Chapter 3 Classification of Elements and Periodicity in Properties

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DEVELOPMENT OF PERIODIC TABLE

The elements have been classified into groups for a systematic study of their properties. Various attempts have been made by scientists from the early 1800's. The first classification was made by **Dobereiner** who formulated 'Triads'. It was followed by **Newland's 'Law of octaves'**. The next development that came was **Mendeleev's periodic table** which classified elements on the basis of their atomic masses.

Moseley showed that atomic number is a more fundamental property of an element than its atomic mass. The Mendeleev's periodic law was then modified to a new law called **Modern Periodic law**, according to which *'The physical and chemical properties of the elements are periodic functions of their atomic numbers'*.

Long form of periodic table is based upon the Modern Periodic law. This is also known as Bohr's table as it is based on Bohr's scheme for the arrangement of various electrons around the nucleus.

The horizontal rows of the periodic table are called 'Periods' while the vertical columns are called 'Groups'. There are 7 periods and 18 groups in the periodic table.

Merits of Long Form of Periodic Table

(i) **Positions of Isotopes and Isobars** - Modern periodic table is based on atomic numbers. Therefore, various isotopes of the same element will occupy the same position in the periodic table. Isobars have to be placed at different positions.

(ii) The positions of actinoids and lanthanoids is more clear now because these have been placed in group 3 and due to paucity of space, these are written at the bottom of the periodic table.

NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBER > 100

A systematic nomenclature has been derived to directly name the element from its atomic number using numerical roots for 0 and numbers 1-9. The roots are put together in order of digits which make up the atomic number and 'ium' is added at the end.

Ex : Name of element with atomic number 101 is Unnilunium, 102 is unnilbium, 103 is unniltrium, etc. and their symbols are Unu, Unb, Unt.

Notation for IUPAC Nomenclature of Elements

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	e

ELECTRONIC CONFIGURATION OF ELEMENTS AND PERIODIC TABLE

An element's location in the periodic table reflects the quantum numbers of the last orbital filled.

Electronic Configuration in Periods

(i) The period indicates the value of n for the outermost or valence shell

(ii) The number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.

(iii) There are 2 elements in 1st period; 8 in the 2nd; 8 in the 3rd; 18 in the 4th; 18 in the 5th; 32 in 6th and 7th period is incomplete.

Groupwise Electronic Configuration

Elements in same vertical column or group have similar valence shell electronic configurations, the same number of e^{-1}s in outer orbitals and similar properties.

For ex: all the group 1 elements have ns^1 valence shell electronic configuration.

s-, *p*-, *d*- AND *f*- BLOCK ELEMENTS

Elements can be classified into four blocks: *s*-block, *p*-block, *d*-block and *f*-block depending upon the type of atomic orbitals that being filled with electrons.

The s-block Elements

(i) General electronic configuration is ns^{1-2}

(ii) Group 1 and Group 2 elements are s-block elements because they have ns^1 and ns^2 outermost electronic configuration.

(iii) They are all reactive metals with low ionisation energy. They lose the outermost e^{-1}s readily to form +1 ion (grp 1) or +2 ion (grp 2).
Their compounds are predominently ionic (except Li and Bi).

(iv) Group 1 elements are known as alkali metals because they react with water to form alkali. Group 2 elements are known as alkaline earth metals because their oxides react with water to form alkali and these are found in the soil or earth. The total number of *s*-block elements are 14.

The p-block Elements

(i) General electronic configuration is $ns^2\ np^{1-6}$

(ii) They comprise of elements from group 13 to 18.

(iii) Group 16 elements are called chalcogens while group 17 elements are called halogens.

(iv) Group 18 elements are the noble gases due to completely filled valence shell. As a result, they are less reactive,

(v) The non-metallic character increases as we move from left to right across a period. Down the group metallic character increases.

(vi) The p-block elements together with s-block elements are called Representative elements.

The d-block Elements

(i) General electronic configuration is $(n-1)d^{1-10}\ ns^{0\text{-}2}$

(ii) They comprise of group 3 to 12. They are all metals.

(iii) They mostly form coloured ions, exhibit variable oxidation states paramagnetism and are used as catalysts.

(iv) They form a bridge between chemically active metals of s-block and less active metals of group 13 and 14 and thus are called 'Transition Elements'.

(v) Zn, Cd and Hg though are *d*-block elements but do not known as transition elements because in these elements *d*-orbitals are fully filled.

The f-block Elements

(i) General electronic configuration is $(n-2)f^{1-14}\ (n-1)\ d^{0-1}\ ns^2$

(ii) They comprise of the two rows of elements at the bottom of periodic Table, called the Lanthanoids and Actinoids.

(iii) These two series of elements are called Inner transition elements.

(iv) They are all metals. The chemistry of early actinoids is more complicated than corresponding lanthanoids due to larger number of oxidation state possible for actinoid elements

(v) The elements after uranium are called Transuranium elements.

Metals, Non-Metals and Metalloids

The elements can be divided into Metals and Non-metals.

Metals

(i) They appear on the left hand side of periodic table.

(ii) They are usually solids at room temperature (except Hg which is a liquid at room temperature)

(iii) They have high m.pts and b.pts, are good conductors of heat and electricity and are malleable and ductile.

Non-metals

(i) They are located at the top right hand side of the periodic table.

(ii) They are usually solids or gases at room temperature with low m.pts and b.pts.

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

(iv) They are brittle and are neither malleable nor ductile.

Metalloids

The elements which lie on the borderline between metals and non-metals show properties that are characteristic of both metals and non-metals. They are called semi-metals or metalloids.

PERIODIC TRENDS IN PROPERTIES OF ELEMENTS

Atomic Radius

It is defined as half the distance between the nuclei of two bonded atoms. It refers to both covalent and metallic radius depending on whether the element is metal or non-metal.

Atomic radii decreases across a period because e^{-1}s are being added into same valence shell so that the effective nuclear charge increases as the atomic number increases resulting in increased attraction of e^{-1}s to the nucleus.

In a group, atomic radius increases. This is because down the group, principal quantum number (n) increases and e^- is being added into new shell. As a result valence e^{-1}s are farther from the nucleus. Thus, nuclear attraction decreases and therefore size increases.

Ionic Radius

The removal of an e^- from an atom results in the formation of cation whereas gain of an e^- leads to an anion.

Ionic radii of elements exhibit the same trend as atomic radii.

A cation is smaller than its parent atom because it has fewer e^{-1}s while its nuclear charge remains the same.

The size of anion is larger than parent atom because addition of one or more e^{-1}s results in increased repulsion among e^{-1}s and a decrease in effective nuclear charge.

Isoelectronic species have different radii due to their different nuclear charges. Cation with greater positive charge has smaller radius due to greater effective nuclear charge. Anion with greater negative charge will have larger radius because the net repulsion of the e^{-1}s will outweigh the nuclear charge and the ion expands in size.

Ionization Enthalpy (IE)

It is the amount of energy required to convert gaseous neutral atom into cation, i.e. $X(g) \longrightarrow X^+(g) + e^-$

Ionization energy is always positive because energy is always required to remove e^{-1}s from an atom.

$IE_3 > IE_2 > IE_1$ This is because it is more difficult to remove an e^- from a positively charged species than from a neutral atom. Down the group, atomic size increases and IE decreases. While across a period, atomic size decreases and IE increases.

Factors governing the Ionization energy

(i) **Nuclear charge:** IE increases with increases in nuclear charge.

(ii) **Atomic size:** IE decreases as atomic radius decreases.

(iii) **Penetrating effect of e^{-1}s:** IE increases as penetration effect of e^{-1}s increases. Within the same shell, penetration effect decreases in the order : $s > p > d > f$. Thus, IE to knock out $s - e^-$ will be higher than $p - e^-$ of the same shell.

(iv) **Shielding or screening effect of inner shell e^{-1}s:** As shielding or screening effect of inner e^{-1}s. increases, IE decreases.

(v) **Effect of exactly half-filled or completely filled orbitals:** More stable the electronic configuration, greater is the IE. This is because of extra stability associated with exactly half-filled or completely filled orbitals due to which more energy is required to remove the e^-. This is the reason why IE of N is more than that of O.

(vi) Noble gases, being stable with completely filled orbital have the highest IE in their respective periods.

Electron Gain Enthalpy (EGE)

It is defined as the energy released when a neutral isolated gaseous atom accepts an extra e^- to form gaseous negative ion, i.e., anion,

i.e. $X(g) + e^- \longrightarrow X^-(g)$

After the addition of one e^-, the atom becomes negatively charged and 2nd e^- is to be added to a negatively charged ion. But the addition of 2nd e^- is opposed by electrostatic repulsion and hence energy has to be supplied for addition of 2nd e^-. Thus, 2nd electron gain enthalpy of an element is positive.

Factors on which EGE depends

(i) **Atomic size:** As size increases EGE becomes less negative.

(ii) **Nuclear charge:** As nuclear charge increases, EGE becomes more negative.

(iii) **Electronic configuration:** Elements having exactly half-filled or completely filled orbitals are very stable and have large positive electron gain enthalpy because they do not accept additional e^- easily.

Variation of EGE in periodic table

(i) Down the group, atomic size increases, EGE becomes less negative. Across a period, atomic size decreases, nuclear charge increases and EGE becomes more negative.

(ii) Halogens have very high negative EGE because they attain stable noble gas electronic configuration by accepting an e^-.

(iii) Noble gases have large positive EGE because the e^- has to enter the next higher shell leading to a very unstable electronic configuration.

(iv) EGE of O or F is less negative than the succeeding element S or Cl. This is because when an e^- is added to O or F, the added e^- goes to smaller n = 2 quantum level and suffers significant repulsion from other e^{-1}s in this level. For n = 3 level (S or Cl), added e^- occupies a larger region of space and $e^- - e^-$ repulsion is much less.

Electronegativity

It is the tendency of an atom of the element to attract the shared pair of e^{-1}s towards itself in a covalent bond. It is represented by X.

The electronegativity of any given element is not constant but depends on the following factors:

(i) **State of hybridization:** *sp*-hybridized carbon is more electronegative than sp^2 hybridized which in turn is more electronegative than sp^3 hybridized carbon.

(iii) **O.S. of the element:** As O.S. of the element increases, electronegativity increases.

(iv) **Nature of substituents attached to the atom: For ex:** C-atom in CF_3I is more electronegative than in CH_3I.

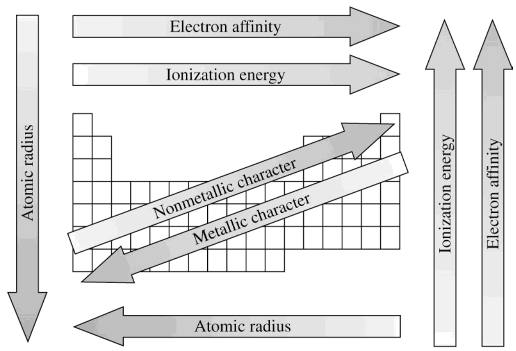
Variation of electronegativity in periodic table

(i) Down the group, atomic radius increases, electronegativity decreases

Across a period, atomic radius and nuclear charge increases, electronegativity increases.

(ii) F is the most electronegative element and caesium is the least electronegative element.

SUMMARY OF TRENDS IN PERIODIC PROPERTIES OF ELEMENTS.



PERIODIC TRENDS AND CHEMICAL REACTIVITY

Chemical reactivity is highest at the two extremes of a period and is lowest in the centre.

Nature of Oxides

If difference of the two electronegativities ($X_O–X_A$) is 2.3 or more then oxide will be basic in nature. Similarly if value of $X_O–X_A$ is slightly lower than 2.3 then oxide will be amphoteric and if value of $X_O–X_A$ is highly lower than 2.3 then oxide will be of acidic nature.

Nature of Hydroxides

According to **Gallis**, if electronegativity of A in a hydroxide (AOH) is more than 1.7 then it will be acidic in nature whereas it will be basic in nature if electronegativity is less than 1.7

Note : Compounds formed from two nonmetals are called binary compounds. Name of more electronegative element is written at the end and 'ide' is suffixed to it. The name of less electronegative element is written before the name of more electronegative element of the formula.

Periodicity of Valence or Oxidation State

O.S. of an element in a particular compound is defined as the charge acquired by its atom on the basis of electronegativity of other atoms in the molecule. The valence of representative elements is usually equal to no. of e^{-1}s in outermost shell.

Variation of oxidation state in periodic table

Across a period, no. of valence e^{-1}s increases from 1 to 8. The valence of elements first increases from 1 to 4 and then decreases to zero.

Down the group, no. of valence e^{-1}s remain the same, and therefore, all elements in a group exhibit the same valence.

Noble gases are zerovalent, i.e., their valence is zero because they are chemically inert.

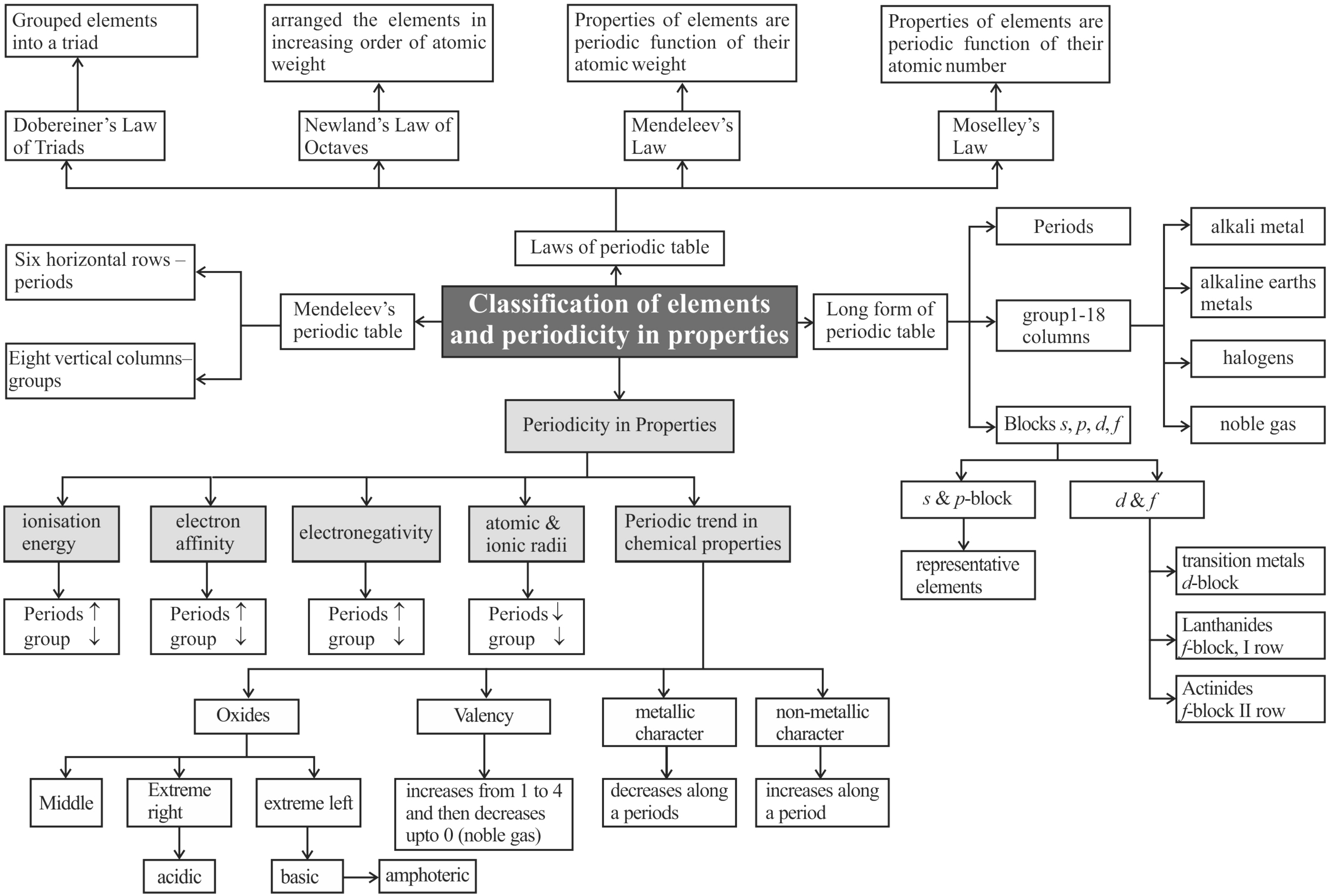
ANOMALOUS PROPERTIES OF SECOND PERIOD ELEMENTS

The first element of each of the groups 1 (Li) and 2(Be) and groups 13-17 (B to F) differs in many respects from other members of its group. Moreover, the behaviour of Li and Be is more similar with the 2nd element of following group i.e. Mg and Al. This sort of similarity is referred to as diagonal relationship in periodic properties.

The anomalous behaviour of these elements is attributed to their (i) small size (ii) large charge / radius ratio (iii) high electronegativity (iv) non-availability of orbitals due to which they cannot expand their covalence beyond 4.

For example : Because of smaller size and higher electronegativity first member of p-block elements displays greater ability to form pπ-pπ multiple bonds to itself ($C=C$, $C\equiv C$, $N\equiv N$) and to other 2nd period elements ($C = O$, $C = N$, $C \equiv N$, $N = O$) compared to subsequent members of same group.

CONCEPT MAP


Grouped elements into a triad
arranged the elements in increasing order of atomic weight
Properties of elements are periodic function of their atomic weight
Properties of elements are periodic function of their atomic number
Dobereiner's Law of Triads
Newland's Law of Octaves
Mendeleev's Law
Moselley's Law
Laws of periodic table
Six horizontal rows – periods
Eight vertical columns– groups
Mendeleev's periodic table
Classification of elements and periodicity in properties
Long form of periodic table
Periods
group1-18 columns
Blocks s, p, d, f
alkali metal
alkaline earths metals
halogens
noble gas
s & p-block
representative elements
d & f
transition metals d-block
Lanthanides f-block, I row
Actinides f-block II row
Periodicity in Properties
ionisation energy
Periods ↑ group ↓
electron affinity
Periods ↑ group ↓
electronegativity
Periods ↑ group ↓
atomic & ionic radii
Periods ↓ group ↓
Periodic trend in chemical properties
Oxides
Middle
Extreme right
acidic
extreme left
basic
amphoteric
Valency
increases from 1 to 4 and then decreases upto 0 (noble gas)
metallic character
decreases along a periods
non-metallic character
increases along a period

EXERCISE - 1

Conceptual Questions

1. Which group of periodic table contains no metal:
(a) IA (b) IIIA
(c) VIIA (d) VIII

2. Which of the following is the atomic number of metal?
(a) 32 (b) 34 (c) 36 (d) 38

3. Which one of these is basic ?
(a) SiO_2 (b) SO_2
(c) CO_2 (d) Na_2O

4. Most acidic oxide is :
(a) Na_2O (b) ZnO
(c) MgO (d) P_2O_5

5. Which of the following metals shows allotropy?
(a) Ca (b) Pb
(c) Sn (d) K

6. The electronic configuration of an element is $1s^2 2s^2 2p^6 3s^2 3p^3$. What is the atomic number of the element, which is just below the above element in the periodic table?
(a) 33 (b) 34
(c) 36 (d) 49

7. An atom has electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$, you will place it in which group?
(a) Fifth (b) Fifteenth
(c) Second (d) Third

8. Which one of the following is an amphoteric oxide ?
(a) Na_2O (b) SO_2
(c) B_2O_3 (d) ZnO

9. The screening effect of 'd' electrons is
(a) Much less than s- electrons
(b) Much more than s- electrons
(c) Equal to s- electrons
(d) Equal to p- electrons

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10. The statement that is not correct for the periodic classification of element is
(a) the properties of elements are the periodic functions of their atomic numbers.
(b) non-metallic elements are lesser in number than metallic elements.
(c) the first ionisation energies of elements along a period do not vary in a regular manner with increase in atomic number.
(d) for transition elements the *d*-subshells are filled with electrons monotonically with increase in atomic number.

11. The elements in which 4*f* orbitals are progressively filled up are called as
(a) Actinoids (b) Transition elements
(c) Lanthanoids (d) Halogens

12. Who developed long form of the periodic table?
(a) Lothar Meyer (b) Neils Bohr
(c) Mendeleev (d) Moseley

13. 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

14. The statement that is not true for the long form of the periodic table is
(a) it reflects the sequence of filling electrons in the order of sub-energy levels *s*, *p*, *d* and *f*.
(b) it helps to predict the stable valence states of the elements
(c) it reflects trends in physical and chemical properties of the elements
(d) it helps to predict the relative ionicity of the bond between any two elements.

15. Among the following elements, the one having the highest ionisation energy is.
(a) $[Ar]3d^{10}, 4s^2\,4p^3$ (b) $[Ne]3s^2\,3p^1$
(c) $[Ne]3s^2\,3p^3$ (d) $[Ne]3s^2\,3p^2$

16. Polarisation power of a cation increases when
(a) size of the cation increases
(b) charge of the cation increases
(c) charge of the cation decreases
(d) it has no relation with its charge or size

17. Which one of the following is not a transition metal ?
(a) Mn (b) Cr
(c) Cu (d) Cd

18. Which is chemically most active non-metal ?
(a) S (b) O
(c) F (d) N

19. Which of the following is non-metallic ?
(a) B (b) Be
(c) Mg (d) Al

20. The only non-metal which is liquid at ordinary temperature is
(a) Hg (b) Br_2
(c) NH_3 (d) None of the above

21. Amphoteric-oxide combinations are in
(a) ZnO, K_2O, SO_3 (b) (b) ZnO, P_2O_5, Cl_2O_7
(c) SnO_2, Al_2O_3, ZnO (d) PbO_2, SnO_2, SO_3

22. The correct order of ionization energies is
(a) $Zn < Cd < Hg$ (b) $Hg < Cd < Zn$
(c) $Ar > Ne > He$ (d) $Cs < Rb < Na$

23. Which one of the following has largest size?
(a) Al (b) Al^{3+}
(c) Al^+ (d) Al^{2+}

24. The ionisation potential order for which set is correct?
(a) $Cs < Li < K$ (b) $Cs > Li > B$
(c) $Li > K > Cs$ (d) $B > Li > K$

25. Which one has least ionisation potential?
(a) Ne (b) N
(c) O (d) F

26. Which one of the following is smallest in size ?
(a) N^{3-} (b) O^{2-}
(c) Na^+ (d) F^-

27. The ionization energy of nitrogen atom is more than that of oxygen atom because of
(a) greater attraction of electrons by the nucleus.
(b) smaller size of nitrogen atom.
(c) more penetrating effect
(d) due to half filled p orbital

28. Sequence of acidic character is
(a) $N_2O_5 > SO_2 > CO > CO_2$ (b) $N_2O_5 > SO_2 > CO_2 > CO$
(c) $SO_2 > CO_2 > CO > N_2O_5$ (d) $SO_2 > N_2O_5 > CO > CO_2$

29. Which is a metalloid ?
(a) Manganese (b) Phosphorus
(c) Oxygen (d) Arsenic

30. Na^+, Mg^{2+}, Al^{3+} and Si^{4+} ions are isoelectronic. The value of ionic radii of these ions would be in the order :
(a) $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$
(b) $Na^+ < Mg^{2+} < Al^{3+} < Si^{4+}$
(c) $Na^+ > Mg^{2+} > Al^{3+} < Si^{4+}$
(d) $Na^+ < Mg^{2+} > Al^{3+} > Si^{4+}$

31. Maximum ionisation potential is of :
(a) Ca (b) Na (c) Be (d) Mg

32. Correct order of first IP among following elements Be, B, C, N, O is
(a) B < Be < C < O < N (b) B < Be < C < N < O
(c) Be < B < C < N < O (d) Be < B < C < O < N

33. Which one of the following ions has the highest value of ionic radius ?
(a) O^{2-} (b) B^{3+} (c) Li^+ (d) F^-

34. An element having electronic configuration $1s^2, 2s^2 2p^6, 3s^2 3p^1$ will form
(a) neutral oxide (b) acidic oxide
(c) basic oxide (d) amphoteric oxide

35. The first ionisation potential in electron volts of nitrogen and oxygen atoms are respectively given by
(a) 14.6, 13.6 (b) 13.6, 14.6
(c) 13.6, 13.6 (d) 14.6, 14.6

36. The correct order of radii is
(a) N < Be < B (b) $F^- < O^{2-} < N^{3-}$
(c) N < Li < K (d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

37. First ionization potential will be maximum for
(a) uranium (b) hydrogen
(c) lithium (d) iron

38. Chloride ion and potassium ion are isoelectronic. Then
(a) their sizes are same
(b) Cl^- ion is bigger than K^+ ion
(c) K^+ ion is relatively bigger
(d) their sizes depend on either cation or anion

39. Ionic radii of
(a) $Ti^{4+} < Mn^{2+}$ (b) $^{35}Cl^- < {}^{37}Cl^-$
(c) $K^+ > Cl^{-1}$ (d) $P^{3+} > P^{5+}$

40. When an electron is removed from an atom, its energy
(a) increases (b) decreases
(c) remains the same (d) none of these

41. Which of the following has minimum melting point?
(a) CsF (b) HCl (c) HF (d) LiF

42. The correct order of ionization energy for carbon, nitrogen and oxygen atoms is:
(a) C > N > O (b) C > N < O
(c) C < N > O (d) C < N < O

43. Which of the following pairs of atomic numbers represents elements belonging to the same group?
(a) 11 and 20 (b) 12 and 30
(c) 13 and 31 (d) 14 and 33

44. The correct order according to size is
(a) $O > O^- > O^{2-}$ (b) $O^- > O^{2-} > O$
(c) $O^{2-} > O^- > O$ (d) $O > O^{2-} > O^-$

45. Which of the following element has maximum, first ionisation potential ?
(a) V (b) Ti (c) Cr (d) Mn

46. The outer electronic configuration of transition elements is
(a) $(n-1)\, s^2\, nd^{1-5}$
(b) $(n+1)\, s^2\, nd^{1-5}$
(c) $(n-1)\, s^2\, p^6\, (n-1)\, d^{1-10}, ns^{0-2}$
(d) $ns^2\, (n+1)\, d^{1-10}$

47. If 19 is the atomic number of an element, then this element will be
(a) a metal with + 3 oxidation state
(b) a metal with + 1 oxidation state
(c) an inert gas
(d) a metal with – 3 oxidation state

48. Highest energy will be absorbed to eject out the electron in the configuration
(a) $1s^2\, 2s^2\, 2p^1$ (b) $1s^2\, 2s^2\, 2p^3$
(c) $1s^2\, 2s^2\, 2p^2$ (d) $1s^2\, 2s^2\, 2p^4$

49. In which of the following process highest energy is absorbed ?
(a) $Cu \rightarrow Cu^+$ (b) $Br \rightarrow Br^-$
(c) $I \rightarrow I^-$ (d) $Li \rightarrow Li^+$

50. Which of the following gaseous atoms has highest value of IE ?
(a) P (b) Si
(c) Mg (d) Al

51. Which ionisation potential *(IP)* in the following equations involves the greatest amount of energy ?
(a) $Na \rightarrow Na^+ + e^-$ (b) $K^+ \rightarrow K^{2+} + e^-$
(c) $C^{2+} \rightarrow C^{3+} + e^-$ (d) $Ca^+ \rightarrow Ca^{2+} + e^-$

52. Arrange S, P, As in order of increasing ionisation energy
(a) S < P < As (b) P < S < As
(c) As < S < P (d) As < P < S

53. Among the following options, the sequence of increasing first ionisation potential will be
(a) B < C < N (b) B > C > N
(c) C < B < N (d) N > C > B

54. As per the modern periodic law, the physical and chemical properties of elements are periodic functions of their
(a) Atomic volume
(b) Electronic configuration
(c) Atomic weight
(d) Atomic size

55. Eka-aluminium and EKa-silicon are known as
(a) Gallium and Germanium
(b) Aluminium and Silicon
(c) Iron and Sulphur
(d) Neutron and Magnesium

56. The correct order of reactivity of halogens is
(a) $F > Cl > Br > I$ (b) $F < Cl > Br < I$
(c) $F < Cl < Br < I$ (d) $F < Cl < Br > I$

57. Which one of the following represents the electronic configuration of the most electropositive element ?
(a) [He] $2s^1$ (b) [Xe] $6s^1$
(c) [He] $2s^2$ (d) [Xe] $6s^2$

58. A group 16 element exists in the monoatomic state in the metallic state. It also exists in two crystalline forms. The metal is
(a) S (b) Po (c) Se (d) Te

59. Electron affinity is positive when :
(a) O changes into O^-
(b) O^- changes into O^{2-}
(c) O changes into O^+
(d) electron affinity is always negative

60. Electron affinity is maximum for
(a) Cl (b) F (c) Br (d) I

61. Pauling's electronegativity values for elements are useful in predicting
(a) polarity of the molecules
(b) position in the E.M.F. series
(c) coordination number
(d) dipole moments.

62. Which of the following is most electronegative?
(a) Lead (b) Silicon
(c) Carbon (d) Tin

63. Variable valency is generally exhibited by
(a) representative elements (b) transition elements
(c) non-metallic elements (d) metallic elements

64. On going from right to left in a period in the periodic table, the electronegativity of the elements
(a) increases
(b) decreases
(c) remains unchanged
(d) decreases first then increases

65. Which one of the following has the highest electronegativity?
(a) Br (b) Cl (c) P (d) Si

66. Fluorine, chlorine, bromine and iodine are placed in the same group 17 of the periodic table because :
(a) they are nonmetals
(b) they are electronegative
(c) their atoms are generally univalent
(d) they have 7 electrons in the outer-most shell of their atom

67. Which is the correct order of electronegativity ?
(a) $F > N < O > C$ (b) $F > N > O > C$
(c) $F > N > O < C$ (d) $F < N < O = C$

68. The electron affinity for the inert gases is
(a) zero (b) high
(c) negative (d) positive

69. Which of the following species has the highest electron affinity ?
(a) F (b) O (c) O^- (d) Na^+

70. An atom with high electronegativity has
(a) large size
(b) high ionisation potential
(c) low electron affinity
(d) low ionisation potential

71. The largest size of the ion is :
(a) Cl^- (b) Ca^{++} (c) K^+ (d) S^{--}

72. Which of the following is the most electronegative?
(a) F (b) He (c) Ne (d) Na

73. The outermost electronic configuration of the most electronegative element is
(a) $ns^2\ np^3$ (b) $ns^2\ np^4$ (c) $ns^2\ np^5$ (d) $ns^2\ np^6$

74. Which of the following sequence correctly represents the decreasing acidic nature of oxides ?
(a) $Li_2O > BeO > B_2O_3 > CO_2 > N_2O_3$
(b) $N_2O_3 > CO_2 > B_2O_3 > BeO > Li_2O$
(c) $CO_2 > N_2O_3 > B_2O_3 > BeO > Li_2O$
(d) $B_2O_3 > CO_2 > N_2O_3 > Li_2O > BeO$

75. Which transition involves maximum amount of energy?
(a) $M^-(g) \longrightarrow M(g) + e$
(b) $M^-(g) \longrightarrow M^+(g) + 2e$
(c) $M^+(g) \longrightarrow M^{2+}(g) + e$
(d) $M^{2+}(g) \longrightarrow M^{3+}(g) + e$

EXERCISE - 2 AAJ KA TOPPER
Applied Questions

1. Which of the following can not be isoelectronic?
(a) two different cations (b) two different anions
(c) cation and anion (d) two different atoms

2. The species with a radius less than that of Ne is
(a) Mg^{2+} (b) F^- (c) O^{2-} (d) K^+

3. The correct order of acidic strength :
(a) $Cl_2O_7 > SO_2 > P_4O_{10}$ (b) $K_2O > CaO > MgO$
(c) $CO_2 > N_2O_5 > SO_3$ (d) $Na_2O > MgO > Al_2O_3$

4. Electron affinity of X would be equal to
(a) electron affinity of X^-
(b) ionization energy of X
(c) ionization energy of X^- with sign reversed
(d) none of these

5. Which group is called buffer group of the periodic table ?
(a) I (b) VII
(c) VIII (d) Zero

6. The pair of elements having approximately equal ionisation potential is
(a) Al, Ga (b) Al, Si
(c) Al, Mg (d) Al, B

7. Which electronic configuration of an element has abnormally high difference between second and third ionization energy?
(a) $1s^2, 2s^2, 2p^6, 3s^1$ (b) $1s^2, 2s^2, 2p^6, 3s^1 3p^1$
(c) $1s^2, 2s^2, 2p^6, 3s^2 3p^2$ (d) $1s^2, 2s^2, 2p^6, 3s^2$

8. Which of the following order is wrong?
(a) $NH_3 < PH_3 < AsH_3$ – Acidic
(b) $Li < Be < B < C$ – First IP
(c) $Al_2O_3 < MgO < Na_2O < K_2O$ – Basic
(d) $Li^+ < Na^+ < K^+ < Cs^+$ – Ionic radius

9. For which of the following processes, enthalpy change is positive
(a) $F_{(g)} + e^- \rightarrow F^-_{(g)}$ (b) $Cl_{(g)} + e^- \rightarrow Cl^-_{(g)}$
(c) $O_{(g)} + 2e^- \rightarrow O^{2-}_{(g)}$ (d) $H_{(g)} + e^- \rightarrow H^-_{(g)}$

10. Arrange the elements with the following electronic configurations in increasing order of electron affinity
(i) $1s^2 s^2 2p^5$ (ii) $1s^2 2s^2 2p^4$
(iii) $1s^2 2s^2 2p^6 3s^2 3p^4$ (iv) $1s^2 2s^2 2p^6 3s^2 3p^5$
(a) (ii) < (iii) < (i) < (iv) (b) (iii) < (ii) < (iv) < (i)
(c) (iii) < (ii) < (i) < (iv) (d) (ii) < (iii) < (iv) < (i)

11. Among Al_2O_3, SiO_2, P_2O_3 and SO_2 the correct order of acid strength is
(a) $Al_2O_3 < SiO_2 < SO_2 < P_2O_3$
(b) $SiO_2 < SO_2 < Al_2O_3 < P_2O_3$
(c) $SO_2 < P_2O_3 < SiO_2 < Al_2O_3$
(d) $Al_2O_3 < SiO_2 < P_2O_3 < SO_2$

12. The formation of the oxide ion $O^{2-}_{(g)}$ requires first an exothermic and then an endothermic step as shown below :

$O(g) + e^- = O^-(g)\ \Delta H^\circ = -142\ kJmol^{-1}$

$O^-(g) + e^- = O^{2-}(g)\ \Delta H^\circ = 844\ kJmol^{-1}$

This is because
(a) O^- ion will tend to resist the addition of another electron
(b) Oxygen has high electron affinity
(c) Oxygen is more elecronegative
(d) O^- ion has comparatively larger size than oxygen atom

13. In which of the following arrangements, the order is NOT according to the property indicated against it?
(a) $Li < Na < K < Rb$:
Increasing metallic radius
(b) $I < Br < F < Cl$:
Increasing electron gain enthalpy (with negative sign)
(c) $B < C < N < O$
Increasing first ionization enthalpy
(d) $Al^{3+} < Mg^{2+} < Na^+ < F^-$
Increasing ionic size

14. Quite a large jump between the values of second and third ionization potentials of an atom would correspond to the electronic configuration
(a) $1s^2 2s^2 2p^6$ (b) $1s^2 2s^2 2p^6 3s^2$
(c) $1s^2 2s^2 2p^6 3s^2 3p^1$ (d) $1s^2 2s^2 2p^6 3s^2 3p^2$

15. Consider the following changes
$A \rightarrow A^+ + e^- : E_1$ and $A^+ \rightarrow A^{2+} + e^- : E_2$
The energy required to pull out the two electrons are E_1 and E_2 respectively. The correct relationship between two energies would be
(a) $E_1 < E_2$ (b) $E_1 = E_2$
(c) $E_1 > E_2$ (d) $E_1 \geq E_2$

16. The incorrect statement among the following is
(a) The first ionization potential of Al is less than the first ionization potential of Mg
(b) The second ionization potential of Mg is greater than the second ionization potential of Na
(c) The first ionization potential of Na is less than the first ionization potential of Mg
(d) The third ionization potential of Mg is greater than the third ionization potential of Al.

17. Successive addition of electronic shells in case of elements of 17th group causes a increase in
(a) electronegativity
(b) ionization energy
(c) ease of formation of unipositive ion
(d) oxidizing power

18. The second ionization potential of an element M is the energy required to
(a) remove one mole of electrons from one mole of gaseous cations of the element
(b) remove one mole of electrons from one mole of gaseous anions
(c) remove one mole of electrons from one mole of monovalent gaseous cations of the element
(d) remove 2 moles of electrons from one mole of gaseous atoms

19. Which of the following ions has the most negative value of enthalpy of interaction with water?
(a) NH_4^+ (b) OH^- (c) H^+ (d) F^-

20. *s*-electrons of the valence shell of some elements show reluctance in bond formation. Such elements are --- and belong to --- :
(a) lighter , *s*-block (b) heavier , *d*-block
(c) heavier , *f*-block (d) heavier , *p*-block

21. Which of the following cations acts as an oxidizing agent?
(a) Ga^{3+} (b) In^{3+} (c) Tl^{1+} (d) Tl^{3+}

22. The first ionization potential 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) $Na > Mg > Al < Si$

23. Correct order of polarising power is
(a) $Cs^+ < K^+ < Mg^{2+} < Al^{3+}$ (b) $K^+ < Cs^+ < Mg^{2+} < Al^{3+}$
(c) $Cs^+ < K^+ < Al^{3+} < Mg^{2+}$ (d) $K^+ < Cs^+ < Al^{3+} < Mg^{2+}$

24. In general, the ionization potentials of elements decreases as one proceeds in the periodic table
(a) bottom → top and right → left
(b) top → bottom and right → left
(c) bottom → top and left → right
(d) top → bottom and left → right

25. Which of the following properties of elements does not exhibit the periodicity ?
(a) Ionization potential (b) Electronegativity
(c) Electronic configuration (d) Neutron to proton ratio

26. Identify the correct order of the size of the following:
(a) $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$
(b) $Ar < Ca^{2+} < K^+ < Cl^- < S^{2-}$
(c) $Ca^{2+} < Ar < K^+ < Cl^- < S^{2-}$
(d) $Ca^{2+} < K^+ < Ar < S^{2-} < Cl^-$

27. Which one of the following ionic species has the greatest proton affinity to form stable compound?
(a) NH_2^- (b) F^-
(c) I^- (d) HS^-

28. The stability of + 1 oxidation state increases in the sequence:
(a) Tl < In < Ga < Al (b) In < Tl < Ga < Al
(c) Ga < In < Al < Tl (d) Al < Ga < In < Tl

29. Amongst the elements with following electronic configurations, which one of them may have the highest ionization energy?
(a) Ne$[3s^2 3p^2]$ (b) Ar $[3d^{10} 4s^2 4p^3]$
(c) Ne $[3s^2 3p^1]$ (d) Ne $[3s^2 3p^3]$

30. Among the elements Ca, Mg, P and Cl, the order of increasing atomic radii is :
(a) Ca < Mg < P < Cl (b) Mg < Ca < Cl < P
(c) Cl < P < Mg < Ca (d) P < Cl < Ca < Mg

31. What is the value of electron gain enthalpy of Na^+ if IE_1 of Na = 5.1 eV ?
(a) –5.1 eV (b) –10.2 eV
(c) +2.55 eV (d) +10.2 eV

32. Following statements regarding the periodic trends of chemical reactivity of the alkali metals and the halogens are given. Which of these statements gives the correct picture?
(a) Chemical reactivity increases with increase in atomic number down the group in both the alkali metals and halogens
(b) In alkali metals the reactivity increases but in the halogens it decreases with increase in atomic number down the group
(c) The reactivity decreases in the alkali metals but increases in the halogens with increase in atomic number down the group
(d) In both the alkali metals and the halogens the chemical reactivity decreases with increase in atomic number down the group

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

34. The correct sequence which shows decreasing order of the ionic radii of the elements is
(a) $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$
(b) $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$
(c) $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$
(d) $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

35. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I, having atomic number 9, 17, 35 and 53 respectively, is :
(a) F > Cl > Br > I (b) Cl > F > Br > I
(c) Br > Cl > I > F (d) I > Br > Cl > F

36. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ?
(a) Ca < S < Ba < Se < Ar (b) S < Se < Ca < Ba < Ar
(c) Ba < Ca < Se < S < Ar (d) Ca < Ba < S < Se < Ar

37. An element of atomic weight 40 has 2, 8, 8, 2 as the electronic configuration. Which one of the following statements regarding this element is not correct
(a) it belongs to II group of the periodic table
(b) it has 20 neutrons
(c) the formula of its oxide is MO_2
(d) it belongs to 4th period of the periodic table

38. Which of the following is not the correct order for the stated property ?
(a) Ba > Sr > Mg; atomic radius
(b) F > O > N : first ionization enthalpy
(c) Cl > F > I; electron affinity
(d) O > Se > Te; electronegativity

39. Which of the following sets has strongest tendency to form anions ?
(a) Ga, In, Tl (b) Na, Mg, Al
(c) N, O, F (d) V, Cr, Mn

40. One of the characteristic properties of non-metals is that they
(a) Are reducing agents
(b) Form basic oxides
(c) Form cations by electron gain
(d) Are electronegative

41. In which of the following electronic configuration an atom has the lowest ionisation enthalpy?
(a) $1s^2 2s^2 2p^3$ (b) $1s^2 2s^2 2p^5 3s^1$
(c) $1s^2 2s^2 2p^6$ (d) $1s^2 2s^2 2p^5$

42. Which of the following represents the correct order of increasing electron gain enthalpy with negative sign for the elements O, S, F and Cl ?
(a) Cl < F < O < S (b) O < S < F < Cl
(c) F < S < O < Cl (d) S < O < CI < F

43. Which is the correct order of ionic sizes (At. No. : Ce = 58, Sn = 50, Yb = 70 and Lu = 71) ?
(a) Ce > Sn > Yb > Lu (b) Sn > Ce > Yb > Lu
(c) Lu > Yb > Sn > Ce (d) Sn > Yb > Ce > Lu

44. The increasing order of the ionic radii of the given isoelectronic species is :
(a) $Cl^-, Ca^{2+}, K^+, S^{2-}$ (b) $S^{2-}, Cl^-, Ca^{2+}, K^+$
(c) $Ca^{2+}, K^+, Cl^-, S^{2-}$ (d) $K^+, S^{2-}, Ca^{2+}, Cl^-$

45. Atom of which of the following elements has the greatest ability to attract electrons?
(a) Silicon (b) Sulphur
(c) Sodium (d) Nitrogen

46. The correct order of decreasing electronegativity values among the elements I-beryllium, II-oxygen, III-nitrogen and IV-magnesium is
(a) II > III > I > IV (b) III > IV > II > I
(c) I > II > III > IV (d) II > III > IV > I

47. The element with positive electron gain enthalpy is
(a) hydrogen (b) sodium
(c) oxygen (d) neon

48. Consider the following statements
I. The radius of an anion is larger than that of the parent atom.
II. The ionization energy generally increases with increasing atomic number in a period.
III. The electronegativity of an element is the tendency of an isolated atom to attract an electron.
Which of the above statements is/are correct?
(a) I alone (b) II alone
(c) I and II (d) II and III

49. The element with atomic number 117 has not been discovered yet. In which family would you place this element if discovered?
(a) Alkali metals (b) Alkaline earth metals
(c) Halogens (d) Noble gases

50. The set representing the correct order for first ionisation potential is
(a) K > Na > Li (b) Be > Mg > Ca
(c) B > C > N (d) Ge > Si > C

EXERCISE - 3
Exemplar & Past Years NEET/AIPMT Questions

Exemplar Questions

1. Consider the isoelectronic species, Na^+, Mg^{2+}, F^- and O^{2-}. The correct order of increasing length of their radii is
(a) $F^- < O^{2-} < Mg^{2+} < Na^+$
(b) $Mg^{2+} < Na^+ < F^- < O^{2-}$
(c) $O^{2-} < F^- < Na^+ < Mg^{2+}$
(d) $O^{2-} < F^- < Mg^{2+} < Na^+$

2. Which of the following is not an actinoid?
(a) Curium (Z = 96) (b) Californium (Z = 98)
(c) Uranium (Z = 92) (d) Terbium (Z = 65)

3. The order of screening effect of electrons of s, p, d and f orbitals of a given shell of an atom on its outer shell electrons is
(a) $s > p > d > f$ (b) $f > d > p > s$
(c) $p < d < s > f$ (d) $f > p > s > d$

4. The first ionisation enthalpies 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

5. The electronic configuration of gadolinium (Atomic number 64) is
(a) $[Xe]\ 4f^3\ 5d^5\ 6s^2$ (b) $[Xe]\ 4f^7\ 5d^2\ 6s^1$
(c) $[Xe]\ 4f^7\ 5d^1\ 6s^2$ (d) $[Xe]\ 4f^8\ 5d^6\ 6s^2$

6. The statement that is not correct for periodic classification of elements is
(a) the properties of elements are periodic function of their atomic numbers.
(b) non-metallic elements are less in number than metallic elements.
(c) for transition elements, the 3*d*-orbitals are filled with electrons after 3*p*-orbitals and before 4*s*-orbitals.
(d) the first ionisation enthalpies of elements generally increase with increase in atomic number as we go along a period.

7. Among halogens, the correct order of amount of energy released in electron gain (electron gain enthalpy) is
(a) F > Cl > Br > I (b) F < Cl < Br < I
(c) F < Cl > Br > I (d) F < Cl < Br < I

8. The period number in the long form of the periodic table is equal to
(a) magnetic quantum number of any element of the period
(b) atomic number of any element of the period
(c) maximum principal quantum number of any element of the period
(d) maximum azimuthal quantum number of any element of the period

9. The elements in which electrons are progessively filled in 4*f*-orbital are called
(a) actinoids (b) transition elements
(c) lanthanoids (d) halogens

10. Which of the following is the correct order of size of the given species
(a) $I > I^- > I^+$ (b) $I^+ > I^- > I$
(c) $I > I^+ > I^-$ (d) $I^- > I > I^+$

11. The formation of oxide ion $O^{2-}(g)$, from oxygen atom requires first an exothermic and then an endothermic step as shown below
$$O\,(g) + e^- \rightarrow O^-\,(g);\ \Delta H^{\ominus} = -141\ kJ\ mol^{-1}$$
$$O^-\,(g) + e^- \rightarrow O^{2-}\,(g);\ \Delta H^{\ominus} = +780\ kJ\ mol^{-1}$$
Thus, process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that
(a) oxygen is more electronegative
(b) addition of electron in oxygen results in larger size of the ion
(c) electron repulsion outweighs the stability gained by achieving noble gas configuration
(d) O^- ion has comparatively smaller size than oxygen atom

12. Comprehension given below is followed by some multiple choice questions. Each question has one correct option. Choose the correct option.
In the modern periodic table, elements are arranged in order of increasing atomic numbers which is related to the electronic configuration. Depending upon the type of orbitals receiving the last electron, the elements in the periodic table have been divided into four blocks, viz *s, p, d* and *f*.
The modern periodic table consists of 7 periods and 18 groups. Each period begins with the filling of a new energy shell. In accordance with the Aufbau principle, the seven periods (1 to 7) have 2, 8, 8, 18, 18, 32 and 32 elements respectively.

The seventh period is still incomplete. To avoid the periodic table being too long, the two series of *f*-block elements, called lanthanoids and actinoids are placed at the bottom of the main body of the periodic table.

(i) The element with atomic number 57 belongs to
(a) *s*-block (b) *p*-block
(c) *d*-block (d) *f*-block

(ii) The last element of the *p*-block in 6th period is represented by the outermost electronic configuration.
(a) $7s^2\ 7p^6$ (b) $5f^{14}\ 6d^{10}\ 7s^2\ 7p^0$
(c) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$ (d) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$

(iii) Which of the elements whose atomic numbers are given below, cannot be accommodated in the present set up of the long form of the periodic table?
(a) 107 (b) 118
(c) 126 (d) 102

(iv) The electronic configuration of the element which is just above the element with atomic number 43 in the same group is ....
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$
(b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^3\ 4p^6$
(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$
(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$

(v) The elements with atomic numbers 35, 53 and 85 are all ....
(a) noble gases (b) halogens
(c) heavy metals (d) light metals

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

NEET/AIPMT (2013-2017) Questions

14. Which one of the following arrangements represents the correct order of least negative to most negative electron gain enthalpy for C, Ca, Al, F and O? ***[NEET Kar. 2013]***
(a) Ca < Al < C < O < F
(b) Al < Ca < O < C < F
(c) Al < O < C < Ca < F
(d) C < F < O < Al < Ca

15. Which of the following orders of ionic radii is correctly represented ? ***[2014]***
(a) $H^- > H^+ > H$ (b) $Na^+ > F^- > O^{2-}$
(c) $F^- > O^{2-} > Na^+$ (d) $Al^{3+} > Mg^{2+} > N^{3-}$

16. The species Ar, K^+ and Ca^{2+} contain the same number of electrons. In which order do their radii increase ? ***[2015]***
(a) $Ca^{2+} < Ar < K^+$ (b) $Ca^{2+} < K^+ < Ar$
(c) $K^+ < Ar < Ca^{2+}$ (d) $Ar < K^+ < Ca^{2+}$

17. The formation of the oxide ion $O^{2-}(g)$, from oxygen atom requires first an exothermic and then an endothermic step as shown below :

$O(g) + e^- \rightarrow O^-(g);\ \Delta_f H^\ominus = -141\ kJ\ mol^{-1}$

$O^-(g) + e^- \rightarrow O^{2-}(g);\ \Delta_f H^\ominus = +780\ kJ\ mol^{-1}$

Thus process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that ***[2015 RS]***
(a) Electron repulsion outweighs the stability gained by achieving noble gas configuration
(b) O^- ion has comparatively smaller size than oxygen atom
(c) Oxygen is more electronegative
(d) Addition of electron in oxygen results in larger size of the ion.

18. In which of the following options the order of arrangement does not agree with the variation of property indicated against it ? ***[2016]***
(a) $Al^{3+} < Mg^{2+} < Na^+ < F^-$ (increasing ionic size)
(b) B < C < N < O (increasing first ionisation enthalpy)
(c) I < Br < Cl < F (increasing electron gain enthalpy)
(d) Li < Na < K < Rb (increasing metallic radius)

19. The element Z = 114 has been discovered recently. It will belong to which of the following family/group and electronic configuration ? ***[2017]***
(a) Carbon family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^2$
(b) Oxygen family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^4$
(c) Nitrogen family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^6$
(d) Halogen family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^5$

Hints & Solutions

EXERCISE - 1

1. (c) Group IA and III A contain mostly metals. Group VIII contains transition elements which are metals. Group VII A contains mostly non-metals (F, Cl, Br).

2. (d) Elements having 1, 2 or 3 electrons in its last shell act as metals.

   $32 = [Ar]\ 3d^{10}\ 4s^2\ p^2$

   $34 = [Ar]\ 3d^{10}\ 4s^2 p^4$

   $36 = [Ar]\ 3d^{10}\ 4s^2 p^6$

   $38 = [Ar]\ 3d^{10}, 4s^2 p^6, 5s^2$

3. (d) Basicity of oxides decreases in a period from left to right. Na_2O is basic oxide, CO_2, SiO_2 and SO_2 are acidic oxides.

   Alternatively, oxides of metals (e.g., Na_2O) are basic, while oxides of non-metals (SO_2, SiO_2 and CO_2) are acidic.

4. (d) Oxides of non metals are acidic in nature. P is a non-metal and its oxides are acidic Rest of the oxides are basic because they are oxides of metals.

5. (c) Allotropy is characteristic property of group 14 elements. All elements of group 14, except Pb, show allotropy.

   Sn has three allotropic forms grey tin, white tin and rhombic.

6. (a) Atomic number of the given element is 15 and it belongs to 5th group. Therefore atomic number of the element below the above element $= 15 + 18 = 33$.

7. (a) The electronic configuration clearly suggest that it is a d-block element (having configuration $(n-1)\ d^{1-10}\ ns^{0-2}$) which starts from III B and goes till II B. Hence with d^3 configuration it would be classified in the fifth group.

8. (d) Na_2O (basic), SO_2 and B_2O_3 (acidic) and ZnO is amphoteric.

9. (a) The screening effect follows the order $s > p > d > f$.

10. (d) The electrons are not filled in d-subshell monotonically with increase in atomic number among transition elements.

11. (c) 12. (b)

13. (c) $ns^2\ p^1$ is the electronic configuration of III A period. Al_2O_3 is amphoteric oxide.

14. (b)

15. (c) I. E. increases across a period and decreases down in a group. So, element with electronic configuration [Ne] $3s^2\ 3p^3$ will have the highest I.E. among the given choices.

I.E. increases →

	grp III A	grp IV A	grp V A
period 3	$[Ne]3s^2 3p^1$	$[Ne]3s^2 3p^2$	$[Ne]3s^2 3p^3$
period 4			$[Ar]3d^1 4s^2 4p^3$

I.E. decreases ↓

16. (b) The tendency of a cation to distort the electron cloud of an anion when it is approaching the anion, is called polarisation power of cation. As polarisation power of cation increases, the covalent character increases. According to Fajan's rule high charge and small size of cation will favour covalency. So, polarisation power of a cation increases with charge of the cation.

17. (d) Cd is not a transition metal among the given options because it do not have incomplete d-subshell either in its atomic state [Cd = $5d^{10}\ 4s^2$] or in its common oxidation state ($Cd^{2+} = 5d^{10}\ 4s^0$).

18. (c) F_2 has highest electronegativity, so it is chemically most active non metal.

19. (a) Metallic character decreases down group and increases along a period.

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20. (b)

21. (c) Basicity of oxides decreases in a period and increases in a group.

    ∴ SnO_2, Al_2O_3 and ZnO are amphoteric oxides.

22. (d)

23. (a) A cation is always smaller in size as compared to corresponding neutral atom. Greater the magnitude of charge, smaller will be size of ion. Following is the correct order of decreasing size $Al^{3+} < Al^{2+} < Al^{+} < Al$.

    ∴ Al has largest size.

24. (c) While moving down in a group, effective nuclear attraction decreases due to addition of new orbits. As a result ionisation potential decreases. Hence, the correct order is Li > K > Cs.

25. (c) Ionisation potential increases while moving in a period.

Group	V	VI	VII	VIII
Element	N	O	F	Ne

    Oxygen (group 6) has low ionisation potential than N (group 5) because of stable configuration of nitrogen (half filled *p*-orbital)

26. (c) They are isoelectronic species.

	N^{3-}	O^{2-}	Na^+	F^-
No. of electrons	10	10	10	10
No. of protons	7	8	11	9

    ∴ Attractive forces are highest in Na^+.

    ∴ Na^+ is smallest in size.

27. (d) According to the general trend of I.E. in a period, it is expected that oxygen atom has higher I.E. than nitrogen atom but nitrogen atom has more stable half filled p-orbitals due to which it has higher I.E. than oxygen atom.

28. (b) The acidic character of non metal oxides increases across a period from left to right and decreases down a group. So, acidic character will follow the order:
oxide of nitrogen > oxides of sulfur > oxides of carbon.
Among oxides of carbon acidic character increases with the oxidation number of carbon. So, $\overset{+4}{C}O_2$ is more acidic than CO. Hence the sequence of acidic character is $N_2O_5 > SO_2 > CO_2 > CO$

29. (d) Arsenic is the only metalloid among the given options. Its small amounts are even very harmful for humans.

30. (a)

Species	Na^+	Mg^{2+}	Al^{3+}	Si^{4+}
Protons	11	12	13	14
Electrons	10	10	10	10

Size of isoelectronic cations decreases with increase in magnitude of nuclear charge
∴ Order of decreasing size is $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$

31. (c) Ionisation potential is amount of energy required to take outermost loosely bonded electron from isolated gaseous atom. Its value decreases in a group and increases along a period. Thus, here Be has highest ionisation potential.

32. (a) $Be - 1s^2 2s^2; B - 1s^2 2s^2 2p^1; C - 1s^2 2s^2 2p^2; N - 1s^2 2s^2 2p^3;$ $O - 1s^2 2s^2 2p^4$. IP increases along the period. But IP of Be > B. Further IP of O < N because atoms with fully or partly filled orbitals are most stable and hence have high ionisation energy.

33. (a) O^{--} and F^- are isoelectronic. Hence have same number of shells, therefore greater the nuclear charge smaller will be the size i.e., $O^{--} > F^-$
further Li^+ and B^{3+} are isoelectronic. therefore
$Li^+ > B^{3+}$
Hence the correct order of atomic size is.
$O^{--} > F^- > Li^+ > B^{3+}$

34. (d) The given electronic configuration represents that it has 3 valency electrons or it can shows a maximum oxidation state of +3 and element with intermediate oxidation states form amphoteric oxides.

35. (a) Ionisation potential of nitrogen is more than that of oxygen. This is because nitrogen has more stable half-filled *p*-orbitals. ($N = 1s^2, 2s^2, 2p^3$, $O = 1s^2, 2s^2, 2p^4$)

36. (b)

37. (b) Due to presence of most penetrating *s*-electron, hydrogen (1s) shows maximum IP out of list.

38. (b)

39. (d) P^{5+} has more effective nuclear charge and smaller size than P^{3+}.

40. (a) Energy is supplied in order to remove electron from atoms. So energy of atom increases when electron is removed from atom.

41. (b) Ionic compounds have high melting point. Greater the ionic character, more is melting point.
HCl has least ionic character because of maximum electronegativity difference between the two constituent elements, H and Cl among CsF, HCl, HF and LiF
∴ HCl has minimum melting point.

42. (c) The ionization energy increases with decrease in size. Further the element having stable configuration has higher ionisation energy than expected. Hence the ionization energy of nitrogen (Z = 7) is more than oxygen (Z = 8) and carbon (Z = 6) because it has half-filled *p*-orbitals.
$$^6C = 1s^2 2s^2 2p^2 ;\ _7N = 1s^2 2s^2 2p^3 ;\ _8O = 1s^2 2s^2 2p^4$$
Hence the correct order should be C < N > O

43. (c)

44. (c)
(i) The anion is always larger in size as compared to corresponding neutral atom.
(ii) Greater the magnitude of negative charge, larger will be the size.
Therefore, the correct order of size is $O^{2-} > O^- > O$

45. (d)

46. (c) $(n-1)\, s^2p^6(n-1)d^{1-10}\, ns^{0-2}$ represents the correct electronic configuration of transition elements among the given choices.

47. (b) As atomic number, number 19 falls within group I of modern periodic table so it is an alkali metal with + 1 oxidation state.

48. (b) Due to high stability of half-filled orbitals.

49. (a) In Cu it has completely filled d-orbital so highest energy is absorbed when it convert in Cu^+ ion.

50. (a) Since, stable half filled configuration.

51. (b) $K^+ \rightarrow K^2 + e^-$. Since e^- is to be removed from stable configuration.

52. (c)

53. (a) 1st I.P. increases from left to right in a period.

54. (b)

55. (a)

56. (a) We know that atomic no. of fluorine (F), chlorine (Cl) Bromine (Br) and Iodine (I) are 9, 17, 35 and 53 respectively. Therefore, correct order of reactivity of halogens is F > Cl > Br > I

57. (b) Electropositive nature increases down the group and decreases across the period.

58. (b) Pollonium is only true metal in group 16. It has two crystalline forms α-form which is cubic and β-form which is rhombohedral.

59. (a) Electron affinity is said to be positive when an atom has spontaneous tendency to accept an electron. When O changes to O^-, energy is released. So, this change has positive electron affinity while all other given changes required to be forced i.e., these require energy to occur.

60. (a) Electron affinity is energy released when electron is added to isolated gaseous atom. Its value decreases down the group. So electron affinity of F should be highest among halogens but due to its smaller size electron affinity of Cl is more than F.
$\therefore$ Cl has highest electron affinity.

61. (a) Pauling scale of electronegativity was helpful in predicting
(i) Nature of bond between two atoms
(ii) Stability of bond
by calculating the difference in electronegativities polarity of bond can be calculated.

62. (c) 63. (b) 64. (b)

65. (b) Electronegativity decreases down the group and increases along a period. Cl lies in 17th group hence more electronegative than P and Si; further it lies above Br, hence more electronegative than Br.

66. (d) Fluorine, chlorine, bromine and iodine are placed in the same group 17 because they have 7 electrons in the outermost shell.

67. (a)

68. (a) Zero, because of the stable electronic configuration the noble gases do not show any force of attraction towards the incoming electron.

69. (a) Halogens have the highest e^- affinity.

70. (b) An atom with high electronegativity has high IP.

71. (c) Chlorine and sulphur are in period three. Potassium and calcium are in period four. As K has radius more than calcium, K^+ ion will have largest size.

72. (a) F, because of its smallest size.

73. (c) Halogens are most electronegative.

74. (b) On passing from left to right in a period acidic character of the normal oxides of the elements increases with increase in electronegativity.

75. (d) The energy involved is ionisation energy (I.E.). Further the 3rd ionisation energy will be greater than the 2nd and 1st.

EXERCISE - 2

1. (d) 2. (a)

3. (a) Acidic character of oxide $\propto$ Non-metallic nature of element.
Non-metallic character increases along the period. Hence order of acidic character is
$Cl_2O_7 > SO_2 > P_4O_{10}$.

4. (c) $X_{(g)} + e^- \rightarrow X^-_{(g)} + x\text{ kJ}$ ....(i)
$X^-_{(g)} \rightarrow X_{(g)} - x\text{ kJ}$ ....(ii)

5. (d) Zero group is also called as buffer group because it is placed between highly electropositive metals (group 1) and highly electronegative non-metals (group 17).

6. (a) In case of Ga there are 10d electrons in the penultimate energy shell which shield the nuclear charge less effectively, the outer electron is held firmly by nucleus. As a result, the ionisation energy remains nearly the same as that of aluminium inspite of the fact that atomic size increases.

7. (d) Abnormally high difference between 2nd and 3rd ionization energy means that the element has two valence electrons, i.e., configuration (d)

8. (b) Along the period, I.P. generally increases but not regularly. Be and B are exceptions. First I.P. increases in moving from left to right in a period, but I.P. of B is lower than Be.

9. (a) Gaining of an electron by a gaseous atom is usually an exothermic process. Gain of second electron by negatively charged species feels a strong repulsion and the energy of the system increases.

10. (a)

11. (d) As the size increases the basic nature of oxides changes to acidic nature i.e., acidic nature increases.

$$\underset{\text{Acidic}}{SO_2 > P_2O_3} > \underset{\text{Weak acidic}}{SiO_2} > \underset{\text{Amphoteric}}{Al_2O_3}$$

SO_2 and P_2O_3 are acidic as their corresponding acids H_2SO_3 and H_3PO_3 are strong acids.

12. (a) O^- ion exerts a force of repulsion on the incoming electron. The energy is required to overcome it.

13. (c) In a period the value of ionisation potential increases from left to right with breaks where the atoms have some what stable configuration. In this case N has half filled stable orbitals. Hence has highest ionisation energy. Thus the correct order is $B < C < O < N$ and not as given in option (c)

14. (b) 3rd ionization involves removal of electron from inert gas configuration $1s^2\, 2s^2\, 2p^6$, hence there would be a large jump between 2nd and 3rd ionization energies.

15. (a) IE_1 is always less than IE_2.

16. (b) IE_2 of Mg is lower than that of Na because in case of Mg^+, 3*s*-electron has to be removed whereas in case of Na^+, an electron is removed from the stable inert gas configuration which is difficult.

17. (c) 18. (c)

19. (c) Amongst the ions carrying same charge, the smallest one will have the greatest hydration energy (most negative).

20. (d)

21. (d) Tl shows the inert pair effect. Hence Tl^+ oxidation state is more stable than Tl^{3+}.

22. (a) 23. (a) 24. (b) 25. (d)

26. (a) For isoelectronic species, size of anion increases as negative charge increases whereas size of cation decreases with increase in positive charge. Further ionic radii of anions is more than that of cations. Thus the correct order is $Ca^{++} < K^+ < Ar < Cl^- < S^{--}$

27. (a) Proton affinity decreases in moving across the period from left to right due to increase in charge, within a group the proton affinities decreases from top to bottom. Nitrogen family > Oxygen family > Halogens

28. (d) The stability of +1 oxidation state increases from aluminium to thallium i.e. Al < Ga < In < Tl
29. (d) The smaller the atomic size, larger is the value of ionisation potential. Further the atoms having half filled or fully filled orbitals are comparitively more stable, hence more energy is required to remove the electron from such atoms.
30. (c) $_{12}Mg$ $_{15}P$ $_{17}Cl$ $_{20}Ca$
160$_p$ 110 99 197 (pm)
Cl < P < Mg < Ca
31. (a) IE_1 of Na = – Electron gain enthalpy of Na^+ = – 5.1 eV.
32. (b) The alkali metals are highly reactive because their first ionisation potential is very low and hence they have great tendency to loses electron to form unipositive ion.
On moving down group- I from Li to Cs ionisation enthalpy decreases hence the reactivity increases. The halogens are most reactive elements due to their low bond dissociation energy, high electron affinity and high enthalpy of hydration of halide ion. However their reactivity decreases with increase in atomic number
33. (b) In hydrides of 15th group elements, basic character decreases on descending the group i.e.
$NH_3 > PH_3 > AsH_3 > SbH_3$.
34. (d) All the given species contains 10 e^- each i.e. isoelectronic.
For isoelectronic species anion having high negative charge is largest in size and the cation having high positive charge is smallest.
35. (b) As we move down in a group electron gain enthalpy becomes less negative because the size of the atom increases and the distance of added electron from the nucleus increases. Negative electron gain enthalpy of F is less than Cl. This is due to the fact that when an electron is added to F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In Cl, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron–electron repulsion. So the correct order is Cl > F > Br > I.
36. (c) On moving along a period from left to right I.E. increases and on moving down a group I.E. decreases.
hence correct order is : Ba < Ca < Se < S < Ar
37. (c) Its valency is 2. So it will form MO type compound.
38. (b) On moving along the period, ionization enthalpy increases.
In second period, the order of ionization enthalpy should be as follows : F > O > N
But N has half-filled structure, therefore, it is more stable than O. That's why its ionization enthalpy is higher than O. Thus, the correct order of IE is
F > N > O.
39. (c) N, O and F (p-block elements) are highly electronegative non metals and will have the strongest tendency to form anions by gaining electrons from metal atoms.
40. (a) Non metals form oxides with oxygen and thus reduce oxides of metals behaving as reducing agents.
41. (b)
42. (b) O < S < F < Cl
Electron gain enthalpy – 141, –200, – 333, – 349 kJ mol^{-1}
43. (b) Correct order of ionic size is Sn > Ce > Yb > Lu.
44. (c)
45. (d) Halogens have very high values of electron gain enthalpies.
46. (a) Electronegativity values of given elements are as follows:
Be – 1.5 (I) Mg – 1.2 (IV)
O – 3.5 (II) N – 3.0 (III)
i.e. II > III > I > IV
47. (d) Noble gases have positive values of electron gain enthalpy because the anion is higher in energy than the isolated atom and electron.
48. (c) The tendency of an atom in a compound to attract a pair of bonded electrons towards itself is known as electronegativity of the atom.
49. (c)
50. (b) The correct order of first ionisation energy is represented by
Be > Mg > Ca
Since on moving down a group atomic size increases due to addition of one extra shell, hence I.E decreases.

EXERCISE - 3 AAJ KA TOPPER

Exemplar Questions

1. (b) In case of isoelectronic species

$$\text{ionic radii} \propto \frac{1}{\text{atomic number}}$$

∴ The correct order of increasing ionic radii will be :

Ionic radii	Mg^{2+} <	Na^+ <	F^- <	O^{2-}
Atomic number	(12)	(11)	(9)	(8)

2. (d) Elements with atomic number, Z = 90 to 103 are called actinoids. Terbium belongs to lanthanoids.
3. (a) For a given shell, screening effect decreases in the order : $s > p > d > f$.
4. (a) Electronic configuration for the given elements will be :
Na = [Ne]$3s^1$, Mg = [Ne]$3s^2$, Al = [Ne]$3s^23p^1$,
Si = [Ne] $3s^23p^2$
Ionisation enthalpy increases along a period but I.E of Mg is higher than Al because of completely filled 3s orbital in Mg.
5. (c) The electronic configuration of Gd (Z = 64) is
[Xe] $4f^7\,5d^1\,6s^2$.
6. (c) In case of transition element, the order of filling of electrons in various orbital is $3p < 4s < 3d$. Thus, 3d orbital is filled only when 4s orbital gets completely filled.

7. (c) As we move in a group from Cl to I, the electron gain enthalpy (i.e., energy released in electron gain) become less and less negative due to corresponding increase in the atomic size.
However, the electron gain enthalpy of F is less negative than that of Cl due to its small size. Thus, the negative electron gain enthalpy among halogens follows the order :
$$F < Cl > Br > I$$
8. (c) As each period starts with the filling of electrons in a new principal quantum number, so, the period number in the long form of the periodic table refers to the maximum principal quantum number (n) of any element in the period.
9. (c) The elements in which electrons are filled in $4f$-orbital are called lanthanoids. Lanthanoids consist of elements from Z = 58 (cerium) to 71 (lutetium).
10. (d) Generally, cations are smaller in size while anions are bigger in size than the neutral atom.
11. (c) O^{2-} has noble gas configuration and isoelectronic with neon but its formation is unfavourable due to strong electronic repulsion between the negatively charged O^- ion and the electron being added.
Thus, the electron repulsion will be more than the stability gained by achieving noble gas configuration.
12. (i) (c) The element with atomic number 57 belongs to d-block element as the last electron enters into the $5d$-orbital against the aufbau principle. This anomalous behaviour can be explained on the basis of greater stability of the xenon (inert gas) core.
After barium (Z = 56), the addition of the next electron should occur in 4f-orbital in accordance with aufbau principle. This will however, tend to destabilize the xenon core (Z = 54), [Kr] ($4d^{10}\,4f^0\,5s^2\,5p^6\,5d^0$) since the $4f$-orbitals lie inside the core. Therefore, the 57th electron prefers to enter $5d$-orbital which lies outside the xenon core and whose energy is only slightly higher than that of $4f$-orbital. Thus, the outer electronic configuration of La(Z = 57) is $5d^1\,6s^2$ rather than the expected $4f^1\,6s^2$.
(ii) (c) Each period starts with the filling of electrons in a new principal energy shell. Therefore, 6th period starts with the filling of $6s$-orbital and ends when $6p$-orbitals are completely filled.
Thus, the outermost electronic configuration of the last element of the p-block in the 6th period is represented by $6s^2\,4f^{14}\,5d^{10}\,6p^6$ or $4f^{14}\,5d^{10}\,6s^2\,6p^6$.
(iii) (c) The long form of the periodic table contain element with atomic number 1 to 118.
(iv) (a) The fifth period begins with Rb (Z = 37) and ends at Xe (Z = 54). Thus, the element with Z = 43 lies in the 5th period. Since, the 4th period has 18 elements, thus, the atomic number of the element which lies immediately above the element with atomic number 43 will be 43 – 18 = 25.
The electronic configuration of the element with Z = 25 is
$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5\,4s^2$ (i.e., Mn).
(v) (b) The elements with atomic numbers 35(36 – 1), 53(54 – 1) and 85(86 – 1), lie in a group before noble gases, i.e., belongs to halogens (group 17).
13. (a) Electronic configuration of given elements indicate that A is a noble gas (i.e., Ne), B is oxygen, C is sodium metal and D is fluorine.
(i) Noble gases have no tendency to gain electrons since all their orbitals are completely filled.
∴ element A will have the least electron gain enthalpy.
(ii) Element D has one electron less and element B has two electrons less than the corresponding noble gas configuration, hence, element D will have the highest electron gain enthalpy in comparison to element B.
(iii) Since, element C has one electron in the s-orbital and need one more electron to complete its configuration, therefore, electron gain enthalpy of C is less than that of element B. So, we can conclude that the electron gain enthalpies of the four elements increases in the order : A < C < B < D.

NEET/AIPMT (2013-2017) Questions

14. (a) As the nuclear charge increases, the force of attraction between the nucleus and the incoming electron increases and hence the elecron gain enthalpy becomes more negative, hence the correct order is
Ca < Al < C < O < F
15. (N) All answers are incorrect.
16. (b) In isoelectronic species the radius decrease with increase in nuclear charge hence increasing order of radius is
$Ca^{+2} < K^+ < Ar$
17. (a) Incoming electrons occupy the smaller n = 2 shell, also negative charge on oxygen (O^-) is another factor due to which incoming electron feel repulsion.
Hence electron repulsion outweigh the stability gained by achieving noble gas configuration.
18. (b&c) The correct order is B < C < O < N
Generaly Ionisation energy increases across a period. But here first I.E. of O is less than the first I.E. of N. This is due to the half-filled $2p$ orbital in N($1s^2$, $2s^2$, $2p^3$) which is more stable than the $2p$ orbital in O ($1s^2$, $2s^2$, $2p^4$)
(c) The correct order of electron affinity is
I < Br < F < Cl
Halogens have high electron affinities which decreases on moving down the group. However, fluorine has lower value than chlorine which is due to its small size and repulsion between the electron added and electrons already present.
19. (a) Z = 114 belong to Group 14, carbon family
Electronic configuration = $[Rn]5f^{14}6d^{10}7s^27p^2$