CHAPTER 2

THE PERIODIC TABLE

2.1 HISTORY OF THE PERIODIC TABLE

The first classification of elements that became the basis of the present day classification was done by a Russian chemist called Dimitri Medelev. In 1869 Mendeleev arranged the elements in increasing atomic weight to form a periodic table. He pointed out that element with similar properties kept reoccurring at regular intervals or periods, ignoring the fact that some elements do not follow that order. Mendeleev placed elements with similar properties in horizontal rows, he called group, while the periods were vertically arranged. He equally left space in his periodic table for elements that were not discovered yet. Apart from spaces left in Mendeleev periodic table, his arrangement is different from the modern arrangement of the periodic table.

From his work, Mendeleev proposed a periodic law which states that the properties of the elements and their compounds vary in a periodic way according to their atomic weight Mendeleev‟s work was based on the fact that he believed that the only property of an element which does not change in cause of a chemical combination was its atomic weight. But this is not true, as it was later discovered from Moseley‟ s work that the atomic number was the only property in an element that does not change in cause of a chemical combination.

2.3 MODERN PERIODIC LAW

Henry Moseley was able to reveal the atomic numbers of several elements from his study of X-ray emitted by elements in a discharge tube. Based on his new discovery and the inconsistencies in the Mendeleev periodic table, Moseley suggested that the elements should be arranged in order of increasing atomic number instead of atomic mass in the periodic table. Moseley also proposed the modern periodic law, which states that the properties of the elements vary in a periodic way according to their atomic numbers.

2.3 INFORMATION FORM THE PERIOD TABLE

In the modern periodic table, classification of elements gives.

1. Elements with similar physical and chemical properties are placed on the same vertical row called group or family
2. Elements are placed in horizontal row called periods in increasing atomic numbers from left where the metals are located to the right with the non-metals
3. The table show the symbol of each element, its atomic number its atomic mass and in some occasion the element‟s electronic configuration
4. The table relates the electronic configuration with the electronic structure and properties of the element.
5. It shows at a glance the exact period and group an element belongs.
6. The periodic table is divided into 8 groups of elements and 7 periods, with the transition metals between the group 2 and 3 elements.
7. The periodic table shows the division of elements into metals on the left side and non-metals on the right hand side of the table.

TABLE 2.1VALENCE SHELL ELECTRONIC CONFIGURATIONS FOR EACH GROUP

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						
s^1	s^2	s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5	s^2p^6

2.4 REACTIVITY TREND IN THE PERIODIC TABLE

With regards to chemical reactivity, a reaction occurs when element loss, gain or share their valence electrons with another element or specie. The ease with which elements loss, gain or share electrons in a chemical reaction measure their reactivity.

Reactivity increases down a metallic group as successive members of the same group increases by one additional energy level of electrons. Since metals react by loss of electrons, the metals down the metallic group lose electrons easier than those on top. Therefore lithium is less reactive than sodium and sodium less reactive than potassium.

On the other hand, non-metals react by gain of electrons, a process called electronegativity. The more electronegative element is one which gains electrons more easily and is more reactive. Reactivity (electronegativity) reduces down the group as successive members of the group increases by one additional energy level of electrons and an increased distance between the valence electrons and the positive nuclear pull. This makes it more difficult for the positive nucleus to attract neighboring electrons into the valence energy level, thereby reducing reactivityas we go down the group. Using the group seven elements as an illustration, fluorine is more reactive followed by chlorine, bromine than iodine.

The period cuts across all the groups in the periodic table. As one moves across a particular period from left to right, there is increase electronegativity of the element. Since the element advance toward elements with more electronegative properties

2.5 GROUP OF ELEMENTS

In going down a particular group, it is observed that the elements have similar chemicalproperties. This property is directly related to the electronic configuration as each group of elements has the same number of valence electrons.

2.5.1 GROUP 1 ELEMENT

They are the most electropositive metals. They have one valence electron which they loss in cause of a chemical reaction.

The members are generally referred to as alkali metals because their hydroxides are soluble. They react vigorously with cold water, forming an alkali solution and liberating hydrogen gas. $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

2.5.2 GROUP 2 ELEMENTS

The Group 2 elements are also called the alkaline earth metals. They are less electropositive than the group one metals.

They react by loss of two electrons thus divalent.
They react with cold water mildly, given a basic solution that is less soluble than the group 1 elements.

$$Mg + 2H_2O \longrightarrow Mg(OH)_2 + H_2$$

2.5.3 GROUP 3 ELEMENTS

They are metals which are less electropositive than the Group 1 and 2 elements. They react by loss of their three valence electrons, thus they are trivalent.

2.5.4 GROUP 4 ELEMENTS

They exhibit non-metallic properties and they are tetravalent because they have four valence electrons. They form covalent bond by sharing of electrons. There is a gradual change from non-metallic properties to metallic properties as one move down the group. They exhibit twooxidation state which are +2 and + 4 oxidation state. The +2 oxidation state increases while the +4 oxidation state decreases on going down the group.

2.5.5 GROUP 5 ELEMENTS

They are non-metals and show two common valences of 3 and 5. They react by gain ofelectrons. Their oxides combine with water to form acid.

2.5.6 GROUP 6 ELEMENTS

Element in group 6 are non-metals. There are electron acceptors and are oxidizing in nature. They have six valence electrons and require two electrons to reach the octet electron configuration

2.5.7 GROUP 7 ELEMENTS

The elements in Group 7 are known as the halogen which means salt former. They are the most reactive non-metals known. They are so reactive that they cannot exist free in nature. They react by gain of one electron.

2.5.8 GROUP 8 ELEMENTS

These elements are also called the noble gases. They are unreactive.

2.5.9 TRANSITION ELEMENTS/METALS

These elements occupy the space between the Group 2 and Group 3 elements on the fourth and fifth period. They are the element with an incomplete d-orbital.

Also we have the special transition metals on the sixth and seventh periods.

1. The lanthanides series; these are the fourteen elements after lanthanum with atomic number from 58-71 in the sixth periods. Members successively add one electron to the 4f sub shell.
2. The actinides series: These are the fourteen elements after actinium with atomic number from 90-103 in the seventh, periods. Members successively add one electron to the 5f sub shell.

2.6 PERIODIC PROPERTIES IN THE PERIODIC TABLE

The period properties in the periodic table are all related to the forces of attraction between the positive nucleus and the negative electrons. These properties include.

1. Ionization energy
2. Electronegativity
3. Electron affinity
4. Atomic radius
5. Ionic radius

2.6.1. Ionizations energy

This is the energy required to remove the most loosely held valence electron or outermost energy level electron of an atom in the gas phase, to form ions. The ionization energy increases across the period and decreases down the group. This is because as the group decend, the energy level increases successively, meaning an increased distanced between the positive nucleus and the outer energy level electrons. The resultant effect is reduced effective nuclear pull on the outermost energy level electrons, making the valence electrons easier to remove with reduced energy.

Ionization energy increases along a period as we go from left to right. Along the period, the energy level remains the same as electrons are added one after another. This causes shrinkage in size as the outer energy level electrons becomes more attracted to the positive nucleus, making it more difficult for electrons to be removed. If the valence electrons become more difficult to remove, it then means more energy is required (ionization energy) for electron removal. The table2.2 below summarizes changes in the first ionization energies.

2.6.2. ELECTRONEGATIVITY

This is the measure of the ability of an atom to attract electrons when it is in a molecule. In another word, when two atoms are covalently bonded, sharing a pair of electrons, the atom that has the ability to attract more of the shared electron pairs is the more electronegative.
Electronegativity increase across the period from left to right. This is because electrons are added to a constant energy level one after another, with no screening or shielding effect. This causes shrinkage in size as the valence electrons become more attracted to the positive nuclear pull, therefore increasing the ability of the atom to attract electrons. As an illustration, in a bond between carbon and oxygen, oxygen will be more able to attract the shared pair of electron carbon.
Electronegativity decreases down the group as repeatedly energy level of electrons is added. The inner energy levels then shield the outer energy level from the attractive pull of the positive nucleus, thus making it difficult for the atom to attract a shared pair of electrons.
Table 2.3 below summarizes the electronegativity values given on the Pauling's scale using a value of 4.0 for fluorine.

TABLE 2.2SUMMARY OF FIRST IONIZATION ENERGY DOWN THE GROUPS AND ACROSS PERIODS

H 1.3							He 2.4
Li 0.5	Be 0.9	B 0.8	C 1.1	N 1.4	O 1.5	F 1.7	Ne 2.2
Na 0.5	Mg 0.7	Al 0.6	Si 0.8	P 1.0	S 1.0	Cl 1.3	Ar 1.5
K 0.4	Ca 0.6						

TABLE 2.3SUMMARIES OF ELECTRONEGATIVITY TRENDS, DOWN THE GROUP AND ACROSS THE PERIOD

H 2.2							He
Li 1.0	Be 1.5	B 2.0	C 2.6	N 3.1	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2	Al 1.5	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar
K 0.8	Ca 1.0						

Note that electronegativity values are not assigned to the noble gases because they commonly donot share electrons

2.6.3. ELECTRON AFFINITY

This is the energy released when an electron is added to a neutral atom in gaseous state to form gaseous ion.

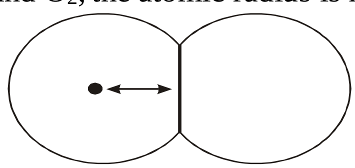
Electron affinity decreases down the group as it become more difficult for electrons to add to an atom. So, less energy will be released in the cause of the reaction as more energy will be used up.

Electron affinity increases with increase in atomic number along the period as the ease for an electron to add to a neutral atom becomes less and less difficult. Therefore using less energy for the reaction and releasing more energy.

2.3.4 ATOMIC RADIUS

An atom does not have a definite diameter; radius, **circumference** or volume because of its wave-like nature. Therefore, its size expressed as its radius is done in three different ways, all in unit of nanometer (nm)

i. Covalent atomic radius: This is the distance from the nucleus to the valence shell when the sharing of electrons covalently bonds the atom. When the covalent bond involves identical atoms e.g. $H_2$and O_2, the atomic radius is half the distance between the two nuclei



ii. Vander Waals radius: This is half the distance between two nuclei when measured from their closest point, when no bond is formed.

iii. Atomic radius of metals: This is the distance between nuclei of atom when arranged in a metal-like crystalline structure.

Atomic radius decrease across a period with increasing atomic number from left to right while it increases down the group. Table2.4 below give a summary of the trend in atomic radius.

TABLE 2.4ATOMIC RADIUS TREND DOWN THE GROUP AND ACROSS THE PERIOD

H 0.030							He
Li 0.123	Be 0.089	B 0.088	C 0.077	N 0.070	O 0.064	F 0.064	Ne
Na 0.157	Mg 0.136	Al 0.125	Si 0.117	P 0.110	S 0.104	Cl 0.099	Ar
K 0.203	Ca 0.174						

Atomic radius values are not given to the noble gases because they do not form covalent bonds.

2.3.5. IONIC RADIUS

Just like the atoms, ions also have a wave like property which makes their size difficult to ascertain. Ions are formed when atom gain or lose electron(s). The ionic radius is the effective distance between the nucleus of the ions and its outer energy level. Usually, ions have eight electrons in the outer energy level.

Metals react by the loss of all their valence electrons, forming positive ions called cation since all the outer energy level electrons are lost the energy level is lost as well, causing the radius of the ion to become smaller that the initial atom. Also, greater positive change exerts a greater attractive force on the electrons, thus causing shrinkage in size

Non-metals on the other hand, react by gain of electrons forming negative ions called anion. The radius of the ion formed becomes bigger than the initial atom. This is because the electron gained fill up the energy level, resulting to a greater negative charge. This reduces the attractive pull of the positive nucleus on the outer energy level electrons. Ionic radius just like atomic radius increases down the group. But it reduces across the metallic period while it increases across the non-metallic period.

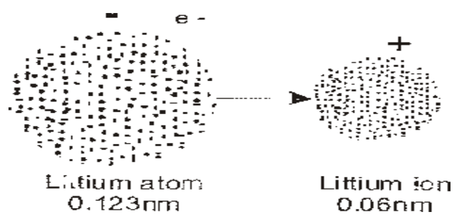


TABLE 2.5 SUMMARY OF THE IONIC RADIUS TREND DOWN THE GROUP AND ACROSS THE PERIOD.

H 0.021							He
Li 0.060	Be 0.031	B 0.020	C 0.015	N 0.017	O 0.140	F 0.136	Ne
Na 0.095	Mg 0.065	Al 0.050	Si 0.041	P 0.212	S 0.184	Cl 0.181	Ar
K 0.133	Ca 0.099						

Reduces **Increases**

Generally,

1. The ionic radii of positive ions are smaller than their corresponding atomic radii because increase in positive charge will have a greater nuclear attractive pull on the valence electrons, causing shrinkage in size.
2 The ionic radii of negative ions are bigger than their corresponding atomic radii because increase in negative charge will have a reduced nuclear attractive pull on the valence electrons. Also the increased electrons will have repelling effect on each other causing

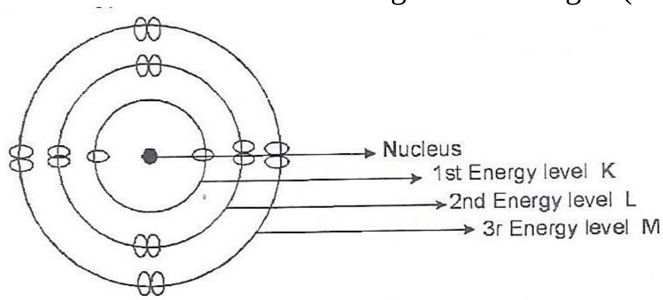
increase in size.

3 For Isoelectronic series (ions and atoms with the same number of electrons), ionic radius decreases as atomic number increase.

ELECTRONIC CONFIGURATION

The electron configuration of an atom is the representation of the arrangement of electrons distributed among the orbital shells and subshells in its ground state. The physical and chemical properties of elements are related to their unique electronic configurations. The valence electrons (electrons in the outermost shell) are the determining factor for the unique chemistry of the element. Electrons fill up the orbital with lower energy level first before occupying orbitals of higher energy level. Also, the first orbital or shell will accommodate a maximum of two electrons after which subsequent orbitals would accommodate a maximum of eight electrons each. Figure 2.1 shows the electronic configuration of argon with 18 electrons However, because of some shortcomings of the Bohr‟s model, the quantum mechanical model is now more valid. The will be treated in the next chapter.

Example 2.1. Do the electronic configuration of argon (atomic number =18).



PRACTICE QUESTIONS

1 a. How is Mendeleev periodic law different from the modern periodic law

b. Mention three information you can get from the periodic table

2. Define the follow and explain their trend down the group and across the period in the periodic table;

i. Electronegativity
ii. Atomic radius
iii. lonisation energy

3Which noble gas and halogen belong to the third period?

4 List the following atoms in order of increasing electronegativity; K, Al, S, CI, and F

5. List the following in order of reactivity Na, K, Be, and Ca.

6. In the following bonded pairs of atoms, state in each case the atom that attracts more of bond to itself (a) H: Br (b) Na:CI (c) Mg and Mg^2+ (d) CI and CI- (e) K and Ca

7. Identify the atom or ion with the larger radius in each of the following pair

(a) S and O

(b) H and Li

(c) Mg and Mg^2

(d) Cl and Cl^{-+}

(e) K and Ca

8. Place the following isoelectronic species in order of increasing atomic radii and explain. Na+, $Mg^{2+,}$ Al^{3+}, $Si^{4+}N^{3-}$, $0^{2-.}$ F^-, Ne

9. Where in the periodic table of the first 20 elements would you expect to find;

(a) The most reactive metals

(b). The most reactive nonmetals

(c). The least reactive metals

(d) The least reactive nonmetals