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CHEM. DEPT.
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SARGODHA

CHAPTER NO:- 1 II YEAR.

PERIODIC CLASSIFICATION.

The systematic arrangement of elements into groups and periods is called PERIODIC CLASSIFICATION. previously this arrangement was based upon atomic weights of elements. It was considered that atoms of different elements have different atomic weights and atoms of same element have same atomic weights (Dalton's postulate). Later on it was proved to be incorrect. Now a days this arrangement is based upon atomic number of elements and their electronic configuration. The elements with increasing atomic numbers are placed in periods and the elements with similar configuration and similar properties have been placed in same groups.

DISCUSS HISTORY OF PERIODIC CLASSIFICATION.

DOBEREINER'S TRIADS :-

The first systematic attempt was made by Dobereiner in 1817. He made some groups of elements. Each group was called "TRIAD", containing three elements each. The three elements present in each group had similar properties and weight of middle element was arithmetic mean of other two.

Let us consider following examples.

Elements	Atomic wts	DISCUSSION
Li	7	The properties of "Li", "Na" and "K" are similar and weight of middle element "Na" is arithmatic mean of "Li" and "K".
Na	$\frac{7+39}{2} = 23$	
K	39	
Cl	35.5	The properties of "Cl", "Br" and "I" are similar and weight of middle element "Br" is arithmetic mean of other two.
Br	$\frac{35.5+127}{2} =$	
I	127	

DEFECTS:- There were few elements which obeyed this rule. Most of the elements cannot be arranged in triads.

Further some atomic weights assigned by Dobereiner were incorrect. However credit goes to Dobereiner because he made a good attempt to classify elements.

NEWLAND LAW OF OCTAVES :- Newland compared physical and chemical properties of elements with notes of music and proposed law of octives which states that: "If elements are arranged in order of increasing atomic weights, every ~~eat~~ eighth element has its properties similar to first one, just like notes of music. Second element will have its properties similar to 9th element, and so on.

Let us consider following elements.

Li Be B C N O F

Na Mg Al Si P S Cl

DEFECTS :- All elements could not be arranged on the basis of this law. It is due to the fact that properties of elements mainly depend upon atomic number and NOT on atomic weights. The law became inapplicable with the discovery of noble gases which occupied viii group. Thus second element now resembles 10^{th} element

MEYER'S ATOMIC VOLUME CURVE :- Lother Meyers plotted the atomic volumes of elements against their (arranged) atomic masses and found that identical elements occupied similar positions on the curves. He arranged elements in order of increasing atomic weights and found that physical and chemical properties are periodic function of atomic weights.

MENDELEEF'S PERIODIC CLASSIFICATION: In 1869 mendeleevs discovered a rule for periodic classification which is known as mendeleev's periodic law. It states that: "The physical and chemical properties of elements are periodic function of their atomic weights. If elements are arranged in order of increasing atomic weights, physical and chemical properties reoccur periodically."

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"Mendeleev's found that if elements arranged with increasing atomic weights, physical and chemical properties repeat after regular intervals these intervals may be small or large. On the basis of this law he arranged elements in such a way that similar elements were falling in the same verticle line called a "group". Mendeleev's periodic table was precise, lucid and formed basis of modern classification of elements.

ADVANTAGES OF MENDELEEV'S PERIODIC TABLE

PREDICTION OF NEW ELEMENTS:- Mendeleev left many vacant spaces in periodic table for elements which were not discovered by that time. For example spaces were left for SCANDIUM, GALLIUM, & GERMANIUM which were not discovered by that time. Later on these elements were discovered and occupied same positions.

PREDICTION OF PROPERTIES OF NEW ELEMENTS:- Mendeleev was not only able to predict new elements but he also predicted properties of these elements before hand. For example GALLIUM was discovered in 1886 but Mendeleev proposed its properties in 1871.

CORRECTION OF ATOMIC WEIGHTS OF ELEMENTS

The arrangement of elements helped to correct atomic weights of elements. For example on the basis of similarities in properties, Mendeleev placed "Be" in second group along with Mg, Ca, Sr and Ba. Previously "Be" had been assigned atomic weight 13.5 and it should have been placed after Carbon. But Mendeleev assigned it atomic weight "9" which justified its position with 2nd Group elements.

The atomic weight of Indium was calculated to be 75.80 but later on corrected as 114.82 by Mendeleev with the help of its position in periodic table.

SYSTEMATIC STUDY:- Mendeleev's periodic table helped a lot for systematic study of elements. By studying the properties of one element in a group, one was able to predict properties of other elements in that group. For example study of properties of Sodium helped chemists to predict properties of other alkali metals.

DEFECTS IN MENDELEEV'S PERIODIC TABLE:-

Mendeleev periodic table was not free of defects. A few defects have been discussed below.

MISFIT PAIRS OF ELEMENTS:- In mendeleevs periodic table elements were arranged in increasing order of atomic wts. For example Ni (59.6) should have been placed before "Co" 59.9 but it was not so in Mendeleevs periodic table. Similarly iodine should preceede "Te"

Argon (40) placed before K (39).

STRUCTURE OF ATOM:- Mendeleevs periodic table gave no idea about structure of atom & electronic configuration.

POSITION OF ISOTOPES:- Isotopes are atoms of same element with different atomic weights. Mendeleev periodic table was unable to ~~justify~~ specify the position of isotopes in periodic table.

POSITION OF COINAGE METALS:- Coinage metals ("Cu, Ag" and "Au") which differ widely from alkali metals, were placed in first group with alkali metals (Li, Na, K, Rb, Cs, Fr). However later on this defect was removed by division of groups into subgroups.

POSITION OF HYDROGEN:- Mendeleevs periodic table was unable to specify position of hydrogen in periodic table.

POSITION OF LANTHANIDES & ACTINIDES:- Lanthanides & Actinides have been placed in periodic table which goes against the periodic ~~table~~ law.

Inspite of its defect, mendeleevs periodic table provides basis of modern classification of elements.

WHAT IS MODERN PERIODIC LAW:- Mosely in 1913 showed that frequency of radiations emitted (X-rays) by bombardment of high energy electrons on target element, depend upon number of positive charges on nucleus of target element. This number of positive charges was called ATOMIC NUMBER, denoted by "Z".

The atomic number was defined as: "The number of protons present in nucleus of atom or number of electrons present outside the nucleus".

MODERN PERIODIC ~~TABLE~~ LAW:- The modern periodic law states that: "The physical and chemical properties of elements are periodic functions of atomic numbers of elements. Thus when elements are arranged in order of increasing or decreasing atomic number, their properties repeat after regular intervals.

MODERN PERIODIC TABLE In modern periodic table recommended by I.U.P.A.C. the elements are arranged in ~~order~~ ascending order of atomic number. The atomic number increases by one unit each step towards right side in a period. The similar elements fall in the same vertical line called Groups. This arrangement according to atomic number removed many defects in Mendaleev's periodic table.

REMOVAL OF MISFITS:-

(i) Argon with atomic weight "40" is placed before Potassium atomic wt "39" because Argon has atomic number "18" and Potassium has atomic number 19.

(ii) Cobalt (At. wt = 58.9) is placed before Nickal (At. wt = 58.6) because Cobalt has atomic number "27" while nickal has atomic number 28.

(iii) Tellurium with atomic number 52 is placed before iodine with atomic number 53.

POSITION OF ISOTOPES: Isotopes have same atomic number but different Atomic wts. The modern periodic table is based upon ~~atom~~ atomic number. Thus each element is represented only once in periodic table.

DISCUSS GROUPS AND PERIODS IN THE MODERN (LONG FORM OF) PERIODIC TABLE:-

The modern periodic table has been formed on the basis of modern periodic law. Which states that. "The physical and chemical properties of elements are periodic function of the atomic number of element. If elements are arranged in order of increasing atomic numbers properties repeat after regular intervals. In the modern periodic table elements has been arranged in order of their increasing atomic numbers.

There are two types of rows in the periodic table. The vertical rows are called GROUPS or COLUMNS. The horizontal rows are called PERIODS

PERIODS:- There are seven periods in the modern periodic table. Their details are given below.

FIRST PERIOD:-

It is very short period of periodic table containing only two elements viz Hydrogen and Helium. with respective atomic numbers "1" and "2". Both these elements has one shell around nucleus. Hydrogen has electronic config $H = 1s^1$ and $He = 1s^2$

There is a large gap of 16-elements between two members of first period. which may be considered as a ~~gap~~ ~~but~~ defect in modern periodic table.

SECOND AND THIRD PERIOD:- The second and third periods have eight elements each. These periods start from an alkali metal and end with a noble gas. These are called SHORT PERIODS.

The second period starts from Lithium (Atomic Number = 3) and end at Neon (Atomic No. 10) The third period start with "Na" (At. No. 11) and end at Argon (At No. 18)

Two of the eight elements present in second and third periods are "s" block elements and

remaining six are p-block elements.
"s"-Block elements have outermost electron in "s" orbital and p-block elements have outermost electron in p-orbital. These are also called Representative elements. The properties of elements of 2nd and 3rd period are almost identical.

FOURTH & FIFTH PERIODS:- (LONG PERIODS).

Fourth and fifth periods have 18 elements each. Two of these 18 are "s" block elements, "6" p-block element and 10 are "d" block elements $\left(\begin{matrix} 2 & + 6 & + 10 & = 18 \\ s & + p & + d & \end{matrix}\right)$

"d" block elements are called TRANSITION ELEMENTS. The fourth period includes first transition series ($_{21}Sc$ to $_{30}Zn$). These elements have "3d" orbital in process of completion. The fifth period includes 2nd transition series ("$_{39}Y$ to Cd"); these have "4d" orbital incomplete. These periods start from alkali metal and end with a noble gas. The metallic character decreases as we proceede from left to right. In these periods properties repeat after 18 elements.

SIXTH PERIOD:- (VERY LONG PERIOD)

It is longest period and contains 32 elements. These are, 2 "s" block elements, 6 "p" block elements, 10 "d" block elements and 14 "f" block elements.
$\left(\begin{matrix} 2 & + 6 & + 10 & + 14 & = 32 \\ s & p & d & f & \end{matrix}\right)$ This period starts from "Cs" and end at Radon (Rn). The 14 "f" block elements are also called "LANTHANIDES" (following Lanthanum). Their "4f" orbital is in process of completion. These are placed at the bottom of periodic table.

SEVENTH PERIOD :- INCOMPLETE PERIOD.

It is an incomplete period starting from "Francium"

It includes :-
Two "s" block elements (Francium & Radium).
Fourteen "f" block elements (Actinium to Lawrencium)
six ~~one~~ "d" block elements (from Unilquadium (104) to Unilennium). The fourteen "f" block elements are called ACTINIDES (from $_{90}Th$ to $_{103}Lw$). The elements beyond Uranium (92) are called Transuranium elements. Actinides have "5f" ~~ds~~ orbital in process of completion.

GROUPS IN PERIODIC TABLE :-

The Vertical columns in periodic Table are called GROUPS. The members of long periods are further divided into two subgroups "A" and "B" subgroups. The normal elements are placed in subgroup "A" and transition elements are placed in subgroups "B". The alkali metals belong to "IA" group. They include "Li, Na, K, Rb, Cs, Fr". The alkaline earth metals are placed in II A group. They include Be, Mg, Ca, Sr, Ba, Ra.
Halogens constitute $\underline{VII}$ A group (F, Cl, Br, I, At).
~~Can~~ Noble gases are placed in $\underline{VIII}$ A group. They include "He", "Ne", "Ar", "Kr", "Xn", "Rn".
The Coinage metals have been placed in I B group. These are "Cu", "Ag" and "Au".
The elements placed in same group have same number of electrons in outermost ~~orbital~~. Their group number ~~of~~ indicates number of electrons in outermost shell. For example all elements of IA group have one electron in their outermost shell. The elements of same group have same chemical properties. However there is a gradual change in properties down a group.

BLOCKS IN PERIODIC TABLE.

Elements of periodic table can be classified into four blocks.

S - BLOCK ELEMENTS.
p - BLOCK ELEMENTS.
d - BLOCK ELEMENTS.
f - BLOCK ELEMENTS.

S-BLOCK ELEMENTS.

S- block elements have their valence elections in S- orbitals. Elements of I-A and II-A sub groups are called S- block elements. Elements of I-A have ns^1 elections and II-A group elements have ns^2 elections in valence shell.

p-BLOCK ELEMENTS.

p- block elements have their valence elections in p- orbitals. Elements of III-A to VIII-A sub- groups are known as p- Block elements.

d-BLOCK ELEMENTS.

These include transition elements. These elements have valence electrons in d- orbitals.

f-BLOCK ELEMENTS.

f - block elements involve LANTHANIDE and ACTINIDES. these elements have valence elections in f- orbital.

METALS, NON METALS AND METALOIDS.

Elements of periodic table are also classified on the basis of their metallic chracter. these are divided into following classes

a - METALS.
b - NONMETALS.
c - METALLOIDS.

METALS:-

Elements having tendency to form positive ions are called metals.
elements of S-block, d- block and some lower elements of p- block are metals.

1 - These have tendency to loose elections
2 - These are good conductor of heat and electricity
3 - Their oxides are basic in nature e.g.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

NON METALS :-

Elements which gain electron and form negative ions are called NON-METALS. All the gases are non-metals.

These are usually poor conductors

They form acidic oxides which yield acid on dissolving in water.

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

METALLOIDS.

Some elements have properties of both metals and non-metals. These elements are called METALLOIDS.

Their oxides have characteristics of both acidic and basic oxides.

lower members of groups, III-A, IV-A and V-A are called metalloids.

POSITION OF HYDROGEN.

DISCUSS POSITION OF HYDROGEN IN PERIODIC TABLE.

Ans. Hydrogen can be placed above alkali metals in group I, above halogens in group VII or above carbon family in group IV. This position is given to it on the basis of its electronic configuration, physical and chemical properties and thermodynamic data. But if we go into details hydrogen seems to be misfit in above three positions.

(1) HYDROGEN ABOVE ALKALI METALS.

Hydrogen is mostly placed above alkali metals group I due to following properties similar to alkali metals.

1 ELECTRONIC CONIGURATION) Hydrogen has electronic configuration $1s^1$ which is similar to alkali metals. All alkali metals have one electron in outermost shell, and hydrogen also has one electron in outermost shell. So hydrogen should be placed over alkali metals in group I.

2 UNIPOSITIVE ION) Alkali metals lose one electron to form unipositive ions(M^+). Hydrogen also loses one electron to form unipositive ions "H^+"

FORMATION OF HALIDES) Alkali metals combine with halogens to form halides with general formula "MX" where "M" has "1+" oxidation state. Hydrogen also combines with halogens to form hydrogen halides "HX" where Hydrogen has oxidation state "1+".

ELECTROLYSIS OF HALIDES) During electrolysis of NaCl sodium metal is obtained on cathode. Similarly during electrolysis of HCl, hydrogen is obtained on cathode.

DIFFERENCES FROM ALKALI METALS

PHYSICAL STATE:- Hydrogen is a gas while alkali metals are solids are room temperature.

2. DIATOMIC MOLECULE Hydrogen exists in the form of diatomic molecules but alkali metals donot form diatomic molecules.

3 VALENCE SHELL} Hydrogen requires one electron to complete its shell (1ST SHELL) while alkali metals require 7 electrons to complete their shell.

4. ELECTRON TRASFER} In hydrogen halides electron is partially shifted from hydrogen to halogen but in alkali metals halides electron is completely transfered from metal to halogen.

5 STABILITY OF UNIPOSITIVE ION}- $H^{\oplus}$ ion cannot exist as such but it gets solvated to form H_3O^+ or $H_9O_4^{\oplus}$ ions. While alkali metal ions can exist independently as Na^+, K^+, $Li^{\oplus}$ etc.

Keeping in view these differences hydrogen should not be placed above alkali metals.

HYDROGEN ABOVE HALOGEN.

Hydrogen may also be placed above halogens in group VII This is due to following similarities of hydrogen with halogens.

DIATOMIC MOLECULES Both hydrogen and halogens form diatomic molecules like H_2, Br_2, Cl_2 etc.

REQUIRE ONE ELECTRON Both hydrogen and halogens require one electron to complete their outermost shell.

FORMATION OF HYDRIDES} Hydrogen combines with alkali metals to form alkali metal hydrides, similar to alkali metal halides, NaH, NaCl etc

STRUCTURE OF HYDRIDES: Alkali metal hydrides and alkali metals halides have similar structure. For example both NaH and NaCl have cubic structure. Thus hydrogen is similar to halogens. Keeping in view these similarities hydrogen should have ~~been~~ placed above halogens in group VII.

DIFFERENCES FROM HALOGENS

(1) $H^{\ominus}$ ION IS NOT STABLE: $H^{\ominus}$ ion is not stable in aqueous solution while $X^{\ominus}$ is stable in aqueous solution.

$$H^{\ominus} + H_2O \longrightarrow H_2 + OH^{\ominus}$$

(2) ELECTRONIC CONFIGURATION: Hydrogen has one electron in its "s" orbital while halogens have seven electrons in outermost shell ns^2np^5 configuration.

(3) BASIC HYDRIDES: Alkali metals hydrides are strongly basic while alkali metal ~~hybrid~~ halides are neutral.

$$NaH + H_2O \longrightarrow NaOH + H_2$$

(4) HYDROGEN IS A GAS: Some halogens like Br_2 are liquid and Iodine is solid while hydrogen is always in gaseous form.

So hydrogen placed over halogens in group VII is not fully justified.

HYDROGEN ABOVE IV GROUP

(5) Hydrogen may also be placed over "C" in group IV due to following similarities.

(1) HALF FILLED VALENCE SHELL: The valance shell of both hydrogen and carbon family is half filled.

(2) ORGANIC COMPOUNDS: In organic compounds hydrogen and "C" exist in close association & form covalent bonds.

(3) THERMODYNAMIC DATA: I.P, E.N, of hydrogen are similar to IV group elements. Thus hydrogen also forms covalent bonds.

BOTH HYDROGEN AND CARBON ARE VERY GOOD REDUCING AGENTS $CuO + H_2 \longrightarrow H_2O + Cu$

$CuO + C \longrightarrow CO + Cu$

Keeping in view these properties hydrogen should be placed above carbon in IV group

DISSIMILARITIES FROM FOURTH GROUP

The major point of difference between hydrogen and other members of IV group is that hydrogen is a gas while other members of IV group are solids at S.T.P.

Further hydrogen is monovalent while other members of fourth group are tetravalent mostly.

Hydrogen has only one valence electron while a IV group elements have four electrons in outer

PERIODIC TRENDS IN PHYSICAL PROPERTIES.

Repetition of physical and chemical properties after regular intervals is called periodicity. According to "modern periodic law" The physical and chemical properties of elements are periodic function of atomic numbers. i.e if elements are arranged in order of increasing atomic numbers, the physical and chemical properties are repeated after regular intervals.

Following physical properties show periodicity.

1. ATOMIC SIZE :-

a. ATOMIC RADIUS:- Atoms are considered as spherical bodies their sizes are generally expressed in terms of atomic radii. atomic radii of atoms are determined when they are linked by some sort of bond.

PERIODICITY:- The atomic radius increase down the group due to increase in atomic number. down the group in each step an extra shell of electron is added. due to this extra shell in each step atomic radius also increases.

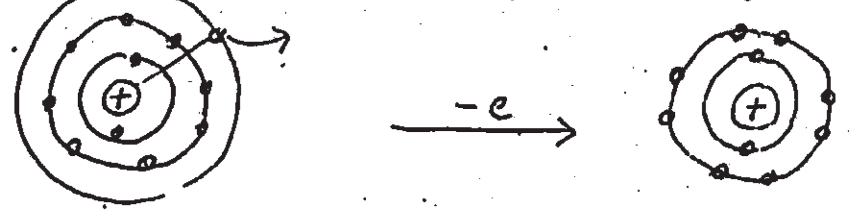
On the other hand atomic radius decrease along the period. As we move from left to right in the period atomic radius goes on decreasing. In a period number of shells remains constant but nuclear charge gradually increases (atomic No. increases) left to right. Thus valance electrons are attracted more powerfully as we procede left to right in a period. It results in decrease in atomic radius.

b. IONIC RADIUS.

An ion is formed by gain or loss of electron from an atom. When an atom looses electron it is converted into positive ion and when an atom

~~Neixiodixity~~:- gains electron it becomes negative ion. Radius of positive ion is always less than the corresponding atom. This may be due to

i- The decrease in number of shells when an electron is removed from an atom to produce positive ion. For example consider formation of Na^+ ion from Na atom

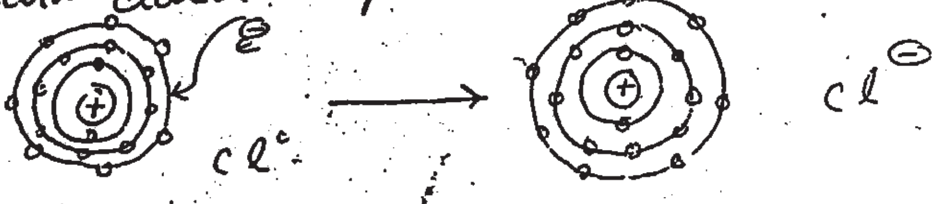


$1s^2 2s^2 2p^6 3s^1$ $\qquad$ $1s^2 2s^2 2p^6$.

ii - Some times number of shells remains the same after removal of electron from an atom even then the size is decreased. It is due to the fact that the same nucleus charge now acts on smaller number of electrons. Thus each electron is attracted more powerfully in a positive ion. Thus size is sequeezed.

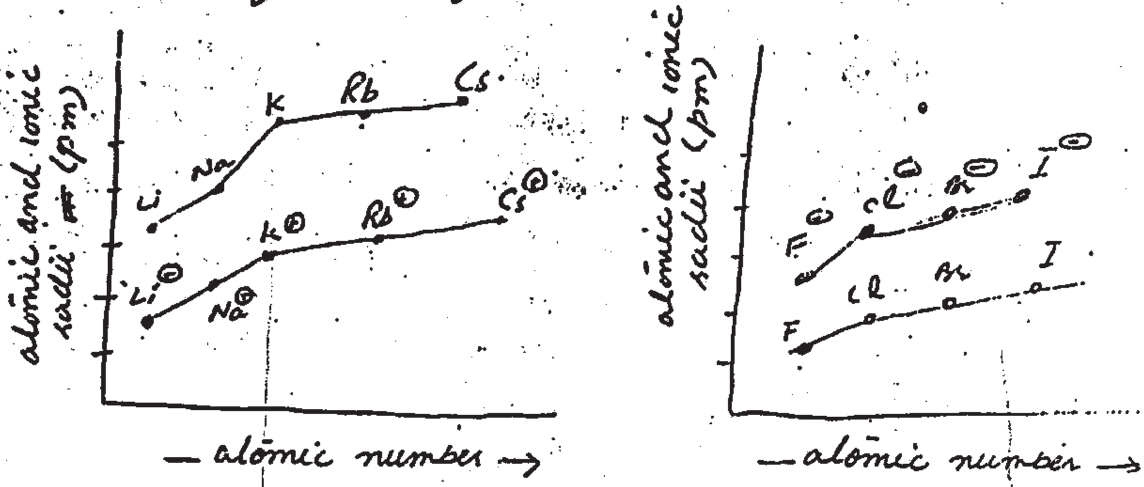
Al $\qquad$ $Al^{\oplus}$.

The NEGATIVE ion is larger than corresponding atom. This is due to the fact that number of electrons are increased. Thus the same nuclear charge now acts on more number of electrons. Thus each electron is less tightly bounded by nucleus. Therefore ionic radii of -ve ion is larger than atomic radii of corresponding atom. Thus $Cl^{\ominus}$ ION HAS LARGER SIZE THAN Cl°. In Cl° there are 17 electrons around nucleus and 17 protons inside the nucleus. Thus each electron is attracted towards the nucleus by a force of 1 Proton. But in $Cl^{\ominus}$ there are 18 electrons being attracted by 17 proton. Thus each electron is less powerfully attracted by the nucleus. Consequently electron cloud expands a little.



Cl° $\qquad$ $Cl^{\ominus}$

In a group, similar charged ions increase in size from top to bottom. while in a period size decrease from left to right.
Variations in atomic and ionic radii of alkali metals and halogens as follows.

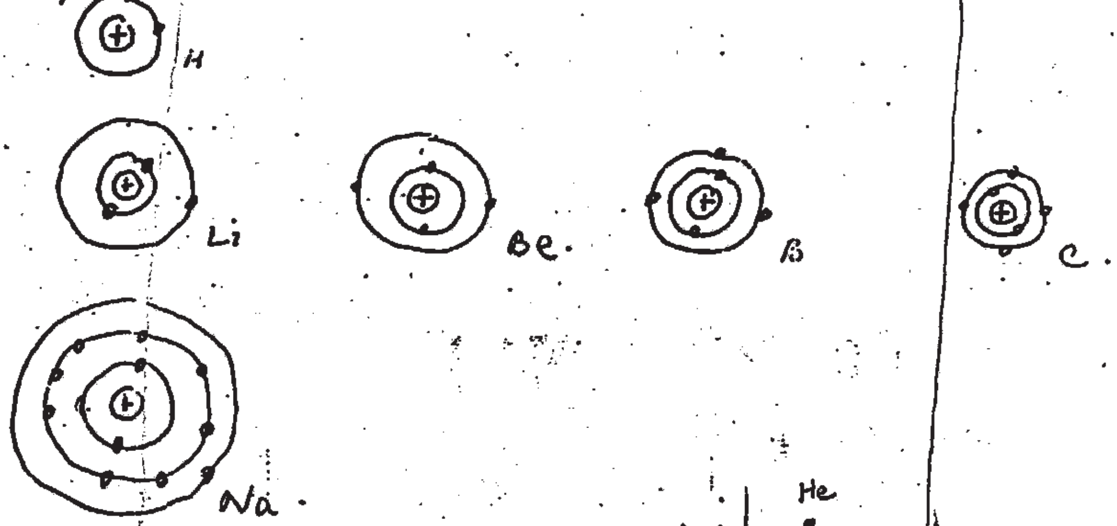


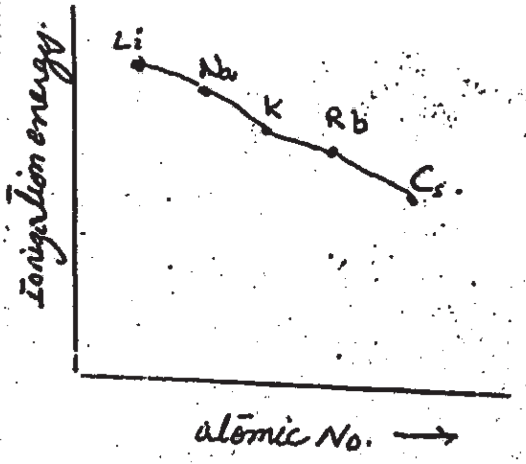
OXIDATION STATE:- The charge which an atom have in the compound is called its oxidation state.
In ionic compounds the number of electrons gained or lost by an atom is called its oxidation state.
For example in case of NaCl oxidation states of Na is +1 and of Cl is −1.
Oxidation state of an element may be the same as its group number. e.g. elements of group I-A to IV-A have same oxidation state as their group number.
Oxidation state ~~may also be~~ is related to No. of valence electrons. e.g P, N, As and Sb show +3 as well as +5 oxidation states.
Similarly in H_2SO_4 sulphur shows oxidation state +6 while in H_2S it is −2. In these elements vacancies in the shell is also related to oxidation state. 2 is the No. of vacancies in the shell.
ELEMENT ALWAYS HAVE ZERO OXIDATION STATE. Elements of VIII-A group have zero oxidation state because there is no vacancies in their valence shell.
Transition elements show oxidation state equal to their group number. But due to greater no. of valance electrons these have variable ox. states

2. IONIZATION ENERGY. It is defined as minimum amount of energy required to remove most loosely bound electron from an atom in gaseous state to produce a positive ion is called IONIZATION POTENTIAL. expressed in KJ/mole.

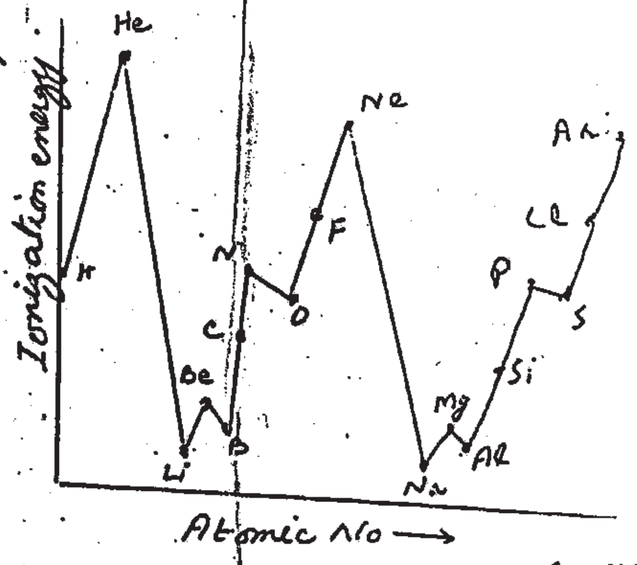
Ionization potential decreases down the group and increase as we proceede from left to right in a period. As we proceede down the group shells are added round and round the nucleas. It increases atomic size and sheilding effect. Both these factors decrease ionization potential.

In a period number of shells around nucleus remain constant, while atomic number ~~decrea~~ increases by one unit each step toward right. Thus atomic size decreases left to right and ionization potential increases.





IONIZATION ENERGY ALONG THE GROUP.



IONIZATION ENERGIES ALONG THE PERIODS

METALLIC CHRACTER.

Elements can be divided into metals non metals and metalloids.

METALS:-

Metallic chracter increases from top to bottom in a group because down the group atomic size increases. So it make removal of electron quite easy. On the other hand metallic chracter decreases as we move from left to right in a period. So elements of I-A and II-A are considered as pure metals. Some elements in the centre and at the bottom of periodic table are also have metallic chracter. Metals have tendency to loose elections. They form basic oxides.

Elements of group VII-A are least metallic in nature.

NON METALS:-

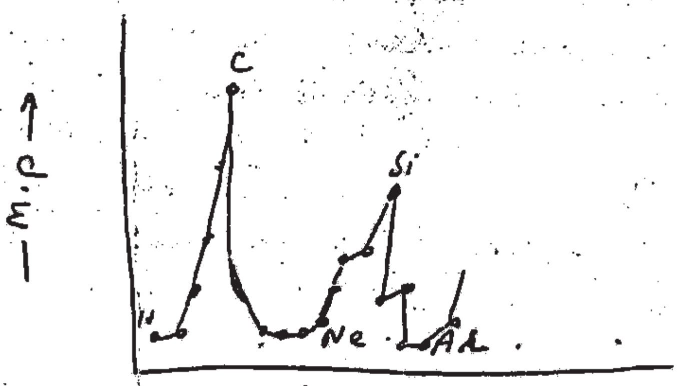
Non metallic chracter is the tendency to gain elections. It decreases as the size of element increases. So elements of VII-A are non-metals. In halogens non metallic chracter also decreases due to increase in atomic size. Fluorine is the most non metallic. This trend can also be found in VA and VI-A groups. These are pure non-metals and exist in gaseous states. While elements at the bottom of these groups are fairly metallic due to the greater size.

MELTING AND BOILING POINTS:-

Melting and boiling points of elements decreases down the group for I-A and II-A groups while increase down the group for VII-A elements. While as we proceed from left to right in a period M.P and B.P increase upto group IV-A and then goes on decreasing upto nobel gases.

ALONG THE PERIODS: M.P and B.P goes on increasing upto carbon family. M.P of I-A elements are lowest. It is due to the fact that-

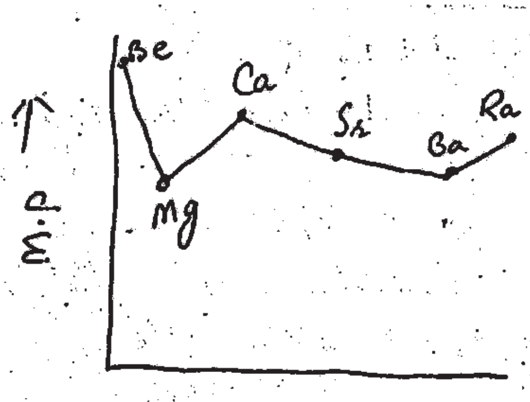
There have only one electron to form bond. M.P of II-A are higher than I-A due to greater No. of bonding electrons. Carbon having maximum No of binding electrons and so it has high M.P. Melting point and Boiling point decrease from IV-A → VII A. These elements exist as small molecules.

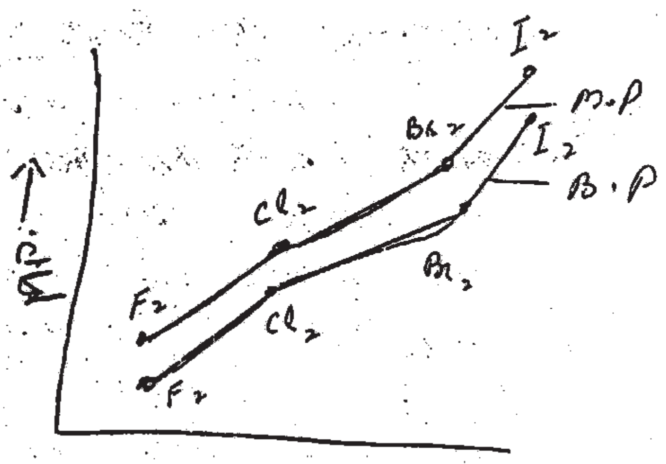


— Atomic No. ⟶

ALONG GROUP.

M.P and B.P of I-A and II-A decrease down the group. Because atomic size increase down the group. Also there is weak binding force b/w the molecules with larger sized atoms than the smaller one. For elements of VII-A M.P and B.P increase down the group. This is because larger molecules have higher polarizability and hence these exert strong force of attraction due to higher polarizability.





II-A group.

VII-A.

ELECTRICAL CONDUCTANCE.

Electrical conductance depend upon No. of free or moveable electrons. Metals can conduct electricity due to presence of lose electrons in valence shell. Electrical conductance of metals in group I-A and II-A decreases from top to bottom. Metals of I-B have very high values of electrical conductance. Non metals on the other hand show low electrical conductance that they can be considered as non-conductors. (V-A and VII-A).

In transition metals no general trend can be assigned.

Carbon, in the form of diamond is non conductor because valence electrons are tetrahedrally bound and canot move freely. while in the form of GRAPHITE it is good conductor. because 1 electron is free to move.

Lower elements of IV are good conductors. Their values are comparable with their counter parts in I-A group.

HYDRATION ENERGY.

It can be defined as.

"The heat absorbed or evolved when one mole of gaseous ions dissolves in water to give an infinite dilute solution: e.g. When one mole of hydrogen ions in gaseous phase are dissolved in water, for infinite dilution, a large amount of heat is liberated.

$$H^{\oplus}_{(g)} + \text{Water} \longrightarrow H^{\oplus}_{(aq)} \qquad \Delta H_{(h)} = -1075\ KJ/mole$$

HYDRATION VALUES OF FEW IONS ARE SHOWN AS.

ION	H_h VALUE ($KJ\ mole^{-1}$)
$Li^{\oplus}$	-499
Na^{+}	-390
Mg^{2+}	-1562
Cl^{1-}	-384
Br^{1-}	-351

Heat of hydration depend upon charge to size ratio of the ions. As charge to size ratio decreases heat of hydration also decreases.

Heat of hydration decreases down the group but goes on increasing as we move from left to right in the period.

It is due to the fact that down the group charge to size ratio decreases so H_h decrease. But along the period H_h increases due to increase in charge to size ratio.

PERIODIC RELATIONSHIP IN COMPOUNDS.

HALIDES:-

"Halides are the binary compounds of Halogens with other elements."

On the basis of nature of bonding halide are classified into COVALENT HALIDE and
IONIC HALIDE.

There is another class b/w the two which acts as a bridge is POLYMERIC HALIDES.

Ionic Halides. are formed by complete transference of $\bar{e}$ from other atom to Halogen. Strongly electro positive elements form ionic halides. These elements have greater electronegativity difference with Halogens.

Halides of group I-A are considered pure ionic halides. The ionic halides are of high M.P solids. These form three dimensional lattice consisting of discrete ions.

Among the pure ionic compound, fluorides have highest lattice energy due to smaller size. It has highest M.P and Boiling points. Which decrease in the order

FLUORIDE > CHLORIDE > BROMIDE > IODIDE.

POLYMERIC Halides.

Halides having partially ionic character are called polymeric Halides. element with less electropositivity form polymeric Halides.

e.g. Be, Ga, Al.

COVALENT HALIDES.

Covalent halides are formed by mutual sharing of electrons b/w Halogen and other atom. On moving from left to right in periodic table electronegativity difference reduces. and trend shifts towards covalent halide.

These halides are usually gases, liquids or these ~~are~~ are low m.p solids. it is due to the fact. Intramolecular forces in covalent halides are weak. Physical properties of these halides depend upon ~~the~~ size and polarizability of the halogen atoms. Iodide being the largest and more polarizable ion possess the strongest van der Waals forces. and thus ~~it~~ have higher M.P and B.P than those of other covalent halides.

Bond character of chlorides of period 3.

Property	NaCl.	$AlCl_3$	PCl_3	S_2Cl_2.
M.P.	808.	195	-92	-80.
BONDING	IONIC	PARTLY IONIC.	POLAR COVALENT	POLAR COVALENT

Bond character also vary as we move from top to bottom in halogens.

Order of decreasing ionic character of halide is

fluoride > Chloride > Bromide > Iodide

e.g. AlF_3 is purly ionic but AlI_3 is ~~somewhat~~ predominently covalent halide.

If an element form more than one Halides. the metal halide in its lower oxidation state tends to be ionic while higher is covalent.

e.g

$PbCl_2$ IONIC. (Pb^{2+})

$PbCl_4$ COVALENT (Pb^{4+}).

HYDRIDES :-

Binary compounds of hydrogen with other elements are called HYDRIDES.
According to nature of bonding, hydrides may be classified into three classes.

IONIC, COVALENT and INTERMEDIATE hydrides.

IONIC HYDRIDES.

Hydrides of alkali metals and heavy members of alkaline earth metals are called ionic hydrides.

1- These are crystalline solid compounds

2- These have high M.P and B.P.

3- Conduct electricity in Molten state

4- They react with H_2O to form H_2 gas.

$$NaH + H_2O \longrightarrow NaOH + H_2\uparrow.$$

5- Their aqueous solution is alkaline due to formation of hydroxal ion.

INTERMEDIATE HYDRIDES.

Hydrides of Be, Mg and transition elements represent the class of intermediate hydrides. Their properties are in between ionic and covalent hydrides. They have polymeric structure.

COVALENT HYDRIDES:-

Hydrides of non metals are covalent in nature

Covalent hydrides are usually gases or volatile liquid.

They are non conductors.

They dissolve in organic solvent.

+ Their bond energies depend upon size and electronegativity of element. Stability of covalent hydrides increase from left to right in a period and decreases from top to bottom in a group.

Fluorine forms most stable hydrides while least stable hydrides are of Thallium.

Elements with electronegativity difference greater than 1.8 form covalent hydrides.
Boiling Points of covalent hydrides generally increase down the group. Except the hydrides like H_2O, HF and NH_3 which have high B.P due to hydrogen bonding.

I-A	II-A	II-B	III-A	IV-A	V-A	VI-A	VII-A
LiH	BeH_2		BH_3	CH_4	NH_3	H_2O	HF
NaH	MgH_2		AlH_3	SiH_4	PH_3	H_2S	HCl
KH	CaH_2	ZnH_2	GaH_3	GeH_4	AsH_3	H_2Se	HBr
RbH	SrH_2	CdH_2	InH_3	SnH_4	SbH_3	H_2Te	HI
CsH	BaH_2			PbH_4			
IONIC		INTERMEDIATE.			COVALENT HYDRIDES.		

HYDRIDES OF ELEMENTS I-A → VII-A AND II-B.

OXIDES :-

Almost all elements of periodic table form one or more binary compound with oxygen called OXIDES.
Nature of oxide depends upon position of elements in periodic table. Oxides can be classified on the basis of their acidic or basic nature.

1 ACIDIC OXIDES
2 BASIC OXIDES
3 AMPHOTARIC OXIDES.

Acidic and basic oxides react with each other to form salt. Amphoteric oxides have properties of both acidic and Basic oxides.

1. BASIC OXIDES:-

Oxides of alkali and alkaline earth metals except beryllium are basic and contain O^{2-} ions. The O^{2-} ion is highly reactive and it takes proton from water to form $OH^{\ominus}$ ion.
Basicity of these oxides increases down the group e.g NaOH is weaker base as compared to KOH.

2. AMPHOTARIC OXIDES:-

Oxides of relatively less electropositive elements such as BeO, Al_2O_3, ZnO are amphoteric and behaves as acids on treatment with strong bases and act as bases towards strong acids.

$$ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$$

$$ZnO + 2NaOH \xrightarrow{H_2O} Na_2[Zn(OH)_4]$$

3. ACIDIC OXIDES:-

Oxides of non metals are usually acidic oxides. for example SO_3 dissolve in H_2O to form acidic oxide.

$$SO_3 + H_2O \longrightarrow H^+ + HSO_4^{1-}$$

Acidity of oxides decreases down the group and increases as we proceed left to right in a period.

In given period oxides progress from strongly basic through weakly basic, amphoteric and weakly acidic to strongly acidic as

$$\underbrace{Na_2O, MgO}_{\text{Basic}}, \underbrace{Al_2O_3}_{\text{amphoteric}}, \underbrace{P_4O_{10}, SO_3, Cl_2O_7}_{\text{acidic}}$$

Basicity of metal oxides increases down the group.

$BeO < MgO < CaO, SrO, BaO$.

Acidity increases with increasing oxidation state:

$MnO < Mn_2O_3 < MnO_2 < Mn_2O_7$

I-A	II-A	II-B	III A	IV-A	V-A	VI-A	VII-A
Li	Be		B	C	N	O	F
Na	Mg		Al	Si	P	S	Cl
K	Ca	Zn	Ga	Ge	As	Se	Br
Rb	Sr	Cd	In	Sn	Sb	Te	I
Cs	Ba	Hg	Tl	Pb	Bi	Po	At
	BASIC			AMPHOTERIC		ACIDIC	

CLASSIFICATION OF OXIDES.