Date: 16/10/2021

Q1. With which quantum number does every period in periodic table begin?

Q2. What are *s*-block elements?

Q3. Write the IUPAC name and symbol of an element having atomic number 118.

Q4. How many elements can be accomodated in the present set up of the long form of periodic table?

Q5. Which important property did Mendeleev use to classify the elements in his periodic table?

Q6. Arrange the elements F, Cl, O and N in the correct order of their chemical reactivity in terms of oxidising property.

Q7. State modern periodic law.

Q8. Write the atomic number of the element present in the third period and seventeenth group of the period table.

Q9. What would be the IUPAC name and symbol for the element with atomic number 120?

Q10. Write the IUPAC name and symbol for the element with atomic number 119.

Q11. What is meant by Lanthanoids and Actinoids?

Q12. To which series do man-made elements belong?

Q13. Give general electronic configurations of *d*-block elements.

Q14. What are *d*-block elements? Why are they called transition metals?

Q15. What are representative elements?

Q16. What are *p*-block elements? Give their general electronic configurations.

Q17. Give general electronic configurations of *s*-block elements.

Q18. Which of the following requires highest energy?

(a) $M(g) \longrightarrow M^{+}(g)$
(b) $M(g) \longrightarrow M^{2+}(g)$
(c) $M(g) \longrightarrow M^{3+}(g)$
(d) $M(g) \longrightarrow M^{4+}(g)$

Q19. Give general electronic configuration of least reative group. Why are they least reative?

Q20. Which orbitals are filled with electrons in 3rd period?

Q21. Give general electronic configuration of *f*-block elements.

Q22. What are inner transition metals? Why are they called rare earth metals?

Q23. Which of the Lanthanoids is man-made element?

Q24. How does ionisation energy vary (a) down the group, (b) along the period from left to right?

Q25. Why do noble gases have bigger atomic size than halogens?

Q26. A and B belong to same group of periodic table. 'A' has higher atomic number than B. Which will have lower ionisation energy and why?

Q27. The problem of position of isotopes in the periodic table was avoided by arranging elements in ascending order of ...

Q28. Chlorine has .... electron affinity than Fluorine.

Q29. When compared to lithium, it is easier to remove valence electron from K (Potassium) because

Q30. Two different elements may have same mass number but not the same ...

Q31. Lanthanoids and actinoids belong to .... block of the periodic table.

Q32. Isoelectronic species have the same number of ....

Q33. Why fluorine has lesser electron gain enthalpy than chlorine?

Q34. Why is second electron gain enthalpy (electron affinity) – ve (i.e., energy is absorbed)?

Q35. Generally, electron gain enthalpy .... on going down a group, and .... on going across the period from left to right.

Q36. Define electron gain enthalpy.

Q37. Among, Li, Na, K, Rb, Cs, the element with lowest ionisation energy is ....

Q38. Name of the radioative element of group 17 is .... and of group 18 is ....

Q39. Among (the non-radioactive) halogens, the element that has the lowest electron affinity is ....

Q40. Which is the smallest among Na^+, Mg^{2+}, Al^{3+} and why?

Q41. How do the basic character and solubility in water vary from $Be(OH)_2$ to $Ba(OH)_2$?

Q42. Why do alkali metals have lowest ionisation energy?

Q43. The first transition series is called .... transition series.

Q44. Second period ends with ....

Q45. The electronic configuration of Re^{3+} is $[Xe]4f^{14}\,5d^4$, the number of unpaired electrons in this ion is ....

Q46. The solubility of alkali metal carbonates .... in going down the group.

Q47. The formula for fluoride of carbon is ....

Q48. Generally, the atomic size along a period gradually decreases due to increase in ....

Q49. Why does the first ionisation energy increase as we go from left to right along a given period of periodic table?

Q50. What is the nature of oxides formed by most of *p*-block elements?

Q51. How does electronegativity vary (a) down the group, (b) across the period from left to right?

Q52. Define (a) metallic radius, (b) van der Waals' radius.

Q53. Which has the largest ionic radius among Ca^{2+}, Mg^{2+}, Ba^{2+}?

Q54. Which out of the N or O has higher negative electron gain enthalpy?

Q55. Atomic number (Z) of an element is 108. Write its electronic configuration and name the group to which it belongs.

Q56. Arrange the following species in increasing order of their size: Mg^{2+}, Al^{3+}, Na^{+}, O^{2-}, F^{-}.

Q57. Predict the position of the element in the periodic table satisfying the electronic configuration $(n-1)\,d^1\,ns^2$ for $n = 4$.

Q58. Out of Na and Mg, which has higher second ionisation energy?

Q59. What is the basic theme of organisation in the periodic table?

Q60. Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that?

Q61. On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

Q62. What is the basic difference in approach between the Mendeleev's Periodic law and the Modern Periodic Law?

Q63. Which of the following statements related to the modern periodic table is incorrect?
- (a) The *p*-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a *p*-shell.
- (b) The *d*-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a *d*-subshell.
- (c) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.
- (d) The block indicates the value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.

Q64. Which element do you think would have been named by
- (a) Lawrence Berkeley laboratory?
- (b) Seaborg's group?

Q65. Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

Q66. In terms of period and group, where would you locate the element with Z = 114?

Q67. Which one of the following statements is incorrect in relation to ionisation enthalpy?
- (a) Ionization enthalpy increases for each successive electron.
- (b) The greatest increase in ionisation enthalpy is experienced on removal of electron from core noble gas configuration.
- (c) End of the valence electrons is marked by a big jump in ionisation enthalpy?
- (d) Removal of electron from orbitals bearing lower *n* value is easier than from orbital having higher *n* value.

Q68. 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.

Q69. 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

Q70. 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

Q71. 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

Q72. Considering the elements F, Cl, O and N, the correct order of their chemical reactivity interms of oxidizing property is:
(a) F > Cl > O > N (b) F > O > Cl > N (c) Cl > F > O > N (d) O > F > N > Cl

Q73. Explain why cations are smaller and anions are larger in radii than their parent atoms.

Q74. Write the general outer electronic configuration of *s*, *p*, *d* and *f*-block elements.

Q75. The first (IE_1) and second (IE_2) ionisation energies (kJ/mol) of a new element desinated by Roman numerals are shown below:

	IE_1	IE_2
I	2372	5251
II	520	7300
III	900	1760
IV	1680	3380

Which of these elements is likely to be (a) a reactive metal, (b) a reactive non-metal, (c) a noble gas, and (d) a metal that forms a binary halide of the formula, AX_2?

Q76. Among the elements Li, K, C, S and Kr, which one is expected to have the lowest first ionisation enthalpy and which one has the highest first ionisation enthalpy?

Q77. Out of O and S, which has higher negative electron gain enthalpy and why?

Q78. Among the elements of the third period Na to Ar, pick out the element:
- (a) with highest first ionisation enthalpy
- (b) with largest atomic radius
- (c) which is most reative non-metal
- (d) which is most reactive metal

Q79. Name a species that will be ioselectronic with each of the following atoms or ions:
(a) Ne (b) Cl^- (c) Ca^{2+} (d) Rb

Q80. A, B, C, D and E have the following electronic configuration:

A: $1s^2\,2s^2\,2p^1$ B: $1s^2\,2s^2\,2p^6\,3s^2\,3p^1$

C: $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$ D: $2s^2\,2s^2\,2p^6\,3s^2\,3p^5$

E: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$

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

Q81. Arrange the following elements in increasing order of non-metallic character: B, C, Si, N and F.

Q82. Arrange the following elements in increasing order of metallic character: B, Al, Mg and K.

Q83. Lanthanoids and actinoids are placed in separate rows at the bottom of periodic table. Explain the reason for this arrangement.

Q84. Arrange the following ions in the order of increasing size:

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

Q85. What does atomic radius and ionic radius really mean to you?

Q86. Why do elements in the same group have similar physical and chemical properties?

Q87. Which of the elements Na, Mg, Si and P would have the greatest difference between the first and second ionisation enthalpies? Briefly explain your answer.

Q88. 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 iodine (d) Element with atomic number 114 and fluorine

(e) Element with atomic number 120 and oxygen.

Q89. Among the second period elements, the actual ionisation enthalpies are in the order

Li < B < Be < C < O < N < F < Ne

Explain, why

(a) Be has higher $\Delta_i H$ than B? (b) O has lower $\Delta_i H$ than N and F?

Q90. 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.

Q91. 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.

Q92. 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.

Q93. What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

(a) F^- (b) Ar (c) Mg^{2+} (d) Rb^+

Q94. How do atomic radius vary in a period and in a group? How do you explain the variation?

Q95. Identify the group and valency of the element having atomic number 119. Also predict the outermost electronic configuration and write the general formula of its oxide.

Q96. All transition elements are *d*-block elements but all *d*-block elements are not transition metals. Explain.

Q97. 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

Q98. What is the basic difference between the terms electron gain enthalpy and electronegativity?

Q99. Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first electron gain enthalpy? Justify your answer.

Q100 Which of the following pairs of elements would have a more negative electron gain enthalpy and electronegativity? (a) F, O (b) Cl, F.

Q101 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?

Q102 What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down in a group?

Q103 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?

Q104 (a) Write the general electronic configuration of *d*-block elements.
(b) What is the oxidation state and covalency of Al in $[AlCl(H_2O)_5]^{2+}$?
(c) Which out of the following will have the most negative electron gain enthalpy and which have the least negative?
P, S, Cl and F.

Q105 Give the reasons 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.

Q106 Give a brief account for the following:
(a) Anions are bigger in size than their parent atom.
(b) Oxygen has lesser first ionization enthalpy than nitrogen.
(c) Fluorine has less negative electron gain enthalpy than chlorine.

Q107 Account for the following:
(a) Ionization enthalpy of nitrogen is more than that of oxygen.
(b) A cation is always smaller than its parent atom.
(c) Noble gases have large positive electron gain enthalpies.

Q108 (a) Assign the position of elements having outer electronic configuration:
(i) $ns^2\ np^4$ for $n = 3$
(ii) $(n-1)d^2\ ns^2$ for $n = 4$
(b) Among the elements B, Al, C and Si
(i) which has the highest first ionization enthalpy?
(ii) which has the most negative electron gain enthalpy?
(iii) which has the largest atomic radius?
(iv) which has the most metallic character? Give reasons.

Q109(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 compound which might be formed by silicon and bromine.

Q110(a) Name the most metallic element in second period and most non-metallic element.
(b) Name the element with (a) largest atomic radius, (b) smallest atomic radius in third period.
(c) Name the element having general electronic configuration $ns^2\ np^4$ in fourth period.

Q111(a) Which is largest in size—
Cu^+, Cu^{2+}, or Cu, and why?
(b) Which element in periodic table has highest I.E. (ionisation energy)?
(c) Which element is more metallic — Mg or Al and why?

Q112(a) How does basic character of oxides and hydroxides down the group in alkali metals change? Why?
(b) How does reducing power of elements vary in Group I?

Q113 Arrange the species in each group in order of increasing ionization energy and give reason:
(a) K^+, Cl^-, Ar (b) Na, Mg, Al (c) C, N, O

Q114 Assign a reason for each of the following statements.
(a) First ionization enthalpy of boron (Z = 5) is slightly less than that of beryllium (Z = 4).
(b) Electron gain enthalpy of F is less negative than chlorine (Cl).
(c) The size of an anion is always larger than that of parent atom.

Q115 Give reason for the following:
(a) Halogens act as good oxidising agent.
(b) Electron gain enthalpy of noble gas is almost zero.
(c) Na and Mg+ have same number of electrons but removal of electron from Mg+ requires more energy.

Q116(a) Write the general electronic configuration of *d*-block elements.
(b) Which of the following elements has most positive electron gain enthalpy?
F, N and Ne
(c) How would you explain the fact that first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher that that of magnesium?

Q117 From each set, choose the atom which has largest ionisation enthalpy and explain your answer:
(a) F, O, N (b) Mg, P, Ar (c) B, Al, Ga

Q118 Among the elements of second period Li to Ne, pick out element:
(a) with the highest first ionisation energy (b) with highest electronegativity
(c) with largest atomic radius (d) that is most reactive non-metal
(e) that is most reactive metal (f) with valency equal to 4.

Q119 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.

Q120 What are the major differences between metals and non-metals?

Q121 Would you expect the first ionisation enthalpies for two isotopes of the same element to be the same or different? Justify your answer.

Q122 Describe the theory associated with the radius of an atom as it

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

Q123 How would you react to the statement that the electronegativity of N on Pauling scale is 3.0 in all the nitrogen compounds?

Q124 Consider the element N, P, O and S and arrange them in order of:

(a) increasing first ionisation enthalpy.

(b) increasing negative electron gain enthalpy

(c) increasing non-metallic character.

Q125 The first (IE_1) and second (IE_2) ionisation enthalpies (kJ Mol^{-1}) of three elements I, II and III are given below

Element	IE_1	IE_2
I	403	2640
II	549	1060
III	1142	2080

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

Q126 Select from each group of species which has the smallest radius stating the appropriate reason:

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

Q127 First member of each representative elements (i.e., *s* and *p*-block elements) show anomalous behaviour. Illustrate with two examples.

Q128 Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionisation enthalpy than nitrogen. Explain.

Q129 Predict the formula of the stable binary compounds that would be formed by the combination of the following pairs of elements.

(a) Lithium and oxygen (b) Magnesium and nitrogen

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

(e) Phosphorus and fluorine (f) Element with atomic number 71 and fluorine

Q130 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	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) [N.C.E.R.T.]

Q131 Assign the position of the element having outer electronic configuration

(a) $ns^2\ np^4$ for $n = 3$ (b) $(n-1)\ d^2\ ns^2$ for $n = 4$, and

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

Q132 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.

Date: 16/10/2021

S1. Principal quantum number

S2. Group 1 and 2 elements belong to *s*-block elements because differentiating electron enters s-orbital.

S3. Ununoctium, Uuo

S4. 118 elements can be accommodated in the present set up of the long form of the periodic table.

S5. Mendeleev used atomic weight to classify the elements in his periodic table.

S6. F, Cl, O and N can be arranged in order of their chemical reactivity in terms of oxidising property as F > O > Cl > N

S7. The properties of elements are periodic functions of their atomic numbers i.e. properties of elements depend upon atomic number.

S8. Element present in the third period and seventeenth group of the periodic table has atomic numbers 17 and its name is chlorine.

S9. IUPAC name-Unbinilium, symbol-Ubn.

S10. Ununennium, Uue.

S11. The fourteen elements after La(57) are called Lanthanoids and the fourteen elements after Ac(89) are called actinoids.

S12. Actinoid series (*f*-block elements).

S13. $(n-1)d^{1-10}\ ns^{0-2}$ is general electronic configuration of *d*-block elements.

S14. Those elements in which differentiating electron enter *d*-orbital are called *d*-block elements. They are also called transition metals because they are less electropositive than *s*-block but more electropositive than *p*-block elements.

S15. *s*-Block and *p*-block elements together are called representative elements.

S16. *p*-Block elements are those in which valence *p*-orbital is progressively filled. Group 13 to 18 are *p*-block elements. General electronic configuration is $ns^2\ np^{1 \text{ to } 6}$.

S17. $ns^{1 \text{ to } 2}$

S18. (d) $M(g) \longrightarrow M^{4+}(g)$ requires highest energy to loose four electrons.

S19. $ns^2\ np^6$ is general electronic configuration of least reactive group because they have stable electronic configuration.

S20. In third period, 3*s* and 3*p* orbitals are filled.

S21. $(n-2)f^{1-14}\,(n-1)d^{0-1}\,ns^2$ is general electronic configuration of *f*-block elements.

S22. Inner transition metals are those in which inner *f*-orbital is progressively filled. *f*-block elements are called inner transition metals. They are also called rare earth metal because they are rarely found in the earth's crust.

S23. Promethium (Pm), atomic number 61 is man-made element.

S24. (a) Ionisation enthalpy decreases down in the group because atomic size increases and effective nuclear charge decreases, less energy is needed to remove electrons.

(b) Ionisation enthalpy increases along the period from left to right due to increase in effective nuclear charge.

S25. Noble gases have bigger atomic size than halogens because van der Waals' radii are bigger than covalent radii and there is interelectronic repulsion in noble gases.

S26. 'A' has lower ionisation energy because it is bigger in size.

S27. Atomic number

S28. higher

S29. Ionisation enthalpy

S30. Atomic number

S31. *f*-Block

S32. Electrons

S33. In fluorine, there is more interelectronic repulsion between valence electrons due to exceptionally small size than chlorine, therefore, it has less electron gain enthalpy as less energy is released on gaining electron.

S34. Second electron is to be added to negatively charged ions which will repel the incoming electron. the energy required to overcome repulsion is more than energy released on gaining electron so, net energy is absorbed.

S35. Decreases, increases

S36. Electron gain enthalpy is defined as energy released when neutral gaseous atom gains 1 electron.

S37. Cs has lowest ionisation energy

S38. Astatine (At), Radon (Rn)

S39. Iodine

S40. Al^{3+} is smallest because it has the highest number of protons (13) among Na^+, Mg^{2+} and Al^{3+} ions, due to which effective nuclear charge is maximum.

S41. Basic character and solubility increases from $Be(OH)_2$ to $Ba(OH)_2$ because hydration energy increases more than lattice energy.

S42. They have largest atomic size, therefore, there is less force of attraction between valence electrons and nucleus.

S43. 3*d*

S44. Neon

S45. Four

S46. Increases

S47. CF_4

S48. Effective nuclear charge

S49. It is due to decrease in atomic size and increase in effective nuclear charge.

S50. They form mostly acidic oxides. Some of them form amphoteric and neutral oxides also.

S51. (a) Electronegativity goes on decreasing down in the group due to increase in atomic size.

(b) It goes on increasing along the period from left to right due to decrease in atomic size.

S52. (a) Metallic radius is half the distance between centres of nuclei of two atoms of metal held together by metallic bond.

(b) van der Waals' radius is half of the distance between centres of nuclei of two atoms held by weak van der Waals' forces of attraction.

S53. Ba^{2+} due to greater number of shells.

S54. O has high electron affinity because 'N' has half filled *p*-orbitals which are more stable.

S55. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 4f^{14}\ 5s^2\ 5p^6\ 5d^{10}\ 5f^{14}\ 6s^2\ 6p^6\ 6d^6\ 7s^2$

It belongs to group 8 of periodic table.

S56. $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

S57. $3d^1\ 4s^2$ belongs to group 3 of periodic table as it has 1 + 2 = 3 electrons in *s* and *d*-orbitals.

S58. Na has higher second ionisation energy because it acquires nearest noble gas configuration after losing one electron.

S59. The basic theme is to study the properties of elements conveniently.

S60. Atomic weight. He realised that some of the elements did not fit with his scheme of classfication if order of atomic weight is strictly followed. He ignored the order of atomic weights and preferred similarity in properties.

S61. $n = 6$ $l = 0, 1, 2, 3$

$n = 6$	$l = 0$	$6s$	2 electrons	2 elements
$n = 6$	$l = 1$	$6p$	6 electrons	6 elements
$n = 6$	$l = 2$	$6d$	10 electrons	10 elements
$n = 6$	$l = 3$	$6f$	14 electrons	14 elements
				32 elements

S62. Mendeleev's Periodic Law was based on the assumption that properties of elements depend upon atomic weight but according to Modern Periodic Law, properties of elements depend upon atomic number.

S63. (b) is incorrect as *d*-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a *d*-subshell.

S64. (a) Berkelium

(b) Seaborgium

S65. Element in 3rd period and 17th group is of atomic no. 17.

S66. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 4f^{14}\ 5s^2\ 5p^6\ 5d^{10}\ 5f^{14}\ 6s^2\ 6p^6\ 6d^{10}\ 7s^2\ 7p^2$. It belongs to 7th period and group 14.

S67. (d) Removal of electron from orbitals bearing lower *n* value is easier than from orbital having higher *n* value.

S68. (a) Nuclear charge (Z)

S69. (c) Nuclear mass

S70. (c) F > N > C > B > Si

S71. (d) K > Mg > Al > B

S72. (b) F > O > Cl > N

S73. A cation is formed by loss of electron(s). The effective nuclear charge in the cation is more than the neutral atom. Hence cations are smaller than the parent atoms. On the other hand, in anion, effective nuclear charge decreases and the electrons experience lesser pull by the nucleus, which lead to increase in size of anions than parent atoms.

S74. ***s*-block:** $ns^{1\text{ to }2}$ where n = 2 to 7

***p*-block:** $ns^2\ np^{1\text{ to }6}$ where n = 2 to 6

***d*-block:** $(n–1)d^{1–10}\ ns^{0–2}$ where n = 4 to 7

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

S75. (a) II, (b) IV, (c) I, (d) III

S76. K has lowest first ionisation enthalpy whereas Kr has highest first ionisation enthalpy.

S77. 'S' has high electron gain enthalpy because in oxygen, there is more inter-electronic repulsion than sulphur, therefore, more energy is released in case of sulphur on gaining electrons.

S78. (a) Ar as highest first ionisation enthalpy.

(b) Na has largest atomic radius (covalent radius).

(c) Cl is most reactive non-metal in third period.

(d) Na is most reactive metal in third period.

S79. (a) Na^+ (b) Ar (c) S^{2-} (iv) Y^{2+}

S80. A and B belong to same group of periodic table because they have same number of valence electrons.

S81. Si, B, C, N, F are in increasing order of non-metallic character.

S82. B, Al, Mg, K are in increasing order of metallic character.

S83. Lanthanoids and actinoids resemble with each other but do not resemble any other group elements.

S84. $Be^{2+} < Mg^{2+} < Na^+ < Cl^- < S^{2-} < Br^-$

S85. Atomic radius gives size of atom whereas ionic radius gives idea about size of ion.

S86. It is because they have same number of valence electrons. Chemical properties depend upon valence electrons. They also have similar physical properties as these elements are largest in respective periods.

S87. Na has greatest difference between first and second ionisation enthalpies because Na^+ has stable electronic configuration, i.e., $1s^2\ 2s^2\ 2p^6$, therefore, it has very high second ionisation energy.

S88. (a) SiO_2 (b) $AlBr_3$

(c) CaI_2 (d) $UuqF_4$

(e) XO where 'X' is element with atomic number 120.

S89. (a) Be (4) $1s^2 2s^2$ and B (5) $1s^2 2s^2 2p^1$

Be has stable electronic configuration, i.e., completely filled *s*-orbital from which removal of electron is difficult as compared to Boron.

(b) N (7) $1s^2\ 2s^2\ 2p^3$

O (8) $1s^2\ 2s^2\ 2p^4$ and

F (9) $1s^2\ 2s^2\ 2p^5$

Oxygen has lower ionisation energy than 'N' because 'N' has stable electronic configuration due to half filled *p*-orbitals.

O (8) has lower I.E. than F (9) because effective nuclear charge is less than that of fluorine.

S90. I.E. = 2.18×10^{-18} J $\times 6.022 \times 10^{23}$ mol^{-1} = $13.12 \times 10^{+5}$ J mol^{-1} = 1.312×10^{6} J mol^{-1}

S91. It is because of comparison purposes, all the atoms should be in same physical state, i.e., gaseous state.

S92. (a) They have same number of electrons.

(b) $Al^{3+} < Mg^{2+} < Na^{+} < F^{-} < O^{2-}$ N^{3-} is increasing order of ionic radii as number of protons i.e., nuclear charge decreases.

S93. Isoelectronic species are those which have same number of electrons.

(a) F– is isoelectronic with Ne

(b) Ar is isoelectronic with Cl^{-}

(c) Mg^{2+} is isoelectronic with Ne and Na^{+}

(d) Rb^{+} is isoelectronic with Kr.

S94. Atomic radii decrease along a period due to increase in effective nuclear charge which increases due to increase in number of electrons and protons successively.

Atomic radii go on increasing down in the group due to increase in number of shells.

S95. It belongs to group 1. Its valency is equal to 1. Its outermost electronic configuration is $8s^1$ as it will belong to 8th period of the periodic table. Formula of its oxide is $(Uue)_2O$ or M_2O.

S96. Those elements which either themselves or their ions have incompletely filled *d*-orbitals are called transition elements. They are all *d*-block elements because '*d*' orbitals are progressively filled.

But those *d*-block elements or their ions which do not have incompletely filled *d*-orbitals are not transition elements e.g., Zn, Cd and Hg are *d*-block elements but not transition metals.

S97. (c) Principal quantum number.

S98. Electron gain enthalpy is energy released when neutral gaseous atom gains one electron. It has absolute value.

Electronegativity is the measure of tendency to attract shared pair of electrons towards itself in covalently bonded molecule. It is arbitrary or relative value.

S99. Second electron affinity is largely +ve because of repulsion between negatively charged ions and second electron to be added. Energy required to overcome repulsion is more than energy released in gaining electron, so net energy is absorbed.

S100(a) F has more negative electron gain enthalpy than oxygen.

(b) Cl has more negative electron gain enthalpy than fluorine.

S101. It is due to poor shielding effect in case of Ga, the effective nuclear charge is more, therefore, I.E. is slightly higher than that of 'Al'.

In case of Tl, effective nuclear charge is increased due to poor shielding effect of *f*-electrons, therefore, its I.E. is higher than that of In.

S102. It is due to increase in atomic size, decrease in effective nuclear charge, i.e., decrease in force of attraction between nucleus and valence electrons.

S103. Mg is smaller than Na, therefore, first ionisation energy of Na is lower than that of Mg.

After losing one electron, Na^+ has stable noble gas electronic configuration, whereas Mg+ does not have noble gas configuration, therefore, 2nd I.E. of Na is higher than that of Mg.

S104. (a) $(n-1)d^{1-10}\ ns^{0-2}$ is the general electronic configuration of *d*-block elements.

(b) The oxidation state of Al in $[AlCl(H_2O)_5]^{2+}$ is +3,

$$x - 1 + 0 = +2$$

$$x = +3$$

The covalency of 'Al' is 6.

(c) Across a period, electron gain enthalpy becomes more negative as we move from left to right. In case of Cl, $n = 3$, added electron occupies a large region of space and electron-electron repulsion is much less. So, it has higher negative value. Therefore, 'Cl' has highest negative electron gain enthalpy and 'P' has least negative electron gain enthalphy.

S105. (a) This is due to the smaller size of F than Cl. The electron-electron repulsion in 2*p* subshell of F is large and incoming electron is not accommodated with the case as is accommodated in a larger 3*p* subshell of Cl.

(b) Anion is formed when a neutral gaseous atom gains electrons. This increases the number of electrons in the anion while its nuclear charge remains the same. Since the same nuclear charge now attracts the greater number of electrons; therefore, the force of attraction of electrons by the nucleous decreases leading to increase in the size of an anio..

(c) N, Z = 7 : 1s2, 2s2, 2p3

O, Z = 8 : 1s2, 2s2, 2p4

In case of N, electron is to be removed from half filled *p*-orbitals; hence more energy is required leading to the higher value of ionization enthalpy of nitrogen.

S106. (a) Anions are formed when a neutral atom gains one or more electrons. Since, the number of electrons increases and the number of protons remains same, the effective nuclear charge decreases which result in decrease in ionic radii.

(b) Due to the half-filled orbital in nitrogen $1s^2\ 2s^2\ 2p^3$, the stability of this configuration is more and ionization energy is higher than oxygen.

(c) Due to the smaller size and seven electrons in its outermost shell, incoming electrons experience less attraction in F. Hence, less energy will be released in case of F than Cl. in other words, we can say that F has less negative electron gain enthalpy value than Cl.

S107.(a) It is due to half filled p-orbitals in 'N' which are more stable than oxygen which does not have half filled *p*-orbital.

(b) It is due to increase in effective nuclear charge due to decreasae in number of electrons but protons remain the same.

(c) It is due to stable electronic configuration of noble gases and because of inter electronic repulsion, addition of electrons is an enthothermic process, as energy is needed to overcome repulsion.

S108.(a) (i) $ns^2\ np^4$ for $n = 3$

Since $n = 3$; the element belongs to the third period. Further, the element belongs to the *p*-block and group number 16.

(ii) $(n - 1)d^2\ ns^2$ for $n = 4$, the element belongs to period number 4 and group number 4.

(b) (i) Ionization energy increases along a period but decreases going down in the group. Hence carbon has highest first ionization enthalpy since it has the smallest size among all.

(ii) Electron gain enthalpy becomes more negative along a period and less negative down a group, thus, carbon has the most negative electron gain enthalpy due to smallest size.

(iii) Atomic radii decrease along a period but increase down the group, thus, Al has the largest atomic radius.

(iv) Metallic character decreases along a period but increases down a group, thus, Al has the maximum metallic character due to largest atomic size and lowest ionization energy.

S109.(a) $Mg^{2+} < Na^+ < F^- < O^{2-}$ is an increasing order of ionic radii.

(b) Be (4) has electronic configuration $1s^2\ 2s^2$ whereas B(5) has $1s^2\ 2s^2\ 2p^1$.

The energy required to remove electron from completely filled 2*s* orbital is higher than the energy required to remove electron from 2*p* orbital.

(c) $SiBr_4$ is a formula of compound which might be formed by Si and Br_2.

S110.(a) Most metallic element is Li and most non-metallic element is F.

(b) (i) Na has largest atomic radius

(ii) Cl has smallest atomic radius.

(c) Se(34): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$ is element in fourth period with general electronic configuration $ns^2\ np^4$.

S111.(a) Cu is largest due to less effective nuclear charge. It has 29 electrons, 29 protons, Cu^+ has 28 electrons and 29 protons, Cu^{2+} has 27 electrons and 29 protons.

(b) He has highest ionisation energy.

(c) Mg is more metallic due to lower ionisation energy.

S112(a) It goes on increasing down the group due to decrease in ionisation energy and increase in metallic character.

(b) Reducing power of 'Li' is highest, Na is lowest, then it goes on increasing:

Li > Cs > Rb > K > Na

S113(a) $Cl^- < Ar < K^+$ because nuclear charge goes on increasing.

(b) Na < Al < Mg because Mg has stable electronic configuration and Na has large atomic size due to which it has lowest ionisation energy.

(c) C < O < N because N has half filled *p*-orbital which is more stable, whereas carbon is larger in size.

S114(a) It is due to completely filled '*s*' orbital in case of Be (Z = 4), which is more stable.

(b) It is due to strong interelectronic repulsion between valence electrons of 'F' atoms as compared to 'Cl' atoms.

(c) Anions are formed by gain of electrons but number of protons remain the same, therefore, effective nuclear charge decreases, that is why the size of anion is larger than neutral atom.

S115(a) Due to highly negative electron gain enthalpy they act as good oxidising agent as they can gain electrons easily.

(b) Electronic configuration of noble gases is such that all sub-shells are completely filled. Hence their electron gain enthalpy is almost zero.

(c) Sodium has eleven electrons and eleven protons but number of protons in Mg^+ are twelve, though it has eleven electrons. Due to higher effective nuclear charge in case of Mg^+, removal of electron from it requires more energy.

S116(a) The general elecetronic configuration of *d*-block elements is $(n-1)d^{1-10}\ ns^{0-2}$.

(b) Ne has most positive electron gain enthalpy due to interelectronic repulsion.

(c) It is because Mg is smaller than Na, therefore, first ionisation energy of Na is lower than Mg but after losing one electron, sodium acquires nearest stable noble gas configuration, therefore, energy required to remove second electron is higher as compared to Mg, that is why, 2nd I.E. of Na is higher.

S117(a) F has highest ionisation enthalpy due to smallest atomic size, therefore, highest effective nuclear charge.

(b) Ar has highest ionisation enthalpy due to stable electronic configuration.

(c) B has highest ionisation enthalpy due to smallest atomic size and, therefore, highest effective nuclear charge.

S118(a) Ne (b) F

(c) Li (Covalent radii) (d) F

(e) Li (f) C

S119. (a) Nitrogen (b) Magnesium

(c) Oxygen (d) Group I

S120.

Metals	Non-metals
(a) They can lose electrons easily	(a) They can gain electrons easily.
(b) They are electropositive.	(b) They are electronegative.
(c) They have mostly 1 to 3 valence electrons.	(c) They have 4 to 8 valence electrons.
(d) They are mostly malleable and ductile.	(d) They are brittle.
(e) They mostly form basic oxides.	(e) They mostly form acidic oxides.
(f) They are good conductors of heat and electricity.	(f) They do not conduct heat and electricity.

S121. They are same because effective nuclear charge is same, Isotopes differ in number of neutrons which do not affect ionisation energy.

S122. (a) When an atom gains electron, its radius increases due to decrease in effective nuclear charge.

(b) When an atom loses electron, it forms cation, which is smaller in size due to greater effective nuclear charge.

S123. It is because nitrogen is larger in size than oxygen and fluorine and therefore less electronegative than oxygen and fluorine.

S124. (a) S < P < O < N is increasing order of first ionisation energy.

(b) N < P < O < S is increasing order of negative electron gain enthalpy.

(c) P < S < N < O is increasing order of non-metallic character.

S125. (a) III is non metal.

(b) I is an alkali metal.

(c) II is alkaline earth metal.

S126. (a) O has smallest radius because it has highest effective nuclear charge.

(b) K^+ has smallest radius due to greater effective nuclear charge.

(c) Cl has smallest radius due to greater effective nuclear charge.

S127. Li is first element of group 1. It shows anomalous behaviour. Lithium forms monoxides. Lithium forms mostly covalent compounds, other form ionic compounds.

B is the first element of group 13. It shows anomalous behaviour. It does not form B^{3+}, others form tripositive ions. It is metalloid whereas others are metals.

It forms acidic oxide whereas others form amphoteric and basic oxides.

S128 Nitrogen has half filled *p*-orbitals ($2p^3$) which are more stable, that is why addition of extra electron needs energy to overcome repulsion, therefore, electron gain enthalpy is positive. Oxygen has 4 electrons in 2*p*-orbitals. It can gain electron easily and electron gan enthalpy is negative.

Oxygen has lower ionisation because after losing one electron it will become stable ($2p^3$), whereas in N, electron needs to be removed from ($2p^3$) which is already stable.

S129 (a) Li_2O (b) Mg_3N_2

(c) AlI_3 (d) SiO_2

(e) PF_5 (f) LuF_3

S130 (a) V is least reactive element due to highest ionisation enthalpy.

(b) II is most reactive metal due to least first ionisation energy.

(c) III is most reactive non-metal due to second highest first ionisation energy.

(d) V is least reactive non-metal because it has +ve electron gain enthalpy and is noble gas due to highest ionisation enthalpy.

(e) VI can form stable binary halide of the formula MX_2 because it has lowest second ionisation enthalpy.

(f) I can form covalent halide of the formula MX[X is halogen] due to high ionisation enthalpy in group I due to small size. It second I.E. is very high, therefore, it is smallest alkali metal.

S131 (a) $ns^2\ np^4$, $3s^23p^4$ belongs to 3rd period and 16th group.

(b) $(n-1)d^2\ ns^2$, $3d^24s^2$ belongs to 4th period and 4th group. The element is Titanium.

(c) $(n-2)f^7\ (n-1)d^1\ ns^2$, i.e., $4f^75d^16s^2$ belongs to lanthanoids. It is Gadolinium with atomic number 64. It is placed at 7th position in the first row at the bottom of periodic table.

S132 It is because in group I, ionisation enthalpy decreases down the group, therefore, reactivity increases as tendency to lose electrons increases.

In Group 17, order of reactivity is F > Cl > Br > I.

It is because F_2 has low bond dissociation energy, high electron affinity and F^- has highest hydration energy, that is why F is most reactive. Oxidising power decreases down in the group, therefore, the reactivity decreases.