Periodic Table

SOLUTIONS

Level – I

DEVELOPMENT OF PERIODIC TABLE

1. The law of triads is applicable to
(a) Hydrogen, oxygen, nitrogen (b) Chlorine, bromine, iodine
(c) Sodium, neon, calcium (d) None

2. Which electronic configurations represent to a transition element?
(a) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^6$ (b) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^1$
(c) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^2, 4s^2$ (d) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6, 4s^2$
$1s^2,2s^2p^6,3s^2p^6d^2,4s^2$ is the electronic configuration of a transitional element as last electron enters into d subshell.

3. An element having electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ is:
(a) An inert gas (b) A transition element
(c) A inner transition element (d) A representative element
As last electron is filled in 3p−orbital so it belongs to p-block. s-block and p-block elements are representative elements.

4. Element A of group III A combines with element B of group VI A. The resulting compound may have the formula
(a) A_2B_3 (b) A_3B_2 (c) A_5B_6 (d) A_6B_5

5. How many elements are present in the fourth period of the modern periodic table?
(a) 8 (b) 10 (c) 18 (d) 32
The number of elements present in the fourth period is 18.
It includes 2s block elements, 6p block elements and 10d block elements.

6. In the fourth period of the periodic table, how many elements have one or more 4*d* electrons?
(a) 2 (b) 18 (c) 0 (d) 6
In fourth period of periodic table, There is no element having one or more 4d electron.

7. The element which is cited as an example to prove the validity of Mendeleev's periodic law is
(a) Indium (b) Helium (c) Gallium (d) None

8. Which one of the following shows paramagnetic character?
(a) Sc^{3+} (b) Fe^{2+} (c) Mn^{7+} (d) Ti^{4+}

9. Which element was named as eka-silicon in Mendeleev classification of elements ?
(a) Germanium (b) Gallium (c) Thallium (d) Selenium
Eka-silicon in Mendeleev's periodic table is known as germanium.

10. The element which has a tendency to show positive and negative oxidation states is :
(a) Lithium (b) Gallium (c) Iodine (d) Cerium
Iodine exhibits both positive and negative oxidation states. It exhibits oxidation states -1, +1, +3, +5 and +7.

11. The non metal which exists in liquid state at room temperature is :
(a) Na (b) Br (c) Mg (d) Ga
Mercury and bromine both are present in liquid state at room temperature but mercury is metal and bromine is non-metal.

So, liquid non-metal at room temperature is bromine.

ATOMIC RADIUS

12. For the element (X), student (a) measured its radius as 102 nm, student (b) as 109 nm and (c) as 100 nm using same apparatus. Their teacher explained that measurements were correct by saying that recorded values by (a), (b) and (c) are:
(a) Crystal, van der Waal and covalent radii
(b) Covalent, crystal and van der Waal radii
(c) van der Waal, ionic and covalentradii
(d) None is correct
As we know,
Van der wall's radius > Metallic radius (crystal) > Covalent radius
So A, B and C are crystal, van der Waals' and covalent radii respectively

13. Isoelectronic species are
(a) CO and NO
(b) N_2 and CO
(c) O_2 and N_2
(d) CO_2 and CO
Isoelectronic species are those species which have same number of electrons. No. of electrons in:
CO: 6+8=14
HF:1+9=10
N_2: 7+7=14
N_2^+ :7+7-1=13
O_2^- :8+8+1=17
CO and N_2 are isoelectronic

14. Which of the following is iso-electronic with carbon atom?
(a) Na^+
(b) Al^{3+}
(c) O^{2-}
(d) N^+
N^+ is isoelectronic to C. Both contain 6 electrons out of which 2 are core electrons and 4 are valence electrons. N has 7 electrons. It loses one electron to form N^+ with 6 electrons.

15. Which one of the following constitutes a group of the isoelectronic species?
(a) C_2^{2-}, O_2^-, CO, NO
(b) NO^+, C_2^{2-}, CN^-, N_2
(c) CN^-, N_2, O_2^{2-}, C_2^{2-}
(d) N_2 , O_2^-, NO^+, CO
Isoelectronic species possess same number of electrons. NO^+, C_2^{2-}, CN^-, N_2
, each have 14 electrons and thus are isoelectronic.

16. The correct order of radii is
(a) N < Be < B
(b) $F^- < O^{2-} < N^{3-}$
(c) Na < Li < K
(d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$
The correct order of atomic radii is $F^- < O^{2-} < N^{3-}$. These are isoelectronic species.
Fluoride ion has the maximum nuclear charge. Hence, it has the maximum attraction of the nucleus for the valence electrons. Hence, it has minimum ionic size. N^{3-} has minimum nuclear charge. Hence, it has minimum attraction of the nucleus for valence electrons. Hence, it has maximum ionic size.

17. The ionic radii of N^{3-}, O^{2-} and F^- are respectively given by:
(a) 1.36, 1.40, 1.71
(b) 1.36, 1.71, 1.40
(c) 1.71, 1.40, 1.36
(d) 1.71, 1.36, 1.40

$\xrightarrow[\text{size decreasing}]{N^{3-}, O^{2-}, F^-}$

z/e increases so size ↓

So N^{3-} has longest radii imporve to O^{2-}

so 1.71 > 1.40 > 1.36

18. Element Hg has 2 oxidation state Hg^{+1} and Hg^{+2}. The correct order of radii of these ions
(a) $Hg^{+} > Hg^{2+}$
(b) $Hg^{+} < Hg^{2+}$
(c) $Hg^{+} = Hg^{2+}$
(d) Can't determine
Hg^{2+} will have smaller radius than Hg^{+} because as the metal Hg loose two electron the effective nuclear charge increase in Hg^{2+} as compased to Hg+ and result in decrease in atomic radius.

19. Ionic radii of
(a) $Ti^{4+} < Mn^{7+}$
(b) $^{35}Cl^- < ^{37}Cl^-$
(c) $K^{+} > Cl^-$
(d) $P^{3+} > P^{5+}$
Ionic radii of $P^{3+} > P^{5+}$
More is the positive charge smaller the size.

$$\text{ionic radii} \propto \frac{1}{\text{charge}}$$

In case of Ti^{4+} & Mn^{7+} $\rightarrow$ Mn^{7+} will be smaller
In $^{35}Cl^-$ and $^{37}Cl^-$ $\rightarrow$ ionic radii remains same for both.
As no. of e^- and no. of p^+ remains same in K^+ and Cl^-, K^+ will be smaller as cation are smaller in size than anions.

20. Cl^- and K^+ are isoelectronic then
(a) Their sizes are same.
(b) Cl^- ion is relatively bigger than K^+ ion.
(c) K^+ ion is bigger than Cl^- ion.
(d) Their sizes depend on other cation and anion.
Potassium ion is smaller because of high nuclear charge.

21. Which possesses the largest radius?
(a) Fe
(b) Fe^{2+}
(c) Fe^{+}
(d) Fe^{3+}

22. Which of the following is largest
(a) Cl^-
(b) S^{2-}
(c) Na^+
(d) F^-
Na^+ & O^{2-} are isoelectronic with Ne. Similarly S^{-2} & Cl^- are iso electronic with Ar. Thus the size of Cl^-S^{-2} is larger than Na^+ & O^{2-} as Ar is larger than Ne and isoelectronic species have similar size. Also among iso electronic species, elements of more negative charge is larger. Hence the largest is S^{-2} i.e.

23. The correct order of atomic size is
(a) Be > C > F > Ne
(b) Be < C < F < Ne
(c) Be > C > F < Ne
(d) F < Ne < Be < C
The size of an element within a period decreases from left to right up to halogens because of the net increase in nuclear charge on the outermost electron. From halogen to noble gases the size increases as the last electron completes the octet and compensates for the increase in nuclear charge. Therefore among B, C, F, Ne, the size will follow the order:

24. Which one of the following is smallest in size?
(a) Na^+ (b) O^{2-} (c) N^{3-} (d) F^-
An isoelectronic series is a group of ions that all have the same number of electrons.
For example, one isoelectronic series could include N^{-3}, O^{2-}, F^-, Na^+. These all have ten electrons. The number of protons, though, increases as atomic number increases, so nuclear charge increases. When we consider effective nuclear charge then, the ions of greater nuclear charge attract those ten electrons more strongly and pull them in more tightly. Therefore, the radii of the ions in an isoelectronic series decrease as nuclear charge (or atomic number) increases.

25. The sizes of X, X^+ and X^- follow the order
(a) $X^+ > X^- > X$ (b) $X^- > X^+ > X$ (c) $X^- > X > X^+$ (d) $X > X^- > X^+$

26. Which of the following has largest size?
(a) Na (b) Na^+ (c) Mg (d) Mg^{2+}
If we move from left to right along a period in the periodic table, size of the elements decrease because effective nuclear charge increases from left to right along a period.So size of Na is bigger than Mg.
Now size of any ion is smaller than it's parent element.So Na has larger size than Na^+ and Mg has larger size than Mg^{2+}.
So decreasing order of size follows as $Na > Mg > Na^+ > Mg^{2+}$.
For information size of each species are given bellow:
Na=154 pm Na^+= 116 pm, Mg=130 Mg^{2+} = 86 Pm.

27. Which of the following is arranged in decreasing order of size?
(a) $Mg^{2+} > Al^{3+} > O^{2-}$ (b) $O^{2-} > Mg^{2+} > Al^{3+}$
(c) $Al^{3+} > Mg^{2+} > O^{2-}$ (d) $Al^{3+} > O^{2-} > Mg^{2+}$
O^{2-} with atomic no. 8 has 8+2 = 10 electrons
Mg^{2+} with atomic no. 12 has 12-2= 10 electrons
Al^{3+} with atomic no. 13 has 13-3= 10 electrons
Since all these ions have 10 electrons in their shell therefore these are iso-electronic species. The more + the charge, the smaller the ionic radius. Remember that − means adding electrons. These electrons go in the outermost shells. Also, when an atom loses electrons, it clings ever more tightly to the ones it has left, further reducing the ionic radius. therefore the order of ionic radii will be: $O^{2-} > Mg^{2+} > Al^{3+}$

28. In which of the following pairs the difference between the covalent radii of two metals is maximum
(a) K, Ca (b) Mn, Fe (c) Co, Ni (d) Y, Zr
It is known,
Covalent radii decrease on moving from left to right and in transition elements, the size variation is not seen that much.
So, answer is K,Ca.

IONIZATION ENERGY

29. The ionisation energy of Al is smaller than that of Mg because
(a) Atomic size of Al > Mg.
(b) Atomic size of Al < Mg.
(c) Penetration of s-subshell electrons in case of Mg is greater than that of P subshell of Al.
(d) Unpredictable.
It can be seen from the electronic configuration that, Al have one unpaired electron in p orbital and Mg have two paired electron in s-orbitals, hence IP of Al is low.

30. IP_1 and IP_2 of Mg are 178 and 348 kcal $mole^{-1}$. The energy required for the reaction
$Mg \longrightarrow Mg^{2+} + 2e^-$ is
(a) +170 kcal/mol (b) +526 kcal/mol (c) –170 kcal/mol (d) –526 kcal/mol

$Mg(g) \longrightarrow Mg^{+2}(g) + 2e^{-}$ is
$Mg \rightarrow Mg^{+} \rightarrow Mg^{+2}$ so energy required for the reaction is 178+348=526 kcal/mol.

31. The IP_1, IP_2, IP_3, IP_4 and IP_5 of an element are 7.1, 14.3, 34.5, 46.8, 162.2 eV respectively. The element is likely to be
(a) Na (b) Si (c) F (d) Ca
IP_1,IP_2,IP_3,IP_4 and IP_5 of an element are 7.1,14.3,34.5,46.8.162.2eV respectively. The element is likely to be Si. The jump in IP values exist in IP5 and thus, removal of fifth electron occurs from inner shell. Thus, the element contains four electrons in its valence shell.
The electronic configuration of silicon is {Ne}$3s^2$,$3p^2$. These 4 electrons need lesser IP than the fifth electron which has to be released from $2p^6$, requires very high IP.

32. An element has successive ionization enthalpies as 940 (first), 2080, 3090, 4140, 7030, 7870, 16000 and19500 kJ mol^{-1}. To which group of the periodic table does this elements belong
(a) 14 (b) 15 (c) 16 (d) 17
The ionization energies difference is more between 6th and 7th energies so it contains 6 electrons in the outer most shell.

33. Amongst the following elements (whose electronic configuration are given below) the one having highest ionization energy is
(a) [Ne] $3s^2$ $3p^1$ (b) [Ne] $3s^2$ $3p^3$
(c) [Ne] $3s^2$ $3p^2$ (d) [Ar] $3d^{10}$ $4s^2$ $4p^3$
The IE increases along a period and decreases down the group. Also, IE of 15 is more than group 16 as group 15 has half-filled p subshell giving extra stability.

34. The 1st ionization energy of Na, Mg, Al and Si are in the order :
(a) Na < Mg > Al < Si (b) Na > Mg > Al > Si
(c) Na < Mg < Al > Si (d) Na > Mg > Al < Si
As we move across the period, nuclear charge increases, atomic size decreases hence, ionization enthalpy increases. For Al ($3s^2 3p^1$), electron has to be removed from partially filled 3p orbital whereas in Mg($3s^2$), electron has to be removed from stable fully filled 3s orbital.
Removal of an electron from stable, fully filled orbital requires more energy than removal of electron from partially filled orbital. Thus, ionisation enthalpy for Mg is greater than ionisation enthalpy for Al. So, the correct order of first ionization enthalpies is: Na<Mg>Al<Si

35. Select correct statement:
(a) More active metals are on the left side of the periodic table
(b) Less active metals are on the left side of the periodic table
(c) Reducing power decreases moving down the group
(d) All are correct statements
The elements on the left side of the periodic table are relatively electron-deficient (i.e., they have few valence electrons), and due to their comparatively low effective nuclear charges (the net positive charge of the protons minus the shielding core electrons below the valence level), their electrostatic hold on these electrons are weak.

36. The ionization energy of boron is less than that of beryllium because:
(a) Beryllium has a higher nuclear charge that boron
(b) Beryllium has a lower nuclear charge than boron
(c) The outermost electron in boron occupies a 2*p*-orbital
(d) The 2*s* and 2*p*orbital of boron are degenerate
B=5: $1s^2 2s^2 2p^1$
Be=4: $1s^2 2s^2$
As we can see from the electronic configuration of Be it is having completely filled outermost shell which is highly stable whereas in B outermost shell consists of one electron which can be removed easily for attaining a stable configuration. Hence ionization energy of B is greater than the ionization energy of Be.

37. Sodium generally does not shown oxidation state of +2, because of its:
(a) High first ionization potential
(b) High second ionization potential
(c) Large ionic radius
(d) High electronegativity

Sodium generally does not show oxidation state of +2, because:
Do to high second ionization potential sodium does not exhibit +2 oxidation state.

38. Which of the following isoelectronic ion has the lowest ionization energy?
(a) K^+
(b) Cl^-
(c) Ca^{2+}
(d) S^{2-}

It is easier to eject an electron from a negatively charged species than a positively charged species. Thus higher the negative charge, easier is the process.

Thus lowest IE among given option is S^{2-}

39. The ionization potentials of Li and K are 5.4 and 4.3 eV respectively. The ionization potential of Na will be:
(a) 9.7 eV
(b) 1.1 eV
(c) 4.9 eV
(d) 5.8 eV

Li, Na, K are elements of some group i.e – IA
In a group from top to bottom I.P value decreases
So, the I.P value of Na will be less than litheium and greater than potassium.
I.P value = Li > Na > K
So, it will be between 5.4 and 4.3 i.e 4.9 ev.

40. The first four I.E. values of an element are 284, 412, 656 and 3210 kJ mol^{-1}. The number of valence electrons in the element are:
(a) One
(b) Two
(c) Three
(d) Four

Considering the first four I.E. values of elements, the $I.I_4$ value of element is high and the difference between $I.E_3$ and $I.E_4$ is very high
I.E2 – I.E1 = 412 – 284 = 128 kJ/mol
I.E3 – I.E2 = 654 – 412 = 242 kJ/mol
I.E4 – I.E3 = 3210 – 656 = 2554 kJ/mol.
The value of 2554 1g/mol is difference between I.E4 and I.E3 suggest that the element is not able to loose its electron after loosing the three electron i.e the element will become stable on loss of three electrons.
So the valancy (or) outermost electrons will be three.

41. Which electronic configuration of an element has abnormally high difference between second and third ionization energy?
(a) $1s^22s^22p^63s^1$
(b) $1s^22s^22p^63s^23p^1$
(c) $1s^22s^22p^6$
(d) $1s^22s^22p^63s^23p^2$

(a) $1s^2\,2s^2\,2p^6\,3s^1 \xrightarrow{+E_1} 1s^2\,2s^2\,2p^6 \xrightarrow{+1E_2} 1s^2\,2s^2\,2p^5 \xrightarrow{+1E_3} 1s^2\,2s^2\,2p^4$

(b) $1s^2\,2s^2\,2p^6\,3s^2\,3p^1 \xrightarrow{+E_1} 1s^2\,2s^2\,2p^6\,3s^2 \xrightarrow{+1E_2} 1s^2\,2s^2\,2p^6\,3s^1 \xrightarrow{+1E_3} 1s^2\,2s^2\,2p^6$

(c) $1s^2\,2s^2\,2p^6 \xrightarrow{+E_1} 1s^2\,2s^2\,2p^5 \xrightarrow{+1E_2} 1s^2\,2s^2\,2p^4 \xrightarrow{1E_3} 1s^2\,2s^2\,2p^3$

(d) $1s^2\,2s^2\,2p^6\,3s^2 \xrightarrow{+E_1} 1s^2\,2s^2\,2p^6\,3s^1 \xrightarrow{+1E_2} 1s^2\,2s^2\,2p^6 \xrightarrow{+1E_3} 1s^2\,2s^2\,2p^5$

In option D, after removal of second valence electron from 3s orbital, the ion formed achieves noble gas configuration. Therefore, to remove the third electron from 2p orbital, a lot of energy is required. Thus, there is an abnormally high difference between second and third ionization enthalpies.

42. Ionization energy of nitrogen is more than oxygen because
(a) Nucleus has more attraction for electrons
(b) Half-filled p-orbital configuration is more stable
(c) Nitrogen atom is bigger than oxygen atom
(d) None

Electronic configuration of Oxygen is
$1s^2 2s^2 2p^4$
Electronic configuration of Nitrogen is
$1s^2 2s^2 2p^3$
Nitrogen has a half filled p subshell which is more stable than partially filled p subshell. Oxygen, on the other hand, can readily loose 1 electron and attain more stable half filled e^- configuration.
That is why, ionisation energy of nitrogen is greater and that of oxygen is lesser as oxygen wants to attain more stable e− configuration by loosing its 1 electron.

43. The second ionization energy is always higher than the first ionization energy because the :
(a) Ion becomes more stable attaining an octet or duplet configurations
(b) Electron is more tightly bound to the nucleus in an ion
(c) Electron is attracted more by the core electrons
(d) None is the correct explanation

In first ionisation energy, we remove electron from a neutral atom but for second ionisation we have to remove electron from a positive atom, where electron are more tightly bounded due to increased attraction force so second ionisation energy is high relative to first IP.

44. Ionisation enthalpy of lithium is 520 kJ mol^{-1} How much energy in joules must be needed to convert all atoms of lithium to ions present in 7 mg of lithium vapours ?
(a) 74.3 kJ (b) 260 kJ (c) 520 J (d) 780 kJ

520 kJ for 1 mole of Li or 6.022×10^{23} atoms
Molar mass of Li = 79/mol.
Then, moles in 7 mg = 7 mg = 10^{-3} moles
79/mol

$\therefore$ Energy required = $\frac{520}{1}\times10^{3}\times10^{-3}\,J$

=520 J

ELECTRON AFFINITY

45. The second electron gain enthalpy of oxygen is:
(a) –140.9 kJ mol^{-1} (b) –200.7 kJ mol^{-1} (c) +780 kJ mol^{-1} (d) 0

46. Identify the least stable ion amongst the following
(a) Li^- (b) Be^- (c) B^- (d) C^-

Half filled & fully filled configurations are more stable due to exchange energy.

47. Ionization potential of Na would be numerically the same as
(a) Electron affinity of Na^+ (b) Electronegativity of Na^+
(c) Electron affinity of He (d) Ionization potential of Mg

Ionization energy: the energy required to remove an electron from a neutral atom. Electron affinity: the energy change when a neutral atom attracts an electron to become a negative ion.
So, ionization potential of Na would be numerically the same as electron affinity of Na+.

48. The process which requires absorption of energy is
(a) $F \rightarrow F^-$ (b) $Cl \rightarrow Cl^-$ (c) $O \rightarrow O^{2-}$ (d) $H \rightarrow H^-$
Second electron affinity is zero for an element since already added electron repels existing electrons. Hence, it becomes difficult for an atom to take up another electron.

49. Arrange N, O and S in order of decreasing electron affinity.
(a) S > O > N (b) O > S > N (c) N > O > S (d) S > N > O
Electron affinity increases as you add more valence electron. That puts oxygen (O) as having more electron affinity than (N). So, O>N
Electron affinity would typically decrease as you move down the periodic table.
But there is a factor in the second period of elements due to the close distance or the orbital from the nucleus so that repulsion of an electron from each other reduce electron affinity.
So, S>O>N

50. Which of the following gains electrons more easily
(a) X^- (Cl^-, Br^-, I^-) (b) O^- (c) H^- (d) Na

51. Which of the following has the maximum electron affinity?
(a) Bromine (b) Iodine (c) Chlorine (d) Fluorine
Chlorine has maximum electron affinity.

52. For which of the following transitions will the electron gain enthalpy?
(a) Formation of O^- from O (b) formation of O^{2-} from O^-
(c) Formation of O^+ from O (d) None

53. The order of first electron affinity of O, S and Se is
(a) S > O > Se (b) S > Se > O (c) Se = O > S (d) S > O = Se O
has an exceptionally smaller value of electron affinity ($-141 kJmol^{-1}$) due to smaller atomic size than sulphur (weaker electron-electron repulsion in 3p-subshell). It is less than Se and Te also.

54. Which of the following represent(s) the correct order of electron affinities
(a) F >Cl> Br =I (b) C < N = Cl< F (c) N < C < O < F (d) C < Si > P < N
Electron affinity increases across a period and decreases down the group.

55. Which element has the highest electron affinity
(a) F (b) Cl (c) Br (d) I
As an exception to the rule, Cl has higher EA as compared to F. Fluorine has very small atomic size and this makes the fluoride anion unstable due to very high charge/mass ratio. F has no d orbitals so it has small atomic size. F has EA less than that of Cl.

56. In which of the following process energy is liberated
(a) $Cl \rightarrow Cl^+ + e^-$ (b) $HCl \rightarrow H^+ + Cl^-$ (c) $Cl + e \rightarrow Cl^-$ (d) $O^- + e \rightarrow O^{2-}$

57. Which of the following species has the highest electron affinity
(a) F^- (b) O (c) O^- (d) Na
Fluoride and chloride ions have a complete octet configuration. Hence, they are not ready to accept an electron. When adding one more electron to the uninegative oxygen anion, electron-electron repulsion will occur and makes the reaction endothermic.

ELECTRONEGATIVITY

58. Which one has more tendency to form covalent compounds ?
(a) Ba (b) Be (c) Mg (d) Ca
All the given elements are from II-A group. In general, moving from top to bottom in a group the ionic character increases and covalent character decreases. Be is at the top in the group having the highest covalent character. Be forms covalent compounds

with Cl,Br & I. Be has 2 valence electrons. Due to its small size, it has high polarizing power & forms a covalent bond.

59. Fluorine is a better oxidising agent than bromine. It is due to
(a) Small size of fluorine (b) Non-metallic character of fluorine
(c) More electronegativity of fluorine (d) More electron repulsion in fluorine

60. For which pair of atoms is the electronegativity difference the greatest?
(a) B, C (b) Li, I (c) K, Cl (d) Se, S

61. If ionisation energy of an atom is 10 eV& EA is 6.8 eV electronegativity of the species on pauling scale.
(a) 4 (b) 3 (c) 5 (d) 6
Given IE=10 eV and EA=6.8 eV
Mulliken electronegativity which is defined as the average of the ionization energy and electron affinity of an atom (both in eV).

$$EN_{Mulliken} = \frac{10+6.8}{2} = 8.4$$

Mullikens electronegativity value is about 2.8 times as large as Paulings electronegativity value.

$$So, EN_{Pauling} = \frac{8.4}{2.8} = 3$$

62. Electronegativity of F on Mulliken's scale is 11.2, what is the electronegativity on the Pauling's scale?
(a) 1 (b) 2 (c) 3 (d) 4

63. Arrange the following in increasing order of their electronegativities –
(a) P<Si<C<F (b) Si<P<F<C (c) Si<P<C<F (d) P<Si<F<C
Electronegativity increases on moving left to right in a period and decreases from top to bottom in a group.
(N and C) and (Si and P) respectively belongs to (n = 2) and (n =3)
∵ Electronegativity of N > electronegativity of C and electronegativity of P > electronegativity of Si.

64. Mullikan's electronegativity is dependent upon
(a) I.P., E.A. (b) Only I.P. (c) Only E.A. (d) None of these
Mulliken proposed that the arithmetic mean of the first ionization energy and the electron affinity should be a measure of the tendency of an atom to attract electrons.

65. A bond with maximum covalent character between non – metallic elements is formed
(a) Between identical atoms
(b) Between chemically similar atoms
(c) Between atoms of widely different electronegativites
(d) Between atoms of the same size
100% covalent character bond will be formed between identical atoms as there is no electronegativity difference between them.

66. Fluorine is more electronegative than nitrogen. The best explanation is that :
(a) The valence electrons in F are on the average, a little closer to the nucleus than in N
(b) The charge on a F nucleus is +9, while that on N nucleus is +7
(c) The valence electrons in F and N are in different shells and thus their energy are greatly different
(d) Electronegativity increases from left to right in each of the periods

67. The electronegativity's of elements A and B are 1.2 and 3.4 units respectively. The type of bond connecting A and B in compound AB is:
(a) Covalent (b) Ionic (c) Coordinate covalent (d) Polar covalent

$$\%\text{ionic character} = \left[1 - e^{\frac{-(X_A - X_B)^2}{4}}\right] \times 100$$

$$X_A = 1.2, X_B = 3.4$$

$$\%\text{ ionic character} = \left[1 - e^{\frac{-(1.2-3.4)^2}{4}}\right] \times 100$$

$$= 71\%$$

68. Electronegativity and electron affinity of an element A are *X* and *Y* respectively. Hence, ionization potentialof *A* is:
(a) $\frac{X+Y}{2}$ (b) 2 X – Y (c) 2Y – X (d) 2 X + Y

69. Pauling's equation for determining the electronegativity of an element is [X_A, X_B = electronegativity values of elements A and B Δ represents polarity of A – B bond]
(a) $X_A - X_B = 0.208\sqrt{\Delta}$ (b) $X_A + X_B = 0.208\sqrt{\Delta}$
(c) $X_A - X_B = 0.208\,\Delta^2$ (d) $X_A - X_B = \sqrt{\Delta}$

MISCELLANEOUS

70. Among the following, which has the maximum hydration energy?
(a) OH^- (b) NH_4^+ (c) F^- (d) H^+

71. The hydration energy of Mg^{2+} ions is lesser than that of:
(a) Al^{3+} (b) Ba^{2+} (c) Na^+ (d) none of these

72. The order in which the following oxides are arranged according to decreasing basic nature is:
(a) Na_2O , Al_2O_3 ,MgO (b) Al_2O_3 , MgO , Na_2O
(c) MgO , Al_2O_3 , Na_2O (d) Na_2O , MgO , Al_2O_3
An oxide is that chemical compound which has a chemical formula containing at least one oxygen atom and one other element. Oxide is a dianion of oxygen. Oxides can be called as binary compounds formed by the reaction of oxygen with other elements. Oxides are classified as acidic, basic, neutral and amphoteric based on their characteristics.

73. An element X occurs in short period having configuration ns^2 np^1. The formula and nature of its oxide is
(a) XO_3, basic (b) XO_3, acidic (c) X_2O_3, Amphoteric (d) X_2O_3, basic
We are given an element having electronic configuration as ns^2np^1, which means that the valency of element X is 3.
We know that the valency of Oxide ion is 2.
Hence, the formula becomes X_2O_3
From the electronic configuration, it is visible that the element belongs to the 3rd group of periodic table and is a metal.
The oxide formed from a metal is a basic oxide.

74. Which element exists as a solid at 25°C and 1 atmospheric pressure among the following?
(a) P (b) Hg (c) Cl (d) Br
Phosphorus exist as solid at 27°C and 1 atmospheric pressure (m.p. of white phosphorus =44°C)

75. Which has maximum stability
(a) $AsCl_3$ (b) $SbCl_3$ (c) $BiCl_3$ (d) Equal
The inertness of s subshell electrons towards the bond formation is called inert pair effect or it can be said as the inactiveness of electrons present in outermost shell (i.e. ns^2) to get unpaired and involve in bond formation is called inert pair effect.
for example:
1) In 13th group, thallium can exhibit +1 and+3 oxidation states but it is stable in +1 oxidation state only due to inert pair effect
2) In 14th group , lead shows both +2 and +4 oxidation states but it is stable in +2 oxidation state due to inert pair effect.
Due to inert-pair effect
Stability of $As^{3+}<Sb^{3+}<Bi^{3+}$

Level – 2

DEVELOPMENT OF PERIODIC TABLE

1. The outer electronic configuration of Gd (At. No. 64)
(a) $4f^8\,5d^6\,6s^2$ (b) $4f^4\,5d^4\,6s^2$ (c) $4f^7\,5d^1\,6s^2$ (d) None
Half-filled stabled configuration will be more favoured to Gd. Arrange the orbitals according to the increasing energy.
The configuration of Gd is $[Xe]4f^7 5d^1 6s^2$.

2. The outer electronic structure of lawrencium (atomic number 103) is :
(a) Rn $5f^{13}7s^{2}7p^{2}$ (b) Rn $5f^{13}6d^{1}7s^{1}7p^{2}$ (c) Rn $5f^{14}7s^{1}7p^{2}$ (d) Rn $5f^{14}6d^{1}7s^{2}$
Lawencium – 103
Lawrencium is an Actinide present in 5f series of f-block elements.
The outermost configuration = $[Rn]\,5f^{14}\,7s^2\,7p^1$

3. Assuming that elements are formed to complete the seventh period, what would be the atomic number of alkaline earth metal of the eighth period.
(a) 113 (b) 120 (c) 119 (d) 106
7th period ends in 118
hence the alkaline earth metal shall belong to the second group, thus its atomic number would be : 120

4. Pd has exceptional electronic configuration $4d^{10}\,5s^0$. It belongs to
(a) 4^{th} group (b) 6^{th} group (c) 10^{th} group (d) None of these

5. Match list - I with list - II and choose the correct anwswer from the code given below

List - I	List - II
(a) Non metal	a. aurum
(b) Half filled d-orbital	b. cerium
(c) Coinage metal	c. chromium
(d) Lanthanide	d. Iodine

Code is -

	(a)	(b)	(c)	(d)
(a)	d	c	a	b
(b)	a	b	c	d
(c)	d	a	c	b
(d)	d	c	b	a

6. False statement for periodic classification of elements is
(a) The properties of the elements are periodic function of their atomic numbers
(b) No. of nonmetallic elements is less than the no. of metallic elements
(c) First ionization energy of elements is not change continuously with increase of atomic no. in a period

(d) d-subshell is filled by directional electron with increasing atomic no. of transition elements.

7. Element with valence shell-electronic configuration as d^5s^1 is placed in:
(a) IA, *s*-block (b) VIA, *s*-block (c) VIB, *s*-block (d) VIB, *d*-block
$(n-1)d^5ns^1$ is a d-block element and it represents group VIB (transition metals) of d-block.

8. M^{3+} has electronic configuration as [Ar] $3d^{10}\,4s^2$, hence it lies in:
(a) *s*-block (b) *p*-block (c) *d*-block (d) *f*-block
M^{3+} has configuration $[Ar]3d^{10}4s^2$
M^{3+} has lose $3e^-$
M has configuration $[Ar]3d^{10}4s^24p^3$
Hence, the element M belongs to p - block.

9. Which of the following has the electronic configuration [Ar] $3d^5$?
(a) Cr (b) Fe^{3+} (c) Mn (d) V
Cr:$[Ar]3d54s^1$
Mn:$[Ar]3d^54s^2$
V:$[Ar]3d^34s^2$
Fe:$[Ar]3d^64s^2$
Fe^{3+}:$[Ar]3d^5$

10. The atoms of the elements belonging to the same group of the periodic table will have :
(a) the same number of protons
(b) the same number of electrons in the valence-shell
(c) the same number of neutrons
(d) the same number of electrons

11. With respect to oxygen maximum valency is shown by:
(a) halogen family (b) oxygen family
(c) nitrogen family (d) boron family
Due to higher no of valence electron, halogen family will show higher valency as:

Element	Compound with oxygen	Valency
Halogen Family	Cl_2O_7	+7
Oxygen Family	SO_3	+6
Nitrogen Family	P_2O_5	+5
Boron Family	B_2O_3	+3

ATOMIC RADIUS

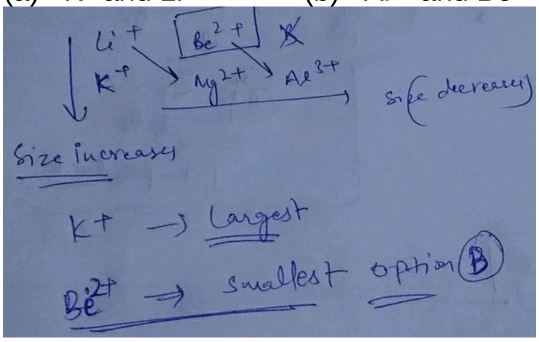
12. Which of the series of elements listed below would have nearly the same atomic radii?
(a) F,Cl,Br,I (b) Na,K,Rb,Cs (c) Li,Be,B,C (d) Fe,Co,Ni,Cu
Fe, Co and Ni are transition metals in the same period where the atomic radii do not vary much. This is due to the extra screening by the 3d electrons which offset the increasing pull by the additional protons going from Fe to Co and Ni on the outer 4s electrons.

13. The radii of F, F^-, O and O^{2-} are in the order of
(a) $O^{2-} > F^- > F > O$ (b) $F^- > O^{2-} > F > O$ (c) $O^{2-} > O > F^- > F$ (d) $O^{2-} > F^- > O > F$
The size of the anion is larger than their parent atom. Also, the more the effective nuclear charge, the lesser is the size. So, the correct order is: $O^{2-}>F^->O>F$.

14. Which is correct in the following
(a) Radius of Cl^- ion is 0.99 Å, while that of Na^+ ion is 1.54 Å
(b) Radius of Cl atom is 0.99 Å while that of Na atom is 1.54 Å
(c) The radius of Cl atom is 0.95 Å while that of Cl^- ion is 0.81 Å
(d) Radius of Na atom is 0.95 Å, while that of Na^+ ion is 1.54 Å
The radius of cation is smaller than the radius of atom. The atomic radius decreases along the period. The radius of anion is greater than the radius of atom. Considering all the factors

15. The correct order of radii is
(a) P<Mg<Al
(b) $Cl^-<S^{2-}<P^{3-}$
(c) $Br^-<F^-<Cl^-$
(d) $Mg^+<Mg^{2+}<Mg$

16. Consider the cations; Li^+, Be^{2+}, Mg^{2+}, K^+ and Al^{3+}. The largest and the smallest ions from this list are respectively
(a) K^+ and Li^+
(b) Al^{3+} and Be^{2+}
(c) Mg^{2+} and Li^+
(d) K^+ and Be^{2+}



17. The set representing the correct order of ionic radius is
(a) $Li^+ > Be^2 > Na^+ > Mg^{2+}$
(b) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$
(c) $Li^+ > Na^+ > Mg^2 > Be^{2+}$
(d) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$
In general, the ionic radius increases on moving from to to bottom in group and decreases on moving from left to right in period So, the correct order is:
$$\underset{0.98\,\text{Å}}{Na^+} > \underset{0.68\,\text{Å}}{Li^+} > \underset{0.65\,\text{Å}}{Mg^{2+}} > \underset{0.62\,\text{Å}}{Be^{2+}}$$

18. Select correct statement (s):
(a) Across a transition series, there is only a small decrease in atomic radius from one element to another due to very small increase in effective nuclear charge
(b) The rate of decrease in the size across the lanthanide series is less than across the first transition series
(c) Both are correct statements
(d) None of the statement is correct
The size of neutral atoms of the *d*-block elements gradually decreases from left to right across a row, due to an increase in the effective nuclear charge (*Z*eff) with increasing atomic number. In addition, the atomic radius increases down a group, just as it does in the *s* and *p* blocks. Because of the *lanthanide contraction*, however, the increase in size between the 3*d* and 4*d* metals is much greater than between the 4*d* and 5*d* metals. Because of the lanthanide contraction, the second- and third-row transition metals are very similar in size.

19. Select correct alternate based on size :
(a) $I^+ < I < I^-$
(b) Fe = Co = Ni
(c) Ni < Cu < Zn
(d) All are correct
As we move along a period radius decreases and as we move down the group, radius increases.
Also as a particular atom accepts electron, its radius increases and as it looses electron its radius decreases so $I^+<I<I^-$.

The size of neutral atoms of the *d*-block elements gradually decreases from left to right across a row, due to an increase in the effective nuclear charge (*Z*eff) with increasing atomic number. In addition, the atomic radius increases down a group, just as it does in the *s* and *p* blocks. Because of the *lanthanide contraction*, however, the increase in size between the 3*d* and 4*d* metals is much greater than between the 4*d* and 5*d* metals. Because of the lanthanide contraction, the second- and third-row transition metals are very similar in size.

20. Which of the following is arranged in order of increasing radius ?
(a) K^+ (aq) < Na^+ (aq) < Li^+(aq)
(b) Na^+ (aq) < K^+ (aq) < Li^+(aq)
(c) K^+ (aq) < Li^+ (aq) < Na^+(aq)
(d) Li^+ (aq) < Na^+ (aq) < K^+(aq)
On moving down the group atomic and ionic radii increases. Li^+, Na^+ and K^+ all belong to group 1. On moving top to bottom they appear in the order.
Li^+, Na^+ and K^+. so, K^+ has the largest ionic radii, followed by Na^+ and Li^+ being the smallest.

21. The correct order of atomic size is
(a) Sc<Ti<V (b) Sc<Y<La (c) Ni>Cu>Zn (d) Ti<Y<Hf
In group 3 size increases as we move down the group due to increase in the number of shell.. So as we move from Sc to La the atomic size increases.

22. Select correct alternate based on size :
(a) $I^+ < I < I^-$
(b) Fe ≈ Co ≈ Ni
(c) Ni < Cu < Zn
(d) All are correct
As we move along a period radius decreases and as we move down the group, radius increases.
Also as a particular atom accepts electron, its radius increases and as it looses electron its radius decreases so $I^+<I<I^-$.
The size of neutral atoms of the *d*-block elements gradually decreases from left to right across a row, due to an increase in the effective nuclear charge (*Z*eff) with increasing atomic number. In addition, the atomic radius increases down a group, just as it does in the *s* and *p* blocks. Because of the *lanthanide contraction*, however, the increase in size between the 3*d* and 4*d* metals is much greater than between the 4*d* and 5*d* metals. Because of the lanthanide contraction, the second- and third-row transition metals are very similar in size.

23. Which of the following is arranged in decreasing order of size ?
(a) $Mg^{2+} > Al^{3+} > O^{2-}$
(b) $O^{2-} > Mg^{2+} > Al^{3+}$
(c) $Al^{3+} > Mg^{2+} > O^{2-}$
(d) $Al^{3+} > O^{2-} > Mg^{2+}$
O^{2-} with atomic no. 8 has 8+2 = 10 electrons
Mg^{2+} with atomic no. 12 has 12-2= 10 electrons
Al^{3+} with atomic no. 13 has 13-3= 10 electrons
Since all these ions have 10 electrons in their shell therefore these are iso-electronic species. The more + the charge, the smaller the ionic radius. Remember that − means adding electrons. These electrons go in the outermost shells. Also, when an atom loses electrons, it clings ever more tightly to the ones it has left, further reducing the ionic radius. therefore the order of ionic radii will be: $O^{2-}>Mg^{2+}>Al^{3+}$

IONIZATION ENERGY

24. Which of the following represent correct order of increasing first I.E for Ca, Ba, S, Se and Ar?
(a) S < Se <Ca< Ba <Ar
(b) Ba <Ca< Se < S <Ar
(c) Ca< Ba < S < Se <Ar
(d) Ca< S < Ba < Se <Ar
Ba<Ca<Se<S<Ar is the correct order of increasing first ionization enthalpy. Ionization enthalpy increases along the period but decreases down the group.
The IE of an element increases as one moves across a period in the periodic table because the electrons are held tighter by the higher effective nuclear charge.

The ionization energy of the elements decreases as one moves down the group because the electrons are held in lower-energy orbitals, away from the nucleus and therefore, are less tightly bound.
Ar has higher IE because it is a noble gas and Ba has the lowest IE as it is in 6 periods and more metallic.

25. Which of the following atoms has least first ionization energy?
(a) Na (b) K (c) Sc (d) Rb
Going down the group, ionization energy decreases and moving across a period ionization energy increases, so scandium has the highest ionization energy.

26. The increasing order of the first ionization enthalpies of the elements B, P, S and F (lowest first) is
(a) F < S < P < B (b) P < S < B < F (c) B < P < S < F (d) B < S < P < F
In general:
(i) First ionization energy ($I.E_1$) or enthalpy, increases along the period from left to right, but due to half-filled configuration of P−atoms, it has higher $I.E_1$ than of S−atom.
(ii) IE_1 decreases as we move down the group from top to bottom.
(iii) More is the size of an atom, less is the IE_1.
Hence, correct order of IE_1 is B<S<P<F.
IE_1 for (in kJ mol^{-1})
B=800
S=999.4
P=1012
F=1680.8.

27. The incorrect statement among the following is
(a) The first ionization enthalpy of Al is less than the first ionization enthalpy of Mg.
(b) The second ionization enthalpy of Mg is greater than the second ionization enthalpy of Na.
(c) The first ionization enthalpy of Na is less than the first ionization enthalpy of Mg.
(d) The third ionization enthalpy of Mg is greater than the third ionization enthalpy of Al.
After removal of an electron sodium acquires stable noble gas configuration. It is difficult to remove electron from stable noble gas configuration species. Therefore, second ionization potential of Mg is less than that of Na.

28. Which represents alkali metals based on $(IE)_1$ and $(IE)_2$ values ?

	$(IE)_1$	$(IE)_2$
(a) *X*	100	110
(b) *Y*	95	120
(c) *Z*	195	500
(d) *M*	200	250

Alkali metals have a very high value of IE_2 because of the octet formation when one electron is removed. Hence IE_2 must be a lot higher than IE_1.

29. Which of the following metal is expected to have the highest third ionization enthalpy?
(a) Cr (Z = 24) (b) V (Z = 23) (c) Mn (Z = 25) (d) Fe (Z = 26)
The electronic configuration of these metals is:
Cr : $[Ar]3d^5 4s^1$
V : $[Ar]3d^3 4s^2$
Mn : $[Ar]3d^5 4s^2$
Fe : $[Ar]3d^6 4s^2$
The third ionization enthalpy of Mn means removal of electron from the stable configuration of $3d^5$. The metal having the highest third ionization enthalpy is Mn.

30. X (g) $\rightarrow X^+$ (g) + e- , ΔH = +720kJ mol^{-1}.
Calculate the amount of energy required to convert 110 mg of '*X*' atom in gaseous state into X^+ion. (Atomic wt. for *X* = 7 g/mol)
(a) 10.4 kJ (b) 12.3 kJ (c) 11.3 kJ (d) 14.5 kJ
For one mole of X (7 gm of X), required energy is 720 kJ/mol.
So, energy required for 110 mg of X =(720/7)×0.110=11.3kJ.

31. Which is the correct order of ionization energies ?
(a) $F^->F>Cl^->Cl$ (b) $F>Cl>Cl^->F^-$ (c) $F^->Cl^->Cl>F$ (d) $F^->Cl^->F>Cl$
The correct order of ionization energies of F^-,Cl^-,F and Cl is $F^-<Cl^-<Cl<F$.
$Cl^->F^-$ because the size of F is less than Cl^-, so the ionization potential will be more for F than Cl^-.
F>Cl because the size of F is less than Cl^-,so the ionization potential will be more for F than Cl. In a group, the ionization energy decreases from top to bottom. The size of the anion is greater than the size of the neutral atom. Due to this, the attraction of the nucleus for the valence electrons is larger in case of neutral atom. Hence, the ionization energy of the neutral atom is larger than that of anion. In case of halogens, the addition of an electron leads to breaking of the noble gas configuration. The ionization energy of fluoride ion is lowest due to electron electron repulsion.

32. Following the transition elements, $(IE)_1$ drops abruptly in Ga, In and Tl. This is due to :
(a) decrease in effective nuclear charge
(b) increases in atomic radius
(c) removal of an electron from the singly occupied *np* orbitals of higher energy than the *ns*-orbitals of
Zn, Cd and Hg
(d) none is correct

33. Which transition involves maximum amount of energy
(a) $X^-_{(g)} \rightarrow X_{(g)} + e$ (b) $X^-_{(g)} \rightarrow X^+ + 2e$
(c) $X^+_{(g)} \rightarrow X^{2+}_{(g)} + e$ (d) $X^{2+}_{(g)} \rightarrow X^{3+}_{(g)} + e$
The transition $X^{2+}_{(g)} \rightarrow X^{3+}_{(g)}$ + e involves maximum amount of energy. It is very difficult to remove an electron from dispositive cation as the effective nuclear charge per electron is maximum.

34. In the following, the element with the highest electropositive character is
(a) Copper (b) Caesium (c) Barium (d) Chromium
Caesium (Cs) is most electropositive in character, as it has the highest tendency to loose an electron from its outermost shell to achieve a stable fully filled configuration.

35. The atomic numbers of Vandium (V), Chromium (Cr), Manganese (Mn) and Iron (Fe) are respectively 23,24,25 and 26. Which one of these may be expected to have the highest second ionization enthalpy
(a) Fe (b) V (c) Cr (d) Mn
The atomic numbers of vanadium, chromium, manganese, and iron are respectively 23, 24, 25, 26. Chromium may be expected to have the highest second ionisation energy. The electronic configuration of chromium is $[Ar]3d^54s^1$
It loses one electron to form $[Ar]3d^54s^0$ in which 3d subshell is half-filled and stable. When another electron is removed, it is removed from half-filled 3d subshell and the stability of 3d subshell is lost. This requires higher energy.

36. Triad - I $[N^{3-},O^-,Na^+]$ Triad - II $[N^+,C^+,O^+]$
Choose the species of lowest IP from triad - I and the species of highest IP from triad - II respectiely
(a) N^{3-} , O^+ (b) Na^+, C^+ (c) N^{3-}, N^- (d) O^- , C^+
In the triad I, N^{3-} has lowest ionization potential due to larger atomic size and lower effective nuclear charge.

In the triad II, O^+ has highest ionization potential due to increased nuclear charge and decreased size.

37. Among the following elements which has the highest ionization energy?
(a) P (b) Si (c) Cl (d) S

38. Amongst the following elements (whose electronic configurations are given below) the one having the highest ionization energy is :
(a) [Ne] $3s^2\ 3p^1$ (b) [Ne] $3s^2\ 3p^3$
(c) [Ne] $3s^2\ 3p^2$ (d) [Ar] $3d^{10}\ 4s^2\ 4p^3$
The IE increases along a period and decreases down the group. Also, IE of 15 is more than group 16 as group 15 has half-filled p subshell giving extra stability.

39. Higher values of ionization energies of the 5*d*-transition elements are consistent with the:
(a) relatively smaller effective nuclear charge
(b) relatively smaller size of their atoms
(c) relatively smaller penetration
(d) all are correct
In the 5d-series of transitions elements, after lanthanum (La), the added electrons go to the next inner 4f orbitals. The 4f electrons have poor shielding effect. As a result, the outermost electrons experience greater nuclear attraction. This leads to higher ionisation energies for the 5d- series of transition elements.

40. The maximum tendency to form the gaseous uni-positive ion is for the element with configuration :
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2$ (b) $1s^2\ 2s^2\ 2p^6\ 3s^1$
(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$ (d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

41. The first ionization enthalpy (in kJ / mol) of Be, B and C atoms are respectively :
(a) 900, 800, 1086 (b) 1086, 800, 900 (c) 800, 900, 1086 (d) 800, 1086, 900

42. Which of the following metals requires radiation of highest frequency to cause emission of electrons?
(a) Na (b) Mg (c) K (d) Ca
Ionisation energies of IIA group elements are greater than IA group.
(Mg,Ca)>(Na,K)
Also, among IIA group elements first ionisation energy gradually decreases as we descend the group.
Mg>Ca>K>Na
Mg requires high-frequency radiation to cause the emission of electrons.

43. The set representing the correct order of first ionization potential is :
(a) K > Na < Li (b) Be > Mg > Ca (c) B > C > N (d) Ge > Si > C

44. Ionization energy of nitrogen is more than oxygen because
(a) Nucleus has more attraction for electrons
(b) Half-filled p-orbital configuration is more stable
(c) nitrogen atom is bigger than oxygen atom
(d) none
Electronic configuration of Oxygen is
$1s^2 2s^2 2p^4$
Electronic configuration of Nitrogen is
$1s^2 2s^2 2p^3$
Nitrogen has a half filled p subshell which is more stable than partially filled p subshell. Oxygen, on the other hand, can readily loose 1 electron and attain more stable half filled e^- configuration.

That is why, ionisation energy of nitrogen is greater and that of oxygen is lesser as oxygen wants to attain more stable e− configuration by loosing its 1 electron.

ELECTRON AFFINITY

45. The increasing order of electron affinity of the electronic configurations of elements is :
(i) $1s^22s^22p^63s^23p^5$ (ii) $1s^22s^22p^3$ (iii) $1s^22s^22p^5$ (iv) $1s^22s^22p^63s^1$
(a) II < IV < III < I (b) I < II < III < IV (c) I < III < II < IV (d) IV < III < II < I
The affinity of electron depends on the outer shell configuration of an element.
From given configurations, III will have highest electron affinity as 2p is more closer to the nucleus and it is deficient of 1 electron to complete its octet. This will be followed by IV where 3s orbital requires 1 electron. II will have less electron affinity than IV because half-filled p orbital is stable. I will have the least electron affinity as 3p is comparatively farther from nucleus.

46. Which of the following processes involves absorption of energy?
(a) $S(g) + e^- \rightarrow S^-(g)$ (b) $S(g) + 2e^- \rightarrow S^{2-}(g)$
(c) $Cl(g) + e^- \rightarrow Cl^-(g)$ (d) none

47. O^{-2} or S^{-2} formation is endothermic because
(a) more stability of O^{-2} or S^{-2}
(b) more energy release due to pairing
(c) electrostatic repulsion outweighs the energy release due to pairing
(d) Both A and B are correct

48. Which of the following statement is correct regarding following process ?
(a) |I.E. of Cl^-|= |E.A. of Cl| (b) |I.E. of Cl|=|E.A of Cl|
(c) |I.E. of Cl^+|=|E.A. of Cl| (d) |I.E. of Cl^+|=|I.E. of Cl|
Ionization Enthalpy is the amount of energy required to remove an electron from the outer most shell of an isolated gaseous atom.
Electron affinity is the amount of energy released when one electron is gained by an ion or atom.
$Cl \longrightarrow Cl^-$ E.A.=-x
$Cl^- \longrightarrow Cl$ I.E.=+x
Thus, magnitude of I.E. and E.A. is same for the above process because same amount of energy is required or released to add an electron in Cl or to remove an electron from Cl^-.
Thus, |I.E. of process(ii)|=|E.A. of process(i)|

49. The increasing order of charge/size ratio of the cations
(a) $Ca^{2+} < Mg^{2+} < Be^{2+} < K^+$ (b) $Mg^{2+} < Be^{2+} < K^+ < Ca^{2+}$
(c) $Be^{2+} < K^+ < Ca^{2+} < Mg^{2+}$ (d) $K^+ < Ca^{2+} < Mg^{2+} < Be^{2+}$
The charge/size ratio of a cation determines its polarizing power.
Higher is the charge and lower is the size, higher will be the charge to size ratios and higher will be the polarizing power of the cation.
The charge of K ion is +1 whereas that of other ions is +2. Thus, K+ has the lowest polarizing power.
For the remaining ions, the decreasing order of the size is $Ca^{2+} > Mg^{2+} > Be^{2+}$.
Hence, the increasing order of polarizing power is $Ca^{2+} < Mg^{2+} < Be^{2+}$.
Hence, the increasing order of the polarizing power of the cationic species is
$K^+ < Ca^{2+} < Mg^{2+} < Be^{2+}$.

50. The electron affinity of chlorine is 3.7eV / atom. How much energy is kcal is released when 2 g of chlorine is completely converted to Cl^- ion in a gaseous state?
(1eV /atom = 23.06 kcal mol^{-1})
(a) 4.8 (b) 2.4 (c) 9.6 (d) None

51. The electron affinities of halogens are F = 322, Cl = 349, Br = 324, I = 295 kJ mol-1. The higher value for Cl as compated to that of F is due to
(a) Weaker electron-electron repulsion in Cl (b) Higher atomic radius of F
(c) Smaller electronegativity of F (d) More vacant p-subshell in Cl
Electron affinity value of Cl is higher than of F as Cl belongs to the 3rd period while F belongs to 2nd period. In Cl, electron-electron repulsion forces are weaker than that of F.

52. Electronic configurations of four elements A, B, C and D are given below:
(1) $1s^2\ 2s^2\ 2p^6$ (2) $1s^2\ 2s^2\ 2p^4$ (3) $1s^2\ 2s^2\ 2p^6\ 3s^1$ (4) $1s^2\ 2s^2\ 2p^5$
Which of the following is the correct order of increasing tendency to gain electron :
(a) 1 < 3 < 2 < 4 (b) 1 < 2 < 3 < 4 (c) 4 < 2 < 3 < 1 (d) 4 < 1 < 2 < 3

53. Which of the following statement is correct ?
(a) Oxygen has more negative electron gain enthalpy than sulphur.
(b) Second electron gain enthalpy of oxygen is positive.
(c) Nitrogen has negative electron gain enthalpy.
(d) Larger is the tendency of an atom to gain an electron, less negative is its electron gain enthalpy

54. Identify the wrong statement in the following :
(a) Amongst isoelectronic species, smaller the positive charge on the cation, smaller is the ionic radius
(b) Amongst isoelectronic species, greater the negative charge on the anion, larger is the ionic radius.
(c) Atomic radius of the elements increases as one moves down the first group of the periodic table.
(d) Atomic radius of the elements decreases as one moves across from left to right in the 2nd period of the periodic table.
Among the isoelectronic species, smaller the positive charge on the cation, higher the ionic radius because the effective nuclear charge increases as the positive charge increases.

ELECTRONEGATIVIY

55. In the compound *M* – O – H, the M – O bond will be broken if :
(a) Δ (E. N.) of *M* and O <Δ (E.N.) of O and H
(b) Δ (E. N.) of *M* and O = Δ (E.N.) of O and H
(c) Δ (E. N.) of *M* and O >Δ (E.N.) of O and H
(d) Cannot be predicated according Δ (E. N.) data
If electronegativity difference of M−O is greater than O−H, then M−O bond will be broken and M+ and OH− ions will form

56. The electronegativity of H, N are 2.1, 3.0 respectively. Calculate percentage ionic character of H – N bond.
(a) 17.24 (b) 8.62 (c) 34.68 (d) None

$$\%\text{Ionic character} = 16(x_A - x_B) + 3.5(x_A - x_B)^2$$
$$x_A = 3.0, x_B = 2.1$$
$$\therefore \%\text{IC} = 16 \times 0.9 + 3.5 \times 0.81$$
$$= 17.24$$

57. $\Delta = X_A - X_B = 2.0$ what is percent ionic character for a covalent molecule A-B
(a) 46 (b) 50 (c) 20 (d) 30

58. X-X bond length is 1.00 and C - C bond length is 1.54 . If electronegativities of X and C are 3.0 and 2.0 respectively the C - X bond length is likely to be
(a) 1.27 (b) 1.28 (c) 1.18 (d) 1.08

$$C - X = r_C + r_X - 0.09(X_C - X_X) = \frac{1.54+1}{2} - 0.09(3-2) = 1.18A^o$$

Thus, C – X bond length is 1.18A°.

59. Fluorine is more electronegative than nitrogen. The best explanation is that :
(a) the valence electrons in F are on the average, a little closer to the nucleus than in N
(b) the charge on a F nucleus is +9, while that on N nucleus is +7
(c) the valence electrons in F and N are in different shells and thus their energy are greatly different
(d) electronegativity increases from left to right in each of the periods

60. The electronegativities of elements A and B are 1.2 and 3.4 units respectively. The type of bond connecting A and B in compound AB is:
(a) covalent (b) ionic
(c) coordinate covalent (d) polar covalent

$$\% \text{ ionic character} = \left[1 - e^{\frac{-(X_A - X_B)^2}{4}}\right] \times 100$$

$X_A = 1.2,\ X_B = 3.4$

$$\% \text{ ionic character} = \left[1 - e^{\frac{-(1.2-3.4)^2}{4}}\right] \times 100$$

$= 71\%$

MISCELLANEOUS

61. In which of the following arrangements, the sequence is not strictly according to the property against it?
(a) $CO_2 < SiO_2 < SnO_2 < PbO_2$: Increasing oxidizing power
(b) $HF < HCl < HBr < HI$: Increasing acid strength
(c) $NH_3 < PH_3 < AsH_3 < SbH_3$: Increasing basic strength
(d) B < C < O <N: Increasing first ionization enthalpy

(a) oxidising power:- ability to release one e^-

but o.p of SnO_2 is $\uparrow$ than PbO_2

$\therefore$ Pb have 4 f which provides less hindrance attraction provided by its protons

$\therefore$ PbO_2 has $\uparrow$ tendency to gain e^- than to oxidise itself than SnO_2

$\therefore C < Si < Pb < Sn \rightarrow$ oxidising power

(b) Basic strength :- $\uparrow$ lone pair $\Rightarrow \uparrow$ basic strength

Group 15

(c) Acidic strength:- More easy dissociation in H^+A^- $\uparrow$ acidic strength.

(d) $B < C < O < N$

atomic number
charge in the order of Boron & oxygen.

62. Which of the following pairs show reverse properties on moving along a period from left to right and from top to down in a group?
(a) Atomic radius and electron affinity
(b) Nuclear charge and electron affinity
(c) Nuclear charge and electronegative character
(d) None of these

Atomic radius decreases from left to right in a period and increases from top to bottom in a group. Similarly, the negative value of electron gain enthalpy decreases in a period and increases in a group. So Atomic radius and electron gain enthalpy show reverse properties on moving along a period from left to right and in the group from top to bottom.

63. Oxidation energy of Li(*s*) to Li^+(*aq*) is least in group IA elements. This is because of :
(a) maximum heat of sublimation of Li(*s*)
(b) maximum heat of hydration of Li^+
(c) less negative heat of hydration of Li^+
(d) maximum ionization energy of Li

64. Which is the most acidic oxide?
(a) Cl_2O (b) Cl_2O_3 (c) Cl_2O_5 (d) Cl_2O_7
Acidic nature depends upon the covalent character. More the covalent character more the acidic nature of the oxides and covalent character depends upon oxidation state of the element.
Now, as the Cl has +7 oxidation state in Cl_2O7 , so it has highest polarising power and highest covalent character so Cl_2O_7 is most covalent , therefore most acidic oxide.

65. Which of the following metals exhibits more than one oxidation state
(a) Na (b) Mg (c) Fe (d) Al
Transition metals have the property of variations in oxidation state, means they can show more than one oxidation state. Among all the given metals, Fe is the transition metal and it can vary its oxidation state from +2 to +7.

66. Among $Al_2O_3,SiO_2<P_2O_3<SO_2$ the correct order of acid strength is
(a) $Al_2O_3 < SiO_2 < P_2O_3 < SO_2$
(b) $SiO_2 < SO_2 < Al_2O_3 < P_2O_3$
(c) $Al_2O_3 < SiO_2 < SO_2 < P_2O_3$
(d) $SO_2 < P_2O_3 < SiO_2 < Al_2O_3$
Acidinc strength depends upon non-metallic character of atom. As we move left to right in a period non-metallic character increases and hence acidity increases.

67. Match list I with list II & then select the correct from the codes given below

List - I	List - II
(a) Increasing atomic size	(a) Cl< O < F
(b) Decreasing atomic radius	(b) B > Be > Li
(c) Increasing electronegativity	(c) Si < Al < Mg
(d) Decreasing effective nuclear charge	(d) N > O > F

Codes

	A	B	C	D
(a)	c	d	a	b
(b)	d	b	c	a
(c)	a	c	b	d
(d)	b	a	d	c

68. Which of the following is the correct statement
(a) Boron is diagonally related to silicon
(b) Elements of third period are known as bridge element
(c) There are sixteen groups and seven periods in extended form of periodic table
(d) Fluorine has higher electron affinity than chlorine

69. Element A,B,C,D belong to the same group. The basic character of their oxides will be in which order if the atomic numbers of A,B,C,D are (Z-x), (Z+2x+2), Z, (Z+x) respectively
(a) A < B > C < D (b) A > B > C > D (c) B > D > C > A (d) B > C > D > A

70. The electronic configuration of four elements L, P, Q and R are given below
$L = 1s^2 , 2s^22p^4$ $Q = 1s^2 , 2s^22p^6 , 3s^23p^5$
$P = 1s^2 , 2s^22p^6 , 3s^1$ $R = 1s^2 , 2s^22p^6 ,3s^2$
The formula of the ionic compounds that can be formed between these elements are
(a) L_2P, RL, PQ, R_2Q
(b) LP, L, PQ, RQ
(c) P_2L, RL, PQ, RQ_2
(d) LP, R_2L, P_2Q, RQ

given, L ($1s^2, 2s^2, 2p^4$) → its valency is 2

P ($1s^2, 2p^6, 3s^1$) → valency → 1

φ ($1s^2, 2s^2, 2p^6, 3s^2 3p^5$) → valency → 1

R ($1s^2, 2s^2 2p^6, 3s^2$) → valency → 2

1) 2P combine with 1L to complete octact, so P_2L

2) 1P with 1φ so, Pφ

3) 2φ with 1R so, Formula is $R\varphi_2$

4) 1R with 1L so, RL

Compound Formula's are :- P_2L, Pφ, $R\varphi_2$, RL

Hence, option (c) is the correct answer.

71. If a = NO_2, b = K_2O, c = ZnO
Arrange the above compounds in decreasing order of there basicity
(a) a, b, c (b) b, c, a (c) c, b, a (d) a, c, b

72. Extent of hydration of Na^+, Mg^{2+}, Al^{3+} is in order :
(a) $Na^+ < Al^{3+} < Mg^{2+}$ (b) $Na^+ < Mg^{2+} < Al^{3+}$
(c) $Al^{3+} < Mg^{2+} < Na^+$ (d) equal
The hydration enthalpies (ie, energies of hydration) of metal ions decreases with increase in ionic radii. Because of the smaller size of Al^{3+} than Mg^{2+} and Na^+, its energy of hydration larger than both these ions.

73. Oxidation energy of Li(*s*) to Li^+(*aq*) is least in group IA elements. This is because of :
(a) maximum heat of sublimation of Li(*s*) (b) maximum heat of hydration of Li^+
(c) less negative heat of hydration of Li^+ (d) maximum ionization energy of Li

74. The screening effect of *d*-electrons is :
(a) equal to the *p*-electrons (b) much more than *p*-electrons
(c) same as *f*-electrons (d) less than *p*-electrons
The screening effect of d- electrons is less than p - electrons. In general, d and f electrons have a poor shielding effect compared to s and p electrons. This is because s and p electrons are close to the nucleus whereas d and f electrons are more diffused (away from the nucleus).

75. In which element shielding effect is not possible?
(a) H (b) Be (c) B (d) N
Shielding effect is the screening of valence electrons by inner electrons. Since H has one electron, it does not have shielding effect.