'O' Level Chemistry
Chapter 6 – Chemical Bonding, Structure and Properties

Bonding between Atoms

- In elements or compounds, the particles are held together by chemical bonds.
- Chemical bond: ***a force holding particles together***
- There are three types of chemical bonding between atoms:
  - Ionic bonding
  - Covalent bonding
  - Metallic bonding
- Elements and compounds formed by **different type of chemical bonding** have **different structures**.
- **Different structures** of substances give rise to **different physical properties**.

Ionic Bonding

- Recall:
  - Metal loses electrons to form positive ions
  - Non-metal gains electrons to form negative ions

  *Ions are formed by **loss / gain of electrons** to obtain the **stable electronic configuration of a noble gas***
- **Ionic bond** is the **strong electrostatic force of attraction** between **oppositely charged ions** in a compound.
- This bonding holds the **positive and negative ions** together to give **ionic compounds**.
- Ionic bonds are formed via **electron transfer**.
  - Usually between metals and non-metals.
  - Metal atoms lose electrons while non-metals gain the electrons lost by the metal atoms.

✓ **Example: Formation of Ionic Bond in Magnesium Chloride**
  - Magnesium atom loses 2 valence electrons
  - Each of the electrons is gained by a chlorine atom.

  Electrons are transferred from magnesium to chlorine
  - Magnesium ion and chloride ions are formed.
  - The positive magnesium ion, Mg^{2+}, and the negative chloride ions, Cl^-, are attracted together by strong electrostatic forces of attraction called the ionic bond.

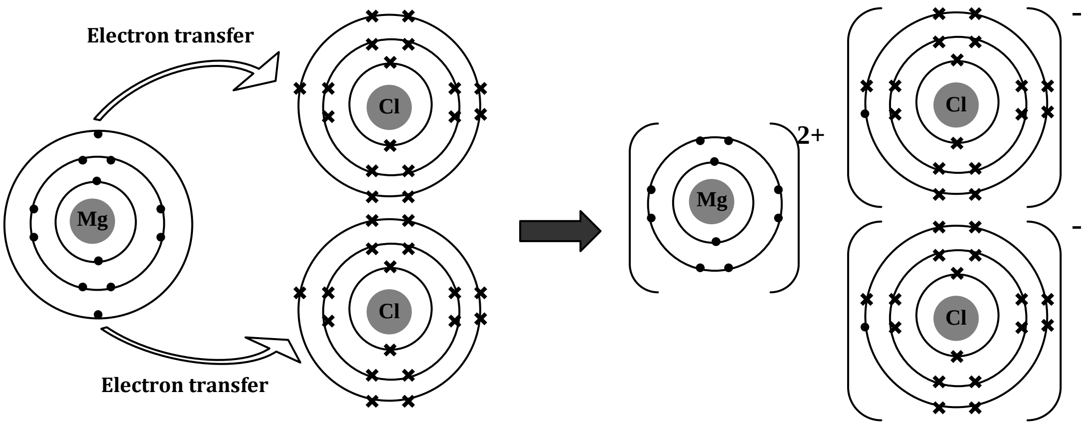


Fig. 1: 'Dot and cross' diagram of ionic bonding in magnesium chloride, $MgCl_2$

Chemical Formulae and Names of Ionic Compounds

- **Simple ions**:
  - Ions formed from a single atom, e.g. Na^+, Cl^-
  - Some metals form more than 1 kind of simple ion,
    - E.g. iron **(II)** ions, Fe^{2+}, and iron **(III)** ions, Fe^{3+}.

*The **roman numerals** indicate the no. of positive charge on the metal ion. Roman numerals are not needed if the metal only forms one ion.*

- **Polyatomic ions**:
  - Ions consist of more than 1 kind of atoms

Name of ion	Chemical formula
Hydroxide	OH^-
Nitrate	NO_3^-
Sulphate	SO_4^{2-}
Carbonate	CO_3^{2-}
Phosphate	PO_4^{3-}

Example: Carbonate ion, CO_3^{2-}
- it consists of 1 carbon atom and 3 oxygen atoms covalently bonded together
- **Total number of electrons in all the atoms is two more than the total number of protons.**
- Hence the group of 4 atoms has a charge of *2-* which is spread over all the 4 atoms giving carbonate a *2-* charge.

- **Writing chemical formulae and naming ionic compounds**
  - The positive and negative **charges must be balanced** in an ionic compound.
  - The **net charge** on the ionic compound must be **0**.
  - E.g. in magnesium chloride, each Mg^{2+} will need two Cl^- ions so that the charges are balanced and the net charge = 0. Hence the formula is written as $MgCl_2$.

There are different ways of balancing charges in ionic compounds.

- **Note on Writing Chemical Formula**
  - The number of a specific ion present in the compound is written as a **subscript**.
  - Metal ions / ammonium ions are **always written first** in the compound name and formula
  - Then, followed by the non-metal ions.
  - If the compound has > 1 polyatomic ion, the ion is placed **in brackets** followed by the **number** of this polyatomic ion present as the **subscript** outside the bracket.
  - If there is only 1 ion present, the number '1' is omitted.

✓ **Example: Writing the chemical formula of Calcium Nitrate**
  - Ca^{2+} and NO_3^- ions are present
  - To balance the charges, one Ca^{2+} ion must be balanced with two NO_3^- ions
  - Hence, the chemical formula is written as $Ca(NO_3)_2$

Exercise 1

Write down the chemical formulae of iron (III) chloride, sodium sulfate, aluminium nitrate, aluminium oxide.

Solution:

Iron (III) chloride (Fe^{3+} and Cl^-) – $FeCl_3$

$(+3) + 3(-1) = 0$

Sodium sulfate (Na^+ and SO_4^{2-}) – Na_2SO_4

$2(+1) + (-2) = 0$

Aluminium nitrate (Al^{3+} and NO_3^-) – $Al(NO_3)_3$

Al^{3+} NO_3^{-} ⇒ $Al(NO_3)_3$

Aluminium oxide (Al^{3+} and O^{2-}) – Al_2O_3

Al^{3+} O^{2-} ⇒ Al_2O_3

Structure of Ionic Compounds

- Ionic compounds consist of **positive ions** and **negative ions**.
- **Large numbers** of these oppositely charged ions are held together by **strong electrostatic attraction**.
- Hence, they are arranged in an **orderly**, **closely packed** manner to form a **giant ionic structure**, also called a **giant lattice**.
- The positive and negative ions are arranged in a **repetitive** manner (regular pattern) throughout the whole structure.

✓ **Physical Properties of Giant Ionic Structure / Giant Lattice**

- Hard *(but brittle)*
  - ➢ Strong ionic bonds
- Crystalline solids with flat sides and regular shapes
  - ➢ Ions are arranged in **straight rows** to form a large structure with **flat sides**
- **High melting and boiling points**
  - ➢ Large amount of energy is required to **break the strong ionic bonds** holding the ions together in a giant lattice structure in order for the ions to separate and move freely.
- Low volatility, no smell
  - ➢ **Strong ionic bonds** hold ions together
  - ➢ Thus, ionic compounds cannot evaporate easily
- Many ionic compounds are soluble in water but insoluble in organic solvents
  - ➢ **Ions attract water molecules**. This disrupts the crystal structure. The ions separate and go into solution.
  - ➢ Ions do not attract molecules of organic solvent and hence **insoluble in organic solvent**

Note: Not all ionic compounds are soluble in water *e.g. lead (II) sulfate and silver chloride salts are insoluble*

- Electrical Conductivity
  - ➢ **Cannot conduct in solid state**
    - ⇒ **Ions cannot move** and so cannot carry electric current
  - ➢ **Can conduct** in **molten** and **aqueous state**
    - ⇒ **Ions can move** freely to carry the electric current

Covalent Bonding

- A covalent bond is a bond formed by **sharing at least a pair of electrons**.
- It is formed **between atoms of non-metals**.
- It can be found **in elements** and **compounds**.
  - Between atoms of same element e.g. H_2, O_2, N_2
  - Between atoms of different elements e.g. CH_4, H_2O ⇒ these are **covalent compounds**
- Sharing of electrons
  - Electrons in the outermost shell are used to form bonds between atoms.
  - Electrons are **shared** in order for the atoms to obtain a ***stable electronic configuration of a noble gas.***

These electrons are called valence electrons. The no. of valence electrons used to form bonds is called its valency or valence number.

- **Covalent bonds in elements**
  - Example: Oxygen molecules
    - ➢ Oxygen atom has 6 valence electrons. It requires 2 more electrons to obtain a stable octet electronic structure.
    - ➢ Hence it shares 2 of its valence electrons with 2 valence electrons from another oxygen atom.
    - ➢ 2 pairs of electrons are shared to form a **double covalent bond** in the oxygen molecule (O = O)
    - ➢ A **diatomic molecule** is formed (Fig. 2).

'–' shows a covalent bond. Hence, '=' shows a ***double covalent bond****.*

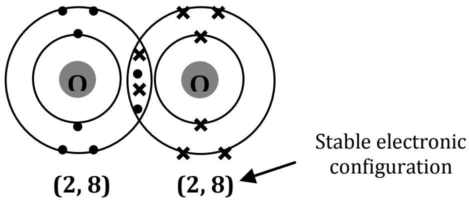


Fig. 2: 'Dot and cross' diagram of covalent bonding in an oxygen molecule

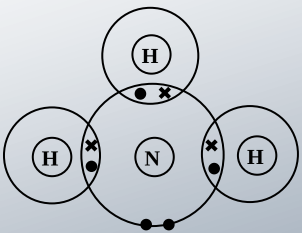
  - Other examples:
    - ➢ 1 pair of shared electrons ⇒ single covalent bond (e.g. in H – H)
    - ➢ 3 pairs of shared electrons ⇒ triple covalent bond (e.g. in N ≡ N)

- **Covalent bonds in compounds**
  - Example: Methane molecules
    - ➢ Carbon has an electronic configuration of 2,4. It needs 4 more electrons to obtain a stable octet configuration in its valence shell.
    - ➢ It needs to share four electrons with the single electron from each of the four hydrogen atoms.
    - ➢ Hence, each methane molecule contains **four C – H single bonds**.
  - Other examples: Water (H_2O), Carbon dioxide (CO_2), Ammonia (NH_3)

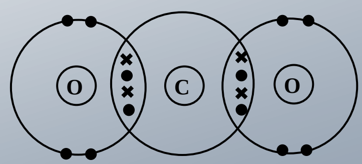
Exercise 2

Draw a 'dot and cross' diagram to show the covalent bonding in ammonia and carbon dioxide.

Solution:



Ammonia



Carbon dioxide

Structure of Covalently Bonded Elements and Compounds

- **Simple Molecular Structure**
  - Found in elements and compounds which consist of ***small molecules.***
  - Within a molecule, atoms are held together by ***strong covalent bonds.***
  - Between molecules, they are attracted by ***weak intermolecular forces.***

Also known as weak ***Van der Waal's forces***

✓ **Physical Properties of Simple Molecular Structure**
  - **Liquids or gases** at room temperature
    - ➢ Weak intermolecular forces ⇒ molecules are not held tightly together ⇒ free to move

Note: During melting or boiling, only the **molecules separate**. The **atoms in the molecules do not separate**.

- **Low** melting and boiling points
  - **Weak intermolecular forces** between molecules are **easily broken**.
  - Strong covalent bonds within each molecule are not broken.
- **Volatile** and evaporates easily to give a smell
  - Due to low boiling points
- Solubility
  - Most are insoluble in water
  - They dissolve in organic solvents
- They do not conduct electricity
  - Molecules **do not contain free moving ions or electrons**.

Note: To conduct electricity, substances must contain either free moving ions or electrons.

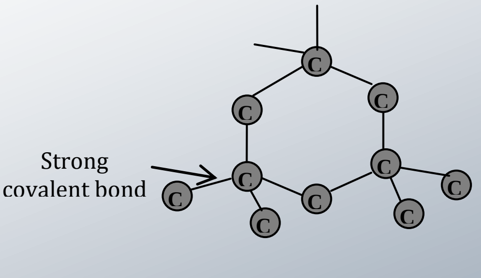
- **Giant Covalent Structure**
  - **Giant network of atoms** are held by **strong covalent bonds throughout the whole structure.**
  - **Diamond**
    - One form of the element carbon
    - Each carbon atom forms strong covalent bonds with four other carbon atoms
    - A **giant network of carbon atoms** each bonded **tetrahedrally** to four others
  - **Silicon Dioxide (SiO_2)**
    - Sand and quartz
    - A giant covalent structure **similar to diamond**
    - Each silicon atom is covalently bonded to four oxygen atoms.
    - Each oxygen atom is covalently bonded to two silicon atoms.
    - There are **no separate SiO_2 molecules**.
  - **Graphite**
    - Another form of element carbon
    - Giant covalent structure is **different from diamond**.
    - Carbon atoms are arranged in **flat layers of hexagons**.
    - **Each layer** is arranged in **rings of 6 carbon atoms** via **strong covalent bonds**.
    - However, the **forces between the layers** are **weak attractive forces** due to the unbonded electrons.

Each carbon atom is bonded to three other carbon atoms. This leaves each carbon atom with an **unbonded electron between the layers**.

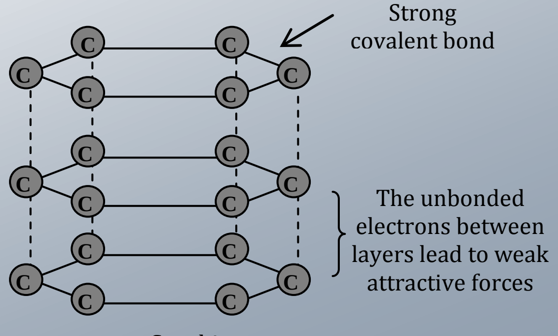
Exercise 3

Sketch the structures of Diamond and Graphite.

Solution:



Diamond



Graphite

✓ Physical Properties of Giant Covalent Structures

- **Solids** at room temperature
- **Hard**
  - ➢ **Many strong covalent bonds** between atoms.
  - ➢ Some are used as abrasives or cutting tools.
  - ➢ **Diamond** is the **hardest natural substance**
    - ⇒ Used to cut other hard solids.
    - ⇒ E.g. saws and diamond-tipped drills

Except Graphite which is soft and slippery.
- Recall the structure of graphite
- Weak forces between layers.
- The layers of atoms can slide past each other easily.
- Used as **lubricants** especially for hot machines.

- **Very high** melting and boiling points
  - ➢ Large numbers of **strong covalent bonds must be broken** to melt and boil the giant covalent structure.
- **Insoluble in water**
- Electrical Conductivity
  - ➢ **Do not conduct electricity (*except graphite)***
    - ⇒ **All** the valence electrons in the atoms are used to form covalent bonds.
    - ⇒ Thus there are **no free electrons** to conduct electricity.
  - ➢ **Graphite can conduct electricity**
    - ⇒ A non-metal which can conduct electricity.
    - ⇒ 1 carbon atom is bonded to 3 other carbon atoms.
    - ⇒ This leaves the carbon atom with one unbonded **mobile** valence electron which can **conduct electricity**.
    - ⇒ Hence, graphite can be used as:
      - Contacts in electric motors
      - The positive terminal in the common dry cell (battery)

Recall the different structures of Diamond and Graphite

Similar to the mobile electrons present in metals.

Metallic Bonding

- Found only in metals.
- Metal atoms have few electrons in the outermost shell.
- Each metal atom gives up its valence electron to become a **positive ion**.
- These **electrons are free to move** and **occupy spaces between the ions**.
- Hence, the positive ions are surrounded by a **'sea of delocalised electrons'**.
- The positive ions and negative electrons are held together by an attractive force called **Metallic bonding**
- A **metallic bond** is the **attractive force** between the **positive metal ions** and the **sea of negatively charged delocalised electrons**.

Free and Mobile electrons

Each positive ion is attracted to a 'sea of electrons' (and **not to a particular electron**)

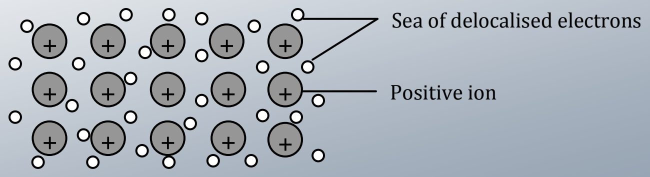
Structure of Metals

- Atoms are closely packed together in regular 3-D patterns to give a **Giant Metallic Lattice**.
- A giant lattice of positive ions in a 'sea of delocalised electrons'.
- There is very little space between atoms.

Exercise 4:

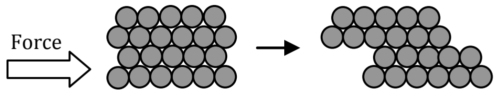
Draw a diagram to show the arrangement of particles in a metal.

Solution:



✓ Properties of Giant Metallic Structure

- **Malleable**
  - ➢ Can be bent without breaking
  - ➢ Can be hammered into different shapes
- **Ductile**
  - ➢ Can be stretched without breaking
  - ➢ Can be drawn into wires



- The layers of atoms can slide over each other easily when a force is applied.
- Metal does not break as the 'sea' of electron holds the atoms together

- **High** melting and boiling points
  - ➢ **Metallic bonds are strong** and a lot of energy is required to separate the atoms.
  - ➢ ***Except mercury and Group I metals***.
- Good conductors of electricity in **both solid and molten states**
  - ➢ The movement of **free mobile electrons** throughout the metal.
- Good conductors of heat
  - ➢ When one end of the metal is heated, the **mobile electrons gain energy** and move faster.
  - ➢ They **collide** with neighbouring atoms and electrons and **heat is transferred**.
  - ➢ The excited electrons in the hotter parts of the metal move to the cooler parts to transfer their heat energy.
  - ➢ Hence, the whole piece of metal becomes hot.

Summary

- *Chemical bond:* ***a force holding particles together***

	Ionic Bonding	Covalent Bonding	Metallic Bonding
Definition	Strong **electrostatic force of attraction** between **oppositely charged ions** in a compound.	Bond formed by the **sharing of at least one pair of electrons** between atoms	The **attractive force** between **positivel charged ions** and negatively charged **free, delocalised electrons.**
Formation of bond	▪ By **electron transfer.** ▪ Between **metals and non-metals** ▪ Metal atom loses e^- while non-metals gain e^-	▪ By **electron sharing** ▪ Between atoms of **non-metals** ▪ Electrons in the outer shell used to form bonds between atoms ⇒ **valency electrons** ▪ No. of valency electrons used to form bonds is called its **valency.**	▪ In the metal, each atom gives up its valence electrons to form positive ior ▪ These electrons are mobile or **delocalised** and formed a **'sea' of electrons.**
Description of particles formed	▪ **Ionic compounds** consisting +ve ions of metals and –ve ions of non-metals. ▪ *metal ions / ammonium ions are always written first* in the compound name and formula ▪ **Simple ions**: e.g. Na^+, Cl^- ▪ **Polyatomic ions**: consist of >1 kind of atom e.g. CO_3^{2-}, NO_3^-	**Covalent bonds in Elements** ▪ Between atoms of same element e.g. H_2, O_2 ▪ 1 pair of shared e^- ⇒ single covalent bond (H – H) ▪ 2 pairs of shared e^- ⇒ double covalent bond (O = O) ▪ 3 pairs of shared e^- ⇒ triple covalent bond (N ≡ N) **Covalent compounds** ▪ Between atoms of different elements e.g. CH_4, H_2O	▪ Positive ions in a 'sea' of mobile electrons. ▪ **Only in metals.**
Structure	**<u>Giant Ionic Lattice</u>** ▪ **Large numbers** of +ve and –ve charged ions arranged in an **orderly manner.** ▪ Ions **held in place** by strong ionic bonds ▪ Repetition of +ve and –ve charged ions **throughout the whole structure.**	**<u>Simple Molecular Structure</u>** ▪ Small covalent molecules ▪ **Strong covalent bonds within molecule** (i.e. between atoms in the molecule) ▪ **Weak forces** of attraction exist **between molecules** **<u>Giant Covalent Structure</u>** ▪ **Giant network of atoms** held by strong covalent bonds throughout structure. e.g. Diamond and Silicon Dioxide ▪ **Graphite: *(different structure)*** - C atoms arranged in flat layers in rings of 6 C atoms via strong covalent bonds. - Weak forces between layers.	**<u>Giant Metallic Structure</u>** ▪ Atoms **closely packed** together in **regular 3D patterns**

Physical Properties				
	Giant Ionic Lattice	**Simple Molecular Structure**	**Giant Covalent Structure**	**Giant Metallic Structure**
Physical state	▪ **Crystalline solid** ▪ **Flat sides** and **Regular shape:** Ions arranged in straight rows ▪ **Hard:** Ions held in place by strong ionic bonds	Most are **liquids or gases** at room temp. ▪ Intermolecular forces between molecules are weak ▪ Molecules are free to move	▪ **Solid** ▪ **Hard**: Many strong covalent bonds between atoms (used as abrasives e.g. Diamond is the hardest natural substance) ▪ ***Except Graphite**: **Soft and Slippery*** → weak forces between layers and can slide past each other easily	▪ **All solid except mercury**. ▪ **Malleable** and **Ductile:** - Layers of atoms slide over each other easily when a force is applied. - Metal does not break as the 'sea' of e^- holds the aton together
Melting and boiling points	**High m.p. and b.p.:** Large amount of heat is required to break the strong ionic bonds	**Low m.p. and b.p.:** Weak intermolecular forces between molecules are broken	**Very High m.p. and b.p.:** Many strong covalent bonds must be broken to melt and boil them	**High m.p. and b.p.:** Metallic bonding in most metals is strong **Except some** e.g. mercury
Volatility	**Not volatile:** Strong ionic bonds hold the ions together	**Volatile** and evaporate easily to give a smell	**Not volatile**	-
Solubility	▪ **Ions attract water molecules** which disrupt the crystal structure → ions separate and **dissolve in water** ▪ Ions do not attract molecules of organic solvent → **insoluble in organic solvent** ***Note: Not all ionic compounds are soluble in water e.g. $PbSO_4$***	▪ Most are insoluble in water. ▪ Dissolves in organic solvent.	▪ **Insoluble in water**	▪ Reaction with water. ▪ Depends on reactivity of the metal.
Electrical conductivity	▪ **Solid** ionic compound **cannot conduct.** ▪ **Molten or aqueous state** can **conduct electricity** as ions can move around freely to carry electric current. (**mobile charges**)	▪ **Do not conduct electricity** → molecules contain no mobile ions or electrons	▪ All valence e^- used to form covalent bonds → no free / mobile e^- → **do not conduct electricity** **Except Graphite**: ▪ Only non-metal that conducts electricity. ▪ 1 C bonded to 3 other C → still has 1 not bonded mobile valence e^- to	▪ Free / mobile e^- throughout the metal → **good conductor of electricity** in both solid and liquid state ▪ Metal heated → mobile e^- gain energy → move faster collide with neighbouring atoms and e^- → heat transferred through whole piece of metal → **good conductor of heat.**

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			conduct electricity.	