Classification of Elements - Periodic Table.

Classification of Elements - The Periodic Table

* Now, we have 118 elements.
* To known about each and every element we should classify the elements.
* The first classification made by Dobenena in 1829.
* He classified the elements based on atomic weight property.
* He classified the elements in Triads.

Dobenener triad law :-

Middle element atomic weight approximately equals to the average of 1st and last elements "atomic weights"

Ex:- $Li = 7$

$Na = ? = 23 \rightarrow \frac{7+39}{2} = \frac{46}{2} = 23.$

$K = 39$

* Second classification made by Newland.

Drawbacks of Dobenener:-

He not classified all the elements.

1. This law is not applicable to high mass element (above 40).
2. This law failed in case of very low mass and very high mass.

Ex:- Filourine, chlorine, Bromine

3. The average of flourine and Bromine is not equals to chlorine atomic mass.
4. He not arrange all the elements into triads.

Newlands Octers law:-

* He classified elements based on atomic weight.
* He arranged the elements in order to increasing the atomic weights.

Octave's law:-

for every 8th elements resembles with 1st element.

Ex:- In our indian music.

$\boxed{Sa}$	ri	ga	ma	pa	da	ni	$\boxed{sa}$
1	2	3	4	5	6	7	8

He arranged the atomic numbers.

$\boxed{H}$	Li	Be	B	C	N	O	$\boxed{F}$	Na
1	2	3	4	5	6	7	8	9

Drawbacks of Newland.

* It is restricted for only 56 element newland discovered elements could not be arranged.
* It is not valid for higher mass elements more than calcium.

Mendeleef's periodic table:-

Periodic law:-

The physical and chemical properties of the elements are the periodic functions of their atomic weight.

* He classified the elements based on atomic weight property.

I. Groups and subgroups:-

1. It has 8groups.
2. Every group divides into 2 subgroups.
3. All the element in the group are same in all properties and valencies

II. Periods:-

* The horizontal rows in Mendeleefs periodic table are called periods.

There are seven periods in the table.

* They are denoted by Arabic numerals.

III. Missing element (or) Predicted element :-

Based on the arrangement of the elements in the table he predicted that some element were missing are left blank spaces at the appropriate places in the table.

* He named those predicted element tentatively by adding the prefix 'eka' (eka is a sanskrit word of numeral one).

1. Eka Aluminium → Gallium.
2. Eka boron → Scandium.
3. Eka silicon → Germanium.

IV. Correction of atomic weight:-

Mendeleef corrected the atomic weight of element like Berilium (Be); Indium (In) and Gold (Au)

Ex:- Be atomic weight
= Equivalent weight x valency
$= 4.5 \times 3 = 13.5$

After correction
Be AW = $4.5 \times 2 = 9$.

V. Anomalous series :-

the changed the order of descending atomic weight like Fe (127.6) I (126.9).

Limitations :

1. In anomalous pairs of elements highest atomic weight placed infront of lower atomic weight Eg: Te, I.

2. Dissimilar elements are placed together :-

Elements with dissimilar were placed in same group as subgroups as A, B.

Eg :- Cl is non-metal placed in VII A and Mn metal placed in VII B.

3. Modern periodic table :-

1. Modern periodic table was given by mosley in 1913.
2. The Modern periodic law stated as the properties of elements are periodic function of their "atomic number".
3. Modern periodic table has eighteen vertical columns known as groups.

Modern Periodic table (or) long form of periodic table:-

* Periodic law:- The physical and chemical properties of the elements are the periodic functions of their atomic number.

H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 3	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Te 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	La 57	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88	Ac 89	Rf 104	Db 105	Sg 106	Bh 107	Hs 108	Mt 109	Ds 110	Rg 111	Cn 112		Fl 114		Lv 116		

Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
Th 90	Pa 91	U 92	NP 93	Pu 94	Am 95	Cm 96	BK 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

Property based on atomic number:-

1. It has 7 periods.
2. It has 18 groups.
3. Ist group consist 2element(hydrogen, Helium].
4. IInd group consist 8 elements.
5. IIIrd group consist 8elements.
6. IVth group consist 18elements.

7. V^{th} group consist 18 elements.
8. VI^{th} group consist 32 elements.
9. VII^{th} group consist (A) incomplete elements (32)
10. I A group consist Alkali metal.
11. II A group consist Alkali earth metal.
12. VII A group consist Halogens.
13. VIII A group consist noble gass (O group elements).

Periodic table divided into '4' types.

1. Representative elements.

s and p blocks are called Representative elements.

s block

H - $1s^1$

Li - $1s^2 2s^2$

Na - $1s^2 2s^2 2p^6 3s^1$

The general electronic configuration is ns^1 or ns^2

p block

B - $1s^2 2s^2 2p^1$

F - $1s^2 2s^2 2p^5$.

The general electronic configuration is ns^2 or np^{1-5}

2. Innert gases (noble gas).

The innert gas elements are Heliom (He-2), Neon (N-10), Argon (Ar-18), Krypton (Kr-36), Xenon (Xe-54), Radon-(Rn-86).

* These are stable elements.
* The elements never participate in any reaction.

* The 'O' group elements are called innert gas.

Zero group:-

He - $1s^2$

Ne - $1s^2$ $\underline{2s^2 \; 2p^6}$.

Ar - $1s^2 2s^2, 2p^6$ $\underline{3s^2 \; 3p^6}$.

The general electronic configuration is $ns^2 np^6$.

3. Transition elements.

The d-Block elements are called Transition elements.

d-Block

Sc - $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$.

The general form of electronic configuration is $ns^2 (n-1)d^{1-10}$

4. Inner Transition elements:-

The f-Block elements are called inner Transition elements. The general form of electron configuration is $ns^2 (n-1) d^{0-1} (n-2) f^{1-14}$.

Physical properties.

Atomic radius or atomic size.

The distance between nucleus to outer most orbit (valency orbit) is called atomic radius

units:- Angstrom(A°)

$1A^\circ = 10^{-8}$ cm

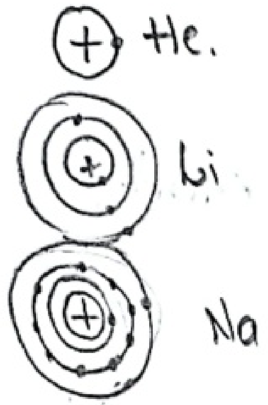
$1A^\circ = 10^{-10}$ m.

Groups:-

In group top to bottom the atomic size increases.

Reason:-

From top to bottom the differentiating electron enter into new shell.



* In Na, Na^+ which has more size?

Na has more size

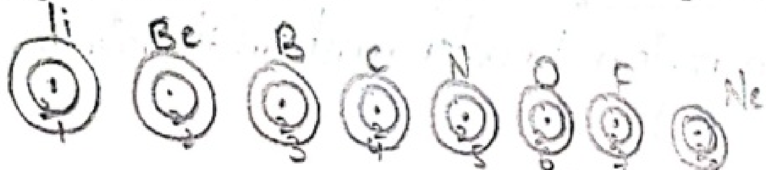
Reason :- Because, when the electrons number decreases to the same proton number the atomic size also decreases.

Period.

In a period from left to right atomic size decreases.

Reasons:-

As we go from left to right the nucleus attraction force increase on outer most electrons so atomic size decrease.



Ionization energy.

The minimum energy required to remove one electron from outer most orbit of an element in gaseous state is called Ionization energy.

units:- electron volts(e.v).

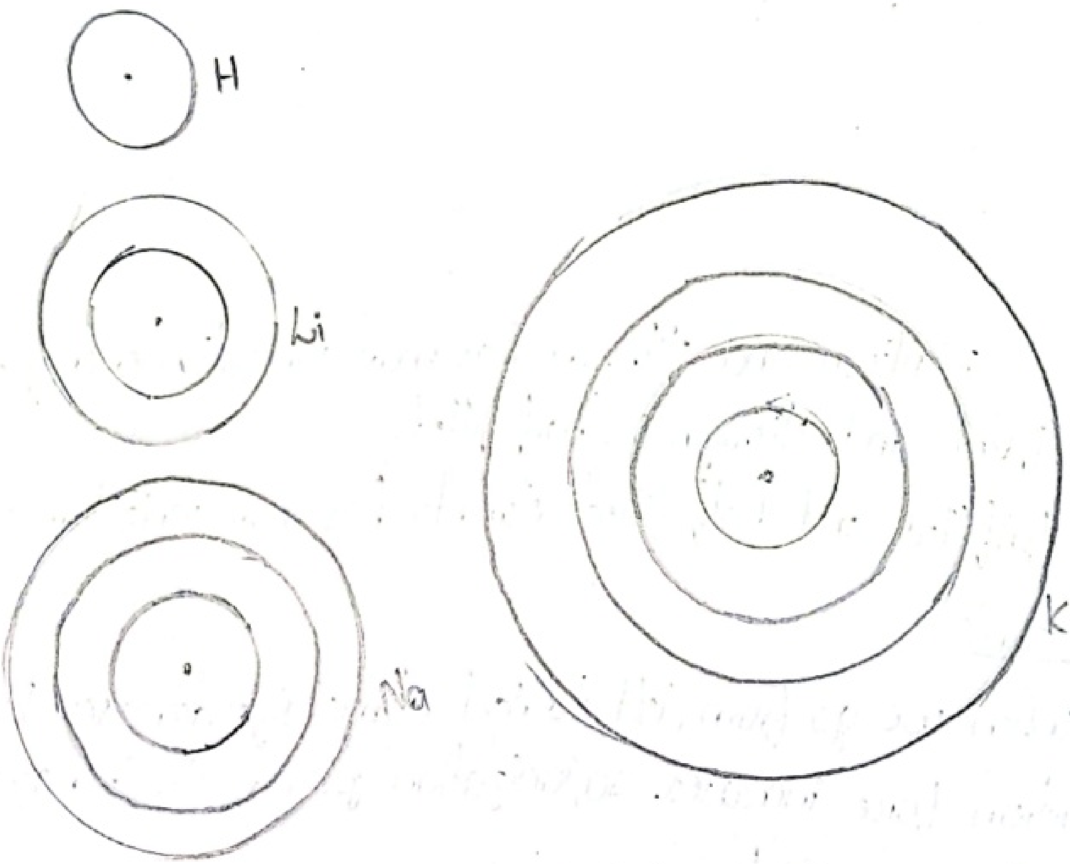
k.callories/mole and k. Joule/mole.

Groups:-

In a group from top to bottom ionization energy decreases.

Reason:-

When from top to bottom atomic size increase. so, the nucleus attraction force decreases. So, ionization energy decreases.

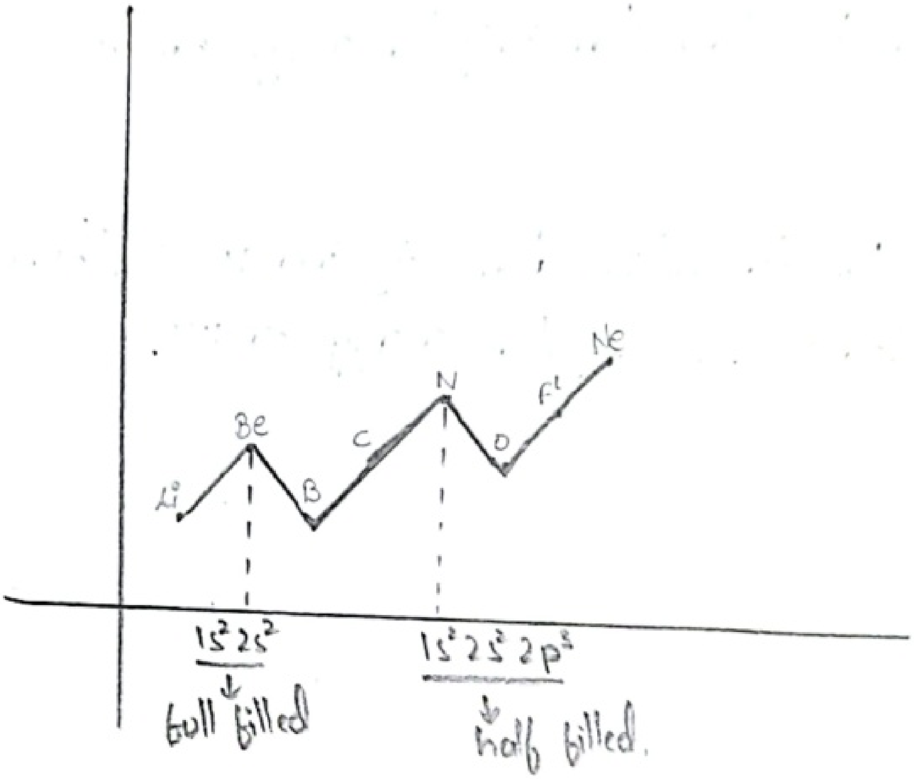


Periods:-

In a period from left to right the ionization energy not follows regular trend.

"Left to right almost increases".

Eg :- II periods.



Berilium and Nitrogen has more ionization energy because Berilium is full filled and Nitrogen is half filled.

* The fulfilled and half filled elements has more ionization energy.

Reason:

when we go from left to right atomic size decreases. So, the nucleus attraction force increases. So, ionization power also increases.

Factors of ionization energy :-

1. Atomic size.
2. Nuclear charge.
3. Shielding effect or screeing effect.
4. Penetration power.
5. Stable electronic configuration.

Factors of Ionization energy:-

1. **Atomic size:-**

When atomic size increase ionization energy decreases. Because nucleus attraction force on valency electrons decreases.

$$\text{Ionization energy} \propto \frac{1}{\text{atomic radius}}$$

2. **Nuclear charge:-**

Nuclear charge increase ionization energy also increases because nuclear attraction force on valency ele is more.

$$\text{Ionization energy} \propto \text{Nuclear charge}$$

3. **Screening effect & Shielding effect:-**

When orbits b/w nucleus to outermost electron acts as screen when screens increases the ionization energy decreases on the valency electron.

$$\text{Ionization energy} \propto \frac{1}{\text{Screening effect}}$$

4. **Penetration power of the orbitals:-**

Orbitals belonging to the same main shell have different penetration power (Ionization) towards the nucleus.

Eg:- $4s > 4p > 4d > 4f$

If it is easier to remove electron from the 4f than 4s because it penetration power is more. ionization energy is also more.

$$\text{Ionization energy} \propto \text{penetration power}$$

5. Stable electronic configuration:-

Atom which has half filled (or) full filled electronic configuration has more stable for that Ionization is very more.

Li<Be<B<C<N<O<F<Ne.

Be(z=4)	B(z=5)	C(z=6)	N(z=7)	Ne(z=10)
$1s^2 2s^2$	$1s^2 2s^2 2p^1$	$1s^2 2s^2 2p^2$	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^6$
			↓ Half filled configuration	↓ full filled configuration.

How does atomic properties varies in group:-

Periodic Property.	Trend in (T.B) groups	Trend in (L.R) Periods.
Atomic size	Increase	Decrease.
Electro positivity	Increase	Decrease
Reduction property	Increase	Decrease
Metallic property	Increase	Decrease.
Non-Metallic Property.	Decrease	Increase.
Electron affinity	Decrease	Increase.
Electro negativity.	Decrease	Increase.
oxidizing property	Decrease	Increase.
Ionization energy	Decrease	Increase.

Question:-

Second I.E of an element is higher than its first Ionization energy : why?

A. The minimum energy to required to remove one electron from outer most orbit of an atom in gaseous state is called the first Ionization energy.

The energy required to remove an electron from uni-positive ion is called the 2nd Ionisation energy.

The 2I.E is greater than 1st I.E because the nucleus attraction force on valency electron is more for uni positive ions (Nucleus).

ex:- $Na \xrightarrow{IE_1} Na^+ + e^-$ [Q energy required].

$Na \xrightarrow{IE_2} Na^{+2} + e^-$ [Q_2 energy required].

($Q_2 < Q_2$).

3. Electron Affinity:- The electron affinity has energy released when one e^- added to outer most orbit of the atom in gaseous state

units :- ev(electronvolts).

Group:- In group top to bottom electron affinity decreases.

Period:- In priod left to right the electron affinity increases.

Questions:-

factors of ionization energy? (or) factors of electron affinity?

A. The electron affinity of an element is the energy released when an electron added to outer most orbit of the atom in gaseous state.

Factors :-

Atomic size :- When atomic size increases ionization energy decreases.

Reason :- Ionization decreases because the nucleus orbit attraction force is less on outer most orbit.

$$\text{Ionzation energy} \propto \frac{1}{\text{atomic size}}$$

ex :- 1. In Li and Na which element has very least energy?

Na has very least Ionization energy.

2. Na and Cl which element has more energy?

Cl has more ionization energy.

Nuclear charge :- When nuclear charge increases I.E energy also increases

Ionization directly propotional to Nuclear change.

$$I.E \propto \text{Nuclear charge}.$$

ex :- mg, si

Si has more I.E because more nuclear charge.

3. Screening Affect :- When screening Affect increases I.E decreases.

$$I.E \propto \frac{1}{\text{screening Affect}}$$

ex :- Li and K

Li has more ionization energy because

Li has less screening effect.

4. Penetration power :- when penetration power increases ionization energy increases.

I.E ∝ penetration power.

ex :- Be $1s^2 2s^2$

B $1s^2 2s^2 2p^1$

Be has more ionization energy.

4s>4p>4d>4f in that from 4f we can remove electro

n easily But 4s has more penetration energy.

Electro positivity :- The property of lose and convert into positive ions is called electron positivity left side elements are called positivity elements. Because it loses its electron's easily.

Group :- from top to bottom the electron positivity will increases.

Period :- from left to right the electron positivity will decreases.

eg. Highest electro positivity is CS

lowest electro positivity is H.

5. Electro Negativity :- The tendency of an element to attract the bounded electron towards it self.

Group :- from top to bottom Electro negativity decreases.

period :- from left to Right Electro negativity increases.

ex. Highest electro negativity is H

lowest electro negativity is C.S.

electro negativity is measured in "pauling scale."

Electro negativity - $\frac{\text{Ionization + Electron Affinity}}{2}$

Oxiding property:- adding of oxygen or remoing of hydrogen or removal of electron.

Group:- In group oxidizing property decreases.

Period:- In period oxidizing property increases.

Redusing property:- adding of hydrogen or adding of water or removal of electron.

Group:- In group redusing property increases.

Period:- In period redusing property decreases.

8. Metallic property:- Which can loss electrons are called metallic property. ex:- $Na \rightarrow Na^{+}$

$mg \rightarrow mg^{+2}$

$Al \rightarrow Al^{+3}$

$k \rightarrow K^{+}$

$Ca \rightarrow Ca^{+2}$

Group:- In group metalic property increases.

Period:- In period metalic property decreases.

9. Non-metallic property:- which can gain the electron are called on-metallic property. ex:- $Cl - Cl^{-}$

$S - S^{-2}$

$O - O^{-2}$

$F - F^{-1}$

Group - In group Non-metallic property decreases

period - In period Non-metallic property increases.