# STRUCTURE & BONDING

# CHEMICAL BONDING AND STRUCTURE

## A. CHEMICAL BONDING

A chemical bond is formed when atoms of the same or different elements **share**, **gain**, **donate or delocalize** their **outer** energy level **electrons** to combine during **chemical reactions** inorder to be **stable**.

Atoms have equal number of negatively charged electrons in the energy levels and positively charged protons in the nucleus.

Atoms are chemically stable if they have filled outer energy level. An energy level is full if it has duplet (2) or octet (8) state in outer energy level.

Noble gases have duplet /octet. All other atoms try to be like noble gases through chemical reactions and forming molecules.

Only electrons in the outer energy level take part in formation of a chemical bond. There are three main types of chemical bonds formed by atoms:

(i) covalent bond

(ii) ionic/electrovalent bond

(iii) metallic bond

### (i)COVALENT BOND

A covalent bond is formed when atoms of the same or different element share some or all the outer energy level electrons to combine during chemical reactions inorder to attain duplet or octet.

A shared pair of electrons is attracted by the nucleus (protons) of the two atoms sharing.

Covalent bonds are mainly formed by non-metals to form molecules. A molecule is a group of atoms of the same or different elements held together by a covalent bond. The number of atoms making a molecule is called **atomicity**. Noble gases are **monatomic** because they are stable and thus do not bond with each other or other atoms. Most other gases are **diatomic**

The more the number of electrons shared, the stronger the covalent bond.

A pair of electrons that do not take part in the formation of a covalent bond is called a **lone pair of electrons.**

Mathematically, the number of electrons to be shared by an atom is equal to the number of electrons remaining for the atom to be stable/attain duplet/octet /have maximum electrons in outer energy level.

**The following diagrams illustrate the formation of covalent bonds:**

a)hydrogen molecule is made up of two hydrogen atoms in the outer energy level each requiring one electron to have a stable duplet.

To show the formation of covalent bonding in the molecule then the following data/information is required;

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | H | H |
| Number of protons/electrons | 1 | 1 |
| Electron configuration/structure | 1: | 1: |
| Number of electron in outer energy level | 1 | 1 |
| Number of electrons remaining to be stable/shared | 1 | 1 |
| Number of electrons not shared(lone pairs) | 0 | 0 |

**Diagram method 1**

**Diagram method 2**

**H** **H**
•x

**Note:**

After bonding the following **intramolecular** forces exist:

(i)the attraction of the shared electrons by both nucleus /protons of the atoms

(ii) the repulsion of the nucleus of one atom on the other.

(iii)balance of the attraction and repulsion is maintained inside/intramolecular/within the molecule as follows;

$E_1$

$P_1$ $P_1$

$E_1$

(iv)Protons($P_1$) from nucleus of atom 1 **repel** protons ($P_2$) from nucleus of atom 2.

(v)Electron ($E_1$) in the energy levels of atom 1 **repel** electron ($E_2$) in the energy levels of atom 2.

(vi) Protons($P_1$) from nucleus of atom 1 **attract** electron ($E_2$) in the energy levels of atom 2.

(vii) protons ($P_2$) from nucleus of atom 2 **attract** electron ($E_2$) in the energy levels of atom 2.

b) Fluorine, chlorine, bromine and iodine molecules are made up also of two atoms sharing the outer energy level electrons to have a stable octet.

To show the formation of covalent bonding in the molecule then the following data/information is required;

(i) fluorine

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | F | F |
| Number of protons/electrons | 9 | 9 |
| Electron configuration/structure | 2:7 | 2:7 |
| Number of electron in outer energy level | 7 | 7 |
| Number of electrons remaining to be stable/shared | 1 | 1 |
| Number of outer electrons not shared( 3-lone pairs) | 6 | 6 |

**Diagram method 1**

**Diagram method 2**

(ii) chlorine

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | Cl | Cl |
| Number of protons/electrons | 17 | 17 |
| Electron configuration/structure | 2:8:7 | 2:8:7 |
| Number of electron in outer energy level | 7 | 7 |
| Number of electrons remaining to be stable/shared | 1 | 1 |
| Number of outer electrons not shared( 3-lone pairs) | 6 | 6 |

**Diagram method 1**

**Diagram method 2**

(iii) Bromine

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | Br | Br |
| Number of protons/electrons | 35 | 35 |
| Electron configuration/structure | 2:8:18:7 | 2:8:18:7 |
| Number of electron in outer energy level | 7 | 7 |
| Number of electrons remaining to be stable/shared | 1 | 1 |
| Number of outer electrons not shared( 3-lone pairs) | 6 | 6 |

**Diagram method 1**

**Diagram method 2**

(iv) Iodine

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | I | I |
| Number of protons/electrons | 53 | 53 |
| Electron configuration/structure | 2:8:18:18:7 | 2:8:18:18:7 |
| Number of electron in outer energy level | 7 | 7 |
| Number of electrons remaining to be stable/shared | 1 | 1 |
| Number of outer electrons not shared( 3-lone pairs) | 6 | 6 |

**Diagram method 1**

**Diagram method 2**

c) Oxygen molecule is made up of two atoms sharing each two outer energy level electrons to have a stable octet as shown below;

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | O | O |
| Number of protons/electrons | 8 | 8 |
| Electron configuration/structure | 2:6 | 2:6 |
| Number of electron in outer energy level | 6 | 6 |
| Number of electrons remaining to be stable/shared | 2 | 2 |
| Number of outer electrons not shared( 2-lone pairs) | 4 | 4 |

**Diagram method 1**

**Diagram method 2**

d) Nitrogen and phosphorus molecule is made up of two atoms sharing each three outer energy level electrons to have a stable octet as shown below;

(i) Nitrogen

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | N | N |
| Number of protons/electrons | 7 | 7 |
| Electron configuration/structure | 2:5 | 2:5 |
| Number of electron in outer energy level | 5 | 5 |
| Number of electrons remaining to be stable/shared | 3 | 3 |
| Number of outer electrons not shared ( 3-lone pairs) | 2 | 2 |

**Diagram method 1**

**Diagram method 2**

(ii) Phosphorus

| | | |
|---|---|---|
| Symbol of atom/element taking part in bonding | P | P |
| Number of protons/electrons | 15 | 15 |
| Electron configuration/structure | 2:8:5 | 2:8:5 |
| Number of electron in outer energy level | 5 | 5 |
| Number of electrons remaining to be stable/shared | 3 | 3 |
| Number of outer electrons not shared ( 3-lone pairs) | 2 | 2 |

**Diagram method 1**

**Diagram method 2**

e) Water molecule is made up of hydrogen and oxygen. Hydrogen requires to share one electron with oxygen to be stable/attain duplet. Oxygen requires to share two electrons to be stable/attain octet. Two hydrogen atoms share with one oxygen atom for both to be stable as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | O | H |
| Number of protons/electrons | 8 | 1 |
| Electron configuration/structure | 2:6 | 1 |
| Number of electron in outer energy level | 6 | 1 |
| Number of electrons remaining to be stable/shared | 2 | 1 |
| Number of electrons not shared( 2-Oxygen lone pairs) | 4 | 0 |

**Diagram method 1**

**Diagram method 2**

f) Ammonia molecule is made up of Hydrogen and Nitrogen. Hydrogen requires to share one electron with Nitrogen to be stable/attain duplet. Nitrogen requires to share three electrons to be stable/attain octet. Three hydrogen atoms share with one nitrogen atom for both to be stable as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | N | H |
| Number of protons/electrons | 7 | 1 |
| Electron configuration/structure | 2:5 | 1: |
| Number of electron in outer energy level | 5 | 1 |
| Number of electrons remaining to be stable/shared | 3 | 1 |
| Number of electrons not shared( 1-Nitrogen lone pairs) | 2 | 0 |

**Diagram method 1**

**Diagram method 2**

g)Carbon(IV) oxide molecule is made up of carbon and oxygen. Carbon requires to share four electrons with oxygen to be stable/attain octet. Oxygen requires to share two electrons to be stable/attain octet. Two oxygen atoms share with one carbon atom for both to be stable as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | O | C |
| Number of protons/electrons | 8 | 6 |
| Electron configuration/structure | 2:6 | 2:4 |
| Number of electron in outer energy level | 6 | 4 |
| Number of electrons remaining to be stable/shared | 2 | 4 |
| 2-lone pairs from each Oxygen atom) | 2 | 0 |

**Diagram method 1**

**Diagram method 2**

h) Methane molecule is made up of hydrogen and carbon. Hydrogen requires sharing one electron with carbon to be stable/attain duplet. Carbon requires sharing four electrons to be stable/attain octet. Four hydrogen atoms share with one carbon atom for both to be stable as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | C | H |
| Number of protons/electrons | 6 | 1 |
| Electron configuration/structure | 2:4 | 1 |
| Number of electron in outer energy level | 4 | 1 |
| Number of electrons remaining to be stable/shared | 4 | 1 |
| Number of electrons not shared ( No lone pairs) | 0 | 0 |

**Diagram method 1**

**Diagram method 2**

i) Tetrachloromethane molecule is made up of chlorine and carbon. Chlorine requires sharing one electron with carbon to be stable/attain octet. Carbon requires sharing four electrons to be stable/attain octet. Four chlorine atoms share with one carbon atom for both to be stable as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | C | Cl |
| Number of protons/electrons | 6 | 17 |
| Electron configuration/structure | 2:4 | 2:8:7 |
| Number of electron in outer energy level | 4 | 7 |
| Number of electrons remaining to be stable/shared | 4 | 1 |
| 3-lone pairs from each Chlorine atom(24 electrons) | 0 | 6 |

**Diagram method 1**

**Diagram method 2**

j) Ethane molecule is made up of six hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Three hydrogen atoms share with one carbon atom while another three hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing a pair of the remaining electrons as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | C | H |
| Number of protons/electrons | 6 | 1 |
| Electron configuration/structure | 2:4 | 1 |
| Number of electron in outer energy level | 4 | 1 |
| Number of electrons remaining to be stable/shared | 4 | 1 |
| Number of electrons not shared( No lone pairs) | 0 | 0 |

**Diagram method 1**

**Diagram method 2**

k) Ethene molecule is made up of four hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Two hydrogen atoms share with one carbon atom while another two hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing two pairs of the remaining electrons as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | C | H |
| Number of protons/electrons | 6 | 1 |
| Electron configuration/structure | 2:4 | 1 |
| Number of electron in outer energy level | 4 | 1 |
| Number of electrons remaining to be stable/shared | 4 | 1 |
| Number of electrons not shared( No lone pairs) | 0 | 0 |

**Diagram method 1**

**Diagram method 2**

l) Ethyne molecule is made up of two hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. One hydrogen atoms share with one carbon atom while another hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing three pairs of the remaining electrons as shown below;

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | C | H |
| Number of protons/electrons | 6 | 1 |
| Electron configuration/structure | 2:4 | 1 |
| Number of electron in outer energy level | 4 | 1 |
| Number of electrons remaining to be stable/shared | 4 | 1 |
| Number of electrons not shared( No lone pairs) | 0 | 0 |

**Diagram method 1**

**Diagram method 2**

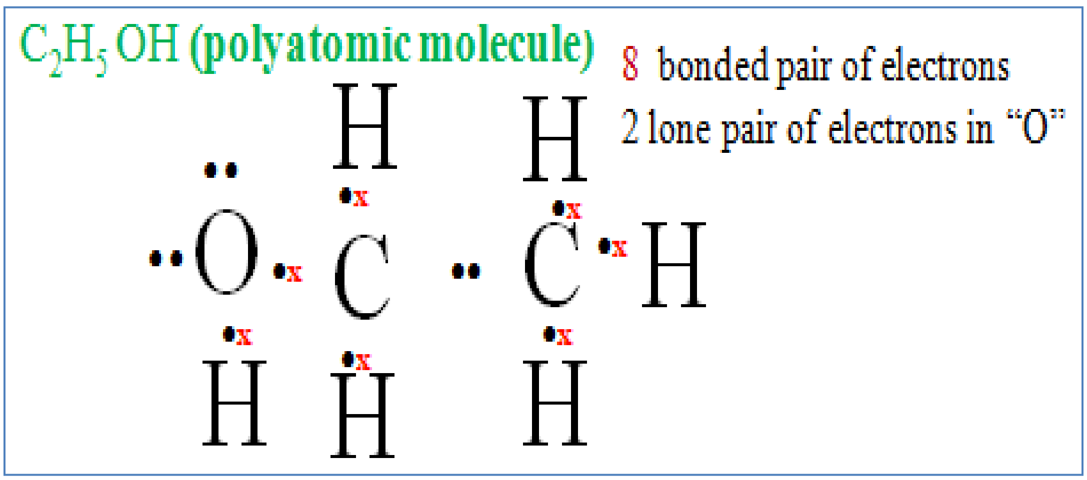
j) Ethanol molecule is made up of six hydrogen one Oxygen

atom two carbon atoms.

Five Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet

Each Carbon uses four electrons to share with "O"and "H"attain octet/duplet.

NB: Oxygen has two lone pairs

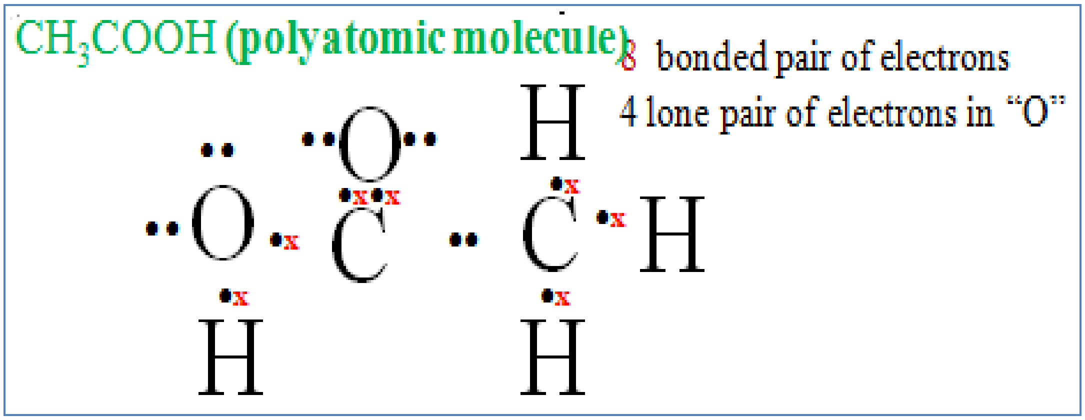



**j)Ethanoic molecule is made up of four hydrogen two Oxygen atom two carbon atoms.**

**Three Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet**

**Each Carbon uses four electrons to share with "O"and "H"attain octet/duplet.**

**NB: Each Oxygen atom has two lone pairs**




By convention (as a rule), a

(i) **single** covalent bond made up of two shared( **a pair**) electrons is represented by a dash(---)

(ii) **double** covalent bond made up of four shared( **two pairs)** electrons is represented by a double dash(==)

(iii) **triple** covalent bond made up of six shared( **three pairs)** electrons is represented by a triple dash(==)

The representation below show the molecules covered in (a) to (k) above:

a) Hydrogen molecule($H_2$) H--H

b) Fluorine molecule($F_2$) F--F

c) Chlorine molecule($Cl_2$) Cl--Cl

d) Bromine molecule($Br_2$) Br--Br

e) Iodine molecule($I_2$) I--I

f) Oxygen molecule($O_2$) O=O

g) Nitrogen molecule($N_2$) N=N

h) Phosphorus molecule($P_2$) P=P

i) Water molecule ($H_2O$) H--O--H

j Ammonia molecule($NH_3$)

```
H--N--H

   H
```

k)Carbon(IV) oxide molecule($CO_2$) O==C==O

l)Methane molecule($CH_4$)

```
   H
H--C--H
   H
```

m)Tetrachloromethane molecule($CCl_4$

```
    Cl
Cl--C--Cl
    Cl
```

n)Ethane molecule($C_2H_6$)

```
   H   H
H--C—C--H
```

H H

p)Ethene molecule($C_2H_4$) H-C==C-H

H H

q)Ethyne molecule($C_2H_6$) H-C—C-H

Dative /coordinate bond

A dative/coordinate bond is a covalent bond formed when a lone pair of electrons is donated then shared to an electron-deficient species/ion/atom.

During dative/coordinate bonding, all the shared pair of electrons are donated by one of the combining/bonding species/ ion/atom.

Like covalent bonding, coordinate /dative bond is mainly formed by non-metals.

**Illustration of coordinate /dative bond**

a)Ammonium ion($NH_4^+$)

The ammonium ion is made up of ammonia ($NH_3$) molecule and hydrogen ($H^+$) ion. ($H^+$) ion has no electrons. $NH_3$ is made up of covalent bonding from Nitrogen and Hydrogen. One lone pair of electrons is present in Nitrogen atom after the bonding. This lone pair is donated and shared with an electron-deficient $H^+$ ion

Diagram method 1

Diagram method 2

b)Phosphine ion ($PH_4^+$)

The Phosphine ion is made up of phosphine($NH_3$) molecule and hydrogen ($H^+$) ion. ($H^+$) ion has no electrons. $PH_3$ is made up of covalent bonding from Phosphorus and Hydrogen. One lone pair of electrons is present in Phosphorus atom. After the bonding this lone pair is donated and shared with the electron-deficient $H^+$ ion

Diagram method 1

Diagram method 2

c) Hydroxonium ($H_3O^+$) ion

The hydroxonium ion is made up of water ($H_2O$) molecule and hydrogen ($H^+$) ion. ($H^+$) ion has no electrons. The $H_2O$ molecule is made up of covalent bonding from Oxygen and Hydrogen. One lone pair of electrons out of the two present in Oxygen atom after the bonding is donated and shared with the electron-deficient $H^+$ ion

Diagram method 1

Diagram method 2

d) Carbon (II) oxide (CO)

Carbon (II) oxide is made up of carbon and Oxygen atoms sharing each two outer electron and not sharing each two electrons. Oxygen with an extra lone pair of electrons donates and share with the carbon atom for both to be stable.

Diagram method 1

Diagram method 2

e) Aluminium (III) chloride ($AlCl_3$/$Al_2Cl_6$)

Aluminium (III) chloride is made up of aluminium and chlorine. One aluminium atom shares its outer electrons with three separate chlorine atoms. All chlorine atoms attain stable octet but aluminium does not. Another molecule of aluminium chloride shares its chlorine lone pair of electrons with the aluminium atom for both to be stable. This type of bond exists only in vapour phase after aluminium chloride sublimes.

Diagram method 1

Diagram method 2

A dative/coordinate bond is by convention represented by an arrow (→) heading from the donor of the shared pair of electrons.

Below is the representation of molecules in the above examples;

a)Ammonium ion.

H

H− N→H

H

b)Phosphine ion

H

H− P→H

H

c)Hydroxonium ion

H− O→H

H

d)Carbon(II) oxide

O→C

d) Aluminium(III)chloride

Cl Cl Cl

Al Al

Cl Cl Cl

## (ii)**IONIC/ELECTROVALENT BOND**

An ionic/electrovalent bond is **extreme** of a covalent bond.

During ionic/electrovalent bonding there is complete **transfer** of valence electrons to one electronegative atom from an electropositive atom.

All metals are electropositive and easily/readily donate/lose their valence electrons.

All non-metals are electronegative and easily/readily gain/acquire extra electrons.

Ionic/electrovalent bonding therefore mainly involves transfer of electrons from metal/metallic radical to non-metallic radical.

When an electropositive atom donates /loses the valence electrons, it forms a positively charged **cation** to attain stable octet/duplet.

When an electronegative atom gains /acquires extra valence electrons, it forms a negatively charged **anion** to attain stable octet/duplet**.**

The **electrostatic attraction force** between the **stable** positively charged cation and the stable negatively charged anion with **opposite** charges constitute the ionic bond.

Like in covalent/dative/coordinate bonding, only the outer energy level electrons take part in the formation of ionic/electrovalent bond

Like in covalent/dative/coordinate bonding, the **more** electrons taking part / involved in the formation of ionic/electrovalent bond, the **stronger** the ionic /electrovalent bond.

## Illustration of ionic /electrovalent bond

a)Sodium chloride(NaCl)

Sodium chloride(NaCl) is formed when a sodium atom donate its outer valence electrons to chlorine atom for both to attain stable octet:

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | Na | Cl |
| Number of protons/electrons | 11 | 17 |
| Electron configuration/structure | 2:8:1 | 2:8:7 |
| Number of electron in outer energy level | 11 | 7 |
| Number of electrons donated and gained to be stable | 1 | 1 |
| New electron configuration/structure | 2:8: | 2:8: |
| Symbol of cation/anion after bonding | $Na^+$ | $Cl^-$ |

**Diagram**

b)Magnesium chloride($MgCl_2$)

Magnesium chloride ($MgCl_2$) is formed when a magnesium atom donate its two outer valence electrons to chlorine atoms. Two chlorine atoms are required to gain each one electron. All the ions (cations and anions) attain stable octet:

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | Mg | Cl |
| Number of protons/electrons | 11 | 17 |
| Electron configuration/structure | 2:8:2 | 2:8:7 |
| Number of electron in outer energy level | 2 | 7 |
| Number of electrons donated and gained to be stable | 2 | 1 |
| New electron configuration/structure | 2:8: | 2:8: |
| Symbol of cation/anion after bonding | $Mg^{2+}$ | $Cl^-$ |

**Diagram**

c)Lithium oxide($Li_2O$)

Lithium oxide($Li_2O$)is formed when a Lithium atom donate its outer valence electrons to Oxygen atom. Two Lithium atoms are required to donate/lose each one electron and attain stable duplet. Oxygen atom acquires the two electrons and attain stable octet:

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | Li | O |
| Number of protons/electrons | 3 | 8 |
| Electron configuration/structure | 2:1 | 2:6 |
| Number of electron in outer energy level | 1 | 6 |
| Number of electrons donated and gained to be stable | 1 | 2 |
| New electron configuration/structure | 2: | 2:8: |
| Symbol of cation/anion after bonding | $Li^+$ | $O^{2-}$ |

**Diagram**

d)Aluminium(III) oxide($Al_2O_3$)

Aluminium(III) oxide($Al_2O_3$)is formed when a Aluminium atom donate its three outer valence electrons to Oxygen atom. Two Aluminium atoms are required to donate/lose each three electron and attain stable octet. Three Oxygen atoms gain/ acquire the six electrons and attain stable octet:

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | Al | O |
| Number of protons/electrons | 13 | 8 |
| Electron configuration/structure | 2:8:3 | 2:6 |
| Number of electron in outer energy level | 3 | 6 |
| Number of electrons donated and gained to be stable | 3 | 2 |
| New electron configuration/structure | 2:8: | 2:8: |
| Symbol of cation/anion after bonding | $Al^{3+}$ | $O^{2-}$ |

**Diagram**

e)Calcium oxide(CaO)

Calcium oxide(CaO)is formed when a Calcium atom donate its two outer valence electrons to Oxygen atom. Both attain stable octet:

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | Ca | O |

| | | |
|---|---|---|
| Number of protons/electrons | 20 | 8 |
| Electron configuration/structure | 2:8:8:2 | 2:6 |
| Number of electron in outer energy level | 2 | 6 |
| Number of electrons donated and gained to be stable | 2 | 2 |
| New electron configuration/structure | 2:8:8: | 2:8: |
| Symbol of cation/anion after bonding | $Ca^{2+}$ | $O^{2-}$ |

**Diagram**

Some compounds can be formed from ionic/electrovalent, covalent and dative/coordinate bonding within their atoms/molecules:

a)Formation of ammonium chloride:

Ammonium chloride is formed from the reaction of ammonia gas and hydrogen chloride gas. Both ammonia and hydrogen chloride gas are formed from covalent bonding. During the reaction of ammonia and hydrogen chloride gas to form Ammonium chloride;

-ammonia forms a dative/coordinate bond with electron deficient $H^+$ ion from Hydrogen chloride to form ammonium ion($NH_4^+$)ion.

-the chloride ion $Cl^-$ and ammonium ion($NH_4^+$)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

b) Dissolution/dissolving of hydrogen chloride:

Hydrogen chloride is formed when hydrogen and chlorine atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When hydrogen chloride gas is dissolved in water;

-water molecules forms a dative/coordinate bond with electron deficient $H^+$ ion from Hydrogen chloride to form hydroxonium ion($H_3O^+$)ion.

-the chloride ion $Cl^-$ and hydroxonium ion($H_3O^+$)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

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c)Dissolution/dissolving of ammonia gas:

Ammmonia gas is formed when hydrogen and Nitrogen atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When Ammonia gas is dissolved in water;

-ammonia forms a dative/coordinate bond with electron deficient $H^+$ ion from a water molecule to form ammonium ion($NH_4^+$)ion.

-the hydroxide ion $OH^-$ and ammonium ion($NH_4^+$)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

(iii)**METALLIC BOND**

A metallic bond is formed when metallic atoms **delocalize** their outer electrons inorder to be stable.

Metals delocalize their outer electrons to form positively charged cation .

The electrostatic attraction force between the metallic cation and the negatively charged electrons constitute the metallic bond.

The more delocalized electrons the stonger the metallic bond.

**<u>Illustration of ionic /electrovalent bond</u>**

a) Sodium (Na) is made of one valence electron. The electron is donated to form $Na^+$ ion. The electron is delocalized /free within many sodium ions.

| | | | |
|---|---|---|---|
| Symbol of atoms/elements taking part in bonding | Na | Na | Na |
| Number of protons/electrons | 11 | 11 | 11 |
| Electron configuration/structure | 2:8:1 | 2:8:1 | 2:8:1 |
| Number of electron in outer energy level | 1 | 1 | 1 |
| Number of electrons delocalized/free within | 1 | 1 | 1 |
| New electron configuration/structure | 2:8: | 2:8: | 2:8: |
| Symbol of cation after metallic bonding | $Na^+$ | $Na^+$ | $Na^+$ |

**Diagram**

(three)Metallic cations **attract**

(three) free/delocalized electrons

b) Aluminium (Al) is made of three valence electron. The three electrons are donated to form $Al^{3+}$ ion. The electrons are delocalized /free within many aluminium ions.

| | | | |
|---|---|---|---|
| Symbol of atoms/elements taking part in bonding | Al | Al | Al |
| Number of protons/electrons | 13 | 13 | 13 |
| Electron configuration/structure | 2:8:3 | 2:8:3 | 2:8:3 |
| Number of electron in outer energy level | 3 | 3 | 3 |
| Number of electrons delocalized/free within | 3 | 3 | 3 |
| New electron configuration/structure | 2:8: | 2:8: | 2:8: |
| Symbol of cation after metallic bonding | $Al^{3+}$ | $Al^{3+}$ | $Al^{3+}$ |

**Diagram**

(three)Metallic cations **attract**

(nine) free/delocalized electrons

c)Calcium (Ca) is made of two valence electron.The two electrons are donated to form $Ca^{2+}$ ion.The electrons are delocalized /free within many Calcium ions.

| | | | |
|---|---|---|---|
| Symbol of atoms/elements taking part in bonding | Ca | Ca | Ca |
| Number of protons/electrons | 20 | 20 | 20 |
| Electron configuration/structure | 2:8:8:2 | 2:8:8:2 | 2:8:8:2 |

| | | | |
|---|---|---|---|
| Number of electron in outer energy level | 2 | 2 | 2 |
| Number of electrons delocalized/free within | 2 | 2 | 2 |
| New electron configuration/structure | 2:8:8: | 2:8:8: | 2:8:8: |
| Symbol of cation after metallic bonding | $Ca^{2+}$ | $Ca^{2+}$ | $Ca^{2+}$ |

**Diagram**

(three)Metallic cations **attract**

(six) free/delocalized electrons

d) Magnesium (Mg) is made of two valence electron. The two electrons are donated to form $Mg^{2+}$ion.The electrons are delocalized /free within many Magnesium ions.

| | | |
|---|---|---|
| Symbol of atoms/elements taking part in bonding | Mg | Mg |
| Number of protons/electrons | 12 | 12 |
| Electron configuration/structure | 2:8:2 | 2:8:2 |
| Number of electron in outer energy level | 2 | 2 |
| Number of electrons delocalized/free within | 2 | 2 |
| New electron configuration/structure | 2:8: | 2:8: |
| Symbol of cation after metallic bonding | $Mg^{2+}$ | $Mg^{2+}$ |

**Diagram**

(two)Metallic cations **attract**

(four) free/delocalized electrons

e)Lithium (Li) is made of one valence electron.The electron is donated to form $Li^+$ ion.The electron is delocalized /free within many Lithium ions.ie;

| | | | | |
|---|---|---|---|---|
| Symbol of atoms/elements taking part in bonding | Li | Li | Li | Li |
| Number of protons/electrons | 3 | 3 | 3 | 3 |
| Electron configuration/structure | 2:1 | 2:1 | 2:1 | 2:1 |
| Number of electron in outer energy level | 1 | 1 | 1 | 1 |
| Number of electrons delocalized/free within 1 | 1 | 1 | 1 | |
| New electron configuration/structure | 2:1: | 2:1: | 2:1: | 2:1: |
| Symbol of cation after metallic bonding | $Li^+$ | $Li^+$ | $Li^+$ | $Li^+$ |

**Diagram**

(four)Metallic cations **attract**

(four) free/delocalized electrons

## B.CHEMICAL STRUCTURE

Chemical structure is the pattern/arrangement of atoms **after** they have bonded. There are two main types of chemical structures:

**(i)simple molecular structure**

**(ii) giant structures**

(i)Simple molecular structure

Simple molecular structure is the pattern formed after atoms of non-metals have **covalently** bonded to form simple molecules.

Molecules are made of atoms joined together by weak intermolecular forces called **Van-der-waals forces.** The Van-der-waals forces hold the **molecules** together while the covalent bonds hold the **atoms** in the molecule.

**Illustration of simple molecular structure**

a)Hydrogen molecule($H_2$)

Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Illustration of simple molecular structure**

a)Hydrogen molecule($H_2$)

Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces

b)Oxygen molecule($O_2$)

Oxygen gas is made up of strong covalent bonds/**intra**molecular forces between each Oxygen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Strong** intramolecular forces/covalent bond

O=O:::: O=O:::: O=O:::: O=O
: :    : :    : :   : :    : :   : :
O=O:::: O=O:::: O=O:::: O=O

**weak** intermolecular forces/van-der-waals forces

c)Iodine molecule($I_2$)

Iodine solid crystals are made up of strong covalent bonds/**intra**molecular forces between each iodine atom making the molecule.Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Strong** intramolecular forces/covalent bond

I--- I:::: I --- I:::: I --- I:::: I --- I
: :    : :    : :   : :    : :   : :   : :
I --- I:::: I --- I:::: I --- I:::: I --- I

**weak** intermolecular forces/van-der-waals forces

d)Carbon(IV) oxide molecule($CO_2$)

Carbon(IV) oxide gas molecule is made up of strong covalent bonds/**intra**molecular forces between each Carbon and oxygen atoms making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

**Strong** intramolecular forces/covalent bond

O=C=O:::: O=C=O:::: O=C=O
: : : : : :
O=C=O:::: O=C=O:::: O=C=O

**weak** intermolecular forces/van-der-waals forces

The following are the main characteristic properties of simple molecular structured compounds:

a)**State**

Most simple molecular substances are gases, liquid or liquids or solid that sublimes or has low boiling/melting points at room temperature ($25^{o}C$) and pressure (atmospheric pressure).

Examples of simple molecular substances include:

-all gases eg Hydrogen, oxygen, nitrogen, carbon (IV) oxide,

**-**Petroleum fractions eg Petrol, paraffin, diesel, wax,

-Solid non-metals eg Sulphur, Iodine

-Water

b) **Low melting/boiling points**

Melting is the process of **weakening** the intermolecular/ van-der-waal forces/ of attraction between the molecules that holding the substance/compound.

Note;

(i)Melting and boiling does not involve weakening/breaking the strong intramolecular force/covalent bonds holding the atoms **in** the molecule.

(ii) Melting and boiling points increase with increase in atomic radius/size of the atoms **making** the molecule as the intermolecular forces / van-der-waal forces of attraction between the molecules increase. e.g.

Iodine has a higher melting/boiling point than chlorine because it has a higher /bigger atomic radius/size than chlorine, making the molecule to have stronger intermolecular force/ van-der-waal forces of attraction between the molecules than chlorine. Iodine is hence a solid and chlorine is a gas.

**(c)Insoluble in water/soluble in organic solvents**

Polar substances dissolve in polar solvents. Water is a polar solvent .Molecular substances do not thus dissolve in water because they are non-polar. They dissolve in non-polar solvents like methylbenzene, benzene, tetrachloromethane or propanone.

**d)Poor conductors of heat and electricity**

Substances with free mobile **ions** or free mobile/delocalized **electrons** conduct electricity. Molecular substances are poor conductors of heat/electricity because their molecules have no free mobile ions/electrons. This makes them very good **insulators**.

<u>Hydrogen bonds</u>

A hydrogen bond is an intermolecular force of attraction in which a very electronegative atom attracts hydrogen atom of another molecule.

The most electronegative elements are Fluorine, Oxygen and Nitrogen .Molecular compounds made up of these elements usually have hydrogen bonds.

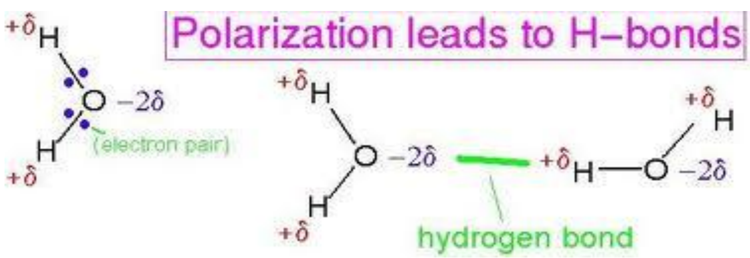
Hydrogen bonds are **stronger** than van-der-waals forces but **weaker** than covalent bonds. Molecular compounds with hydrogen bonds thus have higher melting/boiling points than those with van-der-waals forces.

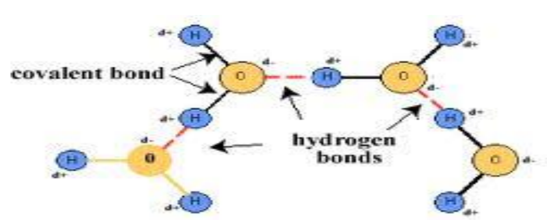
## Illustration of Hydrogen bonding

### a)Water molecule

During formation of covalent bond, the oxygen atom attract/pull the shared electrons more to itself than Hydrogen creating partial negative charges($\delta^-$)in Oxygen and partial positive charges($\delta^+$)in Hydrogen.

Two molecules **attract** each other at the partial charges through Hydrogen bonding.




The hydrogen bonding in water makes it;

(i)a **liquid** with higher boiling and melting point than simple molecular substances with higher molecular mass. e.g. Hydrogen sulphide as in the table below;

Influence of H-bond in water ($H_2O$) in comparison to $H_2S$

| **Substance** | **Water/ $H_2O$** | **Hydrogen sulphide/ $H_2S$** |
|---|---|---|
| Relative molecular mass | 18 | 34 |
| Melting point($^{o}C$) | 0 | -85 |
| Boiling point($^{o}C$) | 100 | -60 |

(ii)have higher **volume** in solid (ice) than liquid (water) and thus ice is less dense than water. Ice therefore floats above liquid water.

**<u>b)Ethanol molecule</u>**

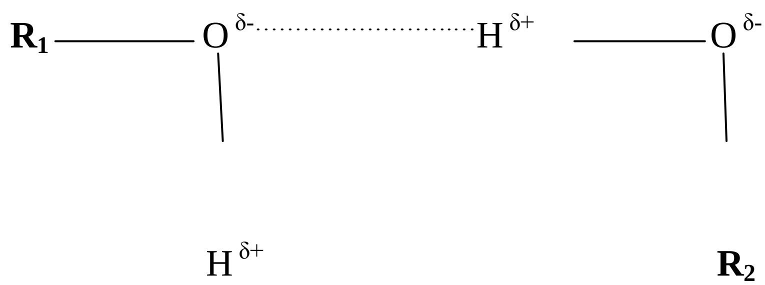
Like in water, the oxygen atom attracts/pulls the shared electrons in the covalent bond more to itself than Hydrogen.

This creates a partial negative charge ($^{\delta-}$) on oxygen and partial positive charge($^{\delta+}$) on hydrogen.

Two ethanol molecules attract each other at the partial charges through Hydrogen bonding forming a **dimme**r.

A dimmer is a molecule formed when two molecules join together as below:

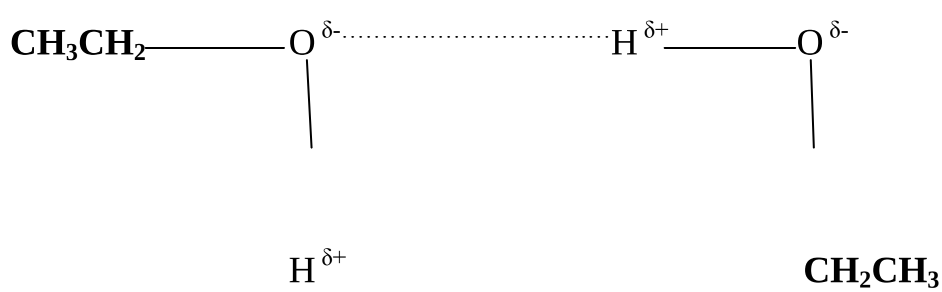
Hydrogen bonds covalent bonds




$R_1$ and $R_2$ are extensions of the molecule.

For ethanol it is made up of $CH_3CH_2$ – to make the structure:

Hydrogen bonds covalent bonds

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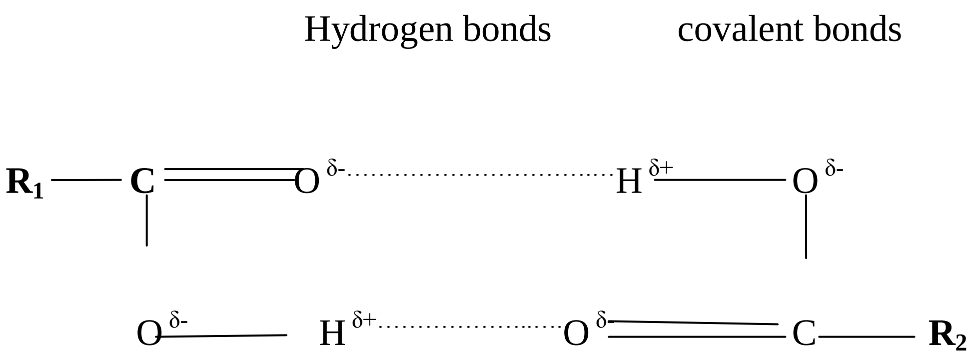
**b)Ethanoic acid molecule**

Like in water and ethanol above, the oxygen atom attracts/pulls the shared electrons in the covalent bond in ethanoic acid more to itself than Hydrogen.

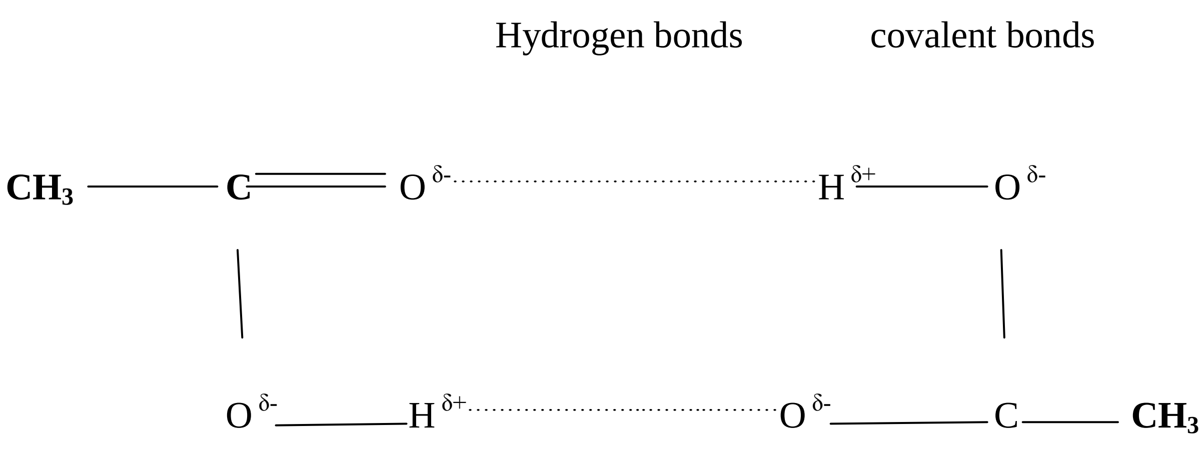
This creates a partial negative charge ($^{\delta-}$)on oxygen and partial positive charge($^{\delta+}$) on hydrogen.

Two ethanoic acid molecules attract each other at the partial charges through Hydrogen-bonding forming a **dimer.**




$R_1$ and $_2$ are extensions of the molecule.

For ethanoic acid the extension is made up of $CH_3$ – to make the structure;



Ethanoic acid like ethanol exists as a dimer.

Ethanoic acid has a **higher** melting/boiling point than ethanol .This is because ethanoic acid has **two/more** hydrogen bond than ethanol.

**d) Proteins and sugars** in living things also have multiple/**complex** hydrogen bonds in their structures.

(ii) Giant structure

This is the pattern formed after substances /atoms /ions bond to form a long chain network.

Giant structures therefore extend in all directions to form a pattern that continues **repeating** itself.

There are **three** main giant structures.

a) giant covalent/atomic structure
b)giant ionic structure

c)giant metallic structure

a) giant covalent/atomic structure

Giant covalent/atomic structure is the pattern formed after atoms have covalently bonded to form long chain pattern consisting of indefinite number of atoms covalently bonded together.

The strong covalent bonds hold all the atoms together to form a very well packed structure. Examples of substances with giant covalent/atomic structure include:

(i) carbon-**diamond**

(ii) carbon-**graphite**

(iii)silicon

(iv) silicon(IV) oxide/sand

Carbon-graphite and carbon-diamond are **allotropes** of carbon.

Allotropy is the existence of an element in more than one stable physical **form** at the same temperature and pressure.

Allotropes are atoms of the same element existing in more than one stable physical form at the same temperature and pressure.

Other elements that exhibit/show allotropy include;

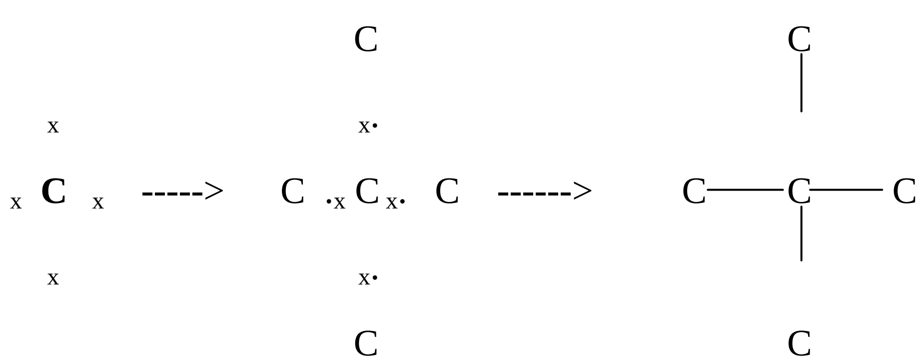
-Sulphur as **monoclinic** sulphur and **rhombic** sulphur

-Phosphorus as **white** phosphorus and **red** phosphorus

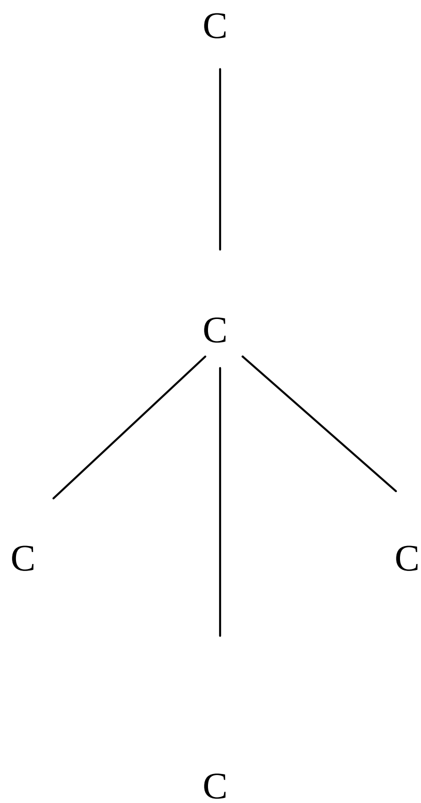
## The structure of carbon-diamond

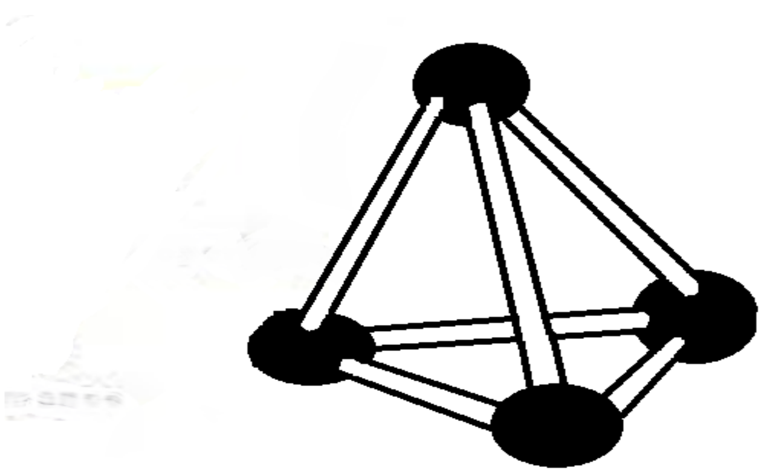
Carbon has four valence electrons. The four valence electrons are used to form covalent bonds.

During the formation of diamond, one carbon atom covalently bond with four other carbon atoms.



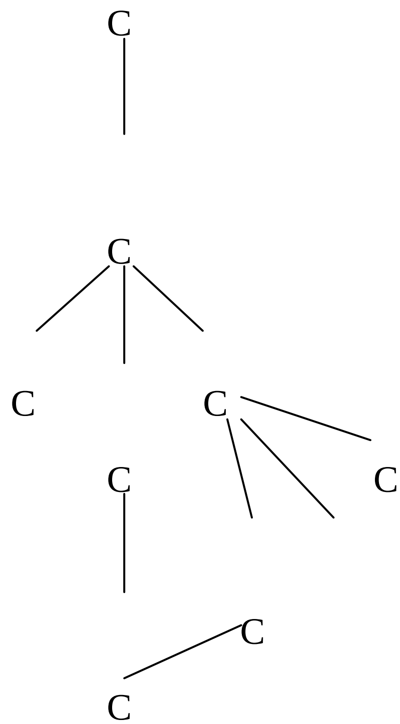
After the bonding, the atoms rearrange to form a regular **tetrahedral** in which **one** carbon is in the **centre** while **four** are at the **apex**/corners.

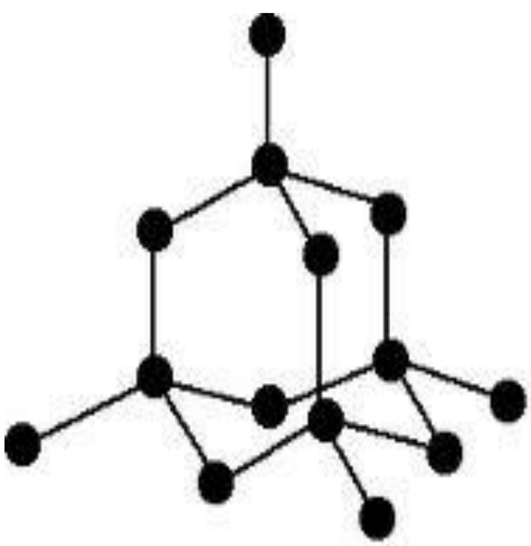


This pattern repeats itself to form a long chain number of atoms covalently bonded together indefinitely. The pattern is therefore called **giant tetrahedral structure.**

It extends in all directions where one atom of carbon is always a centre of four others at the apex/corner of a regular tetrahedral.



The giant tetrahedral structure of carbon-diamond is very well/closely packed and joined/bonded together by strong covalent bond.

This makes carbon-diamond to have the following properties:

**a) High melting/boiling point.**

The giant tetrahedral structure is very well packed and joined together by strong covalent bonds.

This requires a lot of energy/heat to weaken for the element to melt and break for the element to boil.

**b) High density.**

Carbon diamond is the **hardest** known **natural** substance.

This is because the giant tetrahedral structure is a very well packed pattern/structure and joined together by strong covalent bonds.

This makes Carbon diamond be used to make **drill** for drilling boreholes/oil wells.

The giant tetrahedral structure of carbon diamond is a very closely packed pattern /structure such that heat transfer by **conduction** is possible. This makes carbon diamond a **good thermal** conductor.

**c) Poor conductor of electricity.**

Carbon-diamond has **no free/delocalized electrons** within its structure and thus do not conduct electricity.

**d) Insoluble in water.**

Carbon-diamond is insoluble in water because it is non-polar and do not bond with water molecules.

**e) Is abrasive/Rough.**

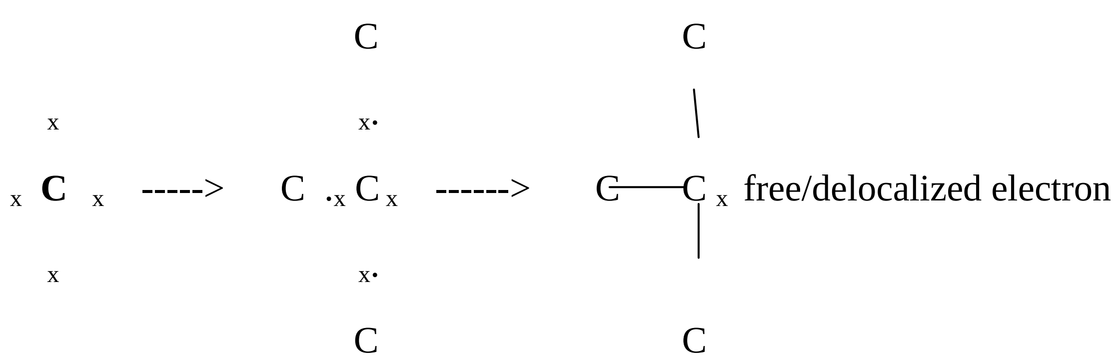
The edges of the closely well packed pattern/structure of Carbon-diamond make its surface rough/abrasive and thus able to smoothen /cut metals and glass.

**f) Have characteristic luster.**

Carbon-diamond has a high **optical dispersion** and thus able to disperse light to different colours .This makes Carbon-diamond one of the most popular **gemstone** for making **jewellery**.

## <u>The structure of carbon-graphite</u>

During the formation of graphite, one carbon atom covalently bond with **three** other carbon atoms leaving **one** free/delocalized electron.



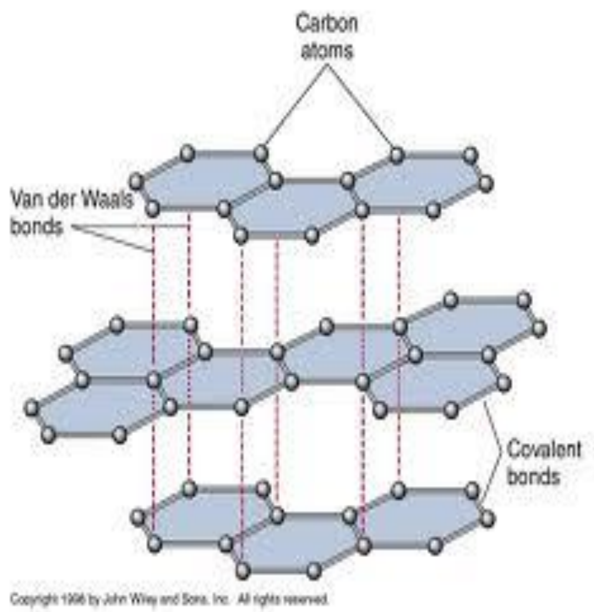

After the bonding, the atoms rearrange and join together to form a regular **hexagon** in which **six** carbon atoms are at the **apex/**corners**.**

The regular hexagon is joined to another in layers on the **same surface** by van-der-waals forces.

Each layer extends to form a plane in all directions.

The fourth valence electron that does not form covalent bonding is free/mobile /delocalized within the layers.

This structure/pattern is called **giant hexagonal planar** structure.




The giant hexagonal planar structure of carbon-graphite is closely packed and joined/bonded together by strong covalent bonds. This makes carbon-graphite to have the following properties:

**a) High melting/boiling point.**

The giant hexagonal planar structure of carbon-graphite is well packed and joined together by strong covalent bonds.

This requires a lot of energy/heat to **weaken** for the element to melt and **break** for the element to boil.

**b) Good conductor of electricity.**

Carbon-graphite has **free/delocalized 4th valence electrons** within its structure and thus conducts electricity.

**c) Insoluble in water.**

Carbon-graphite is insoluble in water because it is **non-polar** and do not bond with water molecules.

**d) Soft.**

Layers of giant hexagonal planar structure of carbon graphite are held together by van-der-waals forces.

The van-der-waals forces easily break when pressed and reform back on releasing/reducing pressure/force thus making graphite soft.

**e) Smooth and slippery.**

When pressed at an **angle** the van-der-waals forces easily break and slide over each other making graphite soft and slippery.

It is thus used as a **dry** lubricant instead of oil.

**f)Some uses of carbon-graphite.**

1. As a dry lubricant- carbon graphite is smooth and slippery and thus better lubricant than oil.Oil **heat up** when reducing friction.

2. Making Lead-pencils- When pressed at an **angle** on paper the van-der-waals forces easily break and slide smoothly over contrasting background producing its characteristic black background.

3. As moderator in nuclear reactors to reduce the rate of decay/disintegration of radioactive nuclides/atoms/isotopes.

4. As electrode in dry/wet cells/battery- carbon graphite is inert and good conductor of electricity. Current is thus able to move from one electrode/terminal to the other in dry and wet cells/batteries. Carbon graphite is also very **cheap**.

b) giant ionic structure

Giant ionic structure is the pattern formed after **ions** have bonded through ionic/electrovalent bonding to form a long chain consisting of **indefinite** number of ions.

The strong ionic/electrovalent bond holds all the **cations** and **anions** together to form a very well packed structure.

Substances with giant ionic structure are mainly crystals of salts e.g. sodium chloride, Magnesium chloride, Sodium iodide, Potassium chloride, copper (II) sulphate(VI).

## The structure of sodium chloride

Sodium chloride is made up of sodium ($Na^+$) and chloride ($Cl^-$)ions.

Sodium ($Na^+$) ion is formed when a sodium atom donate /loose/donate an electron. Chloride ($Cl^-$) ion is formed when a chlorine atom gain /acquire an extra electron from sodium atom.

Many $Na^+$ and $Cl^-$ ions then rearrange such that **one** $Na^+$ ion is surrounded by **six** $Cl^-$ ions and one $Cl^-$ ion is surrounded by six $Na^+$ ions.

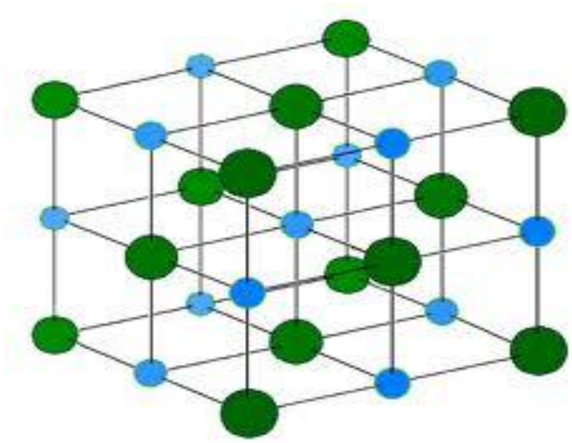
The pattern formed is a **giant cubic structure** where $Cl^-$ ion is sand witched between $Na^+$ ions and the same to $Na^+$ ions.

This pattern forms a **crystal**.

A crystal is a solid form of a substance in which particles are arranged in a definite pattern regularly repeated in three dimensions.

**<u>The structure of sodium chloride</u>**

The giant cubic structure/crystal of sodium chloride is as below;



The giant cubic structure/crystal of sodium chloride is very well packed and joined by strong ionic/electrovalent bonds. This makes sodium chloride and many ionic compounds to have the following properties:

**a) Have high melting /boiling points.**

The giant cubic lattice structure of sodium chloride is very **closely** packed into a crystal that requires a lot of energy/heat to **weaken** and melt/boil. This applies to all crystalline ionic compounds.

**b) Are good conductors of electricity in molten and aqueous state** but **poor conductor of electricity in solid.**

Ionic compounds have **fused** ions in solid crystalline state.

On heating and dissolving in water, the crystal is broken into free mobile ions ($Na^+$ and $Cl^-$ ions).

The free mobile **ions** are responsible for **conduct**ing electricity in ionic compounds in molten and aqueous states.

**c)Soluble in water**

Ionic compounds are **polar** and dissolve in **polar** water molecules.

On dissolving, the crystal breaks to free the fused ions which are then surrounded by water molecules.

b) giant metallic structure

This is the pattern formed after metallic atoms have bonded through metallic bond.

The pattern formed is one where the metallic cations rearrange to form a cubic structure.

The cubic structure is bound together by the free delocalized electrons that move freely within.

The **more** delocalized electrons, the **stronger** the metallic bond.

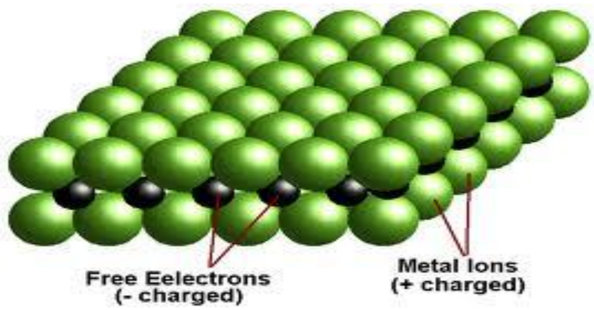
**The structure of sodium and aluminium.**

Sodium has one valence electrons.

Aluminium has three valence electrons.

After delocalizing the valence electrons ,the metal cations ($Na^+$ and $Al^{3+}$) rearrange to the apex /corners of a regular cube that extend in all directions.

The delocalized electrons remain free and mobile as shown below:




The giant cubic structure makes metals to have the following properties:

**a) Have high melting/boiling point**

The giant cubic structure is very well packed and joined/bonded together by the free delocalized electrons.

The more delocalized electrons the higher the melting/boiling point.

The larger/bigger the metallic cation ,the weaker the packing of the cations and thus the lower the melting/boiling point. e.g.

(i) Sodium and potassium have both one valence delocalized electron.

Atomic radius of potassium is larger/bigger than that of sodium and hence less well packed in its metallic structure.

Sodium has therefore a higher melting/boiling point than potassium.

(ii) Sodium has one delocalized electron.

Aluminium has three delocalized electrons.

Atomic radius of sodium is larger/bigger than that of aluminium and hence less well packed in its metallic structure.

Aluminium has therefore a higher melting/boiling point than sodium because of the smaller well packed metallic ($Al^{3+}$)ions and bonded/joined by more/three delocalized electrons.

The table below shows the **comparative** melting/boiling points of some metals:

| **Metal** | **Electronic structure** | **Atomic radius(nM)** | **Melting point($^oC$)** | **Boiling point($^oC$)** |
|---|---|---|---|---|
| Sodium | 2:8:1 | 0.155 | 98 | 890 |
| Potassium | 2:8:8:1 | 0.203 | 64 | 774 |
| Magnesium | 2:8:2 | 0.136 | 651 | 1110 |
| Aluminium | 2:8:3 | 0.125 | 1083 | 2382 |

**b) Good electrical and thermal conductor/electricity.**

All metals are good conductors of heat and electricity including Mercury which is a **liquid**.

The mobile delocalized electrons are free within the giant metallic structure to move from one end to the other transmitting heat/electric current.

The more delocalized electrons the better the thermal/electrical conductivity.

High temperatures/heating lowers the thermal/electrical conductivity of metals because the delocalized electrons vibrate and move randomly hindering transfer of heat

From the table above:

**Compare the electrical conductivity of;**

(i)**Magnesium and sodium**

Magnesium is a better conductor than sodium.

Magnesium has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

(ii)**Potassium and sodium**

Potassium is a better conductor than sodium.

Potassium has bigger/larger atomic radius than sodium. The delocalized electrons are less attracted to the nucleus of the atom and thus more free /mobile and thus better the electrical conductor.

**c) Insoluble in water**

All metals are insoluble in water because they are non polar and thus do not bond with water.

Metals higher in the reactivity/electrochemical series like; Potassium, sodium, Lithium and Calcium **reacts** with **cold** water producing hydrogen gas and forming an **alkaline** solution of their **hydroxides**.ie

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

**Heavy** metal like Magnesium, Aluminium, Iron, Zinc and Lead **react** with **steam/water vapour** to produce hydrogen gas and form the corresponding **oxide**.

| | | | | | | |
|---|---|---|---|---|---|---|
| $Mg(s)$ | + | $H_2O(g)$ | -> | $MgO(s)$ | + | $H_2(g)$ |
| $Fe(s)$ | + | $H_2O(g)$ | -> | $FeO(s)$ | + | $H_2(g)$ |
| $Zn(s)$ | + | $H_2O(g)$ | -> | $ZnO(s)$ | + | $H_2(g)$ |
| $Pb(s)$ | + | $H_2O(g)$ | -> | $PbO(s)$ | + | $H_2(g)$ |
| $2Al(s)$ | + | $3H_2O(g)$ | -> | $Al_2O_3(s)$ | + | $3H_2(g)$ |

Metals **lower** in the reactivity/electrochemical series than hydrogen like; copper, Mercury, Gold Silver and Platinum **do not** react with water/vapour.

**d) Shiny metallic-lustre**

All metals have a shiny **grey** metallic luster <u>except</u> copper which is **brown.**

When exposed to sunlight, the delocalized electrons gain energy, they vibrate on the metal surface scattering light to appear shiny.

With time, most metals corrode and are covered by a layer of the metal oxide.

The delocalized electrons are unable to gain and scatter light and the metal surface tarnishes/become dull.

**e) Ductile and malleable**

All metals are malleable (can be made into thin **sheet**) and ductile (can be made into **wire**.

When beaten/hit/pressed **lengthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a sheet.

When beaten/hit/pressed **lengthwise and bredthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a wire/thin strip.

**f) Have high tensile strength**

Metals are not brittle. The free delocalized electrons bind the metal together when it is bent /coiled at any angle.

The meta thus withstand stress/coiling

**g) Form alloys**

An alloy is a uniform mixture of two or more metals.

Some metals have spaces between their metallic cations which can be occupied by another metal cation with smaller atomic radius.

Common alloys include:

Brass(Zinc and Copper alloy)

Bronze(Copper and Tin alloy)

German silver

**Summary of Bonding and structure**

| | **Simple molecular structure** | **Giant covalent /atomic structure** | **Giant ionic structure** | **Giant metallic structure** |
|---|---|---|---|---|
| **(i)Examples** | $I_2$,$S_8$,HCl,$O_2$,$CH_4$ | Graphite,diamond Si,$SiO_2$ | NaCl, KCl, CaO,$CuSO_4$ | Na,Fe,Cr,Hg,K |
| **Constituent particles making structure** | molecules | Atoms (of non-metals) | Ions (cation and anions) | Atoms (of metals) |
| **Type of substance** | Non-metal element/non-metal molecule/non-metal compound(electronegative elements) | Group IV non-metals and some of their oxides | Metal-non metal compounds(compounds of electropositive and electronegative compounds) | Metallic elements (with low electonegativity and high electropositivity) |
| **Bonding in solid state** | -Strong covalent bonds hold atoms together within separate molecules (intramolecular forces)<br>-Weak van-der-waals forces hold separate molecules together (intermolecular forces) | Atoms are linked through the whole structure by very strong covalent bonds. | Electrostatic attraction of cations and anions link the whole structure through strong ionic bond. | Electrostatic attraction of outer mobile electrons for positive nuclei binds atoms together though metallic bond |
| **Properties (i) Volatility** | -Highly volatile with low melting/boiling point | -Non volatile with very high melting/boiling | -Non volatile with very high melting/boiling | -Non volatile with very high |

| | | | | |
|---|---|---|---|---|
| | -Low latent heat of fusion/vaporization | points<br>-Low latent heat of fusion / vaporization | points<br>-Low latent heat of fusion / vaporization | melting/boiling points<br>-Low latent heat of fusion / vaporization |
| **(ii) State at room temperature /pressure** | Usually gases,volatile liquids or solids that sublimes | solids | solids | Solids except Mercury(liquid) |
| **(iii) Hardness** | Soft and brittle(low tensile strength) | Hard and brittle(low tensile strength) | Hard and brittle(low tensile strength) | Hard, malleable, ductile and have high tensile strength |
| **(iv) Thermal /electrical conductivity** | Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but some dissolve and react to form electrolytes e.g. Hydrogen chloride and ammonia gases. | Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but<br>-Carbon-graphite is a good electrical conductor while<br>-Carbon-diamond is a good thermal conductor. | Poor thermal and electrical conductor when solid. Good thermal and electrical conductor in liquid/molten and aqueous states when the ions are not fused | Good thermal and electrical conductor in solid and liquid/molten states due to the free mobile /delocalized electrons |
| **(v) Solubility** | Insoluble in polar solvents e.g. Water Soluble in non-polar solvents e.g. tetrachloromethane, benzene, methylbenzene | Insoluble in all solvents | Soluble in polar solvents e.g. Water Insoluble in non-polar solvents e.g. tetrachloromethane, benzene, methylbenzene | Insoluble in polar/non-polar colvents.<br>-Some react with polar solvents<br>-Some metal dissolve in other metals to form alloys e.g. Brass is formed when Zinc dissolve in copper. |

## C. PERIODICITY OF BONDING AND STRUCTURE

The periodic table does not classify elements as metals and non-metals. The table arranges

them in terms of atomic numbers.

However, based on structure and bonding of the elements in the periodic table;

(i)-the top right hand corner of about twenty elements are non-metals

(ii)-left of each non-metal is an element which shows characteristics of both metal and non-metal.

These elements are called **semi-metals/metalloids**. They include Boron, silicon, Germanium, Arsenic, and Terullium

(iii)-all other elements in the periodic table are metal.

(iv)-Hydrogen is a non-metal with metallic characteristic/property of donating/losing outer electron to form cation/$H^+$ ion.

(v) –bromine is the only known natural liquid non-metal element at room temperature and pressure.

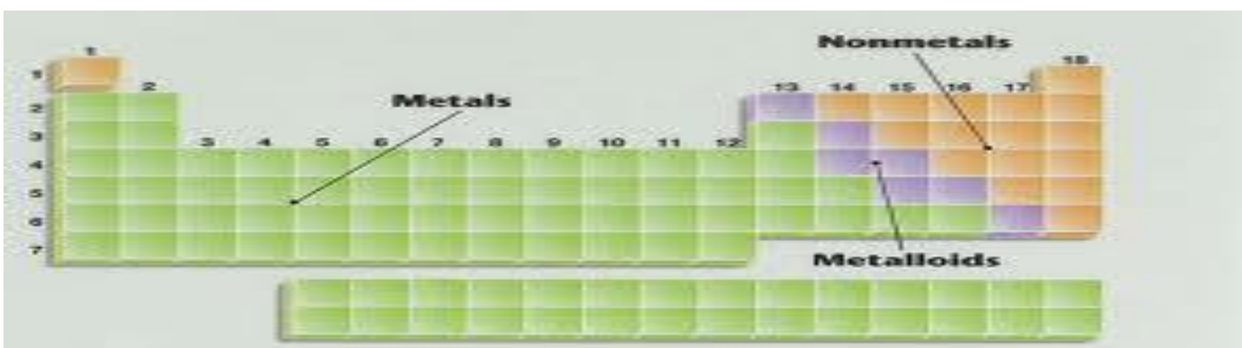
(vi) –mercury is only known natural liquid metal element at room temperature and pressure.

(vii) Carbon-graphite is a semi metals/metalloids. Carbon-diamond is a pure non-metal yet both are allotropes of carbon (same element)

**a) Sketch of the periodic table showing metals ,metalloid and non-metals**

Metals **Metalloids** *Non-metals*

| | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|
| *H* | | | | | | | | | *He* |
| Li | Be | | **B** | **C** | *N* | *O* | *F* | | *Ne* |
| Na | Mg | | Al | **Si** | *P* | *S* | *Cl* | | *Ar* |
| K | Ca | Transition metals | Ga | Ge | **As** | *Se* | *Br* | | *Kr* |
| Rb | Sr | | In | Sn | Sb | **Te** | *I* | | *Xe* |
| Cs | Ba | | Tl | Pb | Bi | Po | **At** | | *Rn* |
| Fr | Ra | | | | | | | | |




**b)Periodicity in the physical properties of elements across period 2 and 3**

Study table I and II below:

Table I(period 2)

| **Property** | **Li** | **Be** | **B** | **C** | **N** | **O** | **F** | **Ne** |
|---|---|---|---|---|---|---|---|---|
| **Melting point($^oC$)** | 180 | 1280 | 2030 | 3700 (graphite) 3550 (diamond) | -210 | -219 | -220 | -250 |
| **Boiling point($^oC$)** | 1330 | 2480 | 3930 | Graphite sublimes 4830 (diamond) | -200 | -180 | -190 | -245 |
| **Density at room temperature ($gcm^{-3}$)** | 0.50 | 1.85 | 2.55 | 2.25 (graphite) 3.53 (diamond) | 0.81 | 0.14 | 0.11 | 0.021 |
| **Type of element** | Metal | Metal | Metal | Metalloid | Non-metal | Non-metal | Non-metal | Non-metal |
| **Chemical structure** | Giant metallic | Giant metallic | Giant atomic/ covalent | Giant atomic/ covalent | Simple molecula or molecule/ $N_2$ | Simple molecula or molecules /$O_2$ | Simple molecula or molecule/$F_2$ | Simple molecula or molecule/Ne |
| **State at room temperature** | Solid | Solid | Solid | Solid | gas | gas | gas | gas |
| **Electron structure** | 2:1 | 2:2 | 2:3 | 2:4 | 2:5 | 2:6 | 2:7 | 2:8 |
| **Valency** | 1 | 2 | 3 | 4 | 3 | 2 | 1 | - |
| **Formular of ion** | $Li^+$ | $Be^{2+}$ | $B^{3+}$ | - | $N^{3-}$ | $O^{2-}$ | $F^-$ | - |

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|

Table II (period 3)

| **Property** | **Na** | **Mg** | **Al** | **Si** | **P(white)** | **S(Rhombic)** | **Cl** | **Ar** |
|---|---|---|---|---|---|---|---|---|
| **Melting point($^oC$)** | 98 | 650 | 660 | 1410 | 44 | 114 | -101 | -189 |
| **Boiling point($^oC$)** | 890 | 1120 | 2450 | 2680 | 280 | 445 | -34 | -186 |
| **Density at room temperature ($gcm^{-3}$)** | 0.97 | 1.74 | 2.70 | 2.33 (graphite) 3.53 (diamond) | 1.82 | 2.07 | 0.157 | 0.011 |
| **Type of element** | Metal | Metal | Metal | Metalloid | Non-metal | Non-metal | Non-metal | Non-metal |
| **Chemical structure** | Giant metallic | Giant metallic | Giant metallic | Giant atomic/ covalent | Simple molecula or molecule/ $P_4$ | Simple molecula or molecules /$S_8$ | Simple molecula or molecule/$Cl_2$ | Simple molecula or molecule /Ar |
| **State at room temperature** | Solid | Solid | Solid | Solid | Solid | Solid | gas | gas |
| **Electron structure** | 2:8:1 | 2:8:2 | 2:8:3 | 2:8:4 | 2:8:5 | 2:8:6 | 2:8:7 | 2:8:8 |

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| **Valency** | 1 | 2 | 3 | 4 | 3 | 2 | 1 | - |
| **Formular of ion** | $Na^{+}$ | $Mg^{2+}$ | $Al^{3+}$ | - | $P^{3-}$ | $S^{2-}$ | $Cl^{-}$ | - |

From table I and II above:

**1. Explain the trend in atomic radius along /across a period in the periodic table**

Observation

Atomic radius of elements in the same period decrease successively across/along a period from left to right.

Explanation

Across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level.

Increase in nuclear charge increases the effective nuclear attraction on the outer energy level pulling it closer to the nucleus successively across the period .e.g.

(i)From the table 1and 2 above, atomic radius of Sodium (0.157nM) is higher than that of Magnesium(0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus than sodium.

(ii)The rate of decrease in the atomic radius become smaller as the atom become heavier e.g. Atomic radius of Magnesium from sodium falls by(0.157nM- 0.137nM) =**0.02**

Atomic radius of Chlorine from sulphur falls by(0.104nM- 0.099nM) =**0.005**

This is because gaining/adding one more proton to 11 already present cause greater proportional change in nuclear attraction power to magnesium than gaining/adding one more proton to 16 already present in sulphur to chlorine.

(iii)Period 3 elements have more energy levels than Period 2 elements. They have therefore bigger/larger atomic radius/size than corresponding period 2 elements in the same group.

**2.Explain the trend in ionic radius along/across a period in the periodic table**

<u>Observation</u>

Ionic radius of elements in the same period decrease successively across/along a period from left to right for the first three elements then increase drastically then slowly successively decrease.

<u>Explanation</u>

Across/along the period from left to right elements change form electron donors/losers (**reducing** agents) to electron acceptors (**oxidizing** agents).

(i)An atom form stable ion by either gaining/acquiring/ accepting extra electron or donating/losing outer electrons.

(ii)Metals form stable ions by donating/losing **all** the outer energy level electrons and thus **also** the outer energy level .i.e.

-Sodium ion has one less energy level than sodium atom. The ion is formed by sodium atom donating/losing (all) the outer energy level electron and thus also the outer energy level making the ion to have smaller ionic radius than atom.

(iii)Ionic radius therefore decrease across/along the period from Lithium to Boron in period 2 and from Sodium to Aluminium in period 3.This is because the number of electrons donated/lost causes increased effective nuclear attraction on remaining electrons /energy levels.

(iv)Non-metals form stable ion by gaining/acquiring/accepting extra electron in the outer energy level. The extra electron/s increases the repulsion among electrons and reduces the effective nuclear attraction on outer energy level. The outer energy level therefore expand/enlarge/increase in order to accommodate the extra repelled electrons .The more electrons gained/accepted/acquired the more repulsion and the more expansion to accommodate them and hence bigger/larger atomic radius. e.g.

-Nitrogen ion has three electrons more than Nitrogen atom. The outer energy level expand/enlarge/increase to accommodate the extra repelled electrons. Nitrogen atom thus has smaller atomic radius than the ionic radius of nitrogen ion.

(v) Ionic radius decrease from group IV onwards from left to right. This because the number of electrons gained to form ion decrease across/along the period from left to right. e.g. Nitrogen ion has bigger/larger ionic radius than Oxygen.

**3.Explain the trend in melting and boiling point of elements in a period in the periodic table.**

<u>Observation</u>

The melting and boiling point of elements rise up to the elements in Group IV(Carbon/Silicon) along/across the period then continuously falls.

Explanation

Melting/boiling points depend on the packing of the structure making the element and the strength of the bond holding the atoms/molecules together.

Across/along the period (2 and 3) the structure changes from giant metallic, giant atomic/covalent to simple molecular.

(i)For metals, the number of delocalized electrons increases across/along the period and hence stronger metallic bond/structure thus requiring a lot of heat/energy to weaken.

The strength of a metallic bond also depends on the atomic radius/size. The melting /boiling point decrease as the atomic radius/size of metals increase due to decreased packing of larger atoms. e.g.

-The melting /boiling point of Lithium is lower than that of Beryllium because Beryllium has two/more delocalized electrons and hence stronger metallic structure/bond.

- The melting /boiling point of Lithium is higher than that of Sodium because the atomic radius/size Lithium is smaller and hence better packed and hence forms stronger metallic structure/bond.

(ii)Carbon-graphite/carbon-diamond in period 2 and Silicon in period 3 form very well packed giant atomic/covalent structures held together by strong covalent bonds. These elements have therefore very high melting/boiling points.

Both Carbon-graphite/ carbon-diamond have smaller atomic radius/size than Silicon in period 3 and thus higher melting/boiling points due to better/closer packing of smaller atoms in their well packed giant atomic/covalent structures.

(ii)Non-metals from group V along/across the period form simple molecules joined by weak intermolecular /van-der-waals force. The weak intermolecular /van-der-waals force require little energy/heat to weaken leading to low melting/boiling points. The strength of the intermolecular /van-der-waals forces decrease with decrease in atomic radius/ size lowering the melting/boiling points along/across the period (and raising the melting/boiling points down the group).e.g.

-The melting /boiling point of Nitrogen is higher than that of Oxygen. This is because the atomic radius/ size of Nitrogen is higher than that of Oxygen and hence stronger intermolecular /van-der-waals forces between Nitrogen molecules.

-The melting /boiling point of Chlorine is higher than that of Fluorine. This is because the atomic radius/ size of Chlorine is higher than that of Fluorine and hence stronger intermolecular /van-der-waals forces between Chlorine molecules.

(iii)Rhombic sulphur exists as a puckered ring of $S_8$atoms which are well packed. Before melting the ring break and join to very long chains that entangle each other causing the **unusually** high melting/boiling point of Rhombic sulphur.

(iv)Both sulphur and phosphorus exists as allotropes.

Sulphur exists as **Rhombic**-sulphur and **monoclinic**-sulphur. Rhombic-sulphur is the stable form of sulphur at room temperature and pressure.

Phosphorus exists as **white-**phosphorus and **red-**phosphorus.

White-phosphorus is the stable form of Phosphorus at room temperature and pressure.

**4. State and explain the trend in density of elements in a period in the periodic table.**

Observation: Density increase upto the elements in group IV then falls across/along the period successively

Explanation:

Density is the mass per unit volume occupied by matter/particles/atoms/molecules of element.

(i)For metals ,the stronger metallic bond and the more delocalized electrons ensure a very well packed giant metallic structure that occupy less volume and thus higher density.

The more the number of delocalized electrons along/across the period, the higher the density. e.g.

(i)Aluminium has a higher density than sodium. This is because aluminium has more /three delocalized electrons than /one sodium thus forms a very well packed giant metallic structure that occupy less volume per given mass/density.

(ii)Carbon-graphite ,carbon-diamond and silicon in group IV form a well packed giant atomic/covalent structure that is continuously joined by strong covalent bonds hence occupy less volume per given mass/density.

Carbon-graphite form a less well packed giant hexagonal planar structure joined by Van-der-waals forces. Its density ($2.25gcm^{-3}$) is therefore less than that of Carbon-diamond($3.53gcm^{-3}$) and silicon($2.33gcm^{-3}$).Both diamond and silicon have giant tetrahedral structure that is better packed. Carbon-diamond has smaller atomic radius/size than silicon. Its density is thus higher because of better packing and subsequently higher density. Carbon-diamond is the hardest known natural substance by having the highest density.

(iii)For non-metals, the strength of the intermolecular /van-der-waals forces decreases with decrease in atomic radius/size along/across the period. This decreases the mass occupied by given volume of atoms in a molecule from group VI onwards. e.g.

Phosphorus has a higher atomic radius/size than chlorine and Argon and thus stronger intermolecular/van-der-waals forces that ensure a given mass of phosphorus occupy less volume than chlorine and neon.

**5.State and explain the trend in thermal/electrical conductivity of elements in a period in the periodic table.**

Observation:

Increase along/across the period from group I, II, and III then decrease in Group IV to drastically decrease in group V to VIII (O).

Explanation

(i)Metals have free delocalized electrons that are responsible for thermal/electrical conductivity.Thermal/electrical conductivity increase with increase in number of delocalized electrons. The thermal conductivity decrease with increase in temperature/heating.

e.g.

Aluminium with three delocalized electrons from each atom in its metallic structure has the highest electrical /thermal conductivity in period 3.

(ii)Carbon-graphite has also free $4^{th}$ valency electrons that are delocalized within its layers of giant hexagonal planar structure. They are responsible for the electrical conductivity of graphite.

(iii)Silicon and carbon diamond do not conduct electricity but conducts heat. With each atom too close to each other in their very well packed giant tetrahedral structure, heat transfer /radiate between the atoms. The thermal conductivity increase with increase in temperature/heating.

(iv)All other non-metals are poor /non-conductor of heat and electricity. They are made of molecules with no free /mobile delocalized electrons in their structure.

**Periodicity of the oxides of elements along/across period 3**

The table below summarizes some properties of the oxides of elements in period 3 of the periodic table.

| **Formular of oxide/ Property** | **$Na_2O$** | **MgO** | **$Al_2O_3$** | **$SiO_2$** | **$P_2O_5$ $P_4O_6$** | **$SO_2$** $SO_3$ | **$Cl_2O_7$ $Cl_2O$** |
|---|---|---|---|---|---|---|---|
| **Melting point($^oC$)** | 1193 | 3075 | 2045 | 1728 | 563 | -76 | -60 |
| **Boiling point($^oC$)** | 1278 | 3601 | 2980 | 2231 | 301 | -10 | -9 |
| **Bond type** | Ionic | Ionic | Ionic | Covalent | Covalent | Covalent | Covalent |
| **Chemical structure** | Giant ionic structure | Giant ionic structure | Giant ionic structure | Giant atomic/ covalent | Simple molecula or molecule | Simple molecula or molecules | Simple molecula or molecule |
| **State at room temperature** | Solid | Solid | Solid | Solid | Solid | gas | Gas (**$Cl_2O_7$** is a liquid) |
| **Nature of Oxide** | Basic/ alkaline | Basic/ alkaline | Amphotellic oxide | 2:8:4 | 2:8:5 | 2:8:6 | 2:8:7 |

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| **Reaction with water** | React to form NaOH /alkaline solution | React to form $MgOH)_2$ /weakly alkaline solution | Don't react with water. | Don't react with water. | React to form $H_2PO_4$ /weakly acidic solution | -$SO_2$ react to form $H_2SO_3$ . $H_2SO_3$ is quickly oxidized to $H_2SO_4$<br>-$SO_2$ react to form $H_2SO_4$/ strongly acidic | **-$Cl_2O_7$ reacts to form $HClO_4$** /weakly acidic solution |
| **Reaction with dilute acids** | Reacts to form salt and water | Reacts to form salt and water | Reacts to form salt and water | No reaction | No reaction | No reaction | No reaction |

1. All the oxides of elements in period 3 <u>except</u> those of sulphur and chlorine are **solids** at room temperature and pressure.

2. Across/along the period, bonding of the oxides changes from **ionic** in sodium oxide magnesium oxide and aluminium oxide (show both ionic and covalent properties) to **covalent** in the rest of the oxides.

3. Across/along the period, the structure of the oxides changes from giant **ionic** structure in sodium oxide, magnesium oxide and aluminium oxide to giant **atomic**/covalent structure in silicon (IV) oxide. The rest of the oxides form simple **molecules**/**molecular** structure.

4. Sodium oxide and magnesium oxide are **basic /alkaline** in nature. Aluminium oxide is **amphotellic** in nature (shows both acidic and basic characteristics). The rest of the oxides are **acidic** in nature.

5. Ionic compounds/oxides have very **high** melting/boiling points because of the strong **electrostatic attraction** joining the giant ionic crystal lattice.

The melting/boiling points increase from sodium oxide to aluminium oxide as the number of electrons involved in bonding increase, increasing the strength of the ionic bond/structure.

6. Silicon (IV) oxide is made of a well packed giant atomic/covalent structure joined by strong covalent bonds.

This results in a solid with very **high** melting/boiling point.

7.Phosphorus (V) oxide, sulphur(IV) oxide/ sulphur (VI) oxide and dichloride heptoxide exist as simple molecules/molecular structure joined by weak van-der-waals/intermolecular forces.

This results in them existing as **low** melting /boiling point solids/gases.

8. Ionic oxide conducts electricity in molten and aqueous states but not in solid.

In solid state the ions are fused/fixed but on heating to molten state and when dissolved in water, the ions are free / mobile.

Sodium oxide, magnesium oxide and aluminium oxide are therefore good conductors in molten and aqueous states.

9. Covalent bonded oxides do not conduct electricity in solid, molten or in aqueous states.

This is because they do not have free / mobile ion. Phosphorus (V) oxide, sulphur(IV) oxide/ sulphur (VI) oxide and dichloride heptoxide are thus non-conductors/insulators.

10. Silicon (IV) oxide is a poor/weak conductor of heat in solid state. This is because it has very closely packed structure for heat to radiate conduct along its structure.

11. Electopositivity decrease across the period while electronegativity increase across the period. The oxides thus become less ionic and more covalent along/across the period.

12.The steady change from giant ionic structure to giant atomic/ covalent structure then simple molecular structure lead to profound differences in the reaction of the oxides with water,acids and alkalis/bases:

(i) Reaction with water

a) Ionic oxides react with water to form alkaline solutions e.g.;

I.Sodium oxide reacts/dissolves in water forming an alkaline solution of sodium hydroxide.

Chemical equation: $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

II. Magnesium oxide slightly/ slowly reacts/dissolves in water forming an alkaline solution of magnesium hydroxide

Chemical equation: $MgO(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq)$

III. Aluminium oxide does reacts/dissolves in water.

b) Non-metallic oxides are acidic. They react with water to form weakly acidic solutions:

I. Phosphorus (V) oxide readily reacts/dissolves in water forming a weak acidic solution of phosphoric (V) acid.

Chemical equation: $P_4O_{10}$ (s) + $6H_2O$ (l) -> $4H_3PO_4$ (aq)

Chemical equation: $P_2O_5$ (s) + $3H_2O$ (l) -> $2H_3PO_4$ (aq)

II. Sulphur (IV) oxide readily reacts/dissolves in water forming a weak acidic solution of sulphuric (IV) acid.

Chemical equation: $SO_2$ (g) + $H_2O$ (l) -> $H_2SO_3$ (aq)

Sulphur (VI) oxide quickly fumes in water to form concentrated sulphuric (VI) acid which is a strong acid.

Chemical equation: $SO_3$ (g) + $H_2O$ (l) -> $H_2SO_4$ (aq)

III. Dichlorine oxide reacts with water to form weak acidic solution of chloric(I) acid/hypochlorous acid.

Chemical equation: $Cl_2O$ (g) + $H_2O$ (l) -> $2HClO$ (aq)

IV. Dichlorine heptoxide reacts with water to form weak acidic solution of chloric(VII) acid.

Chemical equation: $Cl_2O_7$ (l) + $H_2O$ (l) -> $2HClO_4$ (aq)

c) Silicon (IV) oxide **does not** react with water.

It reacts with hot concentrated alkalis forming silicate (IV) salts. e.g.

Silicon (IV) oxide react with hot concentrated sodium hydroxide to form sodium silicate (IV) salt.

Chemical equation: $SiO_2$ (s) + $2NaOH$ (aq) -> $Na_2SiO_3$ (aq) + $H_2O$ (l)

(ii) Reaction with dilute acids

a) Ionic oxides react with dilute acids to form salt and water only. This is a **neutralization** reaction. e.g.

Chemical equation: $Na_2O(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

Chemical equation: $MgO(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$

Chemical equation: $Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$

Aluminium oxide is amphotellic and reacts with **hot concentrated strong** alkalis sodium/potassium hydroxides to form **complex** sodium aluminate(III) and potassium aluminate(III) salt.

Chemical equation: $Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$

Chemical equation: $Al_2O_3(s) + 2KOH(aq) + 3H_2O(l) \rightarrow 2KAl(OH)_4(aq)$

b) Acidic oxides **do not** react with dilute acids.

**c)Periodicity of the Chlorides of elements along/across period 3**

The table below summarizes some properties of the chlorides of elements in period 3 of the periodic table.

| **Formular of chloride/ Property** | **NaCl** | **$MgCl_2$** | **$AlCl_3$** | **$SiCl_4$** | **$PCl_5$** **$PCl_3$** | **$SCl_2$** $S_2Cl_2$ | **$Cl_2$** |
|---|---|---|---|---|---|---|---|
| **Melting point($^oC$)** | 801 | 714 | Sublimes at 180 **$^oC$** | -70 | $PCl_5$ Sublimes at -94 **$^oC$** | -78 | -101 |
| **Boiling point($^oC$)** | 1465 | 1418 | 423(as **$Al_2Cl_6$** vapour | 57 | 74(as **$P_2Cl_6$** Vapour 164 (as **$PCl_5$)** | decomposes at 59 **$^oC$** | -34 |
| **Bond type** | Ionic | Ionic | Ionic/ Covalent/ dative | Covalent | Covalent | Covalent | Covalent |
| **Chemical structure** | Giant ionic | Giant ionic | Molecular / | Simple molecula | Simple molecula | Simple molecula | Simple molecula |

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| | structure | structure | dimerizes | or molecule | or molecule | or molecules | or molecule |
| **State at room temperature** | Solid | Solid | Solid | liquid | Liquid **$PCl_5$ is solid** | liquid | Gas |
| **Nature of Chloride** | Neutral | Neutral | Strongly acidic | Strongly acidic | Strongly acidic | Strongly acidic | Strongly acidic |
| **pH of solution** | 7.0 | 7.0 | 3.0 | 3.0 | 3.0 | 3.0 | 3.0 |
| **Reaction with water** | Dissolve | Dissolve | -Hydrolysed by water -Acidic hydrogen chloride fumes produced | -Hydrolysed by water -Acidic hydrogen chloride fumes produced | Hydrolysed by water -Acidic hydrogen chloride fumes produced | Hydrolysed by water -Acidic hydrogen chloride fumes produced | Forms HCl and HClO |
| **Electrical conductivity in molten/aqueous state** | good | good | poor | nil | nil | nil | nil |

1. Sodium Chloride, Magnesium chloride and aluminium chloride are **solids** at room temperature and pressure.

Silicon(IV) chloride, phosphorus(III)chloride and disulphur dichloride are **liquids.** Phosphorus(V)chloride is a **solid**. Both chlorine and sulphur chloride are **gases.**

2. Across/along the period bonding changes from **ionic** in Sodium Chloride and Magnesium chloride to **covalent** in the rest of the chlorides.

3. Anhydrous aluminium chloride is also a molecular compound .Each aluminium atom is covalently bonded to three chlorine atoms.

In vapour/gaseous phase/state two molecules dimerizes to $Al_2O_6$ molecule through coordinate/dative bonding.

4. Across/along the period the structure changes from giant **ionic** in Sodium Chloride and Magnesium chloride to **simple molecules/molecular structure** in the rest of the chlorides.

5. Ionic chlorides have very high melting /boiling points because of the strong ionic bond/electrostatic attraction between the ions in their crystal lattice.The rest of the chlorides have low melting /boiling points because of the weak van-der-waal /intermolecular forces.

6. Sodium Chloride and Magnesium chloride in molten and aqueous state have free/mobile ions and thus good electrical conductors. Aluminium chloride is a poor conductor. The rest of the chlorides do not conduct because they have no free/mobile **ions**.

7. Ionic chloride form **neutral** solutions with pH =7. These chlorides **ionize**/**dissociate** completely into free cations and anions.i.e;

Sodium Chloride and Magnesium chloride have pH=7 because they are fully/completely ionized/dissociated into free ions.

Chemical equation $NaCl\ (s) \rightarrow Na^+(aq) + Cl^-(aq)$

Chemical equation $MgCl_2\ (s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$

8 Across/along the period from aluminium chloride, **hydrolysis** of the chloride takes place when reacting/dissolved in water.

Hydrolysis is the reaction of a compound when dissolved in water.

a)Aluminium chloride is hydrolyzed by water to form aluminium **hydroxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $AlCl_3$ (s) + $3H_2O$(l)-> $Al(OH)_3$(s) + 3HCl(g)

b)Silicon(IV) chloride is hydrolyzed by water to form silicon(IV)**oxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $SiCl_4$ (l) + $2H_2O$(l)-> $SiO_2$(s) + 4HCl(g)

This reaction is highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

c) Both phosphoric (V) chloride and phosphoric (III) chloride are hydrolyzed by water to form phosphoric (V) **acid** and phosphoric (III) **acid** respectively. Fumes of hydrogen chloride gas are produced. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $PCl_5$ (s) + $4H_2O$(l)-> $H_3PO_4$(aq) + 5HCl(g)

Chemical equation $PCl_3$ (s) + $3H_2O$(l)-> $H_3PO_4$(aq) + 3HCl(g)

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This reaction is also highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

d) Disulphur dichloride similarly hydrolyzes in water to form **yellow** deposits of sulphur and produce a mixture of **sulphur (IV) oxide** and **hydrogen chloride** gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $2S_2Cl_2(l) + 2H_2O(l) \rightarrow 3S(s) + SO_2(g) + 4HCl(g)$

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**D. COMPREHENSIVE REVISION QUESTIONS**

1.The grid below represents periodic table. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| | | | | | | | | A |
| B | | | | G | | H | E | C |
| | J | | I | L | | | | |
| D | ***<u>N</u>*** | | | | | | M | |

(a) (I) Indicate on the grid the position of an element represented by letter N whose electronic configuration of a divalent cation is 2:8:8 . ( 1 mark )

(II) Name the bond formed between D and H react. Explain your answer.(2 marks )

**Ionic/electrovalent**
**D is electropositive thus donates two electrons to electronegative H**
(III) Write an equation for the reaction between B and water. ( 1 mark )

**Chemical equation** **$2B\ (s) + 2H_2O(l) \rightarrow 2BOH(aq) + H_2$ (g)**

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(IV) How do the atomic radii of I and L compare. Explain. ( 2 marks )

(V) In terms of structure and bonding explain why the oxide of G has lower melting point than oxide of L. ( 2 marks )

(b) Study the information given below and answer the question that follow.

| Formula of compound N | NaCl | $MgCl_2$ | $Al_2Cl_6$ | $SiCl_4$ | $PCl_3$ | $SCl_2$ |
|---|---|---|---|---|---|---|
| B.P($^0$C) | 1470 | 1420 | Sublimes | 60 | 75 | 60 |
| M.P($^0$C) | 800 | 710 | At $800^0$C | -70 | 90 | -80 |

( I)Why is the formula of aluminium chloride given as $Al_2Cl_6$ and not $AlCl_3$ ? ( 1 mark )

(II) Give two chlorides that are liquid at room temperature. Give a reason for the answer. (2 marks )

(III) Give a reason why $Al_2Cl_6$ has a lower melting point than $MgCl_2$ although both Al and Mg are metals. (1 mark )

(IV) Which of the chlorides would remain in liquid state for the highest temperature range explain why ? (2 mark )

(Kakamega)

2. a) Study the information given below and answer the questions that follow.

| Element | Atomic radius (nm) | Ionic radius | Formula of oxide | Melting point of oxide (‘C) |
|---|---|---|---|---|

| | | (nm) | | |
|---|---|---|---|---|
| P | 0.364 | 0.421 | $A_2O$ | -119 |
| Q | 0.830 | 0.711 | $BO_2$ | 837 |
| R | 0.592 | 0.485 | $E_2O_3$ | 1466 |
| S | 0.381 | 0.446 | $G_2O_5$ | 242 |
| T | 0.762 | 0.676 | JO | 1054 |

(i) Which elements are non-metals? Give a reason. (2mks)

(ii) Explain why the melting point of the oxide of R is higher than that of the oxide of S. (2mks)

(iii) Give **two** elements that would react vigorously with each other. Explain your answer. (2mks)

b) Study the information in the table below and answer the questions that follow (The letters do not represent the actual symbols of the elements)

| | | Ionization Energy_kJ/Mole | |
|---|---|---|---|
| Element | Electronic configuration | 1st ionization energy | 2nd ionization energy |
| A | 2.2 | 900 | 1800 |
| B | 2.8.2 | 736 | 1450 |
| C | 2.8.8.2 | 590 | 1150 |

(i) What chemical family do the elements A, B and C belong? (1mk)

(ii) What is meant by the term ionization energy? (1mk)

iii) The $2^{nd}$ ionization energy is higher that the $1^{st}$ ionization energy of each. Explain

(1mk)

(iv)When a piece of element C is placed in cold water, it sinks to the bottom and an effervescence of a colourless gas that burns explosively is produced. Use a simple diagram to illustrate how this gas can be collected during this experiment.

(3mks)

3. The grid below represents part of the periodic table. The letters do not represent the actual symbols.

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| | | | | | | | | A |
| B | | | X | G | | Z | E | V |
| | J | | I | L | | T | | |
| D | ***N*** | | | | | | M | |

a) Select the most reactive non-metal.
(1mk)

b) Write the formula of the compound consisting of

I.D and Z only.
(2mk)

II. X and Z

c) Select an element that can form an ion of change +2
(1mk)

d) Which element has the least ionization energy? Explain
(2mks)

e) Suggest with reasons a likely pH value of an aqueous solution of the chlorine of:(3mks)

T.

B

X

f) To which chemical family do the following elements belong? (2mk)

J

V

g) An element K has relative atomic mass of 40.2.It has two isotopes of masses 39 and 42. Calculate the relative abundance of each isotope. (3mks)

4.The grid below shows part of the periodic table study it and answer the questions that follow. The letters do not represent the true symbols.

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| | | | | | | | | |
| | | | | | | A | | |
| | B | | C | | D | | E | |
| F | G | | | | | | | |
| | | | | | | | H | |
| | | | | | | | | |

\(a) Which element forms ions with charge of 2-. Explain (2mks)

(b) What is the nature of the oxide formed by C. (1mk)

(c) How does the reactivity of H compare with that of E. Explain? (2mks)

(d)Write down a balanced equation between B and Chlorine. (1mk)

(e) Explain how the atomic radii of F and G compare. (1mk)

(f) If the oxides of F and D are separately dissolved in water, state and explain the effects of their aqueous solutions on litmus. (3mks)

5. (a) The grid below show part of the periodic table.(The letter do not represent the actual symbols).Use it to answer the questions that follow.

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| **T** | | | | | | | | **Q** |
| | | | | **S** | | **R** | **K** | |
| **A** | **J** | | **Y** | | **U** | | **L** | |
| **W** | | | | | | | **M** | **B** |
| | **C** | | | | | | **N** | |
| **P** | | | | | | | | |
| | | | | | | | | |

(i)Select the most reactive non-metal. (1mk)

(ii)Select an element that forms a divalent cation. (1mk)

(iii)Element **Z** has atomic number 14.Show its position in the grid. (1mk)

(iv)How do the atomic radii of **U** and **J** compare? (2mks)

(v)How do electrical conductivity of **A** and **Y** compare? (2mks)

(vi)How does the boiling point of elements **K, L** and **M** vary? Explain (2mks

(b) The table below gives information on four elements by letters K, L, M and N. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

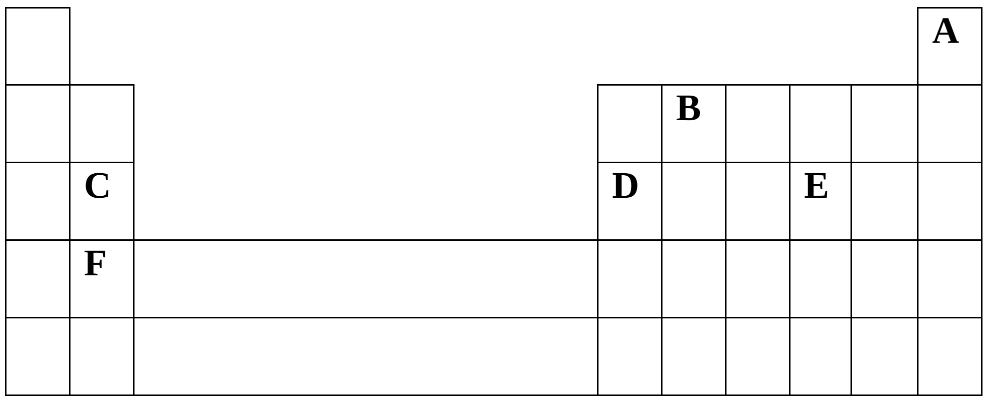
| **Element** | **Electron arrangement** | **Atomic radius** | **Ionic radius** |
|---|---|---|---|
| **K** | **2:8:2** | **0.136** | **0.065** |
| **L** | **2:8:7** | **0.099** | **0.181** |
| **M** | **2:8:8:1** | **0.203** | **0.133** |
| **N** | **2:8:8:2** | **0.174** | **0.099** |

(a) Which two elements have similar properties? Explain. (2mks)

(b) Which element is a non-metal? Explain. (1mk)

(c) Which one of the elements is the strongest reducing agent. (1mk)

6. The grid given below represents part of the periodic table study it and answer the questions that follow. (The letters do not represent the actual symbols of the elements.)




(i) What name is given to the group of elements to which C and F belong? (1mk)

(ii) Which letter represents the element that is the least reactive? (1mk)

(iii) What type of bond is formed when B and E react? Explain (2mks)

(iv)Write formula of the compound formed where elements D and oxygen gas react. (1mk)

(v) On the grid indicate the a tick (√) the position of element G which is in the third period of the periodic table and forms $G^{3-}$ ions.
(1mk)

(b) Study the information in the table below and answer the questions that follow. (The letter do not represents the actual symbols of the substance).

| Substance | Melting point °C | Boiling point °C | Solubility in water | Density at room. Temp/g/cm$^3$ |
|---|---|---|---|---|
| H | -117 | 78.5 | Very soluble | 0.8 |
| J | -78 | -33 | Very soluble | 0.77x $1^{-3}$ |
| K | -23 | 77 | Insoluble | 1.6 |
| L | - 219 | -183 | Slightly Soluable | 1.33 x $10^{-3}$ |

I.(i) Which substance would dissolve in water and could be separated from the solution by fractional distillation.
(1mk)

(ii) Which substances is a liquid at room temperature and when mixed with water two layers would be formed?
(1mk)

II. Which letter represents a substance that is a gas at room temperature and which can be collected ;

(i) Over water?
(1mk)

(ii) By downward displacement of air? Density of air at room temperature = $1.29 \times 10^{-3}$ g/C (1mk)

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