UNIT 3

CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

Answer the questions. (1 Score each)

1. In the modern periodic table elements in a given group have similar chemical properties. Give reason.

Ans: Elements in a given group have same no. of valence electrons. So they have similar chemical properties.

2. Account for the following :

(a) The ionic radius of fluoride ion is 136 pm, while the atomic radius of fluorine (F) is only 64 pm.

(b) The second ionization enthalpy of an element is always greater than that of the first ionization enthalpy.

Ans: (a) This is due to greater electronic repulsion and lesser effective nuclear charge in fluoride ion

(b) This is because it is more difficult to remove an electron from a positively charged ion than from a neutral atom.

3. Among N^{3-}, O^{2-}, F^{-}, Na^{+} and Al^{3+} which one has the smallest size?

Ans; Al^{3+}

4. Give reasons for the following :

a) 'O' has lower ionization enthalpy than N and F.

b) Cl has higher negative electron gain enthalpy than F.

Ans: a) The electronic configuration of Oxygen is $1s^2 2s^2 2p^4$. After the removal of one electron, Oxygen gets the stable half filled electronic configuration. So it has lower ionisation enthalpy

b) Due to larger size and less electron-electron repulsion in chlorine.

5.Which is the acidic oxide among the following?

a) Cl_2O_7 **b) Na_2O** **c) Al_2O_3** **d) CO**

Ans; *a)* Cl_2O_7

6. Justify the following :

a) Ne has positive value for electron gain enthalpy.

b) The electron gain enthalpy of F is lower than that of Cl

c) The size of Al^{3+} is lower than that of F^-

Ans; a) Due to stable octet configuration of Ne.

b) Due to small size and greater electronic repulsion in fluorine.

c) Due to greater effective nuclear charge in Al^{3+} .

7. Transition elements are d-block elements why?

Ans; i) Because in transition elements, the last electron enters in the penultimate d-subshell.

8. Select isoelectronic species from the following:

O^- , F^- , Na^+ , Mg^+

Ans: , F^- and Na^+

9. Define electron gain enthalpy.

Ans; It is the heat change (enthalpy change) occurred when an electron is added to the outer most shell of an isolated gaseous atom.

10. Ionisation enthalpy of Nitrogen is greater than that of oxygen. Give reason

Ans; Due to the stable half filled electronic configuration of Nitrogen.

11. A group of ions are given below. Find one pair which is not isoelectronic.

$Na^+, Al^{3+}, Ca^{2+}, Br^-, F^-$

Ans: Ca^{2+} *and* Br^-

12. Second period elements show anomalous behaviour. Why?

Ans: Due to their smaller size, high electronegativity, large charge to radius ratio and absence of vacant d-orbitals.

13. Why is potassium considered as an s-block element?

Ans: $_{19}K - [Ar]\, 4s^1$. **It's last electron enters the valence s-subshell. So it is considered as s-block element.**

14. Write the general electronic configuration of d-block elements.

Ans; $(n\text{-}1)d^{1\ to\ 10}\ ns^{0\ to\ 2}$

15. In general, ionisation enthalpy increases from left to right across a period. Give reason.

Ans; Due to decrease in atomic size and increase in nuclear charge across a period.

16. Name a numerical scale of electro negativity of elements.

Ans; Pauling's electronegativity scale

17. The values of electron gain enthalpy with atomic number of halogens are given below:

Atomic No	Element	$\Delta_{eg}H$
9	F	– 328
17	Cl	– 349
35	Br	– 325
53	I	– 295

i) Why electron gain enthalpy become less negative from chlorine to iodine?

ii) Chlorine has more negative electron gain enthalpy than Fluorine. Why?

Ans; i) Due to increase in atomic size and screening effect down the group.

ii) Due to larger size and less electron-electron repulsion in chlorine.

18. State the modern periodic law.

Ans; The modern periodic law states that the properties of elements are the periodic functions of their atomic numbers.

19. What is meant by isoelectronic species?

Ans; They are species having same no. of electrons.

20. Select the isoelectronic species from the following.

N, O^{2-}, F^{-}, Al^{2+}, Mg^{2+}, Na^{+}

Ans; O^{2-}, F^{-}, Mg^{2+} and Na^{+}

21. Account for the following:

a) Ionization enthalpy of nitrogen is greater than that of oxygen.

b) Atomic radius decreases from left to right in a period.

Ans; a) Due to the stable half filled electronic configuration of Nitrogen.

b) Along a period, the no. of shells remains the same and the nuclear charge increases one by one. So the atomic radius decreases.

22. Elements have electron gain enthalpy and electronegativity.

a) We are two elements belong to the same group. One of us has the highest electronegativity and other, highest electron gain enthalpy. Identify us.

b) Define electron gain enthalpy?

c) Electron gain enthalpy values of noble gases are zero. Why?

Ans: a) The highest electronegativity - F and the highest electron gain enthalpy – Cl.

b) It is the heat change (enthalpy change) when an electron is added to the outer most shell of an isolated gaseous atom.

c) Due to their stable octet configuration.

Answer the questions. (2 Score each)

23. Chlorine has the most negative electron gain enthalpy'. Justify the statement

Ans; This is because, when an electron is added to F, it enters into the smaller second shell. Due to the smaller size, the electron suffers more repulsion from the other electrons. But for Cl, the incoming electron goes to the larger third shell. So the electronic repulsion is low and hence Cl adds electron more easily than F.

24. Identify the positions of Al (z=13) and S (z=16) in the periodic table with the help of their electronic configurations. Predict the formula of the compound formed between them.

Ans: $_{13}Al$ – [Ne] $3s^2\ 3p^1$, *Period – 3, Group - 13*

$_{16}S$ – [Ne] $3s^2\ 3p^4$, *Period – 3, Group - 16*

Al_2S_3 is the molecular formula.

25. Explain any two factors affecting electron gain enthalpy.

Ans ; Electron gain enthalpy depends on atomic size, nuclear charge, shielding effect etc.

26. In the periodic table, elements are classified into four blocks. Explain any two blocks.

Ans;

s block elements: These are elements in which the last electron enters in the outer most s-sub shell. They include elements of the groups 1 and 2. Their general outer electronic configuration ns^1 or ns^2 or ns^{1-2}

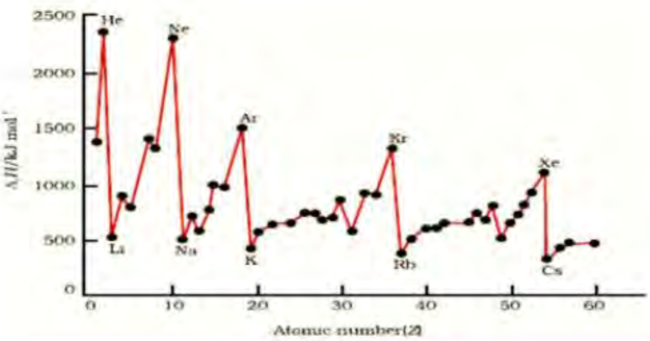
p block elements: These are elements in which the last electron enters in the outer most p-sub shell. They include elements of the groups 13 to 18. Their general outer electronic configuration $ns^2\ np^{1-6}$

d block elements: These are elements in which the last electron enters in the outer most d-sub shell. They include elements of the groups 12 to 17. Their general outer electronic configuration $(n-1)\ d^{1-10}\ ns^{0-2}$

27. First member of a group differs from the rest of the members of the same group why?

Ans; Due to their smaller size, high electronegativity, large charge to radius ratio and absence of vacant d-orbitals.

28. Analyze the following graph



What conclusion can you derive from the graph regarding the first ionization enthalpies of alkali metals and noble gases? Justify your answer.

Ans; In a period, the alkali metals have the least ionisation enthalpy and the noble gases have the most. This is because after the removal of only one electron from the valence shell, alkali metals get the stable completely filled electronic configuration. So they have low ionisation enthalpy. Noble gases have stable octet configuration. So they have high ionisation enthalpy.

29. Aluminium forms $[AlF_6]^{3-}$ whereas boron cannot form $[BF_6]^{3-}$ but forms $[BF_4]^-$ even though both belong to the same group. Explain.

Ans; Due to the presence of vacant d orbitals in Aluminium, Al can extend its covalency beyond 4. So it can form $[AlF_6]^{3-}$ But in Boron, there is no vacant d-orbitals. So its maximum covalency is 4

30. Does the ionization enthalpy decrease along a group? Give reason

Ans; Yes, the ionization enthalpy decreases due to increase in atomic size, decrease in nuclear charge and increase in shielding effect along a group.

31. The first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium. Explain.

Ans; The electronic configuration of sodium is: $_{11}Na - [Ne]\ 3s^1$

After the removal of one electron, Na gets the stable noble gas configuration. So it has lower first ionisation enthalpy and higher second ionisation enthalpy.

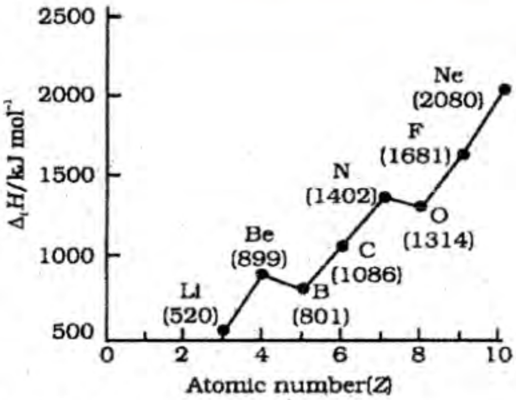
The electronic configuration of magnesium is: $_{12}Mg$ - $[Ne]\ 3s^2$

Mg have a completely filled stable electronic configuration. So it requires higher first ionisation enthalpy . But after that , if it lose one electron, then its Second ionisation enthalpy will be lower.

32. Account for the observation that cations are always smaller than the parent atom while anions are always larger than the parent atom.

Ans; This is because of the greater effective nuclear charge in cations. But in anions, the addition of one or more electrons would result in an increased electronic repulsion and a decrease in effective nuclear charge.

33. Observe the following graph in which the first ionisation enthalpies ($\Delta_i H$) of elements of the second period are plotted against their atomic numbers (Z):



Identify the anomalous values and justify.

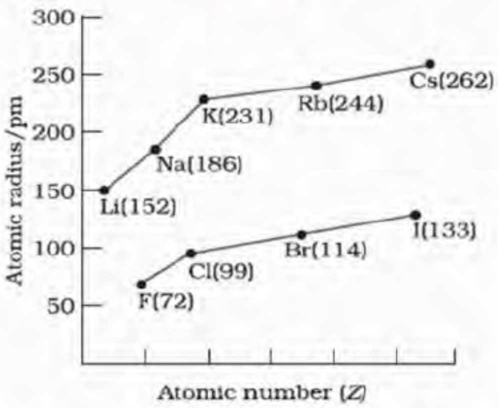
Ans; B and O have lower ionisation enthalpy than expected. This is because after the removal of one electron, B gets the stable fully filled configuration and O gets the stable half filled configuration.

34.Identify the largest and smallest ion given below:

O^{2-} , F^- , Mg^{2+} , Na^+

Ans; *The largest ion is O^{2-} and the smallest is Mg^{2+}.*

35. A graph showing the variation of atomic radius with atomic number for alkali metals is given below.



Comment on the variation of atomic radius with increase in atomic number in a group. Give reason.

Ans; Down a group, the atomic size increases due to increase in no. of shells and shielding effect.