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

Ans.) The basic theme of organization in the periodic table is to classify the elements in periods and groups as per their properties. So, this course of action makes the study of elements and their compounds simple and systematic. In the periodic table, elements with similar properties are placed in the same 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. He organized the components in groups and periods according to the increasing atomic weight. He placed the elements with similar properties in the same group.

However, he did not stick to this arrangement for long. He discovered that if the elements were organized according to their increasing atomic weights, then some of the elements did not fit in with his scheme of classification

Hence, he ignored the order of atomic weights in some cases. For example, 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) along with fluorine, chlorine, bromine because of similarities in properties.

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

Ans.)

Mendeleev's Periodic law	**Modern Periodic Law**
Mendeleev's Periodic Law states that the physical and chemical properties of elements are periodic functions of their atomic weights.	Modern Periodic Law states that the physical and chemical properties of elements are periodic functions of their atomic numbers.

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 outermost shells. Every period starts with the filling of the principal quantum number (n). The value of n 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.

According to Aufbau's principle, electrons are added to various orbitals in order of their increasing energies. Here, the 6d subshell is having much higher energy than the energy of the 7s subshell.

In the sixth period, the electrons can be filled in only 6s, 4f, 5d, and 6p subshells. 6s has 1 orbital, 4f has 7 orbitals, 5d has 5 orbitals, and 6p has 3 orbitals. Hence, there are a sum of 16 (1 + 7 + 5 + 3 =

16) orbitals available. As per Pauli's exclusion principle, each orbital can accommodate a maximum of 2 electrons.

Therefore, sixteen orbitals can accommodate a maximum of 32 electrons.

Thus, the 6th period of the periodic table should have 32 elements.

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

Ans.) Elements with the atomic numbers from Z = 87 to Z = 114 are present in the seventh period of the periodic table. Therefore, the element with Z = 114 is present in the seventh period of the periodic table.

In the seventh period, first 2 elements with Z = 87 and Z= 88 are s-block elements, the next 14 elements except Z = 89 i.e., those with Z = 90 to Z = 103 are f – block elements, and next 10 elements with Z = 89 and Z = 104 to Z = 112 are d-block elements and the elements with Z = 113 to Z = 118 are p-block elements. Hence, the element with Z = 114 is the second p-block element in the seventh period. Thus, the element with Z = 114 is present in the seventh period and fourteenth group of the periodic table.

Therefore,

Period = 7^{th}, Group = 14 and Block = p-block.

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.)

In the third period, we are provided that element. The highest principal quantum number (n) of the element into which the last electron enters is known as the period number. As a result, n=3 for the third period. In addition, the element can be found in the seventeenth group. The elements of the seventeenth group have the following general configuration: ns^2np^5.

As a result, the needed element's overall electronic configuration is $3s^23p^5$ (because n=3).

The element's complete electrical configuration is now: $1s^22s^22p^63s^23p^{5.}$

Now add up the total number of electrons in the element's ground state: 1+2+6+2+5=17 electrons.

We know that an element's atomic number equals the total number of electrons in its ground state.

The element has 17 electrons in its ground state, as determined above, and hence the atomic number of the element is 17. Chlorine is an atomic number 17 element (Cl).

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 elements rely on the number of valence electrons. In periodic table, elements present in the same group have the same number of valence electrons. Therefore, elements present in the same group have similar chemical and physical properties.

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

Ans.) In a molecule, the atomic radius is half the distance between the nuclei of two covalently bound atoms of the same element.

The atomic radius of metals is referred to as the metallic radius. In a crystal lattice, it equals one-half of the distance between two adjacent atoms.

Radius of Ionisation:

The size of an ion, whether it's a cation or an anion, is referred to as its ionic radius.

This is the effective distance between the nucleus of the ions and the ionic bond that it influences.

The size of the cation is always smaller than that of the parent atom, but the size of the anion is always greater.

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 present in a similar valence shell and the atomic number increases 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 increases.

The atomic radius generally increases down a group. This is because, down a group, the principal quantum number(n) increases which results in 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^{2+}

(iv) Rb^{+}

Ans.) Ions and atoms which contain the same number of electrons are called isoelectronic species

(i) F^{-} (Fluorine) ion has 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 (11 – 1 = 10 electrons).

b) Ne ion (10 electrons).

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

(ii) Ar (Argon) has 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 (16 + 2 = 18 electrons).

b) Cl^{-} ion (17 + 1 = 18 electrons).

c) K^{+} ion (19 – 1 = 18 electrons).

(iii) Mg^{2+} (Magnesium) ion has 10 electrons (12 – 2 = 10). Hence, the species which is isoelectronic with Mg^{2+} must also have 10 electrons.

It's some isoelectronic species are

a) Al^{3+} ion (13 – 3 = 10 electrons).

b) Ne ion (10 electrons).

c) F^{-} ion (9 + 1 = 10 electrons).

(iv) Rb^{+} (Rubidium) has 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 (35 + 1 = 36 electrons).

b) Kr ion (36 electrons).

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

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 given species has an equal number of electrons i.e. 10 electrons. Hence, they are isoelectronic species.

(b) The arrangement of the given species in order of their increasing ionic radii is as follows:

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

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

Ans.) A cation has a fewer number of electrons than its parent atom, while its nuclear charge remains the same. Hence, a cation is smaller in size than its parent atom.

On the other hand, an anion has one or more electrons than its parent atom, resulting in an increased repulsion among the electrons and a decrease in the effective nuclear charge. Therefore, an anion is larger in radii than its parent atoms.

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.) The significance of the term **'isolated gaseous atom'** indicates that the atoms in the gaseous phase are so far separated that there does not have any mutual attraction or repulsion interactions present which is an isolated state. Here the value of ionisation enthalpy and electron gain enthalpy are not influenced by the presence of the other atoms.

The significance of the term **'ground state'** means that in an atom, electrons are present in the lowest energy state where they neither lose nor gain an electron. Ionisation enthalpy and electron gain enthalpy are generally expressed with respect to the ground state of an atom only.

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.) Given

Energy of an electron in the ground state of the hydrogen atom = -2.18×10^{-18} J

Thus energy required to remove that electron from the ground state of hydrogen atom is 2.18×10^{-18} J

Hence, ionization enthalpy of atomic hydrogen = 2.18×10^{-18} J

Therefore, the ionization enthalpy of atomic hydrogen in terms of J mol^{-1} = $2.18 \times 10^{-18} \times 6.02 \times 10^{23}$ J mol^{-1}

We get,

= 1.31×10^{6} 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) The electronic configuration of Be is $1s^2 2s^2$ and the electronic configuration of B is $1s^2 2s^2 2p^1$. As it is known that 2p orbital lies far away from the nucleus in comparison to 2s orbital. Thus, the attraction between the 2s electron and nucleus is more in comparison to the attraction between 2p electron and nucleus. Thus, the amount of energy required for the removal of electrons from 2s orbital is more in comparison to 2p electrons. Thus, ionisation enthalpy of Be is higher in comparison to B .

(ii) The electronic configuration of O is $1s^2 2s^2 2p^4$, N is $1s^2 2s^2 2p^3$ and F is $1s^2 2s^2 2p^5$. As it is known that the half filled and fully filled configuration are more stable; thus, the removal of electrons from half-filled N is difficult in comparison to oxygen. Thus, the value of $\Delta_i H$ for N is higher in comparison to O. The nuclear charge over fluorine is more in comparison to oxygen atoms. Thus, its ionisation enthalpy is higher than oxygen. Thus, O has lower $\Delta_i H$ than N and 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 electronic configurations of Na and Mg are:

$_{11}Na$: $1s^2 2s^2 2p^6 3s^1$

$_{12}Mg$: $1s^2 2s^2 2p^6 3s^2$

The first ionization enthalpy of sodium is lower than that of magnesium because the nuclear charge of Na (+11) is lower than that of Mg (+12) and size of Na is lower than Mg. After the loss of first electron, the electronic configuration of Na^+ is $1s^2 2s^1 2p^6$ (has noble gas configuration) and Mg^+ is $1s^2 2s^2 2p^6 3s^1$.

Hence removal of the second electron from Na^+ is very difficult as compared to Mg^+ (electron to be removed from 3s). Thus second ionisation enthalpy of sodium is higher than that of magnesium.

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 following are the reasons that cause the ionisation enthalpy of the major group elements to decrease as we move down the group:

(i) As we progress down the group, the nuclear charge increases.

(ii) As we progress down the group, the atomic size grows as each element gains a new shell.

(iii) The number of inner electrons grows as we descend in a group. The shielding effect on the outer electrons is reduced as a result. The combined effect of atomic size growth and shielding effect is greater than the effect of nuclear charge increase. The force of attraction between the nucleus and the outermost electron is reduced as a result of these events.

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

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.) The following steps can be used to explain the current trend:

(i) As the atom size increases from B to Al, the value of the ionisation enthalpy decreases.

(ii) Moving from Al to Ga, Ga has ten electrons that do not screen like Sulphur and Phosphorus do. As a result, the value of effective nuclear charge increases unexpectedly, resulting in enhanced ionisation energy value.

(iii) When moving from Ga to In and Tl, there are 14 electrons in Tl with a poor shielding effect, increasing the effective nuclear charge and hence the ionisation energy value.

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 present in the same period of the periodic table. An F atom has one proton and one electron more than that of O atom and as an electron is added to the same shell, the atomic size of F is smaller than that of O. Since F contains one proton more than O, its nucleus can attract the incoming electron more strongly in comparison to the nucleus of O atom. Also, F needs only one more electron to achieve the stable noble gas configuration. Hence, the electron gain enthalpy of F is more negative than that of O.

(ii) F and Cl are the elements of the same group in the periodic table. The electron gain enthalpy usually becomes less negative on moving down a group. However, in this case, the value of the electron gain enthalpy of Cl is more negative than that of F. This is because the atomic size of F is smaller than that of Cl. In F, the electron will be added to quantum level n = 2, whereas in Cl, the electron is added to quantum level n = 3. Hence, there are less electron- electron repulsions in Cl and an additional electron can be accommodated easily. Therefore, the electron gain enthalpy of Cl is more negative than 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.) When an electron is added to O atom to form O^- ion, energy is released. Therefore, the first electron gain enthalpy of O is negative.

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

On the other hand, when another electron is added to O^- ion to form O^{2-} ion, energy has to be given out to overcome the strong electronic repulsion between the negatively charged O^- ion and the second electron being added. Hence, the second electron gain enthalpy of oxygen is positive.

$$O_{(g)}^- + e^- \rightarrow O_g^{2-}$$

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

Ans.)

Electron gain enthalpy	Electronegativity
Electron gain enthalpy refers to the tendency of an isolated gaseous atom to accept an additional electron to form a gaseous negative ion.	Electronegativity refers to the tendency of the atom of an element to attract the shared pair of electrons towards it in a covalent bond.

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 of any element depends on the hybridisation state and oxidation state of that element in a particular compound, i.e., the electronegativity of an element varies from compound to compound. For example electronegativity of the N. atom varies as: $sp^3-N < sp^2-N < sp-N$. So, the given statement is not correct.

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

Ans.)

(a)When an atom gains an electron, its size increases. When an electron is added, the number of electrons goes up by one. This results in an increase in repulsion among the electrons. However, the number of protons remains the same. As a result, the effective nuclear charge of the atom decreases and the radius of the atom increases.

(b)When an atom loses an electron, the number of electrons decreases by one while the nuclear charge remains the same. Hence, the interelectronic repulsions in the atom decrease. As a result, the effective nuclear charge increases. Thus, the radius of the atom decreases.

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.) The ionization enthalpy of any atom relies on the number of electrons and protons (nuclear charge) of that atom. Now, the isotopes of an element have an equal number of protons and electrons. Hence, the first ionization enthalpy for two isotopes of the same element should be the same.

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

Ans.)

Metals		Non- metals	
a	These are solids at room temperature except mercury	a	These exist in all three states
b	These are very hard except sodium	b	These are soft except diamond
c	These are malleable and ductile	c	These are brittle and can break down into pieces
d	These are shiny	d	These are non-lustrous except iodine
e	Electropositive in nature	e	Electronegative in nature
f	Have high densities	f	Have low densities
i	Their electronegativity is less.	i	Their electronegativity is more.
j	They are good conductors of heat and electricity	j	They are poor conductors of heat and electricity

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) The electronic configuration of an element having 5 electrons in its outermost subshell should be ns^2np^5. This is the electronic configuration of the halogen group. Hence, the element can be F, Cl, Br, I, At.

(b) An element having two valence electrons will lose two electrons easily to achieve the stable noble gas configuration. The general electronic configuration of such an element will be ns^2. This is the electronic configuration of group 2 elements. The elements present in group 2 are Be, Mg, Ca, Sr, Ba.

(c)An element is likely to gain two electrons if it needs only two electrons to achieve the stable noble gas configuration. Hence, the general electronic configuration of such an element should be ns^2np^4. This is the electronic configuration of the oxygen family.

(d) Elements of group 17 are having metal, non-metal, liquid as well as gas 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.) The tendency of metals to lose electrons increases as we go down a group. So, the reactivity of metals increases down the group. Thus, in group 1, the reactivity follows the order

$$\underrightarrow{\underset{Least Reactive}{Li} < Na < K < Rb < \underset{Most reactive}{Cs}}_{Reactivity increases}$$

The reactivity of non-metals in a group decreases as we go down the group. This is because the tendency to accept electrons decreases down the group.

$$\underrightarrow{\underset{Most Reactive}{F} > Cl > Br > \underset{Least reactive}{I}}_{Reactivity increases}$$

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

Ans.) General outer electronic configuration of s-, p-, d- and f- block elements:

(i) s- Block elements:

ns^{1-2} where n = 2 to 7

(ii) p- Block elements:

ns^2np^{1-6} where n = 2 to 6

(iii) d- Block elements:

$(n-1)d^{1-10}ns^{0-2}$ where n = 4 to 7

(iv) f- Block elements:

$(n-2)f^{0-14}$ $(n-1)d^{0-1}$ ns^2 where n = 6 to 7

Q-30) Assign the position of the element having outer electronic configuration (i) $ns^2\,np^4$ for n=3 (ii) $(n-1)d^2\,ns^2$ 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 = 3, so the element belongs to the third period. It is the 'p-block element' since the last electron occupies the p-orbital. There are 4 electrons in the p-orbital.

Hence, the corresponding group of the element

= Number of s–block groups + Number of d–block groups + Number of p–block groups

= 2 + 10 + 4 = 16

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

(ii) Here n = 4, so the element belongs to the fourth period. It is the 'd-block element' since d- orbitals are incompletely filled. There are 2 electrons in the d-orbital.

Hence, the corresponding group of the element

= Number of s–block groups + Number of d–block groups

= 2 + 2 = 4

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

(iii) Here n = 6, so the element belongs to the sixth period. It is a 'f-block element' since the last electron occupies the f-orbital. It belongs to group 3 of the periodic table since all f-block elements belong to group 3. Its electronic configuration is $[Xe]^{54}$ $4f^7$ $5d^1$ $6s^2$. Hence, its atomic number is 54 + 7 + 2 + 1 = 64. Therefore, the required element is Gadolinium.

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) Element 5 is likely to be the least reactive element since it has the highest first ionization enthalpy ($\Delta_i H_1$) and a positive electron gain enthalpy ($\Delta_{eg}H$).

(b) Element 2 is likely to be the most reactive metal since it has the lowest first ionization enthalpy ($\Delta_i H_1$) and a low negative electron gain enthalpy ($\Delta_{eg}H$).

(c) Element 3 is likely to be the most reactive non metal since it has the highest first ionization enthalpy ($\Delta_i H_1$) and the highest negative electron gain enthalpy ($\Delta_{eg}H$).

(d) Element 5 is likely to be the least reactive non metal since it has a very high first ionization enthalpy ($\Delta_i H_1$) and a positive electron gain enthalpy ($\Delta_{eg}H$).

(e) Element 6 has a low negative electron gain enthalpy ($\Delta_{eg}H$). Hence, it is a metal. Further, it has the lowest second ionization enthalpy ($\Delta_i H_2$). Therefore, it can form a stable binary halide of the formula MX_2 (X = halogen).

(f) Element 1 has the low first ionization energy and high second ionization energy. Hence, it can form a predominantly stable covalent halide of the formula MX (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) Li_2O (lithium oxide).

(b) Mg_3N_2

(c) AlI_3

(d) SiO_2

(e) PF_5 or PF_3

(f) Lutetium (Lu) is the element 71. It has valency 3. Therefore, 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.) In the modern periodic table, each period begins with the filling of new shell. Hence, the period indicates the value of principal quantum number. Thus option (c) is correct.

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

Option (b) is an incorrect statement because the d-block has 10 columns and maximum of 10 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)

The size of an isoelectronic species increases with a decrease in the nuclear charge (Z).

Example: The order of the increasing nuclear charge of F^-, Ne and Na^+ is**:**

$$F^- < \ Ne < \ Na^+$$

Therefore the order of the increasing size of 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.) Option (d) is the incorrect statement

Electrons in orbitals bearing a lower n value are more attracted to the nucleus than electrons in orbitals bearing a higher n value. Therefore, the removal of electrons from orbitals bearing a higher n value is easier than the removal of electrons from orbitals having a lower n value.

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**

Ans.)

(d) K > Mg > Al > B

Reason:

The metallic character of elements decreases from left to right across a period. Hence, the metallic character of Mg is more than that of Al. The metallic character of elements increases down a group. Therefore, the metallic character of Al is more than that of B.

Thus, the correct order of metallic character is K > Mg > Al > B.

Q-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.)

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

Reason:

The non-metallic character of elements increases from left to right across a period. Therefore, the decreasing order of non-metallic character is F > N > C > B.

The non-metallic character of elements decreases down a group. Thus, the decreasing order of non-metallic characters of C and Si are C > Si. However, Si is less non-metallic than B i.e., B > Si.

Thus, the correct order of their non-metallic characters is 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:

The oxidizing character of elements increases from left to right across a period. Thus, we get the decreasing order of oxidizing property as F > O > N.

The oxidizing character of elements decreases down a group. Hence, we get F > Cl.

But, the oxidizing character of O is more than that of Cl i.e., O > Cl.

Therefore, the correct order of chemical reactivity of F, Cl, O, and N in terms of their oxidizing property is F > O > Cl > N.