

UNIT 1: ATOMIC STRUCTURE AND THE PERIODIC TABLE.

Checklist

1. Structure of the atom
2. The subatomic particles;
 - ✓ Protons
 - ✓ Neutrons
 - ✓ Electrons;
3. Atomic number and mass number
4. Isotopes;
5. Energy levels and electron arrangements.
6. The periodic table;
 - ✓ Groups
 - ✓ Periods;
7. Relative atomic mass and isotopes;
8. Ion formation;
 - ✓ Ion;
 - ✓ Cations;
 - ✓ Anions;
 - ✓ Ionization energy;
 - ✓ Electron affinity;
9. Valency and oxidation numbers;
10. Chemical formula;
11. Chemical equations
12. Balancing chemical equations;

The atom:

- Refers to the smallest particle of an element that can take part in a chemical reaction;
- It has an average diameter of 10^{-8} cm with a nucleus of about 10^{-13} cm;

Parts of an atom

- The atom is made of two main parts:
 - ✓ The nucleus
 - ✓ The energy levels;

1. The nucleus:

- Is the positively charged part of an atom;
- The nucleus contains two subatomic particles; neutrons and protons;
- The positive charge is due to presence of protons;
- The nuclei of all atoms contain neutrons except the hydrogen atom;
- The protons and the neutrons are together referred to as the **nucleons**;

2. The energy levels.

- They contain the electrons;
- Electrons are so small and move so fast that their path cannot be traced directly;
- Thus the energy level simply represents the region where the electrons are most likely to be found;

Structure of the atom.

Note:

- The atom can still however be split into smaller particles termed the sub-atomic particles;

The sub-atomic particles.

- Are generally three:
 - ✓ Protons;
 - ✓ Neutrons;
 - ✓ Electrons;

1. Protons.

- Are the positively charged sub-atomic particles;
- Are found in the nucleus and thus form part of the nucleons;
- The number of protons in the nucleus is equal to the number of electrons in the energy levels;

2. Neutrons.

- Are neutrally charged sub-atomic particles found in the nucleus of the atom;
- They are thought to probably prevent the positively charged protons from getting too close to each other;

3. Electrons.

- Are negatively charged sub-atomic particles found in the energy levels;
- The number of electrons in the energy levels is equal to the number of protons in the nucleus;
- This makes the atom to be electrically neutral;

Atomic number and mass number.

Atomic number.

- Refers to the number of protons in the nucleus of an atom;

Examples.

- ✓ Sodium has 11 protons in the nucleus and thus said to have atomic number 11;
- ✓ Chlorine has 17 protons in the nucleus and thus said to have atomic number 17;

Mass number;

- Refers to the sum of the number of protons and neutrons in an atom of an element;

Examples:

- ✓ Sodium has 2 neutrons and 11 protons hence a mass number of 23;
- ✓ Chlorine has 18 neutrons and 17 protons hence a mass number of 35.

Notation of atomic number and mass number;

- Both atomic number and mass number of an element can be written along with the symbol of an element;

Mass number;

- Is conventionally represented as a superscript to the left of the symbol;

Examples:

Sodium; ²³Na;

Magnesium ²⁴Mg;

Atomic number;

- Is conventionally represented as a superscript to the left of the symbol;

Examples:

Sodium; ₁₁Na;

Magnesium ₁₂Mg;

Thus the elements can be conventionally represented as:

Sodium ²³Na

Magnesium ²⁴Mg

Atomic properties of the first 20 elements.

Element	Symbol	Number of electrons	Number of protons	Number Of neutrons	Atomic number	Mass number
Hydrogen	H					
Helium	He					
Lithium	Li					
Beryllium	Be					
Boron	B					
Carbon	C					
Nitrogen	N					
Oxygen	O					
Fluorine	F					
Neon	Ne					
Sodium	Na					
Magnesium	Mg					
Aluminium	Al					
Silicon	Si					
Phosphorus	P					
Sulphur	S					
Chlorine	Cl					
Argon	Ar					
Potassium	K					
Calcium	Ca					

Isotopes.

- Are atoms of the same element with same atomic number but different mass number due to different number of neutrons.

Examples of isotopes.

Element	Isotope	Atomic No.	Number of protons	Number of neutrons	Mass number	Isotopic representation
Hydrogen						
Carbon						
Oxygen						
Chlorine						

Energy levels and electron arrangements.

✓ Energy levels:

- Are definite orbits in an atom that the electrons occupy.
- The energy levels are numbered 1, 2, 3.... starting with the one closest to the nucleus.
- Electrons occupying the same energy level have approximately the same amount of energy.
- Each energy level can only accommodate a given maximum number of electrons.

Maximum number of electrons per energy level

Energy level	Maximum number of electrons
1 st	2
2 nd	8
3 rd	8 (only for the first 20 elements)

Illustrations:

✓ Hydrogen

- It has only one electron and thus this electron occupies the first energy level.
- Since the first energy level is not yet full, hydrogen does not have the second energy level;
- The electron arrangement of hydrogen is thus 1.

✓ Helium:

- Helium is atomic number 2 and has only two electrons, which occupy the first energy level.
- The first energy level is thus completely full, but since there are no other r electrons lithium also has only one energy level;
- The electron arrangement is thus 2.

✓ Chlorine:

- Chlorine has atomic number 17 and thus has 17 electrons;
- The first two electrons occupy the first energy level which is thus completely filled up;
- The remaining 15 electrons occupy the second energy level, which can however accommodate only 8 to be completely filled up;
- Thus the remaining 7 electrons move to the third energy level; which needs 8 to be completely filled up;
- Since the third energy level is not yet full chlorine does not have a fourth energy level;
- The electron arrangement is thus 2.8.7.

Electron arrangement.

- Refers to the distribution of electrons in the energy levels of an atom.

Example: electron arrangement for the first 20 elements.

Element	Symbol	Atomic number	No. of electrons	Electron arrangement
Hydrogen				
Helium				
Lithium				
Beryllium				
Boron				
Carbon				
Nitrogen				
Oxygen				
Fluorine				
Neon				
Sodium				
Magnesium				
Aluminium				
Silicon				

Phosphorus				
Sulphur				
Chlorine				
Argon				
Potassium				
Calcium				

Dot and cross diagrams.

- Is a diagrammatic representation of the electron arrangements in an atom in which the energy levels are represented by concentric lines while electrons are represented by dots or crosses.
- However all electrons are the same regardless of whether they are represented as dots or crosses.

Examples:

Lithium	Magnesium	Aluminium	Carbon
No. of electrons =	No. of electrons	No. of electrons	No. of electrons
No. of protons =	No. of protons	No. of protons	No. of protons
No. of neutrons =	No. of neutrons	No. of neutrons	No. of neutrons

Beryllium	Nitrogen	Chlorine	Argon
No. of electrons =			
No. of protons =			
No. of neutrons =			

Potassium	Boron	Silicon	Phosphorus
No. of electrons =			
No. of protons =			
No. of neutrons =			

The Periodic table

- Is a table showing the arrangements of elements in order of their relative atomic masses.
- Is based on the ideas of Dmitri Ivanovich Medeleev.
- The modern periodic table is based on Meneleev' s periodic law which states that:
 “ The properties of elements are a periodic functions of their relative atomic masses ”
- The modern periodic law itself states that:
 “ The properties of elements are a periodic functions of their atomic numbers ”

Design of the modern periodic table.

It has vertical columns called groups and horizontal rows called periods.

✓ Groups:

- Are the vertical columns of a periodic table.
 - Are eight in number; ad numbered in capital Roman numerals I all through to VIII.
- Note: group VIII is also called group zero because the elements have little tendency to gain o lose electrons during chemical reactions
- Between group 2 and group 3 is a group of elements called the transition metals;

✓ Periods:

- Are the horizontal rows in a periodic table.
- They are 8 in number in a modern periodic table.

✓ Transition metals.

- Are elements that form a shallow rectangle between group II and group III.
- These elements are generally metallic and hence the name “ **transition metals**”
- They are not fitted in any group because they have **variable valencies**.
- Are sometimes called the “ **d-block**” elements.
- They are much less reactive than the elements in groups I and II
- They have some unique characteristics that make them not fit in the 8 groups of the periodic table
- These include:
 - ✓ Have variable valencies; hence show different oxidation states in their compounds;
 - ✓ They form coloured compounds as solids and in aqueous solutions;
 - ✓ Have very high melting and boiling points (than metals in groups I and II).
 - ✓ They do not react with water;
 - ✓ Have very high densities (compared to metals in groups I ad II)

✓ Lanthanides and Actinides.

- They form a block of elements within the transition metals.
- Are sometimes called the **inner transition metals**.
- The lanthanides consist of 14 elements from Cerium (Ce) to Lutetium (Lu).
- The Actinides are the 14 elements from Thorium (Tho) to Lawrencium (Lr).

Placing an elements in the periodic table

- The position of an elements in the periodic table is governed by the atomic number and hence the electron arrangement.

✓ The period:

- The period to which an element belongs is determined by the number of energy levels.
- The number of energy levels is equal to the period to which an elements belongs.

Examples:

Elements	Symbol	Atomic number	Electron arrangement	Number of energy levels	Period
Lithium	Li	3	2.1	2	2
Sodium	Na	11	2.8.1	3	3
Calcium	Ca	20	2.8.8.2	4	4
Nitrogen	N	7	2.5	2	2
Helium	He	2	2	1	1

✓ **Group:**

- The group to which an element belongs to is governed by the number of electrons in the outermost energy level.
- The number of electrons in the outermost energy level is equal to the group to which the element belongs.

Examples:

Elements	Symbol	Atomic number	Electron arrangement	Outermost electrons	Group
Potassium	K	19	2.8.8.1	1	1
Aluminium	Al	13	2.8.3	3	3
Silicon	Si	14	2.8.4	4	4
Oxygen	O	8	2.6	6	6
Chlorine	Cl	17	2.8.7	7	7

Diagram of a full periodic table.

The Periodic Table of the Elements

		Group																																		
		I	II											III	IV	V	VI	VII	0																	
		1 H Hydrogen																2 He Helium																		
		3 Li Lithium		4 Be Beryllium												5 B Boron		6 C Carbon		7 N Nitrogen		8 O Oxygen		9 F Fluorine		10 Ne Neon										
		11 Na Sodium		12 Mg Magnesium												13 Al Aluminium		14 Si Silicon		15 P Phosphorus		16 S Sulphur		17 Cl Chlorine		18 Ar Argon										
GROUP NUMBER	19 K Potassium		20 Ca Calcium		21 Sc Scandium		22 Ti Titanium		23 V Vanadium		24 Cr Chromium		25 Mn Manganese		26 Fe Iron		27 Co Cobalt		28 Ni Nickel		29 Cu Copper		30 Zn Zinc		31 Ga Gallium		32 Ge Germanium		33 As Arsenic		34 Se Selenium		35 Br Bromine		36 Kr Krypton	
	37 Rb Rubidium		38 Sr Strontium		39 Y Yttrium		40 Zr Zirconium		41 Nb Niobium		42 Mo Molybdenum		43 Tc Technetium		44 Ru Ruthenium		45 Rh Rhodium		46 Pd Palladium		47 Ag Silver		48 Cd Cadmium		49 In Indium		50 Sn Tin		51 Sb Antimony		52 Te Tellurium		53 I Iodine		54 Xe Xenon	
	55 Cs Caesium		56 Ba Barium		57 La Lanthanum		72 Hf Hafnium		73 Ta Tantalum		74 W Tungsten		75 Re Rhenium		76 Os Osmium		77 Ir Iridium		78 Pt Platinum		79 Au Gold		80 Hg Mercury		81 Tl Thallium		82 Pb Lead		83 Bi Bismuth		84 Po Polonium		85 At Astatine		86 Rn Radon	
	87 Fr Francium		88 Ra Radium		89 Ac Actinium																															
	103-108 Actinoid series																																			
109-116 Lanthanoid series																																				
117 Ts Tennessine		118 Og Oganesson																																		

Key

A	X	Z
3	86	118
A = relative atomic mass		
X = atomic symbol		
Z = proton (atomic) number		

Note: In the modern periodic table atomic masses are used instead of mass numbers. The atomic masses are preferable because they take care of elements with isotopes unlike mass numbers.

Diagram: part of the periodic table showing the first 20 elements

Relative Atomic Mass and Isotopes.

✓ Introduction:

- The masses of individual atoms of elements are very negligible and thus quite difficult to weigh.
- On average the mass of an atom is approximately 10^{-22} g which cannot be determined by an ordinary laboratory balance.
- For this reason the mass of atom has been expressed relative to that of a chosen standard element hence the term relative atomic mass.
- The initial reference element was hydrogen which was later replaced with oxygen.
- Later the oxygen scale was found unsuitable;

Reason:

- Oxygen exists in several isotopes and thus led to problems when deciding the mass of an oxygen atom.
- For this reason oxygen was replaced with carbon as the reference atom and to date relative atomic masses of elements are based on an atom of **carbon-12** (note that carbon is isotopic and exists as Carbon -12 or carbon-14).

✓ Definition:

Relative atomic mass (R.A.M) of an element refers to the average mass of an atom of the element compared with a twelfth (1/12) of an atom of carbon-12.

$$\text{RAM} = \frac{\text{Average mass of one atom of an element}}{\frac{1}{12} \text{ mass of an atom of carbon-12}}$$

✓ Measurement of Relative atomic mass

- RAM of elements is determined by an instrument called Mass Spectrometer.
- The instrument can also be used to determine the relative abundance of isotopes.
- The use of a mass spectrometer in determining the RAM of elements is called mass spectrometry.

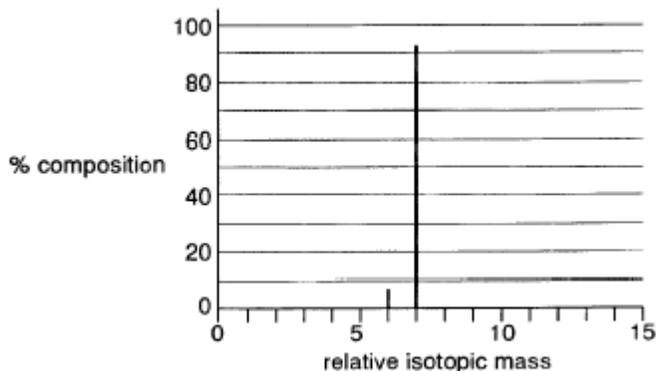
✓ How mass spectrometry works.

- In the mass spectrometer atoms and molecules are converted into ions.
- The ions are then separated as a result of the deflection which occurs in a magnetic field.
- Each ion (from an atom, isotope or molecule) gives a deflection which is amplified into a trace.
- The height of each peak measures the relative abundance of the ion which gives rise to that peak.

Note:

- Generally the relative atomic mass of an element is closest in value to the mass of the most abundant isotope of the element.

Example: Diagram of a spectrometer trace for Lithium



✓ Explanations:

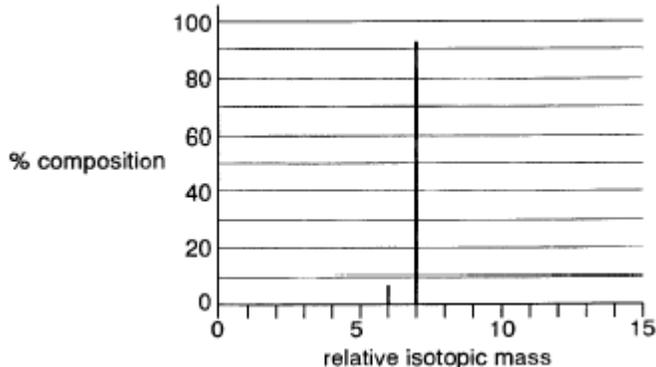
- The trace has two peaks indicating that there are two isotopes for lithium.
- The first peak occurs at a relative isotopic mass of 6 and the second at 7; these are the RAM of the two isotopes respectively.
- The percentage abundance of the isotope with RAM of 6 (${}^6\text{Li}$) is 9 while the RAM of the isotope with RAM 7 (${}^7\text{Li}$) is 91.

Calculating relative atomic masses of isotopic elements.

- Information from a spectrometer trace is usually extracted and used in calculation the relative atomic mass of elements.

Worked examples.

1. The mass spectrum below shows the isotopes present in a sample of lithium.



(i). Use this mass spectrum to help you complete the table below for each lithium isotope in the sample. (3 marks)

Isotope	Percentage composition	Number of	
		Protons	Neutrons
${}^6\text{Li}$			
${}^7\text{Li}$			

(ii). Calculate the relative atomic mass of this lithium sample. Your answer should be given to three significant figures. (3 marks)

2. Element X with atomic number 16 has two isotopes. $\frac{2}{3}$ of ${}^{33}\text{X}$ and $\frac{1}{3}$ of ${}^{30}\text{X}$. What is the relative atomic mass of element X? (2 marks)

3. Calculate the relative atomic mass of an element whose isotopic masses and relative abundances are shown below. (2 marks)

Relative abundance	Isotopic mass.
69	63
31	65

4. A neutral atom of silicon contains 14 electrons, 92% of silicon - 28, 5% silicon – 29 and 3% silicon – 30

(i). What is the atomic number of silicon? (1mark)

(ii). Calculate the relative atomic mass of silicon. (1mark)

5. Oxygen exists naturally as isotopes of mass numbers 16, 17 and 18 in the ratio 96:2:2 respectively. Calculate its R.A.M (2 marks)

6. Calculate the relative atomic mass of potassium from the isotopic composition given below.

Isotope	Relative abundance
^{39}K	93.1
^{40}K	0.01
^{41}K	6.89

7. Sulphur and sulphur compounds are common in the environment.

(a). A sample of sulphur from a volcano contained 88% by mass of ^{32}S and 12% by mass of ^{34}S .

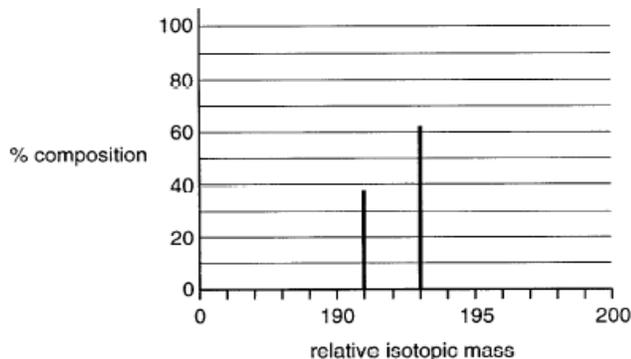
(i). Complete the table below to show the atomic structure of each isotope of sulphur.

Isotope	Number of		
	Protons	Neutrons	Electrons
^{32}S			
^{34}S			

(ii). Define relative atomic mass. (2 marks)

(iii). Calculate the relative atomic mass of the volcanic sulphur. (2 marks)

8. Iridium, atomic number 77, is a very dense metal. Scientists believe that meteorites have deposited virtually all the iridium present on earth. A fragment of a meteorite was analysed using a mass spectrometer and a section of the mass spectrum showing the isotopes present in iridium is shown below.



(a). Explain the term isotopes. (1mark)

(b). Use the mass spectrum to help you complete the table below for each iridium isotope in the meteorite.

Isotope	Percentage composition	Number of	
		Protons	Neutrons
${}^6\text{Ir}$			
${}^7\text{Ir}$			

(c) (i). Define the term relative atomic mass.

(1 mark)

(ii). Calculate the relative atomic mass of the iridium in this meteorite.

(3 marks)

Ion formation.

✓ Introduction:

- Atoms whose outermost energy levels contain the maximum possible number of electrons are said to be stable.
- Thus atoms with energy levels 2, 2.8 and 2.8.8 are said to be stable.
- Electron configuration 2 is said to have a stable duplet state while electron configuration 2.8 and 2.8.8 is said to have a stable octet state.
- These electron configurations resemble those of noble gases and as such they are stable and do not react.
- Atoms without this stability acquire it by either electron gain or electron loss.
- Whether an atom loses or gains electro(s) depend on the number of electrons in the outermost energy level.

✓ Illustration.

- Take the case of sodium.
- Atomic number is 11 with an electron configuration of 2.8.1.
- Thus sodium has two options in to become stable:
 - ✓ to lose the single electron and acquire a stable electron configuration of 2.8.
 - ✓ to gain 7 electrons in its outermost energy level and acquire a stable electron configuration of 2.8.8
- Gaining a single electrons and losing a single electrons requires equal amounts of energy.
- Thus it is cheaper and faster in terms of energy for sodium to lose the single electron in the outermost energy level than to gain 7 electrons into its outermost energy level.
- Thus sodium acquires a stable electron configuration 2.8 by losing the single electron in its outermost energy level.

Diagram

Sodium atom

Sodium ion.

Equation:



Further examples:

Element	Electron arrangement	Options for stability	Best (cheapest) option
Chlorine	2.8.7	2.8 or 2.8.8	2.8.8
Potassium	2.8.8.1	2.8.8 or 2.8.8.8	2.8.8
Aluminium	2.8.3	2.8 or 2.8.8	2.8
Magnesium	2.8.2	2.8 or 2.8.8	2.8
Carbon	2.4	2 or 2.8	2. or 2.8
Oxygen	2.6	2. or 2.8	2.8

✓ **Ion:**

Definition: an ion is a charged particle of an element.

- Are formed when an atom of an element either loses or gains electrons.

✓ **Illustration:**

- For a neutral atom the number of electrons in the energy levels (negative charges) is equal and thus completely balances the number of protons in the nucleus (positive charges).
- Thus the net charge in a neutral atom is zero (0).
- When an atom gains electron(s), the number of electrons becomes higher than the number of protons resulting to a net negative charge hence an ion.
- Oppositely when an atom loses electron(s) the number of protons becomes higher than the number of electrons resulting into a net positive charge hence an ion.
- The charge on the ion is usually indicated as a superscript to the right of the chemical symbol.
- Thus ions are of two types:
 - Cations
 - Anions

✓ **Cations:**

- Are positively charged ions.
- Are formed when atoms lose electrons resulting into the number of protons being higher than the number of electrons.
- Are mostly ions of metallic elements since most metals react by electron loss.

Examples:

(i). Magnesium:

- It has atomic number 12, with electron arrangement 2.8.2.
- It has 12 protons and 12 electrons hence a net charge of “ 0” hence the atom is written simply as Mg.
- It will form its ions by losing the two electrons from the outermost energy level.
- Thus the number of electrons decreases to 10 while the number of protons remains 12.
- This leads to a net charge of +2, giving the ion with the formula Mg^{2+} .

✓ **Diagrammatic illustration:**

(ii). Phosphorus

- It has atomic number 13, with electron arrangement 2.8.3.
- It has 13 protons and 13 electrons hence a net charge of “ 0” hence the atom is written simply as Al.
- It will form its ions by losing three (3) electrons out of the outermost energy level.
- Thus the number of electrons decreases by three to 10 while the number of protons remains 13.
- This leads to a net charge of +3, giving the ion with the formula Al^{+3} .

✓ **Diagrammatic illustration:**

✓ **Anions:**

- Are negatively charged ions.
- Are formed when atoms gain electrons resulting into the number of electrons being higher than the number of protons.
- Are mostly ions of non-metallic elements since most non-metals ionize (react) by electron gain.

Examples.

(i). Chlorine:

- It has atomic number 17, with electron arrangement 2.8.7.
- It has 17 protons and 17 electrons hence a net charge of “ 0” hence the atom is written simply as Cl.
- It will form its ions by gaining a single electron into the outermost energy level.
- Thus the number of electrons increases to 18 while the number of protons remains 17.
- This leads to a net charge of -1, giving the ion with the formula Cl⁻.

✓ **Diagrammatic illustration:**

Note: the electrons gained must be represented by a different notation from the initial electrons in the atom. E.g. if the initial electrons are represented with crosses (x) then the gained electrons should be represented by dots (.) and vice versa.

(ii). Phosphorus

- It has atomic number 15, with electron arrangement 2.8.5.
- It has 15 protons and 15 electrons hence a net charge of “ 0” hence the atom is written simply as P.
- It will form its ions by gaining three (3) electrons into the outermost energy level.
- Thus the number of electrons increases by three to 18 while the number of protons remains 15.
- This leads to a net charge of -3, giving the ion with the formula P⁻³.

✓ **Diagrammatic illustration:**

Electron transfer during chemical reactions.

- Atoms react either by electron gain or electron loss.
- Generally metals react by electron gain while non-metals react by electron loss.

✓ **Illustration:**

- Consider the reaction between sodium and chlorine.
- Sodium attains stability // reacts by losing the single electron from its outermost energy level.
- Chlorine attains stability // reacts by gaining a single electron into its outermost energy level.
- Thus during the reaction between the two elements the single electron lost by the sodium atom to form the sodium ion is the same one gained by the chlorine atom to form the chloride ion.

Some basic concepts.

✓ **Valence electrons:**

- Refers to the number electrons in the outermost energy level.

Examples:

- Calcium, with electron arrangement 2.8.8.2 has 2 valence electrons.
- Oxygen with electron arrangement 2.6 has 6 valence electrons.
- Phosphorus with electron arrangement 2.8.5 has 5 valence electrons.

✓ **Valency:**

- Refers to the number of electrons an atom loses or gains during a chemical reaction.
- Valency is also known as the **combining power** of an element.

Examples:

- Calcium, with electron arrangement 2.8.8.2 loses 2 electrons during chemical reactions and hence has a valency of 2.
- Oxygen with electron arrangement 2.6 gains 2 electrons during chemical reactions and thus has a valency of 2.
- Phosphorus with electron arrangement 2.8.5 has gains 3 electrons during chemical reactions and hence has a valency of 3.
- Aluminium, with electron arrangement 2.8.3 loses 3 electrons during chemical reactions and hence has a valency of 3.

Note: Some elements have variable valencies and are usually termed the transitional elements (metals)

Examples:

- Iron can have valency 2 or 3;
- Copper can have valency 1 or 2
- Lead can have valency 2 or 4.

Summary on valencies of common elements.

	Valency 1	Valency 2	Valency 3
Metals	Sodium Potassium	Calcium Barium Magnesium Zinc Iron Lead Copper	Aluminium Iron
Non-metals	Nitrogen Chlorine Fluorine Hydrogen	Nitrogen Oxygen Sulphur	Nitrogen Phosphorus

Radicals.

- Are groups of atoms with a net charge that exist and react as a unit during chemical reactions.
- Radicals also have a valency, which is equivalent to the value of its charge.

Summary on valencies of some common radicals.

	Valency 1	Valency 2	Valency 3
Radicals	Ammonium (NH_4^+) Hydroxide (OH^-) Nitrate (NO_3^-) Hydrogen carbonate (HCO_3^-) Hydrogen sulphate (HSO_4^-)	Carbonate (CO_3^{2-}) Sulphate (SO_4^{2-}) Sulphite (SO_3^{2-})	Phosphate (PO_4^{3-})

Oxidation number.

- Refers to the number of electrons an atom loses or gains during a chemical reaction.
- In writing the oxidation number the sign (+ or -) to show gain or loss is written followed by the number of electrons lost or gained respectively.

Illustration.

- Atoms are electrically neutral and are thus assigned an oxidation state of 0 since the number of protons in the nucleus is equal to the number of electrons in the energy levels.
- However when atoms react they either lose or gain electrons and thus acquire a new state.
- This new state is a new oxidation state and the atom thus acquires a new oxidation number

Examples:

Atom	E. arrangement	Ion formula	Valency	Oxidation number
Sodium	2.8.1	Na^+	1	+1
Magnesium	2.8.2	Mg^{2+}	2	+2
Aluminium	2.8.3	Al^{3+}	3	+3
Nitrogen	2.5	N^{3-}	3	-3
Sulphur	2.8.6	S^{2-}	2	-2
Chlorine	2.8.7	Cl^-	1	-1

Further examples:

Particle	Oxidation number
Copper metal, Cu	0
Lead (II) ion, Pb ²⁺	+2
Bromide ion, Br ⁻	-1
Aluminium ion, Al ²⁺	+2
Sulphide ion, S ²⁻	-2
Magnesium metal, Mg	0

✓ **Note: oxidation number (state) and charge of an element.**

- **Oxidation state** is written with the positive or the negative sign coming before the element.

Examples: -2, 3, +1, -1 etc.

- **Charge** on an element is written as a superscript of the element with the number coming before the positive or the negative sign

Examples: Mg²⁺, Al³⁺, Na⁺, Cl⁻ etc.

Chemical formulae.

- Refers to a representation of a chemical substance using chemical symbols.

- In a single atom it is equivalent to the chemical symbol of the element.

- In a compound it shows the constituent elements and the proportions in which they are combined.

✓ **Deriving the chemical formula of compounds.**

- In order to write the correct formula of a compound the following must be known:

- The symbols of the constituent elements or radicals.
- The valencies of the elements or radicals

- The chemical formula should start with the element which is more likely to lose electron (s) followed by the element that is more likely to gain.

Worked examples.

1. Deriving the formula of sodium chloride.

Elements	Sodium	Chlorine
Formula	Na	Cl
valencies	1	1
balancing	x1	x1

Balancing ratios as subscripts: Na₁Cl₁

Formula: NaCl

Explanation:

- For sodium to combine with chlorine to form sodium chloride, sodium loses an electron while chlorine gains an electron.

- Thus every sodium atom needs only a single chlorine atom for both to be fully stable

Note:

- When the balancing ratio // subscript is 1, it is usually not written since the symbol of the element itself represents a single atom.

2. Deriving the formula of magnesium chloride.

Elements	Magnesium	Chlorine
Formula	Mg	Cl
valencies	2	1
balancing	X1	X2

Balancing ratios as subscripts: Mg_1Cl_2

Formula: $MgCl_2$

3. Deriving the formula of magnesium oxide.

Elements	Magnesium	Oxygen
Formula	Mg	O
valencies	2	2
balancing	X1	X1

Balancing ratios as subscripts: Mg_1O_1

Formula: MgO

Further worked examples.

1. Derive the chemical formula of each of the following compounds.

- ✓ Calcium fluoride
- ✓ Carbon (II) oxide
- ✓ Carbon (IV) oxide
- ✓ Aluminium nitrate
- ✓ Calcium hydrogen carbonate

2. Complete the table below for elements A, B and C

Element	Valency	Chemical formula of various compounds					
		hydroxides	Sulphates	carbonates	nitrates	phosphates	Hydrogen carbonates
A	1						
B	2						
B	3						

Chemical equations.

- Refers to representations of a chemical reaction by means of chemical symbols and formula.

✓ Key features of a chemical equation.

- The correct formulae of the reactants are on the left of the equation.
- The correct formulae of the products are on the right of the equation.
- The reactants and products are separated by an arrow pointing to the right.
- The state symbols of the reactants and products must be stated as subscripts to the right of the symbols
- The number of each atom on the reactants side must be equal to the number of the same atom on the products side.

Example.

Reaction between hot copper metals and oxygen gas.

✓ **Word equation:** Copper + oxygen → Copper (II) oxide.

✓ **Chemical equation:** $2\text{Cu}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{CuO}_{(s)}$.

Balancing chemical equations.

- A chemical equation is only valid if it is balanced.
- A chemical equation is said to be balanced if the number of each atom on the reactants side is equal to that on the products side.
- This is because atoms are neither created nor destroyed during a chemical reaction.

Rules and guidelines in balancing chemical equations.

✓ *Step I: Write the chemical equation in words.*

Example: Copper metal + oxygen gas.

✓ *Step II: Write the correct formulae of both reactants and products*

Example: $\text{Cu} + \text{O}_2 \rightarrow \text{CuO}$

✓ *Step III: Check whether the number of atoms of each element on the reactants side is equal to that on the products side.*

- If equal proceed to step (V);
- If not equal proceed to step (IV).

Example: $\text{Cu} + \text{O}_2 \rightarrow \text{CuO}$

- In this case there are two oxygen atoms on the reactants side yet there is only one oxygen atom on the products side. Thus we proceed to step IV

✓ *Step IV: Multiply the chemical formula containing the unbalanced atoms with the lowest common multiple.*

Example: $\text{Cu} + \text{O}_2 \rightarrow \text{CuO}$

- In this case the chemical formula with the unbalanced atom is CuO on the products side. We thus multiply it by 2.
- The new equation now reads: **$\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$**

✓ *Step V: check again to ensure that all atoms are balanced.*

- If all atoms are balanced proceed to step VI.
- If not then repeat step IV until all atoms are balanced.

Example: $\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$

- In this case multiplying CuO by 2 offsets the balancing of Cu; which is now unbalanced!
- We therefore repeat step IV in order to balance Cu.
- There is only 1 Cu atom on the reactants side yet there are 2 Cu atoms on the products side.
- We thus multiply the formula with the unbalanced atom (s) by the lowest common multiple, in this case 2.
- The new equation at this step thus becomes: $2\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$
- We then repeat step V; in this case all atoms are now balanced.

✓ *Step VI: The physical states of the reactants and the products are then indicated.*

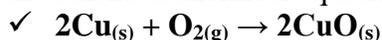
- If this is not done the chemical equation is considered incorrect.

✓ **Types of state symbols.**

- There four main state symbols.
 - ✓ Solid; denoted as **(s)**
 - ✓ Liquid; denoted as **(l)**
 - ✓ Aqueous (in solution in water); denoted as **(aq)**
 - ✓ Gaseous; denoted as **(g)**
- In a chemical equation the state symbols are written with their denotations as subscripts to the right of the chemical formulae.

Example: $2\text{Cu}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{CuO}_{(s)}$

Thus the balanced chemical equation for the reaction between copper metal and oxygen is:



Worked examples:

1. Balance equations for each of the following reactions.

- ✓ Sodium hydroxide and dilute hydrochloric acid
- ✓ Zinc oxide and dilute sulphuric (VI) acid
- ✓ Zinc metal and dilute nitric (V) acid
- ✓ Calcium carbonate and dilute sulphuric (VI) acid
- ✓ Sodium and water

2. Balance each of the following equations.

- ✓ $\text{Mg}_{(s)} + \text{HCl}_{(aq)} \rightarrow \text{MgCl}_{2(aq)} + \text{H}_{2(g)}$
- ✓ $\text{Na}_{(s)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{NaOH}_{(aq)} + \text{H}_{2(g)}$
- ✓ $\text{NaOH}_{(aq)} + \text{H}_2\text{SO}_{4(aq)} \rightarrow \text{Na}_2\text{SO}_{4(aq)} + \text{H}_2\text{O}_{(l)}$
- ✓ $\text{CuCO}_{3(s)} + \text{HNO}_{3(aq)} \rightarrow \text{Cu}(\text{NO}_3)_{2(aq)} + \text{CO}_{2(g)} + \text{H}_2\text{O}_{(l)}$

- ✓ $\text{H}_2\text{S}_{(g)} + \text{O}_{2(g)} \rightarrow \text{O}_{2(g)} + \text{H}_2\text{O}_{(l)}$
- ✓ $\text{C}_2\text{H}_{6(g)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + \text{H}_2\text{O}_{(l)}$
- ✓ $\text{Pb}(\text{NO}_3)_{2(s)} \rightarrow \text{PbO}_{(s)} + \text{NO}_{2(g)} + \text{O}_{2(g)}$
- ✓ $\text{Fe}_{(s)} + \text{Cl}_{2(g)} \rightarrow \text{FeCl}_{3(s)}$
- ✓ $\text{Al}_{(s)} + \text{H}_2\text{SO}_{4(aq)} \rightarrow \text{Al}_2(\text{SO}_4)_{3(aq)} + \text{H}_{2(g)}$

UNIT 2: CHEMICAL FAMILIES; PATTERNS AND PROPERTIES.

Checklist.

- Meaning of chemical families.

- Main chemical families;

✓ Alkali metals

- Meaning and members
- Trends down the group
 - Atomic radius
 - Ionic radius
 - Energy levels
- Physical properties
 - Appearance
 - Ease of cutting
 - Melting and boiling [points
 - Electrical conductivity
 - 1st ionization energy
- Chemical properties
 - Burning in air
 - Exposure to air
 - Reaction with water
 - Reaction with chlorine
- Similarity of ions and formula of compounds of alkali metals
- Uses of alkali metals.

✓ Alkaline earth metals

- Meaning and members
- Trends down the group
 - Atomic radius
 - Ionic radius
 - Energy levels
- Physical properties
 - Melting and boiling [points
 - Electrical conductivity
 - 1st ionization energy
 - 2nd ionization energy
- Chemical properties
 - Reaction with air and water
 - Reaction with steam
 - Reaction with chlorine
 - Reaction with dilute acids
- Similarity of ions and formula of compounds of alkaline earth metals
- Uses of alkaline earth metals.

✓ The Halogens

- Meaning and members
- Trends down the group
 - Atomic radius
 - Ionic radius
 - Energy levels
- Physical properties
 - Preparation and properties of chlorine
 - Appearance and physical states of halogens at room temperature
 - Melting and boiling points
 - Electrical conductivity
- Chemical properties
 - Ion formation
 - Electron affinity
 - Reaction with metals

- Reaction with water
 - Reactivity trend of halogens.
 - Similarity of ions and formula of compounds of alkali metals
 - Uses of alkali metals.
- ✓ The noble gases
 - Meaning and members
 - Trends down the group
 - Atomic radius
 - Ionic radius
 - Energy levels
 - Physical properties
 - Melting and boiling [points
 - Electrical conductivity
 - 1st ionization energy
 - Chemical properties
 - Uses of the noble gases

Introduction:

- Elements are classified and hence positioned in the periodic table based on the number of valence electrons and the number of energy levels.
- The number of valence electrons is equal to the group to which the element belongs; while the number of energy levels is equal to the period to which the element belongs.
- Elements in the **same group** are said to belong to the same **chemical family**.

- Thus a chemical family refers to a group of elements in the same number of valence electrons and hence in the same group of the periodic table.

Characteristics of a chemical family:

- have same number of valence electrons;
- Show a uniform gradation in physical properties;
- have similar chemical properties;

Main chemical families.

- Four main chemical families will be studied in this section.
 - ✓ The Alkali metals
 - ✓ The Alkaline earth metals
 - ✓ The halogens
 - ✓ The noble gases

1. The Alkali metals.

- Are the elements with one valence electron and hence in group I of the periodic table.
- All are metallic in nature.
- The members of the family in order down the group is as follows:
 - ✓ Lithium
 - ✓ Sodium
 - ✓ Potassium
 - ✓ Rubidium
 - ✓ Caesium
 - ✓ Francium

Electron arrangement of the first three alkali metals.

Elements	Electron arrangement
Lithium	2.1
Sodium	2.8.1
Potassium	2.8.8.1

Diagram: Part of periodic table showing the alkali metals

Gradation in properties of alkali metals.

1. Atomic and ionic radius.

✓ **Atomic radius:**

- Refers to the distance between the centre of the nucleus of an atom and the outermost energy level occupied by electron(s)

✓ **Ionic radius:**

- Refers to the distance between the centre of the nucleus of an ion and the outermost energy level occupied by electron(s)

Trend:

- The ionic radius and the atomic radius of alkali metals increase down the group.

Reason:

- There is an increase in the number of energy levels down the group from lithium to Francium.

Illustration:

- Lithium (2.1) has only 2 energy levels; sodium (2.8.1) has 3 energy levels while potassium (2.8.8.1) has 4 energy levels.

- Thus the outermost electron in potassium is further from the nucleus than the outermost electron in sodium and lithium.

2. Atomic and ionic radius of the same element.

- For the same alkali metals the atomic radius is larger than the ionic radius.

Reason:

- Alkali metals form ions by losing the valence electron, leading to the loss of an entire outermost energy level. Thus the atoms have more energy levels than the corresponding ion hence a larger radius in the atom than in the ion.

Illustration

- Potassium atom has electron arrangement of 2.8.8.1 hence 4 energy levels.

- During ion formation potassium reacts by losing the single valence electron to acquire a new electron arrangement of 2.8.8 hence 3 energy levels.

- Thus the ion has a smaller radius than the atom.

Diagrammatically: Potassium atom and potassium ion

Summary: changes in atomic and ionic radius among alkali metals.

Element	Symbol	Atomic number	Atomic radius (nm)	Ionic radius (nm)
Lithium	Li	3	0.133	0.060
Sodium	Na	11	0.157	0.095
Potassium	K	19	0.203	0.133

Physical properties of alkali metals.

1. Appearance.

- Alkali metals have metallic luster when freshly cut. This refers to a shiny appearance on the cut surface.
- This surface however tarnishes due to reaction on exposure to air.

2. Ease of cutting.

- They are soft and easy to cut.
- The softness and ease of cutting increase down the group.

Reason:

- Alkali metals have giant metallic structures held together by metallic bonds.
- Metallic bond is due to attraction between the positively charged nucleus of one atom and the electrons in the outermost energy level of the next atom.
- Thus the force of attraction is stronger in smaller atoms than in larger atoms.
- The increase in atomic radius down the group implies that the strength of metallic bonds also decrease down the group (hence ease of cutting and softness).

3. They have relatively low melting and boiling points (in comparison to other metals).

Reason: they have relatively weaker metallic bonds.

- The melting and boiling points decrease down the group.

Reason: - The size of the atoms increase down the group due to increasing number of energy levels hence decrease in the strength of the metallic bonds (down the group).

4. Electrical conductivity.

- Alkali metals are good conductors of heat and electricity.

Reason: they have delocalized electrons in the outermost energy level.

- The electrical conductivity is similar for all alkali metals.

Reason: all alkali metals have the same number of delocalized electrons (a single electron) in the outermost energy level.

Note:

- In metals the electrons in the outermost energy level do not remain in one fixed position. They move randomly throughout the metallic structure and are thus said to be delocalised.

5. Ionization energy.

Ionization energy is the minimum energy required to remove an electron from the outermost energy level of an atom in its gaseous state.

- The number of ionization energies an element may have is equivalent to the number of valence electrons.
- The **first ionization energy** is the minimum energy required to remove the first electron from the outermost energy level of an atom in its gaseous state.
- The first ionization among alkali metals decreases down the group.

Reason:

- The effective force of attraction of on the outermost electron by the positive nucleus decreases with increasing atomic size and distance from the nucleus.
- Note that the atomic radius increases down the group due to increase in the number of energy levels.

Summary on physical properties of Alkali metals.

Element	Appearance	Ease of cutting	Melting point (°C)	Boiling point (°C)	Electrical conductivity	Atomic radius (nm)	Ionic radius (nm)
Lithium	Silver white	Slightly hard	180	1330	Good	0.133	520
Sodium	Shiny white	Easy	98	890	Good	0.157	496
Potassium	Shiny grey	Easy	64	774	Good	0.203	419

Chemical properties of alkali metals.

1. Reaction with air.

- When exposed to air alkali metals react with atmospheric moisture to form the corresponding metal hydroxide and hydrogen gas.

✓ **General equation:**

Metal + Water → Metal hydroxide + hydrogen gas.

- The metal hydroxide further reacts with atmospheric carbon (IV) oxide to form hydrated metal carbonate.

✓ **General equation:**

Metal hydroxide + carbon (IV) oxide → Hydrated metal carbonate.

Examples:

✓ **Lithium**

With moisture: $2\text{Li}_{(s)} + 2\text{H}_2\text{O}_{(l)} \rightarrow 2\text{LiOH}_{(aq)} + \text{H}_{2(g)}$;

Then with carbon (IV) oxide: $2\text{LiOH}_{(aq)} + \text{CO}_{2(g)} \rightarrow \text{Li}_2\text{CO}_3 \cdot \text{H}_2\text{O}_{(s)}$;

✓ **Sodium**

With moisture: $2\text{Na}_{(s)} + 2\text{H}_2\text{O}_{(l)} \rightarrow 2\text{NaOH}_{(aq)} + \text{H}_{2(g)}$;

Then with carbon (IV) oxide: $2\text{NaOH}_{(aq)} + \text{CO}_{2(g)} \rightarrow \text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}_{(s)}$;

✓ **Potassium**

With moisture: $2\text{K}_{(s)} + 2\text{H}_2\text{O}_{(l)} \rightarrow 2\text{KOH}_{(aq)} + \text{H}_{2(g)}$;

Then with carbon (IV) oxide: $2\text{KOH}_{(aq)} + \text{CO}_{2(g)} \rightarrow \text{K}_2\text{CO}_3 \cdot \text{H}_2\text{O}_{(s)}$;

2. Burning in air.

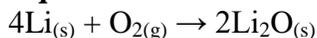
- Alkali metals burn in air with characteristic flame colours to form corresponding metal oxides.

Examples:

✓ **Lithium.**

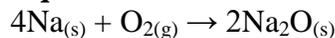
- Burns in air to form lithium oxide as the only product.

Equation:

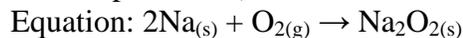


✓ **Sodium.**

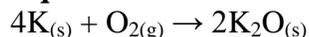
- Burns in air to with a **yellow flame** to form sodium oxide as the only product.

Equation:**Note:**

- When burned in air enriched with oxygen or pure oxygen sodium burns with a yellow flame to form sodium peroxide (instead of sodium oxide).

**✓ Potassium.**

- Burns in air with a **lilac flame** to form potassium oxide as the only product.

Equation:**3. Reaction with water.**

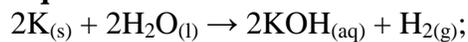
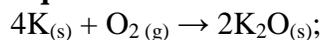
- Alkali metals react with water to form the corresponding hydroxides and hydrogen gas.

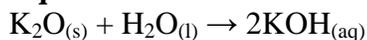
Examples:**✓ Potassium.****Procedure:**

- A small piece of potassium metal is cut and dropped into a trough containing water;
- The resultant solution is tested with litmus paper;

Diagram of apparatus:**Observations and explanations:**

- The metal floats on the water surface; because it is less dense than water;
- A hissing sound is produced; due to production of hydrogen gas;
- It **explosively** melts into a silvery ball then disappears because reaction between water and potassium is exothermic (produces heat). The resultant heat melts the potassium due to its low melting point.
- It darts on the surface; due to propulsion by hydrogen;
- The metal bursts into a lilac flame; because hydrogen explodes into a flame which then burns the small quantities potassium vapour produced during the reaction;
- The resultant solution turns blue; because potassium hydroxide solution formed is a strong base;

(b). Reaction equations.**Equation I****Equation II**

Equation III:**Effect of resultant solution on litmus paper;**

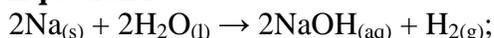
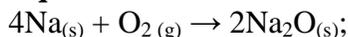
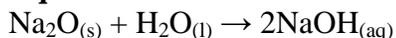
- Litmus paper turns blue; sodium hydroxide formed is highly soluble in water; releasing a large number of hydroxyl ions which result into alkaline conditions // high pH;

✓ Sodium.**Procedure:**

- A small piece of sodium metal is cut and dropped into a trough containing water;
- The resultant solution is tested with litmus paper;

Diagram of apparatus:**Observations and explanations:**

- The metal floats on the water surface; because it is less dense than water;
- A hissing sound is produced; due to production of hydrogen gas;
- It **vigorously** melts into a silvery ball then disappears because reaction between water and sodium is exothermic (produces heat). The resultant heat melts the sodium due to its low melting point.
- It darts on the surface; due to propulsion by hydrogen;
- The metal may burst into a golden yellow flame; because hydrogen may explode into a flame which then burns the sodium;
- The resultant solution turns blue; because sodium hydroxide solution formed is a strong base;

(b). Reaction equations.**Equation I****Equation II****Equation III:****Effect of resultant solution on litmus paper;**

- Litmus paper turns blue; sodium hydroxide formed is highly soluble in water; releasing a large number of hydroxyl ions which result into alkaline conditions // high pH;

✓ **Lithium.**

Procedure:

- A small piece of lithium metal is cut and dropped into a trough containing water;
- The resultant solution is tested with litmus paper;

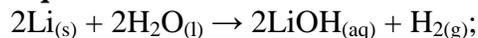
Diagram of apparatus:

Observations and explanations:

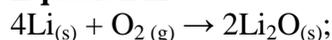
- The metal floats on the water surface; because it is less dense than water;
- A hissing sound is produced; due to production of hydrogen gas;
- It reacts **less vigorously** than sodium and does not melt since the melting point of lithium is relatively higher.
- It darts on the surface; due to propulsion by hydrogen;
- The gas does not ignite spontaneously;
- The resultant solution turns blue; because lithium hydroxide solution formed is a strong base;

(b). Reaction equations.

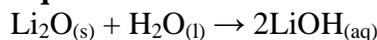
Equation I



Equation II



Equation III:



Effect of resultant solution on litmus paper;

- Litmus paper turns blue; sodium hydroxide formed is highly soluble in water; releasing a large number of hydroxyl ions which result into alkaline conditions // high pH;

Summary on reaction rate.

Metal	Reaction rate
Lithium	Vigorous
Sodium	More vigorous
Potassium	Explosive

4. Reaction with chlorine

- All alkali metals react with chlorine to form corresponding metal chlorides.

✓ **General procedure:**

- A small piece of the alkali metal is cut and placed in a deflagrating spoon;
- It is then warmed and quickly lowered into a gas jar containing chlorine.

✓ **Precautions.**

- The experiment should be done in a working fume chamber or in the open

Reason: Chlorine gas is poisonous.

- This experiment should not be attempted in the laboratory with potassium

Reason: the reaction is too explosive and very dangerous.

✓ **Observations:**

Sodium:

- The metal bursts into a yellow flame.
- White fumes of sodium chloride are formed.

Equation: $2\text{Na}_{(s)} + \text{Cl}_{2(g)} \rightarrow 2\text{NaCl}_{(s)}$

Lithium:

- The metal reacts less vigorously than sodium without bursting into a flame.
- White fumes of lithium chloride are formed.

Equation: $2\text{Li}_{(s)} + \text{Cl}_{2(g)} \rightarrow 2\text{LiCl}_{(s)}$

Potassium:

- The metal bursts into a lilac flame.
- White fumes of potassium chloride are formed.

Equation: $2\text{K}_{(s)} + \text{Cl}_{2(g)} \rightarrow 2\text{KCl}_{(s)}$

The general trend in reactivity of alkali metals.

- The reactivity of alkali metals increase down the group.

Reason:

- Alkali metals react by losing the single valence electron. The ease of loss of the valence electron increases down the group due to decrease in attraction of the valence electron towards the positive nucleus. This in turn is due to the increase in atomic radius down the group as a result of increase in number of energy levels.

Similarity of ions and formulae of some compounds of Alkali metals.

- Alkali metals have a similar charge on their ions since they all have a single valence electron.
- Thus they have the same general formula in their ions and compounds.

Examples.

Alkali metal ion	Hydroxide (OH ⁻)	Oxides (O ²⁻)	Chloride (Cl ⁻)	Sulphates (SO ₄ ²⁻)
Li ⁺				
Na ⁺				
K ⁺				

Note:

- Alkali metals are usually not found as free elements; but rather in their combined states in the earth' s crust.

Reason:

- They have high reactivities.

Uses of Alkali metals and their compounds.

1. Sodium is used in the manufacture of sodium cyanide for gold extraction.
2. Lithium is used in the manufacture of special high strength glasses and ceramics.
3. Lithium compounds are used in the manufacture of dry cells for use in mobile phones, laptops, stopwatches, and zero emission electric vehicles.
4. A molten mixture of sodium and potassium is used as a coolant in nuclear reactors.
5. Sodium vapour is used to produce the yellow glow in street and advertisement lights.
6. Molten sodium is used as a reducing agent in the extraction of titanium.



7. Sodium chloride is used as a food additive.
8. A mixture of sodium hydroxide (caustic soda) and carbon disulphide is used in the manufacture of artificial silk called rayon.

The Alkaline Earth metals.

- Are the elements with two valence electrons and hence in group II of the periodic table.
- All are metallic in nature.
- The members of the family in order down the group is as follows:
 - ✓ Beryllium
 - ✓ Magnesium
 - ✓ Calcium
 - ✓ Strontium
 - ✓ Barium
 - ✓ Radium

Electron arrangement of the first three alkaline earth metals.

Elements	Electron arrangement
Beryllium	2.2
Magnesium	2.8.2
Calcium	2.8.8.2

Diagram: Part of periodic table showing the alkaline earth metals

Dot and cross diagrams for the first three Alkaline Earth metals

Gradation in properties of alkali metals.

1. Atomic and ionic radius.

✓ **Atomic radius:**

- Refers to the distance between the centre of the nucleus of an atom and the outermost energy level occupied by electron(s)

✓ **Ionic radius:**

- Refers to the distance between the centre of the nucleus of an ion and the outermost energy level occupied by electron(s)

Trend:

- The ionic radius and the atomic radius of alkaline earth metals increase down the group.

Reason:

- There is an increase in the number of energy levels down the group from Beryllium to radium.

Illustration:

- Beryllium (2.2) has only 2 energy levels; Magnesium (2.8.2) has 3 energy levels while calcium (2.8.8.2) has 4 energy levels.

- Thus the outermost electron in calcium is further from the nucleus than the outermost electron in magnesium and beryllium.

2. Atomic and ionic radius of the same element.

- For the same alkaline earth metal the atomic radius is larger than the ionic radius.

Reason:

- Alkaline earth metals form ions by losing the valence electrons, leading to the loss of an entire outermost energy level. Thus the atoms have more energy levels than the corresponding ion hence a larger radius in the atom than in the ion.

Illustration

- Calcium atom has electron arrangement of 2.8.8.2 hence 4 energy levels.

- During ion formation it reacts by losing the 2 valence electrons to acquire a new electron arrangement of 2.8.8 hence 3 energy levels.

- Thus the ion has a smaller radius than the atom.

Diagrammatically: calcium atom and calcium ion

Summary: changes in atomic and ionic radius among alkali metals.

Element	Symbol	Atomic number	Electron arrangement	Atomic radius (nm)	Ionic radius (nm)
Beryllium	Be	4	2.2	0.089	0.031
Magnesium	Mg	12	2.8.2	0.136	0.065
Calcium	Ca	20	2.8.8.2	0.174	0.099

Physical properties of alkali metals.

1. Appearance.

- Alkaline earth metals acquire a metallic luster when polished. This refers to a shiny appearance on the cut surface.
- They however lose this metallic luster when exposed to air due to oxidation.

✓ **Note:**

- The purpose of polishing alkaline earth metals before using them in experiment is to remove the oxide coating that slows down and prevents them from reacting.

2. Ease of cutting.

- Magnesium is hard to cut with a knife; but is however **malleable** and **ductile**.
- Calcium cannot also be cut with a knife because it is **brittle**.

Note:

✓ **Malleability:**

- Refers to the ability of a material to be hammered into sheets.

Example: Iron sheets are possible to be made because iron metal is malleable.

✓ **Ductility:**

- The ability of a material to be rolled into wires.

Example: electric cables are made of aluminium because aluminium metal is ductile.

✓ **Brittle:**

- Refers to a substance which is hard and likely to break.

3. They have relatively high melting and boiling points in comparison to alkali metals.

Reason: they have relatively stronger metallic bonds (than alkali metals).

- The melting and boiling points decrease down the group.

Reason: -

- The size of the atoms increase down the group due to increasing number of energy levels. As the atomic radius increase the force of attraction between the positive nucleus and the delocalized electrons decrease. This leads to a decrease in the strength of the metallic bonds (down the group).

4. Electrical conductivity.

- Alkaline earth metals are good conductors of heat and electricity.

Reason: they have delocalized electrons in the outermost energy level.

- The electrical conductivity is similar for all the alkaline earth metals.

Reason: all alkaline earth metals have the same number of delocalized electron (two valence electrons) in the outermost energy level.

5. Ionization energy.

Ionization energy is the minimum energy required to remove an electron from the outermost energy level of an atom in its gaseous state.

- The number of ionization energies an element may have is equivalent to the number of valence electrons.

- Thus alkaline earth metals have two ionization energies.

- The **first ionization energy** is the minimum energy required to remove the first electron from the outermost energy level of an atom in its gaseous state.

- The **second ionization energy** is the minimum energy required to remove the second electron from the outermost energy level of an atom in its gaseous state.

- The first and second ionization energies among alkaline earth metals decreases down the group.

Reason:

- The effective force of attraction of on the outermost electron by the positive nucleus decreases with increasing atomic size and distance from the nucleus.

- Note that the atomic radius increases down the group due to increase in the number of energy levels.

✓ **Variation between 1st and 2nd Ionization energies.**

- The first ionization energy is always lower than the second ionization energy for the same element.

Reason:

- After losing the first electron from an atom, the overall positive charge holds the remaining electrons more firmly. Thus removing a second electron from the ion requires more energy than the first electron

Note:

- The third ionization energy will also be higher than the second ionization energy for the same reason.

Example:

Magnesium.

✓ First ionization energy: $\text{Mg}_{(g)} \rightarrow \text{Mg}^+_{(g)} + e^-$ (1st I.E = 736 kJ per mole)

✓ Second ionization energy: $\text{Mg}^+_{(g)} \rightarrow \text{Mg}^{2+}_{(g)} + e^-$ (2nd I.E = 1450 kJ per mole)

Summary on physical properties of alkaline earth metals.

Element	Atomic number	Melting point (°C)	Boiling point (°C)	Atomic radius (nm)	1 st I.E (kJmol ⁻¹)	2 nd I.E (kJmol ⁻¹)
Beryllium	4	1280	2450	0.089	900	1800
Magnesium	12	650	1110	0.136	736	1450
Calcium	20	850	1140	0.174	590	1150
Strontium	38	789	1330	0.210	550	1060
Barium	56	725	1140	0.220	503	970

Chemical properties of Alkaline earth metals.

1. Burning alkaline earth metals in air.

- Alkaline earth metals react burn in air to form corresponding oxides.

- More reactive alkaline earth metals may also react with atmospheric nitrogen to form corresponding nitrides.

Examples:

✓ **Magnesium.**

- Burns in air with a blinding brilliant flame forming a white solid.

- The white solid is a mixture of magnesium oxide and magnesium nitride.

Equations:

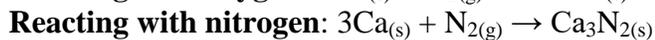
Reacting with oxygen: $2\text{Mg}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{MgO}_{(s)}$

Reacting with nitrogen: $3\text{Mg}_{(s)} + \text{N}_{2(g)} \rightarrow \text{Mg}_3\text{N}_{2(s)}$

✓ **Calcium.**

- Burns in air with a faint orange flame forming a white solid.

- The white solid is a mixture of calcium oxide and calcium nitride.

Equations:**Note:**

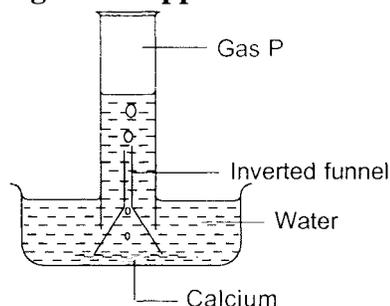
- The trend in the reactivity of alkaline earth metals when burning in air is not clear; due to the oxide coating on the calcium that tends to slow down the reaction of calcium in air.
- For this reason it is important to polish the surfaces of alkaline earth metals before using them in experiments.

2. Reaction of alkaline earth metals with cold water.

- Alkaline earth metals react slowly with cold water to form corresponding hydroxides and hydrogen gas.

Examples✓ **Magnesium:**

- Reacts slowly with water to form magnesium hydroxide and hydrogen gas.
- The reaction is very slow and the amount of hydrogen gas evolved is very low hence the hydrogen gas bubbles stick on the surface of the metal.
- The magnesium hydroxide formed dissolves slightly in water to form magnesium hydroxide.
- Thus the resultant solution is slightly alkaline.

Equation:✓ **Calcium:****Diagram of apparatus:****Observations and explanations:**

- Calcium sinks to the bottom of the beaker; because it is denser than water;
- Slow effervescence of a colourless gas; due to slow evolution of hydrogen gas;
- Soapy solution formed; due to formation of alkaline calcium hydroxide;
- A white suspension is formed; because calcium hydroxide is slightly soluble in water;

Reaction equation:**Effect of resultant solution on litmus paper;**

- Litmus paper slowly turns blue; calcium hydroxide formed is slightly soluble in water; releasing a small number of hydroxyl ions which result into alkaline conditions // high pH;

3. Reaction with steam.

- Alkaline earth metals react with steam to produce corresponding metal oxide and hydrogen gas.
- The reactivity with steam is faster and more vigorous for each alkaline earth metal as compared to reaction with cold water.

Examples:

✓ Magnesium

Procedure:

- A small amount of wet sand is put at the bottom of a boiling tube;
- A small piece of magnesium ribbon is cleaned and put in the middle of the combustion tube;
- The wet sand is heated gently first then the magnesium ribbon is heated strongly until it glows.

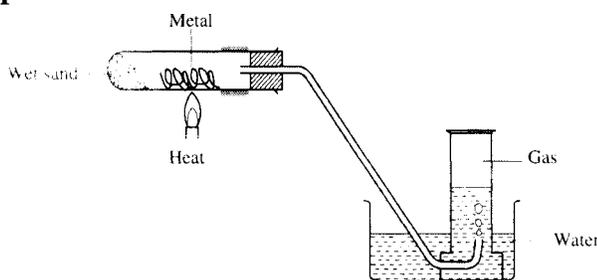
Reason:

- To generate steam that drives out the air that would otherwise react with the magnesium (preventing reaction with steam)
- The delivery tube is removed from the water before heating stops.

Reason:

- To prevent sucking back (of the gas) as the apparatus cools
- The gas produced is tested using a burning splint.

Diagram of apparatus:

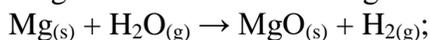


Observations and explanations.

- Magnesium burns with a white blinding flame;
- Grey solid (magnesium) forms a white solid; due to formation of magnesium oxide;
- Evolution of a colourless gas that produces a pop sound when exposed to a burning splint; confirming it is hydrogen;

Reaction equation.

Magnesium + Steam \rightarrow Magnesium oxide + Hydrogen gas;



✓ Calcium

- Reaction between calcium and steam would produce calcium oxide and hydrogen gas.
- However the reaction is too explosive to be done under laboratory conditions.

4. Reaction with chlorine.

- Alkaline earth metals react with chlorine to form corresponding chlorides as the only products.

Condition: presence of heat, hence the metal must be heated first.

Precaution: reaction should be done in a working fume chamber because chlorine gas is poisonous).

Examples:✓ **Magnesium.****Procedure:**

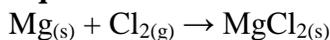
- A piece of burning magnesium is lowered into a gas jar containing chlorine.

Observations:

- The metal continues to burn with a brilliant white flame.
- They grey solid forms a white powder.

Explanation.

- Reaction between magnesium and chlorine is exothermic.
- The heat produced keeps the metal burning; thus facilitates the reaction between magnesium and chlorine to form magnesium chloride, which is the white powder.

Equation:✓ **Magnesium.****Procedure:**

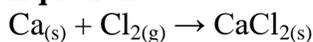
- A piece of burning calcium is lowered into a gas jar containing chlorine.

Observations and explanations.

- The metal burns shortly with an orange flame but soon smolders off.
- There is no steady reaction between calcium and chlorine.

Reason:

- When calcium is heated a coating of the metal oxide is formed first which prevents further reaction between the metal and chlorine.
- However under suitable conditions calcium reacts with chlorine to form a white powder of calcium chloride.

Equation:**5. Reaction of alkaline earth metals and dilute acids**

- Generally alkaline earth metals react with dilute acids to form salts and hydrogen gas.

Examples:✓ **With hydrochloric acid.**▪ **Beryllium:**

- When a piece of beryllium is dropped into a beaker containing hydrochloric acid, there is effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of beryllium chloride remains in the test tube // boiling tube.



▪ **Magnesium:**

- When a piece of magnesium is dropped into a beaker containing hydrochloric acid, there is effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of magnesium chloride remains in the test tube // boiling tube.



▪ **Calcium:**

- When a piece of calcium is dropped into a beaker containing hydrochloric acid, there is effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of calcium chloride remains in the test tube // boiling tube.



✓ **With sulphuric (VI) acid.**

▪ **Beryllium:**

- When a piece of beryllium is dropped into a beaker containing sulphuric (VI) acid, there is effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of beryllium sulphate remains in the test tube // boiling tube.



▪ **Magnesium:**

- When a piece of magnesium is dropped into a beaker containing sulphuric (VI) acid, there is effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of magnesium sulphate remains in the test tube // boiling tube.



▪ **Calcium:**

- When a piece of calcium is dropped into a beaker containing sulphuric (VI) acid, there is slight effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- The reaction however quickly stops and there is formation of a white precipitate in the test tube or boiling tube.

Reason:

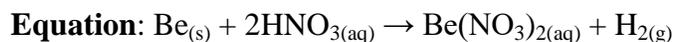
- When calcium reacts with dilute sulphuric (VI) acid there is formation of calcium sulphate which is insoluble hence the formation of a white precipitate. The insoluble calcium sulphate coats the surface of the (calcium) metal stopping further reaction.



✓ **With nitric (V) acid.**

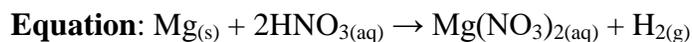
▪ **Beryllium:**

- When a piece of beryllium is dropped into a beaker containing nitric (V) acid, there is effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of beryllium nitrate remains in the test tube // boiling tube.



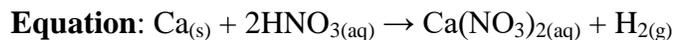
▪ **Magnesium:**

- When a piece of magnesium is dropped into a beaker containing nitric (V) acid, there is effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of magnesium nitrate remains in the test tube // boiling tube.



▪ **Calcium:**

- When a piece of calcium is dropped into a beaker containing nitric (V) acid, there is slight effervescence of a colourless gas.
- The colourless gas produces a pop sound when exposed to a burning splint confirming that it is hydrogen gas.
- A colourless solution of calcium nitrate remains in the test tube // boiling tube.



Trend in reactivity in alkaline earth metals.

- The reactivity of the alkaline earth metals increases down the group.

Reason:

- The atomic radius increases down the group due to increasing number of energy levels. The attraction of the two valence electrons towards the positive nucleus thus decreases down the group. Since alkaline earth metals react by losing electrons the ease with which the valence electrons are lost therefore decreases down the group

Similarity of ions and formulae of some compounds of Alkaline earth metals.

- Alkaline earth metals have the same valency (2) and a similar charge (2+) on their ions since they all have a single valence electron.
- Thus they have the same general formula in their ions and compounds.

Examples.

Alkali metal ion	Hydroxide (OH ⁻)	Oxides (O ²⁻)	Chloride (Cl ⁻)	Sulphates (SO ₄ ²⁻)
Be ²⁺				
Mg ²⁺				
Ca ²⁺				

Uses of some alkaline earth metals and their compounds

1. Magnesium is used in the manufacture of magnesium hydroxide which is used as an anti-acid medicine.

Reason: magnesium hydroxide is a non-toxic base.

2. A low-density alloy of magnesium and aluminium is used in aeroplanes construction.

3. Hydrated calcium sulphate (plaster of Paris) is used in hospitals to set fractured bones.

4. Cement is made by heating a mixture of calcium carbonate (limestone), clay and sand.

5. Calcium carbonate is used in the extraction of iron.

6. Calcium oxide (quicklime) is added to acidic soils to raise pH for agricultural purposes.

7. Calcium nitrate is used as a nitrogenous fertilizer.

8. Magnesium oxide is used in the lining of furnaces.

9. Barium sulphate is used in the diagnosis of ulcers.

10. Barium nitrate is used to produce the green flame in fireworks.

11. Calcium carbonate is mixed with oil to make putty.

The Halogens.

- Are the elements with seven valence electrons and hence in group VII of the periodic table.
- All are non-metallic in nature.
- The members of the family in order down the group is as follows:
 - ✓ Fluorine.
 - ✓ Chlorine
 - ✓ Bromine
 - ✓ Iodine
 - ✓ Astatine

Electron arrangement of the first three halogens.

Elements	Electron arrangement
Fluorine	2.7
Chlorine	2.8.7
Bromine	2.8.18.7

Note:

- Halogen is a derivative of two Greek words: halo and gen.

Diagram: Part of periodic table showing the halogens

Gradation in properties of halogens.

1. Atomic and ionic radius.

✓ **Atomic radius:**

- Refers to the distance between the centre of the nucleus of an atom and the outermost energy level occupied by electron(s)

✓ **Ionic radius:**

- Refers to the distance between the centre of the nucleus of an ion and the outermost energy level occupied by electron(s)

Trend:

- The ionic radius and the atomic radius of halogens increase down the group.

Reason:

- There is an increase in the number of energy levels down the group from Fluorine to iodine.

Illustration:

- Fluorine (2.7) has only 2 energy levels; Chlorine (2.8.7) has 3 energy levels while bromine (2.8.18.7) has 4 energy levels.
- Thus the outermost electron in bromine is further from the nucleus than the outermost electron in chlorine and fluorine.

2. Atomic and ionic radius of the same element.

- For the same halogen the ionic radius is larger than the atomic radius.

Reason:

- Halogens form ions by gaining (an) electron into the outermost energy level. This increases the electron-electron repulsion in the outermost energy level hence decreasing the nuclear attraction of the outermost electrons towards the nucleus.

Illustration

- Chlorine atom has electron arrangement of 2.8.7 hence 3 energy levels.
- During ion formation it reacts by gaining an electron into the third energy level to acquire a new electron arrangement of 2.8.8.
- In the atom 17 protons are attracting 17 electrons; while in the chloride ion there are 17 protons attracting 18 electrons in the outermost energy level.
- Thus in the ion the effect of the positive nucleus is lower.
- This is due to increased repulsive effect between the existing electrons in the outermost energy level and the incoming electron (electron-electro repulsion)

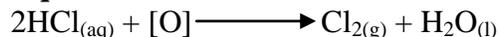
Diagrammatically: chlorine atom and chloride ion**Summary: changes in atomic and ionic radius among alkali metals.**

Element	Symbol	Atomic number	Electron arrangement	Atomic radius (nm)	Ionic radius (nm)
Fluorine	F	9	2.7	0.064	0.136
Chlorine	Cl	17	2.8.7	0.099	0.181
Bromine	Br	35	2.8.18.7	0.114	0.195
Iodine	I	53	2.8.18.18.7	0.133	0.216

Laboratory preparation of chlorine.

Note: It is usually prepared by oxidation of concentrated hydrochloric acid by removal of hydrogen.

Equation:



- The [O] is from a substance containing oxygen (an oxidizing agent).
- The main oxidizing agents normally used for preparation of chlorine are:
 - ✓ Potassium manganate (VII); KMnO_4 .
 - ✓ Manganese (IV) oxide; MnO_2

(a). Preparation of chlorine from MnO_2 and HCl .

(i). Apparatus:

(ii). Conditions:

- Heating;
- Presence of an oxidizing agent; in this case it is manganese (IV) oxide.

(iii). Procedure:

- Hydrochloric acid is reacted with manganese (IV) oxide (dropwise);

Equation:



(iv). Explanation:

- Manganese (IV) oxide oxidizes hydrochloric acid by removing hydrogen resulting into chlorine.
- The manganese (IV) oxide is reduced to water and manganese chloride.
- The resultant chlorine gas is passed through a bottle containing water.

Reason:

- To remove hydrogen chloride fumes (gas) which is very soluble in water.
- Next it is passed through concentrated sulphuric acid or anhydrous calcium chloride; to dry the gas.

(v). Collection:

- (a). Wet chlorine is collected over brine (saturated sodium chloride solution) or hot water.

Reason:

- It does not dissolve in brine and is less soluble in water

- (b). Dry chlorine is collected by downward delivery (upward displacement of air)

Reason:

- It is denser than air (2.5 times).

Note:

- Chlorine may also be dried by adding calcium chloride to the jar of chlorine.

(c). The first bottle must contain water and the second concentrated sulphuric acid.

Reason:

- If the gas is first passed through concentrated sulphuric acid in the first bottle then to the water; it will be made wet again.

Physical properties of Halogens.

1. Physical state and appearance at room temperature.

Halogen	State and appearance
Fluorine	Pale yellow gas
Chlorine	Green-yellow gas
Bromine	Volatile brown liquid
Iodine	Shiny dark grey solid.

2. Solubility

(a). In water

Experiment: To investigate solubility of halogens in water.

✓ **Procedure:**

- A boiling tube containing chlorine gas is inverted into a beaker containing water.
- The experiment is repeated with fluorine, bromine and a few crystals of iodine.

✓ **Diagram:**

✓ **Observations:**

✓ **Fluorine:**

- The level of solution rises in the boiling tube.
- The pale yellow colour of fluorine disappears.

✓ **Fluorine:**

- The level of solution rises in the boiling tube.
- The green-yellow colour of fluorine disappears.

✓ **Bromine**

- The level of solution rises in the boiling tube.
- The brown colour of fluorine disappears.

✓ **Iodine:**

- The level of solution remains the same in the boiling tube.
- The shiny dark grey crystals remain in the beaker.

- ✓ The rise in water level is higher in fluorine than in chlorine while the rise in chlorine is higher than in bromine.

Diagrams: observations after some time

✓ **Explanations:**

- Fluorine, chlorine and bromine are all soluble in water, while iodine is insoluble in water.
- When a boiling tube containing the soluble halogens is inverted into a beaker containing water, the halogen dissolves in the water.
- This creates a partial vacuum and the water in the beaker thus rises to occupy the space left by the dissolved gas.
- The halogens dissolve in water to form acidic solutions.
- The more soluble the halogen, the higher the rise in water level in the boiling tube.
- Thus the solubility of halogens decreases down the group from fluorine to bromine.

Reason:

- All halogens have molecular structures with weak van der Waals forces between the molecules. During the dissolution the Van der Waals must be broken. The strength of Van der Waals increases as the atomic size and hence the molecular size increases which occurs down the group.

(b). In tetrachloromethane.

- The same procedure (of dissolving halogens in water) is followed with tetrachloromethane.

✓ **Observations:**

- All halogens are soluble in tetrachloromethane.
- The solubility of each halogen is higher in tetrachloromethane than in water.

Reason: Halogens are molecular thus non-polar and thus are more soluble in polar organic solvents like tetrachloromethane than in polar solvents like water.

3. Effect of heat.

- Fluorine and chlorine are gases at room temperature and simply expand and increase in volume when heated.
- Bromine is a brown liquid at room temperature and evolves brown fumes when heated.
- Iodine exists as shiny dark grey solid at room temperature and sublimes to give brown fumes when heated.

4. Electrical conductivity.

- All halogens are non-conductors of heat and electricity.

Reason:

- Halogens are molecular and thus have neither delocalized electrons nor free mobile ions for electrical conductivity.

5. Trend in melting and boiling points.

- The melting and boiling points of halogens increase down the group.

Reason:

- Halogens exist as diatomic molecules and thus have molecular structures;
- The molecules are held together by intermolecular forces called the Van der Waals which have to be broken during melting and boiling;
- The strength of the Van der Waals increases as the size of the atoms and hence molecules increases which occurs down the group.

Summary on some physical properties of halogens>

Halogen	Formula	Atomic number	Appearance	Melting point (°C)	Boiling point (°C)
Fluorine	F	9	Pale yellow gas	-238	-188
Chlorine	Cl	17	Green yellow gas	-101	-35
Bromine	Br	35	Brown liquid	-7	59
Iodine	I	53	Shiny dark grey solid	114	184

Chemical properties of Halogens.

Note: It is not easy for non-metals like halogens to form cations.

Reason: the ionization energy (amount of energy required to lose an electron(s) from the outermost energy level of a gaseous atom) is very large.

- Thus most non-metals react by forming anions (negatively charged ions) by electron loss.

1. Ion formation.

- Halogens react by gaining a single electron into the outermost energy level to form a stable electron configuration and corresponding anions.
- During ion formation by **electron loss** energy is **released**, and the energy change for this process is called **electron affinity**

Note: Definition.

- **Electron affinity is thus the energy released when an atom in gaseous state gains an electron to form a negative ion (anion)**

✓ Trend in electron affinity of halogens:

- Generally the electron affinity of halogens decreases down the group.

Reason:

- The size of the atoms increases (due to increasing number of energy levels) down the group and thus the force of attraction of the electrons in the outermost energy level towards the nucleus decreases.
- Thus down the group the ease with which electrons are gained decreases and the faster the ease of electron gain, the more the energy released hence the more the electron affinity.

Summary:

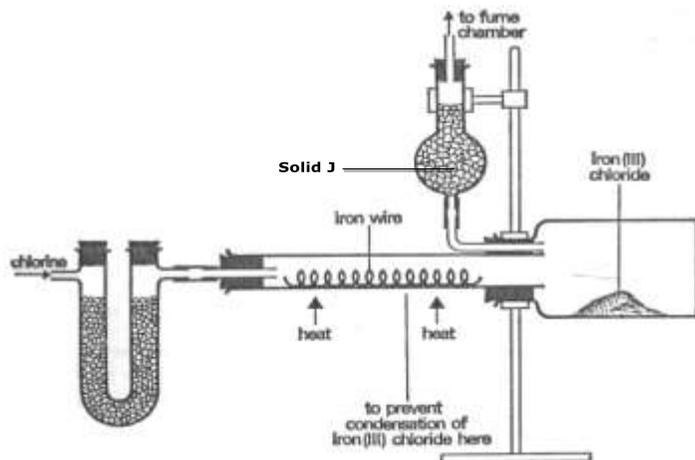
Element	Ionization equation	Electron affinity (kJ per mole)
Fluorine	$F_{(g)} + e \rightarrow F^{-}_{(g)}$	-322
Chlorine	$Cl_{(g)} + e \rightarrow Cl^{-}_{(g)}$	-349
Bromine	$Br_{(g)} + e \rightarrow Br^{-}_{(g)}$	-325
Iodine	$I_{(g)} + e \rightarrow I^{-}_{(g)}$	-295

2. Reaction with metals.

(a). Chlorine.

✓ Reaction of chlorine with iron.

(i). Apparatus.



Note:

- The set up can also be modified by using sodium hydroxide to absorb excess chlorine gas as shown below.

(ii). Precaution.

- Experiment should be done in a functional fume cupboard or in the open.

Reason:

- Chlorine gas is poisonous and will thus be harmful to the human body.

(iii). Procedure:

- A stream of dry chlorine gas is passed over heated iron wool as per the diagram.

(iv). Conditions.

➤ Chlorine gas has to be dry (done by the anhydrous calcium chloride in the U-tube)

Reason:

To prevent **hydration** hence **oxidation** of iron (which will then form $\text{Fe}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$) hence preventing reaction between iron and chlorine.

➤ Iron metal must be **hot**; and this is done by heating.

Reason:

To provide **activation energy** i.e. the minimum kinetic energy which the reactants must have to form products.

➤ Anhydrous calcium chloride.

- In the U-tube; to **dry** the chlorine gas.
- In the thistle funnel; to prevent **atmospheric water vapour** (moisture) from getting into the apparatus and hence reacting with iron (III) chloride.

Note: In the guard tube, calcium oxide is preferable to anhydrous calcium chloride.

Reason:

- Other than preventing atmospheric water vapour from getting into the flask with iron wool; it also absorbs excess chlorine thus preventing environmental pollution

(v). Observations:

- Iron metal glows red-hot.
- Red brown fumes ($\text{FeCl}_{3(g)}$) are formed in the combustion tube.
- A black solid ($\text{FeCl}_{3(s)}$) is collected in the flask.

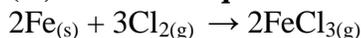
Note:

- Iron (III) chloride cannot be easily collected in the combustion tube.

Reason:

- It sublimes when heated and hence the hotter combustion tube causes it to sublime and its vapour is collected on the cooler parts of the flask.

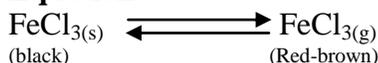
(vi). Reaction equation.



(vii). Conclusion.

- Iron (III) chloride sublimes on heating; the black solid changes to red-brown fumes on heating.

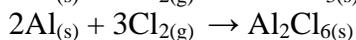
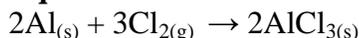
Equation:



✓ **Reaction with aluminium.**

- Chlorine reacts with aluminium in a similar manner to iron to form aluminium chloride.

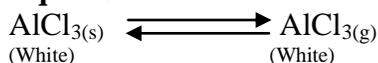
Equation:



Note:

- Aluminium chloride also **sublimes** on heating.

Equation:



✓ **Reaction with burning magnesium.**

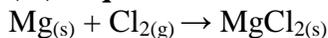
(i). Procedure:

- Burning magnesium is lowered into a gas jar of chlorine gas.

(ii). Observations:

- The magnesium continues to **burn** with a **bright blinding flame**;
- Formation of **white fumes** (MgCl_2); which cools into a white powder.

(iii). Equation:



- Generally chlorine reacts with most metals when hot to form corresponding chlorides.

Note:

Where a metal forms two chlorides when it reacts with chlorine, the higher chloride is usually formed.

Reason:

The higher chloride is **stable**. This explains why reactions of chlorine with iron result into iron (III) chloride and not iron (II) chloride.

(b). Bromine.

✓ **Reaction of bromine with iron**

Apparatus

(ii). Precaution.

- Experiment should be done in a functional fume cupboard or in the open.

Reason:

- Bromine gas is poisonous and will thus be harmful to the human body.

(iii). Procedure:

- Bromine liquid is heated to generate bromine vapour (fumes).

- The iron wool is then heated and a stream of the bromine gas is passed over heated iron wool as per the diagram.

(iv). Conditions.

➤ Bromine must be heated to generate fumes before heating the iron.

Reason:

So that bromine vapour will drive out air from the boiling tube to prevent **oxidation** of iron with oxygen which would otherwise prevent reaction between iron and bromine.

➤ Iron metal must be **hot**; and this is done by heating.

Reason:

To provide **activation energy** i.e. the minimum kinetic energy which the reactants must have to form products.

(v). Observations:

- Brown fumes of bromine are produced on heating bromine liquid.

- The iron wool glows red-hot upon heating

- The iron wool forms dark red crystals (of iron (III) bromide)

(vi). Reaction equation.

Word: Iron + bromine → Iron (III) bromide.

Chemical: $2\text{Fe}_{(s)} + 3\text{Br}_{2(g)} \rightarrow 2\text{FeBr}_{3(s)}$
Dark red crystals.

✓ **Reaction of bromine with zinc**

- Using the same set up bromine also reacts with zinc to form zinc bromide.

Equation: $\text{Zn}_{(s)} + \text{Br}_{2(g)} \rightarrow \text{ZnBr}_{2(s)}$.

✓ **Reaction of bromine with magnesium**

- Using the same set up bromine also reacts with zinc to form magnesium bromide.

Equation: $\text{Mg}_{(s)} + \text{Br}_{2(g)} \rightarrow \text{MgBr}_{2(s)}$.

✓ **Reaction of bromine with sodium**

- Using the same set up bromine also reacts with zinc to form sodium bromide.

Equation: $2\text{Na}_{(s)} + \text{Br}_{2(g)} \rightarrow 2\text{NaBr}_{(s)}$.

(c). Iodine.

✓ **Reaction of iodine with iron**

Apparatus

(ii). Precaution.

- Experiment should be done in a functional fume cupboard or in the open.

Reason:

- Iodine is poisonous and will thus be harmful to the human body.

(iii). Procedure:

- Iodine crystals are heated to generate iodine vapour (fumes).

- The iron wool is then heated and a stream of the iodine gas is passed over heated iron wool as per the diagram.

(iv). Conditions.

➤ Iodine must be heated to generate fumes before heating the iron.

Reason:

So that iodine vapour will drive out air from the boiling tube to prevent **oxidation** of iron with oxygen which would otherwise prevent reaction between iron and iodine.

➤ Iron metal must be **hot**; and this is done by heating.

Reason:

To provide **activation energy** i.e. the minimum kinetic energy which the reactants must have to form products.

(v). Observations:

- Purple vapour (fumes) of iodine is produced on heating iodine crystals.
- The iron wool glows red-hot upon heating
- The iron wool forms grey black crystals (of iron (II) iodide)

(vi). Reaction equation.

Word: Iron + iodine → Iron (II) iodide.

Chemical: $\text{Fe}_{(s)} + \text{I}_{2(g)} \rightarrow \text{FeI}_{2(s)}$
Grey-black crystals.

✓ **Reaction of Iodine with zinc**

- Using the same set up bromine also reacts with zinc to form zinc iodide.

Equation: $\text{Zn}_{(s)} + \text{I}_{2(g)} \rightarrow \text{ZnI}_{2(s)}$.

✓ **Reaction of Iodine with magnesium**

- Using the same set up bromine also reacts with zinc to form magnesium iodide.

Equation: $\text{Mg}_{(s)} + \text{I}_{2(g)} \rightarrow \text{MgI}_{2(s)}$.

✓ **Reaction of Iodine with sodium**

- Using the same set up bromine also reacts with zinc to form sodium iodide.

Equation: $2\text{Na}_{(s)} + \text{I}_{2(g)} \rightarrow 2\text{NaI}_{(s)}$.

Note:

- The reactivity of chlorine with metals is more vigorous than that of bromine, which is more than that of iodine.
- Thus reactivity of halogens with metals decreases down the group.

Reason:

- Halogens react by gaining electrons; the ease of gaining electrons decrease down the group as the atomic size increases which leads to progressive decrease in the force of attraction of electrons in the outermost energy levels by the positively charged nucleus.

3. Reaction with water.

- Halogens that dissolve in water form a mixture of two acids.

Reaction of chlorine with water.

- Chlorine dissolves in water to form chlorine water, which is a mixture of two acids: chloric (I) acid and hydrochloric acid.

Equation:

$\text{Cl}_{2(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{HOCl}_{(aq)} + \text{HCl}_{(aq)}$
Chloric (I) acid Hydrochloric acid

✓ **Effects of sunlight on chlorine water.**

(i). Procedure:

- Chlorine water is made by dissolving the gas in water.
- A long tube filled with chlorine water is inverted over a beaker containing water.
- It is then exposed to sunlight (bright light) as shown below.

(ii). Apparatus:

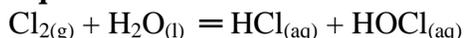
(iii). Observations:

- After sometime a gas collects in the tube and on applying a glowing splint, the splint is rekindles showing that the gas collected is oxygen.

(iv). Explanation:

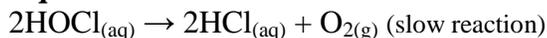
- Chlorine water has two components.

Equation:

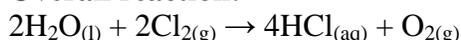


- The HOCl being unstable will dissolve on exposure to sunlight, giving out oxygen.

Equation:



Overall reaction:



✓ **Effect of chlorine water on litmus papers**

(i). Procedure:

- A strip of blue and a strip of red litmus papers are dropped into chlorine water in a beaker.

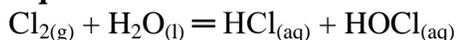
(ii). Observations:

- The blue litmus paper turns red; then both litmus papers are decolourised.

(iii). Explanations.

- Chlorine water contains a mixture of chloric (I) acid and hydrochloric acid.

Equation:

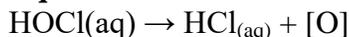


- The two acids cause the litmus paper to turn red (from blue) while the red litmus paper remained red.

- The litmus papers are then decolourised due to bleaching action of chlorine water, which is due to the activity of chloric (I) acid.

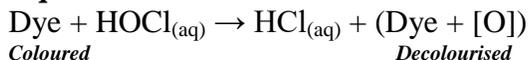
- The chloric (I) acid is unstable and thus decomposes to give hydrochloric acid and oxygen atom.

Equation:



- The oxygen atom combines with the chemical of the natural dye in the litmus thus decolourising it.

Equation:



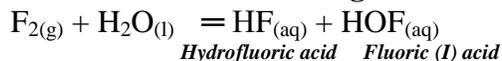
Coloured

Decolourised

Note: solutions of bromine and fluorine in water will behave in a similar manner.

Examples:

✓ **Fluorine dissolving in water**



- The mixture is fluorine water

✓ **Bromine dissolving in water**



- The mixture is bromine water.

Some uses of halogens and their compounds.

1. Fluorine is a raw material in the preparation of a synthetic fibre known as polytetrafluoroethane.
2. Some compounds of fluorine are added to water and some tooth pastes in small quantities to reduce tooth decay.
3. Fluorine is used in the manufacture of hydrogen fluoride used to engrave words or pictures on glass.
4. Chlorine is used to make bleaches used in paper and textile industries.
5. Chlorine is added to water to kill micro-organisms in water treatment works.
6. Chlorine is used in the manufacture of a plastic known as polyvinylchloride (PVC).
7. Chlorine is used in large scale manufacture of hydrochloric acid.
8. Bromine is used in the manufacture of silver bromide which is used to make the light sensitive photographic paper and films.
9. A solution of iodine in alcohol (tincture of iodine) is used as a disinfectant.

The Noble gases.

- Are the elements with the maximum possible number of valence electrons and hence in group VIII of the periodic table.
- All are gaseous in nature.
- The members of the family in order down the group is as follows:
 - ✓ Helium
 - ✓ Neon
 - ✓ Argon
 - ✓ Krypton
 - ✓ Xenon
 - ✓ radon
- They are found as free atoms in nature and form about 1% of air with argon being the most abundant

Electron arrangement of the first three noble gases.

Elements	Electron arrangement
Helium	2.
Neon	2.8.
Argon	2.8.8.

- Note:** - Helium with only two electrons has one occupied energy level; which is completely full and hence is said to have a duplet electron configuration
- The rest of the noble gases have eight electrons in their outermost occupied energy level hence are said to have the octet electron configuration.

Diagram: Part of periodic table showing the noble gases

Dot and cross diagrams for the first three Alkaline Earth metals

Physical properties of noble gases

1. Atomic and ionic radius.

✓ **Atomic radius:**

- Refers to the distance between the centre of the nucleus of an atom and the outermost energy level occupied by electron(s)

✓ **Ionic radius:**

- Refers to the distance between the centre of the nucleus of an ion and the outermost energy level occupied by electron(s)

Trend:

- The atomic radius of noble gases increase down the group.

Reason:

- There is an increase in the number of energy levels down the group from Helium to radon.

Illustration:

- Helium (2.) has only 1 energy level; Neon (2.8) has 2 energy levels while Argon (2.8.8) has 3 energy levels.

- Thus the outermost electron in Argon is further from the nucleus than the outermost electron in neon and helium.

Note:

- Under normal circumstance the noble gases neither lose nor gain electrons and are thus stable hence unreactive.

- Due to this they rarely form ions and hence comparative studies of ionic radii among noble gases are not feasible.

2. Ionization energies.

- Noble gases have very high ionization energies.

Reason:

- Noble gases are colourless monoatomic gases with very stable electron arrangements, either duplet (helium) or octet (the rest).

- The nuclear attraction force of electrons in the outermost energy level towards the nucleus is thus very strong.

Trend:

- The ionization energy decrease down the noble gases group.

Reason:

- As the size of the atoms increase down the group (due to increase in the number of energy levels), the force with which the electrons in the outermost energy levels are attracted towards the nucleus decrease hence increase in the ease of electron loss (from the outermost energy level).

3. Melting and boiling points.

- Generally noble gases have very low melting and boiling points.

Reason:

- They exist as monoatomic gases with simple molecular structures with weak van der Waals forces that are easy to break.

Trend:

- Melting and boiling points among the noble gases increase down the group.

Reason:

- Noble gases exist as monoatomic molecules and thus have simple molecular structures;

- The molecules (atoms) are held together by intermolecular (inter-atomic) forces called the Van der Waals which have to be broken during melting and boiling;

- The strength of the Van der Waals increases as the size of the atoms and hence molecules increases which occurs down the group.

Summary: changes in atomic and ionic radius among alkali metals.

Element	Symbol	Atomic number	Atomic radius (nm)	1 st Ionization energy (kJmol ⁻¹)	Melting point (°C)	Boiling point (°C)
Helium	He	2	0.128	2372	-270	-269
Neon	Ne	10	0.160	2080	-249	-246
Argon	Ar	18	0.192	1520	-189	-186
Krypton	Kr	36	0.197	1350	-157	-152
Xenon	Xe	54	0.217	1170	-112	-108

Chemical properties of Noble gases.

- Generally the noble gases neither lose nor gain electrons and are thus stable hence unreactive.

- Due to this they rarely form ions and hence have no feasible chemical reactions.

- However xenon and radon with very large atomic radii and smaller ionization energies take part in some reactions and thus display some chemical properties.

Uses of some noble gases.

Note: The application of noble gases is ironically centered on their inert nature.

1. Argon is used in light bulbs to provide an inert environment to prevent oxidation.

2. Argon is used as an insulator in arch-welding.

3. Neon gas is used in streets and advertisement lights.

4. Helium mixed with oxygen is used in deep sea diving and mountaineering.

5. Helium mixed with oxygen also is used in hospitals for breathing in patients with respiratory problems and those undergoing certain forms of surgery.

6. Helium can be used instead of hydrogen in balloons for meteorological research.

Reason:

- Hydrogen is explosive in presence of air and may explode causing serious accidents.

- Helium is less dense than hydrogen.

7. Helium is used in thermometers for measurements of very low temperatures.

8. Liquid helium is used to keep certain metal alloys at temperatures low enough for them to become superconductors.

UNIT 3: PROPERTIES AND TRENDS ACROSS A PERIOD.

Checklist.

1. Introduction
2. Trends in physical properties of elements in period 3
 - ✓ Electrical conductivity
 - ✓ Atomic and ionic radii
 - ✓ Melting and boiling points
3. Trends in chemical properties of elements in period 3
 - ✓ Reaction with oxygen
 - ✓ Reaction with water
 - ✓ Reaction with acids

Introduction:

- A period is a vertical row of elements in the periodic table.
- Elements in the same period have same number of energy levels.
- There are 7 periods in the periodic table except for lanthanides and actinides which are not assigned periods.
- Periods 1 – 3 have fewer elements because they lack the d-block elements and have only the s-block elements and the p-block elements.

Note:

- ✓ s-block elements: group 1 and 2
- ✓ d-block elements: transitional elements
- ✓ p-block elements: groups III to VIII.

- To understand trends and properties across a typical period of the periodic table, we shall use period 3 as the reference.

The Period three of the periodic table.

- Consists of elements with three energy levels.
- Consists of the 8 elements from sodium to argon.
- It is only made of s-block and p-block elements and lacks any element in the d-block group of elements.

Part of periodic table showing period 3 of the periodic table.**Trends in physical properties of elements in period 3.****1. Electrical conductivity.**

- Sodium, magnesium and aluminium are good conductors of electricity.

Reason:

They all have giant metallic structures with delocalized electrons which conduct electricity;

- The electrical conductivity increases from sodium to aluminium.

Reason:

Electrical conductivity increases with increase in the number of delocalized electrons; thus aluminium with the highest number of delocalized electrons (3) in each atom will have the highest electrical conductivity;

- In the metals the electrical conductivity decreases with increase in temperature.

Reason:

Increase in temperature distorts the alignment of electrons thus preventing their easy flow and hence poor conductivity;

- Phosphorus, sulphur, chlorine and argon do not conduct electric current.

Reason:

They all have molecular structures and all the electrons in the atoms are used in bonding; thus they lack delocalized electron or free ions for electrical conductivity.

- Silicon conducts electric current, and its electrical conductivity increases with increase in temperatures.

Reason:

- It is a semi-conductor; making it a very unique element in this period.

Note:

- A semi-conductor is a crystalline material which only conducts electricity under certain conditions.

2. The atomic and ionic radii.

- The atomic radii of period 3 elements decrease across the period.

Reason.

- For the same number of energy levels the number of protons in the nucleus increases across the period; this leads to the increase in nuclear charge while the shielding effect remains the same hence decrease in atomic radius across the period.

3. Melting and boiling points

- Sodium, magnesium and aluminium have very high melting and boiling points.

✓ Reason:

- They have **giant metallic structures** with **strong metallic bonds** which need a lot of energy to break.

- The boiling and melting points increase from sodium to aluminium.

✓ Reason:

- As you move across the period from sodium to aluminium, the **nuclear charge increases** while the **energy levels remain the same** hence decrease in atomic radius; the smaller the atomic radius (for metals) the stronger the metallic bonds.

- Silicon, though a non-metal also has a very high melting and boiling points.

✓ Reason:

- Silicon has a **giant atomic structure** with **strong covalent bonds** throughout the structure, which need a lot of heat energy to break.

- Phosphorus, Sulphur, Chlorine and argon have low melting and boiling points.

✓ Reason:

- They all have **molecular structures** with strong covalent bonds between the atoms (except in argon) but **weak van der Waals forces** between the molecules which are easy to break.

- Note that argon exist as atoms and thus a monoatomic molecule.

- The melting and boiling points decreases from phosphorus to argon.

✓ Reason.

- As we move across the period from phosphorus to argon, the size of the atoms decreases leading to smaller atoms and hence molecules, which lead to decrease in the strength of the van der Waals (across the period)

Note:

- Phosphorus and sulphur exists as solids at room temperature while chlorine and argon exists as gases at room temperature.

Reason: phosphorus and sulphur have giant molecular structures while chlorine and argon have simple molecular structures.

-

Summary: Some physical properties of elements in period 3.

Property	Na	Mg	Al	Si	P (white)	S (monoclinic)	Cl	Ar
Physical state and appearance	Silver	Silver solid	Silver solid	Black solid	White solid	Yellow solid	Green yellow gas	Colourless gas
Electron arrangement	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	5 or 5	6	7	8
Atomic radius	0.157	0.136	0.125	0.117	0.110	0.104	0.09	0.192
MP (°C)	98	650	660	1410	44	119	-101	-189
BP (°C)	890	1110	2470	2360	280	445	-35	-186
Structure	Giant metallic	Giant metallic	Giant metallic	Giant atomic	Molecular	Molecular	Simple molecular	Simple molecular /atomic
Bond type	metallic	metallic	metallic	Covalent	Covalent	Covalent	Covalent	Van der waals

Trends in chemical properties of the elements in period 3.

Note: Trends in reactivity.

- The reactivity among the metallic elements decreases across the period from sodium to aluminium.

Reason:

- There is a continuous increase in nuclear charge from sodium to aluminium which leads to increase in ionization energies hence increasing difficulty in removing an electron from the outermost energy level.

- Among the non-metallic elements, the reactivity increases across the period from phosphorus to chlorine.

Reason:

- There is increase in nuclear charge from phosphorus to chlorine, hence an increase in ease of electron gain (electro affinity) since non-metals react by gaining electrons.

- Argon is unreactive and can only react under very special conditions.

Reason:

- It's a noble gas with a stable octet configuration.

1. Reaction of period 3 elements with oxygen.

- All period three elements react with (burn in) oxygen with the exception of argon.

Experiment: To investigate the reactions between period 3 elements and oxygen

Procedure:

- A small piece of the element is placed in a deflagrating spoon and warmed gently until it catches fire.

- It is then lowered into a gas jar full of oxygen.

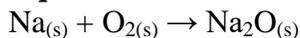
- The flame colour and the colour of the product are noted.

- 10 cm³ of water containing universal or litmus indicator is added into the gas jar with the products.

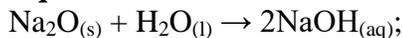
✓ Sodium:

- Burns vigorously in oxygen with a **golden yellow flame**; to produce white solid of sodium oxide.

Equation:



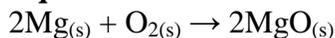
- The resultant sodium oxide dissolves in water to form sodium hydroxide.

Equation:

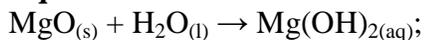
- The sodium hydroxide is alkaline in nature and thus turns litmus indicator blue;

✓ **Magnesium:**

- Burns vigorously in oxygen with a **bright white light**; to produce white solid of magnesium oxide.

Equation:

- The resultant magnesium oxide is slightly soluble in water to form magnesium hydroxide.

Equation:

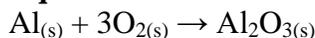
- The magnesium hydroxide is alkaline in nature and thus turns litmus indicator blue;

✓ **Aluminium:**

Note:

- Aluminium (foil) is usually coated with a thin layer of aluminium oxide, Al_2O_3 ; which prevents reaction with the oxygen.

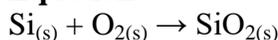
- When polished, it reacts slowly with oxygen to form a white solid of aluminium oxide.

Equation:

- The resultant aluminium (III) oxide is insoluble in water.

✓ **Silicon:**

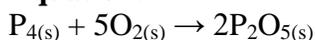
- Silicon powder can only burn in oxygen at very high temperatures of about 450°C to form solid silicon (IV) oxide.

Equation:

- The resultant silicon (IV) oxide is insoluble in water.

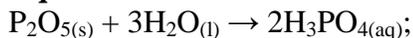
✓ **Phosphorus.**

- Burns readily in oxygen with a **bright orange flame**; to produce dense white fumes (solid) of phosphorus (V) oxide

Equation:

Note: sulphur exists and therefore reacts as molecules of P_4 .

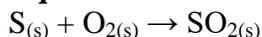
- The resultant phosphorus (V) oxide readily dissolves in water to form phosphoric (V) acid.

Equation:

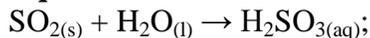
- The phosphoric acid is acidic in nature and thus turns litmus indicator pink / red;

✓ **Sulphur.**

- Burns in oxygen with a **blue flame**; to form a colourless gas of sulphur (IV) oxide

Equation:

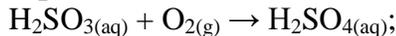
- The resultant sulphur (IV) oxide readily dissolves in water to form sulphuric (IV) acid.

Equation:

- The phosphoric acid is acidic in nature and thus turns litmus indicator pink / red;

- The sulphuric (IV) acid is unstable and thus easily gets oxidized by oxygen to the stable sulphuric (VI) acid.

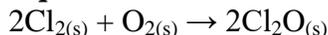
Equation:



✓ **Chlorine.**

- Burns in oxygen only under certain conditions to form acidic oxides.

Equation:



✓ **Argon.**

- Argon is unreactive.

Conclusion:

- Metallic elements burn in oxygen to form basic oxides.
- Non-metallic oxides burn in oxygen to form acidic oxides.

2. Reaction of period 3 elements with water.

✓ **Sodium.**

Procedure:

- A small piece of sodium metal is cut and dropped into a trough containing water;
- The resultant solution is tested with litmus paper;

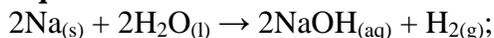
Diagram of apparatus:

Observations and explanations:

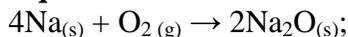
- The metal floats on the water surface; because it is less dense than water;
- A hissing sound is produced; due to production of hydrogen gas;
- It vigorously melts into a silvery ball then disappears because reaction between water and sodium is exothermic (produces heat). The resultant heat melts the sodium due to its low melting point.
- It darts on the surface; due to propulsion by hydrogen;
- The metal may burst into a golden yellow flame; because hydrogen may explode into a flame which then burns the sodium;
- The resultant solution turns blue; because sodium hydroxide solution formed is a strong base;

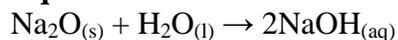
Reaction equations.

Equation I



Equation II



Equation III:**Effect of resultant solution on litmus paper;**

- Litmus paper turns blue; sodium hydroxide formed is highly soluble in water; releasing a large number of hydroxyl ions which result into alkaline conditions // high pH;

✓ Magnesium.**Procedure:**

- A small piece of magnesium ribbon is cut and dropped into a trough containing water;
- The resultant solution is tested with litmus paper;

Diagram of apparatus:**Observations and explanations:**

- The metal sinks into the water surface; because it is denser than water;
- It reacts slowly with water leading to slow evolution of hydrogen gas.
- The resultant solution turns blue; because sodium hydroxide solution formed is a strong base;

Reaction equation.**Effect of resultant solution on litmus paper;**

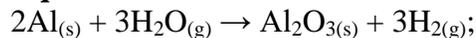
- Litmus paper turns blue; magnesium hydroxide formed is highly soluble in water; releasing a large number of hydroxyl ions which result into alkaline conditions // high pH;

✓ Aluminium.

- Aluminium does not normally react with cold water or steam.

Reason:

- Aluminium is usually coated with a thin coating of aluminium oxide which prevents further reaction.
- However at very high temperatures of about 700°C steam can react with aluminium to form aluminium oxide and hydrogen gas.

Equation:**Note:**

- Due to its inability to react with water and air aluminium is preferable for making cooking utensils like sufurias and cooking pans.

✓ **Silicon, phosphorus and sulphur.**

- These non-metals do not displace hydrogen from water and thus do not react with water.

✓ **Chlorine.**

- Dissolves in water to form chlorine water, which is a mixture of chloric (I) acid and hydrochloric acid.

Equation:



3. Reaction of period 3 elements with acids

Procedure:

- A piece of the element is dropped into 5 cm³ of an acid in a test tube.
- Any gas produced is tested.

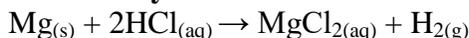
✓ **Sodium:**

- Reacts explosively with acids to form salts and hydrogen and thus reactions of sodium with acids should not be tried in the laboratory.

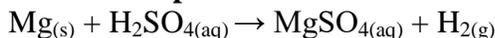
✓ **Magnesium:**

- Reacts with both dilute hydrochloric and dilute sulphuric acid to form magnesium salts and hydrogen gas.

▪ **With hydrochloric acid:**



▪ **With sulphuric acid:**



✓ **Aluminium:**

Note:

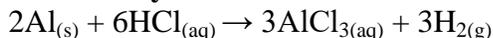
- It does not readily react with dilute acids.

Reason:

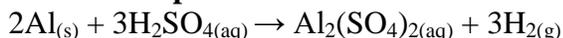
- This is due to presence of a thin aluminium oxide coating that prevents contact hence reaction with the acids.

- When polished to remove the oxide coating it reacts with both dilute hydrochloric and dilute sulphuric acid to form aluminium salts and hydrogen gas.

▪ **With hydrochloric acid:**



▪ **With sulphuric acid:**



✓ **Silicon, phosphorus, sulphur and chlorine.**

- They do not react with dilute acids.

Summary: Chemical properties of period 3 elements.

Element	Na	Mg	Al	Si	P	S	Cl
Reaction with air or oxygen	Readily reacts with air. - Burns brightly in oxygen with a golden yellow flame to form Na_2O	Reacts slowly with air. - Burns in oxygen with a bright white flame to form MgO	- Forms a protective coating of Al_2O_3 when it burns in oxygen.	Si powder burns at temperatures above 950°C to form SiO_2	White phosphorus smolders in air; - P burns in air with a bright orange flame to form P_2O_3 and P_2O_5	Burns in air or oxygen with a blue flame to form SO_2 gas	No reaction with air or oxygen under normal conditions.
Reaction with water	Reacts vigorously to produce H_2 and NaOH	Slow reaction with cold water to form $\text{Mg}(\text{OH})_2$ and H_2 ; - Reacts faster with steam to form MgO and H_2	No reaction	No reaction	No reaction	No reaction	Dissolves in water to form chlorine water
Reaction with dilute acids	Violent reaction giving out H_2 and a sodium salt	Rapid evolution of H_2 gas and a Mg salt is formed	Reacts slowly to give H_2 and an Al salt	No reaction	No reaction	No reaction	No reaction

UNIT4: STRUCTURE AND BONDING.

1. Meaning of structure and bond.

2. Nature of the chemical bond.

3. Types of bonds.

✓ Ionic bonds

- Meaning
- Formation
- Dot ad cross diagrams.
- Giant Ionic structures
 - Examples
 - Meaning
 - Properties of giant ionic structures

✓ Covalent and dative bonds

- Meaning
- Formation
- Dot ad cross diagrams.
- Giant covalent (atomic) structures
 - Examples
 - Meaning
 - Properties of giant covalent structures
 - Details of Diamond, graphite and silicon (IV) oxide
- Molecular (simple and giant) structures
 - Examples
 - Meaning
 - Hydrogen bond and Van der waals
 - Properties of molecular structures

✓ Metallic bonds

- Meaning
- Formation
- Dot ad cross diagrams.
- Giant Metallic structures
 - Examples
 - Meaning
 - Properties of giant metallic structures

3. Bond types across period 3

✓ Oxides of period 3 elements

- Nature
- Melting and boiling points
- General trends in bond types

✓ Chlorides of period 3 elements

- Nature
- Melting and boiling points
- General trends in bond types

4. Summary on bond characteristics

1. Meaning of structure and bond.

✓ Bond:

- The mutual force of attraction that holds particles together when atoms (similar or different) combine during chemical reactions.

✓ Structure:

- A regular pattern of particles in a substance held together by chemical bonds.

2. Nature of the chemical bond.

- Atoms are made up of energy levels and nucleus.

- The energy levels contain electrons while the nucleus contains protons and neutrons.

- The electrons are negatively charged, protons are positively charged while neutrons are electrically neutral.

- The noble gases are chemically inert since their outermost energy levels are completely filled with the maximum possible number of electrons.

- Thus noble gases have either a stable duplet state (2) like in helium or a table octet configuration (2.8 or 2.8.8) as in neon and argon respectively.

- Other atoms are unstable because the outermost energy levels are not yet completely filled with the maximum possible number of electrons.

- To attain the stable duplet or octet noble gas configuration, such atoms lose, gain or share their valence electrons.

- It is the act of losing, gaining or sharing valence electrons that lead to the chemical bonds.

- When atoms gain or lose valence electron(s) they become charged forming anions and cations respectively.

- Particles of the same charge repel each other while particles of different charges attract one another.

3. Types of bonds.

- There are three main types of chemical bonds:

✓ Ionic bonds

✓ Covalent bond

✓ Metallic bond

(a). Ionic bonds / electrovalent bond.

✓ Meaning

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

Formation of an ionic bond.

- Formed due to complete transfer of electrons from one atom to another; and **mainly** formed between a metal and a non-metal.

- This occurs in a bid for both atoms to acquire a stable noble gas configuration.

- One atom loses all its valence electrons thus forming a cation (positively charged ions).

- The other atom gains all the lost valence electrons forming an anion (negatively charged ion).

- The cation and the anion are oppositely charged and thus develop a mutual force of attraction between them which is the ionic / electrovalent bond.

Illustration: formation of ionic bond between sodium and chlorine to form sodium chloride.

- Sodium metal (atomic number 11) has electronic configuration 2.8.1 and thus unstable with 1 valence electron.
- Chlorine gas (atomic number 17) has electronic configuration 2.8.7 and thus unstable with 7 valence electrons.
- Sodium is more electropositive (the tendency to lose electrons to form cations) than chlorine while chlorine is more electronegative (the tendency to gain electrons to form anions).
- Sodium loses its single valence electron to form sodium ion with electronic configuration 2.8 and a net charge of +1 (Na^+)
- Chlorine atom accepts the single electron lost by sodium to form a chloride ion with electronic configuration 2.8.8 with a net charge of -1 (Cl^-).
- The positively charged sodium ion and the negatively charged chloride ion attract each other.
- The electrostatic forces of attraction develop between the two oppositely charged ions and this constitutes the ionic bond.
- Compounds formed due to ionic bonding are thus called **ionic compounds**.

Diagrammatically:**Examples of ionic compounds.**

- ✓ Sodium chloride.
- ✓ Potassium fluoride.
- ✓ Magnesium oxide
- ✓ Aluminium (III) oxide.

Dot and cross diagrams for ionic compounds

1. Potassium fluoride.
2. Magnesium oxide
3. Sodium oxide.
4. Magnesium chloride.
5. Aluminium (III) oxide.

Giant Ionic structures

- Ionic bonding results into one type of structure, the giant ionic structure.
- This is a type of structure in which all ions are bonded with strong ionic bonds throughout the structure.
- Each ion in the giant ionic structure is surrounded by several others resulting into giant pattern of several ions, hence giant ionic structure.
- Most ionic substances with the giant ionic structure are crystalline in nature, made up crystals.

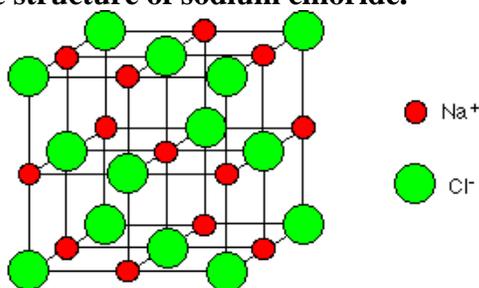
✓ **Note:**

A crystal is a solid form of a substance in which the particles are arranged in a definite pattern repeated regularly in 3 dimensions.

Illustration of the giant ionic structure: sodium chloride structure.

- The NaCl structure consists of many Na^+ and Cl^- arranged and packed in a regular pattern.
- Each Na^+ is surrounded by six Cl^- that are equidistant from it.
- Similarly each Cl^- is surrounded by six Na^+ that are equidistant from it.
- This pattern occurs repeatedly in all directions.
- The result is a giant of ions in all directions hence giant ionic structure.

Diagram of the cubic structure of sodium chloride.



Properties of giant ionic structures.

1. They are hard and brittle.

✓ **Reason:**

- Ionic solids are hard because each ion is held in the crystal by strong attractions from the oppositely charged ions around it.
- They are brittle and thus may be split cleanly (cleaved) using a sharp-edged razor.

✓ **Explanation:**

- When a crystal is tapped sharply along a particular plane it is possible to displace one layer of ions relative to the next.
- Due to the displacement, ions of similar charge come together leading to repulsive forces between the portions of the crystals.
- This forces the two portions of the crystals to split apart.

Diagram: crystal cleavage in ionic compounds.

2. They have high melting and boiling points.

✓ Reason:

- They have strong electrostatic forces / ionic bonds / electrovalent bonds between the oppositely charged ions throughout the structure which require large amounts of energy to break.

3. Solubility.

(i). *They are soluble in polar solvents like water, ethanol and acetone (propanone)*

✓ Reason:

- Water contains highly polar molecules.
- The positive ends of the polar water molecules are attracted to the negative ions in the crystal, and the negative ends of the water molecules are attracted to the positive ions in the crystal.
- This results to the formation of ion-solvent bonds which leads to release of energy.
- This energy is sufficient to cause the detachment of ions from the crystal lattice hence dissolution.

Note:

- This detachment of ions is called **solvation**, and the energy required for this is called **solvation energy**.
- Where the solvent is water the ions are said to have been **hydrated**, and the energy involved in the process is called **hydration energy**.

Diagrams: hydrated positive and negative ions.

(ii). *They are insoluble in non-polar organic solvents like tetrachloromethane, benzene and hexane.*

✓ Reason:

- Non-polar molecules are held together by weak intermolecular forces, the **Van der Waals** forces.
- The Van der Waals are much smaller in magnitude compared to the ionic bonds in the ionic solid crystal lattice.
- Thus the ion-ion interactions in the ionic solid are stronger than the solvent-solvent interactions in the solvent or the solvent-ion interactions between the solid and the solvent.
- Thus the non-polar solvent molecules cannot penetrate the ionic lattice to cause solvation.

4. Electrical conductivity.

- Ionic substances do not conduct electric current in **solid state**.

✓ Reason:

The ions are held in static positions in the solid crystal lattice and thus cannot move to conduct electric current.

- They conduct electric current in **molten** and **solution** (aqueous) states.

✓ Reason:

- In molten and aqueous states the ions are free and mobile and thus move about conducting electric current.

Gradation in properties of some ionic compounds of sodium.

Property of compound	Compound of sodium			
	Sodium fluoride	Sodium chloride	Sodium bromide	Sodium iodide
Solubility in water	Soluble	Soluble	Soluble	Soluble
Melting point (°C)	993	801	747	661
Boiling point (°C)	1695	1413	1390	1304
Electrical conductivity	Solid	Does not	Does not	Does not
	Molten / solution	Conducts	Conducts	Conducts

Note:

- Solubility of the compounds decrease from sodium fluoride to sodium iodide.
- Melting and boiling points decrease from sodium fluoride to sodium iodide.

2. The covalent bond.

✓ Meaning:

- Refers to a bond formed when two atoms of the same or of different elements share electrons to become stable.
- Formation of covalent bond between atoms (similar or dissimilar) result to the formation of a **molecule**.
- Covalent bonds are usually formed by the association of non-metals.

✓ Note:

- A **molecule** is a group of atoms (two or more) of the same or different elements that are held together by strong covalent bonds.

Formation of a covalent bond.

- Covalent bonding is brought about by the facts that the electro-positivity and the electro-negativity of the elements involved are very close.
- For that reason, none of the atoms can completely lose its valence electrons to the next atom.
- For this reason, both atoms donate electrons which are then shared between them.
- Both atoms thus attain a stable noble gas (duplet or octet) configuration.

Illustrations:

✓ Formation of chlorine molecule.

- Each chloride atom has electronic configuration 2.8.7 and thus need to gain a single electron in the outermost energy level to attain a stable noble gas configuration.
- Since both chloride atoms have same electro-negativities, none will easily lose an electron to the other.
- For this reason, both donate the number of electrons required by the other atom (in this case 1), which they share between them.
- Thus chlorine molecule is formed by sharing 2 electrons between two chlorine atoms, hence a **single covalent bond**.

Diagram: formation of chlorine molecule.

Dot and cross diagram for chlorine molecule.

Formation of oxygen molecule.

- Each oxygen atom has electronic configuration 2.8.6 and thus need to gain 2 electrons into the outermost energy level to attain a stable noble gas configuration.
- Since both oxygen atoms have same electro-negativities, none will easily lose an electron to the other.
- For this reason, both donate the number of electrons required by the other atom (in this case 2), which they share between them.
- Thus oxygen molecule is formed by sharing 4 electrons (2 from each atom) between two chlorine atoms, hence a **double covalent bond**.

Diagram: the oxygen molecule.

Note:

- A single covalent bond is represented in dot (.) and cross (x) diagrams using two dots, two crosses, a dot and a cross or a single line (–) between the atoms involved in the bond.
- Thus a single covalent bond like in chlorine can be represented as Cl – Cl, a double covalent bond like in oxygen can be represented as O = O while a triple covalent bond like in nitrogen can be represented as N ≡ N.

Dot (.) and cross (x) diagrams for various covalent compounds.

1. Hydrogen, H₂
2. Hydrogen chloride, HCl
3. Nitrogen, N₂
4. Water
5. Carbon (IV) oxide, CO₂
6. Ammonia gas, NH₃

7. Phosphene, PH_3
8. Methane, CH_4
9. Ethane, CH_2CH_2
10. Ethyne, C_2H_2
10. Ethanol, $\text{C}_2\text{H}_5\text{OH}$
12. Bromoethane, $\text{C}_2\text{H}_4\text{Br}$.

The coordinate bond.

- Refers to a type of covalent bond in which the shared pair of electrons forming the bond is contributed by only one of the atoms forming the bond.
- It is also called the dative bond.

Examples:

✓ **Formation of ammonium ion.**

- Occurs when an ammonia gas molecule combines with a hydrogen ion (proton).
- All the atoms in the ammonia molecule have a stable noble gas configuration and thus the molecule is stable.
- However the nitrogen in the ammonia molecule has a lone pair of electrons (electrons that have not yet been used in bond formation)
- The hydrogen ion has lost its outermost single valence electron to form the hydrogen ion.
- Thus the hydrogen ion has no electron(s) in its outermost energy level.
- To be stable the hydrogen needs two electrons in its outermost energy level.
- The hydrogen ion thus accepts bonds with the lone pair (2) of electrons in the nitrogen of the ammonia molecule forming a dative bond.
- The total number of electrons in the ammonium ion is 11 while the total number of electrons is 10 leading to a net positive charge of +1

Note:

- In a dot (.) and cross (x) diagram where the covalent bond is represented by horizontal lines (—), the dative // coordinate bond is represented by an arrow (→) pointing the atom that “accepts” the electrons.

Diagram: formation of ammonium ion.

Formation of hydroxonium ion (H_3O^+)

- Occurs when a water molecule combines with a hydrogen ion (proton).
- All the atoms in the water molecule have a stable noble gas configuration and thus the molecule is stable.
- However the oxygen in the ammonia molecule has 2 lone pairs of electrons // four electrons (electrons that have not yet been used in bond formation)
- The hydrogen ion has lost its outermost single valence electron to form the hydrogen ion.
- Thus the hydrogen ion has no electron(s) in its outermost energy level.
- To be stable the hydrogen needs two electrons in its outermost energy level.
- The hydrogen ion thus accepts and bonds with two of the four electrons in the oxygen of the water molecule forming a dative bond // coordinate bond.
- The total number of electrons in the hydroxonium ion is 11 while the total number of electrons is 10 leading to a net positive charge of +1

Diagram: formation of the hydroxonium ion.

Note:

- The hydroxonium ion (H_3O^+) can still further react with another hydrogen ion to form another ion of the formula H_4O^{2+} .
- This is due to the presence of a single lone pair of electrons in the structure of the hydroxonium ion (H_3O^+).

Diagram: formation of H_4O^{2+} from H_3O^+ and H^+

Note:

- The H_4O^{2+} cannot however react further since all the valence electrons in all its atoms have been used in bonding leaving no lone pairs.

Further examples of dative covalent bonds in other compounds.

1. Carbon (IV) oxide
2. Formation of PH_4^+
3. Aluminium chlorine dimer (Al_2Cl_6).
4. Ammonia-aluminium chloride vapours complex, $\text{AlCl}_3 \cdot \text{NH}_3$

Note:

- Substances with covalent bonds form two main types of structures:
 - ✓ Molecular structures
 - ✓ Giant atomic // Giant covalent structures.

1. The Molecular structures.

- Refers to a structure in which covalent bonds holds atoms together to form molecules and the resultant molecules are held together by intermolecular forces.
- Substance with molecular structures are usually gases or liquids at room temperature.
- Few solids such as sulphur, iodine, fats, sugar, naphthalene and paraffin wax also have molecular structures.
- The intermolecular forces in molecular structures are of two types:
 - ✓ Van der Waals
 - ✓ Hydrogen bonds

(a). The Van der Waals.

- Are the weakest form of intermolecular forces due to induced dipole- induced dipole attractions between molecules.
- As the size of the molecule increases, the number of constituent electrons increases leading to increase in strength of the induced dipole – induced dipole interactions.
- The strength of the Van der Waals thus increases as the molecular size increases.

Note:

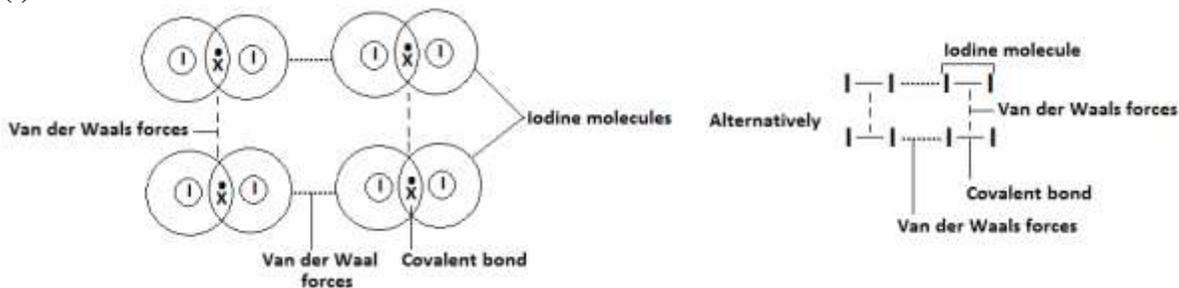
This explains the trend in boiling points for the halogens, which increase from fluorine to iodine.

✓ Reason:

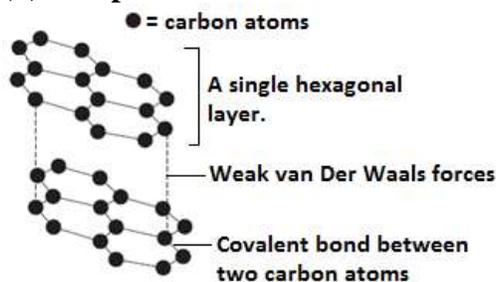
- From fluorine to iodine the size of the atoms and hence the molecules increases leading to increase in molecular masses that lead to stronger induced dipole – induced dipole interactions hence increasing strength of the Van der Waals (from F_2 to I_2)
- This also applies for the increase in boiling points for the homologous series of alkanes.

Diagram: Illustration of Van der Waals forces

(i). Iodine



(ii). Graphite.



(b). The hydrogen bonds.

- Is an intermolecular force in which the electropositive hydrogen atom of one molecule is attracted to an electronegative atom of another molecule.
- The essential requirements for the formation of a hydrogen bond (H – bond) are:
 - ✓ A hydrogen atom attached to a highly electronegative atom.
 - ✓ An unshared pair of electrons on the electronegative atom.

Note:

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

The formation of hydrogen bonds

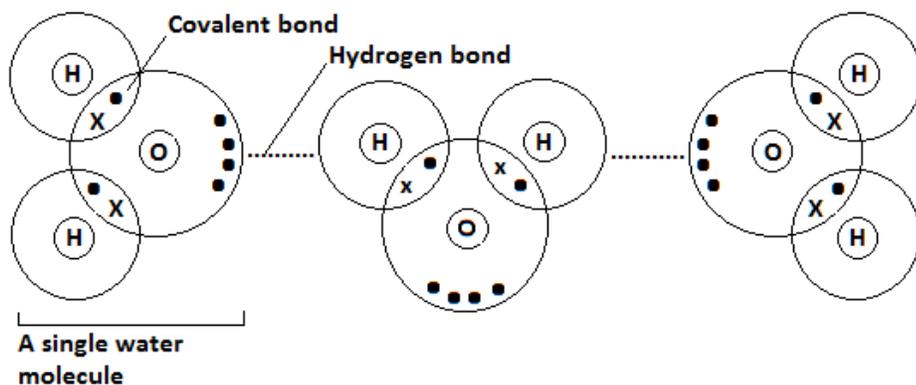
- Occurs when hydrogen atom is bonded to a highly electronegative atom like nitrogen, oxygen and fluorine.
- The electrons in the covalent bond (between hydrogen and the more electronegative atom) are drawn towards the electronegative atom.
- Hydrogen atom has no electrons other than the one it contributes to the covalent bond, which is also being pulled away from it (by the more electronegative atom).
- Hydrogen atom has no outer energy level of electrons making the single proton in the nucleus unusually “bare” .
- The proton is thus readily available for any form of **dipole-dipole attractions**.
- The “bare” proton of the hydrogen atom thus attracts the more electronegative atom (e.g. N, O and F) on either side.
- It thus exerts an attractive force on the more electronegative atom hence bonding them together.
- The two (more electronegative) larger atoms are drawn closer with a hydrogen atom effectively buried in their electron clouds.
- This constitutes the **Hydrogen bond**.

Illustration: formation of hydrogen bonds in water.

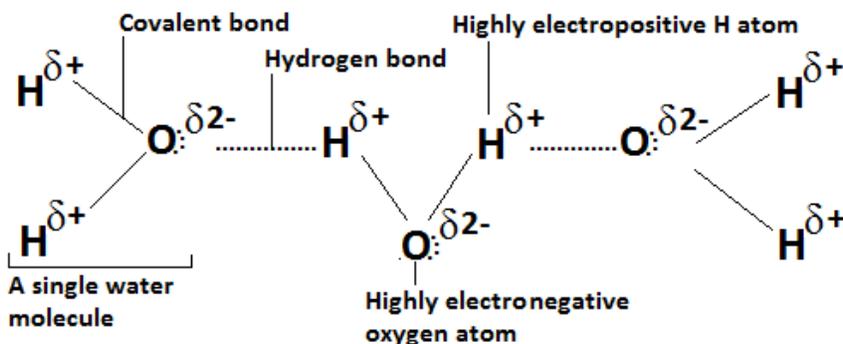
- In water two hydrogen atoms are bonded to an oxygen atom which highly electronegative.
- The electrons in the covalent bond (between each hydrogen and oxygen atom) are drawn towards the more electronegative oxygen atom.
- The hydrogen atoms have no electrons other than their share of those in the covalent bond, which are also being pulled away from them by the oxygen atom.
- Hydrogen atoms have no outer energy level of electrons making the single proton in their nucleus unusually “bare” .
- The proton is thus readily available for any form of **dipole-dipole attractions** (with oxygen in this case).
- The “bare” proton of each of the hydrogen atoms in one molecule thus attracts the more electronegative oxygen atom of the neighboring molecule.

- Each hydrogen atom thus exerts an attractive force on the oxygen atom of the next molecule hence bonding them together.
- The electronegative oxygen atom of one molecule is drawn to the electropositive hydrogen atom of the next molecule with the hydrogen atoms effectively buried in the electron clouds of oxygen.
- This constitutes the **Hydrogen bond** in the water molecule.

Diagram: hydrogen bonds in water molecules.



Alternatively:



Note:

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

Effect of hydrogen bonding on properties of molecular substances.

- Hydrogen bonding tends to disrupt the gradation in physical properties of molecular substances in relation to molecular weights.
- The effect on molecular masses on the melting and boiling points only apply when the intermolecular force is the same.

Examples.

1. Both ethanol (C_2H_5OH) and dimethyl ether (C_2H_6O) have the same relative molecular mass of 46. However the boiling point of ethanol is higher at $78.5^\circ C$ than that of dimethyl ether at only $-24^\circ C$.

Reason:

- Even though both have molecular structures with covalent bonds between the atoms, the intermolecular forces in ethanol are hydrogen bonds which are much stronger than the intermolecular forces in dimethyl ether which are weak Van der Waals forces.

2. Ethanol (C₂H₅OH) has a molecular mass of 46 while butane (C₄H₁₀) has a molecular mass of 58, yet the boiling point of ethanol is higher than that of butane.

Reason:

- Both have molecular structures with covalent bonds between the atoms. However, the intermolecular forces in ethanol are hydrogen bonds which are much stronger (and require more energy to break) than the intermolecular forces in dimethyl ether which are weak Van der Waals forces.

- Molecular substances are generally insoluble in polar solvents like water. However those with hydrogen bonding as the intermolecular forces are soluble in water since the hydrogen bonding confers them some polarity.

Examples:

- Sugar, ethanol, ethanoic acid etc

Properties of molecular structures.

1. Melting and boiling points.

- They generally have low melting and boiling points.

Reason:

- Although the atoms forming the molecules are held together by strong covalent bonds, the intermolecular forces are weak Van der Waals forces which require low amounts of energy to break.

2. Heat and electrical conductivity

- They are poor conductors of heat and electricity at any state.

Reason:

- They have neither delocalized electrons nor free mobile ions for electrical conductivity.

3. Solubility

- Molecular substances are generally insoluble in polar solvents like water but soluble in non-polar organic solvents like benzene.

Reason:

- In polar solvents like water there are strong water – water attractions which are considerably stronger than the intermolecular forces (Van der Waals) attractions or molecule – water (solvent) attractions, making the molecules unable to penetrate the water (solvent) structure for dissolution to occur.

- In non-polar solvents like benzene, the benzene-benzene attractions are similar in strength to the intermolecular forces or the molecule – benzene (solvent) attractions, enabling the molecules to penetrate the solvent thus allowing dissolution.

Note:

- Molecular substances with hydrogen bonds as the intermolecular forces are soluble in polar solvents like water.

Reason:

- The hydrogen bonds in the molecules are equal in strength to the water – water interactions which are also hydrogen bonds, thus the molecules are able to penetrate the structure of water leading to solvation // dissolution // hydration i.e. they are polar like the water molecules.

Summary: Properties of some molecular substances.

Property of compound	Molecular substance					
	Sugar (Sucrose)	Naphthalene	Iodine	Rhombic sulphur	Water	Hydrogen sulphide
Solubility in water	Soluble	Insoluble	Insoluble	Insoluble	-	Slightly soluble
Molecular mass	183	128	186	256	18	34
Melting point (°C)	186	82	113	114	0	-85
Boiling point (°C)	-	218	183	444	100	-60
Electrical conductivity	Solid	Does not	Does not	Does not	Does not	Does not
	Molten / solution	Does not	Does not	Does not	Does not	Does not

2. Giant covalent structures // Giant atomic structures.

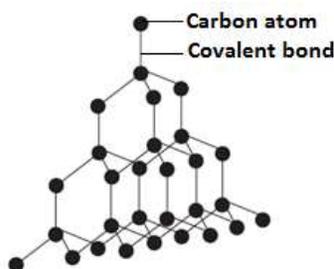
- Are molecular substances in which atoms are linked throughout the whole structure by very strong covalent bonds from one atom to the next.
- The result is an indefinite number of atoms which are all covalently bonded together.
- This pattern occurs repeatedly throughout the structure leading to a “giant” of atoms all covalently bonded.

Examples:

(a). Diamond.

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

Diagram: Structure of diamond:



Properties of diamond.

i. Have high melting and boiling points.

Reason:

- It has a giant atomic structure with strong covalent bonds throughout the structure which require large amounts of energy to break.

ii. It is insoluble in water.

Reason:

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

iii. Does not conduct heat and electricity.

Reason:

- Each carbon atom in the structure of diamond is bonded to four others hence uses all its four valence electrons in bonding and thus lacks any delocalized electrons for electrical conductivity.

iv. It is the hardest substance known.

Reason:

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

(b). Graphite.

- Is also an **allotrope** of carbon.

- In graphite, each carbon atom is bonded to three other carbon atoms by strong covalent bonds.

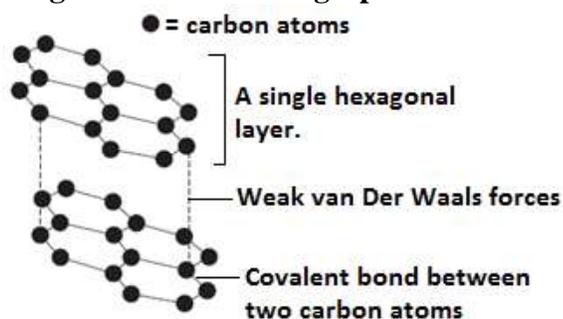
- Since each carbon atom is bonded to only three others, only three of the four valence electrons in each carbon are used in bonding hence presence delocalized electrons in the structure of graphite.

- This explains the electrical conductivity of graphite.

- The carbon atoms in graphite are covalently bonded into hexagonal layers, which are joined to each other by weak Van der Waals forces.

- The presence of weak Van der Waals forces explains the slippery nature of graphite.

Diagram: Structure of graphite.



Properties of Graphite.

i. Have high melting and boiling points.

Reason:

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

ii. Insoluble in water.

Reason:

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

iii. It is a good conductor heat and electricity.

Reason:

- Each carbon atom in the structure of graphite is bonded to three others hence uses only three of its four valence electrons in bonding. This leads to presence of delocalized electrons in the structure of graphite which conducts heat and electricity.

iv. It is soft and slippery.

Reason:

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

(c). Silicon (IV) oxide.

- Is a covalent compound of silicon and oxygen.
- Silicon has four electrons in its outermost energy level while oxygen has six.
- Each silicon atom is bonded to four oxygen atoms by strong covalent bonds.
- Each oxygen atom is bonded to two silicon atoms by strong covalent bonds.
- This means there are no delocalized electrons in the structure of silicon (IV) oxide making it unable to conduct electric current.
- The silicon and oxygen atoms are all covalently bonded together (by strong covalent bonds) in a repeated manner leading to a giant of covalent bonds throughout the structure.
- The extra ordinarily strong covalent bonds in silicon (IV) oxide throuout the structure contribute to the very high melting (1728°C) and boiling (2231°C) points.

Diagram: structure of silicon (IV) oxide.

Properties of giant atomic // giant covalent structures.

- Have high melting and boiling points
- They are non-conductors of heat and electricity with the exception of graphite
- They are insoluble in water
- Most are generally very hard, with the exception of graphite

- ✓ **Note:** for explanations of the above properties, refer to individual explanations of each compound.

3. The metallic bond.

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

Formation of a metallic bond.

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

Diagram: model of the giant metallic structure.

Properties of the giant metallic structure.

1. They have high melting and boiling points.

Reason:

- They have strong metallic bonds throughout the structure which need large amounts of energy to break.
- The strength of the metallic bond increases with decrease in atomic size, as well as with increase in the number of delocalized electrons.
- Thus metals with smaller atoms and more delocalized electrons tend to have stronger metallic bonds hence higher melting and boiling points.
- This explains why the melting and boiling points of metallic elements in period three increase from sodium to aluminium.

2. They are good conductors of heat and electricity.

Reason:

- They have delocalized electrons which heat and electric current.

Note:

- The electrical conductivity of metals increases with increase in the number of delocalized electrons in each atom in the structure.
- This explains why aluminium metal is a better conductor of heat and electricity than both magnesium and sodium.

3. They are insoluble in water.

Reason:

- There are no dipoles in the giant metallic structure and are thus non-polar, so cannot dissolve in polar water molecules.

Summary: some physical properties of metals

Metal	Valency	Melting point (°C)	Boiling point (°C)	Atomic radii (nm)	Electrical conductivity
Lithium	1	180	1330	0.133	Good
Sodium	1	98	890	0.155	Good
Potassium	1	64	774	0.203	Good
Magnesium	2	651	1110	0.136	Good
Aluminium	3	1083	2582	0.125	Good

Summary: Comparing various types of structures.

Attribute	Giant metallic	Giant atomic / giant covalent	Giant ionic	Giant metallic
1. Structure i. Examples	Na, Fe, Cu.	Diamond, SiC, SiO ₂ .	Ca ²⁺ O ²⁻ , (K ⁺) ₂ SO ₄ ²⁻ , Na ⁺ Cl ⁻ ,	I ₂ , S ₈ , C ₁₀ H ₈ , HCl, CH ₄ .
ii. Constituent particles	Atoms	Atoms	Ions	Molecules
iii. Type of substance compound	Metal element with low electronegativity	Non-metal element in group IV or its compound.	Metal/non-metal compound (a compound of elements with a large difference in electronegativity)	Non-metal element or non-metal/non-metal compound (elements with high electronegativity)
2. Bonding: In the solid	Attraction of outer mobile electrons for positive nuclei binds atoms together by strong metallic bonds	Atoms are linked through the whole structure by very strong covalent bonds from one atom to the next.	Attraction of positive ions for negative ions results in strong ionic bonds	Strong covalent bonds hold atoms together within the separate molecules; separate molecules are held together by weak intermolecular forces.
3. Properties. i. Volatility: State at room temp.	Non-volatile. Very high melting and very high boiling points. Usually solid	Non-volatile. Very high melting and very high boiling points. Solid	Non-volatile. Very high melting and very high boiling points. Solid	Volatile. Low melting and low boiling points. Usually gases or volatile liquids
ii. Hardness // malleability	Hard, yet malleable.	Very hard and brittle	Hard and brittle	Soft
iii. Conductivity:	Good conductors when solid or molten.	Non-conductors in any state (except graphite)	Non-conductors when solid; good conductors when molten or in aqueous solution (electrolytes)	Non-conductors when solid, molten and in aqueous solution. (A few like HCl react with water to form electrolytes)
iv. Solubility:	Insoluble in polar and non-polar solvents but soluble in molten metals	Insoluble in all solvents	Soluble in polar solvents (e.g. H ₂ O), insoluble in polar solvents like tetrachloromethane, CCl ₄ .	Polar molecules e.g. HCl are soluble in polar solvents like water; but insoluble in non-polar solvents like CCl ₄ and vice-versa.

Types of bonds across a period.

- The number of valence electrons play an important role in determination of chemical bonding.
- Across a period in the periodic table, the nature of the bonds varies from metallic to covalent.
- The structure also thus varies from giant metallic to simple molecular.
- Thus similar compounds of the elements in period 3 will also exhibit variation in bond types, structures and properties.

Variation in bond types in oxides of period three elements.

Oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₂	Cl ₂ O ₇
Physical state	Solid	Solid	Solid	Solid	Solid	Gas	Gas
M.P (°C)	1193	3075	2045	1728	563	-76	-60
B.P (°C)	1278	3601	2980	2231	301	-10	-9
Structure	Giant ionic	Giant ionic	Giant ionic	Giant atomic	Molecular	Molecular	Molecular
Bonding	Ionic	Ionic	Ionic	Covalent	Covalent with Van der Waals	Covalent with Van der Waals	Covalent with Van der Waals
Nature of oxide	Basic (alkaline)	Basic (weakly alkaline)	Amphoteric	Acidic	Acidic	Acidic	Acidic
Solubility in water	Dissolves to forms an alkaline solution	Dissolves to forms an alkaline solution	Insoluble	Insoluble	Dissolves in water to form acidic solution.	Dissolves in water to form acidic solution.	Dissolves in water to form acidic solution.
Reaction with 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

Note:

1. - The melting and boiling points of the magnesium oxide is higher than that of sodium oxide

Reason:

- Both have giant ionic structures. However the electrostatic forces of attraction between magnesium ions and oxide ions are stronger due to the fact that magnesium ion has a charge of +2 and is smaller in size than the sodium ion.

Effect of bond types of properties of chlorides of period 3 elements

- Most period 3 elements form stable chlorides.
- The trend in bond types, structures and properties of chlorides of period 3 elements show variation across the period.

Properties:

i. Reaction with water.

Procedure:

- A test tube is half filled with water and initial temperature of the water recorded.
- A spatula end full of sodium chloride is added and stirred until it dissolves.
- The highest temperature attained when all the solid dissolves is recorded and the temperature change calculated.
- Two – three drops of universal indicator are added and the pH of the solution noted and recorded.

Observations:

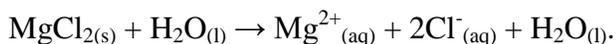
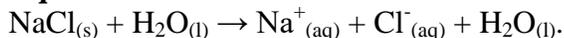
Chloride	Observations		
	Solubility	Temperature change (°C)	pH of solution
Sodium chloride	Dissolves	Drop in temperature	7
Magnesium chloride	Dissolves	Slight increase	6.5
Aluminium chloride	Hydrolyzed	Increases	3
Silicon (IV) oxide	Hydrolyzed	Increases	2
Phosphorus (III) chloride	Hydrolyzed	Increases	2
Phosphorus (V) chloride	Hydrolyzed	Increases	2

Explanations.

✓ Sodium and magnesium chlorides.

- Sodium chloride dissolves in water causing a slight drop in temperature.
- Magnesium dissolves readily with a small increase in temperature.
- Both chlorides are ionic and when added to water there is an immediate attraction of polar water molecules for ions in the chlorides.
- The solid thus readily dissolves forming aquated ions such as $\text{Na}^+_{(\text{aq})}$ and $\text{Cl}^-_{(\text{aq})}$.
- These are separate metal and non-metal ions surrounded by polar water molecules.
- Since there is no production of either H^+ or OH^- ions the solutions are neutral.

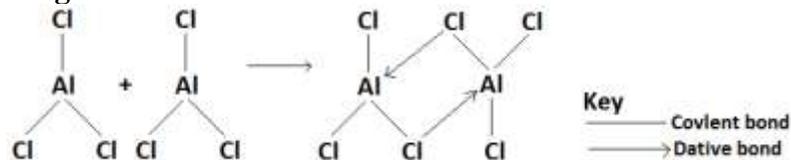
Equations:



✓ Anhydrous aluminium chloride.

- It exists in molecular form as a dimeric molecule of Al_2Cl_6 .
- The dimeric molecule is formed when aluminium chloride (AlCl_3) molecules vapour condense and combine forming larger molecules of Al_2Cl_6 .

Diagram: Formation of a dimer in aluminium chloride.

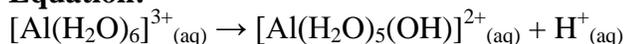


- When added to water aluminium chloride is hydrolyzed to form an acidic solution.
- By so doing it behaves like a covalent chloride rather than an ionic chloride.
- The hydrolysis is an exothermic reaction accompanied by release of heat hence the increase in temperature.

Explanations.

- The hydrolysis is due to the very small but highly charged aluminium ion, Al^{3+} .
- The Al^{3+} draws electrons away from its surrounding water molecules and causes them to give up H^+ ions.
- This reaction usually involves aluminium ions combining with six water molecules to form hexa-aqua-aluminium (III) ions which dissociate to give H^+ .

Equation:



Note:

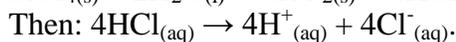
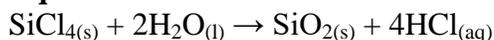
This reaction can be **simplified** as follows:

- Aluminium chloride reacting with water and hydrolyzing to give $\text{HCl}_{(\text{aq})}$ as one of the products.
 - ✓ **Equation:** $\text{Al}_2\text{Cl}_{6(\text{s})} + 6\text{H}_2\text{O}_{(\text{l})} \rightarrow 2\text{Al}(\text{OH})_{3(\text{aq})} + 6\text{HCl}_{(\text{aq})}$
- Then the HCl dissociates to give H^+ and Cl^-
 - ✓ **Equation:** $6\text{HCl}_{(\text{aq})} \rightarrow 6\text{H}^+_{(\text{aq})} + 6\text{Cl}^-_{(\text{aq})}$
- It is the resultant H^+ that confer the solution its acidic properties.

✓ Silicon (IV) chloride

- Undergoes hydrolysis in water in an exothermic reaction producing a lot of heat.
- The products of the hydrolysis are silicon (IV) oxide solid and hydrogen chloride gas.
- The hydrogen chloride gas immediately dissolves in the water to form hydrochloric acid.
- The hydrochloric acid dissociated to liberate H^+ which leads to acidic conditions.

Equations:

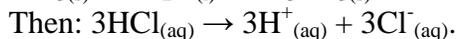
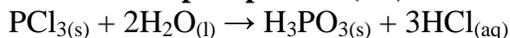


✓ Phosphorus (III) chloride and phosphorus (V) chloride

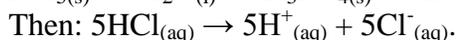
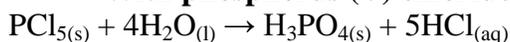
- Both undergo hydrolysis in water in an exothermic reaction producing a lot of heat.
- The products of the hydrolysis are phosphorus (III) acid and phosphoric (V) acid respectively, and hydrogen chloride gas.
- The hydrogen chloride gas immediately dissolves in the water to form hydrochloric acid.
- The hydrochloric acid dissociated to liberate H^+ which leads to acidic conditions.

Equations:

✓ With phosphorus (III) chloride:



✓ With phosphorus (V) chloride:



Trends in bond types and properties of chlorides of elements in period 3

Property	Period 3 chloride					
	NaCl	MgCl ₂	Al ₂ Cl ₆	SiCl ₄	PCl ₃ & PCl ₅	SCl ₂
Formula	NaCl	MgCl ₂	Al ₂ Cl ₆	SiCl ₄	PCl ₃ & PCl ₅	SCl ₂
Physical state at RT	Solid	Solid	Solid	Liquid	Liquid	Liquid
M.P (°C)	801	714	Sublimes at 180°C	-70	Sublimes at -94°C	-78
B.P (°C)	1467	1437	-	57	-	Decomposes at 57°C
Conductivity	Good	Good	V. poor	nil	nil	Nil
Structure	Giant ionic	Giant ionic	Molecular dimer	Molecular	Molecular	Molecular
Bond type	Ionic	Ionic	Ionic / covalent	Covalent	Covalent	Covalent
Effect on water	Readily dissolves	Readily dissolves	Hydrolyzed to give HCl fumes	Hydrolyzed to give HCl fumes	Hydrolyzed to give HCl fumes	Hydrolyzed to give HCl fumes
pH of solution	7	6.5	3	2	2	2

Summary: Characteristics of bonds

Property	Substances with		
	Covalent bonds	Ionic bonds	Metallic bonds
Electrical conductivity	Non-conductors except graphite	- Solids do not conduct. - Aqueous solutions and molten state conduct	- Conducts
Thermal conductivity	- Non-conductors except graphite	- Do not conduct	- Conducts
Melting point (oC)	- Low for molecular substances - High for giant atomic structures	- Usually high.	- Generally high
Boiling point (oC)	- Low for molecular substances - High for giant atomic structures	- Usually high	- Generally high
Solubility	- Generally insoluble in water but soluble in organic solvents	- Generally soluble in water	- Some metals react with water

UNIT 5: SALTS.

1. Meaning.

2. Types of salts:

- ✓ Normal salts
- ✓ Acid salts
- ✓ Basic salts
- ✓ Double salts.

3. Solubility of salts in water

- ✓ Sulphates
- ✓ Chlorides
- ✓ Nitrates
- ✓ Carbonates
- ✓ Sodium, potassium and ammonium salts

4. Solubility of bases in water.

5. Obtaining crystals

6. Preparation of salts

- ✓ Insoluble salt
 - Double decomposition
- ✓ Soluble salts

A sodium, potassium and ammonium salts

A non-potassium, sodium and ammonium salt.

That reacts with water

Direct synthesis

That does not react with water

Acid and metal method

Acid + base method

Neutralization

7. Uses of salts

8. Action of heat on salts

- ✓ Carbonates
- ✓ Nitrates
- ✓ Sulphates.