

## BONDING AND STRUCTURES

### Introduction:

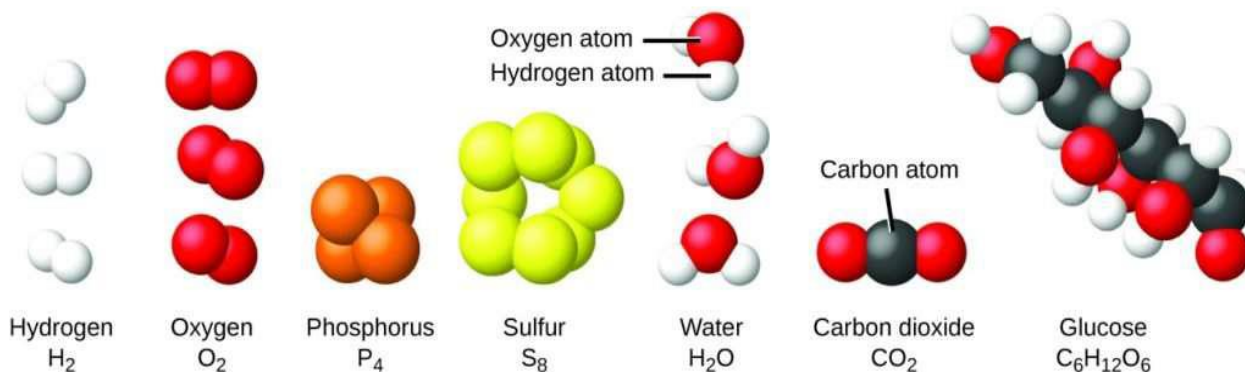
Chemical bonding is one of the basic fundamentals of chemistry that explain other concepts such as chemical reactions. An atom consists of the nucleus containing protons and neutrons and electrons in certain energy levels rotating around the nucleus. In chemical bonding, only the valence electrons (electrons located in the outermost energy levels) of an atom are involved.

Chemical bonding is the fundamental force that holds atoms, ions, or molecules together to form compounds. It's the glue that dictates how elements combine and interact, ultimately explaining macroscopic properties and chemical reactions.



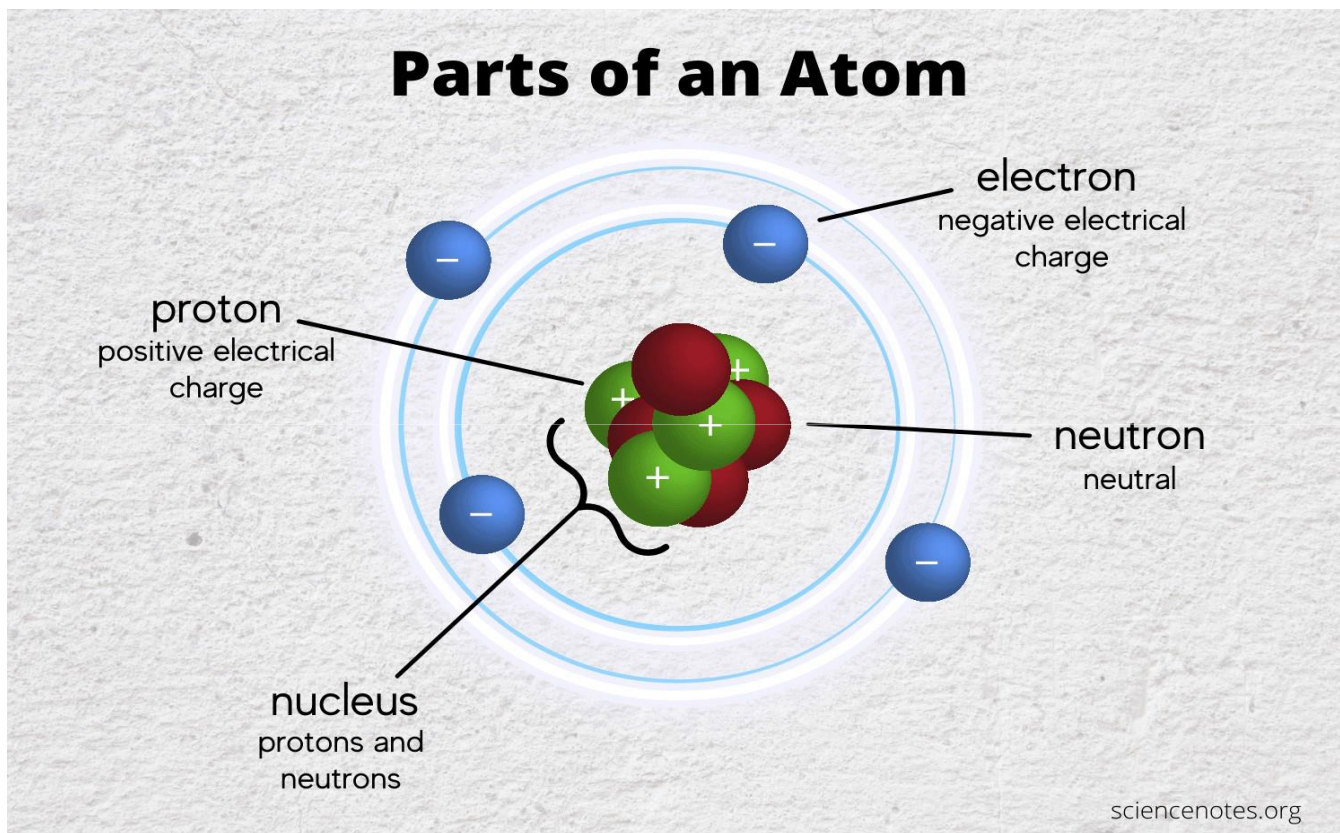
### Why do Atoms Bond?

Atoms bond to achieve a more stable electron configuration, typically resembling that of the noble gases (Group 8 elements), the electronic theory of valence/electronic theory of chemical bonding. Noble gases have a full valence electron shell (usually eight valence electrons, except for Helium which has two), making them very stable and unreactive. Atoms tend to gain, lose, or share valence electrons to attain this stable arrangement, often following the **Octet Rule** *which is tendency to have eight electrons in the outermost energy level.*



### Involvement of Electrons:

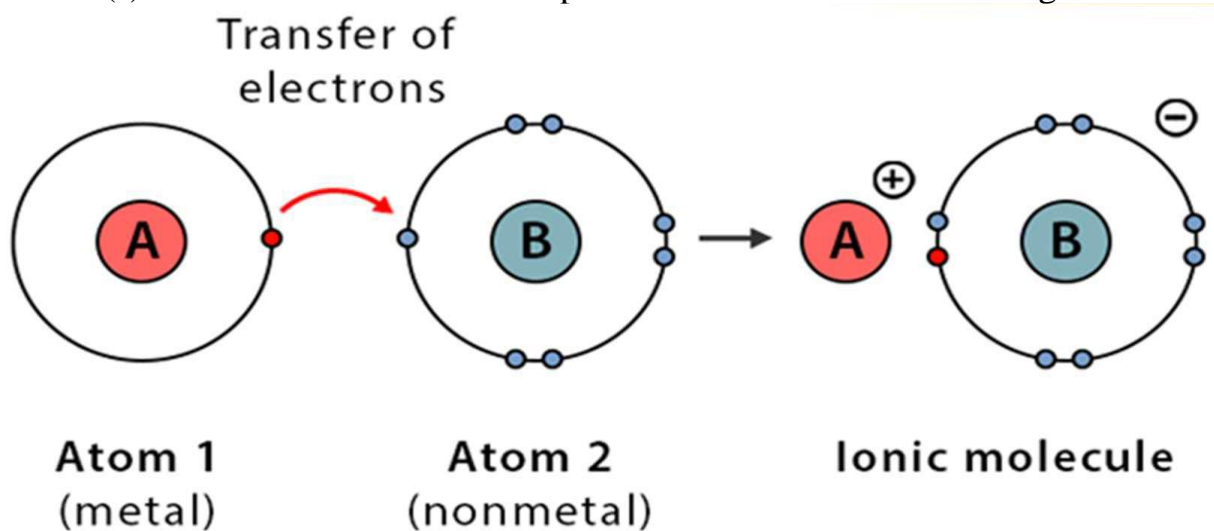
Normally, only the valence electrons, that is the electrons in the outermost energy level are the one that participate in chemical bonding. The inner core electrons and the nucleus (containing protons and neutrons) are generally not directly involved in forming bonds, although the nucleus's positive charge influences the attraction for valence electrons.



### How do Atoms Bond?

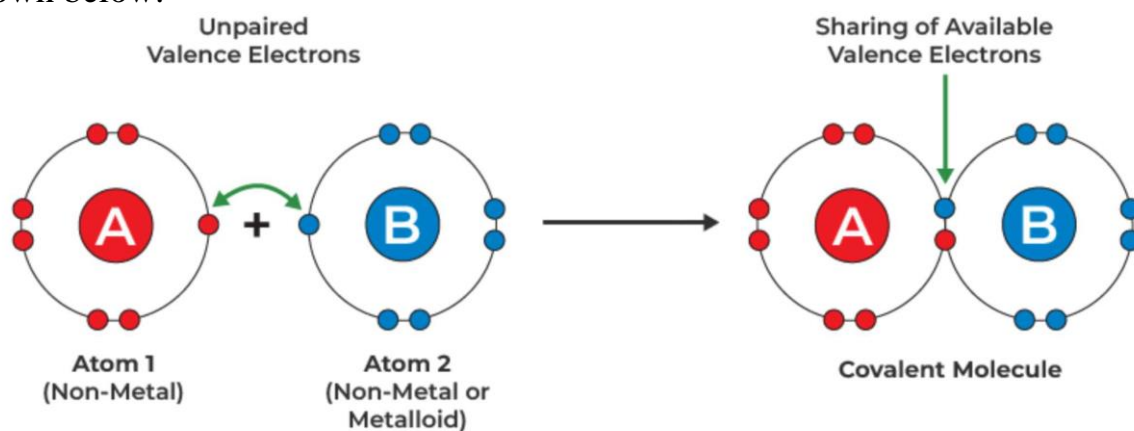
There are three primary ways atoms achieve stability through bonding: i.e.

1. **Electron Transfer**: One atom gives electrons to another or there is complete transfer of electron(s) from one to the next. Example in formation of ionic bonding.



## 2. Electron Sharing:

Atoms bond by sharing the valence electrons leading to formation of covalent bonding as shown below.



## 3. Electrostatic Attraction in a 'Sea' of Electrons:

Here positive ions are attracted to a pool of delocalized valence electrons as shown in the formation of metallic bonding.

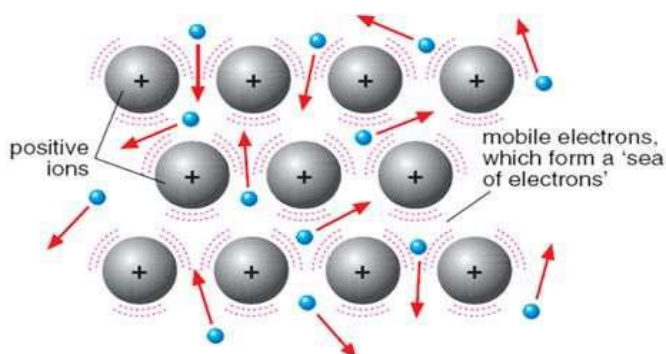


Figure 6.27 Bonding in metals

## Lewis Dot Symbols:

The American chemist G.N. Lewis devised a system of symbols called Lewis electron dot symbols that can be used to predict the number of bonds formed by most elements in their compounds. Each Lewis Dot symbol consists of the chemical symbol of the element surrounded by dots. The dots represent the valence electrons of an element.

**The following steps can be followed to draw Lewis Dot symbols of different elements**

1. Identify the number of valence electrons for the element. These are the total number of electrons in the outermost energy level. They can be known by the electronic configuration
2. Write the correct chemical symbol of the element
3. Put one dot in all directions; up, down, left and right of the element symbol
4. If some electrons are left, begin doubling the dots

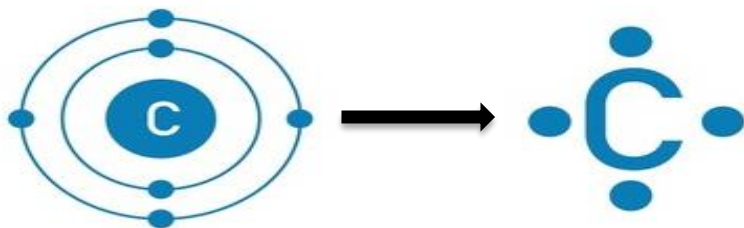
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From the Lewis Dot symbol of each element, we can predict the number of bonds an element can form during bonding.

*Number of bonds formed by an element = the number of unpaired electrons present in its Lewis Dot symbol*

**Example:**

1. Carbon (Group 4) has 4 valence electrons. Its Lewis symbol is



Therefore, carbon has 4 unpaired electrons and therefore forms 4 bonds.

2. Chlorine a group 7 element has 7 valence electrons. Its Lewis symbol is



chlorine has 1 unpaired electron (valence) and therefore, forms 1 bond.

### Activity 1

Draw the Lewis Dot symbols for the elements Boron, Carbon, Nitrogen and Oxygen

<i>Element</i>	<i>Atomic number</i>	<i>Electronic configuration</i>	<i>Number of valence electron</i>	<i>Lewis dot symbol</i>	<i>Number of bonds that can be formed</i>
<i>Boron</i>					
<i>Carbon</i>					
<i>Nitrogen</i>					
<i>oxygen</i>					

## INTRODUCTION AND FORMATION OF IONIC BONDING

An ionic bond, also known as an electrovalent bond, is formed by the complete transfer of one or more valence electrons from one atom to another.

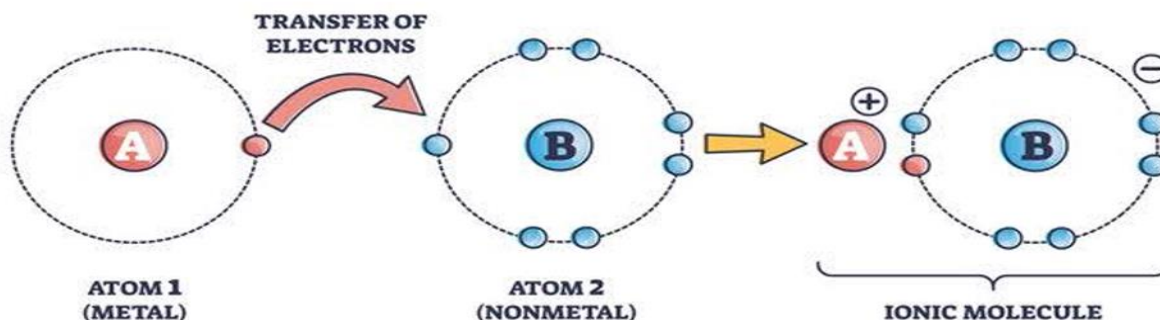
This type of bonding typically occurs between elements with a large difference in electronegativity: Usually a metal (low electronegativity, tends to lose electrons) and a non-metal (high electronegativity, tends to gain electrons).

### Formation of ionic bonding:

The metal atom loses valence electrons to become a positively charged ion called a cation. It is considered electropositive.

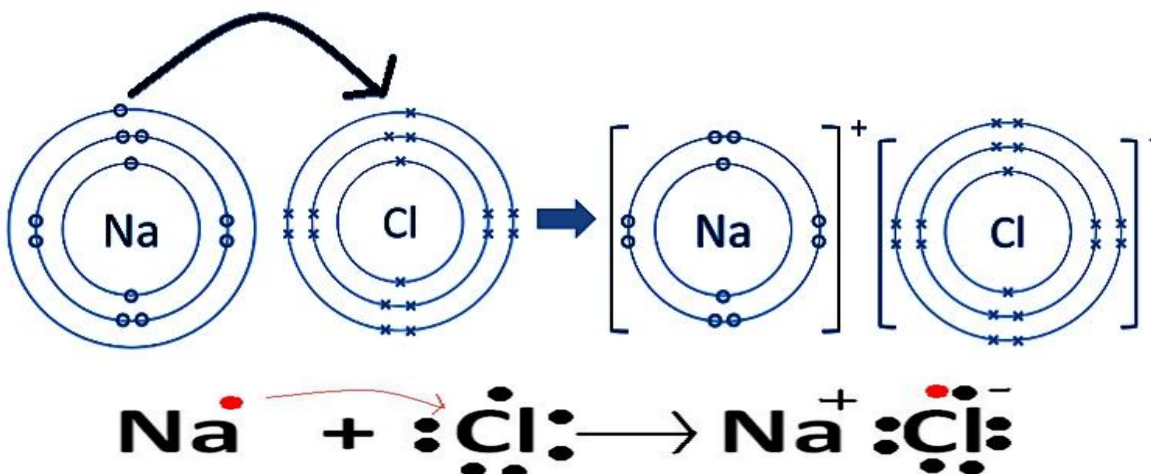
The non-metal atom gains these electrons to become a negatively charged ion called an anion. It is considered electronegative.

*The cation and anion formed are held together by strong electrostatic forces of attraction.*



### Illustrative Examples

#### 1. Formation of Sodium Chloride (NaCl):



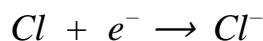
Sodium (Na), a Group 1 metal, has electron configuration  $1s^2 2s^2 2p^6 3s^1$ . It has 1 valence electron. Chlorine (Cl), a Group 7 non-metal, has electron configuration  $1s^2 2s^2 2p^6 3s^2 3p^5$ . It has 7 valence electrons.

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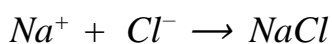
Na readily loses its  $3s^1$  electron to achieve a stable  $1s^22s^22p^6$  configuration (like Neon), forming a  $Na^+$  cation:



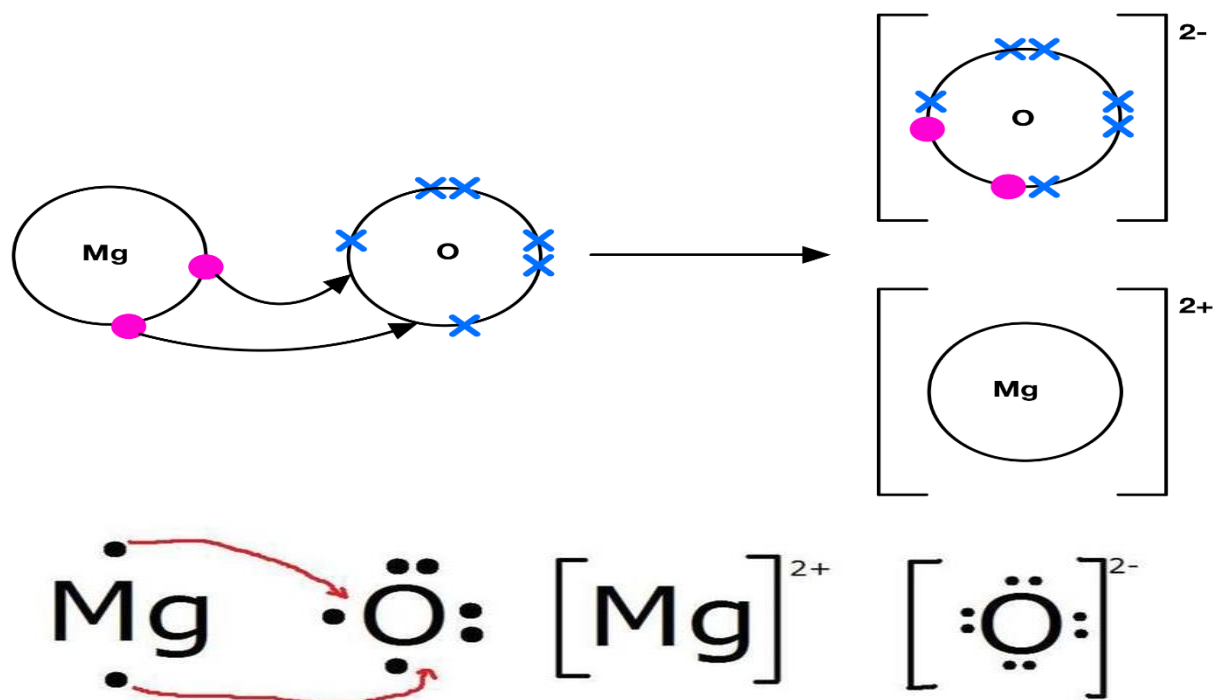
Cl readily gains this electron to achieve a stable  $1s^22s^22p^63s^23p^6$  configuration (like Argon), forming a  $Cl^-$  anion:



The resulting  $Na^+$  and  $Cl^-$  ions are then held together by strong electrostatic attraction:



### 2. Formation of Magnesium Oxide (MgO):



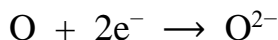
Magnesium (Mg), a Group 2 metal, has electron configuration  $1s^22s^22p^63s^2$ . It has 2 valence electrons. Oxygen (O), a Group 6 non-metal, has electron configuration  $1s^22s^22p^4$ . It has 6 valence electrons.

Mg loses its two  $3s^2$  electrons to become a stable  $Mg^{2+}$  cation with electronic configuration  $1s^22s^22p^6$ :

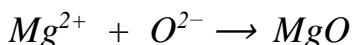


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Oxygen gains these two electrons to become a stable  $O^{2-}$  anion with electronic configuration  $1s^2 2s^2 2p^6$ :



The  $Mg^{2+}$  and  $O^{2-}$  ions are held by strong electrostatic attraction:

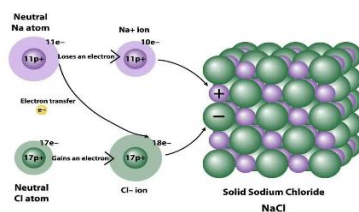
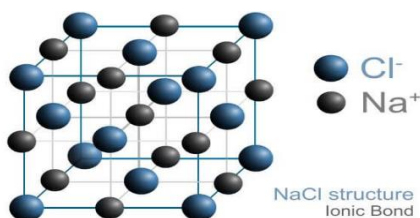


### Structure of ionic compound

Ionic compounds do not exist as discrete molecules. Instead, the cations and anions arrange themselves in a regular, repeating three-dimensional structure called a giant ionic lattice. In this lattice, each ion is surrounded by a specific number of oppositely charged ions, known as its coordination number. The strong electrostatic forces extend throughout the entire crystal.

### Examples

#### Sodium chloride



### Properties of ionic compounds:

#### 1. High Melting and Boiling Points:

Breaking down the giant ionic lattice requires overcoming the very strong electrostatic forces of attraction between the ions. This requires a large amount of thermal energy; hence ionic compounds have high melting and boiling points.

#### 2. Electrical Conductivity:

Ionic solids do not conduct electricity. The ions are held in fixed positions within the rigid lattice and are not free to move and carry charge.

Ionic compounds conduct electricity when they are melted or dissolved in polar solvents (like water). In these states, the ions are no longer held in fixed positions and therefore become mobile. These mobile ions can then move towards electrodes of opposite charge, carrying electric current.

#### 3. Solubility:

Many ionic compounds are soluble in polar solvents. Polar solvent molecules like water, with  $\delta^+$  and  $\delta^-$  ends are attracted to the ions on the surface of the crystal. These ion-solvent attractions (hydration) are strong enough to overcome the electrostatic forces holding the

*lattice together, pulling the ions into solution where they become surrounded by solvent molecules.*

*Ionic compounds are generally insoluble in non-polar solvents. Non-polar solvent molecules have little or no charge separation and cannot effectively interact with or separate the charged ions in the lattice. The ion-ion interactions in the crystal are much stronger than any potential ion-solvent interactions.*

#### 4. Physical State:

*Ionic compounds exist as crystalline solids at room temperature due to the strong electrostatic forces locking the ions into a fixed, ordered lattice structure.*

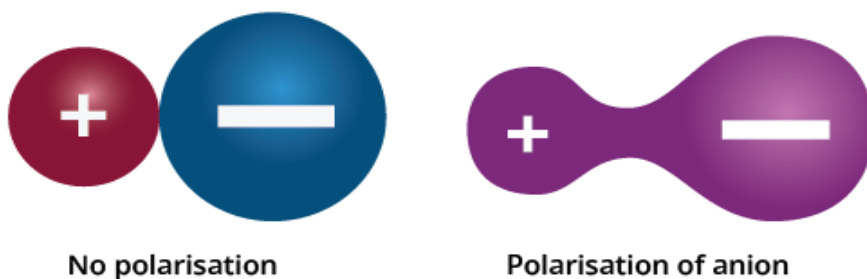
*Ionic crystals are brittle. If a force is applied that causes layers of ions to shift, ions of the same charge may come into close proximity. The strong electrostatic repulsion between these like-charged ions causes the crystal to cleave or break.*

### Covalent Character in Ionic Compounds

While considered ionic, some compounds exhibit properties suggesting a degree of covalent character. **This is explained by:**

#### 1. Polarising Power of the Cation:

The ability of the positive cation to distort the electron cloud of the negative anion. A highly polarising cation pulls the anion's electron density towards itself, leading to some sharing of electrons (covalent character). The greater the polarising power of the cation, the greater the tendency of the cation to form a covalent bond.



#### Polarising power increases with:

*-Higher charge on the cation, The larger the positive charge on the ion, the greater the attraction of the valence electrons. Example  $Al^{3+}$  ion therefore has a higher polarising power than the  $Na^+$  due to its higher charge.*

*-Smaller size of the cation, The smaller the ionic radius, the higher the polarising power of the cation. Example  $Al^{3+}$  ion therefore has a higher polarising power than the  $Na^+$  due to its smaller ionic radius.*

*-Charge Density: The ratio of ionic charge to ionic radius. The higher the charge density, the higher the polarising power of the cation. Example; charge density increases from  $Na^+$  to  $Mg^{2+}$  to  $Al^{3+}$ , correlating with increasing covalent character in their chlorides from  $NaCl$ , to  $MgCl_2$ , to  $AlCl_3$ .*

## 2. Polarisability of the Anion:

The ease with which the electron cloud of the anion can be distorted by the cation. The smaller the anion, the lower it is polarizable and the larger the anion, the easier it is polarized. Larger anions have their valence electrons further from the nucleus and are less tightly held, making them more easily polarizable. Polarisability increases with increasing anion size e.g.,  $I^- > Br^- > Cl^- > F^-$ .

### Effect of covalent characters on Properties of ionic compounds:

Compounds with significant covalent character tend to have lower melting points and may be more soluble in non-polar solvents than purely ionic compounds e.g.,  $AlCl_3$  sublimes at a relatively low temperature and dissolves in methylbenzene, unlike  $NaCl$ .

### Activities

1. The Ministry of Works and Transport in Uganda is piloting a road improvement project in the Karamoja sub-region, where dust storms and dry, loose soils make travel difficult. To stabilize the roads and control dust, engineers are testing different salts including sodium chloride, magnesium chloride, and aluminium chloride as additives to bind the soil. The team must also consider that road temperatures in the region can soar during the dry season, sometimes exceeding  $45^\circ C$  on the surface. You are part of a secondary school science club tasked with advising the engineers on which compound would perform best under these harsh conditions.

You are provided with the following data:

Chloride	$NaCl$	$MgCl_2$	$AlCl_3$
Melting Point ( $^\circ C$ )	801	712	180

Task:

State and explain the trend in the melting points of the chlorides of the three metals.

Melting point decrease in the order Sodium chloride ( $801^\circ C$ ) > Magnesium chloride ( $712^\circ C$ ) > Aluminium chloride ( $180^\circ C$ )

Explanation:

Sodium chloride and magnesium chloride have higher melting point because they have giant ionic structures held together by strong electrostatic forces between oppositely charged ions. Magnesium chloride has a slightly lower melting point than sodium chloride due to a higher charge density of  $Mg^{2+}$ , which can distort the chloride ion cloud slightly, introducing a bit of covalent character.

Aluminium chloride has a significantly lower melting point because it is largely covalent and exists as discrete  $Al_2Cl_6$  molecules at room temperature, which are held together by weak Van der Waals forces, not strong ionic bonds.

Therefore, the melting point decreases from  $NaCl$  to  $AlCl_3$  as the bonding changes from pure ionic to covalent, and the structure changes from giant ionic lattice to simple molecular.

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Based on the melting points and bonding structure of the compounds, recommend which chloride(s) would be more stable and effective for use in soil stabilization and dust control in hot climates like Karamoja, and which ones might decompose or evaporate more easily.

*Sodium chloride and magnesium chloride are more suitable for use in dust control and soil stabilization in hot areas like Karamoja. NaCl and MgCl<sub>2</sub> have high melting points (801°C and 712°C), which means they remain thermally stable even when road temperatures rise to 45°C or more. Their ionic lattices make them durable and less likely to degrade, which is crucial for long-term road performance.*

*Aluminium chloride is not suitable because its covalent nature and low melting point (180°C) makes it likely to sublime or decompose in the high surface temperatures common in such regions making it ineffective.*

Explain how the bonding type and lattice structure of each chloride relates to its melting point and suitability in long-term road construction in Uganda.

*Sodium chloride and magnesium chloride consist of giant ionic lattices with strong electrostatic attractions between metal cations (Na<sup>+</sup>, Mg<sup>2+</sup>) and chloride anions (Cl<sup>-</sup>).*

*These strong ionic bonds require a lot of energy to break, hence their high melting points and suitability for high-temperature environments.*

*In contrast, aluminium chloride is mostly covalent. It forms discrete molecular units (Al<sub>2</sub>Cl<sub>6</sub>), with weak intermolecular forces that are easy to break, explaining its low melting point.*

*This means ionic compounds like NaCl and MgCl<sub>2</sub> are more stable and heat-resistant, making them better suited for applications such as dust control and soil stabilization in Uganda's hot, dry climates.*

*Understanding the bonding type and structure helps engineers choose materials that will not decompose or lose effectiveness under environmental stress.*

2. A fishing community on Buvuma Island along Lake Victoria traditionally preserves fish by rubbing salt on the surface before sun-drying. A visiting student suggested soaking fish in oil mixed with salt to enhance preservation, but the fish spoiled quickly.

Later, a science teacher explained that salt (sodium chloride) did not dissolve in oil, which prevented it from penetrating the fish and stopping bacterial growth resulting into its spoilage quickly.

Task:

Using your knowledge of bonding and solubility, explain why salt dissolves in water (present in fish tissues) but not in oil (a non-polar substance like benzene).

Based on your explanation, clarify why the traditional method using dry salt works better than the oil-based method.

Suggest how this knowledge can be used to improve food preservation or hygiene practices in rural Ugandan settings without access to refrigeration.

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### Applications of the concept of ionic bonding in our daily life:

Ionic bonding, a fundamental chemical concept involving the complete transfer of valence electrons between atoms to form oppositely charged ions held together by electrostatic attraction, is responsible for the existence of many substances we encounter and utilize daily. The unique properties of ionic compounds, such as high melting and boiling points, hardness, and ability to conduct electricity when dissolved or molten, directly influence their applications in various aspects of our lives.

1. One of the most common examples is sodium chloride (NaCl), commonly known as table salt. The ionic bond between sodium cations ( $\text{Na}^+$ ) and chloride anions ( $\text{Cl}^-$ ) forms this essential compound. In our daily lives, sodium chloride is crucial for:

- ✓ *It is used as a flavor enhancer and a preservative, inhibiting the growth of bacteria and extending the shelf life of food products.*
- ✓ *Sodium and chloride ions play vital roles in nerve impulse transmission, muscle contraction, and maintaining fluid balance.*
- ✓ *Large quantities of sodium chloride are used to de-ice roads and sidewalks in colder climates due to its ability to lower the freezing point of water.*
- ✓ *It is a raw material in the production of chlorine and sodium hydroxide, which are used in numerous industrial processes, including manufacturing plastics, paper, and detergents.*

2. Another common ionic compound is calcium carbonate ( $\text{CaCO}_3$ ), found naturally in rocks like limestone, marble, and chalk, as well as in seashells and eggshells. Its applications stemming from its ionic nature include:

- ✓ *It is a primary component of cement, concrete, and mortar, serving as a crucial building material. Marble and limestone are also used as decorative and structural elements.*
- ✓ *Calcium carbonate is used in some antacids to neutralize excess stomach acid, providing relief from heartburn and indigestion.*
- ✓ *It is used to reduce soil acidity and provide calcium, an essential nutrient for plant growth.*
- ✓ *Calcium carbonate is an ingredient in toothpaste as an abrasive to help remove plaque and in some paints, plastics, and paper as a filler and whitener.*

3. Potassium chloride (KCl) is another ionic compound with several daily applications:

- ✓ *It is often used as a sodium-free substitute for table salt for individuals looking to reduce their sodium intake.*
- ✓ *Potassium is a vital nutrient for plant growth, and potassium chloride is a common component of fertilizers.*
- ✓ *It is used to treat and prevent low blood potassium levels (hypokalemia).*

4. Magnesium oxide ( $\text{MgO}$ ), formed from magnesium cations ( $\text{Mg}^{2+}$ ) and oxide anions ( $\text{O}^{2-}$ ), also finds its way into daily use:

- ✓ *It is used as an antacid and a laxative medicine.*

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✓ *Due to its high melting point, it is used in the lining of furnaces and kilns.*

5. Calcium oxide (CaO), also known as quicklime, is produced by heating calcium carbonate and is used in:

- *It is a key ingredient in the production of cement and mortar.*
- *It is used to treat acidic soils.*
- *It is used in the purification of sugar and in the manufacturing of paper and glass.*

### Examples of substances used in our daily life



## METALLIC BONDING

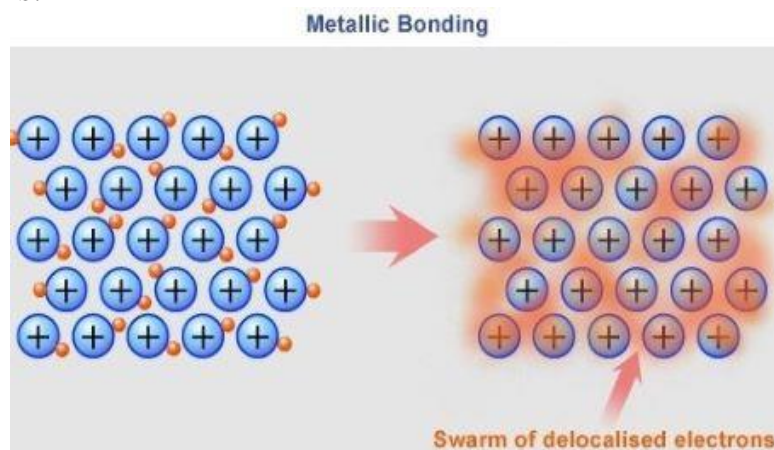
Metallic bonding is the electrostatic force of attraction between positively charged metal ions and a 'sea' of delocalized valence electrons.

Occurs exclusively between atoms of metallic elements. Metal atoms are arranged in a regular, repeating pattern called a crystal lattice. Each metal atom typically has one or more loosely held valence electrons.

In the metallic structure, the valence electrons of each metal atom become delocalized. This means they are no longer associated with any specific atom. These delocalized electrons form an "electron sea" that permeates the entire metal lattice. Once the valence electrons are delocalized, the metal atoms become positively charged ions (cations). These metal cations are embedded in the sea of delocalized electrons.

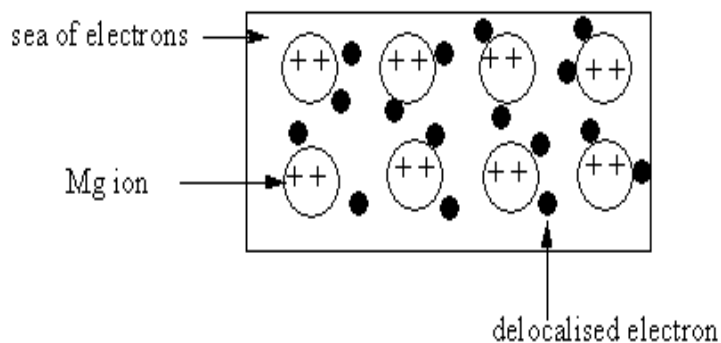
The metallic bond is formed through the electrostatic attraction between the positively charged metal cations and the negatively charged, delocalized electrons. This attraction holds the metal ions together and gives the metal its cohesive structure.

The metallic bond is the strong, non-directional electrostatic attraction between the positively charged metal ions and the collectively shared, negatively charged cloud of delocalized electrons.



## Formation of metallic bonding

### 1. Magnesium metal



## 2. Iron metal

Iron (Fe): In metallic iron, positively charged iron (Fe) ions are held together by a "sea" of delocalized electrons, allowing for high electrical conductivity.



## Properties of Metallic Bonds and its Applications

### 1. High Melting and Boiling Points:

Significant energy is required to overcome the strong electrostatic attraction between the positive metal ions and the delocalized electron sea, leading to high melting and boiling points.

### 2. Electrical Conductivity:

Metals are excellent conductors of electricity in both the solid and liquid (molten) states. The delocalized electrons are free to move throughout the structure. When a voltage is applied, these electrons move towards the positive terminal, constituting an electric current. Conductivity decreases with increasing temperature because increased thermal vibrations of the metal ions impede the flow of electrons.

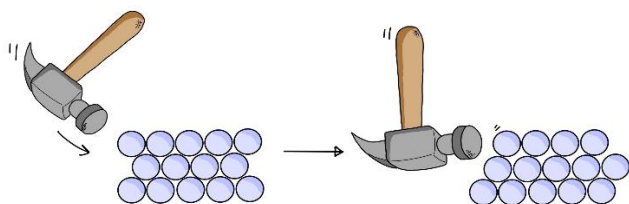
### 3. Thermal Conductivity:

Metals are good conductors of heat. Heat energy is efficiently transferred through the metal by the movement and collisions of the highly mobile delocalized electrons, as well as by vibrations of the metal ions.

### 4. Malleability and Ductility:

Metals are malleable (can be hammered into sheets) and ductile (can be drawn into wires). When a force is applied, the layers of positive metal ions can slide past one another. The delocalized electron sea acts like a mobile glue, maintaining the attraction between the ions even as their positions change, preventing the structure from shattering (unlike ionic solids).

## Illustration



### 5. Lustre:

Metals have a characteristic shiny appearance (lustre). This is because the delocalized electrons can absorb and re-emit light over a range of wavelengths.

### 6. Physical State:

Most metals are solids at room temperature (except mercury) due to the strength of the metallic bond.

### 7. Solubility:

Metals are generally insoluble in polar and non-polar solvents because the strong metallic bonds are much stronger than any potential interactions with solvent molecules.

## Strength of Metallic Bond:

The strength of the metallic bond, and thus properties like melting point, is influenced by:

### 1. Charge of the Metal Ion:

Higher positive charge on the ion means a stronger attraction to the electron sea (e.g., Group 2 metals like Mg have higher melting points than Group 1 metals like Na because Mg forms  $Mg^{2+}$  ions and contributes 2 electrons per atom, while Na forms  $Na^+$  ions and contributes 1 electron per atom).

### 2. Number of Delocalized Electrons per Atom:

More electrons contributed to the sea results in a greater overall attraction (as seen in the Group 1 vs Group 2 comparison). Transition metals often have higher melting points than main group metals because they can involve d-electrons in bonding, contributing more electrons to the electron sea. Example Calcium with electronic configuration  $[Ar]4s^2$  has lower melting point (842 °C) as compared to Chromium with electronic configuration  $[Ar]4s^13d^5$ , (1857 °C), because calcium contributes its two valence electrons from the 4s orbital to the sea of delocalized electrons. So, each Calcium atom contributes 2 electrons, while Chromium's higher melting point is due to contributing both 4s and 3d electrons to the delocalized sea compared to Calcium. A larger number of delocalized electrons per atom leads to stronger attraction between the positive ions and the electron sea.

### 3. Size of the Metal Ion:

Smaller metal ions allow the delocalized electrons to be closer to the positive nuclei, resulting in stronger attraction.

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### Applications:

The unique properties of metals – their ability to conduct electricity and heat, their malleability and ductility, their strength, and their lustrous appearance – all stem directly from the 'sea' of delocalized electrons and the electrostatic attraction to the positive metal ions.

### Electrical Wiring:

Metals like copper and aluminium are excellent conductors of electricity because their delocalized electrons can move freely, carrying electric current. From the power lines bringing electricity to our homes and businesses to the wires inside our phones, computers, and appliances, metallic bonding's conductivity is fundamental to modern life. Think about charging your phone or simply turning on a light – you're relying on metallic bonding.

### Cookware:

Metals like aluminium, iron, and stainless steel are used to make pots, pans, and cooking utensils because they are good conductors of heat. The delocalized electrons efficiently transfer thermal energy from the stove or fire throughout the pot, ensuring food cooks evenly. Their high melting points (a result of strong metallic bonds) mean they can withstand cooking temperatures without melting.

### Construction and Infrastructure:

Steel (mostly iron with carbon) is widely used in building structures, bridges, and reinforcement (like rebar used in concrete) due to its strength, hardness, and durability. These properties come from the strong forces within the metallic lattice. Aluminium is used for roofing sheets due to its light weight, malleability (can be formed into sheets), and corrosion resistance.

### Tools and Machinery:

Shovels, hoes, pangas, hammers, nuts, bolts, and parts of vehicles and industrial machinery are made from metals and alloys (like steel, iron, brass). Their strength, hardness, malleability (to be shaped), and durability, all linked to metallic bonding, make them essential for agriculture, construction, manufacturing, and transport.

### Vehicles:

The frames, engines, wheels, and many components of vehicles are made from metals like steel and aluminium. The strength and rigidity of the metallic structure are crucial for safety and performance. The electrical systems rely on metal wiring (conductivity).

### Coins and Currency:

Coins are made from metal alloys like brass. The metallic bonding contributes to their durability, malleability (can be stamped into shape), and resistance to corrosion, making them suitable for everyday transactions.

### Jewelry:

Precious metals like gold, silver, and platinum are used for jewelry due to their lustre (how they reflect light, related to the delocalized electrons), malleability (can be shaped into intricate designs), and relative inertness (don't easily corrode).

## Examples of common metals



### Sample items

1. A new classroom block is being constructed at Alero Secondary School in Nwoya District. The school has to choose between using metal sheets and plastic roofing sheets. The metal sheets are heavier but strong, shiny, and resistant to breaking, even when hammered into shape. The school engineer explains that metallic bonding gives metals these useful properties.

Tasks:

Using your knowledge of metallic bonding, explain why metals can be hammered into shapes without breaking.

*Metals are malleable because their atoms are arranged in layers that can slide over each other when a force is applied.*

*The presence of delocalised electrons from metallic bonding allows the layers to move without breaking the structure.*

Why are metals stronger and more durable than plastics in construction?

*Metals have strong metallic bonds due to the attraction between metal cations and the sea of delocalised electrons.*

*Plastics have weaker intermolecular forces, making them easier to deform, melt, or break under stress.*

How does metallic bonding contribute to the choice of roofing materials in Uganda's hot and rainy climate?

*Metallic bonding gives metals high melting points, strength, and resistance to weathering. This makes them suitable for Uganda's hot sun and heavy rains, as they can withstand expansion, contraction, and impact.*

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2. In a village near Lira Town, a solar energy project is being set up to power homes. The technicians insist on using copper wires instead of iron or plastic strips for electrical wiring, citing copper's excellent conductivity and ability to withstand heat.

Tasks:

Explain why copper is a good conductor of electricity, using your knowledge of metallic bonding.

Why would plastic strips not work as substitutes for metal in wiring?

Suggest another metallic material that could be used for wiring and justify your choice based on its bonding and properties.

3. A blacksmith in Masindi District is famous for making strong, durable hoes and machetes from scrap metal. His tools remain tough even after repeated heating and hammering. A science club from a nearby school visits his workshop and learns that these properties are due to metallic bonding.

Tasks:

Explain how metallic bonding allows metals to stay strong at high temperatures during forging.

Why do the tools not shatter when hammered, unlike glass?

How can this knowledge of metallic bonding help young innovators in Uganda's metal fabrication industry?

## COVALENT BONDING AND ITS FORMATION

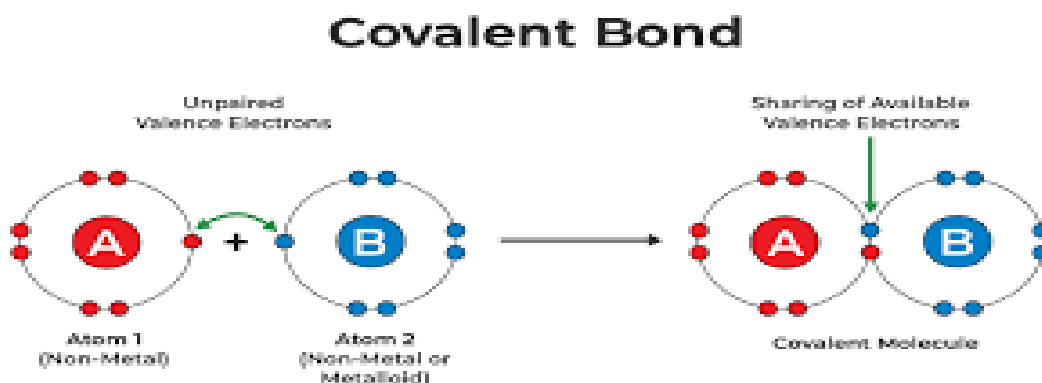
A covalent bond is formed by the sharing of valence electrons between two atoms. Instead of transferring electrons completely, atoms contribute one or more valence electrons each to form shared pairs of electrons that are attracted to the nuclei of both atoms. These shared pairs constitute the covalent bond.

Atoms share electrons to achieve a stable electron configuration, typically fulfilling the Octet Rule (having eight valence electrons in their outer shell, including the shared electrons). Hydrogen is an exception, aiming for a duet (two valence electrons).

Covalent bonds are commonly represented by lines between the bonded atoms. A single line represents one shared pair (a single bond), a double line represents two shared pairs (a double bond), and a triple line represents three shared pairs (a triple bond). Unshared valence electrons (those not involved in bonding) are shown as dots around the atoms and are called lone pairs.

Covalent bonds primarily form between non-metallic elements, although they can also occur between a non-metal and certain metals with high electronegativity or high charge density (leading to covalent character in otherwise ionic compounds, as discussed earlier).

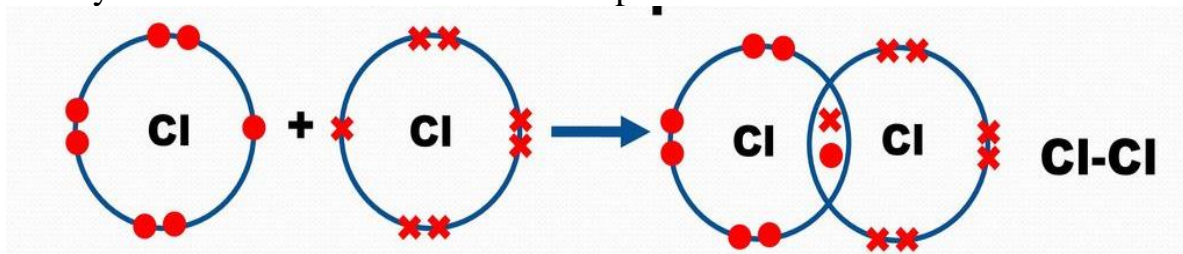
Illustration:



### Examples

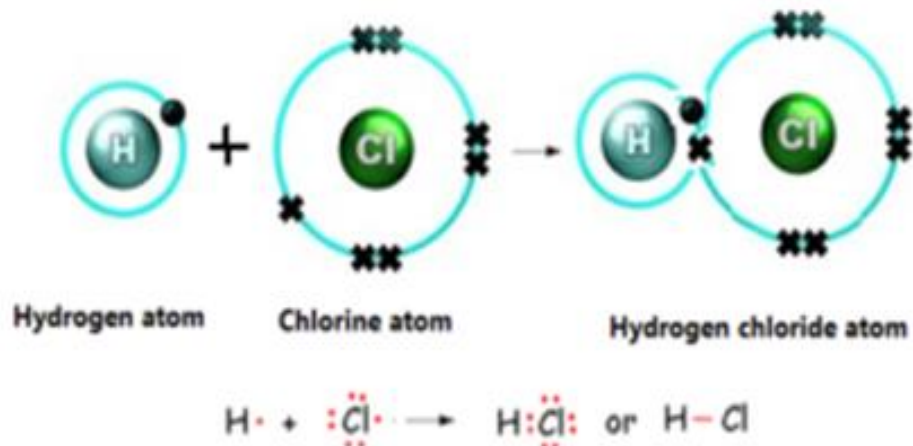
#### (i) Formation of chlorine molecule

The chlorine atom has electronic configuration,  $1s^2 2s^2 2p^6 3s^2 3p^5$ . The chlorine molecule is formed by each of the chlorine atom sharing one of its valence electrons with another atom. The two atoms approach each other closely for their atomic orbitals to overlap.



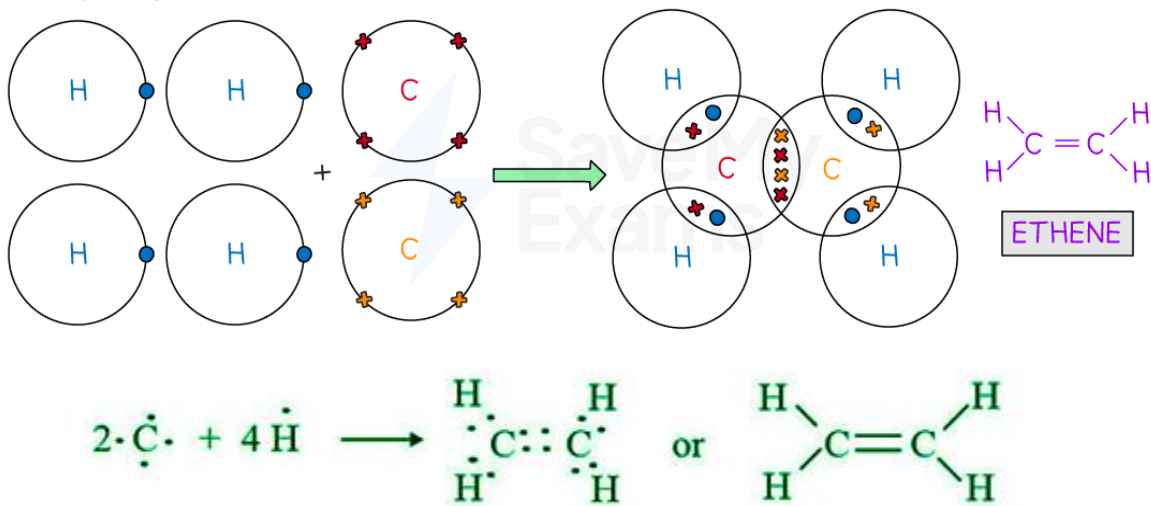
(ii) Formation of hydrogen chloride molecule

The hydrogen atom has electronic configuration  $1s^1$  and the chlorine atom has electronic configuration,  $1s^2 2s^2 2p^6 3s^2 3p^5$ . Hydrogen contributes its only electron and chlorine contributes one of its valence electrons. The electron pair is shared to make hydrogen have a full s sub energy level and chlorine to have a full octet.



(iii) Formation of ethene

In an ethene molecule, each carbon atom shares one pair of electrons with a hydrogen atom and two pairs of electrons with the other carbon atom. When atoms share two pairs of electrons, a double bond is formed between them. Each carbon atom forms single bonds to two hydrogen atoms and a double bond to the other carbon atom.

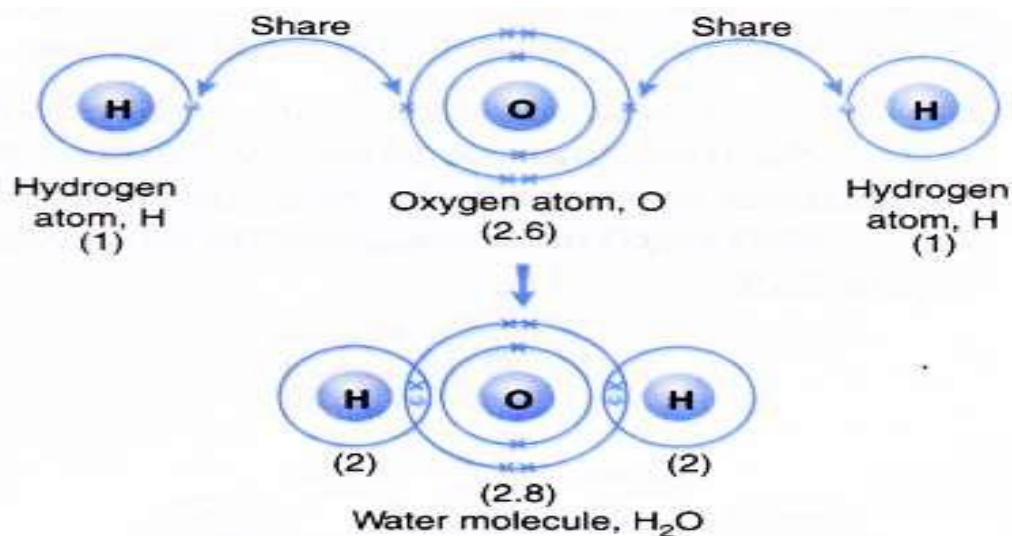


iv) formation of water.

An O atom has 6 valency electrons and needs 2 electrons, each, to form a noble gas configuration.

It bonds with 2 hydrogen atoms where each contributes one electron

Hence, each share the amount of electrons each is short of, in this case – 2 electrons, to form a stable molecule.



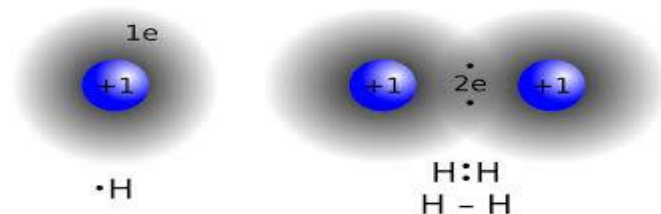
## Types of Covalent Bonds

### 1. Normal Covalent Bond:

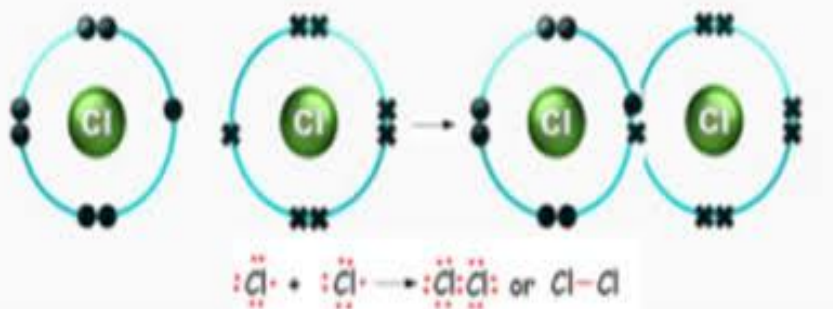
In a normal covalent bond, each bonded atom contributes one electron to form the shared pair. This typically occurs between atoms with similar or identical electronegativities.

Examples:

H-H in H<sub>2</sub> gas



Cl-Cl in Cl<sub>2</sub> molecule



### 2. Polar Covalent Bond:

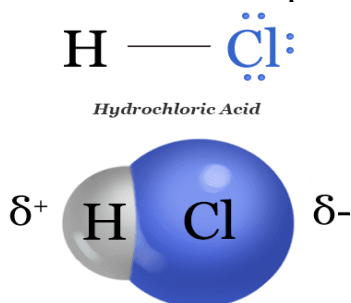
This type of covalent bond forms between two atoms with a significant difference in electronegativity. The shared electron pair is not equally shared; it is pulled closer towards the more electronegative atom. This unequal sharing creates a bond dipole, where the more electronegative atom gains a partial negative charge ( $\delta^-$ ) and the less electronegative atom gains a partial positive charge ( $\delta^+$ ).

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Examples:

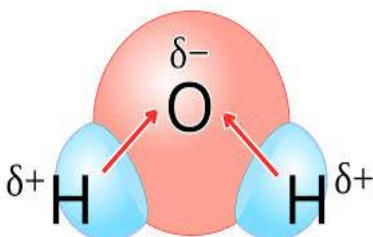
H-Cl in HCl

Cl is more electronegative than H, so the electrons are pulled towards Cl, creating  $H^{\delta+}-Cl^{\delta-}$ ,

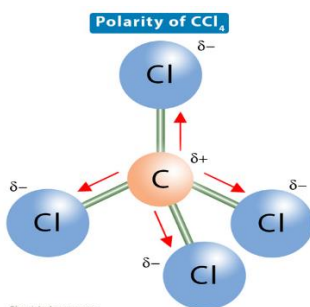


Other examples include:

H-O in  $H_2O$ ,  $H^{\delta+}-O^{\delta-}$



C-Cl in  $CCl_4$ .  $C^{\delta+}-Cl^{\delta-}$ .



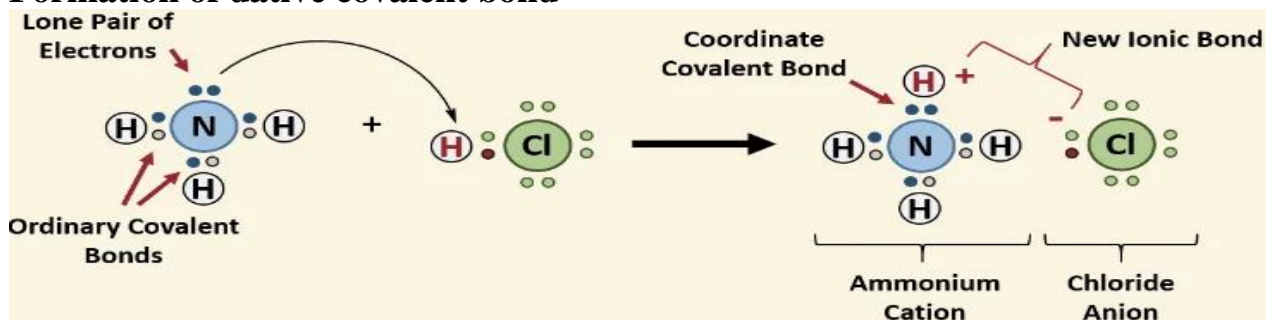
The magnitude of the bond dipole is related to the electronegativity difference and the distance between the nuclei (dipole moment).

### 3. Dative Covalent Bond (Coordinate Covalent Bond):

In a dative covalent bond, both electrons in the shared pair are contributed by only one of the bonded atoms. The atom contributing the electron pair is the donor atom, and it must have a lone pair of electrons available. The atom accepting the electron pair is the acceptor atom, and it must have a vacant orbital to accommodate the electron pair.

Dative bonds are often represented by an arrow ( $\rightarrow$ ) pointing from the donor atom to the acceptor atom. Once formed, a dative bond is indistinguishable from a normal covalent bond in terms of its properties.

## Formation of dative covalent bond

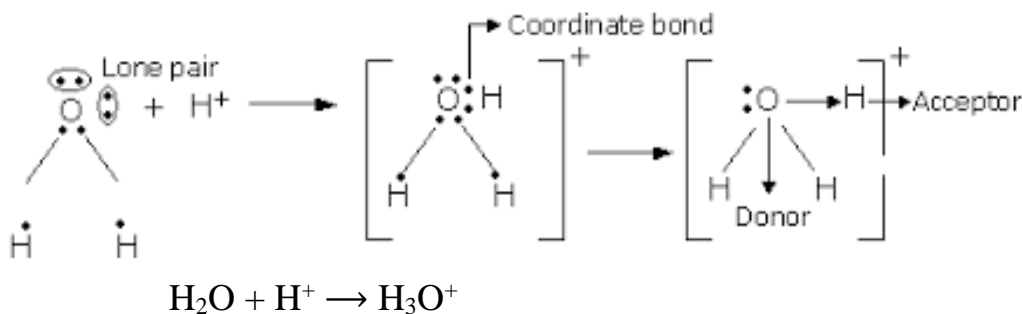


## Examples

### 1. Formation of Hydroxonium Ion ( $\text{H}_3\text{O}^+$ ):

A water molecule ( $\text{H}_2\text{O}$ ) has oxygen with two lone pairs. A hydrogen ion ( $\text{H}^+$ ) has no electrons (empty 1s orbital). The oxygen atom donates a lone pair to the  $\text{H}^+$  ion.

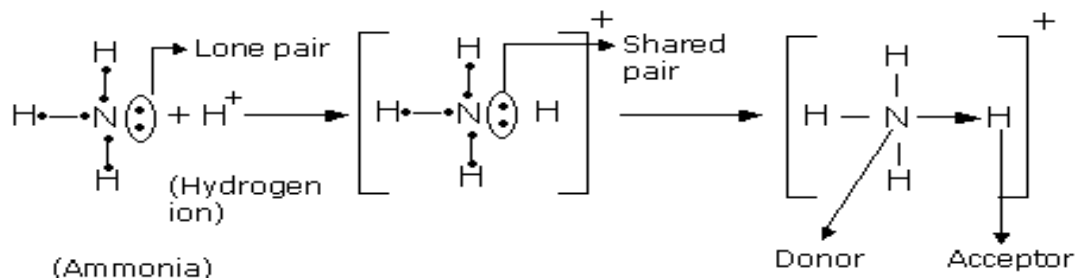
Illustration



### 2. Formation of Ammonium Ion ( $\text{NH}_4^+$ ):

An ammonia molecule ( $\text{NH}_3$ ) has nitrogen with one lone pair. A hydrogen ion ( $\text{H}^+$ ) has an empty 1s orbital. The nitrogen atom donates its lone pair to the  $\text{H}^+$  ion.

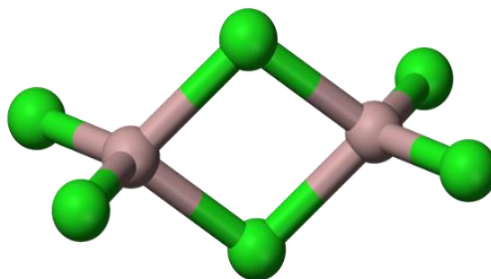
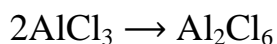
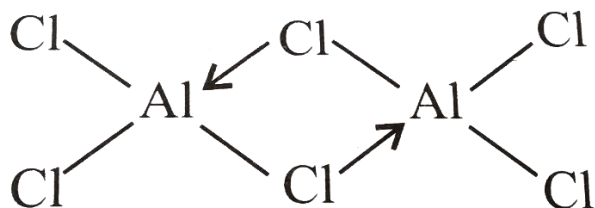
Illustration



### 3. Formation of Dimerised Aluminium Chloride ( $\text{Al}_2\text{Cl}_6$ ):

In  $\text{AlCl}_3$ , Aluminium has an incomplete octet (6 valence electrons). A chlorine atom from another  $\text{AlCl}_3$  molecule can donate a lone pair to the Aluminium atom with the vacant orbital.

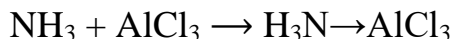
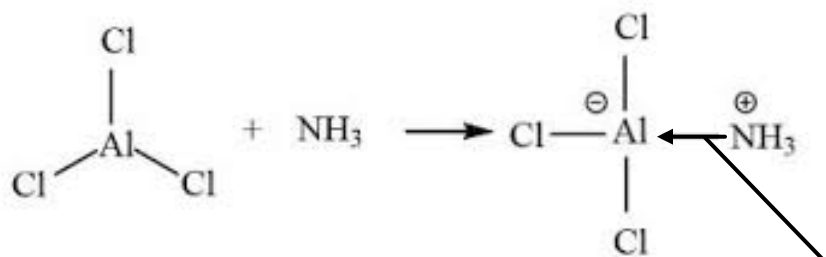
**Illustration**



**4. Reaction between Ammonia and Anhydrous Aluminium Chloride:**

Similar to the dimerisation, the nitrogen in  $\text{NH}_3$  donates a lone pair to the electron-deficient Aluminium in  $\text{AlCl}_3$ .

**Illustration**

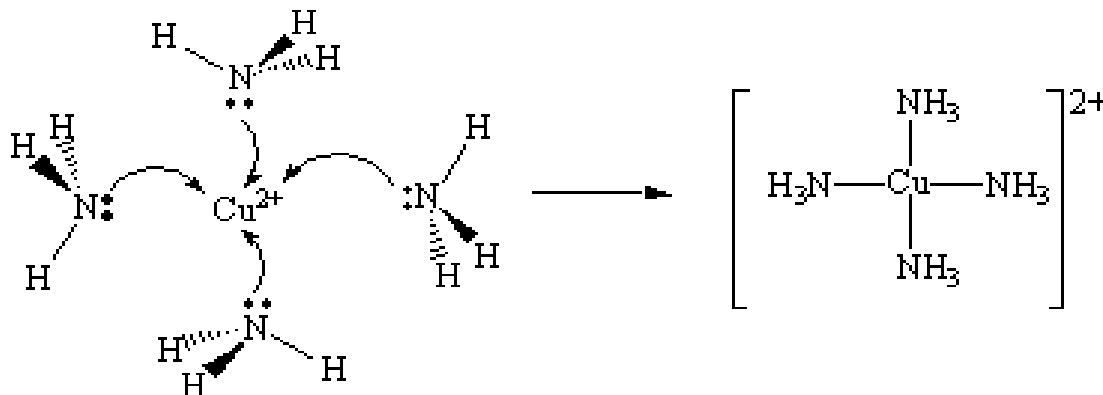


**5. Transition Metal Complexes:**

Ligands (molecules or ions with lone pairs) form dative bonds with central metal ions (which have vacant orbitals).

**Example:** Tetraamminecopper (II) ion,  $[\text{Cu}(\text{NH}_3)_4]^{2+}$ .

**Illustration**



## STRUCTURES OF COVALENT COMPOUNDS.

Covalent substances exhibit two main types of structures, leading to vastly different properties:

### 1. Simple Molecular Structures:

Consist of discrete molecules. The atoms within each molecule are held together by strong covalent bonds (intramolecular forces). However, the forces *between* these molecules (intermolecular forces) are relatively weak (van der Waals forces and hydrogen bonds). The intermolecular forces in molecular structures are of two types

- ✓ Van der waal forces
- ✓ Hydrogen bonds

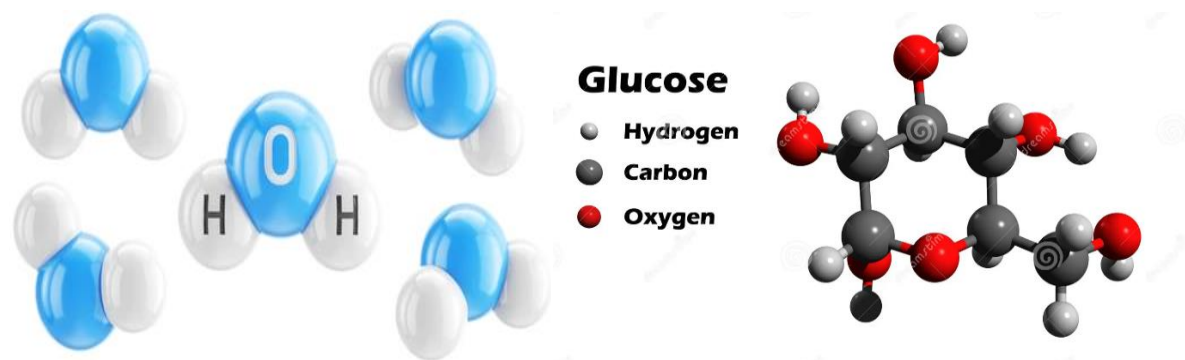
### The van der waals forces

Are the weakest form of intermolecular forces due to induce dipole-induce dipole attraction between molecules

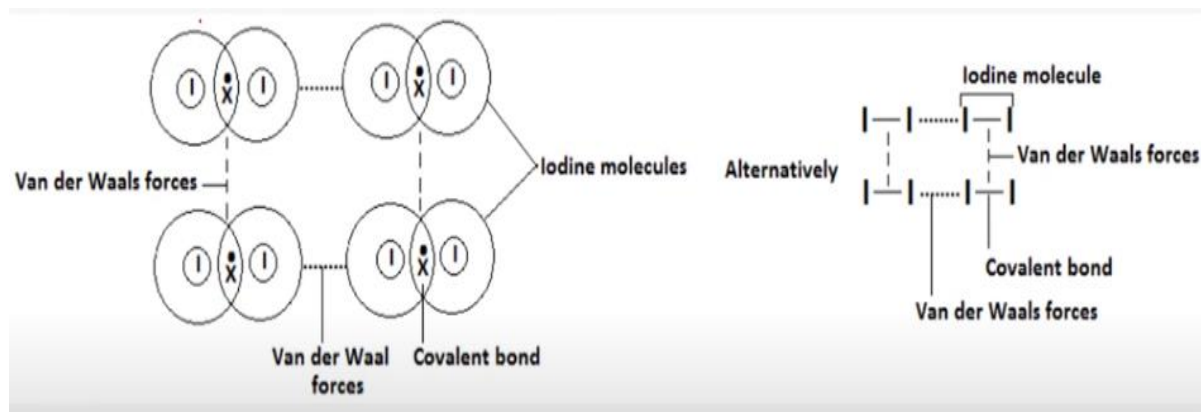
As the size of the molecule increases the number of constituent electron increases leading to increase in the strength of the induced dipole-induced dipole interactions

The strength of the van der waals thus increases as the molecular size increases

**Example water molecules, glucose etc.**



### Iodine molecule



## Hydrogen bonds

Is an intermolecular force in which the electropositive hydrogen atom of one molecule is attracted to an electronegative atom of another molecule

The essential requirements for the formation of a hydrogen bond are

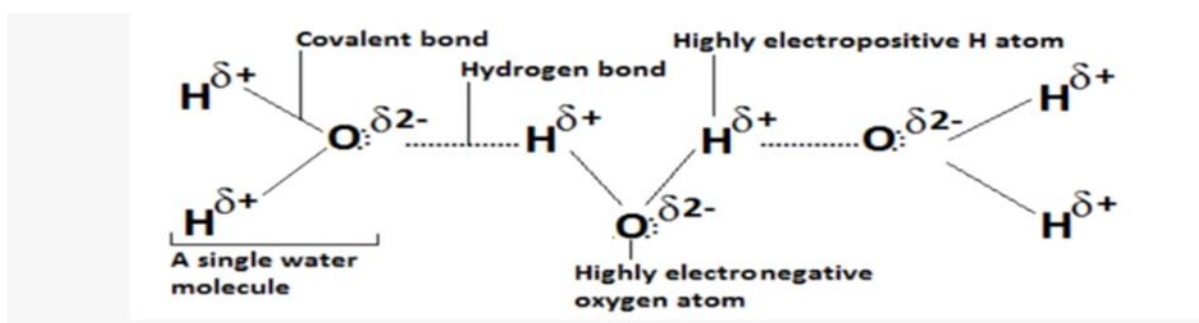
A hydrogen atom attached to a highly electronegative atom

An unshared pair of electrons on the electronegative atom

This explains why hydrogen bonds are common in molecules in which hydrogen are bonded to highly electronegative atom like nitrogen, oxygen and fluorine

## Example

### Formation of hydrogen bonds in water



Other compounds with hydrogen bond include ethanol, ammonia, and hydrogen fluoride. Hydrogen bonds are much stronger than weak van der Waals forces but still weaker than the covalent bonds.

Hydrogen bonding tends to disrupt the gradation in physical properties of molecular substances in relation to molecular weight.

The effects on molecular masses on the melting and boiling points only apply when the intermolecular forces are the same.

## Example

Both ethanol and dimethyl ether have the same relative molecular mass of 46, however the boiling point of ethanol is higher at  $78.5^\circ\text{C}$  than that of dimethyl ether at  $-24^\circ\text{C}$ .

- ✓ Both have molecular structure with covalent bonds between the atoms. However, the intermolecular forces in ethanol are hydrogen bonds which are much stronger and require more energy to break than the intermolecular forces in dimethyl ether which are weaker van der Waals forces.

Molecular structures are generally insoluble in polar solvents like water. However, those with hydrogen bonding as the intermolecular forces are soluble in water since hydrogen bonding confers them some polarity. Examples include sugar, ethanol, ethanoic acid etc.

## Types of hydrogen bonding

**(i) Intermolecular hydrogen bonding:** This is the type of hydrogen bond that occurs between two or more similar or different molecules. Intermolecular hydrogen bonds occur in the following examples;

### (a) Hydrogen fluoride, water and ammonia molecules

#### ✓ Hydrogen fluoride

If two polar hydrogen fluoride molecules are close enough, there is an attraction between the positive end of one molecule and the negative end of the other molecule. Since the attraction between the hydrogen atom in one molecule and the fluorine atom in the other molecule is stronger than the repulsion between the two hydrogen atoms and between the two fluorine atoms, the two molecules are bonded together by a hydrogen bond.

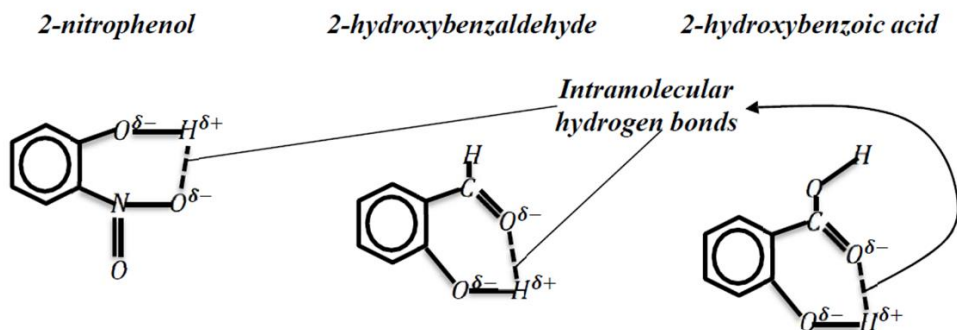
This explains why hydrogen fluoride is a liquid whereas other group VII hydrides are gases at room temperature; hydrogen fluoride has a higher boiling point than any other hydride of group VII.



**Intermolecular hydrogen bond**

### (ii) Intramolecular hydrogen bonding

This is the type of hydrogen bonding that occurs within the same molecule. An example is the intramolecular hydrogen bond formed in 2-nitrophenol, 2-chlorophenol, 2-hydroxybenzaldehyde and 2-hydroxybenzoic acid



## ACTIVITY

1. AIB Logistics Company Uganda is planning to import various hydrogen halides for industrial use. The company's chemical consultant has advised the importation team to transport hydrogen fluoride (HF) in a special liquid containment system, while hydrogen chloride (HCl), hydrogen bromide (HBr), and hydrogen iodide (HI) can safely be transported in gas cylinders. This recommendation has puzzled the importation team. They are wondering

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why there is a difference in the mode of transportation, yet all of these compounds are hydrides of halogens.

### Task:

As a Chemistry student and advisor, use your understanding of bonding and molecular structure to help the importation team understand why hydrogen fluoride (HF) requires different handling compared to the other hydrogen halides.

2. During a science exhibition in Kiryandongo District, a group of Senior Four students presented an experiment where they placed ice cubes in a glass of water. Visitors were surprised to observe that the ice cubes floated on water. One visitor asked, "Since ice is made of the same substance as water, why doesn't it sink?"

The students became curious and decided to consult you, a Chemistry student, to help them understand this phenomenon.

### Task:

Using your knowledge of bonding, molecular structure, and density, explain why ice floats on water, even though both are forms of the same substance.

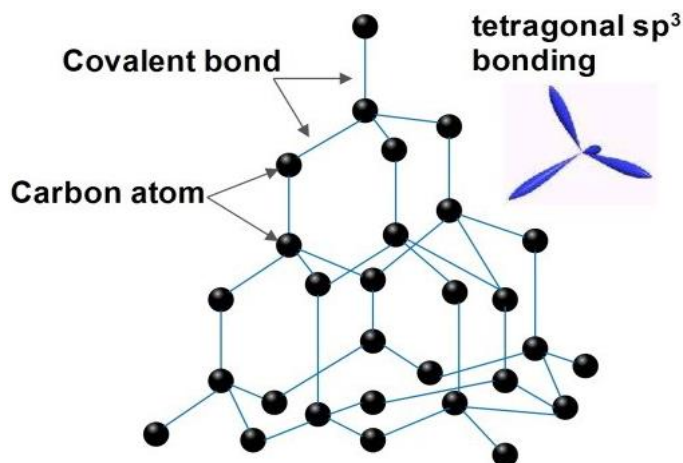
## 2. Giant Covalent Structures:

Consist of a vast, continuous three-dimensional network where atoms are linked by strong covalent bonds throughout the entire crystal. There are no discrete molecules. Examples include diamond, graphite, and silicon dioxide ( $\text{SiO}_2$ ).

### Diamond

- ✓ Is an allotropes of carbon
- ✓ Allotropes are different crystalline forms of the same element in the same physical states
- ✓ In diamond each carbon atom is bonded to four other carbon atoms by strong covalent bond
- ✓ The carbon atoms in diamond are covalently bonded into an octahedral pattern, which repeats itself in all directions resulting into a giant covalent structure
- ✓ Since each carbon atom is bonded to four others, all four valence electrons in each carbon are used in bonding hence no delocalised electrons in the structure of diamond.
- ✓ Diamonds is the hardest known substance due to the fact that all atoms are covalently bonded together and are closely packed together

### Structure of diamond



### Properties of diamond

#### Have high melting and boiling points

- ✓ It has a giant covalent structure with strong covalent bonds throughout the structure which requires large amounts of energy to break

#### It is insoluble in water

- ✓ It is non-polar and thus cannot dissolve in polar water molecules since there are no intermolecular interactions which would facilitate penetration into water molecules for dissolution to occur

#### Does not conduct heat and electricity

- ✓ Each carbon atom in the structure is bonded to four others hence uses all its valence electrons in bonding and thus lacks any delocalised electrons for heat and electrical conductivity

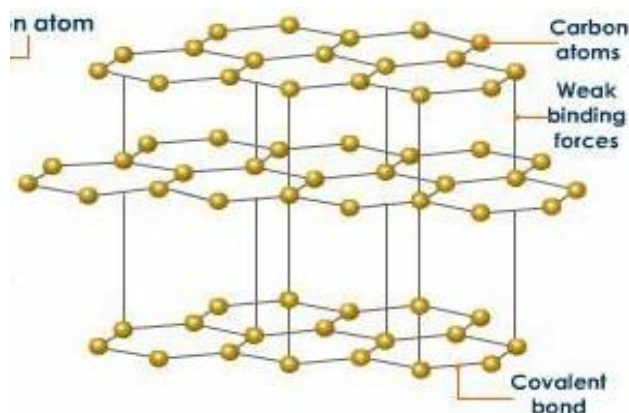
#### It is the hardest known substance

- ✓ All carbon atoms are compactly bonded in a continuous octahedral pattern with strong covalent bonds throughout the structure which is very difficult to break

### Graphite

- ✓ Is also an allotrope of carbon
- ✓ In graphite each carbon atom is bonded to three carbon atoms by strong bonds
- ✓ Since each carbon atom is bonded to only three others, only three of the four valence electrons in each carbon are used in bonding hence there is a presence of delocalised electrons in the structure of graphite
- ✓ The carbon atoms in graphite are covalently bonded into hexagonal layers which are joined to each other by weak van der Waals forces
- ✓ The presence of weak van der Waals forces explains the slippery nature of graphite

### Structure of graphite



## Properties

### Have high melting and boiling points

- ✓ It has a giant covalent structure with strong covalent bonds throughout the hexagonal layers which require large amounts of energy to break. Even though there are van der waal forces between the layers the effects of large number of covalent bonds still contribute to high melting and boiling points in graphite

### Insoluble in water

- ✓ It is non-polar and thus cannot dissolve in polar water molecules since there are no intermolecular interactions which would facilitate penetration into the water molecules for dissolution to occur

### It is a good conductor of heat and electricity

- ✓ Each carbon atom in the graphite is bonded to three others hence only three of its four valence electrons are in bonding. This led to the presence of delocalised electrons in the structure of graphite which conduct electricity

### It is soft and slippery

- ✓ The carbon atoms in graphite are covalently bonded into hexagonal layers which are joined to each other by weak van der waal forces
- ✓ The weak van der waal forces easily slide over each other when pressed hence the soft and slippery feel.

## General Properties of Simple Molecular Substances:

### 1. Low Melting and Boiling Points:

Due to the weak intermolecular forces between molecules. Only a small amount of energy is needed to overcome these forces and separate the molecules, allowing the substance to melt or boil.

### 2. Electrical Conductivity:

Generally, covalent compound does not conduct electricity in any state (solid, liquid, or gas). They lack free mobile ions or delocalized electrons to carry charge. Except if the substance reacts with water to form ions e.g., HCl gas in water forms  $H^+$  and  $Cl^-$  ions which have free mobile ions thus conduct electricity or in giant covalent structures like graphite.

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### 3. Solubility:

Most covalent substances are insoluble in polar solvents like water but soluble in non-polar solvents like benzene, ethanol or propanone. In polar solvents, the solvent-solvent interaction is much greater than interaction between the covalent molecules in the solute or the interaction between the covalent molecules and the polar molecules in the solution. Covalent compounds tend to dissolve in non-polar solvents because van der Waals' forces are the binding forces in both cases thus making them more soluble in non-polar solvent than in polar solvent

### 4. Physical State:

Most of covalent substances with simple molecular structures exist as gases or liquids at room temperature, because the weak intermolecular forces are easily overcome by thermal energy, keeping the molecules relatively far apart. Some with larger molecules or stronger intermolecular forces are soft solids e.g., sulfur, iodine, sugar.

### General Properties of Giant Covalent Structures:

#### 1. Very High Melting and Boiling Points:

Most covalent compounds with giant covalent structures have strong covalent bonds that extend throughout the entire compound. This requires a higher amount of energy, resulting in extremely high melting and boiling points.

#### 2. Hardness:

Generally, covalent compounds with giant covalent structures are very hard and rigid due to the strong, extensive network of covalent bonds e.g., diamond is one of the hardest known covalent substances.

#### 3. Electrical Conductivity:

Most are electrical insulators because all valence electrons are localized in strong covalent bonds and are not free to move e.g., diamond, SiO<sub>2</sub>. Except graphite where; each carbon atom forms covalent bonds to three others in layers, leaving one valence electron per atom delocalized within each layer, allowing it to conduct electricity along the layers.

### Activity

1. A Ugandan agrochemical company based in Lira District is developing a new insecticide spray to help cotton farmers fight bollworms. The active ingredient must dissolve well in methylbenzene, the solvent used in the spray to allow for even distribution on plant leaves. Two chemical options are tested: aluminium chloride and sodium chloride. After testing, only aluminium chloride dissolved in methylbenzene.

You are a senior secondary science student invited to join a regional science fair on sustainable farming. The agrochemical company presents this challenge to students to help them understand the science behind formulation.

Task:

- Explain why aluminium chloride dissolves in methylbenzene, while sodium chloride does not, using your knowledge of bonding and solubility.

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*Aluminium chloride is a covalent compound, especially in its anhydrous form. It consists of discrete  $Al_2Cl_6$  molecules with non-polar covalent character, allowing it to dissolve in non-polar organic solvents like methylbenzene.*

*On the other hand, sodium chloride is a strongly ionic compound, made up of  $Na^+$  and  $Cl^-$  ions held in a giant ionic lattice. It only dissolves well in polar solvents like water, but not in non-polar solvents like methylbenzene.*

*This difference is due to the solubility rule: "like dissolves like" – non-polar compounds dissolve in non-polar solvents, while ionic/polar compounds dissolve in polar solvents.*

**b) Advise the company on which types of compounds (ionic or covalent) are more suitable for formulation in organic solvents, and why.**

*The company should use covalent compounds or compounds with non-polar covalent character when formulating products in organic solvents such as methylbenzene. This is because covalent compounds are more likely to dissolve in non-polar organic solvents, ensuring that the active ingredient is evenly distributed when sprayed on crops.*

*Using ionic compounds like sodium chloride would be ineffective in such formulations, as they remain insoluble and may settle at the bottom, leading to uneven pesticide application.*

**c) Discuss how this knowledge can help improve agricultural practices in Uganda, especially in designing more effective pest control methods for rural farmers.**

*Understanding the relationship between bonding and solubility helps agrochemical companies develop effective pesticides and herbicides that are well absorbed and spread on plant surfaces.*

*For Ugandan farmers, especially in cotton and maize-growing areas, this ensures that pest control sprays are effective, reduce crop loss, and improve yields.*

*It also avoids wastage of resources—using inappropriate compounds that don't dissolve would lead to ineffective pest control, economic losses, and possibly harm to the environment if excessive amounts are applied without effect.*

*Knowledge of chemical compatibility leads to cost-effective, environmentally friendly, and science-based agriculture, supporting sustainable farming practices in rural communities.*

2. In Bweyale Town Council, a local youth group makes liquid soap using ingredients like ethanol, water, and organic oils. During a training, a chemistry teacher explained that ethanol and water are covalent compounds. One student asked why ethanol could dissolve in water while oil could not, even though both are covalent substances.

### Tasks

- Explain why ethanol can dissolve in water, yet oil cannot, even though both substances contain covalent bonds.
- Discuss how the structure and polarity of covalent compounds affect their solubility in water.

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- c) Suggest how understanding the properties of covalent compounds can help improve the quality of the soap produced by the group.

3. At Cornerstone Secondary School, the science department recently received new samples of Period 3 elements for use in practical chemistry lessons. The lab technician noted difficulties in storing phosphorus, sulfur, chlorine, and argon. Phosphorus and sulfur were stored as solids at room temperature, chlorine was kept in a special pressurized gas cylinder, while argon was stored in a sealed, cooled gas tank. The technician shared a table with the chemistry students showing the melting points of the elements:

Element	Phosphorus	Sulphur	Chlorine	Argon
Melting Point (°C)	44.2	115.2	-101.5	-189.4

### Tasks

- Using the data in the table, explain why phosphorus and sulfur are solids at room temperature, whereas chlorine and argon are gases.
- Relate the structure and type of bonding in each element to their melting points.
- Based on your explanation, advise the school technician on how best to store each of these elements safely in the school laboratory.

4. A group of Senior 5 learners at Nile High School were investigating materials used in everyday life. During a visit to a mechanic's garage, they observe a dark powder being used to reduce friction between moving metal parts in machines. The mechanic explains that this powder contains graphite, the same material found in pencil leads.

The students notice that both pencil lead and the powder feel smooth and slippery to touch, and they start wondering why graphite behaves this way, unlike other forms of carbon like diamond.

### Tasks:

- Using your knowledge of atomic structure and bonding, explain why graphite is soft and slippery.
- How does this property of graphite make it useful in?
  - Pencil making?
  - Reducing friction in machines (as a lubricant)?
- Suggest one other local application in Uganda where the slippery nature of graphite might be useful, and explain why.

### Applications of the concept of covalent bonding in our daily life:

Covalent bonding, where atoms share electrons, is fundamental to many substances we use daily. Water ( $\text{H}_2\text{O}$ ), the air we breathe ( $\text{O}_2$ ,  $\text{N}_2$ ), and even the food we eat (like sugars and proteins) are all examples of covalent compounds. These compounds are essential for life, fuel, and various household products.

1. Many of the most crucial molecules for life are held together by covalent bonds. Water ( $\text{H}_2\text{O}$ ), the solvent of life, is a prime example of a polar covalent molecule. The shared electrons are pulled more strongly by the oxygen atom, leading to a partial negative charge on oxygen and partial positive charges on hydrogen. This polarity allows water to dissolve many substances and participate in numerous biological reactions. Carbon dioxide ( $\text{CO}_2$ ), a nonpolar covalent molecule, is vital for photosynthesis, the process by which plants produce food. The oxygen we breathe ( $\text{O}_2$ ) and the nitrogen that makes up the majority of the atmosphere ( $\text{N}_2$ ) are also covalently bonded molecules.

2. The vast majority of organic compounds, which form the basis of living organisms and many synthetic materials, feature covalent bonds between carbon atoms and other elements like hydrogen, oxygen, nitrogen, and sulfur.

- ✓ *The carbohydrates, fats, and proteins that make up our food are all complex organic molecules held together by extensive covalent bonding.*
- ✓ *The diverse range of plastics and synthetic fibers we use daily, such as polyethylene, polypropylene, nylon, and polyester, are polymers. These large molecules are formed by the repetitive linking of smaller molecular units (monomers) through covalent bonds, resulting in materials with varied properties like strength, flexibility, and durability.*
- ✓ *Fossil fuels like natural gas (primarily methane,  $\text{CH}_4$ ), petroleum, and coal are composed of hydrocarbons, which are molecules containing carbon and hydrogen atoms linked by covalent bonds. The combustion of these covalent compounds releases energy that we harness for power and transportation.*
- ✓ *The cellulose and lignin that make up wood and paper are naturally occurring polymers formed through covalent bonding.*
- ✓ *Many everyday household products are covalent compounds.*
- ✓ *Components in soaps, detergents, and disinfectants often contain covalent molecules designed to interact with grease, dirt, and microbes. Ammonia ( $\text{NH}_3$ ) and bleach (sodium hypochlorite,  $\text{NaClO}$ , which has covalent character in the hypochlorite ion) are common examples.*
- ✓ *The active ingredients in most pharmaceuticals are covalent compounds meticulously designed to interact with specific biological targets in the body to treat illnesses and alleviate symptoms. Aspirin, antibiotics, and countless other drugs rely on the precise arrangement of atoms connected by covalent bonds.*

## Examples of covalent compounds in our daily life



### SHAPES AND STRUCTURES OF MOLECULES AND IONS

Sarah when setting a dining table for her family. She places chairs around the table in a way that everyone has enough space and no one's elbow bumps into the next person. Sometimes she space them evenly around a circular table while other time she arrange them differently around a rectangular one, depending on how many people are eating.

Atoms in a molecules and ions just like sarah's family around the table and their electrons which can be bonding and lone pairs just like the family members elbows.tend to **repel each other and arrange themselves as far apart as possible around the central atom** to reduce repulsion Just like sarah's family members prefer to spread out to minimize discomfort during dining.

This idea is the foundation of the **Valence Shell Electron Pair Repulsion (VSEPR) theory**, which helps us predict the **three-dimensional shapes** of molecules and ions.

### UNDERSTANDING VALENCE SHELL ELECTRON PAIR REPULSION THEORY (VSEPR)

The Valence Shell Electron Pair Repulsion (VSEPR) theory stands as a foundational model in chemistry, providing a framework for the prediction of the three-dimensional arrangement of atoms within molecules. The theory posits that the geometry of a molecule is primarily determined by the repulsion between pairs of valence electrons surrounding a central atom. These electron pairs, whether involved in bonding or existing as lone pairs, tend to arrange themselves in three-dimensional space to maximize the distance between them, thereby minimizing the repulsive forces and achieving a state of minimal energy and maximal stability for the molecule

The Valence Shell Electron Pair Repulsion (VSEPR) theory states that:

Electron pairs (both bonding and lone pairs) around a central atom in a molecule or ion arrange themselves as far apart as possible in three-dimensional space in order to minimize repulsion between them.

### Lewis Structures

But how do we know how many bonding and lone pairs are present around a central atom in the first place?

That's where **Lewis structures** come in. A Lewis structure is like a blueprint which help us visualize electron arrangements

A Lewis structure (also called a Lewis dot diagram) is a simplified two-dimensional representation of a molecule or ion that shows:

- The arrangement of atoms (how they are bonded together),
- The valence electrons (outermost electrons) of the atoms,
- Bonding pairs of electrons (shared between atoms as single, double, or triple bonds),
- Lone pairs (non-bonding electrons located on individual atoms).

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By drawing the Lewis structure of a molecule or ion, we can identify how many regions of electron density surround the central atom.

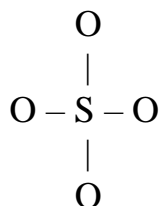
### STEPS FOR DRAWING LEWIS STRUCTURES

1. Write the skeleton structure of the molecule or ion. The central atom is usually the least electronegative element in the compound and then place a bonding pair of electrons between each pair of adjacent atoms or draw a line to indicate a single bond.

#### Example

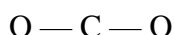
##### 1: Sulphate ion, $\text{SO}_4^{2-}$

Skeleton structure:



##### 2. carbon dioxide

Skeleton structure:



Carbon (C) and sulphur (S) is less electronegative than oxygen, so it goes in the center as central atoms.

Arrange the two oxygen atoms on either side of carbon and four on either side of sulphur.

Connect each O to C with a single bond initially.

Each single line represents a bonding pair (2 electrons).

2. Determine the total number of Valence electrons for the atoms in the molecule or ion. For a neutral molecule, this is the sum of valence electrons on each atom. For a negatively charged ion, add one electron for each negative charge. For a positively charged ion, subtract one electron for each positive charge.

#### Example $\text{CO}_2$

Carbon (C) = 4 valence electrons

Oxygen (O) = 6 valence electrons  $\times 2 = 12$

Molecule is neutral no added or removed electrons.

Total = 4 + 12 = 16 valence electrons

#### Example: $\text{SO}_4^{2-}$

S = 6 valence electrons

O = 6 valence electrons  $\times 4 = 24$

Charge =  $-2$  add 2 electrons

Total = 6 + 24 + 2 = 32 valence electrons

3. Subtract two valence electrons for each bond formed in the skeleton structure.

Each bond uses 2 electrons. Subtract these from the total to know how many electrons remain for lone pairs.

#### $\text{SO}_4^{2-}$

4 S—O bonds = 4  $\times$  2 = 8 electrons used

**Remaining = 32 – 8 = 24 electrons**

#### $\text{CO}_2$

2 single bonds in the skeleton (C—O and C—O)

Each bond = 2 electrons

2 bonds  $\times$  2 electrons = 4 electrons

**Remaining = 16 – 4 = 12 electrons**

4. Beginning with the terminal atoms, add enough electrons of the remaining electrons as valence electrons to satisfy the octets of the atoms (two for hydrogen). These electrons are the lone pairs.

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### Example CO<sub>2</sub>

- Start adding the remaining 12 electrons to the **two O atoms** to complete their octets.
- Each O already has 1 bond (2 electrons), so each need 6 more electrons (3 lone pairs):
- O (left): 3 lone pairs = 6 electrons
- O (right): 3 lone pairs = 6 electrons
- Total used = 6 + 6 = 12 electrons
- **Remaining = 12 – 12 = 0 electrons**

### Example SO<sub>4</sub><sup>2-</sup>

- Each O already has one bond (2 electrons), so each need **6 more electrons (3 lone pairs)**:
- 4 O atoms × 6 electrons = 24 electrons
- Remaining = 0**

5. The remaining electrons are then placed on the central atom (some central atoms can accommodate more than eight electrons)

- After satisfying terminal atoms, **place leftover electrons on the central atom.**
- Some elements in **Period 3 or higher (like P, S, Cl)** can expand their octet.

#### ➤ SO<sub>4</sub><sup>2-</sup>

- All electrons used in step 4.
- S already has 8 electrons (from 4 bonds), so no more electrons to place.

#### ➤ CO<sub>2</sub>

No electrons are left.

Carbon currently has 2 single bonds = 4 electrons only, not a full octet.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet. This will not change the number of electrons on the terminal atoms

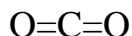
### Example

- Carbon has only 4 electrons (from 2 single bonds).
- It needs 8 → lacking 4 electrons.

→ Convert one lone pair from each oxygen into a bonding pair with carbon.

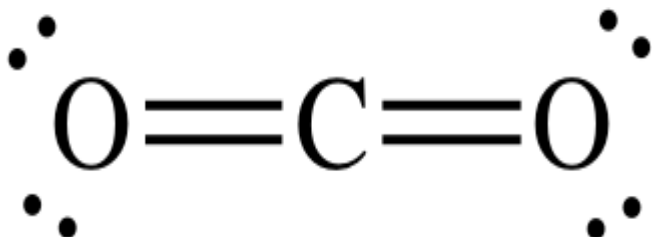
→ Form double bonds between C and both O atoms.

### New structure:

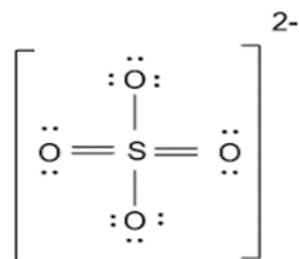


- Each O now has 2 lone pairs, and carbon has 4 shared pairs (8 electrons)

Final Lewis structure of carbon dioxide



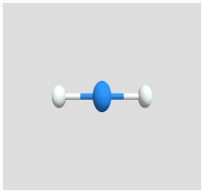
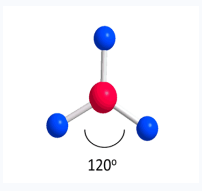
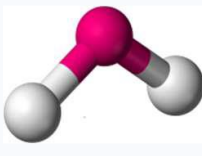
Final Lewis structure of sulphate ion




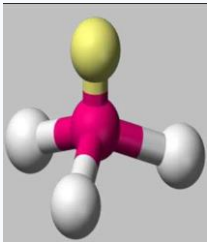
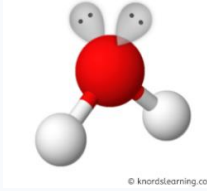
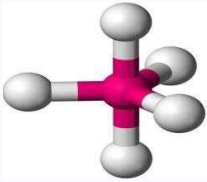

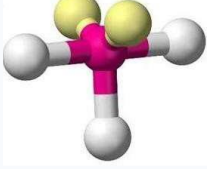
**PREDICTING MOLECULAR SHAPES BASED ON VSEPR THEORY**

A Lewis structure is like a blueprint—it shows how atoms are bonded together and the arrangement of valence electrons (both shared and unshared). By drawing the Lewis structure of a molecule or ion, we can identify how many regions of electron density surround the central atom.


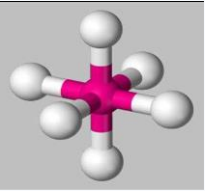
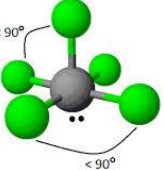
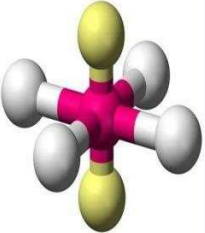
Once we know that, we apply **VSEPR theory** to determine how these regions arrange themselves in space, leading to molecular geometries like **linear, trigonal planar, tetrahedral, trigonal bipyramidal, or octahedral**—and even bent or trigonal pyramidal shapes when lone pairs are involved.

Total valence electron	Bonding Pairs	Lone Pairs	Shape of molecule	3D presentation	Ideal Bond Angle	Example
4	2	0	Linear		180°	BeCl <sub>2</sub> , CO <sub>2</sub> , HCN, CS <sub>2</sub> , Ag(NH <sub>3</sub> ) <sub>2</sub> <sup>+</sup> , Ag(CN) <sub>2</sub> <sup>-</sup> , [CuCl <sub>2</sub> ] <sup>-</sup>
6	3	0	Trigonal Planar		120°	BCl <sub>3</sub> , BF <sub>3</sub> , SO <sub>3</sub> , HCHO, NO <sub>3</sub> <sup>-</sup> , CO <sub>3</sub> <sup>2-</sup> , HCO <sub>3</sub> <sup>-</sup>
6	2	1	Bent or V-shaped		<120°	SnCl <sub>2</sub> , PbCl <sub>2</sub> , GeCl <sub>2</sub> , SO <sub>2</sub> , O <sub>3</sub> , NOCl, NO <sub>2</sub> <sup>-</sup>

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8	4	0	Tetrahedral		109.5°	CH <sub>4</sub> , CCl <sub>4</sub> , SiCl <sub>4</sub> , GeCl <sub>4</sub> , SnCl <sub>4</sub> , SiH <sub>4</sub> , GeH <sub>4</sub> , SnH <sub>4</sub> , CHCl <sub>3</sub> , H <sub>2</sub> SO <sub>4</sub> , H <sub>3</sub> PO <sub>4</sub> , HClO <sub>4</sub> , SO <sub>2</sub> Cl <sub>2</sub> , POCl <sub>3</sub> , NH <sub>4</sub> <sup>+</sup> , SO <sub>4</sub> <sup>2-</sup> , PO <sub>4</sub> <sup>3-</sup> , ClO <sub>4</sub> <sup>-</sup> , CrO <sub>4</sub> <sup>2-</sup> , MnO <sub>4</sub> <sup>2-</sup> , Zn(NH <sub>3</sub> ) <sub>4</sub> <sup>2+</sup>
8	3	1	Trigonal Pyramidal		<109.5°	NH <sub>3</sub> , PCl <sub>3</sub> , PF <sub>3</sub> , PH <sub>3</sub> , H <sub>2</sub> SO <sub>3</sub> , (CH <sub>3</sub> ) <sub>3</sub> N, H <sub>3</sub> PO <sub>3</sub> , HClO <sub>3</sub> , SO <sub>3</sub> <sup>2-</sup> , H <sub>3</sub> O <sup>+</sup> , ClO <sub>3</sub> <sup>-</sup>
8	2	2	Bent or V-shaped		<109.5°	H <sub>2</sub> O, H <sub>2</sub> S, Cl <sub>2</sub> O, ClO <sub>2</sub> <sup>-</sup> , OF <sub>2</sub> , SCl <sub>2</sub> , HOCl
10	5	0	Trigonal Bipyramidal		120°, 90°	PCl <sub>5</sub> , PF <sub>5</sub> , PBr <sub>5</sub> , PI <sub>5</sub> , Fe(CO) <sub>5</sub>
10	4	1	Irregular Tetrahedral		Angles vary	SF <sub>4</sub>
10	3	2	T-shaped		Approx. 90°	ClF <sub>3</sub> , ICl <sub>3</sub>

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10	2	3	Linear		180°	$I_3^-$ , $ICl_2^-$
12	6	0	Octahedral		90°	$SF_6$ , $SCl_6$ , $SBr_6$ , $Cr(NH_3)_6^{3+}$ , $Co(NH_3)_6^{3+}$ , $Fe(CN)_6^{4-}$ , $Al(H_2O)_6^{3+}$
12	5	1	Square Pyramidal		Approx. 90°	$IF_5$ , $BrF_5$
12	4	2	Square Planar		90°	$ICl_4^-$

Studying molecular and ionic shapes helps chemists understand, predict, and control the behavior and interactions of substances in various environments from biological systems like Enzymes and substrates "lock and key" mechanism, function of Biomolecules like proteins, DNA, and hormones, to industrial processes like in designing catalysts and optimizing reaction conditions. It provides a foundation for innovation in medicine, materials, and environmental science like the behavior of pollutants such as dioxins or CFCs.

### Activity

Draw the structure and name the shapes of the following molecules or ions

Molecules or ions	Structure	Shape
$H_2O$		
$PCl_3$		
$CCl_4$		
$NO_3^-$		
$CO_3^{2-}$		

## THE EFFECT OF LONE PAIRS ON MOLECULAR GEOMETRY

Lone pairs are pairs of valence electrons that are not shared between atom and do not participate in bonding, however they occupy space and strongly influence the shape of molecules

### Repulsion strength

- Lone pair – lone pair > lone pair – bond pair > bond pair – bond pair
- Lone pairs push bonding pairs closer together altering the bond angle and thus changing the expected geometry of molecules or ions
- The presence of lone pair electron decreases bond angle between bonded atoms due to increased repulsion this greatly affect the molecular and ionic shapes

### Example

1. Water molecules takes a bent shape due to the presence of 2 lone pair electrons and 2 bonding pairs reducing the bond angle between the hydrogen bonded to oxygen as the central atom in water from  $109.5^\circ$  to  $104.5^\circ$

The bent shape gives water its polarity making it an excellent solvent for ionic and polar substances

2. Ammonia molecules takes a trigonal pyramidal shape since the central atom nitrogen forms 3 bonding pairs with hydrogen atom leaving one lone pair.

The trigonal pyramidal shape allows ammonia to act as an electron donor in fertilizer and industrial reactions

3. Molecules with lone pairs tend to be polar and hydrophilic this determines their used in everyday life like making detergents a cleaning agents and also biological buffers among others

4. Transition metals often have ligands with lone pairs that donate electron to form coordinate bond essential in industrial catalysis example Haber process for preparation of ammonia.

### Activity

1. During a Chemistry practical session at Mvara Senior Secondary School, Mr. Owor instructed his Senior Six class to investigate the molecular structures of two bromide compounds—beryllium bromide ( $\text{BeBr}_2$ ) and tin (II) bromide ( $\text{SnBr}_2$ ).

Each group was given the molecular formulas and asked to model the shapes using ball-and-stick kits. One group noticed that  $\text{BeBr}_2$  appeared linear in their model and had bond angles around  $180^\circ$ , while another group, modeling  $\text{SnBr}_2$ , formed a bent structure and recorded bond angles less than  $120^\circ$  using a molecular geometry app.

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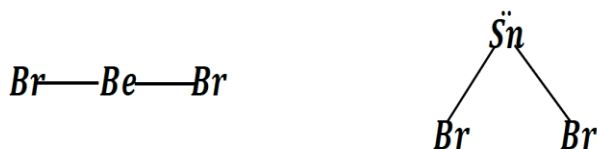
Confused by the differences, the students asked Mr. Owor why the bond angles and shapes varied so significantly even though both compounds had the same type of halogen atoms attached.

### Tasks

Using your knowledge of VSEPR theory, explain the actual molecular geometry and bond angles expected in the two-bromide compound.

### Responses

*Beryllium chloride molecule has two bonding pairs of electrons and no lone pair. The two bonding pairs repel each other greatly to form a linear molecule with a bond angle of 180°. Tin (II) bromide molecule has two bonding pairs of electrons and one lone pair. The two bonding pairs of electrons repel each other but the lone pair-bond pair repulsion is greater. This reduces the bond angle making the molecule V-shaped.*



2. At Muni University's outreach STEM fair, Senior Six students from Maracha Town Council were tasked with designing 3D molecular models of common polyatomic ions. Among the ions given were sulphate (SO<sub>4</sub><sup>2-</sup>) and Sulphite (SO<sub>3</sub><sup>2-</sup>).

Two groups presented different geometrical models:

- Group A modeled sulphate as a symmetrical structure with all oxygen atoms evenly spaced.
- Group B modeled Sulphite with one oxygen atom appearing slightly pushed downward, creating a structure that was not completely symmetrical.

When asked to explain their choices, Group B mentioned that their app indicated the presence of a lone pair on the central atom in Sulphite. Some students were puzzled as to why both ions contain a central sulphur atom, yet their shapes were so different.

### Tasks

Using VSEPR theory, explain why the two ions adopt different shapes yet they all have the same central atom.

### Responses

*The sulphate ion has four bonding pairs and no lone pairs. The four bonding pairs of electrons mutually repel each other, resulting into a bond angle of 109° making the molecule tetrahedral.*

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The Sulphite ion possesses three bonding pairs of electrons and one lone pair. The three bonding pairs of electrons repel each other but the lone pair-bond pair repulsion is greater. This reduces the bond angle making the molecule trigonal pyramidal.



3. At Cornerstone Chemistry Club, two students, Lydia and Brian, were discussing the shapes of different sulphur compounds. Lydia observed that both sulphur dioxide (SO<sub>2</sub>) and sulphur dichloride (SCl<sub>2</sub>) appear to have V-shaped (bent) molecular geometries. However, she was surprised to learn from their molecular modeling kit that SO<sub>2</sub> has a bond angle of about 120°, while SCl<sub>2</sub> has a smaller bond angle of about 109°.

Confused, Lydia asked Brian, “If both molecules are V-shaped, why is there such a difference in their bond angles?”

### Task

As a Chemistry student, help Lydia and Brian understand the reason for the difference in bond angles.

### Reponses

*The Sulphur dioxide molecule has two bonding pairs and one lone pair. The two bonding pairs of electrons repel each other but the lone pair-bond pair repulsion is greater. This makes the bond angle 120°.*

*The Sulphur dichloride molecule also possesses two bonding pairs of electrons but has two lone pairs. Since the Sulphur dichloride molecule has more lone pairs than the Sulphur dioxide molecule, there exists lone pair-lone pair repulsion, and the lone pair-bond pair repulsion is also greater than in Sulphur dioxide. This reduces the bond angle in Sulphur dichloride.*



## THE EFFECT OF ELECTRONEGATIVITY ON SIZE OF BOND ANGLES

Electronegativity is a fundamental chemical property that describes an atom's tendency or ability to attract shared electrons towards itself in a chemical bond.

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Electronegativity plays a significant role in determining bond angles, particularly in molecules with a central atom and surrounding atoms. The effect depends on whether the central atom or the surrounding atoms have higher electronegativity.

**Higher Electronegativity of the Central Atom = Larger Bond Angle:** When the central atom is more electronegative, it pulls the shared bonding electron pairs closer to itself. This increased electron density around the central atom leads to greater repulsion between the bonding pairs. To minimize this repulsion, the bond angle tends to increase, pushing the bonded atoms further apart.

- *Example: In  $\text{NH}_3$  (ammonia) vs.  $\text{PH}_3$  (phosphine), Nitrogen is more electronegative than Phosphorus. This results in  $\text{NH}_3$  having a bond angle of approximately  $107^\circ$ , which is larger than the bond angle of  $\text{PH}_3$  (approximately  $93.5^\circ$ )*

### Activity

During a Chemistry practical lesson at Cornerstone Secondary School, the teacher demonstrated the use of a molecular model kit to show the structures of ammonia ( $\text{NH}_3$ ) and phosphine ( $\text{PH}_3$ ). While both molecules appeared to have pyramidal shapes, the students were surprised to learn that ammonia has a bond angle of about  $107^\circ$ , while phosphine has a much smaller bond angle of about  $94^\circ$ .

John, a curious student, asked, "If both molecules are shaped the same way, why is the bond angle in phosphine so much smaller than in ammonia?"

### Task

As a Chemistry student, help John understand the reason behind this observation.

**Higher Electronegativity of the Substituent Atoms = Smaller Bond Angle:** When the substituent atoms are more electronegative, they pull the shared bonding electron pairs away from the central atom and towards themselves. This reduces the electron density around the central atom in the bonding region. With less electron density concentrated near the central atom, the repulsion between the bonding pairs decreases, allowing the bond angle to become smaller.

- *Example: Consider  $\text{NH}_3$  and  $\text{NF}_3$ . Fluorine is significantly more electronegative than Hydrogen. In  $\text{NF}_3$ , the highly electronegative fluorine atoms pull electron density away from the nitrogen, reducing the repulsion between the N-F bonds. This results in a smaller bond angle in  $\text{NF}_3$  (around  $102.5^\circ$ ) compared to  $\text{NH}_3$  (around  $107^\circ$ ).*

### Activity

During a chemistry class discussion, when asked by the chemistry teacher, Sarah showcased her research on the shapes of common phosphorus compounds. She compared phosphorus trifluoride ( $\text{PF}_3$ ) and phosphorus trichloride ( $\text{PCl}_3$ ), both of which have a pyramidal molecular shape.

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One student, Aviko, reported that  $\text{PF}_3$  has a bond angle of about  $96^\circ$ , while  $\text{PCl}_3$  has a slightly larger bond angle of about  $100^\circ$ . This puzzled her classmate Judges, who asked, "Aren't both compounds formed by phosphorus bonded to three halogen atoms? Why then is the bond angle in  $\text{PF}_3$  smaller than in  $\text{PCl}_3$ ?"

### Task

As a Chemistry student, analyze and explain the reason for the difference in their bond angles.

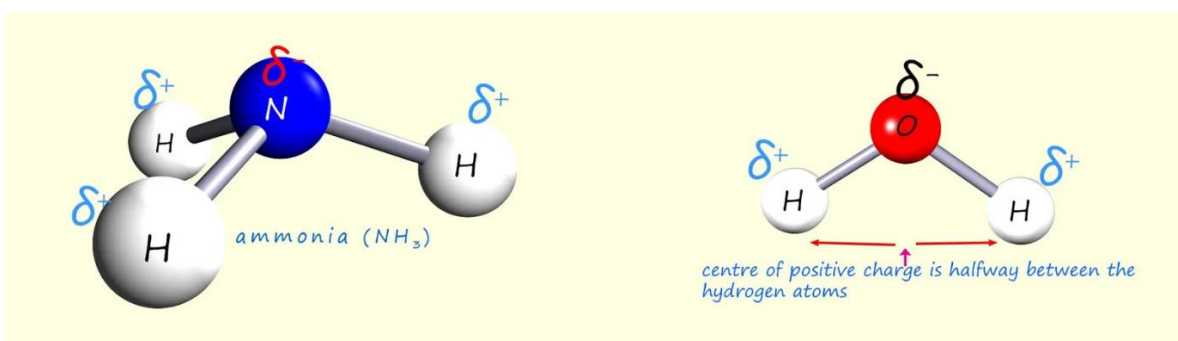
## POLARITY OF MOLECULES

Molecular polarity is a crucial concept in chemistry that describes the uneven distribution of electron density within a molecule, leading to the formation of partial positive and negative charges (poles). This uneven distribution results in a net dipole moment. Molecules with a net dipole moment are called polar molecules, while those with no net dipole moment are nonpolar molecules.

The polarity of a molecule is determined by two main factors:

### 1. Bond Polarity (Electronegativity Difference):

- **Electronegativity** is an atom's ability to attract electrons in a covalent bond.
- When two atoms with different electronegativities form a covalent bond, the electrons are pulled closer to the more electronegative atom. This creates a polar covalent bond, where one atom has a partial negative charge ( $\delta^-$ ) and the other has a partial positive charge ( $\delta^+$ ).
- The greater the electronegativity difference between the bonded atoms, the more polar the bond. E.g.,

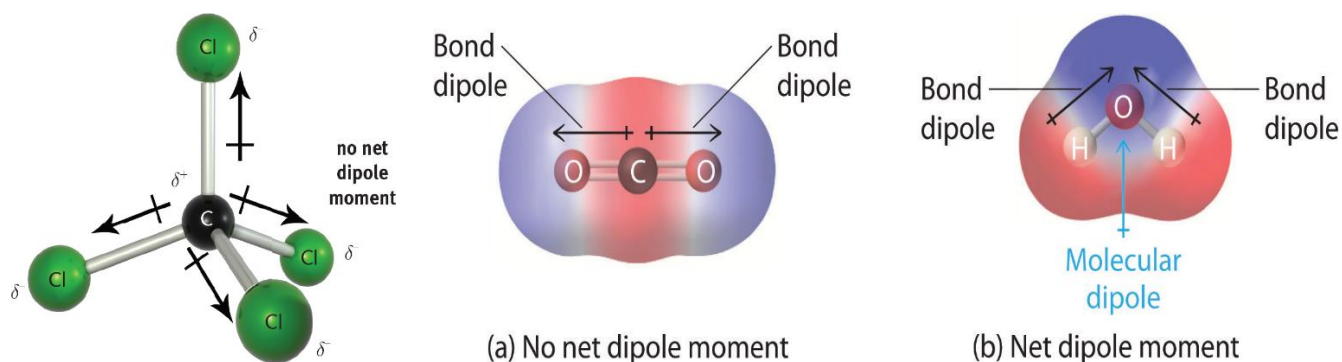


- **Nonpolar covalent bonds** occur when atoms have similar or identical electronegativities, leading to an equal sharing of electrons (e.g.,  $\text{O}_2$ ,  $\text{Cl}_2$ ).



## 2. Molecular Geometry (Shape of the Molecule):

- Even if a molecule contains polar bonds, it can still be nonpolar overall if its molecular geometry causes the individual bond dipoles to cancel each other out.
- Think of bond dipoles as **vectors**, having both magnitude (strength of polarity) and direction.
- If the vectors are arranged symmetrically and point in opposite directions, they can cancel each other out, resulting in a net dipole moment of zero.
- If the vectors do not cancel each other out due to an asymmetrical arrangement, the molecule will have a net dipole moment and be polar.



### Activity

1. During a Chemistry group discussion, students were task with analyzing different molecules to determine their polarity. One group looked at carbon tetrachloride ( $\text{CCl}_4$ ) and used an electronegativity table to confirm that the carbon-chlorine ( $\text{C}-\text{Cl}$ ) bond is polar. To test their understanding, the group performed a solubility experiment, where they attempted to mix  $\text{CCl}_4$  with water and with an organic solvent like hexane. They observed that  $\text{CCl}_4$  did not mix with water (a polar solvent) but dissolved well in hexane (a non-polar solvent).

Confused, one of the students, Musa, asked, “How can carbon tetrachloride be non-polar if all the bonds are polar?”

### Task

As a Chemistry student, help Musa and his team understand this apparent contradiction.

2. At a beverage production plant in Arua, Senior 5 Chemistry students on an industrial tour observed the use of carbon dioxide in the carbonation of soft drinks. The plant engineer explained that  $\text{CO}_2$  is preferred because it dissolves easily in non-polar flavoring agents used in the drinks and can also diffuse rapidly through airtight plastic bottles made of non-polar polymers.

This explanation puzzled **Viola**, one of the students, who reflected on their recent Chemistry lessons and remarked:

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“But the carbon–oxygen bonds in CO<sub>2</sub> are polar. Yet the engineer said the whole molecule behaves like a non-polar substance in so many of its applications. How is that possible?”

### Task

As a Chemistry student, help Viola understand engineer’s explanation

### IMPORTANCE OF POLARITY IN EVERYDAY USAGE OF MOLECULES AND IONS

- *Water, the universal solvent is a highly polar molecule due with its bent shape and the significant electronegativity difference between oxygen and hydrogen. Its polarity makes it an excellent solvent for many polar and ionic substances. This is why Salt dissolves in water vital in cooking, Sugar in making tea*
- *Soaps and detergents work because they have both polar and nonpolar parts. The nonpolar part can interact with grease and oil (nonpolar dirt), while the polar part interacts with water, allowing the dirt to be washed away.*
- *Polarity directly influences the types and strengths of intermolecular forces between molecules, which in turn dictate a substance's physical properties like boiling and melting points. This is why water polar molecule boils at 100°C while methane, a nonpolar molecule of similar size, boils at –161.5°C.*
- *Polarity is absolutely fundamental to life processes like cell membranes are primarily composed of phospholipids, which are amphipathic molecules with both a polar head and a nonpolar tail. This dual nature causes phospholipids to spontaneously form a bilayer in water, with the polar heads facing the aqueous environment and the nonpolar tails tucked inside, forming the barrier that defines a cell.*
- *The sequence of amino acids in a protein, and their individual polar or nonpolar characteristics, dictates how the protein folds into its unique three-dimensional shape. This shape is crucial for its function e.g., as an enzyme, a transport molecule, or a structural component.*
- *The polarity of a drug molecule significantly affects how it interacts with biological targets and how it's absorbed, distributed, metabolized, and excreted by the body.*

### Activity of integration