DEPARTMENT OF EDUCATION

GRADE 11

CHEMISTRY

MODULE 2

CHEMICAL AND METALLIC BONDING

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Writer
Honeylet Akiatan

Content Editors
Science Department
Subject Review Committee

Language Editor
Dr. Steven Winduo

Course Format Editor
Joydee Mabbagu

GRADE 11

CHEMISTRY

MODULE 2

CHEMICAL AND METALLIC BONDING

IN THIS MODULE YOU WILL LEARN ABOUT:

11.2.1: ATOMIC STRUCTURE

11.2.2: PERIODIC TABLE

11.2.3: BONDING AND STRUCTURE

Acknowledgement

We acknowledge the contributions of all secondary teachers who in one way or another have helped to develop this Course.

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DIANA TEIT AKIS
PRINCIPAL

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SECRETARY'S MESSAGE

Achieving a better future by individual students and their families, communities or the nation as a whole, depends on the kind of curriculum and the way it is delivered.

This course is a part of the new Flexible, Open and Distance Education curriculum. The learning outcomes are student-centred and allows for them to be demonstrated and assessed.

It maintains the rationale, goals, aims and principles of the national curriculum and identifies the knowledge, skills, attitudes and values that students should achieve.

This is a provision by Flexible, Open and Distance Education as an alternative pathway of formal education.

The course promotes Papua New Guinea values and beliefs which are found in our Constitution, Government Policies and Reports. It is developed in line with the National Education Plan (2005 -2014) and addresses an increase in the number of school leavers affected by the lack of access into secondary and higher educational institutions.

Flexible, Open and Distance Education curriculum is guided by the Department of Education's Mission which is fivefold:

- To facilitate and promote the integral development of every individual
- To develop and encourage an education system that satisfies the requirements of Papua New Guinea and its people
- To establish, preserve and improve standards of education throughout Papua New Guinea
- To make the benefits of such education available as widely as possible to all of the people
- To make the education accessible to the poor and physically, mentally and socially handicapped as well as to those who are educationally disadvantaged.

The college is enhanced through this course to provide alternative and comparable pathways for students and adults to complete their education through a one system, two pathways and same outcomes.

It is our vision that Papua New Guineans' harness all appropriate and affordable technologies to pursue this program.

I commend all the teachers, curriculum writers and instructional designers who have contributed towards the development of this course.

UKE KOMBRA, PhD
Secretary for Education

MODULE 2: CHEMICAL AND METALLIC BONDING

Introduction

We see different substances around us which are either elements or compounds. You also know that atoms of the same or different elements may combine. When atoms of the same elements combine, we get molecules of the elements. We get compounds; however, when atoms of different elements combine.
Have you ever thought why atoms combine at all? In this lesson, we will find an answer to this question. We will explain what a chemical bond is and then discuss various types of atoms of different elements and variation in the periodic properties of elements. Configuration of atoms, the electronegativity difference between bonded atoms. Such as, ionic compounds result from a transfer of electrons from one atom to another which produces ions, oppositely charged atoms, that attract. Covalent compounds share a pair of opposite spin electrons; equal sharing gives a non-polar bond and unequal sharing a polar bond. Exploration of covalent bonding begins with the empirical rules of Lewis that allow for the construction of simple two-dimensional models of covalent compounds.

Learning outcomes

After going through the Module, you are expected to:

- demonstrate an understanding of the patterns of arrangement of electrons in atoms and describe the electron arrangement in terms of shells and sub shells (s, p, d and f).
- explain trends and relationships between elements in terms of atomic structure and bonding.
- explain that the position of an element in the periodic table reflects its electron configuration.
- draw diagrams (shells or Lewis dot) to show how ions can be formed by atoms gaining or losing electrons.
- draw electronic shell diagrams to show the formation of ionic and covalent compounds.
- differentiate between ionic, covalent, and metallic bonds.
- explain the formation of coordinate (dative) bonds.
- discuss the allotropes of carbon (graphite and diamond).
- describe the concept of metallic bonding, using diagrams.
- explain the polarity and formation of hydrogen bond.

Time Frame

10 weeks

This module should be completed within 10 weeks.

If you set an average of 3 hours per day, you should be able to complete the module comfortably by the end of the assigned week.

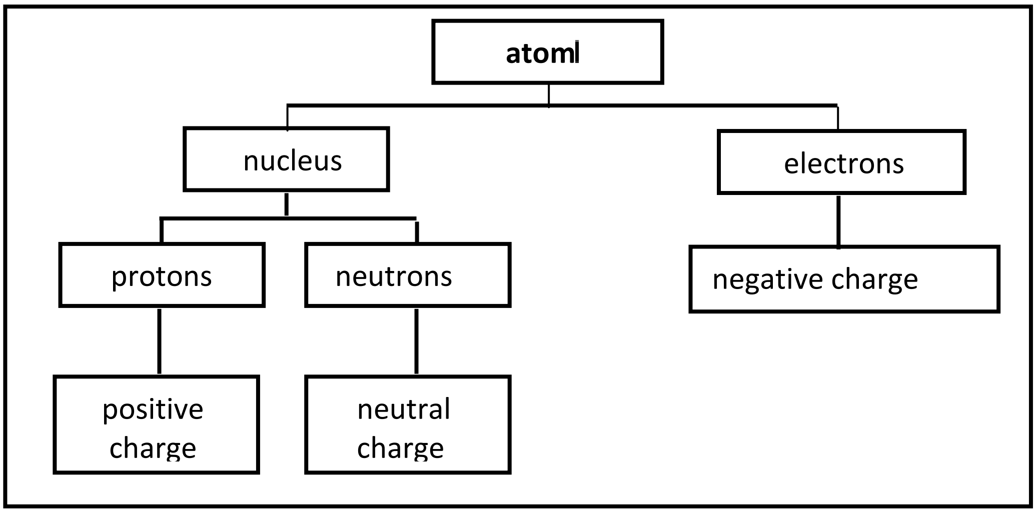
Try to do all the learning activities and compare your answers with the ones provided at the end of the module. If you do not get a particular exercise right in the first attempt, you should not get discouraged but instead, go back and attempt it again. If you still do not get it right after several attempts then you should seek help from your friend or even your tutor.

DO NOT LEAVE ANY QUESTION UN-ANSWERED.

Terminologies

Allotrope	Is the same element with different physical properties.
Atom	Is the smallest particle of an element. Atoms of different elements may also combine into systems called **molecules**, which are the smallest modules of chemical compounds. The total number of **protons (positively charge)** in a given atom determines the **atomic number** of an element.
Atomic mass or mass number	Is the sum of an atom's protons and neutrons that are always expressed in whole numbers.
Atomic number	Is the number of protons (also number of electrons) that the atom contains.
Electron configuration	The distribution of electrons among the orbitals of an atom.
Isotopes	Are an atom of the same element with same number of protons but different number of neutrons.
Mass number	Is the number of protons plus number of neutrons.
Number of neutron	Is the mass number minus atomic number.
Periodic table	Is a chart of all the known elements in order of increasing atomic number.
Periodic trend	Is a property that changes as you move across a period or down a group of the periodic table.
Relative isotopic mass	Is the mass of an atom of the isotope relative to the mass of an atom.

11.2.1 Atomic Structure

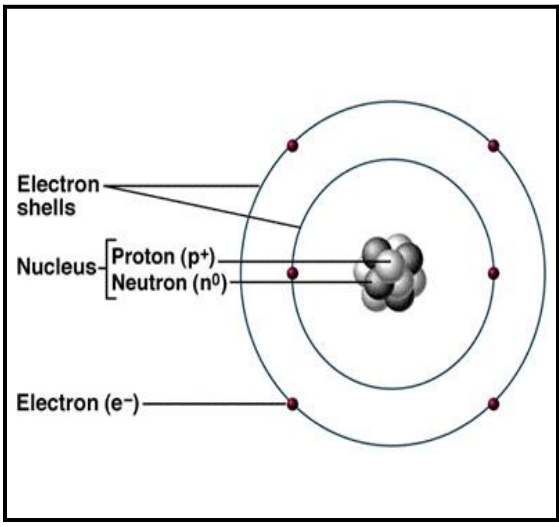


Inside the atom

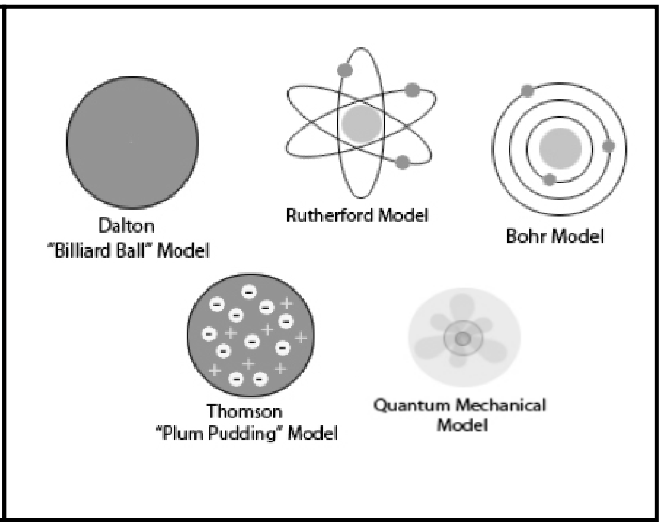
This topic explains the structure of matter. Matter is composed of tiny particles called atom. **An atom** is the smallest particle of an element. Atoms of different elements may also combine into systems called **molecules**, which are the smallest modules of chemical compounds. These are also considered as the ultimate building blocks of matter.

What are atoms made of? What makes one type of atom different from another? A careful study of the atom shows that it has a small, but dense core called the **nucleus.**

The nucleus is composed of **protons** the positively charged particle and the **neutrons**, the particle with no charge. Around the nucleus is the **electron**, the negatively charged particle.



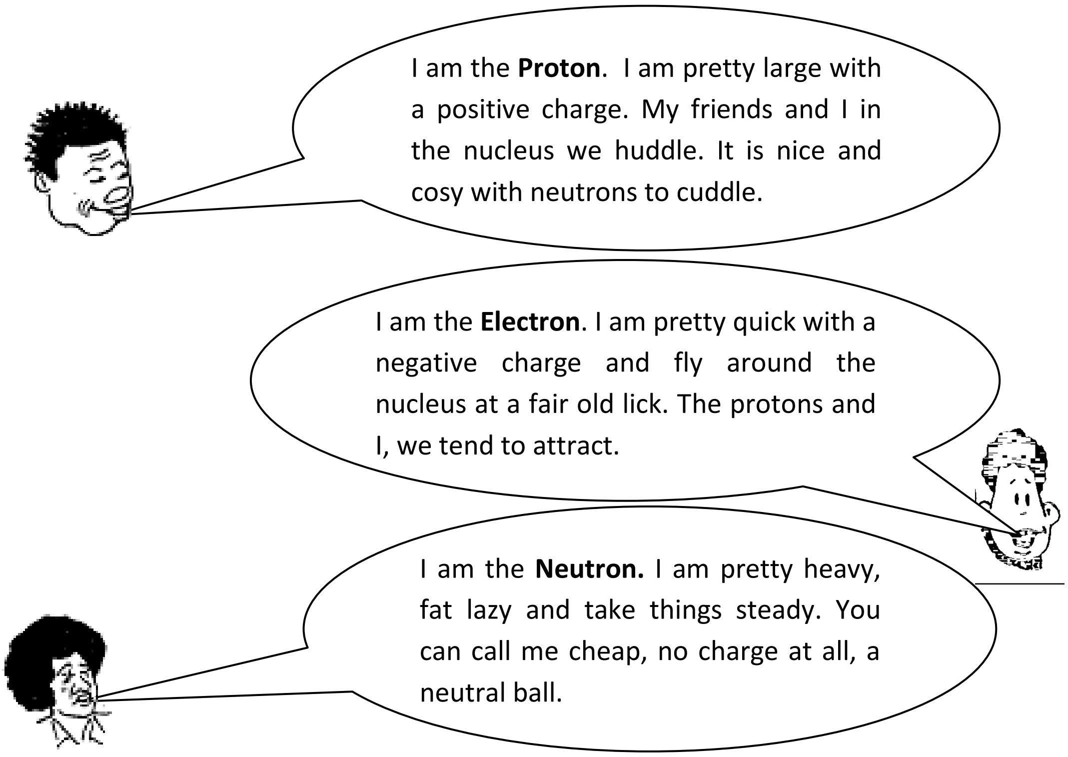
The atomic structure



Different Atomic

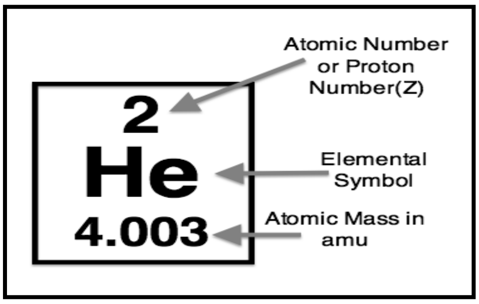
The table shows the charge and mass of the particles in an atom.

Particle	Charge	Mass	Discoverer
Proton	1+	1	Eugene Goldstein
Neutron	0	1	James Chadwick
Electron	1-	0 (almost)	J.J Thomson
Nucleus			Ernest Rutherford



The atom is the element helium (He). Its position in the periodic table is second, and it has the atomic number 2.

Each element on the periodic table has a different atomic number. The atomic number represents the number of protons an of a particular elements contains in its nucleus. The number of protons equals the number of electrons.

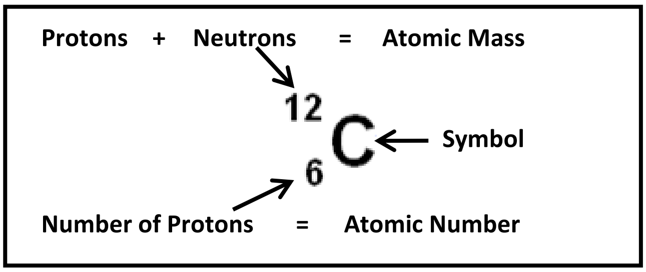


Position of atomic mass and atomic number in Periodic Table

The total number of **protons** in a given atom determines the **atomic number** of an element. The atomic number is the number of protons (positively charged elementary particles) in the nucleus of one of its atoms. If the atom is electrically neutral, the same number of electrons is present, since the number of protons is equal to the number of electrons.

The **atomic mass or mass number** is the sum of an atom's protons and neutrons that are always expressed in whole numbers.

For example, $^{12}_{6}$**C** indicates a carbon atom of atomic mass 12 and atomic number 6, the difference being equal to the number of neutrons in the nucleus. This means for $^{12}_{6}$**C,** there are 6 protons and 6 neutrons. It follows that it has 6 electrons, too.



(In other presentations of The Modern Periodic Table, the superscript is the atomic number and the subscript is the atomic mass. The atomic number is always less than the atomic mass).

Number of neutrons + Number of protons = Mass number

Atomic number	**=**	**number of protons (also number of electrons)**
Mass number	**=**	**number of protons plus number of neutrons**
Number of neutrons	**=**	**mass number minus atomic number**

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 1

30 minutes

Answer the following questions:

1. What are the sub- particles found inside the atom?
   - i. ___________
   - ii. ___________
   - iii. ___________

2. Complete the table (by showing their masses and charge.)

Sub – atomic particles	charge	Mass (atomic mass module)
Proton		
	0	1
		0 (almost)

3. Give the number of protons, electrons and neutrons in the atom below:

a) N
number of protons ___________
number of electrons ___________
number of neutrons ___________

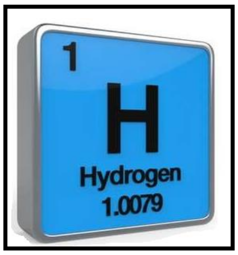
b) K
number of protons ___________
number of electrons ___________
number of neutrons ___________

c) Ne
number of protons ___________
number of electrons ___________
number of neutrons ___________

Thank you for completing your learning activity 1. Check your work. Answers are at the end of this module.

Relative Atomic Mass

Each type of atoms has a different mass. However, the mass of an atom is too small to measure. So we compare their masses on a scale which gives the lightest element of all, hydrogen, a mass of 1. This gives us the number called the **relative atomic mass.** (R.A.M). Sometimes given the symbol $\mathbf{A_r}$**.**



Let us look, more closely at the hydrogen atom.

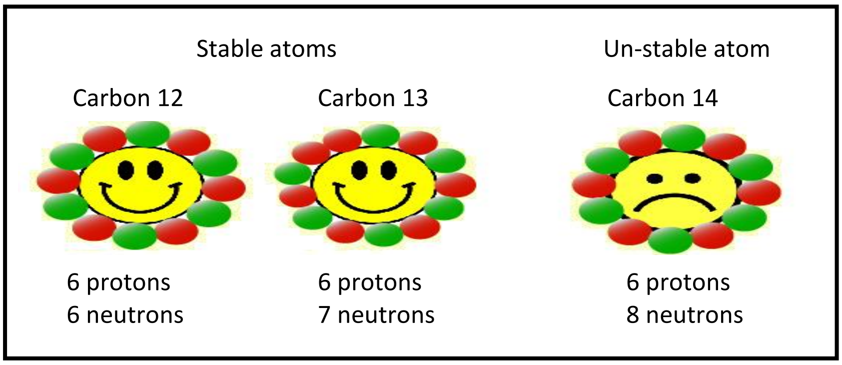
Hydrogen's atomic number is 1. Its mass number is also 1. This means that hydrogen atom has 1 proton, 1 electron, but no neutrons.

You can see why it is so light!

You might think that there is no point having relative atomic mass and mass number. Both seem to tell us how heavy atom is.

What are isotopes?

Isotopes are an atom of the same element with same number of protons but different number of neutrons.



Why carbon-14 is unstable?

If an atom has too many neutrons in its nucleus, it is unstable and will change into a stable form. To date a sample, scientists calculate how much time would be required for the unstable atoms in the sample to change into a stable form.

For example, most carbon atoms are stable because they have only six or seven neutrons in their nuclei (carbon-12 and carbon-13, above). But some carbon atoms have too many neutrons and are unstable (carbon-14)

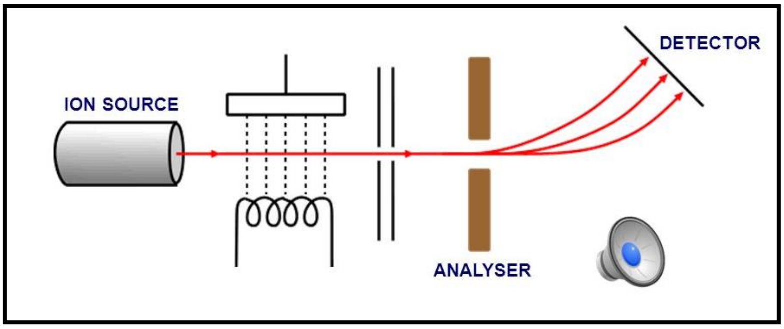
Relative isotopic mass is the mass of an atom of the isotope relative to the mass of an atom.

Relative Isotopic Mass =mass of one atom of that isotope relative to the mass of one atom of ^{12}C = 12 units exactly.

Most elements, like chlorine, are a mixture of isotopes. Details of isotopes of some common elements are shown in the table below.

Element	Isotopes	Relative isotopic mass	Abundance (%)
Hydrogen	^{1}H	1.008	99.986
	^{2}H	2.014	0.014
	^{3}H	3.016	0.0001
Carbon	^{12}C	12	98.888
	^{13}C	13.003	1.112
	^{14}C	14.003	Approx.10-^{10}
Chlorine	^{35}Cl	34.969	75.0
	^{37}Cl	36.966	25.

Isotopic compositions of some elements



Relative atomic masses of elements can be obtained using an instrument called a **mass spectrometer.** This separates the individual isotopes in a sample of element being measured and determines the mass of each isotope relative and the relative abundance of the isotopes. This information presented graphically is called **mass spectrum**.

A mass spectrometer consists of an ion source, an analyser and a detector. Particles must be ionized so they can be accelerated and deflected.

Mass spectrometry has been described as the smallest scale in the world, not because of the mass spectrometre's size but because of the size of what it weighs the molecules.

For example: Chlorine has two isotopes ^{37}Cl with the the abundance of 25% and ^{35} Cl with 75% abundance. Using this information provided by a mass spectrum we can now calculate the relative atomic mass of chlorine.

The total mass of 100 atoms of chlorine:

$$= \frac{\textbf{isotope 1 x abundance (\%) + isotope 2 x abundance (\%)}}{\textbf{100}}$$

$$= \frac{\textbf{35 x 75\% + 37 x 25\%}}{\textbf{100}} \textbf{ = 35.5}$$

Relative atomic mass of chlorine is 35.5

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 2

30 minutes

Answer the following questions:

1. Define the following:
   i. Isotope ____
   ii. Mass spectrometer ____

2. Using the different isotopes and abundance, calculate the relative atomic mass of the following:

   i. ^{63}Cu with 69%, ^{65}Cu with 31%.

   ii. ^{16}O with 99.79%, ^{17}O with 0.04%, ^{18}O with 0.20%.

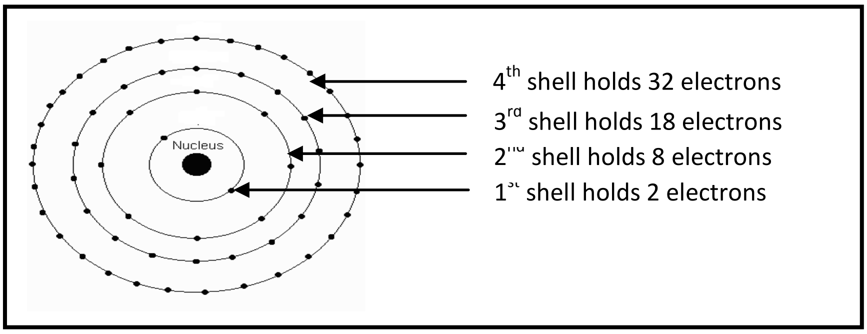
Thank you for completing your learning activity 2. Check your work. Answers are at the end of this module.

Electrons in Orbit

Electron shells

The electrons in the atom are arranged in shells around the nucleus.

- The shells are sometimes called orbital's or energy levels.
- The **first shell** can hold just **2 electrons.**
- The **second shells** can hold up **to 8 electrons.**
- The **third shell** can also hold **8 electrons.**



Electron Shells

1. The distribution of electrons among the orbitals of an atom is called **electron configuration.**
2. The **s, p, d, and f** are called **sublevels**; they are smaller "subdivisions" of energy within the primary levels.
3. You refer to different energy levels using a number for the primary level plus a letter for the sublevel; for example, you might speak of an electron in a "3p" **state** or **orbital**.

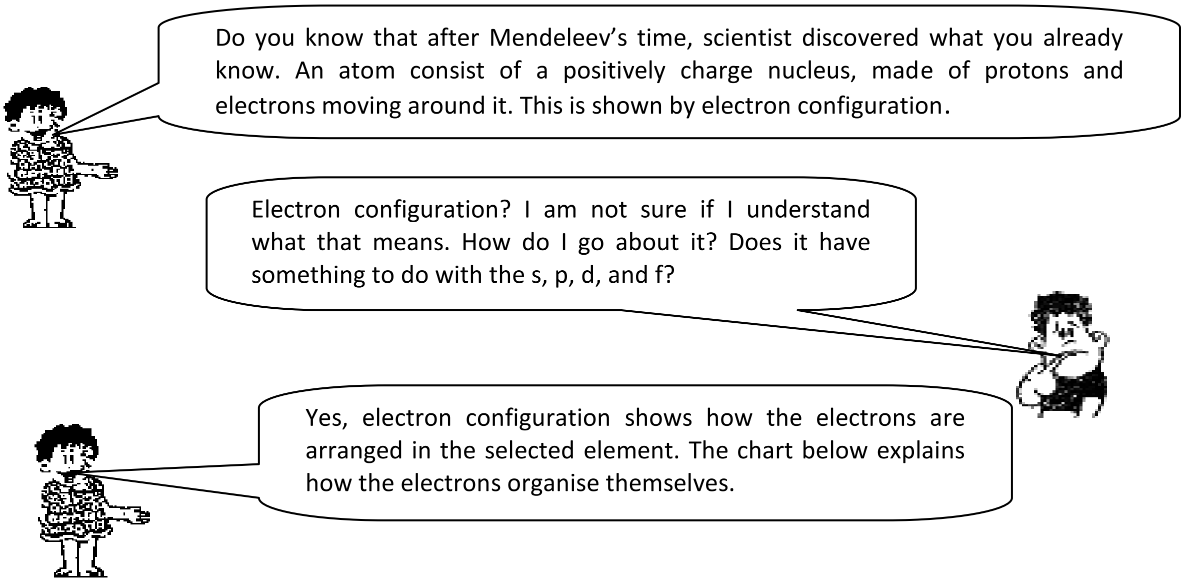
   Each primary level has one more sublevel than the one below: the first primary level has only **s** orbitals; the second has **s** and **p**, the third **s, p**, and d.

   s – sharp **2** electrons to fill
   p – principal **6** electrons to fill
   d – diffuse **10** electrons to fill
   f – fundamental **14** electrons to fill

Total number of electrons in an energy level can be found by the formula:

$$\boxed{\textbf{number of electrons} = 2n^2}$$

Where **n** is the number of energy level
1st = 2
2nd = 8 = (2+6)
3rd = 18 = (2 +2+ 6 = 10)
4th = 32 = (2 + 6 + 10 + 14)



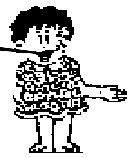
$1s^1$																	$1s^2$
$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$3d^5$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^{10}$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^4$	$4d^5$	$4d^5$	$4d^7$	$4d^8$	$4d^{10}$	$4d^{10}$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
$6s^1$	$6s^2$		$5d^2$	$5d^3$	$5d^4$	$5d^5$	$5d^6$	$5d^7$	$5d^9$	$5d^{10}$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
$7s^1$	$7s^2$		$6d^2$	$6d^3$	$6d^4$	$6d^5$	$6d^6$	$6d^7$	$6d^8$	$6d^{10}$	$6d^{10}$	$7p^1$	$7p^2$	$7p^3$	$7p^4$	$7p^5$	$7p^6$
			$5d^1$	$4f^1$	$4f^3$	$4f^4$	$4f^5$	$4f^6$	$4f^7$	$4f^7$	$4f^9$	$4f^{10}$	$4f^{11}$	$4f^{12}$	$4f^{13}$	$4f^{14}$	$4f^{14}$
			$6d^1$	$6d^2$	$5f^2$	$5f^3$	$5f^4$	$5f^6$	$5f^7$	$5f^7$	$5f^9$	$5f^{10}$	$5f^{11}$	$5f^{12}$	$5f^{13}$	$5f^{14}$	$5f^{14}$

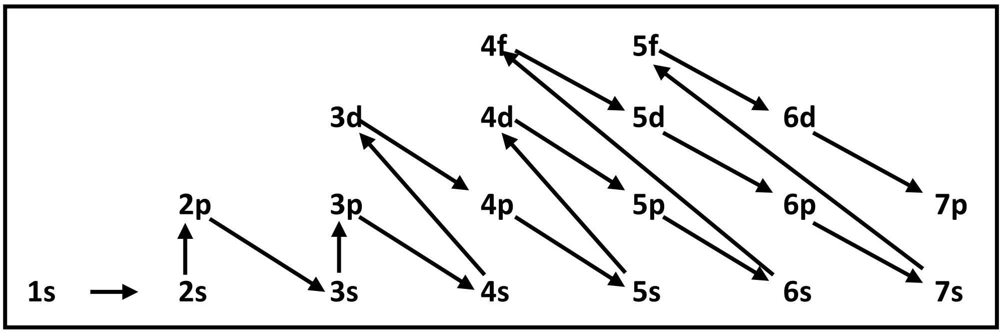
Chart of organised electrons

It say that groups I and II belong to **s** block, group III to VIII belong to the **p** block, transition metals to the **d** block and the lanthanides and actinides to the **f** block.

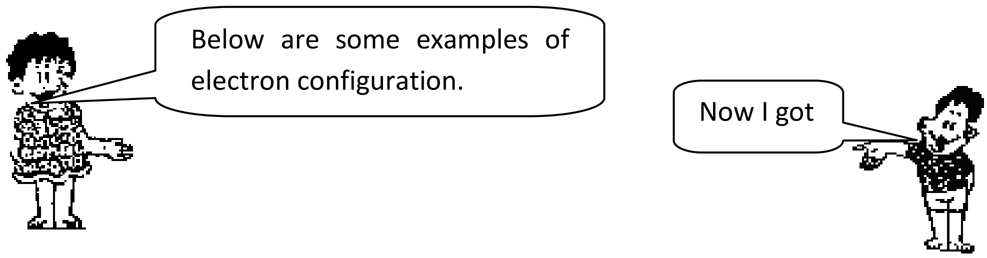
But, how will I do it? Can you give me a pattern to make it easier and faster?

Very easy! Just follow the arrows.





Electronic configuration



Elements	Symbol	Atomic Number	Electron Configuration	Period	Group
1.Hydrogen	H	1	$1s^1$	1	IA
2.Lithium	Li	3	$1s^22s^1$	2	IA
3.Berylium	Be	4	$1s^22s^2$	2	IIA
4.Carbon	C	6	$1s^22s^22p^2$	2	IVA
5.Nitrogen	N	7	$1s^22s^22p^3$	2	VA
6.Aluminum	Al	13	$1s^22s^22p^63s^23p^1$	3	IIIA
7.Sulphur	S	16	$1s^22s^22p^63s^23p^4$	3	VIA
8.Potassium	K	19	$1s^22s^22p^63s^23p^64s^1$	4	IA
9.Bromine	Br	35	$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$	4	VIIA

Analysis of the Table
Take note of the coefficient of the last energy level. Notice also the exponent or the number on the superscript of the letter of the last energy level. The coefficient represents the series or the period of the element and the superscript or the exponent represents the number of electrons in the outermost energy level or simply the valence electrons. It is also the Family or the Group Number of the element.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 3

30 minutes

Answer the following questions:

1. Define electron configuration.
____
____

2. Draw the electron configuration of the following using the **s, p, d and f.**

a) Na ________________________________
b) Ca ________________________________
c) Cl ________________________________
d) Zn ________________________________

e) F ____

f) Al ____

Thank you for completing your learning activity 3. Check your work. Answers are at the end of this module.

12.2.2 Periodic Table

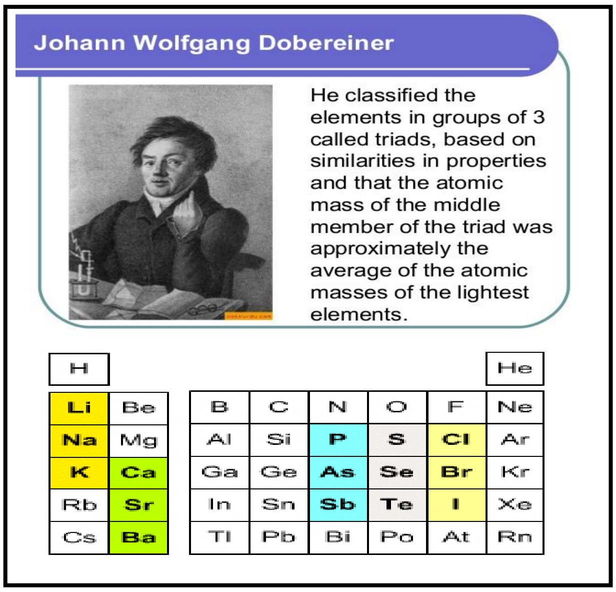
The Development of Periodic Table

- 400 BC Democritus suggest that all things are made of particles.
- 1805 John Dalton's atomic theory. Atoms of the same element are alike. They combine to make compounds.
- 1909 Ernest Rutherford discover the proton.
- 1911 Ernest Rutherford discover the nucleus.
- 1913 Neil's Bohr suggests those electrons are found in shells around the nucleus.
- 1932 James Chadwick proves that neutrons exist.

The early years of the 19th century witnessed a rapid development in chemistry. The art of distinguishing similarities and differences among atoms prompted scientists to devise a way of arranging the elements. Relationships were discerned more readily among the compounds than among the elements. Thus, the classification of elements lagged many years behind the classification of compounds.

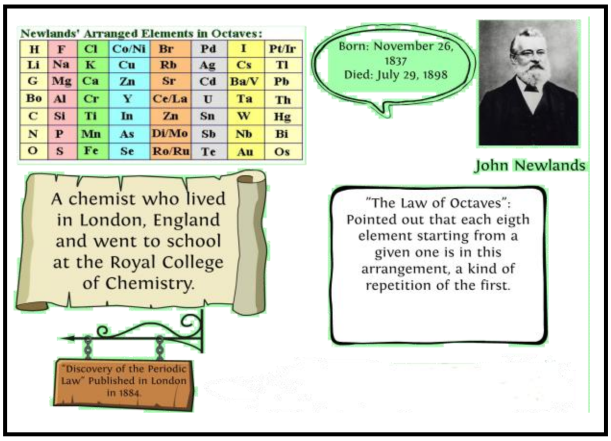
In fact, no general agreement had been reached among chemists as to the classification of elements for nearly half a century after the systems of classification of compounds had been established.

It was in 1817 when **Johann Wolfgang Döbereiner** showed that the atomic weight of strontium lies midway between those of calcium and barium. Some years later he showed that other such **"triads"** exist (chlorine, bromine, and iodine [halogens] and lithium, sodium, and potassium [alkali metals]).



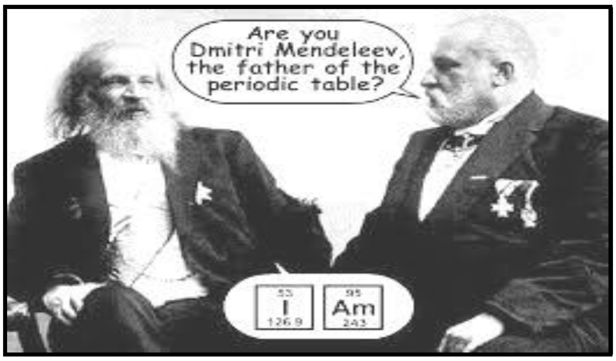
He also showed that similar relationships extended further than the triads of elements, fluorine being added to the halogens, and magnesium to alkaline-earth metals. Oxygen, sulfur, selenium, and tellurium were classed as one family, and nitrogen, phosphorus, arsenic, antimony, and bismuth as another family of elements.

Another way of classifying the elements was later proposed by **John Alexander Reina Newlands** in 1864. He proposed that elements be classified in the order of increasing atomic weights, the elements being assigned ordinal numbers from one upward and divided into seven groups, with each group having properties closely related to the first seven of the elements then known: hydrogen, lithium, beryllium, boron, carbon, nitrogen, and oxygen.



Mendeleev's version of the periodic table

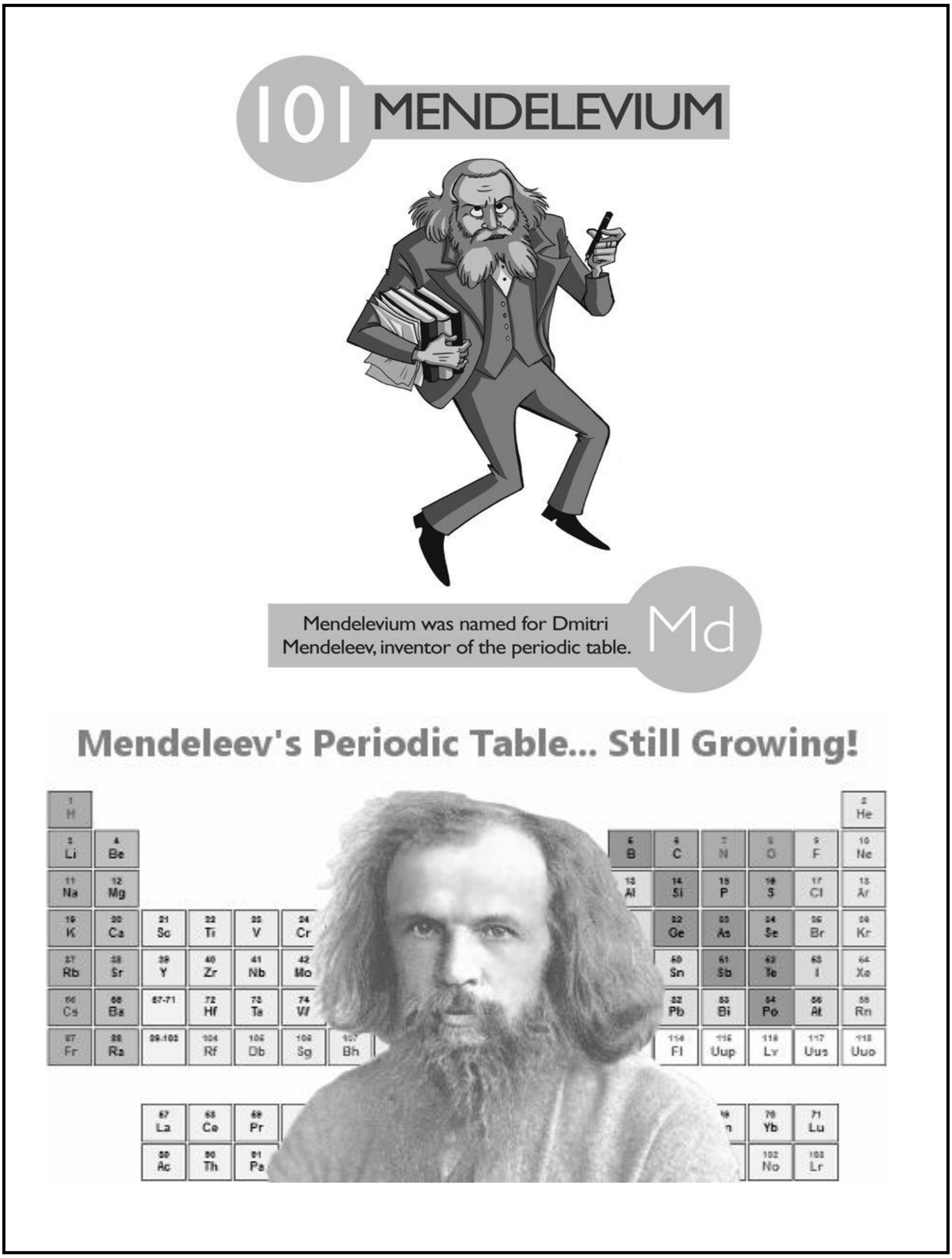
The rows **1** to **7** are called **periods**. The columns **I** on the left to **0** on the right are known as **groups**. Elements with similar properties fall into vertical columns (groups) and horizontal rows (periods), which form the table. Elements within the groups have similar valences. Mendeleev left spaces in his table for elements not yet discovered.

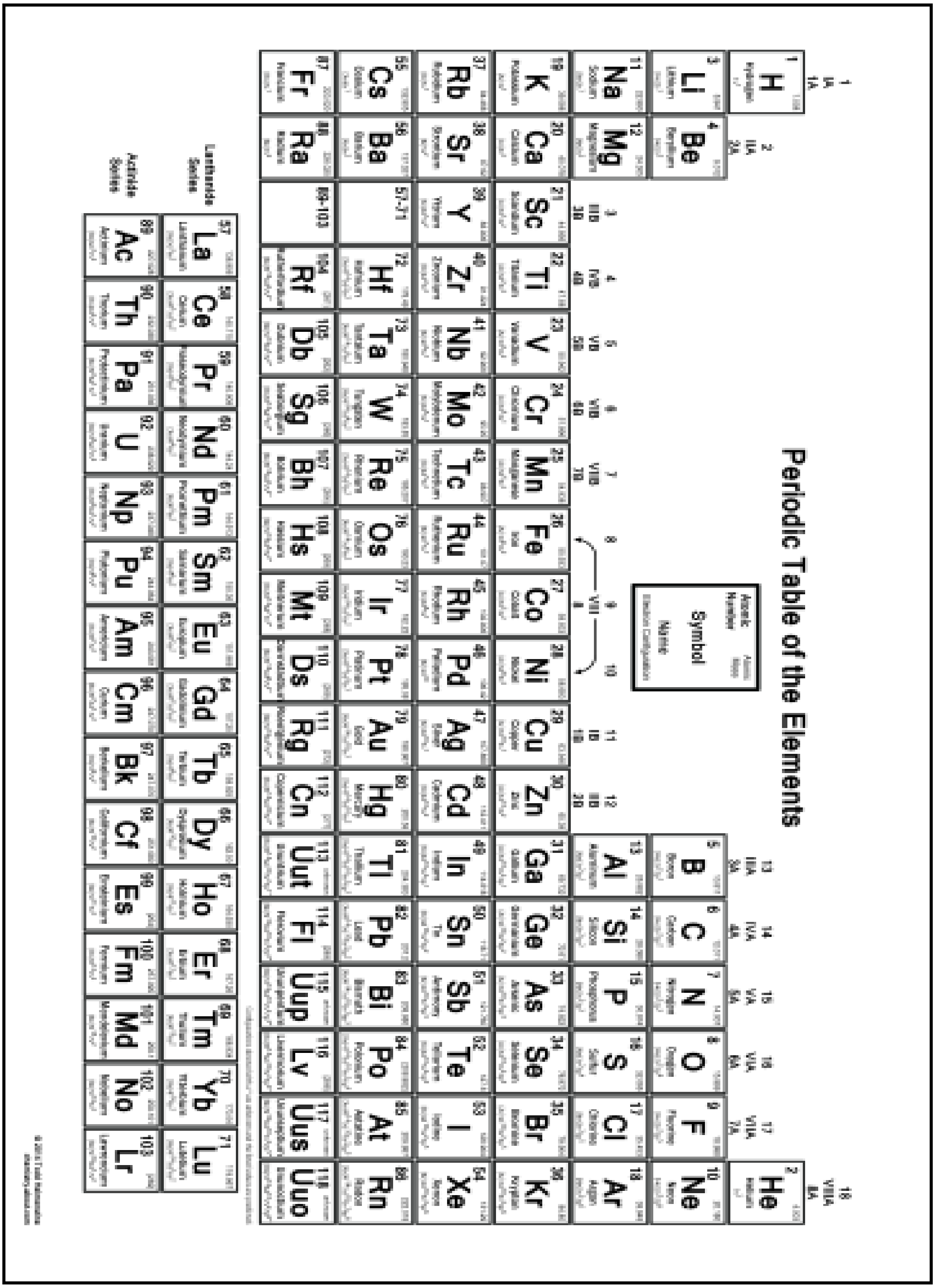


He also predicted what properties these undiscovered elements would have. Between 1875 and 1886, the elements gallium, scandium and germanium were discovered. They all fitted into the positions predicted by Mendeleev. As a result, the periodic law gained universal acceptance. The elements in the same row have something in common.

All of the elements in a **period** have the same number of electron **shells**. Both elements in the top row (the first period) have one shell for their electrons. All elements in the second period have two electron shells. The number of shells increases as you go down the table.The columns in the table is called Groups.

The elements in a group have the same number of electrons in their outer shell. So, all elements in Group I have one electron in their outer shells. The elements in Group II have two electrons in the outer shell.


101 MENDELEVIUM
Mendelevium was named for Dmitri Mendeleev, inventor of the periodic table.
Md
Mendeleev's Periodic Table... Still Growing!


Periodic Table of the Elements
Lanthanide Series
Actinide Series

Study the terms below for you to be familiar with the modern periodic table:

- **Periodic table** is a chart of all the known elements in order of increasing atomic number. The table puts elements into groups with similar characteristics, allowing us to recognize trends over the whole array of elements.
- **Atom** is the smallest module of a substance that still has all the properties of that substance. In most cases, an atom consists of protons, neutrons, and electrons. The protons and neutrons are found in the center of the atom, called the atomic nucleus and the electrons orbit or circle around the center of the nucleus in paths called orbitals.
- **Atomic number** of an atom is equal to the number of protons that the atom contains. Atoms can have differing numbers of neutrons and electrons while still retaining the original characteristic properties of that atom. However, if an atom gains or loses a proton, in essence, it changes its atomic number and becomes an entirely new atom with new characteristics.
- **Atomic mass** of an atom is a measure of how much mass an atom has. The atomic mass is calculated by adding the number of protons and neutrons together. Atomic masses are not listed as whole numbers on the periodic table because atoms can come in forms with different amounts of neutrons.

The Elements in the Periodic Table

How are elements named?

Chemists have developed a unique system of symbols and notation designed to simplify the writing of chemical symbols, formula, and reactions. This system also shows the mathematical relations of atoms and reacting chemicals, the way atoms are put together to form complex molecules. Elements, in most cases he was able to use the first letter of the name of the element as its symbol: **O** stood for **oxygen**, **C** for **carbon**, **H** for **hydrogen**, and so on. Two letters are used to distinguish between elements that have the same initial letter: **N** for **nitrogen**, **Ne** for **neon**, and **Ni** for **nickel**. Sometimes the symbol is derived from the Latin name of the element, example **gold** (aurum) is **Au, iron** (ferrum) is **Fe,** and **lead** (plumbum) is **Pb**. Whenever two letters are used for an element, the first letter is capitalized but the second is not. Thus, the element **cobalt, Co**, is distinguished from the compound **carbon monoxide, CO**

Some Elements	**Familiar Place or Name**	**Symbol of Element**
Californium	California	Cf
Einsteinium	Albert Einstein	Es
Nobelium	Alfred Nobel (Nobel Prize)	No
Neptunium	Neptune	Ne
Plutonium	Pluto	Pu
Americium	America	Am
Berkelium	Berkeley, California	Bk

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 4 **30 minutes**

Answer the following questions:

1. Define the following terms:

   i. Atom ____

   ii. Atomic mass ____

   iii. Atomic number ____

2. In what year the neutron was discovered? ______________________________
3. Who discovered the electron? ______________________________

Thank you for completing your learning activity 4. Check your work. Answers are at the end of this module.

Periodic Table Variations

The Periodic Table and the Elements

The periodic table is composed of 7 (rows) or periods and 8 major groups or (columns). Elements in a group have similar properties particularly those elements in groups: Group I (The Alkali metals) Group II (the Alkaline earth metals) Group VII the (Halogens) and Group VIII the (Noble gases).

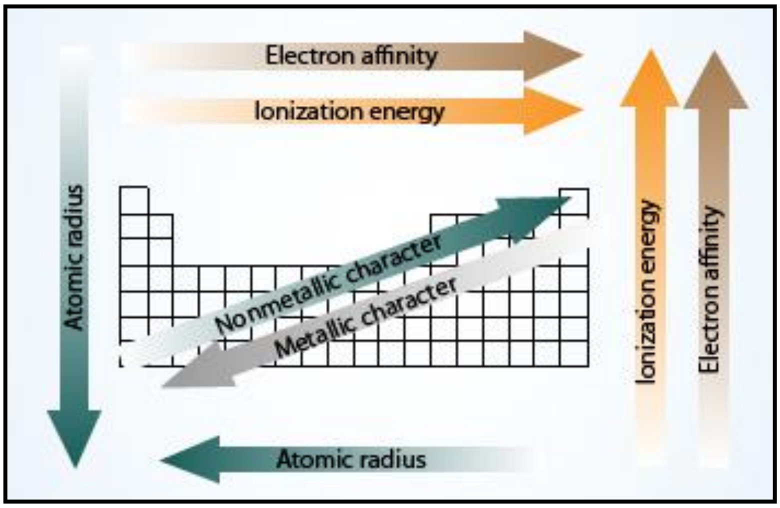
1. The elements are given symbols devised by John Jacob Berzelius. An element is named after its discoverer, place of discovery, first letter of the name of the element, first and the second letter for those having the same first letter and some are after their Latin names. The elements are grouped into Group A and B Group by the INTERNATIONAL UNION OF PURE AND APPLIED CHEMISTRY (**IUPAC**)

2. Elements in the Periodic Table are also grouped according to metals, non-metals and metalloids. **Metals** are lustrous, malleable and ductile. They are good conductors of heat. Metals are found on the left side of the periodic table. **Non-metals** have a many set of properties. They are found on the upper right side of the periodic table. **Metalloids** or semimetals possess the properties of both the metals and the non-metals.

Periodic Trends
A periodic trend is a property that changes as you move across a period or down a group of the periodic table.

The table below shows the example of some elements.

Metals	Symbol	Non- metals	Symbol	Metalloids	Symbol
Sodium	Na	Carbon	C	Silicon	Si
Copper	Cu	Oxygen	O	Arsenic	As
Calcium	Ca	Fluorine	F	Antimony	Sb
Magnesium	Mg	Chlorine	Cl	Germanium	Ge



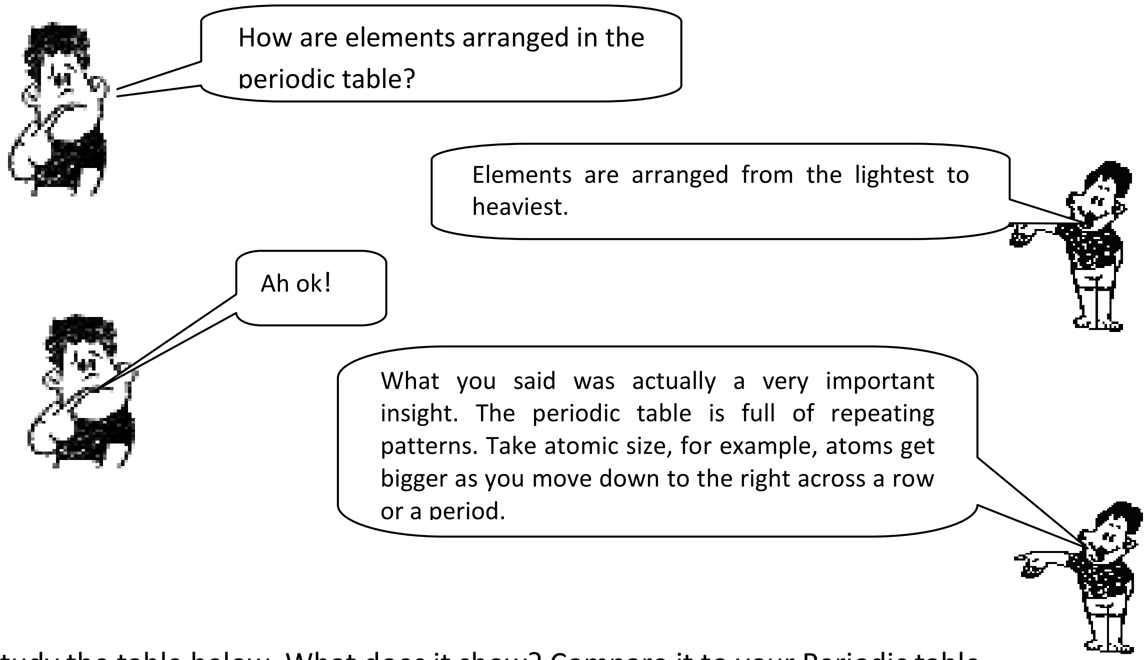
The Periodic Trends

The following are the trends of elements in the periodic table.
1. Atomic radius increases as you move down a group and decreases as you move across.
2. Ionization energy is the energy needed to remove an electron to form a positive ion.
3. Electron affinity is the energy change that occurs when an atom gains an electron to form a negative ion.
4. An element's electronegativity reflects its attraction for electrons in a chemical bond.
5. The **s, p, d, and f** are called **sublevels**; they are smaller "subdivisions" of energy within the primary levels. You refer to different energy levels using a number for the primary level plus a letter for the sublevel; for example, you might speak of an electron in a "3p" **state** or **orbital**. Each primary level has one more sublevel than the one below, the first primary level has only s. Orbitals; the second has s and p, the third s, p, and d, and so forth.

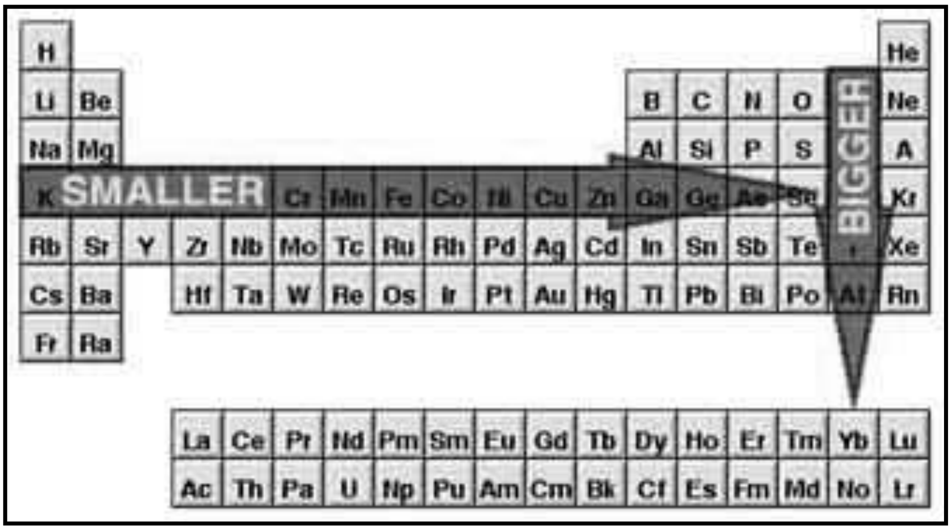
Pattern and trends in the Periodic Table

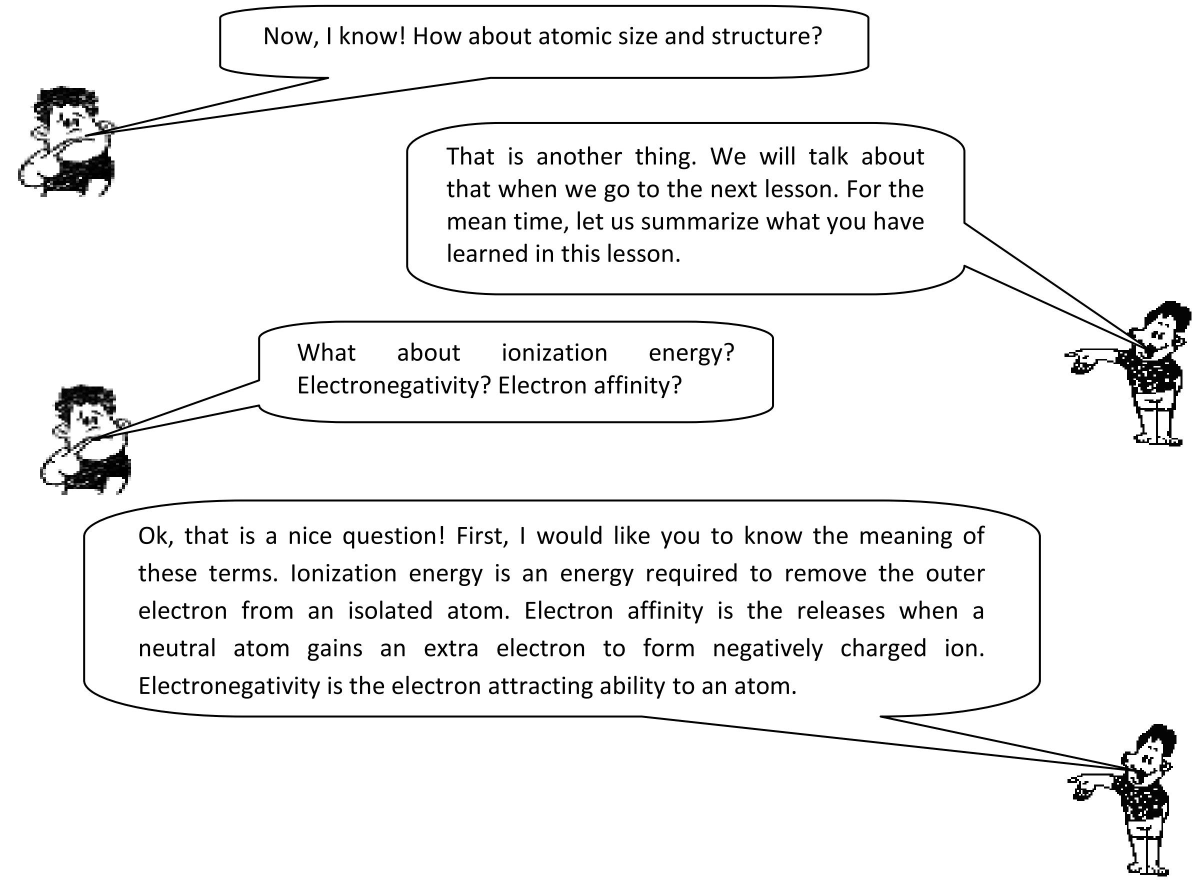
This lesson will focus on the arrangement of the elements in the periodic table based on their properties.

Read the comic strips below. You will need your own periodic table for this activity.



Study the table below. What does it show? Compare it to your Periodic table.





For horizontal arrangement of elements, there are two distinct trends in the atomic radius of the elements in the periodic table.

1. Atoms get larger going down a group (vertical arrangement or column).
2. Atoms get smaller moving from left to right across each period.

Example of electronegativity values of some of the lighter elements.

Group I	**II**	**III**	**IV**	**V**	**VI**	**VII**
Li	Be	B	C	N	O	F
1.0	**1.6**	**2.0**	**2.5**	**3.0**	**3.5**	**4.0**
Na	Mg	Al	Si	P	S	Cl
0.9	**1.3**	**1.6**	**1.9**	**2.2**	**2.6**	**3.2**

The atomic radius is affected by the ionization energy, electronegativity and electron affinity. As you go across from left to right, ionization energy, electro negativity and electron affinity increase, thus the atomic radius decreases. From top to bottom, ionization energy, electronegativity and electron affinity decreases, thus the atomic radius increases.

The table below shows the example of position of electron configuration.

Elements	Electron configuration	Group number	Period number
Sodium, Na	11- 2,8,1	I	3
Iron, Fe	13 – 2,8,3	III	2
Oxygen, O	8 - 2,6	VI	1
Sulphur,S	16 – 2,8,6	IV	3
Silver, Ag	47 – 2,8,18,18,1	I	5
Iodine, I	53- 2,8,18,18,7	VII	5
Radon, Rn	86 – 2,8,18,32,18,8	VIII	6

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 5

30 minutes

Answer the following questions:

1. Define the following terms:

   i. Metalloid ______________________________
   ii. Ionization energy ______________________________
   III. Electron affinity ______________________________
   iv. Electronegativity ______________________________

2. Give the two distinct trends in the atomic radius of the elements in the periodic table.
   I. ______________________________
   ii. ______________________________

3. How are elements arranged in the periodic table?
   ____
   ____

Thank you for completing your learning activity 5. Check your work. Answers are at the end of this module.

The group

These elements are organized using the **Periodic Table**. A periodic table is a classification and tabulation of the elements in the order of their atomic numbers and atomic masses that show the elements' chemical and physical properties.

How are elements grouped?

Take note of the colour-coding in the periodic table. Elements are grouped according to metals, nonmetals, and metalloids. **Metals** are solid, malleable, ductile and good conductors of heat. They also possess luster. The only liquid metal is mercury; (**Hg**).**Nonmetals** can be solids, liquids or gases. The only liquid nonmetal is bromine, (**Br**). In-between metals and nonmetals that lie along either side of the zigzag line of the periodic table are the **metalloids**. Some of these elements like boron (**B**) and silicon (**Si**) are used as semiconductors.

How are elements classified?

Elements are classified based on their positions or locations in the periodic table.

Group I **The Alkali Metals**

Group 1 elements are soft silvery metals. They react strongly with water. The further down the group you go, the more violent this reaction is. These alkali metals are usually stored under oil to protect them from moisture and oxygen. They all have one electron in their outer shells. In a chemical reaction an alkali metal atom loses this single electron. It achieves the stable electron structure of the noble gases.

Group II **The Alkaline Earth Metals**

This group consists of all metals that occur naturally in compound form. They are obtained from mineral ores and form alkaline solutions. These are less reactive than alkali metals.

Group III **The Aluminum Group**

The elements in this group are fairly reactive. The group is composed of four metals and one metalloid which is boron.

Group IV **The Carbon Group**

This group is composed of elements having varied properties because their metallic property increases from top to bottom meaning the top line, which is carbon, is a nonmetal while silicon and germanium are metalloids, and tin and lead are metals.

Group V **The Nitrogen Group**

Like the elements in group IV A, this group also consists of metals, nonmetal and metalloids.

Group VI The Oxygen Group

This group is called the oxygen group since oxygen is the top line element. It is composed of three nonmetals, namely: oxygen, sulfur and selenium, one metalloid,(tellurium) and one metal (polonium)

Group VII The Halogens

This group is composed of entirely non-metals. The term **halogens** come from the Greek word (**hals***)* which means salt and **gene***s* which means forming. Halogens group are called (**salt former).**

Group VIII The Noble Gases

This group is composed of stable gases otherwise known as the non-reactive or inert elements.

The transition elements

The elements in the middle of the table are called transition elements. They are all metals and so they are also called transition metals. A group was devised by the **International Union of Applied and Pure Chemistry (IUPAC)** to eliminate confusion.

The table below shows the example of trends across a period in the periodic table. (Period 3)

Group	I	II	III	IV	V	VI	VII	0
Element	Sodium	Magnesium	Aluminium	Silicon	Phosphorus	Sulphur	Chlorine	Argon
Valence electrons	1	2	3	4	5	6	7	8
Element is a…..	Metal	Metal	Metal	Metal -loid	Non-metal	Non-metal	Non-metal	Non-metal
Reactivity	high →			low →			high	Unre-active
Melting point/°C	98	649	660	1410	590	119	-101	-189
Boiling point/°C	883	1107	2467	2355	Ignites	445	-35	-186
Oxide is….	basic →		amphoteric →				acidic	-
Typical compound	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PH_3	H_2S	HCl	-
Valency shown in that compound	1	2	3	4	3	2	1	-

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 6

30 minutes

Answer the following questions:

1. What is the other name for the group I elements? ____________________________

2. Explain why the noble gases are unreactive.
____

3. Draw the atomic structure of the following elements and identify what group and period they belong?

a) Zn

Group______ period ______

b) Al

Group______ period ______

c) F

Group______ period ______

Thank you for completing your learning activity 6. Check your work. Answers are at the end of this module.

11.2.3 Bonding and Structure

Chemical bonding

You have read about the electronic configuration of atoms of different elements and variation in the periodic properties of elements. Chemical bonds which join the atoms together will give various types of substances. The discussion would also highlight how these bonds are formed.

The properties of substances depend on the nature of bonds present between their atoms. In this lesson you will learn that sodium chloride, the common salt and washing soda dissolve in water whereas methane gas or napthalene do not. This is because the type of bonds is present between them are different. In addition to the difference in solubility, these two types of compounds differ in other properties.

Why do atoms combine?

The answer to this question is hidden in the electronic configurations of the noble gases. It was found that noble gases namely: helium, neon, argon, krypton, xenon and radon did not react with other elements to form compounds. They are not reactive. They were also called inert gases due to their non-reactive nature.

Thus, it was thought that these noble gases lacked reactivity because of their stable electronic configuration. When we write the electronic configurations of the noble gases (see table below), we find that all of them have 8 electrons in their outermost shell except helium.

Name	Symbol	Atomic number	Electron configuration	Number of electrons in the outermost shell
Helium	He	**2**	**2**	**2**
Neon	Ne	**10**	**2,8**	**8**
Argon	Ar	**18**	**2,8,8**	**8**
Krypton	Kr	**36**	**2,8,18,8**	**8**
Xenon	Xe	**54**	**2,8,18,18,8**	**8**
Radon	Ra	**86**	**2,8,18,32,18,8**	**8**

Electronic configuration of noble gases

Atoms having 8 electrons in their outermost shell are very stable. They did not form compounds. It was also observed that other atoms such as hydrogen, sodium, chlorine do not have 8 electrons in their outermost shell undergo chemical reactions. They can be stable by combining with each other and attain the above configurations of noble gases.

Example, 8 electrons (or 2 electrons in case of helium) are in their outermost shells. Thus, atoms tend to attain a configuration in which they have 8 electrons in their outermost shells.

This is the basic cause of chemical bonding. This attainment of eight electrons for stable structure is called the **octet rule**. The octet rule explains the chemical bonding in many compounds.

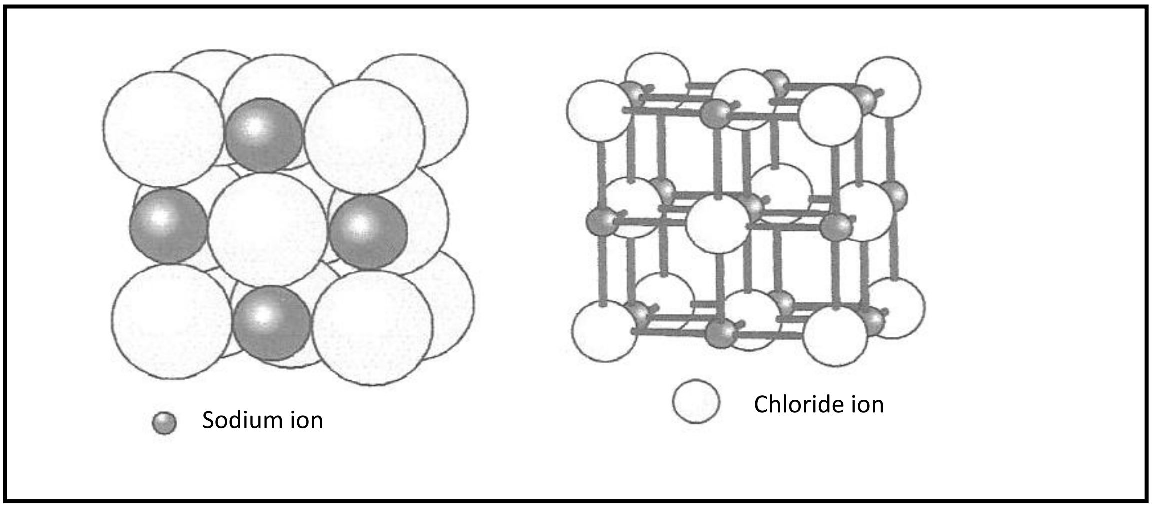
Atoms are held together in compounds by the forces of attraction which result information of **chemical bonds**. The formation of chemical bonds results in the lowering of energy which is less than the energy of the individual atoms. The resulting compound is lower in energy as compared to sum of energies of the reacting atom/molecule and hence is more stable. Thus, stability of the compound formed is an important factor in the formation of chemical bonds.

Types of chemical bonds

1. The Ionic bond

The chemical bond formed by transfer of electron from a metal to a non- metal is known as **ionic or electrovalent bond.**

For example, when sodium metal and chlorine gas are brought into contact, they react violently and we obtain sodium chloride. This reaction is shown below:



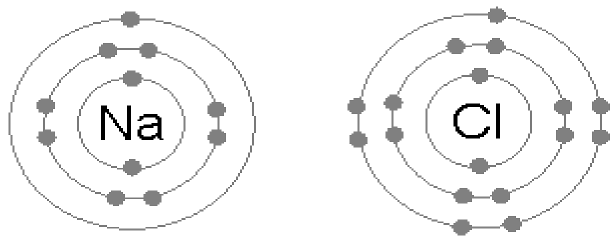
Reaction of sodium chloride ions

$$2Na(s) + Cl_2\,(g) \rightarrow 2NaCl(s)$$

The bonding in sodium chloride can be understood as follows:
Sodium (Na) has the atomic number 11 and we can write its electronic configuration as 2,8 and 1 . It has one electron in its outermost (M) shell. If it loses this electron, it is left with 10 electrons and becomes positively charged. Such a positively charged ion is called a **cation**. The cation in this case is called sodium ion, Na^+.

Na (s)	+ Cl (g)	$\longrightarrow$	$Na^+\ Cl^-$	$\longrightarrow$	NaCl (S)
2, 8, 1	2, 8, 7		2, 8		2, 8, 8

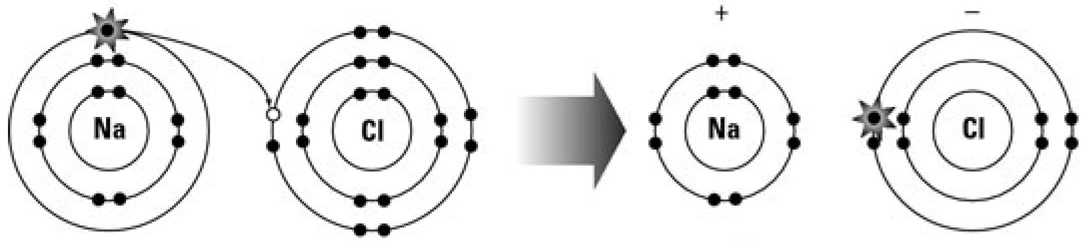
Proton transfer



Na atom. It has 1 electron in its outer shell

Cl atom. It has 7 electrons in its outer shell

The atom of sodium and chlorine



Sodium gives its outer electron to chlorine

The ion of sodium and chlorine.

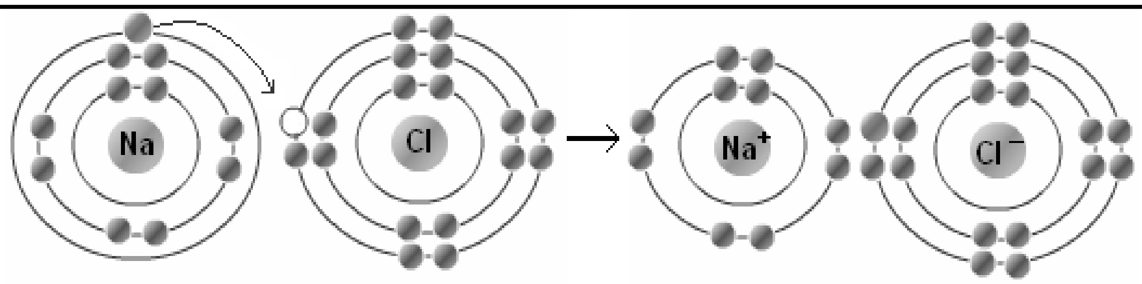
Remember that all atoms are neutral. They have an equal number of positive protons and negative electrons. So the charges cancel out, the electrons and protons in both atoms no longer balance.

Let's add up the charges:

Sodium 10 electrons $= 10^-$
11 Protons $= \underline{11^+}$
1^+

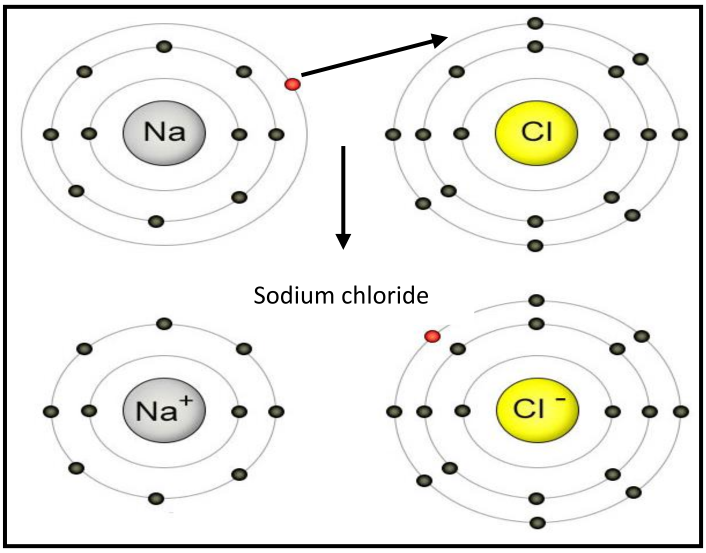
Chlorine 18 electrons $= 18^-$
17 protons $= \underline{17^+}$
1^-

The atoms, which we now call **ions,** become charged: Na^+ and Cl^-



Na= 2,8,1 Na= 2,8,7 Na= 2,8 Na= 2,8,8

The ionic compound of sodium chloride.

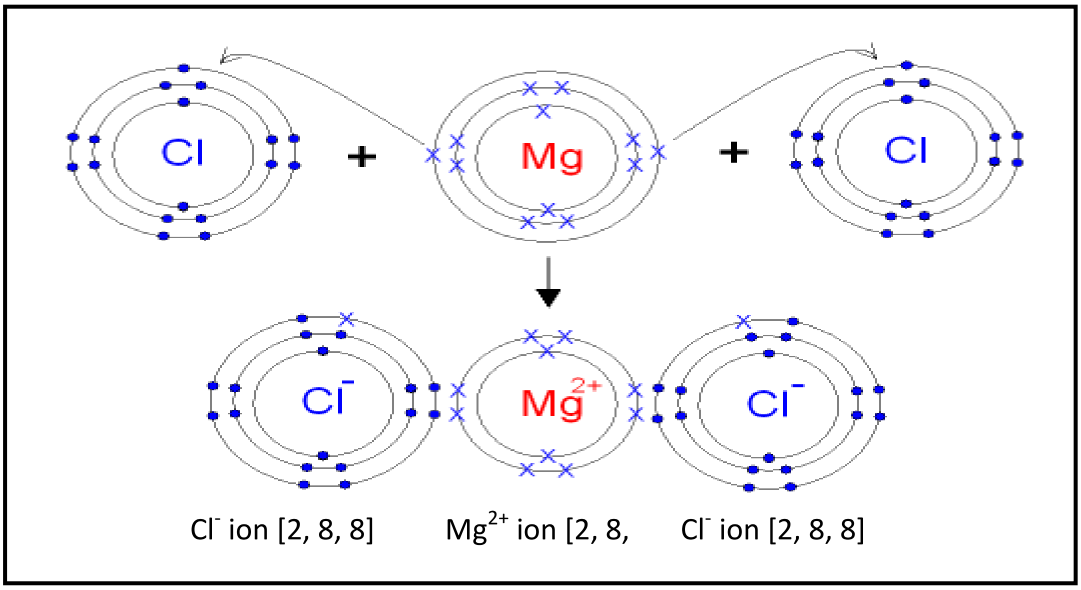


Compound of sodium chloride.

Another example of ionic bonding is between Magnesium and Chlorine atom. We know that in the first example between sodium and chlorine the atoms satisfy each other – one to one. So the formula of Sodium chloride is NaCl. How about Magnesium? We know that it has 2 electrons in the outer shell. What do you think will happen?

Look at the diagram below:

$$Mg^{2+} + 2Cl^{-} \longrightarrow MgCl_2$$



You can see now that each magnesium atom can `satisfy' 2 chlorine atoms.

Properties of ionic compounds

Since the ionic compounds contain ions (cations and anions) which are held together by the strong electrostatic forces of attraction, they show the following general characteristic properties:

(a) Physical State

Ionic compounds are crystalline solids. In the crystal, the ions are arranged in a regular pattern. The ionic compounds are hard and brittle in nature.

(b) Melting and boiling points

Ionic compounds have high melting and boiling points. The melting point of sodium chloride is 1074 K (801°C) and its boiling point is 1686K (1413°C). The melting and boiling points of ionic compounds are high because of the strong electrostatic forces of attraction present between the ions.

Thus, it requires a lot of thermal energy to overcome these forces of attraction. The thermal energy given to the ionic compounds is used to overcome the strong force of attractions present in the cations and anions in an ionic crystal. Remember that the crystal has a three dimensional regular arrangement of cations and anions which is called **crystal lattice**. On heating, the breaking of this crystal lattice leads to the molten state of the ionic compound in which the cations and anions are free to move.

(c) Electrical Conductivity

Ionic compounds conduct electricity in their molten state and in aqueous solutions. Since ions are free to move in the molten state, they can carry current from one electrode to another in a cell.

Thus, ions can conduct electricity in molten state. However, in solid state, such a movement of ions is not possible as they occupy fixed positions in the crystal lattice. Hence in solid state, ionic compounds do not conduct electricity. In aqueous solution, water is used as a solvent to dissolve ionic compounds. It weakens the electrostatic forces of attraction present among the ions. When these forces are weakened, the ions become free to move.

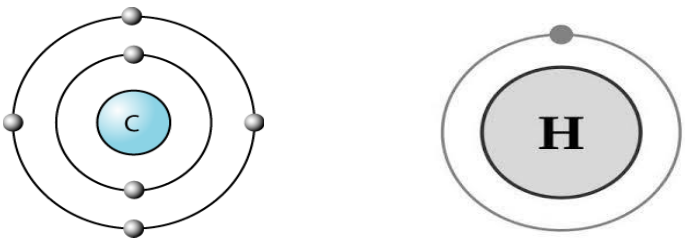
2. The Covalent Bond

In this section, we will study about another kind of bonding called **covalent bonding.** Covalent bonding is helpful in understanding the formation of molecules. In the previous lesson you studied about molecules having similar atoms such as hydrogen (H_2), chlorine (Cl_2), oxygen (O_2), and nitrogen (N_2). Molecules of elements whereas those containing different atom like hydrochloric acid HCl, nitric acid NH_3, methane CH_4, carbon dioxide CO_2 are molecules of compounds.

Let us now see how these molecules are formed. Let us consider the formation of hydrogen molecule (H_2).

The hydrogen atom has one electron. It can attain the electronic configuration of the noble gas helium by sharing one electron of another hydrogen atom. When the two hydrogen atoms come closer, there is an attraction between the electrons of one atom and the proton of another and there are repulsions between the electrons as well as the protons of the two hydrogen atoms. In the beginning, covalent bond forms by sharing electrons between non-metals.

Two non- metals atoms can both gain electrons by sharing. If their outer shells overlap, they share some of each other's electrons. This can `satisfy' both atoms.
Let us look at the smallest hydrocarbon, Methane (CH_4).



Carbon atom has 4 electrons in its outer shell. It needs to get full outer shell.

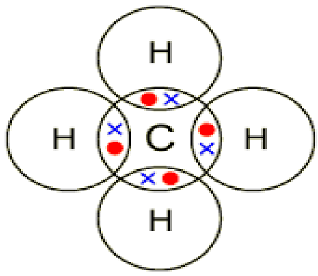
Hydrogen has just 1 electron. If it can gain 1 more electron, it will fill its shell.

Atoms of carbon and hydrogen.

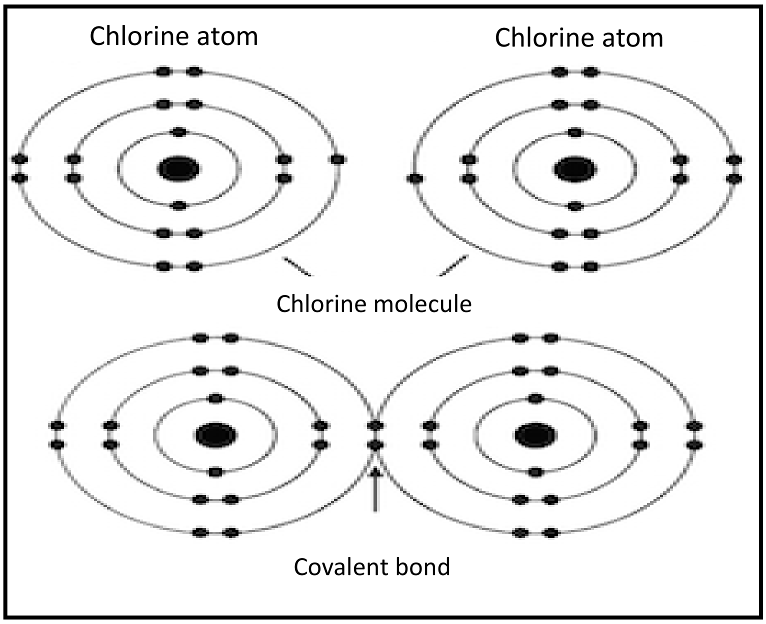
Count the electrons in the outer shell of each carbon and hydrogen atom in methane. Are they full?

Can you see that the carbon now has 8 electrons in its outer shell? And hydrogen has 2?

It means that all atoms of hydrogen and carbon are stable.



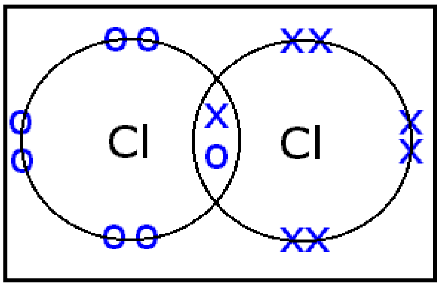
Let us look at another example of covalent bonding. **Chlorine gas does not exist as single** atoms. Remember that a chlorine atom has 7 electrons in its outer shell. It needs 1 more electron to fill it.



Ways to show the covalent bond between Cl_2. The right – hand is a `dot and cross' diagram. (showing the outer electrons only.)

Like many gases, it is a **diatomic**. Two chlorine atoms bond together to make a Cl_2 molecules.

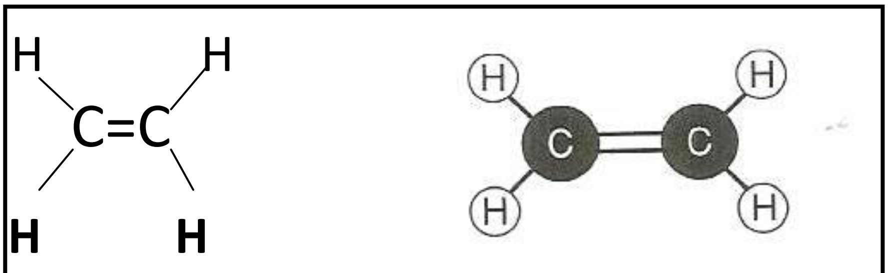
Look at the diagram below of Cl_2 molecules:



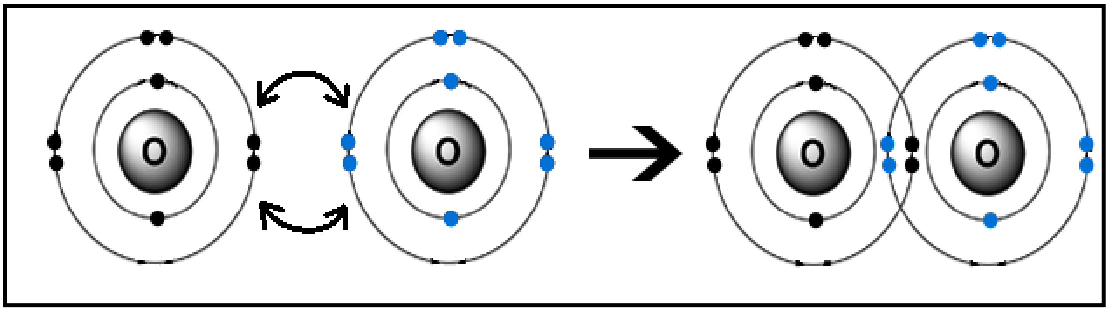
Chlorine molecule

Double bond
Do you recall the molecules ethene? Its formula is C_2H_4. We said that its carbon atoms joined by a double bond.

Let us look at the diagram below:



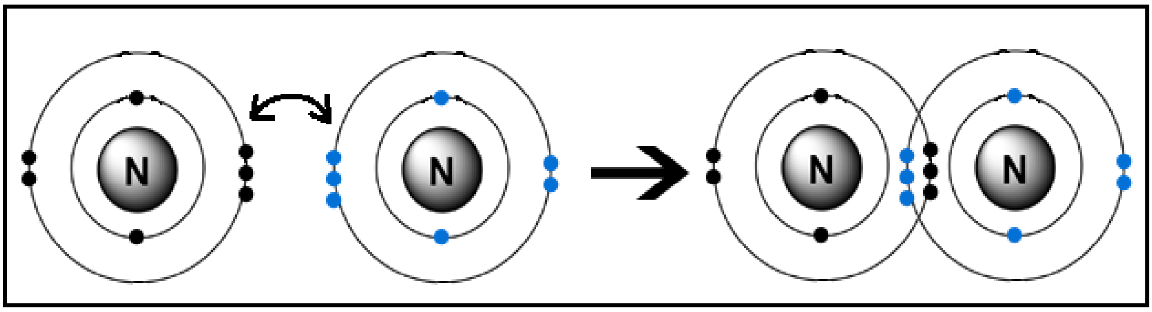
Ethane has a double covalent bond between its carbon atoms



Oxygen is another molecule contains a double bond.

Triple bond
Another type of covalent bond is the triple bond. Can you think of elements that form a triple bond?

Example: nitrogen molecule



Nitrogen molecule showing the outer shell electrons.

Properties of covalent substances

The covalent compounds consist of molecules which are electrically neutral in nature. The forces of attraction present between the molecules are less strong as compared to the forces present in ionic compounds. Therefore, the properties of the covalent compounds are different from those of the ionic compounds. The characteristic properties of covalent compounds are given below:

a. Physical state

Because of the weak forces of attraction present between discrete molecules, called intermolecular forces, the covalent compounds exist as a gas or a liquid or a solid. For example O_2, N_2, CO_2 are gases; water and CCl_4 are liquids and iodine is a solid.

b. Melting and boiling points

As the forces of attraction between the molecules are weak in nature, a small amount of energy is sufficient to overcome them. Hence, the melting points and boiling points of covalent compounds are lower than those of ionic compounds. For example, melting point of naphthalene which is a covalent compound is 353 K (80°C).Similarly, the boiling point of carbon tetrachloride which is another covalent liquid compound is 350 K (77°C).

c. Electrical conductivity

The covalent compounds contain neutral molecules and do not have charged species such as ions or electrons which can carry charge. Therefore, these compounds do not conduct electricity and are called poor conductors of electricity.

Giant covalent structures

The great variety of life on Earth depends on carbon's ability to form covalent bonds with itself. As the element, carbon atoms can bond to millions of other carbon atoms in both **diamond and graphite**, the allotropes of carbon. **Allotrope is the same element with different physical properties.**

The structures, properties and uses of diamond and graphite.

Diamond	Graphite
It occurs naturally in free state.	It occurs naturally and manufactured artificially.
It is the hardest natural substance known.	It is soft and greasy to touch.
It has high relative density (3.5).	Its relative density is 2.3.
It has transparent and has refractive.	It is black in colour and opaque.
It is non-conductor of heat and electricity.	It is a good conductor of heat and electricity.
It burns in air at 900^{o}C to give carbon dioxide.	It burns in air at 700-800^{o}C to give carbon dioxide.
It occurs as octahedral crystals.	It occurs as hexagonal crystals.
It is insoluble in all substance.	It is insoluble in all ordinary substance.

Carbon in the form of diamonds

Diamond's hardness makes it very useful material. It is a form from carbon atoms only.

Uses of diamonds:

(i) Diamonds are valuable gemstones. Larger and purer the diamond, the more valuable it is.

(ii) Diamonds are used in jewellery such as rings, earrings, pendants, bangles or necklaces.

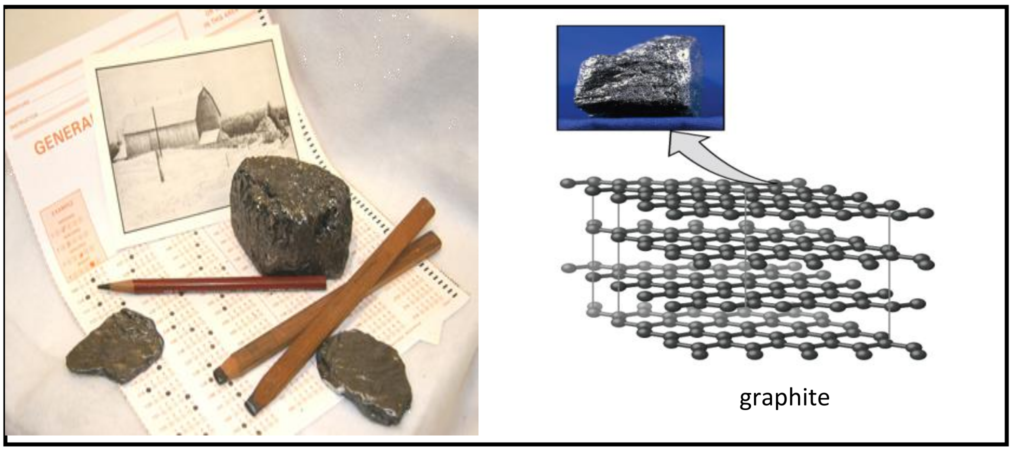
(iii) Smaller pieces of diamonds are used for cutting glass and drilling rocks.

(iv) Only a diamond can cut another diamond. Diamond dust is used for polishing diamonds and precious stones.

(v) In the future, diamonds may be used for surgical tools, medical devices, and prosthethic human joints.

Carbon in the form of graphite

Another form of carbon is graphite. If you touch a lump of graphite, it feels smooth and slippery. One use of graphite is the pencil leads. If you use pencil to draw, as you move it across your paper it flakes off, leaving a trail of carbon atoms.



Some other uses of graphite:

(i) Graphite being smooth and slippery, is used in making lubricants for use in machinery, motorcar engines and even bicycle chains.

(ii) Graphite being chemically unreactive and a conductor of electricity, is used in making electrodes for use in electrolysis and in dry cells.

(iii) Due to its very high melting point, graphite is used as a heat insulator. It is used to coat the nose of a space shuttle.

3. Metallic bonding

Think of some of the things around your home that are made of metal. Did you include all the wiring, any radiators, your hot – water tank, or your cutlery and pans?

Different kinds of metals

Do you know which metals these things are made from? Which properties make metals good for these uses?

In the last topic you have learnt ionic and covalent bonding. The atoms in a metal are held together in a different way.

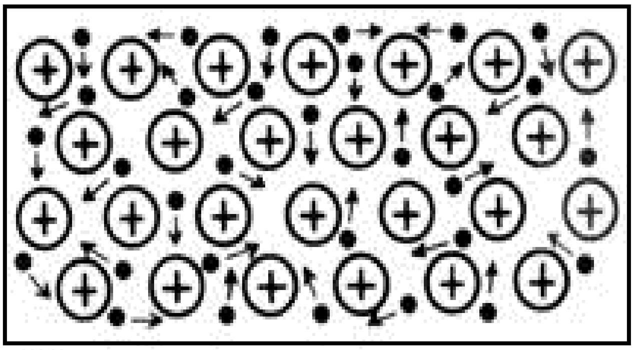
Do you remember all the properties of metals? Any ideas we have about the bonding and structure of metals must be able to explain their properties.

Properties of metals

- Have high melting and boiling points.
- Conduct electricity and heat.
- Are hard and dense.
- Can be hammered into shapes (they are malleable).
- Can be drawn out into wires (they are ductile).

We believe that metal atoms (or ions) are held together by a **sea of electrons.**

Look at the diagram below:

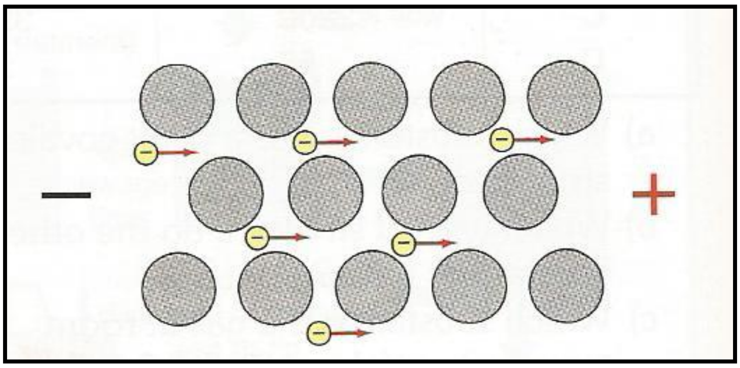


Metal ions in their sea of electrons.

Each metal atom gives up one or more of its electrons into the `sea' or `cloud' of electrons. The electrons can drift about in the metal. These free electrons explain how electricity can pass through solid metals.

What happens when one end of the metal is made up of positive and the other end negative?

The diagram below represents the movement of electrons.



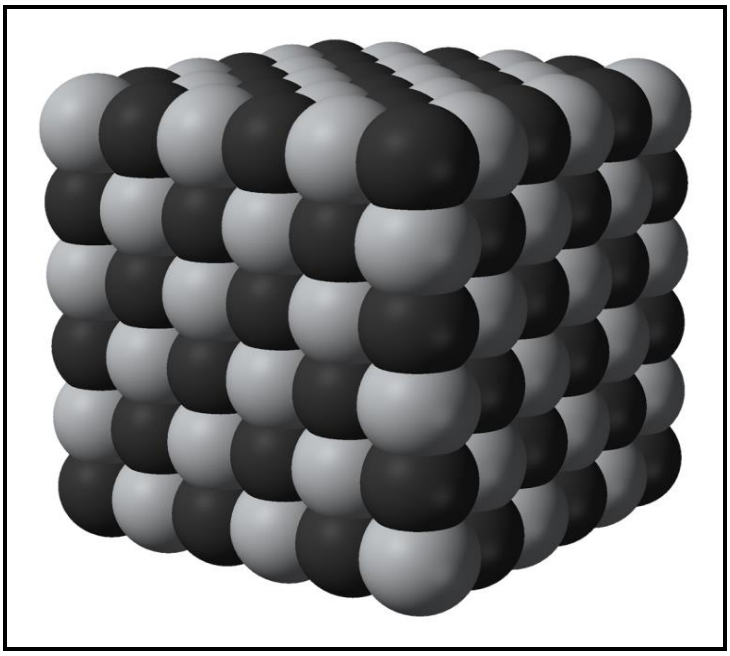
Electrons move towards the positive charge. These free electrons can also transfer heat through metals quickly.

Structure of the metals

As you know, most metals are dense. This suggests that their atoms must **be packed closely** together. They also have high melting points. And their atoms are arranged in **giant structures.** Metals are giant structures of atoms held together by metallic bonds. **Giant** implies that large but variable numbers of atoms are involved - depending on the size of the bit of metal.

Metals form giant structures in which electrons in the outer shells of the metal atoms are free to move. The metallic bond is the force of attractions between these free electrons and metal ions.

Metallic bonds are strong, so metals can maintain a regular structure and usually have high melting and boiling points. Metals are good conductors of electricity.



Giant structure of metal.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 7

40 minutes

Answer the following questions:

1. Name two ions present in sodium chloride.
   I. ______________

   ii ______________

2. How many shells are present in Na^+ ion? ______________

3. What is the number of electrons present in Cl^- ion? ______________

4. How are covalent bonds formed?
   ____

5. Draw the outermost shell electron diagram of the following compound:

i. O_2

ii. HCl

iii. Cl_2

iv. N_2

Thank you for completing your learning activity 7. Check your work. Answers are at the end of this module.

REVISE WELL USING THE MAIN POINTS ON THE NEXT PAGE.

SUMMARY

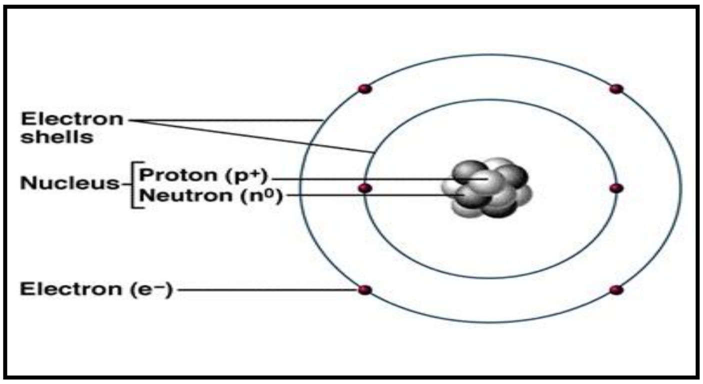
You will now revise this module before doing Assessment 6. Here are the main points to help you revise. Refer to the module topic if you need more information.

Atom contains **proton, neutrons and electrons.**

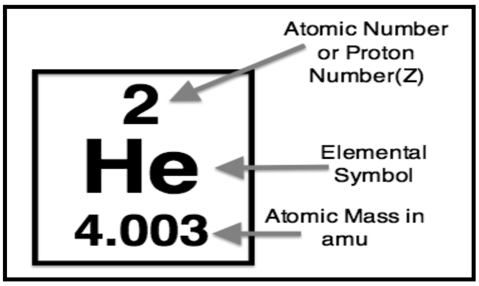
The **nucleus** is composed of **protons;** the positively charged particle and the **neutrons**, the particle with no charge. Around the nucleus is the **electron**, the negatively charged particle.

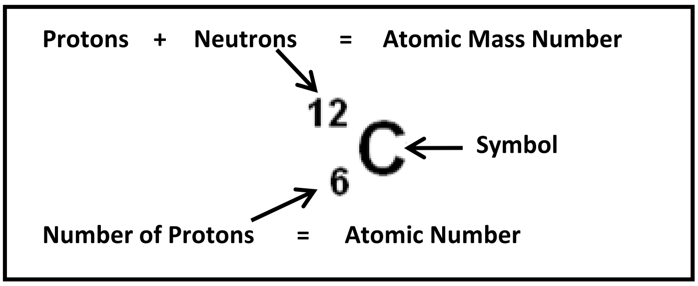
Particle	Charge	Mass (in atomic mass module)
Proton	1+	1
Neutron	0	1
electron	1-	0 (almost)

- Proton and neutrons have the same mass.
- The electrons are so light that you can ignore their mass.



The atomic structure





Position of atomic mass and atomic number in the Periodic Table.

If we add up the RAM as in the formula, we get the **Relative Formula Mass**

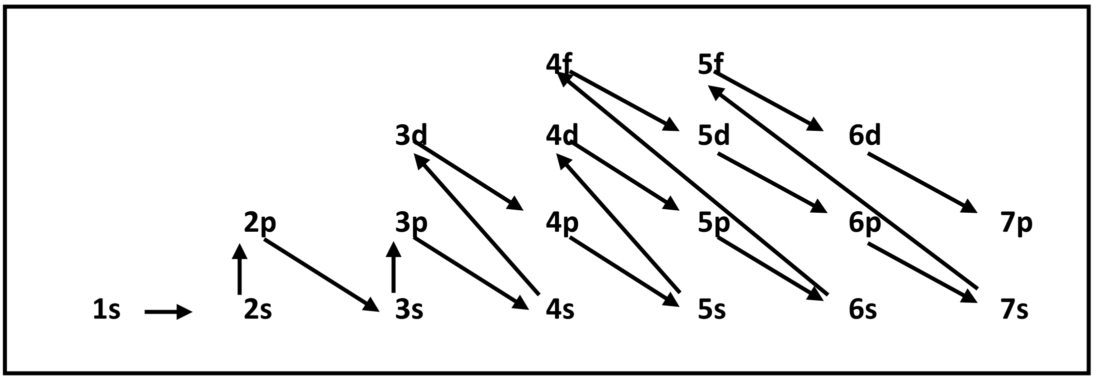
Electrons in orbit

- The electrons orbit the nucleus in shells
- The 1st shell can hold 2 electrons
- The 2nd shell can hold 8 electrons. As can the 3rd shell

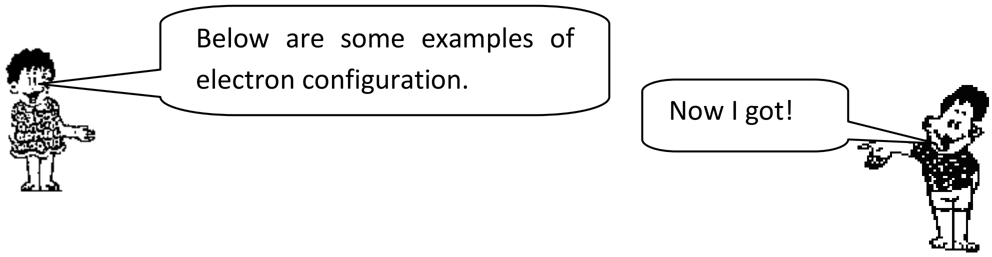
The distribution of electrons among the orbitals of an atom is its **electron configuration.**
S – sharp **2** electrons to fill
p – principal **6** electrons to fill
d – diffuse **10** electrons to fill
f – fundamental **14** electrons to fill

$1s^1$																	$1s^2$
$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$3d^5$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^{10}$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^4$	$4d^5$	$4d^5$	$4d^7$	$4d^8$	$4d^{10}$	$4d^{10}$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
$6s^1$	$6s^2$		$5d^2$	$5d^3$	$5d^4$	$5d^5$	$5d^6$	$5d^7$	$5d^9$	$5d^{10}$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
$7s^1$	$7s^2$		$6d^2$	$6d^3$	$6d^4$	$6d^5$	$6d^6$	$6d^7$	$6d^8$	$6d^{10}$	$6d^{10}$	$7p^1$	$7p^2$	$7p^3$	$7p^4$	$7p^5$	$7p^6$
			$5d^1$	$4f^1$	$4f^3$	$4f^4$	$4f^5$	$4f^6$	$4f^7$	$4f^7$	$4f^9$	$4f^{10}$	$4f^{11}$	$4f^{12}$	$4f^{13}$	$4f^{14}$	$4f^{14}$
			$6d^1$	$6d^2$	$5f^2$	$5f^3$	$5f^4$	$5f^6$	$5f^7$	$5f^7$	$5f^9$	$5f^{10}$	$5f^{11}$	$5f^{12}$	$5f^{13}$	$5f^{14}$	$5f^{14}$

Total number of electrons in an energy level can be found by the formula:
number of electrons = $2n^2$
Where **n** is the number of energy level;
1st = 2
2nd = 8 = (2+6)
3rd = 18 = (2 + +2+ 6 = 10)
4th = 32 = (2 + 6 + 10 + 14)



Electronic configuration



Elements	Atomic Number	Electron Configuration	Period	Group
1.Hydrogen, H	1	$1s^1$	1	I A
2.Lithium, Li	3	$1s^22s^1$	2	I A
3.Berylium, Be	4	$1s^22s^2$	2	II A
4.Carbon, C	6	$1s^22s^22p^2$	2	IV A
5.Nitrogen, N	7	$1s^22s^22p^3$	2	V A
6.Aluminum, Al	13	$1s^22s^22p^63s^23p^1$	3	III A
7.Sulfur, S	16	$1s^22s^22p^63s^23p^4$	3	VI A
8.Potassium, K	19	$1s^22s^22p^63s^23p^64s^1$	4	I A
9.Bromine, Br	35	$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$	4	VII A

Periodic Table

The Development of Periodic table

Johann Wolfgang Döbereiner showed that the atomic weight of strontium lies midway between those of calcium and barium. Some years later he showed that other such **"triads"** exist (chlorine, bromine, and iodine [halogens] and lithium, sodium, and potassium [alkali metals]). He also showed that similar relationships extended further than the triads of elements, fluorine being added to the halogens, and magnesium to alkaline-earth metals. Oxygen, sulfur, selenium, and tellurium were classed as one family, and nitrogen, phosphorus, arsenic, antimony, and bismuth as another family of elements.

1863, A.E.B, De Chancourtois proposed a classification of the elements based on the new values of atomic weights given by Stanislao Cannizzaro's system of 1858. De Chancourtois plotted the atomic weights on the surface of a cylinder with a circumference of 16 modules, corresponding to the approximate atomic weight of oxygen. The resulting helical curve brought closely.

400 BC Democritus suggest that all things are made of particles.

1805 John Dalton's atomic theory. Atoms of the same element are alike. They combine to make compounds.

1909 Ernest Rutherford discover the proton.

1911 Ernest Rutherford discover the nucleus.

1913 Neil's Bohr suggests those electrons are found in shells around the nucleus.

1932 James Chadwick proves that neutrons exist.

Periodic table

The periodic table is a chart of all the known elements in order of increasing atomic number. The table puts elements into groups with similar characteristics, allowing us to recognize trends over the whole array of elements.

Atom

An atom is the smallest module of a substance that still has all the properties of that substance.

Atomic number

The atomic number of an atom is equal to the number of protons that the atom contains.

Atomic weight/mass

The atomic weight of an atom is a measure of how much mass an atom has. The atomic weight is calculated by adding the number of protons and neutrons together.

Periodic Variations

The groups (Group 1 to 8)

The periodic table is composed of 7 rows or periods and 18 major groups or columns. Elements in a group have similar properties particularly those elements in four of the groups: Group IA – The alkali metals; Group II A - the alkaline earth metals; Group VII A – the halogens and Group VIII A – the noble gases.

Group IA The Alkali Metals

Group 1 elements are soft and silvery metals. They react strongly with water.

Group IIA The Alkaline Earth Metals

This group consists of all metals that occur naturally in compound form. They are obtained from mineral ores and form alkaline solutions. These are less reactive than alkali metals.

Group IIIA The Aluminum Group

The elements in this group are fairly reactive. The group is composed of four metals and one metalloid which is boron.

Group IVA The Carbon Group

This group is composed of elements having varied properties because their metallic property increases from top to bottom meaning the top line.

Group VA The Nitrogen Group

Like the elements in group IV A, this group also consists of metals, nonmetal and metalloids.

Group VIA The Oxygen Group

This group is called the oxygen group since oxygen is the top line element.

Group VIIA **The Halogens**

This group is composed of entirely nonmetals. The term **halogen** comes from the Greek word **hals** which means salt and *genes* which means forming. Halogens group are called **salt formers**.

Group VIIIA **The Noble Gases**

This group is composed of stable gases otherwise known as the non-reactive or inert elements.

Bonding and Structures

Chemical bonding

The chemical bond formed by transfer of electron from a metal to a non- metal is known as ionic or electrovalent bond. Ionic bonds form between metals and non- metals.

Properties of Ionic Compounds

Physical State

Ionic compounds are crystalline solids. In the crystal, the ions are arranged in a regular fashion. The ionic compounds are hard and brittle in nature.

Melting and Boiling points

Ionic compounds have high melting and boiling points.

Electrical Conductivity

Ionic compounds conduct electricity in their molten state and in aqueous solutions.

Covalent bond forms by sharing electrons between non-metals.
Covalent substances can have either

a) Giant covalent structure or
b) Simple molecular structure

Properties of Covalent Substances

The covalent compounds consist of molecules which are electrically neutral in nature.The forces of attraction present between the molecules are less strong as compared to the forces present in ionic compounds.

Physical State

Because of the weak forces of attraction present between discrete molecules, called intermolecular forces, the covalent compounds exist as a gas or a liquid or a solid.

Melting and Boiling Points

As the forces of attraction between the molecules are weak in nature, a small amount of energy is sufficient to overcome them. Hence, the melting points and boiling points of covalent compounds are lower than those of ionic compounds.

Electrical Conductivity
The covalent compounds contain neutral molecules and do not have charged species such as ions or electrons which can carry charge. Therefore, these compounds do not conduct electricity and are called poor conductors of electricity.

Metallic bonding are bonded together by sea of electrons. The electrons are free to drift between atoms. They move in one direction when electric current flow. Atoms are arranged in giant structures.

Properties of the different types of chemical

	Ionic	**Metallic**	**Covalent**	**Covalent**
Type of structure	Giant	Giant	Giant	Simple molecules
Are ions present?	Yes	Yes	No	No
Are delocalized present?	No	Yes	No (except in graphite)	No
How strong is the chemical bond?	Strong	Strong	Very strong	Very strong
High/Low melting point	High	High	Very high	Low
Conductor of electricity	When molten/in solution	Yes	No (except graphite	No

NOW YOU MUST COMPLETE ASSESSMENT 2 AND RETURN IT TO THE PROVINCIAL CENTRE CO-ORDINATOR.

ANSWERS TO LEARNING ACTIVITIES 1 - 7

Learning Activity 1

1. a) Proton
   b) Electron
   c) Neutron

2.

Sub – atomic particles	**charge**	**Mass (atomic mass module)**
Proton	**1+**	**1**
neutron	0	1
electron	**1**-	0 (almost)

3.

i. number of protons 7

number of electrons 7

number of neutrons 7

ii. number of protons 19

number of electrons 19

number of neutrons 20

Iii. number of protons 10

number of electrons 10

number of neutrons 10

Learning Activity 2

1. i. Atom of the same element with the same number of protons but different number of neutrons.
   ii. Used to separate the different atoms or isotopes of an element.

2. i. $\frac{63 \times 69 + 65 \times 31}{100}$

   = 63.62

   ii. $\frac{16 \times 99.79 + 17 \times 0.04 + 18 \times 0.20}{100}$

   = 16.0092

Learning Activity 3

1. Show how the electrons are arranged in the selected element and the distribution of electrons among the orbital's of an atom.

2. Draw the electron configuration of the following using the **s, p, d and f.**

   i. Na = $1s^2 2s^2 2p^6 3s^1$

   ii. Ca = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

   iii. Cl = $1s^2 2s^2 2p^6 3s^2 3p^5$

   iv. Zn = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

   v. F = $1s^2 2s^2 2p^5$

   vi. Al = $1s^2 2s^2 2p^6 3s^2 3p^1$

Learning Activity 4

1. i. Is the smallest module of a substance.
   ii. Is the total number of protons and neutrons.
   iii. The total number of protons in the nucleus of an atom.

2. 1932

3. Neil's Bohr

Learning Activity 5

1. i. Possess the properties of both the metals and non-metals.
   ii. The energy required to remove the outer electron from an isolated atom.
   iii. Is the energy released when neutral atoms gain an extra electron to form negatively charged ion.
   iv. The electron attracting ability to an atom.

2. i. Atoms get larger going down a group.
   ii. Atoms get smaller moving from left to right across each period.

3. Elements are arranged from the lightest to heaviest.

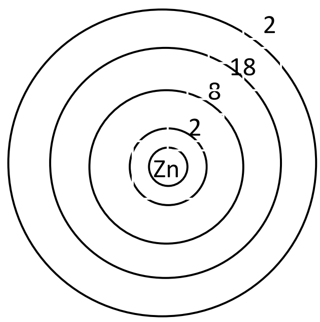
Learning Activity 6

1. Alkali group

2. Because their atoms have already have a stable outer most electron shell.

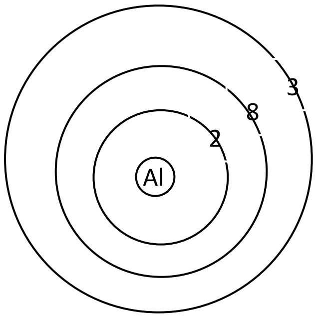
3. Draw the atomic structure of the following and elements and identify what group and period they belong?

a) Zn



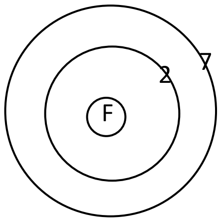
Group **2** Period **4**

b) Al



Group **3** Period **3**

c) F



Group **7** Period **2**

Learning Activity 7

1. Name the ions present in sodium chloride.
   i. Chloride
   ii. Sodium

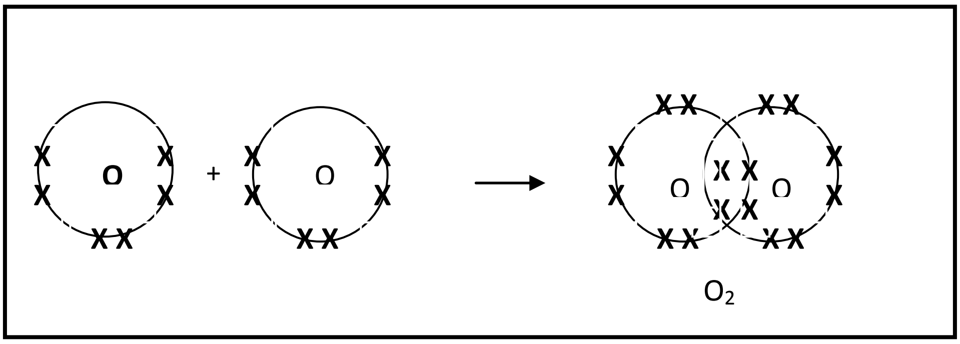
2. $\boxed{2}$

3. 18

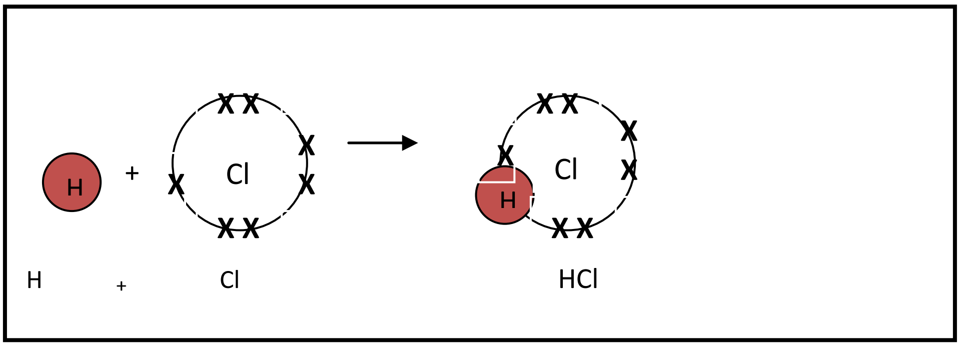
4. By sharing of electrons.

5.

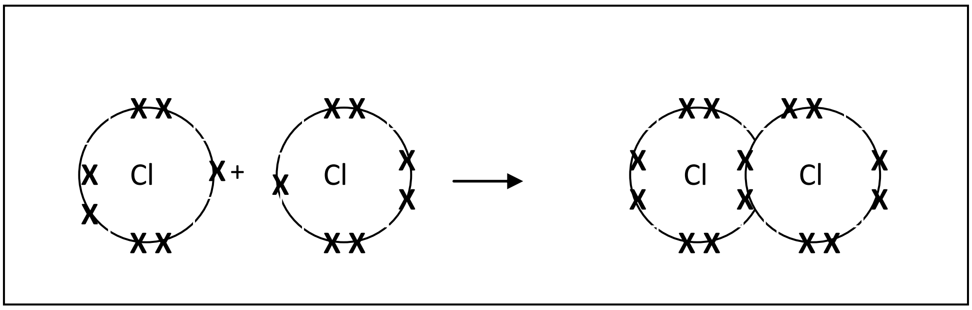
i. O_2,



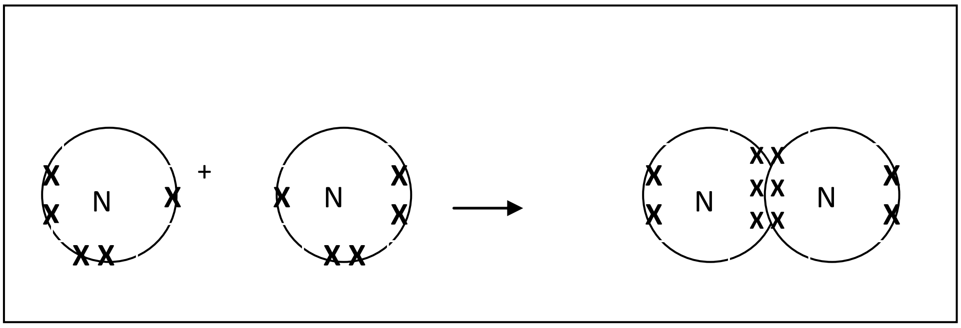
ii. HCl,



ii. Cl_2



Iv . N_2.



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	3. Personal Development
	4. Social Science
	5. Science
	6. Making a Living
Grades 9 and 10	1. English
	2. Mathematics
	3. Personal Development
	4. Science
	5. Social Science
	6. Business Studies
	7. Design and Technology- Computing
Grades 11 and 12	1. English – Applied English/Language& Literature
	2. Mathematics – General / Advance
	3. Science – Biology/Chemistry/Physics
	4. Social Science – History/Geography/Economics
	5. Personal Development
	6. Business Studies
	7. Information & Communication Technology

REMEMBER:

- For Grades 7 and 8, you are required to do all six (6) subjects.
- For Grades 9 and 10, you must complete five (5) subjects and one (1) optional to be certified. Business Studies and Design & Technology – Computing are optional.
- For Grades 11 and 12, you are required to complete seven (7) out of thirteen (13) subjects to be certified.

Your Provincial Coordinator or Supervisor will give you more information regarding each subject and course.

Notes: You must seek advice from your Provincial Coordinator regarding the recommended courses in each stream. Options should be discussed carefully before choosing the stream when enrolling into Grade 11. FODE will certify for the successful completion of seven subjects in Grade 12.

GRADES 11 & 12 COURSE PROGRAMMES			
No	**Science**	**Humanities**	**Business**
1	Applied English	Language & Literature	Language & Literature/Applied English
2	General / Advance Mathematics	General / Advance Mathematics	General / Advance Mathematics
3	Personal Development	Personal Development	Personal Development
4	Biology	Biology/Physics/Chemistry	Biology/Physics/Chemistry
5	Chemistry/ Physics	Geography	Economics/Geography/History
6	Geography/History/Economics	History / Economics	Business Studies
7	ICT	ICT	ICT

CERTIFICATE IN MATRICULATION STUDIES		
No	**Compulsory Courses**	**Optional Courses**
1	English 1	**Science Stream:** Biology, Chemistry and Physics
2	English 2	**Social Science Stream:** Geography, Intro to Economics and Asia and the Modern World
3	Mathematics 1	
4	Mathematics 2	
5	History of Science & Technology	

REMEMBER:

You must successfully complete 8 courses: 5 compulsory and 3 optional.