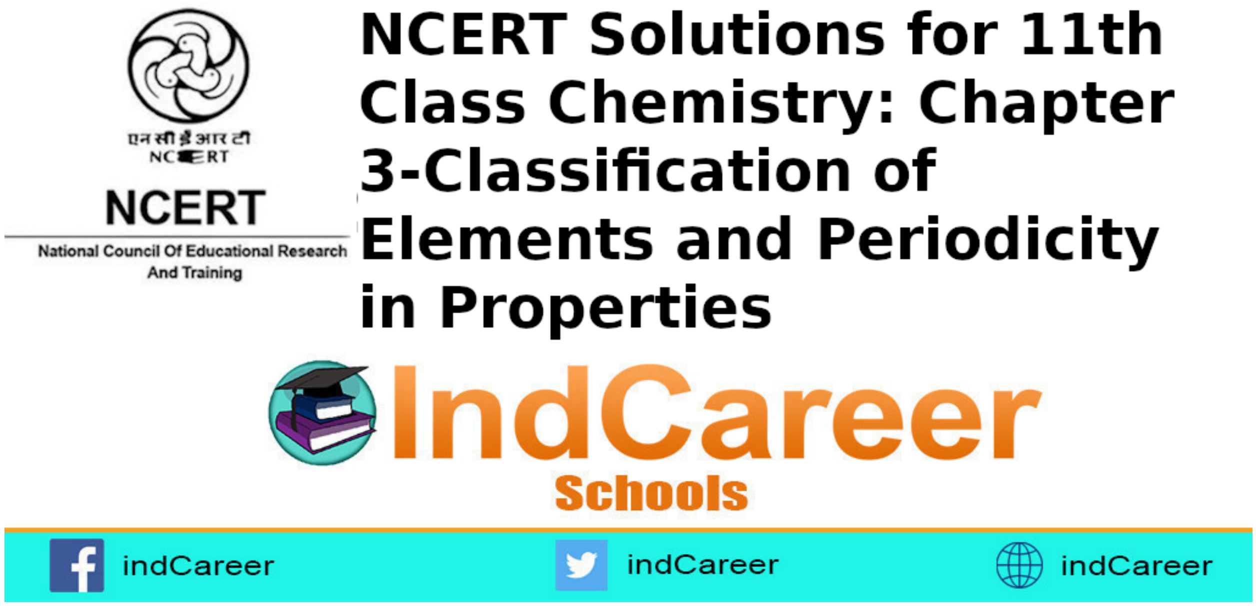

NCERT Solutions for 11th Class Chemistry: Chapter 3-Classification of Elements and Periodicity in Properties

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NCERT Solutions for 11th Class Chemistry: Chapter 3-Classification of Elements and Periodicity in Properties

NCERT 11th Chemistry Chapter 3, class 11 Chemistry Chapter 3 solutions

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

Answer: The basic theme of organisation of elements in the periodic table is to simplify and systematize the study of the properties of all the elements and millions of their compounds. This has made the study simple because the properties of elements are now studied in form of groups rather than individually.

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

Answer: Mendeleev used atomic weight as the basis of classification of elements in the periodic table. He did stick to it and classify elements into groups and periods.

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

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

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

Answer: The sixth period corresponds to sixth shell. The orbitals present in this shell are 6s, 4f, 5p, and 6d. The maximum number of electrons which can be present in these sub¬shell is 2 + 14 + 6 + 10 = 32. Since the number of elements in a period corresponds to the number of electrons in the shells, therefore, sixth period should have a maximum of 32 elements.

Question 5. In terms of period and group where will you locate the element with z = 114?

Answer: Period – 7 and Group -14 Block-p.

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

Answer: The element is chlorine (Cl) with 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 atomic number (z) = 103

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

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

Answer: The elements in a group have same valence shell electronic configuration and hence have similar physical and chemical properties.

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

Answer: Atomic radius. The distance from the centre of nucleus to the outermost shell o electrons in the atom of any element is called its atomic radius. It refers to both covalen or metallic radius depending on whether the element is a non-metal or a metal.

Ionic radius. The Ionic radii can be estimated by measuring the distances between cations and anions in ionic crystals.

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

Answer: Within a group Atomic radius increases down the group.

Reason. This is due to continuous increases in the number of electronic shells or orbit numbers in the structure of atoms of the elements down a group.

Variation across period.

Atomic Radii. From left to right across a period atomic radii generally decreases due

to increase in effective nuclear charge from left to right across a period.

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Question 11. What do you understand by isoelectronic species? Name a species that tvill be iso electronic with each of the following atoms or ions.

(i) F^-(ii) Ar (iii) Mg^{2+}(iv) Rb^+

Answer: Isoelectronic species are those species (atoms/ions) which have same number of

electrons. The isoelectronic species are:

(i)Na^+ (iii) Na^+

(ii)K^+ (iv) Sr^{2+}

Question 12. Consider the following species:

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

(a) What is common in them?

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

Answer: (a) All of them are isoelectronic in nature and have 10 electrons each.

(b) In isoelectronic species, greater the nuclear charge, lesser will be the atomic or ionic radius.

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

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

Answer: A cation is smaller than the parent atom because it has fewer electrons while its nuclear

charge remains the same. The size of anion will be larger than that of parent atom

because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.

Question 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:

- Significance of term 'isolated gaseous atom'. The atoms in the gaseous state are far separated in the sense that they do not have any mutual attractive and repulsive interactions. These are therefore regarded as isolated atoms. In this state the value of ionization enthalpy and electron gain enthalpy are not influenced by the presence of the other atoms. It is not possible to express these when the atoms are in the ; liquid or solid state due to the presence of inter atomic forces.
- Significance of ground state. Ground state of the atom represents the normal – energy state of an atom. It means electrons in a particular atom are in the lowest energy state

and they neither lose nor gain electron. Both ionisation enthalpy and I electron gain enthalpy are generally expressed with respect to the ground state ofan atom only.

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Question 15. Energy of an electron in the ground state of the hydrogen atom is- 2.18 x 10^{-18} J.Calculate the ionization enthalpy of atomic hydrogen in terms of $JMol^{-1}$.[Hint: Apply the idea of mole concept to derive the answer],

Answer: The ionisation enthalpy is for 1 mole atoms. Therefore, ground state energy of the , atoms may be expressed as E (ground state) = (– 2.18 x 10^{-18} J) x(6.022 x 10^{23} mol^{-1})= -1.312 x 10^{6} J mol^{-1}

Ionisation enthalpy =E_{∞}_E ground state

= 0-(-1.312 x $10^{6}mol^{-1}$)

= 1.312 x 10^{6} J mol^{-1}.

Question 16. Among the second period elements, the actual ionization enthalpies are in the order: Li <B< Be <C<0<KI<F< Ne

Explain why

(i) Be has higher $\Delta_i H_1$than B ?

(ii) O has lower $\Delta_i H_1$ than N and F?

Answer: (i) In case of Be ($1s^2$ $2s^2$) the outermost electron is present in 2s-orbital while in B ($1s^2$ $2s^2$ $2p^1$) it is present in 2p-orbital. Since 2s – electrons are more strongly attracted by the nucleus than 2p-electrons, therefore, lesser amount of energy is required to knock out a 2p-electron than a 2s – electron. Consequently, At of Be is higher than that $\Delta_i H_1$ of B.

(ii) The electronic configuration of

N_7 = $1s^2$ $2s^2$ $2p_x^1$ $2p_y^1$ $2p_z^1$

O_8 =$1s^2$ $2s^2$ $2p_x^1$ $2p_y^1$ $2p_z^1$

We can see that in case of nitrogen 2p-orbitals are exactly half filled. Therefore, it is difficult to remove an electron from N than from O. As a result $\Delta_i H_1$ of N is higher than that of O.

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 and Mg are

Na = $1s^2\ 2s^2\ 2p^6\ 3s^1$

Mg = $1s^2\ 2s^2\ 2p^6\ 3s^2$

First electron in both cases has to be removed from 3s-orbital but the nuclear charge of Na (+ 11) is lower than that of Mg (+ 12) therefore first ionization energy of sodium is lower than that of magnesium.

After the loss of first electron, the electronic configuration of

Na^+ = $1s^2\ 2s^2\ 2p^6$

Mg^+ = $1s^2\ 2s^2\ 2p^6\ 3s^1$

Here electron is to be removed from inert (neon) gas configuration which is very stable and hence removal of second electron requires more energy in comparison to Mg.

Therefore, second ionization enthalpy of sodium is higher than that of magnesium.

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

Answer: Atomic size. With the increase in atomic size, the number of electron shells increase. Therefore, the force that binds the electrons with the nucleus decreases. The ionization enthalpy thus decreases with the increase in atomic size.

Screening or shielding effect of inner shell electron. With the addition of new shells, the number of inner electron shells which shield the valence electrons increases. As a result, the force of attraction of the nucleus for the valence electrons further decreases and hence the ionization enthalpy decreases.

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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 decrease in $\Delta_i H_1$ value from B to Al is due to the bigger size of Al.

In Ga there is 10 3d electrons which do not screen as is done by S and P electrons. Therefore, there is an unexpected increase in the magnitude of effective nuclear charge resulting in increased $\Delta_i H_1$ values. The same is with into Tl. The later has fourteen Δf electrons with very poor shielding effect. This also increases, the effective nuclear charge thus the value of $\Delta_i H_1$ increases.

Question 20. Which of the following pairs of elements would have a move negative electron gain enthalpy? (i) O or F (ii) F or Cl.

Answer: (i) O or F. Both O and F lie in 2nd period. As we move from O to F the atomic size decreases.

Due to smaller size of F nuclear charge increases.

Further, gain of one electron by

F —> F^-

F~ ion has inert gas configuration, While the gain of one electron by

0->O^-

gives CT ion which does not have stable inert gas configuration, consequently, the energy released is much higher in going from

F ->F^-

than going from O —>O^-

In other words electron gain enthalpy of F is much more negative than that of oxygen.

(ii) The negative electron gain enthalpy of Cl (Δ eg H = – 349 kj mol-1) is more than that of F (Δ eg H = – 328 kJ mol $^{-1}$).

The reason for the deviation is due to the smaller size of F. Due to its small size, the electron repulsions in the relatively compact 2p-subshell are comparatively large and hence the attraction for incoming electron is less as in the case of Cl.

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: For oxygen atom:

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

$O^-(g) + e^- \longrightarrow O^{2-}(g)$ ($\Delta_{eg} H = +780$ kJ mol $^{-1}$)

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.

Question 22. What is basic difference between the terms electron gain enthalpy and electro negativity?

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

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

Ans. On Pauling scale, the electronegativity of nitrogen, (3.0) indicates that it is sufficiently electronegative. But it is not correct to say that the electronegativity of nitrogen in all the compounds is 3. It depends upon its state of hybridisation in a particular compound, greater the percentage of s-character, more will be the electronegativity of the element. Thus, the electronegativity of nitrogen increases in moving from SP^3 hybridised orbitals to SP hybridised orbitals i.e., as $SP^3 < SP^2 < SP$.

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

(a) gains an electron (b) loses an electron ?

Answer:

- Gain of an electron leads to the formation of an anion. The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among electrons and decrease in effective nuclear charge.

This the ionic radius of fluoride ion (F^-) is 136 pm whereas atomic radius of Fluorine (F) is only 64 pm.

- Loss of an electron from an atom results in the formation of a cation. A cation is smaller than its parent atom because it has fomer electrons while its nuclear charge remains the same. For example, The atomic radius of sodium (Na) is 186 pm and atomic radius of sodium ion (Na^+) = 95 pm.

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

Answer: Ionization enthalpy, among other things, depends upon the electronic configuration (number of electrons) and nuclear charge (number of protons). Since isotopes of an element have the same electronic configuration and same nuclear charge, they have same ionization enthalpy.

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

Answer:

Metals	*Non-Metals*
1. Have a strong tendency to lose electrons to form cations.	1. Non-metals have a strong tendency to accept electrons to form anions.
2. Metals are strong reducing agents.	2. Non-metals are strong oxidising agent.
3. Metals have low ionization enthalpies.	3. Non-metals have high ionization enthalpies.
4. Metals form basic oxides and ionic compounds.	4. Non-metals form acidic oxides and covalent compounds.

Question 27. Use periodic table to answer the following questions:

(a) Identify the element with five electrons in the outer subshell.

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

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

Answer: (a) Element belonging to nitrogen family (group 15) e.g., nitrogen.

(b) Element belonging to alkaline earth family (group 2) e.g., magnesium.

(c) Element belonging to oxygen family (group 16) e.g., oxygen.

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

Answer: The elements of Group I have only one electron in their respective valence shells and thus have a strong tendency to lose this electron. The tendency to lose electrons in turn, depends upon the ionization enthalpy. Since the ionization enthalpy decreases down the group therefore, the reactivity of group 1 elements increases in the same order Li < Na < K < Rb < Cs. In contrast, the elements of group 17 have seven electrons in their respective valence shells and thus have strong tendency to accept one more electron to make stable configuration. It is linked with electron gain enthalpy and electronegativity. Since both of them decreases down the group, the reactivity therefore decreases.

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

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

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

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

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

Question 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?

Answer: (i) n = 3

Thus element belong to 3rd period, p-block element.

Since the valence shell contains = 6 electrons, group No = 10 + 6 = 16 configuration =$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^4$ element name is sulphur.

(ii) n = 4

Means element belongs to 4th period belongs to group 4 as in the valence shell (2 + 2) = 4 electrons.

Electronic configuration.=$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$, and the element name is Titanium (T_i).

(iii) n = 6

” Means the element belongs to 6th period. Last electron goes to the f-orbital, element is from f-block.

group = 3

The element is gadolinium (z = 64)

Complete electronic configuration =$[X_e]\ 4f^7\ 5d^1\ 6s^2$.

Question 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:

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

Which of the above 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) The element V has highest first ionization enthalpy ($\Delta_i H_1$) and positive electron gain enthalpy ($\Delta_{eg}H$) and hence it is the least reactive element. Since inert gases have positive$\Delta_{eg}H$, therefore, the element-V must be an inert gas. The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg}H$match that of He.

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

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

(d) The element IV has a high negative electron gain enthalpy ($\Delta_{eg}H$) but not so high first ionization enthalpy ($\Delta_{eg}H$). Therefore, it is the least reactive non-metal. The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg}H$ match that of I (Iodine).

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

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

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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 iodine(d) Silicon and oxygen

(e) Phosphorous pentafluoride (f) Element 71 and fluorine.

Answer: (a) $Li0_2$(Lithium oxide) (b) Mg_3N_2(Magnesium nitride)

(c) AlI_3(Aluminium iodide) (d) $Si0_2$ (Silicon dioxide)

(e) Phosphorous and fluorine (f) Z = 71

The element is Lutenium (Lu). Electronic configuration[X_e] 4 f^7 $5d^1$ $6s^2$.with fluorine it will form a binary compound = Lu F3.

Question 33. In the modem periodic table, the period indicates the value of

(a)atomic number (b) mass number (c) principal quantum number (d) azimuthal quantum number?

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

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

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

(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: Statement (b) is incorrect.

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

(a) Valence principal quantum number (n)

(b) Nuclear charge (Z)

(c) Nuclear mass

(d) Number of core electrons.

Answer: (c) Nuclear mass.

Question 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: (a) Nuclear charge (Z).

Question 37. Which 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 electrons 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: (d) is incorrect.

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

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: In a period, the non-metallic character increases from left to right. Thus, among B, C, N and F, non-metallic character decreases in the order: F > N > C > B. However, within a group, non-metallic character decreases from top to bottom. Thus, C is more non-metallic than Si.

Therefore, the correct sequence of decreasing non-metallic character is: F > N > C > B > Si, i.e., option (c) is correct.

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: Within a period, the oxidising character increases from left to right. Therefore, among F, O and N, oxidising power decreases in the order: F > O > N. However, within a group, oxidising power decreases from top to bottom. Thus, F is a stronger oxidising agent than Cl. Further because O is more electronegative than Cl, therefore, O is a stronger oxidising agent than Cl. Thus, overall decreasing order of oxidising power is: F > O > Cl > N, i.e., option (b) is correct.

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Chapterwise NCERT Solutions for Class 11 Chemistry:

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- Chapter 2-Structure of Atom
- Chapter 3-Classification of Elements and Periodicity in Properties
- Chapter 4-Chemical Bonding and Molecular Structure
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- Chapter 6-Thermodynamics
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