Chapter – 3 (Classifications of Elements and periodicity)

Exercise Questions:

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

Answer:

The basic theme of organization of elements in the periodic table is to classify the elements in period and groups according to their properties. This arrangement 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'

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

Answer:

Mendeleev was the first to develop a periodic table & he gave a law called Mendeleev periodic law which states that the physical & chemical properties of the elements are a periodic function of their atomic masses. On the basis of this law he developed a Mendeleev periodic table, where he arranged the elements in his periodic table ordered by atomic weight or mass. He arranged the elements in periods and groups in order of their 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 found out that if the elements were arranged strictly in order of their increasing atomic weights, then some elements did not fit within this scheme of classification.

Therefore, he ignored the order of atomic weights in some cases. For example, the atomic weight of iodine is lower than that of tellurium. Still Mendeleev placed tellurium (in Group VI) before iodine (in Group VII) simply because iodine's properties are so similar tofluorine, chlorine, and bromine

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

Answer:

Mendeleev's periodic law states that the physical and chemical properties or the elements are periodic function of their atomic weight. On the other hand, the Modern periodic law states that the physical and chemical properties of the elements are the periodic function of their atomic numbers.

Question: 4 On the basic of quantum numbers, justify that the sixth period of the

periodic table should have 32 elements.

Answer:

Sixth period corresponds to n = 6. In this period 16 orbitals, viz, one 6 s, seven 4f, five 5d and three 6p orbitals are filled. These sixteen orbitals can accommodate 32 elements. So, there are 32 elements in

the sixth period

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

Answer:

Elements with atomic numbers from Z = 87 to Z = 114 are present in the 7th period of the periodic table. Thus, the element with Z = 114(Flerovium) with atomic weight 289 and a poor metal is present in the 7th period & 14th group of the periodic table

In the 7th period, first two elements with Z = 87 and Z= 88 are s-block elements, the next 14 elements excluding Z = 89 i.e., those with Z = 90 – 103 are f – block elements, ten elements with Z = 89 and Z = 104 – 112 are d – block elements, and the elements with Z = 113 – 118 are p – block elements. Therefore, the element with Z = 114 is the second p – block element in the 7th period

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

Answer:

There are two elements in the 1^{st} period and eight elements in the second period. The third period starts with the element with z = 11. Now there are eight elements in the third period. Thus, the third period ends with the element with z = 18 I,e, the element in the 18^{th} group of the 3^{rd} period has z = 18. Hence, the element in the 17^{th} group of the third period has atomic number z = 17.

Question: 7 Which element do you think would have been named by

I) Lawrence Berkeley Laboratory

II) Seaborg's group?

Answer:

i.) Lawrencium (Lr) with z = 103 and Berkelium (Bk) with z = 97

ii.) Seaborgium (Sg) with z = 106.

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

Answer:
The physical and chemical properties of an elements depend on the no. of valence electrons. Elements present in the same group have the same number of valence electrons. Therefore, elements present in the same group have similar physical and chemical properties.

Question: 9 What does atomic radius and ionic radius really mean to you?

Answer:
Atomic radius & ionic radius are the periodic properties which are directly or indirectly related to the electronic configuration of their atoms & shows gradation on moving down a group or along a period.
Atomic radius is defined as the distance from the centre of the nucleus to the outer most shell containing the electrons. It measures the size of an atom.It is of 3 types:
A) Covalent radius- It is the one half of the distance between the centres of the nuclei of two adjacent similar atoms joined to each other by single covalent bond.
Covalent radius = inter nuclear distance in the bonded atoms/ 2

B) Metallic radius- It is defined as half the distance between the centres of the nuclei of two adjacent atoms in the metallic crystal

C) Van der waal's radius- It is defined as one half of the inter nuclear distance between 2 similar adjacent atoms belonging to the two neighbouring molecules of the same substance in the solid state.

Ionic radius means the radius of an ion (cation or anion). It is defined as the distance from the centre of the nucleus of the ion upto which it exerts its influence on the electron cloud.The ionic radii can be calculated by measuring the distances between the cations and anions in ionic crystals.

Since a cation is formed by removing an electron from an atom, the cation has fewer electrons than the parent atom resulting in an increase in the effective nuclear charge. Thus, a cation is smaller than the parent atom. For example, the ionic radius of Na^+ ion is 95 pm, whereas the atomic radius of Na atom is 186 pm. On the other hand, an anion is larger in size than its parent atom. This is because an anion has the same nuclear charge, but more electrons than the parent atom resulting in an increased repulsion among the electrons and a decrease in the effective nuclear charge. For example, the ionic radius of F^- ion is 136 pm, whereas the atomic radius of F atom is 64 pm.

Question: 10 How do atomic radius vary in a period and in a group? How do you explain the variation?

Answer:
Atomic radius generally decreases from left to right across a period. This is because within a period, the outer electrons are present in the same valence shell and the atomic no. increases from left to right

across a period, resulting in an increased nuclear charge. As a result, the attraction of electrons to the nucleus increased.
On the other hand, the atomic radius generally increase down a group. This is because down a group, the principal quantum number increases which results in an increase of the distance b/w the nucleus and valence electrons.

Question: 11 What do you understand by isoelectric species? Name a species that will be isoelectric with each of the following atoms or ions.

- **I) F-**
- **II) Ar**
- **III) Mg^2+**
- **IV) Rb^+**

Answer:
Isoelectronic species/ions/atoms are the species which have same number of electrons but different magnitude of nuclear charges & belongs to different atoms or ions. The isoelectronic ions with greater nuclear charge will have small size as compared to the ion with smaller nuclear charge.

(i) F^- ion has 9 + 1 = 10 electrons. Thus, the species isoelectronic with it will also have 10 electrons. Some of its isoelectronic species are Na^+ ion (11 – 1 = 10 electrons), Ne (10 electrons), O^{2-} ion (8 + 2 = 10 electrons), and Al^{3+} ion (13 – 3 = 10 electrons).

(ii) Ar has 18 electrons. Thus, the species isoelectronic with it will also have 18 electrons. Some of its isoelectronic species are S^{2-} ion (16 + 2 = 18 electrons), Cl^- ion (17 + 1 = 18 electrons), K^+ ion (19 – 1 = 18 electrons), and Ca^{2+} ion (20 – 2 = 18 electrons).

(iii) Mg^{2+} ion has 12 – 2 = 10 electrons. Thus, the species isoelectronic with it will also have 10 electrons. Some of its isoelectronic species are F^- ion (9 + 1 = 10 electrons), Ne (10 electrons), O^{2-} ion (8 + 2 = 10 electrons), and Al^{3+} ion (13 – 3 = 10 electrons).

(iv) Rb^+ ion has 37 – 1 = 36 electrons. Thus, the species isoelectronic with it will also have 36 electrons. Some of its isoelectronic species are Br^- ion (35 + 1 = 36 electrons), Kr (36 electrons), and Sr^{2+} ion (38 – 2 = 36 electrons).

Question: 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 the order of increasing ionic radii.**

Answer:

(a) Each of the given species (ions) has the same number of electrons (10 electrons). Hence, the given species are isoelectronic, i.e

N^{3-} has 7+3 = 10 electrons

O^{2-} has 8+2= 10 electrons

F^{-} has 9+1 = 10 electrons

Na^{+} has 11-1 = 10 electrons

Mg^{2+} has 12-2 = 10 electrons

Al^{3+} has 13-3= 10 electrons

(b) The ionic radii of isoelectronic species increases with a decrease in the magnitudes of nuclear charge.

The arrangement of the given species in order of their increasing nuclear charge is as follows:

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

Nuclear charge = +7 +8 +9 +11 +12 +13

Therefore, the arrangement of the given species in order of their increasing ionic radii is as follows:

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

Question: 13 Explain why cation smaller and anions larger in radii than their parent atoms?

Answer:

Cations are smaller & anions larger in radii than their parent atoms because there is a loss of electrons from valence shell in the case of cation, so decreases the shell number and in the case of anion there is addition of electrons. which cause decreases in size of cation and increases the size of anion.

Question: 14 What is the significance of the terms – 'isolated gaseous atom' and 'ground state' while defining the ionization enthalpy and electron gain enthalpy?

Answer:

Ionization enthalpy is the minimum amount of energy which is needed to remove the most loosly bounded electron from a neutral isolated gaseous atom to form a cation. The cations are formed when the neutral atoms loses electrons. But for losing electrons they, should be in isolated gaseous form . Although the atoms are widely separated in the gaseous state, there are some amounts of attractive forces among the atoms. To determine the ionization enthalpy, it is impossible to isolate a single atom.

But, the force of attraction can be further reduced by lowering the pressure. For this reason, the term 'isolated gaseous atom' is used in the definition of ionization enthalpy.

Ground state of an atom refers to the most stable state of an atom. If an isolated gaseous atom is in its ground state, then less amount energy would be required to remove an electron from it. Therefore, for comparison purposes, ionization enthalpy and electron gain enthalpy must be determined for an 'isolated gaseous atom' and its 'ground

Question: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.

Answer:

The energy of an electron in the ground state of hydrogen atom is – 2.18×10^{-18}J. Therefore, the energy required to remove that electron from the ground state of hydrogen atom is 2.18×10^{-18} J.

: Ionisation enthalpy of atomic hydrogen = 2.18×10^{-18}J.

Hence, ionization enthalpy of atomic hydrogen in terms of J/ mol

= $2.18 \times 10^{-18} \times 6.02 \times 10^{23}$ J/ mol = 1.31×10^{6}Jmol^{-1}.

Question: 16 Among the second period elements the actual ionization enthalpy are in the order Li < B < Be < C < O < N < F < Ne.

I) Be has higher Δi H than B

II) O has lower Δi H than N and F?

Answer:

(a).Be has higher first ionization enthalpy than B because Be has more stable electronic configuration $1s^2 2s^2$ than B $1s^2 2s^2 2p^1$.

(b). ΔI H1 of O is expected to be more than that of N but is actually lesser because the electronic configuration of N is more symmetrical as well as stable in comparison to O. Δi H1 of O is less than that of F because the ionization enthalpy in general increases along a period

Question: 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?

Answer:

Electronic configuration of Na is $1s^2 2s^2 2p^6 3s^1$. After losing one electron from its outermost shell, sodium easily attains stable electronic configuration ($1s^2 2s^2 2p^6$), while magnesium does not lose its electron easily due to presence of two electrons in s-orbital ($ls^2 2s^2 2p^6 3s^2$). Hence first ionisation energy of sodium is less than magnesium.

When one electron is removed from Na and Mg, their configurations become $1s^2 2s^2 2p^6$ and $1s^2 2s^2 2p^6 3s^1$ respectively. Now it is easier to remove one electron from 3s of Mg + than $2p^6$ of Na+. Hence, second ionisation energy of Mg is less than Na.

Question: 18 What are the various factors due to which the organization enthalpy of the main group elements tends to decrease down a group?

Answer:

The minimum amount of energy which is needed to remove the most loosely bound electron from a neutral isolated atom to form a cation is called as ionization energy or enthalpy.

The factors responsible for the ionization enthalpy of the main group elements to decrease down a group are listed below:

(i) Increase in the atomic size of elements: As we move down a group, the number of main energy shells(n) increases. As a result, the atomic size also increases gradually on moving down a group. As the distance of the valence electrons from the nucleus increases, the electrons are not held very strongly. Thus, they can be removed easily. Hence, on moving down a group, ionization energy decreases.

(ii) Increase in the shielding effect: The number of inner shells of electrons increases on moving down a group. Therefore, the shielding of the valence electrons from the nucleus by the inner core electrons increases down a group. As a result, the valence electrons are not held very tightly by the nucleus. Hence, the energy required to remove a valence electron decreases down a group

Question: 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?

Answer:

The given trend can be explained by the following steps:

(i) Moving from B to Al, there is an increase in the size of the atom as a result decrease in the value of ionization enthalpy.

(ii) Moving from Al to Ga, there are 10 electrons in Ga which do not screen as is done by Sulphur and Phosphorus. Therefore, there is an unexpected increase in the value of effective nuclear charge resulting in increased ionization energy value.

(iii) Moving from Ga to In and Tl, there are 14 electrons in Tl with very poor shielding effect, which increases the effective nuclear charge thus the value of ionization energy increases.

Question: 20 Which of the following pairs of elements would have a more negative electron gain enthalpy?

I) O or F

II) F and Cl

Answer:

(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 O and as an electron is being added to the same shell, the atomic size of F is smaller than that of O. As 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 attain the stable noble gas configuration. Hence, the electron gain enthalpy of F is more negative than that of O.

(ii) F and Cl belong to the same group of 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, but in Cl, the electron is added to quantum level n = 3. Therefore, there are less electron-electron repulsions in Cl and an additional electron can be accommodated easily. Hence, the electron gain enthalpy of Cl is more negative than that of F.

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

Answer:

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

$$O_{(g)} - e^- \rightarrow O^-_{(g)}$$

On the other hand, when an electron is added to O^- ion to form O^{2-} ion, energy has to given out in order to overcome the strong electronic repulsion. Thus, the second electron gain enthalpy of O is positive.

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

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

Answer:

Electron gain enthalpy is the measure of the tendency of an isolated gaseous atom to accept an electron, whereas electronegativity is the measure of the tendency of an atom in a chemical compound to attract a shared pair of electrons.

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

Answer:

Electronegativity of an element is a variable property. It is different in different compounds. Hence, the statement, which says that the electronegativity of N on Pauling scale is 3.0 in all nitrogen compounds is incorrect. The electronegativity of nitrogen is different in NH3 and NO2.

Question: 24 Describe the theory associated with the radius of an atoms as it

a) **gain in electron**

b) **loses an electron**

Answer:

a.) When an atom gains an electron, its size increases. When an electron is added, the number of electrons goes up 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 atoms increases.

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

Question: 25 Would you expect the ionization enthalpies for two isotopes of the same element to be the same or different? Justify your answer.

Answer:

The ionization enthalpy of an atom depends on the number of electrons and protons of the atom. Now, the isotopes of an element have the same no. of electrons and protons. Therefore, the first ionization enthalpy for the two isotopes of the same element should be the same.

Question: 26 What are the differences between metals and non-metals?

Answer:

Metals	No – metals
Metals can lose electrons easily	They cannot lose electrons easily.
They cannot gain ekectrons easily.	Non – metals can gain electrons easily.
They generally form ionic compounds	They generally form covalent compounds.
Metal oxides are basic in nature.	Non- metallic oxides are acidic in nature.
They have low ionization enthalpy.	They have high ionization enthalpy.
Metals are less electronegative. They are rather electropositive elements.	Non – metals are electronegative
Metals have a high reducing power.	They have a low reducing power.

Question: 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.**

Answer:

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

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

c.) An element is likely to gain two electrons if it needs only two electrons to attain the noble gas configuration. Thus the general electronic configuration of such elements is ns^2np^4.this is the electronic configuration of oxygen family.

d.) Group 17 has metal, non- metal, liquid as well as gas at room temperature.

Question: 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.

Answer:

In the case of group 1, we have ordered Li < Na < K < Rb <Cs because there is only one valence electron and thus have a strong tendency to lose this electron. The tendency to lose electrons in turn also depends upon the ionization enthalpy. Moving down the group there is a decrease in the ionization enthalpy, therefore, increasing reactivity down the group.

In the case of group 17, we have the order F > CI > Br > I because there are seven electrons present in the valence shells and thus have a strong tendency to accept electrons to make it a stable noble gas electronic configuration. So, moving down we have a decrease in both electron gain enthalpy and electronegativity. Hence there is a decrease in the reactivity also.

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

Answer:

Elements	General outer electronic configuration
S – block	ns^{1-2}, where n = 2 – 7
P – block	ns^2np^{1-6}, where n = 2 – 6
D - block	$(n – 1)d^{1-10}ns^{0-2}$, where n = 4-7
F - block	$(n – 2)f^{1-14}(n – 1)d^{0-10}ns^2$, where n = 6-7

Question: 30 Assign the position of the elements having outer electronic configuration

I) ns^2np^4 for n = 3

II) (n – 1) d^2ns^2 for n = 4

III) (n – 2) f^7 (n – 1) d^1ns^2 for n = 6, in the periodic table.

Answer:

(i) Since n = 3, the element belongs to the 3rd period. It is a p–block element since the last electron occupies the p–orbital.

There are four electrons in the p–orbital. Thus, the corresponding group of the element

= Number of s–block groups($3s^2$) + number of d–block groups([Ne]10 + number of p–electrons($3p^4$)

= 2 + 10 + 4

= 16

Therefore, the element belongs to the 3rd period and 16th group of the periodic table. Hence, the element is Sulphur ($[Ne]^{10}3s^23p^4$)

(ii) Since n = 4, the element belongs to the 4th period. It is a d–block element as d–orbitals are incompletely filled.

There are 2 electrons in the d–orbital.

Thus, the corresponding group of the element

= Number of s–block groups($4s^2$) + number of d–block groups($3d^2$)

= 2 + 2

= 4

Therefore, it is a 4th period and 4th group element. Hence, the element is Titanium ($[Ar]^{18}3d^2 4s^2$.

(iii) Since n = 6, the element is present in the 6th period. It is an f –block element as 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 $[Xe]^{54}4f^6 5d^1 6s^2$. Thus, its atomic number is 54 + 7 + 2 + 1 = 64. Hence, the element is Gadolinium

Question: 31 The first ($\Delta_1 H_1$) and the second ($\Delta_1 H_2$) 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	ΔH_1	ΔH_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 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)?

Answer:

a.) Element V is likely to be the least reactive element. This is because it has the highest ionization enthalpy and a positive electron gain enthalpy.

b.) Element II is likely to be the most reactive metal as it has the lowest first ionization enthalpy and a low negative electron gain enthalpy.

c.) Element III is likely to be the most reactive non – metal as it has the high first ionization enthalpy and highest negative electron gain enthalpy.

d.) Element V is likely to be the least reactive non – metal since it has a very high first ionization enthalpy and a positive electron gain enthalpy.

e.) Element VI has a low negative electron gain enthalpy. Thus, it is a metal. Further, halide of the formula MX2.

f.) Element I has low first ionisation enthalpy and high second ionisation energy. Therefore, it can form a predominantly stable covalent halide of formula MX.

Question: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 iodide

d.) Silicon and oxygen

e.) Phosphorous and fluorine

f.) Element 71 and fluorine

Answer:

a.) Li2O

b.) Mg3N2

c.) AlI3

d.) SiO2

e.) PF_3 or PF_5

f.) The elements with the atomic number 71 is Lutetium (Lu). It has valency 3. Hence, the formula of the compound is LuF_3.

Question: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

Answer:

The value of the principal quantum number (n) for the outermost shell or the valence shell indicates a period in the Modern periodic table.

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

a.) The p – block has 6 column, 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 value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.

Answer:

The d – block has 10 columns because a maximum of 10 electrons can occupy all the orbitals in a d subshell.

Question:35 Anything that influences the valence electron will affect the chemistry of electrons. Which one of the following factor does not affect the valence shell?

a.) **Valence principal quantum number.**

b.) **Nuclear charge (z)**

c.) **Nuclear mass**

d.) **Number of core electrons.**

Answer:

Nuclear mass does not affect the valence shell.

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

a.) **Nuclear charge (z)**

b.) **Valence principal quantum number**

c.) **Electron – electron interaction in the outer orbitals**

d.) **None of the factors because their size in the same.**

Answer:

The size of an isoelectronic species increases with a decrease in the nuclear charge (Z). For example, the order of the increasing nuclear charge of F^-, Ne, and Na^+ is as follows:

	F	$<$	Ne	$<$	Na^+
Z	9		10		11

Therefore, the order of the increasing size of F^-, Ne and Na^+ is as follows:

$Na^+ < Ne < F^-$

Question:37 Which one of the following statements is incorrect in the relation to ionization enthalpy?

a.) **Ionization enthalpy**

b.) **The greatest increase in ionization enthalpy is experienced on the removal of electron from core noble gas configuration.**

c.) **End of valence electrons is marked by a big jump in ionization enthalpy.**

d.)

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

Answer:

Electrons in orbitals bearing a lower n value are more attracted to the nucleus than electrons in orbitals bearing a higher n value. Hence, 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. (option d)

Question: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**

Answer:

The metallic character of the elements decreases from left to right across a period. Thus, the metallic character of Mg is more than that of Al.

The metallic character of elements increase down a group. Thus, the metallic character of Al is more than that of B.

Hence, the correct order of metallic character is K > MG > Al > B.

Question: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**

Answer:

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

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

Question: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 > N**
- **c.) Cl > F > O > N**
- **d.) O > F > N > Cl.**

Answer:

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

Again, the oxidising character of elements decreases down a group. Thus, we get F > Cl. However, the oxidising character of O is more than that of Cl i.e. O > Cl.

Hence, the correct order of chemical reactivity of F > O > Cl > N.