and to two hydrogen atoms to form three sigma bonds. This leaves the 2p$_y$ and 2p$_z$ atomic orbitals (2 electrons) on each carbon atom and these overlap in pairs to form two pi bonds. The two pi bonds and the sigma bond between the two carbon atoms means that the two carbon atoms are bound more tightly than in alkenes like ethene.

$H-C\equiv C-H$

🧩 **Summary:** The orbital approach shows that:

- Covalent bonds are formed by **overlap of orbitals.**
- There are sigma **(σ) and** pi **(π) bonds.**
- Type of bond depends on **how orbitals overlap.**
- Hybridization explains the **geometry of molecules.**

🧠**Why the Orbital Approach Matters.**

- Explains **bond strength and length.**
- Predicts **molecular shapes.**
- Clarifies **reactivity and polarity.**
- Helps understand advanced concepts like resonance.

Bond Length and Bond Strength of Covalent Bonds

Covalent bond length is the distance between nuclei of two covalently bonded atoms.

Covalent bond strength is the measure of the ability of a covalent bond to withstand bond breakage.

Factors that affect bond length include:

(a) Atomic (Covalent) radius.

The larger the atomic radius, the longer the covalent bond. Hence the weaker the bond and the lower the bond energy.

This because, when atomic radius increases, the effective nuclear charge becomes weaker thus the nuclear attraction for the electrons shared between the atoms reduces. This increases the spreading out of the bonded electrons and the bond length increases. Hence the covalent bond becomes weaker and the lower the bond energy. e.g. the bond strength is in the order; Cl-Cl > Br-Br > I-I

(b) The number of bonding electrons i.e number of electrons involved in formation of the covalent bond.

The higher the number of electrons in a covalent bond, the shorter the covalent bond. Hence the stronger the bond and the higher the bond energy.

This is because although the nuclei of the covalently bonded atoms repel one another due to the fact that both are positively charged, each nucleus on the other hand is attracted to the electrons shared between the bonded atoms. Thus, the higher the number of shared electrons, the stronger their attraction from the nuclei of the covalently bonded atoms and the closer the nuclei approach each other.

Therefore, the stronger the covalent bond and the higher the bond energy. e.g. the bond strength of carbon-to-carbon triple, double and single covalent bonds are in the order; C≡C > C=C > C-C respectively.

Type of covalent bond	*Bond length (nm)*	*Bond energy (kJmol^{-1})*
Carbon to carbon single bond	0.154	348
Carbon to carbon double bond	0.134	615
Carbon to carbon triple bond	0.120	835

(c) Electronegativity difference between the two covalently bonded atoms i.e. Bond polarity.

The greater the difference in electronegativity between the two covalently bonded atoms, the more polar the covalent bond. This increases the ionic character of the bond and in turn increases the bond strength. e.g. the bond strength is in the order; H-F > O-H > N-H.

However, the covalent bond becomes weaker if the polarity is shared out between a greater number of bonds. e.g. the phosphorus to chlorine bond is weaker in phosphorus(V) chloride than in phosphorus(III) chloride.

N.B: This factor does not usually outweigh the factor of atomic radius.

Polarization in Covalent Bonds (Bond Polarity)

When a covalent bond is formed between two similar non-metallic atoms, the electrons in the covalent bond are shared equally. However, this is not true when a covalent bond is formed between two different types of non-metallic atoms. In this case, the more electronegative atom attracts bonding electrons more towards itself. This creates a partial negative charge on the more electronegative atom because there is greater amount of the negatively charged electrons around it and on the other hand, a partial positive charge is created on the less electronegative element because bonding electrons have moved away from it. A combination of the partial positive charge and the partial negative charge on the two non-metals in the covalent bond is termed as a dipole within the molecule.

A **dipole** is a covalent bond containing an atom with a partial positive charge and another atom with a partial negative charge due to their difference in electronegativities. The covalent bond becomes polar.

Hence a **polar covalent bond** is a covalent bond with a dipole and has a significant ionic character. This means that the two shared electrons are closer to one of the atoms than the other, creating an imbalance of charge. Dipoles give rise to dipole-dipole interactions between molecules also known as intermolecular forces of attraction. The electronegativity difference between the two atoms in these covalent bonds to have a significant ionic character ranges from 0.3 to 1.7.

Consider a covalent bond formed between atoms A and B. If A is more electronegative than B, it will withdraw electrons from B. Atom A will then acquire a partial negative charge leaving atom B with a partial positive charge. This separation of charge constitutes bond polarity and the two polar atoms are referred to as a dipole.

Note that; when individual bonds within a molecule are polar, the molecule as a whole can however be non-polar. This is because the bond pairs are symmetrically arranged and the dipole moments cancel out. i.e if the centre of all the negative charges and the centre of all the positive charges coincide and hence cancel out. i.e. dipole moment cancels out (the mathematical product of the separation of the ends of a dipole and the magnitude of the charges). This is observed with all particles with no lone pairs at the central atom i.e those that are linear, trigonal planar, tetrahedral, octahedral, etc. Examples include:

(a) In tetrachloromethane (CCl_4), its bonds are polar but the molecule is non-polar. The carbon to chlorine bonds are polar because there is a high electronegativity difference between the chlorine atom and the carbon atom. (i.e. because the chlorine atom is more electronegative than the carbon atom, it attracts bonding electrons more towards itself. The chlorine atom thus acquires a partial negative charge leaving the carbon atom with a partial positive charge. This leads to the distortion of electron distribution between the carbon atom and the chlorine atom. This unequal distribution of electrons constitutes bond polarity).

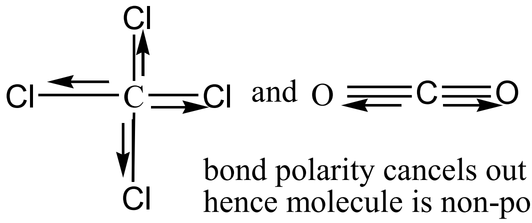
However, the due to absence of lone pairs on the carbon atom, the carbon to chlorine bonds are symmetrical as the chlorine atoms are tetrahedrally arranged around the central carbon atom. Thus, the dipole moments (bond polarity effects) cancel out each other so that the whole molecule remains non-polar.

(b) In the carbondioxide molecule, CO_2 has polar bonds but the entire molecule is non-polar.

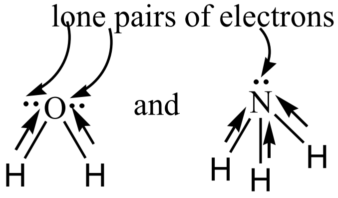
The carbon to oxygen bond is polar because the oxygen atom is more electronegative than the carbon atom. However, due to absence of lone pairs on the carbon atom, the bonds are symmetrical as the oxygen atoms are linearly arranged around the central carbon atom. Thus, bond polarity effects cancel out each other so that the whole molecule remains non-polar.

On the other hand, molecules such as water (H_2O) and ammonia (NH_3) are polar since the dipole moments do not cancel each other. The molecules have lone pairs of electrons on the central atom which repel the bond pairs strongly such that the bonded atoms are not symmetrical.

Illustrations for examples are shown below;



bond polarity cancels out
hence molecule is non-polar



bond polarity doesnt cancel out
hence molecule is polar

N.B: The tendency of bonds being polar but the entire molecule being non-polar is observed with all those molecules whose central atom lacks a lone pair and it is at the same time bonded to atoms of the same type of element.

General characteristics of covalent compounds

(a) They are made up of particles known as molecules.

Covalent compounds are formed when **non-metal atoms share electrons** to achieve stability.

The smallest independent unit of a covalent compound is called a **molecule** thus; a molecule is a group of atoms **held together by covalent bonds (shared electrons)**.

These molecules exist as **separate, discrete units**, not as a giant structure.

Examples include;

- The water (H_2O) molecule → one oxygen atom shares electrons with two hydrogen atoms
- The carbon dioxide (CO_2) molecule → one carbon atom shares electrons with two oxygen atoms

👉 This is different from ionic compounds, which form **giant lattices of ions**, not molecules.

(b) They are soluble in organic (non-polar) solvents but not in water.

Covalent compounds dissolve based on the rule: 👉 **"Like dissolves like."**

Thus, **covalent compounds** dissolve in **non-polar (organic) solvents** like benzene, ether, etc.

This is because they consist of the intermolecular Van der Waals forces of attraction which are similar.

On the other hand, **covalent compounds** are **insoluble in water** because it is a **polar (ionic) solvent**, and hence the polar water molecules cannot attract non-polar molecules strongly.

However, polar covalent molecules (those with uneven charge distribution) can dissolve in water.

Example: Hydrogen chloride (HCl) dissolves in water because it is **polar**.

(c) They do not conduct electricity (i.e. they are non-electrolytes).

Electric current flows when there are **free-moving charged particles (ions or electrons)**.

This is because the **covalent compounds** exist as **neutral molecules**, but not ions.

Therefore, they **do not conduct electricity** in solid, molten, or dissolved states and termed as **non-electrolytes**.

However, some **polar covalent compounds** (like hydrogen chloride, HCl) **ionize in water**.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

The ions then can move freely, and so the solution **conducts electricity**.

Such substances behave as **electrolytes in aqueous solution only**.

(d) They are often gases, volatile liquids, or soft (waxy) solids.

This is because the covalent compounds have **weak Van der Waals forces of attraction between molecules**.

These forces are much weaker than ionic or covalent bonds and as a result the molecules can **easily move apart** to establish a liquid or gaseous state.

Many covalent substances exist as; **Gases** (e.g., oxygen, CO_2), **Volatile liquids** (e.g., ethanol) and **Soft solids** (e.g., iodine, wax).

👉 "Volatile" means they **evaporate easily**.

(e) They have low melting and boiling points.

Melting or boiling involves **separating molecules**, but not breaking covalent bonds inside the molecule.

The covalent compounds have **intermolecular Van der Waals forces of attraction** which are **weak** and as such a small amount of energy is needed to separate molecules.

Therefore, covalent compounds have **low melting points** and have **low boiling points**

For Example; Water boils at 100°C (relatively low compared to ionic compounds like sodium chloride, which melts above 800°C)

Sample Items

1. A senior three class at a Kololo Secondary School in Uganda is carrying out a practical activity during their Chemistry lesson on molecular structure and polarity. The teacher provides two colorless liquids labeled A and B. The learners are told that liquid A is trichloromethane ($CHCl_3$) and liquid B is tetrachloromethane (CCl_4).

During the experiment, a charged plastic rod is brought close to a thin stream of each liquid as it flows from a burette. The learners observe that the stream of liquid A bends toward the charged rod, while the stream of liquid B continues to flow straight down without any deflection.

The teacher then asks the learners to recall their knowledge of molecular shape, bond polarity, and distribution of charge within molecules to interpret these observations.

Task: As a chemistry learner, use your understanding of molecular structure and polarity to explain why liquid A is deflected by the charged rod whereas liquid B is not.

Possible Response:

In trichloromethane ($CHCl_3$), the carbon to chlorine bonds are polar hence constituting dipoles because there is a large electronegativity difference between the chlorine atoms and the carbon atom. However, the carbon to hydrogen bond is not polar and does not constitute a dipole because the carbon and hydrogen atoms have similar electronegativities and the electronegativity difference is negligible. Although the hydrogen and chlorine atoms are symmetrical due to absence of lone pairs on the carbon atom, the dipole moments of the carbon to chlorine bonds do not readily cancel out because there is no opposite dipole on the carbon to hydrogen bond. Consequently, the trichloromethane molecule is polar and becomes deflected by a charged rod.

However, in tetrachloromethane (CCl_4), the carbon to chlorine bonds are polar hence constituting dipoles because there is a high electronegativity difference between the chlorine atoms and the carbon atom.

Due to absence of a lone pair on the carbon atom, the carbon to chlorine bonds are symmetrical as the chlorine atoms are tetrahedrally arranged around the central carbon atom. Thus, the dipole moments (bond polarity effects) cancel out each other so that the whole molecule remains non-polar. Consequently, the tetrachloromethane molecule is nonpolar and is not deflected by a charged rod.

2. A science teacher in an urban school in Uganda is guiding learners through a lesson on chemical bonding and molecular structure. During the lesson, the teacher presents the molecule trichloroiodomethane (CCl_3I) and asks learners to analyze its shape and distribution of electrons. The teacher explains that even though the molecule contains several highly electronegative atoms, its overall polarity is not very strong.

To help learners apply their understanding, the teacher poses a challenge: students are asked to examine the arrangement of chlorine and iodine atoms around the central carbon atom, consider differences in electronegativity, and determine how these factors influence the distribution of charge within the molecule.

As part of the discussion, one student suggests that the molecule might be nonpolar because it appears symmetrical, while another argues that the presence of iodine could create an imbalance in charge distribution.

Task: As a chemistry student, use your knowledge of molecular geometry, electronegativity, and bond polarity to explain why trichloroiodomethane (CCl_3I) is considered slightly polar.

Possible Response:

In trichloroiodomethane, the carbon to iodine bond and carbon to chlorine bonds are polar hence constituting dipoles because there is a large electronegativity difference between the iodine atom and carbon atom and then the chlorine atoms and the carbon atom. However, the carbon to chlorine bonds and carbon to iodine bonds have different magnitudes of the dipole moments.

Although the bonds are symmetrical, due to absence of lone pairs on the carbon atom, the dipole moments do not effectively cancel out and the molecule becomes slightly polar.

3. At a secondary school in Uganda, a group of Senior Four learners are carrying out a practical investigation in their chemistry laboratory. Their teacher provides them with two colourless liquid samples labeled **Substance A** and **Substance B**. The learners are informed that both substances have the same molecular formula, $C_2H_2Cl_2$, but differ in their structural arrangement.

During the experiment, the learners carefully heat each substance and record their boiling points. They observe that **Substance A boils at 333 K**, while **Substance B boils at 305 K** (approx.). This difference surprises the learners since both substances contain the same types and numbers of atoms.

The teacher then explains that the two substances are **cis-1,2-dichloroethene** and **trans-1,2-dichloroethene**, which are examples of geometric isomers. The learners are asked to relate the observed difference in boiling points to the structure and intermolecular forces present in each substance.

Task: As a chemistry student, use your knowledge of chemical bonding and intermolecular forces to explain why one of the isomers has a higher boiling point than the other.

Possible Response:

The carbon to chlorine bonds in trans-1,2-dichloroethene and cis-1,2-dichloroethene are polar hence constituting dipoles because there is a large electronegativity difference between the chlorine and the carbon atoms.

However, the carbon to chlorine bonds in trans-1,2-dichloroethene have their dipole moments acting in opposite directions and cancel each other such that the net dipole is zero. Thus, the molecule is nonpolar.

On the other hand, the carbon to chlorine bonds in cis-1,2-dichloroethene have their dipole moments not acting in opposite directions and the molecule is polar.

The polarity of the cis-1,2-dichloroethene creates stronger intermolecular Van der Waal's forces of attraction between its molecules than those in the nonpolar trans-1,2-dichloroethene. Hence more heat energy is required to break the Van der Waal's forces of attraction in cis-1,2-dichloroethene leading to a higher boiling point.

Short scenario-based items about covalent compounds

Item 1: Mary adds sugar to her tea and stirs it until it completely dissolves. She asks why sugar dissolves so easily in water even though it is not an ionic compound.

Suggested Response:

Sugar is mainly sucrose which is a **covalent compound** that contains carbon, hydrogen and oxygen atoms. The oxygen atom being highly electronegative forms a strongly (highly) polar oxygen to hydrogen bond in each sucrose (sugar) molecule.

Sugar dissolves in water because the highly polar hydroxyl groups it contains form **hydrogen bonds** with water molecules. Water is a polar solvent which can interact with the polar parts of the sugar molecule, allowing it to dissolve.

Item 2: James heats to warm his food in a plastic container. He wonders why the container doesn't conduct electricity and heat like a metal.

Suggested Response:

Plastics are made of **covalent compounds** (usually organic polymers) that contain carbon and hydrogen atoms in which all their unpaired electrons are utilized in bonding. Covalent compounds typically do **not have free-moving (delocalised) ions or electrons**, and hence, they **cannot conduct electricity and are also poor conductors of heat**. This makes them safe for microwave use.

Item 3: Rita pours cooking oil into water, but the two don't mix. She wonders why this happens?

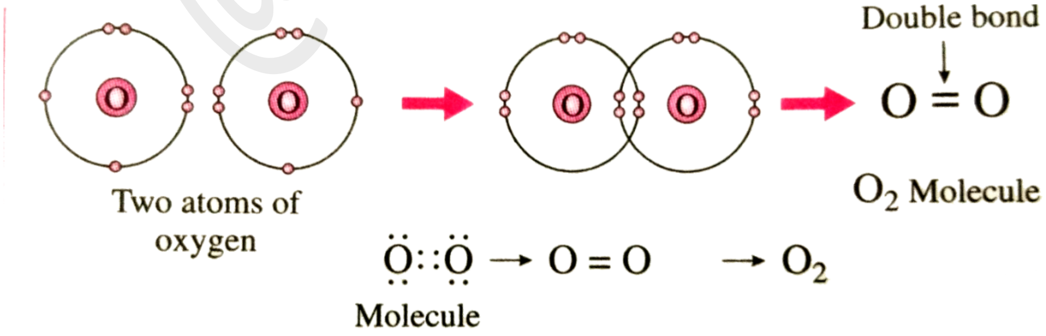
Suggested Response:

Cooking oil is made of **non-polar covalent compounds**, while water is polar. Since **"like dissolves like"**, non-polar covalent molecules don't mix with polar molecules like water.

Item 4: John learns that humans breathe in air containing the oxygen molecules which are utilized in respiration. He asks what kind of bond holds oxygen atoms together in the oxygen, O_2 molecule.

Suggested Response:

The oxygen molecule (O_2) consists of two oxygen atoms joined by a **double covalent bond**, where they share **four electrons** (two pairs) to achieve stable electron configurations.



Item 5: Maria uses vinegar in her salad and asks what makes vinegar, a liquid with a strong unpleasant smell.

Suggested Response:

Vinegar contains acetic (ethanoic) acid, CH_3COOH a carboxylic acid which is a **small covalent molecule** and has a polar functional group. It has a **low boiling point** due to relatively weak intermolecular Vand der Waal's forces, which can easily be broken and make it evaporate easily and produce a strong smell.

Item 6: David notices that candle wax doesn't dissolve in water. He wonders why this happens?

Suggested Response:

Candle wax is composed **covalent hydrocarbons** which are **non-polar**. Thus, the hydrocarbon molecules do **not interact well with polar water molecules**. As a result, wax will not dissolve in water.

Item 7: Lucy opens a soda bottle and sees bubbles forming. What is the bonding nature of the gas escaping the soda bottle?

Suggested Response:

The gas is carbon dioxide (CO_2), a **covalent compound** made of carbon doubly bonded to two oxygen atoms. It's a **linear molecule** which is **non-polar**, with weak intermolecular Van der Waal's forces of attraction. These forces can more readily be broken which is why it easily escapes into the air when pressure is released.

Item 8: When paper burns, it produces smoke and ash. What happens to the covalent compounds in the paper?

Suggested Response:

Paper is made of **cellulose** (a polymer of glucose) which is a covalent compound containing carbon, hydrogen and oxygen atoms.

When paper burns, the covalent compounds within the paper, primarily cellulose, undergo a chemical reaction called combustion. This reaction involves breaking down the **covalent bonds** in the complex molecules into new simpler molecules, releasing energy in the form of heat and light.

New and more stable compounds like carbon dioxide, water vapour, and various other gases are formed together with particles of carbon, which contribute to the smoke and ash.

This is a **chemical change** involving both covalent bond breaking and covalent bond formation.

COORDINATE COVALENT OR DATIVE BONDING

General objectives; By the end of the subtopic, one should be able to:

- Define coordinate-covalent bonding.
- Explain the formation of coordinate-covalent bonds using examples.
- Describe the properties of compounds containing coordinate-covalent bonds.
- Relate coordinate-covalent bonding to real-life scenarios.

Introduction to Coordinate-Covalent Bonding

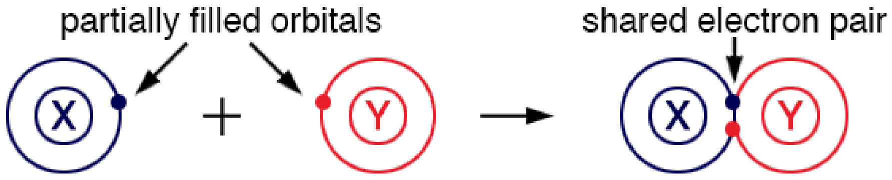
Typically, each atom contributes one or more electrons to the formation of the covalent bond. When the electrons shared between two atoms are donated by only one atom involved in bond formation, the resulting bond is known as a **coordinate covalent bond**.

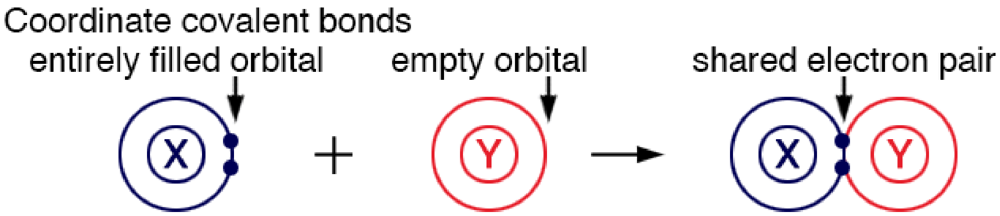
i.e. A **coordinate covalent bond or simply a coordinate bond** (also called a **dative bond**) is a type of covalent bond in which **both electrons** in the shared pair come **from one of the bonding atoms**.

In a coordinate covalent bond, an atom with a filled orbital shares its two electrons with an empty orbital of another atom.

Key difference from a regular covalent bond: In a normal covalent bond, each atom donates one electron to the bond. In a coordinate covalent bond, **only one atom donates both electrons**, but once formed, the bond behaves like a normal covalent bond.

Typical covalent bonds





Two conditions are required for dative bonding to effectively take place;

(a) Presence of a lone pair (non-bonded pair) of electrons on a donor atom or an anion or a central atom of a molecule.

(b) Presence of an acceptor atom with an empty orbital to accommodate the lone pair of electrons.

The properties of coordinate bonds formed are indistinguishable from those of the covalent bonds.

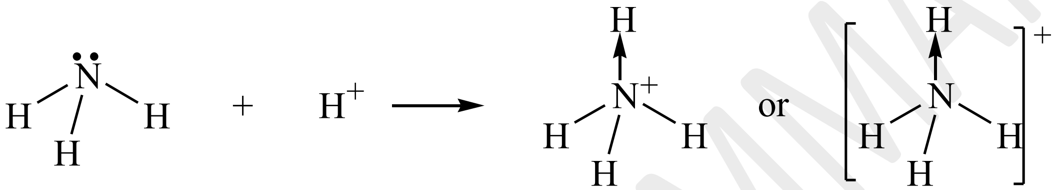
A coordinate bond is shown or represented using **an arrow** pointing from the donor to the acceptor atom.

Examples of substances formed through coordinate bonding include;

(a) Formation of ammonium ion, NH_4^+ when the ammonia molecule combines with a proton.

The nitrogen atom in the ammonia molecule has a lone pair of electrons while the proton has got a vacant 1s-orbital that can accommodate the electrons.

The lone pair of electrons from the nitrogen atom in the ammonia molecule is accommodated by the vacant 1s-orbital in the proton to form a coordinate bond.

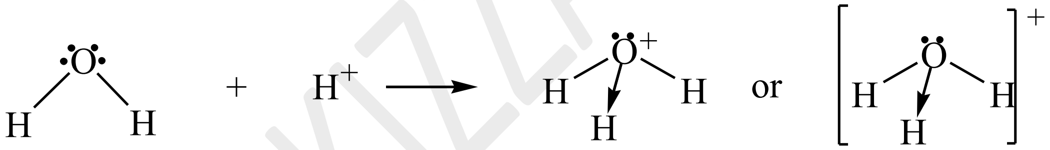


i.e $NH_3 + H^+ \rightarrow NH_4^+$

(b) Formation of hydroxonium ion, H_3O^+ when a water molecule combines with a proton.

The oxygen atom in the water molecule has 2 lone pairs of electrons while the proton has got a vacant 1s-orbital that can accommodate the electrons.

One lone pair of electrons from the oxygen atom in the water molecule is accommodated by the vacant 1s-orbital in the proton to form a coordinate bond.



i.e $H_2O + H^+ \rightarrow H_3O^+$

(c) Reaction between ammonia and boron trichloride to form ammonium boron trichloride.

The nitrogen atom in the ammonia molecule has a lone pair of electrons while the boron atom in boron trichloride has got a vacant 2p-orbital that can accommodate the electrons.

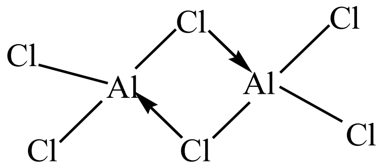
The lone pair of electrons from the nitrogen atom in ammonia molecule is accommodated by the vacant 2p-orbital in the boron atom in boron trichloride to form a coordinate bond.

$$NH_3 + BCl_3 \longrightarrow \left[H_3N \rightarrow BCl_3\right]$$

(d) Dimerization of aluminium and beryllium halides such as:

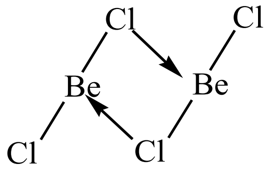
(i) Aluminium chloride, Al_2Cl_6.

Each chorine atom has three lone pairs of electrons while each aluminium atom has got vacant orbitals in its valence shell that can accommodate the lone pair of electrons. One of the lone pairs of electrons from a chlorine atom in aluminium chloride is accommodated by a vacant 3p-orbital in an aluminium atom of another aluminium chloride unit to form a coordinate bond. i.e. $2AlCl_3 \rightarrow Al_2Cl_6$



(ii) Beryllium chloride, Be_2Cl_4.

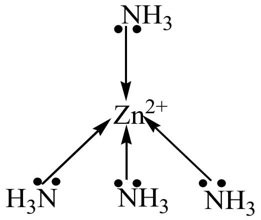
Each chorine atom has 3 lone pairs of electrons while each beryllium atom has got vacant orbitals in its valence shell that can accommodate the lone pair of electrons. One of the lone pairs of electrons from a chlorine atom of a molecule is accommodated by a vacant 2p-orbital in a beryllium atom of another beryllium chloride to form a coordinate bond. i.e. $2BeCl_2 \rightarrow Be_2Cl_4$



(e) Formation of complex ions.

Some metal cations have vacant orbitals (usually d-orbitals for the first series of transition metals) which they use to accommodate the electrons. These metals cations attract lone pairs of electrons from molecules or ions (called ligands) and accommodate them in their vacant orbitals to form coordinate bonds resulting into complex ions. e.g. formation of the tetraamminezinc ion, $[Zn(NH_3)_4]^{2+}$

Due to its small ionic radius, the zinc ion has a high charge density. Thus, the zinc ion strongly attracts lone pairs of electrons from ammonia molecules and accommodates them in its vacant 4s and 4p orbitals to form coordinate bonds in the resulting complex ion. $Zn^{2+}_{(aq)} + 4NH_{3(aq)} \rightarrow [Zn(NH_3)_4]^{2+}_{(aq)}$



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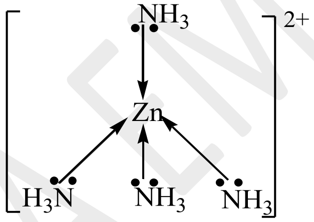


📌 Remember:

- Once formed, coordinate bonds are indistinguishable from other covalent bonds in terms of strength and behaviour.
- Represent coordinate bonds with an arrow (→) from donor to acceptor.

Short Sample Items:

Item 1: Hemoglobin and Oxygen Transport.

(a) A patient is experiencing shortness of breath, and the doctor explains that oxygen is carried in the blood by hemoglobin. How is dative bonding involved in the transport of oxygen by hemoglobin?

Suggested Response:

Dative bonding is involved when oxygen binds to the iron (Fe^{2+}) center in the haeme group of haemoglobin present in blood.

The oxygen molecule acts as the donor of a lone pair of electrons to the iron (Fe^{2+}) ion, which accepts the electron pair. This forms a coordinate covalent bond with the Fe^{2+} ion, facilitating oxygen transport in the blood stream.

Item 2: Water Treatment – Metal Complexes.

In some water treatment facilities, metal ions like copper(II) ion, Cu^{2+} are used to remove impurities. How does dative bonding play a role in this process?

Suggested Response:

Dative bonding occurs when ligands (such as ammonia or water molecules) donate lone pairs of electrons to the Cu^{2+} ion, forming complex ions. These complexes can interact with and neutralize impurities, aiding in water purification.

Item 3: Formation of Ammonium Ion in Household Cleaners and Fertilizers.

(a) A farmer uses a fertilizer containing ammonium ions (NH_4^+). How is dative bonding involved in the formation of the ammonium ion?

Suggested Response:

The ammonium ion (NH_4^+) is formed when an ammonia molecule (NH_3) donates a lone pair of electrons to a hydrogen ion (H^+), forming a dative covalent bond. This creates the positively charged ammonium ion found in many fertilizers.

Item 4: Carbon Monoxide Poisoning.

Carbon monoxide (CO) can bind to hemoglobin more strongly than oxygen. Why is this binding so strong in terms of dative bonding?

Suggested Response:

Carbon monoxide donates a lone pair of electrons from its carbon atom to the iron ion in hemoglobin, forming a very stable dative bond. This bond is stronger than the one formed with oxygen, preventing oxygen transport and leading to poisoning.

Item 5: Dative Bonds in Everyday Products (e.g., Toothpaste).

Many toothpastes contain fluoride ions (F^-). In some cases, these ions can form complexes with metal ions like tin (Sn^{2+}). What type of bond is formed, and what is the role of the fluoride ion?

Suggested Response:

The fluoride ion forms a dative bond by donating a lone pair of electrons to the metal ion (Sn^{2+}), helping to form stable complexes that can protect the enamel and reduce cavities.

INTERMOLECULAR FORCES OF ATTRACTION

Intermolecular forces of attraction are the weak forces that exist between molecules, holding them together in different states of matter. Understanding these forces helps to explain everyday phenomena such as why water forms droplets, why some substances evaporate quickly, and why solids have fixed shapes.

Unlike strong chemical bonds that hold atoms together within a molecule, intermolecular forces act between separate molecules. These forces vary in strength and include types such as Van der Waals forces, dipole-dipole interactions, and hydrogen bonding. Although they are weaker than covalent or ionic bonds, they play a crucial role in determining the physical properties of substances, including boiling point, melting point, and solubility.

By studying intermolecular forces, one develops a deeper understanding of how matter behaves in real-life situations, making it easier to connect scientific concepts to practical applications in their environment.

Molecules are formed by both covalent compounds and non-metals.

Covalent bonds have directional properties since they are polar and hence intermolecular forces exist between dipoles of covalent compounds. The major intermolecular forces of attraction are either *Van der Waal's forces* or *Hydrogen bonds*. They arise from electrostatic attraction of the nuclei of one atom of a molecule for electrons of an atom of a different molecule.

The strength of the intermolecular forces of attraction will determine whether a molecule is a solid, liquid or gas at room temperature.

Intermolecular forces are attractions between molecules (not within them), and they vary in strength and origin.

Remember that; ionic bonds, covalent bonds, metallic bonds and dative bonds are stronger than the intermolecular forces of attraction.

THE VAN DER WAAL'S FORCES

General objectives; By the end of the subtopic, one should be able to:

- Define Van der Waal's forces.
- Explain the formation of Van der Waal's forces using examples.

- Describe the properties of compounds containing Van der Waal's forces.
- Relate Van der Waal's forces to real-life situations.

Introduction to Van der Waal's forces

Van der Waal's forces are short range *weak forces of attraction* between molecules (atoms or covalent compounds that exist in a free and separate state).

Van der Waals forces are **weak intermolecular forces** that occur **between molecules or atoms** due to temporary or permanent **dipoles**. These forces are **not chemical bonds** (like covalent or ionic), but **physical attractions**.

Hence Van der Waal's forces are always significant (recognized) when the molecules are closer to one another. They are the most common intermolecular forces of attraction.

Types of Van der Waal's forces:

These include; the Dipole-to-Dipole interactions, the London Dispersion forces and the Dipole-Induced Dipole Forces.

(a) **Dipole to dipole (orientation)** forces / interactions (Keesom 1912).

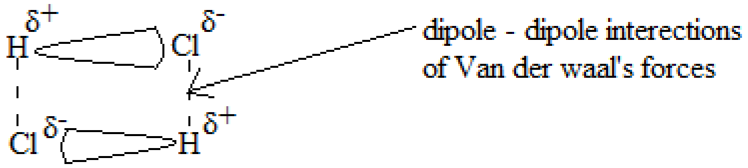
- They exist between **polar molecules** that have permanent dipoles.
- The molecules align so that the **positive end** of one molecule attracts the **negative end** of another molecule.
- They are stronger than the London dispersion forces but weaker than hydrogen bonds.

If a covalent compound has a bond with a permanent dipole, the positively charged end of the dipole of one molecule will attract the negatively charged end of another molecule and the molecules will orient themselves accordingly.

e.g. The hydrogen chloride, HCl molecule for illustration.

Owing to different electron attracting powers of chlorine and hydrogen (i.e chlorine being more electronegative than hydrogen), the concentration of electrons around the chlorine atom is greater than the concentration around the hydrogen atom and an uneven charge distribution is produced. The chlorine end of the molecule acquires a partial negative charge and the hydrogen end acquires a partial positive charge. This combination of charges is referred to as a dipole.

A *permanent dipole is formed because the two atoms have an established permanent difference in electronegativity.*



The electron distribution is equivalent to the separation of the charge.

From the above figure, the attraction between unlike charges and the repulsion of like charges causes the dipoles to arrange themselves so that unlike charges face each other. In this position, the molecules have an attraction to one another through inter molecular forces.

In summary; a large difference in electronegativity between two bonded atoms will cause a permanent charge separation, or a permanent dipole in a covalent molecule or ion. Two or more molecules or ions with permanent dipoles can interact within dipole-dipole interactions. The bonding electrons in a molecule or ion will, on average be closer to the more electronegative atom more frequently than the less electronegative one, giving rise to partial charges on each atom, and causing electrostatic forces between molecules or ions.

(b) **Dipole to induced dipole (induction)** forces / interactions (Debye 1920)

- They occur when a **polar molecule** induces a dipole in a **non-polar molecule**.
- They are weaker than dipole-dipole forces.

A covalent compound (molecule) with a covalent bond that has a permanent dipole can induce a dipole in another non polar molecule and an attraction between the two will be attained.

(c) **London Dispersion** forces (London 1930) **(Instantaneous Dipole-Induced Dipole)**

- They **occur in all molecules**, especially those that are **non-polar**.
- They arise due to **temporary fluctuations** in electron density, creating temporary dipoles.
- They are the **weakest** of all Van der Waals forces.
- Their strength increases with:
  - Larger molecular size (more electrons).
  - Greater surface area.

Momentary distortions of the electron orbitals (**temporary, random shifts in how electrons are distributed around an atom or molecule**) can produce instantaneous fluctuating dipoles. These pulsations are synchronized to produce attractive forces between non-polar covalent molecules or atoms (in case of noble gases) known as dispersion forces.

In other words, the average distribution of charge in non-polar molecules (e.g. methane, etc.) or atoms (e.g. noble gases) over a period of time is not uniform.

In these atoms or molecules, electrons aren't fixed but they are constantly in motion. Because of this, their distribution around the nucleus is not perfectly even all the time.

Because electrons move, they may at one instant be slightly accumulated on one part of the molecule and as a consequence, a small *temporary dipole* will occur.

This temporary dipole in one molecule can induce opposite (attractive) dipoles in the surrounding molecules. It does so because the negative (or positive) charge in a portion of one molecule will distort the electron cloud of an adjacent molecule causing opposite charges also temporarily to develop there.

These temporary dipoles change continuously but the net result of their existence is to produce attractive forces between non-polar molecules and thus make possible existence of their liquid and solid states.

In summary; the London dispersion force arises due to instantaneous temporary dipoles in neighbouring molecules. As the negative charge of the electron is not uniform around the whole atom, there is always a charge imbalance. This small charge will induce a corresponding dipole in a nearby molecule; causing an attraction between the two dipoles. When the electron then moves to another part of the electron cloud, the previous attraction is broken.

i.e. electrons are always moving, a temporary imbalance in electrons occurs, this creates a **temporary dipole**, and this distortion affects nearby particles and attraction follows between molecules.

Factors that affect the strength of the Van der Waal's forces

The *strength of the Van der Waal's forces depends* on the following factors:

(a) The molecular mass. The larger the atomic mass (radius), the higher the molecular mass and the greater the surface area of contact between the molecules. Thus, there are more Van der Waal's forces formed between the molecules and strength of the Van der Waal's forces of attraction increases.

Note that:

The ease with which the charge distribution in a molecule is distorted by an external force is called **polarizability**.

The greater the polarizability of the molecule, the more easily its electron cloud is distorted to give a temporary dipole and hence the greater the strength of London dispersion forces (intermolecular Van der Waal's forces).

Therefore, larger atoms or molecules also have **more electrons** (a greater electron cloud). The more the electrons, the greater the polarizability because their bonding electrons are further from their nucleus and can the more easily distorted to form stronger temporary dipoles. This also increases the strength of the **Van der forces**, making attraction between the molecules stronger.

Therefore, the strength of these intermolecular forces tends to increase with increasing molecular mass which also increases with increase in atomic mass or radius.

This factor helps to understand or explain the trend in the melting points or boiling points of most nonmetals in the Periodic Table and their covalent compounds. e.g., boiling points of a group of molecules such as halogens Fluorine, Bromine, Chlorine and Iodine molecules increase with increasing molecular mass.

(b) The molecular shape (surface area). This deals with the difference in the arrangement of same atoms creating different molecules. The longer and flat the molecules, the greater the surface area of contact between the molecules. Thus, there are more Van der Waal's forces formed between the molecules and strength of the Van der Waal's forces of attraction increases.

Longer, flat, or extended molecules have **greater surface area for contact**, and hence leading to stronger intermolecular attractions.

e.g., Straight-chain hydrocarbons have stronger Van der Waals forces than their branched isomers.

In this case, the branching in hydrocarbons or other organic compounds reduces the surface area of contact between the molecules with either the same or similar molecular mass. This in turn decreases the number and strength of the Van der Waal's forces of attraction leading to lower values of the melting or boiling points.

(c) Bond polarity: The more polar the molecule, the stronger the dipole interactions (the Van der Waal's forces).

Conclusion

Van der Waals forces may be weak, but they are **crucial in many physical and biological processes**. They explain why gases condense, why insects can walk on water, and why molecular structures behave as they do. Understanding them helps in fields like **medicine, materials science, and environmental chemistry**.

Sample Items

1. Simeon sees a gecko lizard walk effortlessly on walls and ceilings without falling. He wonders how this is possible without the lizard having glue. Help Simeon understand this observation.

Suggested Response:

Geckos use Van der Waals forces through millions of tiny hair-like structures (setae) on their feet. These structures increase contact with surfaces, allowing more weak Van der Waals forces to form. Thus, the Van der Waal's forces collectively provide stronger adhesion, enabling them to stick and walk on vertical or even inverted surfaces.

2. During manufacture of oxygen in an air separation plant, gases like oxygen and nitrogen are liquefied. How are Van der Waals forces involved?

Suggested Response:

Van der Waals forces are weak intermolecular attractions that allow nonpolar gases such as nitrogen and oxygen to come close enough to condense into liquids.

These forces become significant under high pressure and low temperature, which are the conditions required in industrial gas liquefaction processes.

To liquefy air, it is **compressed** and then **cooled** to very low temperatures (below −183 °C for oxygen and −196 °C for nitrogen).

At low temperatures, gas molecules **lose kinetic energy**, and **Van der Waals forces become strong enough to pull molecules closer together**, causing **condensation into liquid form**.

The **difference in strength** of these forces between the two gases results in **different boiling points**, which makes it possible to **separate them by fractional distillation**.

3. Emma notices that 2 plastic sheets sometimes stick together without glue. Help Emma understand what causes this.

Suggested Response:

This sticking is due to Van der Waals forces between the long polymer chains of the plastic. These weak attractions are enough to cause temporary adhesion between surfaces like plastic bags or cling film.

When the plastic sheets are pressed together, especially if they are **smooth and clean**, the molecules on the surfaces come very close to each other.

At this close range, **temporary attractions** form between the molecules of one sheet and those of the other and these are Van der Waals forces.

Although weak individually, when many such interactions happen over a large surface area (like a plastic sheet), they can create enough attraction to make the sheets **stick together** lightly.

Also, **static electricity** can sometimes play a role. When plastic sheets are rubbed or peeled apart, they can become **electrostatically charged**, leading to an attractive force between them.

4. When Julian sprays perfume in one corner of the room, the fragrance spreads quickly throughout. What role do Van der Waals forces play here?

Suggested Response:

The perfume molecules, which are often nonpolar or weakly polar, experience Van der Waals interactions as they move and interact with air molecules. These weak attractions help in dispersion and diffusion of the fragrance throughout the room.

5. Why does argon have a higher boiling point than helium, even though both are noble gases?

Suggested Response:

As the size and molar mass of noble gas atoms increase, Van der Waals (London dispersion) forces also increase. Argon has a larger molecular mass and thus has stronger dispersion forces than helium, which leads to a higher boiling point.

Situation	**Type of Van der Waals Force**	**Explanation**
1. Noble gases liquefying (e.g., Argon, Neon)	London Dispersion	Despite being monoatomic and non-polar, they can be liquefied due to weak dispersion forces.
2. Geckos climbing walls	London Dispersion	Microscopic hairs on their feet interact with wall molecules via dispersion forces, allowing adhesion.
3. Perfume smell spreading in air	Weak intermolecular forces	The molecules are non-polar and spread because only weak Van der Waals forces hold them together.
4. Cooking oil floating on water	Dipole-induced dipole (or lack thereof)	Oil is non-polar, water is polar; weak interactions prevent mixing.
5. Plastic wrap clinging to containers	Dipole-induced dipole	Wrap may pick up static charge, allowing weak attractions to containers.
6. Liquids like acetone evaporating quickly	Dipole-Dipole + Weak London Forces	Lower intermolecular forces allow molecules to escape easily into gas phase.

Assignment Item:

Imagine a group of Senior Four learners in a Rubaga secondary school participating in a laboratory lesson on elements in Group VII of the periodic table. Their teacher brings samples and data charts for the halogens—fluorine, chlorine, bromine, and iodine—and asks the students to observe their physical properties.

During the discussion, one student notices that fluorine and chlorine are gases at room temperature, bromine is a liquid, and iodine is a solid. The teacher then provides data showing the boiling points of these elements and asks the class to analyze the trend. The learners observe that the boiling points increase steadily from fluorine (F_2) to iodine (I_2).

Curious about this pattern, the teacher challenges the students to think beyond memorization and relate the trend to the structure and behavior of the halogen molecules.

Task: As one chemistry student in the class, use your knowledge of intermolecular forces and atomic structure to explain why the boiling points of the halogens increase in the order $F_2 < Cl_2 < Br_2 < I_2$.

THE HYDROGEN BONDS

General objectives; By the end of the subtopic, one should be able to:

- Define the term hydrogen bond.
- Explain the formation of a hydrogen bond using examples.
- Describe the properties of compounds containing hydrogen bonds.
- Relate hydrogen bonds to real-life scenarios.

Introduction to hydrogen bonds

Hydrogen bonding is a special type of **intermolecular force**—that is, a force of attraction between molecules. It plays an important role in many substances you encounter in everyday life, especially water.

A **hydrogen bond** is a *strong* dipole to dipole *intermolecular* or intramolecular force of attraction which occurs between molecules.

It occurs when a hydrogen atom bonded to a highly electronegative atom in a molecule becomes strongly polar and therefore strongly attracts an adjacent highly electronegative atom with a lone pair of electrons in another molecule or the same molecule.

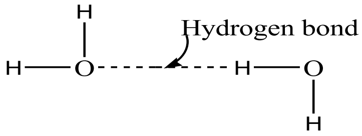
The highly electronegative atom that forms a hydrogen bond may be fluorine, oxygen and nitrogen.

Causes of a hydrogen bond (How a hydrogen bond arises):

The hydrogen atom has a *small atomic radius*, and the positive charge is more concentrated and exposed. Fluorine, oxygen and nitrogen have a *small atomic radius* and a *high electronegativity.* Thus, the fluorine to hydrogen, oxygen to hydrogen and nitrogen to hydrogen bonds are *strongly (highly) polar*. It is therefore very effective for the hydrogen atom in one molecule to form a strong link with other molecules. Hydrogen atoms in this type of situation are found to make particularly *strong intermolecular bonds* with a fluorine, oxygen or nitrogen atom in a neighbouring molecule called a *hydrogen bond*.

A hydrogen atom bond is electrostatic and is usually represented by dotted lines.

i.e.



In summary; a hydrogen bond is effectively a strong example of an interaction between two permanent dipoles. The large difference in electronegativity between hydrogen and any of fluorine, nitrogen and oxygen, coupled with their lone pairs of electrons cause strong electrostatic forces between molecules. Hydrogen bonds are responsible for the high boiling points of water and ammonia with respect to their heavier analogues.

There are two types of hydrogen bonds:

(a) Intramolecular hydrogen bonds. These occur within the same molecule.

(b) Intermolecular hydrogen bonds. These occur between two similar or different molecules and this is the most common type.

Note that; a hydrogen bond is weaker than a simple ionic or covalent bond but it is stronger than the ordinary Van der Waal's forces of attraction.

Factors that influence the strength of a hydrogen bond

The strength of a **hydrogen bond** is determined by several factors, which influence how strongly the hydrogen atom is attracted to an electronegative atom in another molecule or within the same molecule. Here are the **main factors**:

1. **Electronegativity of the Atoms Involved.**

Hydrogen bonds form between hydrogen and highly electronegative atoms **like** fluorine (F), oxygen (O), and nitrogen (N)**.**

Higher electronegativity difference **causes** stronger partial charges **which also lead to a** stronger hydrogen bond**.** The electronegativity decreases in the following order F > O > N.

Hence in terms of electronegativity, the H-F bonds **tend to form** stronger hydrogen bonds **than the H-O bonds** and the **H-O bonds** form stronger hydrogen bonds than **H-N bonds.**

2. **Polarity of the Molecules.**

Molecules with **greater polarity** (large difference in electronegativity between atoms) form **stronger hydrogen bonds**.

A highly polarized **donor–acceptor pair** increases electrostatic attraction, strengthening the hydrogen bond.

3. **Distance Between Donor and Acceptor.**

Hydrogen bonds are **stronger when the donor (X–H) and acceptor (Y)** atoms are **closer**.

Optimal hydrogen bonding occurs when the **H···Y distance is short**, typically **1.5 – 2.5 Ångströms**.

4. **Bond Angle.**

The **linear arrangement** of atoms (**X–H···Y**) leads to **stronger hydrogen bonds**.

A hydrogen bond is strongest when the angle is **close to 180°**, which allows **maximum orbital overlap**.

5. **Solvent Effects.**

In **aqueous environments**, water can **compete** with and **disrupt** hydrogen bonding between solutes.
In contrast, in **non-polar solvents**, hydrogen bonds are usually **stronger** because there is less competition.

6. **Charge on the Donor or Acceptor**

If either the donor or acceptor is **charged** (e.g., an anion or cation), hydrogen bonding is **significantly stronger**.
Example: **O–H···O⁻** is much stronger than a neutral O–H···O.

7. **Number of Hydrogen Bonds (Cooperative Effect)**

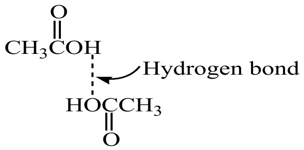
When **multiple hydrogen bonds** form in a **network** (as in water or proteins), the overall strength is increased due to **cooperative stabilization**.

<u>Effects of hydrogen bonding on properties of some compounds</u>

1. In non-polar solvents (such as benzene, methyl benzene, camphor, etc), organic acids are found to have a relative molecular mass which is twice the expected value.

This is because the oxygen atom has a small atomic radius and is highly electronegative. This makes the oxygen to hydrogen bond in ethanoic acid strongly polar and thus two molecules of ethanoic acid are held together by strong hydrogen bonds forming a dimer.

e.g.



The observed relative formula mass of ethanoic acid is 120 and not 60 due to hydrogen bonding.

2. Ice is less dense and floats on water.

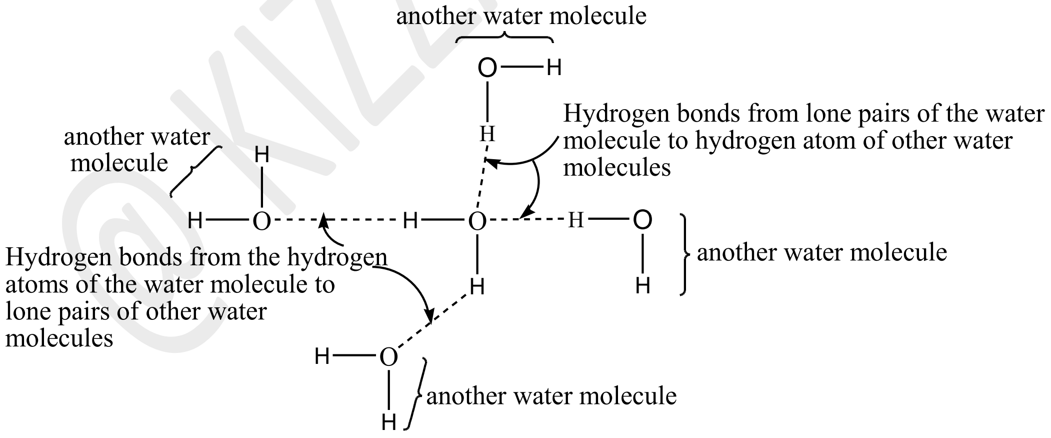
The oxygen atom has a small atomic radius and is highly electronegative. This makes the oxygen to hydrogen bond in water strongly polar.

In the crystal lattice (ice), the molecules are closely packed as lower thermal energy freezes and locks them into an ordered tetrahedral pattern that maximises hydrogen bonding.

Therefore, each water molecule forms:

(a) Two hydrogen bonds by virtue of (due to) the slight positive charges on its hydrogen atom being attracted to the non-bonding pairs (lone pairs) of electrons being associated with other oxygen atoms of other water molecules.

(b) Two hydrogen bonds by virtue of its non-bonding (lone pair) electron pairs on the oxygen atom being associated/attracted to the slight positive charges on the hydrogen atoms of other water molecules.



The four hydrogen bonds formed create an approximately tetrahedral structure around each oxygen atom. This creates in turn an open structure with spaces in between occupying a larger volume and the density of the ice formed will be lower hence floating on water.

When the ice melts, the crystal lattice crumbles as the hydrogen bonds are progressively broken. This facilitates formation of close packing in the liquid water and hence a decrease in the volume of water.

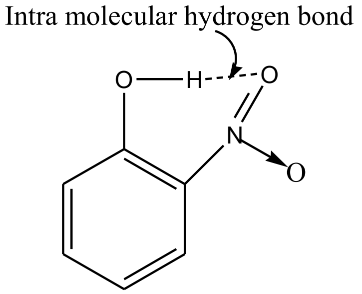
Thus, a given mass of ice occupies a bigger volume than the same mass of liquid water.

Note that; Liquid water cannot maintain four hydrogen bonds per molecule. The molecules in the liquid state move freely due to higher thermal energy and hydrogen bonds constantly break and reform, leading to an average of fewer than four bonds per molecule.

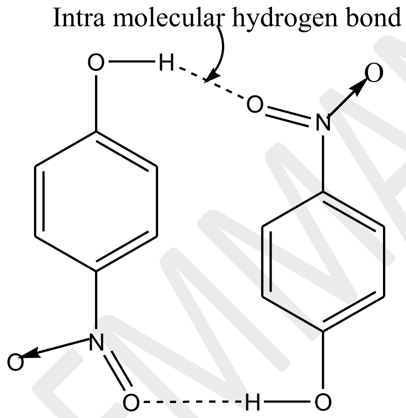
3. The boiling point of 2–nitrophenol is lower than the boiling point of 4–nitrophenol.

This is because the oxygen atom has a small atomic radius and is highly electronegative. This makes the oxygen to hydrogen bond in the hydroxyl group strongly polar. Thus, 2–nitrophenol possesses an internal (intramolecular) hydrogen bond formed by the interaction of the hydrogen atom of the hydroxyl group (OH) with the oxygen atom of the nitro (NO_2) group in the same molecule due to the closeness of the two groups. Consequently, there is formation weaker Van der Waal's forces of attraction between molecules of 2–nitrophenol hence having a lower boiling point.

On the other hand, 4–nitrophenol associates via intermolecular hydrogen bonding. This is because the two groups are not close to each other in the same molecule but are closer between two different molecules. The hydrogen bonds are stronger intermolecular forces than the Vander Waal's forces. Thus, extra energy is needed to overcome these forces hence having a higher boiling point.



In the 2-nitrophenol molecule



In the 4-nitrophenol molecules

4. The boiling points of ammonia NH_3, water H_2O and hydrogen fluoride HF are abnormally the highest as compared to other hydrides of group V, VI and VII respectively in the periodic table.

This is because the nitrogen, oxygen and fluorine atoms have small atomic radii and are highly electronegative. This makes the nitrogen to hydrogen, oxygen to hydrogen and fluorine to hydrogen bonds in the hydrides strongly polar. Thus, the hydrides of nitrogen (NH_3), Oxygen (H_2O) and fluorine (HF) associate via strong intermolecular hydrogen bonding. Therefore, extra energy is needed to overcome the hydrogen bond.

Other hydrides associate via the weak intermolecular Vander Waal's forces which require little energy to be broken.

5. Alcohols, amines, etc are soluble in water, yet water is an inorganic solvent and they are organic compounds.

This is because the oxygen atom in alcohols and the nitrogen atom in amines have small atomic radii and are highly electronegative. This makes the oxygen to hydrogen bond in alcohols and the nitrogen to hydrogen in amines strongly polar. Thus, alcohols and amines can form strong intermolecular hydrogen bonds with the water molecules.

6. Hydrated copper(II) sulphate, **$CuSO_4.5H_2O$** i.e. copper(II) sulphate pentahydrate is a *blue* crystalline salt. On heating to about 100^0C, the pentahydrate is converted into the monohydrate.

$$CuSO_4.5H_2O_{(s)} \longrightarrow CuSO_4.H_2O_{(s)} + 4H_2O_{(l)}$$

The on further heating to about 250^0C, the anhydrous copper(II) sulphate is formed.

$$CuSO_4.H_2O_{(s)} \longrightarrow CuSO_{4(s)} + H_2O_{(l)}$$

This shows that one of the molecules of water of crystallization is different from the other four.

In the crystal lattice, the fifth water molecule is not directly bonded to the copper(II) ion but is held in its position by strong intermolecular hydrogen bonds between water molecules and sulphate ions of two octahedral groups.

Sample Item:

At National Curriculum Development Centre, science teachers are encouraging learners to relate chemistry concepts to real-life situations.

During a laboratory activity at a secondary school in Kampala, a group of students is investigating how different substances interact with water. The teacher provides them with two samples: blue crystals of hydrated copper(II) sulphate and a white powder of anhydrous copper(II) sulphate.

The students carefully add each substance separately into beakers containing water and observe the temperature changes using thermometers. They notice that when the blue hydrated copper(II) sulphate dissolves, the temperature of the solution drops, making the beaker feel cold. However, when the white anhydrous copper(II) sulphate dissolves, the temperature rises, and the beaker feels warm.

Curious about the difference, the students begin discussing their observations and try to relate them to energy changes during dissolving.

Task: As a chemistry learner, use your knowledge of energy changes in chemical processes to explain why hydrated copper(II) sulphate dissolves endothermically while anhydrous copper(II) sulphate dissolves exothermically.

Suggested Response:

In the solid copper(II) sulphate pentahydrate, the **fifth water molecule is held in position by intermolecular hydrogen bonds** between the water molecules and sulphate ions of two octahedral groups.

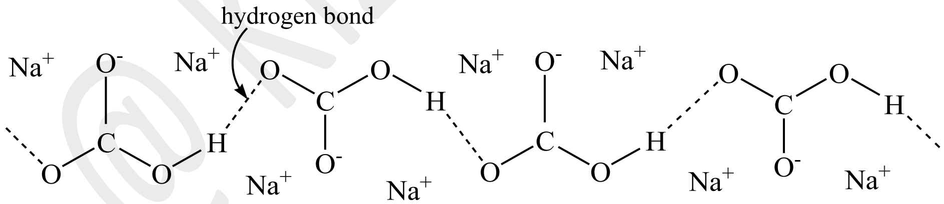
The **heat energy that must be absorbed to overcome the hydrogen bonding and also to break the ionic lattice is greater than the heat energy released when the ions are completely surrounded by water molecules** (make the ions be solvated by water molecules).

This makes the overall process endothermic and therefore when the pentahydrate crystals are dissolved in water, heat energy is absorbed.

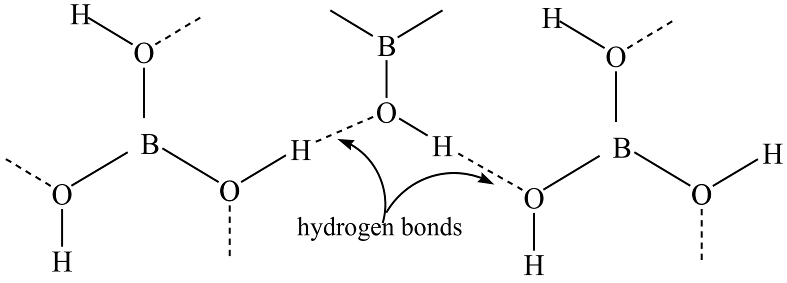
However, when the solid anhydrous copper(II) sulphate is dissolved in water, heat energy is given out (i.e. the process is exothermic). This is because, the **heat energy that must be absorbed to break the ionic lattice is less than the heat energy released when the ions are completely surrounded by water molecules**. i.e. the formation of hydrated ions is easier and the hydration energy is greater than the lattice energy as the solid is being dissolved hence the overall enthalpy change is an exothermic process.

7. Sodium hydrogen carbonate crystals, $NaHCO_3$ do not easily dissolve in water. They dissolve in water endothermically.

In the crystal, the hydrogen carbonate ions are linked together into infinite chains by hydrogen bonds. The chains are held laterally by sodium ions between them.



8. Boric acid crystals, H_3BO_3 do not easily dissolve in water. They dissolve in water endothermically. They are made up of infinite layers of $B(OH)_3$ units linked together by hydrogen bonds.



Experimental Determination of a Type of Bond

This is achieved by identifying the properties of the type of bond.
e.g. determination of the type of bond in the chloride of sodium.
A sample of sodium chloride is placed in a clean and dry test-tube and heated over a bunsen flame. The sample does not melt implying that it has a high melting point.
Another sample of sodium chloride is added to water in a clean container and the mixture stirred. The sample dissolves to form an aqueous solution implying that it is a polar compound.
The aqueous solution is then connected to an electric current on which a bulb is also connected. The bulb produces light implying that an aqueous solution of the sample is an electrolyte and hence contains mobile ions.
In conclusion, the solid sample has a high melting point, readily dissolves in water and acts as an electrolyte. These are properties typical of ionic compounds implying that the type of bond in the chloride of sodium is ionic (electrovalent).

DRAWING STRUCTURES OF SIMPLE COVALENT COMPOUNDS AND IONS

Octet Rule of Valency:

Atoms often gain, lose or share electrons to achieve the same number of electrons as the noble gases closest to them in the periodic table. Because of the fact that all noble gases (except Helium) can have eight electrons in the valence shell, many atoms undergoing a chemical reaction end up with eight valence electrons. This is known as the **Octet Rule of Valency**.

Note that; the hydrogen atom does not follow the octet rule. This is because it gains one electron to form an electronic structure similar to that of helium which is a **duplet**.

Rules of drawing Lewis structures:

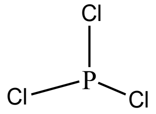
These are used to establish the number of bond pairs and lone pairs associated with a central atom in a particular specie that is responsible for determining the shape.

1. Sum up the valence electrons from all the atoms in the molecule or ion. If the species is an ion, add an electron for each negative charge or subtract an electron for each positive charge.

Identify the number of electron pairs involved.
e.g. (i) PCl_3 (phosphorus trichloride) one P atom = 5 electrons
three Cl atoms = (7 x 3) = 21 electrons
Total = 21 + 5 = 26 electrons = 13 pairs of electrons
(ii) NO_2^- (nitrite ion) one N atom = 5 electrons
two O atoms = (6 x 2) = $12e^-$
Total = 5 + 12 + 1 = $18e^-$ = 9 electron pairs
(iii) NO_2^+ (nitronium ion) one N atom = $5e^-$
two O atoms = (6 x 2) = $12e^-$
Total = 5 + 12 -1 = $16e^-$ = 8 electron pairs

2. Distribute the bond pairs. This is achieved by writing symbols of the atoms involved so as to show which atoms are connected to each other. The atom with the least number is usually taken as the central atom and then draw a single bond between each pair of atoms bonded to the central atom and this becomes a *bond pair*.

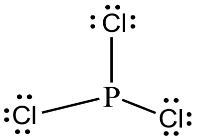
e.g.



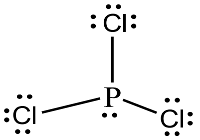
3. Distribute the lone pairs. This is achieved by completing the octets of the atoms bonded to the central atom using *lone pairs*. Remember that hydrogen has a maximum of two electrons (it is a duplet).

Count the total number of electron pairs (lone pairs and bond pairs) distributed onto the atoms and then compare them with the previously calculated number of electron pairs to ascertain whether they have been completely distributed. However, if the previously calculated number of electron pairs are still more than the counted electron pairs, then the extra electron pairs are taken as lone pairs of electrons for the central atom.

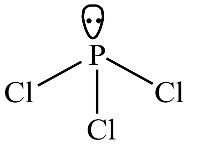
e.g.



A total of 12 electron pairs (3 bond pairs and 9 lone pairs on the oxygen atoms) have been distributed out of the 13 electron pairs previously calculated. Then the remaining 1 electron pair is a lone pair of electrons on the central phosphorus atom which also gives it an octet.



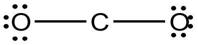
Hence the phosphorus atom has 3 bond pairs and 1 lone pair in phosphorus trichloride which gives a *trigonal pyramidal shape*.



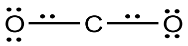
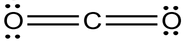
4. If there are not enough electrons to give the central atom an octet, try out multiple bonds i.e. double or triple bonds.

e.g. In CO_2 molecule, one C atom = 4 electrons two O atoms = 2 x 6 = 12 electrons

4 + 12 = 16 electrons = 8 electron pairs



A total of 8 electron pairs (2 bond pairs and 6 lone pairs on the oxygen atoms) have been distributed out of the 8 electron pairs previously calculated. However, carbon has four electrons in the structure and does not form an octet. Thus, a lone pair on each of the two oxygen atoms is converted into a bond pair to form double bonds.

i.e.  which becomes 

Hence the carbon atom has 2 bond pairs and no lone pair which gives carbondioxide a *linear shape*.

$O=C=O$

Exceptions of the octet rule:

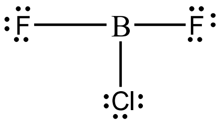
The octet rule works well for the elements in the 2nd period of the Periodic Table. This is because these elements have only 2s and 2p orbitals available for bonding and a total of which eight electrons fill these orbitals.

The exceptions of the octet rule are of three types;

(a) Molecules with an odd number of electrons.

The majority of elements have an even number of total electrons and complete sharing of spins occurs. However, in a few molecules such as ClO_2, NO, NO_2, the number of electrons is odd. Complete pairing is impossible and an octet around each atom cannot be achieved.

(b) Molecules in which an atom has less than an octet. This is most encountered in compounds of Boron and Beryllium. e.g. Boron trifluoride, BF_3



(c) Molecules in which an atom has more than an octet.

Elements of period 3 and beyond have orbitals for bonding. These elements can accommodate more than eight electrons in their outermost shells and therefore can form more than four covalent bonds.

e.g. Phosphorus can form PCl_3 and PCl_5 but nitrogen can only form NCl_3.

This can be explained comparing the spin configuration of phosphorus and nitrogen in their ground state and possible excited state.

Predicting the geometry of molecules or ions (Shapes of molecules or ions)

We can predict the shapes of molecules or ions using the Valence Shell Electron Pair Repulsion (VSEPR) theory. The valence shell electron pair repulsion theory predicts the shape and bond angles of molecules. It states that “electron pairs around a central atom will arrange themselves to minimize repulsion”. The ideas put forward are;

(a) Electron pairs around the central atom are distributed so as to minimize their mutual repulsions.

(b) Orbitals of lone pairs exert a bigger repulsion than those of bond pairs.

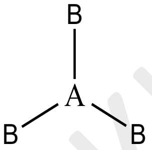
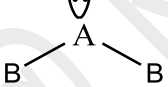
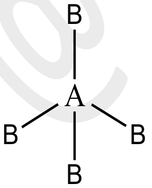
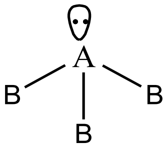
This is because the charge cloud distribution of the lone pair is closer to the central atom and more spread out sideways than the charge cloud of the bonded pair of electrons.

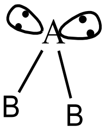
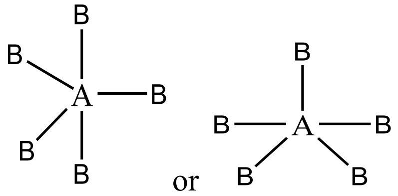
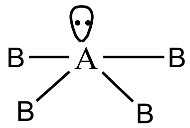
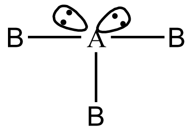
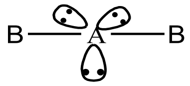
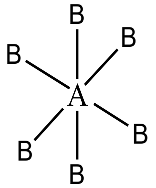
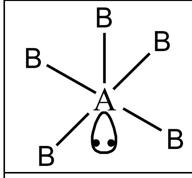
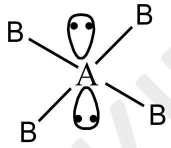
As a result, lone pair electrons exert a large repulsive effect than the bonded pair electrons do. The bonded pair electrons move closer to one another and away from lone pairs to minimize repulsion. This means that repulsion is in order;

Lone Pair / Lone Pair > Lone Pair / Bond Pair > Bond Pair / Bond Pair.

The following steps are used in predicting molecular geometries using the VSEPR theory model;

- Sketch the Lewis dot structure of the molecule or ion.
- Count the total number of electron pairs around the central atom and arrange them in a way to minimize the electron pair repulsion.
- Describe the molecular geometry in terms of the angular arrangement of the bonding pairs.
- The angular arrangement of the bond pairs corresponds to the angular arrangement of the bonded atoms.
- Note that the VSEPR theory model when applied to predict the molecular shapes, multiple bonds should be treated as though it were a single electron pair.
- Electron pair geometries and molecular shapes for molecules with two or four electron pairs around the central atom are given below;

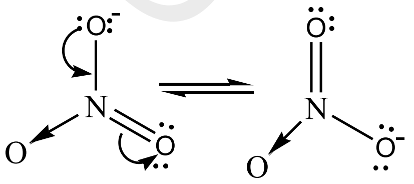
Number of electron pairs	Electron pair geometry (Structure shape)	Bond pairs	Lone pairs	Molecular geometry (Name of the shape)
2		2	0	Linear
3		3	0	Trigonal planar
		2	1	Bent or V-shaped
4		4	0	Tetrahedral
		3	1	Trigonal (triangular) pyramidal

		2	2	Bent or V-shaped
5		5	0	Trigonal bipyramidal
		4	1	See saw
		3	2	T-shape
		2	3	Linear
6		6	0	Octahedral
		5	1	Square pyramidal
		4	2	Square planar

RESONANCE

In chemistry, **resonance** is a way of describing delocalized electrons within some molecules or ions where bonding cannot be expressed by one single Lewis structure. This is observed in most negatively charged ions containing two or more oxygen atoms. Consider the examples below:

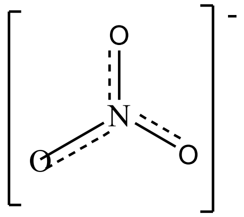
(a) the nitrate ion



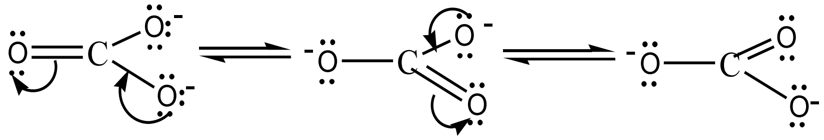
- The three structures only differ in the position double bonds but have the same energy.
- Giving one of the structures as a representation of the nitrate ion gives a wrong impression.
- The nitrate ion is taken to be a blend of all the above structures. This blending of structures is called **resonance**.

- The structure resulting from resonance is called the **resonance hybrid** of the contributing structures called **canonical forms**. It is used as a representation of the nitrate ion.

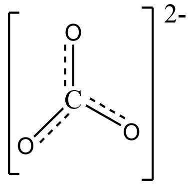
i.e.



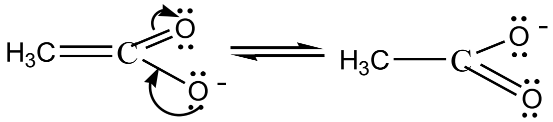
(b) the carbonate ion



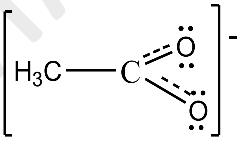
The resonance hybrid is



(c) the CH_3COO^- ion



The resonance hybrid is



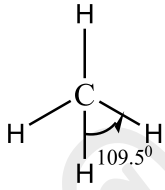
Factors influencing (affecting) bond angles include:

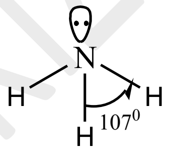
(a) Effect of lone pairs on the bond angle.

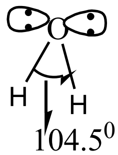
Lone pairs exert a greater repulsion on the adjacent electron pairs and thus tend to compress the angle between bonding pairs.

Hence the lone pair to lone pair repulsion effect is greater than the lone pair to bond pair repulsion effect. The lone pair to bond pair repulsion effect is in turn greater than the bond pair to bond pair repulsion effect.

e.g. (i) the bond angles in methane, ammonia and water are 109.5^0, 107^0 and 104.5^0 respectively.





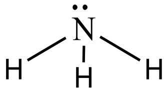


- The bond angle in ammonia (NH_3) is approximately 107 degrees, which is slightly less than the ideal tetrahedral angle of 109.5 degrees (the tetrahedral arrangement puts into consideration all the four electron pairs around the central atom i.e. both bond pairs and the lone pair). This is due to the presence of a lone pair of electrons on the nitrogen atom in ammonia which repels the bonding pairs more strongly. This pushes the bonding pairs closer together and reducing the H-N-H bond angle from the ideal 109.5° to approximately 107°.

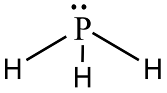
- The bond angle in water (H_2O) is approximately 104.5 degrees, which is less than the ideal tetrahedral angle of 109.5 degrees (the tetrahedral arrangement puts into consideration all the four electron pairs around the central atom i.e. both bond pairs and lone pairs). This is due to the presence of two lone pair of electrons on the oxygen atom in water which repel the bonding pairs more strongly. This pushes the bonding pairs closer together and reducing the H-O-H bond angle from the ideal 109.5° to approximately 104.5°.

(ii) the bond angles in carbondioxide, sulphurdioxide and water are 180^0, 119^0 and 104.5^0 respectively.

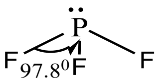
The bond angle in sulphur dioxide is approximately 119 degrees. This is because the molecule has a bent or V-shaped geometry due to the presence of a lone pair of electrons on the sulphur atom.
In sulphurdioxide, the sulphur atom is the central atom, and it has two bonding pairs (with oxygen) and one lone pair of electrons.
If the lone pair was a bond pair, the ideal bond angle would be 120 degrees (trigonal planar). However, the stronger lone pair repulsion pushes the bond pairs closer together causing the molecule to adopt a bent or V-shaped geometry and reduces the bond angle to approximately 119 degrees.
(b) Effect of electronegativity on the bond angle.
(i) If surrounding atoms are less electronegative, more electron density stays closer to the central atom, increasing bond pair to bond pair repulsion and often increasing the bond angle.

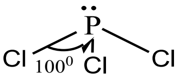
Consider, ammonia (nitrogen trihydride) 

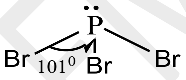
has a bond angle of 107^0.

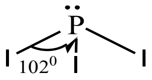
and phosphine (phosphorus trihydride) 

has a bond angle of 93.2^0.
The bond angles are different because the nitrogen atom is more electronegative compared to the phosphorus atom. So, the shared (bonded) electrons are closer to the central nitrogen atom in ammonia, NH_3 than they are to the central phosphorus atom in phosphine, PH_3.
This results into stronger (increased) repulsions between the nitrogen to hydrogen bond pairs in ammonia than the phosphorus to hydrogen bond pairs in phosphine hence creating a larger bond angle.
(ii) If surrounding atoms are more electronegative, they pull bonding electrons towards themselves. This reduces electron density near the central atom, weakening bond pair to bond pair repulsion and hence reducing the bond angle.

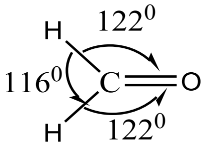
Consider 







Fluorine is the most electronegative and therefore pulls electrons from phosphorus most strongly. Chlorine is the 2nd most electronegative followed by bromine and iodine in the least.
i.e. F > Cl > Br > I
At a certain distance from phosphorus, there's repulsion. However, as electrons move out, the repulsion forces also decrease and the bond angle will correspondingly be larger.
Thus, the shared (bonded) electrons move furthest away from the central phosphorus atom in phosphorus trifluoride, PF_3 than they are to the central phosphorus atom in phosphorus trihalides.
This results into weaker (decreased) repulsions between the phosphorus to fluorine bond pairs in phosphorus trifluoride than other phosphorus trihalide bond pairs hence creating a smaller bond angle.
(c) Effect of multiple bonds on bond the angle.
Multiple bonds (double or triple bonds) have a higher electron density than single bonds.
This makes the region of space around them bulkier (electron cloud is more concentrated), so they push single bonds away more strongly.

e.g in methanal 

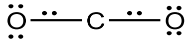
The carbonyl, C=O (carbon to oxygen) double bond has a higher electron density than the carbon to hydrogen single bond. Thus, the double bond pair more strongly repels the single bond pair. This reduces the expected angle of 120° between the two single bonds to 116° while the angle between the double bond and single bond is increased to 122°.

Therefore, multiple bonds, like lone pairs exert a greater repulsive effect on adjacent electron pairs than single bonds due to existence of a high electron cloud.

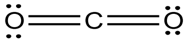
Question: The bond angle of the compounds H_2O, H_2S, SrH_2 and TrH_2 decreases in the following order;
$H_2O > H_2S > SrH_2 > TrH_2$ Explain

Examples:
Sketch the shapes of the following molecules.

1. CO_2 molecule, C = 4 electrons 2O = 2x6 = 12 electrons
4 + 12 = 16 electrons = 8 electron pairs

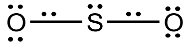


which becomes

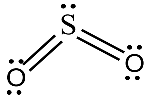


Two bond pairs and no lone pair on the lewis structure makes a **linear shape**.

2. SO_2 S = $6e^-$ 2 O = (2 x 6) = $12e^-$
6 + 12 = $18e^-$ = 9 electron pairs

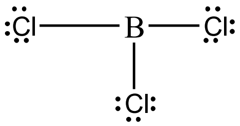


which becomes

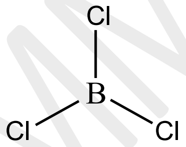


Two bond pairs and one lone pair on the lewis structure makes a **v-shape or bent shape**.

3. BCl_3 B = $3e^-$ 3 Cl = (3 x 7) = $21e^-$
3 + 21 = $24e^-$ = 12 electron pairs

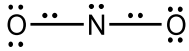


which becomes

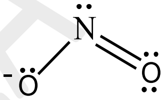


Three bond pairs and no lone pair on the lewis structure makes a **trigonal planar shape**.

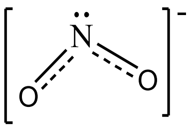
4. NO_2^- N = $5e^-$ 2 O = (2 x 6) = $12e^-$
5 + 12 + 1 = $18e^-$ = 9 electron pairs



which becomes

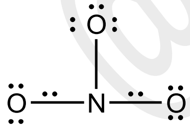


and finally

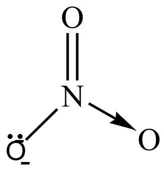


Two bond pairs and one lone pair on the lewis structure makes a **V-shape or bent shape**.

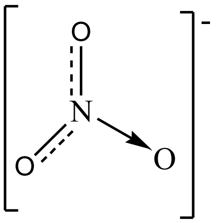
5. NO_3^- N = $5e^-$ 3 O = (3 x 6) = $18e^-$
5 + 18 + 1 = $24e^-$ = 12 electron pairs



which becomes

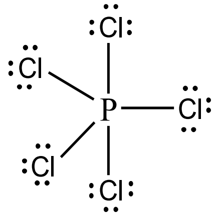


and finally

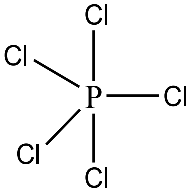


Three bond pairs and no lone pair on the lewis structure makes a **trigonal planar shape**.

6. PCl_5 P = $5e^-$ 5 Cl = (5 x 7) = $35e^-$
5 + 35 = 40 electrons = 20 electron pairs

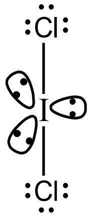


which forms

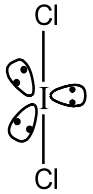


Five bond pairs and no lone pair on the lewis structure makes a **trigonal bipyramidal shape**.

7. ICl_2^- I = $7e^- + e^- = 8e^-$ 2Cl = $(2 \times 7) = 14e^-$
$8 + 14 = 22e^- = 11$ electron pairs

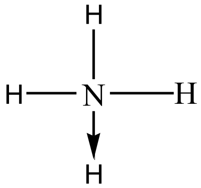


which forms

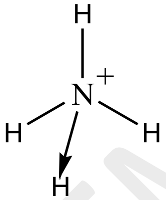


Two bond pairs and three lone pairs on the lewis structure makes a **V-shape or bent shape**.

8. NH_4^+ N = $5e^- - e^- = 4e^-$ 4 H = $(1 \times 4) = 4e^-$
$4 + 4 = 8e^- = 4$ electron pairs

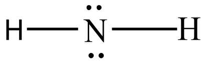


which becomes

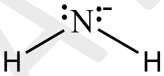


Four bond pairs and no lone pair on the lewis structure makes a **tetrahedral shape**.

9. NH_2^- N = 5e 2 H = $(1 \times 2) = 2e^-$
$5 + 2 + 1 = 8e^- = 4$ electron pairs



which becomes



Two bond pairs and two lone pairs on the lewis structure makes a **V-shape or bent shape**.

Sample Item:

At Kibuli Secondary School, a Senior Four chemistry class is carrying out a discussion on molecular shapes during a lesson on chemical bonding. Their teacher, Mr.Okello, divides the learners into groups and gives them molecular models of common substances, including water (H_2O) and ammonia (NH_3). Each group is asked to observe the shapes and estimate the bond angles using the models.

During the activity, one group notices that the shape of the water molecule appears more "bent" compared to the ammonia molecule, which looks more "pyramidal." After measuring with their model kits, they estimate the bond angle in water to be about 105° and that in ammonia to be about 107°. Curious about this difference, the learners begin discussing possible reasons.

One student suggests that the number of hydrogen atoms might be responsible, while another thinks the central atom could influence the shape. Mr. Okello then challenges the class to think deeper about the arrangement of electron pairs around the central atoms and how this affects molecular geometry.

As a learner in this class, use your understanding of electron pair repulsion and molecular structure to explain why the bond angle in water is smaller than that in ammonia, even though both molecules have central atoms bonded to hydrogen atoms.

Suggested Response:

In water, the oxygen atom has ***two lone pairs and two bond pairs***. The lone pairs strongly repel the bonds to form a ***v-shape*** (bent shape).

In ammonia, the nitrogen atom has ***one lone pair and three bond pairs***. The lone pair repels repel the three bonds strongly to form a ***trigonal pyramidal shape***.

However, *the two lone pairs in water cause a* ***greater repulsion of bonds*** *than the one lone pair* in ammonia hence reducing the bond angle.

Assignment Item:

A group of Senior Four learners from a rural secondary school in Wakiso District visited a nearby water treatment and industrial research facility as part of their chemistry field study. During the visit, they observed that different chemical substances are used in processes such as water purification, gas treatment, and industrial synthesis.

The guide explained that the **shape of molecules and ions** plays a key role in determining how these substances behave, including how they dissolve, react, and interact with pollutants. The learners became curious about how scientists determine these shapes using bonding theories such as the Valence Shell Electron Pair Repulsion (VSEPR) theory.

Back in class, their teacher asked them to investigate the molecular geometry of some of the substances they encountered during the trip.

Task: Using your knowledge of chemical bonding and VSEPR theory:

(i) Draw the possible structures of the following chemical species.

(ii) Name the molecular shape of each species.

(iii) State the bond angles where applicable.

The substances to be analyzed are:

(a) $BeCl_2$

(b) HCN

(c) $HgCl_2$

(d) SO_4^{2-}

(e) SO_3

(f) NCl_3

STRUCTURE OF SOLIDS (CRYSTAL LATTICES)

Solids are substances where particles (atoms, ions, or molecules) are closely packed in a fixed arrangement. The way these particles are arranged and bonded determines the solid's properties, such as melting point, hardness, and electrical conductivity.

Solid lattices are of two types; (i) Simple structures (ii) Giant structures

Crystals are a type of solid material composed of atoms or groups of atoms that are arranged in a three-dimensional pattern that is very ordered. In a crystal, the groups of atoms are repetitive at evenly spaced intervals, all maintaining their orientation to one another. In other words, the geometric shape of a crystal is highly symmetrical. The word 'symmetrical,' means simply a perfect proportion and balance of atoms in a crystal.

A **lattice** is the regular arrangement of atoms, ions or molecules in a crystal. These atoms or groups of atoms are commonly referred to as points within a crystal lattice site. Thus, a crystal lattice site is one containing a series of points arranged in a specific pattern with high symmetry. These points don't tell the position of an atom in a crystal but are simply points 'in space' oriented in such a way to build a lattice or structure.

(1) SIMPLE STRUCTURES (LATTICES)

They include either simple molecular structures or simple atomic structures.

- **Description:**

Simple molecular or atomic lattice is a crystal **lattice** made discrete **molecules** bonded together by intermolecular forces including the weak Van der Waal's forces and the relatively strong hydrogen bonds.

These are exhibited by most covalent compounds as simple molecular structures and most nonmetals such as noble gases as simple atomic structures since they are monoatomic substances.

- **Properties:**
  - They have low melting and boiling points.
  - They are usually soft.
  - They are non-electrolytes except in a few simple polar molecules.
  - They are often gases or liquids at room temperature (some are soft solids).
- **Explanation:**

The molecules themselves are stable, but the forces between them especially the van der Waals forces are weak, and so little energy is needed to separate them. A few molecules form hydrogen bonds.

(2) GIANT STRUCTURES (LATTICES)

The term giant lattice is used to describe the large continuous arrangement of particles (atoms or ions). The binding forces between the particles may be directional (covalent like in diamond) or non-directional (in metals and ionic compounds).

These are categorized into three;

- Giant metallic structures.
- Giant ionic or crystalline structures
- Giant covalent structures or molecular structures

(a) Giant Metallic Structures.

◆ **Description:**

These are established by metals and consist of a **lattice of positive ions** surrounded by a “sea” of **delocalized electrons**.

The bonding forces within a metal arise from the delocalization of the valence electrons of the metal atoms. The metal atoms (ions) are arranged with the delocalised valence electrons which are free to move within the metal. The atoms (ions) in these metals are assumed to be spherical and identical in size and they adopt close packed structures that form either a regular ordered arrangement or an irregular distorted arrangement.

◆ **Properties:**

- They have high melting and boiling points.
- They are good conductors of electricity and heat.
- They are malleable (can be hammered into shape).
- They are ductile (can be drawn into wires).

◆ **Explanation:**

The **metallic bond** (electrostatic force of attraction between positive ions and delocalised electrons) is strong and also allows electrons to move freely, enabling conductivity.

(b) Giant Ionic Structures.

◆ **Description:**

These are made of **positive and negative ions** arranged in a regular lattice, held together by **strong electrostatic forces**.

These are exhibited by compounds formed from metals and nonmetals and consist of ionic compounds of the type MX. e.g. NaCl.

The ionic compound consists of a symmetrical array of many positive ions surrounding the negative ion and vice versa. Usually, each ion is surrounded by the greatest number of oppositely charged ions called the coordination number. This is so in order to maintain electrical neutrality and to ensure maximum lattice energy.

The binding force in an ionic solid is the strong ionic bond which is an electrostatic force of attraction between the oppositely charged ions.

The main factors which determine the type of arrangement of ions adopted by an ionic compound are:

(i) the relative charges on the ions which dictates the relative numbers of ions present.

(ii) the relative ionic radius.

◆ **Properties:**

- They have high melting and boiling points.
- They are hard but brittle.
- They are electrolytes. Conduct electricity **only when molten or dissolved in water.**
- They are usually soluble in water.

◆ **Explanation:**

The strong ionic bonds require a lot of energy to break. However, when melted or dissolved, ions can move freely, allowing electrical conduction.

(c) Giant covalent (atomic or molecular) structure.

◆ **Description**

These consist of a **network of atoms bonded together by strong covalent bonds** throughout the entire structure.

The giant atomic structure is shown by a few non-metals particularly carbon (diamond and graphite), silicon and germanium while the giant molecular structure is shown by a few covalent compounds particularly silicon dioxide, germanium dioxide.

(i) Diamond. This has a tetrahedral arrangement. Each carbon atom is bonded to four others by covalent bonds giving a coordinate number of 4. The C-C bond length is 0.154nm which is the normal length of a single covalent bond between two carbon atoms.

It is very strong because it has very many strong covalent bonds infinitely tetrahedrally arranged in a giant three-dimensional structure. Hence it is used to cut glass and makes drill tips (rock borers). It is also used as abrasive for smoothing very hard materials.

Silicon has a similar arrangement.

(ii) Graphite is an example of a layer structured solid. Within each layer, each carbon atom is covalently bonded in a planar arrangement to three others. The many bonds are arranged as hexagonal rings. The layers are held together by weak Van der Waal's forces and because these forces are weak, the layers slide over one another. Thus, used as a lubricant.

Graphite has delocalized mobile electrons within its layers hence conducts electricity.

The C-C bond length is 0.142nm which suggests that there is delocalization of the fourth electron that results in the extra bonding. However, the extra bonding is not as powerful as that in benzene in which the bond length between the carbon atoms is 0.139nm.

(iii) Silicon dioxide has a tetrahedral arrangement with each of the silicon atoms surrounded by four oxygen atoms to form a giant three-dimensional structure similar to that of diamond.

◆ **Properties:**

- They have very high melting points and boiling points.
- They are very hard (e.g., diamond).
- They do not conduct electricity except graphite.
- They are insoluble in water.

◆ **Explanation:**

All atoms are linked by short strong covalent bonds in a highly ordered arrangement that makes them to be closely packed, so a lot of energy is required to break the structure.

SUMMARY TABLE:

Structure Type	**Bonding Type**	**Melting Point**	**Conductivity**	**Example**
Simple Molecular or Atomic	Weak intermolecular Van der Waals forces and hydrogen bonds.	Low	No	CO_2
Giant Metallic	Metallic bonds	High	Yes	Copper
Giant Ionic	Ionic bonds	High	Molten/aqueous only	NaCl
Giant Covalent or Molecular	Covalent bonds	Very high	No (except graphite)	Diamond

Key Idea:

The **type of bonding and structure** determines the properties of a solid. Stronger bonds → higher melting point and greater hardness.