NCERT TEXTUAL EXERCISE (SOLVED)

1. What is the basic theme of organisation in the periodic table?

Ans. The basic theme of organisation of elements in the periodic table is to simplify and systematise the study of the properties of all the elements and millions of their compounds.

2. Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that ?

Ans. Mendeleev used atomic weight as the basis of classification of elements in the periodic table. He arranged the known elements in order of increasing atomic weights grouping together elements with similar properties.

3. What is the basic difference in approach between Mendeleev's Periodic Law and the Modern Periodic Law ?

Ans. The basic difference in approach between Mendeleev's periodic Law and Modern Periodic Law is the change in basis of classification of elements from atomic weight to atomic number.

4. On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

Ans. In 6th period, electrons can be filled in only 6s, 4f, 5d and 6p-sub-shells whose energies increase in the order : 6s < 4f < 5d< 6p. Now s-sub-shell has one, p-sub-shell has three, d-sub-shell has five and f-subshell has seven orbitals. Hence, in all, there are 16 (1 + 3 + 5 + 7) orbitals that can be filled in this period. Therefore, 16 orbitals, at the maximum, can have 32 electrons and hence sixth period has 32 elements.

5. In terms of period and group, where would you locate the element with Z = 114?

Ans. The filling of the 6th period ends at $_{86}$Rn. Thereafter, the filling of 7th period starts. Therefore, after $_{86}$Rn, the next two elements with $Z = 87$ and $Z = 88$ are s-block elements, the next fourteen, i.e., $Z = 90–103$ are *f*-block elements, the next ten, i.e., $Z = 104 –112$ are d-block elements and the last six, i.e., $Z = 113 –118$ are p-block elements. Therefore, the element $Z = 114$ is the second p-block element (i.e., group 14) of the 7th period. Thus, the location of the element with $Z = 114$ in the period table is

Period = 7th, Block : *p,* Group : 14

6. Write the atomic number of the element in the third period and seventeenth group of the periodic table.

Ans. In the third period, the filling up of only 3s- and 3p-orbitals occurs. Therefore, in this period there are only two s-and six p-block elements. Since third period starts with $Z = 11$ and ends at $Z = 18$, therefore, elements with $Z = 11$

and $Z = 12$ are s-block elements. The next six elements with $Z = 13$ to 18 are p-block elements and belong to groups 13, 14, 15, 16, 17 and 18. Therefore, the element which will lie in seventeenth group will have $Z = 12 + 5 = 17$

7. Which element do you think would have been named by

(i) Lawrence Berkalay Laboratory

(ii) Seaborg's group?

Ans. (i) Lawrencium ($Z = 103$) and Berkeliuin ($Z = 97$)

(ii) Seaborgium ($Z = 106$)

8. Why do elements in the same group have similar physical and chemical properties?

Ans. Elements in the same group have same valence electron and hence have similar physical and chemical properties.

9. What does atomic radius or ionic radius mean to you?

Ans. Atomic radius literally means size of the atom. It can be measured either by X-ray or by spectroscopic methods. In case of non-metals, atomic radius is called covalent radius. It is defined as one-half the distance between the nuclei of two covalently bonded atoms of the same element in a molecule. For example, the internuclear distance between two chlorine atoms in chlorine molecule is 198 pm. Therefore, the covalent radius of chlorine atom is 198/2 =99 pm (0.99 Å). In case of metals, atomic radius is called metallic radius. It is defined as one-half the distance between the two adjacent atoms in the crystal lattice. For example, the distance between two adjacent copper atoms in solid copper is 256 pm, therefore, the metallic radius of copper is 256/2 = 128 pm (1.28 Å). Ionic radius means size of the ion. An ion can be a cation or an anion. The size of a cation is always smaller than that of the parent atom because loss of one or more electrons increases the effective nuclear charge. As a result, force of attraction increases and hence the ionic size decreases. On the other hand, the size of the anion is larger than the parent atom because the addition of one or more electrons decreases the effective nuclear charge. As a result, the force of attraction decreases and hence, the ionic size increases. For example, the ionic radius of Na^+ is 95 pm while the atomic radius of sodium is 186 pm. On the other hand, ionic radius of fluoride ion is 136 pm whereas the atomic radius of fluorine atom is only 72 pm.

10. How do atomic radii vary in a period and in a group ? How do you explain the variation?

Ans. The atomic radius increases down the group. This is because a new energy shell (i.e., principal quantum number increases by unity) is added at each

succeeding element while the number of electrons in the valence shell remains to be the same. As a result, the force of attraction of the nucleus for the valence electrons decreases and hence the atomic size increases. In contrast, the atomic size decreases as we move from left to right in a period. This is because that within a period the outer electrons remain in the same shell but the nuclear charge increases by one unit at each succeeding element. Due to this increased nuclear charge, the attraction of the nucleus for the outer electrons increases and hence the atomic size decreases.

11. What do you understand by iso-electronic species? Name the species that will be iso-electronic with each of the following atoms or ions.

(i) F^- (ii) Ar

(iii) Mg^{2+} (iii) Rb^+

Ans. Ions of different elements which have the same number of electrons but different magnitude of the nuclear charge are called isoelectronic ions.

(i) F^- has 10 (9 + 1) electrons. Therefore, species N^{3-}, O^{2-}, Ne, Na^+, Mg^{2+}, Al^{3+} each one of which contains 10 electrons, are isoelectronic with it.

(ii) Ar has 18 electrons. Therefore, the species, P^{3-}, S^{2-}, Cl^-, K^+, Ca^{2+} etc. each one of which contains 18 electrons, are isoelectronic with it.

(iii) Mg^{2+} has 10 (12 – 2) electrons, therefore, the species N^{3-}, O^{2-}, F^-, Ne, Na^+, Al^{3+}, etc. each one of which contains 10 electrons, are isoelectronic with it.

(iv) Rb^+ has 36 (37 – 1) electrons. Therefore, the species, Br^-, Kr and Sr^{2+} each one of which has 36 electrons, are isoelectronic with it.

12. Consider the following species: N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, and Al^{3+}.

(i) What is common in them?

(ii) Arrange them in order of increasing ionic radii ?

Ans. (i) Each one of these ions contains 10 electrons and hence, all are isoelectronic ions.

(ii) The ionic radii of isoelectronic ions decrease with the increase in the magnitude of the nuclear charge. For example, consider the isoelectronic ions:

N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+} and Al^{3+}.

All these ions have 10 electrons but their nuclear charges increase in the order:

N^{3-} (+7), O^{2-} (+8), F^- (+9), Na^+ (+11), Mg^{2+} (+12) and Al^{3+} (+13).

Therefore, their ionic radii decrease in the order:

$N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$.

13. Explain why cations are smaller and anions are larger in radii than their parent atoms.

Ans. The ionic radius of a cation is always smaller than the parent atom because the loss of one or more electrons increases the effective nuclear charge. As a result, the force of attraction of nucleus for the electrons increases and hence, the ionic radii decrease. In contrast, the ionic radius of an anion is always larger than its parent atom because the addition of one or more electrons decreases the effective nuclear charge. As a result, the force of attraction of the nucleus for the electrons decreases and hence, the ionic radii increase.

14. What is the significance of the terms 'isolated gaseous atom' and 'ground state' while defining the ionisation enthalpy and electron gain enthalpy?

Ans. (i) In the gaseous state, the atoms are widely separated, therefore, these inter-atomic forces are minimum. Further, since it is not possible to isolate a single atom for the purpose of determination of its ionisation enthalpy, therefore, the inter-atomic distances are further reduced by carrying out the measurement at a low pressure of the gaseous atom. It is because of these reasons, that the term isolated gaseous atom has been included in the definition of ionisation enthalpy.

(ii) The term ground state means that the atom must be present in the most stable state, i.e., the ground state. The reason being that when the isolated gaseous atom is in the excited state, lesser amount of energy will be released when it gets converted into gaseous anion after accepting an electron.

15. Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18} J atom^{-1}. Calculate the ionisation enthalpy of atomic hydrogen in terms of kJ mol^{-1}.

Ans. The energy required to remove an electron in the ground state of hydrogen atom

$= -$ (its energy in the ground state) $= -(-2.18 \times 10^{-18}\ J) = 2.18 \times 10^{-18}$ J

Ionisation energy per mole of hydrogen atoms

$$= \frac{2.18 \times 10^{-18} \times 6.02 \times 10^{23}}{1000}$$

$= 1312.36$ kJ mol^{-1} $= 1312.36 \times 10^{3}$ J mol^{-1}

16. Among the second period elements, the actual ionisation energies are in the order : Li < B < Be < C < O < N < F < Ne Explain why ?

Explain why

(i) Be has higher Δi H than B

(ii) O has lower Δi H than N and F ?

Ans. (i) In case of Be ($1s^2 2s^2$) the outermost electron is present in 2s-orbital while in B ($1\ s^2\ 2\ s^2\ 2p^1$) it is present in 2p-orbital. Since 2s-electrons are more strongly attracted by the nucleus than 2p-electrons, therefore, lesser amount of energy is required to knock out a 2p-electron than a 2s-electron. Consequently, $\Delta_i H$ of Be is higher than that $\Delta_i H$ of B.

(ii) The electronic configuration of N ($1\ s^2\ 2\ s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$) in which 2p-orbitals are exactly half-filled is more stable than the electronic configuration of O ($1s^2\ 2s^2\ 2p_x^2$, $2p_y^1$, $2p_z^1$) in which the 2p-orbitals are neither half-filled nor completely filled. Therefore, it is difficult to remove an electron from N than from O. As a result, $\Delta_i H$ of N is higher than that of O. Further, the electronic configuration of F is $1\ s^2\ 2\ s^2\ 2p_x^2\ 2p_y^2\ 2p_z^1$. Because of higher nuclear charge (+9), the first ionisation enthalpy of F is higher than that of O.

17. 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.

Ans. The electronic configurations of Na and Mg are:

Na : $1\ s^2\ 2s^2\ 2p^6\ 3s^1$ and Mg: $1\ s^2\ 2s^2\ 2p^6\ 3s^2$

Thus, the first electron in both the cases has to be removed from the 3s-orbital but the nuclear charge of Na (+11) is lower than that of Mg (+12), therefore, the first ionisation energy of sodium is lower than that of magnesium. After the loss of first electron, the electronic configuration of Na^+ is $1s^2\ 2s^2\ 2p^6$. Here, the electron is to be removed from inert (neon) gas configuration which is very stable and hence, removal of second electron from sodium is very difficult. However, in case of magnesium, after the loss of first electron, the electronic configuration of Mg^+ is $1s^2 2s^2\ 2p^6\ 3s^1$. Here, the electron is to be removed from a 3s orbital which is much easier than to remove an electron from inert gas configuration. Therefore, the second ionisation enthalpy of sodium is higher than that of magnesium.

18. What are the various factors due to which the ionisation enthalpy of the main group elements tends to decrease down the group?

Ans. Within the main group elements, the ionisation enthalpy decreases regularly as we move down the group due to the following two factors:

(i) **Atomic size:** On moving down the group, the atomic size increases gradually due to the addition of one new principal energy shell at each succeeding element. Consequently, the force of attraction of the nucleus for the valence electrons decreases and hence the ionisation enthalpy decreases.

(ii) **Screening effect:** With the addition of new shells, the number of inner electron shells which shield the valence electrons increase. 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 decreases.

19. The first ionisation enthalpy values (in kJ mol^{-1}) of group 13 elements are :

B	Al	Ga	In	Tl
801	577	579	558	589

How will you explain this deviation from the general trend?

Ans. On moving down the group 13 from B to Al, the ionisation enthalpy decreases as expected due to an increase in atomic size and screening effect which outweighs, the effect of increased nuclear charge.

However $\Delta_i H$ of Ga is only slightly higher (2 kJ mol^{-1}) than that of Al while that of Tl is much higher than those of Al, Ga and In. These deviations can be explained as follows:

Al follows immediately after. s-block elements while Ga and In follow after d-block elements and Tl after d- and f-block elements. These extra d- and f-electrons do not shield (or screen) the outer shell-electrons from the nucleus very effectively. As a result, the valence electrons remain more tightly held by the nucleus and hence, larger amount of energy is needed for their removal. This explains why Ga has higher ionisation enthalpy than Al. Further on moving down the group from Ga to In, the increased shielding effect (due to the presence of additional 4d-electrons) outweighs the effect of increased nuclear charge (49 – 31 = 18 units) and hence, the $\Delta_i H_1$ of In is lower than that of Ga. Thereafter, the effect of increased nuclear charge (81 – 49 = 32 units) outweighs the shielding effect due to the presence of additional 4f- and 5d-electrons and hence, the $\Delta_i H_1$ of Tl is higher than that of In.

20. Which of the following pairs of elements would have a more negative electron gain enthalpy

(i) O or F **(ii) F or Cl**

Ans. (i) As we move from O to F, the atomic size decreases and the nuclear charge increases. Both these factors tend to increase the attraction of the nucleus for the incoming electron and hence electron gain enthalpy becomes more negative i.e., electron gain enthalpy of F is much more negative (–328 kJ mol^{-1}) than that of oxygen (–141 kJ mol^{-1}).

(ii) The reason for this deviation is the small size of F atom. Due to its small size, the electron–electron repulsions in the relatively compact 2p-sub-shell are comparatively large, and hence, the incoming electron is not accepted with the same ease as is the case with larger Cl atom. Consequently, electron gain enthalpy of Cl is more negative than that of F.

21. Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first. Justify your answer.

Ans. The second electron gain enthalpy of O is positive as explained below:

When an electron is added to O atom to form O^{-1} ion, energy is released. Thus, first electron gain enthalpy of O is negative.

$O(g) + e^{-}(g) \rightarrow O^{-1}(g)$; $\Delta_{eg}H = -141$ kJ mol^{-1}

But when another electron is added to O^{-1} to form O^{-2} ion, energy is absorbed to overcome the strong electrostatic repulsion between the negatively charged O^{-1} ion and the second electron being added. Thus, the second electron gain enthalpy of oxygen is positive.

$O^{-1}(g) + e^{-}(g) \rightarrow O^{-2}(g)$; $\Delta_{eg}H = +78$ kJ mol^{-1}

22. What is the basic difference between the terms electron gain enthalpy and electronegativity?

Ans. Both electron gain enthalpy and electro negativity refer to the tendency of the atom of an element to attract electrons. Whereas electron gain enthalpy refers to the tendency of an isolated gaseous atom to accept an additional electron to form a negative ion, electro negativity refers to the tendency of the atom of an element to attract the shared pair of electrons towards it in a covalent bond.

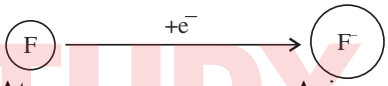
23. How would you react to the statement that the electronegativity of N on Pauling scale is 3.0 in all the nitrogen compounds.

Ans. The electro-negativity of any given atom is not constant. Therefore, the statement that the electro-negativity of N on Pauling scale is 3.0 in all nitrogen compounds is wrong. Actually, electro-negativity varies with the state of hybridisation and the oxidation state of the element. The electro-negativity increases as the percentage of s-character of a hybrid orbital increases or the

oxidation state of the element increases. For example, the electro negativity of N increases as we move from ,sp^3–sp^2–sp–hybrid orbital. Similarly, the electro-negativity of N in NO_2 where oxidation state of N is +4 is higher than in NO where the oxidation state of N is +2.

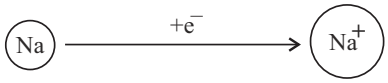
24 . Describe the theory associated with the radius of an atom as it (a) gains electron (b) loses electron.

Ans. (a) When a neutral atom gains one electron to form an anion, its radius increases. The reason being that the number of electrons in the anion increases while its nuclear charge remains the same as the parent atom. Since the same nuclear charge now attracts greater number of electrons, therefore, force of attraction of the nucleus on the electrons of all the shells decreases (i.e., effective nuclear charge decreases) and hence the electron cloud expands.



Atom	Anion
9 Electrons	10 Electrons
+ 9 Nuclear charge	+ 9 Nuclear charge
Ionic size: 64 pm	136 pm

(b) When a neutral atom loses one electron to form a cation, its atomic radius decreases. The reason being that the number of electrons in the cation decreases while its nuclear charge remains the same as the parent atom. Since the same nuclear charge now attracts lesser number of electrons, therefore, the force of attraction of the nucleus on the electrons of all the shells–increases (i.e., effective nuclear charges increases) and hence the size of cation decreases. Thus



Atom	Cation
11 Electrons	10 Electrons
+ 11 Nuclear charge	+ 11 Nuclear charge
Ionic size: 156 pm	95 pm

25. Would you expect the first ionisation enthalpies of two isotopes of the same element to be same or different? Justify your answer.

Ans. Two isotopes of the same element have same no. of electrons, same nuclear charge, same size and hence they are expected to have same ionisation

enthalpy.

26. What are major differences between metals and non-metals ?

Ans. Elements which have a strong tendency to lose electrons to form cations are called metals while those which have a strong tendency to accept electrons to form, anions are called non-metals. Thus, metals are strong reducing agents, they have low ionisation enthalpies, have less negative electron gain enthalpies, low electro-negativity, and form basic oxides and ionic compounds.

Non–metals, on the other hand, are strong oxidising agents, they have high ionisation enthalpies, have high negative electron gain enthalpies, high electro-negatively and form acidic oxides and covalent compounds.

27. Use the periodic table to answer the following questions. Identify an element with five electrons in the outer sub-shell.

(a) Identity an element with five electrons in the outer sub-shell.

(b) Identify an element that would tend to lose two electrons.

(c) Identify an element that would tend to gain two electrons.

(d) Identify the group having metal, non-metal, liquid as well as gas at room temperature.

Ans. (a) The general electronic configuration of the elements having five electrons in the outer sub-shell is $ns^2\ np^5$. This electronic configuration is characteristic of elements of group 17, i.e., halogens and their examples are F, Cl, Br, I, At, etc.

(b) The elements which have a tendency to lose two electrons must have two electrons in the valence shell. Therefore, their general configuration should be ns^2. This electronic configuration is characteristic of group 2 elements, i.e., alkaline earth metals and their examples are Mg, Ca, Sr, Ba, etc.

(c) The elements which have a tendency to accept two electrons must have six electrons in the valence shell. Therefore, their general electronic configuration is $ns^2\ np^4$. This electronic configuration is a characteristic of group 16 elements and their examples are O and S.

(d) A metal which is liquid at room temperature is mercury. It is a transition metal and belongs to group 12. A non-metal which is a gas at room temperature is nitrogen (group 15), oxygen (group 16), fluorine, chlorine (group 17) and inert gases (group 18).

A non-metal which is a liquid at room temperature is bromine (group

17).

28 The increasing order of reactivity among group 1 elements is Li < Na < K < Rb < Cs whereas that of group 17 is F > Cl > Br > I. Explain.

Ans. The elements of group 1 have only one electron in their respective valence shells and thus, have a strong tendency to lose this electron. The tendency to lose electrons, in turn, depends upon the ionisation enthalpy. Since the ionisation enthalpy decreases down the group, therefore, the reactivity of group 1 elements increases in the same order: Li < Na < K < Rb < Cs. In contrast, the elements of group 17, have seven electrons in their respective valence shells and thus, have a strong tendency to accept one more electron. The tendency to accept electrons, in turn, depends upon their electrode potentials. Since the electrode potentials of group 17 elements decrease in the order: F (+2.87 V) > Cl (1.36 V), Br (1.08 V) and I (+0.53V), therefore, their reactivity also decrease in the same order :

F > Cl > Br > I.

29. Write the general electronic configuration of s-, p-, d- and f-block elements :

Ans. (i) s-Block elements: ns^{1-2} where $n = 2 - 7$

(ii) p-Block elements: $ns^2\ np^{1-6}$ where $n = 2 - 6$

(iii) d-Block elements: $(n-1)\ d^{1-10}\ ns^{0-2}$ where $n = 4 - 7$

(iv) f-Block elements: $(n-2)f^{0-14}\ (n-1)\ d^{0-1}\ ns^2$ here $n = 6 - 7$

30. Assign the position of the element having outer electronic configuration,

(i) $ns^2\ np^4$ for $n = 3$ (ii) $(n-1)\ d^2\ n\ s^2$ for $n = 4$ and

(iii) $(n-2)f^7\ (n-1)\ d^1\ n\ s^2$ for $n = 6$ in the periodic table.

Ans. (i) When n = 3 , it suggests that the element belongs to third period. Since the last electron enters the p-orbital, therefore, the given element is a p-block element. Further since the valence shell contains 6 (2 + 4) electrons, therefore, group number of the element = 10 + no. of electrons in the valence shell = 10 + 6 = 16. The complete electronic configuration of the element is $1s^2 2s^2 2p^6\ 3s^2 3p^4$ and the element is S (sulphur)

(ii) $n = 4$ suggests that the element lies in the 4th period.

Since the d-orbitals are incomplete, therefore, it is d-block element. The group number of the element = no. of d-electrons +no. of s-electrons = 2 + 2 = 4. Thus, the elements lies in group 4 and 4th period. The complete electronic configuration of the element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$ and the element in Ti (titanium).

(iii) $n = 6$ means that the element lies in the sixth period. Since the last electron goes to the f-orbital,therefore, the element is a f–block element. All f-block elements lie in group 3. The complete electronic configuration of the element is [Xe] $4f^7\ 5d^1\ 6s^2$. The atomic number of the element 54 + 7 + 1 + 2 = 64 and the element Gd (gadolinium).

31. The first ($\Delta_i H_1$) and the ($\Delta_i H_2$) ionization enthalpies (in kJ mol^{-1}) and the ($\Delta_{eg} H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below:

Element	$\Delta_i H_1$	$\Delta_i H_2$	$\Delta_{eg} H$
I	520	7300	–60
II	419	3051	–48
III	1681	3374	–328
IV	1008	1846	–295
V	2372	5251	+48
VI	738	1451	–40

Which of the above element is likely to be:

(a) the least reactive metal

(b) the most reactive metal

(c) the most reactive non-metal

(d) the least reactive non-metal

(e) the metal which can form a stable binary halide of the formula MX_2 (X = halogen).

(f) the metal which can form predominantly stable covalent halide of the formula MX (X = halogen) ?

Ans. (a) The element V has highest first ionisation enthalpy (Δi H1) and positive electron gain enthalpy (Δeg H) and hence, it is the least reactive element. Since inert gases have positive $\Delta i\ H_1$, therefore, the element V must be an inert gas. The values of $\Delta i\ H_1$, $\Delta i\ H_2$ and $\Delta_{eg}\ H$ match that of He.

(b) The element II which has the least first ionisation enthalpy ($\Delta i\ H_1$) and a low negative electron gain enthalpy ($\Delta_{eg}\ H$) is the most reactive metal. The values of $\Delta_i\ H_1$, $\Delta_i H_2$ and $\Delta_{eg} H$ match that of K (potassium).

(c) The element III which has high first ionisation enthalpy ($\Delta_1 H_1$) and a very high negative electron gain enthalpy ($\Delta_{eg}\ H$) is the most reactive non-metal. The values of $\Delta i H_1$, $\Delta i H_2$ and $\Delta_{eg} H$ match that

of F (fluorine).

(d) The element IV has a high negative electron gain enthalpy (Δ_{eg} H) but not so high first ionisation enthalpy ($\Delta i\ H_1$). Therefore, it is the least reactive non-metal. The values of $\Delta i\ H_1$, ΔiH_2 and $\Delta_{eg}H$ match that of I (Iodine).

(e) The element VI has low first ionisation enthalpy ($\Delta i\ H_1$) but higher than that of alkali metals. Therefore, it appears that the element is an alkaline earth metal and hence will form binary halide of the formula MX_2 (where X = halogen). The values of $\Delta i\ H_1$, $\Delta i\ H_2$ and $\Delta_{eg}H$ match that of Mg (magnesium).

(f) The element I has low first($\Delta i\ H_1$) but a very high second ionisation enthalpy ($\Delta i\ H_2$), therefore, it must be an alkali metal. Since the metal forms a predominantly stable covalent halide of the formula MX (X = halogen), therefore, the alkali metal must be least reactive. The values of $\Delta i\ H_1$, $\Delta i\ H_2$ and $\Delta_{eg}H$ match that of Li (lithium).

32. Predict the formula of the stable binary compounds that would be formed by the combination of the following pairs of elements.

(a) Lithium and oxygen **(b) Magnesium and nitrogen**

(c) Aluminium and iodine **(d) Silicon and oxygen**

(e) Phosphorus and fluorine **(f) Element 71 and fluorine.**

Ans. (a) Lithium is an alkali metal (Group 1). It has only one electron in the valence shell, therefore, its valency is 1. Oxygen is a group 16 element with a valency of 2. Therefore, formula of the compound formed would be Li_2O (lithium oxide).

(b) Magnesium is an alkaline earth metal (Group 2) and hence, has a valency of 2. Nitrogen is a group 15 element with a valency of 8 – 5 = 3. Thus, the formula of the compound formed would be Mg_3N_2 (magnesium nitride).

(c) Aluminium is group 13 element with a valency of 3 while iodine is a halogen (group 17) with a valency of 1. Therefore, the formula of the compound formed would be $A1I_3$ (aluminium iodide).

(d) Silicon is a group 14 element with a valency of 4 while oxygen is a group 16 element with a valency of 2. Hence, the formula of the compound formed is SiO_2 (silicon dioxide).

(e) Phosphorus is a group 15 element with a valency of 3 or 5 while fluorine is a group 17 element with a valency of 1. Hence, the formula of the compound formed would be PF_3 or PF_5.

(f) Element with atomic number 71 is a lanthanoid called lutetium (Lu). Its common valency is 3. Fluorine is a group 17 (halogen) element with a valency of 1. Therefore, the formula of the compound formed would be LuF_3 (lutetium fluoride).

33 In the modern periodic table, the period indicates the value of

(a) atomic number **(b) mass number**

(c) principal quantum number **(d) azimuthal quantum number**

Ans. In the modern periodic table, each period begins with the filling of a new shell. Therefore, the period indicates the value of principal quantum number. Thus, option (c) is correct.

34. Which of the following statements related to the modern periodic table is incorrect?

(a) The p-block has six columns, because a maximum of 6 electrons can occupy all the orbitals in a p-sub-shell.

(b) The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-sub-shell.

(c) Each block contains a number of columns equal to the number of electrons that can occupy that sub-shell.

(d) The block indicates value of azimuthal quantum number (l) for the last sub-shell that received electrons in building up the electronic configuration.

Ans. Statement (b) is incorrect while other statements are correct. The correct statement is: the d-block has 10 columns, because a maximum of 10 electrons can occupy all the orbitals in a d-sub-shell.

35 Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell

(a) Valence principal quantum number (*n*)

(b) Nuclear charge (*Z*)

(c) Nuclear mass

(d) Number of core electrons.

Ans. Nuclear mass does not affect the valence shell because nucleus consists of protons and neutrons. Where protons, i.e., nuclear charge affects the valence shell but neutrons do not. Thus, option (c) is wrong.

36 The size of isoelectronic species – F^-, Ne and Na^+ is affected by

(a) nuclear charge (*Z*)

(b) valence principal quantum number (*n*)

(c) electron–electron interaction in the outer orbital

(d) none of factors because their size is the same.

Ans. The size of the isoelectronic ions depends upon the nuclear charge (*Z*). As the nuclear charge increases the size decreases. For example, F^- (+9) > Ne (+10) > Na^+ (+11). Therefore, statement (a) is correct while all other statements are wrong.

37 Which of the following statements is incorrect in relation to ionisation enthalpy?

(a) Ionisation enthalpy increases for each successive electron.

(b) The greatest increase in ionisation enthalpy is experienced on removal of electrons from core noble gas configuration.

(c) End of valence electrons is marked by a big jump in ionisation enthalpy.

(d) Removal of electron from orbitals bearing lower n value is easier than from orbital having higher n value.

Ans. Statement (d) is incorrect. The correct statement is: Removal of electron from orbitals bearing lower *n* value is difficult than from orbital having higher n value. All other statements are correct.

38 Considering the elements B, Al, Mg and K, the correct order of their metallic character is:

(a) B > Al > Mg > K **(b) Al > Mg > B > K**

(c) Mg > Al > K > B **(d) K > Mg > Al > B**

Ans. In a period, metallic character increases as we move from right to left. Therefore, metallic character of K, Mg and Al decreases in the order: K > Mg > Al. However, within a group, the metallic character, increases from top to bottom. Thus, Al is more metallic than B. Therefore, the correct sequence of decreasing metallic character is: K > Mg > Al > B, i.e., option (d) is correct.

39 Considering the elements B, C, N, F and Si, the correct order of their non-metallic character is

(a) B > C > Si > N > F **(b) Si > C > B > N > F**

(c) F > N > C > B > Si **(d) F > N > C > Si > B**

Ans. In a period, the non-metallic character decreases from right to left. Thus, among B, C, N and F, non-metallic character decreases in the order: F > N > C > B. However, within a group, non-metallic character decreases from top to bottom. Thus, C is more non-metallic than Si. Therefore, the correct sequence of decreasing non-metallic character is : F > N > C > B > Si, i.e.,

option (c) is correct.

40 Considering the elements F, Cl, O and N, the correct order of their chemical reactivity in terms of oxidising property is:

(a) F > Cl > O > N **(b) F > O > Cl > M**

(c) Cl > F > O > N **(d) O > F > N > Cl**

Ans. Within a period, the oxidising character increases from left to right. Therefore, among F, O and N, oxidising power decreases in the order: F > O > N. However, within a group, oxidising power decreases from top to bottom. Thus, F is a stronger oxidising agent than Cl. Further because O is more electronegative than Cl, therefore, O is a stronger oxidising agent than Cl. Thus, over all decreasing order of oxidising power is :

F > O > Cl > N, i.e., option (b) is correct.