Q-1) What is the basic theme of organization in the periodic table?

Ans.) It is to characterize the elements in periods and groups as per their properties. So, this course of action makes the investigation of elements and compounds of elements in a simple and methodical way. In this periodic table, elements with comparative properties are set in a similar group.

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

Ans.) Mendeleev organised the components in his periodic table, according to the order of their atomic weight. Mendeleev organized the components in groups and periods according to the increasing atomic weight. Mendeleev set the elements which are having comparative properties in similar groups.

Nonetheless, he didn't adhere to an arrangement that he gave for long. He discovered that if the elements were organized according to their increasing atomic weights, then a few elements did not match within this plan of characterization.

In this manner, he overlooked the order of atomic weights now and again. For instance, the atomic mass of iodine is lower than the atomic mass of tellurium.

Still, Mendeleev set tellurium (in Group 6) ahead of iodine (in Group 7) essentially in light of the fact that iodine's properties are so comparable to fluorine, chlorine, and bromine.

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

Ans.)

Mendeleev's Approach for periodic law	**Modern approach for the periodic law**
Chemical properties and physical properties of the elements are the periodic functions of the atomic mass of the corresponding elements.	Chemical properties and physical properties of the elements are the periodic functions of the atomic numbers of the corresponding elements.

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

Ans.) In a periodic table containing elements, a period shows the value of a principal quantum number (n) for the furthest shells. Every period starts with the filling with the principal quantum number (n). And ***n***'s value for the 6th period is equal to 6. Now, for n = 6, the azimuthal quantum number (l) can have "0, 1, 2, 3, 4" values.

As indicated by Aufbau's rule, electrons will be added to various orbitals according to their increasing energies. Here, the 6d subshell is having much higher energy than the energy of 7s subshell.

In the sixth period, the electrons can occupy in just 6s, 4f, 5d, and 6 p subshells. 6s is having 1 orbital, 4f is having 7 orbitals, 5d is having 5 orbitals, and 6p is having 3 orbitals. In this way, there are a sum of 16 (1 + 7 + 5 + 3 = 16) orbitals accessible. As indicated by Pauli's exclusion, one orbital can only accommodate at max 2 electrons.

Hence, sixteen orbitals can have 32 electrons.

Subsequently, the 6th period of the period table ought to have 32 elements.

Q-5) In terms of period and group where would you locate the element with Z =114?

Ans.) Elements whose atomic number is from Z = 87 to Z = 114 are available in the seventh period of the periodic table. Therefore, the element having Z = 114 is available in the seventh period in the periodic table.

In the seventh period, initial 2 elements with Z = 87 and Z= 88 are the elements of s-block and the following 14 elements except Z = 89 i.e., those from Z = 90 to Z = 103 are elements of f – block, and next 10 elements from Z = 89 and Z = 104 to Z = 112 are elements of d-block, next the elements from Z = 113 to Z = 118 are elements of p-block. In this manner, the element Z = 114 is the 2^{nd} element of p-block in the seventh period of the periodic table.

Therefore, the element Z = 114 is available in the seventh period and fourth group in the periodic table.

Q-6) What is the atomic number of element keeping in mind both the cases given below;

1. Element is in 3^{rd} period of the periodic table.

2. Element is in 17^{th} group of the periodic table.

Ans.) The first period is having 2 elements and the second period is having 8 elements. So, the third period begins with element Z = 11. Presently, the third period contains 8 elements. So, the 18^{th} element is the last element of the third period and this 18^{th} element is present in the 18^{th} group. Thus, the element in the seventeenth group of the 3rd period is having atomic number 17 i.e. Z = 17.

Q-7) Which element do you think would have been named by
(i) Lawrence Berkeley Laboratory
(ii) Seaborg's group?

Ans.)

(i) Lawrencium (Lr) which has an atomic number, Z = 103 and Berkelium (Bk) which has an atomic number, Z = 97

(ii) Seaborgium (Sg) which has an atomic number, Z = 106

Q-8) Why do elements in the same group have similar physical and chemical properties?

Ans.) The chemical and physical properties of any elements rely on the number of valence electrons. In periodic table elements are in the same group are having the same quantity of valence electrons. This is why elements present in the same group are having similar chemical and physical properties.

Q-9) What does atomic radius and ionic radius really mean to you?

Ans.) The radius of an atom is known as the atomic radius. It quantifies the size of an atom. On the chance that the element is a metal, then its radius is termed as metallic radius, and if an element is a non-metal, then its radius is termed as covalent radius. The metallic radius can be calculated as the inter-nuclear distance between two molecules divided by 2. For instance, the inter-nuclear distance between two adjoining copper atoms is 256 pm in solid copper.

Metallic radius of copper = $\frac{256}{2}pm = \ 128pm$

Covalent radius can be measured as the interatomic distance between 2 atoms when they are together by a solitary bond in a covalent atom. For instance, the interatomic distance between 2 chlorine atoms of chlorine molecule = 198 pm.

Covalent radius of copper = $\frac{198}{2}pm = \ 99pm$

The radius of an ion (cation or anion) is known as ionic radius. Ionic radius is computed by measuring the inter-ionic distance between the cation and anion in an ionic crystal. Since cations are created by expelling an electron from the outermost orbit of an atom, thus cation has less electrons compared to parent atom which results in increased effective nuclear charge.

In this way, a cation is small in size than the parent atom. For instance, the ionic radius of Na^+ ion (sodium ion) = 95 pm, while the atomic radius of Na (sodium) atom = 186 pm. An anion is bigger in size than the parent atom. It is because an anion is having the same nuclear charge, yet more number of electrons compared to the parent atom which results in increased repulsion within atom among the electrons which also results in decreased effective nuclear charge. For instance, ionic radius of F^- (fluorine ion) = 136 pm, while the atomic radius of F (fluorine) atom = 64 pm.

Q-10) How do atomic radius vary in a period and in a group? How do you explain the variation?

Ans.) Atomic radius declines as we move from left to right in a period. It happens because in a period, the external electrons are available in a similar valence shell so, the atomic number increments from left to right in a period, which results in an increase in the effective nuclear charge. Therefore, the attraction of electrons towards the nucleus is increased.

Also, atomic radius declines as we move from top to bottom in the group. It happens because as we move down in a group then there is an increase in principal quantum number(n) which brings about an increase in the distance between the nucleus and the valence electrons.

Q-11) What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

(i) F^-

(ii) Ar

(iii) Mg^+

(iv) Rb^+

Ans.) Ions and atoms which are having equal numbers of the electrons are called the isoelectronic species.

(i) F^- (Fluorine) ion is having 10 electrons (9 + 1 = 10). Hence, the species which is isoelectronic with F^- must also have 10 electrons.

It's some isoelectronic species are

a) Na^+ ion it is also having 10 electrons (11 – 1 = 10).

b)Ne ion it is also having 10 electrons.

c) Al^{3+} ion it is also having 10 electrons (13 – 3 = 10).

(ii) Ar (Argon) is having 18 electrons. Hence, the species which is isoelectronic with Ar must also have 18 electrons.

It's some isoelectronic species are

a) S^{2-} ion it is also having 18 electrons (16 + 2 = 18).

b) Cl^- ion it is also having 18 electrons (17 + 1 = 18).

c) K^{+} ion it is also having 18 electrons (19 – 1 = 18).

(iii) Mg^{+} (Magnesium) ion is having 11 electrons (12 – 1 = 11). Hence, the species which is isoelectronic with Mg^{+} must also have 11 electrons.

It's some isoelectronic species are

a) Al^{2+} ion it is also having 11 electrons (13 – 2 = 11).

b)Na ion it is also having 11electrons.

c) Si^{3+} ion it is also having 11 electrons (14 – 3 = 11).

(iv) Rb^{+} (Rubidium) is having 36 electrons (37 – 1 = 36). Hence, the species which is isoelectronic with Rb^{+} must also have 36 electrons.

It's some isoelectronic species are

a) Br^{-} ion it is also having 36 electrons (35 + 1 = 36).

b)Kr ion it is also having 36 electrons.

c) Sr^{2+} ion it is also having 36 electrons (38 – 2 = 36).

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

(a) What is common in them?
(b) Arrange them in the order of increasing ionic radii.

Ans.)

(a) The species that are given are having an equal number of electrons i.e. 10 electrons. So, they are isoelectronic species.

(b) Arrangement of the given ions according to their increasing order of ionic radii is:

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

Q-13) Explain why cation is smaller and anions larger in radii than their parent atoms?

Ans.) Cations are formed by expelling an electron from the outermost orbit of an atom, thus cation has less electrons compared to parent atom which results in increased effective nuclear charge but the total nuclear charge remains same which results in the increased attraction of electrons towards nucleus than that of the parent atom. Thus, cations are having smaller radii than that of their parent atom.

Anions are formed by gaining an electron in the outermost orbit of an atom. Thus anion has more electrons compared to parent atom, which results in decreased effective nuclear charge but the total nuclear charge remains the same which results in the increased distance the nucleus and the valence electrons as the attraction of electrons towards nucleus decreases than that of the parent atom. Thus, anions are having larger radii than that of their parent atom.

Q-14) What is the significance of the terms — 'isolated gaseous atom' and 'ground state' while defining the ionization enthalpy and electron gain enthalpy?

Ans.) "Ionization enthalpy is the energy that is required to expel an electron from an isolated gaseous atom in the ground state". Despite the fact that in the gaseous state the atoms are generally widely separated, there are a few measures of attractive forces between the atoms. To find the ionization enthalpy of any ion, it is difficult to isolate a solitary atom. This attractive force can be further diminished by bringing down the pressure. Hence, the term "isolated gaseous atom" is utilized as a part of the meaning of ionization enthalpy.

An atom's ground state is the most stable state. Less energy is required to expel an electron if an isolated gaseous atom is present in the ground state. In this way, for the purpose of comparison, electron gain enthalpy and ionization enthalpy must be calculated for an "isolated gaseous atom" and its "ground state".

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

Ans.) Here it is given that, electron of hydrogen is having $-2.18 * 10^{-18} J$ in ground state.

Thus, $-2.18 * 10^{-18} J$ amount of energy will be required to expel an electron from ground state in

H(hydrogen) – atom.

∴ For Hydrogen atom the Ionization of enthalpy = $-2.18 * 10^{-18} J$

Thus, ionization enthalpy of a Hydrogen atom in $J\ mol^{-1}$

= $-2.18 \times\ 10^{-18} \times\ 6.02 \times\ 10^{23} J\ mol^{-1}$

Q-16) Among the second period elements the actual ionization enthalpies are in the order Li < B < Be < C < O < N < F < Ne.

Explain Why

(i) Be has higher $\Delta_i H$ than B

(ii) O has lower $\Delta_i H$ than N and F?

Ans.)

(i) During the ionization process, the electron that can be expelled from Be(beryllium) – atom is 2s – electron, but the electron that can be expelled from boron is 2p – electron.

The attractive force between a 2s – electron and nucleus are higher than between a 2s – electron and nucleus.

Thus, the energy required to expel 2s –electron is higher than the energy required to expel 2p –electron.

Thus, $\Delta_i H$ for Be is higher than $\Delta_i H$ than B.

(ii) In nitrogen, there are three 2p-electrons and all of these 3 occupy 3 distinct atomic orbitals. While in oxygen 2 out of 4, 2p – electrons occupy the same 2p-orbital, so the repulsion between the electrons in the oxygen atom increases.

Thus, the energy required to expel 2^{nd} 2p –electron in oxygen atom is higher than the energy required to expel 4^{th} 2p –electron in nitrogen atom.

Thus, $\Delta_i H$ for O is lower than $\Delta_i H$ of N.

The fluorine atom is having one proton and one electron more than that in the oxygen atom. As the electron is added to a similar shell, the increment in an attractive force between nucleus and electron (as a proton is added) is higher than the increment in the repulsive force between electron-electron(as an electron is added). Thus, valence electrons in the fluorine atom experience a higher effective nuclear charge compared to that, which is experienced by an electron of the oxygen atom. Thus, the energy required to expel an electron from fluorine is higher than the energy required to expel an electron from oxygen.

Thus, $\Delta_i H$ for O is lower than $\Delta_i H$ of F.

Q-17) How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium?

Ans.) The 1st ionization enthalpy of magnesium is higher than 1st ionization enthalpy of sodium because,

1. Magnesium is having greater atomic size than sodium.

2. Magnesium is having higher effective nuclear charge than sodium.

Thus, energy required to expel an electron from sodium is lower than that in magnesium. Thus, the 1st ionization enthalpy of magnesium is higher than 1st ionization enthalpy of sodium.

The 2nd ionization enthalpy of magnesium is lower than 2nd ionization enthalpy of sodium is because after expelling an electron, there is still 1 electron remaining in the 3s-orbital of magnesium, whereas sodium achieves stable inert gas configuration after expelling an electron. So, magnesium still requires to expel 1 electron to achieve stable inert gas configuration.

Thus, energy required to expel 2nd electron from magnesium is lower than that in sodium. Thus, the 2nd ionization enthalpy of magnesium is lower than 2nd ionization enthalpy of sodium.

Q-18) What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down a group?

Ans.) The factors because of which in elements of the main group the ionization enthalpy decreases when we move down the group are given below:

1. “Increase in the shielding effect”: Inner shells increases as we move down the group. Thus, the shielding effect of valence electrons increases by inner core electrons from the nucleus. Thus, the attractive force on electrons towards the nucleus is very strong. So, the energy required to expel a valence electron decreases as we move down the group.

2. “Increase in atomic size of elements”: Inner shells increases as we move down the group. Thus, the atomic size increases as we move down the group. Also, the distance between the valence electron and nucleus of an atom, as a result, the electrons are not strongly bounded. So, valence electrons can be expelled easily. Thus, the energy required to expel a valence electron decreases as we move down the group.

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

B	**Al**	**Ga**	**In**	**Tl**
801	577	579	558	589

How would you explain this deviation from the general trend?

Ans.) Inner shells increase as we move down the group. Thus, the shielding effect of valence electrons increases by inner core electrons from the nucleus. Thus, the attractive force on electrons towards the nucleus is very strong. So, ionization enthalpy decreases as we move down the group. Hence for elements of group 13 the ionization enthalpy decreases as we move down from B to Al.

Here, Ga is having high ionization enthalpy than that of Al. This is because Al comes after the s-blocks elements, while Ga comes after the d-blocks elements. The shielding that is provided by electrons of d-block elements is not effective. So, the valence electrons are not shielded effectively. Thus, valence electrons in Ga atom experience higher effective nuclear charge compared to Al.

Further on moving down from Ga to In, the value of ionization enthalpy is decreased because of the increase in the shielding effect and an increase in atomic size.

But, Tl is having high ionization enthalpy than that of In. This is because Tl comes after the '4f and 5d electrons'. The shielding that is provided by these '4f and 5d electrons' is not effective. So, the valence electrons are not shielded effectively. Thus, valence electrons in Tl atom experience higher effective nuclear charge compared to In.

Q-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) O and F are the elements of the same period in the periodic table. An F-atom is having 1 electron and 1 proton more than that of O-atom as the electron is added in the same shell, thus the atomic size of O-atom is larger than F-atom. As O-atom is having 1 proton less than F-atom. So, the nucleus of O-atom cannot attract an incoming electron that strongly as that of an F-atom. Also, F-atom requires only 1 electron to achieve a stable inert gas configuration. So, the electron affinity of F(Fluorine) is more negative than that of O(oxygen).

(ii) F and Cl are the elements of the same group in the periodic table. On moving down the group the electron affinity becomes less negative. Here, the value of electron affinity of F is less negative than that of Cl. It is because the atomic size of Cl is larger than that of F. In Cl, the electron will be added to n = 3 quantum level, whereas in F, the electron will be added to n = 2 quantum level. Thus, as the electron-electron repulsion is reduced in Cl so an extra electron can easily be accommodated. So, the electron affinity of Cl is more negative compared to that of F.

Q-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.) O^- ion is formed when O-atom gains one electron and the energy is being released during this process.

So the 1st electron affinity for O-atom is negative.

$$O_{(g)} + e^- \rightarrow O_g^-$$

If an electron is added in O^- ion then it forms O^{2-} ion, the energy is required to be given to counter the strong electronic repulsions. So, the 2nd electron affinity of O-atom is positive.

$$O^-_{(g)} + e^- \rightarrow O^{2-}_g$$

Q-22) What is the basic difference between the terms electron gain enthalpy and electronegativity?

Ans.)

Electron gain enthalpy	Electronegativity
Tendency to gain electrons for an isolated gaseous atom is its electron gain enthalpy.	Tendency to attract the shared pairs of electrons for an atom which is in chemical compound is its electronegativity.

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

Ans.) Electronegativity is a variable property of any element. Electronegativity is different for different compounds. Hence, the given statement "The electronegativity of Nitrogen on Pauling scale measures 3.0 for all nitrogen compounds" is incorrect. This is because the electronegativity of Nitrogen is different in NO_2 and NH_3.

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

Ans.)

(a) As an atom expels a single electron, so the quantity of electron decreases by 1 but the nuclear charge does not change. Thus, the electron-electron repulsion decreases in an atom. So, there is an increase in the effective nuclear charge. Thus, there is a decrease in the radius of an atom.

(b) There is an increase in the size of an atom when it gains an electron. As it gains an electron then the quantity of electrons raises by 1. Thus, the electron-electron repulsion increases in an atom. There is an increase in effective nuclear charge as the quantity of proton remains unchanged. Thus, there is an increase in the radius of an atom.

Q-25) Would you expect the first ionization enthalpies for two isotopes of the same element to be the same or different? Justify your answer

Ans.) Ionization enthalpy of any atom relies on its number of protons and electrons. But, the isotopes of any element have an equal number of electrons and protons. Thus, the 1st ionization enthalpy of two isotopes of a single element is the same.

Q-26) What are the major differences between metals and non-metals?

Ans.)

Metals		**Non- metals**	
a	They can easily expel an electron	a	They cannot easily expel an electron.
b	They cannot easily receive an electron.	b	They can easily receive an electron.
c	They form ionic compounds.	c	They form covalent compounds.
d	Their oxides are having basic nature.	d	Their oxides are having acidic nature.
e	Their ionization enthalpies are low.	e	Their ionization enthalpies are high.
f	Their electron affinity is less negative.	f	Their electron affinity is highly negative.
i	Their electronegativity is less.	i	Their electronegativity is more.
j	Their reducing power is high.	j	Their reducing power is low.

Q-27) Use the periodic table to answer the following questions.
(a) Identify an element with five electrons in the outer subshell.
(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 the room temperature

Ans.)

(a) Element with 5 electrons in outer subshell is having electronic configuration ns^2np^5. Halogen group is having the same electronic configuration. So, the elements can be At, I, Br, Cl and F.

(b) The elements that tend to gain 2 electrons are those who can achieve stable inert gas configuration by gaining these electrons. They are having electronic configuration ns^2np^4. Oxygen containing group is having the same electronic configuration.

(c) The elements that tend to expel 2 electrons are those who can achieve stable inert gas configuration by expelling these electrons. They are having electronic configuration ns^2. 2nd Group elements are having the same electronic configuration. So, the elements can be Ba, Sr, Ca, Mg, Be.

(d) Elements of group 17 are having liquid, gas, metal and non-metals as well at room temperature.

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

Ans.) For group 1

- There is only 1 valance electron in the elements of group 1. So, their tendency is to expel this electron in order to achieve a stable inert gas configuration. As we move down in group 1 the ionization enthalpies of the elements decreases. Thus, less energy will be required to expel the valence electron. As a result reactivity of elements increases as we move down in group 1.

For group 17

- In the elements of group 17, there is a requirement of only 1 electron in order to achieve stable inert gas configuration. So, their tendency is to gain this 1 electron. As we move down in group 17 the ionization enthalpies of the elements increases. Thus, more energy will be required to expel the valence electron. As a result reactivity of elements decreases as we move down in group 17.

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

Ans.)

Element	General electronic configuration
s-block	ns^{1-2} ; n = 2 to 7
p-block	ns^2np^{1-6} ; n = 2 to 6
d-block	$(n-1)d^{1-10}ns^{0-2}$; n = 4 to 7
f-block	$(n-2)f^{1-14}d^{0-10}ns^2$; n = 6 to 7

Q-30) Assign the position of the element having outer electronic configuration (i) ns^2np^4 for n=3 (ii) (n-1) d^2ns^2 for n=4, and (iii) (n-2) f^7 (n-1) d^1 ns^2 for n=6, in the periodic table.

Ans.)

(i) Here, n = 6 so the element is in the sixth period. The element is an 'f-block element' because the last electron enters in the f-orbital. As the f-block elements are in the third group. They are having electronic configuration

$[Xe]4f^7 5d^1 6s^2$. So, the atomic number can be calculated as 54 + 7 + 2 + 1 = 64. Thus, the required element is Gadolinium.

(ii) Here, n = 3 so the element is in the third period. The element is in 'p-block element' because the last electron enters in the p-orbital. It contains 4 electrons in p-orbital.

For a group of the element

= No. of s – block groups + No. of d – block groups + No. of p – block groups

= 2 + 10 + 4 = 16

Thus, the given element is in the third period and sixteenth the group in the periodic table. Thus, the required element is Sulphur.

(iii) Here, n = 4 so the element is in the fourth period. The element is in 'd-block element' because the last electron enters in the d-orbital but this orbital is incompletely filled. It contains 2 electrons in d-orbital.

For a group of the element

= No. of s – block groups + No. of d – block groups

= 2 + 2 = 4

Thus, the given element is in the fourth period and fourth the group in the periodic table. Thus, the required element is Titanium.

Q-31) The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) 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:

Elements	$(\Delta_i H_1)$	$(\Delta_i H_2)$	$(\Delta_{eg} H)$
1	520	7300	-60
2	419	3051	-48
3	1681	3374	-328
4	1008	1846	-295

5	2372	5251	+48
6	738	1451	-40

Which of the above elements is likely to be:

(a) the least reactive element.
(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 a predominantly stable covalent halide of the formula MX (X=halogen)?

Ans.)

(a) As element 5 is having highest 1st ionization enthalpy and is having a positive electron affinity, so it is the most reactive among all the elements given.

(b) As element 2 is having lowest 1st ionization enthalpy and is having a negative electron affinity, so it is the least reactive among all the elements given.

(c) As element 5 is having high 1st ionization enthalpy and is having a positive electron affinity, so it is the least reactive non-metal.

(d) As element 3 is having high 1st ionization enthalpy and is having a negative electron affinity, so it is the most reactive non-metal.

(e) As element 1 is having low 1st ionization enthalpy and high 2nd ionization enthalpy. Thus, metal can easily form predominantly stable covalent halide having formula MX; X = halogen.

(f) As element 6 is having low 2nd ionization enthalpy and is having a negative electron affinity so, it is a metal. Thus, metal can easily form stable binary halide having formula MX_2; X = halogen.

Q-32) Predict the formulas 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) LiO_2

(b) Mg_3N_2

(c) AlI_3

(d) SiO_2

(e) PF_5 or PF_3

(f) Lutetium (Lu) is the element 71. It is having valency 3. Thus, the required formula is LuF_3.

Q-33) In the modern periodic table, the period indicates the value of:
(a) atomic number
(b) atomic mass
(c) principal quantum number
(d) azimuthal quantum number.

Ans.)

(c) The period in Modern periodic table indicates the value of 'n' i.e. a principal quantum number.

Q-34) Which of the following statements related to the modern periodic table is incorrect?
(a) The p-block has 6 columns because a maximum of 6 electrons can occupy all the orbitals in a p-shell.
(b) The d-block has 8 columns because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.
(c) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.
(d) The block indicates the value of an azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.

Ans.)

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

Q-35) " Anything that influences the valence electrons will affect the chemistry of the element". Which of the factors given below is not affecting the valence shell?

(a) Valence Principal quantum number (n)

(b) Nuclear charge (Z)

(c) Nuclear mass

(d) Number of core electrons

Ans.)

(c) Nuclear mass

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

Ans.)

(a) Nuclear charge (Z)

Because for isoelectronic species:

The atomic size decreases with an increase in nuclear charge (Z).

e.g. the arrangement according to increasing order of nuclear charge (Z) for $F^{-}, Ne\ and\ Na^{+}$ is:

$$F^{-} < Ne < Na^{+}$$

And the arrangement according to increasing order of atomic size for $F^{-}, Ne\ and\ Na^{+}$ is:

$$Na^{+} < Ne < F^{-}$$

Q-37) Which one of the following statements is incorrect in relation to ionization enthalpy?
(a) Ionization enthalpy increases for each successive electron.
(b) The greatest increase in ionization enthalpy is experienced on the removal of an electron from core noble gas configuration.
(c) End of valence electrons is marked by a big jump in ionization enthalpy.
(d) Removal of an electron from orbitals bearing lower n value is easier than from orbital having higher n value

Ans.)

(d) is the incorrect statement

For orbitals having a lower value of ‘n’ the removal of an electron is easy compared to the orbitals having a higher value of ‘n’.

Because the electrons of orbitals having a lower value of ‘n’ are highly attracted to the nucleus than that of the electrons of orbitals having a higher value of ‘n’

Q-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**

(d) K > Mg > Al > B

Reason:

As we move from left and right in a period the metallic character of the elements decreases. Thus, Mg > Al.

As we move down from a group the metallic character of the elements decreases. Thus, Al > B.

From the above two statements, it can be stated that K > Mg.

Thus, K > Mg > Al > B

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

B > C > Si > N > F

Si > C > B > N > F

F > N > C > B > Si

F > N > C > Si > B

Ans.)

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

Reason:

As we move from left and right in a period the non-metallic characteristic of the elements decreases. Thus, F > N > C > B.

As we move down from a group the metallic character of the elements decreases. Thus, C > Si.

From the above two statements, it can be stated that B > Si.

Thus, F > N > C > B > Si

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

(a) F > Cl > O > N

(b) F > O > Cl > N

(c) Cl > F > O > N

(d) O > F > N > Cl

Ans.)

(b) F > O > Cl > N

Reason:

As we move from left and right in a period the non-metallic characteristic of the elements increases. Thus, F > O > N.