

ATOMIC STRUCTURE (9 Periods)

An atom is the smallest indivisible particle of an element which can exist on its own and take part in a chemical reaction. According to the work carried out by Rutherford and other scientists, atoms are made up of three sub-atomic particles which differ in their mass, charge and location within the atom.

Sub-atomic particles of an atom

- i. Electrons (e)
- ii. Neutrons (n)
- iii. Protons (p)

An atom is made of the tiny nucleus at its centre which contains the protons and neutrons whereas the electrons are found in a large volume of space outside the nucleus of an atom. The mass of an atom is mainly concentrated in the nucleus because the neutrons and protons are found in it. Neutrons and protons are called **nucleons or nuclear particles** because they are found in the nucleus of an atom.

Protons: They are found inside the nucleus of an atom. Protons are positively charged (They have a charge of +1) and have a mass of 1. The nucleus of an atom is positively charged because it contains a number of protons. Protons being positively charged repel one another; however, this repulsion is minimized by the presence of the neutrons.

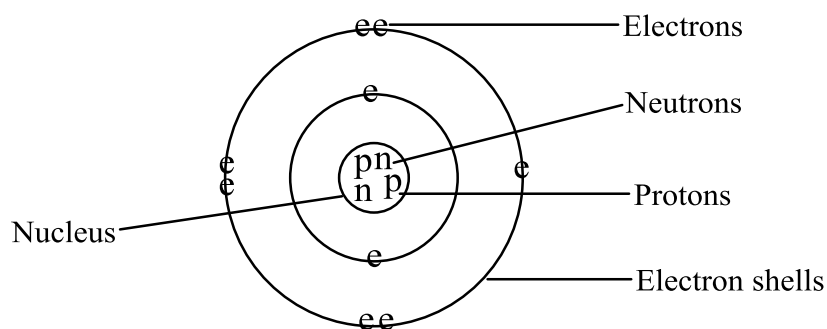
Neutrons: The neutrons are also found in the nucleus and they help minimize repulsive forces that would be between protons. They have a mass of 1 which is approximately equal to that of a proton and have no charge (are neutral).

Electrons: Electrons are found in a large volume of space around the nucleus revolving in pairs in definite paths called energy levels or electron shells or orbitals. They are negatively charged particles (They have a charge of -1) and have a mass of $\frac{1}{1840} \approx 0$.

Summary of characteristics of sub-atomic particles

| Particle and symbol | Mass | Charge | Location in an atom |
|---------------------|------------------|--------|----------------------------------|
| Proton (p) | 1 | +1 | Nucleus |
| Neutron (n) | 1 | 0 | Nucleus |
| Electron (e) | $\frac{1}{1840}$ | -1 | Energy levels around the nucleus |

ARRANGEMENT OF SUB-ATOMIC PARTICLES IN AN ATOM



TERMS RELATED TO ATOMIC STRUCTURE

a) Atomic number (Z)

This is the number of protons in the nucleus of an atom. It is normally denoted by letter Z. In a neutral atom (one which has not lost or gained electrons), the number of protons is always equal to the number of electrons. Since the protons have a charge of positive one and the electrons have a charge of negative one, the overall charge of the atom becomes zero and hence the atom becomes neutral.

b) Mass number or Atomic mass (A)

This is the sum of protons and neutrons in the nucleus of an atom of an element. It is normally denoted by letter A. This implies that; Mass number = protons + neutrons.

For example lithium with three (3) protons and four (4) neutrons, has atomic number of 3 and the mass number of 7 ($3p + 4n$).

Table showing number of protons, neutrons and electrons of some selected elements

| Atom | No. of protons (Atomic number) | No. of electrons | No. of neutrons | Mass number |
|-----------|-----------------------------------|------------------|-----------------|-------------|
| Hydrogen | 1 | 1 | 0 | 1 |
| Carbon | 6 | 6 | 6 | 12 |
| Flourine | 9 | 9 | 10 | 19 |
| Sodium | 11 | 11 | 12 | 23 |
| Magnesium | 12 | 12 | 12 | 24 |
| Calcium | 20 | 20 | 20 | 40 |

c) Isotopes

Isotopes are different atoms of the same element having the same atomic number but different mass numbers. The difference between the atomic and mass numbers in an atom is due to the difference in the number of neutrons.

OR: Isotopes are different atoms of the same element with the same number of protons but different number of neutrons.

Isotopy: This is the existence of different atoms of the same element with the same number of protons but different number of neutrons. **OR:** Is the existence of different atoms of the same element with the same atomic number but different mass numbers.

The elements that exhibit isotopy include;

- i. Carbon. It has three isotopes which can be represented as ${}^1_6\text{C}$, ${}^{13}_6\text{C}$ and ${}^{14}_6\text{C}$.
- ii. Chlorine. It has two isotopes which are represented as ${}^{35}_{17}\text{Cl}$ and ${}^{37}_{17}\text{Cl}$.
- iii. Hydrogen. It has three isotopes i.e hydrogen, deuterium and tritium. The isotopes are represented as ${}^1_1\text{H}$, ${}^2_1\text{H}$ and ${}^3_1\text{H}$.
- iv. Oxygen. It has two isotopes, oxygen-16 (${}^{16}_8\text{O}$) and oxygen-18 (${}^{18}_8\text{O}$)

NB: In general, an atom is represented as ${}^A_Z\text{X}$. Where X is the symbol of the atom, A is the mass number or Atomic mass and Z is the atomic number.

Mass number minus atomic number (A – Z) is equal to the number of neutrons.

d) Relative Atomic Mass (R.A.M)

The relative atomic mass is the mass of an atom of an element relative to a twelfth of the mass of an atom of carbon-12-isotope.

$$\text{i.e. RAM of an element} = \frac{\text{Mass of one atom of an element}}{\frac{1}{12} \text{ of the mass of one atom of carbon-12-isotope}}$$

Since the relative atomic mass is simply a ratio, it has **no units** since the units cancel out.

Relative atomic masses of some elements are not whole numbers because they are isotopic and have varying quantities such as;

- i. Mass numbers of isotopes.
- ii. Relative abundance of isotopes.

For example a sample of chlorine gas contains two isotopes chlorine-35 and chlorine-37 whose relative abundance is 75% and 25% respectively. The relative atomic mass of chlorine can be obtained as;

$$\begin{aligned} \text{R.A.M of chlorine} &= \frac{(\text{Mass} \times \% \text{ composition of chlorine 35}) + (\text{Mass} \times \% \text{ composition of chlorine 37})}{100} \\ &= \frac{(35 \times 75) + (37 \times 25)}{100} \\ &= 35.5 \end{aligned}$$

ELECTRONIC CONFIGURATION AND ELECTRONIC STRUCTURE

Electronic configuration

Electronic configuration is the arrangement of electrons in the electronic shells/ orbitals or energy levels of an atom. Each energy level can hold a fixed number of electrons. The first energy level takes a maximum of two (2) electrons, the second energy level takes a maximum of eight (8) electrons and the third energy level takes a maximum of eighteen (18) electrons. It should be noted that after adding eight electrons to the third energy level, any extra electron added goes into the fourth energy level and not in the third energy level as expected. The reason for this is beyond this course. ***Therefore for this course, the third energy level should be taken***

to accommodate a maximum of eight electrons and any remaining electrons are added to the fourth energy level.

The table showing the symbol, number of electrons, number of electrons in each shell and electronic configuration of the first twenty (20) elements is shown below.

| Element | Symbol | No. of electrons | No. of electrons in each shell | | | | Electronic Configuration |
|------------|--------|------------------|--------------------------------|-----------------|-----------------|-----------------|--------------------------|
| | | | 1 st | 2 nd | 3 rd | 4 th | |
| Hydrogen | H | 1 | 1 | | | | 1 |
| Helium | He | 2 | 2 | | | | 2 |
| Lithium | Li | 3 | 2 | 1 | | | 2: 1 |
| Beryllium | Be | 4 | 2 | 2 | | | 2: 2 |
| Boron | B | 5 | 2 | 3 | | | 2: 3 |
| Carbon | C | 6 | 2 | 4 | | | 2: 4 |
| Nitrogen | N | 7 | 2 | 5 | | | 2: 5 |
| Oxygen | O | 8 | 2 | 6 | | | 2: 6 |
| Fluorine | F | 9 | 2 | 7 | | | 2: 7 |
| Neon | Ne | 10 | 2 | 8 | | | 2: 8 |
| Sodium | Na | 11 | 2 | 8 | 1 | | 2: 8: 1 |
| Magnesium | Mg | 12 | 2 | 8 | 2 | | 2: 8: 2 |
| Aluminium | Al | 13 | 2 | 8 | 3 | | 2: 8: 3 |
| Silicon | Si | 14 | 2 | 8 | 4 | | 2: 8: 4 |
| Phosphorus | P | 15 | 2 | 8 | 5 | | 2: 8: 5 |
| Sulphur | S | 16 | 2 | 8 | 6 | | 2: 8: 6 |
| Chlorine | Cl | 17 | 2 | 8 | 7 | | 2: 8: 7 |
| Argon | Ar | 18 | 2 | 8 | 8 | | 2: 8: 8 |
| Potassium | K | 19 | 2 | 8 | 8 | 1 | 2: 8: 8: 1 |
| Calcium | Ca | 20 | 2 | 8 | 8 | 2 | 2: 8: 8: 2 |

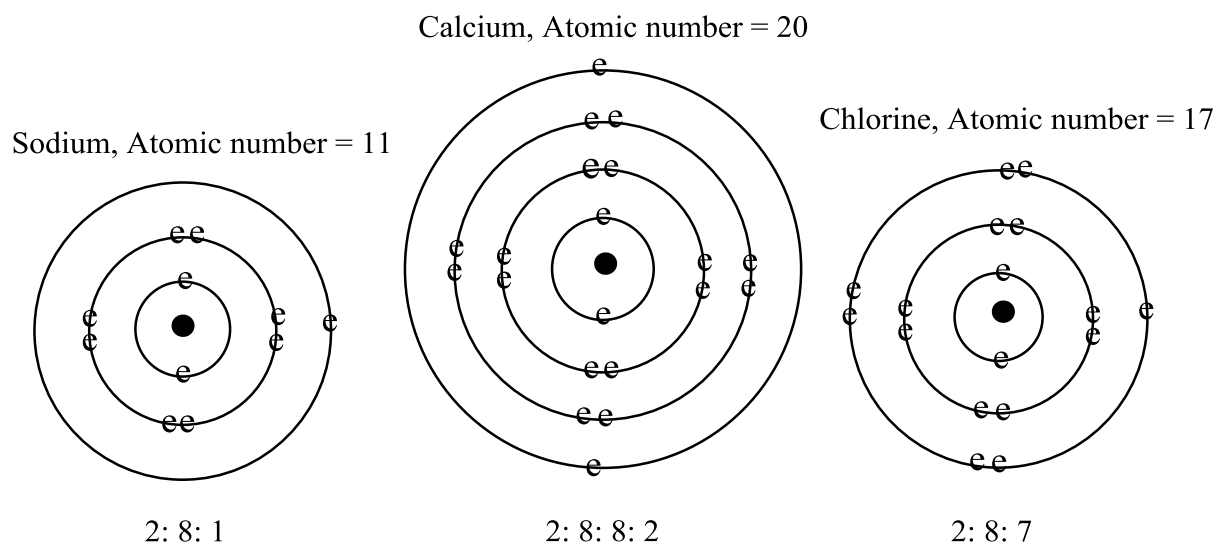
The electronic configuration can be used to determine the group and period in which an element belongs or falls.

- The group in which an element belongs is determined by the number of electrons in its outer most shell and it must be written in Romans.
- The period in which an element falls corresponds with the number of electronic shells occupied by electrons.

Electronic structure

The electronic structure is the diagrammatic representation of the arrangement of electrons in the energy levels of an atom.

Structure of atoms



Question: Draw the electronic structures of the first 20 elements in the periodic table.

THE PERIODIC TABLE (3 Periods)

Periodic table

This is a table that shows the arrangement of elements in order of increasing atomic number with elements placed in vertical columns according to the number of electrons in the outer most shells and in horizontal rows according to the number of energy levels occupied by electrons of an element. The horizontal rows are called periods and the vertical columns are called groups. Elements with the same number of electrons in the outer most shell are placed in the same group while those with same number of orbitals/ energy levels occupied by electrons are placed in the same group. Groups are written using **Roman numerals** while periods are written using **Arabic numerals**.

History of the periodic table

The first periodic table was brought forward in 1869 by a Russian chemist called Dmitri Mendeleev who was a professor of chemistry at St Petersburg. He arranged elements in order of increasing atomic mass/ mass number and left gaps for elements that were not yet discovered by that time. He then arranged the elements in columns such that those with similar properties such as fluorine, chlorine, bromine etc. appeared in the same row. The gaps of elements in this periodic table encouraged Mendeleev and other chemists to search for new elements which are today presented in a modern periodic table in which elements are arranged in order of increasing atomic number.

ARRANGEMENT OF THE FIRST 20 ELEMENTS IN THE PERIODIC TABLE

| | I | II | | III | IV | V | VI | VII | VIII/O |
|----------|---|---|---------------------|---|---|--|--|---|---|
| Period 1 | $\begin{matrix} 1 \\ 1 \\ \text{H} \end{matrix}$ | | | | | | | | $\begin{matrix} 2 \\ 2 \\ \text{He} \end{matrix}$ |
| Period 2 | $\begin{matrix} 3 \\ 3 \\ \text{Li} \end{matrix}$ | $\begin{matrix} 4 \\ 4 \\ \text{Be} \end{matrix}$ | | $\begin{matrix} 5 \\ 5 \\ \text{B} \end{matrix}$ | $\begin{matrix} 6 \\ 6 \\ \text{C} \end{matrix}$ | $\begin{matrix} 7 \\ 7 \\ \text{N} \end{matrix}$ | $\begin{matrix} 8 \\ 8 \\ \text{O} \end{matrix}$ | $\begin{matrix} 9 \\ 9 \\ \text{F} \end{matrix}$ | $\begin{matrix} 10 \\ 10 \\ \text{Ne} \end{matrix}$ |
| Period 3 | $\begin{matrix} 11 \\ 11 \\ \text{Na} \end{matrix}$ | $\begin{matrix} 12 \\ 12 \\ \text{Mg} \end{matrix}$ | | $\begin{matrix} 13 \\ 13 \\ \text{Al} \end{matrix}$ | $\begin{matrix} 14 \\ 14 \\ \text{Si} \end{matrix}$ | $\begin{matrix} 15 \\ 15 \\ \text{P} \end{matrix}$ | $\begin{matrix} 16 \\ 16 \\ \text{S} \end{matrix}$ | $\begin{matrix} 17 \\ 17 \\ \text{Cl} \end{matrix}$ | $\begin{matrix} 18 \\ 18 \\ \text{Ar} \end{matrix}$ |
| Period 4 | $\begin{matrix} 19 \\ 19 \\ \text{K} \end{matrix}$ | $\begin{matrix} 20 \\ 20 \\ \text{Ca} \end{matrix}$ | Transition elements | | | | | | |

$\begin{matrix} 16 \\ 8 \\ \text{O} \end{matrix}$

 → Mass number
 → Symbol
 → Atomic number

Note: A modern periodic table has 8 groups and 7 periods. However for this course we shall only consider the first 20 elements.

Metals, non-metals and noble gases

The table above shows that the metals are on the left whereas non-metals are on the right. Semi-metals or metalloids are on the boundary (border line) between metals and non-metals e.g. Boron, Silicon and so on. Generally, metals have 1 to 3 electrons in their outer most shells (i.e. they belong to group I, II and III) whereas non- metals have 4 to 7 electrons in their outer most shell (i.e. they belong to group IV, V, VI and VII).

Noble gases also called inert gases belong to group O or VIII of the periodic table. Elements in this group have their outer most shells filled up with electrons.

Transition metals

These elements form a separate block in the middle of the periodic table. They include; copper, zinc, iron, lead etc. Elements in this group are characterised by having two or more valencies (e.g. copper has valencies 1 and 2), forming coloured compounds (iron (III) chloride is brown in colour) etc.

The reactivity of the elements

Metals react by losing electrons whereas non-metals react by gaining electrons. The ease at which metals lose or non-metals gain electrons determines their reactivity. If a metal readily/easily loses electrons then it becomes most reactive. On the other the most reactive non-metal easily gains electrons.

The most reactive metals are in group I (Alkali metals). The reactivity of metals increases down the group because of the ease with which they lose the electrons down the group. Therefore, potassium is the most reactive metal in the table above. However, the reactivity of metals decreases across a particular period. For example considering elements in period 2 of the periodic table, lithium is the most reactive followed by beryllium and so on.

For non-metals, the most reactive are group VII elements (halogens). The reactivity of non-metals decreases down the group hence fluorine is the most reactive non-metal. However, it increases from right to left across the period. The elements in group O are unreactive and are therefore called inert gases.

Groups in the periodic table

The vertical columns in the periodic table are called groups and it contains elements with similar properties. A modern periodic table has eight groups some of which have specific names for example alkali metals (Group I), alkaline earth metals (Group II), halogens (Group VII) and noble gases (Group VIII/O). The general characteristics of the groups are;

- i. Elements in the same group have the same number of electrons in their outer most shells. For example oxygen and sulphur have electronic configurations of 2: 6 and 2: 8: 6, respectively. Both elements have six electrons in their outer most shells and therefore belong to group VI of the periodic table.
- ii. The atomic size of the elements increases down the group as more electron shells are added.
- iii. Elements in the same group have similar chemical properties because they have the same number of electrons in their outermost shells.
- iv. The elements in the same group have a common valency. For example, group I elements have a valency of 1 because they lose one electron in order to obtain a full octet.

NB: The valency of elements in group I to group III is equal to the group number where as the valency of elements in group IV to group VIII equals to 8 minus the group number of the element.

Summary

| | | | | | | | | |
|----------------|---|----|-----|----|---|----|-----|------|
| Group | I | II | III | IV | V | VI | VII | VIII |
| Valency | 1 | 2 | 3 | 4 | 3 | 2 | 1 | 0 |

Periods in the periodic table

Periods are the horizontal rows and they have the following characteristics;

- i. Elements in the same period have the same number of energy levels occupied by electrons.
- ii. The atomic size of elements decreases across the period from the left to the right due to the strong nuclear attraction for the outermost electrons.
- iii. Across the period, each element has one electron more than the previous element.

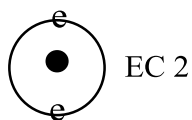
IONS AND IONIC COMPOUNDS (6 Periods)

An ion is a charged atom. Ions can either be positive or negative. Positively charged ions are called **cations** whereas negatively charged ions are called **anions**.

The concept of duplet and octet

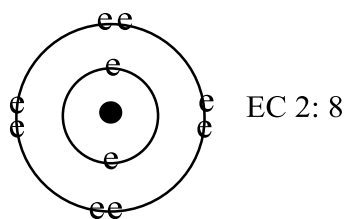
Noble gases like helium, neon, argon etc. have completely filled outermost energy level. The characteristic feature is an outermost shell which has an electron duplet (have 2 electrons in the outer most shell) for the case of helium and an electron octet (have 8 electrons in the outer most shell) for the case of neon and argon. This makes noble gases stable and reluctant to react.

Helium, Atomic number = 2



Complete outermost energy level in helium, 2 (Duplet)

Neon, Atomic number = 10



Complete outermost energy level in neon, 8 (Octet)

Other atoms however do not have complete outermost energy levels and seek to react by either losing or gaining electrons in order to obtain stable octets or duplets.

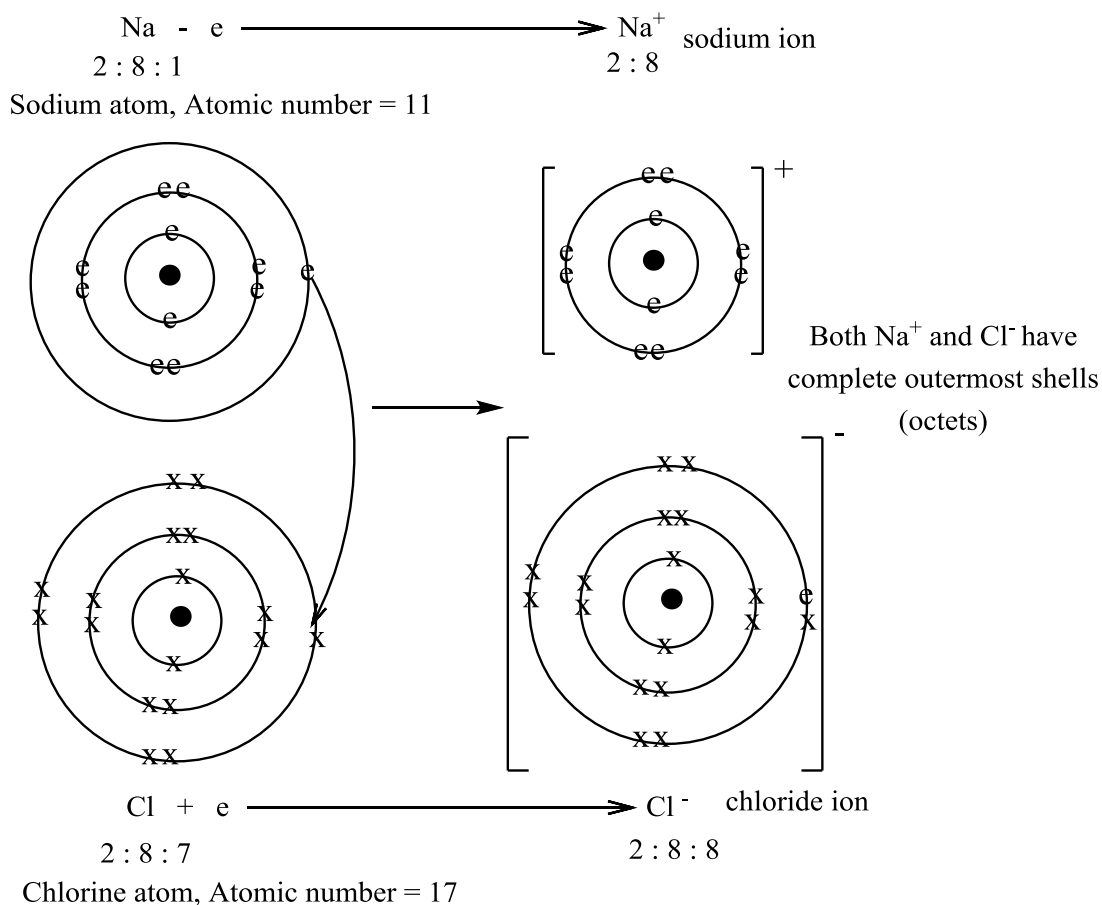
VALENCY

This is the number of electrons shared, gained or lost by an atom of an element in order to obtain a stable noble gas electronic configuration (in order to obtain a full outer most shell or stable octet or duplet)

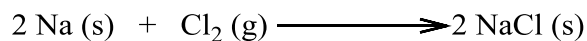
| Element | Symbol | Valency |
|------------|--------|-------------|
| Hydrogen | H | 1 |
| Helium | He | - |
| Lithium | Li | 1 |
| Beryllium | Be | 2 |
| Boron | B | 3 |
| Carbon | C | 4 |
| Nitrogen | N | 3 |
| Oxygen | O | 2 |
| Fluorine | F | 1 |
| Neon | Ne | - |
| Sodium | Na | 1 |
| Magnesium | Mg | 2 |
| Aluminium | Al | 3 |
| Silicon | Si | 4 |
| Phosphorus | P | 3 or 5 |
| Sulphur | S | 2 |
| Chlorine | Cl | 1 |
| Argon | Ar | - |
| Potassium | K | 1 |
| Calcium | Ca | 2 |
| Bromine | Br | 1 |
| Iodine | I | 1 |
| Zinc | Zn | 2 |
| Iron | Fe | 2 or 3 |
| Lead | Pb | 2 or 4 |
| Copper | Cu | 1 or 2 |
| Silver | Ag | 1 |
| Barium | Ba | 2 |
| Manganese | Mn | 2 or 3 or 4 |
| Gold | Au | - |

Reaction between sodium and chlorine

Sodium burns in chlorine gas forming a white solid of sodium chloride. During the reaction, sodium atom gives up its single outermost electron to chlorine atom forming sodium ion and chloride ion with stable octets.



In both sodium atom and sodium ion, there are eleven protons (+11) but after the reaction sodium atom loses one electron forming sodium ion which will now have ten electrons (-10) giving sodium ion an overall charge of +1. The sodium ion is indicated by symbol Na^+ . In both chlorine atom and chloride ion there are seventeen protons (+17) but after the reaction, the chlorine atom gains an electron from sodium atom forming a chloride ion which will now have eighteen electrons (-18) giving the chloride ion an overall charge of -1. The chloride ion is indicated by the symbol Cl^- . The chemical formula of sodium chloride is NaCl (Na^+ , Cl^-) and the equation of reaction is,



QUESTION: Describe the reaction between magnesium and oxygen.

EXERCISE

1. The number of protons, neutrons and electrons of atoms W, X, Y and Z are shown in the table below.

| Atom | Protons | Neutrons | Electrons |
|------|---------|----------|-----------|
| W | 6 | 6 | 6 |
| X | 9 | 9 | 10 |
| Y | 12 | 12 | 12 |
| Z | 19 | 20 | 18 |

- a) Which one of the atoms is;
- A cation.
 - An anion.
- b) Write the electronic configuration of Z.
- c) State the valency of W? Give a reason for your answer.
2. The number of protons and neutrons of atoms A, B, C and D are shown in the table below.

| Atom | Protons | Neutrons |
|------|---------|----------|
| A | 6 | 6 |
| B | 20 | 20 |
| C | 6 | 8 |
| D | 17 | 20 |

- a) Which of the atoms are isotopes? (Give a reason for your answer)
- b) Which one of the atoms is of an element in group 2 of the periodic table? (Give a reason for your answer)
- c) Write the formula of the most common ion of D.
3. The electronic configuration of X is 2:8:5.
- Write the following of the most common ion of X.
 - To which period of the periodic table does X belong? Give a reason for your answer.
 - Element X reacts with element Y whose atomic number is 11.
 - Write the electronic configuration of atom Y
 - State whether Y is a metal, non-metal or noble gas.

RADICALS

Is the group of atoms chemically combined together that exist in several compounds but they do not exist on their own. The table below shows radicals and their valencies.

| Radical | Formula | Valency |
|--------------------|------------------------------|---------|
| Ammonium | NH_4^+ | 1 |
| Carbonate | CO_3^{2-} | 2 |
| Sulphate | SO_4^{2-} | 2 |
| Nitrate | NO_3^- | 1 |
| Phosphate | PO_4^{3-} | 3 |
| Sulphite | SO_3^{2-} | 2 |
| Nitrite | NO_2^- | 1 |
| Manganate (VII) | MnO_4^- | 1 |
| Hydrogen carbonate | HCO_3^- | 1 |
| Hydrogen sulphate | HSO_4^- | 1 |
| Hydrogen Sulphite | HSO_3^- | 1 |
| Hydroxide | OH^- | 1 |
| Chlorate | ClO_3^- | 1 |
| Chloride | Cl^- | 1 |
| Bromate | BrO_3^- | 1 |
| Dichromate | $\text{Cr}_2\text{O}_7^{2-}$ | 2 |
| Chromate | CrO_4^{2-} | 2 |

FORMULAE OF IONIC COMPOUNDS

A formula is a representation of a compound by means of chemical symbols of the elements and the number of atoms of each element that make up a compound.

Interpretation of a formula

The subscript i.e. the number written at the right hand bottom of a symbol of a given element indicates the number of atoms of that element in the given compound e.g. Iron (III) oxide, Fe_2O_3 has 2 atoms of iron and 3 atoms of oxygen.

The number on the right hand bottom outside the brackets enclosing a radical indicates the number of radicals present in the compound e.g. Ammonium Sulphate $(\text{NH}_4)_2\text{SO}_4$ has 1 sulphate radical and 2 Ammonium radicals.

The number in front of a formula indicates the number of moles of that compound e.g. 4CO_2 means that there are 4 moles of carbon dioxide.

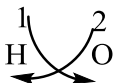
Writing chemical formula of compounds

Write down the symbols of the elements and/or radicals in the same order as they appear in the name of the compound and place their valencies on top. Cross interchange the valencies as shown below such that the negative and positive charges are balanced in the compound.

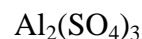
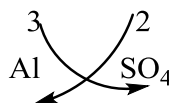
Examples

Write the chemical formula of the following compounds

1. Water



2. Aluminium Sulphate



Exercise: Write the formulae of the following compounds.

| S/N | Name of compound | Formula | S/N | Name of compound | Formula |
|-----|----------------------------|---------|-----|---------------------|---------|
| 1. | Sodium hydroxide | | 9. | Copper (II) nitrate | |
| 2. | Calcium hydrogen carbonate | | 10. | Iron (III) chloride | |
| 3. | Lead (II) oxide | | 11. | Zinc carbonate | |
| 4. | Magnesium chloride | | 12. | Ammonium nitrite | |
| 5. | Magnesium sulphate | | 13. | Aluminium oxide | |
| 6. | Potassium chloride | | 14. | Lead (IV) oxide | |
| 7. | Calcium oxide | | 15. | Ammonium nitrate | |
| 8. | Aluminium fluoride | | 16. | Copper (I) oxide | |

ATOMICITY

Atomicity is the number of atoms in a given molecule or compound. The table below shows atomicity of some elements

| Element | Atomicity |
|---------------------------|-------------|
| Chlorine, Cl ₂ | Diatomic |
| Neon, Ne | Mono atomic |
| Oxygen, O ₂ | Diatomic |
| Magnesium, Mg | Mono atomic |

Steps followed in determining atomicities of compounds.

- Identify the number of moles in the given compound or molecule.
- Determine the number of atoms of each element in the compound.
- Multiply the number of atoms of each element by the number of moles and add up the products to get atomicity.

Examples

Find the atomicities of the following compounds.

- NH₃ Atomicity = 1+3 =4
- 4Al₂O₃ Atomicity = (4×2) + (4×3) = 20
- 2CO₂ Atomicity = (2×1) + (2×2)= 6
- Ca₃(PO₄)₂ Atomicity = (1×3) + (1×2) + (4×2) = 13
- Na₂CO₃ Atomicity = (1×2) + (1) + (1×3) = 6
- Mg(HCO₃)₂ Atomicity = (1×1) + (1×2) + (1×2) + (3×2) = 11

Exercise

Find the atomicities of the following compounds

- 2Pb(NO₃)₂
- 5Na₂O
- 6K₂SO₄
- Calcium hydrogen carbonate
- Zinc sulphate

CHEMICAL FAMILIES (9 Periods)

The elements in the same group have similar physical and chemical properties because they have the same number of electrons in their outer most shells. For this reason, they are sometimes called ‘‘a chemical family’’ of elements. Some of the examples of chemical families include; alkali metals (Group I elements), alkaline earth metals (Group II elements), halogens (Group VII elements) and noble gases.

THE ALKALI METALS (Group I elements)

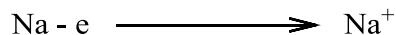
Group I elements in the periodic table are called alkali metals because they react with water forming alkaline solutions. They include; lithium (Li), sodium (Na), potassium (K) and so on and all have a valency of 1 because they have one electron in their outer most shells which is lost during chemical reaction.

Lithium Li 2, 1

Sodium Na 2, 8, 1

Potassium K 2, 8, 8, 1

They are stored or kept under oil/ paraffin because they are highly reactive with water and air. The high reactivity (high electro positivity) is because alkali metals have **one electron** in their outer most shell which is readily lost during chemical combination. The reactivity of group I metals increases down the group because the outer electron distance from the nucleus increases, hence it is less attracted by the nucleus requiring less energy to lose the valency electron.



Physical properties of alkali metals

1. They are all metals. This is because they contain one electron in their outer most shells which is lost during a chemical reaction.
2. They are soft silvery and can be cut easily with a knife.
3. They are good conductors of heat and electricity.
4. They are malleable i.e. can be hammered into sheets.
5. They are kept under oil to prevent them from reacting with ar.

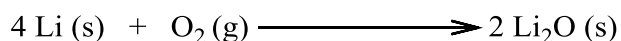
6. They have low densities and float on water. The densities increase down the group.
7. They have low melting and boiling points compared to other metals. This is because they form weak metallic bonds. Their melting and boiling points decrease down the group.
8. They are ductile i.e. they can be drawn into thin wires.

Chemical properties of alkali metals

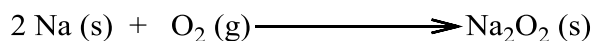
a) Reaction with oxygen (air)

Alkali metals tarnish when exposed to air forming metal oxides and peroxides. These oxides dissolve in water forming colourless solutions of metal hydroxides (alkaline solutions).

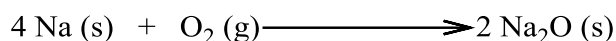
- i. Lithium burns in air with a scarlet flame forming a white solid called lithium oxide.



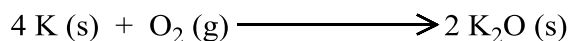
- ii. Sodium burns in air with a yellow flame forming a yellow solid called sodium peroxide.



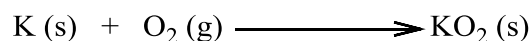
However, in limited supply of oxygen, sodium burns with a yellow flame forming a white solid called sodium oxide.



- iii. Potassium burns with a purple/ lilac flame forming a white solid called potassium oxide.



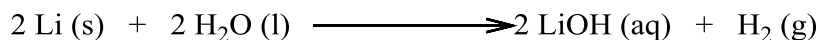
Potassium also burns in air with a purple/ lilac flame forming potassium superoxide.



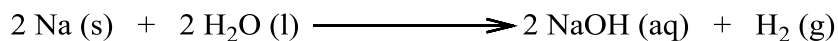
b) Reaction with water.

The alkali metals react violently with water. **Experiment:** Place a piece of a metal in a trough of cold water.

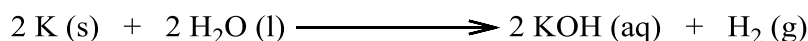
- i. Lithium reacts steadily with cold water exothermically without melting forming a colourless solution of lithium hydroxide (alkaline solution) and bubbles of a colourless gas (hydrogen gas) are evolved.



- ii. Sodium metal floats and immediately melts into a silvery ball that darts on the water surface with a hissing sound with effervescence of a colourless gas (hydrogen gas)



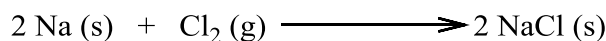
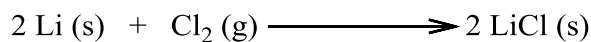
- iii. Potassium reacts violently than sodium. Potassium floats and melts into a silvery ball that darts on the water surface with a hissing sound with evolution of bubbles of a colourless gas, hydrogen gas which explodes and burns with a purple/ lilac flame.



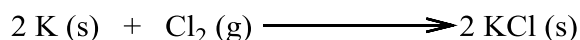
NB: The alkaline solutions formed turn red litmus paper to blue and turn phenolphthalein indicator pink.

- c) Reaction with chlorine gas. **Experiment:** Lower a piece of burning lithium or sodium into a jar of chlorine gas.

- i. Lithium or sodium continues burning with a scarlet flame or bright yellow flame respectively forming white solids of lithium chloride or sodium chloride respectively.



- ii. Potassium reacts explosively with chlorine gas forming potassium chloride which is a white solid.



Uses of sodium and potassium

- i. Sodium is used in the manufacture of sodium cyanide which is used in the extraction of gold.
- ii. Sodium is used in street lights that give off a very intense yellow light.
- iii. Sodium- potassium alloy is used as a coolant in nuclear reactions.
- iv. Sodium is used in the manufacture of sodium peroxide.

- v. Preparation of tetraethyl-lead which is used as an anti-knock in petrol.

ALKALINE EARTH METALS (Group II elements)

Group II elements in the periodic table are called alkaline earth metals because they exist in nature in the earth's crust as carbonates, silicates and sulphates. They include **beryllium (Be)**, **magnesium (Mg)**, **calcium (Ca)** and so on.

| | | | |
|-----------|----|------------|--|
| Beryllium | Be | 2, 2 | ↓ Reactivity increases down the group |
| Magnesium | Mg | 2, 8, 2 | |
| Calcium | Ca | 2, 8, 8, 2 | |

The elements are all metals but are less reactive (less electropositive) than group I elements. The elements have a common valency of 2. This is because they lose 2 electrons in their outer most shells to obtain a stable electronic configuration. They therefore form positively charged cations denoted as Mg^{2+} .

Physical properties of alkaline earth metals

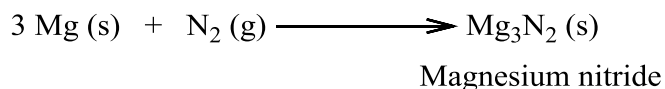
- Polished surfaces of metals are shiny and silvery.
- They have higher density than alkali metals. They don't float on water instead they sink.
- They have higher melting and boiling points compared to alkali metals.
- They are metallic, malleable, ductile and lustrous.
- They are good thermal and electric conductors.

Chemical properties of alkaline earth metals

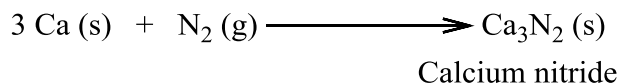
The elements in this group are chemically reactive but not as reactive as group I elements.

a) Reaction with nitrogen gas

Magnesium burns with nitrogen gas with a bright white flame to forming a white solid called magnesium nitride.

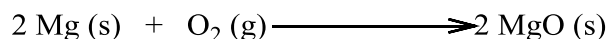


Calcium burns with a bright red flame with nitrogen gas to form calcium nitride which is a white solid.

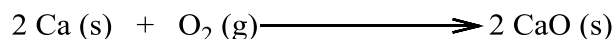


b) Reaction with oxygen gas

- i. Magnesium burns with oxygen with a bright white flame to form a white solid called magnesium oxide.



- ii. Calcium burns with oxygen with a bright red flame to form a white solid called calcium oxide. This reaction is less vigorous because the metal surface reacts so quickly with oxygen that it protects the inside of the metal from reacting.

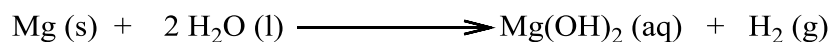


NB: Beryllium does not dissolve in water because it is covalent.

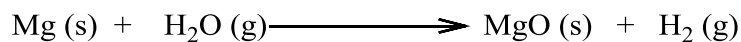
c) Reaction with water.

The reaction of group II elements with water depends on the individual elements.

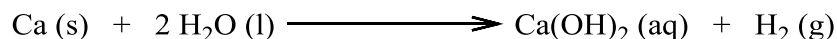
- i. Beryllium does not react with water or steam because of its covalent character.
- ii. Magnesium reacts very slowly with cold water producing bubbles of a colourless gas (hydrogen gas) and a cloudy solution of magnesium hydroxide which turns red litmus paper blue.



However, magnesium vigorously burns with steam when heated with a white bright flame to form a white solid called magnesium oxide.

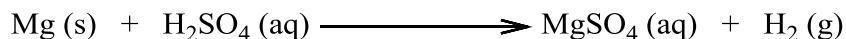


- iii. Calcium reacts quietly/moderately with cold water producing bubbles of a colourless gas (hydrogen gas) and a cloudy solution of calcium hydroxide. The solution appears cloudy because calcium hydroxide is sparingly soluble. The resultant solution is alkaline and therefore turns a red litmus paper blue.

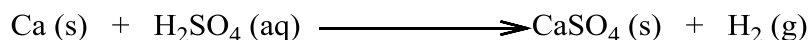


d) Reaction with dilute acids.

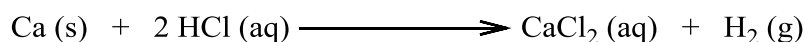
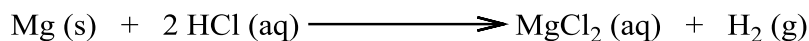
Magnesium reacts with dilute sulphuric acid to form magnesium sulphate and hydrogen gas.



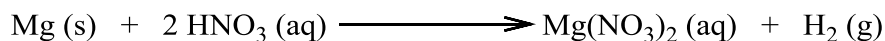
The reaction of calcium with dilute sulphuric acid is slower. This is because it forms calcium sulphate, a white insoluble salt which forms a coating on the calcium metal thereby preventing the reaction from proceeding in a short time.



With dilute hydrochloric acid, both magnesium and calcium reacts forming colourless solutions of metal chlorides and bubbles of a colourless gas (hydrogen gas). Calcium reacts more quickly in hydrochloric acid than magnesium.



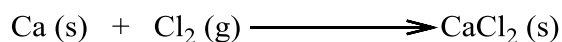
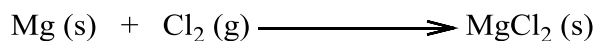
Only magnesium reacts with very dilute nitric acid to form magnesium nitrate and hydrogen gas.



NB: A Thin oxide layer covers beryllium metal preventing attack by acids. However when powdered, beryllium reacts with acids forming respective salts and hydrogen gas.

e) Reaction with chlorine

Both magnesium and calcium burn in chlorine gas on heating with a bright white flame and a bright red flame forming white solids of magnesium chloride and calcium chloride respectively. However the reaction of chlorine is slower in this case again because the metal surface quickly reacts with chlorine forming a coating which protects the inside of the metal from reacting.



Uses of alkaline earth metals

| Metal | Uses |
|------------------|---|
| Beryllium | <ul style="list-style-type: none">- Alloy with copper used in making some parts of computers.- Used in transmission of x-rays. |
| Magnesium | <ul style="list-style-type: none">- Magnesium is used in the manufacture of drugs e.g. milk of magnesia.- Magnesium powder is used in fireworks.- Used in alloys like with aluminium used to make aircrafts, car engine casing etc. |
| Calcium | <ul style="list-style-type: none">- Calcium is used in making animal feeds.- Calcium is used as a reducing agent in the extraction of thorium.- Calcium is used for making tooth paste because it strengthens the teeth. |

HALOGENS (Group VII Elements)

Group VII elements in the periodic table are called halogens meaning salt makers because they combine readily with metals to form salts. They usually exist as diatomic molecules and include; fluorine, chlorine, bromine and iodine. Halogens are the most reactive non-metals. This is due to their willingness to accept electrons. The willingness to accept the electrons decreases down the group.

Fluorine therefore, is the most reactive non-metal and this makes its properties and those of its compounds different from those of other elements in the group. Halogens have seven electrons in their outer most shell hence lack only one electron to form stable octets.

| Element | Electronic Configuration | Symbol | Formula | Appearance |
|----------|--------------------------|--------|-----------------|-----------------------------------|
| Flourine | 2: 7 | F | F ₂ | Greenish yellow gas |
| Chlorine | 2: 8: 7 | Cl | Cl ₂ | Greenish yellow gas |
| Bromine | | Br | Br ₂ | Red liquid (Reddish brown vapour) |
| Iodine | | I | I ₂ | Black solid (Violet vapour) |

Physical properties of halogens

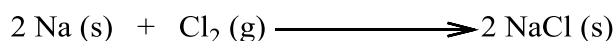
- i. Fluorine is a pale yellow/ greenish-yellow gas, chlorine is greenish-yellow gas, bromine is a reddish brown liquid and iodine is a black solid at room temperature.
- ii. They are all non-metals existing as diatomic molecules.
- iii. The melting point of the elements increases with increase in atomic number hence iodine has the highest melting point.
- iv. Chlorine and bromine are very poisonous. Therefore their preparation should be carried out in a fume cupboard and only in small quantities.
- v. When solid iodine is heated, it sublimes to form a violet vapour which condenses on cooling forms a black solid.

Chemical properties of halogens

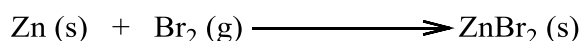
a) Reaction with metals

Halogens directly react with metals to forming halides (salts). The salts formed are ionic and the ionic nature decreases down the group.

Heated sodium continues burning with a bright yellow flame forming white solids of sodium chloride.

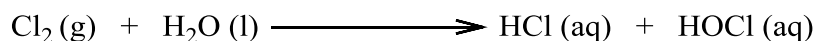


Heated zinc reacts with chlorine gas, bromine fumes or iodine vapour forming white solids of zinc (II) chloride, zinc (II) bromide and zinc (II) iodide respectively.

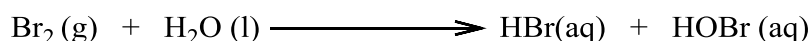


b) Reaction with water

Chlorine gas slightly dissolves in water forming a pale yellow solution containing a mixture of hydrochloric acid (HCl) and hypochlorous acid/ chloric (I) acid (HOCl). The acidic solution formed turns a blue litmus paper red and then bleaches it.



Bromine vapour slightly dissolves in water forming a brown solution containing a mixture of hydrobromic acid (HBr) and hypobromous acid/ bromic (I) acid (HOBr).



NB: Iodine is almost insoluble in water.

c) Reaction with alkalis

Halogens react with cold dilute or hot concentrated potassium hydroxide or sodium hydroxide solutions to form their respective salts and water.

Chlorine reacts with cold dilute sodium hydroxide solution to form **a colourless solution** containing a mixture of sodium chloride, sodium chlorate (I) (bleaching agent) and water.



Uses of halogens

| Halogen | Uses |
|-----------------|--|
| Fluorine | <ul style="list-style-type: none">- Manufacture of plastics.- Added to water and tooth paste to avoid teeth decay. |
| Chlorine | <ul style="list-style-type: none">- Chlorine is used in the treatment of water because it is poisonous- Manufacture of solvents like chloroform used in hospitals.- Manufacture of hydrochloric acid.- Manufacture of plastics like polyvinyl chloride. |
| Bromine | <ul style="list-style-type: none">- Making anti-knock compounds.- Manufacture of fumigants. |
| Iodine | <ul style="list-style-type: none">- Is a constituent of common salt used to prevent goitre.- Iodine solution can be applied on fresh wounds to kills bacteria and prevent infection. |

GROUP VIII or O (Noble/ Rare gases)

The elements in this group are called noble/ rare gases. The elements are chemically unreactive because their outer most shells are filled up with maximum number of electrons. Examples of noble gases include; helium, neon, argon, xenon and so on.

| | | |
|--------|----|---------|
| Helium | He | 2 |
| Argon | Ar | 2, 8, 8 |

Since the elements have stable octets or duplets, they can neither lose nor gain electrons. They are therefore chemically inert and do not take part in chemical reactions.

Uses of noble gases

| Noble gas | Uses |
|--------------------------|--|
| Helium | <ul style="list-style-type: none">- Filling meteorological balloons since it is light.- Provides necessary pressure for liquid fuel rockets.- Cooling medium in nuclear reactions. |
| Neon | <ul style="list-style-type: none">- Making advertisement lights.- a mixture of neon and helium is used in making gas lasers. |
| Argon | <ul style="list-style-type: none">- Argon is used to fill light bulbs and fluorescent tubes.- Used as an inert gas shield for arc welding and cutting. |
| Krypton and xenon | <ul style="list-style-type: none">- Used in filling light bulbs used in projectors. |