KEMENTERIAN PENDIDIKAN TINGGI
JABATAN PENDIDIKAN POLITEKNIK DAN KOLEJ KOMUNITI

POLITEKNIK MALAYSIA
TUN SYED NASIR

REACTING TO CHEMISTRY

A HANDS ON LEARNING EXPERIENCE

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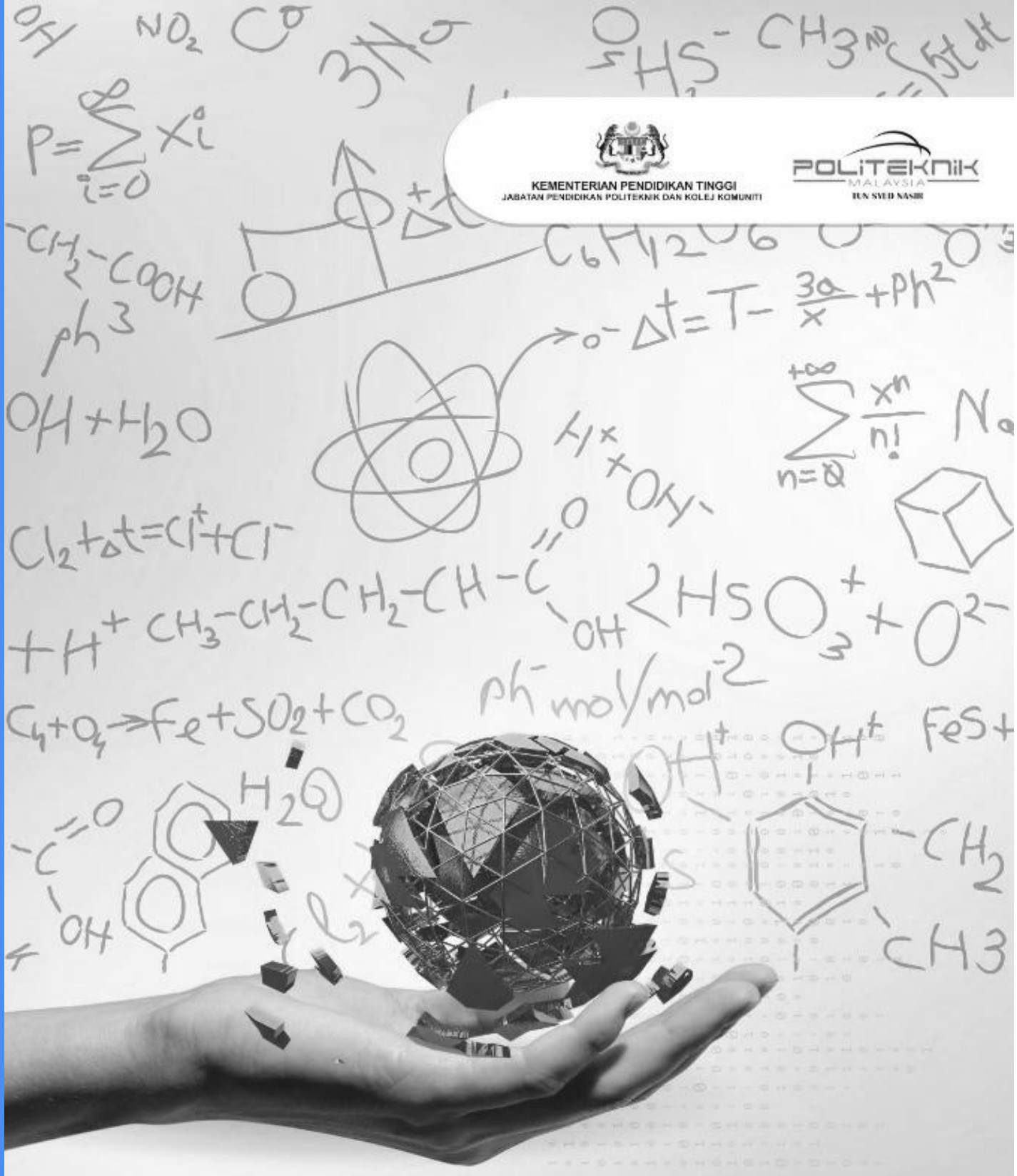

REACTING TO CHEMISTRY

A HANDS ON LEARNING EXPERIENCE

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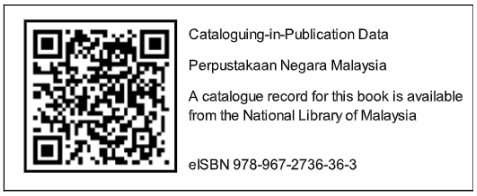
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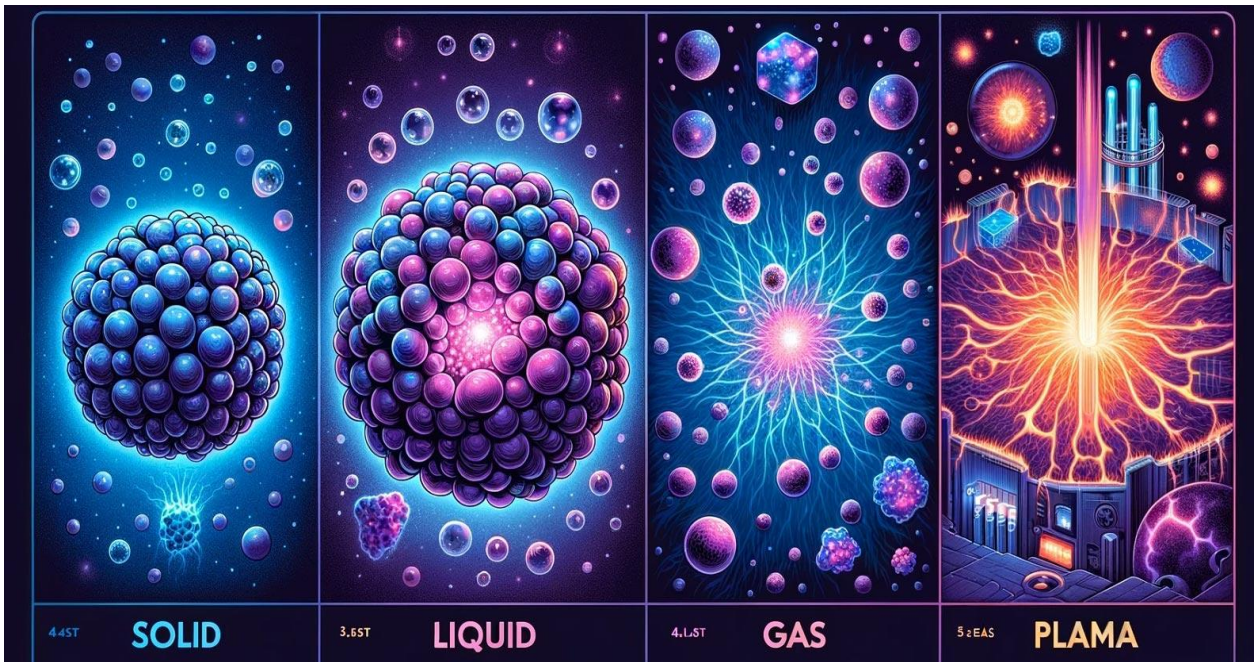
For persons who want to hand-on learning experience of chemistry

and

For students at higher education (pre-university level, diploma level, degree level)

CONTENTS

CHEMICAL BONDS



TOPIC 1: MATTER

LEARNING OUTCOMES:

1.1 Explain the concept of matter
1.2 Explain the mole concept
1.3 Explain a chemical formula

1.1 Explain the concept of matter

1.1.1 Definition

- Any material that has mass and occupies space
- According to theory of conservation of matter, matter cannot be destroyed but can be changed from one state to another
- Construction unit of all material around us
-

1.1.2 Particle Theory of Matter

- Matter consists of tiny and discrete particles that are constantly in motion.
- These particles are held together to different degrees by forces of attraction.
- The particles in a certain substance may be **atoms**, **molecules** or **ions**

Table 1.1: Classification of particle

ATOMS	MOLECULES	IONS
Most basic units for any element	Are neutral particles	Electrically charged particles.
Neutral particles (that is not electrically charged) particles.	Can be formed from atoms in small groups or big groups	Positively charged ions = cations Formed when electrically neutral atoms/ molecules lose electrons.
Capable of taking part in chemical reactions e.g: carbon atom combine with oxygen atoms to form carbon dioxide	Consist of two or more atoms, of the same kind (eg: O_2) or of different kinds (eg: CO_2)	Negatively charged ions = anions Formed when electrically neutral atoms/ molecules gain electrons.
Exist independently e.g sodium and potassium atoms	Exist independently, for example the nitrogen molecule (N_2) and the glucose molecule ($C_6H_{12}O_6$)	Example: Na - 1e (lose) → Na+ (cations) Cl + 1e (gain) → Cl^- (anions)

Table 1.2: Classification of matter

<table>
<tr><th>ELEMENTS</th><th>COMPOUND</th><th>MIXTURE</th></tr>
<tr><td>Simplest form of a chemical substance and cannot be broken down any further into other substances regardless of any chemical or physical process.</td><td>Pure substances which are formed by chemically combining 2 or more elements.

Can divided into 2 kinds: covalent or ionic compound</td><td rowspan="3">A mixture is a physical combination of two or more different components.</td></tr>
<tr><td rowspan="2">Exist as solids, liquids or gases depending on temperature and pressure.</td><td>Covalent compounds made up of particles in the form of molecules.
eg;H_20, CO_2, and C_2H_5OH</td></tr>
<tr><td>Ionic compounds made up of charged particles (cations and anion)
eg: Na+Cl-, Mg2+O2-</td></tr>
</table>

1.1.3 Physical State of Matter

- Matter can exist in three forms: solid, liquid and gas
- The differences between those three forms can be determining from Table 1.3 below.

Example: Different Phase of Copper Molecules

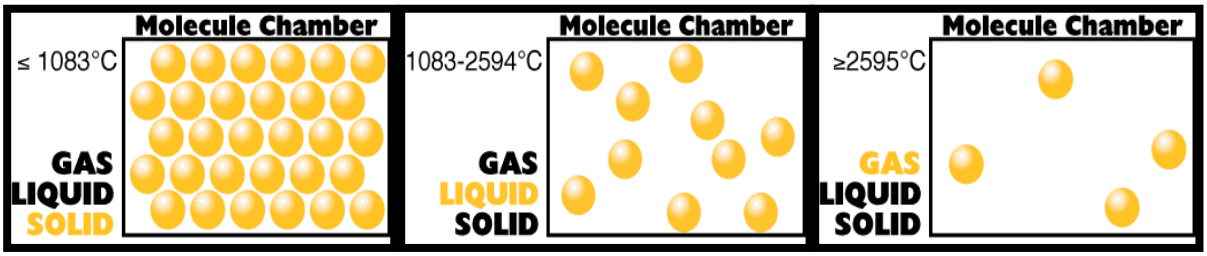
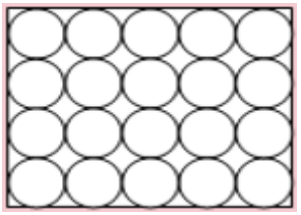
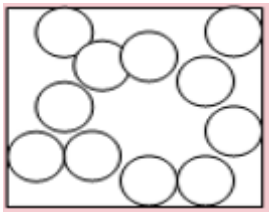
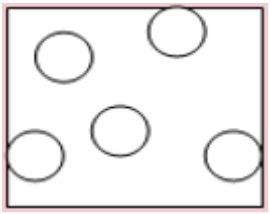


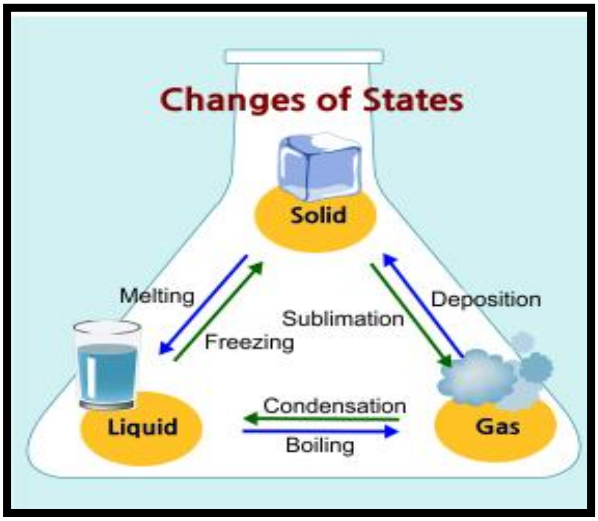
Figure 1.1: Copper molecule's phases

Table 1.3: Comparison of physical state of matter

	Solid	Liquid	Gas
Arrangement of particles	 Densely packed and ordered arrangement of particles: Particles held in fixed position	 Arrangement of particles less ordered but particles still in contact.	 **Particles unarranged and widely spaced.**
Forces of attraction between particles	Very strong and cannot be overcome because of low energy content of particles.	Moderately strong and easily overcome because particles at higher energy state.	**Very weak**
Freedom of motion	No free motion, only vibration and rotation	Free motion but limited to slow speed and within a small area	**Move freely at high speed**
Rate of diffusion of particles	Lowest because particles have nofree motion.	Moderately high because particles have free, though limited motion.	**Highest because particles move freely at high speed.**
Energy content	Lowest	Moderately High	**Highest**
Density	Highest	Moderately High	**Lowest**
Compressibility	Almost incompressible	Poor compressibility	**Easily compressed**
Melting and boiling points	**Highest**	**Low**	**Very Low**

1.1.4 Change State of Matter

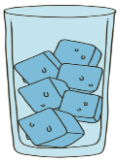
Matter cannot be destroyed but can changed from one state to another



Change of state	Name of process	Transfer of energy
Solid → Liquid	Melting	Heat energy is absorbed
Liquid → Solid	Freezing	Heat energy is release
Liquid → Gas	Boiling/Evaporation	Heat energy is absorbed
Gas → Liquid	Condensation	Heat energy is release
Solid → Gas	Sublimation	Heat energy is absorbed
Gas → Solid	Sublimation	Heat energy is release

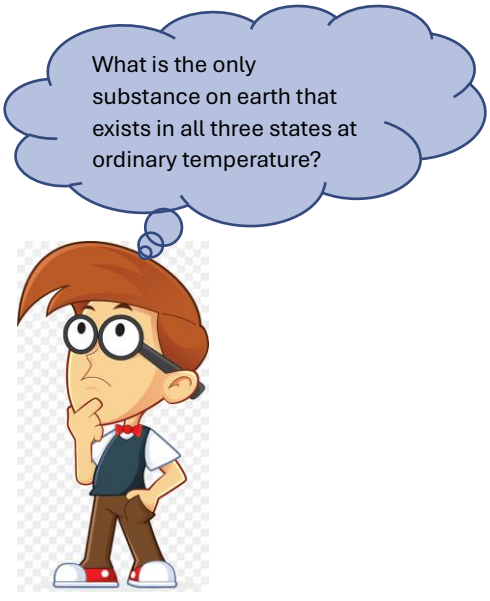
Figure 1.2: Changing process of matter

The melting point of a substance is the temperature at which it changes from a solid to a liquid



The freezing point of a substance is the temperature at which it changes from a liquid to a solid.

The boiling point of a substance is the temperature at which it changes from a liquid to a gas.



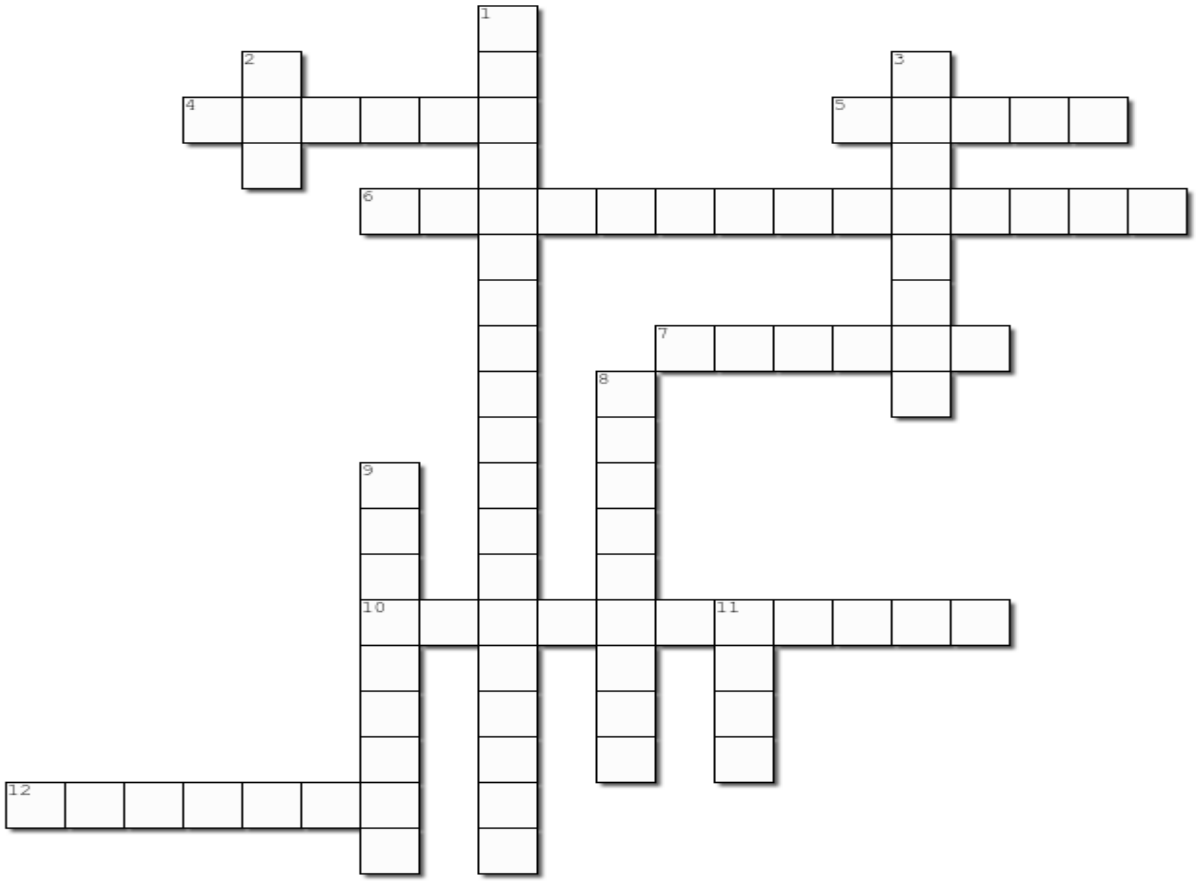
Click below link to watch the summarize video of topic 1.1.

https://youtu.be/oIYyeFhZ7eE

Check your understanding by completing this puzzle below.

Name:_____

Complete the crossword puzzle below



Created using the Crossword Maker on TheTeachersCorner.net

Across

4. consists of tiny and discrete particles that are constantly in motion
5. No free motion
6. temperature at which it changes from a liquid to a solid.
7. Negatively charged ions
10. when liquid absorbed energy
12. a physical combination of two or more different components

Down

1. particles are held together to different degrees by _____ .
2. Move freely at high speed
3. Pure substances which are formed by chemically combining 2 or more elements.
8. , matter cannot be _____ but can be changed from one state to another
9. Can be formed from atoms in small groups or big groups
11. Most basic units for any element

1.2 Explain the mole concept

1.2.1 Relative Atomic Mass (RAM)

- Mass of atoms is very small. So we have to use relative atomic mass.
- The mass of a proton (1.0074 u) is almost the same as the mass of neutron (1.0089 u) while the mass of electron is very small.
- Therefore, the RAM of an element is considered to be the same as its nucleon number.
- For example, Carbon-12 atom contains 6 protons, 6 neutrons and 6 electrons. Its relative atomic mass is defined exactly as 12.
- Every other atom is compared to this.

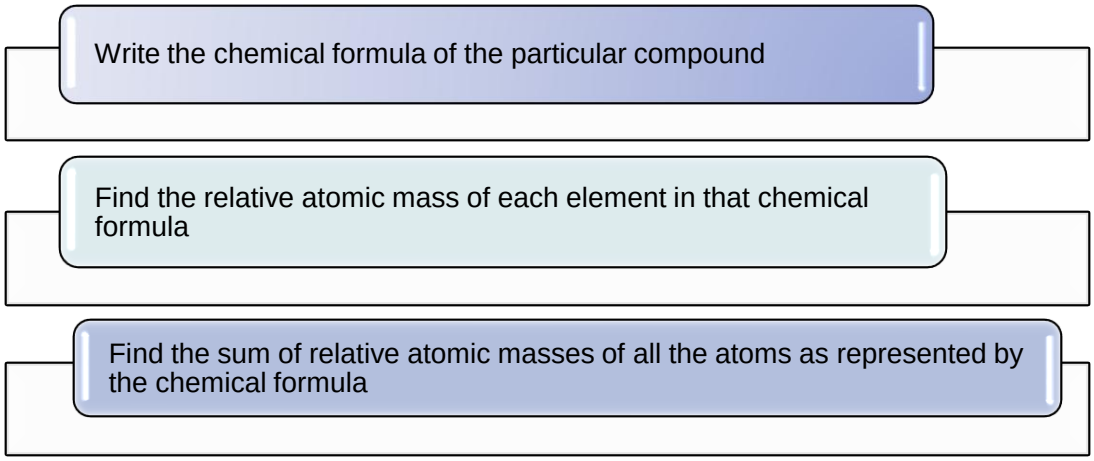
$$\text{R. A. M of an element} = \frac{\text{Mass of one atom of the element}}{1/12 \times \text{mass of one carbon-12 atom}}$$

1.2.2 Relative Molecular Mass (RMM)

- The average or sum of the relative atomic masses of all the atoms present in one molecule of the substance**.**

$$\text{RMM} = \frac{\text{mass of 1 molecule of the element or compound}}{1/12 \times \text{mass of 1 atom of carbon-12}}$$

Steps in calculating the RMM:-



Example: Calculate the RMM of O_2, NaCl and $CuSO_4$

Given relative atomic mass:

H = 1, C = 12, O = 16, Na = 23, Mg = 24, Cl = 35.5, Br = 80,Cu=64, S=32

Solution:

O_2 = O + O = 16 + 16 = **32**

NaCl = Na + Cl = 23 + 35.5 = **58.5**

$CuSO_4$ = Cu + S + O + O + O + O = 64 + 32 + 16 + 16 + 16 + 16 = **160**

1.2.3 Concept of Mole

- One mole of any substance is the amount of substance that contains 6.02 x 10^{23} particles

$$\text{Number of moles} = \frac{\text{mass in gram}}{\text{RMM in grams}} \quad (1)$$

- This number, 6.02 x 10^{23} mol^{-1} is called the Avogadro number or the Avogadro constant (N_A)

$$\text{Number of molecules} = \text{Number of moles} \times (N_A) \quad (2)$$

$$\text{Number of atom} = \text{Number of moles} \times (N_A) \times \text{number of atom present} \quad (3)$$

Example: Given in an experiment use 8g $CuSO_4$, calculate: -

a) **Number of mole**
b) **Number of molecules $CuSO_4$**
c) **Number of atom Cu, S and O**

Solution: Firstly, identified number of mole using equation 1,

a) Number of mole

Relative molecular mass of $CuSO_4$ = 60 + 32 + 4 (16) =160 g/mole

Number of mole found in 8 g $CuSO_4$:

Solution: Next, identified number of molecules using equation 2,

b) Number of molecules

Number of molecule found in 8 g $CuSO_4$:

= (0.05 mole) x (6.02×10^{23} $mole^{-1}$)

Solution: Finally, identified number of each atom using equation 3,

c) Number of atom Cu

Number of atom Cu found in 8 g $CuSO_4$:

= (0.05 mole) x (6.02×10^{23} $mole^{-1}$) x (1 atom of Cu)

= **3.01×10^{22} atom of Cu**

d) Number of atom S

Number of atom S found in 8 g $CuSO_4$:

= (0.05 mole) x (6.02×10^{23} $mole^{-1}$) x (1 atom of S)

1.2.4 Molar Volume (MV)

- Molar volume or molecular volume is the volume occupied by one mole of gas at a particular temperature and pressure. At standard temperature and pressure (s.t.p), the temperature is at zero degree Celcius (0°C) and the pressure is at 1 atm.

1 molar volume = 22.4dm^3 at standard temperature and pressure (s.t.p) of any gas

or

1 molar volume = 24dm^3 at room temperature and pressure

Example: Calculate: -

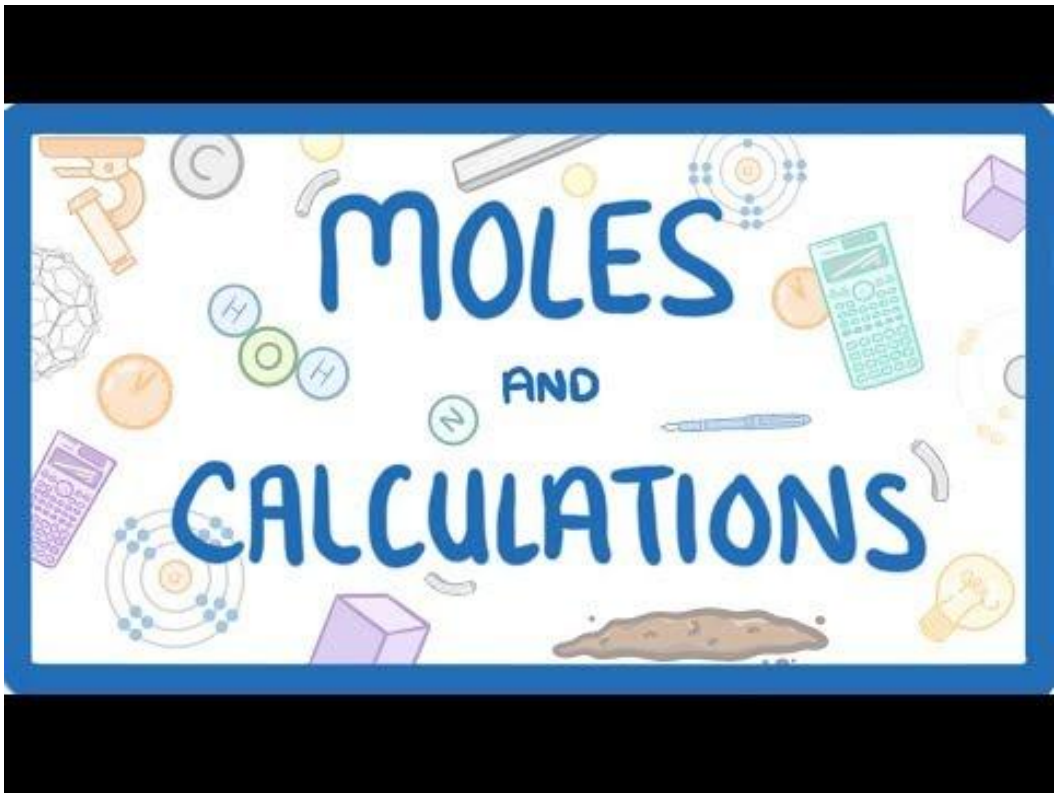
a) **Volume in dm^3 at s.t.p for 34 g of ammonia gas, NH_3**
b) **Volume in dm^3 at room temperature and pressure for 8 g of sulphur dioxide, SO_2**

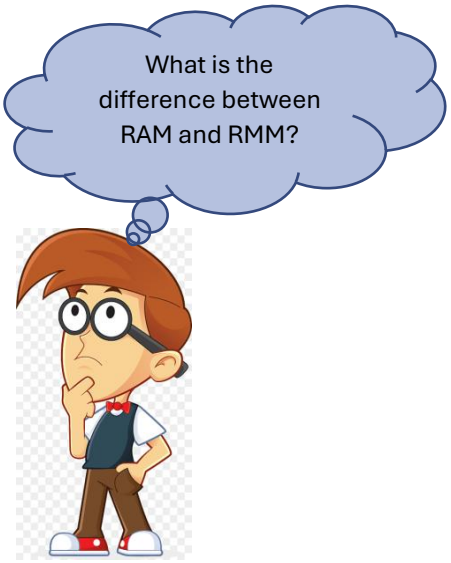
Solution: Firstly, identified number of mole using equation 1, then multiply with MV

a) Number of volume of 34 g of NH_3 Relative molecular mass of NH_3 = 14 + 3 (1) =17 g/mole Number of mole found in 34 g NH_3:	Number of volume of NH_3: = (2 mole) x (22.4 dm^3 at s.t.p) = **44.8 dm^3 at s.t.p**

Solution: Firstly, identified number of mole using equation 1, then multiply with MV

b) Number of volume of 8 g of SO_2 Relative molecular mass of SO_2 = 32 + 2 (16) =64 g/mole Number of mole found in 8 g SO_2:	Number of volume of NH_3: = (0.125 mole) x (24 dm^3) = **3.0 dm^3 at room temperature**





Scan below QR codes to answering the quizzes.

QUIZ 1

QUIZ 2

1.3 Explain the chemical formula

1.3.1 Definition

- The chemical formula of an element shows the symbol representing that element and the number of atoms in the element

- For example,

    - Chemical formula of magnesium – Mg

    - Chemical formula of chlorine - Cl_2

    - Chemical formula of bromine - Br_2

    - Chemical formula of sulphuric acid - H_2SO_4

- There are three ways to write the chemical formula: Empirical Formula, Molecular Formula and Structural Formula.

- It is possible for different compound to have same empirical formula. For example, benzene (C_6H_6) and acetylene (C_2H_2) both have same empirical formula CH.

MOLECULAR FORMULA

The chemical formula which shows the **actual number** of the atoms of the elements that combine to form the compound

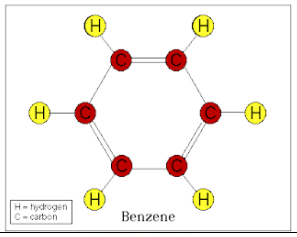
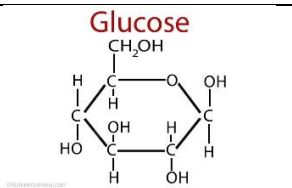
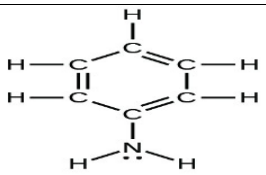
EMPIRICAL FORMULA

The chemical formula which shows the **smallest whole number ratio** of the atoms of the elements that combine to form the compound

STRUCTURAL FORMULA

A formula which shows the **arrangement of atoms** in the molecule of a compound.

Table 1.4: Comparing molecular formula and empirical formula

Name of compound	Molecular formula	Empirical formula	Lowest ratio of element	Structural formula
Benzene	C_6H_6	CH	1:1	
Acetylene	C_2H_2	CH	1:1	 ethyne (acetylene)
Glucose	$C_6H_{12}O_6$	CH_2O	1:2:1	
Aniline	C_6H_7N	C_6H_7N	6:7:1	

1.3.1 Determining Empirical Formula

- Empirical formula can be determining from:
  - percentage (%) composition
  - mass of each component element in the compound
- Following is the guideline to write the empirical formula:
  - **Step 1**
    - assume a definite starting quality (usually 100.0 g) of the compound, if the percentage composition is given.
  - **Step 2**
    - ensure the RAM of each elements
  - **Step 3**
    - Divide % composition or mass composition of each element by its RAM to get no of moles.
  - **Step 4**
    - Calculate simplest whole number ratio for each element
  - **Step 5**
    - Determine the empirical formula

Example: Determine the empirical formula of the compound.

A nitrogen compound contains 82.3% nitrogen, and the rest of it is hydrogen. What is the empirical formula of the compound?
(Relative atomic mass: H = 1; N = 14)

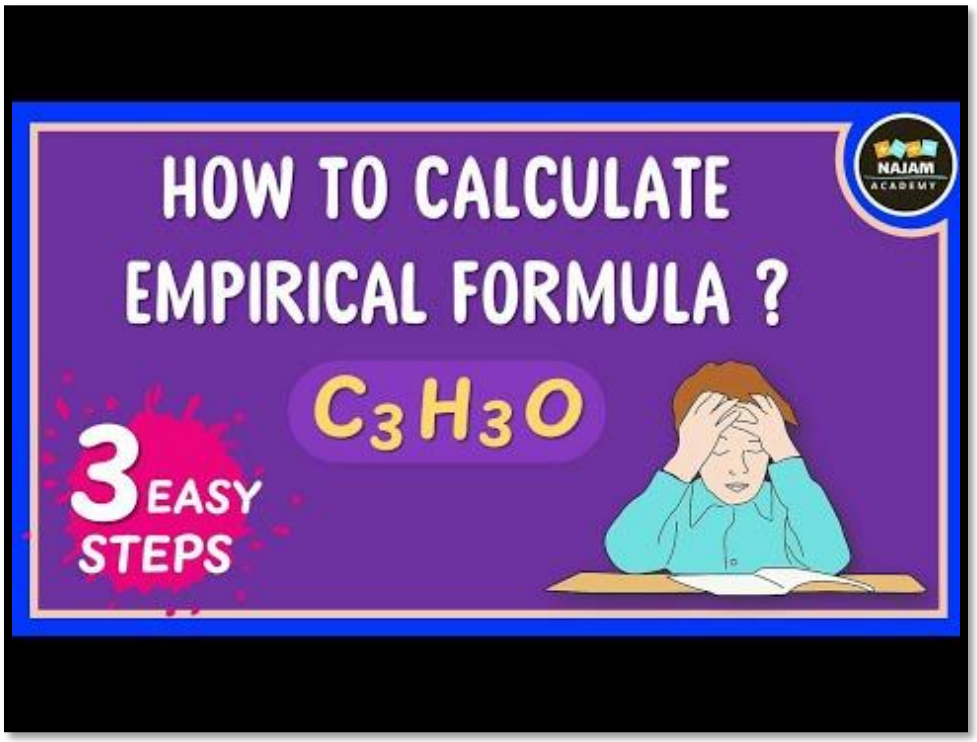
Solution

	Element N	Element H
% mass	82.3	17.7
Mass in 100 g	82.3 g	17.7 g
Number of mole atom = $\frac{\text{mass}}{\text{R.A.M}}$	$\frac{82.3}{14} = 5.88$	$\frac{17.7}{1} = 17.7$
Simplest Ratio	$\frac{5.88}{5.88} = 1$	$\frac{17.7}{5.88} = 3$
Empirical formula	N_1H_3 or NH_3	

Example: Determine the molecular formula of the compound.

The empirical formula of benzene is CH and its relative molecular mass is 78	**Substitute the RAM of each element**
Solution:	n (12) + n (1) = 78
	13 n = 78
$(CH)_n = 78$ where n = a whole number	n = 78 / 13
	n = 6

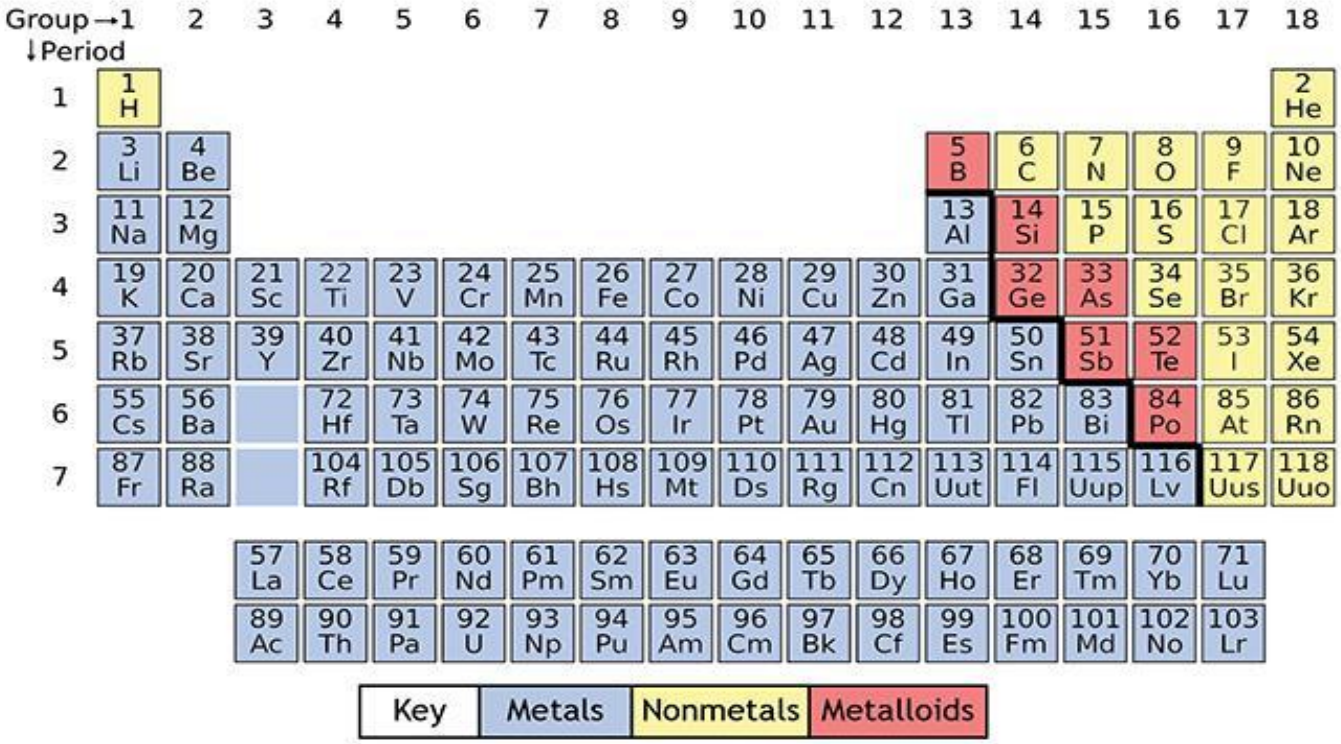
Please watch below video for more understanding.



Test your understanding by answering the quiz below.

https://take.quiz-maker.com/QFKDUFKSM

CHAPTER 2:
PERIODIC TABLE



LEARNING OUTCOMES

2.1 Explain the periodic table and atomic structure.
2.2 Explain the electron configuration.
2.3 Explain the properties of groups and periods in the periodic table of elements.

2.1 Explain the periodic table and atomic structure.

- Periodic table is a table that lists all known elements.
- All elements are arranged in a particular order by their proton number in increasing order.
- This table holds 109 elements:
  - arranged according to increasing proton number.
  - Arranged in specific groups and periods.

Scan this barcode to know what periodic table is!

Niels Bohr

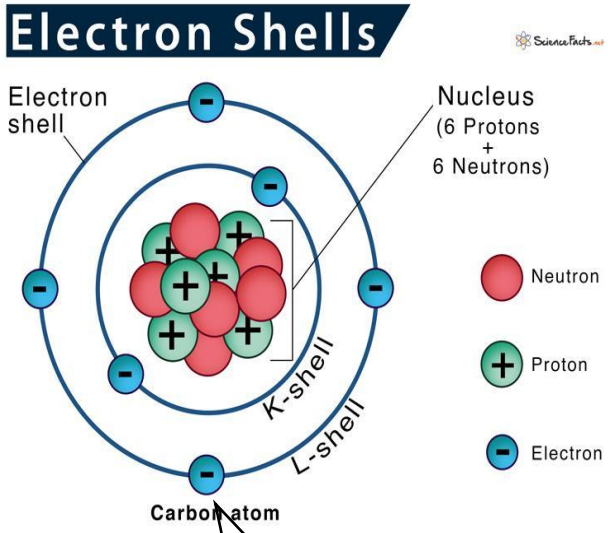
- In 1913, Niels Bohr proposed an improvement to Rutherford's Model:
- Electron's move in definite orbits around the nucleus, like planets around the sun.

2.2 Explain the electron configuration.

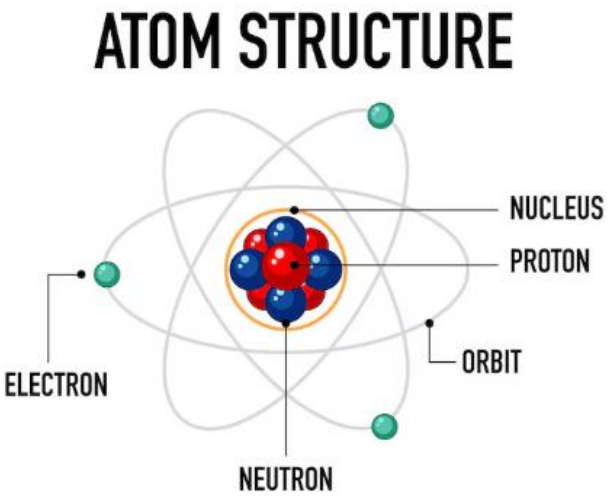
THE THEORY

- Electrons revolve around nucleus in specific orbits
- Orbits associated with definite energies
- Orbits called shells or energy levels
- Shells represented by letters K, L, M, N, etc.
- Maximum number of electrons per shell: 2n^2 (n = shell number)

- K shell (n=1): 2 electrons maximum
- L shell (n=2): 8 electrons maximum
- M shell (n=3): 18 electrons maximum
- N shell (n=4): 32 electrons maximum
- Outermost shell: Maximum of 8 electrons



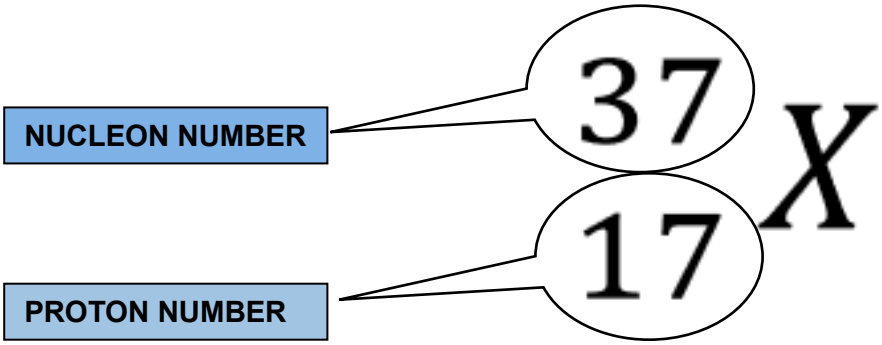
ARRANGEMENT OF ELECTRON IN ATOM



VALENCE ELECTRON-The outmost shell which has electrons and which is also the furthest from the .

- **Proton number (Z) / Atomic number / Number of electron**
  - total number of protons in the nucleus of 1 atom of that element

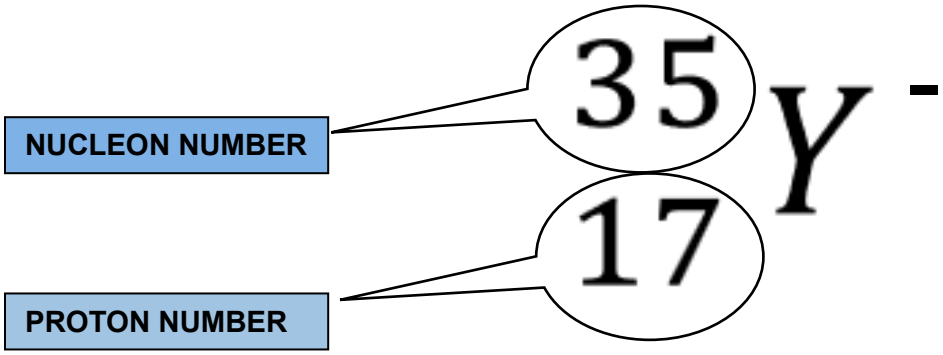
- **Nucleon number (A) / Mass number**
  - total number of protons and neutrons in the nucleus of 1 atom of that element.



NEUTRON NUMBER = Nucleon number –Proton number

= 37 -17

= 20 (because the element is neutral



NEUTRON NUMBER = Nucleon number – Proton number

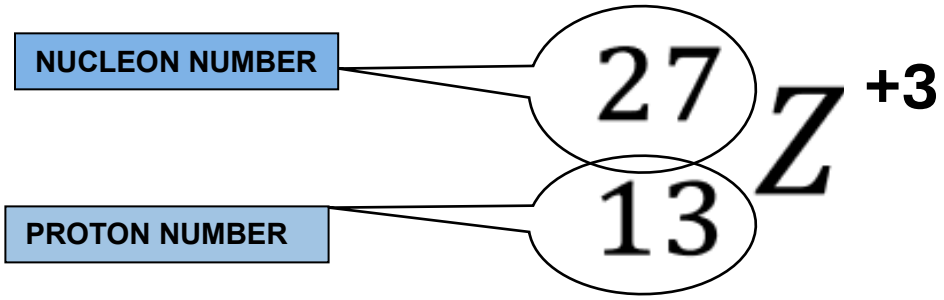
= 35 -17

= 18

ELECTRON NUMBER = Proton number – charge carried by the element

= 17 – (-1)

= 18



NEUTRON NUMBER = Nucleon number – Proton number

= 27 -13

= 14

ELECTRON NUMBER = Proton number – charge carried by the element

= 13 – (+3)

= 10

Scan this barcode to test your understanding about nucleon and proton/electron number

WHAT ARE ISOTOPES?

- Isotopes are atoms of the same element having *different masses*, due to varying numbers of neutrons.
- It has same number of protons but different number of neutrons
- Soddy won the Nobel Prize in Chemistry in 1921 for his work with isotopes and radioactive materials.
- Examples: carbon-12, carbon-14 and uranium-235

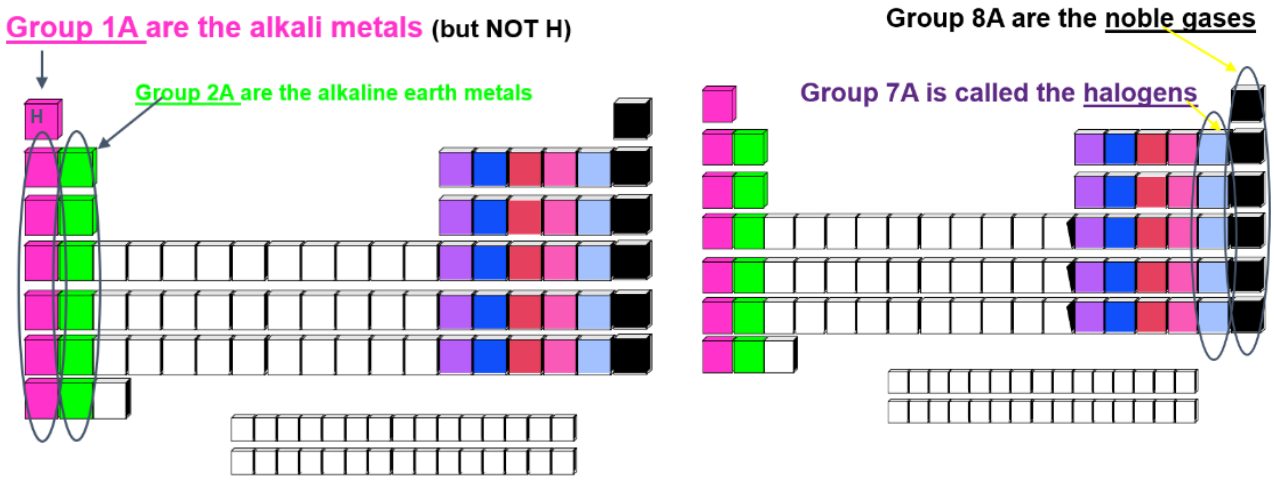
Question: Discuss the similarities and differences between isotopes

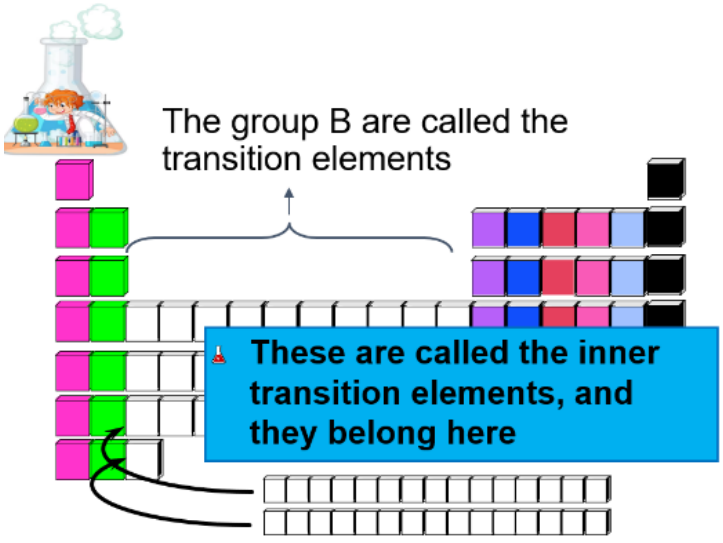
SIMILARITIES	DIFFERENCES

2.3 Explain the properties of groups and periods in the periodic table of elements

GROUP 1A/1	GROUP 2A/2	GROUP 7A/17	GROUP 8A/18
Forms a "base" (or alkali) when reacting with water	alkaline earth metals Also form bases with water; do not dissolve well, hence "earth metals"	halogens Means "salt-forming"	Previously called "**inert gases**" because they rarely take part in a reaction; very stable = don't react Noble gases have an electron configuration that has the outer s and p sublevels completely full
-Solids at room temperature -Grey in colour & have shiny silvery surfaces when freshly cut -Good conductors of heat & electricity -Strong reducing agents		-First two elements F_2 and Cl_2 are gases, Br^2 is a liquid & At_2 is a solid -Have its own colour -The colour darker down the group -Do not conduct electricity and heat -All halogens in gaseous form are pungent & poisonous	All inert gases are non-metals which have the following properties: -Low melting and boiling points because of the Van der Waals' forces of attraction is very weak -Low densities -Insoluble in water -Do not conduct electricity or heat

Properties	Going down group 1 (Alkali Metals)	Going down group 17 (Halogens)	Going down group 18 (Inert gases)
Atomic radius and atomic size	Increase	Increase	Increase
Densities	Increase	Increase	Increase
Attractive force between atoms	Decrease	Increase	-
Melting points & boiling points	Decrease	Increase	Increase
Electropositivity	Increase	-	-
Electronegativity	-	Decrease	-
Reactivity	Increase	Decrease	-





Test your understanding on Topic 2!

TOPIC 3:
CHEMICAL BONDS

LEARNING OUTCOMES:

3.1 Explain the concept of chemical bond
3.2 Explain the formation of ionic bond
3.3 Explain the formation of covalent bond
3.4 Explain the intramolecular and intermolecular forces

3.1 Explain the concept of chemical bond

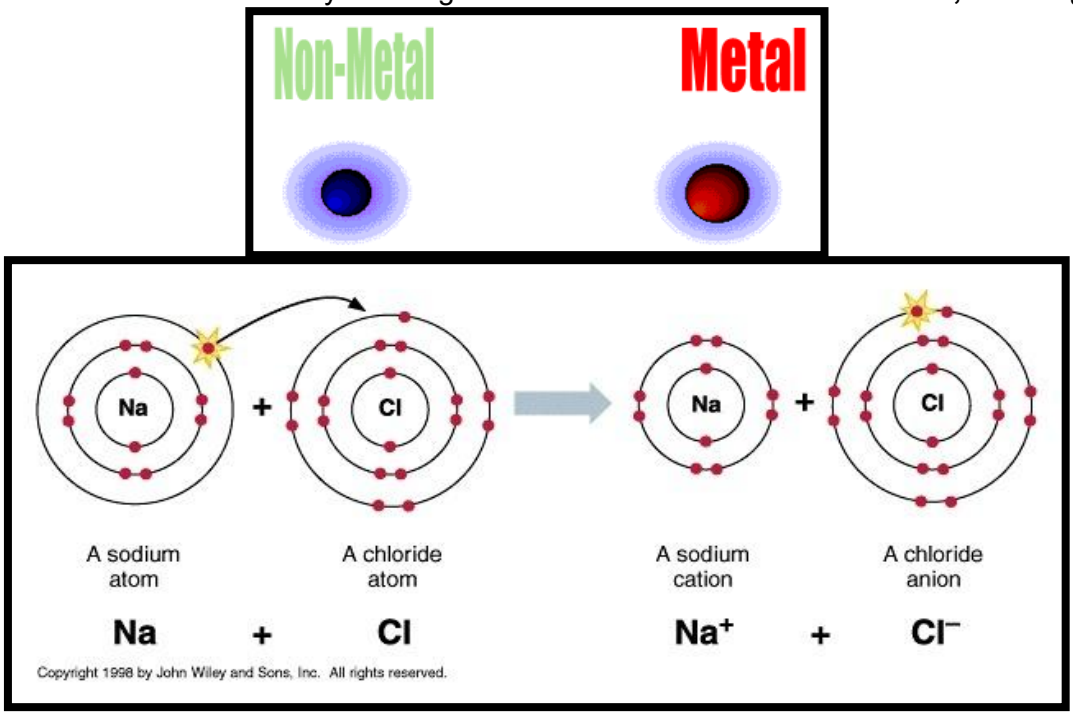
3.1.1 Definition

- When two atoms are joined together to make a chemical compound, the force of attraction between the two atoms is referred as a chemical bond.
- Only valence electrons are involved in chemical bonding
- The valence electron are the outer most shell electrons,
- There are 2 types of chemical bonding: ionic bond and covalent bond

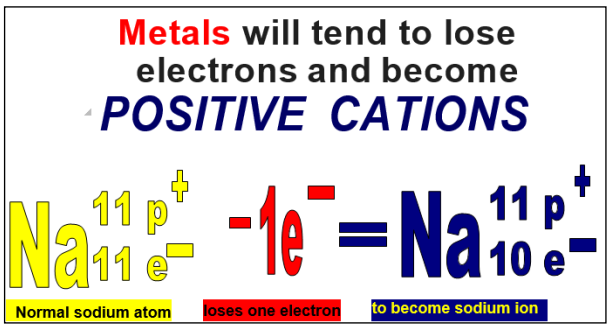
3.2 Explain the formation of ionic bond

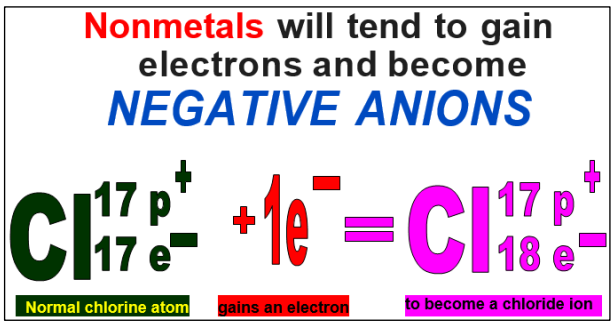
3.2.1 Ionic Bonding

- An ionic (electrovalent) bond is formed as a result of the transfer of one or more electron(s) from **metal atom** to a **non-metal atom** until both the ions have achieved **octet structure.**
- Neutral atoms come near each other. Electron(s) are transferred from the metal atom to the non-metal atom. They stick together because **of electrostatic forces, like magnets.**

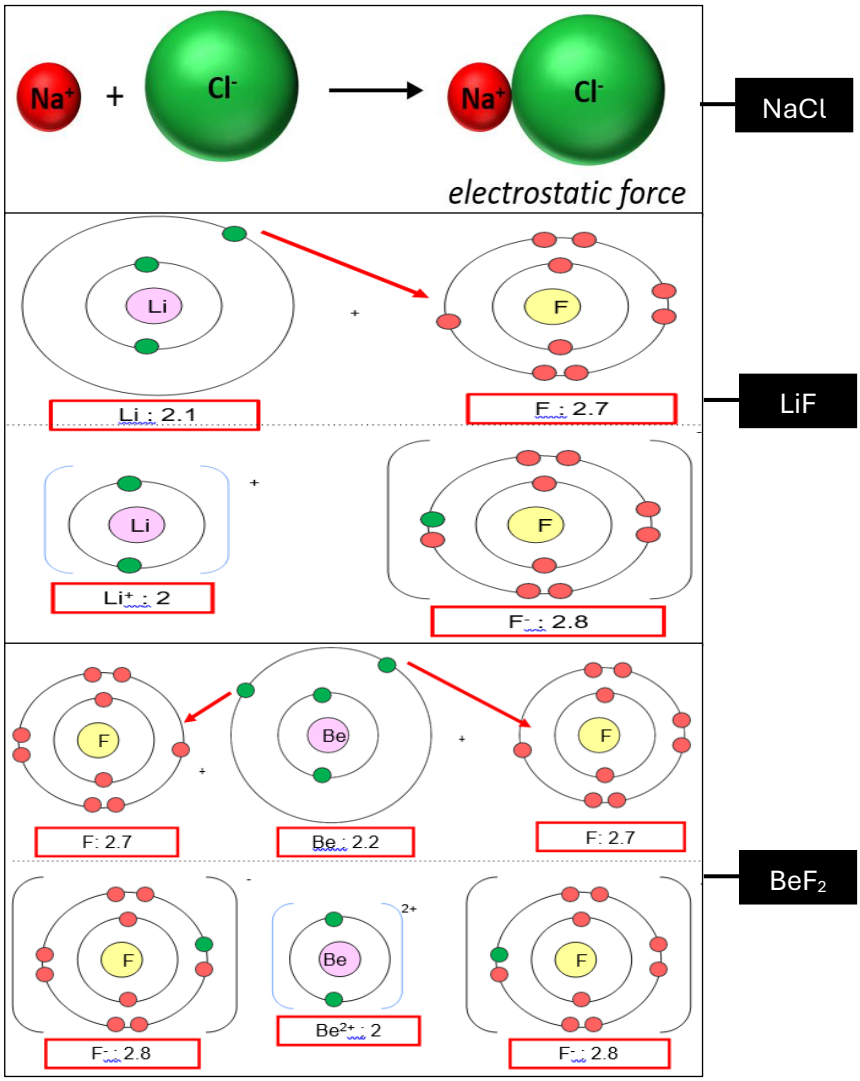


Example: Ionic bonding in NaCl





Example of compound that formed through ionic bonding.



3.2.2 Characteristic of ionic compounds

- Ionic bond stronger than covalent bond
- Structure is rigid – ions occupy fixed position in the ionic lattice
- Crystalline structure – the cations and anions stick together like a magnet
- A regular repeating arrangement of ions in the solid
- High melting points - because of strong forces between ions
- Many soluble in water but not in nonpolar liquid
- Low volatility
- Ionic solids are brittle – strong repulsion breaks crystal apart
- If the ions are made mobile by dissolving the ionic compound in water or melting it, it can conduct electricity

For more understanding, please watch the below video by clicking the link given to see how ionic bond is formed.

https://www.youtube.com/watch?v=5EwmedLuRmw

3.3 Explain the formation of covalent bond

3.3.1 Covalent Bonding

- Covalent bonding is a bond in which two atoms share a pair of electrons.
- When an atom of one **non-metal shares** one or more electrons with an atom of another **non-metal,** so both atoms end up with eight valences electrons/ to achieve **octet configuration.**
- Atoms in covalent bonding are held together by **electrostatic attraction** between the shared electrons and the nuclei of the atoms involved.
- Atoms that are bonded through covalent bonds produce a **molecule.**
- Atoms can form **single, double or triple covalent bond.**

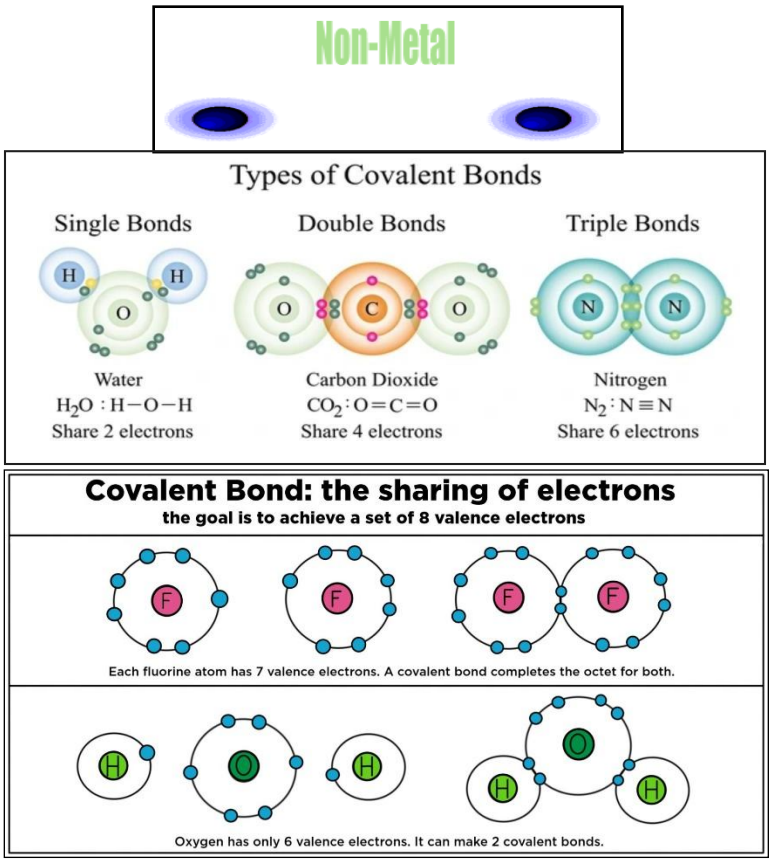
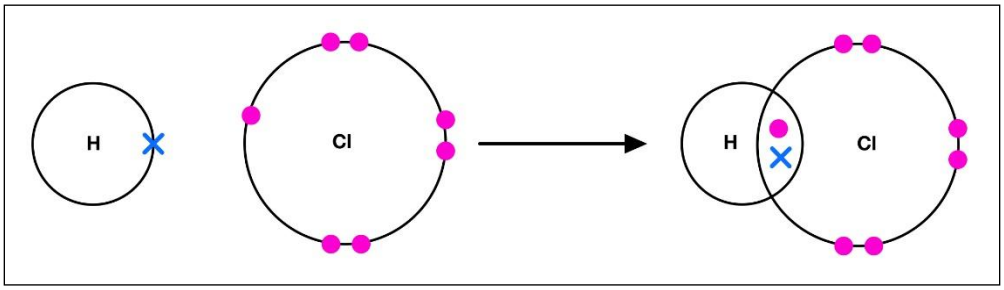


Figure 3.2

Example: Covalent bonding in HCl (a sharing of two valence electrons)



3.3.2 Step to draw the bond between atoms

i. Add up all the valence electrons.
ii. Count up the total number of electrons to make all atoms happy.
iii. Subtract.
iv. Divide by 2
v. Tells you how many bonds - draw them.
vi. Fill in the rest of the valence electrons to fill atoms up.

Example:

- HCN - C is central atom
- N - has 5 valence electrons wants 8
- C - has 4 valence electrons wants 8
- H - has 1 valence electrons wants 2
- HCN has 5 +4+1 = 10
- HCN wants 8+8+2 = 18
- (18-10)/2= 4 bonds
- 3 atoms with 4 bonds -will require multiple bonds - not to H

Solution:

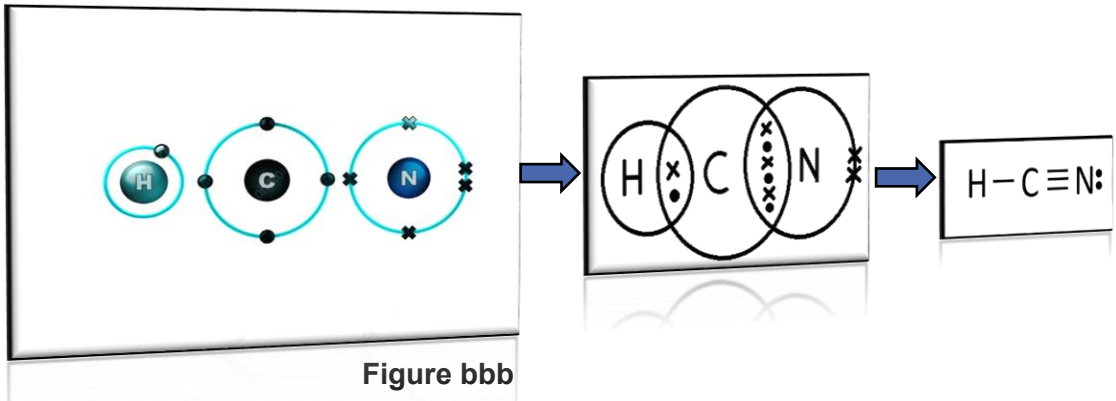


Figure bbb

3.3.3 Polar Covalent Bonding and Electronegativity

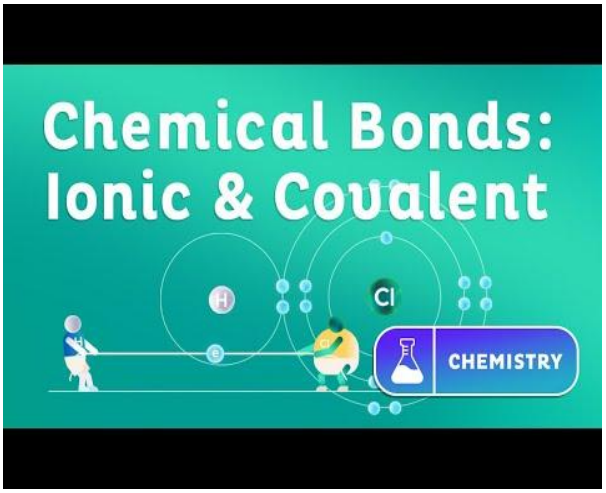
- **Two identical atoms**, which have the same electronegativity, the electrons are shared equally. This is a **non-polar covalent bond.**
- When **two different atoms** are connected, the more electronegativity element has a stronger attraction for the electron, the atoms may not be shared equally. This is a **polar covalent bond.**
- Electronegativity is a measurement of how strongly the atoms attract electrons in a bond.
- The bigger the electronegativity difference the more polar the bond.
- The range electronegativity value is (0.0 - 0.3 Covalent non-polar), (0.3 - 1.67 Covalent polar) and (>1.67 Ionic bond).

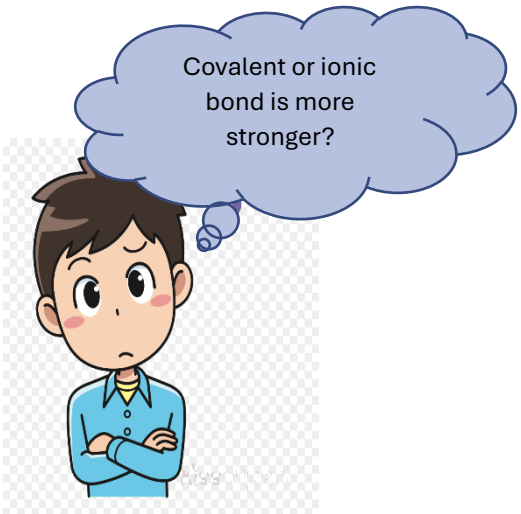
3.3.4 Differences Between Ionic Compounds and Covalent Compounds

Table 3.44

Ionic Compounds	Covalent Compounds
Crystalline solids	Gases, liquids or solids
High melting and boiling point	Low melting and boiling point
Conduct electricity when melted	Poor electricity conductors
Many soluble in water but not in non-polar liquid	Many soluble in non-polar liquid (organic solvent) but not in water

For more understanding, please watch the below video to see how covalent bond is formed.

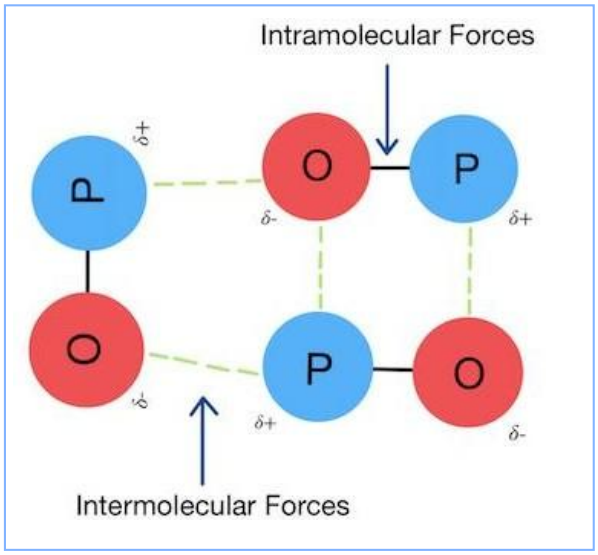


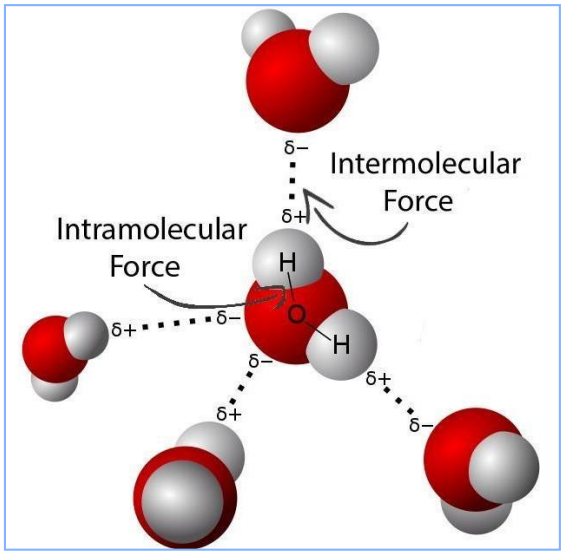


3.4 Explain the intramolecular and intermolecular forces

3.3.1 Definition

- Intramolecular are attractive forces that hold atoms together within a molecule. These are stronger and usually called chemical bonds
- Intermolecular forces are formed from the attraction between molecules
- Intermolecular forces are weaker than intramolecular forces
- Example of intermolecular forces is van der Waal force.





3.3.2 Van der Waal Force

- Van der Waals Forces are intermolecular forces
- The forces are due to the attractions between the partial positive and partial negative electrical charges between molecules
- These forces are affected by the distance between the molecules (like how when you bring magnets closer together, the attraction is greater)
- The weakest of the intermolecular forces
- There are two types of van der Waal Forces:
    - **London dispersion forces**

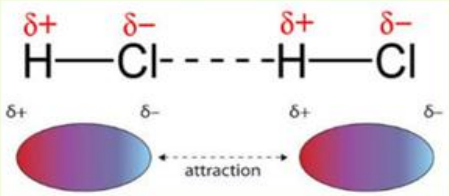
      The weakest force of the Van der Waals and occurs between any two atoms in a molecule (note: this is a temporary force!)

      Due to constant motion of its electron, an atom can develop a temperory dipole moment.
    - **Dipole – dipole interaction**

      Occurs between the positive end of one polar molecule and the negative end of another polar molecule

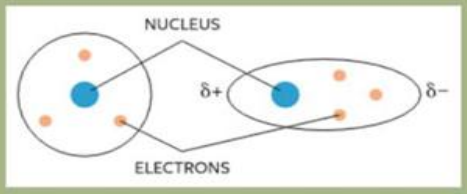
      Molecules with larger dipole moments will have a stronger dipole-dipole interaction.

Dipole-Dipole Forces



Attractive forces that occur between polar molecules

London Dispersion Forces



Intermolecular forces that occur between atoms and between nonpolar molecules as a result of the motion of electrons

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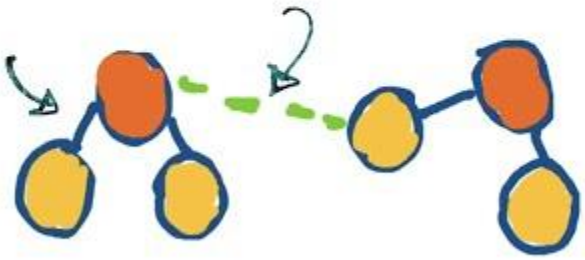
Intramolecular Vs. Intermolecular Forces



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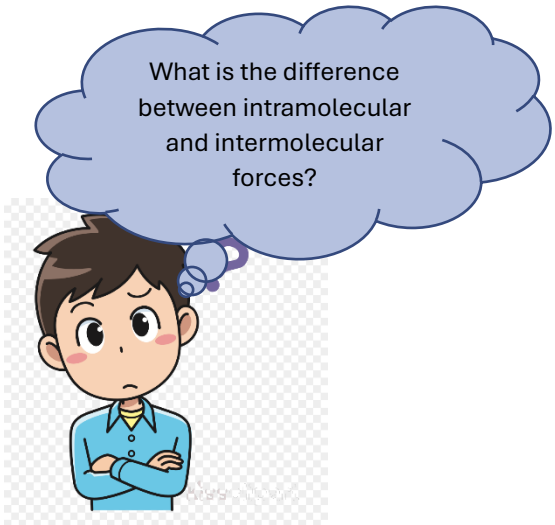

PHOTO CREDITS

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References

Ang, C. (2018). A Complete Guide to Gce O Level Pure Chemistry. Fairfield Book Publishers.

Chang, R. (2002). Chemistry. . Mcgraw Hill..

Lim, Y., S. & Yip , K., H. (2009). Pre U Text STPM Chemistry. Pearson Malaysia.

Lim, Y., S. & Yip , K., H. (2009). Pre U Text STPM Physical Chemistry. Pearson Malaysia.

Norbani, A. (1998). Kimia Fizikal Asas Matrikulasi.Fajar Bakti. (ISBN 978-0-00-749775-1)

Slowinsk, E., J.,Wolsey, W., C. & Rossi, R., C. (2012). Chemical Principles in The Laboratory. Brooks/Cole, Cengage Learning.

Timberlake, K., C. (2010). Basic Chemistry (3rd edition) .Prentice Hall.

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