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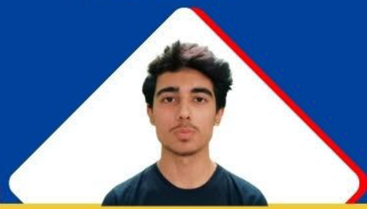

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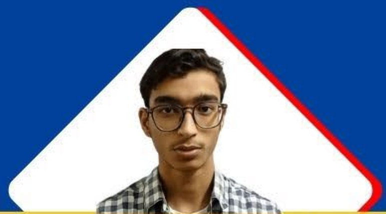

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Important Questions

Classification of Elements and Periodicity in Properties

Classification of Elements and Periodicity in Properties

Q1. How many groups and periods were there in the original Mendeleev's Periodic Table?

Q2. State the Modern Periodic Law.

Q3. What is the main features of long form of periodic table?

Q4. How many groups and periods are there in the long form of the periodic table?

Q5. Write the electronic configuration of scandium having atomic number 21.

Q6. Out of nitrogen and oxygen, which has higher value of first ionization enthalpy?

Q7. Explain the term electron gain enthalpy.

Q8. Why is the electron gain enthalpy of nitrogen is zero?

Q9. How do electronegativity of elements behave in a period?

Q10. In K and K^+, which one would have larger size?

Q11. In Br and Br^-, which one would have larger size?

Q12. In O^{2-} and F^-, which one would have larger size?

Q13. In $Li^+$and Na^+, which one would have larger size?

Q14. What is meant by isoelectronic species?

Q15. Arrange the following elements in the increasing order of metallic character:

Q16. Arrange the following elements in the increasing order of non-metallic character:

Q17. Predict the position of the elements in the periodic table satisfying the electron configuration $(n-1)d^1ns^2$ for $n = 4$.

Q18. Elements A, B, C, D and E have the following electronic configurations:

A: $1s^22s^22p^1$ B: $1s^22s^22p^63s^23p^1$

C: $1s^22s^22p^63s^23p^3$ D: $1s^22s^22p^63s^23p^5$

E: $1s^22s^22p^63s^23p^64s^2$

Which among these will belong to the same group in the periodic table?

Q19. What would be the IUPAC name and symbol for the element with atomic number 120?

Q20. Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

Q21. The first ionization enthalpy (Δ_iH) values of the third period elements, Na, Mg, and Si are respectively 496,737 and 786 kJ mol^{-1}. Predict whether the first Δ_iH value for Al will be more close to 575 or 760 kJ mol^{-1} Justify your answer.

Q22. Explain why first ionization enthalpy of nitrogen is higher than that of elements on left and right hand side in the same period (i.e., carbon and oxygen).

Q23. Among the elements Li, K, Ca, S and Kr, which one has the lowest first ionization enthalpy and which one has the highest first ionization enthalpy?

Q24. Which of the following pairs of elements would you expect to have lower first ionization enthalpy ? (i) Cl or F, (ii) Cl or S, (iii) K or Ar and (iv) Kr or Xe.

Q25. Second and third ionization enthalpies of an element are always greater than its first ionization enthalpy. Explain.

Q26. Electron gain enthalpy values of inert gases are zero. Why?

Q27. Electron gain enthalpy of halogens are high? Explain.

Q28. Define the term electronegativity. Explain.

Q29. Electronegativity values of inert gases are zero. Explain.

Q30. Give the formula of a species that will be isoelectronic with the following atoms or ions: (i) Ne (ii) Cl^-(iii) Ca^{2+} (iv) Rb^+.

Q31. Account for the fact that 4th period has eighteen and not eight elements.

Q32. How would you justify the presence of 18 elements in the 5 th period of the periodic table?

Q33. The elements Z = 117 and 120 have not yet been discovered.

In which family/group would you place these elements and also give the electronic configuration in each case?

Q34. Considering the atomic number and position in the periodic table, arrange the following elements in the increasing order of metallic character: Si, Be, Mg, Na, P.

Q35. Lanthanides and actinides are placed in separate rows at the bottom of the periodic table. Explain the reason for this arrangement.

Q36. The elements Z = 107 and Z = 109 have been made recently; element Z = 108 has not yet been made. Indicate the groups in which you will place the above elements.

Q37. Explain why the electron gain enthalpy of fluorine is less negative than that of chlorine.

Q38. The formula of the oxide is M_2O.

Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionization enthalpy than nitrogen. Explain

Q39. How would you explain the fact that first ionisation enthalpy of sodium magnesium but its second ionisation enthalpy is higher than that of magnesium?

Q40. Explain the following:

(a) Electronegativity of elements increases on moving from left to right in the periodic table.

(b) Ionization enthalpy decreases from top to bottom in a group.

Q41. Among the second period elements, the actual ionization enthalpies are in the order:

$$\mathrm{Li} < \mathrm{B} < \mathrm{Be} < \mathrm{C}$$

Explain why Be has higher $\Delta_i H$ than B.

Q42. 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?

Q43. Would you expect the first ionization enthalpies of the isotopes of the same elements to be same or different? Justify your answer.

Q44. Among the elements of the second period Li to Ne, pick out the element

(i) with the highest first ionization enthalpy,

(ii) with the highest electronegativity,

(iii) with the largest atomic radius,

(iv) that is the most reactive non-metal,

(v) that is the most reactive metal.

Q45. Predict the density of Cs from the density of the following elements :

K0.86 $\mathrm{g\ cm^{-3}}$, Ca 1.548 $\mathrm{g\ cm^{-3}}$, Sc 2.991 $\mathrm{g\ cm^{-3}}$, Rb 1.532 $\mathrm{g\ cm^{-3}}$, Sr 2.68 $\mathrm{g\ cm^{-3}}$, Cs, Y4.34 $\mathrm{gcm^{-3}}$, Ba3.51 $\mathrm{g\ cm^{-3}}$, La6.16 $\mathrm{g\ cm^{-3}}$.

Q46. Formation of O^-ion is accompanied by release of energy while that of O^{2-} ion by absorption of energy. Explain.

Q47. Using the periodic table, predict the formulae of compounds which might be formed by the following pairs of elements :

(a) silicon and bromine (b) aluminium and sulphur.

Q48. The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionization enthalpies (kJmol $^{-1}$) of the three elements I, II and III are given below :

	I	II	III
$\Delta_i H_1$	403	549	1142
$\Delta_i H_2$	2640	1060	2080

Identify the element which is likely to be (a) a non-metal (b) an alkali metal (c) an alkaline earth metal.

Q49. Predict the formulae of the stable binary compounds that would be formed by the following pairs of elements:

(a) silicon and oxygen (b) aluminium and bromine (c) calcium and iodide

(d) element 114 and fluorine (c) element 120 and oxygen.

Q50. Among the elements B, Al, C and Si

(a) which has the highest first ionization energy?

(b) which has the most negative electron gain enthalpy?

(c) which has the largest atomic radius ?

(d) which has the most metallic character?

Q51. Consider the elements N, P, O and S and arrange them in order of

(a) increasing first ionization enthalpy.

(b) increasing negative electron gain enthalpy.

(c) increasing non-metallic character.

Q52. Write the IUPAC nomenclature of elements from atomic number 104 to 109.

Q53. Show by a chemical reaction with water that Na_2O is a basic oxide and Cl_2O_7 is an acidic oride

Q54. Are the oxidation state and covalency of Al in $[AlCl(H_2O)_5]^{2+}$ same?

Q55. Among the elements of the third period Na to Ar pick out the element

(i) with the highest first ionization enthalpy.

(ii) with the largest atomic radius.

(iii) that is the most reactive non-metal.

(iv) that is the most reactive metal.

Q56. Explain screening or shielding effect.

Q57. Arrange the following elements in the increasing order of metallic character.

$$Si, Be, Mg, Na, P.$$

Q58. Choose the ion with the largest ionic size: $F^-, Mg^{2+}, Na^+, O^{2-}$.

Q59. The following species are isoelectronic with noble gas argon:

$Ar, K^+, S^{2-}, Cl^-, Ca^{2+}$. Arrange them in order of increasing size.

Q60. Arrange the following ions in the order of increasing size :

$$Be^{2+}, Cl^-, S^{2-}, Na^+, Mg^{2+}, Br^-$$

Q61. Select from each group the species which has the smallest radius:

(a) O, O^-, O^{2-}

(b) K^+, Sr^{2+}, Ar (c) Si, P, Cl.

Q62. Which of the following species will have the largest and smallest size?

$$Mg, Mg^{2+}, Al^{-Al^{3+}}$$

Q63. Give the reason for the following :

(a) Electron gain enthalpy of fluorine is less negative than that of chlorine.

(b) Anionic radius is always more than that of neutral atom.

(c) Ionization enthalpy of nitrogen is more than that of oxygen.

Q64. (a) Arrange the following ions in the order of increasing ionic radii : Na^+, Mg^{2+}, F^-, O^{2-}

(b) Explain why Be has higher ionization enthalpy than B.

(c) Predict the formula of the compound which might be formed by silicon and bromine.

Q65. What are the s-block elements? Write the characteristic properties of s-block elements.

Q66. What are the p-block elements? Write the characteristic properties of p-block elements.

Q67. What are the d-block elements? Write the characteristic properties of d-block elements.

Q68. What are the f-block elements? Write the characteristic properties of f-block elements.

Q69. Which of the following pairs would have a larger size ? Explain.

(i) K or K^+ (ii) Br or Br^- (iii) O^{2-} or F^-

(iv) Li^+or Na^+ (v) P or As (vi) Na^+or Mg^{2+}.

Q70. What do you understand by the term atomic radius? How do atomic sizes vary in a group and in a period? Give reasons for the variations.

Q71. Which of the elements Na, Mg, Si and P would have the greatest difference between the first and second ionization enthalpies? Briefly explain your answer.

Q72. Define the term ionization enthalpy. Why is second ionization enthalpy always greater than the first ionization enthalpy? On what factors ionization enthalpy depends?

Q73. Explain the term electron gain enthalpy. How does it vary in a group and in a period?

Q74. Which of the following will have the most negative electron gain enthalpy and which the least negative? P, S, Cl, F. Explain your answer.

Q75. The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionization energies ($kJmol^{-1}$) of a few elements designated by Roman numerals are shown below:

Element	$\Delta_i H_1$	$\Delta_i H_2$
I	2372	5251
II	520	7300
III	900	1760
IV	1680	3380

Which of the above element is likely to be :

(i) a reactive metal (ii) a reactive non-metal

(iii) a noble gas

(iv) a metal that forms a stable binary halide of the formula, AX_2, (X = halogen) ?

Answers

1. Ans. 8 groups and 7 periods.

2. Ans. The physical and chemical properties of elements are in the periodic functions of their atomic numbers.

3. Ans. The main features of the long form of periodic table is that the whole table is divided into four blocks.

(i) s-block (Representative elements)

(ii) p-block (Representative elements)

(iii) d-block (Transition elements)

(iv) f-block (Inner transition elements)

4. Ans. 18 groups and 7 periods.

5. Ans. The atomic configuration of scandium is :

6. Ans. According to general behavior oxygen should have higher value of first ionization energy than nitrogen. But nitrogen has higher value of first ionization than oxygen.

7. Ans. Electron gain enthalpy is defined as the enthalpy released when a neutral gaseous atom takes up an extra electron to form an anion.

8. Ans. Because of the higher symmetrical electronic configuration of nitrogen, its atom has no tendency to take up the electron.

9. Ans. The electronegativity increases when we move from left to right in a period.

10. Ans. K (Potassium).

11. Ans. Br^-(Bromide ion).

12. Ans. O^2 has larger size

13. Ans. Na^+(Sodium ion).

14. Ans. Atoms, molecules or ions having the same number of total electrons are called isoelectronic species. Example: $Ar \rightarrow (2+8)$ and $S^{2-} \rightarrow (2+6)+2$ are isoelectronic having total 10 electrons each.

15. Ans. $B < Mg < Al < K$. B, Al, Mg, K

16. Ans. $B < Si < C < N < F$. B, C, Si, N, F

17. Ans. For $n = 4$, the configuration would be $3d^1 4s^2$. This will be placed in 3rd group of the periodic table.

18. Ans. A : Boron, B : Aluminium, C : Phosphorus, D : Chlorine, E : Calcium.

A and B belong to the same group in the periodic table.

19. Ans. The roots for 1, 2, 0 are un, bi and nil and hence the name and symbol respectively would be unbinilium (Ubn).

20. Ans. The element is chlorine with atomic number 17.

21. Ans. It will be more close to 575 kJ mol^{-1}. The value of Al should be lower than that of Mg because of effective shield of 3p-electrons from the nucleus by 3 -electrons.

22. Ans. Electronic configurations of C is $1s^2 2s^2 2p^2$, N is $1s^2 2s^2 2p^3$ and O is $1s^2 2s^2 2p^4$. Out of the three nitrogen has exactly half-filled 2p-orbitals.

Since it is difficult to remove an electron from exactly half-filled 2p-orbitals of nitrogen compared to others, hence N has higher ionization enthalpy than elements on its sides, which do not have half-filled 2p-orbitals.

23. Ans. Potassium has the lowest first ionization enthalpy (ionization enthalpy decreases in going down the group). Krypton has the highest first ionization enthalpy (ionization enthalpy increases in moving right in a period).

24. Ans. (i) Cl has lower ionization enthalpy than F.

(ii) S has lower ionization enthalpy than Cl.

(iii) K has lower ionization enthalpy than Ar.

(iv) Xe has lower ionization enthalpy than Kr.

25. Ans. Second and third ionization enthalpies of an element are always greater than its first ionization enthalpy, because it becomes more difficult to remove an electron from positively charged ions left.

Hence $E_3 > E_2 > E_1$.

$$M(g) \xrightarrow{+E_1} M^+(g) + e^-$$
$$M^+(g) \xrightarrow{+E_2} M^{2+}(g) + e^-$$
$$M^{2+}(g) \xrightarrow{+E_3} M^{3+}(g) + e^-$$

26. Ans. In inert gases, all the valency orbitals are fully filled and do not accept any more electrons because of extra stability associated with these orbitals. Addition of any electron has to go to a next higher energy levels at which the attractive pull is negligible. In other words, such atoms cannot hold the additional electron and hence possess zero electron affinity.

27. Ans. Electron gain enthalpy of halogens are highest in its period. This is because halogen atom is one electron less of the inert gas configuration and hence, it readily accepts one electron to form the stable halide having inert gas configuration.

28. Ans. It is the tendency of an atom to attract electron(s) towards itself to form a covalent compound. In a group, electronegativity decreases from top to bottom due to increase in the atomic size. In a period, electronegativity increases from left to right due to decrease in size and increase in nuclear charge. Because of a large difference in electronegativity between two atoms in a compound, the bond develops ionic or polar character.

29. Ans. Electronegativity values of inert gases are zero, as they ordinarily do not form bonds atom

30. Ans. (i) $Na^+: 1s^2 2s^2 2p^6$ [same as Ne]

(ii) $Ar: 1s^2 2s^2 2p^6 3s^2 3p^6$ [same as Cl^-]

(iii) $Ar: 1s^2 2s^2 2p^6 3s^2 3p^6$ [same as Ca^{2+}]

(iv) $Kr: 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$ [same as Rb^+].

31. Ans. The 4th period starts at potassium (K) with the filling of 4 s-orbital. However, before the filling of 4p-orbitals, the 3d-orbitals with 10 electrons having energies less than 4p-orbitals, are filled to give 10 elements of 3d transition series. Next to 3d-orbitals, 4p-orbitals are filled and this filling up of 4p-orbitals is complete at krypton (Kr). Hence, the 4 th period has eighteen (two in 4s, ten in 3d and six in 4p-orbitals) and not eight elements.

32. Ans. When $n = 5, l = 0,1,2,3$. The order in which the energy of the available orbitals $4d, 5s$ and $5p$ increases is $5s < 4d < 5p$. The total number of orbitals available are 9.

The maximum number of electrons that can be accommodated is 18; and therefore, 18 elements are there in the 5th period.

33. Ans. From the periodic table we find that element with $Z = 117$, would belong to the halogen family (Group 17) and the electronic configuration would be [Rn] $5f^{14} 6d^{10} 7s^2 7p^5$.

The element with $Z = 120$, will be placed in Group 2 (alkaline earth metals), and will have the electronic configuration [Uuo] $8\ s^2$.

34. Ans. Metallic character increases down a group and decreases along a period as we move from left to right. Hence the order of increasing metallic character is : $P < Si < Be < Mg < Na$.

35. Ans. Lanthanides and actinides have been placed in two separate rows at the bottom of the periodic table. (i) to save space, (ii) to keep the elements of similar properties in a single column, and (iii) because their inner 4f and 5f orbitals are filled respectively.

36. Ans. Element $Z = 107$, will be placed in group $7, Z = 108$ will be placed in Group 8, and $Z = 109$ will be placed in Group 9.

37. Ans. The added electron in fluorine goes to second quantum level. Due to small size of fluorine, it experiences repulsion from other electrons much more in comparison to that in chlorine because in chlorine, the electron is added to third quantum level in which larger space is available for movement.

38. Ans. The outermost electronic configuration of

nitrogen to any of the 2p orbitals requires energy because p-oritals are half-filled. Addition of extrates stable configuration i.e., $2p^3$ after removing one electron. one electron.

39. Ans. After removing 1 electron from sodium atom, the ion formed acquires the configuration of inert gas, neon. The second electron is removed from one of the 2p orbitals which filled i.e., have a total of 6 electrons and are closer to the nucleus.

40. Ans. (a) Electronegativity increases on moving from left to right in a period because of decrease in size of the atom and increase in nuclear charge.

(b) Ionisation enthalpy decreases from top to bottom in a group because of increase in atomic size.

41. Ans. The ionization enthalpy besides other factors depends upon the type of electron to be removed from the same principal shell. In Be (electronic configuration $1s^2, 2s^2$), the outermost electron is present in 2s orbital while in B (electronic configuration $1s^2, 2s^2, 2p^1$), it is present in 2p orbitals. Since 2s electrons are more strongly attracted to the nucleus compared to 2p electrons, lesser energy is required to remove a 2p electron as compared to 2s electron. Therefore $\Delta_i H$ of Be is higher than that of B.

42. Ans. Sodium has electronic configuration $1s^2 2s^2 2p^6 3s^1$ and magnesium has electronic configuration $1s^2 2s^2 2p^6 3s^2$. Here, in both cases the first electron is to be removed from 3s-orbital but the nuclear charge of $Na(+11)$ is lower than that of $Mg(+12)$. Therefore, the first ionization enthalpy of sodium is lower than that of magnesium. After the loss of one electron from sodium, electronic configuration of sodium becomes completely filled i.e., $1s^2 2s^2 2p^6$ and after the loss of one electron configuration becomes $1s^2 2s^2 sp^6 3s^1$ i.c., half-filled. Completely filled configuration is always more stable like in case of noble gases. Hence second ionization enthalpy of sodium is greater than that of magnesium.

43. Ans. Two isotopes of the same element have the same atomic number and the same number of electrons. The force of attraction between the outer electrons and the nucleus is the same. Hence they the same ionization enthalpy.

44. Ans. (i) Neon (Ne), (ii) Fluorine (F), (iii) Lithium (Li), (iv) Fluorine (F) and (v) Lithium (Li).

45. Ans. The density of Cs must be more than $Rb 1.532 \text{ g cm}^{-3}$, but less than $Sr 2.68 \text{ g cm}^{-3}$. It should be nearly $\frac{1}{2}[1.532 + 2.680] = 2.106 \text{ g cm}^{-3}$.

46. Ans. In the formation of $O^-(g)$ ion, $O(g)$ accepts an electron from outside $[O(g) + e^- \rightarrow O^- + e^- + q]$, and in the process energy equal to electron affinity is realised. Further when $O^-(g)$ accept another electron to form $O^{2-}(g)$ ion, $[O^-(g) + e^- \rightarrow O^{2-}(g) - q]$, the second incoming electron is repelled by $O^-(g)$ ion. Thus to overcome the force of repulsion, absorption of energy occurs.

47. Ans. (a) Silicon, is a group 14 elements with a valence of 4 and bromine belongs to the halogen family with a valence of 1 . Hence, the formula of the compound formed would be $SiBr_4$.

(b) Aluminium belongs to Group 13 with a valence of 3 and sulphur belongs to Group 16 elements with a valence of 2 . Hence, the formula of the compound formed would be Al_2S_3.

48. Ans. I : an alkali metal II : an alkaline earth metal III : a non-metal.

49. Ans. (a) SiO_2 (b) $AlBr_3$(c)CaI_2(d) (Uuq) F_4 (Uuq will be placed in 14th group and will have electronic configuration $7s^2 7p^2$) (c) (Ubn)O (Ubn will be placed in group 2 and will have electronic configuration (Uuo) $8s^2$

50. Ans. (a) C, (b) B (c) Si, (d) Al.

51. Ans. (a) $S < O < P < N$. (b) $N < P < O < S$. (c) $P < N < O < S$.

52. Ans. The nomenclature is based on Latin words for the atomic number of elements. Th names are derived directly from atomic number using the numerical roots for 0 and number 1 – 9. The roots are strung together in the order of digits which makes up the atomic number an 'ium' is added at the end. Example : For element having atomic number 104, the roots for 1, 0 and 4 are un, nil, quand and hence, the name Unnilquadium (Unq).

Similarly :

105	Unnilpentium	(Unp)
106	Unnilhexium	(Unh)
107	Unnilseptium	(Uns)
108	Unniloctium	(Uno)
109	Unnilennium	(Une).

53. Ans. Na_2O with water forms a strong base whereas Cl_2O_7 forms strong acid.

$$Na_2O + H_2O \longrightarrow 2NaOH \text{ (strong base)}$$
$$Cl_2O_7 + H_2O \longrightarrow 2HClO_4 \text{ (strong acid).}$$

54. Ans. No. The oxidation state of Al is +3 and the covalency is 6.

55. Ans. (i) Argon. (ii) Sodium. (iii) Bromine. (iv) Sodium.

56. Ans. A valence electron in a multi-electron atom is pulled by the nucleus and repelled by the other electrons in the core (inner shells). Thus, the effective pull on the electron will be the pull due to positive nucleus, less the repulsion due to core electrons. The effective repulsive effect of the core electrons is called the screening or shielding effect. For example, the 2s electron in lithium is shielded by the inner core of 1 s electrons. As a result, valence electron experiences a net positive charge less than +3 . In general shielding is effective when the orbitals in the inner shells are completely filled.

57. Ans. Metallic character increases down the group and decreases along a period as we move from left to right. Hence, the order of increasing metallic character would be :

$$P < Si < Be < Mg < Na.$$

58. Ans. All the ions given above are isoelectronic having 10 electrons, although the nuclear charge in $F^- = +9, Mg^{2+} = +12, Na^+ = +11$ and $O^{2-} = +8$.
Thus, all these ions have identical extra-nuclear structures, but their nuclear charges are different. Largest ion should have least nuclear charge. Since O^{2-} possesses least nuclear charge +8 , so the attraction between its nucleus and electrons is least and O^{2-} ion is largest in size.
59. Ans. The order would be : $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$. All are isoelectronic having 18 electrons, but the nuclear charge on $Ar = 18,\ K^+ = 19,\ S^{2-} = 16, Cl^- = 17$ and $Ca^{2+} = 20$. Greater the nuclear charge, greater will be the attraction between nucleus and electrons hence smaller the size or smaller the nuclear charge, lesser will be the attraction between nucleus and electrons hence greater the size.
60. Ans. $Be^{2+} < Mg^{2+} < Na^+ < Cl^- < S^{2-} < Br^-$.
61. Ans. (a) O (b) K^+ (c) Cl.
62. Ans. Atomic radii decrease across a period. Cations are smaller than their parent ^2E_R Among isoelectronic species, the one with the larger positive nuclear charge will have a smaller radius. Hence the largest species is Mg; the smallest one is Al^{3+}.
63. Ans. (a) Electron gain enthalpy of fluorine is less negative (energy released is given negative sign) than that of chlorine. This is because fluorine is a smaller atom. The extra electron added faces repulsion due to greater electron density in fluorine as compared to chlorine.
(b) Anionic radius is always greater than that of the neutral atom. This is because an anion is formed by the addition of electron/s. The e/p ratio changes on addition of electrons. Electrons outnumber the protons. The hold of the nucleus on the electrons decreases. In other words, the size of the anion becomes greater than that of the atom.
(c) Nitrogen has more stable configuration $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$ which is half-filled compared to the configuration of oxygen which is $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$. Hence, greater energy is required to remove an electron from the stable atom of N. Thus, ionization energy of nitrogen is greater than that of oxygen.
64. Ans. (a) Increasing order of ionic radii is :

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

(b) In beryllium, the electron removed during ionization is an s-electron whereas the electron removed from boron during ionization is p-electron. The penetration of s-election to the nucleus is more than the penetration of p-electron. Hence, greater amount of energy is required to remove the s-electron from beryllium. Therefore Be has higher ionization enthalpy than B.

(c) Silicon is group 14 element with a valence of 4 . Bromine belongs to halogen family with a valence of Hence the formula of the compound formed will be $SiBr_4$.

65. Ans. The elements of groups I and II (earlier group IA and IIA) are known as s-block elements, because in these elements the last electron enters the s-subshell. They have the configuration ns^1 or ns^2 in their outermost shell.

Characteristic properties of s-block elements :

(i) Group I elements have +1 valency and group II elements have +2 valency.

(ii) They are metallic and ductile.

(iii) They are good conductors of heat and electricity.

(iv) They have low ionization potential.

(v) They are good reducing agents.

(vi) They are diamagnetic and form colourless compounds except chromates, permanganate and dichromates.

(vii) They are, in general, prepared by the electrolysis of their fused salt.

(viii) As their last shell contains only 1 or 2 electrons, chemically they are very reactive.

Note: Hydrogen does not have above properties.

66. Ans. The elements of group 13,14, 15, 16, 17 (earlier groups IIIA, IVA, VA, VIA, VIIA) are known as p-block elements, because they receive the last electron in the p-subshell. Their outermost shell has the general configuration ns^2np^{1-5} (i.e., depending upon the group to which the element belongs). The elements of the s and p-blocks are also collectively called representative elements.

Characteristic properties of p-block elements :

(i) They have variable valencies and oxidation states, e.g., P(III) and P(V).

(ii) Some of them are non-metals, e.g., F, Cl, Br, I and some of them are metalloids, e.g., B, Si, Ge, As, Sb and Te.

(iii) They (except the metals) are bad conductors of heat and electricity.

(iv) The non-metals form acidic oxides, e.g., NO_2, P_2O_5, etc.

(v) They form ionic as well as covalent compounds.

67. Ans. The elements of groups 3 to 12 (earlier I B to VII B and (VIII) groups), the elements lying between s and p-blocks, are known as d-block elements, because the last added electron enters the dsubshell. The general electronic configuration for the last two energy levels of these elements may be given as $(n-1)s^2(n-1)p^6(n-1)d^{1-9}ns^{1-2}$. The outermost orbital remains incomplete whereas the d-orbitals of the penultimate (i.e., the outermost but one shell) are progressively filled. The d-block elements are also known as transition elements because their properties are intermediate between those of s and p-block elements.

Characteristic properties of d-block elements :

(i) They have variable valencies and oxidation states, e.g., $\mathrm{Fe(II), Fe(III), Mn(II), Mn(IV), Mn(VI), Mn(VII), Cu(I), Cu(II)}$.

(ii) They are malleable and ductile.

(iii) They are good conductors of heat and electricity.

(iv) They are paramagnetic since they contain unpaired electrons.

(v) They form coloured complexes.

(vi) They have good catalytic properties.

(vii) They are hard and have high density and melting point.

(viii) They have higher ionization potential than s-block elements. Note: Those d-block elements which have unfilled or partially filled d orbitals or formeptions in which the d orbitals are only partially filled are called transition elements or metals. Except are $\mathrm{Zn, Cd, Hg}$ (or their ions) which have completely filled d orbitals.

68. Ans. Two sets of elements (each having 14 elements) which belong to the 6th (Lanthanides and 7th (Actinides) periods of outermost orbitals, i.e., n and $(n-1)$ remain incompletely filled. The general electronic configuration for the last three shells may be given as :

$$(n-2)f^{0-14}(n-1)d^{0-1}ns^2$$

In lanthanides, 4f-orbitals and in actinides 5f-orbitals are progressively filled up. f block $_0$ elements are placed at the bottom of the periodic table.

Characteristic properties of f-block elements:

In these elements the f-subshell is pre-penultimate and any change in its electronic population is well screened from neighbouring atoms by the penultimate and the outermost shells, which have almost the same electronic configuration in all the elements. Consequentely, all of them have almost same chemical properties.

69. Ans. (i) Size of K is larger than K^+. Both K and K^+have a nuclear charge of +19 , however, K^+ion has one electron less than the atom, due to the removal of one complete shell (4s). Because of lesser number of electrons in the extra-nuclear part of K^+there will be more attraction between the nucleus and the electrons as compared to K. Hence the result.

(ii) Size of Br^-ion is greater than that of Br. Both Br and Br^-ion possess a nuclear charge of +35 , but Br^-ion has one electron more than Br atom. Greater number of electrons in the extranuclear part of Br^-ion, provides greater inter electronic repulsion. Hence the result.

(iii) Size of O^{2-} ion is larger than that of F^-ion. The nuclear charge on O^{2-} and $F^-$are 8 and 9 respectively although they are isoelectronic having 8 electrons. The force of attraction between the nuclear and electrons in $F^-$is more than that of O^{2-} ion. Hence O^{2-} is larger than F^-ion.

(iv) Na^+ion is larger in size than Li^+ion. Electronic configuration of Li^+ion $= 1s^2$ whereas that of Na^+ion $= 1s^2 2s^2 2p^6$. Thus, Na^+ion has one more shell than Li^+ion. Hence, Na^+ion has a larger size than Li^+.

(v) As atom has a larger size than P atom. As has +33 nuclear charge, whereas P has +15 nuclear charge.

Addition of new shells 4s and 4p in As increases its size, but its increased nuclear charge decreases its size. However, the effect of increase in size due to the addition of a new shell is much more than the decrease in size because of increased nuclear charge. Hence the result.

(vi) Na^+ion has a larger size than Mg^{2+} ion. The nuclear charge on $Na^+$and Mg^{2+} are +11 and +12 respectively, but they are isoelectronic having 10 electrons. Having more nuclear charge on Mg^{2+} than Na^+, the force of attraction between the nucleus and electrons in Mg^{2+} ion is more than that of Na^+ion. Hence the result.

70. Ans. Atomic radius is one-half of the distance between the nuclei of two identical atoms in a molecule bonded by a single bond. In a period, as we move from left to right, the number of shells remain the same but as more electrons are added the nuclear charge increases. This results in an increase in the nuclear attraction for the electrons, which in turn brings about a decrease in the radius of the elements in a period. Thus, the radius decreases along a period (from left to right), e.g., $r_{Na} > r_{Mg} > r_{Al} > r_{Si} > r_P >$... etc. However, in a group from top to bottom, the number of shells increases, therefore, the radius also increases, e.g., $r_{Li} < r_{Na} < r_K < r_{Rb} < \cdots$ etc.

71. Ans. The first ionization enthalpy is the energy needed to remove an electron from a neutral atom to produce univalent cation:

$$X(g) \rightarrow X^+(g) + e^-$$

On the other hand, the second ionization enthalpy is the energy needed to remove from a +vely charged ion to produce a divalent cation :

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

In a univalent cation X^+fewer electrons are attracted by more +ve charges of the nucleus. Therefore, more energy is required to remove an electron from the univalent cation than that from a neutral atom. Hence, second ionization enthalpy of Magnesium is higher than its first ionization enthalpy. Hence, Mg would have the greatest difference between the first and second ionization enthalpies.

72. Ans. Ionization enthalpy of an element is defined as the energy required to remove an electron from the last shell of an isolated atom in the gaseous state.

$$X(g) \rightarrow X^-(g) + e^-(g)$$

The energy required to remove the first electron is known as first ionization enthalpy and the second electron as second ionization enthalpy and so on. The second ionization enthalpy is always greater than the first ionization enthalpy because once an electron is removed, it becomes a positive ion and its nucleus has increased attraction for electrons. This makes it more difficult to remove the second electron. Ionization enthalpy decreases from top to bottom in a group and increases from left to right in a period. Thus, Cs has the lowest ionization enthalpy and fluorine has the highest ionization enthalpy.

The ionization enthalpy depends on two factors (i) atomic size and (ii) screening effect.

73. Ans. It is defined as the enthalpy released when a neutral gaseous atom takes up an extra electron to form an anion

$$X(g) + e^- \rightarrow X^-(g)$$

The value of electron gain enthalpy depends upon atomic size, nuclear charge etc. With increase in size, the electron gain enthalpy decreases as the nuclear attraction decreases. Thus, electron gain enthalpy decreases in going from top to bottom in a group (F > Cl > Br > I). In a period from left to right the electron gain enthalpy increases due to increase in nuclear charge. Thus halogens have very high electron affinity. However, the electron affinity of fluorine is less than that of chlorine. This is because fluorine has small atomic size (only two shells). Also electron repulsion is more in case of fluorine because $1s^2$ is very near to $2p^5$.

74. Ans. Electron gain enthalpy generally becomes more negative across a period as we move from left to right. Within a group, electron gain enthalpy becomes less negative down a group. However, adding an electron to the 2p-orbital leads to greater repulsion than adding an electron to the larger 3p-orbital. Hence, the element with most negative electron gain enthalpy is chlorine, the one with the least negative electron gain enthalpy is phosphorus.

75. Ans. (i) II is a reactive metal.

(ii) IV is a reactive non-metal.

(iii) I is a noble gas.

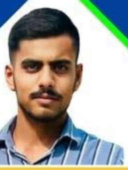
(iv) III forms a stable binary halide of the formula AX_2.

TOPPER'S CHOICE

IIT-JEE MAIN & ADVANCE | MEDICAL-NEET | COMMERCE

12th CBSE RESULTS 2023

Harshdeep

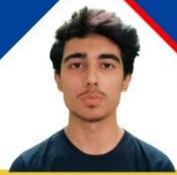


PHY - 88
CHM - 95
BIO - 90

Nidhi

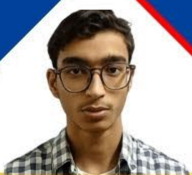
PHY -95
CHM -99
MATH -96

Krish



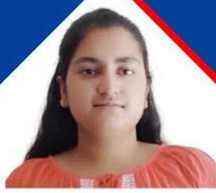
PHY - 86
CHM - 93
MATH - 95

Abhinav



PHY - 85
CHM - 98
BIO - 90

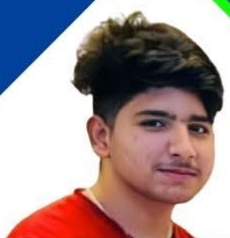
Vanshika



PHY - 84
CHM - 92
MATH - 93

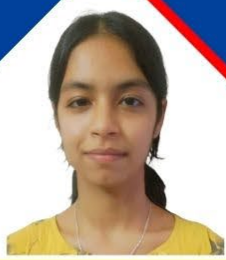
10th CBSE RESULTS 2023

Shivam



ENG - 96
SCI - 95
MATH - 94

Gargi



ENG - 94
MATH - 90

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