3

Classification *of* Elements and Periodicity *in* Properties

Multiple Choice Questions (MCQs)

Q. 1 Consider the isoelectronic species, Na^+, Mg^{2+}, F^- and O^{2-}. The correct order of increasing length of their radii is

(a) $F^- < O^{2-} < Mg^{2+} < Na^+$ (b) $Mg^{2+} < Na^+ < F^- < O^{2-}$
(c) $O^{2-} < F^- < Na^+ < Mg^{2+}$ (d) $O^{2-} < F^- < Mg^{2+} < Na^+$

Ans. *(b)* In case of isoelectronic species ionic radii $\propto \frac{1}{\text{atomic number}}$

The ionic radii increases as the positive charge decreases or the negative charge increases.

Ion	Mg^{2+}	$< Na^+$	$< F^-$	$< O^{2-}$
Atomic number	(12)	(11)	(9)	(8)

Q. 2 Which of the following is not an actinoid?

(a) Curium ($Z = 96$) (b) Californium ($Z = 98$)
(c) Uranium ($Z = 92$) (d) Terbium ($Z = 65$)

Ans. *(d)* Elements with atomic number, $Z = 90$ to 103 are called actinoids. Thus, terbium ($Z = 65$) is not an actinoid. Terbium belong to lanthanoids.

Q. 3 The order of screening effect of electrons of *s, p, d* and *f* orbitals of a given shell of an atom on its outer shell electrons is

(a) $s > p > d > f$ (b) $f > d > p > s$ (c) $p < d < s > f$ (d) $f > p > s > d$

Thinking Process

To solve this question, keep in mind that shielding effect represent the repulsive force felt by the valence shell from the electrons presents in the inner shells.

Ans. *(a)* For the same shell, screening effect decreases in the order $s > p > d > f$

Q. 4 The first ionisation enthalpies of Na, Mg, Al and Si are in the order

(a) Na < Mg > Al < Si (b) Na > Mg > Al > Si

(c) Na < Mg < Al < Si (d) Na > Mg > Al < Si

Ans. *(a) Follow the following steps to solve out such problems*

Steps	Method	Apply
Step I	Write the electronic configuration to find position in the periodic table	$_{11}Na = [Ne]\,3s^1$, $_{12}Mg = [Ne]\,3s^2$ $_{13}Al = [Ne]\,3s^2\,3p^1$, $_{14}Si = [Ne]\,3s^2\,3p^2$
Step II	Arrange them in the order as they are in the periodic table	11 12 13 14 Na Mg Al Si
Step III	Follow the general trend and also keep in mind the exception	The IP increases along a period from left to right but IP of Mg is higher than that of Al due to completely filled 3*s* orbital in Mg.
Step IV	On the above basis find the order	The order of IP is Na < Mg > Al < Si. Thus, option (a) is the correct.

Q. 5 The electronic configuration of gadolinium (Atomic number 64) is

(a) $[Xe]\,4f^3\,5d^5\,6s^2$ (b) $[Xe]\,4f^7\,5d^2\,6s^1$

(c) $[Xe]\,4f^7\,5d^1\,6s^2$ (d) $[Xe]\,4f^8\,5d^6\,6s^2$

Ans. *(c)* The electronic configuration of La ($Z = 57$) is $[Xe]\,5d^1\,6s^2$. Therefore, further addition of electrons occurs in the lower energy 4*f*-orbital till it is exactly half-filled at Eu ($Z = 63$) Thus, the electronic configuration of Eu is $[Xe]\,4f^7\,6s^2$. Thereafter, addition of next electron does not occur in the more stable exactly half-filled $4f^7$ shell but occurs in the little higher energy 5*d*-orbital. Thus, the electronic configuration of Gd ($Z = 64$) is $[Xe]\,4f^7\,5d^1\,6s^2$.

Q. 6 The statement that is not correct for periodic classification of elements is

(a) The properties of elements are periodic function of their atomic numbers

(b) Non-metallic elements are less in number than metallic elements

(c) For transition elements, the 3*d*-orbitals are filled with electrons after 3*p*-orbitals and before 4*s*-orbitals

(d) The first ionisation enthalpies of elements generally increase with increase in atomic number as we go along a period

Ans. *(c)* In case of transition elements (or any elements), the order of filling of electrons in various orbital is $3p < 4s < 3d$. Thus, 3*d* orbital is filled when 4*s* orbital gets completely filled.

Q. 7 Among halogens, the correct order of amount of energy released in electron gain (electron gain enthalpy) is

(a) F > Cl > Br > I (b) F < Cl < Br < I (c) F < Cl > Br > I (d) F < Cl < Br < I

Ans. *(c)* As we move from Cl to I, the electron gain enthalpy (*i.e.*, energy released in electron gain) become less and less negative due to a corresponding increase in the atomic size.

However, the electron gain enthalpy of F is less negative than that of Cl due to its small size. *Thus, the negative electron gain enthalpy follows the order*

F < Cl > Br > I

Q. 8 The period number in the long form of the periodic table is equal to

(a) magnetic quantum number of any element of the period
(b) atomic number of any element of the period
(c) maximum principal quantum number of any element of the period
(d) maximum azimuthal quantum number of any element of the period

Ans. *(c)* Since each period starts with the filling of electrons in a new principal quantum number, therefore, the period number in the long form of the periodic table refers to the maximum principal quantum number of any element in the period.

Period number = maximum n of any element (where, n = principal quantum number).

Q. 9 The elements in which electrons are progressively filled in $4f$-orbital are called

(a) actinoids (b) transition elements
(c) lanthanoids (d) halogens

Ans. *(c)* The elements in which electrons are progressively filled in $4f$-orbital are called lanthanoids. Lanthanoids consist of elements from Z = 58 (cerium) to 71 (lutetium).

Q. 10 Which of the following is the correct order of size of the given species

(a) $I > I^- > I^+$ (b) $I^+ > I^- > I$
(c) $I > I^+ > I^-$ (d) $I^- > I > I^+$

Ans. *(d)* Anion formed after the gain of electron to the neutral atom and cation formed after the lose of electron from outer shell. Hence, cation has smaller size but anion has bigger size than its neutral atom. Thus, $I^- > I > I^+$.

Q. 11 The formation of oxide ion O^{2-} (g), from oxygen atom requires first an exothermic and then an endothermic step as shown below

$$O\,(g) + e^- \rightarrow O^-\,(g);\ \Delta H^s = -\,141\text{ kJ mol}^{-1}$$

$$O^-\,(g) + e^- \rightarrow O^2\,(g);\ \Delta H^s = +\,780\text{kJ mol}^{-1}$$

Thus, process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that

(a) oxygen is more electronegative
(b) addition of electron in oxygen results in larger size of the ion
(c) electron repulsion outweighs the stability gained by achieving noble gas configuration
(d) O^-ion has comparatively smaller size than oxygen atom

Ans. *(c)* Although O^{2-} has noble gas configuration isoelectronic with neon but its formation is unfavourable due to the strong electronic repulsion between the negatively charged O^- ion and the second electron being added.

Hence, the electron repulsion outweighs the stability gained by achieving noble gas configuration.

Q. 12 Comprehension given below is followed by some multiple choice questions. Each question has one correct option. Choose the correct option. In the modern periodic table, elements are arranged in order of increasing atomic numbers which is related to the electronic configuration. Depending upon the type of orbitals receiving the last electron, the elements in the periodic table have been divided into four blocks, *viz* *s*, *p*, *d* and *f*.

The modern periodic table consists of 7 periods and 18 groups. Each period begins with the filling of a new energy shell. In accordance with the Aufbau principle, the seven periods (1 to 7) have 2, 8, 8, 18, 18, 32 and 32 elements respectively.

The seventh period is still incomplete. To avoid the periodic table being too long, the two series of *f*-block elements, called lanthanoids and actinoids are placed at the bottom of the main body of the periodic table

(i) The element with atomic number 57 belongs to

(a) *s* - block (b) *p* - block (c) *d* - block (d) *f* - block

Ans. *(c)* The element with atomic number 57 belongs to *d*-block element as the last electron enters the 5*d*-orbital against the aufbau principle. This anomalous behaviour can be explained on the basis of greater stability of the xenon (inert gas) core.

After barium (Z = 56) , the addition of the next electron (*i.e.*, 57th) should occur in 4*f*-orbital in accordance with aufbau principle. This will however, tend to destabilize the xenon core (Z = 54), [Kr] ($4d^{10}\ 4f^0\ 5s^2\ 5p^6\ 5d^0$) since the 4*f*-orbitals lie inside the core.

Therefore, the 57th electron prefers to enter 5*d*-orbital which lies outside the xenon core and whose energy is only slightly higher than that of 4*f*-orbital. In doing so, the stability conferred on the atom due to xenon core more than compensates the slight instability caused by the addition of one electron to the higher energy 5*d*- orbital instead of the lower energy 4*f*- orbital.

Thus, the outer electronic configuration of La(Z =57) is $5d^1\ 6s^2$ rather than the expected $4f^1\ 6s^2$.

(ii) The last element of the *p*-block in 6th period is represented by the outermost electronic configuration.

(a) $7s^2\ 7p^6$ (b) $5f^{14}\ 6d^{10}\ 7s^2\ 7p^0$
(c) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$ (d) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$

Ans. *(c)* Each period starts with the filling of electrons in a new principal energy shell. Therefore, 6th period starts with the filling of 6*s*-orbital and ends when 6 *p*-orbitals are completely filled.

In between 4*f* and 5*d*-orbitals are filled in accordance with aufbau principle. Thus, the outmost electronic configuration of the last element of the *p*-block in the 6th period is $6s^2\ 4f^{14}\ 5d^{10}\ 6p^6$ or $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$.

(iii) Which of the elements whose atomic numbers are given below, cannot be accommodated in the present set up of the long form of the periodic table?

(a)107 (b) 118 (c) 126 (d) 102

Ans. *(c)* The long form of the periodic table contain element with atomic number 1 to 118.

(iv) The electronic configuration of the element which is just above the element with atomic number 43 in the same group is ....

(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

(b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^3\ 4p^6$

(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$

(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$

Ans. *(a)* The fifth period begins with Rb ($Z = 37$) and ends at Xe ($Z = 54$). Thus, the element with $Z = 43$ lies in the 5th period. Since, the 4th period has 18 elements, therefore, the atomic number of the element which lies immediately above the element with atomic number 43 is $43 - 18 = 25$.

Now, the electronic configuration of element with $Z = 25$ is

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2 \text{ (i.e., Mn).}$$

(v) The elements with atomic numbers 35, 53 and 85 are all ....

(a) noble gases
(b) halogens
(c) heavy metals
(d) light metals

Ans. *(b)* Each period ends with a noble gas. The atomic number of noble gases (*i.e.*, group 18 elements) are 2, 10, 18, 36, 54 and 86. Therefore, elements with atomic numbers 35 (36 – 1), 53 (54 – 1), and 85 (86 – 1), lie in a group before noble gases, *i.e.*, halogens (group 17) elements.

Thus, the elements with atomic number 35, 53 and 85 are all belongs to halogens.

Q. 13 Electronic configuration of four elements *A, B, C,* and *D* are given below

A. $1s^2\ 2s^2\ 2p^6$
B. $1s^2\ 2s^2\ 2p^4$
C. $1s^2\ 2s^2\ 2p^6\ 3s^1$
D. $1s^2\ 2s^2\ 2p^5$

Which of the following is the correct order of increasing tendency to gain electron?

(a) $A < C < B < D$
(b) $A < B < C < D$
(c) $D < B < C < A$
(d) $D < A < B < C$

Ans. *(a)* Electronic configuration of elements indicate that *A* is a noble gas (*i.e.*, Ne), *B* is oxygen (group 16), *C* is sodium metal (group 1) and *D* is fluorine (group 17).

(i) Noble gases have no tendency to gain electrons since all their orbitals are completely filled. Thus, element *A* has the least electron gain enthalpy.

(ii) Since, element *D* has one electron less and element *B* has two electrons less than the corresponding noble gas configuration, hence, element *D* has the highest electron, gain enthalpy followed by element *B*

(iii) Since, element C has one electron in the *s*-orbital and hence needs one more electron to complete it, therefore, electron gain enthalpy of *C* is less than that of element *B*. Combining all the facts given above, the electron gain enthalpies of the four elements increase in the order $A < C < B < D$.

Multiple Choice Questions (More Than One Options)

Q. 14 Which of the following elements can show covalencey greater than 4?

(a) Be (b) P (c) S (d) B

Ans. *(b, c)*

Elements Be and B lie in the 2nd period. They can have a maximum of 8 electrons in the valence shell. In other words, they can have a maximum covalency of 8/2 = 4.

However, elements P and S have vacant *d*- orbitals in their respective valence shells and hence can accommodate more than 8 electrons in their respective valence shell. In other words, they can show a covalency of more than 4.

Q. 15 Those elements impart colour to the flame on heating in it, the atoms of which require low energy for the ionisation (*i.e.*, absorb energy in the visible region of spectrum). The elements of which of the following groups will impart colour to the flame ?

(a) 2 (b) 13 (c) 1 (d) 17

Ans. *(a, c)*

The elements of group 1 (alkali metals) and group 2 (alkaline earth metals) have 1 and 2 electrons respectively in their valence shells and hence have low ionisation energies. In other words, elements of group 1 and 2 imparts colour to the flame.

Group 1	Colour	Group 2	Colour
Li	Crimson	Ca	Brick-red
Na	Yellow	Sr	Crimson red
K	Pale violet	–	–
Rb	Red violet	Ba	Apple green
Cs	Blue	Ra	Crimson

Q. 16 Which of the following sequences contain atomic numbers of only representative elements?

(a) 3, 33, 53, 87 (b) 2, 10, 22, 36

(c) 7, 17, 25, 37, 48 (d) 9, 35, 51, 88

Ans. *(a, d)*

Elements of *s* and *p*-block elements are called representative elements. Elements of *d*-block (Z = 21-30; 39-48; 57 and 72-80; 89 and 104-112) are called transition elements while those of *f*-block (with Z = 58-71 and Z = 90-103) are called inner transition elements.

With reference to this division, elements listed under option (a) and option (*d*) are all representative elements.

Q. 17 Which of the following elements will gain one electron more readily in comparison to other elements of their group?

(a) S(g) (b) Na(g) (c) O(g) (d) Cl(g)

Ans. *(a, d)*

Chlorine has the highest tendency to gain one electron because by doing so, it acquires the stable electronic configuration of the nearest noble gas, *i.e.*, argon. Sulphur and oxygen belong to group 16 but the size of oxygen is much smaller than that of sulphur.

As a result, when an electron is added to them, the electron-electron repulsions in the smaller 2*p*- subshell of oxygen are comparatively stronger than those present in the bigger 3*p*-subshell of sulphur. Therefore, S has a higher tendency to gain an electron than O.

Na, on the other hand, has only one electron in the valence shell and hence has a strong tendency to lose rather than gain one electron.

Q. 18 Which of the following statements are correct?

(a) Helium has the highest first ionisation enthalpy in the periodic table
(b) Chlorine has less negative electron gain enthalpy than fluorine
(c) Mercury and bromine are liquids at room temperature
(d) In any period, atomic radius of alkali metal is the highest

Ans. *(a, c, d)*

Chlorine has more negative electron gain enthalpy than fluorine. Therefore, all other given statements are correct.

Q. 19 Which of the following sets contain only isoelectronic ions?

(a) Zn^{2+}, Ca^{2+}, Ga^{3+}, Al^{3+}
(b) K^+, Ca^{2+}, Sc^{3+}, Cl^-
(c) P^{3-}, S^{2-}, Cl^-, K^+
(d) Ti^{4+}, Ar, Cr^{3+}, V^{5+}

Thinking Process

Isoelectronic represents those species which have same number of electrons.

Ans. *(b, c)*

(a) Zn^{2+} $(30-2=28)$, Ca^{2+} $(20-2=18)$, Ga^{3+} $(31-3=28)$, Al^{3+} $(13-3=10)$.
These species have different number of electrons and hence are not isoelectronic ions.

(b) K^+ $(19-1=18)$, Ca^{2+} $(20-2=18)$, Sc^{3+} $(21-3=18)$, Cl^- $(17+1=18)$.
These are all isoelectronic ions since each one of them has 18 electrons.

(c) P^{3-} $(15+3=18)$, S^{2-} $(16+2=18)$, Cl^- $(17+1=18)$, K^+ $(19-1=18)$.
These are all isoelectronic ions since each one of them has 18 electrons.

(d) Ti^{4+} $(22-4=18)$, Ar(18), Cr^{3+} $(24-3=21)$, V^{5+} $(23-5=18)$.
These have different number of electrons and hence are not isoelectronic ions.

Q. 20 In which of the following options order of arrangement does not agree with the variation of property indicated against it?

(a) $Al^{3+} < Mg^{2+} < Na^+ < F^-$ (Increasing ionic size)
(b) $B < C < N < O$ (Increasing first ionisation enthalpy)
(c) $I < Br < Cl < F$ (Increasing electron gain enthalpy)
(d) $Li < Na < K < Rb$ (Increasing metallic radius)

Thinking Process

(i) The ionic size increases as the positive charge on the cation decreases or the negative charge on the anion increases.
(ii) First ionisation enthalpy increases from left to right in the periodic table.
(iii) Electron gain enthalpy increases as the electronegativity of the atom increases.
(iv) The metallic character increases as the size of the metal atom increases.

Ans. *(b, c)*

Due to greater stability of the half filled electronic configuration of nitrogen, its ionisation enthalpy is higher than that of oxygen. Thus, option (b) is incorrect.

Due to stronger electron-electron repulsions in the small size of flourine the negative electron gain enthalpy of fluorine is lower than that of chlorine. Hence, option (c) is incorrect.

Q. 21 Which of the following have no unit?

(a) Electronegativity (b) Electron gain enthalpy
(c) Ionisation enthalpy (d) Metallic character

Ans. ***(a, d)***

Electronegativity and metallic character have no units while electron gain enthalpy and ionisation enthalpy have units of kJ mol^{-1}.

Q. 22 Ionic radii vary in

(a) inverse proportion to the effective nuclear charge
(b) inverse proportion to the square of effective nuclear charge
(c) direct proportion to the screening effect
(d) direct proportion to the square of screening effect

Ans. ***(a, c)***

Ionic radius decreases as the effective nuclear charge increases.

$$\text{Ionic radius} \propto \frac{1}{\text{effective nuclear charge}}$$

Further, ionic radius increases as the screening effect increases.

$$\text{Ionic radius} \propto \text{screening effect}$$

Q. 23 An element belongs to 3rd period and group 13 of the periodic table. Which of the following properties will be shown by the element?

(a) Good conductor of electricity (b) Liquid, metallic
(c) Solid, metallic (d) Solid, non-metallic

Ans. ***(a, c)***

Except boron, all elements of groups 13 are metallic. These exists as solid. Being metallic in nature, aluminium is good conductor of electricity.

Short Answer Type Questions

Q. 24 Explain why the electron gain enthalpy of fluorine is less negative than that of chlorine?

Ans. Electron gain enthalpy of F is less negative than that of Cl because when an electron is added to F, the added electron goes to the smaller $n = 2$ quantum level and suffers repulsion from other electrons present in this level.

In case of Cl, the added electron goes to the larger $n = 3$ quantum level and suffers much less repulsion from other electrons.

Q. 25 All transition elements are d- block elements, but all d- block elements are not transition elements. Explain.

Ans. Elements in which the last electron enters in the d-orbitals, are called d-block elements or transition elements. These elements have the general outer electronic configuration $(n-1)d^{1-10}ns^{0-2}$. Zn, Cd and Hg having the electronic configuration $(n-1)d^{10}ns^2$ do not show most of the properties of transition elements.

The d-orbitals in these elements are completely filled in the ground state as well as in their common oxidation states. Therefore, they are not regarded as transition elements.

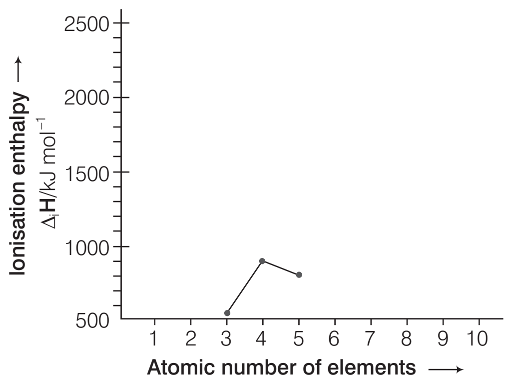
Thus, on the basis of properties, all transition elements are d-block elements but on the basis of electronic configuration, all d-block elements are not transition elements.

Q. 26 Identify the group and valency of the element having atomic number 119. Also predict the outermost electronic configuration and write the general formula of its oxide.

Ans. The present set up of the Long form of the periodic table can accommodate maximum 118 elements. Thus, in accordance with aufbau principle, the filling of 8s-orbital will occur. In other words 119th electron will enter 8s-orbital. As such its outmost electronic configuration will be $8s^1$.

Since, it has only one electron in the valence shell, *i.e.*, 8s, therefore, its valency will be 1 and it will lie in the group IA along with alkali metals and the formula of its oxide will be M_2O where M represents the element.

Q. 27 Ionisation enthalpies of elements of second period are given below Ionisation enthalpy/k cal mol^{-1}: 520, 899, 801, 1086, 1402, 1314, 1681, 2080. Match the correct enthalpy with the elements and complete the graph given in figure. Also write symbols of elements with their atomic number.



Ans. To match the correct enthalpy with the elements and to complete the graph the following points are taken into consideration. As we move from left to right across a period, the ionisation enthalpy keeps on increasing due to increased nuclear charge and simultaneous decrease in atomic radius.

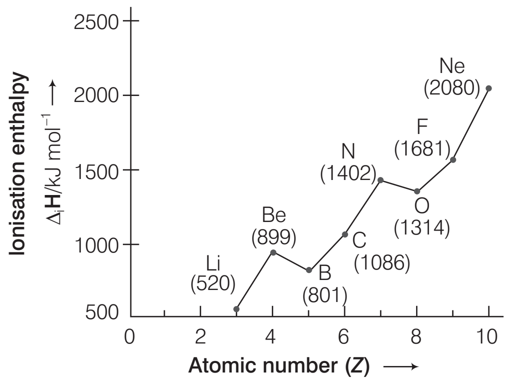
However, there are some exceptions given below

(a) In spite of increased nuclear charge, the first ionisation enthalpy of B is lower than that of Be. This is due to the presence of fully filled 2s orbital of Be [$1s^2\ 2s^2$] which is a stable electronic arrangement. Thus, higher energy is required to knock out the electron from fully filled 2s orbitals. While B [$1s^2\ 2s^2\ 2P^1$] contains valence electrons in 2s and 2p orbitals. It can easily lose its one e^- from 2p orbital in order to achieve noble gas configuration. Thus, first ionisation enthalpy of B is lower than that of Be.

Since, the electrons in 2s-orbital are more tightly held by the nucleus than these present in 2p-orbital, therefore, ionisation enthalpy of B is lower than that of Be.

(b) The first ionisation enthalpy of N is higher than that of O though the nuclear charge of O is higher than that of N. This is due to the reason that in case of N, the electron is to be removed from a more stable exactly half-filled electronic configuration $(1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1)$ which is not present in O $(1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1)$.

Therefore, the first ionisation enthalpy of N is higher than that of O. *The symbols of elements along with their atomic numbers are given in the following graph*



Q. 28 Among the elements B, Al, C and Si,
(a) which element has the highest first ionisation enthalpy?
(b) which element has the most metallic character?
Justify your answer in each case.

Ans. *The placing of elements are as*

Period	Group-13	Group-14
2nd Period	Boron	Carbon
3rd Period	Aluminium	Silicon

(a) Ionisation enthalpy increases along a period (as we move from left to right in a period) with decrease in atomic size and decreases down the group with increase in atomic size. Hence, carbon has the highest first ionisation enthalpy.

(b) Metallic character decreases across a period but increases on moving down the group. Hence, aluminium has the most metallic character.

Q. 29 Write four characteristic properties of *p*-block elements.

Ans. *The four important characteristic properties of p-block elements are the following*

(a) *p*-Block elements include both metals and non-metals but the number of non-metals is much higher than that of metals. Further, the metallic character increases from top to bottom within a group and non-metallic character increases from left to right along a period in this block.

(b) Their ionisation enthalpies are relatively higher as compared to *s*-block elements.

(c) They mostly form covalent compounds.

(d) Some of them show more than one (variable) oxidation states in their compounds. Their oxidising character increases from left to right in a period and reducing character increases from top to bottom in a group.

Q. 30 **Choose the correct order of atomic radii of fluorine and neon (in pm) out of the options given below and justify your answer.**

(a) 72, 160 (b) 160, 160 (c) 72, 72 (d) 160, 72

Ans. *(a)* Atomic radius of F is expressed in terms of covalent radius while, atomic radius of neon is usually expressed in terms of van der Waals' radius. van der Waals' radius of an element is always larger than its covalent radius.

Therefore, atomic radius of F is smaller than atomic radius of

Ne (F = 72 pm, Ne = 160 pm).

Q. 31 **Illustrate by taking examples of transition elements and non-transition elements that oxidation states of elements are largely based on electronic configuration.**

Ans. Oxidation state of an element depends upon the electrons present in the outermost shell or eight minus the number of valence shell electrons (outermost shell electrons). *e.g.,*

Alkali metals (Group 1 elements) General valence shell electronic configuration —ns^1; Oxidation state = + 1.

Alkaline earth metals (Group 2 elements) General valence shell electronic configuration —ns^2; Oxidation state = + 2 .

Alkali metals and alkaline earth metals belong to *s*-block elements and elements of group 13 to group 18 are known as *p*-block elements.

Group 13 elements General valence shell electronic configuration —$ns^2\ np^1$; Oxidation states = + 3 and + 1.

Group 14 elements General valence shell electronic configuration —$ns^2\ np^2$; Oxidation states = + 4 and + 2.

Group 15 elements General valence shell electronic configuration —$ns^2\ np^3$; Oxidation states = – 3, + 3 and + 5. Nitrogen shows + 1, + 2, + 4 oxidation states also.

Group 16 elements General valence shell electronic configuration —$ns^2\ np^4$; Oxidation states = – 2, + 2, + 4 and + 6.

Group 17 elements General valence shell electronic configuration —ns^2np^5; Oxidation states = – 1. Cl, Br and I also show + 1, + 3, + 5 and + 7 oxidation states.

Group 18 elements General valence shell configuration —ns^2np^6. Oxidation state = zero.

Transition elements or d-block elements General electronic configuration – $(n-1)d^{1-10}\,ns^{1-2}$. These elements show variable oxidation states due to involvement of not only *ns* electrons but *d* or *f*-electrons (inner-transition elements) as well. Their most common oxidation states are + 2 and + 3.

Q. 32 **Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionisation enthalpy than nitrogen. Explain.**

Ans. Electronic configuration of $_7N = 1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$. Nitrogen has stable configuration because *p*-orbital is half-filled. Therefore, addition of extra electron to any of the *p*-orbital requires energy.

Electronic configuration of $_8O = 1s^2, 2s^2, 2p_x^2, 2p_y^1, 2p_z^1$. Oxygen has $2p^4$ electrons, so process of adding an electron to the *p*-orbital is exothermic.

Oxygen has lower ionisation enthalpy than nitrogen because by removing one electron from $2p$-orbital, oxygen acquires stable configuration, *i.e.,* $2p^3$. On the other hand, in case of nitrogen it is not easy to remove one of the three $2p$-electrons due to its stable configuration.

Q. 33 First member of each group of representative elements (*i.e., s* and *p*-block elements) shows anomalous behaviour. Illustrate with two examples.

Ans. First member of each group of representative elements (*i.e., s*- and *p*-block elements) shows anomalous behaviour due to (i) small size (ii) high ionisation enthalpy (iii) high electronegativity and (iv) absence of *d*- orbitals. *e.g.*, in *s*-block elements, lithium shows anomalous behaviour from rest of the alkali metals.

(a) Compounds of lithium have significant covalent character. While compounds of other alkali metals are predominantly ionic.

(b) Lithium reacts with nitrogen to form lithium nitride while other alkali metals do not form nitrides.

In *p*-block elements, first member of each group has four orbitals, one 2*s*- orbital and three 2*p*-orbitals in their valence shell. So, these elements show a maximum covalency of four while other members of the same group or different group show a maximum covalency beyond four due to availability of vacant *d*- orbitals.

Q. 34 *p*-block elements form acidic, basic and amphoteric oxides. Explain each property by giving two examples and also write the reactions of these oxides with water.

Ans. In *p*-block, when we move from left to right in a period, the acidic character of the oxides increases due to increase in electronegativity. *e.g.,*

(i) **2nd period** $B_2O_3 < CO_2 < N_2O_3$ acidic nature increases.

(ii) **3rd period** $Al_2O_3 < SiO_2 < P_4O_{10} < SO_3 < Cl_2O_7$ acidic character increases.

On moving down the group, acidic character decreases and basic character increases. *e.g.,*

(a) **Nature of oxides of 13 group elements**

$$\underset{\text{Weakly acidic}}{B_2O_3} \quad \underbrace{Al_2O_3 \quad Ga_2O_3}_{\text{Amphoteric}} \quad \underset{\text{Basic}}{In_2O_3} \quad \underset{\text{Strongly basic}}{Tl_2O}$$

(b) **Nature of oxides of 15 group elements**

$$\underset{\text{Strongly acidic}}{N_2O_5} \quad \underset{\text{Moderately acidic}}{P_4O_{10}} \quad \underset{\text{Amphoteric}}{As_4O_{10}} \quad \underset{\text{Amphoteric}}{Sb_4O_{10}} \quad \underset{\text{Basic}}{Bi_2O_3}$$

Among the oxides of same element, higher the oxidation state of the element, stronger is the acid. *e.g.,* SO_3 is a stronger acid than SO_2.

B_2O_3 is weakly acidic and on dissolution in water, it forms orthoboric acid. Orthoboric acid does not act as a protonic acid (it does not ionise) but acts as a weak Lewis acid.

$$\underset{\text{Boron trioxide}}{B_2O_3} + 3H_2O \rightleftharpoons \underset{\text{Orthoboric acid}}{2H_3BO_3}$$

$$B(OH)_3 + H{-}OH \longrightarrow [B(OH)_4]^- + H^+$$

Al_2O_3 is amphoteric in nature. It is insoluble in water but dissolves in alkalies and reacts with acids.

$$\underset{\text{Aluminium trioxide}}{Al_2O_3} + 2NaOH \xrightarrow{\Delta} \underset{\text{Sodium meta aluminate}}{2NaAlO_2} + H_2O$$

$$\longleftarrow Al_2O_3 + 6HCl \xrightarrow{\Delta} \underset{\text{Aluminium chloride}}{2AlCl_3} + 3H_2O$$

Tl_2O is as basic as NaOH due to its lower oxidation state (+ 1).

$$Tl_2O + 2HCl \longrightarrow 2TlCl + H_2O$$

P_4O_{10} on reaction with water gives orthophosphoric acid

$$\underset{\text{Phosphorus pentoxide}}{P_4O_{10}} + 6H_2O \longrightarrow \underset{\text{Orthophosphoric acid}}{4H_3PO_4}$$

Cl_2O_7 is strongly acidic in nature and on dissolution in water, it gives perchloric acid.

$$\underset{\text{Dichlorine heptoxide}}{Cl_2O_7} + H_2O \longrightarrow \underset{\text{Perchloric acid}}{2HClO_4}$$

Q. 35 **How would you explain the fact that the first ionisation enthalpy of sodium is lower than that of magnesium but its second ionisation enthalpy is higher than that of magnesium?**

Thinking Process

The species having exactly half-filled or fully filled orbitals have extra ordinarily high ionisation enthalpies.

Ans. First ionisation enthalpy of sodium (Na = $1s^2, 2s^2, 2p^6, 3s^1$) is lower than that of magnesium (Mg = $1s^2, 2s^2\ 2p^6, 3s^2$) because the electron to be removed in both the cases is from 3*s*-orbital but the nuclear charge is lower in Na than that of magnesium.

After the removal of first electron Na^+ acquires inert gas (Ne) configuration ($Na^+ = 1s^2, 2s^2, 2p^6$) and hence, removal of second electron from sodium is difficult.

While in case of magnesium, after the removal of first electron, the electronic configuration of Mg^+ is $1s^2, 2s^2, 2p^6, 3s^1$. In this case $3s^1$ electron is easy to remove in comparison to remove an electron from inert gas configuration. Therefore, IE_2 of Na is higher than that of magnesium.

Q. 36 **What do you understand by exothermic reaction and endothermic reaction? Give one example of each type.**

Ans. **Exothermic reactions** Reactions which are accompanied by evolution of heat are called exothermic reactions. The quantity of heat produced is shown either along with the products with a '+ ' sign or in terms if ΔH with a '–' sign. *e.g.*,

$$C(s) + O_2(g) \longrightarrow CO_2(g) + 393.5 \text{ kJ}$$

$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O\ (l)\ ; \Delta H = -285.8 \text{ kJ mol}^{-1}$$

Endothermic reactions Reactions which proceed with absorption of heat are called endothermic reactions. The quantity of heat absorbed is shown either along with the products with a '–' sign or in terms of ΔH with a '+' sign *e.g.*,

$$C(s) + H_2O(g) \longrightarrow CO\ (g) + H_2(g) - 131.4 \text{ kJ}$$

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3\ (g)\ ; \Delta H = +92.4 \text{ kJ mol}^{-1}$$

Q. 37 **Arrange the elements N, P, O and S in the order of**

(i) increasing first ionisation enthalpy.

(ii) increasing non-metallic character.

Give reason for the arrangement assigned.

Ans. *The placing of elements are as*

Period	Group 15	Group 16
2nd period	N	O
3rd period	P	S

(i) Ionisation enthalpy of nitrogen ($_7N = 1s^2, 2s^2, 2p^3$) is greater than oxygen ($_8O = 1s^2, 2s^2, 2p^4$) due to extra stable exactly half-filled $2p$-orbitals. Similarly, ionisation enthalpy of phosphorus ($_{15}P = 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$) is greater than sulphur ($_{16}S = 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$).

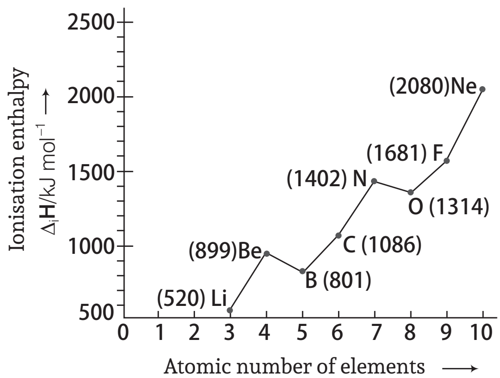
On moving down the group, ionisation enthalpy decreases with increasing atomic size. So, the order is

$$S < P < O < N \rightarrow \text{First ionisation enthalpy increases.}$$

(ii) Non-metallic character across a period (left to right) increases but on moving down the group it decreases. So, the order is

$$P < S < N < O \rightarrow \text{Non-metallic character increases.}$$

Q. 38 **Explain the deviation in ionisation enthalpy of some elements from the general trend by using given figure.**



Ans. There is deviation of ionisation enthalpy of some elements from the general trend as shown in figure. The first ionisation enthalpy of B is lower than that of Be and in case of nitrogen, the first ionisation enthalpy is higher than that of O. (*Also, refer to Q. 27*)

Q. 39 **Explain the following**

(a) Electronegativity of elements increase on moving from left to right in the periodic table.

(b) Ionisation enthalpy decrease in a group from top to bottom.

Ans. (a) Across the period, the nuclear charge increases and the atomic radius decreases. As a result, the tendency of the atom of an element to attract the shared pair of electrons towards itself increases and hence the electronegativity of the element increases. *e.g.,* electronegativity of the elements of the 2nd period increases regularly from left to right as follows Li (1.0), Be (1.5), B (2.0) , C (2.5), N (3.0), O (3.5) and .

(b) *The ionisation enthalpy decreases regularly as we move from top to bottom, as explained below*

(i) On moving down a group from top to bottom, the atomic size increases gradually due to the addition of a new principal energy shell at each succeding element. As a result, the distance between the nucleus and the valence shell increases.

In other words, the force of attraction of the nucleus for the valence electrons decreases and hence the ionisation enthalpy should decrease.

(ii) With the addition of new shells, the number of inner shell which shield the valence electrons from the nucleus increases. In other words, the shielding effect or the screening effect increases.

As a result, the force of attraction of the nucleus for the valence electrons further decreases and hence the ionisation enthalpy should decrease.

(iii) Further, in a group from top to bottom nuclear charge increases with increase in atomic number. As a result, the force of attraction of the nucleus for the valence electrons increases and hence the ionisation enthalpy should increase.

The combined effect of the increase in atomic size and screening effect more than compensate the effect of the increased nuclear charge. Consequently, the valence electrons become less and less firmly held by the nucleus and hence the ionisation enthalpy gradually decreases as we move down the group.

Q. 40 How does the metallic and non-metallic character vary on moving from left to right in a period?

Ans. As we move from left to right in a period, the number of valence electrons increases by one at each succeeding element but the number of shells remains same. Due to this effective nuclear charge increases. More is the effective nuclear charge, more is the attraction between nuclei and electron.

Hence, the tendency of the element to lose electrons decreases, this results in decrease in metallic character. Furthermore, the tendency of an element to gain electrons increases with increase in effective nuclear charge, so non-metallic character increases on moving from left to right in a period.

Q. 41 The radius of Na^+ cation is less than that of Na atom. Give reason.

Ans. When an atom loses an electron to form cation, its radius decreases. In a cation, per electron nuclear forces increases due to decrease in number of electrons. As a result of this, effective nuclear charge increases and the radius of cation decreases. *e.g.,* ionic radius of Na^+ is smaller than the radius of its parent atom Na.

	Na $\longrightarrow$	$Na^+ + 1e^-$
Electrons	11	10
Nuclear charge	11	11
Ionic size	186pm	95pm

Q. 42 Among alkali metals which element do you expect to be least electronegative and why?

Ans. On moving down the group, electronegativity decreases because atomic size increases. Fr has the largest size, therefore it is least electronegative.

Matching The Columns

Q. 43 Match the correct atomic radius with the element.

Element	Atomic radius (pm)
Be	74
C	88
O	111
B	77
N	66

Thinking Process

(i) All the element given in the question belong to the second period of the periodic table.

(ii) Atomic radius is the distance from the centre of the nucleus to the point upto which the density of the electron cloud (i.e., probability of finding the electron) is maximum.

Ans. All the given elements are of same period and along a period, atomic radii decreases because effective nuclear charge increases. Thus, the order of atomic radii is

O < N < C < B < Be or, Be = 11 pm, O = 66 pm, C = 77 pm, B = 88 pm, N = 74 pm.

Q. 44 Match the correct ionisation enthalpies and electron gain enthalpies of the following elements.

	Elements		ΔH_1	ΔH_2	$\Delta_{eg} H$
(i)	Most reactive non-metal	A.	419	3051	− 48
(ii)	Most reactive metal	B.	1681	3374	− 328
(iii)	Least reactive element	C.	738	1451	− 40
(iv)	Metal forming binary halide	D.	2372	5251	+ 48

Ans. (i) Most reactive non-metal has high $\Delta_i H_1$ and $\Delta_i H_2$ and most negative $\Delta_{eg} H$. Therefore, the element is *B*.

(ii) Most reactive metal has low $\Delta_i H_1$ and high $\Delta_i H_2$ (because the second electron has to be lost from noble gas configuration) and has small negative $\Delta_{eg} H$. Therefore, the element is *A*.

(iii) Noble gases are the least reactive elements. They have very high $\Delta_i H_1$ and $\Delta_i H_2$ and have positive $\Delta_{eg} H$ values. Thus, the element is *D*.

(iv) Metal forming binary halides are alkaline earth metals. They have $\Delta_i H_1$ and $\Delta_i H_2$ values little higher than those of most reactive metals (such as *A*) and have comparatively slightly less negative $\Delta_{eg} H$ values. Thus, the element is *C*.

Q. 45 Electronic configuration of some elements is given in Column I and their electron gain enthalpies are given in Column II. Match the electronic configuration with electron gain enthalpy.

Column I (Electronic configuration)		Column II (Electron gain enthalpy/ $kJ\,mol^{-1}$)
A.	$1s^2\ 2s^2\ 2p^6$	-53
B.	$1s^2\ 2s^2\ 2p^6\ 3s^1$	-328
C.	$1s^2\ 2s^2\ 2p^5$	-141
D.	$1s^2\ 2s^2\ 2p^4$	$+48$

Ans. **A.** $\rightarrow$ (4) **B.** $\rightarrow$ (1) **C.** $\rightarrow$ (2) **D.** $\rightarrow$ (3)

A. This electronic configuration corresponds to the noble gas *i.e.*, neon. Since, noble gases have $+\Delta_{eg} H$ values, therefore, electronic configuration (A) corresponds to the $\Delta_{eg} H = +48\ kJ\ mol^{-1}$.

B. This electronic configuration corresponds to the alkali metal *i.e.*, potassium. Alkali metals have small negative $\Delta_{eg} H$ values, hence, electronic configuration (B) corresponds to $\Delta_{eg} H = -53\ kJ\ mol^{-1}$.

C. This electronic configuration corresponds to the halogen *i.e.*, fluorine. Since, halogens have high negative $\Delta_{eg} H$ values, therefore, electronic configuration (C) corresponds to $\Delta_{eg} H = 328\ kJ\ mol^{-1}$.

D. This electronic configuration corresponds to the chalcogen *i.e.*, oxygen. Since, chalcogens have $\Delta_{eg} H$ values less negative than those of halogens, therefore, electronic configuration (D) corresponds to $\Delta_{eg} H = -\ 141 kJmol^{-1}$.

Assertion and Reason

In the following questions a statement of Assertion (A) *followed by a statement of Reason* (R) *is given. Choose the correct option out of the choices given below in each question.*

Q. 46 Assertion (A) Generally, ionisation enthalpy increases from left to right in a period.

Reason (R) When successive electrons are added to the orbitals in the same principal quantum level, the shielding effect of inner core of electrons does not increase very much to compensate for the increased attraction of the electron to the nucleus.

(a) Assertion is correct statement and reason is wrong statement.
(b) Assertion and reason both are correct statements and reason is correct explanation of Assertion.
(c) Assertion and reason both are wrong statements.
(d) Assertion is wrong statement and reason is correct statement.

Ans. *(b)* Assertion and reason both are correct statements and reason is correct explanation of assertion. Ionisation enthalpy increases along a period because effective nuclear charge increases and atomic size decreases.

Q. 47 Assertion (A) Boron has a smaller first ionisation enthalpy than beryllium.

Reason (R) The penetration of 2*s* electron to the nucleus is more than the 2*p* electron hence 2*p* electron is more shielded by the inner core of electrons than the 2*s* electrons.

(a) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

(b) Assertion is correct statement but reason is wrong statement.

(c) Assertion and reason both are correct statements and reason is correct explanation for Assertion.

(d) Assertion and reason both are wrong statements.

Ans. *(c)* Assertion and reason both are correct statements and reason is correct explanation for assertion.

Boron has a smaller first ionisation enthalpy than beryllium because the penetration of a 2*s* electron to the nucleus is more than the 2*p* electron. Hence, 2*p* electron is more shielded by the inner core of electron than the 2*s* electron.

Q. 48 Assertion (A) Electron gain enthalpy becomes less negative as we go down a group.

Reason (R) Size of the atom increases on going down the group and the added electron would be farther from the nucleus.

(a) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

(b) Assertion and reason both are correct statements and reason is correct explanation for assertion.

(c) Assertion and reason both are wrong statements.

(d) Assertion is wrong statement but reason is correct statement.

Ans. *(b)* Assertion and reason both are correct statements and reason is correct explanation for assertion.

Electron gain enthalpy becomes less negative as the size of an atom increases down the group. This is because within a group screening effect increases on going downward and the added electron would be farther away from the nucleus.

Long Answer Type Questions

Q. 49 Discuss the factors affecting electron gain enthalpy and the trend in its variation in the periodic table.

Ans. Electron gain enthalpy of an element is equal to the energy released when an electron is added to valence shell of an isolated gaseous atom.

$$A(g) + e^- \longrightarrow A^-(g);\ \Delta_{eg} H = \text{negative}$$

Factors affecting electron gain enthalpy

(i) **Effective nuclear charge** Electron gain enthalpy increases with increase in effective nuclear charge because attraction of nucleus towards incoming electron increases.

(ii) **Size of an atom** Electron gain enthalpy decreases with increase in the size of valence shell.

(iii) **Type of subshell** More closer is the subshell to the nucleus, easier is the addition of electron in that subshell.

Electron gain enthalpy (in decreasing order) for addition of electron in different subshell (n-same) is $s > p > d > f$

(iv) **Nature of configuration** Half-filled and completely-filled subshell have stable configuration, so addition of electron in them is not energetically favourable.

Variation in the periodic table As a general rule, electron gain enthalpy becomes more and more negative with increase in the atomic number across a period. The effective nuclear charge increases from left to right across a period and consequently it will be easier to add an electron to a smaller atom.

Electron gain enthalpy becomes less negative as we go down a group because the size of the atom increases and the added electron would be farther from the nucleus.

Electron gain enthalpy of O or F is less than that of the succeeding element (S or Cl) because the added electron goes to the smaller $n = 2$ level and suffers repulsion from other electrons present in this level. For the $n = 3$ level (S or Cl), the added electron occupies a larger region of space and suffers much less repulsion from electrons present in this level.

Q. 50 **Define ionisation enthalpy. Discuss the factors affecting ionisation enthalpy of the elements and its trends in the periodic table.**

Ans. **Ionisation enthalpy** The minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom so as to convert it into a gaseous cation is called its ionisation enthalpy. It is represented by $\Delta_i H$.

Factors affecting ionisation enthalpy of the elements

Ionisation enthalpy depends upon the following factors

(i) **Nuclear charge** The ionisation enthalpy increases with increase in nuclear charge. This is due to the fact that with increase in nuclear charge, the electrons of the outer shell are more firmly held by the nucleus and thus greater energy is required to pull out an electron from the atom.

e.g., the ionisation enthalpy increases as we move along a period from left to right due to increased nuclear charge.

Element of 2nd period	Li	Be	B	C	N	O	F	Ne
Nuclear charge	+3	+4	+5	+6	+7	+8	+9	+10
First ionisation enthalpy (kJ mol^{-1})	520	899	801	1086	1402	1314	1681	2080

(ii) **Atomic size or radius** Ionisation enthalpy decreases as the atomic size increases. As the distance of the outer electrons from the nucleus increases with increase in atomic radius, the attractive force on the outer electron decreases.

As a result, outer electrons are held less firmly and hence lesser amount of energy is required to knock them out. Thus, ionisation enthalpy decreases with increase in atomic size. Ionisation enthalpy is found to decrease on moving down a group

Element (alkali metals)	Li	Na	K	Rb	Cs
First ionisation enthalpies (kJ mol^{-1})	520	496	419	403	374

(iii) **Penetration effect of the electrons** Ionisation enthalpy increases as the penetration effect of the electrons increases. It is well known fact that in case of multielectron atoms, the electrons of the *s*-orbital has the maximum probability of being found near the nucleus and this probability goes on decreasing in case of *p*, *d* and *f*-orbitals of the same shell.

In other words, *s*-electrons of any shell are more penetrating towards the nucleus than *p*-electrons the same shell. Thus, within the same shell, the penetration effect decreases in the order $s>p>d>f$

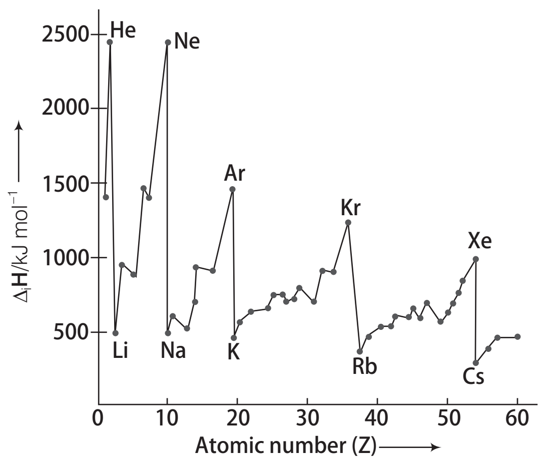
e.g., First ionisation enthalpy of aluminium is lower than that of magnesium. This is due to the fact that in case of aluminium ($1s^2 2s^2 2p^6 3s^2 3p_x^1$), we have to pull out a *p*-electron to form Al^+ ion whereas in case of magnesium ($1s^2 2s^2 2p^6 3s^2$) we have to remove an *s*-electron of the same energy shell to produce Mg^+ ion.

(iv) **Shielding** *or* **screening effect of inner shell electrons** As the shielding or the screening effect of the inner electrons increases, the ionisation enthalpy decreases. Consequently, the force of attraction by the nucleus for the valence shell electrons decreases and hence the ionisation enthalpy decreases.

(v) **Effect of arrangement of electrons** If an atom contains exactly half filled or completely filled orbitals then such an arrangement has extra stability. Therefore, the removal of an electron from such an atom requires more energy than expected.

e.g., Be ($1s^2 2s^2$) has higher ionisation enthalpy than B ($1s^2 2s^2 2p^1$) and N($1s^2 2s^2 2p_x^6 2p_y^1 2p_z^1$) has higher ionisation enthalpy than O ($1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$). In general, as we move from left to right in a period, the ionisation enthalpy increases with increasing atomic numbers.

The ionisation enthalpies keep on decreasing regularly as we move down a group from one element to the other.



$(\Delta_i H)$

Q. 51 Justify the given statement with suitable examples–"the properties of the elements are a periodic function of their atomic numbers".

Ans. There are numerous physical properties of elements such as melting points, boiling points, heats of fusion and vaporisation, energy of atomisation, etc., which show periodic variations.

The cause of periodicity in properties is the repetition of similar outer electronic configurations after certain regular intervals. *e.g.,* all the elements of 1st group (alkali metals) have similar outer electronic configuration, *i.e.,* ns^1.

$$_3Li = 1s^2, 2s^1$$
$$_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$$
$$_{19}K = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$$
$$_{37}Rb = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 5s^1$$
$$_{55}Cs = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2, 5p^6, 6s^1$$
$$_{87}Fr = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 4f^{14}$$
$$5s^2, 5p^6, 5d^{10}, 6s^2, 6p^6, 7s^1$$

Therefore, due to similar outermost shell electronic configuration all alkali metals have similar properties. *e.g.,* sodium and potassium both are soft and reactive metals. They all form basic oxides and their basic character increases down the group. They all form unipositive ion by the lose of one electron.

Similarly, all the elements of 17th group (halogens) have similar outermost shell electronic configuration, *i.e.,* ns^2np^5 and thus possess similar properties.

$$_9F = 1s^2, 2s^2, 2p^5$$
$$_{17}Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$$
$$_{35}Br = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^5$$
$$_{53}I = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2, 5p^5$$
$$_{85}At = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10},$$
$$4f^{14}, 5s^2, 5p^6, 5d^{10}, 6s^2, 6p^5$$

Q. 52 Write down the outermost electronic configurations of alkali metals. How will you justify their placement in group 1 of the periodic table?

Ans. All the elements of group IA (or I), *i.e.,* alkali metals have the similar outer electronic configuration, *i.e.,* ns^1 where n refers to the number of principal shell. *These electronic configurations are given below*

Symbol	Atomic number	Electronic configuration
Li	3	$1s^2 2s^1$ or [He] $2s^1$
Na	11	$1s^2 2s^2 2p^6 3s^1$ or [Ne]$3s^1$
K	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ *or* [Ar] $4s^1$
Rb	37	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ *or* [Kr] $5s^1$
Cs	55	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$ or [Xe] $6s^1$
Fr	87	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14}$ $5s^2 5p^6 5d^{10} 6s^2 6p^6 7s^1$ or [Rn] $7s^1$.

Hence, placement of all these elements in group 1 of the periodic table because of similarity in electronic configuration and all the elements have similar properties.

Q. 53 **Write the drawbacks in Mendeleef's periodic table that led to its modification.**

Ans. *The main drawbacks of Mendeleef's periodic table are*

(i) Some elements having similar properties were placed in different groups whereas some elements having dissimilar properties were placed in the same group.

e.g., alkali metals such as Li, Na, K, etc., (IA group) are grouped together with coinage metals such as Cu, Ag, Au (IB group) though their properties are quite different. Chemically similar elements such as Cu(IB group) and Hg (IIB group) have been placed in different groups.

(ii) Some elements with higher atomic weights are placed before the elements with lower atomic weights in order to maintain the similar chemical nature of elements.

i.e., $^{39.9}_{18}Ar$ and $^{39.1}_{19}K$; $^{58.9}_{27}Co$ and $^{58.7}_{28}Ni$, etc.

(iii) Isotopes did not find any place in the periodic table. However, according to Mendeleef's classification, these should be placed at different places in the periodic table.

(All the above three defects were however removed when modern periodic law based on atomic number was given).

(iv) Position of hydrogen in the periodic table is not fixed but is controversial.

(v) Position of elements of group VIII could not be made clear which have been arranged in three triads without any justification.

(vi) It could not explain the even and odd series in IV, V and VI long periods.

(vii) Lanthanides and actinides which were discovered later on have not been given proper positions in the main frame of periodic table.

Q. 54 **In what manner is the long from of periodic table better than Mendeleef's periodic table? Explain with examples.**

Ans. The long form of the periodic table is better than Mendeleef's periodic table because it classifies the elements on the basis of electronic configurations of their atoms.
The characteristics of this table are

(i) The table consists of 9 vertical columns, called the groups and 7 horizontal rows, called the periods.

(ii) The groups are marked 0 to VIII out of which group I to VII are subdivided into subgroups A and B.

(iii) The group IA elements (Li,Na,K,Rb,Cs and Fr) are known as alkali metals and the group IIA elements (Be,Mg,Ca,Sr,Ba and Ra) are known as alkaline earth metals. Elements in group VIII A (F,Cl,Br,I and At) are called halogens and elements in group VIII (He, Ne, Ar, Kr, Xe and Rn) are called noble gases or rare gases.

(iv) The group VIII has there similar elements placed together in one place. These are called transition triads, *e.g.,* Fe, Co and Ni, Ru, Rh and Pd; Os, Ir and Pt etc.

(v) In the 6th and 7th period, 14 elements present called as lanthanides and actinides respectively.

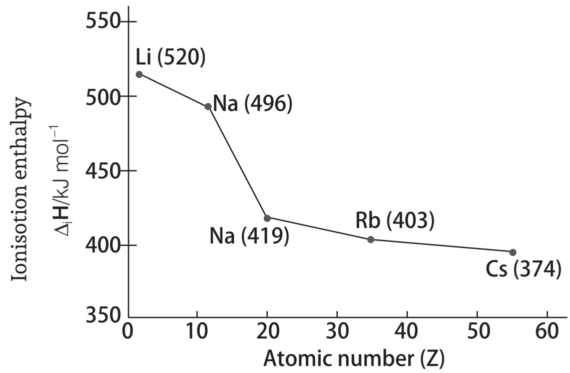
(vi) Based on their electronic configuration, elements have been grouped into *s* – ,*p*–, *d*- and *f*-blocks. This has helped us to understand their properties more easily.

(vii) There is gradual change in properties seen from one end to the other.

Q. 55 Discuss and compare the trend in ionisation enthalpy of the elements of group 1 with those of group 17 elements.

Ans. The ionisation enthalpies decreases regularly as we move down a group from one element to the other. This is evident from the values of the first ionisation enthalpies of the elements of group 1 (alkali metals) and group 17 elements as given in table and figure.

Group 1	First ionisation enthalpies ($kJ\ mol^{-1}$)	Group 17	First ionisation enthalpies ($kJ\ mol^{-1}$)
H	1312	F	1681
Li	520	Cl	1255
Na	496	Br	1142
K	419	I	1009
Rb	403	At	917
Cs	374		



Given trend can be easily explained on the basis of increasing atomic size and screening effect as follows

(i) On moving down the group, the atomic size increases gradually due to the addition of one new principal energy shell at each succeeding element. Hence, the distance of the valence electrons from the nucleus increases.

Consequently, the force of attraction by the nucleus for the valence electrons decreases and hence the ionisation enthalpy should decrease.

(ii) With the addition of new shells, the shielding or the screening effect increases. As a result, the force of attraction of the nucleus for the valence electrons further decreases and hence the ionisation enthalpy should decrease.

(iii) Nuclear charge increases with increase in atomic number. As a result, the force of attraction by the nucleus for the valence electrons should increase and accordingly the ionisation enthalpy should increase.

The combined effect of the increase in the atomic size and the screening effect more than compensates the effect of the increased nuclear charges. Consequently, the valence electrons become less and less firmly held by the nucleus and hence the ionisation enthalpies gradually decrease as move down the group.