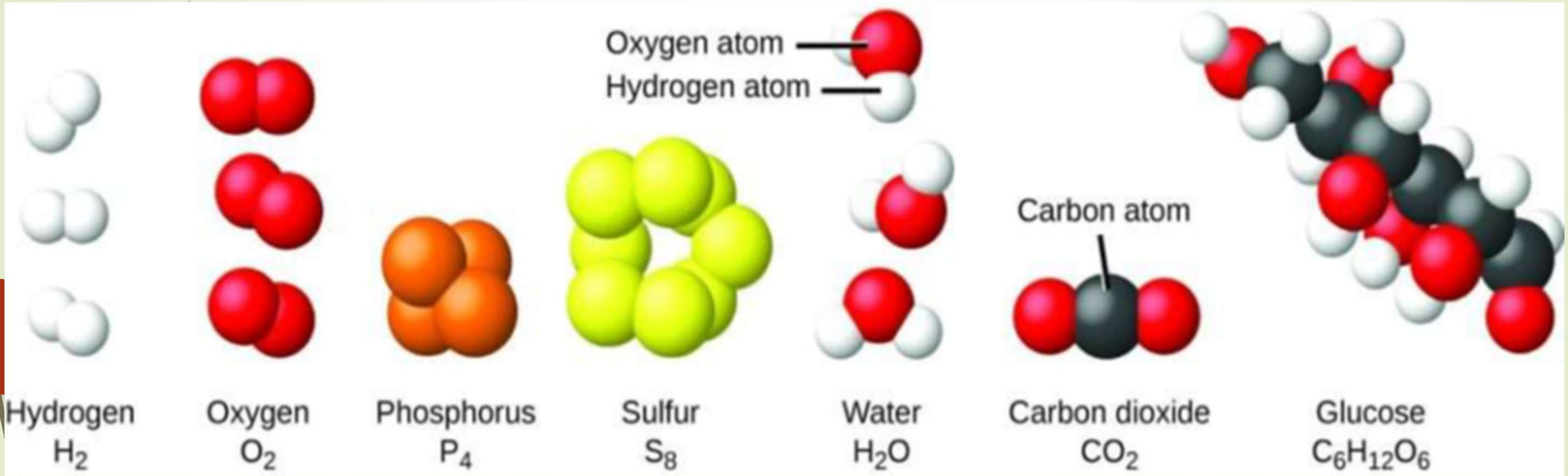
CHEMICAL BONDING AND STRUCTURES



TR SAMUEL

COMPETENCY: **The learner analyses the types of chemical bonds and molecular structures, and relates them to the properties and uses of substances in real-life contexts.**

SUB-TOPIC 3.1: Formation of Ionic and Metallic Bonds

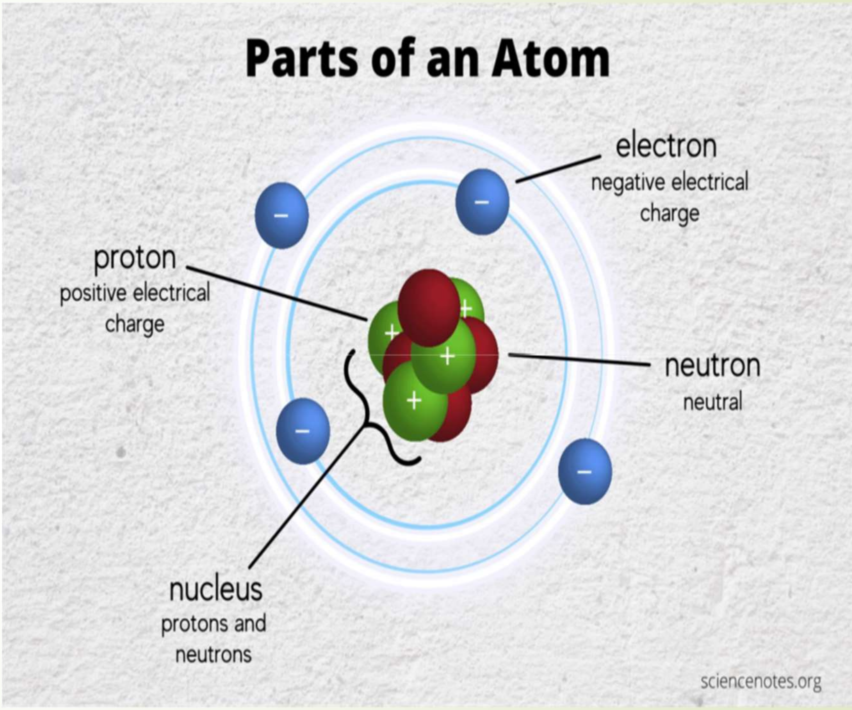
- **Learning Outcomes**

The learner should be able to:

- ✓ **a) justify the formation of ionic and metallic bonds based on electron transfer and electrostatic forces. (u, s)**
- ✓ **b) evaluate the properties of ionic and metallic compounds in relation to bond strength and structural composition. (u)**

INTRODUCTION

- Atoms are made up of energy levels and nucleus
- The energy levels contain electrons while the nucleus contains proton and neutrons
- The electrons are negatively charged, protons are positively charged while neutrons are electrically neutral
- To attain the stable octet noble gas configuration such atom, lose, gain or share valence electrons
- It is the act of losing, gaining or sharing valence electron that leads to chemical bonds
- When atom gain or lose valence electron(s) they become charged anions and cations respectively
- Particles of the same charge repel each other while particles of different charges attract one another

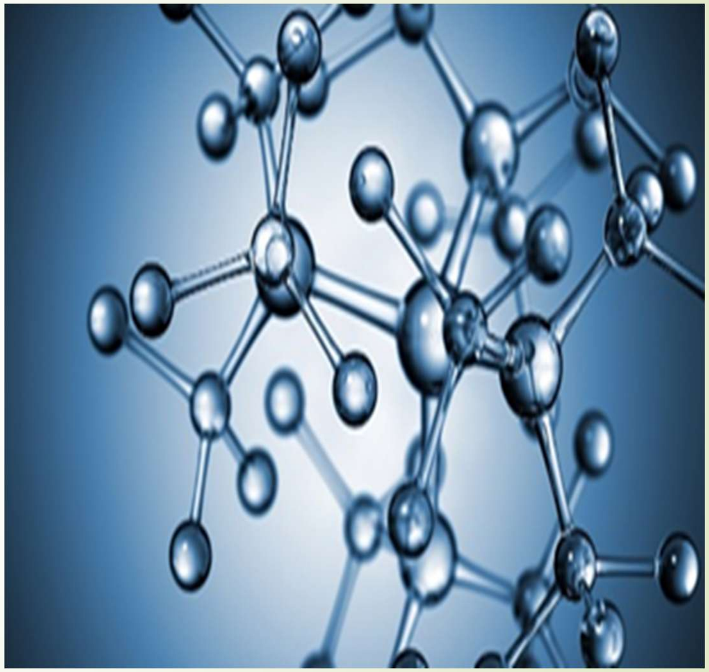


THE CONCEPT OF BONDING AND STRUCTURES

- 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.

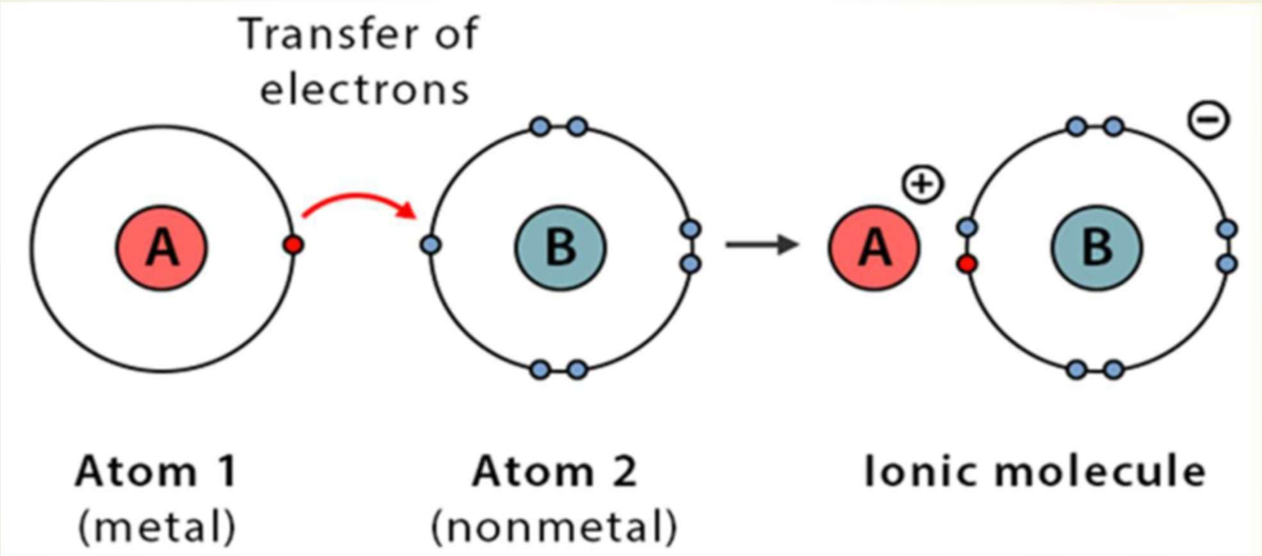
Chemical bonding is the fundamental force that holds atoms, ions, or molecules together to form compounds. It's the glue that dictates how elements combine and interact, ultimately explaining properties and chemical reactions.

- **Bond** is the mutual force of attraction that holds particles together when atoms (similar or different) combine during chemical reactions
- **Structure** refers to a regular pattern of particles in substance held together by a chemical bond



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.



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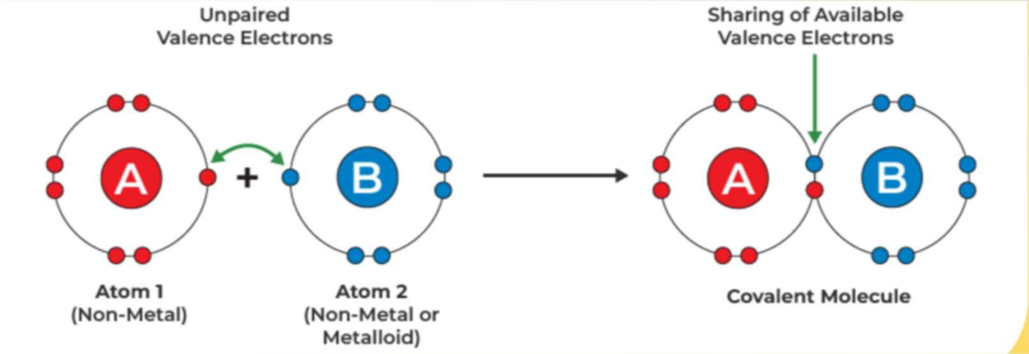
2. Electron Sharing:

- 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.

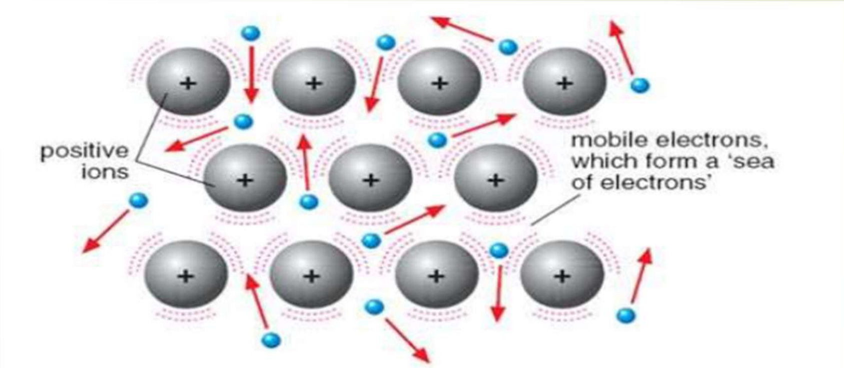


Figure 6.27 Bonding in metals

TYPES OF BONDS

Basically, there are three main types of chemical bonds

- Ionic,
- covalent and,
- metallic bonds

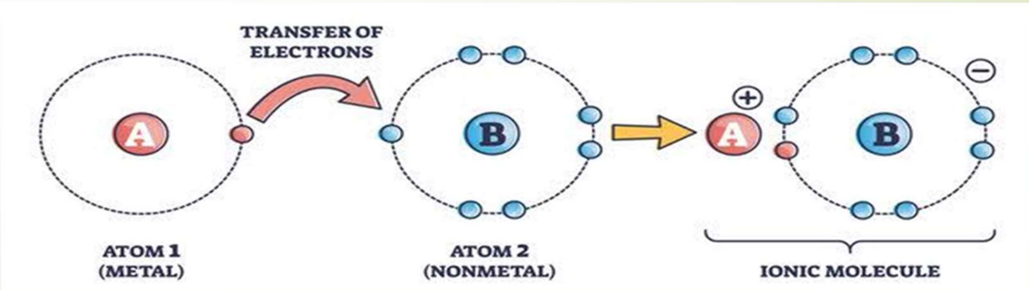
IONIC BONDS/ELECTROVALENT BONDS

Definition:

- Is a bond formed due to complete transfer of electrons from one atom to another resulting into two oppositely charged ions

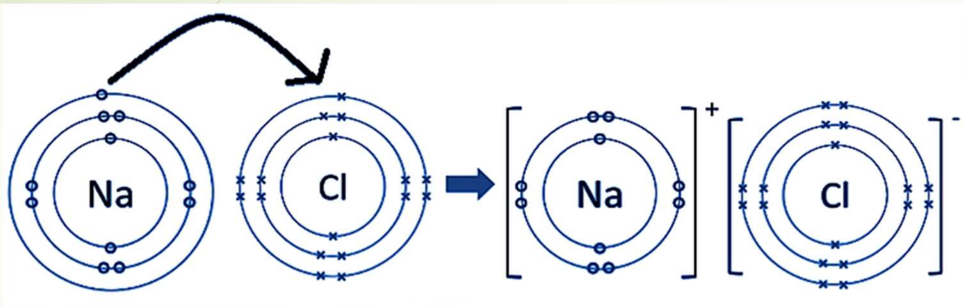
Formation of ionic bond

- Formed due to complete transfer of electron from one atom to another, and mainly formed between a metal and a non-metal.
- This occurs in a bid for both atoms to acquire a stable noble gas configuration
- One atom loses all its valence electrons, thus forming cation (positively charged ions)
- The other atom gains all the lost valence electrons forming an anion (negatively charged ions)
- The cation and the anion are oppositely charged and thus develop a mutual force of attraction between them which is the ionic/electrovalent bond



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.
Na readily loses its $3s^1$ electron to achieve a stable $1s^2 2s^2 2p^6$ configuration (like Neon), forming a Na^+ cation:

$$Na \longrightarrow Na^+ + e^-$$

Cl readily gains this electron to achieve a stable $1s^2 2s^2 2p^6 3s^2 3p^6$ configuration (like Argon), forming a Cl^- anion:

$$Cl + e^- \longrightarrow Cl^-$$

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

$$Na^+ + Cl^- \longrightarrow NaCl$$

2. Formation of Magnesium Oxide (MgO):

- Magnesium (Mg), a Group 2 metal, has electron configuration $1s^2 2s^2 2p^6 3s^2$. It has 2 valence electrons.

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

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

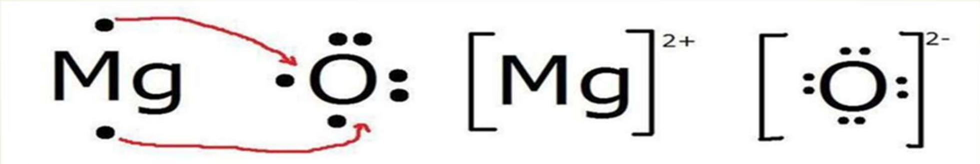
$$Mg \longrightarrow Mg^{2+} + 2e^-$$

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

$$O + 2e^- \longrightarrow O^{2-}$$

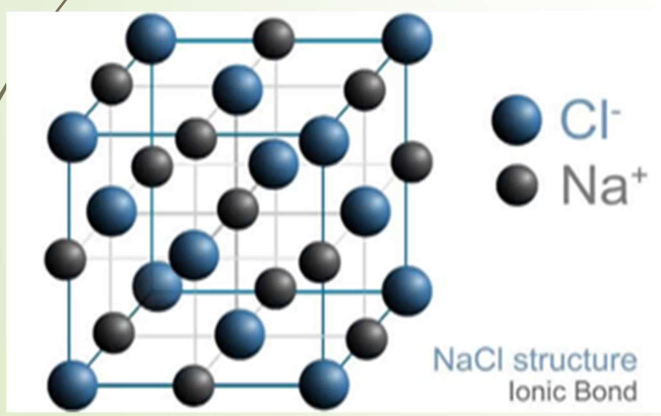
The Mg^{2+} and O^{2-} ions are held by strong electrostatic attraction:

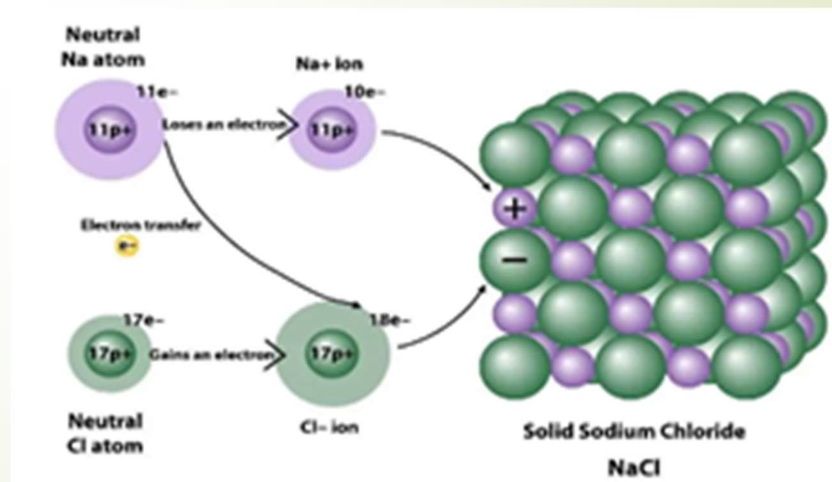
$$Mg^{2+} + O^{2-} \longrightarrow MgO$$



Structure of ionic compound

- Ionic bonding results into one type of structure, the **giant ionic structure**
- This is a type of structure in which all ions are bonded with strong ionic bonds throughout the structure
- Each ion in the giant ionic structure is surrounded by several others resulting into giant pattern of several ions hence giant ionic structure
- Most ionic substances with giant ionic structure are crystalline in nature, made up of crystals
- A crystal is a solid form of a substance in which the particles are arranged in a definite pattern repeated regularly in 3 dimension





PROPERTIES OF IONIC COMPOUND

1. They are hard and brittle

- Ionic solids are hard because each ion is held in the crystal by strong attractions from the oppositely charged ions around it
- They are brittle and thus may be split cleanly (cleaved) using sharp edged razor
- When a crystal is tapped sharply along a particular plane it is possible to displace one layer of ions relative to the next
- Due to the displacement, ions of similar charge come together leading to repulsive forces between the portion of the crystals
- This forces the two portions of the crystals to split apart

2. High melting and boiling points

- They have strong electrostatic forces between the oppositely charged ions throughout the structure which requires large amount of energy to break

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3. Solubility

- They are soluble in polar solvent like water, ethanol, and propanone
- Water contains highly polar molecules
- The positive ends of the polar water molecules are attracted to the negative ions in the crystal and the negative ends of the water molecules are attracted to the positive ions in the crystal
- This results to the formation of ion-solvent bond which leads to release of energy
- This energy is sufficient to cause the detachment of ions from crystal lattice hence dissolution
- This detachment of ions is called solvation, and the energy required for this is called solvation energy
- Where the solvent is water, the ions are said to have been hydrated and the energy involved in this process is called hydration energy
- They are insoluble in non-polar solvent like tetrachloromethane, benzene and hexane
- Non polar molecules are held together by weak intermolecular forces, the van der Waals forces
- The van der Waals are much smaller in magnitude compared to ionic bond in ionic solid crystal lattice
- Thus, the ion-ion interaction in the ionic solid are stronger than the solvent-solvent interactions in the solvent or the solvent ion interaction between the solid and the solvent
- Thus, the non-polar solvent molecules cannot penetrate the ionic lattice to cause solvation

4. Electrical conductivity

- Ionic substances do not conduct electric current in solid state
- The ions are held in static positions in the solid crystal lattice and thus cannot move to conduct electric current
- They conduct electricity in molten and solution (aqueous) state
- In molten and aqueous states, the ions are free and mobile and thus move about conducting electric current

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.

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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°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 (°C)	801	712	180

Task:

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

b) 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.

c) 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.

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:

a) 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).

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

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

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 (Na^+) and chloride anions (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 ($CaCO_3$)**, 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.*

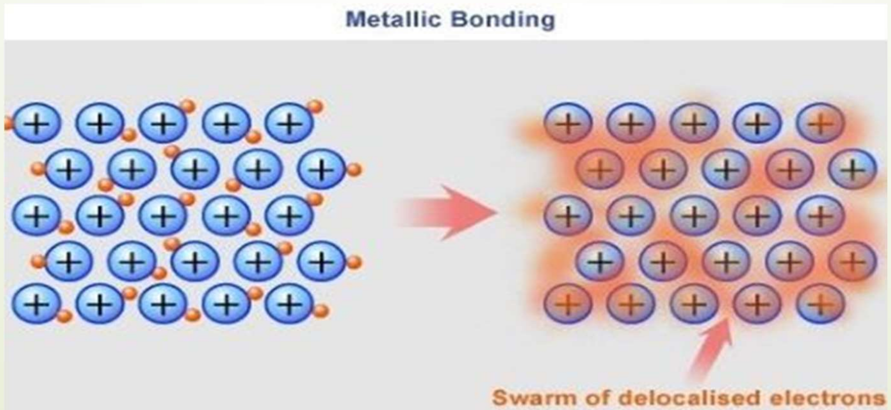
Examples of ionic substances used in our daily life

Metallic bonding

- Is a bond formed due to electrostatic attraction between the positively charged nuclei and the negatively charged delocalised electrons that hold atom together

Formation of metallic bonds

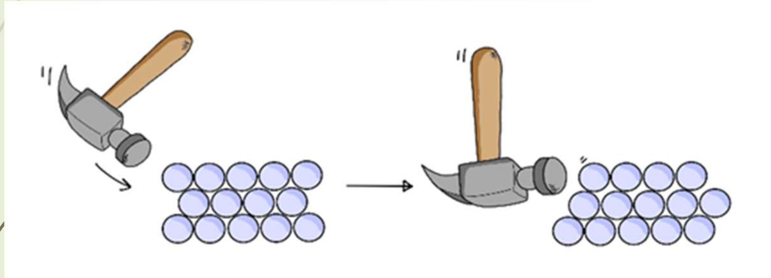
- In a metal there are usually many atoms surrounding any one atom
- The valence electron of any one atom are therefore mutually attracted to many nuclei
- This leads to a situation in which the positive nuclei appear to be immersed in a sea of mobile electrons
- The sea of mobile electrons are said to be delocalised which explain s the ability of the substance s with metallic bonds to conduct electric current
- This pattern of positive nuclei in sea of electrons is repeated many times throughout the structure leading to a giant metallic structure, the only structure due to metallic bonds



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. 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.

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).

Examples of common metals

Activity

- 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:**

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

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

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

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:

a) Explain why copper is a good conductor of electricity, using your knowledge of metallic bonding.
b) Why would plastic strips not work as substitutes for metal in wiring?
c) 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.

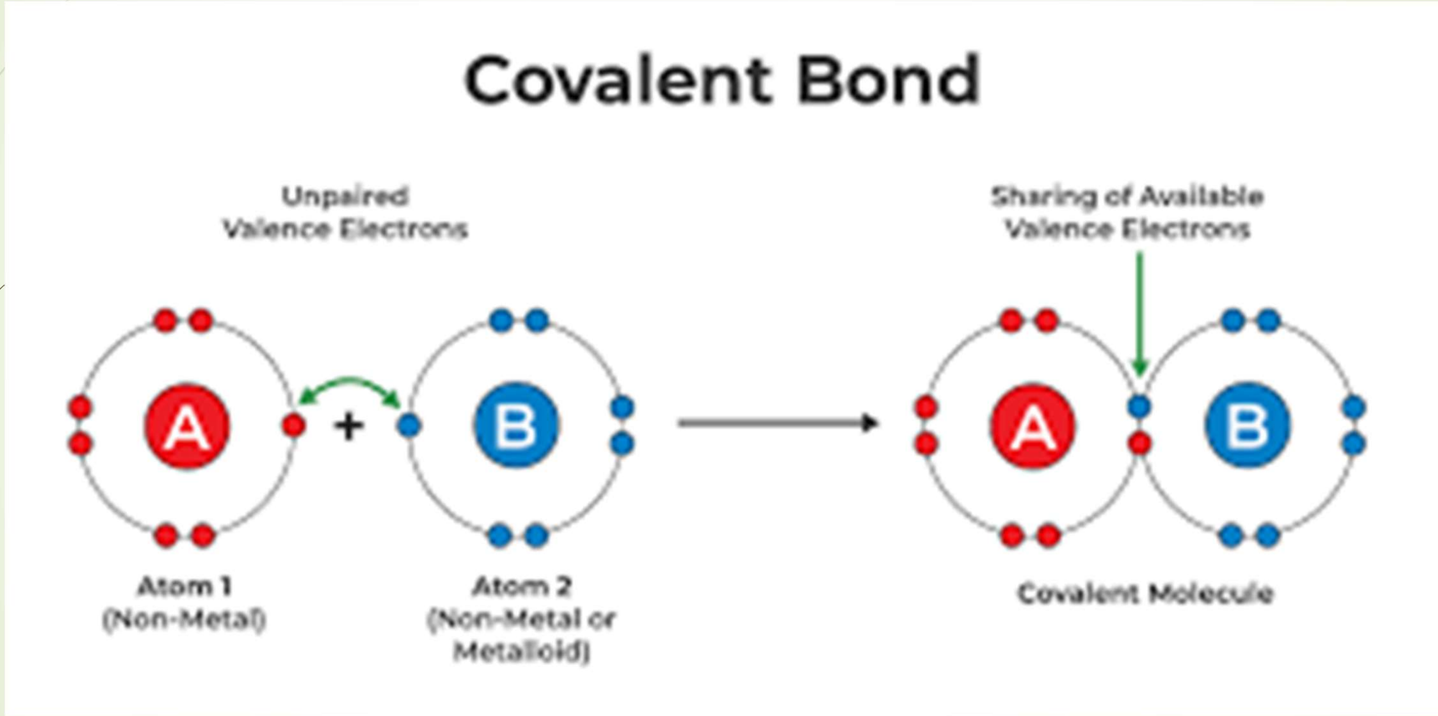
Tasks:

a) Explain how metallic bonding allows metals to stay strong at high temperatures during forging.
b) Why do the tools not shatter when hammered, unlike glass?
c) How can this knowledge of metallic bonding help young innovators in Uganda's metal fabrication industry?

COVALENT BONDING

- Refers to a bond formed when two atoms of the same or different elements share electrons to become stable
- Formation of covalent bond between atom (similar or different) results to the formation of a molecules
- A molecule is a group of atoms (two or more) of the same or different elements that are held by covalent bonds
- 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).

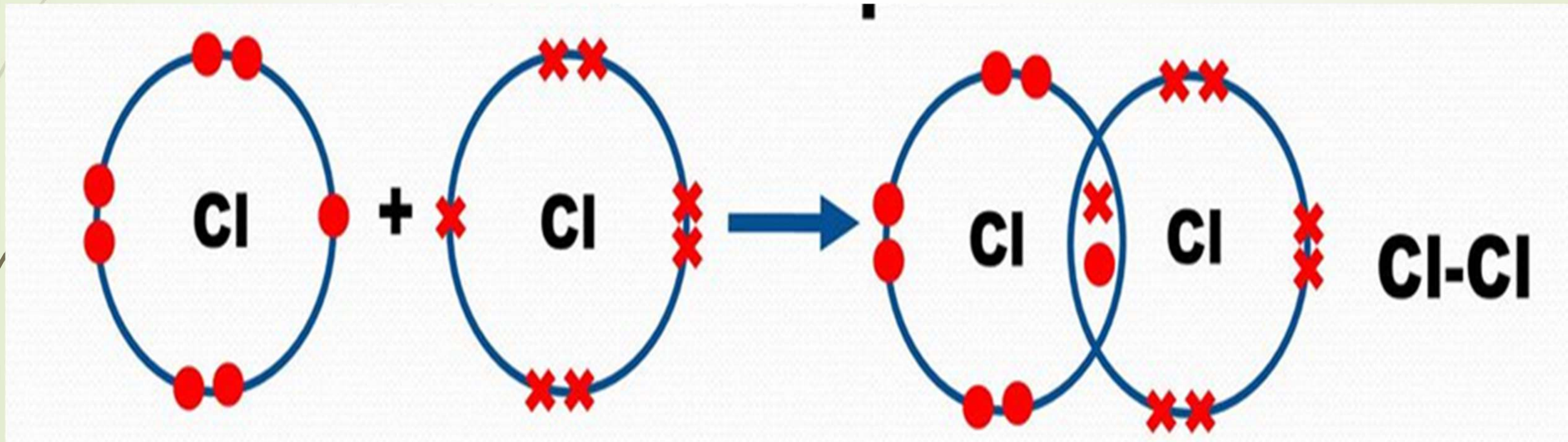
Illustration:



Examples

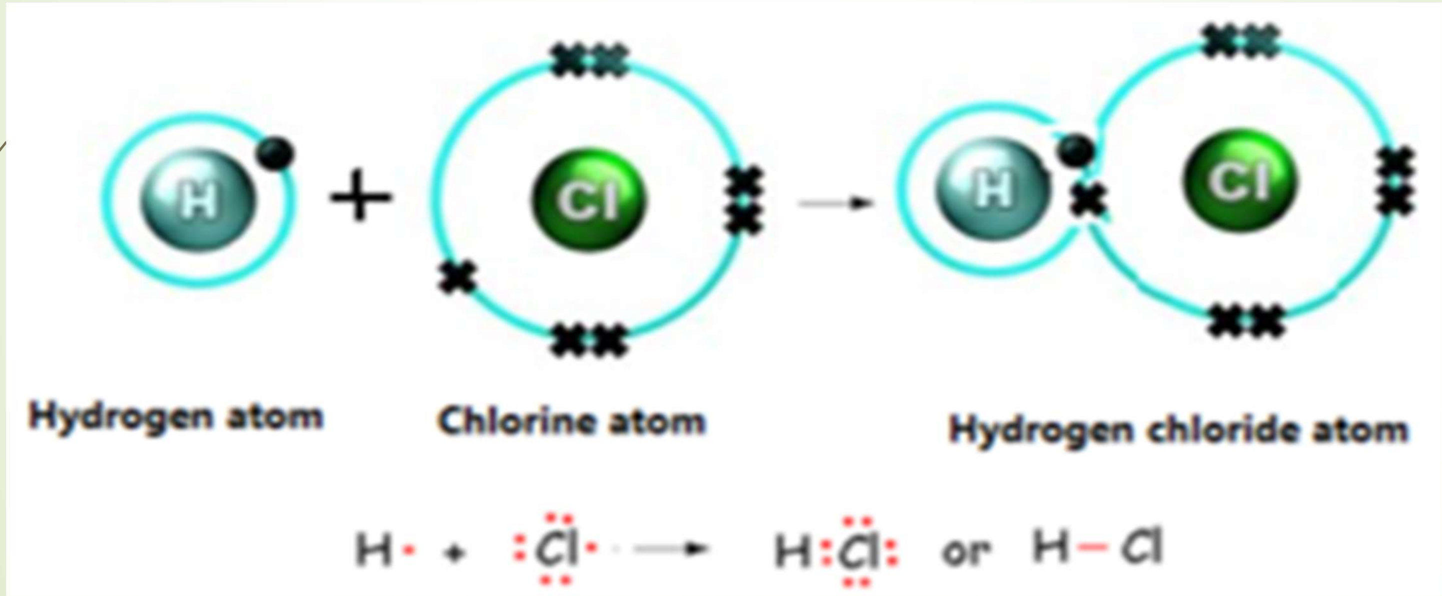
(i) Formation of chlorine molecule

- 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.



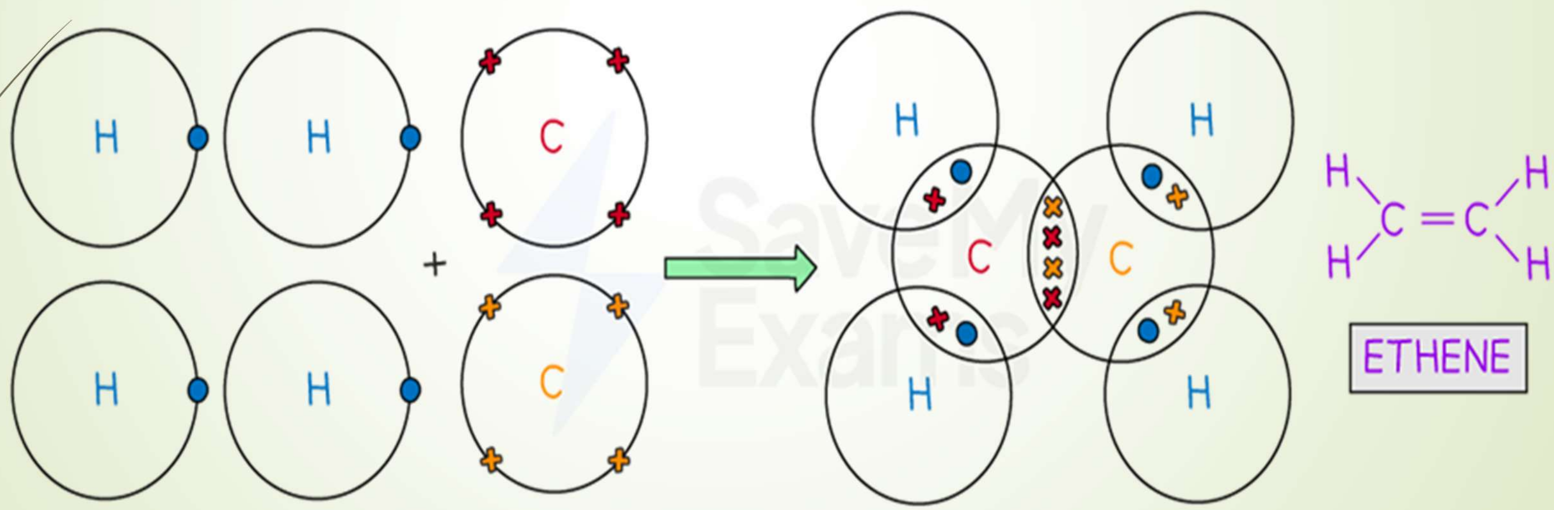
(ii) Formation of hydrogen chloride molecule

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

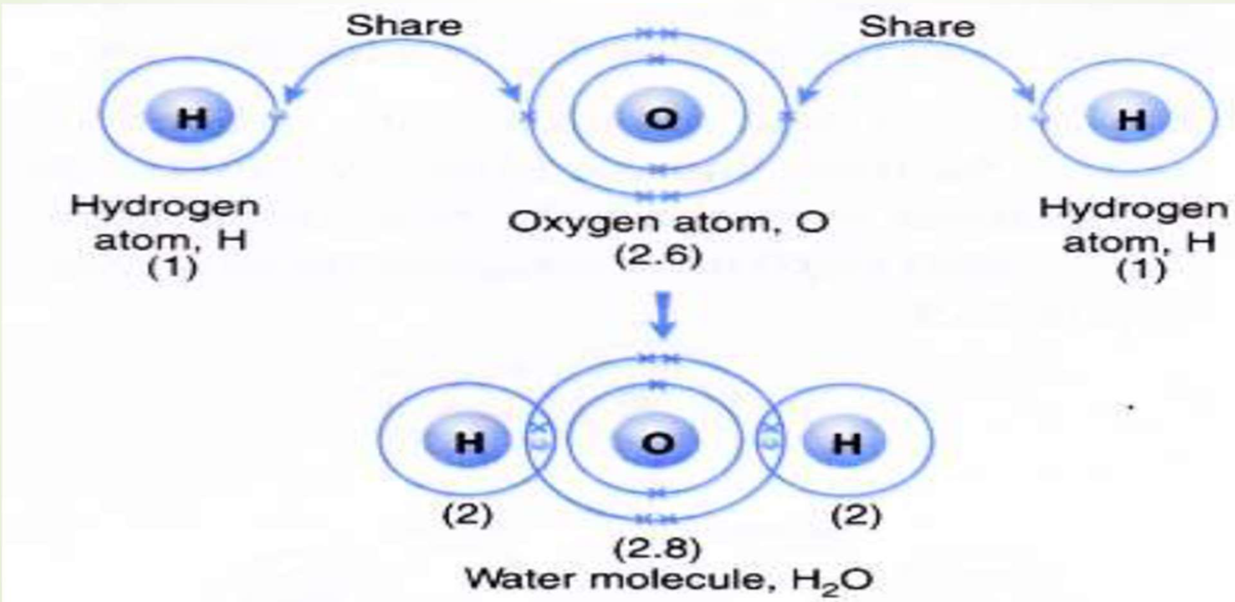


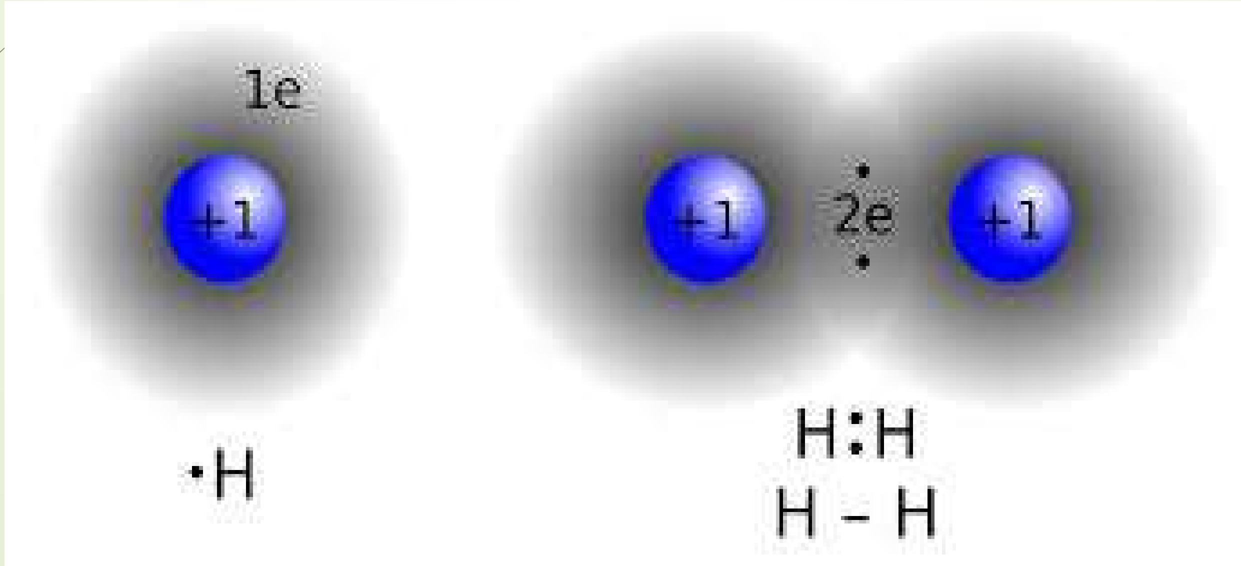
Figure *Formation of a water molecule, H_2O*

Types of covalent bonding

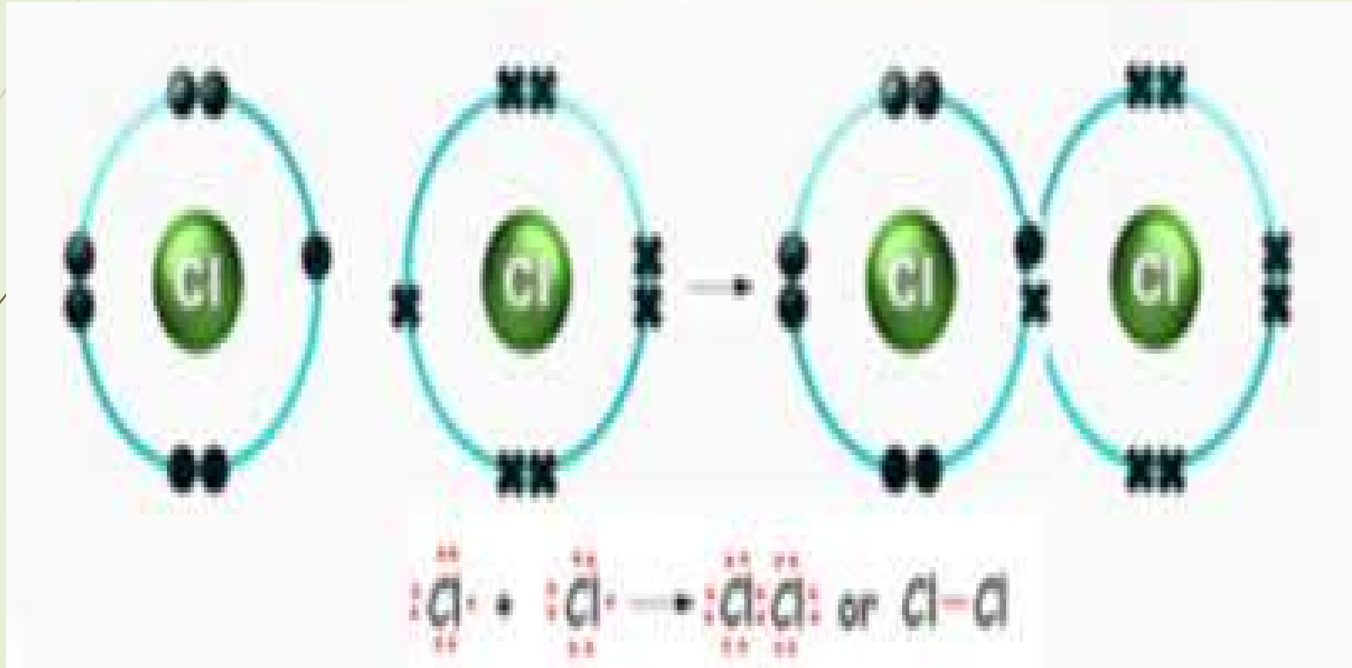
- 1. Normal Covalent Bond:
- In a normal covalent bond, each bonded atom contributes one electron to form the shared pair. This typically occurs between atoms with similar or identical electronegativities.

Examples:

- H-H in H_2 gas

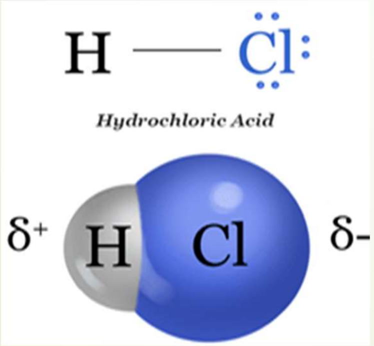


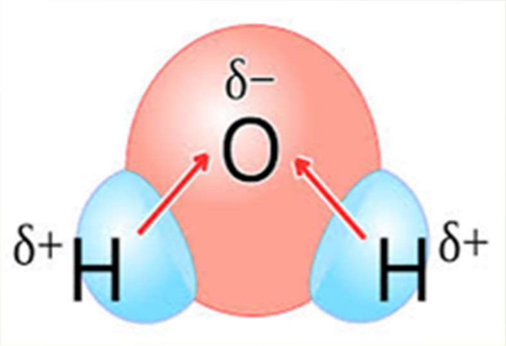
Cl-Cl in Cl_2 molecule



2. Polar Covalent Bond:

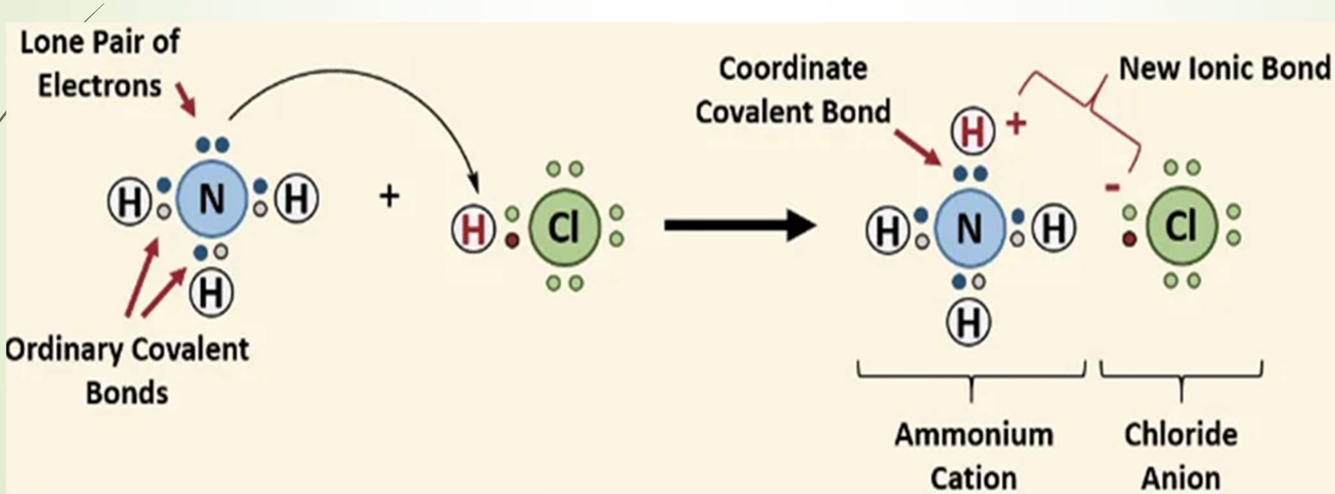
- This type of covalent bond forms between two atoms with a significant difference in electronegativity. The shared electron pair is not equally shared; it is pulled closer towards the more electronegative atom. This unequal sharing creates a bond dipole, where the more electronegative atom gains a partial negative charge (δ−) and the less electronegative atom gains a partial positive charge (δ+).
- Examples:
- H-Cl in HCl
- Cl is more electronegative than H, so the electrons are pulled towards Cl, creating $H^{\delta+}-Cl^{\delta-}$,
- 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 (→) 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.

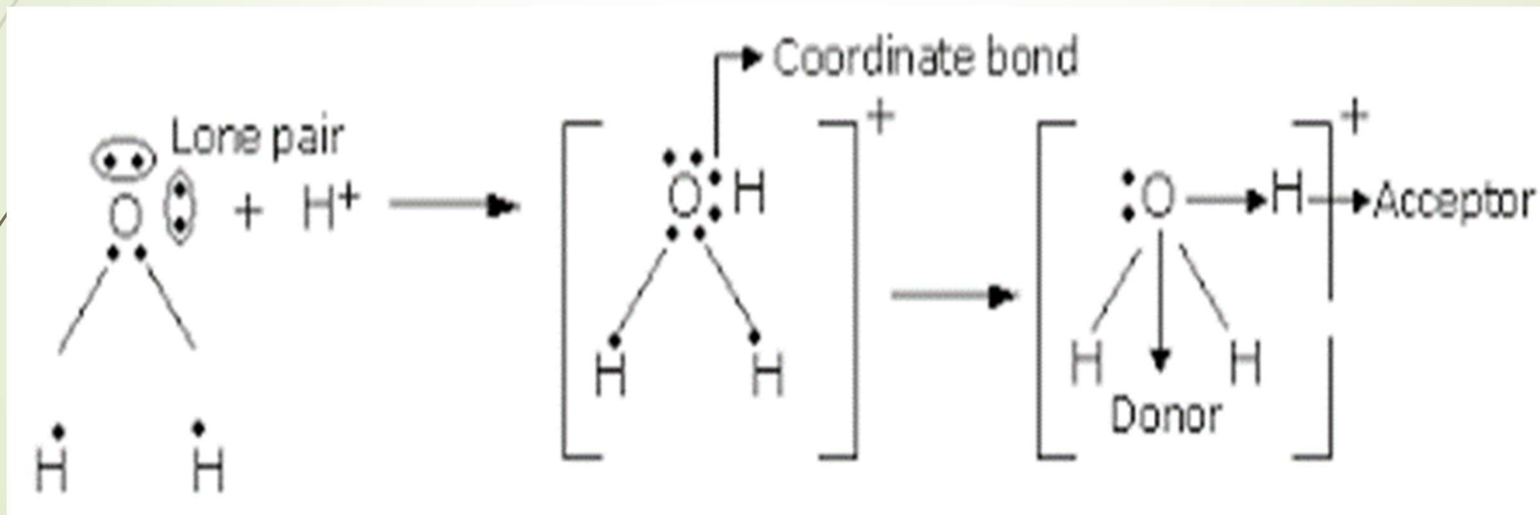


Examples

- 1. Formation of Hydroxonium Ion (H_3O^+):

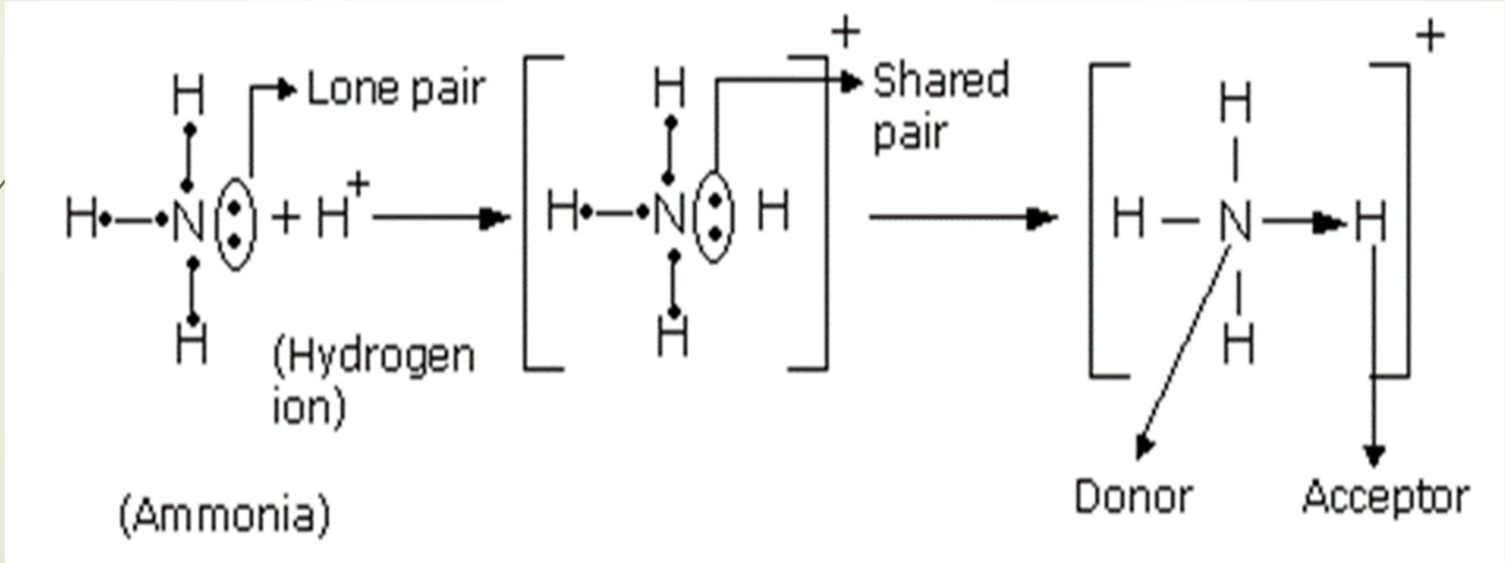
A water molecule (H_2O) has oxygen with two lone pairs. A hydrogen ion (H^+) has no electrons (empty 1s orbital). The oxygen atom donates a lone pair to the H^+ ion.

- **Illustration**



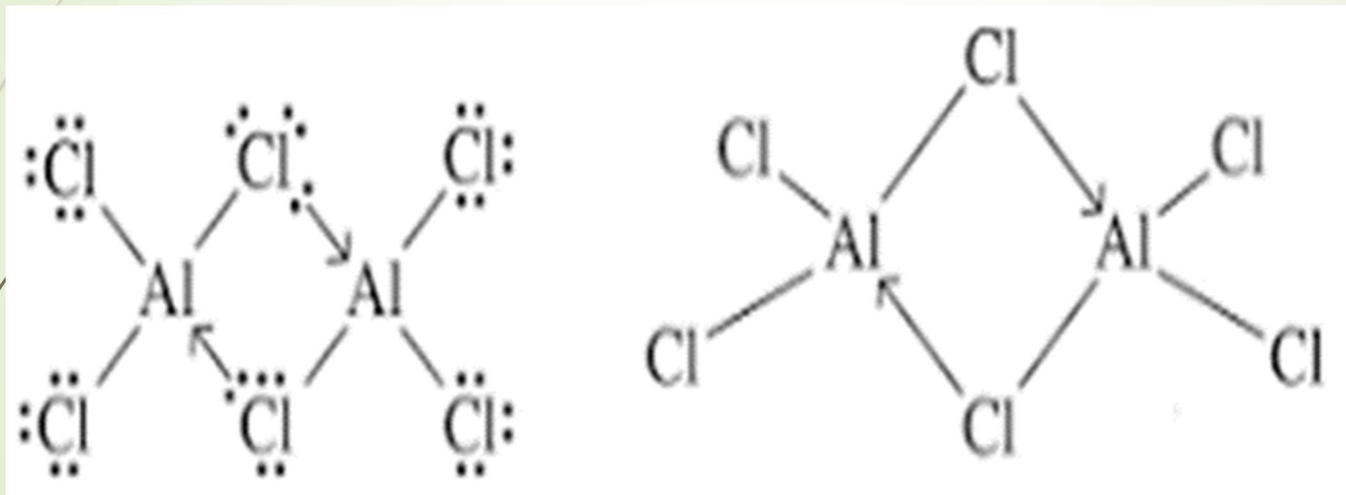
2. Formation of Ammonium Ion (NH_4^+):

- An ammonia molecule (NH_3) has nitrogen with one lone pair. A hydrogen ion (H^+) has an empty 1s orbital. The nitrogen atom donates its lone pair to the H^+ ion.
- **Illustration**



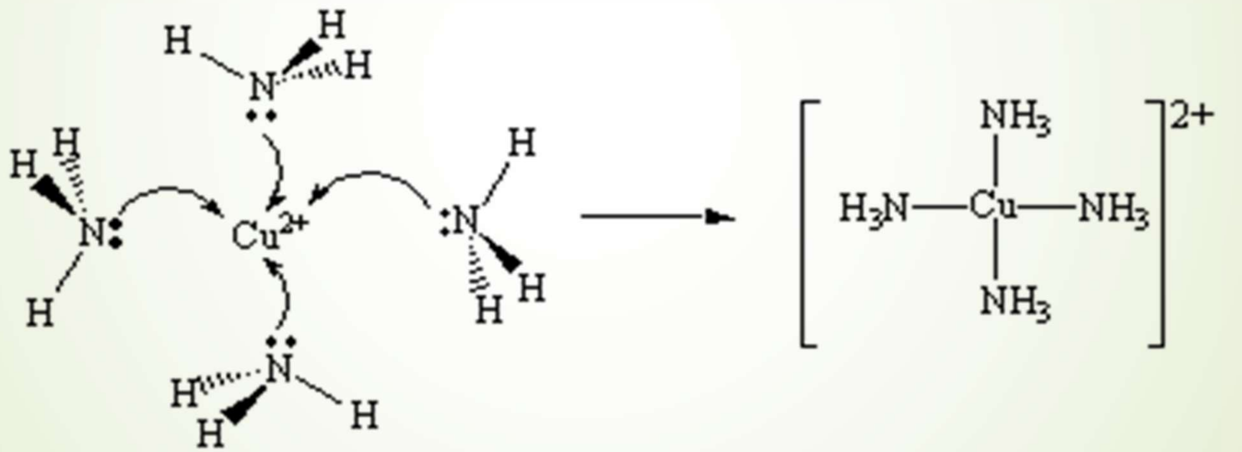
Formation of Dimerised Aluminium Chloride (Al_2Cl_6):

- In $AlCl_3$, Aluminium has an incomplete octet (6 valence electrons). A chlorine atom from another $AlCl_3$ molecule can donate a lone pair to the Aluminium atom with the vacant orbital.
- **Illustration**



Transition Metal Complexes:

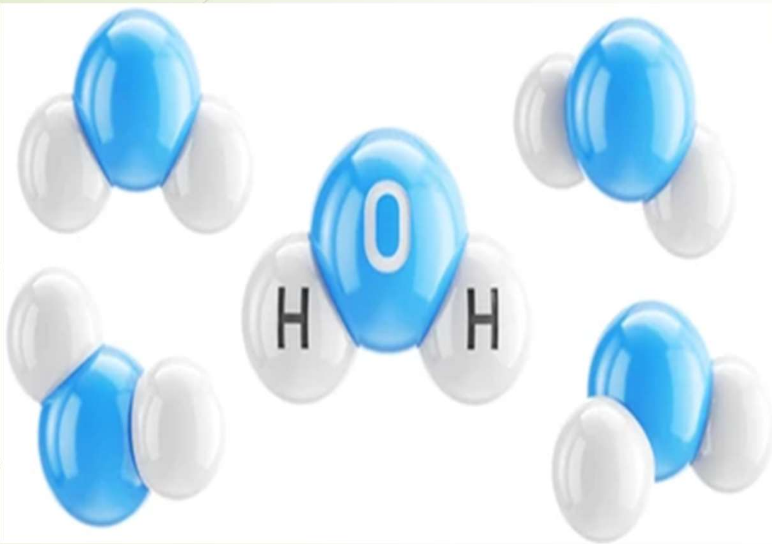
- Ligands (molecules or ions with lone pairs) form dative bonds with central metal ions (which have vacant orbitals).
- Example: Tetraamminecopper (II) ion, $[Cu(NH_3)_4]^{2+}$.
- **Illustration**

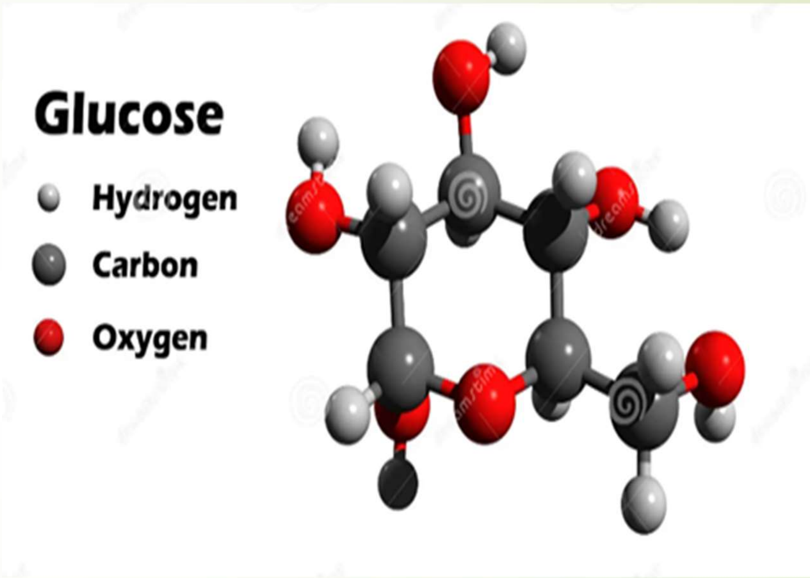


STRUCTURES OF COVALENT COMPOUNDS

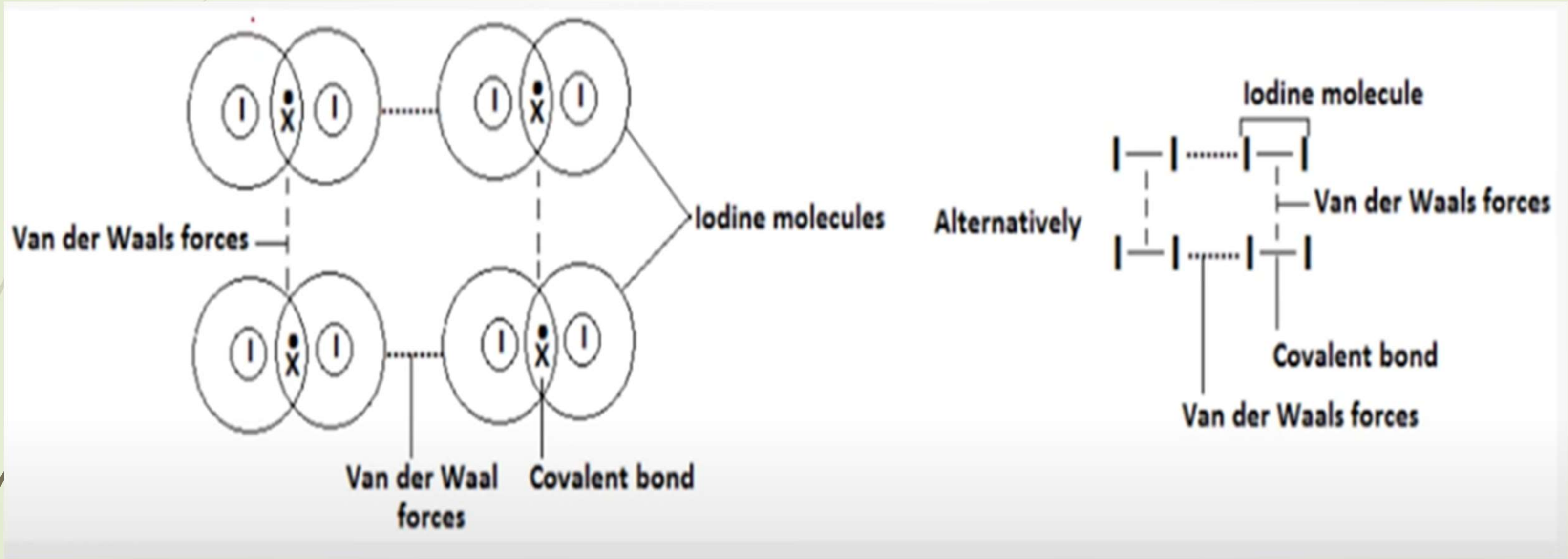
- Covalent substances exhibit two main types of structures, leading to vastly different properties:
- 1. Simple Molecular Structures:
- Consist of discrete molecules. The atoms within each molecule are held together by strong covalent bonds (intramolecular forces). However, the forces *between* these molecules (intermolecular forces) are relatively weak (van der Waals forces and hydrogen bonds).
- The intermolecular forces in molecular structures are ot 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

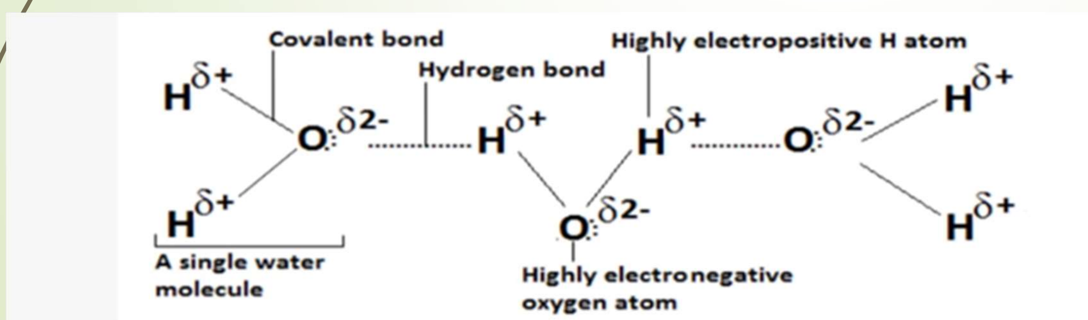
The essential requirements for the formation of a hydrogen bond are

- 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 bond are much stronger than weak van der waal force 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

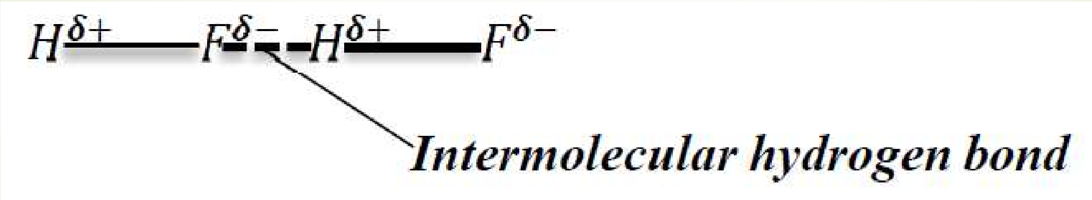
- Both ethanol and dimethyl ether have the same relative molecular mass of 46, however the boiling point of ethanol is higher at 78.5^{o}c than that of dimethyl ether at -24^{o}c
  - *Both have molecular structure with covalent bonds between the atoms. However, the intermolecular forces in ethanol are 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 confer 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;

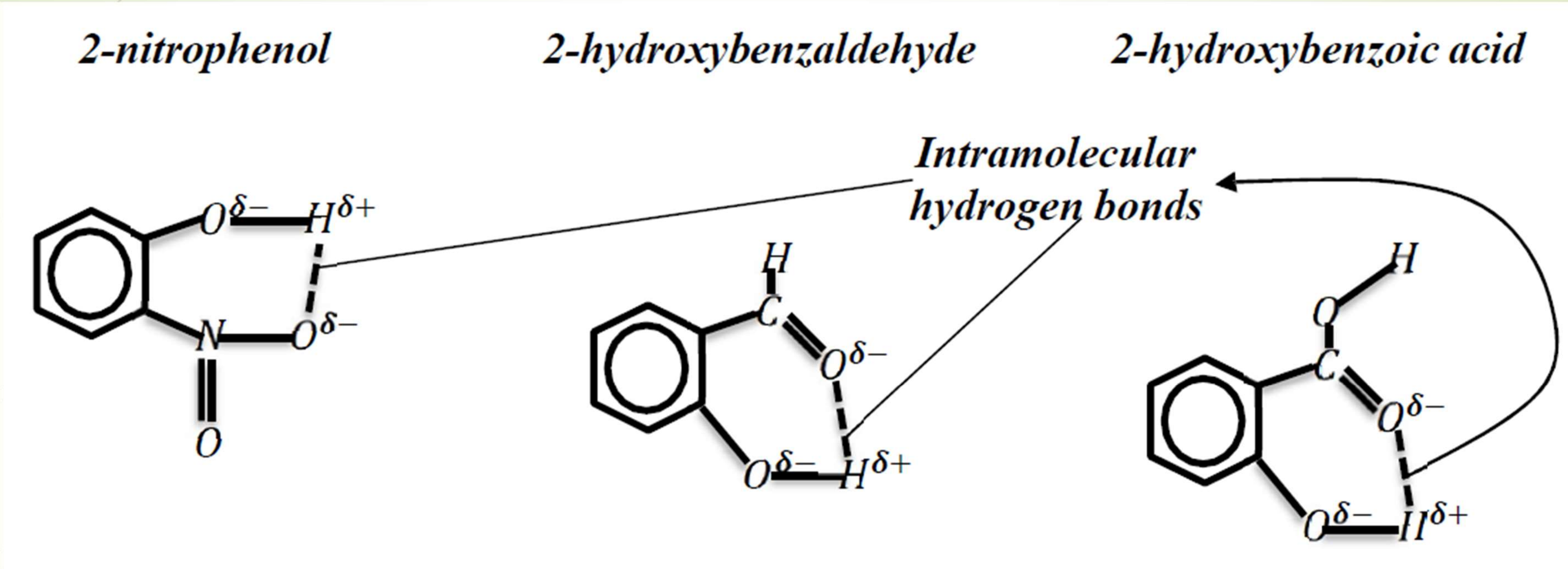
(a) Hydrogen fluoride , water and ammonia molecules

- 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.



(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

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 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.

ACTIVITY

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—H_2O.

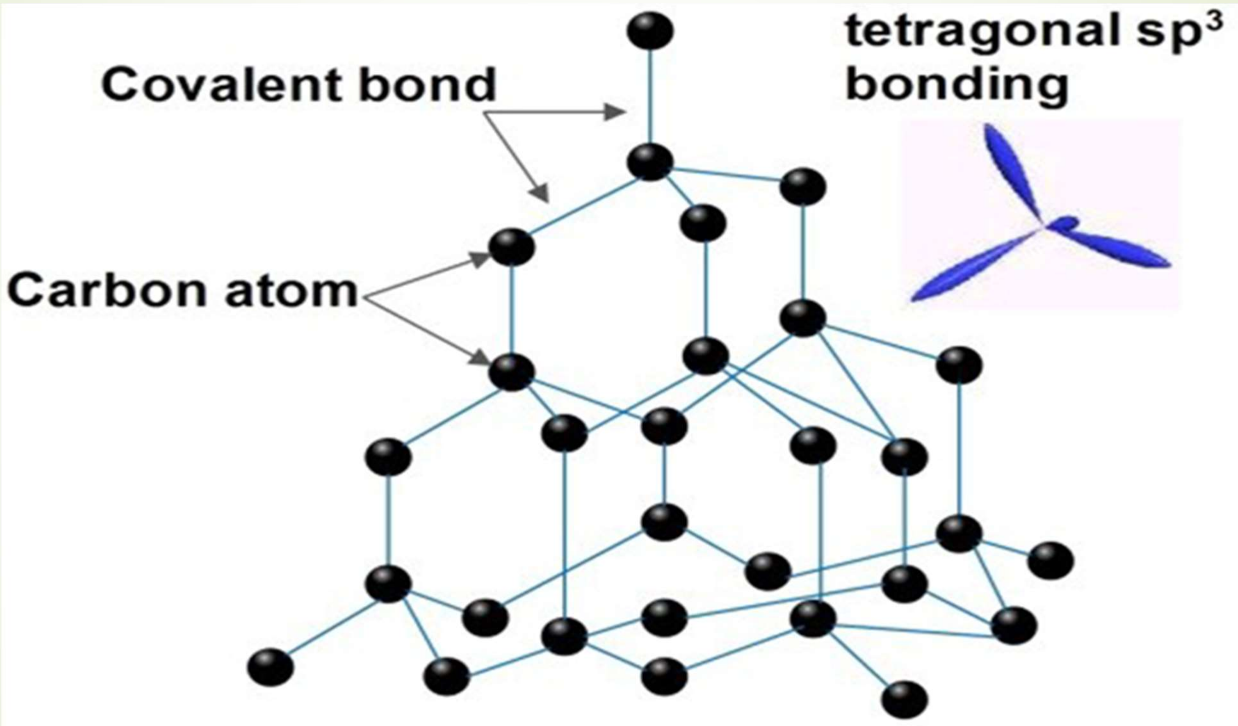
2. Giant Covalent Structures (Network Covalent):

- 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 (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

Illustration



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 dissolves in polar water molecules since there are no intermolecular interaction which would facilitate penetration into water molecule 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 lack any delocalised electrons for heat and electrical conductivity

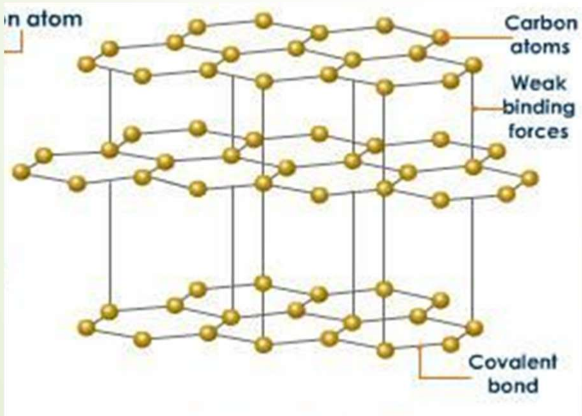
- **It is the hardest known substance**

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 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 presence delocalised electron 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 waal forces
- The presence of weak van der waal forces explain the slippery nature of graphite

Illustration



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 dissolves 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 in bonded to three others hence only three of its four valence electrons 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 feels.

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.

- 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_2. **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.

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 (H_2O), the air we breathe (O_2, 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 (H_2O), 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 (CO_2), a nonpolar covalent molecule, is vital for photosynthesis, the process by which plants produce food. The oxygen we breathe (O_2) and the nitrogen that makes up the majority of the atmosphere (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, 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 (NH_3) and bleach (sodium hypochlorite, 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.*

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:**

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

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

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

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**

a) Using the data in the table, explain why phosphorus and sulfur are solids at room temperature, whereas chlorine and argon are gases.

b) Relate the structure and type of bonding in each element to their melting points.

c) Based on your explanation, advise the school technician on how best to store each of these elements safely in the school laboratory.

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. At its core, 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 VSEPR model is particularly effective for predicting the shapes of simple and symmetric molecules where a central atom is clearly identifiable.

- The ability to predict molecular shapes is of paramount importance in chemistry because the three-dimensional arrangement of atoms in a molecule dictates many of its fundamental physical and chemical properties. These properties include a molecule's reactivity, polarity, color, and even its biological activity. For instance, the shape of a molecule influences how it interacts with other molecules, which in turn affects macroscopic properties such as boiling point and solubility. Furthermore, in biological systems, the specific three-dimensional structure of molecules like enzymes is critical for their function, as it determines how they bind to substrates and catalyze reactions. VSEPR theory provides insights that go beyond the two-dimensional representations offered by Lewis structures. While Lewis structures illustrate the connectivity of atoms and the presence of lone pairs

Principles of VSEPR Theory

- The foundation of VSEPR theory lies in understanding the role of valence shell electrons and how their arrangement around a central atom dictates molecular geometry. The theory hinges on the concept that electron pairs, both those involved in chemical bonds and those existing as lone pairs, will position themselves to minimize the electrostatic repulsion between them.

- The focus of VSEPR theory is on the **valence shell**, which represents the outermost electron-occupied shell of an atom. These are the electrons that are primarily involved in forming chemical bonds. The number of valence electrons an atom possesses is crucial as it determines the total number of electron pairs that will surround the central atom in a molecule. To apply VSEPR theory effectively, it is often necessary to first determine the Lewis structure of the molecule or ion. The Lewis structure provides a visual representation of the valence electrons, showing how atoms are connected through bonds and indicating the presence of any non-bonding electron pairs, also known as lone pairs. This structural information is the starting point for predicting the molecule's three-dimensional shape using VSEPR theory.

Lewis structures

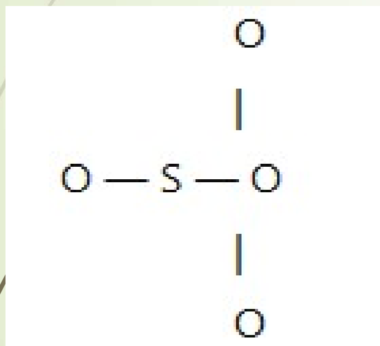
- A **Lewis structure** (also called Lewis dot diagram or electron dot structure) is a representation of a molecule or ion that shows:
- **Valence electrons** of atoms as dots,
- **Shared pairs (bonding pairs)** as lines,
- **Lone pairs (non-bonding electrons)** as dots around atoms.

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, SO_4^{2-}**

Skeleton structure:



(S is central because it is less electronegative than oxygen.)

2. carbon dioxide

Skeleton structure:

O — C — O

Carbon (C) is less electronegative than oxygen, so it goes in the center.

Arrange the two oxygen atoms on either side of carbon.

Connect each O to C with a single bond initially.

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

Steps for Drawing Lewis Structures – Cont'd

- 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. This can begot from the Lewis Dot symbol of the atom or its electronic configuration. For a negatively charged ion, add one electron for each negative charge. For a positively charged ion, subtract one electron for each positive charge.
- Add up the valence electrons from each atom.
- Add electrons for negative charges, subtract for positive charges.
- Example CO_2
  - Carbon (C) = **4 valence electrons**
  - Oxygen (O) = **6 valence electrons × 2 = 12**
  - Molecule is neutral → no added or removed electrons.
- **Total = 4 + 12 = 16 valence electrons**

Example: SO_4^{2-}

S = 6 valence electrons

O = 6 valence electrons × 4 = 24

Charge = –2 → add 2 electrons

Total = 6 + 24 + 2 = 32 valence electrons

Steps for Drawing Lewis Structures – Cont'd

- 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.

SO_4^{2-}

4 S—O bonds = 4 × 2 = 8 electrons used

Remaining = 32 – 8 = 24 electrons

CO_2

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

Each bond = 2 electrons

2 bonds × 2 electrons = 4 electrons

Remaining = 16 – 4 = 12 electrons

Steps for Drawing Lewis Structures – Cont'd

- 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.

Example CO_2

- 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 needs 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_4^{2-}

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

Steps for Drawing Lewis Structures – Cont'd

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_4^{2-}**

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

- **CO_2**

No electrons are left.

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

Steps for Drawing Lewis Structures – Cont’d

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

- If the central atom has fewer than 8 electrons, convert **lone pairs** from terminal atoms into **double or triple bonds** until the octet is satisfied.

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:
O=C=O

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

Steps for Drawing Lewis Structures – Cont'd

Final lewis structure of carbon dioxide