Summary

Development of modern Periodic Table:

(a) Dobereiner's Triads: He arranged similar elements in the groups of three elements called as triads, in which the atomic mass of the central element was merely the arithmetic mean of atomic masses of other two elements or all the three elements possessed nearly the same atomic masses.

(b) Newland's Law of Octave: He was the first to correlate the chemical properties of the elements with their atomic masses. According to him if the elements are arranged in the order of their increasing atomic masses the eight element starting from given one is similar in properties to the first one. The arrangement of elements is called as Newland's law of Octave.

(c) Lother Meyer's Classification: He determined the atomic volumes by dividing atomic with their densities in solid states.

he plotted a graph between atomic masses against their respective atomic volumes for a number of elements. He found the observations; (i) elements with similar properties occupied similar positions on the curve, (ii) alkali metals having larger atomic volumes occupied the crests, (iii) transitions elements occupied the troughs, (iv) the halogens occupied the ascending portions of the curve before the inert gases and (v) alkaline earth metals occupied the positions at about the mid points of the descending portions of the curve. On the basis of these observations he concluded that the atomic volumes (a physical property) of the elements are the periodic functions of their atomic masses

(d) Mendeleev's Periodic Table:

Mendeleev's Periodic Law

According to him the physical and chemical properties of the elements are the periodic functions of their atomic masses.

This table was divided into nine vertical columns groups and seven horizontal rows called periods.

Periods	Number of Elements	Called as
$(1)^{st}$ n= 1	2	Very short period
$(2)^{nd}$ n = 2	8	Short period
$(3)^{rd}$ n = 3	8	Short period
$(4)^{th}$ n = 4	18	Long period
$(5)^{th}$ n = 5	18	Long period
$(6)^{th}$ n = 6	32	Very long period
$(7)^{th}$ n = 7	32	Very long period

The groups were numbers as I, II, III, IV, V, VI, VII, VIII and Zero group

Merits of Mendeleev's Periodic table:

- It has simplified and systematised the study of elements and their compounds.
- It has helped in predicting the discovery of new elements on the basis of the blank spaces given in its periodic table.

Demerits in Mendeleev's Periodic Table:

- Position of hydrogen is uncertain. It has been placed in IA and VIIA groups because of its resemblance with both the groups.
- No separate positions were given to isotopes.
- Anomalous positions of lanthanides and actinides in periodic table.
- Order of increasing atomic weights is not strictly followed in the arrangement of elements in the periodic table. For example Ar(39.94) is placed before K(39.08) and Te(127.6) is placed before I(126.9).
- Similar elements were placed in different groups e.g. Cu in IB and Hg in IIB and similarly the elements with different properties were placed in same groups e.g. alkali metals in IA and coinage metals in IB.
- It didn't explained the causes of periodicity.

(e) Long form of the Periodic Table or Moseley's Periodic Table:

Modern Periodic Law (Moseley's Periodic Law):

Physical and chemical properties of the elements are the periodic functions of their atomic number. If the elements are arranged in order of their increasing atomic number, after a regular interval, elements with similar properties are repeated.

Periodicity:

The repetition of the properties of elements after regular intervals when the elements are arranged in the order of increasing atomic number is called periodicity.

Cause of Periodicity:

The periodic repetition of the properties of the elements is due to the recurrence of similar valence shell electronic configurations after certain regular intervals. For example, alkali metals have same valence shell electronic configuration ns^1, therefore, have similar properties.

The modern periodic table consists of horizontal rows (periods) and vertical column (groups).

Periods:

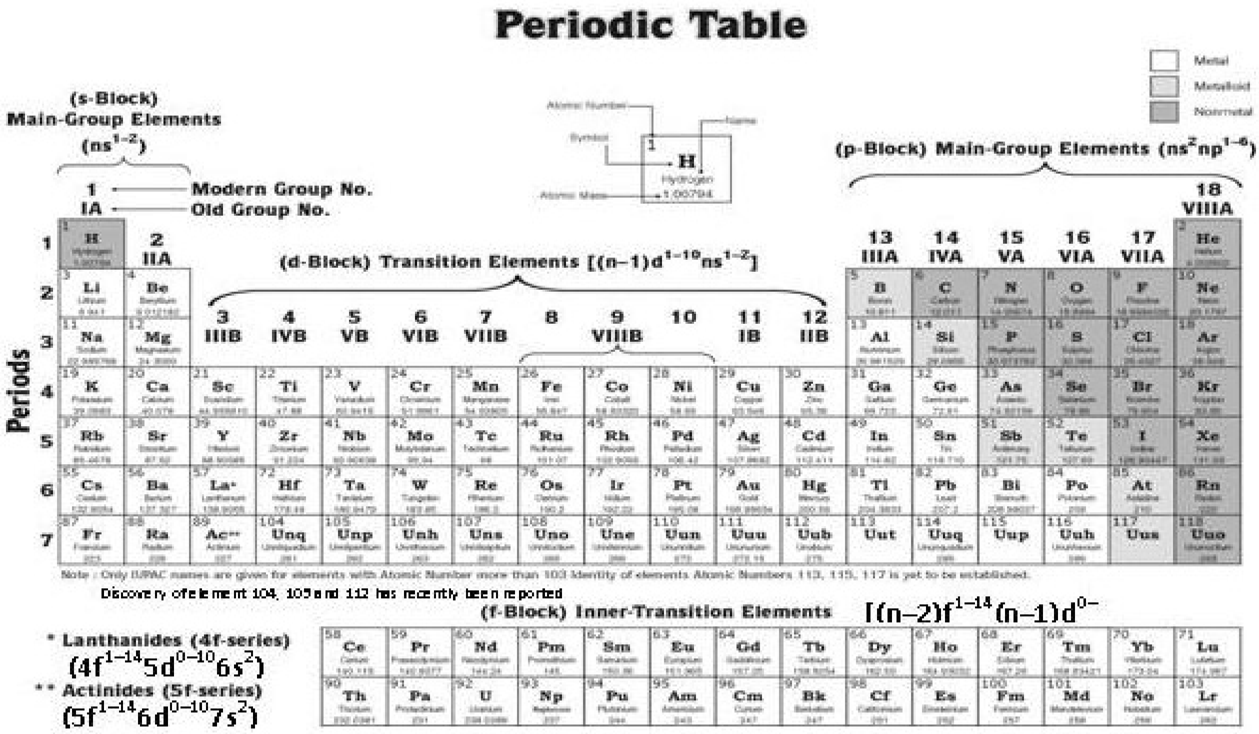
There are seven periods numbered as 1, 2, 3, 4, 5, 6 and 7.

- Each period consists of a series of elements having same valence shell.
- Each period corresponds to a particular principal quantum number of the valence shell present in it.
- Each period starts with an alkali metal having outermost electronic configuration as ns^1.
- Each period ends with a noble gas with outermost electronic configuration ns^2np^6 except helium having outermost electronic configuration as $1s^2$.
- Each period starts with the filling of new energy level.
- The number of elements in each period is twice the number of atomic orbitals available in energy level that is being filled.

Groups:

There are eighteen groups numbered as 1, 2, 3, 4, 5, …… 13, 14, 15, 16, 17, 18.

Group consists of a series of elements having similar valence shell electronic configuration.



Classification of the Elements:

It is based on the type of orbitals which relieves the differentiating electron (i.e., last electron).

(a) s-block elements

When shells upto (n – 1) are completely filled and the last electron enters the s-orbital of the outermost (nth) shell, the elements of this class are called s-block elements.

Group 1 and 2 elements constitute the s-block. General electronic configuration is [inert gas] ns^{1-2} s-block elements lie on the extreme left of the periodic table.

(b) p-block elements

When shells upto (n – 1) are completely filled and differentiating electron enters the p-orbital of the nth orbit, elements of this class are called p-block elements.

Group 13 to 18 elements constitute the p-block. General electronic configuration is [inert gas] ns^2np^{1-6} p-block elements lie on the extreme right of the periodic table. This block includes some metals, all nonmetals and metalloids. s-block and p-block elements are collectively called normal or representative elements.

(c) d-Block elements

When outermost (nth) and penultimate shells (n – 1)th shells are incompletely filled and differentiating electron enters the (n – 1) d orbitals (i.e., d-orbital of penultimate shell) then elements of this class are called d-block elements. Group 3 to 12 elements constitute the d-block. General electronic configuration is [inert gas] $(n-1)\ d^{1-10}\ ns^{1-2}$ (except, palladium which has valence shell electron configuration $4d^{10}5s^0$). All the transition elements are metals and most of them form coloured complexes or ions.

(d) f-Block elements

When n, (n – 1) and (n – 2) shells are incompletely filled and last electron enters into f-orbital of antepenultimate i.e., (n – 2)th shell, elements of this class are called f-block elements. General electronic configuration is (n – 2) f^{1-14} (n – 1) d^{0-1} ns^2. All f-block elements belong to 3rd group.

The elements of f-blocks have been classified into two series. (1) Ist inner transition or 4 f-series, contains 14 elements $_{58}Ce$ to $_{71}Lu$. Filling of electrons takes place in 4f subshell. (2). IInd inner transition or f-series, contains 14 elements $_{90}{}^{Th}$ to $_{103}Lr$. Filling of electrons takes place in 5f subshell.

Prediction of period, group and block:

- Period of an element corresponds to the principal quantum number of the valence shell.
- The block of an element corresponds to the type of subshell which receives the last electron.
- The group is predicted from the number of electrons in the valence shell or/and penultimate shell as follows.

(a) For s-block elements; Group no. = the no. of valence electrons

(b) For p-block elements; Group no. = 10 + no. of valence electrons

(c) For d-block elements; Group no. = no. of electrons in (n – 1)d sub shell + no. of electrons in valence shell.

Metals and non-metals:

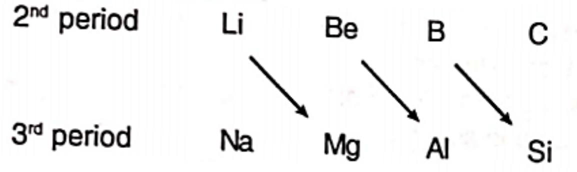
- The metals are characterised by their nature of readily giving up the electron(s) and from shinning lustre. Metals comprises more than 78% of all known elements and appear on the left hand side of the periodic table. Metals are usually solids at room temperature (except mercury, gallium). They have high melting and boiling points and are good conductors of heat and electricity. Oxides of metals are generally basic in nature (some metals in their higher oxidation state form acid oxides e.g. CrO_3).
- Nonmetals do not lose electrons but take up electrons to form corresponding anions. Nonmetals are located at the top right hand side of the periodic table. Nonmetals are usually solids, liquids or gases at room temperature with low melting and boiling points. They are poor conductors of heat and electricity. Oxides of non-metals are generally acidic in nature.

Metalloids (Semi metals):

It can be understood from the periodic table that nonmetallic character increases as we move from left to right across a row. It has been found that some elements which lie at the border of metallic and nonmetallic behavior, possess the properties that are characteristic of both metals and nonmetals. These elements are called semi metals or metalloids. The metalloids comprise of the elements B, Si,Ge, As, Sb and Te.

Diagonal relationship:

Some elements of certain groups of 2nd period resemble much in properties with the elements of third period of next group i.e. elements of second and third period are diagonally related in properties. This phenomenon is known as diagonal relationship.



Diagonal relationship arises because of:

(i) Effective nuclear charge:

The effective nuclear charge (Z_{eff}) is the charge actually felt by the valence electron. Z_{eff} 1S given by $Z_{eff} = Z - \sigma$, (where Z is the actual nuclear charge (atomic number of the element) and σ is the shielding (screening) (constant). The value of σ i.e. shielding effect can be determined using the Slater's rules.

(ii) Atomic radius:

(A) Covalent radius : It is one-half of the distance between the centres of two nuclei (of like atoms) bonded by a single covalent bond. Covalent radius is generally used for nonmetals.

(B) Vander Waal's radius (Collision radius): It is one-half of the internuclear distance between two adjacent atoms in two nearest neighbouring molecules of the substance in solid State.

(C) Metallic radius (Crystal radius): It is one-half of the distance between the nuclei of two adjacent metal atoms in the metallic crystal f lattice.

- Thus, the covalent, vander Wall's and metallic radius magnitude wise follows the order,

  $r_{covalent} < r_{crystal} < r_{vander\ Walls}$

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Variation in a Period	**Variation in a group**
In a period left to right:	In a group top to bottom:
Nuclear charge (Z) increases by one unit	Nuclear charge (Z) increases by more than one unit
Effective nuclear charge (Z_{eff}) also increases	Effective nuclear charge (Z_{eff}) almost remains constant because of increased screening effect of inner shells electrons.
But number of orbitals (n) remains constant	But number of orbitals (n) increases.
As a result, the electrons are pulled closer to the nucleus by the increases Z_{eff}. $r_n \propto \frac{1}{Z^*}$ Hence atomic radii decreases with increases in atomic number in a period from left to right.	The effect of increased numbero f atomic shells overweighs the effect of increased nuclear charge. As a result of this the size of atom increases from top to bottom in a given group.

The atomic radius of inert gas (zero group) is given largest in a period because it is represented as vander wall's radius which is generally larger than the covalent radius. The vander Waal's radius of inert gases also increases from top to bottom in a group.

(iii) Ionic radius:

The effective distance from the centre of nucleus of the ion up to which it has an influence in the ionic bond is called ionic radius.

Cation	Anion
It is formed by the lose of one or more electrons from the valence shell of an atom of an element. Cations are smaller than the parent atoms because, (i) the whole of the outer shell of electrons is usually removed. (ii) in a cation, the number of positive charges on the nucleus is greater than number of orbital electrons leading to increased inward pull of remaining electrons causing contraction in size of the ion.	It is formed by the gain of one or more electrons in the valence shell of an atom of an element. Anions are larger than the parent atoms because (i) anion is formed by gain of one or more electrons in the neutral atom and thus number if electrons increases but magnitude of nuclear charge remains the same. (ii) nuclear charge per electrons is thus reduced and the electrons cloud is held less tightly by the nucleus leading to the expansion of the outer shell. Thus size of anion is increased.

- The species containing the same number of electrons but differ in the magnitude of their nuclear charges are called as isoelectronic species. For example, N^{3-}, O^{2-}, F^-, Ne, Na^+, Mg^{2+} and Al^{3+} are all isoelectronic species with same number of electrons (i.e 10) but different nuclear charges of +7, +8, +9, +10, +11, +12 and +13 respectively.

  Within a series of isoelectronic species as the nuclear charge increases, the force of attraction by the nucleus on the electrons also increases. As a result, the ionic radii of isoelectronic species decrease with increases in the magnitude of nuclear charges.

(iv) Ionisation Energy:

Ionisation energy (IE) is defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form a cation.

$$M(g) \xrightarrow{(IE_1)} M^+(g) + e^- \text{ ; } M^+(g) + IE_2 \rightarrow M^{2+}(g) + e^-$$

$$M^{2+}(g) + IE_3 \rightarrow M^{+3}(g) + e^-$$

IE_1, IE_2 and IE_3 are the I^{st}, II^{nd} and III^{rd} ionization energies t remove electron from a neutral atom, monovalent and divalent cations respectively. In general, $(IE)_1 < (IE)_2 < (IE)_3 <$.........

- **Factors Influencing Ionisation energy**

  Variation in ionization energies in a period and group may or not be regular and can be influenced by the following factors.

- **(A) Size of the Atom:** Ionisation energy decreases with increase in atomic size
- **(B) Nuclear Charge:** The ionisation energy increases with increase in the nuclear charge.
- **(C) Shielding or screening effect:** The electrons in the inner shells act as a screen or shield between the nucleus and the electrons in the outer most shell. This is called shielding effect. The larger the number of electrons in the inner Shells, greater is the screening effect and smaller the force of attraction and thus ionization energy (IE) decreases.
- **(D) Penetration effect of the electron:** Penetration effect of the electrons follows the order s > p > d > f for, the same energy level. Higher the penetration of electron higher will be the ionisation energy.
- **(E) Electronic Configuration:** If an atom has exactly half-filled or completely filled orbitals, then such an arrangement has extra stability. The removal of an electron from such an atom requires more energy then expected. Metallic or electropositive character of elements increases as the value of ionisation energy decreases. The relative reactivity of the metals in gaseous phase increases with the decrease in ionisation energy.

The reducing power of elements in gaseous phase increases as the value of ionisation energy decreases. Amongst alkali metals, the lithium is strongest reducing agent in aqueous solution.

(v) Electron Gain Enthalpy:

The electron gain enthalpy $\Delta_{eg} H^{\ominus}$, is the change in standard molar enthalpy when a neutral gaseous atom gains an electron to form an anion.

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

Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion. Electron gain may be either exothermic or endothermic depending on the elements.

When an electron is added to the atom and the energy is released, the electron gain enthalpy is negative and when energy is needed to add an electron to the atom, the electron gain enthalpy is positive.

The second electron gain enthalpy, the enthalpy change for the addition of a second electron to an initially neutral atom, invariably positive because the electron repulsion out weighs the nuclear attraction.

- Group 17 elements (halogens) have very high negative electron gain enthalpies (i.e. high electron affinity) because they can attain stable noble gas electronic configuration by picking up an electron.
- Across a period, with increase in atomic number, electron gain enthalpy becomes more negative because left to right across a period effective nuclear charge increases and consequently it will be easier to add an electron to a small atom.
- As we move in a group from top to bottom, electron gain enthalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus.
- Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher energy level leading to a very unstable electronic configuration.

- Negative electron gain enthalpy of O or F is less than S or Cl. This is due to the fact that when an electron is added to O or 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 S or 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.
- Electron gain enthalpies of alkaline earth metals are very less or positive because the extra electron is to be added to completely filled s-orbitals in their valence shells.
- Nitrogen has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.
- (i) Electron affinity $\propto \frac{1}{\text{Atomic size}}$

  (ii) Electron affinity $\propto$ Effective nuclear charge (Z_{eff})

  (iii) Electron affinity $\propto \frac{1}{\text{Screening effect}}$

  (iv) stability of half filled and completely filled orbitals of a subshell is comparatively more and the addition of an extra electron to such an system is difficult and hence the electron affinity value decreases.

(vi) Electronegativity:

Electronegativity is a measure of the tendency of an element to attract shared electrons towards itself in a covalently bonded molecules.

The magnitude of electronegativity of an element depends upon its ionisation potential & electron affinity. Higher ionisation potential & electron affinity values indicate higher electronegativity value.

(a) Pauling's scale: Linus Pauling developed a method for calculating relative electronegativities of most elements. According to Pauling

$$\Delta = X_A - X_B = 0.208\sqrt{E._{A-B} - \sqrt{E_{A-A} \times E_{B-B}}}$$

E_{A-B} = Bond enthalpy/Bond energy of A – B bond.

E_{A-A} = Bond energy of A – A bond

E_{B-B} = Bond energy of B – B bond

(All bond energies are in kcal/mol)

$$\Delta = X_A - X_B = 0.1017\sqrt{E._{A-B} - \sqrt{E_{A-A} \times E_{B-B}}}$$

All bond energies are in kJ/mol.

(b) Mulliken's scale: Electronegativity χ can be regarded as the average of the ionisation energy (IE) and the electron affinity (EA) of an atom (both expressed in electron volts).

$$\chi_M = \frac{IE + EA}{2}$$

Paulings's electronegativity χ_P is related to Mulliken's electronegativity χ_M as given below.

$$\chi_P = 1.35(\chi_M)^{1/2} - 1.37$$

Mulliken's values were about 2.8 times larger than the Pauling's values.

(vii) Periodicity of Valence or Oxidation States:

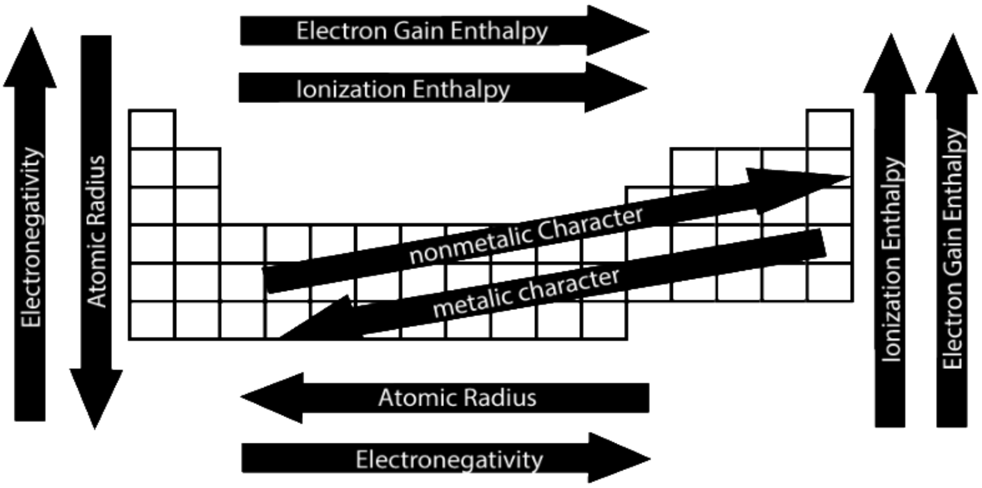
The valence of representative elements is usually (though not necessarily) equal to the number of electrons in the outer most orbitals and/or equal to eight minus the number ofouter most electrons. Nowadays the term oxidation state is frequently used for valence. There are many elements which exhibit variable valence. This is particularly characteristic of transition y elements and actinoids.

(viii) Periodic Trends and Chemical Reactivity:

The ionization enthalpy of the extreme left element in a period is the least and the electron gain enthalpy of the element on the extreme right is the highest negative (except noble gases which having completely filled shells have rather positive electron gain enthalpy values). This results in high chemical reactivity at the two extremes and the lowest in the centre. The loss and gain of electron can be related with the reducing and oxidizing behaviour of the elements respectively. However, it can also be directly related to the metallic and nonmetallic character of elements. Thus, the metallic character of an element, which is highest at the extremely left decreases and the nonmetallic character increases while moving from left to right across the period. The normal oxide formed by the element on extreme left is the most basic (e.g. Na_2O), whereas that formed by the element on extreme right is the acidic (e.g. Cl_2O_7). Oxides of elements in the centre are amphoteric (e.g. Al_2O_3, As_2O_3) or neutral (e.g., CO, NO, N_2O).

- In a group, basic nature of oxides increases or acidic nature decreases. Oxides of the metals are generally basic and oxides of the nonmetals are acidic. The oxides of the metalloids are generally amphoteric in nature. The oxides of Be, Al, Zn, Sn, As, Pb and Sb are amphoteric.
- In a period the nature of the oxides varies from basic to acidic.

Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Strongly basic	Basic	amphoteric	Weakly acidic	Acidic	Acidic	Strongly acidic



Practice Questions

1. The ionic radii (in Å) of N^{3-}, $O2^{-}$ and F^{-} respectively are (2015)
(a) 1.36, 1.40 and 1.71
(b) 1.36, 1.71 and 1.40
(c) 1.71, 1.40 and 1.36
(d) 1.71,1.36 and 1.40

2. Which one of the following alkaline earth metal sulphates, has its hydration enthalpy greater than its lattice enthalpy? (2015)
(a) $CaSO_4$
(b) $BeSO_4$
(c) $BaSO_4$
(d) $SrSO_4$

3. Which among the following is the most reactive? (2015)
(a) Cl_2
(b) Br_2
(c) I_2
(d) ICI

4. Which one has the highest boiling point?
(a) He
(b) Ne
(c) Kr
(d) Xe

5. The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be
(2013)
(a) –2.55 eV
(b) –5.1 eV
(c) –10.2 eV
(d) +2.55 eV

6. Which of the following represents the correct order of increasing first ionisation enthalpy for Ca, Ba, S, Se and Ar? (2013)
(a) $Ca < S < Ba < Se < Ar$
(b) $S < Se < Ca < Ba < Ar$
(c) $Ba < Ca < Se < S < Ar$
(d) $Ca < Ba < S < Se < Ar$

7. Identify the least stable ion amongst the following. (2002)
(a) Li^+
(b) Be^-
(c) B^-
(d) C^-

8. The set representing the correct order of first ionisation potential is (2001)
(a) $K > Na > Li$
(b) $Be > Mg > Ca$
(c) $B > C > N$

(d) Ge > Si > C

9. The correct order of radii is (2000)
(a) N < Be < B
(b) $F^- < O^{2-} < N^{3-}$
(c) Na < Li < K
(d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

10. The incorrect statement among the following (1997)
(a) The first ionisation potential of A1 is less than the first ionisation potential of Mg
(b) The second ionisation potential of Mg is greater than the second ionisation potential of Na
(c) The first ionisation potential of Na is less than the first ionisation potential of Mg
(d) The third ionisation potential of Mg is greater than third ionisation potential of Na

11. Which of the following has the maximum number of unpaired electrons? (1996)
(a) Mg^{2+}
(b) Ti^{3+}
(c) V^{3+}
(d) Fe^{2+}

12. Which has most stable +2 oxidation state? (1995)
(a) Sn
(b) Pb
(c) Fe
(d) Ag

13. Amongst the following elements (whose electronic configurations are given below), the one having the highest ionisation energy is (1990)
(a) $[Ne]\ 3s^2\ 3p^1$
(b) $[Ne]\ 3s^2\ 3p^3$
(c) $[Ne]\ 3s^2\ 3p^2$
(d) $[Ar]\ 3d^{10}\ 4s^2\ 4p^3$

14. Which one of the following is the smallest in size? (1989)
(a) N^{3-}
(b) O^{2-}
(c) F^-
(d) Na^+

15. The first ionisation potential of Na, Mg, Al and Si are in the order (1988)
(a) Na < Mg > Al < Si
(b) Na > Mg > Al > Si
(c) Na < Mg < Al > Si
(d) Na > Mg > Al < Si

16. The electronegativity of the following elements increases in the order (1987)
(a) C, N, Si, P
(b) N, Si, C, P

(c) Si, P, C, N
(d) P, Si, N, C

17. Atomic radii of fluorine and neon in Angstrom units are respectively given by (1987)
(a) 0.72, 1.60
(b) 1.60, 1.60
(c) 0.72, 0.72
(d) None of these

18. The first ionisation potential in electron volts of nitrogen and oxygen atoms are respectively given by (1987)
(a) 14.6, 13.6
(b) 13.6, 14.6
(c) 13.6, 13.6
(d) 14.6, 14.6

19. The hydration energy of Mg^{2+} is larger than that of (1984)
(a) Al^{3+}
(b) Na^{+}
(c) Be^{2+}
(d) Mg^{3+}

20. The element with the highest first ionisation potential is (1982)
(a) boron
(b) carbon
(c) nitrogen
(d) oxygen

21. An element with atomic number 107 has recently been discovered. Its block, group number, period and outershell electronic configuration respectively are:
(a) s-block, group 2, period 6, $6s^2$
(b) p-block, group 13, period 5, $5s^2\ 5p^4$
(c) d-block, group 7, Period 7, $7s^2$
(d) f-block, group 3, Period 6, $6s^2$

22. ln which one of the following pairs the radius of the second species is greater than that of first?
(a) Ti^{4+}, Mn^{7+}
(b) Na, Mg
(c) Cl^{-}, K^{+}
(d) P^{5+}, P^{3+}

23. The chemical formula of Aluminium plumbate is -
(a) $Al_3(PbO_3)_3$
(b) $Al_2(PbO_3)_3$
(c) $Al_2(PbO_2)_3$
(d) $Al_3(PbO_2)_3$

24. The chemical name of K_2MnO_4 is -

(a) Potassium Permangnate
(b) Potassium magnate
(c) Potassium metamagnate
(d) Potassium magnite

25. The chemical name of NaH_2PO_4 is
(a) Sodium dihydrogen phosphite
(b) Sodium dihydrogen phosphide
(c) Sodium hydrogen phosphate
(d) Sodium dihydrogen phosphate

26. The five successive ionisation energies of an element 'X' are 800,1427, 2658, 25024 and 32824 KJ mole^{-1} respectively. The valency of 'X' is:
(a) 1
(b) 2
(c) 3
(d) 4

27. For an element 'A', the first ionisation energy will be numerically equal to:
(a) EA of A^+
(b) EA of A^{2+}
(c) IE of A^{2+}
(d) none of these

28. The first ionisation energy of Al is smaller than that of Mg because :
(a) the atomic number of Al is greater than that of Mg.
(b) the atomic size of Al is less than that of Mg.
(c) Penetration of s-subshell electrons in case of Mg is greater than that of p-subshell in Al
(d) Mg has incompletely filled s-orbital.

29. In which of the following pairs, the first member has higher first ionization energy?
(a) N, O
(b) B, Be
(c) Al, Ga
(d) CI, F

30. Which of the following species will have the smallest size ?
(a) Li^+
(b) Mg^{2+}
(c) Al^{3+}
(d) K^+

31. The chemical formula of Permanganic acid is -
(a) $HMnO_4$
(b) H_2MnO_4
(c) $HMnO_3$
(d) H_2MnO_3

32. Which of the following is the correct order of ionisation enthalpy?
(a) $Te^{2-} < I^- < Cs^+ < Ba^{2+}$
(b) $I^- < Te^{2-} < Cs^+ < Ba^{2+}$
(c) $Te^{2-} < Cs^+ < I^- < Ba^{2+}$
(d) $Ba^{2+} < Cs^+ < I^- < Te^{2-}$

33. The statement that is not correct for the periodic classification of elements 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.

34. In the modern periodic table, the block indicates the value of ……………… for the last subshell that received electrons in building up the electronic configuration. Fill in the blank with appropriate option.
(a) Atomic number
(b) Azimuthal quantum number
(c) Principal quantum number
(d) Atomic mass

35. Element with electronic configuration as $[Ar]^{18}\ 3d^5 4s^2$ is placed in:
(a) 1st group, s-block
(b) 2nd group, s-block
(c) 5th group, d-block
(d) 7th group, d-block

36. Which set does not show correct matching ?

(a) Sc^{3+} [Ne] $3s^2\ 3P^6$	zero group
(b) Fe^{2+} [Ar] $3d^6$	8th group
(c) Cr [Ar] $3d^5\ 4s^1$	6th group

(d) All of the above

37. Atomic radii of F & Ne in Angstrom are respectively given by :
(a) 0.72, 1.60
(b) 1.60, 1.60
(c) 0.72, 0.72
(d) 1.60, 0.72

38. which of the following element has maximum first ionisation energy ?
(a) V
(b) Ti
(c) Sc
(d) Mn

39. 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

40. Which of the following relation is correct with respect to first (I) and second (II) ionization energies of sodium and magnesium ?
(a) $I_{mg} = II_{Na}$
(b) $I_{Na} > I_{Mg}$
(c) $II_{mg} > II_{Na}$
(d) $II_{Na} > II_{mg}$

41. First, second & third ionization energies are 737 ,1 045 & 7733 KJ/mol respectively. The element can be:
(a) Na
(b) B
(c) Al
(d) Mg

42. For which of the following process, the value of electron gain enthalpy is positive ?
(a) $F(g)+e \rightarrow F^{-}(g)$
(b) $Cl(g)+e^{-} \rightarrow Cl^{-}(g)$
(c) $Be(g)+e^{-} \rightarrow Be^{-}(g)$
(d) $B(g) + e \rightarrow B^{-}(g)$

43. Which is the correct property mentioned -
(a) $Fe^{+} < Fe^{2+} < Fe^{3+}$ — size
(b) $Fe^{+} < Fe^{2+} < Fe^{3+}$ — ionisation energy
(c) B < Be < C — size
(d) N < O < F — ionisation energy

44. 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) 2
(d) 1

45. Fluorine has the highest electronegativity among the $ns^2\ np^5$ group on the Pauling scale, but the electron affinity of fluorine is less than that of chlorine because:
(a) the atomic number of fluorine is less than that of chlorine.
(b) fluorine being the first member of the family behaves in an unusual manner.
(c) chlorine can accommodate an electron better than fluorine by utilising its vacant 3d-orbital.
(d) small size, high electron density and an increased electron repulsion makes addition of an electron to fluorine less favourable than that in the case of chlorine in isolated stage

46. The order of first electron affinity of O, S and Se is
(a) O > S > Se
(b) S > Se > O
(c) Se > O > S
(d) S > O > Se

47. Electron addition will be easier in :-
(a) O

(b) O^+
(c) O^{2+}
(d) O^{2-}

48. Electron gain enthalpy of first element in the following pairs is higher.
(a) Cl, F
(b) O, S
(c) O, F
(d) S, Cl

49. Following the Mulliken scale, what parameters are required to evaluate electronegativity ?
(a) Only electronegativity
(b) Only electron affinity
(c) electron affinity and ionization energy
(d) Ionic potential and electronegativity

50. Which one of the following oxides is neutral?
(a) CO
(b) SnO_2
(c) ZnO
(d) SiO_2

51. Which of the following orders is correct?
(a) F > N > C > Si > Ga – non –metallic character.
(b) F > Cl > O > N – oxidising property.
(c) S > Se > Te > O – electron affinity value
(d) All of these

52. Sodium sulphate is soluble in water whereas barium sulphate is sparingly soluble because
(a) the hydrogen energy of sodium sulphate is more than its lattice energy
(b) the lattice energy of barium sulphate is less than its hydrogen energy
(c) the lattice energy has no role to play in solubality
(d) the hydration energy of sodium sulphate is less than its lattice energy

53. ln which of the following compound Mn shows minimum radius?
(a) MnO_3
(b) $KMnO_4$
(c) MnO
(d) None of these

54. The ionisation energy of A(g) is similar to in terms of magnitude -
(a) electron affinity of $A^+(g)$
(b) electron affinity of A(g)
(c) ionisation energy of $A^+(g)$
(d) ionisation energy of $A^{2+}(g)$

55. Which of the following statements is INCORRECT?
(a) Generally the radius trend and the ionization energy trend across a period are opposites.

(b) Metallic and covalent radii of potassium are 2.3Å and 2.03Å respectively.
(c) Amongst Li^-, Be^-, B^- and C^-, Li^- is least stable ion
(d) Atomic and ionic radii of Niobium and Tantalum are almost same

56. Electron affinity is the:
(a) Energy absorbed when an electron is added to an isolated atom in the gaseous state
(b) Energy released when an electron is added to an isolated atom in the gaseous state
(c) Energy required to take out an electron from an isolated gaseous atom
(d) Power of an atom to attract an electron to itself

57. If x, y and z are electronegativity, ionisation potential and electron-affinity respectively. Then the electron affinity (z) in the terms of electronegativity (x) and ionisation potential (y) will be :
(a) $x=\frac{y+z}{2}$
(b) $z=\frac{x-y}{2}$
(c) $z=\frac{x^2-y^2}{2}$
(d) $z = 2x + y$

58. Which of the following statements is incorrect?
(a) ln the long form of periodic table, the number of period indicates the value of principal quantum number.
(b) There are four d-block series comprising of total 40 elements in the long form of periodic table
(c) s-block, d-block and f-block elements are metals.
(d) All p-block elements are non-metal.

59. If the same element is forming oxides in different oxidation states then :
(a) that oxide will be neutral in nature in which element will be in its highest oxidation state
(b) that oxide will be highest acidic in nature in which element will be in its highest oxidation state
(c) that oxide will be amphoteric in nature in which element will be in its highest oxidation state.
(d) that oxide will be highly basic in nature in which element will be in its highest oxidation state

60. Which of the following statement is correct ?
(a) Ionisation energies of elements decrease along the period.
(b) Ionisation energies of the IIA group elements are less than that of the corresponding IIIA group elements
(c) Ionisation energies of group 15 elements are less than that of the corresponding group 16 elements.
(d) Ionisation energy of Ga is greater than Al.

61. The dominating factor responsible for the decreasing ionisation energies of the elements on moving down the group is:
(a) atomic radius
(b) type of electron to be removed
(c) the valence shell electron configuration
(d) all of these

62. Which of the following order is not correct?
(a) IE(l) of Be > IE(l) of B but IE(II) of Be < IE(II) of B
(b) IE(l) of Be < IE(l) of B but IE(II) of Be < IE(II) of B

(c) IE(II) of O > IE(II) of N
(d) IE(l) of Mg > IE(l) of Al

63. The correct order of the metallic character is :
(a) Na > Mg >Al > S
(b) Mg > Na >Al > S
(c) Al > Mg > Na > S
(d) Si > Al > Na > Mg

64. The correct order of the non-metallic character is :
(a) B > C > N > F
(b) C > B > N > F
(c) F > N > C > B
(d) F > N > B > C

65. Which of the following statement is incorrect ?
(a) Oxide of aluminium (Al_2O_3), and arsenic (As_2O_3) are amphoteric.
(b) Oxide of chlorine (Cl_2O_7) is less acidic than oxide of nitrogen (N_2O_5).
(c) Oxide of carbon (CO_2) is more acidic than oxide of silica (SiO_2).
(d) The correct increasing order of basic character of various oxides is $H_2O < CuO < MgO < CaO$.

66. Higher values of ionisation energies of the 5d-transition elements are consistent with the -
(a) relatively smaller effective nuclear charge
(b) relatively smaller size of their atoms
(c) relatively smaller penetration effect of inner orbitals
(d) all of the above

67. The correct decreasing order of size of isoelectronic species is -
(a) $Se^{-2} > Br^{-} > Kr > Rb^{+} > Sr^{+2}$
(b) $S^{-2} > Cl^{-} > K^{+} > Ar > Ca^{+2}$
(c) $N^{-3} > O^{-2} > Ne > F^{-} > Ca^{+2}$
(d) $F^{-} > Ne > Na^{+} > Al^{+3} > Mg^{+2}$

68. Which one of the following is not the representative element?
(a) Fe
(b) K
(c) Ba
(d) N

69. An element of atomic mass 40 has 2, 8,8,2 as the electronic configuration. Which one of the following statements regarding this element is not correct?
(a) It forms basic oxide
(b) It belongs to II A group
(c) It belongs to IV period
(d) It forms an acidic oxide

70. Ionic radii is/are :
(a) directly proportional to effective nuclear charge

(b) directly proportional to square of effective nuclear charge
(c) inversely proportional to effective nuclear charge
(d) inversely proportional to square of effective nuclear charge

71. Which of the following is not a correct match ?

(a) Cr^-, P^{3-}, Ar	Isoelectronics
(b) Size of Mo = size of W	Lanthanide contraction
(c) IP of ‘Be’ > IP of ‘B’	Penetration effect
(d) Size of Ne > size of F	Due to complete octet of Ne

72. Find the atomic number of element belonging to 4th period and 17th group in Modern periodic table:
(a) 17
(b) 25
(c) 59
(d) 35

73. Correct order of the property indicated below -

(a) Na > Al > Mg > SI	IE_1
(b) Cl > S > O > F	ΔHeg_1
(c) F > O > Cl > S	EN
(d) Cl > O > F > S	EA_1

74. Consider the following points:
(i) Cs is the strongest reducing agent in IA group element
(ii) $Be(OH)_2$ is amphoteric
(iii) The density of potassium is less than sodium
(iv) In alkali metals Li, Na, K and Rb, lithium has the minimum value of M.P.
Correct statement are:
(a) (i) & (ii) are correct
(b) (i), (ii) & (iii) are correct
(c) (ii) & (iii) are correct
(d) (ii), (iii) & (iv) are correct

75. The electronic configurations of four elements are given below. Element with highest order of magnitude of electron affinity is.
(a) $2s^2 2p^3$ (N)
(b) $3s^z 3p^5$ (Cl)
(c) $2s^2 2p^4$ (O)
(d) $3s^2 3p^4$ (S)

Answers:

1. (c) 2. (b) 3. (d) 4. (d) 5. (b)
6. (c) 7. (b) 8. (b) 9. (b) 10. (b)
11. (d) 12. (b) 13. (b) 14. (d) 15. (a)
16. (c) 17. (a) 18. (a) 19. (b) 20. (c)
21. (c) 22. (d) 23. (b) 24. (b) 25. (d)

26. (c) 27. (a) 28. (c) 29. (a) 30. (c)
31. (a) 32. (a) 33. (d) 34. (b) 35. (d)
36. (a) 37. (a) 38. (d) 39. (b) 40. (d)
41. (d) 42. (c) 43. (b) 44. (b) 45. (d)
46. (b) 47. (c) 48. (a) 49. (c) 50. (a)
51. (d) 52. (a) 53. (b) 54. (a) 55. (c)
56. (b) 57. (a) 58. (d) 59. (b) 60. (d)
61. (a) 62. (b) 63. (a) 64. (c) 65. (b)
66. (b) 67. (a) 68. (a) 69. (d) 70. (d)
71. (d) 72. (d) 73. (c) 74. (d) 75. (b)

“Detail solutions are mentioned in the content library”