Chapter#1
Periodic Classification of Elements and Periodicity

1. Define periodic table.

Ans: The periodic table provides a basic framework to study the periodic behaviour of physical and chemical properties of elements as well as their compounds.

2. What is the contribution of Al-Razi in the field of chemistry?

Ans: Al-Razi classified the substances on the basis of their physical and chemical properties.

3. What are Dobereiner's triads?

Ans: Dobereiner, a German chemist in 1829, arranged the known elements in a group called Triads, as each contained three elements with similar properties. According to Dobereiner's Law of Triads,

"If the atomic mass of 1st and 3rd element is known, then the atomic mass of 2nd (middle) element can be calculated by taking the average of other two elements". For example,

Li	Na	K
7	$\boxed{23}$	39

$$\frac{7+39}{2} = 23$$

4. What are Newland's Octaves?

Ans: Newland, who was an English chemist, in 1864, classified 62 elements, known at that time, in increasing order of their atomic masses. He noticed that *every eighth element had some properties in common with the first one*. The principle on which this classification is based was called the Law of Octaves.

5. Define Mendeleev's and modern periodic law.

Ans: **Mendeleev' Periodic law:**

"If the elements are arranged in ascending order of their atomic masses, their chemical properties repeat in a periodic manner".

Modern periodic law:
"If the elements are arranged in ascending order of their atomic numbers, their chemical properties repeat in a periodic manner".

6. What are the drawbacks in Mendeleev's periodic table? OR Give two defects of Mendeleev's periodic table.

Ans: Following are the drawbacks in Mendeleev's periodic table:

1. The elements were arranged in increasing order of their atomic masses.
2. Another confusion in Mendeleev's table was that elements like Be, Mg, Ca, Sr, Ba and Zn, Cd, Hg were placed in a single vertical group, while according to their properties they belonged to two different categories. The same was true for so many other elements placed in the same vertical group.

7. What are the improvements in Mendeleev's periodic table?

Ans: Following are the improvements made in Mendeleev's periodic table:

1. After the discovery of atomic number by Moseley in 1911, it was noticed that elements could be classified more satisfactorily by using their atomic numbers rather than their atomic masses.
2. The periodic table was improved by arranging the elements in ascending order of their atomic numbers instead of their atomic masses.
3. Another improvement was the addition of an extra group (group VIIIA) at the extreme right of the periodic table. This group contains noble gases, which had not been discovered in Mendeleev's time.
4. In modern periodic table, the confusion of placing elements with different properties in same group was removed by dividing the elements in two types of vertical groups, A and

B. In modern periodic table, Be, Mg, Ca, Sr and Ba are placed in group IIA and Zn, Cd, Hg in group IIB.

8. Zn,Cd,Hg were placed along with alkaline earth metals in Mendeleev's periodic table. How this confusion was removed in Modern periodic table?

Ans: In modern periodic table, the confusion of placing elements with different properties in same group was removed by dividing the elements in two types of vertical groups, A and B. In modern periodic table, Be, Mg, Ca, Sr and Ba are placed in group IIA and Zn, Cd, Hg in group IIB.

9. Define groups and periods.

Ans: **Groups**

Elements with similar properties are placed in vertical columns called Groups. There are eight groups, which are usually numbered by Roman numerals I to VIII.

Points to remember (not part of definition)

Each group is divided into two sub-groups designated as A and B subgroups. The sub-groups, containing the representative or normal elements are labeled as A subgroups, whereas B subgroup contain less typical elements, called transition elements and are arranged in the centre of the periodic table.

Periods

The horizontal rows of the periodic table are called Periods. There are seven periods in the periodic table.

10. Which period is the shortest one in the periodic table?

Ans: The period 1 is the shortest one in the periodic table. It contains only two elements, hydrogen and helium.

11. Tell about short periods in the periodic table.

Ans: The periods 2 and 3 contain eight elements each and are called short periods. All the elements in these periods are representative elements and belong to A subgroup. In these periods, every eighth element resembles in properties with the first element. Lithium and beryllium in the 2^{nd}period resemble in most of their properties with sodium and magnesium of the 3^{rd}period, respectively. Similarly, boron and aluminium both show oxidation state of +3, fluorine in 2^{nd}period has close resemblances with chlorine of 3^{rd} period.

Elements of second period: Li, Be, B, C, N, O, F, Ne

Elements of third period: Na, Mg, Al, Si, P, S, Cl, Ar

12. Tell about long periods in the periodic table.

Ans: The periods 4 and 5 are called long periods. Each long period consists of eighteen elements. Out of these, eight are representative elements belonging to A subgroup similar to second and third periods. Whereas the other ten elements, placed in the centre of the table belong to B subgroups and are known as transition elements. In these periods, the repetition of properties among the elements occurs after 18 elements. As after $_{19}K$ (having atomic number 19) the next element with similar properties is $_{37}Rb$. The period 6 is also a long period, which contains thirty-two elements. In this period there are eight representative elements, ten transition elements and a new set of fourteen elements called Lanthanides as they start after $_{57}La$. Lanthanides have remarkably similar properties and are usually shown separately at the bottom of the periodic table.

Elements of 4^{th} period: K, Ca, Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr

Elements of 5^{th} period: Rb, Sr, Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd, In, Sn, Sb, Te, I, Xe

Elements of 6^{th} period: Cs, Ba, La, Hf, Ta, W, Re, Os, Ir, Pt, Au, Hg, Tl, Pb, Bi, Po, At, Rn

In sixth period Lanthanides: La (57)-Lu (71)

13. Which period is called incomplete in the periodic table?

Ans: The period 7 is incomplete so far. It contains only two normal elements $_{87}Fr$ and $_{88}Ra$, ten transition elements and fourteen inner transition elements. The inner transition elements of this period are called Actinides, as they follow $_{89}Ac$-$_{103}Lr$. The actinides are also shown at the bottom of the periodic table under the Lanthanides.

14. What are rare earth elements?

Ans: Lanthanides and actinides are called rare earth elements as they are found in a very small amount in the earth's crust.

15. Tell about blocks in the periodic table.

Ans:

- IA and IIA subgroups are called s-block elements because their valence electrons are available in s orbital.
- The elements of IIIA to VIIIA subgroups (except He) are known as p-block elements as their valence electrons are present in p orbital.
- Similarly in transition elements, electrons in d-orbital are responsible for their valency hence they are called d-block elements.
- For Lanthanides and Actinides valence electrons are present in f- orbital hence these elements are called f-block elements.

16. How the classification of elements in different blocks helps in understanding their chemistry?

Ans: The classification in blocks (s, p, d, f) is based upon the valence orbital of the element involved in chemical bonding. So this classification is quite useful in understanding the chemical properties of elements, especially, the concept of valency as oxidation state.

17. What are alkali metals? Why they are called so?

Ans: Group IA elements are called alkali metals because of their property to form strong alkalies with water.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

18. What are alkaline earth metals? Why they are called so?

Ans: Group IIA elements are called alkaline earth metals because of their presence in the earth's crust and alkaline character.

19. Why halogens are called so?

Ans: Group VIIA elements are called halogens because of their salt forming properties when halogens are treated with metals.

20. Why noble gases are called so?

Ans: The gases of group VIIIA are least reactive as their outermost shells are complete (ns^2np^6). They cannot gain, lose or share electrons. That is why, they are called noble gases.

21. Write name and symbol of an element from s block that has zero oxidation state. Also write its electronic configuration.

Ans: Helium (He) is the element from s block that has zero oxidation state. The electronic configuration is $1s^2$.

22. Discuss the position of metals, non-metals and metalloids in the periodic table.

Ans: The elements on the left hand side, in the centre and at the bottom of the periodic table are metals, while the non-metals are in the upper right corner of the table. Some elements, especially lower members of groups, III A, IVA and VA have properties of both metals as well as non-metals. These elements are called semi-metals or metalloids. In the periodic table elements of groups IVA to VIIIA, at the top right hand corner above the stepped-line, are non-metals. The elements just under the "steps' such as Si, As, and Te are the metalloids. All the remaining elements, except hydrogen, are metals.

23. What are transition elements?

Ans: Transition elements may be defined as those elements which have partially filled d or f-subshells in atomic state or in any of their commonly occurring oxidation states.

The d-block and the f-block elements are called transition elements because they are located between the s and p-block elements and their properties are in transition between the metallic elements of the s-block and nonmetallic elements of the p-block.

24. d and f-block elements are called transition elements. Why?

Ans: The d-block and the f-block elements are called transition elements because they are located between the s and p-block elements and their properties are in transition between the metallic elements of the s-block and non-metallic elements of the p-block.

25. What are typical and non-typical transition elements?

Ans: The elements of group IIB and group IIIB are referred to as non-typical transition elements as some of their properties are like representative elements and the elements in the remaining transition series are called typical transition elements.

26. What are inner transition elements and outer transition elements?

Ans: The f-block elements i.e. lanthanides and actinides are called inner transition elements and the d-block elements are called outer transition elements.

27. What are coinage metals?

Ans: The elements of group IB i.e. Cu, Ag, Au are called coinage metals as they are used in making coins.

28. Discuss position of hydrogen with respect to Group IA.

Ans: Position of hydrogen with respect to IA is discussed below:

Similarities

1. Like alkali metals hydrogen atom has one electron in Is sub-shell, which it can lose to form H^+
2. Both hydrogen and alkali metals have a strong tendency to combine with electronegative elements such as halogens.
3. Similar to alkali metals hydrogen also forms ionic compounds, which dissociate in water.

Dissimilarities

1. Hydrogen is a nonmetal in true sense. It does not lose electron as easily as most of the alkali metals do.
2. Unlike alkali metals molecular hydrogen exists in open atmosphere.

29. Discuss position of hydrogen with respect to Group VIIA.

Ans: Position of hydrogen with respect to VIIA is discussed below:

Similarities

1. Hydrogen is a gas like most of the halogens and is stable in diatomic form such as F_2, Cl_2 and Br_2.
2. As required by halogens, hydrogen also needs one electron to complete its outermost shell.
3. By accepting one electron hydrogen forms H^- (Hydride ion) similar to F^-, Cl^- and Br^-.
4. Both hydrogen and halogens form stable ionic compounds with alkali metals.

Dissimilarities

1. By losing its only electron, hydrogen forms H^+ but halogens do not form positive ions.

2. Combining with oxygen, hydrogen forms very stable oxides while halogens lack this property.

30. Discuss position of hydrogen with respect to IVA Group.

Ans: Position of hydrogen with respect to IVA is discussed below:

Similarities

1. Valence shell of hydrogen is half filled like those of group IVA elements.
2. Both, hydrogen and group IV elements combine with other elements through covalent bonding.
3. Like carbon, hydrogen also possesses remarkable reducing properties.

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

$$SnO_2(s) + C(s) \longrightarrow Sn(s) + CO_2(g)$$

Dissimilarities

1. Carbon and silicon form long chain compounds, when their atoms combine with each other, while hydrogen does not form such compounds.
2. Similarly, carbon can simultaneously form bonds with more than one elements, whereas, hydrogen due to having only one electron can combine with only one element at a time.

31. Discuss reducing properties of hydrogen with respect to IVA.

Ans: Like carbon, hydrogen also possesses remarkable reducing properties.

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

$$SnO_2(s) + C(s) \longrightarrow Sn(s) + CO_2(g)$$

32. Define halides. Give examples.

Ans: Halides are the binary compounds which halogens form with other elements. For example, NaCl, KBr, NaI.

33. How are halides classified?

Ans: Halides can be classified into two general classes: ionic and covalent. In between the two, there is another class of halides in which the halogen atom acts as a bridge between the two atoms of the other element. Such halides are termed as "Polymeric" halides.

34. How are ionic halides formed? Give their properties.

Ans: Strongly electropositive elements, having greater electronegativity difference with halogen atom, form ionic halides. The halides of group IA are considered purely ionic compounds, which are high melting point solids. Such halides have three-dimensional lattices consisting of discrete ions.

35. Which ionic halides have the highest melting point and why?

Ans: Among the pure ionic compounds, the fluorides have the highest lattice energies due to the small size of fluoride ion. Thus, for ionic halides, the fluorides have the highest melting and boiling points which decrease in the order:

fluoride > chloride > bromide > iodide.

36. Which elements give polymeric halides and what are their properties?

Ans: Less electropositive elements, such as Be, Ga and Al form polymeric halides having partly ionic bonding with layer or chain lattices.

37. What is the trend of halides from left to right in the periodic table?

Ans: On moving across the periodic table from left to right, the electronegativity difference reduces and the trend shifts towards covalent halides. Ionic halides are formed by elements on the right side of the periodic table (IA and IIA). For example, in 3rd period:

i. NaCl and $MgCl_2$ are ionic halide.

ii. $AlCl_3$ is polymeric halide.

iii. $SiCl_4$, PCl_3, S_2Cl_2 are covalent in nature.

38. What are the properties of covalent halides?

Ans: As the intermolecular forces in covalent halide molecules are weak van der Waal's forces so they are often gases, liquids or low melting point solids. Physical properties of covalent halides are influenced by the size and polarizability of the halogen atom. Iodides, as being the largest and more polarizable ions, possess the strongest van der Waal's forces and have higher melting and boiling points than those of other covalent halides.

39. What is the trend of halides from top to bottom in the periodic table?

Ans: In general, for a metal the order of decreasing ionic character of the halides is:

fluoride > chloride > bromide > iodide.

40. AlF_3 has higher melting point than AlI_3. Why?

Ans: AlF_3 is purely ionic compound having melting point 1290°C and fairly a good conductor, whereas AlI_3 is predominantly covalent with melting point 198°C and electrically a non-conductor.

41. Why $PbCl_2$ is ionic but $PbCl_4$ is fairly covalent compound?

Ans: When a metal forms more than one halide, the halides in which metal has lower oxidation state tends to be ionic while that in higher oxidation state is covalent. Similarly, high polarizing power of Pb^{+4} as compared to Pb^{+2} makes $PbCl_2$ mainly ionic, but, $PbCl_4$ fairly covalent.

42. How oxidation state tells about ionic or covalent nature of halides?

Ans: In case of an element forming more than one halides the metal halide in its lower oxidation state tends to be ionic, while that in the higher oxidation state is covalent. For example, $PbCl_2$ is mainly ionic and $PbCl_4$ is fairly covalent. This can again be explained by the high polarizing power of Pb^{4+} as compared to that of Pb^{2+}.

43. What are hydrides?

Ans: The binary compounds of hydrogen with other elements are called hydrides. For example, NaH, H_2O, H_2S.

44. What are hydrides? What is the trend of boiling points of hydrides of group VIA down the group?

Ans: The binary compounds of hydrogen with other elements are called hydrides. For example, NaH, H_2O, H_2S.The boiling point of hydrides of group VIA increases down the group except H_2O which is due to hydrogen bonding and has higher boiling point than might be expected.

Group VIA (Hydrides)	Boiling point (°C)
H_2O	100
H_2S	-60.3
H_2Se	-42
H_2Te	-2

45. How hydrides are classified?

Ans: Hydrides may be broadly classified into three classes: ionic, covalent and intermediate.

46. Why alkali metals give ionic hydrides?

Ans: Alkali metals are more electropositive than hydrogen. They have strong tendency to lose electron and form a uni-positive ion. This electron is accepted by hydrogen to form a hydride ion (H^-). These cations and anions then combine to form ionic bond. That's why alkali metals give ionic hydrides. For example, Na^+H^- and K^+H^-.

47. Which elements form ionic hydrides and what are their properties?

Ans: The elements of group IA and the heavier members of group IIA form ionic hydrides, which contain H^- (Hydride) ion. These hydrides are crystalline solid compounds, with high melting and boiling points, which conduct electricity in molten state.

48. What is the trend of hydrides from left to right in the periodic table?

Ans: The elements on the left side of the periodic table (IA and IIA) form ionic hydrides. The tendency towards covalent character increases by moving from left to right in the periodic table. Hydrides of beryllium and magnesium represent the class of intermediate hydrides. Their properties are in between the ionic and covalent hydrides.

49. What are the properties of covalent hydrides?

Ans: The covalent hydrides are usually gases or volatile liquids. They are non-conductors and dissolve in organic solvents. Their bond energies depend on the size and the electronegativity of the element. Stability of covalent hydrides increases from left to right in a period and decreases from top to bottom in a group.

50. The boiling point of covalent hydrides generally increases except the first members of group VA-VIIA which have higher boiling point than the other group members. Why?

Ans: The boiling points of covalent hydrides generally increase on descending a group except the hydrides like H_2O, HF and NH_3 which, due to hydrogen bonding, have higher boiling points than might be expected.

51. Define oxides. Give examples.

Ans: The compounds which oxygen forms with other elements are called oxides. For example, P_2O_5, Na_2O, CO_2.

52. How are oxides classified?

Ans: Oxides can be classified in more than one ways: based upon the type of bonding they have as well as their acidic or basic character. On the basis of acidic and basic character they are categorized as acidic, basic and amphoteric oxides.

53. What are basic oxides?

Ans: Metals form basic oxides like IA and IIA. When oxides of metals are dissolved in water they give bases.

For example:

$$Na_2O\ (s) + H_2O\ (l) \longrightarrow 2NaOH$$

54. What are acidic oxides?

Ans: Non-metals give acidic oxides like C, N, P and S. When oxides of non-metals are dissolved in water they form acids. For example:

$$SO_3\ (g) + H_2O\ (l) \longrightarrow H_2SO_4\ (aq)$$

55. What are amphoteric oxides?

Ans: The oxides having both acidic and basic properties are called amphoteric oxides. Oxides of relatively less electropositive elements, such as BeO, Al_2O_3, Bi_2O_3 and ZnO are amphoteric and behave as acids towards strong bases and as bases towards strong acids.

$$ZnO\ (s) + H_2SO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + H_2O\ (l)$$

$$ZnO\ (s) + 2NaOH\ (aq) + H_2O\ (l) \longrightarrow Na_2[Zn(OH)_4]\ (aq)$$

56. What happens when acidic and basic oxides combine with each other?

Ans: Basic oxides and acidic oxides react with one another to give salts e.g,

$$Na_2O + SO_3 \longrightarrow Na_2SO_4$$

57. Although both sodium and phosphorus are present in the same period of the periodic table yet their oxides are different in nature, Na_2O is basic while P_2O_5 is acidic in character. Justify.

Ans: Sodium is a metal while phosphorus is a non-metal. Metals give basic oxides while non-metals give acidic oxides.

$$Na_2O + H_2O \longrightarrow 2NaOH\ (base)$$

$$P_2O_5 + 3H_2O \longrightarrow 2H_3PO_4$$

58. What is the trend of oxides in the periodic table?

Ans: In a given period, the oxides progress from strongly basic through weakly basic, amphoteric, and weakly acidic to strongly acidic, e.g. Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_3, Cl_2O_7.The basicity of main group metal oxides increases on descending a group of the periodic table, (e.g. $BeO<MgO<CaO<SrO<BaO$), though the reverse trend is observed in the transition metal oxides.

59. What is the effect of oxidation state of metal on nature of oxide?

Ans: The oxidation state of the metal also affects the acid/base character of its oxide. The acidity increases with increasing oxidation state (e.g. the acidity of $MnO < Mn_2O_3 < MnO_2 < Mn_2O_7$).

60. Define hydration energy. Give an example.

Ans: The hydration energy is the heat absorbed or evolved when one mole of gaseous ions dissolve in water to give an infinitely dilute solution. For example, when one mole of gaseous hydrogen ions is dissolved in water to give an infinitely dilute solution and a large amount of heat is liberated:

$$H^+ (g) + H_2O (l) \longrightarrow H_3O^+ (aq) \qquad \Delta H_h = -1075 \text{ kJ mol}^{-1}$$

61. On which factor does hydration energy depends?

Ans: Hydration energies highly depend upon charge to size ratio of the ions. For a given set of ions, for example of group IA, charge to size ratio decreases from top to bottom in a group, the hydration energy also decreases in the same fashion. On the contrary, the hydration energy increases significantly by moving from left to right in a period as the charge to size ratio increases, as found in the metal ions of third period.

62. Hydration energies of ions are in the following order. $Al^{+3}>Mg^{+2}>Na^{+1}$. Justify it.

Ans: Hydration energy depends upon the charge to size ratio, greater the charge to size ratio greater the hydration energy. Hence, in the given order ($Al^{+3}>Mg^{+2}>Na^{+1}$), Al^{+3} has greater charge to size ratio than $Mg^{+2}$and Na^{+1}. That's why, the hydration energies are in this order:

$Al^{+3} > Mg^{+2} > Na^{+1}$
$-4613 > -1903 > -510$

63. What is the trend of electrical conductance in the periodic table?

Ans: The electrical conductance of metals in groups IA and IIA, generally increases from top to bottom. Metals of group IB, which are known as coinage metals, have extraordinary high values of electrical conductance. Non-metals especially of groups VIA and VIIA show such low electrical conductance that they can be considered as non-conductors. In the series of transition metals, the values of electrical conductance vary so abruptly that no general trend can be assigned to them. The lower elements of group IVA, tin and lead, are fairly good conductors and their values of electrical conductivity are comparable with those of their counterparts in group IA.

64. Why metallic character increases from top to bottom in a group of metal?

Ans: Metallic character increases from top to bottom in a group of metal because atomic size increases, shielding effect increases and effective nuclear charge decreases from top to bottom in a group. As a result, the removal of electrons from the outermost shell becomes easier as hold of nucleus on the outermost shell electrons decreases. Therefore, metallic character increases from top to bottom in a group of metals.

65. Why do metals conduct electricity? OR Why the metals are good conductors?

Ans: Metals are good conductors of heat and electricity due to the presence of free moving electrons. These free electrons act as carriers of heat and electricity from one end of the metal to the other. Greater the number of free moving electrons greater is the electrical conductance.

66. Why graphite is a conductor whereas diamond is a non-conductor?

Ans: Carbon, in the form of diamond is non-conductor because all of its valence electrons are tetrahedrally bound and unable to move freely, while in the form of graphite, carbon is fairly a good conductor because one of its four valence electrons is relatively free to move.

67. Define oxidation state.

Ans: The oxidation state of an atom in a compound is defined as the charge with the sign +ve or -ve, which it would carry in the compound. For example Na^{+1}, Cl^{-1}, F^{-1}.

68. How the oxidation state is linked with group number?

Ans: The oxidation state of a typical element is directly or indirectly related to the group number to which the element belongs in the periodic table. The elements of group IA to IVA have the same oxidation states as their group numbers are. Just as B, Al and Ga belong to group IIIA, hence, they always show oxidation state of +3. So, for the elements of these groups, the oxidation state is same as the number of electrons present in the valence shells of the elements.

69. When the oxidation state is not same as group number?

Ans: For the elements of group VA, the oxidation states are either the number of electrons present in the valence shell (which is same as their group number) or the number of vacancies available in these shells. For example, N, P, As and Sb frequently show +3 as well as +5 oxidation states. Elements of group VIA show almost similar behaviour. In H_2SO_4, sulphur shows the oxidation state of +6, which is the number of electrons in its outermost shell whereas its oxidation state is -2 in H_2S, which is the number of vacancies in the shell. In group VIIA elements oxidation state is mostly - 1, which is again the number of vacancies in their outermost shells. Group VIIIA elements, which are also called zero group elements, usually show zero oxidation state because there is no vacancy in their outermost shells.

70. Oxidation states usually remain same in a group. Why? OR The oxidation states vary in a period but remain almost constant in a group. Justify.

Ans: The number of electrons in the outermost shells goes on changing in period from left to right, so oxidation states go on changing but, the number of electrons in the outermost shell remains same in a group so the oxidation states remain the same. The process of un- pairing of electrons may happen in a group and oxidation states may change, especially, in case of groups of transition elements.

71. Why the oxidation states of noble gases are usually zero?

Ans: The oxidation state of an element is directly or indirectly related to the number of its valance electrons or the number of vacancies available in its valence shell. In case of noble gases, their outermost shells are completely filled with electrons and no vacancy is available in their outermost shells (ns^2np^6). Thus, these gases usually show zero oxidation state. That's why they are often called zero group elements.

72. Why transition elements show variable oxidation state?

Ans: Transition elements, which are shown in B sub-groups of the periodic table, also show the oxidation states equal to their group number as it can be seen for Cu(I), Zn(II), V(V), Cr(VI) and Mn (VII). But due to greater number of valence electrons available in partly filled d-orbitals these elements usually show more than one oxidation states in their compounds.

73. Why melting point and boiling point increases up to the middle of the periodic table and then decreases as we move from left to right?

Ans: Across the short periods, the melting and boiling points of elements increase with the number of valence electrons up to group IVA and then decrease up to the noble gases. Since carbon has the maximum number of binding electrons, thus, it has a very high melting point in diamond in which each carbon is bound to four other carbon atoms. In general, the elements which exist as giant covalent structures have very high melting points.

An important change occurs when we move from group IVA to groups VA, VIA, VIIA as the lighter elements of these groups exist as small, covalent molecules, rather than as three dimensional lattices. For instance, nitrogen, oxygen and fluorine exist as individual molecules which have very weak intermolecular forces between them. Consequently, their melting and boiling points are extremely low.

74. Why IA elements have lower melting point than IIA elements?

Ans: The melting points of group IA elements are low because each atom in them provides only one electron to form a bond with other atom. Melting points of group IIA elements are considerably higher than those of group IA elements because each atom in them provides two binding electrons.

75. Explain the variation in melting points along the short periods.

Ans: Melting points in short periods increase up to the middle and then decrease.

i. Increase in melting point is due to decrease in atomic size and increase in inter-atomic forces.

ii. After group IV-A lighter elements of these groups exist as small, discrete, covalent molecules and have weak inter-molecular forces as a result of which melting points also decrease.

76. What is the trend of melting and boiling point down the group in the periodic table?

Ans: The melting and boiling points of IA and IIA group elements decreases from top to bottom due to the increase in their atomic sizes. The binding forces present between large sized atoms are relatively weaker as compared to those between smaller atoms. For elements of group VIIA, which exist in the form of molecules, the melting and boiling points increase down the group. This is because large molecules exert stronger force of attraction due to their higher polarizabilities.

77. Why the melting and boiling points of halogens increase down the group?

Ans: For elements of group VIIA, which exist in the form of molecules, the melting and boiling points increase down the group. This is because large molecules exert stronger force of attraction due to their higher polarizabilities.

78. What is the trend of metallic and non-metallic character in the periodic table?

Ans: Metallic character increases from top to bottom in a given group of elements as the tendency to lose electrons increases from top to bottom because atomic size increases, effective nuclear charge decreases and shielding effect increases. Therefore, hold of nucleus on outer shell electrons decreases and it becomes easier to remove electrons from the outermost shell. On the contrary, it decreases from left to right across a period as the tendency to gain electrons increases from left to right, shielding effect remains constant, effective nuclear charge increases and atomic size decreases.

79. Define electron affinity. Write an example.

Ans: The electron affinity is the energy released or absorbed, when an electron is added to a gaseous atom to form a negative ion. For example,

$$F\,(g) + e^- \longrightarrow F^-\,(g) \quad E.A = -337\ kJ\ mol^{-1}$$

80. Why the second value of ionization energy is shown with a positive sign while the first value is shown with a negative sign?

Ans: Energy is usually released when electronegative elements absorb the first electron and electron affinity in such cases is expressed in negative figures as in the case of halogens. When a second electron is added to a uni-negative ion, the incoming electron is repelled by the already present electron and energy is absorbed in this process.

$$O\,(g) + e^- \longrightarrow O^-\,(g) \qquad E.A_1 = -141\ kJ\ mol^{-1}$$

$$O^-\,(g) + e^- \longrightarrow O^{2-}\,(g) \qquad E.A_2 = +780\ kJ\ mol^{-1}$$

81. What is the trend of electron affinity in the periodic table?

Ans: Electron affinity increases from left to right in periods of the periodic table because effective nuclear charge increases, shielding effect remains constant, size of atom decreases so hold of nucleus on outermost shell electrons increases.

Electron affinity decreases from top to bottom in the periodic table because effective nuclear charge decreases, shielding effect increases, size of atom increases so hold of nucleus on outermost shell electrons decreases.

82. Define ionization energy.

Ans: The ionization energy of an element is the minimum quantity of energy which is required to remove an electron from the outermost shell of its isolated gaseous atom in its ground state. For example,

$$Na\ (g) \longrightarrow Na^{+}\ (g) + e^{-} \quad \Delta H_i = +496\ kJ\ mol^{-1}$$

83. What is the trend of ionization energy in the periodic table? OR Why ionization energy decreases down the group and increases along a period?

Ans: Ionization energy decreases from top to bottom in a group because shielding effect increases, effective nuclear charge decreases and size of atom increases so less energy is required to remove electron from the outermost shell as the hold of nucleus on the outer shell electron decreases. Ionization energy increases from left to right in a period as the shielding effect remains constant, effective nuclear charge increases and size of atom decreases so more energy is required to remove electron from the outermost shell as the hold of nucleus on the outer shell electron increases.

84. Define atomic radius.

Ans: Half of the distance between the centers of two bonded atoms is considered to be the radius of the atom.

85. What is the trend of atomic radius in the periodic table?

Ans: Atomic radius decreases from left to right in periods of the periodic table as the shielding effect remains constant and effective nuclear charge increases so hold of nucleus on outermost shell electrons increases.

Atomic radius increases from top to bottom in a group because shielding effect increases and effective nuclear charge decreases so hold of nucleus on outermost shell electrons decreases. The trend of ionic radius is same.

86. Define Lanthanide contraction.

Ans: The lanthanide contraction is the greater-than-expected decrease in ionic radii of the elements of lanthanide series from atomic number 57, lanthanum, to 71, lutetium, which results in smaller than otherwise expected ionic radii for the subsequent elements starting with hafnium 72. It is due to the involvement of f sub-shells.

OR

The gradual or progressive decrease in the atomic size of the elements in the lanthanide series is significant and is called lanthanide contraction. Same decrease is observed in actinide series. This is due to poor shielding effect of f sub-shell which is being gradually filled along the series.

87. Lanthanide contraction controls the atomic sizes of elements of 6th and 7th periods. Justify.

Ans: In Lanthanide and Actinide series there is a gradual decrease in the atomic size from left to right due to increased nuclear charge. The slow reduction in atomic size in both these series is called Lanthanide contraction. This contraction controls the size of 6th and 7th period elements.

88. Why size of cation is smaller than the parent atom?

Ans: The removal of electrons causes an imbalance in proton-electron ratio. Due to the greater attraction of the nuclear charge, the remaining electrons of the ion are drawn closer to the nucleus. A positive ion is always smaller than the neutral atom from which it is derived. The radius of Na is 157pm and the radius of Na^+ is 95pm.

89. Why size of anion is larger than the parent atom?

Ans: A negative ion is always bigger than its parent atom. The reason is that addition of one or more electrons in the shell of a neutral atom enhances repulsion between the electrons causing expansion of the shell. Thus, the radius of fluorine atom is 72pm and that of the fluoride ion (F^-) is 136pm.

90. What are isoelectronic species?

Ans: Ions with the same electronic configuration are called isoelectronic species. For example, Na^{+1}(10 electrons), Ne (10 electrons), Mg^{2+} (10 electrons), O^{2-}(10 electrons), N^{3-}(10 electrons), etc.