2 TABLE 1

...is periodic law? Describe the main characteristic of long form of periodic table.

Jewel 02MT

attempt -

...periodic law :- The periodic law stated by Mandeleef in 1869 was the first successful attempt in the classification of elements and their comparative study. According to this law ;

"The physical and chemical properties of elements are a periodic function of their atomic weights."

After inventing isotope Mosley proved that the chemical and physical properties are more depending on the electronic configuration of the elements. So, the Mandeleef's periodic law was represented by Mosley as follows ;

"The physical and chemical properties of elements are a periodic function of their atomic numbers." i.e., if the elements are arranged in the increasing order of their atomic numbers, the properties of the elements (i.e., similar elements) are repeated after definite regular intervals or periods.

Main characteristics of the long form of periodic table :-

Out of the various periodic tables the extended long form of periodic table is most simple. The main characteristics of the long form of periodic table are as follows ;

Groups: – The vertical columns shown in the table are called groups or families or simply columns.

i) There 18 vertical columns which are IA, IIA, IIIB, IVB, VB, VIB, VIIB and three columns of group VIII, IB, IIB, IIIA, IVA, VA, VIA, VIIA, zero.

ii) Elements of groups IA and IIA are called s-block elements and the elements of IIIA, IVA, VA, VIA, VIIA and zero (with helium inspite of its electronic configuration $1s^2$) groups are called p-block elements. The elements of s-block and p-block except zero group have their outermost shell incomplete while each of their inner shell is complete. These elements are called normal or representative elements. These elements consist of some metals, all nonmetals and metalloids. IA and IIA are called left portion and IIIA to VIIA and zero groups are called the right portion.

iii) Elements of group zero have all their shell completely filled i.e., ns^2np^6 where n is the maximum principal quantum number. These elements are called inert gas elements or noble gases. These are placed at the extreme right of the table.

iv) Elements of groups IIIB (only Sc, Y, La and ...), IVB, VB, VIB, VIIB, VIII, IB and IIB have their two outermost shells [i.e., nth and (n-1)th] are partly filled while the remaining remaining inner shells [1st, 2nd - - - - $(n-2)$th] are completely filled. The common electronic configuration of these elements are $(n-1)d^{1-10} ns^2$. These elements are called d-block elements. These are also called $\boxed{\text{transition}}$ elements. These elements are placed in the middle portion of the periodic table. All these elements are metals.

v) Two groups of 14 elements lying in group IIIB [Ce(58) to Lu(71) and Th(90) to Lw(103)] have thier three outermost shells [$(n-2)$th, $(n-1)$th and nth shell] are partly filled while the remaining inner shells [1st, 2nd - - - $(n-3)$th shell] are completely filled. These are called Lanthanides and Actinides respectively and have been placed at the bottom of the table. Lanthanides series have the incomplete 4f orbital and Actinides series have the incomplete 5f orbital. These elements are also called f-block elements.

Periods. — The horizontal rows shown in the periodic table are called periods. Long form of periodic table consists of 7 ~~columns~~ periods.

i) 1st period consists of two elements which are H(1) and He(2).

ii) Second period ($n=2$) : There are two shells in the elements of this period. It has 8 elements in this period. These are Li (3) to Ne (10).

iii) Third period ($n=3$) : This period also has 8 elements. These are Na(11) to Ar (18).

Second and third periods are called short period.

iv) Fourth period ($n=4$) : This period contains 18 elements. These are K (19) to Kr (36).

v) Fifth period ($n=5$) : This period has 18 elements also. These are Rb(37) to ~~to~~ Xe(54).

Fourth and fifth periods are called long period.

vi) Sixth period ($n=6$) : This period has 32 elements. These are Cs (55) to Lu(71) Rn(86). This period is called longest period. 6th period also includes 14 rare earths or Lanthanides.

vii) Seventh period ($n=7$) : This period is an incomplete period and at present it consists of 19 elements which are Fr (87) to Ha(105) Hs(109). All these elements are radioactive. This period also includes 14 actinides [Th(90) to Lw (103)]. The elements after U (92) are called transuranic elements. These elements are the result of atomic research and hence are synthetic elements. Lanthanides and actinides are placed at the bottom of the periodic table.

viii) The elements of 3 columns of group VIII are ...ilar in chemical and physical properties.

M ...

Describe the classification of elements according to electronic configuration.

On the basis of electronic configuration the elements are classified into two ways.

a) Differentiating classification.
b) Bohr classification.

a) Differentiating classification :—

According to this classification elements ...e classified into s, p, d and f blocks depending the nature of atomic orbital in which the last differentiating (or extra or additive) electron ...ters.

(i) s- block elements :— The elements in which differentiating electron enters the ns orbital are ...led s-block element. The elements of groups IA ...IIA [alkali metal, alkaline earth metal] belong to this block. The valence shell configuration of these elements varies from $ns^1 - ns^2$. The members of s-block elements lie on the extreme left of the periodic table.

(i) p-block elements :- The elements in which p-orbitals are being progressively filled are called p-block elements. The elements of the groups IIIA, IVA, VA, VIA, VIIA and zero (with the possible exception of helium whose configuration is $1s^2$) are the members of p-block. The valence-shell electronic configuration of these elements vary from ns^2p^1 to ns^2p^6. These elements lie at the extreme right of the periodic table and consists of some metals, all non-metals, metalloids and noble gases.

(ii) d-block elements (transition elements) : The elements in which the last electron enters the $(n-1)d$ orbital of the $(n-1)$th main shell are called d-block elements. The elements of the groups IIIB – VIIB and VIII are the members of this block. The valence-shell electronic configuration of the atoms of d-block elements can be represented by $(n-1)d^{1-10}ns^{0,1,2}$. These elements are placed in the middle of the periodic table i.e., between s and p-block. And all these are metals. These elements are also called transition elements.

The d-block elements are classified into four series corresponding to the filling of 3d, 4d, 5d and 6d-orbitals of $(n-1)$th main shell.

...p IIIB, IVB, VB, VIB, VIIB, VIII, IB and IIB belong to this type. Thus we see that these elements are located between electropositive elements of group IA and IIA (s-block elements) and electronegative elements of group IIIA to VIIA (p-block elements).

outermost shell
remaining inner shells

iv) Inner transitional elements : In the atoms of these elements, three outermost shell [nth, (n-1)th and (n-2)th shells] are partly filled while the remaining inner shells [1st, 2nd, 3rd ---- (n-3)th shell] are completely filled. Lanthanides [Ce(58) to Lu(71)] and Actinides [Th(90) to Lw(103)] belong to this type. These elements lie in 6th and 7th period respectively. The complete electronic configuration of these elements can be represented as follows :

Lanthanides : $1s^2 2s^2p^6 3s^2p^6d^{10} 4s^2p^6d^{10}f^{x} 5s^2p^6d^{y} 6s^2$

[where, $x = 1, 2 \cdots 14$ and $y = 0, 1$]

Actinides : $1s^2 2s^2p^6 3s^2p^6d^{10} 4s^2p^6d^{10}f^{14} 5s^2p^6d^{10}f^{x} 6s^2p^6d^{y} 7s^2$

[where, $x = 1, 2, 3 \cdots 14$ and $y = 0, 1, 2$]

The configuration given above clearly show that these elements have not only partially filled $(n-1)d$ orbitals but also underlying partially filled $(n-2)f$ orbital, ns orbital is completely filled. Thus the ante-penultimate shell (i.e., $(n-2)$th shell) is being expanded from 18 to 32 by the addition of electrons to $(n-2)f$ orbital.

In case of lanthanides $(n-2)f$ i.e. 5f orbital is being progressively filled. Elements within each of these series are very similar in chemical properties. These elements are located in group IIIB and have been given a separate place at the bottom of the periodic table.

Write down the defects of long form of periodic table.

Although the long form of periodic table is superior to Mendeleef's periodic table in many respects it retains some of the defects as such. For example

i) The position of hydrogen: The problem of the position of hydrogen in this table has not been solved completely. Although the electronic configuration of hydrogen is the same as the valence-shell configuration of alkali metals in the sense that both have only one electron ($H = 1s^1$, alkali metals $= ns^1$), their properties are not the same.

ii) Position of helium : Chemically helium is an inert gas but the remaining members of the inert gas group posses the valence-shell configuration $ns^2 p^6$ against that of helium. ($He \rightarrow 1s^2$). Further, since helium and alkaline earth metals have similar electronic configuration ($He \rightarrow 1s^2$, alkaline earth metals $\rightarrow ns^2$), helium should be placed in the group of alkaline earth metals. But this cannot be done as the properties of helium are altogether different from those of alkaline earth metals.

iii). Position of Lanthanides and Actinides : This periodic table is unable to include Lanthanides and Actinides in its main body.

iv) The arrangement is unable to reflect the electronic configuration of many elements.

periodic

What are the advantages or merits/superiority of the long form of periodic table over Mendeleef's periodic table ?

The long form of periodic table has a number of advantages or merits over the Mendeleef's periodic table in the following respects : –

(i) The arrangement of elements is based on a mo[re] fundamental property namely atomic number.

(ii) In this table the position of an element is rela[ted] to the electronic configuration of its atom. Thu[s] each group contains elements with similar electronic co[n]figuration and hence similar properties. For example all the valence shell alkali metals have similar valence shell electronic configuration viz. ns^1 configuration and hence have similar properties.

Alkali metals	Complete electronic configuration	Valence shell configurati[on]
Li(3)	$1s^2 2s^1$	$2s^1$
Na(11)	$1s^2 2s^2 2p^6 3s^1$	$3s^1$
K(19)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	$4s^1$
Rb(37)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$	$5s^1$
Cs(55)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$	$6s^1$
Fr(87)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6s^2 6p^6 7s^1$	$7s^1$

(iii) There is a gradual change in properties of the elements with the increase in their atomic numbers.

(iv) The inert gases having completely filled elect[ron] shells have been placed at the end of the ea[ch] period. Such a location of the inert gases ... present completion of each period.

(V) In this form of periodic table the elements of [the] two sub-groups have been placed separately and [th]us dissimilar elements do not fall together.

(VI) It provides a clear demarcation of different [t]ypes of the elements like active metals, transition [m]etals, non-metals, metalloids, inert gases lanthanides [a]nd actinides. The elements of group IA and IIA [a]re active metals and are located at the extreme [le]ft of the periodic table. The transition metals are [f]ound in the middle portion of the periodic table. These [ar]e also called d-block elements. Inert gases, metalloids [an]d non-metals are found in the right portion of the [pe]riodic table. These elements are also called p-block [ele]ments. Lanthanides and actinides are placed at the [bo]ttom of the periodic table. The elements of these [ser]ies are called f-block elements.

(VII) Long form of periodic table is easier to [re]member, understand and reproduce than the Mandeleef's periodic table.

(II) It is s-block element.

(III) Hydrogen is a strong reducing agent like alkali metals. reduction

(IV) Hydrogen and alkali metals both react with halogen to produce corresponding halide. halo halogen

On the other hand, there are some arguments on behalf of putting hydrogen in group VIIA of the periodic table. They are as follows :

(I) Hydrogen is a gas like halogens.

(II) It's molecule is diatomic.

(III) It is substituted by halogens from organic compounds.

(IV) Hydrogen is a non metal like halogens (F, Cl, Br, I

(v) Hydrogen forms hydrides with metals viz. LiH, NaH, CaH2 etc. These hydrides may be compared with metal halide like LiX, NaX and CaX_2. If electricity is passed into the solution of these hydride, hydrogen is produced at anode. So, hydrogen shows electronegative characteristics in metal hydrides.

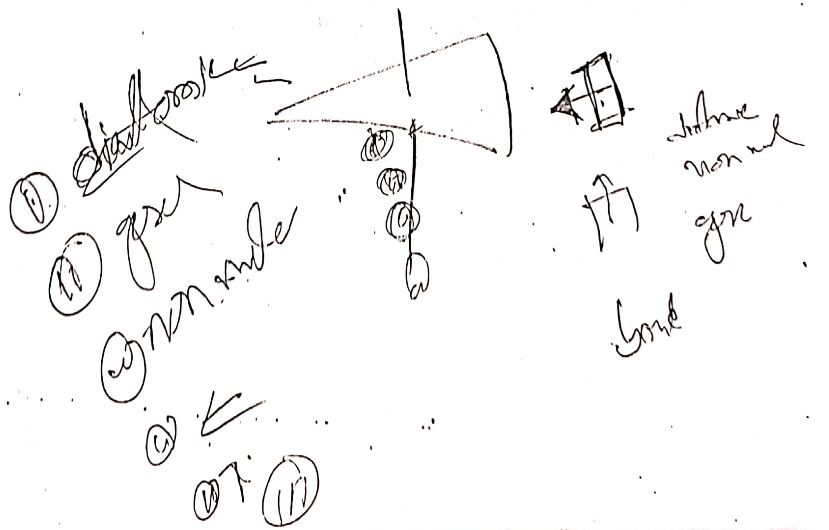
(vi) It is harder to liquify hydrogen than flourine. So, hydrogen may be placed at the top of flourine.

(vii) ~~Hydrogen forms covalent compounds with C, Si (i.e., CH_4, SiH_4) like halogen (i.e., CCl_4, $SiCl_4$)~~

(viii) ~~Hydrogen has one electron less compare nearest inert gas like halogens.~~

From above discussion it is clear that hydrogen behaves as electropositive, electronegative ~~to form~~ or nutral to form compound.

So, hydrogen is placed at the top of the periodic table without acquiring in any group.



Periodic properties

Q: What is periodicity and magic number?

Ans. "The repetition of the elements with similar properties at certain regular intervals ~~in the order of increasing atomic number~~ as in the periodic table is called periodicity ~~of~~ or periodic properties" and the numbers 2, 8, 18 and 32 are called magic number.

Q: What is periodic properties? Write the names some periodic properties.

"The properties which depend on the electronic configuration of the atom of the elements and change gradually with the variation of atomic number in the same periods or groups, are called the periodic properties of the elements."

= in the same periods or groups

The periodic properties of the elements are given below:

1. Atomic radius and volume.
2. Ionic radius.
3. Density.
4. Ionisation energy or ionisation potential.
5. Electron affinity.
6. Electronegativity.

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~~8. Melting point.~~
~~8. Boiling point.~~
9. Valency and oxidation number.
10. Bond energy.
11. Hardness.
12. Magnetic behaviour.
~~13.~~ Thermal conductivity.
14. Parachor.
15. ~~Refra~~ Refractive index.
~~16.~~ Heat of formation.

~~Heat of formation~~

What is atomic volume or gram atomic volume?

Atomic volume or gram atomic volume:—

Atomic volume is defined as the volume in c.c. occupied by one gram atom of the element in the solid state and hence is commonly called gram atomic volume. It is obtained by dividing the atomic weight of the element by its density. i.e.,

$$\text{Atomic volume} = \frac{\text{atomic weight}}{\text{density}}.$$

In other words atomic volume is the volume in occupied by 6.023×10^{23} atoms of the element.

by a single covalent bond.) The covalent radii are essentially the single bond covalent radii (SBCR). Le:

First consider a molecule, A_2 having two like atoms, A and A. If in this molecule the two like atoms are regarded as effective spheres in close contact with each other, then according to the definition of covalent radius, the distance between the nuclei of the atoms, A and A, d_{A-A}, should be equal to the sum of SBCR of A, i.e.,

$$d_{A-A} = r_A + r_A$$

or, $r_A = \frac{d_{A-A}}{2}$ where, r_A = SBCR of atom A.

Now consider a molecule, AB having two different atoms A and B. ~~If we know the internuclear distance, d_{A-B} and SBCR of the other atom~~. Thus if r_A and r_B are the SBCR of the atoms A and B, then;

$$d_{A-B} = r_A + r_B$$

when $x_A = x_B$.

What do you mean by atomic or metallic or crystal radius?

This type of radius is defined as one-half of the distance between the nuclie of two adjacent metal atoms in the metallic close-packed crystal lattice in which metal exhibits a coordination number of 12, e.g., the internuclear distance between two adjacent Na atoms in a crystal of sodium metal is 3.8 A° only, the atomic radius of Na metal = $3.8/2$ A° = 1.9 A°.

Metallic radius are about 10 to 15% higher than the single bond covalent radii.

Q) What is ionic radius ? state with example.

Ans. The ionic radius is defined as the distance between the nucleus of an ion and the point where the nucleus exerts its influence on the electron cloud.

যে প্রভাব প্রদর্শন করে

The internuclear distance between two ions is equal to ~~sum~~ the sum of the radius of the ions of which the ionic ~~...~~ crystal is composed. Thus in $C^+ A^-$ ionic crystal shows that

$$d(C^+ - A^-) = r(C^+) + r(A^-)$$

$$R = r(C^+) + r(A^-)$$

Q) What is vander Waals radius ?

Ans. vander Waals radius is half of the distance between the nuclii of two non-bonded ~~...~~ neighbouring atoms of two adjacent molecules.

In the crystaline form of benzine hexachlorid ($C_6H_6Cl_6$), the molecules are arranged, so that the shortest distance between chlorine nuclii on different molecules is 3.6 A°. Thus in this molecule the vander Waals radius of chlorine is $\frac{3.6}{2} = 1.8$ A° (half the inter atomic distance).

What is the comparison between the vander Waals and covalent radius ?

This comparison can be made by considering chlorine atom illustrated with reference to the following figure.

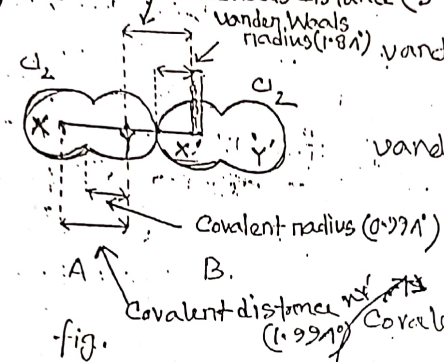

fig.

vander Waals ~~radius~~ distance $= X'Y = 3\cdot60\,A^\circ$

vander Waals radius $= \dfrac{X'Y}{2}$

$= \dfrac{3\cdot60}{2} = 1\cdot80\,A^\circ$

Covalent radius $= \dfrac{XY}{2} = \dfrac{1\cdot99}{2} = 0\cdot99\,A^\circ$

Internuclear distance $= XY = 1\cdot99\,A^\circ$

It is evident from this figure that half of the distance between the nuclii Y and X' of the two non-bonded neighbouring chlorine atoms of adjacent molecules A and B is the vander Waals radius of chlorine atom, Half of the distance between the two nuclii, X and Y in the same molecule (say, A molecule) is the covalent radius of chlorine atom. As can be seen from the figure, covalent radius is smaller than the vander Waals radius.

Explain why; Ionic radius decreases as atomic number increases ~~ato~~ along a period?

We know that the number of shells in all the elements of a given period remains the same but the value of effective nuclear charge, as calculated by Slater's rule, increases with the increase of atomic number from left to right in that period of the periodic table. The increased effective nuclear charge pulls the electron cloud of the atom nearer to the nucleus and thus the size of the atoms decreases from left to right in the same period of the periodic table.

This fact can be illustrated by the table as shown below;

Period 2	Li	Be	B	C	N
atomic number	3	4	5	6	7
atomic radius (A°)	1.23	0.90	0.82	0.77	0.75

[Ex]plain why atomic radius increases as the atomic [n]umber increases along a group ?

On moving downwards in a group from top to bottom the atomic radius of the elements increases with the increase of atomic number.

We have seen that, on descending a group the magnitude of the effective nuclear charge acting on the outermost-shell electron of the elements remain the same (the first element is the typical case). So, the concept of effective nuclear charge can not be used to explain the successive increase in the atomic radius of a given group. However, on proceeding in a group the number of shell or principle quantum number increases from 2 to 6 and hence the outermost-shell electrons get farther and farther away from the nucleus and hence atomic radii increases on descending in the same group. [F]or example, the atomic radius of alkali metals increases on proceeding from Li to Cs as shown below ;

[A]lkali [m]etals	Effective nuclear charge	No. of shell or, principle quantum no.	atomic radius (Å)	Electronic configuration
Li (3)	1.30	2	1.23	2, 1
[N]a (11)	2.20	3	1.54	2, 8, 1
[K] (19)	2.20	4	2.03	2, 8, 8, 1
[R]b (37)	2.20	5	2.16	2, 8, 8, 18, 8, 1
[Cs] (55)	2.20	6	2.35	2, 8, 18, 18, 8, 1

1. Why noble gas elements have larger atomic radius?

Ans:

Covalent bond is formed between ~~to~~ two atoms by overlapping each other, on the other hand, by vander Waals forces atoms/molecules come closer to each other but do not enter into overlapping. Consequently the vander Waals radius of an element is larger than its covalent radius.

Noble gases normally do not combine or overlap with each other and with another atoms. So, covalent radius is imposible for noble gases. They are chemically inert. They never loss or gain electrons at ordinary state to form cation or anion. So, ionic radii is imposible for the noble gases. The noble gas molecules come closer by a force and this force is vander-Waals force. For this reason the only vander Waals radius is present for noble gases and atomic radius of noble gases is represented by vander Waals radius. ~~But~~ vander Waals radius is larger. So the noble gas elements have larger atomic radius.

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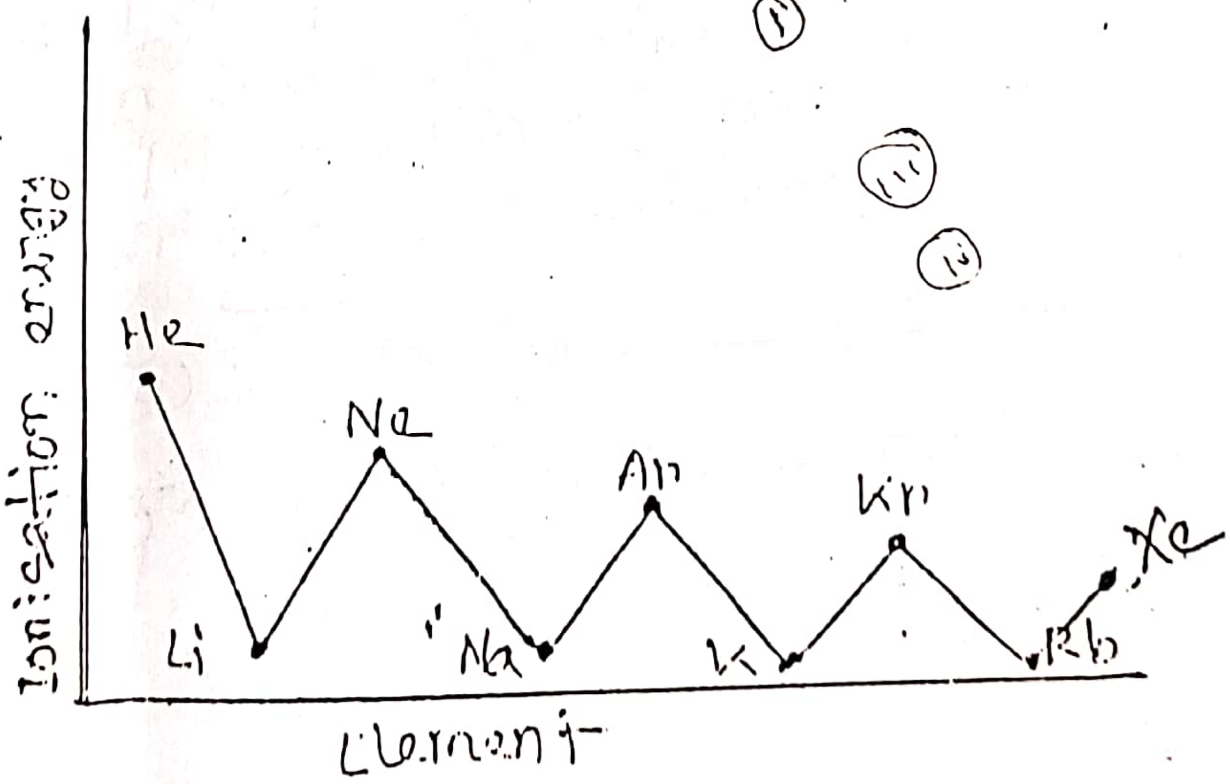
The ionisation energy for various elements are ~~represent~~ expressed by the following ~~graph~~ graph.



Def - The same electronic configuration the IP will increased with the increasing order of their own positive charge of ~~their~~ nucleus.

Explain :- On moving along a period ionisation energy of elements increase from left to right and on moving on a group ionisation energy of elements decrease from top to bottom.

In a period :—

~~Amongst~~ atoms with similar outer electronic configuration, the ionisation energy increase with the increase in nuclear charge and the value generally increases in moving from left to right in a period.

2nd period	: Li	Be	B	C	N	O	F	Ne
Nu. charge	: +3	+4	+5	+6	+7	+8	+9	+10
I.E (ev)	: 5.4	9.3	8.3	11.3	14.5	13.6	17.4	21.6

exception   exception

→ increasing

EN of Ne

The increase in values of ionisation energy from left to right is due to the fact that with the increase in nuclear charge the electrostatic attraction holding the outer electrons towards the nucleus increases and to remove the electron is comparatively more difficult.

In a group :—

On the other hand, on moving along a group upwards, as the value of the valence electron increases, the ionisation energy decreases.

IIA group : —	Be	Mg	Ca	Sr	Ba	Ra
of innershells : —	1	2	3	4	5	6
I.(ev) : —	9.3	7.6	6.1	5.7	5.2	

→ decreasing →

... energy can be explained as follows on the basis of the concept of $\boxed{\text{effective nuclear charge}}$ ... the values of succe-

We know that the cations are formed by removing the electrons from the atom one by one.

$$\underset{5}{M(g)} \xrightarrow{-e^-} \underset{4}{M^{+1}(g)} \xrightarrow{-e^-} M^{+2}(g) \xrightarrow{-e^-} \underset{2}{M^{+n}(g)}$$

Thus a given cation has lesser number of elec-trons than its ~~parent~~ atom. ~~or cation~~ With the decrease of the number of electrons, the magnitude of the effec-tive nuclear ~~increase~~ charge increase from $M(g)$ to $M(g)^{n+}$. With the increase of the effective nuclear charge, the magnitude of the force of attraction between the nucleus and the outermost shell electron cloud increase ~~could~~

also increase as shown below : —

Atom / cation : $M(g)$ $\quad M(g)^{+}$ $\quad M(g)^{+2}$ - - - - $M(g)^{+n}$

effective nuclear charge ⟶ increasing - - - ->

Force of attraction ⟶ increasing ⟶

With the increase of force of attraction the magnitude of energy to remove the outermost shell electron increase, i.e., ionisation energy also goes on increasing from $M(g)$ to $M(g)^{n+}$ and hence,

$$I_1 < I_2 < I_3 \text{ - - - - and so on}$$

Q. Why the ionisation energy of Cl is greater than that of Br ?

Ans. We know that the greater is the atomic radius the lower will be the ionisation energy because the distance between the valence electron and positive nucleus increases with the increase of atomic radius. Thus, the removal of electron become very easy and the ionisation energy decrease. The electronic configuration of Cl and Br are as below : —

$Cl\,(17) \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^5$

$Br\,(35) \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

In the case of Cl n = 3 for the valence electron but n = 4 in case of Br. Therefore the radius of valence electron of Cl is smaller than that of Br. So, the ionisation energy of Cl is greater than that of Br.

Q. Why the ionisation energy of N is greater than that O ?

Ans. The electronic configuration of N is $1s^2 2s^2 2p^3$ and that of O is $1s^2 2s^2 2p^4$. So, in case of N it is difficult to remove an electron from the half filled $2p^3$ orbital while in case of O it is easier to remove an electron from the partial filled $2p^4$ orbital. For this reason the ionisation energy of N is greater than that of oxygen.

Similarly, the ionisation energy of Be is greater than that of B and ionisation energy of Mg is greater than that of Al.

Stable

N half filled $2p^3$

O partial filled

Be B

Mg Al

Q. Explain why :— Group '0' elements are called inert gas ?

Ans. In the atoms of '0' group elements the outer-most shell (which has ns^2p^6 configuration is) completely filled; He(2) which has $1s^2$ configuration is an exceptional '0' group element. ns^2p^6 configuration is a most stable configuration and hence these elements generally do not take part into ordinary chemical reaction. These elements have maximum ionisation energy and hence it is generally impossible to convert them into cation by removing an electron. On the other hand, the electron affinity of these elements are zero and hence they have no tendency to

Stable inter

(i) R → cat

(ii) O

(iii) no tendency

... state) to convert it into a gaseous anion, is called the electron affinity of that element."

Thus the electron affinity of an atom M(g) can be defined by the following process:

$M(g) + e^-(g) \longrightarrow M^-(g) +$ Electron affinity $(-E)$

Isolated gaseous atom — gaseous anion

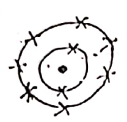
Electron affinity defined above is strictly called first electron affinity. Hence it is represented as $-E_1$ where, negative sign represents the release of energy. It is an exothermic process. If we add one more electron then energy will absorbed then the process is called endothermic process.

$M^-(g) + e^-(g) +$ energy absorbed $\longrightarrow M^{2-}(g)$

Explain why :- the electron affinity (E.A) of chlorine is greater than that of flourine ?

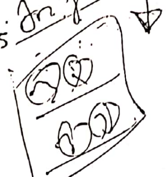
$\boxed{\text{Generally}}$ in a group of periodic table the electron affinity decreases from top to bottom. By this conception F has more electron affinity than Cl. But actually the electron affinity of Cl is greater than that of F. The size of F atom is the main factor for this.

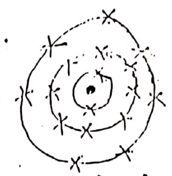
The electronic configuration of F and Cl are as follows ;



$F_9 \rightarrow \underbrace{1s^2}_{1}\ \underbrace{2s^2 2p^5}_{2}$

By this concept in general





$Cl_{17} \rightarrow \underbrace{1s^2}_{1}\ \underbrace{2s^2 2p^6}_{2}\ \underbrace{3s^2 3p^5}_{3}$

$\boxed{\text{electron density}}$

In the case of F there are 7 electrons outermost quantum no. (2). The Cl has also 7 electrons. But the quantum number is 3. So, the electron density in the outermost shell of F is greater than Cl.

This is why, the electron affinity of F is lower than that of Cl.

electron density

outermost shell F/Cl outermost shell

F/Cl $7/2$ - outer most shell

density

$7/3$ outermost shell

outermost shell

for F

Q. Explain why ~~the~~ The electron affinities of Be and ... is zero?

Ans. Beryllium and magnesium have filled s-orbitals ($1s^2 2s^2$) and the new electron will have to go to the p-orbital which has considerable high energy and as such the electron affinity is zero. Similarly for the inert gases, the s and p orbitals are completely filled and the incoming electron has to go to with ~~larger~~ value of principle quantum number n.

Explain about the exceptional case of ~~ele~~ the electron affinity of nitrogen and ~~pos~~ phosphorus.

IP

Write a short note on ionisation energy or ionisation potential.

$\Delta x \times \Delta p \geq \frac{h}{4\pi}$

Ionisation energy:—

The amount of energy required to remove the most loosely bound electron (i.e., outermost shell electron) from isolated neutral gaseous atom of an element in its lowest energy state (i.e., ground state) to convert it into a gaseous cation, is called the ionisation energy or ionisation potential of that element. It is generally represented as I.E.

$$\underset{\text{Isolated gaseous atom}}{M(g)} + I.E. \longrightarrow \underset{\text{Gaseous cation}}{M^+(g)} + e^-(g)$$ removed infinite energy away from the nucleus

This ionisation energy is usually measured in electron volt (eV) or kilocalories (kcal). The electrons are removed by stages one by one. The energy corresponding to the removal of the first electron is known as the first ionisation energy. Then when one more electron is removed, we have second, similarly stepwise we have third, fourth . . . . ionisation energies. Thus,

$$M(g) + \text{first I.E.}\ (I_1) \longrightarrow M^+(g) + e^-(g)$$

$$M^+(g) + \text{Second I.E.}\ (I_2) \longrightarrow M^{2+}(g) + e^-(g)$$

Ionisation energy depends on some particular conditions; they are as follows;

Relative — আপেক্ষিক, সম্পর্ক যুক্ত, সংযুক্ত

tendency — ঝোঁ

attract — আকর্ষ

প্রলুব্ধ

Write a short note on electronegativity.

Electronegativity :—

When two different atoms are bonded together by a covalent in a molecule participating in the formation of covalent bond, the electron pair forming the covalent bond is not shared equally by both the atoms. The electron pair lies nearer to one atom than the other atom. The relative tendency of a bonded atom in a molecule to attract the shared electron pair towards itself is called the electronegativity of the atom.

Electronegativity depends on the number of outer-most shell electron. Those elements are more electronegative in which the outer-most shell electrons are nearer to the maximum holding capacity. So, the elements of halogen family are most electronegative.

Electronegativity is a periodic property. In the same period the electronegativity increases from left to right of the long form of periodic table. Because atomic radius decreases gradually with the increase of atomic number, as a result the attraction of nucleus to the outer-most shell increases. On the other hand, in the same group the atomic radius of the elements increase from top to bottom of the periodic table. For this reason, the electronegativity of the elements decrease from top to bottom of the periodic table.

Q. Write a short-note on diagonal relationship.

Ans. Diagonal ~~relat~~ ~~relationship~~ is the resemblance of the properties of the elements of 2nd period with the diagonally opposite members lying in and period."

Three diagonal pairs viz Li - Mg, Be - Al, B - Si are shown below:

Group → / Period ↓	IA	IIA	IIIA	IVA
1st	H			
2nd	Li	Be	B	C
3rd	Na	Mg	Al	Si

Cause of diagonal relationship :— The two members of a given diagonal pair show diagonal relationship (i.e., similarities in properties) on the basis of the followings;

① Electronegativity :— Diagonal relationship may be explained in the light of change of electronegativity across a period and on descending a group. Electronegativity increases from left to right in the same period and decreases on descending a group in the periodic table.

Starting from Li, and moving one step to right we reach Be which is little more electronegative than Li. From Be moving one step down we reach Mg which is slightly less electronegative than Be. So, Li and Mg being both slightly less electronegative than Be have almost same value of electronegativity and