CHAPTER 3

Classification of Elements and Periodicity in Properties

CLASS 11TH

NCERT EXERCISE AND SOLUTIONS - CHEMISTRY

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

ANSWER:-

The main purpose of organizing elements in the periodic table is to classify them into periods and groups based on their properties. This arrangement simplifies and systematizes the study of elements and their compounds.

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

ANSWER:-

Ans Mendeleev's key method for classifying elements was based on atomic mass. He arranged all the known elements according to their atomic masses. Later, he grouped elements with similar properties together, rather than strictly following their atomic mass order.

Q. 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 of elements are periodic functions of their atomic weights. On the other hand, the 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.

ANSWER:-

In the sixth period (n = 6) of the periodic table, the 6s, 4f, 5d, and 6p orbitals are filled in order of increasing energy. There are a total of 16 orbitals, each of which can hold a maximum of 2 electrons. Therefore, the sixth period can accommodate a maximum of 32 elements.

Q. 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 is present in the 7th period 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. Thus, the element with Z = 114 is present in the 7th period and 14th group of the periodic table.

Q. 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 1st period and eight elements in the 2nd period., The third period starts with the element with Z = 11. Now, there are eight elements in the third period. Thus, the 3rd period ends with the element with Z = 18 i.e., the element in the 18thgroup of the third period has Z = 18. Hence, the element in the 17th group of the third period has atomic number Z = 17.

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

(i) Lawrence Berkeley Laboratory

(ii) Seaborg's group?

ANSWER:-

(i) Lawrencium (Z = 103).

(ii) Seaborgium (Z = 106)

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

ANSWER:-

Elements in the same group have similar electronic configurations, meaning they have the same number of electrons in their valence shell. As a result, they exhibit similar physical and chemical properties.

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

ANSWER:-

Atomic radius is the distance from the center of the nucleus to the outermost shell that contains electrons. Ionic radius, on the other hand, refers to the size of an atom's ion in a crystal lattice, defined as half the distance between two ions that are just touching each other.

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

ANSWER:-

In a period, as the atomic number increases from left to right, the atomic radius decreases. This is due to the increase in nuclear charge, which enhances the attraction of the nucleus for the valence electrons. In a group, as you move from top to bottom, the ionic radius increases because a new energy level is added with each successive element, while the number of valence electrons remains the same.

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^+

ANSWER:-

Isoelectronic species contain same number of electrons.

(i) Both F^- and O^{2-} contain 10 electrons each. Hence, they are isolectronic.

(ii) Both Ar and Cl^- contains 18 electrons each. Hence, they are isoelectronic.

(iii) Both Mg^{2+} and Na^+ contain 10 electrons each. Hence, they are isoelectronic.

(iv) Both Rb^+ and Kr contain 36 electrons. Hence, they are isoelectronic

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.

ANSWER:-

(a) They all have 10 electrons i.e they are isoelectronic species.

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

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

ANSWER:-

When a gaseous atom loses one or more electrons to form a cation, the nuclear charge remains unchanged, but the number of electrons decreases. As a result, the same nuclear charge acts on fewer electrons, increasing the effective nuclear charge per electron. This causes the electrons to be more strongly attracted and pulled towards the nucleus, leading to a decrease in size.

When a gaseous atom gains one or more electrons to form an anion, the nuclear charge remains the same, but the number of electrons increases. In this case, the same nuclear charge acts on more electrons, decreasing the effective nuclear charge per electron. As a result, the electrons are less strongly attracted and are pushed away from the nucleus, causing the size to increase.

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

Hint : Requirements for comparison purposes

ANSWER:-

Isolated gaseous atom means that the atom should be free from other atom in the gaseous state.

No energy should be required to separate it further from other atoms.

Ground state means the lowest energy state possible for that atom. These terms are used for comparison purposes.

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}.

Hint: Apply the idea of mole concept to derive the answer

ANSWER:-

It is given that the energy of an electron in the ground state of the 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.

Ionization enthalpy of atomic hydrogen = 2.18×10^{-18} J

Hence, 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} = 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 Δi H than B

(ii) O has lower Δi H than N and F?

ANSWER:-

(i) The electronic configuration of Be is $1s^2\ 2s^2$, and the electronic configuration of B is $1s^2\ 2s^2\ 2p^1$. Be has a higher ionization enthalpy (ΔiH) than B for the following reasons:

(a) The electronic configuration of Be is more stable than that of B because the 2s orbital is completely filled in Be.

(b) When ionizing Be, a s electron is removed, while in B, p electron is removed. The 2s electron is more strongly attracted to the nucleus than the 2p electron due to its greater penetration. As a result, the 2p electron is more shielded, and the nucleus has a stronger attraction for the 2s electron. Therefore, it requires more energy to remove the 2s electron than the 2p electron, making Be's ionization enthalpy higher than that of B.

(ii) The electronic configuration of oxygen is $1s^2\ 2s^2\ 2p^4$. In the 2p orbital, there are 4 electrons, with 2 of them occupying the same 2p orbital. This increases electron repulsion. Nitrogen (N) has a stable half-filled 2p configuration, while fluorine (F) has a greater nuclear charge. As a result, the ionization enthalpy of oxygen is lower than that of nitrogen and fluorine.

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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?

ANSWER:-

The first ionization enthalpy of sodium (Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$) is lower than that of magnesium (Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$) because, although the electron to be removed in both cases is from the 3s-orbital, the nuclear charge of magnesium is greater than that of sodium. After the removal of the first electron, Na^+ acquires the inert gas configuration of neon (Na^+: $1s^2\ 2s^2\ 2p^6$). The $3s^1$ electron in sodium is easier to remove compared to removing an electron from the stable inert gas configuration of Na^+. Therefore, the second ionization enthalpy (IE_2) 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?

ANSWER:-

Within the main group elements, the ionization enthalpy decreases progressively as we move down the group. This trend is explained by the following factors:

(i) Atomic size: As we move down the group, the atomic size increases due to the addition of a new main energy shell (n).

(ii) Shielding effect: The shielding effect on the outermost electron increases because the number of inner electrons rises.

(iii) Nuclear charge: The nuclear charge increases as we move from top to bottom in a group.

However, the combined effects of the increase in atomic size and the shielding effect are more significant than the increase in nuclear charge. As a result, the outermost electron is less tightly bound to the nucleus as we move down the group. This leads to a gradual decrease in ionization enthalpies within a 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?

ANSWER:-

As we move down the group, the ionization enthalpy generally decreases, which is true for elements like B and Al. However, the ionization enthalpy of Ga is unexpectedly higher than that of Al. Ga has 10d electrons in its inner shell, which are less penetrating and provide weaker shielding compared to s and p electrons. As a result, the outer electron in Ga is held more strongly by the nucleus, leading to a slight increase in ionization enthalpy.

A similar increase in ionization enthalpy is observed from In to Tl, as Tl contains 14f electrons in its inner shell. These 14f electrons offer poor shielding, which results in a stronger attraction between the outer electrons and the nucleus.

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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

ANSWER:-

Electron gain enthalpy is the energy change that occurs when an electron is added to a neutral atom in the gaseous state. A negative electron gain enthalpy indicates that energy is released when an electron is added, which is generally favorable.

(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 D. 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, but in Cl, the electron is added to quantum level. 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.

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.

ANSWER:-

$$O_{(g)} + e^- \rightarrow O^-_{(g)} \left(\Delta_{eg}H = -141 \text{ kJ mol}^{-1}\right)$$

$$O^-_{(g)} + e^- \rightarrow O^{2-}_{(g)} \left(\Delta_{eg}H = +780 \text{ kJ mol}^{-1}\right)$$

The first electron gain enthalpy of oxygen is negative because energy is released when a gaseous atom accepts an electron to form monovalent anion. The second electron gain enthalpy is positive because energy is needed to overcome the force of repulsion between monovalent anion and second incoming electron.

Q. 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.

Q. 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 N is different in is different in NH_3 and NO_2.

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Q. 24. Describe the theory associated with the radius of an atom as it

(a) gains an electron

(b) loses an electron.

ANSWER:-

(a) When an atom gains an electron, its size increases. Adding an electron increases the total number of electrons, which leads to greater repulsion among them. However, the number of protons remains unchanged, causing a decrease in the atom's effective nuclear charge. This reduction in nuclear attraction results in an expansion of the atom's radius.

(b) When an atom loses an electron, the number of electrons decreases while the nuclear charge stays constant. As a result, the repulsive forces between the remaining electrons decrease. With less electron-electron repulsion, the effective nuclear charge becomes stronger, leading to a reduction in the atom's radius.

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.

ANSWER:-

Ionization enthalpy is influenced by the electronic configuration (number of electrons) and the nuclear charge (number of protons). Since isotopes of an element share the same electronic configuration and nuclear charge, they are expected to have identical ionization enthalpies.

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

ANSWER:-

Metals have strong tendency to lose electrons to form cations. Metals are strong reducing agents with low ionization enthalpies, low negative electron enthalpies and low electronegativities. They form basic oxides and ionic compounds.

Non-metals have strong tendency to accept electrons to form anions. Non-metals are strong oxidizing agents with high ionisation enthalpies, high negative electron gain enthalpies and high electronegativities. They form acidic oxides and covalent compounds.

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.

ANSWER:-

(a) Chromium (Z = 24) contains five electrons in 3d subshell.

(b) Magnesium (Z = 12), an alkaline earth metal, can readily lose two electrons to form Mg^{2+}.

(c) Oxygen (Z = 8) gains 2 electrons to form O^{2-}.

(d) Halogens (group 17) contains metal (iodine), non metals (F,Cl and Br), liquid bromine and gases.

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

ANSWER:-

Group 1 elements have one electron in their valence shell and a strong tendency to lose this electron. The ease of losing the valence electron is influenced by ionization enthalpy, which decreases as you move down the group. As a result, the reactivity of Group 1 elements increases in the following order:

Li < Na < K < Rb < Cs.

Group 17 elements have seven electrons in their valence shell and a strong tendency to gain one more electron. This tendency depends on their electrode potentials, which decrease from fluorine to iodine. Consequently, the reactivity of Group 17 elements decreases in the order:

F > Cl > Br > I.

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

ANSWER:-

The general outer electronic configuration of s block elements is $ns^{(1-2)}$.

The general outer electronic configuration of p block elements is $ns^2np^{(1-6)}$.

The general outer electronic configuration of d block elements is $(n-1)d^{(1-10)}ns^{(0-2)}$.

The general outer electronic configuration of f– block elements is $(n-2)f^{(0-14)}(n-1)d^{(0-1)}ns^2$.

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

ANSWER:-

(i) For n = 3, the period in which the element belongs is third. The electronic configuration is $3s^23p^4$ and the element belongs to p block. The group number of the element is 10+ number of electrons in the valence shell.

= 10 + 6 = 16

Thus, the element belongs to third period and sixteenth group.

(ii) For n = 4, the period in which the element belongs is fourth. The electronic configuration is $3d^24s^2$ and the element belongs to d block. The group number of the element is = number of electrons in the (n − 1)d subshell + number of electrons in ns subshell = 2 + 2 = 4

Thus, the element belongs to fourth period and fourth group.

(iii) For n = 6, the period in which the element belongs is sixth. The electronic configuration is $4f^75d^16s^2$ and the element belongs to f block. All f block elements belong to third group.

Thus, the element belongs to sixth period and third group.

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Q. 31. **The first ($\Delta_i H1$) 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	ΔH_1	ΔH_2	Δeg 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) and

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

ANSWER:-

(a) The least reactive element will be V because it has the highest first ionization enthalpy and positive electron gain enthalpy so the addition and removal of electrons from the outermost orbital will be tough.

(b) The most reactive metal will be element II because it has the lowest first ionization enthalpy and the low negative electron gain enthalpy so the removal of electrons from the outermost orbital will be easy and hence it will be reactive.

(c) The most reactive non-metal will be element III because it has high first ionization enthalpy and highest negative electron gain enthalpy. So, the addition of electrons in the outermost orbital will be easy.

(d) the least reactive non-metal will be element V because it has the highest first ionization enthalpy and positive electron gain enthalpy.

(e) The low negative electron gain enthalpy element is VI, so it is a metal. And the second ionization enthalpy of this element is also low. Therefore it can form a stable binary halide with the formula MX_2(X = halogen).

(f) To form a predominantly stable covalent halide of the formula MX (X = Halogen) element should have low first ionization enthalpy and high second ionization enthalpy. Thus the element I will be the answer.

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**

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

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

ANSWER:-

(a) Alkali metal lithium (with 1 valence electron) combine with group 16 element oxygen (having valence of 2) to form Li_2O.

(b) Alkaline earth metal magnesium (with 2 valence electrons) combines with group 5 element nitrogen (having valence 3) to form Mg_3N_2.

(c) Group 13 element aluminum (with 3 valence electrons) combines with group 17 element iodine (a halogen with valence 1) to form AlI_3.

(d) Group 14 element silicon (with a valence of 4) combines with group 16 element oxygen (having valence 2) to form SiO_2.

(e) Group 15 element phosphorous (with valence 3 or 5) combines with group 17 element fluorine (a halogen with valence 1) to form PF_3 or PF_5.

(f) Element 71 with 3 electrons in the valence shell ($4f^1 45d^1 6s^2$) combines with group 17 element fluorine (a halogen with valence 1) to form EF_3 where E is the symbol of the element.

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.**

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.

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 value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.**

ANSWER:-

The option (b) represents incorrect statement. It states that "The d-block has 8 columns because a maximum of 8 electrons can occupy all the orbitals in a d-subshell."

The correct statement is "(b) The d-block has 10 columns because a maximum of 10 electrons can occupy all the orbitals in a d-subshell"

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

(a) **Valence principal quantum number (n)**
(b) **Nuclear charge (Z)**
(c) **Nuclear mass**
(d) **Number of core electron**

ANSWER:-

(c) Nuclear mass does not affect the valence shell because nucleus consists of protons and neutrons. Protons i.e., nuclear charge affect the valence shell but neutrons do not.

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.**

ANSWER:-

The size of isoelectronic species is primarily affected by the nuclear charge (Z). Therefore, the correct answer is (a) nuclear charge (Z).

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 removal of electron from core noble gas configuration.**
(c) **End of valence electrons is marked by a big jump in ionization enthalpy.**
(d) **Removal of electron from orbitals bearing lower n value is easier than from orbital having higher n value.**

ANSWER:-

Statement 1: Ionization enthalpy increases for each successive electron.

Analysis: This statement is correct. As more electrons are removed, the remaining electrons experience less repulsion and more effective nuclear charge, making it harder to remove additional electrons.

Statement 2: The greatest increase in ionization enthalpy is experienced on removal of an electron from a core noble gas configuration.

Analysis: This statement is also correct. Noble gases have a stable electron configuration, and removing an electron from such a stable configuration requires significantly more energy

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Statement 3: End of valence electrons is marked by a big jump in ionization enthalpy.

Analysis: This statement is correct. When all valence electrons are removed, the next electron removed comes from a more stable, filled inner shell, resulting in a large increase in ionization energy.

Statement 4: Removal of an electron from orbitals bearing lower n value is easier than from orbitals having higher n value.

Analysis: This statement is incorrect. Electrons in orbitals with lower principal quantum numbers (n) are closer to the nucleus and experience a stronger nuclear attraction. Therefore, it is harder to remove them compared to electrons in orbitals with higher n values.

Option 4 is the correct answer.

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**

ANSWER:-

The correct order of metallic character for the elements B, Al, Mg, and K is:

K > Mg > Al > B (Option d)

(a) Potassium (K): Being in Group 1, it has the highest metallic character as it is the furthest left and down the group.

(b) Magnesium (Mg): Group 2, so it has a lower metallic character than K but higher than Al and B.

(c) Aluminium (Al): Group 13, it has a lower metallic character than Mg due to its position in the periodic table.

(d) Boron (B): Also in Group 13 but in Period 2, it has the lowest metallic character among these elements.

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**

ANSWER:-

In a period as we move from left to right, non metallic character increases because of decrease in the size of the atom which leads to increase in tendency to accept electron from left to right. Even in a group, non metallic character decreases due to increase in size which leads to decrease in tendency to accept electron from top to bottom.

Hence, correct order will be F > N > C > B > Si (option d).

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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**

ANSWER:-

The oxidizing character of elements increases from left to right across a period because of the presence of vacant d orbitals in their valence shells. Thus, we get the decreasing order of oxidizing property as F > O > N.

Again, the oxidizing character of elements decreases down a group. Thus, we get F > Cl.

Since Oxygen is more electronegative than Cl, O is a stronger oxidizing agent than Cl. Also, F is the best electronegative element. Hence, it is the best oxidizing agent.

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

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

Therefore, the correct option is (b) F > O > Cl > N

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