UNIT (04)

CLASSIFICATION OF ELEMENTS AND PERIODIC TABLE

↳ The table of element in which element are placed with their similar properties together in the same vertical column and this similar element are separated from one another is called periodic table.

Mendeleev's periodic law:-

↳ In 1869, Russian chemist Dmitri Mendeleev classified the element on the basis of increasing atomic weight and proposed a law called Mendeleev's periodic law.
It states that," The physical and chemical properties of elements are periodic function of their atomic weight."

★ Main features of Mendeleev's periodic table:-

* Only 63 elements are classified in the Mendeleev periodic table which consists of 7 horizontal rows and 9 vertical column. The horizontal rows are called periods and vertical column are called groups.

1) Short period :-
↳ The first 3 period are called short periods. Since they contain few elements. The first period contain 2 elements and 2^{nd} & 3^{rd} period contain 8 electrons.

2) Long period :-
↳ 4^{th}, 5^{th} & 6^{th} periods are called long periods which contain 18 elements in 4^{th} & 5^{th} period and 32 elements in 6^{th} period. 7^{th} period is long but incomplete (19 elements).

3) Groups :-
↳ There are overall 9 groups which are I to VIII and 0 groups.

4) Subgroup :-
↳ Each groups from 1 to 7 are divided into Subgroups A & B. For e.g: Li, Na, K, Rb, Cs etc. constitude the subgroup A and the elements Cu, Ag & Au constitude the B Subgroup of the group first.

5) Group 8th :-
↳ Group 8th consist of three(3) set each one contain 3 elements. The group '0' consists of inert or nobel gases. These nobel gases are inert and inactive.

V.V.Imp.

Disadvantage/ defect/ Anomalies of Mendeleev's periodic table:-

↳ The main defect in the Mendeleev's periodic table are as follows:-

[1] position of hydrogen (H):-

↳ Hydrogen has dual character it ressembles with alkali metal as well as helogen group. Therefore, position of hydrogen is not correctly defined.

[2] Anomalous pair/ Inversion in law:-

↳ The Certain elements of higher atomic weight are placed before the elements of lower atomic weight.

e.g. Argon [Atomic mass 39.9] is placed before potassium [atomic weight 39.1]. Similarly, cobalt [Atomic mass 58.9] is placed before Nickel (Atomic mass 58.7). etc. This position were not justified.

[3] position of Isotope:-

↳ Isotope of same element have different mass number. Therefore, they should be kept in different groups but they are kept in same group.

[4] position of Lanthanides and Actinides:-

↳ 14 elements after Lanthenum [Ce, 58 – Lu, 71] are known as Lanthanides. And 14 elements after actinium [Th, 90 – Lw 103] are known as actinides. They are placed outside the buttom of the periodic table without proper region.

[5] Separation of Similar element but grouping of dissimilar element:-

↳ The element having Similar properties are separated. For e.g. Cu, Hg, Au, etc. but element having dissimilar properties ar placed in a same group.

6) Cause of ~~periodicity~~ periodicity:-

↳ Mendeleev's periodic law doesn't explain about main cause of periodicity.

Advantages of Mendeleev's periodic table:-

* Systematic study of element
* prediction of new element.
* Correction of faulty atomic weight

Modern periodic law:-

↳ In 1912, Moseley proposed a new periodic law on the basis of atomic number called modern periodic law. It states that "The physical and chemical properties of element are periodic function of their atomic number."

periodicity:-

↳ The reoccurance of element with similar properties after certain regular interval when these are arranged in increasing of their atomic number is called periodicity.

eg:

Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

Cause of periodicity:-

↳ From the electronic configuration of element, all elements having is due to reoccurance of Similar electronic Configuration. If elements are arranged in order of their increasing atomic number. Elements coming at interval of 2,8,8,18,18, 32,32 and will have Similar properties. These number are Called Magic Numbers.

Main feature of Modern periodic law:-

[1] It has seven horizontal rows are Called periods.
i) The first period is the shortest and consists of 2 elements (H-He).
ii) The second period (Li-Ne) and third period (Na-Ar) are short period having 8 elements each.
iii) The fourth period (K-Kr) and fifth period (Rb-Xe) are long periods, each contains 18 elements.
iv) Sixth period (Cs-Rn) is longest period and Consists of 32 elements.
v) The Seventh period (Fr---) Contains 19 elements and is incomplete period. Elements beyond Uranium are Synthetic and are called trans-Uranic. They are radioactive.

[2] Group:-
↳ They are all together 18 vertical Column and divided into 9 groups. [Group-I] — [Group-VII] are sub-divided into Subgroup A & B. Consists of 14 Column. The elements Lanthanides and Actinides have been placed in separated row at the bottom of the periodic table.

[3] Classification of element into s, p, d & f blocks.

s-block element :-

* Last electrons drops into s-orbital / s-subshell.
* They are kept at the left of periodic table & form ionic compound
* Alkali metals (ns^1) & alkaline earth metals (ns^2) and hydrogen are kept under s-block.
* Their general electronic configuration is ns^{1-2}.
* They are soft metals having low melting and boiling point.
* They are highly electropositive and have low ionization energy.

p-block element :-

* last electrons drops into p-orbital / p-subshell.
* 'IIIA' to 'VIIA' and '0' group elements are kept under p-block.
* Their general outer electronic configuration is ns^2np^{1-6}
* They are kept at the right side of periodic table.
* They are mostly electronegative in nature.
* They form ionic as well as covalent compount.

[s & p block elements are called representative elements.]

d-block element :-

* Last electron drops into d-orbital / d-subshell.
* IB to VIIB and VIII group elements are kept under d-block.
* Their general outer electronic configuration is $(n-1)d^{1-10}ns^{1-2}$.
* They are called transition elements.
* They are kept at the centre between s & p block element.

* They form ionic as well as covalent compounds.
* They show variable oxidation state.
* They form colour compound.

f-block element:-

* Last electron drops into f-orbital/f-subshell.
* Their general outer electronic configuration is $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$
* They kept below the periodic table.
* They are called inner transition elements.
* They are heavy metals having high melting & boiling point.

Advantages of Modern periodic law:-

[1] Atomic number is more fundamental property than atomic mass. So it is more superior to Mendeleev's periodic classification.

[2] The study of element is more systematic and convenient by classifying them into s, p, d & f block elements.

[3] The position of hydrogen is kept along alkali metals (IA) according to electronic configuration.

[4] Mendeleev's anomalous pairs of elections are corrected by Modern periodic table.

[5] Isotopes have same atomic ~~mass~~ number so they are kept in same group.

[6] According to this law similar groups elements are kept together and dissimilar groups elements are kept/separated in different groups.

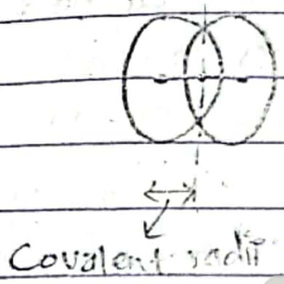
Disadvantage of Modern periodic table:-

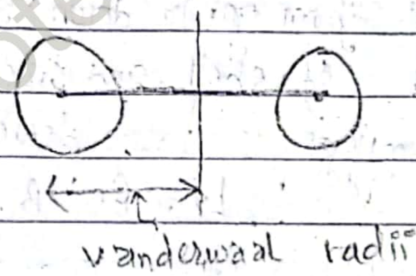
1) Helium should be kept in s-block from its outermost electronic configuration but it is placed in p-block which is not explained by Henry Moseley.

periodic properties or atomic properties of element:

* Atomic radii (Atomic size)

↳ It is the distance of nucleus to outermost shell, atomic radii are Covalent radii, metalic radii, Vanderwaal radii.

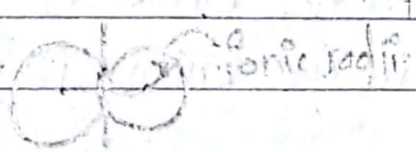




$r_{cov.} = \frac{1}{2}$ [Internuclear distance between two bonded atom]

$r_{van.} = \frac{1}{2}$ [Internuclear distance between two non-bonded nearest neighbour's atoms]

Ionic radii:-



Note:-

[Covalent bond > Metalic bond > Vanderwaal bond]:

[1] # In period from left to right atomic radii decreases due to increase in nuclear charge and electron is added to same shell. Therefore, electron are attracted more strongly by nucleus due to which size of atom decreases

Nuclear charge

period $\rightarrow$ Increases $\rightarrow$

Li Be B C N O

Na

group ↓ K

Rb

decreases ↓ Cs

(2) In group from up to down atomic radii increases due to decrease in number of shell. and size also increases.

atomic radii

decreases $\rightarrow$

Li Be B C N O F

Na

K

Increases ↓ Rb

Cs

Fr

[3] In ionic radii, the size of cation is smaller than size of parent atom.

e.g. $Na > Na^+$, $Ca^{2+} < Ca$

Again, the size of anion is greater than size of parent atom.

e.g. IR; $Cl^- > Cl$

IR; $S^{2-} > S$

Isoelectronic Species:-

↳ The species having same no. of electron but different nuclear charge are called isoelectronic species. The attractive force between electrons and nucleus increases when the nuclear charge increases as a result decrease of atomic radii.

e.g. $C^{4-} > N^{3-} > O^{--} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+} > Si^{4+}$

Valency:-

↳ In a group valency of element is same, this is due to the presence of same number of valence electron.

IA group

Valency I.

H
Li
Na
K
Rb
Cs
Fr

↳ In a period, valency of element first increases from one to four then decreases to '0'.

elements	Li	Be	B	C	N	O	F	Ne
Valency	1	2	3	4	-3	-2	-1	0

Ionization Energy [Ionisation Energy] [I.E] [I.P]:

↳ It is defined as the amount of energy required to remove the most loosely held electron from the isolated gaseous atom and formation of cation.

e.g. ↳ $M(g) + IE \longrightarrow M^{+}(g) + ne^-$

OR, $M(g) - ne^- \xrightarrow{I.E.} M^{+}(g)$

↳ $Na + IE \longrightarrow Na^{++} + e^-$

OR, $Na - e^- \xrightarrow{IE} Na^{++}$

☆ In a group, I.E. decreases on moving from up(top) to down(bottom).

☆ In a period, IE increase from left to right.

I.E. — Increases →

Li Be B C N O F Ne

Na

K

Rb

Cs

Fr

Decreases ↓

Successive I.E.:-

↳ The energy required to remove successive electrons from cation is called successive I.E.

e.g. $M(g) + I.E_1 \longrightarrow M^{+}(g) + e^-$

$M^{+}(g) + I.E_2 \longrightarrow M^{++}(g) + e^-$

$M^{++}(g) + I.E_3 \longrightarrow M^{+++}(g) + e^-$

&

$Al \xrightarrow{IE_1} Al^{+} + e^-$

$Al^{+} \xrightarrow{IE_2} Al^{++} + e^-$

$Al^{++} \xrightarrow{IE_3} Al^{+++} + e^-$

The order of successive Ionization energy is;

$\boxed{(IE)_{III} > (IE)_{II} > (IE)_{I}}$

Factor affecting the I.E.

(1) Atomic Size:-

↳ Ionization energy decreases with increase in atomic size because larger the distance from nucleus lesser will be the attraction between nucleus and valence electrons then valence electron can be easily removed.

i.e,
$$IE \propto \frac{1}{\text{Atomic Size}}$$

(2) Nuclear Charge:-

↳ If the nuclear charge increases the force of attraction between nucleus and electron increases. Therefore, loss of electron becomes difficult. Thus, ionization energy increases.

i.e, $IE \propto \text{Nuclear charge}$

(3) Screening effect [Shielding effect]:-

↳ The outermost electron are screened/shielded from the nucleus by the inner electron. This phenomenon is called screening effect.

If the number of electrons in the inner shell is large, the screening effect will be larger. As a result, the attractive interaction between nucleus and valence electron will be less. Thus, Ionization energy decreases.

i.e,
$$I.E. \propto \frac{1}{\text{Screening effect}}$$

[4] Electronic configuration:-

↳ Half filled or full filled orbital are more stable than other. Hence more energy is needed to remove in electron from such atoms. Thus more stable the electronic configuration greater will be the ionisation energy. That's why the I.E. of '0' group elements (He, Ne, Ar, etc) & Be ($1s^2 2s^2$) have higher I.E.

IE_1 Be > B [$2s^2 > 2p^1$]

IE_1 N > O [$2p^3 > 2p^4$]

Electron Affinity [E.A.]

↳ The amount of energy released when an electron is added to in isolated gaseous atom to for anion is called electron affinity.

e.g. $X(g) + e^- \longrightarrow X^{\ominus}(g) + Energy$

The energy releasted is electron affinity.

$F(g) + e^- \longrightarrow F^{\ominus}(g) + Energy$

Factor effecting electron affinity:-

↳

[1] Atomic size:-

↳ On moving from top(up) to bottom(down) in a group size of atoms increases due to which tendency of atom to attract incoming electron decreases. therefore electron affinity decreases.

[2] Nuclear charge:-

↳ On moving from left to right in a period nuclear charge increases & size of atom decreases due to which tendency of atom to attract incoming electron increases.

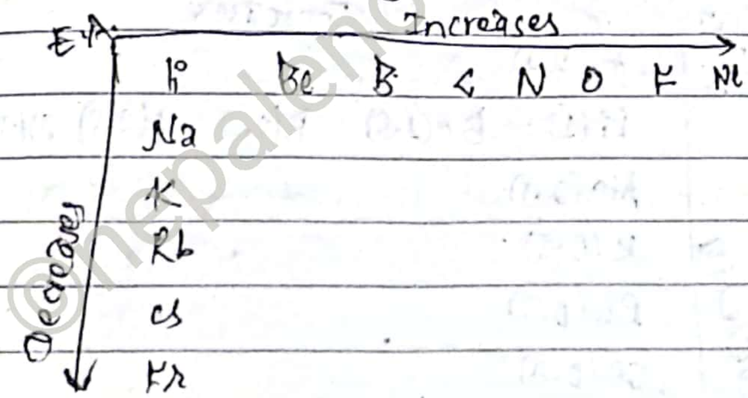
[3] Electronic configuration:-

↳ If the atom has stable electronic configuration lesser will be its tendency to accept electron and lower will be the value of its electron affinity.

↳ In a group electron affinity of an element decreases from up to down

↳ In a period electron affinity of an element increases from left to right

e.g.



↳ Nobel gaseous has '0' electron affinity:-

This is because of the highly stable configuration.

Electronegativity:-

↳ The tendency of an atom of an element to attract a shared pair of electrons towards itself is called electronegativity of the element.

$$\text{Electronegativity [E.N.]} = \frac{IE/IP + EA}{2}$$

It depends upon atomic size, number of inner shell electron, nature of element etc.

↳ In a group, E.N. of elements decreases from top to bottom while
↳ In a period E.N of elements increases.

E.N. on pauling scale

E.N. Increases →

H(2.1)
Li(1) Be(1.5) B(2.1) C(2.5) N(3) O(3.5) F(4) Ne
Na(0.9)
K(0.8)
Rb(0.8)
Cs(0.7)
Fr(0.7)

Decreases ↓